Chemical Energetics and Change

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Preface

First Edition

This is a collection of lecture notes for chem 20B at UCLA, from when I taught during the Winter quarter of 2018 and Spring quarter of 2017, as well as chem 20BH (honors version) Winter quarter of 2015. Chem 20B ("Energetics and Change") is the second quarter of general chemistry for physical sciences students. Chem 20B mainly follows the second half of the textbook by Oxtoby, Gillis and Campion, “Principles of Modern Chemistry” [1]. The lecture notes were originally based on certain chapters of the book (hence the similar structure and organization), but evolved over time to provide more in-depth coverage of selected topics. This document is in the form of “lecture notes” rather than a textbook; the former being less wordy than the latter. The word *energetics* here refers to all aspects of energy, as it relates to chemical transformations and more generally speaking, physical interactions and the conversion of energy between thermodynamical systems.

The pre-requisites are calculus (differential and integral) as well as high-school or freshman physics. For example I assume you know introductory concepts in classical mechanics and electromagnetism, such as forces, Newton’s laws, Coulomb’s law, etc. I did include exercises to review and/or perfect these skills. Other concepts such as probability and partial differentiation, which will likely be new concepts to most freshman students, are introduced as needed, either in the various chapters or in the “Appendix” chapter. Practice may be found by working through end-of-chapter problems, whose solutions are provided for self-study. Some students may find it helpful to supplement these class-notes with an introductory textbook on
general chemistry. The latter typically contain more extensive discussions, worked examples, as well as end-of-chapter practice problems.

I have assembled the notes into a single PDF file during the summer of 2018, with the help of some UCLA students that offered to draw and redraw some of the figures. Many thanks go to Ellee Vikram and Emily Hays (two former chem 20B students) for drawing the illustrations. Ellee has also kindly proofread the entire book for typos and overall logical flow and organization. Any remaining errors are mine. These notes are made available on the internet as a downloadable PDF file, so they can remain accessible to anyone anywhere, free of charge.

For those wishing to pursue more in-depth studies, more advanced textbooks exist. To learn more about intermolecular forces, see “Intermolecular and surface forces” by Israelachvili [2] or “The molecular theory of gases and liquids” by Hirschfelder [3]. For statistical mechanics, McQuarrie has authored an excellent textbook (“Statistical mechanics”) [4]. McQuarrie and Simon’s “Molecular thermodynamics” [5] provides excellent coverage of both equilibrium and non-equilibrium thermodynamics. Kjelstrup and Bedeaux (“Non-equilibrium thermodynamics for engineers”) [6] as well as De Groot and Mazur (“Non-equilibrium thermodynamics”) [7] are also recommended. To learn more about probability, the book by A.N. Shiryaev (“Probability”) [8] is outstanding. For chemical kinetics, see Ancheyta’s “Chemical reaction kinetics” [9] as well as Chorkendorff and Niemantsverdriet (“Concepts of modern catalysis and kinetics”) [10].

Finally, a few words about notation. For work, the unusual notation $\Delta W$ is used rather than $w$, to emphasize that it is obtained from infinitesimal quantities, i.e. $\Delta W = \int \delta W$ (along some path). Same for heat transfer, the notation $\Delta Q$ is used instead of $q$ because $\Delta Q = \int \delta Q$ (along some path). The “Capital Delta” notation is used to emphasize that those quantities are obtained by integrating a differential 1-form along a given path. I use boldface and arrows to denote vectors ($\mathbf{E}$ and $\vec{E}$); both notations mean the same thing. Instead of $\ln$ the natural log is written $\log$ to avoid confusion with $\ln$. Integration volume is often written $\text{vol}$ to avoid confusion with $V$, the thermodynamic system volume, when the two are different. The shorthand notation $\partial_x$ is sometimes used for the partial derivative $\frac{\partial}{\partial x}$. Finally, $\nabla f$ denotes the gradient of a function $f$, i.e. the vector $\nabla f = (\partial_x f, \partial_y f, \partial_z f)$.

Louis Bouchard
Los Angeles, CA
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Gases

1.1. The Ideal Gas

The famous ideal gas law, \( PV = nRT \) (\( P \), pressure; \( V \), volume; \( n \), number of moles; \( R \), gas constant; \( T \), temperature), originated in 18th research by Boyle, Avogadro, and Charles. The ideal gas consists of an ensemble of \( N \) classical molecules which are moving randomly (due to thermal motion at positive temperature, \( T > 0 \)) and do not interact with each other. \( N \) and \( n \) are related by the Avogadro’s constant, \( N_A \), which counts the number of atoms in 12 grams of \(^{12}\text{C} \) isotope:

\[
N \equiv n \cdot N_A \approx N \cdot (6.022 \times 10^{23} \text{ molecules/mol}).
\]

The statement that the particles do not interact with each other implies that the gas density is low enough, so that intermolecular forces are not important. No interactions means no collisions. No collisions implies that the cross-sectional area of the particles is very small. In Fig. 1.1 the molecules should be thought of as as point particles. At low densities, the average distance, \( \langle r_{12} \rangle \), between nearest neighbor molecules (labelled “1” and “2” here)

\[
\langle r_{12} \rangle \sim \sqrt[3]{\frac{V}{N}},
\]

is large, much larger than the molecular diameter. \( N/V \) is called the number density. At higher densities, the assumption of non-interacting particles is no longer valid and we must invoke principles of kinetic molecular theory to describe the behavior of the gas. At very low temperatures (\( \ll 1 \text{ K} \)), quantum effects become important and the ideal gas law breaks down [11].
1. Gases

While the properties of liquids and solids can be quite different among different substances, the properties of different types of gases are very similar to each other, especially in the low density limit, where the ideal gas law applies. The ideal gas law reflects the view that the properties of a gas \((P, V, n, T)\) do not depend on the type of gas. This implies, for example, that 1 mole of ideal gas at 300 K and 1 atm occupies exactly the same volume regardless of the type of gas (e.g., oxygen, nitrogen, xenon, etc.). It also implies that the total pressure in a mixture of different types of non-reacting gases is the sum of the partial pressures of each constituent gas (Dalton’s law, see Section 1.2.2).

1.1. Equation of State. A thermodynamic system is described by an equation of state, which describes a relationship between the variables that describe the state of matter. It is a relation of the form:

\[ f(n, P, T, V) = 0, \]

where \(n\) is the number of moles of a substance, \(P\) is the pressure of the gas, \(T\) is its temperature and \(V\) is the volume occupied by the substance. For a given set of physical conditions (e.g. pressure, volume, temperature, etc.) a state function is useful for analyzing the behavior of the system (gases, liquids and solids).
The function $f$ is not necessarily smooth\(^1\) and may contain discontinuities, especially at the boundaries between different phases of matter. Phase transitions that are “abrupt” will give rise to discontinuities in $f$. The ideal gas law ($PV = nRT$), which can be put in this form as $f(n, P, T, V) = PV - nRT = 0$, appears to be a smooth function of its variables. However, this is only because this analytical expression describes a gas state only; it does not describe the transition to the liquid or solid phases. This can be seen by plotting the function (we do this in Section 1.4.4). In Section 1.4.2 we discuss the equation of state for a Van Der Waals (VDW) gas, which can describe a transition from gas to liquid phase.

Real substances will undergo phase transitions and in principle, $f$, is defined over all possible states of the substance, and may include discontinuities at the boundaries of phase transitions. We note that the equation of state is often obtained numerically from a discrete set of experimental data points and may not always exist in the form of a mathematical formula (analytical expression).

1.1.2. Description of the Ideal Gas. In the context of the ideal gas law, it is sufficient to define the ideal gas as a set of noninteracting classical molecules. The state of the gas is described by its macroscopic properties: $T, n, P, V$. Particles have zero size (no collisions, no intermolecular interactions); collisions will be discussed later, in the context of kinetic theory. The substance (gas) is characterized by macroscopic variables such as $T$ (temperature), $P$ (pressure), $N$ (number of particles) and $V$ (volume of container). The mass of the particles plays no role in the description of the gas. The equation of state for the ideal gas can be derived using three important concepts: Boyle’s law, Charles’ law and Avogadro’s law. If we think of the $n, P, T, V$ space as 4-dimensional, the equation of state adds a constraint which reduces the dimensionality of the space to 3. The ideal gas molecule is required to “live” in that 3-dimensional space. For fixed $n$, we can take $n/V$ (density) to be a single variable. In that case, $n/V, P, T$ is a 3-dimensional space and the ideal gas law constraint reduces it to a 2-dimensional space (surface), as you will see in Section 1.4.4.

1.1.3. Boyle’s Law (Robert William Boyle, 1627-1691). In the 17th century Robert Boyle trapped a gas within a glass tube with one end sealed. Mercury was added to the tube to trap the gas. Any changes in the gas volume could be measured as changes in the mercury level. The glass tube was J-shaped, with a long end and a short end. The difference in mercury levels on both ends of the tube enabled him to determine the gas pressure

---

\(^1\)A smooth function is a function that can be differentiated with respect to its variables as many times as desired, and all of its derivatives are continuous.
(assuming that the ambient pressure to which the other end of the tube was in contact which remained constant). Boyle’s 1662 J-tube experiment is illustrated in Fig. 1.2.

![Figure 1.2. Boyle’s law describes the inverse proportionality between pressure and volume ($P \propto 1/V$). Left: When the heights of mercury on the two sides of the tube are the same, the pressure of the confined gas must equal that of the atmosphere, 1 atm or 760 mm Hg. Right: Boyle then added mercury to the J-tube. This caused the pressure of the gas to be increased by the number of millimeters of mercury in the height difference $h$. The compression of the gas causes it to occupy a smaller volume.

Boyle’s law: In his experiments, Boyle noted that the gas volume varied inversely with pressure

$$P \propto \frac{1}{V}, \quad \text{at constant } T, n$$

$P$: pressure of the gas (force/area)

$V$: volume of the gas within the container

This relationship is independent of the gas type. In the case of the J-tube, the gas of interest is trapped at the sealed end by mercury.

An equivalent way to express this law is:

$$PV = \text{constant}.$$ 

This constant does not change as long as the temperature and number of gas and mercury molecules are held constant. This means that:

$$P_1V_1 = \text{constant} = P_2V_2$$
where $P_1, V_1$ denote the pressure and volume (respectively) of the system in state 1 (i.e. at time 1), whereas $P_2, V_2$ are pressure and volume of the system in state 2 (i.e. at time 2). For this to hold, everything else must be held constant. This constant is also independent of the type of gas.

Example: Consider a rigid sphere filled with air. At time 1 its volume is 50 L and the internal air has pressure 1 bar. The number of molecules is 2 mol, which remains constant because the sphere does not leak. At time 2, we manage to compress the size of the rigid sphere by half so that its volume is now 25 L. What is the pressure inside? (Temperature and number of molecules are held constant.) Answer: It will be 2 bars because $P_1V_1 = P_2V_2$ or $P_2 = P_1V_1/V_2 = (1 \text{ bar})(50 \text{ L}/25 \text{ L}) = 2 \text{ bars}$.  

### 1.1.4. Charles’ Law (Jacques Alexandre César Charles, 1746-1823).

Charle’s law, also known as the law of Charles and Gay-Lussac, involves the concept of temperature. Charles (and later, Gay-Lussac, in similar experiments) found experimentally that different gases expand to the same extent over the same temperature interval. This established a linear relationship between volume and temperature:

| Charles’ law: The gas volume varies in proportion with temperature |
| $V \propto T$ at constant $P, n$ |
| $T$: temperature of the gas |
| $V$: volume of the gas within the container |

The constant of proportionality depends on the pressure and number of molecules of gas present, but not on the type of gas.

Figure 1.3 illustrates the experimental finding that the volume of a gas confined at constant pressure increases as the temperature increases. Since $V \propto T$ we can write $V/T = \text{constant}$. The constant is independent of gas type. It does not change as long as the pressure and number of molecules are kept fixed. From this, we can then write

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}.$$

### 1.1.5. Avogadro’s Law (Lorenzo Romano Avogadro, 1776-1856).

Avogadro’s law states that the volume is proportional to the number of molecules $N$ provided that temperature and pressure are held constant:

$$V \propto N.$$

You can imagine inflating a balloon by blowing two identical puffs of air consecutively. The first puff ($N$ molecules) will inflate the balloon from volume 0 to volume $V_1$. This volume $V_1$ corresponds to $N$ molecules. The
second puff, because it is identical to the first one, adds \( N \) more molecules so that \( N_{\text{final}} = N + N = 2N \). Its volume will have doubled, \( V_2 = 2V_1 \).

If the balloon is spherical, doubling the volume corresponds to a radius or diameter increase of \( (2)^{1/3} \approx 1.26 \) (26% higher), since \( V = \frac{4}{3} \pi r^3 \).

Avogadro’s law is also known as the extensive property of \( V \). An extensive property is often defined as one which scales with the size of the system, i.e. \( V \) doubles if \( N \) doubles (and vice-versa).

1.1.6. Some Definitions. We will make frequent use of the following terminology, especially in the context of thermodynamics:

- **Macroscopic property**: properties which involve many molecules.
- **Microscopic property**: properties which involve one or few molecules.
- **Extensive property**: properties which depend on the size of the system (e.g. total mass, volume, number of moles of substance).
- **Intensive property**: properties which are not extensive (e.g., temperature, pressure, density, molar volume, molar mass).

1.1.7. Combining All Three Laws to Get Ideal Gas Law. We summarize the above laws:

- \( P \propto \frac{1}{V} \) at constant \( T, N \) (Boyle’s law)
- \( V \propto T \) at constant \( P, N \) (Charles’ law)
- \( V \propto N \) at constant \( T, P \) (extensive property of \( V \))

If we combine these laws together we get:

\[
V \propto \frac{NT}{P}
\]
1.1. The Ideal Gas

The proportionality constant is known as Boltzmann’s constant, \( k_B \). We get the ideal gas law:

<table>
<thead>
<tr>
<th>Ideal gas law: Equation of state for the ideal gas:</th>
</tr>
</thead>
<tbody>
<tr>
<td>( PV = Nk_B T )</td>
</tr>
</tbody>
</table>

- **\( P \):** Pressure of the gas (force/area)
- **\( T \):** temperature of the gas (Kelvin scale: \( 0^\circ C = 273.15 \text{ K} \))
- **\( N \):** number of particles in the container
- **\( V \):** volume of the container
- **\( k_B = R/N_A \):** Boltzmann’s constant (\( k_B = 1.38 \times 10^{-23} \text{ J/K} \))

We can also express the ideal gas law in terms of the number of moles (\( n \): number of moles) by using the relationship (1 mole = \( N_A = 6.022 \times 10^{23} \) molecules)

\[
nR = Nk_B
\]

where \( R = 8.31 \text{ J/mole.K} \) (ideal gas constant). Note: since \( k_B = nR/N \), plugging \( n = 1 \) mol and \( N = N_A \) molecules gives \( k_B = 8.31/(6.022 \times 10^{23}) = 1.38 \times 10^{-23} \text{ J/K} \).

<table>
<thead>
<tr>
<th>Ideal gas law (if working in moles): The alternative form is:</th>
</tr>
</thead>
<tbody>
<tr>
<td>( PV = nRT ).</td>
</tr>
</tbody>
</table>

1.1.8. Units of Pressure. It is very important to know the units of pressure and how they are related to each other. The unit of pressure in the SI system is the pascal (Pa), defined as a force of one Newton per square meter. The conversion between atm, Pa, and torr is as follows: 1 atm = 101,325 Pa = 760 torr. Another common unit of pressure is the psi (pounds per square inch): 1 atm = 14.69595 psi. Another unit is bar: 1 bar = 100 kPa (not to be confused with atm, where 1 atm = 101.325 kPa).

1.1.9. Gases at Standard Conditions. Chemists define standard conditions as standard temperature and pressure (STP):

STP: \( 0^\circ C \) (273.15 K) and 1 atm (760 torr).

Under these conditions, the volume of 1 mol of ideal gas is called the standard molar volume:

standard molar volume = 22.414 L, or 22.4 L (to 3 significant figures).

Indeed, this follows from the ideal gas law,

\[
V = \frac{nRT}{P} = \frac{(1 \text{ mol})(0.082057 \text{ L.atm/mol/K})(273.15 \text{ K})}{(1 \text{ atm})} = 22.414 \text{ L}.
\]

This particular value of \( R \) and its units (0.082057 L.atm/mol/K) are explained at the end of Section 1.4.2.
1.1.10. **Ideal Gas Law Does Not Depend on Gas Type.** You may ask, Why don’t these properties depend on the type of molecule/atom in the gas? For example, why isn’t pressure (at fixed $n, V, T$), which is force per unit area, higher for a heavy atom than a light one? This has to do with the velocity of molecules at a given temperature, which is lower for heavy atoms.

The force exerted by a molecule colliding against a wall is related to the transfer of momentum $\Delta p$, via $F = \Delta p/\Delta t$, where $\Delta t$ is a time interval. The change in momentum is $2mv$, where $m$ is the molecular mass and $v$ is the molecule’s velocity. Let’s take $v$ to be an average velocity. In a cubic container of side length $L$, $\Delta t$ is equal to $\Delta t = 2L/v$, so that $P = F/L^2 \propto mv^2$.

The independence of $P$ on molecular mass means, for example, that $m_{O_2}v_{O_2}^2 = m_{He}v_{He}^2$, where $m_{He} = 4$ amu and $m_{O_2} = 32$ amu. Thus, if the mass is 8 times larger, the velocity must be $\sqrt{8}$ times slower.

Oxygen molecules at room temperature have an average speed of about 480 m/s. This means that helium molecules at room temperature must have an average speed of approximately $\sqrt{8}$ times 480 $\approx$ 1,360 m/s. This is the only way that the ideal gas law can hold.

We will see later (in Section 1.3.9) that the kinetic theory of gases tells us that the speed of gas molecules is inversely proportional to the square root of the molecular mass, e.g.

$$v_p = \sqrt{\frac{2k_B T}{m}}, \quad \overline{v} = \sqrt{\frac{8k_B T}{\pi m}}, \quad v_{rms} = \sqrt{\frac{3k_B T}{m}}.$$  

where $v_p$ is the most probable speed, $\overline{v}$ is the average speed and $v_{rms}$ is the root-mean-square speed.

1.1.11. **Ideal Gas Example: Heating Gas at Constant Volume.** Suppose that we trap a gas inside a container of fixed volume $V$. The temperature is $T_1 = 293$ K and the pressure $P_1$ is 1 atm. There is a pressure gauge attached to the volume. Now we place the container in boiling water ($T_2 = 393$ K). What is the pressure gauge showing?

**Answer:** We can invoke the ideal gas law before and after, and use the fact that the ideal gas constant $R$ is the same both before and after, i.e. $PV = nRT$ is written as $R = PV/nT$. Then: $P_1V_1/n_1T_1 = P_2V_2/n_2T_2$. Since the number of moles doesn’t change (container is leak-proof) and the volume remains fixed, we can cancel out the $n$’s and the $V$’s: $P_1/T_1 = P_2/T_2$. Solving for $P_2$ gives $P_2 = P_1(T_2/T_1)$. Inserting the numbers we get:

$$P_2 = (1 \text{ atm}) \left(\frac{393}{293}\right) \approx (1 \text{ atm})(1.34) = 1.34 \text{ atm.}$$
1.1. The Ideal Gas

Note: we used an equation of state (namely, the ideal gas law). An equation of state means that the state of the gas only depends on the current values of its thermodynamic variables (e.g., $P, T, V, n$), and not on their histories. Fortunately, the solution of this problem does not depend on how fast or slow the heating takes place.

1.1.12. Phase Diagrams. A phase diagram is a plot which shows the different phases of a given substance as function of the conditions (macroscopic variables, such as $T, P, V$). The substance is in thermal equilibrium. On the phase diagram you will find regions corresponding to the different phases (e.g. solid, liquid, gas), as well as “phase boundary” lines between the different phases. These lines mark conditions under which multiple phases can coexist at equilibrium. Phase transitions occur along lines of equilibrium.

The triple point marks the condition where the substance coexists as a solid, liquid and a gas simultaneously. Different substances have different triple points. The critical point is the point where the substance may transition to the supercritical fluid state. To be in the supercritical state, the temperature and pressure must be at or above the critical temperature and pressure, respectively. Different substances have different critical points. Each substance has a different phase diagram, the characteristics of which are determined by intermolecular forces. Intermolecular forces are discussed in Chapter 2.

A phase diagram should describe real substances. Typically, phase diagrams are obtained from experimental measurements. For any substance, there exists an “equation of state” describing a constraint between $P, T$ and $N/V$. This equation of state can be obtained from experimental data\(^2\). A hypothetical phase diagram is shown in Fig. 1.4.

Do not confuse Fig. 1.4 with the ideal gas law, as the ideal gas law does not describe solids or liquids. The ideal gas law applies only to gases in the low-density limit (e.g., the bottom-right region of $P − T$ diagrams such as Fig. 1.4. For example, take a fixed value of the temperature $T$. Consider the compression of a gas inside a cylinder by a piston. Isothermal compression means that we move along a vertical line ($T$ fixed) from bottom to the top on the diagram of Fig. 1.4. At low gas densities, $PV = constant$ (Boyle’s law), and the gas pressure $P$ increases as $V$ decreases. As $P$ increases, the gas eventually undergoes a phase transition from gas to liquid or gas to solid (depending on the value of $T$). The ideal gas law does not describe the transition to the liquid, nor does it describe the state of the liquid itself.

\(^2\)Some equations of state, such as the Van Der Waals equation of state, are obtained from theoretical arguments.
For fixed $N$, an equation of state $f(P, T, N/V) = 0$ (such as $PV = Nk_BT$, i.e. $P - k_BT N/V = 0$) is a constraint that defines a 2D surface in the 3D space of $T$, $P$ and $N/V$ ($N/V$=number density). Thus, phase diagrams can also be represented as surfaces, such as the one shown in Fig. 1.5. In these diagrams, “volume” refers to the specific volume (ratio of a substance’s volume to its mass). Specific volume is the inverse of mass density.

### 1.1.13. Example: Pressurized CO$_2$ in Tank.

Suppose that we have a gas cylinder (40 cm high, 15 cm diameter) at ambient temperature filled with CO$_2$. The cylinder contains 10 lbs of CO$_2$. The questions we ask is whether the tank is filled with liquid CO$_2$ or gaseous CO$_2$. And if so, what is the pressure?

**Answer:** The tank is a cylinder whose volume is $V = (40 \text{ cm})(\pi r^2) = (0.4 \text{ m}) (\pi (0.075 \text{ m})^2) \approx 7.1 \times 10^{-3} \text{ m}^3$. How many moles do we have in the tank? The mass of CO$_2$ is $m=10 \text{ lbs} = 4,500 \text{ g}$. (1 lb is 450 g.) The molecular mass of CO$_2$ is 12+32=44 g/mol. Thus, $n = 4,500 \text{ g}/(44 \text{ g/mol}) \approx 10^2 \text{ mol}$. Ambient temperature is 293 K. If this were a gas, we would have $P = nRT/V \approx 3 \times 10^7 \text{ Pa} \approx 300 \text{ atm} \approx 300 \text{ bar}$. This is a very high pressure. We suspect that there is probably no gas inside the tank because at that high pressure, the CO$_2$ probably becomes a liquid.
The phase diagram for CO$_2$ is shown in Fig. 1.6. This graph shows that we cannot have a pure gas in the tank. At 293 K, increasing the pressure leads to a liquid. We have either a liquid or a supercritical fluid, depending on the exact position of the vertical dashed line. (Although I looked up the critical point of CO$_2$, and found it to be 73 atm and 31°C, which suggests we have a liquid since our temperature is not high enough.) $P$ inside the tank is most definitely not equal to 300 atm because the gas eventually turns into a liquid (phase transition).

1.2. Mixtures of Gases (ideal)

1.2.1. Partial Pressures. A consequence of the ideal gas law is that the total pressure of a gas mixture, $P_{tot}$, is a sum of the partial pressures of each individual gas components, i.e.,

$$P_{tot} = n_{tot} \frac{RT}{V} = (n_1 + n_2 + \cdots + n_N) \frac{RT}{V} = P_1 + \cdots + P_N$$

$$= X_1 P_{tot} + \cdots + X_N P_{tot},$$

where $n_i$ is the number of moles of species $i$, $n_{tot} = n_1 + \cdots + n_N$ is the total number of moles in the mixture, $X_i = n_i/n_{tot}$ is the mole fraction of species $i$ and $P_i = n_i RT/V = X_i P_{tot}$ is the partial pressure of species $i$. This result follows because the ideal gas law only cares about the number of moles (at fixed $T,P$), not the nature of the molecules.
Let’s look at an example. Suppose that a gas mixture consists of $n_{N_2}$ moles of $N_2$ and $n_{O_2}$ moles of $O_2$. The **mole fraction** $X_{N_2}$ of $N_2$ is defined as:

$$X_{N_2} = \frac{n_{N_2}}{n_{N_2} + n_{O_2}} = \frac{n_{N_2}}{n_{tot}}.$$

Similarly, the mole fraction $X_{O_2}$ of $O_2$ is defined as:

$$X_{O_2} = \frac{n_{O_2}}{n_{N_2} + n_{O_2}} = \frac{n_{O_2}}{n_{tot}}.$$

In particular, the **partial pressure** of $N_2$ within the gas container is given by:

$$P_{N_2} = n_{N_2} \frac{RT}{V} = X_{N_2} P_{tot},$$

where $P_{tot} = P_{N_2} + P_{O_2}$ is the total pressure in the container. The second equality follows from:

$$X_{N_2} P_{tot} = \left(\frac{n_{N_2}}{n_{tot}}\right) \left(\frac{n_{tot} \frac{RT}{V}}{V}\right) = n_{N_2} \frac{RT}{V}.$$

Similarly, the partial pressure of $O_2$ within the gas container is:

$$P_{O_2} = n_{O_2} \frac{RT}{V} = X_{O_2} P_{tot}.$$

### 1.2.2. Dalton’s Law For Gas Mixtures.

The result from the previous section is known as Dalton’s law. Dalton’s law says that the total pressure of a gas mixture is the sum of the pressures exerted by the individual gases. This is illustrated in Fig. 1.7 for the case of two gases mixed together into one tank. Note that the total volume is the same in all three containers. The total pressure is the sum of the partial pressures:
1.3. Translational Molecular Motion

1.3.1. Estimating Pressure in Terms of Wall Collisions. Pressure arises from collisions of gas molecules against a wall (Fig. 1.8). An ensemble of \( N \) gas molecules is held in a container of volume \( V = L^3 \) at temperature \( T \). Consider elastic collisions of gas molecules against the shaded wall whose area is \( A = L^2 \) as shown below. Only the \( v_x \) component of the molecule’s velocity matters, as orthogonal components \( (v_y, v_z) \) do not lead to collisions against the shaded wall.

When a gas molecule collides with the wall in an elastic collision, its \( x \) velocity component \( v_x \) is reversed \( \rightarrow -v_x \). Consider a molecule which is initially near the left wall, moves to the right, collides with the shaded wall, then bounces back to the left wall. Its momentum change is:

\[
\Delta p_x = p_{x, before} - p_{x, after} = p_{x, before} - (-p_{x, before}) = 2p_{x, before} = 2mv_x.
\]
Figure 1.8. Pressure on the wall is force per unit area exerted by the molecules colliding on the wall and abruptly changing their momentum upon collision.

\[ \Delta t = \frac{2L}{v_x}. \]

Force exerted on the shaded wall is obtained from Newton’s second law

\[ F_x = \frac{\Delta p_x}{\Delta t} \approx \frac{\Delta p_x}{2L/v_x} = \frac{2mv_x}{L}. \]

The total force on the wall arising from \( N \) molecules is

\[ F_{x,tot} = \sum_{i=1}^{N} F_{x,i} = \sum_{i} \frac{m}{L}v_{x,i}^2 = \frac{Nm}{L} \left( \frac{1}{N} \sum_{i} v_{x,i}^2 \right) = \frac{Nm\overline{v_x^2}}{L}. \]

The bar denotes the average over \( N \) molecules. If there is no preference for the velocity direction (isotropic motion) we have that

\[ \overline{v^2} = \overline{v_x^2} + \overline{v_y^2} + \overline{v_z^2} = 3\overline{v_x^2} \]

where \( \bar{v} = (v_x, v_y, v_z) \). Therefore, \( \overline{v_x^2} = \frac{1}{3}\overline{v^2} \).

**Pressure and Kinetic Energy:** The pressure \( P \) on the shaded wall is force, \( F = F_{x,tot} \), divided by the wall area \( (L^2) \):

\[ P = \frac{F_{x,tot}}{L^2} = \frac{Nm\overline{v_x^2}}{3L^3}. \]

Since \( V = L^3 \), this can be written in the form of an ideal gas law:

\[ PV = \frac{1}{3} Nmv^2. \]
If we denote the average kinetic energy per molecule as
\[ K_{av} = \frac{1}{2} m v^2 \]
we see that the pressure, a macroscopic property, is related to the average translational kinetic energy per molecule, which is a microscopic property.

\[ P = \frac{2}{3} \frac{N K_{av}}{L^3} = \frac{2}{3} \frac{N K_{av}}{V} \quad \text{or} \quad PV = \frac{2}{3} N K_{av}. \] (1.1)

1.3.2. Temperature and Kinetic Energy of Molecules. Molecular motions can be related to temperature of the gas if we invoke the ideal gas law:

\[ PV = N k_B T. \] (1.2)

**Temperature and Kinetic Energy:** Comparison of (1.1) and (1.2) yields the important result that the average molecular kinetic energy is proportional to the absolute temperature:

\[ T = \frac{mv^2}{3k_B} = \frac{2}{3} \frac{K_{av}}{k_B}. \]

Thus, temperature is a measure of the average kinetic energy of the atoms in a monatomic gas. (This is valid only for a monatomic gas because the kinetic energy, \( K \), would be different than \( \frac{1}{2}mv^2 \) in the case of molecules, which possess internal rotational and translational motions.)

1.3.3. Derivation of Ideal Gas Law. To “prove” or derive the ideal gas law from kinetic theory\(^3\) we need to show that \( \overline{v^2} \) is proportional to temperature:

\[ \overline{v^2} = \int_0^\infty v^2 f(v)dv \propto T, \]

where \( f(v)dv \) is the probability that a molecule chosen at random in the gas will have speed in the range \([v, v + dv]\). Ludwig Boltzmann derived the result

\[ f(v) = \left( \frac{m}{2\pi k_B T} \right)^{3/2} 4\pi v^2 \exp \left( -\frac{mv^2}{2 k_B T} \right) \quad \{ \text{Maxwell-Boltzmann speed distribution} \} \]

\(^3\)The combination of the results \( T = \frac{2}{3} \frac{K_{av}}{k_B} \) and \( P = \frac{2}{3} \frac{N K_{av}}{V} \) does not lead to the ideal gas law, \( PV = N k_B T \), because we’ve had to invoke the ideal gas law in order to obtain the relationship between temperature and kinetic energy, \( T = \frac{2}{3} \frac{K_{av}}{k_B} \).
for the speed of a particle of mass \( m \), where the speed \( v \) is \( \sqrt{v_x^2 + v_y^2 + v_z^2} \) and the velocity is \( \vec{v} = (v_x, v_y, v_z) \). Then

\[
K_{av} = \frac{1}{2} mv^2 = \frac{m}{2} \overline{v^2} = \frac{m}{2} \int_0^\infty v^2 f(v)dv = \frac{3}{2} k_B T.
\]

This equality is called the equipartition theorem. This theorem states that at thermal equilibrium, every particle has exactly the same average kinetic energy, \((3/2)k_B T\). Boltzmann thus established a definition of temperature (macroscopic quantity) in terms of the kinetic energy of the gas particles (microscopic quantities), \( T = (2/3)K_{av}/k_B \).

1.3.4. Root-Mean-Square Speed. In the previous section we found that \( \frac{1}{2} mv^2 = \frac{3}{2} k_B T \). Cancelling out the 1/2, solving for \( \overline{v^2} \) and taking the square root gives the root-mean-square speed:

\[
\overline{v_{rms}} = \sqrt{\overline{v^2}} = \sqrt{\frac{3k_B T}{m}} = \sqrt{\frac{3RT}{M}}.
\]

where \( M \) is the molar mass.

1.3.5. Interpretation of Maxwellian Density. The velocity of the center-of-mass of a molecule can be denoted by \( \vec{V} = (V_x, V_y, V_z) \) (capital \( V \)). We view this velocity as a random variable, in the sense that the act of picking a molecule at random within an ensemble is a random experiment. When a molecule is selected, we then have a deterministic quantity, which we denoted by a lowercase \( \vec{v} = (v_x, v_y, v_z) \). The probability that the random variable \( \vec{V} \) will take on values in the range \( V_x \in [v_{1x}^1, v_{2x}^2], V_y \in [v_{1y}^1, v_{2y}^2], V_z \in [v_{1z}^1, v_{2z}^2] \) is given by the integral:

\[
\mathbb{P} \left( V_x \in [v_{1x}^1, v_{2x}^2], V_y \in [v_{1y}^1, v_{2y}^2], V_z \in [v_{1z}^1, v_{2z}^2] \right) = \int_{v_{1x}^1}^{v_{2x}^2} dv_x \int_{v_{1y}^1}^{v_{2y}^2} dv_y \int_{v_{1z}^1}^{v_{2z}^2} dv_z f_{\vec{v}}(\vec{v}),
\]

where \( f_{\vec{v}}(\vec{v}) = f_{\vec{V}}(v_x, v_y, v_z) \) if the probability density function for \( \vec{V} \).

---

\(^4\)Let \( a = m/2k_B T \). From integral tables, \( \int_\infty^\infty e^{-ax^2}dx = \sqrt{\pi/a} \), differentiating both sides twice with respect to \( a \) gives:

\[
\frac{d^2}{da^2} \int_\infty^\infty e^{-ax^2}dx = \int_\infty^\infty x^4e^{-ax^2}dx = \sqrt{\pi}(-1/2)(-3/2)a^{-5/2} = (3/4)\sqrt{\pi}a^{-5/2}.
\]

Changing the limits of integration to \( \int_0^\infty \) leads to a factor of 1/2 and we have:

\[
\int_0^\infty x^4e^{-ax^2}dx = (3/8)\sqrt{\pi}a^{-5/2}.
\]

Now we can easily integrate:

\[
\overline{v^2} = \int_0^\infty v^2 f(v)dv = \left( \frac{m}{2\pi k_B T} \right)^{3/2} \frac{3}{4\pi} \int_0^\infty v^4 \exp(-av^2)dv
\]

\[
= \left( \frac{m}{2\pi k_B T} \right)^{3/2} \frac{3\sqrt{\pi}}{4\pi} \left( \frac{m}{2k_B T} \right)^{-5/2} = \frac{3k_B T}{m}. 
\]

Learn more about Gaussian integrals: [https://en.wikipedia.org/wiki/Gaussian_integral](https://en.wikipedia.org/wiki/Gaussian_integral)
The left hand side denotes a joint probability, which is the probability of the simultaneous occurrence of all three events $V_x \in [v^1_x, v^2_x]$, $V_y \in [v^1_y, v^2_y]$, and $V_z \in [v^1_z, v^2_z]$. If we take the intervals to be infinitesimally small, i.e. $V_x \in [v_x, v_x + dv_x]$, $V_y \in [v_y, v_y + dv_y]$, $V_z \in [v_z, v_z + dv_z]$, then the integrals reduce to the following\(^5\)

$$\mathbb{P}(V_x \in [v_x, v_x + dv_x], V_y \in [v_y, v_y + dv_y], V_z \in [v_z, v_z + dv_z]) = f_{\vec{V}}(\vec{v})dv_xdv_ydv_z.$$  

This quantity is the probability of finding a molecule whose $x$ component of velocity, $V_x$, lies in the range $[v_x, v_x + dv_x]$ (centered near $v_x$, with width $dv_x$), and similarly for $V_y$ and $V_z$. This can be written in terms of the volume element $dV$ (the symbol $\times$ denotes Cartesian product):

$$dV = [v_x, v_x + dv_x] \times [v_y, v_y + dv_y] \times [v_z, v_z + dv_z].$$

Thus,

$$f_{\vec{V}}(\vec{v})dv_xdv_ydv_z = \mathbb{P}(\vec{V} \in dV).$$

In 1D the situation is even simpler. The quantity

$$\mathbb{P}(V_x \in [v_1, v_2]) = \int_{v_1}^{v_2} f_{V_x}(v_x)dv_x$$

denotes the probability that the $x$ component of the velocity of a molecule chosen at random will be in the interval $[v_1, v_2]$. The probability density function can be obtained from this probability by taking the interval $[v_1, v_2]$ to be an infinitesimally small interval, e.g. $[v_x, v_x + dv_x]$, in which case we have:

$$\mathbb{P}(V_x \in [v_x, v_x + dv_x]) = \int_{v_x}^{v_x+dv_x} f_{V_x}(v_x)dv_x = f_{V_x}(v_x)dv_x.$$

### 1.3.6. Derivation of Maxwellian Distribution

We will show how the speed distribution is obtained from the velocity distribution. Boltzmann has shown that the velocity distribution is

$$f_{\vec{V}}(\vec{v}) = f_{V_x}(v_x)f_{V_y}(v_y)f_{V_z}(v_z) = \left(\frac{m}{2\pi k_B T}\right)^{3/2} \exp\left(-\frac{mv^2}{2k_B T}\right),$$

He did this by assuming that the kinetic energies $\epsilon = \frac{1}{2}mv^2$ of a gas molecule are distributed according to the Gibbs distribution (Boltzmann factor):

$$p(x) = A e^{-\epsilon(x)/k_B T},$$

where $x$ denotes the state of the system (in our case, $x$ is the velocity $v$), $\epsilon(x)$ is the energy of the state $x$, $A$ is a normalization constant such that $\int p(x)dx = 1$. (We will derive this Gibbs distribution later in the course,

---

\(^5\)In other words, we made use of $\int f(x)dx = f(y)dy$. To prove this, simply write the integrals as Riemannian sums. There is only 1 term in the summation.
using the first law of thermodynamics.) For \(\epsilon(v) = \frac{1}{2}mv^2\), we shall denote this distribution as \(f_{\vec{v}}(\vec{v})\):

**Maxwell velocity distribution:** The Maxwell velocity distribution

\[
f_{\vec{v}}(\vec{v}) = \left(\frac{m}{2\pi k_B T}\right)^{3/2} \exp\left(-\frac{mv^2}{2k_B T}\right),
\]

describes the probability distribution of the velocities of gas molecules in an ideal gas. The quantity \(f_{\vec{v}}(\vec{v})d^3\vec{v} = f_{\vec{v}}(\vec{v})dv_x dv_y dv_z\) is the probability that you will find a molecule whose velocity lies within the volume element \(d^3\vec{v}\) centered at \(\vec{v}\), i.e. where the \(x\) component of velocity is in the interval \([v_x, v_x + dv_x]\), its \(y\) component is in the interval \([v_y, v_y + dv_y]\) and its \(z\) component, in \([v_z, v_z + dv_z]\).

1.3.7. Maxwellian is a Product of Gaussians. In the previous section we wrote that the velocity distribution factors into a product, \(f_{\vec{v}}(\vec{v}) = f_{\vec{v}_x}(v_x)f_{\vec{v}_y}(v_y)f_{\vec{v}_z}(v_z)\). This is due to the independence of the directions \(x\), \(y\) and \(z\). The velocity of a gas molecule is described with a vector \(\vec{v} = (v_x, v_y, v_z)\). In a gas the velocity components \(v_i\) \((i = x, y, z)\) average to zero, \(\overline{v_i} = 0\), so that negative values are equally likely to occur as positive values in an isotropic gas. To say that the gas is “isotropic” means that \(f_{\vec{v}}(\vec{v})\) is only a function of the magnitude \(|\vec{v}|\), or equivalently, \(v^2 = v_x^2 + v_y^2 + v_z^2\). Thus, we can write \(f_{\vec{v}}(\vec{v})\) as \(f_{\vec{v}}(|\vec{v}|)\). I prefer to use \(f_{\vec{v}}(v^2) = f_{\vec{v}}(v_x^2 + v_y^2 + v_z^2)\) instead, since \(v^2\) also depends only on the magnitude of \(\vec{v}\).

Also, the \(x, y, z\) components of the velocity are statistically independent of each other. Thus, \(f_{\vec{v}}(v^2) = f_{\vec{v}}(v_x^2 + v_y^2 + v_z^2)\) can be factored into a product of three independent functions, one for each coordinate:

\[
f_{\vec{v}}(v^2) = f_{\vec{v}}(v_x^2 + v_y^2 + v_z^2) = f_{V_x}(v_x^2)f_{V_y}(v_y^2)f_{V_z}(v_z^2),
\]

where \(f_{V_x}, f_{V_y}\) and \(f_{V_z}\) are the same function because motion along \(x\) follows the same behavior as motion along \(y\) or \(z\).

The only mathematical function which fulfills this condition is the exponential, since \(\exp(-av^2) = \exp(-av_x^2)\exp(-av_y^2)\exp(-av_z^2)\), where \(a > 0\) is a constant. We need \(a > 0\) so that the \(f_{V_i}\) \((i = x, y, z)\) integrate to 1. (If \(a < 0\) the function \(f_{V_i}\) would blow up at infinity and is not integrable.) Thus these probability density functions are all Gaussian. If \(f_{\vec{v}}\) is to be a probability distribution it must be nonnegative \(f_{\vec{v}} \geq 0\) and normalized:

\[
\int f_{\vec{v}}(v^2)d^3\vec{v} = 1.
\]

A normalization constant \(c\) can be determined from the normalization condition

\[
1 = \int_{-\infty}^{\infty} f_{V_i}(v_i)dv_i = c\int_{-\infty}^{\infty}\exp(-av_i^2)dv_i (i = x, y, z),
\]

leading to \(c = \sqrt{a/\pi}\). By assuming that the molecular kinetic energies are distributed according to the Gibbs distribution (Boltzmann factor), Ludwig Boltzmann
found that $a$ is related to the average kinetic energy per particle, $K_{av} = \frac{3}{2}k_BT = \frac{1}{2}mv^2$. Namely, $a = m/2k_BT$.

1.3.8. Maxwell-Boltzmann Speed Distribution. In the previous section, we saw that each $f_{V_i}(v_i)$ ($i = x, y, z$) has the form of a Gaussian distribution centered at 0. However, the speed $v = \sqrt{v_x^2 + v_y^2 + v_z^2}$ does not average to zero at finite temperatures because each component is squared, turning negative values into positive ones.

Let’s now find the probability of observing a molecule with speed in the range $[v, v + dv]$. We must “add” the probability $f_{\vec{v}}(\vec{v})$ over all velocities in a spherical shell of radius $v$ and thickness $dv$, as illustrated in Fig. 1.9.

This shell has volume $4\pi v^2 dv$, and $f_{\vec{v}}(\vec{v})$ has the same “constant” value over this entire shell. The result is obtained trivially by multiplying $f(\vec{v})$ by $4\pi v^2$.

**Maxwell speed distribution:** The Maxwell speed distribution is:

$$f(v) = f_{\vec{v}}(\vec{v})4\pi v^2 = \left(\frac{m}{2\pi k_BT}\right)^{3/2} 4\pi v^2 e^{-\frac{mv^2}{2k_BT}}.$$  

The product $f(v)dv$ gives the probability of finding a molecule with speed in the range $[v, v + dv]$. We note that the units of $f(v)$ are $s/m$ because the integral $\int_0^{\infty} f(v)dv = 1$. Thus, since $dv$ has units $m/s$, $f(v)$ must have the inverse units, $s/m$.

This distribution is no longer Gaussian, but is skewed toward nonnegative values, as seen in Fig. 1.10.
Another way to derive this is to transform to spherical coordinates in velocity space, \((v, \theta, \phi)\). Denoting, \(V = |\vec{V}| = \sqrt{V_x^2 + V_y^2 + V_z^2}\) (\(V\) is also a random variable, since it is constructed from the components \(V_x, V_y\) and \(V_z\), which are themselves random variables), the probability \(\mathbb{P}(V \in [v_1, v_2])\) is

\[
\mathbb{P}(V \in [v_1, v_2]) = \int_{(\vec{v} | v_1 \leq v \leq v_2)} dv_x dv_y dv_z f_{\vec{V}}(\vec{v})
\]

\[
= \int_{v_1}^{v_2} v^2 dv \int_0^{2\pi} \! d\phi \int_0^{\pi} \! \sin \theta d\theta f_{\vec{V}}(\vec{v}) = (4\pi) \int_{v_1}^{v_2} v^2 dv f_{\vec{V}}(\vec{v})
\]

since \(f_{\vec{V}}(\vec{v})\) is independent of \(\theta, \phi\) (by definition, it only depends on the magnitude \(|\vec{v}|\)). In particular, if we choose the interval \([v_1, v_2]\) to be \([v, v+dv]\), then

\[
\mathbb{P}(V \in [v, v+dv]) = 4\pi v^2 f_{\vec{V}}(\vec{v}) dv = \left(\frac{m}{2\pi k_B T}\right)^{3/2} \frac{mv^2 e^{-\frac{mv^2}{2k_BT}}}{4\pi v^2} dv.
\]

1.3.9. How Fast do Molecules Travel? There are three different types of average speeds we may be interested in, depending on the context. The most probable speed is obtained by setting \(df(v)/dv = 0\) and solving for \(v\), i.e. \(df/dv \propto (2v - v^2 m/2k_BT) = 0\). The result is:

\[
v_p = \sqrt{\frac{2k_BT}{m}}.
\]
The mean value of the speed is

\[ \overline{v} = \sqrt{\frac{8k_B T}{\pi m}}. \]  

The root-mean-square speed is (derived in Section 1.3.4):

\[ v_{\text{rms}} = \sqrt{\frac{3k_B T}{m}}. \]

We see that: \( v_p < \overline{v} < v_{\text{rms}} \). These formulae can be expressed in terms of molar mass \( \mathcal{M} \) instead by simply replacing \( k_B/m \) by \( R/\mathcal{M} \). These quantities are illustrated in Fig. 1.11 relative to the probability distribution (Eq. 1.5).

Geometrically, the most probable speed corresponds to the peak of this distribution whereas the mean speed is the velocity at which the areas to its left and to its right are equal. And you will recall from the previous lecture that the RMS speed is related to temperature and average kinetic energy of a molecule.

1.3.9.1. Sample Calculation. Calculate the most probable speed, \( v_p \), for \( \text{H}_2 \) gas at room temperature \( (T = 300 \text{ K}) \).

**Strategy:** Invoke the formula in molar units, \( v_p = (2RT/\mathcal{M})^{1/2} \) and plug in \( \mathcal{M}=2 \text{ g/mol} \).

**Solution:**

\[ v_p = \sqrt{\frac{2(8.31 \text{ J/mol.K})(300 \text{ K})}{(2 \text{ g/mol})} \frac{1000 \text{ g}}{1 \text{ kg}}} \sim 1580 \text{ m/s}. \]
1.3.10. Dependence of Speed Distribution on $T$ and $m$. For a fixed temperature $T$ the Maxwell speed distribution (Eq. 1.4) for heavy gases (e.g. $^{132}\text{Xe}$) is shifted to the left because of the low velocities, whereas the distribution for light gases (e.g. $^4\text{He}$) is shifted to the right because of higher velocities. This is illustrated in Fig. 1.12.

The area under the curve, however, remains constant, which explains the flattening of the curves with light gases. If we fix the gas type (e.g. $\text{N}_2$) the distribution is skewed to the left for cold gases and to the right for warmer gases. This is illustrated in Fig. 1.13.
1.4. Real Gases

In this section we look at the case where the ideal gas law is inadequate.

1.4.1. Deviations From the Ideal Gas Law. If we compress a gas, its density increases and deviations from the ideal gas law are expected (Fig. 1.14).

![Compression of a gas](image)

Figure 1.14. Compression of a gas leads to higher density, where intermolecular interactions become more important.

The ideal gas law becomes less accurate when intermolecular interactions are important. Such a situation occurs at high pressures and low temperatures. At high pressures, the molecules are closer to each other and interact more. At low temperatures their velocities are reduced and intermolecular collisions last longer.

Consider two different gases, He and HF. In both cases, we have 0.1 mol of each gas, and each gas is placed in a cubic container of volume 2 L. The temperature of both containers is held at 500 K. Experimentally, we measure the pressures to be $P_{\text{He}} = 2.05\ \text{atm}$ and $P_{\text{HF}} = 2.03\ \text{atm}$. The ideal gas law predicts that the pressure will be (recall that 1 atm = 101.325 Pa)

$$P_{\text{ideal}} = \frac{nRT}{V} = \frac{(0.1 \text{ mol})(8.31 \text{ J/mol.K})(500 \text{ K})}{(2 \text{ L})(0.001 \text{ m}^3/\text{L})} = 207.750 \text{ kPa} = 2.05\ \text{atm},$$

or\(^6\)

$$P_{\text{ideal}} = \frac{nRT}{V} = \frac{(0.1 \text{ mol})(0.08206 \text{ L.atm/K.mol})(500 \text{ K})}{2 \text{ L}} = 2.05\ \text{atm}.$$

Comparing this to the experimentally measured pressure values we conclude that He gas behaves like an ideal gas under those conditions whereas HF deviates from an ideal gas. This is due to the nature of the intermolecular

\(^6\)You can find values of the gas constant $R$ expressed in many different convenient units at the URL https://en.wikipedia.org/wiki/Gas_constant
interactions. In He, such interactions are small, limited to VDW forces, where as HF is a polar molecule and dipole-dipole interactions are significant. This experiment can be repeated for numerous additional gases and we obtain the results:

<table>
<thead>
<tr>
<th>Molecule</th>
<th>$P_{\text{real}}$ (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO$_2$</td>
<td>2.06</td>
</tr>
<tr>
<td>Ar</td>
<td>2.05</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>2.02</td>
</tr>
<tr>
<td>SF$_6$</td>
<td>2.08</td>
</tr>
<tr>
<td>CH$_2$Cl$_2$</td>
<td>2.04</td>
</tr>
</tbody>
</table>

Repulsive forces increase the pressure compared to the ideal gas law. Repulsive forces can arise when the molecules are ions. They also arise because of the finite volume of the molecules. Indeed, molecules take up space, leading to a “reduced volume” in which gas molecules can diffuse. These molecules also hit the walls harder due to their finite-size, compared to ideal gas molecules which have size zero. Thus, $P$ is larger.

Attractive forces exist when molecules have opposite charges, or dipole moments (permanent or induced), or VDW interactions. This has the effect of reducing the pressure of the gas.

1.4.2. **Van Der Waals Gas.** A simple model for a non-ideal gas is the VDW gas. It accounts for the volume occupied by the particles and for the attractive forces between pairs of particles. Although it performs slightly better than the ideal gas law, its range of applicability is limited, making it mainly a pedagogical tool.

The starting point is the ideal gas law, $PV = nRT$, where $P$ is the “ideal gas pressure” and $V$ is the “ideal gas volume”. The volume is set equal to the container volume $\tilde{V}$ less the volume occupied by the gas molecules, $nb$:

$$V = \tilde{V} - nb,$$

where $b$ is the particle volume occupied by one mole of gas. The ideal gas pressure, $P$, is equal to the measured pressure $\tilde{P}$ plus a correction term, $an^2/V^2$:

$$P = \tilde{P} + a \frac{n^2}{V^2}$$

where $a > 0$ for attractive intermolecular forces and $a < 0$ for repulsive forces. The result is (dropping the tilde notation):
Van Der Waals equation of state: The VDW equation of state is
\[
\left( P + \frac{n^2 a}{V^2} \right)(V - nb) = nRT
\]
where \( a > 0 \) for attractive intermolecular interactions and \( a < 0 \) for repulsive forces. \( b \) is always positive because it represents volume. Rearranging this equation
\[
P_{vdw} = \frac{nRT}{(V - nb)} - \frac{n^2 a}{V^2}
\]
makes it obvious that the effect of the term \( \frac{n^2 a}{V^2} \) is to reduce the pressure in the system (when forces are attractive, \( a > 0 \)). In general, both \( a \) and \( b \) depend on temperature.

The excluded volume for one mole of particles is:
\[
b = N_A \cdot 4 \times \left( \frac{4\pi r^3}{3} \right),
\]
where \( r \) is the particle radius and \( N_A \) is Avogadro’s number. The factor of 4 arises because the radius of the sphere is taken to be 2\( r \) (instead of \( r \)), which is the distance of closest approach between two spheres’ centers (Fig. 1.15).

![Figure 1.15. Molecules as hard spheres. The distance of closest approach is \( d \) (as measured from center to center).](image)

This gives a factor of \( 2^3 = 8 \), but we then divide by 2 to avoid over-counting (adding up excluded volumes for each sphere would over-count this total excluded volume by a factor of two).

The quantity \( (n/V)^2 \) is proportional to \( \rho^2 \), the square of the gas density. This reflects the fact that pressure drops occur because of pairwise encounters between molecules.

The VDW equation of state describes a first order phase transition from liquid to vapor. Several isotherms are shown in Fig. 1.16.

\( T_c \) is the critical temperature. At \( T_c \), the local maximum and minimum coalesce into a single inflection point where \( (\partial^2 P/\partial V^2)_T = 0 \), in addition...
**Figure 1.16.** VDW equation of state illustrated in the form of isotherms in the $P,V$ plane. The critical temperature ($T_c$) is indicated on this graph; see Section 1.4.3. The Maxwell construction is shown in dark blue.

To $(\partial P/\partial V)_T = 0$. The critical point is the maximum temperature and pressure at which distinct liquid and gas phases exist. In Section 1.4.3 we show how to express the critical point in terms of the substance $(a,b)$. The quantity

$$\kappa = -\frac{1}{V} \frac{\partial V}{\partial P} \bigg|_T \approx -\frac{1}{V} \frac{\Delta V}{\Delta P}$$

at constant $T$ is called *isothermal compressibility*. For a normal substance, pressure should drop with increasing volume (think Boyle’s law, $P \propto 1/V$), leading to $\kappa > 0$. However, for the VDW gas we see that $\kappa < 0$ in the oscillatory region of Fig. 1.16. Such behavior is not physical. The Maxwell construction was invented to avoid non-physical behavior where $P$ increases with $V$. This construction is shown in the blue isotherm of Fig. 1.16: the oscillatory region of the VDW isotherm is deleted and replaced by a horizontal straight line. The vertical position of the horizontal straight line is such that the two blue areas are equal. The horizontal line indicates the phase transition from liquid to gas and is characterized by infinite isothermal compressibility, $\kappa = -(1/V)(\partial V/\partial P) \to \infty$. Experimentally measured values of the isothermal compressibility for different substances are shown in Table 1.1.

The VDW constants $a$ and $b$ depend on the gas type. $a$ is a measure of the attractive forces between the molecules. $b$ is due to the finite volume of the molecules and to their general incompressibility. Table 1.2 provides examples for several gases.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Compressibility (Pa$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$T/\degree C$</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>20</td>
</tr>
</tbody>
</table>
1.4. Real Gases

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>a</th>
<th>b</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetone</td>
<td>CH₃COOH</td>
<td>17.59</td>
<td>0.1068</td>
</tr>
<tr>
<td>Acetone</td>
<td>(CH₃)₂CO</td>
<td>13.91</td>
<td>0.0994</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
<td>4.170</td>
<td>0.03707</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>1.345</td>
<td>0.03219</td>
</tr>
<tr>
<td>Benzene</td>
<td>C₆H₆</td>
<td>18.00</td>
<td>0.1154</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>CO₂</td>
<td>3.592</td>
<td>0.04267</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>CO</td>
<td>1.485</td>
<td>0.03985</td>
</tr>
<tr>
<td>Chlorobenzene</td>
<td>C₆H₅Cl</td>
<td>25.43</td>
<td>0.1453</td>
</tr>
<tr>
<td>Cyclohexane</td>
<td>C₆H₁₂</td>
<td>22.81</td>
<td>0.1424</td>
</tr>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td>5.489</td>
<td>0.06380</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C₂H₅OH</td>
<td>12.02</td>
<td>0.08407</td>
</tr>
<tr>
<td>Helium</td>
<td>He</td>
<td>0.03412</td>
<td>0.02370</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>0.2444</td>
<td>0.02661</td>
</tr>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td>2.253</td>
<td>0.04278</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>1.390</td>
<td>0.03913</td>
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<tr>
<td>Nitrous oxide</td>
<td>N₂O</td>
<td>3.782</td>
<td>0.04415</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>1.360</td>
<td>0.03183</td>
</tr>
<tr>
<td>Phosphine</td>
<td>PH₃</td>
<td>4.631</td>
<td>0.05156</td>
</tr>
</tbody>
</table>

Table 1.1. Isothermal compressibility, $\kappa$, of selected compounds. $\kappa$ depends on temperature and pressure. Source: CRC Handbook of chemistry and physics [12].
Note: to use the values from Table 1.2, the pressure must be in atmospheres, temperature should be in Kelvin and volume must be in liters. To use the VDW formula with these constants, $R$ must be 0.08206 L.atm/mol/K. The value of $R$ that is appropriate depends on the units you are working with. Here are some commonly used values of $R$ you may encounter in the literature:

- 0.082057 L atm mol$^{-1}$ K$^{-1}$
- 62.364 L Torr mol$^{-1}$ K$^{-1}$
- 8.3145 m$^3$ Pa mol$^{-1}$ K$^{-1}$
- 8.3145 J mol$^{-1}$ K$^{-1}$

These values of $R$ are all equal to each other. They are obtained by converting units. For example, 8.3145 J mol$^{-1}$ K$^{-1}$ is equal to 8.3145 m$^3$ Pa mol$^{-1}$ K$^{-1}$ because 1 J = 1 m$^3$ Pa. Likewise, if we multiply 8.3145 m$^3$ Pa mol$^{-1}$ K$^{-1}$ by (1 atm/101,325 Pa) and by (1 L/0.001 m$^3$), we obtain 0.082057 L atm mol$^{-1}$ K$^{-1}$.

1.4.3. Critical Point. Differentiating the VDW pressure

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

with respect to $V$, we get

$$-nRT(V - nb)^{-2} + 2n^2aV^{-3}.$$  

Differentiating again gives

$$2nRT(V - nb)^{-3} - 6n^2aV^{-4}.$$  

Setting these two expressions equal to zero and solving for $T$ and $V$ gives the critical temperature and volume, respectively. Multiply

$$-nRT(V - nb)^{-2} + 2n^2aV^{-3} = 0$$

by $2/(V - nb)$ to get

$$-2nRT(V - nb)^{-3} + 4n^2aV^{-3}(V - nb)^{-1} = 0.$$  

Then add the second equation

$$2nRT(V - nb)^{-3} - 6n^2aV^{-4} = 0$$

to get (cancelling terms with $T$),

$$-6an^2V^{-4} + 4an^2V^{-3}(V - nb)^{-1}.$$  

Then multiply by $V^3/(2an^2)$ to get, $2(V - nb)^{-1} = 3V^{-1}$, which is solved to get $V_c = 3nb$.

To get $T_c$ we substitute $V_c$ into any of the two equations (say the first one), which gives

$$-nRT(3nb - nb)^{-2} + 2an^2(3nb)^{-3} = 0,$$
1.4. Real Gases

or \( RT/4 = 2a/(27b) \), from which \( T_c = 8a/(27bR) \).

To get \( P_c \) we substitute \( V_c \) and \( T_c \) into the VDW equation and solve for \( P_c \),

\[
P_c = \frac{nRT_c}{V_c - nb} - aV_c^2 = \frac{nR(84/(27bR))}{3nb - nb - an^2/(3nb)^2},
\]

which simplifies to \( P_c = a/(27b^2) \). In Section 1.4.4, we show how the critical point can be used to express the VDW equation of state in dimensionless form.

1.4.4. Dimensionless VDW Eq. of State. In the previous section we have found the critical point:

\[
P_c = a/(27b^2), \quad T_c = 8a/(27bR), \quad V_c = 3nb.
\]

We can define the reduced parameters:

\[
P' = \frac{P}{P_c}, \quad T' = \frac{T}{T_c}, \quad V' = \frac{V}{V_c}.
\]

Solve for \( P, T, \) and \( V \) and substitute into

\[
\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT,
\]

to get:

\[
\left( P' \frac{a}{27b^2} + \frac{an^2}{(V'3nb)^2} \right) (V'3nb - nb) = nR \frac{8aT'}{27bR}.
\]

This simplifies to:

\[
\left( P' + \frac{3}{V'^3} \right) \left( V' - \frac{1}{3} \right) = \frac{8}{3} T'.
\]

This dimensionless form is useful for graphing the VDW equation. As a function of \( V', T', P' \), it is independent of the type of gas (the dependence on gas type has been absorbed in the definitions of \( V', T', P' \)).

1.4.5. P-V-T Diagrams for VDW Gas. For fixed number of moles, the equation of state is a function of three variables \( (P, V, T) \). In 3D space, this relation between these variables represents a constraint, i.e. a 2D surface. The equation of state plotted as a 2D surface is called P-V-T diagram. Figure 1.17 compares surface plots of the equation of state for ideal vs VDW gas.

In Fig. 1.17 the “isothermals” (\( T = constant \)) are shown in the horizontal plane for the ideal gas. For the VDW gas, “isobars” (\( P = constant \)) are shown in the horizontal plane. For the ideal gas law, the isothermals are \( P \propto 1/V \) whereas for the VDW gas the relationship between \( T \) and \( V \) is more complicated.
1.4.6. Virial Expansion. The virial expansion of the pressure of a many-particle system in powers of the density provides systematic corrections to the ideal gas law. Denoting the gas density as \( \rho = N/V \) the ideal gas law reads:

\[
P V = N k_B T \quad \rightarrow \quad \frac{P}{k_B T} = \frac{N}{V} \equiv \rho.
\]  

(ideal gas law)

We know that the ideal gas law holds at low pressures, or equivalently (for fixed \( T \)), at low densities. As the gas density is increased, we expect some deviations from the ideal gas law behavior. These deviations can be described by replacing the right hand side by a Taylor series in the gas density.

\[
\text{Virial expansion:} \quad \frac{P}{k_B T} = \rho + B_2(T) \rho^2 + B_3(T) \rho^3 + \ldots
\]

The ideal gas law is then viewed as the low density limit (\( \rho \to 0 \)) of the virial expansion. The virial coefficients \( B_i(T) \) depend on temperature and describe the interactions between the particles. The calculation of virial coefficients from knowledge of the intermolecular potentials is done using methods of statistical mechanics.

The first correction term, \( B_2(T) \rho^2 \), describes pairwise collisions between molecules. It is proportional to \( \rho^2 \) because two-body collisions depend on the product of the probability of finding molecule 1 in some volume element times the probability of finding molecule 2 in the same volume element. For a uniform gas, each probability is proportional to the gas density; their product is \( \rho^2 \). The second correction term, \( B_3(T) \rho^3 \), describes three-body collisions. The probability of a three-body encounter is proportional to \( \rho^3 \).
1.5. Diffusion

The latter are extremely rare events and this term is typically much smaller than the previous term, $B_2(T)\rho^2$.

1.4.7. VDW eq. of State vs Virial Expansion, Boyle Temperature.

Comparison of the virial expansion:

$$\frac{P}{k_B T} = \rho + B_2(T)\rho^2 + B_3(T)\rho^3 + \ldots$$

where $\rho = N/V$, with the VDW equation of state:

$$\left( P + \frac{n^2a}{V^2} \right)(V - nb) = nRT.$$  

This leads to the virial coefficients of the VDW gas:

$$B_2(T) = \frac{b}{N_A} - \frac{a}{N_A^2 k_B T}, \quad B_3(T) = \frac{b^2}{N_A^3}, \quad \text{etc.}$$

This is left to the reader as an exercise. The point at which $B_2(T) = 0$ is called the Boyle temperature:

$$T_B = \frac{a}{bk_B N_A} = \frac{a}{bR}.$$  

It is at the Boyle temperature that the attractive and repulsive forces acting on the gas particles balance out. At that temperature, the non-ideal gas behaves like an ideal gas over some range of pressures.

1.5. Diffusion

Molecular diffusion, often simply called diffusion, is the thermal motion of all (liquid or gas) particles at temperatures above absolute zero. The rate of this movement is a function of temperature, viscosity of the fluid and the size (mass) of the particles. Diffusion explains the net flux of molecules from a region of higher concentration to one of lower concentration. Once the concentrations are equal the molecules continue to move, but since there is no concentration gradient the process of molecular diffusion has ceased and is instead governed by the process of self-diffusion, originating from the random motion of the molecules. The result of diffusion is a gradual mixing of material such that the distribution of molecules is uniform. Since the molecules are still in motion, but an equilibrium has been established, the end result of molecular diffusion is called a “dynamic equilibrium”. In a phase with uniform temperature, absent external net forces acting on the particles, the diffusion process will eventually result in complete mixing.

---

7Hint: 1) expand the VDW eq. into powers of $\rho$; 2) identify the coefficients of like powers of $\rho$. 


1.5.1. Einstein’s Theory of Brownian Motion (1905). A translation of Einstein’s paper from German to English can be found here:

http://www.relatividad.org/einstein_brownian05.pdf

Suppose that we have a probability density function\(^8\) \(\phi(\Delta)\) which describes the probability that a molecule will undergo a displacement \(\Delta\) during a short time interval \(\tau\). The medium contains \(n\) particles and during this time interval the number of particles, \(dn\), which experience a shift between \(\Delta\) and \(\Delta + d\Delta\) is:

\[
\frac{dn}{n} = \phi(\Delta)d\Delta = \left\{ \begin{array}{l}
\text{probability that a particle at } x \text{ at time } t \text{ will be in the range } [x + \Delta, x + \Delta + d\Delta] \text{ at time } t + \tau
\end{array} \right\}
\]

We assume that \(\phi\) is normalized, \(\int \phi(\Delta)d\Delta = 1\), and symmetric (probability of a displacement to the left equals the probability of a displacement to the right):

\(\phi(\Delta) = \phi(-\Delta)\).

Let \(c(x, t)\) be the number of particles per unit volume at position \(x\) and time \(t\). When going from time \(t\) to time \(t + \tau\) the particle density at \(x\) is a result of particles which undergo a displacement \(\Delta\) from another location:

\[
c(x, t + \tau) = \int_{-\infty}^{\infty} c(x + \Delta, t)\phi(\Delta)d\Delta. \quad (*)
\]

Next we assume that \(\phi\) extends over short distances only and that beyond \(\Delta\) its value is small. Thus, we can Taylor expand \(c\) in both space and time:

\[
c(x, t + \tau) = c(x, t) + \tau \frac{\partial c}{\partial t} + \ldots
\]

\[
c(x + \Delta, t) = c(x, t) + \Delta \frac{\partial c(x, t)}{\partial x} \frac{\Delta^2}{2!} \frac{\partial^2 c(x, t)}{\partial x^2} + \ldots
\]

Substituting into (\(\ast\)) we get (all limits of integration are over all space):

\[
c + \tau \frac{\partial c}{\partial t} = c \int_{\mathbb{R}} \phi(\Delta)d\Delta + \frac{\partial c}{\partial x} \int_{\mathbb{R}} \Delta \phi(\Delta)d\Delta + \frac{\partial^2 c}{\partial x^2} \int_{\mathbb{R}} \frac{\Delta^2}{2!} \phi(\Delta)d\Delta.
\]

Because \(\phi(\Delta) = \phi(-\Delta)\) (even function), the second, fourth, etc. terms on the right hand side vanish. This is because those terms feature an odd function (\(\Delta, \Delta^3, \Delta^5\), etc.) times an even function (\(\phi(\Delta)\)), which is odd.

---

\(^8\)A probability density function is one which is nonnegative (\(\phi \geq 0\)) and normalized \(\int_{-\infty}^{\infty} \phi(\Delta)d\Delta = 1\).
The integral of an odd function over all space is zero: \[ \int_{\mathbb{R}} odd \times even = \int_{\mathbb{R}} odd = 0. \] This is illustrated in Fig. 1.18 for the special case of \[ \phi(\Delta) = \left(\frac{1}{\sqrt{2\pi}}\right) \exp\left(-\frac{\Delta^2}{2}\right). \]

Using \[ \int \phi(\Delta) d\Delta = 1 \] for the first term and setting \[ \frac{1}{\tau} \int_{\mathbb{R}} \frac{\Delta^2}{2} \phi(\Delta) d\Delta \equiv D, \quad \text{ (diffusion coefficient)} \] we get that \( c(x, t) \) must satisfy:

\[ (1.10) \quad \frac{\partial c}{\partial t} = D \frac{\partial^2 c}{\partial x^2}. \]

\( c(x, t) \) is a solution to this equation for given initial and boundary conditions. For initial condition \( c(x, 0) = n\delta(x) \) (all particles are at the origin at time \( t = 0 \)) and no physical boundaries (the case of unrestricted diffusion), the solution to the 1D diffusion equation is

\[ c(x, t) = \frac{n}{\sqrt{4\pi Dt}} e^{-x^2/4Dt} \]

where \( \sqrt{x^2} = \sqrt{2Dt} \) is the root-mean-square displacement. You can check that this is indeed a solution by direct substitution into equation 1.10.

---

9If you don’t like \( \Delta \), use \( x \):

\[ \int_{\mathbb{R}} x\phi(x) dx = 0, \quad \int_{\mathbb{R}} x^3\phi(x) dx = 0, \quad \int_{\mathbb{R}} x^5\phi(x) dx = 0, \quad \text{ etc.} \]
Calculation of the value of $D$ requires knowledge of $\phi$ which the above analysis does not provide. For a gas, we can obtain an estimate for $D$ using simple arguments about molecular collisions.

The solution $c(x, t)$ is a function of two variables, $x$ and $t$. For fixed $t$, this can be thought as the concentration profile as function of space ($x$). The time dependence means that the concentration profile evolves over time. For the case where all the particles are found at $x = 0$ at time $t = 0$, the time-evolution simply describes the diffusional spreading of particles over space until the concentration is uniform. If there are no boundaries (i.e. space extends to infinity), the limit $t \to \infty$ results in a spatially uniform concentration (of zero).

One way to represent the solution $c(x, t)$ is a two-dimensional surface in 3D space (see Fig. 1.19). For fixed $t$, the function of $x$ always describes a Gaussian (bell-shaped) profile. At time $t = 0$, it is sharply peaked near the origin ($x = 0$); however, its width increases over time ($t > 0$).

**Figure 1.19.** Concentration of particles as function of time ($t$) and space ($x$), then all particles are at the origin ($x = 0$) at time $t = 0$.

### 1.5.2. Diffusion Coefficient: Estimating the Value of $D$.

The Einstein model for diffusion does include a prescription to calculate the diffusion coefficient,

$$ D = \frac{\overline{\Delta^2}}{2\tau}, $$

where $\Delta$ is the particle displacement along some direction and $\overline{\Delta^2}$ is the mean square displacement. To calculate it, we require knowledge of the PDF $\phi(\Delta)$. But $\phi(\Delta)$ may not necessarily be known to us.
Kinetic theory can be used to estimate $D$: First we note that from the formula, $D = \frac{\Delta x^2}{2\tau}$, $D$ has units of m$^2$/s. The collision rate $Z_1$ among molecules in a gas should be proportional to gas density $N/V$, molecular cross-sectional area $\pi d^2$ (where $d = 2r$ instead of $r$ as the hard spheres cannot overlap) and mean speed $\bar{v}$:

$$Z_1 = \sqrt{2} \frac{N}{V} \pi d^2 \bar{v}.$$  

The factor $\sqrt{2}$ comes from a more detailed calculation. We can express $\bar{v}$ in terms of the temperature using

$$\bar{v} = \sqrt{\frac{8k_B T}{\pi m}}$$

to get:

$$Z_1 = 4 \frac{N}{V} d^2 \sqrt{\frac{\pi k_B T}{m}}.$$  

Then the mean free path $\lambda$ (average distance travelled between collisions)

$$\lambda = \frac{\bar{v}}{Z_1^{-1}} = \frac{\bar{v}}{\sqrt{2}(N/V)\pi d^2 \bar{v}} = \frac{1}{\sqrt{2} \pi d^2 N/V}.$$  

We then have for the diffusion coefficient

\[ D = \frac{3\pi}{16} \lambda \bar{v} = \frac{3\pi}{16} \sqrt{\frac{8k_B T}{\pi m}} \cdot \sqrt{\frac{8k_B T}{\pi m}} = \frac{3}{8} \sqrt{\frac{k_B T}{\pi m}} \cdot \frac{1}{d^2 N/V}. \]

The numerical constants shown arise from a more detailed theoretical analysis. The main trends are that if $T$ increases the diffusion is faster whereas if $m$, $d$ or $N/V$ increase diffusion is slower. We note that $D \sim \sqrt{T}$ here, whereas $D \sim T$ for the Stokes-Einstein model:

\[ D = 2\gamma k_B T/M, \]

where $\gamma$ is a friction coefficient and $M$ is the mass of the Brownian particle. The former is for a gas, whereas the latter is for a Brownian particle in a liquid. Example calculations:

1.5.2.1. Example 1. Calculate the collision frequency for a molecule of nitrogen gas at 1 atm and 300 K (room temperature). Take the diameter of $N_2$ to be 370 picometers.

\[ \text{Solution: The density } \frac{N}{V} = \frac{N_A P}{RT} \text{ (ideal gas law) is} \]

\[ \frac{(6.022 \times 10^{23} \text{ mol}^{-1})(1 \text{ atm})}{(0.08206 \text{ atm mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 2.46 \times 10^{22} \text{L}^{-1} = 2.45 \times 10^{25} \text{m}^{-3}. \]
To calculate the collision frequency, we need the mean speed:

\[
\bar{v} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8(8.3145 \text{ J.mol}^{-1}\text{K}^{-1})(300 \text{ K})}{\pi (28.00 \times 10^{-3} \text{ kg/mol})}} = 445 \text{ m.s}^{-1},
\]

where \(M\) is the molar mass. Then,

\[
Z_1 = \sqrt{2\pi(2.45 \times 10^{25} \text{ m}^{-3})(3.70 \times 10^{-10} \text{ m})^2(445 \text{ m.s}^{-1})} = 6.63 \times 10^9 \text{ s}^{-1}.
\]

1.5.2.2. Example 2. Calculate the mean free path and the diffusion constant for the nitrogen molecules of the previous example.

**Solution:** The mean free path is

\[
\lambda = \frac{1}{\sqrt{2\pi(3.70 \times 10^{-10} \text{ m})^2(2.45 \times 10^{25} \text{ m}^{-3})}} = 6.68 \times 10^{-8} \text{ m}.
\]

The diffusion constant is

\[
D = \frac{3\pi}{16} \bar{v} = \frac{3\pi}{16}(6.68 \times 10^{-8} \text{ m})(445 \text{ m.s}^{-1}) = 1.75 \times 10^{-5} \text{ m}^2\text{s}^{-1}.
\]

1.5.3. Mean-Square Displacement. If we denote the displacement along \(x\) by \(\Delta_x\) (and similarly for \(y\) and \(z\)), we have seen that the root-mean-square (RMS) displacement along \(x\) is related to the diffusion coefficient \(D\) by:

\[
D = \frac{\Delta_x^2}{2\tau} \quad \text{or} \quad \Delta_x^2 = 2D\tau.
\]

This relation is valid for displacement in one dimension. In three dimensions, the length of the displacement is \(\Delta_r = \sqrt{\Delta_x^2 + \Delta_y^2 + \Delta_z^2}\), and by linearity of the averaging operator\(^{10}\) we have

\[
\Delta_r^2 = \Delta_x^2 + \Delta_y^2 + \Delta_z^2 = \Delta_x^2 + \Delta_y^2 + \Delta_z^2 = 6D\tau.
\]

The last equality follows from the assumption that the mean-square displacement along \(x\) is the same (on average) as the displacements along \(y\) and \(z\), i.e. \(\Delta_x^2 = \Delta_y^2 = \Delta_z^2 = 2D\tau\). This is the assumption of isotropic diffusion, which is correct in the absence of external fields or physical barriers to diffusion. In \(d\)-dimensions, the mean-square displacement is \(2dD\tau\).

In Section 1.5.1, we have seen that in the case where all diffusing particles are located at the origin \(x = 0\) at time \(t = 0\), their distribution (concentration as function of \(x\)) at some later time \(t > 0\) will be Gaussian centered at \(x = 0\) and width proportional to \(\sqrt{2D\tau}\). This distribution is a statistical concept: the Gaussian function describes the behavior of a very large number of particles.

---

\(^{10}\)The average of a function, \(f(X)\) of a random variable \(X\), is the integral \(f(X) = \int f(x)p(x)dx\), where \(p(x)\) is the probability density of \(X\) and the integral is over the range of allowed values of \(X\). Thus, the averaging operation is a linear operation because the integral is a linear operator, i.e. \(af(X) + bg(X) = af(X) + bg(X)\), where \(a\) and \(b\) are constants and \(f, g\) are functions.
The RMS displacement, which is proportional to \( \sqrt{t} \), should be contrasted with ballistic motion, for which a projectile thrown at velocity \( v \) undergoes a displacement \( vt \), i.e. displacement depends on \( t \) rather than \( \sqrt{t} \). The \( \sqrt{t} \) dependence reflects the fact that random motion leads to much back-and-forth, in such a way that the RMS displacement increases over time but the growth is slower. An example path for a single particle\(^{11} \) is shown in Fig. 1.20.

**Figure 1.20.** A particle undergoes random motion. A sample path is shown here. The displacement, \( \Delta \vec{r} \), is the vector that connects the start and end points.

### 1.5.4. Diffusion vs Self-Diffusion.

The theory of diffusion we have been discussing is called self-diffusion, i.e. the diffusion of gas molecules within the same gas type, i.e. oxygen molecules in a gas of oxygen. Diffusion is a more general concept: it generally refers to the diffusion of species 1 within species 2. For example, the diffusion of hydrogen atoms through copper, or the diffusion of oxygen impurity molecules through nitrogen gas.

### 1.6. Problems

**Problem 1.** Consider a monatomic gas with atoms whose diameter is \( d = 3.16 \times 10^{-10} \) m. At what pressure does the mean free path of the gas atoms become comparable with the diameter of an atom if \( T = 300 \) K? Calculate the diffusion constant at this pressure and comment on its magnitude. Assume that the gas obeys the ideal gas law even at these high pressures.

**Solution.** The diameter of Kr atoms is \( d = 3.16 \times 10^{-10} \) m. When the mean free path \( \lambda = \frac{1}{\sqrt{2\pi d^2 N/V}} \) is comparable to the diameter of atom, \( N/V \) should be around \( 7.12 \times 10^{27} \)/m\(^3\). From the ideal gas law, \( N/V = \frac{P}{k_B T} \), \( P = 2.95 \times 10^7 \) Pa. The diffusion constant is

\[
D = \frac{3}{8} \sqrt{\frac{RT}{\pi M d^2 N/V}} = \frac{3}{8} \sqrt{\frac{8.314 \times 300}{\pi \times (83.789 \cdot 10^{-3}) \cdot (3.16 \cdot 10^{-10})^2 (7.12 \cdot 10^{27})} \]

\(^{11}\)This picture shows a single trajectory (of a single molecule). The concept of RMS displacement is statistical, i.e. it is obtained from the random diffusion of many particles.
which evaluates to $D = 5.1342 \times 10^{-8} \text{ m}^2/\text{s}$. This value is comparable to diffusion in solids. Diffusion coefficients in solids are in the range $10^{-8} - 10^{-30} \text{ m}^2/\text{s}$.

**Problem 2.** Consider a VDW gas at the critical point. Obtain expressions for the critical pressure, molar volume, and temperature in terms of constants $a$ and $b$. The critical point is defined as the point on the $P - V_m$ curve where the first and second derivatives of the pressure with respect to molar volume vanish, i.e. \( \frac{\partial P}{\partial V_m} \bigg|_c = 0 \) and \( \frac{\partial^2 P}{\partial V_m^2} \bigg|_c = 0 \).

**Solution.** This problem was solved in Section 1.4.3. Based on the VDW equation

\[
P = \frac{RT}{V_m - b} - \frac{a}{V_m^2}.
\]

The first derivative of pressure equals zero

\[
\frac{\partial P}{\partial V_m} \bigg|_c = -\frac{RT_c}{(V_m - b)^2} + 2a \frac{1}{V_m^3} \bigg|_c = 0.
\]

The second derivative of pressure equals zero

\[
\frac{\partial^2 P}{\partial V_m^2} \bigg|_c = \frac{2RT_c}{(V_m - b)^3} - 6a \frac{1}{V_m^4} \bigg|_c = 0.
\]

Solving the above equations then substituting $V_c$ and $T_c$ into the VDW equation to solve for $p_c$ gives (see Section 1.4.3 for details):

\[
V_c = 3b, \quad T_c = \frac{8a}{27bR}, \quad P_c = \frac{a}{27b^2}.
\]

**Problem 3.** A 20 liter container is filled with 1 mol of Kr atoms at a temperature of 400 K. (a) Use the VDW equation of state to compute the pressure of the system. (b) Find the pressure once more but using the ideal gas law. (c) Calculate the percent error between the results. (d) Now calculate the VDW and ideal-gas pressures if the volume of the container was 4 liters. (e) Find the percent error between the results. (f) Compare the percent errors you have obtained and explain any differences. For Kr the VDW parameter are $a = 2.325 \text{ L}^2 \text{ bar mol}^{-2}$ and $b = 0.0396 \text{ L mol}^{-1}$.

**Solution.** For (a), we use $P = \frac{RT}{V_n - b} - \frac{a}{V_n^2}$. This gives:

\[
P = \frac{(0.082057 \text{ L atm K}^{-1} \text{ mol}^{-1})(400 \text{ K})}{20 \text{ L mol}^{-1} - 0.0396 \text{ L mol}^{-1}} - \frac{2.295 \text{ L}^2 \text{ atm mol}^{-2}}{(20 \text{ L mol}^{-1})^2} = 1.639 \text{ atm}.
\]
1.6. Problems

For (b) we use \( P = \frac{RT}{V_n} \). This gives:
\[
P = \frac{(0.082057 \text{ L atm K}^{-1} \text{ mol}^{-1})(400 \text{ K})}{20 \text{ L mol}^{-1}} = 1.641 \text{ atm}
\]
For (c),
\[
\% \text{ Error} = \left| \frac{1.639 - 1.641}{1.641} \right| \times 100 = 0.122\%
\]
For (d),
\[
P = \frac{(0.082057 \text{ L atm K}^{-1} \text{ mol}^{-1})(400 \text{ K})}{4 \text{ L mol}^{-1} - 0.0396 \text{ L mol}^{-1} - \frac{2.295 \text{ L}^2 \text{ atm mol}^{-2}}{(4 \text{ L mol}^{-1})^2}} = 8.144 \text{ atm}
\]
For (e),
\[
P = \frac{(0.082057 \text{ L atm K}^{-1} \text{ mol}^{-1})(400 \text{ K})}{20 \text{ L mol}^{-1}} = 8.206 \text{ atm}
\]

**Problem 4.** There are 25 grams of He and 56 grams of Ar gas mixed in a 20 L container. The temperature of the mixture is 300 K. Find the partial pressure of each gas and total pressure inside the container. Assume the ideal-gas law is valid for both gases.

**Solution.**
\[
n_{\text{total}} = \frac{25 \text{ g}}{4.0026 \text{ g mol}^{-1}} + \frac{56 \text{ g}}{39.948 \text{ g mol}^{-1}} = 7.65 \text{ mol}^{-1}
\]
\[
P = \frac{(7.65 \text{ mol}^{-1})(0.082057 \text{ L atm K}^{-1} \text{ mol}^{-1})(300 \text{ K})}{20 \text{ L}} = 9.42 \text{ atm}
\]
\[
P_{\text{He}} = \frac{25 \text{ g}}{7.65 \text{ mol}^{-1}}(9.42 \text{ atm}) = 7.69 \text{ atm}
\]
\[
P_{\text{Ar}} = 9.42 \text{ atm} - 7.69 \text{ atm} = 1.73 \text{ atm}
\]

**Problem 5.** You have two ideal gases, M and N, in a container of fixed volume. The mole fraction of M is \( X_M = \frac{1}{3} \). Initially the pressure was equal to \( P_i \). Later a total of two moles of one of the gases are added to the container, while holding the temperature constant. The new pressure is equal to \( \frac{11}{9} P_i \). Compute the number of moles of M and N that were initially present in the container.

**Solution.**
\[
P_f = \frac{11}{9} P_i, \quad \frac{11}{9} \left( n_M + n_N \right) \frac{RT}{V} = \frac{(n_M + n_N + 2)RT}{V},
\]
\[ n_M + n_N + 2 = \frac{11}{9} (n_M + n_N), \quad \frac{1}{3} = \frac{n_M}{n_M + n_N} \rightarrow n_N = 2n_M, \]
\[ 3n_M + 2 = \frac{11}{9} (3n_M), \quad n_M = 3 \text{ mol}, \quad n_N = 6 \text{ mol} \]

**Problem 6.** Suppose that the diameter of a gas molecule is 0.4 nm. Next, draw a sphere around each molecule, whose diameter is approximately equal to the average distance between neighboring molecules (assuming a spatially uniform gas). What is the sphere diameter in units of molecular diameters at 0°C and 1 bar.

**Solution.**

\[
\frac{V}{n} = \frac{RT}{P} = \frac{PV}{nRT} = \frac{(8.31)(273)}{100,000} \approx 0.023 \text{ m}^3/\text{mol}
\]

\[
V_{\text{molecule}} = \frac{V}{nN_A} = \frac{0.023}{6.022 \times 10^{23}} \approx 3.8 \times 10^{-26} \text{ m}^3 = 38 \text{ nm}^3
\]

\[ V_{\text{sphere}} = \frac{4\pi r^3}{3} \]

The answer is:

\[
2 \left( \frac{3V_{\text{molecule}}}{4\pi} \right) d \approx \frac{2 \sqrt{9.07}}{0.4} \approx 10.4.
\]

**Problem 7.** Calculate the RMS speed of oxygen molecules having a typical kinetic energy of 10 kJ mol\(^{-1}\). At what temperature would this be the RMS speed?

**Solution.**

\[ \bar{v}^2 \frac{1}{2} = \sqrt{\frac{3RT}{M}}, \quad \text{KE} = \frac{3}{2} RT, \quad \bar{v}^2 \frac{1}{2} = \sqrt{\frac{2\text{KE}}{M}} = 7.9 \times 10^2 \text{ ms}^{-1}, \]
\[ T = \frac{2\text{KE}}{3R} = 802 \text{ K} \]

**Problem 8.** An ideal gas A has a temperature \( T_A \) and molar mass \( M_A \). A second gas B has a molar mass \( M_B \). At what temperature \( T_B \) will the gases have the same average speed?

**Solution.**

\[ \bar{v}_A = \sqrt{\frac{8RT_A}{\pi M_A}} \]
\[ \bar{v}_A = \bar{v}_B \]
\[ \frac{T_A}{M_A} = \frac{T_B}{M_B} \rightarrow T_B = T_A \left( \frac{M_B}{M_A} \right) \]
Problem 9. The variance is useful because it measures how far data points are spread relative to their mean value. Let $X$ be a random variable with discrete distribution $\{P(x_i)\}$ $(i = 1, \ldots, N)$. The variance is defined as $\sigma^2 \equiv \text{var}(X) = (X - \bar{X})^2 = \sum_{i=1}^{N} P(x_i)(x_i - \bar{x})^2$. (a) Expanding the square in $(X - \bar{X})^2$, and invoking the linearity property of expectation value (i.e. $aX + bY = a\bar{X} + b\bar{Y}$, where $Y = X^2$ here), show that the variance is also equal to $\bar{X}^2 - (\bar{X})^2$. (b) Instead of the variance, consider a slightly different quantity, $X - \bar{X}$. Calculate the expectation value of $X - \bar{X}$. Explain why the latter is not a very useful quantity.

Solution. 

$$X - \bar{X} = \sum_{i=1}^{N} P(x_i)(x_i - \bar{X}) = \sum_{i=1}^{N} P(x_i)x_i - \sum_{i=1}^{N} P(x_i)\bar{X} = \bar{x} - \bar{X} = 0$$

$X - \bar{X}$ is not a useful quantity because it is always zero.

Problem 10. Consider a volume $V$ of an ideal gas inside a selectively permeable piston. What would the volume be if all of the other extensive properties of the gas double and all the intensive properties become a third of their initial values?

Solution. Extensive properties are $V, n$. Intensive properties are $P, T$.

$$V = \frac{nRT}{P}, \quad V_f = \frac{2nR\frac{1}{3}T}{\frac{1}{3}P}, \quad V_f = 2\frac{nRT}{P}, \quad V_f = 2V$$

Problem 11. The partial pressure of oxygen in a mixture of oxygen and hydrogen is 0.200 atm, and that of hydrogen is 0.800 atm.

a) How many molecules of oxygen are in a 1.500-L container of this mixture at 40°C?
b) If a spark is introduced into the container, how many grams of water will be produced?

Solution. a) total pressure = 0.200 atm + 0.800 atm = 1.000 atm

$$n = \frac{PV}{RT}$$

$n = 0.05840$ mol of total gas

$$X_{O_2} = \frac{0.200 \text{ atm}}{1.000 \text{ atm}} = 0.200$$

$n_{O_2} = X_{O_2}n = 0.0117$ moles of $O_2$

b) $2 \text{H}_2 + \text{O}_2 = 2 \text{H}_2\text{O}$

$$X_{\text{H}_2} = \frac{0.800 \text{ atm}}{1.000 \text{ atm}} = 0.800$$
\[ n_{\text{H}_2} = X_{\text{H}_2} n = 0.0467 \text{ moles of } \text{H}_2 \]
0.0467 moles of \( \text{H}_2 \) / 2 = 0.0234 moles > 0.0117 moles \( \text{O}_2 \)
\( \text{O}_2 \) is thus the limiting reagent
0.0117 moles \times 2 = 0.0234 moles \( \text{H}_2\text{O} \)
0.0234 moles \times 18.0152 \text{ g/mol} = 0.422 \text{ g } \text{H}_2\text{O} \]

**Problem 12.** Researchers recently reported the first optical atomic trap. In this device, tightly focused beams of laser light replace the physical walls of conventional containers, creating a type of trap where the motion of the atoms is substantially reduced. The laser beams briefly (0.5 s) exert enough pressure to confine 500 sodium atoms in a volume of \( 1.0 \times 10^{-15} \text{ m}^3 \). The temperature of this gas can be cooled down to 0.00024 K (cooling means that the kinetic energy is lowered by the laser trap), a pretty low temperature. Compute the root-mean-square speed of atoms in this confinement.

**Solution.** \( M_{\text{NA}} = 22.9898 \text{ g/mol} = 0.0229898 \text{ kg/mol} \)

\[ v_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.31 \text{ J/(mol } \cdot \text{ K}) \times 0.00024 \text{ K}}{0.02298 \text{ kg/mol}}} = 0.51 \text{ m/s} \]

**Problem 13.** Calculate the average (mean or expectation value) of the function \( \sin(X) \) over the interval \([0,1]\), where \( X \) is a Gaussian-distributed random variable with mean 0 and variance 1, i.e. its pdf is \( p(x) = \frac{1}{\sqrt{2\pi}} e^{-x^2/2} \).

(a) Write down an expression for the mean value of \( \sin(X) \), namely \( \sin(X) \).

(b) Evaluate it numerically. (Hint: feel free to use a web site such as https://www.wolframalpha.com to get the result.)

**Solution.** (a) The integral we need to compute is:

\[ \overline{\sin(X)} = \int_{0}^{1} \frac{1}{\sqrt{2\pi}} e^{-x^2/2} \sin(x) dx \]

(b) In Wolfram Alpha, we can type

Integrate \((1/\sqrt(2*Pi))*\sin(x)\*exp(-x^2/2)\) from 0 to 1

in the box to get the result: 0.145352.

**Problem 14.** Show by direct calculation (integration) that the mean speed of an ideal gas molecule in 3D is \( \overline{v} = \sqrt{\frac{8k_B T}{\pi m}} = \sqrt{\frac{8RT}{\pi M}} \). (Hint: consult tables of Gaussian integrals.)

**Solution.** We need to evaluate the integral \( \overline{v} = \int_{0}^{\infty} v f(v) dv \), where \( f(v) \) is the Maxwell-Boltzmann speed distribution,

\[ f(v) = \left( \frac{m}{2\pi k_B T} \right)^{3/2} 4\pi v^2 e^{-mv^2/2k_B T} \]
From tables of Gaussian integrals, e.g. https://en.wikipedia.org/wiki/Gaussian_integral, we invoke:
\[
\int_0^\infty x^{2n+1} e^{-ax^2} \, dx = \frac{n!}{2a^{n+1}} = \frac{1}{2a^2} \left( \frac{2k_B T}{m} \right)^{2n+1}
\]
with \( n = 1 \) and \( a = m/2k_B T \). With this, we obtain \( \bar{v} = \sqrt{\frac{8k_B T}{\pi m}} \).

For the problems below, you are asked to calculate probabilities. In all cases, this amounts to computing the integral of the PDF over the range of interest. (To get the result numerically, you can use Wolfram Alpha.) For example, let \( X \) be a random variable taking values in the range \((-\infty, \infty)\). By definition, the PDF is a non-negative function \( p(x) \geq 0 \) that is normalized to 1, i.e. \( \int_{-\infty}^{\infty} p(x) \, dx = 1 \). Suppose you are asked what is the probability that \( X \) takes values in the range \([-1, 1]\), then you simply integrate over all \( x \) values in that range:
\[
P(-1 \leq X \leq 1) = \int_{-1}^{1} p(x) \, dx
\]
and knowing \( p(x) \), you would integrate to get the result. If you are asked what is the probability that \( X \) takes values less than 10, this would be an integral of \( p(x) \) over the set \((-\infty, 10]\):
\[
P(-\infty < X \leq 10) = \int_{-\infty}^{10} p(x) \, dx
\]
Likewise, the probability of \( X \) taking negative values is:
\[
P(X < 0) = \int_{-\infty}^{0} p(x) \, dx
\]
If you want to know the probability that \( X \) takes a value less than -1 but greater than -2, or greater than 3 and less than 4, i.e. in the range \([-1, -2] \cup [3, 4]\), then:
\[
P(-2 \leq X \leq -1 \text{ or } 3 \leq X \leq 4) = \int_{-2}^{-1} p(x) \, dx + \int_{3}^{4} p(x) \, dx
\]
You may also find the following page useful:
https://en.wikipedia.org/wiki/Probability_density_function

**Problem 15.** Show by direct integration that the Gaussian (normal) PDF
\[
p(x) = \frac{1}{\sqrt{2\pi \sigma^2}} e^{-\frac{(x-\mu)^2}{2\sigma^2}}
\]
is normalized to 1. You can use analytical or numerical techniques, as long as you explain what you did. The Gaussian (bell-shaped distribution) random variables are frequently encountered in the physical sciences.
Solution. Shifting the origin to 0 (i.e. substitution of variable \( y = x - \mu \)) and using the Gaussian integral:

\[
\int_{-\infty}^{\infty} e^{-y^2/2\sigma^2} \, dy = \sqrt{\pi} \sigma^2
\]

which cancels the prefactor \( \frac{1}{\sqrt{2\pi}\sigma^2} \) to give 1.

Problem 16. Use the Maxwell velocity distribution to calculate the probability of an oxygen (O\(_2\)) molecule (a molecule chosen randomly) would have an \( x \) component of velocity between 10 and 20 m/s at room temperature and standard pressure (a). Does the pressure matter? Why? (b) At 20 K. (c) Calculate the probability of an oxygen (O\(_2\)) molecule would have an \( x \) component of velocity between -20 and -10 m/s at room temperature and standard pressure. (d) Calculate the probability of an oxygen (O\(_2\)) molecule would have an \( x \) component of velocity between -20 and 20 m/s at room temperature and standard pressure. (e) What is the probability of a molecule having an \( x \) velocity component exactly equal to 20 m/s? (Hint: you may use Wolfram Alpha to get the numerical result, but clearly explain what you did.)

Solution. The velocity distribution (PDF) for \( v_x \) is:

\[
f_{v_x}(v_x) = \left( \frac{m}{2\pi k_B T} \right)^{1/2} e^{-mv_x^2/2k_B T} = \left( \frac{M}{2\pi RT} \right)^{1/2} e^{-Mv_x^2/2RT}
\]

where \( m \) is the molecular mass and \( M \) is the molar mass (\( M=0.032 \) kg/mol for O\(_2\)). Pressure does not matter because ideal gas assumes no intermolecular interactions. Plug in \( T=293 \) K for (a); the factor \( \sqrt{\frac{M}{2\pi k_B T}} \) evaluates numerically to \( \frac{1.45}{\sqrt{0.032/(2\pi \times 8.31 \times 293)}} \approx 1.45 \times 10^{-3} \). For \( T=20 \) K it is \( \frac{1.45}{\sqrt{0.032/(2\pi \times 8.31 \times 20)}} \approx 5.54 \times 10^{-3} \). In Wolfram, type:

integrate 1.45e-3*exp(-0.032*v^2/(2*8.31*293)) from 10 to 20

and get 0.014479. For (b) we use \( T=20 \) K for (b), and integrate from 10 to 20 m/s:

integrate 5.54e-3*exp(-0.032*v^2/(2*8.31*20)) from 10 to 20

and get 0.0541713. For (c) integrate from -20 to -10 at RT:

integrate 1.45e-3*exp(-0.032*v^2/(2*8.31*293)) from -20 to -10

and get 0.0144778. For (d), from -20 to 20:

integrate 1.45e-3*exp(-0.032*v^2/(2*8.31*293)) from -20 to 20

and get 0.0579492. For (e), the probability is zero, because the integral has no width (the length of the interval of integration is zero).

Problem 17. Use the Maxwell speed distribution to calculate the probability an oxygen (O\(_2\)) molecule would have a speed between 10 and 20 m/s at room temperature and standard pressure (a). (b) What is the probability of the molecule having a speed between -20 and -10 m/s? (c) Probability of
speed between 0 and 1000 m/s? (d) What is the probability of a molecule having exactly the speed 20 m/s? 

**Solution.** The Maxwell speed distribution is:

\[ f_V(v) = \left( \frac{M}{2\pi RT} \right)^{3/2} 4\pi v^2 e^{-Mv^2/2RT}. \]

The numerical factor is \((4\pi)(0.032/(2\pi \times 8.31 \times 293))^{3/2} = 3.8 \times 10^{-8}\). For (a), in Wolfram Alpha we type

\[
\text{integrate } 3.8e-8*v^2*exp(-0.032*v^2/(2*8.31*293)) \text{ from 10 to 20}
\]

and get 0.000089. For (b) if we naively type

\[
\text{integrate } 3.8e-8*v^2*exp(-0.032*v^2/(2*8.31*293)) \text{ from -20 to -10}
\]

we get 0.00496 (same as in (a)). However, negative speeds are not allowed, so the speed distribution is not defined for negative speeds. Thus, this probability is zero. For (c),

\[
\text{integrate } 3.8e-8*v^2*exp(-0.032*v^2/(2*8.31*293)) \text{ from 0 to 1000}
\]

we get 0.995253. For (d), the probability is zero. ■

**Problem 18.** Consider molecules in thermal equilibrium that can diffuse only in 1 direction (say, \(x\)). The Maxwellian distribution (PDF) for the velocity component \(v_x\) was given in class:

\[ f_{v_x}(v_x) = \sqrt{\frac{m}{2\pi k_B T}} e^{-mv_x^2/2k_BT} \]

The meaning of this PDF is that the product \(f_{v_x}(v_x)dv_x\) gives the probability that a molecule chosen at random will have an \(x\) component of velocity between \(v_x\) and \(v_x + dv_x\). Recall that in 3D the speed is \(v \equiv \|\vec{v}\| = \sqrt{v_x^2 + v_y^2 + v_z^2}\). We computed the speed distribution by integrating the 3D velocity distribution over a spherical shell of volume \(4\pi v^2dv\). The speed distribution (PDF), \(f(v)\), has the meaning that the product \(f(v)dv\) gives the probability that a molecule chosen at random will have a speed between \(v\) and \(v + dv\). However, in 1 dimension, the speed is \(v \equiv \|\vec{v}\| = |v_x|\). Calculate the speed distribution in the 1D case. Check that your result is normalized.

**Solution.** The velocity distribution integrated over the set of points \(v_x\) such that \(v \equiv \|\vec{v}\| = |v_x|\). There are two such points: \(v_x = v\) and \(v_x = -v\). Denoting \(V\) the *random variable* that corresponds to speed, the speed
distribution, $f(v)dv = P(v < V < v + dv)$, is then the sum of two terms:

$$P(v < V < v + dv) = \int_{v_x:v < V < v + dv} f_{v_x}(v_x)dv_x$$

$$= \left[ \sqrt{\frac{m}{2\pi k_B T}} e^{-m(v^2)/2k_B T} + \sqrt{\frac{m}{2\pi k_B T}} e^{-mv^2/2k_B T} \right] dv$$

$$= \sqrt{\frac{m}{2\pi k_B T}} e^{-mv^2/2k_B T} dv$$

where the notation $\{v_x : v < V < v + dv\}$ means we integrate over all points $v_x \in (-\infty, \infty)$ such that the speed $V$ is between $v$ and $v + dv$. Normalization is easy to check by integrating from 0 to $\infty$; you will see that the factor of 2 is essential. 

Problem 19. Calculate the speed distribution in 2D, $P(v < V < v + dv) = f(v)dv$. The 2D velocity distribution is:

$$f_{v_x}(v_x)f_{v_y}(v_y) = \left( \frac{m}{2\pi k_B T} \right) e^{-m(v^2_x + v^2_y)/2k_B T}.$$ 

The meaning of this velocity distribution is that the product $f_{v_x}(v_x)f_{v_y}(v_y)dv_xdv_y$ gives the probability that a molecule chosen randomly will have $x$ velocity component in the range $[v_x, v_x + dv_x]$ and $y$ velocity component in the range $[v_y, v_y + dv_y]$. (Hint: integrate $f_{v_x}(v_x)f_{v_y}(v_y)dv_xdv_y$ over the set of points $(v_x, v_y)$ such that $v < V < v + dv$; it is easiest to use polar coordinates.) Check that the resulting PDF is normalized.

Solution. In polar coordinates the area element is $vdv\theta$. Then, $v = \sqrt{v_x^2 + v_y^2}$ and

$$P(v < V < v + dv) = \int_{\{(v_x,v_y):v < V < v + dv\}} f_{v_x}(v_x)f_{v_y}(v_y)dv_xdv_y$$

$$= \left( \frac{m}{2\pi k_B T} \right) \int_0^{2\pi} d\theta \int_v^{v + dv} ve^{-mv^2/2k_B T} dv$$

$$= \left( \frac{m}{2\pi k_B T} \right) 2\pi ve^{-mv^2/2k_B T} dv.$$ 

To check for normalization, we borrow the result $\int_0^\infty xe^{-ax^2} dx = 1/(2a)$ from tables of integrals. Integrating $f(v)dv$ from 0 to $\infty$ gives:

$$\int_0^\infty f(v)dv = \left( \frac{m}{2\pi k_B T} \right) 2\pi \int_0^\infty ve^{-mv^2/2k_B T} dv = \left( \frac{m}{2\pi k_B T} \right) 2\pi \frac{2k_B T}{2m} = 1.$$
Problem 20. Now that you have done the 1D and 2D cases, go back and study how the 3D case was done (see Section 1.3.8). What is the difference between 1D, 2D and 3D cases?

Solution. \( f(v) \) is proportional to \( v^n e^{-mv^2/2k_BT} \), where \( n = 0, 1, 2 \) (for 1D, 2D and 3D, respectively). The numerical coefficient also follows a pattern related to dimensionality. Namely, we have \( \left( \frac{m}{2\pi k_BT} \right)^{n/2} \), \( n = 1, 2, 3 \) as well as the other numerical factor, which is either 2, \( 2\pi \) or \( 4\pi \), depending on whether we have 1D, 2D or 3D, respectively.

Problem 21. Show that for an ideal gas, the most probable speed is \( u_p = \sqrt{\frac{2k_BT}{m}} \) and the average speed is \( \bar{u} = \sqrt{\frac{8k_BT}{\pi m}} \). Recall the Maxwell-Boltzmann speed distribution is

\[
    f(u)du = \left( \frac{m}{2\pi k_BT} \right)^{3/2} 4\pi u^2 \exp \left( -\frac{mu^2}{2k_BT} \right) du.
\]

Note: here we used \( u = \sqrt{v_x^2 + v_y^2 + v_z^2} \) for speed to avoid confusion with velocity \( v \) which we use in the next problem.

Solution. The most probable speed the value of \( u \) at the point where the first derivative of the distribution is zero. Let \( a = \frac{m}{2\pi k_BT} \). Dropping the constants, since we are equating the derivative to zero,

\[
    \left( \frac{m^{3/2}}{2\pi k_BT} \right)^{3/2} 4\pi u^2 \exp \left( -\frac{mu^2}{2k_BT} \right) = 0.
\]

We denote this particular value of \( u \) as \( u_p \). Therefore,

\[
    1 - au^2 = 0 \quad \text{or} \quad u_p = \sqrt{\frac{T}{a}} = \sqrt{\frac{2k_BT}{m}}.
\]

To calculate \( \bar{u} \), the integral that we need to solve is \( \int_0^\infty u^3 \exp(-au^2)\) \( du \). Since the remaining terms are constants that we will multiply in later. We do a change of variables by setting \( w = u^2 \) which also gives \( dw = 2u\,du \) and then do integration by parts.

\[
    \int_0^\infty u^3 \exp(-au^2)\,du = \frac{1}{2} \int_0^\infty w \exp(-aw)\,dw
\]

\[
    = -\frac{1}{2a} e^{-aw}\bigg|_0^\infty + \frac{1}{2a} \int_0^\infty e^{-aw}\,dw = \frac{1}{2a^2}.
\]

Now we restore the constants and substitute for \( a \),

\[
    \overline{u} = \sqrt{\frac{8k_BT}{\pi m}}.
\]

Problem 22. Show that \( v^2 = v_x^2 + v_y^2 + v_z^2 = \overline{v_x^2} + \overline{v_y^2} + \overline{v_z^2} \).
Solution. By linearity of the integral
\[
\overline{v^2} = v_x^2 + v_y^2 + v_z^2 = \sum_{\text{directions}} \left( \frac{m}{2\pi k_B T} \right)^{-3/2} \int_{\mathbb{R}^3} (v_x^2 + v_y^2 + v_z^2) \exp\left( -\frac{mv^2}{2k_B T} \right) d^3\mathbf{v}
\]
\[
= \left( \frac{m}{2\pi k_B T} \right)^{3/2} \int_{\mathbb{R}^3} v_x^2 \exp\left( -\frac{mv^2}{2k_B T} \right) d^3\mathbf{v} + \int_{\mathbb{R}^3} v_y^2 \exp\left( -\frac{mv^2}{2k_B T} \right) d^3\mathbf{v} + \int_{\mathbb{R}^3} v_z^2 \exp\left( -\frac{mv^2}{2k_B T} \right) d^3\mathbf{v}
\]
where the averaging in the last line is done over \( \int_{\mathbb{R}^3} \cdots d^3\mathbf{v} \). This averaging can also be done along a single variable, as explained below.

For the \( \overline{v_x^2} \) term, integration with respect to \( y \) and \( z \) is immediate, since the PDFs are normalized, i.e. using \( \int_x \left( \frac{m}{2\pi k_B T} \right)^{1/2} \exp\left( -\frac{mv^2}{2k_B T} \right) dv_y = 1 \) (and the analogous result for \( v_z \)):
\[
\left( \frac{m}{2\pi k_B T} \right)^{3/2} \int_{\mathbb{R}^3} v_x^2 \exp\left( -\frac{mv^2}{2k_B T} \right) d^3\mathbf{v} = \left( \frac{m}{2\pi k_B T} \right)^{3/2} \int_{-\infty}^{\infty} v_x^2 \exp\left( -\frac{mv^2}{2k_B T} \right) dv_x,
\]
which equals \( \langle v_x^2 \rangle \). The angle brackets in \( \langle v_x^2 \rangle \) denote averaging along a single variable \( \int_x \cdots dv_x \). Thus,
\[
\overline{v^2} = \langle v_x^2 \rangle + \langle v_y^2 \rangle + \langle v_z^2 \rangle
\]

\[\blacksquare\]

Problem 23. Show that each degree of freedom of a gas particle adds \( \frac{1}{2} k_B T \) to the average kinetic energy of the gas, \( K_{av} = \frac{1}{2} m \overline{v^2} \), where \( v^2 = v_x^2 + v_y^2 + v_z^2 \). In other words, show that
\[
\langle K_x \rangle = \langle K_y \rangle = \langle K_z \rangle = \frac{1}{2} k_B T
\]
where \( K_{av} = \langle K_x \rangle + \langle K_y \rangle + \langle K_z \rangle \) and \( K_x = \frac{1}{2} mv_x^2 \), \( K_y = \frac{1}{2} mv_y^2 \) and \( K_z = \frac{1}{2} mv_z^2 \). The angle brackets denote averaging along a single direction, e.g. \( \langle K_x \rangle = \int \frac{1}{2} mv_x^2 \cdot f_{V_x}(v_x) dv_x \).

Solution.
\[
\langle K_x \rangle = \frac{1}{2} m \langle v_x^2 \rangle
\]
We have just shown that \( \overline{v^2} = \langle v_x^2 \rangle + \langle v_y^2 \rangle + \langle v_z^2 \rangle \) so the corresponding equality for kinetic energy is also true. We now just need to calculate \( \langle v_x^2 \rangle \).
\[
\langle v_x^2 \rangle = \left( \frac{m}{2\pi k_B T} \right)^{3/2} \int_{-\infty}^{\infty} v_x^2 \exp\left( -\frac{mv_x^2}{2k_B T} \right) dv_x
\]
Let us switch to a more convenient notation:
\[
\int_{-\infty}^{\infty} v_x^2 \exp \left( -\frac{mv_x^2}{2k_BT} \right) dv_x = \int_{-\infty}^{\infty} x^2 \exp(-ax^2)dx = -\frac{d}{da} \int_{-\infty}^{\infty} \exp(-ax^2)dx
\]
which equals to \(-\frac{d}{da} \sqrt{\frac{\pi}{a}}\), according to the well-known result for Gaussian integration. Then,
\[
-\frac{d}{da} \sqrt{\frac{\pi}{a}} = \frac{1}{2} \sqrt{\frac{\pi}{a^3}} = \frac{\sqrt{\pi}}{2} \left( \frac{2k_BT}{m} \right)^{\frac{3}{2}}.
\]
Finally,
\[
\langle K_x \rangle = \frac{1}{2} m \langle v_x^2 \rangle = \frac{1}{2} k_BT.
\]
Same for \(K_y\) and \(K_z\). Therefore, we see that \(K_{av}\), which equals \(\frac{1}{2}mv_x^2 = \frac{3}{2}k_BT\), is made up of 3 contributions, each equalling \(\frac{1}{2}k_BT\). This is called the equipartition theorem.

Problem 24. In Section 1.4.6, we expressed the equation of state for a gas as a Taylor expansion in terms of gas density, \(\rho = \frac{N}{V}\). That is,
\[
P = \frac{RT}{\rho} = \rho + B_2(T)\rho^2 + B_3(T)\rho^3 + \ldots
\]
This expansion is called the virial expansion. \(B_n(T)\) is the \(n^{th}\) virial coefficient. The virial expansion is a generalization of the ideal gas law \(\frac{P}{RT} = \rho\) that includes higher powers of \(\rho\) to account for the gas behavior at higher densities. (The ideal gas law is only valid in the low density limit.)

(a) Show that for the VDW equation of state,
\[
B_2(T) = \frac{b}{N_A} - \frac{a}{N_A^2 k_BT}.
\]
(b) The temperature for which \(B_2(T) = 0\) is called Boyle’s temperature, \(T_B\). Find the expression for \(T_B\) for a VDW gas. Explain what is the Boyle temperature and discuss what are practical applications of the Boyle’s temperature.

Solution. Let us compare the virial expansion:
\[
\frac{P}{k_BT} = \rho + B_2(T)\rho^2 + B_3(T)\rho^3 + \ldots
\]
where \(\rho = N/V\), with the VDW equation of state:
\[
\left( P + \frac{n^2a}{V^2} \right) (V - nb) = nRT.
\]
Expanding the product:
\[
PV - Pnb + \frac{n^2a}{V} - \frac{n^3ab}{V^2} = nRT.
\]
Dividing throughout by $RT(V - nb)$:

$$\frac{P}{RT} = \frac{n}{(V - nb)} + \frac{n^3 ab}{V^2 RT(V - nb)} - \frac{n^2 a}{VRT(V - nb)}.$$  

Writing $\tilde{\rho} = n/V$:

$$\frac{P}{RT} = \frac{\tilde{\rho}}{(1 - b \tilde{\rho})} + \frac{\tilde{\rho}^3 ab}{RT(1 - b \tilde{\rho})} - \frac{\tilde{\rho}^2 a}{RT(1 - b \tilde{\rho})}.$$  

Using the geometric series $\frac{1}{1 - r} = 1 + r + r^2 + \ldots$ with $r = b \tilde{\rho}$ we get:

$$\frac{P}{RT} = \left\{ \tilde{\rho} + \tilde{\rho}^3 ab \frac{RT}{\tilde{\rho}} - \tilde{\rho}^2 a \frac{RT}{\tilde{\rho}} \right\} (1 + b \tilde{\rho} + b^2 \tilde{\rho}^2 + \ldots).$$  

Expanding,

$$\frac{P}{RT} = \tilde{\rho} + \rho^2 \left( b - \frac{a}{RT} \right) + \rho^3 \left( b^2 \right) + \ldots$$  

Finally, we replace $R = N_A k_B$ and $\tilde{\rho} = n/V = N/(N_A V) = \rho/N_A$

$$\frac{P}{k_B T} = \rho + \rho^2 \left( \frac{b}{N_A} - \frac{a}{N_A^2 k_B T} \right) + \rho^3 \left( \frac{b^2}{N_A^3} \right) + \ldots$$  

(1.14)

Identifying the coefficients of like powers of $\rho$ in (1.14) and (1.13), we find:

$$B_2(T) = \frac{b}{N_A} - \frac{a}{N_A^2 k_B T}, \quad B_3(T) = \frac{b^2}{N_A^3}, \quad \text{etc.}$$  

The point at which $B_2(T) = 0$ is called the Boyle temperature:

$$T_B = \frac{a}{b k_B N_A} = \frac{a}{b R}.$$  

It is at the Boyle temperature that the attractive and repulsive forces acting on the gas particles balance out. At that temperature, the non-ideal gas behaves like an ideal gas over some range of pressures. (It’s an ideal gas as far as the second virial coefficient is concerned; there are of course, higher-order terms that do not enforce the ideal-gas behavior, although these terms get progressively smaller in magnitude.)

Applications of Boyle temperature: gas cylinders and compression pumps are industrially important. If a gas in a mechanical cylinder liquefies it is a problem. The mechanical device needs to be above the Boyle temperature.

In thermodynamics, the Joule-Thomson (JT) effect describes the temperature change of real gases when forced through a valve or porous plug while being insulated from heat exchange with the environment. Whether the gas will cool or heat up depends on whether attractive or repulsive forces predominate at the temperature in which the forced expansion occurs. The
temperature at which this reversal occurs is known as the inversion temperature $T_i$ which is defined as when the JT coefficient $\mu_{JT}$ becomes zero.

$$\mu_{JT} = \left. \frac{\partial T}{\partial P} \right|_H = 0$$

Solving this expression for the VDW gas (undergoing an isenthalpic process) gives $T_i = 2T_b$ and is used to determine whether an isenthalpic expansion of a real gas will cause it to cool or heat up. Thus, the concept of the Boyle temperature is not only important in refrigeration and air conditioning but also industrially for gas liquefaction. One such process of turning gases into liquids is the Hampson-Linde cycle which needs a positive $\mu_{JT}$ to operate. In other words, a gas must be below its inversion temperature to be liquefied by this process, which is true for all real gases at room temperature except for helium, hydrogen, and neon. If a gas undergoes a JT expansion at a temperature above its $T_i$, then it will actually heat up. ■

**Problem 25.** Derive expressions for the pressure, molar volume, and temperature of a VDW gas at its critical point in terms of VDW constants $a$ and $b$. The critical point is defined as the point on the $P - V_m$ curve where the first and second derivatives of the pressure with respect to molar volume vanish, i.e. $\left. \frac{\partial P}{\partial V_m} \right|_c = 0$ and $\left. \frac{\partial^2 P}{\partial V_m^2} \right|_c = 0$.

**Solution.**

$$P = \frac{RT}{V_m - b} - \frac{a}{V_m^2}$$

$$\left. \frac{\partial P}{\partial V_m} \right|_c = -\frac{RT_c}{(V_m - b)^2} + \frac{2a}{V_m^3} = 0$$

$$\left. \frac{\partial^2 P}{\partial V_m^2} \right|_c = \frac{2RT_c}{(V_m - b)^3} - \frac{6a}{V_m^4} = 0$$

Solving the above expressions for $V_c$ and $T_c$ then substituting into the vdw equation of state to solve for $P_c$ results in

$$V_c = 3b, \quad T_c = \frac{8a}{27bR}, \quad P_c = \frac{a}{27b^2}.$$  ■

**Problem 26.** The diffusion coefficient of CCl$_4$ in heptane at 25.0 °C is 3.89 x 10$^{-9}$ m$^2$ s$^{-1}$. How long will it take for a CCl$_4$ molecule to undergo a mean-square displacement of 5.5 mm?

**Solution.** In 3D, the mean-square displacement is

$$\Delta r^2 = \Delta x^2 + \Delta y^2 + \Delta z^2 = 6Dt$$
Substituting values

\[ t = \frac{(0.0055 \text{ m})^2}{6(3.89 \times 10^{-9} \text{ m}^2 \text{ s}^{-1})} = 1.30 \times 10^3 \text{ s} \]

Problem 27. Calculate the pressure necessary for the mean free path \( \lambda \) of argon at 25.0 °C to become comparable to the size of the 1.00 L vessel that contains it (in other words, when \( \lambda = V^{1/3} \)). Take the radius of an Ar atom to be 188 pm. Assume argon behaves as an ideal gas.

Solution. The mean free path can be written as

\[ \lambda = \frac{1}{\sqrt{2\pi d^2 N/V}} \]

The number density of an ideal gas can be written as

\[ \frac{N}{V} = \frac{P}{k_B T} \]

Combining the above expressions, substituting values, and solving for \( P \)

\[ P = \frac{k_B T}{\sqrt{2\pi d^2 \lambda}} = \frac{(1.381 \times 10^{-23} \text{ J/K})(298.15 \text{ K})}{\sqrt{2\pi(2 \times 188 \times 10^{-12} \text{ m})^2(0.100 \text{ m})}} = 0.0656 \text{ Pa} \]
Intermolecular Interactions

Intermolecular forces are very important in chemistry because they determine the properties of dense matter such as the nature of phase transitions. For example, the boiling point depends on the strength of intermolecular interactions: tightly bound solvent molecules typically have a higher boiling point. In a (dense) gas, intermolecular interactions may manifest themselves in deviations from the ideal gas law at higher pressures. In a liquid, they determine the spatial arrangement of solvent molecules or the interaction of solvent molecules with solutes. In a crystalline solid, these forces hold the crystal together.

2.1. Electrostatics

The electric field \( E \) is a vector field whose SI units are N/C (Newtons per Coulomb) or V/m (Volts per meter). It gives rise to the force exerted on a charged particle, \( q \) (units: Coulomb), according to the Lorentz force law:

\[
F = q(E + v \times B),
\]

(\( v \): particle velocity; \( B \): magnetic field, assumed zero here.)

The electric field of a point charge falls off as \( 1/R^2 \), where \( R \) is the distance to the charge, whereas the vector field \( E \) points along the radial direction as shown in Fig. 2.1. The fact that the field falls off with distance is illustrated by the lower density of \( E \)-field lines away from the charge. For a charge of opposite sign, the direction of the arrows is reversed.
2. Intermolecular Interactions

Figure 2.1. Electric field lines of a negatively charged point charge are radial and point towards the charge.

Figure 2.2. An electric dipole consists of opposite charges separated by a distance. The electric field flows from the positive to the negative charge.

For two point charges, the electric field lines are created at the site of the positive charge and extend toward the negative charge, as shown in Fig. 2.2. The density of lines represents the strength of the electric field locally. For two charges of the same sign, the field lines form a repulsive pattern (Fig. 2.3).

For a parallel plate capacitor (two parallel electrodes, with opposite charges), the field lines are parallel, as shown in Fig. 2.4. The figure also shows that an electron traveling initially to the right will be attracted toward the positive electrode, according to the Lorentz law: $F = qE = (-e)E$, i.e., $F$ points downward if $E$ points upward. $e$ is the elementary charge, $1.602176 \times 10^{-19}$ C. The charge of the proton is $e$ and the charge of the electron is $-e$.

2.1.1. Electric Dipole. Let $q > 0$ be an electric charge. A permanent electric dipole moment (EDM), $p$, is a physical arrangement of a positive
2.1. Electrostatics

Figure 2.3. Two electric charges of the same sign lead to repulsion. The electric field lines still flow from positive to negative charges, except that there are no negative charges here, so the lines go to infinity where the charge is eventually zero.

Figure 2.4. Parallel plate capacitor. Electric field lines flow from positive to negative charges (upward, in this example). A charge $q$ enters the capacitor from the left with initial velocity $\vec{v}$ along $x$. The dotted line shows the downward accelerating trajectory of the charge, implying that $q < 0$.

Figure 2.5. The EDM is a vector, $\mathbf{p} = q\mathbf{d}$, that points from the negative ($-q$) to the positive ($+q$) charge ($q > 0$). Its length is $q$ times $|d|$.

$(+q)$ and a negative ($-q$) charge separated by a distance $d$. The dipole moment is defined as $\mathbf{p} = q\mathbf{d}$. Its SI units are charge times distance [Coulomb × meter]. It is a vector pointing along the intermolecular vector $\mathbf{d}$. By convention, we take it to point from the negative to the positive charge. This is illustrated in Fig. 2.5.

The CGS unit for dipole moment is Debye. 1 D $\approx 3.336 \times 10^{-30}$ C·m. The Debye is defined to be equal to $10^{-10}$ esu·Å, where 1 esu (electrostatic units) $\approx 3.336 \times 10^{-10}$ C. Inter-atomic distances within molecules are typically measured in units of angstrom (Å), where 1 Å $= 10^{-10}$ m.
Dipole moments in a molecule can be permanent or induced. $\text{H}_2\text{O}$, for example, has a permanent EDM. Molecules without a permanent EDM can still be polarized by the application of an electric field (see Section 2.1.2). A consequence of $\mathbf{p} \neq 0$ is that the electric dipole will want to orient itself along the electric field lines as a consequence of the Lorentz force $\mathbf{F} = q\mathbf{E}$ which leads to a torque $\tau = \mathbf{p} \times \mathbf{E}$ ($\times$: cross product\(^1\)) on the dipole, as shown in Fig. 2.6.

\(^1\)The cross product $\mathbf{x} \times \mathbf{y}$ of two vectors $\mathbf{x} = (x_1, x_2, x_3)$ and $\mathbf{y} = (y_1, y_2, y_3)$ is defined as the vector obtained from the determinant of the matrix

\[
\begin{vmatrix}
\hat{x} & \hat{y} & \hat{z} \\
x_1 & x_2 & x_3 \\
y_1 & y_2 & y_3
\end{vmatrix} = \hat{x}(x_2y_3 - x_3y_2) + \hat{y}(x_3y_1 - x_1y_3) + \hat{z}(x_1y_2 - x_2y_1).
\]
Dipole moments can occur between two ions in an ionic bond or between atoms in a covalent bond; dipole moments arise from differences in electronegativity. The larger the difference in electronegativity, the larger the dipole moment. The distance between the charge separation is also a deciding factor into the size of the dipole moment. The dipole moment is a measure of the polarity of the molecule. Electronegativities can be found online, e.g., http://webelements.com. Such a table is reproduced here (Fig. 2.7).

In Fig. 2.8 a molecule of HCl can be represented as having a small net negative charge on the Cl end, balanced by a small net positive charge on the H end. The overall charge is zero, however, the distribution of charges is spatially-dependent. The forces between two HCl molecules depend on their orientations. (a) The oppositely charged ends (green arrows) are closer

---

Its magnitude is $|x||y|\sin \theta$. $\theta$ is the angle between $x$ and $y$. Its geometric meaning is the area of the parallelogram traced with sides $x$ and $y$. It is a vector pointing in a direction perpendicular to the plane spanned by $x$ and $y$. $x \times y = 0$ means the two vectors are parallel to each other.
than the ends with the same charge (red arrows). This gives a net attractive force. (b) Here, the opposite is true, and the net force is repulsive.

![Attraction and Repulsion Diagram](image)

**Figure 2.10.** Permanent EDM in H$_2$O leads to attraction or repulsion between pairs of molecules, depending on the relative orientation of these molecules. Four different examples are shown here.

**Table 2.1.** Dipole moments for various molecules. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Compound</th>
<th>Dipole Moment (Debye)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>9.0 (meas. in gas phase)</td>
</tr>
<tr>
<td>CH$_3$Cl</td>
<td>1.87</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>1.85</td>
</tr>
<tr>
<td>NH$_3$</td>
<td>1.47</td>
</tr>
<tr>
<td>CO$_2$</td>
<td>0</td>
</tr>
<tr>
<td>CCl$_4$</td>
<td>0</td>
</tr>
</tbody>
</table>

A popular example of permanent EDM is the water molecule H$_2$O, which has partial positive charge on the hydrogen and partial negative charges on its oxygen, leading to the formation of hydrogen bonds in water (see Fig. 2.9). The molecule H$_2$O is said to be polar because it possesses a permanent electric dipole moment indicated by the arrow in Fig. 2.11. This interaction can be interpreted in terms of Coulomb’s law (like charges repel each other whereas unlike charges attract each other), as shown in Fig. 2.10.

Table 2.1 lists dipole moments of several popular compounds.
2.1. Electrostatics

\[ \delta^+ \quad \vec{p} \quad \delta^+ \]

**Figure 2.11.** Net electric dipole moment of a water molecule forms as the vector sum of two dipole moments, each along the two H-O bonds, pointing from O to H, since O is more electronegative. According to our convention for the direction of the EDM, the vector $\vec{p} = q\vec{d}$ points upward, from the negative partial charge to the positive side.

Tables 2.2, 2.3 and 2.4 are some values of permanent dipole moments of selected molecules. The unit is the Debye (D). Recall that 1 D=$3.336 \times 10^{-30}$ C·m. We note that hydrogen-bonding molecules can have different dipole moments in the gas, liquid, and solid phases, as well as in different solvents.

**Table 2.2.** Dipole moments of molecules. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Molecules</th>
<th>Molecule</th>
<th>Dipole Moment (D)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkanes</td>
<td>$\text{C}_n\text{H}_m$ (benzene)</td>
<td>0</td>
</tr>
<tr>
<td></td>
<td>$\text{C}<em>9\text{H}</em>{2n+1}\text{OH}$ (alcohols)</td>
<td>1.7</td>
</tr>
<tr>
<td>CCl₄</td>
<td>$\text{C}<em>6\text{H}</em>{11}\text{OH}$ (cyclohexanol)</td>
<td>1.7</td>
</tr>
<tr>
<td>CO₂</td>
<td>OMCTS</td>
<td>0.42</td>
</tr>
<tr>
<td>CO</td>
<td>CH₃COOH (acetic acid)</td>
<td>1.7</td>
</tr>
<tr>
<td>CHCl₃ (chloroform)</td>
<td>C₂H₄O (ethylene oxide)</td>
<td>1.9</td>
</tr>
<tr>
<td>HCl</td>
<td>CH₂COCH₃ (acetone)</td>
<td>2.9</td>
</tr>
<tr>
<td>HF</td>
<td>HCONH₂ (formamide)</td>
<td>3.7</td>
</tr>
<tr>
<td>NH₃</td>
<td>C₆H₅OH (phenol)</td>
<td>1.5</td>
</tr>
<tr>
<td>CH₃Cl</td>
<td>C₆H₅NH₂ (aniline)</td>
<td>1.5</td>
</tr>
<tr>
<td>NaCl</td>
<td>C₆H₅Cl (chlorobenzene)</td>
<td>1.8</td>
</tr>
<tr>
<td>CsCl</td>
<td>C₆H₅NO₂ (nitrobenzene)</td>
<td>4.2</td>
</tr>
</tbody>
</table>

**Table 2.3.** Bond Moments. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Bond Moments</th>
<th>C-H⁺ 0.4</th>
<th>C-C 0</th>
<th>C⁺-Cl 1.5-1.7</th>
</tr>
</thead>
<tbody>
<tr>
<td>N-H⁺</td>
<td>1.31</td>
<td>C=C 0</td>
<td>N⁺-O 0.3</td>
</tr>
<tr>
<td>O-H⁺</td>
<td>1.51</td>
<td>C⁺-N 0.22 C⁺=O 2.3-2.7</td>
<td></td>
</tr>
<tr>
<td>F-H⁺</td>
<td>1.94</td>
<td>C⁺-O 0.74 N⁺=O 2.0</td>
<td></td>
</tr>
</tbody>
</table>

**Table 2.4.** Bond Moments. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Group Moments</th>
<th>C⁺CH₃ 0.4</th>
<th>C⁺COOH 1.7</th>
<th>Adenine ~ 3</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>C⁺OH 1.65</td>
<td>C⁺OCH₃ 1.3</td>
<td>Thymine ~ 4</td>
</tr>
<tr>
<td></td>
<td>C⁺NH₂ 1.2-1.5</td>
<td>C⁺NO₂ 3.1-3.8</td>
<td>Guanine ~ 7</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Cytosine ~ 8</td>
</tr>
</tbody>
</table>
2.1.2. Electric Polarization. Suppose that we have an electrically neutral molecule placed in an electric field $\mathbf{E}$. It can acquire a dipole moment $\mathbf{p}$ through the action of the electric field by distorting its electron cloud. This is the case regardless of whether or not the molecule has a permanent EDM. The application of an electric field to a molecule gives rise to an induced EDM whose magnitude is proportional to $\mathbf{p}$:

$$\mathbf{p} = \epsilon_0 \alpha \mathbf{E},$$

where $\mathbf{E}$ is the “local” electric field at the site of the molecule, $\epsilon_0 = 8.85 \times 10^{-12}$ Farad per meter (F/m) is the permittivity of free space and $\alpha$ is the molecular polarizability (or simply, “polarizability”). The latter is a property that is specific to the molecule. Polarizabilities for several molecules can be found in Table 2.5.

Question: What must be the SI units of $\alpha$?

Solution: From $\mathbf{p} = \epsilon_0 \alpha \mathbf{E}$, $\alpha = \mathbf{p}/(\epsilon_0 \mathbf{E})$, where $E$ is V/m, $p$ is C.m and $\epsilon_0$ is F/m. Thus, $\alpha$ has units C.m/(F.V/m/m) or $m^3$, since Farad is defined as Coulombs per Volt.

Alternate Units: Alternatively, we absorb $\epsilon_0$ into $\alpha$, keeping SI units, and write $\mathbf{p} = \alpha \mathbf{E}$ instead of $\mathbf{p} = \epsilon_0 \alpha \mathbf{E}$. In this case, $\alpha$ has SI units of $C.m^2/V$.

Closely Related Units: A related set of units (which differs from the previous one by $4\pi$), which we shall use in Section 2.6, is $\mathbf{p} = \alpha \mathbf{E}$, where $\alpha$ is given in units of $(4\pi\epsilon_0)^{\frac{1}{3}} = 1.11 \times 10^{-40} \text{C.m}^2/\text{V}$. (See Table 2.5.)

Example 2.1. Equation $\mathbf{p} = \alpha \mathbf{E}$ is to be used with the data from Table 2.5. Let’s look at the case of CCl$_4$, which has a polarizability of 10.5 and units of $(4\pi\epsilon_0)^{\frac{1}{3}}$. Suppose the electric field is 110 V between two electrodes separated 1 cm apart, i.e. $E=\frac{110 \text{ V}}{1 \text{ cm}}=11,000 \text{ V/m}$.

$$\mathbf{p} = \alpha \mathbf{E} = (4\pi\epsilon_0)^{\frac{1}{3}} (10.5) (110 \text{ V})$$

$$= (1.11 \times 10^{-40} \text{ C.m}^2/\text{V}) (10.5) (11,000 \text{ V/m})$$

$$= 1.282 \times 10^{-35} \text{ C.m}$$

Because our formula uses SI units, the units of $\mathbf{p}$ are automatically C.m, as long as $\alpha$ and $E$ are in SI units. The SI units of $E$ are V/m. The SI units of $\alpha$ are C.m$^2$/V (and if polarizability values from Table 2.5 are used, we must multiply by the scaling factor $1.11 \times 10^{-40}$). This magnitude can be compared to a permanent EDM. While CCl$_4$ has no permanent EDM, other molecules such as NaCl or H$_2$ do. According to Table 2.1, H$_2$O has

---

Sometimes this is written in terms of the polarization density, $\mathbf{P}$ (dipole moment per unit volume), where $\mathbf{P} = \chi_e \epsilon_0 \mathbf{E}$, where $\chi_e$ is the electric susceptibility. $\mathbf{p}$ is obtained from $\mathbf{P}$ by multiplying the latter by volume. $\chi_e$ and $\alpha$ are related by $\chi_e = \alpha N/V$, where $N/V$ is the number of molecules per unit volume.
2.1. Electrostatics

Table 2.5. Electronic Polarizabilities of Molecules. Polarizabilities here are given in volume units of $(4\pi\epsilon_0)\lambda^3 = (4\pi\epsilon_0)10^{-30} \text{ m}^3 = 1.11 \times 10^{-40} \text{ C}^2 \text{ m}^2 \text{ J}^{-1}$. 1 Joule = 1 Coulomb times Volt (J=C.V). The numerical value of $4\pi\epsilon_0$ is $4\pi \times 8.854 \times 10^{-12} \sim 1.11 \times 10^{-10}$. Note that when molecules are dissolved in a solvent medium, their polarizability can change by up to 10%. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Atoms and Molecules</th>
<th>He</th>
<th>NH₃</th>
<th>CH₂=CH₂</th>
<th>C₂H₆</th>
<th>H₂O</th>
<th>HCl</th>
<th>Cl₂</th>
<th>CO₂</th>
<th>CH₃OH</th>
<th>C₆H₆</th>
<th>CO</th>
<th>Xe</th>
<th>CCl₄</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>0.20</td>
<td>2.3</td>
<td>4.3</td>
<td></td>
<td>0.81</td>
<td>2.6</td>
<td>4.5</td>
<td>1.60</td>
<td>3.2</td>
<td>10.3</td>
<td>1.95</td>
<td>4.0</td>
<td>10.5</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Bond Polarizabilities</th>
<th>C-C aliphatic</th>
<th>C-H</th>
<th>C-Cl</th>
<th>C≡C aromatic</th>
<th>O-H</th>
<th>C-F</th>
<th>Si-Si</th>
<th>C≡C</th>
<th>C=O</th>
<th>Si-H</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>0.48</td>
<td>0.65</td>
<td>2.60</td>
<td>1.07</td>
<td>0.73</td>
<td>0.73</td>
<td>2.24</td>
<td>2.39</td>
<td>1.36</td>
<td>1.27</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Molecules Groups</th>
<th>C-O-H</th>
<th>-CH₂-</th>
<th>CF₃</th>
<th>C-O-C</th>
<th>CH₃</th>
<th>Si-O-Si</th>
<th>C-NH₂</th>
<th>-CF₂-</th>
<th>Si-OH</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1.3</td>
<td>1.84</td>
<td>2.4</td>
<td>1.1</td>
<td>2.0</td>
<td>1.4</td>
<td>2.0</td>
<td>1.6</td>
<td></td>
</tr>
</tbody>
</table>

A permanent EDM of 1.85 Debye. Since 1 D $\approx 3.3465 \times 10^{-30}$ C.m, its permanent EDM in SI units is $6.2 \times 10^{-30}$ C.m.

**Example 2.2.** What is the magnitude of the electric field strength required to induce an EDM in H₂O of the same magnitude as its permanent EDM?  **Solution:** According to the data from Table 2.1, H₂O has a permanent EDM of 1.85 Debye, which correspond to $6.2 \times 10^{-30}$ C.m in SI units. From Table 2.5, the polarizability of H₂O is in the range 1.45-1.48, with units of $1.11 \times 10^{-40}$ C.m$^2$/V. Let’s take 1.48. Equating the polarizations

$6.2 \times 10^{-30}$ C.m $= (1.11 \times 10^{-40}$ C.m$^2$/V)$(1.48)E$

and solving for $E$ gives:

$E = \frac{6.2 \times 10^{-30} \text{ C.m}}{(1.11 \times 10^{-40} \text{ C.m}^2/\text{V})(1.48)} = 38 \times 10^9 \text{ V/m}$.

**Example 2.3.** How does one generate an electric field of magnitude $38 \times 10^9$ V/m?  **Solution:** One way is to apply a voltage across electrodes that are very close to each other. For example, if we apply 110 V across electrodes separated by 3 nm, we will get an electric field of this magnitude. In practice, this is difficult to achieve. On the other hand, let’s look at the electric field of an ion. A point charge $Q$ generates an electric field of magnitude
$E = Q/(4\pi \varepsilon_0 r^2)$ some distance $r$ away from the charge. Let's take $Q = e$ (elementary charge). Solving for $r$:

$$r = \sqrt{\frac{1}{4\pi \varepsilon_0} \frac{Q}{E}} = \sqrt{\left(9 \times 10^9 \text{ N.m}^2/\text{C}^2\right) \left(1.6 \times 10^{-19} \text{ C}\right) \frac{38 \times 10^9 \text{ V/m}}{\text{N.m}^2/\text{C}^2}} \approx 2 \times 10^{-10} \text{ m} = 2 \text{ Å}.$$ 

Thus, in close proximity of an ion, such an electric field can be found.

Note: The units are m because we are using a formula in SI units and all our physical quantities have been provided in SI units. If you are unconvinced, you can work out the units yourself: inside the square root we have $\text{C.}(\text{N.m}^2/\text{C}^2)/(\text{V/m})$, which equals $\text{N.m}^3/(\text{C.V})$. However, 1 J=1 C.V. Thus, our units are $\text{N.m}^3/\text{J}$, whereas 1 J=1 N.m. This leaves us with $\text{m}^2$. Taking the square root gives m.

### 2.1.3. Origin of the Coulomb’s Law.

The Maxwell equations in free space are:

$$\nabla \cdot E = \frac{\rho}{\varepsilon_0}, \ \nabla \cdot B = 0, \ \nabla \times E = -\frac{\partial B}{\partial t}, \ \nabla \times B = \mu_0 \left( J + \varepsilon_0 \frac{\partial E}{\partial t} \right)$$

The various quantities are: $J$, current density (units: Amp`eres per square meter $\text{A/m}^2$); $\mu_0$, permeability of free space (value: $4\pi \times 10^{-7} \text{ V.s}/(\text{A.m})$ or N/A$^2$); $\rho$, charge density (units: Coulomb per cubic meter $\text{C/m}^3$).

The Coulomb’s law is a consequence of the first Maxwell equation, which we integrate over a volume $\text{vol}$ enclosing an electrical charge $Q$:

$$\int_{\text{vol}} \nabla \cdot E \, d^3r = \int_{\text{vol}} \frac{\rho}{\varepsilon_0} \, d^3r = \frac{Q}{\varepsilon_0}.$$ 

Here,

$$Q = \int_{\text{vol}} \rho(r) \, d^3r.$$ 

Consider the electric field of a single charged particle placed at the origin. Using the Gauss divergence theorem, we may rewrite the left hand side as a surface integral:

$$\int_{\partial\text{vol}} E \cdot ds = \frac{Q}{\varepsilon_0},$$

where $ds$ is a surface area element and $\partial\text{vol}$ denotes the boundary of the volume $\text{vol}$. For example, if $\text{vol}$ is a sphere, then $\partial\text{vol}$ is a spherical shell.

---

$^3$Let $E$ be a $C^1$ (differentiable, with continuous derivative) vector field. Let $\Omega \subset \mathbb{R}^3$ be a volume. The Gauss divergence theorem states that the volume integral can be reduced to a surface integral:

$$\int_{\text{vol}} (\nabla \cdot E) \, d^3r = \int_{\partial\text{vol}} E \cdot ds.$$ 

The surface area element $ds$ is a vector pointing outward normal to the surface at each point. Another way to write the integrand of the second integral is $E \cdot \hat{n} \, ds$, where $\hat{n}$ is the unit vector that is outward normal to the surface. $ds$ is an area element.
2.1. Electrostatics

The electric field of a single charged particle is radial and does not depend on the spherical angles $\theta, \phi$.

Let $vol$ be a sphere centered at the origin with radius $R$. Then, spherical coordinates

$$\int_{vol} f(r) \, d^3r = \int_0^\pi d\theta \int_0^{2\pi} d\phi \int_0^R df(\theta, \phi, r) \sin \theta$$

render the integral trivial and the electric field is

$$E = \frac{Q}{4\pi \epsilon_0} \frac{1}{r^2} \hat{r}.$$ 

Using the Lorentz law with $B = 0$, the force is:

$$F = qE = k \frac{qQ}{r^2} \hat{r},$$

where

$$k = \frac{1}{4\pi \epsilon_0} = 9 \cdot 10^9 \text{Nm}^2/\text{C}^2.$$ 

This relationship is called the Coulomb’s law. Since $F \propto 1/R^2$, we also say it’s an inverse square law. Work is the line integral of force over distance. Let us integrate along a line/curve starting at the point $P_i$ and ending at $P_f$. $P_i$ is usually taken to be a “reference” point at infinity whereas $P_f$ is the point of interest (where $V$ is to be evaluated). We find:

$$V(R) = - \int_{P_i}^{P_f} F \cdot dr = -k \int_{\infty}^R \frac{qQ}{r^2} \, dr = k \frac{qQ}{R}.$$ 

This is often written ($q = q_1, Q = q_2$):

$$V(R) = k \frac{q_1 q_2}{R}.$$ 

The negative sign (Eq. 2.1) describes the work we do against the electrical force to move the charge from point $P_i$ (say, at infinity) to point $P_f$; i.e., we have to apply a negative force $-F$.

Here, $R$ is the distance vector between the two ions, as indicated in Fig. 2.12. In 3D, $R \equiv |R| \equiv \sqrt{X^2 + Y^2 + Z^2}$, where $X, Y, Z$ are the components of the vector $R$.\(^5\)

\(^4\)For consistency, you can check that the force — the negative gradient of the potential energy $V$ — gives the correct result:

$$F = -\nabla V = k \frac{q_1 q_2}{R^2} \hat{r}$$

where $\hat{r} \equiv R / |R|$.
2. Intermolecular Interactions

2.1.4. Interacting Charge Distributions. Suppose now that instead of two interacting point charges, we have two non-overlapping, continuous charge distributions (\( \rho_A \) and \( \rho_B \)), as shown in Fig. 2.13 (\( \mathbf{R} = \mathbf{r}_B - \mathbf{r}_A \)).

The vectors \( \mathbf{r}_B \) and \( \mathbf{r}_A \) are center-of-mass (\( \text{com} \)) coordinates for each charge distribution. Namely, if \( \mu_A(\mathbf{r}) \) is the mass density for the charge distribution \( \rho_A \), its center of mass is defined by

\[
\mathbf{r}_A = \frac{1}{M_A} \int_{\text{vol}} \mu_A(\mathbf{r}) d^3\mathbf{r}
\]

where \( M_A \) is the total mass of \( A \), i.e. \( M_A = \int_{\text{vol}} \mu_A(\mathbf{r}) d^3\mathbf{r} \). As an example, suppose that the mass distribution\(^6\) describes a point mass \( m_1 \) at \( \mathbf{r}_1 \) and another point mass \( m_2 \) located at \( \mathbf{r}_2 \). The center of mass is then:

\[
\mathbf{r}_{\text{com}} = \frac{1}{m_1 + m_2} (m_1 \mathbf{r}_1 + m_2 \mathbf{r}_2).
\]

\(^6\)In terms of Dirac delta functions (see A.23), the mass density is \( \mu(\mathbf{r}) = m_1 \delta(\mathbf{r} - \mathbf{r}_1) + m_2 \delta(\mathbf{r} - \mathbf{r}_2) \). The total mass is \( \int_{\mathbb{R}^3} \mu(\mathbf{r}) d^3\mathbf{r} = m_1 + m_2 \). The center of mass is at \( \mathbf{r}_{\text{com}} = \frac{1}{m_1 + m_2} \int_{\mathbb{R}^3} \mu(\mathbf{r}) d^3\mathbf{r} = \frac{1}{m_1 + m_2} (m_1 \mathbf{r}_1 + m_2 \mathbf{r}_2) \).
The Coulomb energy describing the interaction of these two charge distributions is (c.f. Eq. 2.2):

\[
V_{AB}(R) = k \int_{\mathbb{R}^3} \int_{\mathbb{R}^3} \frac{\rho_A(r')\rho_B(r'')}{|r'' - r'|} d^3r' d^3r''.
\]

Now suppose that \( R \) is much larger than the radius of \( \rho_A \) and \( \rho_B \) (far field approximation). Let us rewrite the denominator \(|r'' - r'|\) as:

\[
|\overrightarrow{R} - \overrightarrow{r}' + \overrightarrow{r}''| = |\overrightarrow{R} - \overrightarrow{r}' + \overrightarrow{r}''| - \overrightarrow{R}
\]

and since \( R \gg |\overrightarrow{r}_B - \overrightarrow{r}''| \) and \( R \gg |\overrightarrow{r}_A - \overrightarrow{r}'| \) we may Taylor expand (see Section A.11) this factor about the point \( 1/R \), using the small parameter \( \epsilon = \overrightarrow{r}_A - \overrightarrow{r}' - (\overrightarrow{r}_B - \overrightarrow{r}'') \): (here, we use the shorthand notation \( \partial_\alpha \equiv \partial/\partial R_\alpha \) as well as summation over repeated indices; see Section A.10)

\[
\frac{1}{|\overrightarrow{R} + \epsilon|} = \frac{1}{R} + \epsilon_\alpha \partial_\alpha \left( \frac{1}{R} \right) + \frac{1}{2} \epsilon_\alpha \epsilon_\beta \partial_\alpha \partial_\beta \left( \frac{1}{R} \right) + O(\epsilon^3).
\]

Let us keep zeroth and first moments for now since we want to discuss only the lowest order interactions (ion-ion, ion-dipole and dipole-dipole), i.e. we keep cross terms \((r')(r'')\) and lower order \((r')\) and \((r'')\). Terms such as \((r')^2\) or \((r'')^2\) and higher order describe quadrupole, octupole and higher order moments. Substitution of \( \epsilon \) gives

\[
\frac{1}{|\overrightarrow{R} + \epsilon|} = \frac{1}{R} + \epsilon_\alpha (r_A - r' - (r_B - r'')) \partial_\alpha \left( \frac{1}{R} \right)
\]

\[
+ \frac{1}{2} \left\{ -(r_A - r')_\alpha (r_B - r'')_\beta - (r_B - r')_\alpha (r_A - r'')_\beta \right\} \partial_\alpha \partial_\beta \left( \frac{1}{R} \right) + O(\epsilon^3)
\]

where the derivatives are:

\[
\partial_\alpha \left( \frac{1}{R} \right) = -R_\alpha \frac{1}{R^3}, \quad \partial_\alpha \partial_\beta \left( \frac{1}{R} \right) = \partial_\alpha \left( -\frac{R_\beta}{R^3} \right) = \frac{3R_\alpha R_\beta}{R^5} - \delta_\alpha_\beta \frac{1}{R^3}.
\]

The function \( \delta_\alpha_\beta \) is called Kronecker delta. It is defined as 1 when \( \alpha = \beta \) and 0 when \( \alpha \neq \beta \). The Kronecker delta arose because we encountered a situation that involved computing derivatives of the type: \( \partial y/\partial x = 1 \) and \( \partial y/\partial x = 0 \), etc.

The integral

\[
V_{AB}(R) = k \int_{\mathbb{R}^3} \int_{\mathbb{R}^3} \rho_A(r')\rho_B(r'')|\overrightarrow{R} + \epsilon|^{-1} d^3r' d^3r''
\]
evaluates to:

\[
\frac{k}{R} \left[ \frac{q_A q_B}{R} - \frac{q_A p_B \cdot R}{R^3} + \frac{q_B p_A \cdot R}{R^3} + \frac{(p_A \cdot p_B - 3(p_A \cdot R)(p_B \cdot R))}{R^5} \right] + \ldots
\]

We have made the following abbreviations:

- \( q_A = \int \rho_A(r')d^3r' \), \( q_B = \int \rho_B(r'')d^3r'' \) (monopole)
- \( p_A = \int \rho_A(r')(r' - \vec{r}_A)d^3r' \), \( p_B = \int \rho_B(r'')(r'' - \vec{r}_B)d^3r'' \) (dipole)

This multipole expansion leads to a series of terms:

- \((+)(+): \text{ion ion } 1/R\)
- \((+)(+ -): \text{ion dipole } 1/R^2\)
- \((+ -)(+ -): \text{dipole dipole } 1/R^3\)
- \((+)(+ - - +): \text{ion quadrupole } 1/R^3\)
- \((+ -)(+ - - +): \text{dipole quadrupole } 1/R^4\)

For those who think we may have done this too quickly, here is a more detailed derivation. What we did is substitute Eq. (2.3) into the integral

\[
k \int_{\mathbb{R}^3} \int_{\mathbb{R}^3} \rho_A(r')\rho_B(r'') |\mathbf{R} + \mathbf{e}|^{-1}d^3r'd^3r''
\]

in place of \(|\mathbf{R} + \mathbf{e}|^{-1}\). The first term in Eq. (2.3) is \(1/R\). Its contribution to the integral is:

\[
k \int_{\mathbb{R}^3} \int_{\mathbb{R}^3} \frac{1}{R} \rho_A(r')\rho_B(r'') \frac{d^3r'd^3r''}{q_A} = \frac{k}{R} \int_{\mathbb{R}^3} \rho_A(r')d^3r' \cdot \int_{\mathbb{R}^3} \rho_B(r'')d^3r'' = \frac{kq_A q_B}{R}.
\]

which is the term we sought. Next, according to Eq. (2.3) the second term in the expansion of \(|\mathbf{R} + \mathbf{e}|^{-1}\) is

\[
\left(\vec{r}_A - r' - (\vec{r}_B - r'')\right)_\alpha \partial_\alpha \left(\frac{1}{R}\right)
\]

where \(\partial_\alpha \left(\frac{1}{R}\right) = -R_\alpha/R^3\). Plugging this term into the integral gives

\[
-k \int_{\mathbb{R}^3} \int_{\mathbb{R}^3} \rho_A(r')\rho_B(r'')(\vec{r}_A - r' - (\vec{r}_B - r'')) \alpha \left(\frac{R_\alpha}{R^3}\right) d^3r'd^3r''.
\]

Next, we pull the term \(R_\alpha/R^3\) outside the integral, since it does not depend on the two variables of integration. We absorb the minus sign by reversing the order of \(\vec{r}_A - r' \rightarrow r' - \vec{r}_A\) and \(\vec{r}_B - r'' \rightarrow r'' - \vec{r}_B\). Finally, we split the
2.2. Intermolecular Interactions and Their Ranges

Interactions between electrostatic charges are described by the Coulomb’s law. So far we have considered electric charge distributions and interactions between charge distributions. The charge distributions can be molecules. Thus, the concepts introduced in the previous section can be applied to describe interactions between pairs of molecules. We often have that the molecular radius is small compared to $R$. For a dipole, this gives rise to the so-called “point dipole approximation; the separation between the two charges of the dipole is much smaller than $R$.
2. Intermolecular Interactions

Molecules consist of many electrical charges. The spatial arrangement of charges on a molecule may include monopole (total charge), dipole moment, quadrupole moment, etc. Some molecules can dissociate into ions, in which case the ions have nonzero charge. Even if the total charge of a molecule is zero, there can be dipole, quadrupole or higher-order multipole moments, depending on how charges are spatially arranged. These spatial arrangements interact among each other, as described in the previous section by the multipole expansion.

The range of interaction is determined by how rapidly the interaction energy falls off as function of the intermolecular separation $R$. Slowly decreasing dependences of the form $1/R$, $1/R^2$ and $1/R^3$ are considered “long-range interactions” because they extend far in space, much farther than the molecular dimensions. $1/R^4$, $1/R^5$ and $1/R^6$ etc. are considered “short-range interactions” because they fall off rapidly with $R$.

2.2.1. Ion-Ion (Charge-Charge) Interaction. In Section 2.1.4 the lowest-order term describes the interaction of net charges $q_1,q_2$ (of molecules “1” and “2”) according to Coulomb’s law, $V(R) = k q_1 q_2 / R$, where $k \approx 9 \times 10^9 \text{ N.m}^2/\text{C}^2$. The interaction is isotropic (spherically symmetric), as there is no angle dependence, only a radial ($R$) dependence.

Example 2.4. NaCl is made of two ions: for Na$^+$ and Cl$^-$ we take $q_1 = -e$ and $q_2 = e$, with $e = 1.6 \times 10^{-19}$ C. In the gas phase the interatomic separation is 0.236 nm (2.36 Å). The interaction energy is then:

$$V(R) = k \frac{q_1 q_2}{R} = -(9 \times 10^9 \text{ N.m}^2/\text{C}^2) \frac{(1.6 \times 10^{-19} \text{ C})^2}{0.236 \times 10^{-9} \text{ m}} \approx -10^{-18} \text{ J}.$$ 

Since 1 eV = $1.6\times10^{-19}$ J, the magnitude $10^{-18}$ J $\sim$ 6.1 eV. This “binding energy” can be viewed as the energy required to break this molecule apart.

The Coulomb interaction is said to be a long-range force because it falls off slowly with distance ($\sim 1/R$). Ion-ion forces are relevant to electrolyte solutions. Its magnitude can be as strong as a covalent bond. We note that the Coulomb force vector, $\mathbf{F}$, is directed along the axis of the two ions; so the force does not depend on molecular orientation.

To have ionic interactions, we need a metal attached to a non-metal. Let’s look at the example of NaCl. Sodium is the metal. When NaCl is dissolved in solution, we get the ions Na$^+$ and Cl$^-$. Here, the metal is sodium, and the non-metal is chlorine (Fig. 2.14). The structure of NaCl crystal is shown in Fig. 2.15.
2.3. Charge-Dipole Interaction

In Section 2.1.4, the next order term is the ion-dipole interaction, which depends on the angle between the dipole and the ion-dipole vector. When atoms in a molecule share electrons unequally, they create what is called a dipole moment. This occurs when one atom is more electronegative than another, resulting in that atom pulling more tightly on the shared pair of electrons, or when one atom has a lone pair of electrons and the difference of electronegativity vector points in the same way. One of the most common examples is the water molecule (Fig. 2.11), made up of one oxygen atom and two hydrogen atoms. The differences in electronegativity and lone electrons give oxygen a partial negative charge and each hydrogen a partial positive charge.

The interaction between a polar solvent molecule, such as water, and a dissolved ion is the most common case of ion-dipole interaction. This interaction energy scales as $1/R^2$, where $R$ is the distance between the ion and
the dipole (with the dipole assumed to be a “point dipole”). Figure 2.16 shows dissolved Na and Cl ions interacting with water dipoles. Positive ions (Na\(^+\)) are attracted by the negative end of the dipole (the oxygen side of the H\(_2\)O molecule) and repelled by the positive end (H atoms); thus, the (Na\(^+\)) ion is surrounded by a shell of water molecules whose oxygen (O) atoms are near the cation and whose hydrogen (H) atoms point outward into the solution (away from the cation). Likewise, the anion (here, Cl\(^-\)) attracts the hydrogen atoms of the H\(_2\)O molecules, which are more positively charged than the oxygen atoms. The H\(_2\)O molecules are oriented around the anion in such a way that their negative ends (here, the O atoms of H\(_2\)O) point away from the anion, and into the solution.

![Figure 2.16. Ion-dipole interaction between Na\(^+\) and Cl\(^-\) ions and surrounding water molecules.](image)

**Water Interacts With Dissolved Ions:** The water molecules have dipole moments; thus, the oxygen (O) atoms bear small, negative charges, whereas the hydrogen (H) atoms bear small, positive charges. (a) Positive ions are attracted to neighboring water molecules in aqueous solution by ion-dipole forces. (b) Negative ions form hydrogen bonds with water, with a nearly linear bond from O to H to the anion. (See Fig. 2.16.)

### 2.3.1. Electrostatic Origin

This interaction was derived in Section 2.1.4. Let us look at it from a different point of view. Consider a charge \(Q\) interacting with an electric dipole \(\mathbf{p} = q\mathbf{d}\) as shown in Fig. 2.17. The interaction energy is obtained from the Coulomb’s law: (self-interaction energy of the

\(^7\)“Point dipole” in the present context means that the distance between the ion and any of the dipole’s charges is much larger than the length of the dipole itself (charge separation).
2.3. Charge-Dipole Interaction

\[ V(R) = -kqQ \left( \frac{1}{AB} - \frac{1}{AC} \right), \]

where (in the last step we make the \textit{far-field} approximation \( R \gg d \), such that only the first term inside the square root is kept)

\[ AB = \sqrt{(R - \frac{1}{2}d \cos \theta)^2 + \left( \frac{1}{2}d \sin \theta \right)^2} \approx R - \frac{1}{2}d \cos \theta, \]

\[ AC = \sqrt{(R + \frac{1}{2}d \cos \theta)^2 + \left( \frac{1}{2}d \sin \theta \right)^2} \approx R + \frac{1}{2}d \cos \theta. \]

Insert this into the expression for \( V \) and get:

\[ V(R) = -kqQ \left( \frac{1}{R - \frac{1}{2}d \cos \theta} - \frac{1}{R + \frac{1}{2}d \cos \theta} \right) \]

\[ = -kqQ \left( \frac{d \cos \theta}{R^2 - \frac{1}{4}d^2 \cos^2 \theta} \right) \approx -\frac{kqQd \cos \theta}{R^2}. \]

Thus, the charge-dipole energy is \( V(R) = -\frac{kQp \cos \theta}{R^2} \), which drops off as \( R^{-2} \) (inverse square law). Here, \( \cos \theta = \hat{R} \cdot \hat{p} \equiv \hat{R} \cdot \hat{p}/(||R||p) \).

**Example 2.5.** Suppose we have an aqueous solution with dissolved NaCl. What is the interaction energy between Na\(^+\) ions and water molecules? **Solutions:** This is an ion-dipole interaction. The Na\(^+\) and Cl\(^-\) ions are surrounded by water molecules. For sodium, complexes of the type \([\text{Na(H}_2\text{O)}\_8]^+\)

---

8Notice that we have not included a term for the Coulomb interaction of the dipole’s charges \(+q\) and \(-q\). This interaction is called the \textit{self-energy} (of the dipole). Its inclusion causes conceptual difficulties, especially in the limit of a point dipole, where the self-energy causes singularities. We often sweep the self-energy under the rug and ignore it. For molecules, we are content with saying that they have finite size and the idealization of a point dipole is not needed. In this case, the self-energy contributes a constant term independent of \( R \). (The zero of the energy is not important.)
form (with 8 water molecules bound to \( \text{Na}^+ \) ions, on average), where the average \( \text{Na}-\text{O} \) distance is approximately 250 pm. According to Table 2.1 the permanent EDM of \( \text{H}_2\text{O} \) molecules is 1.85 D, where 1 D = \( 3.336 \times 10^{-30} \) C.m.

Using the formula, 
\[
V(R) = -\frac{kQ_Q\cos \theta}{R^2},
\]
we find that the interaction energy for each \( \text{Na}-\text{O} \) bond is:
\[
V(R) = -(9 \times 10^9 \text{ N.m}^2) \frac{(1.6 \times 10^{-19} \text{ C})(1.85 \times 3.336 \times 10^{-30} \text{ C.m})}{(250 \times 10^{-12} \text{ m})^2} \cos \theta
\]
\[
= -(1.42 \times 10^{-19} \text{ J}) \cos \theta
\]
and since 1 eV = \( 1.6 \times 10^{-19} \) J, this energy is \( \sim 0.9 \) eV, about 7 times weaker than the interaction energy of \( \text{NaCl} \). The factor \( \cos \theta \) has to do with orientation of the dipole relative to the ion-dipole vector. In solution, molecules tumble and this factor is averaged by rotations. For fixed orientation, and referring to Figure 2.17, we see that \( \theta = 0 \) (dipole pointing away from the positive charge) has the lowest possible energy because the negative end of the dipole is closest to the \( \text{Na}^+ \) ion. When \( \theta = \pi \), \( \cos \pi = -1 \) and we have a state of high energy because \( \text{Na}^+ \) repels the positive end of the dipole. When \( \theta = \pi/2 \) or \( 3\pi/2 \), \( \cos \theta = 0 \) and the interaction energy is zero (half-way point between the high energy and low energy states).

2.4. Dipole-Dipole Interaction

The dipole-dipole interaction can be thought of as describing the electrostatic coupling between the charges making up the dipoles. The dipole-dipole interaction falls off as \( \sim 1/R^3 \), which means that the force one dipole 1 due to dipole 2 falls off as \( 1/R^4 \). This interaction energy is shorter range than the Coulomb interaction, but is still considered long-range.\(^9\) The dipole-dipole interaction depends on the relative orientation of the two dipoles.

2.4.1. Electrostatic Origin.\(^9\) This interaction was derived in Section 2.1.4. Let us look at a slightly different method of derivation. Consider two dipoles 
\[
\mathbf{p}_A = q_A \mathbf{d}_A, \quad \mathbf{p}_B = q_B \mathbf{d}_B
\]
interacting electrostatically, as shown in Fig. 2.18.

The interaction energy is obtained from Coulomb’s law applied to all the charge pairs in the system \( \mathbf{R} = \mathbf{r}_B - \mathbf{r}_A \):
\[
V(R) = \frac{k(-q_A)(-q_B)}{|\mathbf{R}|} + \frac{k(-q_A)q_B}{|\mathbf{R} + \mathbf{d}_B|} + \frac{kq_A(-q_B)}{|\mathbf{R} - \mathbf{d}_A|} + \frac{kq_Aq_B}{|\mathbf{R} + \mathbf{d}_B - \mathbf{d}_A|}.
\]

\(^9\)Long range means that when integrated over all space, the integral of \( V(R) \) is finite and doesn’t depend on the volume (which we denote here as \( \text{vol} \)). \( 1/r^3 \) is long range because the integral \( \int_{\text{vol}} (1/r^3) d^3r \) in spherical coordinates is \( (4\pi) \int_{\text{vol}} (1/r^3) r^2 dr \sim \int (1/r) dr \sim \log(\text{vol}) \), which depends on volume (\( \text{vol} \)). Integrals whose values depend on the volume of the region of integration are called conditionally convergent.
2.4. Dipole-Dipole Interaction

We then take the limit

\[ |\mathbf{R}| \equiv |\mathbf{r}_A - \mathbf{r}_B| \gg |\mathbf{d}_A|, |\mathbf{d}_B|, \]

i.e., so we can neglect terms that are higher order in \( |\mathbf{d}_A|/|\mathbf{R}| \) and \( |\mathbf{d}_A|/|\mathbf{R}| \), and expand the square roots as

\[ (1 + x)^{-1/2} \approx 1 - \frac{x}{2} + \frac{3x^2}{8} + \ldots. \]

After some algebra, we get the famous result: (left as a homework exercise)

\[ V(\mathbf{R}) = k \left\{ \frac{\mathbf{p}_A \cdot \mathbf{p}_B}{\mathbf{R}^3} - \frac{3(\mathbf{p}_A \cdot \mathbf{R})(\mathbf{p}_B \cdot \mathbf{R})}{\mathbf{R}^5} \right\}. \]

**Figure 2.18.** Two EDMs coupled electrostatically.

**Figure 2.19.** The relative orientation of two dipoles can be described using 3 angles: \( \theta_1, \theta_2 \) and \( \phi \). Both \( \theta_1 \) and \( \theta_2 \) are measured in-plane, whereas \( \phi \) is measured out-of-plane.

---

10The first term \( \frac{1}{|\mathbf{r}_A - \mathbf{r}_B|} = \frac{1}{|\mathbf{R}|} = \frac{1}{R} \) does not need to be expanded. The remaining three terms are the ones to be expanded. For example, the second term becomes:

\[ \frac{1}{|\mathbf{R} + \mathbf{d}_B|} = \frac{1}{|\mathbf{R} + \mathbf{d}_B|} = \frac{1}{\sqrt{\mathbf{R}^2 + \mathbf{d}_B \cdot \mathbf{d}_B + \mathbf{d}_B^2}} = \frac{1}{R} \frac{1}{\sqrt{1 + 2\mathbf{R} \cdot \mathbf{d}_B/\mathbf{R}^2 + \mathbf{d}_B^2/\mathbf{R}^2}} \]

\[ \approx \frac{1}{R} \left[ 1 - \frac{(2\mathbf{R} \cdot \mathbf{d}_B/\mathbf{R}^2 + \mathbf{d}_B^2/\mathbf{R}^2)/2 + 3(2\mathbf{R} \cdot \mathbf{d}_B/\mathbf{R}^2 + \mathbf{d}_B^2/\mathbf{R}^2)^2/8 + \ldots}{2} \right] \]

where in the last line we have used the expansion \( (1 + x)^{-1/2} \approx 1 - \frac{x}{2} + \frac{3x^2}{8} + \ldots. \)
2.4.2. Spherical Coordinates. Equation (2.4) describes two dipoles \( p_A \) and \( p_B \) interacting according to the geometry shown in Fig. 2.19. Spherical coordinates will be useful later when we calculate angle-averaged interactions (see Fig. 2.20).

Using the identity:

\[
\cos \gamma = \cos \theta_1 \cos \theta_2 + \sin \theta_1 \sin \theta_2 \cos (\varphi_1 - \varphi_2)
\]

\( V \) becomes:

\[
V(R) = k \left\{ \frac{p_A p_B}{R^3} \left[ -2 \cos \theta_1 \cos \theta_2 + \sin \theta_1 \sin \theta_2 \cos \phi \right] \right\}.
\]

This follows from writing the first term as:

\[
(2.5) \quad \frac{p_A \cdot p_B}{R^3} = \frac{p_A p_B \cos \gamma}{R^3} = \frac{p_A p_B}{R^3} \left\{ \cos \theta_1 \cos \theta_2 + \sin \theta_1 \sin \theta_2 \cos \phi \right\}
\]

and the second term as

\[
(2.6) \quad \frac{(p_A \cdot \hat{R})(p_B \cdot \hat{R})}{R^3} = \frac{(p_A \cdot \hat{\mathbf{r}})(p_B \cdot \hat{\mathbf{r}})}{R^3} = \frac{p_A p_B \cos \theta_1 \cos \theta_2}{R^3}
\]

where \( \hat{\mathbf{r}} = \mathbf{R}/|\mathbf{R}| \), and collecting the terms.

**Example 2.6.** Suppose we have two water molecules separated 3 Å apart and interacting (only) via the dipole-dipole interaction. What is the magnitude of this interaction? **Solution:** According to Table 2.1 the permanent EDM of H\(_2\)O molecules is 1.85 D, where 1 D=3.336 \( \times 10^{-30} \) C.m. The magnitude of the dipole-dipole interaction\(^{11}\) is on the order of:

\[
V(R) = k \frac{p_A p_B}{R^3} = (9 \times 10^9 \text{ N.m}^2/\text{C}^2) \left( \frac{1.85 \times 3.336 \times 10^{-30} \text{ C.m}}{5 \times 10^{-10} \text{ m}^3} \right)^2
\]

\[
\approx 1.3 \times 10^{-20} \text{ J} \approx 0.08 \text{ eV}
\]

\(^{11}\)This formula is exact when \( \hat{p}_A \) and \( \hat{p}_B \) are parallel to each other (setting \( \gamma = 0 \) in Eq. 2.5), and when \( \hat{p}_A \) and \( \hat{p}_B \) are both perpendicular to \( \hat{R} \) (setting \( \theta_1 = \theta_2 = 0 \) in Eq. 2.6).
It is known from experiments that the hydrogen bond energy in water is 23.3 kJ/mol. The conversion to eV goes like this:

\[
(23.3 \times 10^3 \text{ J mol}^{-1}) \left( \frac{1 \text{ eV}}{1.6 \times 10^{-19} \text{ J}} \right) \left( \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \right) = 0.242 \text{ eV},
\]

i.e. the conversion factor is 1 kJ/mol = 0.01037 eV. The dipole-dipole interaction only partially accounts for the interaction energy between pairs of water molecules. Covalent character is also present in the hydrogen bond.

2.5. Hydrogen Bonding

Hydrogen bonding is commonly found in water, between pairs of water molecules (Fig. 2.21), where the H atom bonds to the O atom of a neighboring molecule. The bond is partly ionic, due to the partial positive charge of the H atom and partial negative charge of the O atom, and partly covalent, due to the sharing of electrons from the oxygen’s lone pair.

![Hydrogen bonding between two water molecules.](image)

Figure 2.21. Hydrogen bonding between two water molecules.

Hydrogen (H) bonds can be thought of as especially strong dipole-dipole interactions involving the strongest electronegative elements of the periodic table. A hydrogen bond (H-bond) is the interaction between two polar groups that occurs when a hydrogen (H) atom covalently bound to a small, highly electronegative atom such as nitrogen (N), oxygen (O), or fluorine (F) experiences the electric field of another highly electronegative atom nearby. (Remember: FON.)

The molecule itself must possess either a N-H, N-O or N-F bond for this to work. As an example, H$_2$O can undergo hydrogen bonding because it has H-O bonds. The H-bond itself will occur between pairs of neighboring H$_2$O molecules. On the other hand, CH$_3$F does not H-bond because there are no H-F bonds (CH$_3$F only possesses C-F and H-C bonds).

The H-bond is strong because the dipole is strong. The dipole is strong due to the large difference in electronegativity; the H-atom does not possess any core electrons to shield its positive charge.

Hydrogen bonding is typically stronger than VDW interactions, but weaker than covalent or ionic bonds. H-bonds can be intramolecular (within the same molecule) or intermolecular (between different molecules). Intermolecular hydrogen bonding is responsible for the high boiling point of water.
(100°C) compared to the other group 16 hydrides that have much weaker hydrogen bonds. Intramolecular hydrogen bonding is partly responsible for the secondary and tertiary structures of proteins and nucleic acids. It also plays an important role in the structure of polymers, both synthetic and natural.

![Figure 2.22. Examples of hydrogen bonding.](image)

![Figure 2.23. Effect of hydrogen bonding on boiling point. Group 15 (orange), 16 (blue) and 17 (purple) elements are compared.](image)
The examples of NH$_3$, H$_2$O and HF are shown in Fig. 2.22. Hydrogen bonding can also occur between different types of molecules, as shown in the last two lines in Fig. 2.22 (the case of intermolecular NH and HO bonding). Finally, we note that the boiling point is strongly affected by hydrogen bonding, as shown in Fig. 2.23. If H$_2$O did not possess the ability to form hydrogen bonds, its boiling point would likely be below -50°C.

2.6. Ion-Induced Dipole

Ion-dipole interactions can also be found in non-polar solvents. We call these “ion-induced dipole” or “charge-induced dipole”. In that case, it is the ion’s electric field that induces a dipole moment in neighboring solvent molecules (electric polarization was explained in Section 2.1.2). This situation is illustrated in Fig. 2.24. We note that molecules can be polarized regardless of whether or not they already possess an EDM. In the case where the molecule has an EDM, the polarization would be an additional (albeit smaller) effect. So the ion-induced dipole interaction is present regardless of the existence of a permanent EDM.

Figure 2.24. Ion-induced dipole interaction.

When a molecule of polarizability $\alpha$ is at distance $R$ from an ion of charge $ze$, the electric field of the ion $E = ze/4\pi\epsilon_0 R^2$ (electric charge is often written as a multiple of the elementary charge: $q = ze$) will induce in the molecule a dipole moment of

$$p = \alpha E = \frac{\alpha ze}{4\pi\epsilon_0 R^2}$$

Here, $\alpha$ is given in units of $(4\pi\epsilon_0)\text{Å}^3 = (4\pi\epsilon_0)10^{-30} \text{ m}^3 = 1.11\times10^{-40} \text{ C}^2\text{m}^2\text{J}^{-1} = 1.11\times10^{-40} \text{ C}\text{m}^2/V$ since 1 Joule = 1 Coulomb times Volt (1 J=1 C.V). (See Table 2.5.)

Let us look at the magnitude of this interaction. We shall estimate the distance by which the electron cloud of a methane (CH$_4$) molecule is shifted relative to the center of the molecule due to the presence of a bare sodium ion whose center is 0.4 nm from the center of the molecule.
For a monovalent ion such as Na\(^+\), the electric field at a distance of 0.4 nm from its center is
\[
E = \frac{e}{4\pi\epsilon_0 R^2} = \frac{(1.602 \times 10^{-19})}{(4 \times 3.142 \times 8.854 \times 10^{-12})(0.4 \times 10^{-9})^2} = 9.0 \times 10^9 \text{ V/m}
\]
The induced dipole moment on a methane molecule (see Table 2.5 for polarizability of CH\(_4\)) is therefore,
\[
p = \alpha E = 4\pi\epsilon_0 (2.6 \times 10^{-30})(9.0 \times 10^9) = 2.60 \times 10^{-30} \text{ C.m}
\]
Table 2.5
\[
= 2.60 \times 10^{-30}/3.336 \times 10^{-30} = 0.78 \text{ D}.
\]
Since a dipole moment is charge\(\times\)distance, this corresponds to a unit charge separation in the molecule of \(d = p/e = 0.016\) nm, which is about 8% of the molecular radius of methane (of 0.2 nm).

From this calculation, the induced dipole moment can be on the order of 1 Debye, which is quite large. We may therefore anticipate that the forces associated with induced dipole moments can also be large.

The scenario is the following. An ion is adjacent to a molecule, including a dipole moment \(p\). The electric field of the induced dipole (which will act on the ion) is\(^{12}\)
\[
E(R) = \frac{-2p}{4\pi\epsilon_0 R^3} = -\frac{2\alpha E}{4\pi\epsilon_0 R^3} = -\frac{2\alpha(ze)}{(4\pi\epsilon_0)^2 R^5}
\]
so that the attractive force and energy are:
\[
F(R) = (ze)E(R) = -\frac{2\alpha(ze)^2}{(4\pi\epsilon_0)^2 R^5},
\]
\[
V(R) = -\int_{R}^{\infty} F(r)dr = -\frac{\alpha(ze)^2}{2(4\pi\epsilon_0)^2 R^4} = -\frac{1}{2} \alpha E^2.
\]
where \(E(R) = ze/2\pi\epsilon_0 R^2\) is the field acting on the molecule. Thus,
\[
V(R) = \boxed{-\frac{\alpha(ze)^2}{2(4\pi\epsilon_0)^2 R^4}}
\]
The situation is illustrated in Fig. 2.25.

2.7. Ion-Quadrupole Interaction: Cation-\(\pi\)

An example of the ion-quadrupole interaction (\(V(R) \sim 1/R^3\)) is the cation-\(\pi\) interaction, which involves an electron-rich \(\pi\) system (e.g. benzene, ethylene, acetylene) and an adjacent cation (e.g. Li\(^+\), Na\(^+\)). This interaction can be as strong as hydrogen bonding and salt bridges in proteins. Cation-\(\pi\)

\(^{12}\)The electric field of a dipole is \(E = \frac{1}{4\pi\epsilon_0} \left(\frac{3\hat{p} \cdot \hat{r}}{R^3} \right)\).
2.8. Quadrupole-Quadrupole: \(\pi\)-Stacking

interactions play important roles in protein structure, molecular recognition and enzyme catalysis.

Take the example of benzene. It has no permanent dipole moment because all dipoles (CH bonds) are weak and their interactions vanish due to symmetry. However, the \(\pi\) electrons above and below the benzene ring create a negative charge, which counterbalances the positive charge in the plane of the benzene ring. This leads to a quadrupole moment. The negative \(\pi\) electrons interact with the cation nearby, leading to the cation-\(\pi\) interaction, which is a type of ion-quadrupole interaction. See Fig. 2.26.

```
\[\text{Figure 2.25.} \quad \text{Ion-induced dipole interaction.}\]
```

\[\text{Figure 2.26.} \quad \text{Ion-quadrupole interaction: benzene interacting with Na}^+ \text{ion (left). Benzene possesses an electric quadrupole moment (right).}\]

2.8. Quadrupole-Quadrupole: \(\pi\)-Stacking

\(\pi\)-stacking is an example of quadrupole-quadrupole interaction. The benzene dimer is capable of quadrupole-quadrupole interaction because each benzene ring contains \(\pi\) electron clouds. These interactions are important in nucleobase stacking within DNA and RNA molecules, protein folding and molecular recognition. Three different conformations of the benzene dimer are shown in Fig. 2.27.
2.9. Van Der Waals Force

Van Der Waals (VDW) interactions, which behave as \( V(R) \sim 1/R^6 \), describe “dynamical effects” arising from quantum fluctuations, induced dipoles and molecular tumbling. \( 1/R^6 \) interactions include 3 contributions:

- **London dispersion forces**: Of quantum mechanical origin, they are a type of “induced dipole-induced dipole” interaction whereby dynamic motions of charge distributions around the molecule lead to a temporary dipole moment. Such a temporary dipole on one molecule will induce a temporary dipole in the other molecule. Example (He): For a non-polar atom such as helium, the time average of its dipole moment is zero, but at any instant there exists a finite dipole moment given by the instantaneous positions of the electrons about the nucleus. This instantaneous dipole generates an electric field that polarizes any nearby neutral atom, inducing a dipole moment in it. The resulting interaction between the two dipoles gives rise to an instantaneous attractive force between the two atoms, and the time average of this force is finite. The overall dependence is \( 1/R^6 \):

  \[
  V(R) = \frac{3}{2} \frac{\alpha_1 \alpha_2}{(4 \pi \epsilon_0)^2 R^6} \frac{I_1 I_2}{(I_1 + I_2)}
  \]

  where \( \alpha_1 \) and \( \alpha_2 \) are the polarizabilities of the two molecules, \( I_1 \) and \( I_2 \) are the first ionization potentials of the molecules,

- **Keesom force**: When molecules have permanent dipole moments and tumble, the Keesom force (derived in Section 2.10) is defined as the motionally-averaged dipole-dipole interaction. While static dipole-dipole interactions behave as \( V(R) \sim 1/R^3 \), the motionally-averaged interaction is weaker and \( V(R) \sim 1/R^6 \). It is the angular dependent part of the dipole-dipole interaction that causes a weakening of the dipole-dipole interaction.

- **Debye force**: A molecule with a permanent dipole moment (such as HCl or H\(_2\)O) adjacent to a non-polar molecule can induce a dipole moment in the non-polar molecule. These two dipoles (permanent dipole-induced dipole) interact together, giving rise to the Debye force (see Section 2.12).
The strength of the induced dipole depends on the polarizability of the non-polar molecule.

By far, London dispersion forces comprise the largest and most important contribution to VDW forces. Because of this, VDW forces are often referred to as “London dispersion forces” interchangeably. In reality, Keesom and Debye forces can also contribute to the $1/R^6$ dependence.\(^{13}\)

Approximate contributions of the dispersion to the total intermolecular interaction energy are shown in Table 2.6.

<table>
<thead>
<tr>
<th>Molecule pair</th>
<th>% of the total energy of interaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ne-Ne</td>
<td>100</td>
</tr>
<tr>
<td>CH(_4)-CH(_4)</td>
<td>100</td>
</tr>
<tr>
<td>HCl-HCl</td>
<td>86</td>
</tr>
<tr>
<td>HBr-HBr</td>
<td>96</td>
</tr>
<tr>
<td>HI-HI</td>
<td>99</td>
</tr>
<tr>
<td>CH(_3)Cl-CH(_3)Cl</td>
<td>68</td>
</tr>
<tr>
<td>NH(_3)-NH(_3)</td>
<td>57</td>
</tr>
<tr>
<td>H(_2)O-H(_2)O</td>
<td>24</td>
</tr>
<tr>
<td>HCl-HI</td>
<td>96</td>
</tr>
<tr>
<td>H(_2)O-CH(_4)</td>
<td>87</td>
</tr>
</tbody>
</table>


### 2.9.1. Factors Affecting the Strength of London Dispersion Forces.

Dispersion forces are a mutual interaction between the polarizable charge distributions on two separate molecules, and they are always attractive. Very often, dynamic motions of charge around the molecule lead to a temporary dipole moment. Such a temporary dipole on one molecule will induce a temporary dipole in the other molecule. These transient, fluctuating dipoles attract one another in much the same way as do permanent dipoles.

Figure 2.28 provides a simple view of the source of this interaction. The polarizability increases with the number of electrons in the atom or molecule. Heavier atoms or molecules interact more strongly by dispersion forces than do lighter ones. These electrons are less strongly bound than those of the lighter elements, because they are shielded from the full attraction of the nucleus by intervening electrons. Consequently, they are more easily distorted by external fields of neighboring dipoles. Dispersion forces provide the attractive, $1/R^6$, term in the Lennard-Jones potential.

\(^{13}\)Keesom and Debye interactions are not always present because they require permanent dipoles and sufficient polarizability (in the case of Debye). Not every molecule meet such criteria.
2. Intermolecular Interactions

A fluctuation of the electron distribution on one atom induces a corresponding temporary dipole moment on a neighboring atom. The two dipole moments interact to give a net attractive force, called a “dispersion force.”

There are three main factors affecting the strength of London dispersion forces:

- **Contact area**: the closer we can bring two molecules together, get their electron clouds to touch, the stronger the dispersion force will be. For example, \( n \)-pentane has stronger dispersion forces than neopentane (Fig. 2.29), even though they have the same number and types of atoms. The \( n \)-pentane is shaped like a column (as opposed to spheres, as is the case for neopentane) and makes better contact with other like molecules. Thus, the boiling point of \( n \)-pentane will be higher than neopentane.

\[
\begin{align*}
\text{n-pentane} & \quad \text{vs.} \quad \text{neopentane} \\
\text{BP} & = 36.1^\circ C \\
\text{BP} & = 9.5^\circ C
\end{align*}
\]

- **Polarizability**: the more electrons we have, the stronger the polarizability of the molecule. Thus, molecular size matters. Think about proteins (very large molecules). Much of their 3D shapes originate from London dispersion forces. Another example is halogens going down the rows: F, Cl, Br, I. At ambient temperatures, F and Cl are gases, Br is a liquid and Iodine is a solid. As the atoms get bigger, they have more electrons, and the stronger the dispersion forces are.

**Example**: Silk Fibroin. Fibroin is the main protein in silk from moths and spiders. It consists of antiparallel \( \beta \)-sheet structure (Fig. 2.30). It...
has small side chains (Ala and Gly) that allow the close packing of sheets. Its structure is stabilized by: 1) hydrogen bonding within sheets and 2) London dispersion interactions between the sheets:

![Figure 2.30. Structure of silk fibroin.](image)

Note: for proteins, several interactions matter. These include: 1) the hydrophobic effect, 2) hydrogen bonding (peptide bond N-H and C=O interact to yield $\alpha$-helices and $\beta$-sheets), 3) London dispersion and 4) electrostatic interactions, such as interactions between permanently charged groups, e.g., salt bridges stabilize the protein.

- Pi-bonding: molecules with $\pi$-bonding will exhibit stronger London dispersion forces (Fig. 2.31).

![Figure 2.31. $\pi$-bonding.](image)

### 2.10. Thermal Averaging of Intermolecular Interactions

When the interaction energy falls below $k_B T$ the dipoles can rotate more or less freely. We are then interested in the Boltzmann average of these
interactions over the angles. At room temperature we have

\[ k_B T \approx 0.03 \text{ eV} \sim 4 \cdot 10^{-21} \text{ J}. \]

We define the thermal average energy by angle-averaging the Boltzmann factor corresponding to some interaction energy \( V(R, \Omega) \) (Ref. [2]):

\[ e^{-V(R)/k_B T} \equiv \langle e^{-V(R,\Omega)/k_B T} \rangle = \frac{\int e^{-V(R,\Omega)/k_B T} d\Omega}{\int d\Omega} \]

In spherical coordinates the solid angle element is \( d\Omega = \sin \theta d\theta d\phi \), where \( \theta, \phi \) are polar and azimuthal angles, respectively. Thus,

\[ e^{-V(R)/k_B T} \equiv \langle e^{-V(R,\Omega)/k_B T} \rangle = \frac{1}{4\pi} \int_0^{2\pi} d\phi \int_0^\pi e^{-V(R,\theta,\phi)/k_B T} \sin \theta d\theta \]

When calculating such integrals the following results are useful\(^{14}\):

\[
\langle \cos^2 \theta \rangle = \frac{1}{4\pi} \int_0^\pi \cos^2 \theta \sin \theta d\theta \int_0^{2\pi} d\phi = \frac{1}{2} \int_{-1}^1 x^2 dx = \frac{1}{3}.
\]

\[
\langle \sin^2 \theta \rangle = \frac{1}{4\pi} \int_0^\pi \sin^2 \theta \sin \theta d\theta \int_0^{2\pi} d\phi = \frac{2}{3}
\]

\[
\langle \sin^2 \phi \rangle = \langle \cos^2 \phi \rangle = \frac{1}{2}
\]

\[
\langle \sin \theta \rangle = \langle \cos \theta \rangle = \langle \sin \theta \cos \theta \rangle = 0
\]

\[
\langle \sin \phi \rangle = \langle \cos \phi \rangle = \langle \sin \phi \cos \phi \rangle = 0
\]

Averaging over the angles implies that the system (e.g. charge-dipole or dipole-dipole) samples all possible angles over some time period\(^{15}\).

When \( V \) is less than \( k_B T \) we may Taylor expand the exponential

\[ e^{-V(R)/k_B T} = 1 - \frac{V(R)}{k_B T} + \cdots = \left( 1 - \frac{V(R,\Omega)}{k_B T} \right) + \frac{1}{2} \left( \frac{V(R,\Omega)}{k_B T} \right)^2 - \cdots \]

\(^{14}\)These are obtained using trigonometric identities or tables of integration. For example, the second relation follows from

\[
\int \sin^3(ax) dx = \frac{\cos 3ax}{12a} - \frac{3\cos ax}{4a} + C.
\]

\(^{15}\)Instead of computing the time-average of some quantity, we invoke the idea of the ensemble average through the ergodic theorem

\[ \frac{1}{T} \int_0^T \langle \cdot \rangle dt \approx \frac{1}{\text{vol}} \int \langle \cdot \rangle d^3r, \]

where \( \text{vol} = \text{volume} \) (of some region of space comprising the ensemble of molecules). The left hand side is the time average of some quantity and the right hand side is the average of that quantity across an ensemble of molecules. The validity of the ergodic theorem is discussed in statistical mechanics and will not concern us here except to say that the ensemble average is used here by invoking the Boltzmann distribution.
Thus,
\[ V(R) = \left\langle V(R, \Omega) - \frac{V(R, \Omega)^2}{2k_B T} + \ldots \right\rangle \]

Example: 1) Charge-dipole interaction: we recall the geometric picture for this interaction (Fig. 2.17) and the expression for its interaction energy
\[ V(R) = -\frac{q}{4\pi \epsilon_0 R^2} \cos \theta. \]

Performing the thermal averaging we find:
\[ V(R) = \left\langle -\frac{q}{4\pi \epsilon_0 R^2} \cos \theta \right\rangle = -\frac{q^2 p^2}{6(4\pi \epsilon_0)^2 k_B T R^4} \sim \frac{1}{R^4} \]
for \( k_B T > \frac{q}{4\pi \epsilon_0 R^2} \).

2) Dipole-dipole interaction: we recall the geometric picture for this interaction (Fig. 2.19) and its interaction energy in spherical coordinates (see Section 2.4.2 for derivation)
\[ V(R) = \frac{k p_A p_B}{R^3} \left\{ -2 \cos \theta_1 \cos \theta_2 + \sin \theta_1 \sin \theta_2 \cos \phi \right\}. \]

The first order term vanishes since \( \langle \cos \theta_1 \rangle = 0 \), etc. The second order term does not vanish and equals
\[ V(R) = \left\langle -\frac{k^2 p_A^2 p_B^2}{2 R^6 k_B T} \left( \frac{4 \cos^2 \theta_1 \cos^2 \theta_2}{1/3} + \sin^2 \theta_1 \sin^2 \theta_2 \cos^2 \phi \right) + \ldots \right\rangle \]
\[ = -\frac{k^2 p_A^2 p_B^2}{3 k_B T R^6} \]
for \( k_B T > \frac{k p_A p_B}{R^3} \). Note: we did not include the cross term
\[ -4 \cos \theta_1 \cos \theta_2 \sin \theta_1 \sin \theta_2 \cos \phi \]
since it is zero (because \( \langle \cos \theta_1 \rangle = 0 \), etc.). This is the Keesom interaction and it scales as \( \sim 1/R^6 \). The Keesom interaction is 1 of 3 different contributions to the VDW interaction.

2.11. Keesom Interaction (permanent dipole-permanent dipole, thermally averaged)

The Keesom interaction, which was derived in Section 2.10, is due to the interaction of permanent dipoles. It is temperature dependent. Such interactions consist of attractive interactions between an ensemble of dipoles that are averaged over different rotational orientations of the dipoles. It is assumed that the molecules are constantly rotating and are never get locked. The energy of a Keesom interaction depends on the inverse sixth power of the
distance, unlike the interaction energy of two spatially fixed dipoles, which depends on the inverse third power of the distance. The Keesom interaction can only occur among molecules that possess permanent dipole moments, i.e., two polar molecules. Also, Keesom interactions are very weak VDW interactions and do not occur in aqueous solutions that contain electrolytes. The angle-averaged interaction is given by:

\[ V(R) = \frac{-m_1^2 m_2^2}{24 \pi^2 \epsilon_0^2 \epsilon^2 (k_B T) R^6}, \]

where \( m \) is the charge per unit length and \( \epsilon \) is the dielectric constant of surrounding medium (\( \epsilon = 1 \) for vacuum).

This force is inversely proportional to temperature. At low temperatures the dipoles want to be in the energetically favorable “ground state” whereas at high temperatures they are found in random orientations (where the forces cancel out on average), causing a weakening of the force.

2.12. Debye Force (Dipole-Induced Dipole)

The interaction of a polar molecule with a nonpolar molecule is analogous to the ion-induced dipole interaction. Suppose that we have a fixed dipole \( p \) with orientation \( \theta \) with respect to the line joining it to a polarizable molecule, the electric field of the dipole acting on the molecule is (omitting the angular factor and other factors), \( E \propto R^{-3} \), which means that the interaction energy is

\[ (2.7) \quad V(R, \theta) = -\frac{1}{2} \alpha E^2 \propto \frac{\alpha}{R^6} \]

For two molecules possessing permanent dipole moment \( p_1 \) and \( p_2 \) and polarizabilities \( \alpha_1 \) and \( \alpha_2 \), their net dipole-induced dipole energy is

\[ V(R) = -\frac{[p_1^2 \alpha_2 + p_2^2 \alpha_1]}{(4\pi\epsilon_0)^2 R^6} \]

This Debye interaction is one of three inverse sixth power contributions to the total VDW interaction energy between molecules.

The Debye force is an angle averaged interaction because it arises from interactions between rotating permanent dipoles and from the polarizability of atoms and molecules (induced dipoles). These induced dipoles occur when one molecule with a permanent dipole repels another molecule’s electrons. A molecule with permanent dipole can induce a dipole in a similar neighboring molecule and cause mutual attraction. Debye forces cannot occur between atoms, because atoms do not possess permanent electric dipole moments. The forces between induced and permanent dipoles are not as temperature dependent as Keesom interactions because the induced dipole is free to shift...
and rotate around the non-polar molecule. The Debye induction effects and Keesom orientation effects are termed polar interactions.

One example of an induction-interaction between permanent dipole and induced dipole is the interaction between HCl and Ar. In this system, Ar experiences a dipole as its electrons are attracted (to the H side of HCl) or repelled (from the Cl side) by HCl.

This kind of interaction can be expected between any polar molecule and non-polar/symmetrical molecule. The induction-interaction force is far weaker than dipole-dipole interaction, but can be stronger than the London dispersion force, depending on the strength of the dipoles involved.

### 2.13. Boltzmann-Weighted Interactions

Here is an alternative approach to deriving thermally averaged interaction energies. It gives almost the same result as the previous method. We all agree that the interaction energy, $V(R, \Omega)$, should be averaged over the angles in some way (i.e. over the unit sphere). We also know that some angles are more likely than others, because certain configurations of the dipole-dipole interaction are lower in energy than others. It makes sense to weigh these angles more heavily in the average. This can be done using the Boltzmann factor:

$$V(R) = \frac{\langle V(R, \Omega)e^{-V(R,\Omega)/k_B T} \rangle}{\langle e^{-V(R,\Omega)/k_B T} \rangle} \equiv \int \frac{V(R, \Omega)e^{-V(R,\Omega)/k_B T} d\Omega}{\int e^{-V(R,\Omega)/k_B T} d\Omega},$$

where the denominator, $\langle e^{-V(R,\Omega)/k_B T} \rangle$, is a normalization factor.

Here, $\Omega$ denotes any set of spherical angles, depending on the nature of the problem we want to solve. The ion-dipole interaction, for example, requires only one such set of angles whereas the dipole-dipole interaction requires two sets of spherical angles. Suppose that there are two sets of angles involved, $\Omega = (\Omega_1, \Omega_2) = (\theta_1, \phi_1, \theta_2, \phi_2)$ then $\Omega$ is shorthand notation for all spherical angles. The integral is over all unit spheres associated with each set of angles. For example, if $\Omega = (\Omega_1, \Omega_2) = (\theta_1, \phi_1, \theta_2, \phi_2)$, then $\int d\Omega$ stands for:

$$\int d\Omega = \int_{0}^{\pi} \sin \theta_1 d\theta_1 \int_{0}^{\pi} \sin \theta_2 d\theta_2 \int_{0}^{2\pi} d\phi_1 \int_{0}^{2\pi} d\phi_2.$$

In the high-temperature regime ($V < k_B T$) we may employ a Taylor expansion for the exponential. The denominator approaches 1 if we keep only the zeroth order term $e^{-V(R,\Omega)/k_B T} \approx 1$. The numerator is expanded in powers of $V/k_B T$:

$$V(R) = \langle V(R, \Omega)e^{-V(R,\Omega)/k_B T} \rangle \approx \langle V(R, \Omega) \left( 1 - \frac{V(R, \Omega)}{k_B T} + \ldots \right) \rangle$$
2. Intermolecular Interactions

Substituting the dipole-dipole energy (see Section 2.4.2 for derivation)

\[ V(R) = \left( k p_A p_B / R^3 \right) \{ -2 \cos \theta_1 \cos \theta_2 + \sin \theta_1 \sin \theta_2 \cos \phi \} \]

into this expression, the lowest order term \( \langle V(R, \Omega) \rangle \) vanishes for the same reasons as before\textsuperscript{16}. The next order term gives

\[
V(R) = - \left( \frac{V(R, \Omega)^2}{k_B T} + \ldots \right)
= - \frac{k^2 p_A^2 p_B^2}{(k_B T) R^6} \left\langle 4 \cos^2 \theta_1 \cos^2 \theta_2 + 0 + \sin^2 \theta_1 \sin^2 \theta_2 \cos^2 \phi \right\rangle
\]

So,

\[ V(R) = - \frac{2k^2 p_A^2 p_B^2}{3(k_B T) R^6} \]

We are off by a factor of 2 but otherwise have obtained the correct formula with the \( 1/R^6 \) dependence.

2.14. \( k_B T \) vs Interaction Strength

In the previous section, the Boltzmann factor is used to weigh the angles according to their energies. Angles that lead to lower energy configurations weigh more heavily: the Boltzmann factor, \( e^{-V/k_B T} \), involves comparing the interaction energy strength \( V(R, \Omega) \) for a given angle configuration \( \Omega = (\theta, \phi) \) to \( k_B T \). If \( V(R, \Omega) \) is small compared to the magnitude of \( k_B T \) (0.03 eV at room temperature), then the Boltzmann factor is large. Otherwise, its magnitude is small (compared to 1).

2.15. Relative Strengths of Intermolecular Forces

Table 2.7 gives typical dissociation energies \( (E_d) \) for different types of bonds. It is best to think of these values as “typical” values, i.e. orders of magnitude. The actual strengths will vary depending on the molecules involved. Ionic bonding and covalent bonding will always be stronger than intermolecular forces in any given substance. To determine whether ionic bonding occurs, we must have a metal bound to a non-metal, e.g. NaCl.

The main intermolecular forces are summarized in Table 2.8.

\textsuperscript{16}Recall: \( \langle \cos^2 \theta \rangle = 1/3, \langle \sin^2 \theta \rangle = 2/3, \langle \sin^2 \phi \rangle = \langle \cos^2 \phi \rangle = 1/2 \) and \( \langle \sin \theta \rangle = \langle \cos \theta \rangle = 0, \langle \sin \theta \cos \theta \rangle = 0, \langle \sin \phi \rangle = \langle \cos \phi \rangle = 0, \langle \sin \phi \cos \phi \rangle = 0 \)
Table 2.7. Typical dissociation energies for different types of chemical bonds. 1 cal = 4.184 J. 1 eV = 1.6×10^{−19} J.

<table>
<thead>
<tr>
<th>Bond type</th>
<th>$E_d$ (kcal/mol)</th>
<th>$E_d$ (kJ/mol)</th>
<th>$E_d$ (eV)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Interatomic (intramolecular) bonds</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ionic lattice (solid)</td>
<td>250-4,000</td>
<td>1,046-16,736</td>
<td>11-173</td>
</tr>
<tr>
<td>Ionic (molecular)</td>
<td>80-251</td>
<td>335-1,050</td>
<td>3.5-10.9</td>
</tr>
<tr>
<td>Covalent bond</td>
<td>15-260</td>
<td>63-1,088</td>
<td>0.653-11.2</td>
</tr>
<tr>
<td>Metallic</td>
<td>26.3-83.7</td>
<td>110-350</td>
<td>1.14-3.63</td>
</tr>
<tr>
<td>Intermolecular bonds</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen bond</td>
<td>1-12</td>
<td>4.2-50.2</td>
<td>0.04-0.52</td>
</tr>
<tr>
<td>Dipole-dipole (excluding H-bond)</td>
<td>0.5-5.0</td>
<td>2.1-21</td>
<td>0.022-0.218</td>
</tr>
<tr>
<td>London dispersion</td>
<td>&lt;1 to 15</td>
<td>&lt;4.2-62.8</td>
<td>&lt;0.04-0.65</td>
</tr>
<tr>
<td>Dipole-induced dipole</td>
<td>∼0.48</td>
<td>∼2.0</td>
<td>∼0.021</td>
</tr>
</tbody>
</table>

2.16. Hard Sphere and Lennard-Jones Potentials

Insights into the physics of liquids, gases and solids have often come from computer simulations. Some of the very first scientific computer simulations back in the 1950s studied the hard-sphere (HS) liquid.

The hard-sphere potential is defined as:

$$V(R) = \begin{cases} +\infty & R < a \\ 0 & R \geq a \end{cases}$$

where $a$ is the distance of closest approach, twice the molecular radius and $R = |\mathbf{R} - \mathbf{R}'|$. This potential only models the repulsive core, but no attractive tail. Because of the abrupt transition at $r = a$ it describes perfectly non-overlapping spheres. The HS potential has proved valuable in the study of sphere packings. HS exhibit a fluid-solid phase transition at sufficiently high densities. An early surprise was that the HS system has a first-order freezing transition.

The simplest liquid models\(^\text{17}\) are defined by pair potentials, $V(\mathbf{R})$. If $r_{ij} = |\mathbf{r}_i - \mathbf{r}_j|$ is the distance between particles $i$ and $j$, the potential energy $V$ as a function of all particle coordinates is given by

$$V(\mathbf{r}_1, \ldots, \mathbf{r}_N) = \sum_{i<j} V(r_{ij}).$$

\(^{17}\)We call them “liquid models” because in gases we mostly neglect collisions; thus the intermolecular potentials are not so important. In solids, these spherically symmetric models are not suitable because they fail to describe the many different possible spatially ordered arrangements in crystals. More sophisticated models are needed which describe reduced symmetry.
### Table 2.8. Most common intermolecular forces. Source: Israelachvili [2]

<table>
<thead>
<tr>
<th>Type of Interaction</th>
<th>Interaction Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>Covalent, Metallic</td>
<td>Complicated, short range</td>
</tr>
<tr>
<td>Charge-Charge</td>
<td>$+Q_1Q_2/4\pi\epsilon_0 r$ (Coulomb Energy)</td>
</tr>
<tr>
<td>Charge-Dipole (fixed)</td>
<td>$-Q\rho \cos \theta /4\pi\epsilon_0 r^2$</td>
</tr>
<tr>
<td>Charge-Dipole (freely rotating)</td>
<td>$-Q^2\rho^2/(6(4\pi\epsilon_0)^2 kr^4)$</td>
</tr>
<tr>
<td>Dipole-Dipole (fixed)</td>
<td>$</td>
</tr><tr>
<td>ho_1\rho_2(2\cos \theta_1 \cos \theta_2 - \sin \theta_1 \sin \theta_2) /4\pi\epsilon_0 r^2$</td>
<td></td>
</tr>
<tr>
<td>Dipole-Dipole (freely rotating)</td>
<td>$-\rho_1^2 \rho_2^2 /3(4\pi\epsilon_0)^2 kr^4$ (Keesom energy)</td>
</tr>
<tr>
<td>Charge-Non-polar</td>
<td>$-Q^2\rho/2(4\pi\epsilon_0)^2 r^4$</td>
</tr>
<tr>
<td>Dipole-Non-polar (fixed)</td>
<td>$-\rho^2 \rho (1 + 3\cos^2 \theta/2(4\pi\epsilon_0)^2 r^6$</td>
</tr>
<tr>
<td>Dipole-Non-polar (rotating)</td>
<td>$-\rho^2 \rho (4\pi\epsilon_0)^2 r^6$ (Debye energy)</td>
</tr>
<tr>
<td>Non-polar - Non-polar</td>
<td>$-\hbar \nu v_0^2 / ((4\pi\epsilon_0)^2 r^3$ (London Dispersion energy)</td>
</tr>
<tr>
<td>Hydrogen Bond</td>
<td>Complicated, short range, proportional to $-1/r^2$</td>
</tr>
</tbody>
</table>

Such systems are traditionally referred to as ‘simple liquids’ because they describe point particles with isotropic interactions, i.e. with a mathematically simple Hamiltonian.

In many simulations one uses ad hoc pair potentials. The most famous one is the Lennard-Jones (LJ) pair potential (1924) defined by

$$V_{LJ}(R) = 4\varepsilon \left[ \left( \frac{\sigma}{R} \right)^{12} - \left( \frac{\sigma}{R} \right)^6 \right],$$

where $\varepsilon$ is the well depth$^{18}$, $\sigma$ is the distance at which $V(R)$ passes through zero. The attractive part, which describes VDW forces, is proportional to

$^{18}$The point where the slope is zero gives:

$$\frac{dV}{dR} = 4\varepsilon \left[ 12(\sigma/R)^{11}(-\sigma/R^2) - 6(\sigma/R)^5(-\sigma/R^2) \right] = 0,$$
2.16. Hard Sphere and Lennard-Jones Potentials

\[ R^{-6} \] whereas the repulsive part, which describes Pauli repulsion, is proportional to \( R^{-12} \) (Fig. 2.32).

![Lennard-Jones potential](image)

Figure 2.32. Lennard-Jones potential.

(Recall that the force \( F = -dV/dR \) in 1D or \( F = -\nabla V \) in 3D.) The sixth-power term does not represent other kinds of bonding (non-VDW) well. The twelfth-power term appearing in the potential is chosen for its ease of calculation for simulations (by squaring the sixth-power term) and is not physically based. The LJ potential is a model for intermolecular forces designed to describe the interaction between neutral atoms or molecules.

The minimum value of \( V_{LJ}(R) \) is \(-\varepsilon\), which is found at \( R = 2^{1/6}\sigma \) (check this by finding \( R \) such that \( dV_{LJ}/dR = 0 \)). Commonly simulated simple liquid models are: the LJ pair potential (sometimes cut off at the minimum, sometimes with other exponents than 12 and 6), the purely repulsive inverse-power law pair potentials \( V(R) \propto (R/\sigma)^{-n} \), the Yukawa “screened Coulomb” pair potential \( V(R) \propto \exp(-R/\sigma)/R \), the Morse pair potential that is a difference of two exponentials, \( V(R) = D_{e}(e^{-2a(R-r_e)} - 2e^{-a(R-r_e)}) \), and the HS pair potential.

2.16.1. PE Curves for Different Molecules (but same # of electrons). The relative strengths and effective ranges of several intermolecular forces are illustrated in Fig. 2.33, which shows how the potential energy

and solving for \( R \) leads to \( R = \sigma 2^{1/6} \). Substitution of this \( R \) value into the Lennard-Jones equation gives \( V(\sigma 2^{1/6}) = -\varepsilon \).
(PE) depends on the intermolecular separation (center-to-center distance) for several pairs of ions, atoms, and molecules.

The potentials illustrated in Fig. 2.33 include Coulomb \((R^{-1})\), dipole-dipole \((R^{-3})\), dispersion \((R^{-6})\), and repulsive \((R^{-12})\) potentials. The species shown were chosen so that the interacting atoms, ions, or molecules have the same number of electrons (Ar, Cl, K, HCl). For comparison, the covalent bond (intramolecular force) for \(\text{Cl}_2\) is also shown.

The ion-ion interaction of K with Cl is the strongest (stronger even than the covalent interaction in \(\text{Cl}_2\)), followed by the interaction between two HCl molecules (dipole-dipole and dispersion) and the Ar-Ar interaction (dispersion only).

Notice in Fig. 2.33 the dramatically different depths of the wells (several orders of magnitude), as well as the distinctly different distances at which the minima occur.

![Figure 2.33. Potential energy curves showing attractive and repulsive forces for different types of interactions.](image)

The potential energy of a pair of atoms, ions, or molecules depends on the distance between the members of the pair. Here, the potential energy at large separations (to the right side of the graph) is arbitrarily set to zero by convention. As pairs of particles approach each other, the potential energy becomes negative because attractive forces come into effect. The lowest point in each curve occurs at the distance where attractive and repulsive forces exactly balance. The relative potential energy values at these minima measure the relative strength of the attractive forces in the various cases.
illustrated. Note the shallow potential energy minimum for HCl and Ar. The inset of Fig. 2.33 shows these same two curves with the vertical scale expanded by a factor of 100.

2.16.2. Are CO₂ and H₂O Polar or Nonpolar? To determine whether dipole-dipole interactions are significant, we need to look at two things: 1) the dipole moment of each bond and 2) the geometry of the molecule. The net dipole moment, which determined if a molecule is polar or not, is the vector sum of all its constituent dipole moments (i.e. we vectorially add up the dipole moments of all bonds). When two atoms involved in a bond have a significant difference in electronegativity, the bond is polar, with a dipole moment that points toward the least electronegative atom (according to our convention). For example, the C=O bond is polar because the difference in electronegativity is \( \Delta E_N = 3.5 - 2.5 = 1.0 \), which is significant. In the C=O bond, the most electronegative atom is O; thus the dipole moment points from O to C.

CO₂ has covalent bonds, each of which have a dipole moment, but its overall dipole moment is zero due to cancellation of the two dipole moment vectors, which point in opposite directions. The molecule is therefore nonpolar (Fig. 2.34). Only VDW forces exist. The molecule has no net dipole moment because even though CO bonds are polar, they cancel each other out due to the geometry of the molecule.

**Figure 2.34.** CO₂ has no permanent EDM (by symmetry) whereas H₂O does.

CO₂ is linear because the C atom has a complete octet of valence electrons arising from covalent bonding to its two neighboring O atoms. H₂O is bent because the octet around the O atom features 2 lone electron pairs.

2.16.3. Are BF₃ and NH₃ Polar or Nonpolar? Boron trifluoride (BF₃) is perfectly nonpolar because the dipole vectors for each bond vectorially cancel out. Ammonia (NH₃) has an overall dipole because the vectors point somewhat in one direction (Fig. 2.35).
2. Intermolecular Interactions

2.16.3.1. Ammonia ($\text{NH}_3$). Ammonia is nonplanar because of the presence of a lone pair of electrons on the nitrogen atom. This can be seen by drawing the Lewis structure (Fig. 2.36). Ammonia can also do hydrogen bonding since it has N-H bonds.

$$\text{H} : \text{N} : \text{H}$$

Figure 2.36. Lewis structure of $\text{NH}_3$.

2.16.3.2. Boron Trifluoride ($\text{BF}_3$). To see that $\text{BF}_3$ is planar we draw its Lewis structure (Fig. 2.37). Boron has 3 valence electrons, whereas fluorine has 7. We have a total of $3 + 3(7) = 24$ valence electrons, which is a multiple of 8. There are no lone pairs on the central boron atom. Because fluorine is more electronegative than boron, boron will see its 3 electrons pulled away toward the F atoms.

Figure 2.37. Lewis structure of $\text{BF}_3$.

There are no ionic interactions here. There are no hydrogen atoms, so there cannot be any hydrogen bonds. This molecule is non-polar, because the 3 orientations of the dipole moments relative to the boron atom all cancel out vectorially.

<table>
<thead>
<tr>
<th>ionic</th>
<th>H-bond</th>
<th>dipole</th>
<th>dispersion</th>
</tr>
</thead>
<tbody>
<tr>
<td>no</td>
<td>no</td>
<td>no</td>
<td>yes</td>
</tr>
</tbody>
</table>
2.16. Are CF$_4$ and CH$_3$F Polar or Nonpolar? CF$_4$ is nonpolar because of geometry (Fig. 2.38). CH$_3$F has an overall dipole because of the CF bond. However, it does not participate in hydrogen bonding since it has no H-F bonds. For CF$_4$ we have:

<table>
<thead>
<tr>
<th>ionic</th>
<th>H-bond</th>
<th>dipole</th>
<th>dispersion</th>
</tr>
</thead>
<tbody>
<tr>
<td>no</td>
<td>no</td>
<td>no</td>
<td>yes</td>
</tr>
</tbody>
</table>

For CH$_3$F we have:

<table>
<thead>
<tr>
<th>ionic</th>
<th>H-bond</th>
<th>dipole</th>
<th>dispersion</th>
</tr>
</thead>
<tbody>
<tr>
<td>no</td>
<td>no</td>
<td>yes</td>
<td>yes</td>
</tr>
</tbody>
</table>

![Figure 2.38. Structure of CF$_4$ and CH$_3$F. CF$_4$ is nonpolar because the vector sum of all bond dipoles is zero. CH$_3$F is polar due to its lower symmetry. The direction of the arrow indicates the orientation of the permanent EDM.](image)

2.16.5. Phosphorous Trichloride (PCl$_3$). Phosphorous trichloride has the Lewis structure and 3D representation shown in Fig. 2.39. For PCl$_3$, phosphorous (P) has 5 valence electrons whereas chlorine (Cl) has 7, for a total of 5 + 3(7) = 26. The highest multiple of 8 is 24. The difference is 26 - 24 = 2, which tells us that there is 1 lone pair of 2 electrons on the central phosphorus atom. Because Cl has 7 valence electrons, it needs 1 more to get to 8, Cl wants to form 1 bond to acquire this extra electron. Therefore, we draw 3 lone pairs on each chlorine (Cl) atom so that each Cl atom has 8 electrons. If we count all valence electrons in the structure, we have a total of 26.

![Figure 2.39. Structure of phosphorous trichloride.](image)
In $\text{PCl}_3$ we don’t have a metal attached to a non-metal. Thus, there are no ionic interactions. There are no hydrogen bonding interactions because there is no hydrogen atom. However, we do have dipolar interactions because the molecule is polar: P has a lone pair of electrons whereas Cl has 3 (P atom is more positively charged than Cl), leading to a dipole moment from each Cl atom (negatively charged) to the P atom (positively charged). Their vector sum (net dipole moment) points upward in Fig. 2.39. Finally, dispersion forces are always present.

<table>
<thead>
<tr>
<th>ionic</th>
<th>H-bond</th>
<th>dipole</th>
<th>dispersion</th>
</tr>
</thead>
<tbody>
<tr>
<td>no</td>
<td>no</td>
<td>yes</td>
<td>yes</td>
</tr>
</tbody>
</table>

2.16.6. Boiling Points (B.P.) of Alkanes: Effects of Molar Mass. If we have only VDW interactions among molecules, it is important to look at other factors, such as number of electrons (polarizability), contact area and $\pi$-bonding. Consider molecules that are all linear but whose molar mass differs. In Fig. 2.40, the boiling points of $n$-alkanes increase with molar mass because the dispersion force is stronger due to increased contact area between dimers.

![Boiling point (°C) vs. No. of carbon atoms](image)

Figure 2.40. Boiling point of several alkanes.

2.16.7. B.P. of $n$-pentane vs Neopentane: Effects of Molecular Shape. Molar mass is not the only factor. Molecular shape also matters. For $n$-pentane, the boiling point (B.P.) is higher than neopentane because $n$-pentane makes contact with its neighbors using a larger contact area, increasing the number of interactions possible.
For neopentane the contact area is much smaller because the molecules are nearly spherical, therefore decreasing the boiling point by reducing the number of interactions possible:

2.16.8. Dipole-Dipole Examples. For CH\(_4\) the molecule is nonpolar because of its symmetry. We note that each C-H bond is very weakly polar, due to the very small difference in electronegativity \(\Delta E_N = 2.5 - 2.1 = 0.4\) between C and H.

For CCl\(_4\), the molecule is also nonpolar for the same reason (geometry). Here, the difference in electronegativity in the C-Cl bonds is larger, \(\Delta E_N = 3.0 - 2.1 = 0.9\).

Chloroform (CHCl\(_3\)) is polar because of its lower symmetry. The net dipole moment points along the C-H bond direction. This is shown below:
2. Intermolecular Interactions

CH$_4$ is nonpolar (due to symmetry)

CCl$_4$ is nonpolar (due to symmetry)

CHCl$_3$ is polar (C-H axis)

For C$_2$H$_2$F$_2$, there are 3 possible isomers. Of the three shown, the first two are polar while the third one is nonpolar:

While CH$_2$FCH$_2$F looks nonpolar as drawn, a C-C single bond can rotate freely. Thus, there are times when the dipole moments do not cancel each other out. This residual (time-averaged) dipole moment is non-zero and the molecule is polar on the average.

2.16.9. Examples that Involve Hydrogen Bonding. H$_2$O is more polar than H$_2$S. In fact, it can undergo hydrogen bonding with other water molecules.
In Fig. 2.41, both molecules have the same atoms (C₂H₆O), but different arrangements. Dimethyl ether is polar but its boiling point is lower than ethanol, which has a much higher boiling point because it undergoes hydrogen bonding (due to the terminal O-H).

**Figure 2.41.** Effect of intermolecular bonding on boiling and melting points. (Temperatures shown are in Celsius degrees.)

### 2.16.10. Combining it All: Ordering Compounds Based on Boiling Points.

Another typical question you may be asked to solve involves ordering compounds in order of their boiling points. In the example below, we have: NaCl, ethane, ethanol, difluoromethane.

You may be asked to arrange these molecules in order of increasing B.P. The first step is to identify the dominant intermolecular interaction:

- NaCl: ionic
- CH₃CH₃: nonpolar
- CH₃CH₂OH: H-bond
- CH₂F₂: dipolar

Note: CH₂F₂ does not form H-bonds. It’s a tetrahedral structure where all H and F atoms are bound to the carbon (Fig. 2.42).

**Figure 2.42.** Structure of CH₂F₂.

Ordering according to B.P. gives:
2. Intermolecular Interactions

\[ \text{CH}_3\text{CH}_3 < \text{CH}_2\text{F}_2 < \text{CH}_3\text{CH}_2\text{OH} < \text{NaCl} \]
\((-89^\circ\text{C}) \quad (-52^\circ\text{C}) \quad (78^\circ\text{C}) \quad (1413^\circ\text{C})\]

2.16.10.1. Boiling Point: Periodic Table Trends. Substances with weaker intermolecular forces lead to lower boiling points, \( T_b \). For example, ionic liquids have high boiling points because of the strong Coulomb interaction among ions. \( \text{NaCl} \) has \( T_b \approx 1,686 \text{ K} \). Helium boils at \( \sim 4 \text{ K} \). Xenon has a higher boiling point than helium because it has a higher polarizability. Polar liquids (e.g. \( \text{HCl} \)) have higher boiling points than non-polar liquids (e.g. \( \text{N}_2 \)). See Fig. 2.43.

2.16.10.2. Example. State which intermolecular forces dominate among pairs of molecules of the following substances: (a) \( \text{F}_2(\text{s}) \), (b) \( \text{HBr}(\text{l}) \), (c) \( \text{NH}_4\text{Cl}(\text{s}) \).

Solution: (a) Since \( \text{F}_2 \) is nonpolar, VDW forces dominate. (b) Since \( \text{HBr} \) has an EDM, dipole-dipole dominates. \( \text{HBr} \) does not form H-bonds. (c) Within \( \text{NH}_4\text{Cl}(\text{s}) \) there are ion-ion forces (between \( \text{NH}_4^+ \) and \( \text{Cl}^- \) ions). Between pairs of \( \text{NH}_4\text{Cl} \) molecules, there can be H-bonds (thanks to the N-H bonds). \( \text{NH}_4\text{Cl} \) also has a permanent EDM, so there are dipole-dipole forces. And of course, VDW forces. The strongest intermolecular forces would be the H-bonding.

2.17. Problems

Problem 28. Compare the magnitude of the gravitational force on a molecule to typical intermolecular forces. For the latter, assume an electrostatic
(Coulomb) interaction between two elementary charges separated 1 nm apart. There are 2 components to the gravitational force: the attraction of the Earth, the attraction from nearest-neighbor molecules. Take the distance between nearest-neighbor molecules to be 1 nm. Why do we discuss intermolecular forces in this chapter, but not gravity?

Solution. The gravitational force between two objects is \( F = \frac{G m_1 m_2}{R^2} \), where \( R \) is the distance between the centers of mass of the two objects, \( m_1, m_2 \) are the masses of objects 1 and 2 (respectively) and \( G \) is the gravitational constant, whose magnitude is \( 6.67 \times 10^{-11} \text{ N.m}^2/\text{kg}^2 \). We first consider the gravitational force exerted on a molecule by planet Earth. Near the surface of the Earth, \( F = mg \), where \( g \approx 9.81 \text{ m/s}^2 \). One proton weighs \( 1.672 \times 10^{-27} \text{ kg} \). Then, \( F \approx 1.64 \times 10^{-26} \text{ N} \).

Second, we consider the gravitational attractive forces between two nearby molecules. For two hydrogen atoms separated \( R = 1 \text{ nm} \) apart, this force is \( F = (6.67 \times 10^{-11})(1.672 \times 10^{-27})^2 \cdot (1 \times 10^{-9})^{-2} \approx 1.86 \times 10^{-46} \text{ N} \).

For electrostatic interactions, the forces are calculated from the Coulomb’s law, which states that the force experienced by two charges \( q_1, q_2 \) separated by a distance \( R \) has magnitude \( F = (4\pi\varepsilon_0)^{-1}(q_1 q_2)R^{-2} \). Electrostatic charges of ions are on the order of \( 1.6 \times 10^{-19} \text{ C} \). Typical intermolecular distances are a few Angstroms. Take \( R = 1 \text{ nm} \) for example. Then, \( F = (1.11 \times 10^{-10})(1.6 \times 10^{-19})^2 \cdot (10^{-9})^{-2} \approx 2.8 \times 10^{-30} \text{ N} \).

We conclude that intermolecular electrostatic forces are 16 orders of magnitude smaller than intermolecular gravitational forces. Thus, intermolecular gravitational forces can be neglected. On the other hand, the Earth’s gravitational force is 4 orders of magnitude stronger than electrostatic forces. But unlike intermolecular forces, which have random directions, the Earth’s gravitational field always points downward and has a net effect of causing substances to flow.

### Problem 29

Calculate the force applied by one atom on another according to a Lennard-Jones potential. Evaluate the force at the distances of 3.0, 3.8, 4.5 Å for oxygen. (For oxygen \( \sigma = 3.58 \text{ Å} \) and \( \varepsilon = 1.622 \times 10^{-21} \text{ J} \).) Is the force attractive or repulsive at these points?

Solution. Starting from:

\[
V_{LJ}(R) = 4\varepsilon \left[ \left( \frac{\sigma}{R} \right)^{12} - \left( \frac{\sigma}{R} \right)^6 \right]
\]

Take the (-ve) derivative of the LJ potential with respect to \( r \) to get the force

\[
F_{LJ}(R) = -\frac{dV}{dR} = -4\varepsilon \left( \frac{6\sigma^6}{R^7} - \frac{12\sigma^{12}}{R^{13}} \right)
\]
For 3.0 Å: $1.8 \times 10^{-9}$ J/m, which is repulsive. For 3.8 Å: $2.9 \times 10^{-11}$ J/m, which is repulsive. For 4.5 Å: $-1.1 \times 10^{-11}$ J/m, which is attractive. [Note: the unit J/m is equivalent to Newtons.]

**Problem 30.** How can NaCl crystal dissolve in water when ionic forces in the NaCl crystal (the “lattice energy” that holds the NaCl crystal together is approximately 8.2 eV; the ion-ion bond strength in NaCl is 1.52 eV) are stronger than ion-dipole forces? For the ion-dipole force between water molecules and the ions in the crystal, use the formula (SI units)

$$V(R) = -\frac{Qp}{4\pi\epsilon R^2},$$

where $R$ is the distance between the ion and the dipole (we assume a “point dipole”), $p$ is the electric dipole moment and $Q$ is the ion charge (an electron has charge $Q = -1.602 \times 10^{-19}$ C). The ion separation in NaCl crystal is 0.28 nm. For water molecules, $p=1.85$ Debye or $6.17 \times 10^{-30}$ C.m in SI units, since 1 D=3.336 $\times 10^{-30}$ C.m. $\epsilon$ is the dielectric permittivity. For vacuum, $\epsilon = 8.85 \times 10^{-12}$ F/m. Explain how it is possible for water to dissolve the NaCl crystal in spite of the fact that the ion-dipole interaction is weaker than the ion-ion interaction that holds the crystal together (since we know this to be true from experiments).

**Solution.** For vacuum, $1/4\pi\epsilon_0 = 8.99 \times 10^9$ N.m$^2$.C$^{-2}$. For an ion-dipole distance of 2.5 Å, the interaction strength is

$$\frac{(1.85 \times 3.336 \times 10^{-30})(1.6 \times 10^{-19})(8.99 \times 10^9)}{(2.5 \times 10^{-10})^2} = 1.42 \times 10^{-19} \text{ J} = 0.88 \text{ eV}$$

If we consider the lattice energy of NaCl (8.2 eV), approximately 9 water molecules are needed (on average) to break apart each ion in the NaCl structure. On the other hand, if we consider the ion-ion bond strength of NaCl (1.52 eV), about 2 water molecules are needed per ion. (So the actual number of water molecules per ion of NaCl is between 2 and 9.)

**Problem 31.** Imagine that NaCl crystal is completely dissolved in water. The more NaCl we dissolve, the more Na$^+$ and Cl$^-$ ions are present in water. The solution consists of metal aquo complexes with the formula [Na(H$_2$O)$_8$]$^+$, with Na-O distance of 250 pm. The chloride ions are also strongly solvated, each being surrounded by an average of 6 molecules of water. When solvated the Na and Cl ions are stable. There will come a point where too many ions present in aqueous solution will lead to precipitation (Na$^+$ and Cl$^-$ ions will become so close to each other, in terms of average separation, that the formation of NaCl will be preferred over dissolution). What is this concentration, and how does it compare with the solubility limit of NaCl in water (357 mg/mL at 25°C)? (Hint: consider the point
where the ion-ion interaction energy becomes stronger than the ion-dipole interaction energy.)

\[ \text{Solution. Notation: We use } V \text{ to denote volume and } V(R) \text{ to denote intermolecular interaction strength. Let us assume } n = 7 \text{ water molecules on average per hydrated ion complex, so we can treat the case of Na and Cl simultaneously. (It's no problem to take } n = 8 \text{ and } n = 6; \text{ I am doing this simplification only to save space.) This calculation can only be approximate since we disregard the screening of charges and the dielectric constant of the medium. We also neglect any dynamical effects and neglect the hydrogen bonding between the inner and outer hydration layers.} \]

It is instructive to look at two different limits. We start with the worst possible case scenario. Let us consider the size of water molecule, 2.75 Å, and atomic radius of Na\(^+\) (1.02 Å) and Cl\(^-\) (1.81 Å). The worse case scenario is an ion-dipole separation of \( R = 2.75 + 1.81 = 4.56 \) Å. You can think of this as a rough attempt at modeling hydrogen bonding between the inner and outer hydration layers; the net effect being to “pull” inner water molecules away from the ion. Then, \( V(R) = \frac{-Q_p}{4\pi\epsilon R^2} \) evaluates numerically to

\[
8.99 \times 10^9 \cdot \frac{(1.85 \times 3.336 \times 10^{-30})(1.6 \times 10^{-19})}{(456 \times 10^{-12})^2} = 4.23 \times 10^{-20} \text{ J,}
\]

Multiplying by 7 and solving for \( R \) in the Coulomb’s law:

\[
7 \times (4.23 \times 10^{-20} \text{ J}) = \frac{q_1 q_2}{4\pi\epsilon R} = 8.99 \times 10^9 \cdot \frac{(1.6 \times 10^{-19})^2}{R}
\]

which gives \( R = 7.7 \) Å. Let’s round this up to 1 nm. For inter-ion separation of 1 nm, the concentration of ions, \( \frac{N}{V} = 1 \left/ \frac{4}{3} \pi R^3 \right, \) is

\[
= 5.2 \times 10^{26} \text{ ions/m}^3 = 868 \text{ mol/m}^3 \cdot (0.001 \text{ m}^3/\text{L}) = 0.868 \text{ mol/L.}
\]

We want to compare this to the solubility of NaCl in water (359 g/L). The molar mass of Na of Cl ions is 25 or 35 g/mol, respectively (29 g/mol avg.). Then,

\[
0.868 \text{ mol/L} \times 29 \text{ g/mol} = 25 \text{ g/L.}
\]

which is too low. This is ok because it is only an estimate for the worst case scenario.

Let us use instead the data provided in the problem (\( R = 250 \) pm for Na-O distance). The ion-dipole energy between ion and water, \( V(R) = -\frac{Q_p}{4\pi\epsilon R^2} \),

is
\[ = 8.99 \times 10^9 \times \frac{(1.85 \times 3.336 \times 10^{-30})(1.6 \times 10^{-19})}{(250 \times 10^{-12})^2} = 1.42 \times 10^{-19} \text{ J} = 0.88 \text{ eV}, \]
and since we have \( n = 7 \) such water molecules bound to the ion, the total binding energy is 6.16 eV. Our task is to find \( R \) (distance between Na and Cl ions) such that the ion-ion interaction energy is larger than 6.16 eV. Setting \( E(\text{ion} - \text{ion}) = 6.16 \text{ eV} \) and invoking Coulomb’s law:

\[ 7 \times (1.42 \times 10^{-19} \text{ J}) = \frac{q_1 q_2}{4\pi\varepsilon R} = 8.99 \times 10^9 \times \frac{(1.6 \times 10^{-19})^2}{R} \]

which gives \( R \) of approximately 2.3 Å. Then, \( \frac{N}{V} = 1 \sqrt{\frac{4\pi R^3}{3}} \) is

\[ = 1.96 \times 10^{28} \text{ ions/m}^3 = 32,583 \text{ mol/m}^3 \times (0.001 \text{ m}^3/\text{L}) = 32.6 \text{ mol/L}, \]
where

\[ 32.6 \text{ mol/L} \times 29 \text{ g/mol} = 945 \text{ g/L}. \]
This overshoots the actual solubility limit of NaCl (359 g/L), but is off only by a factor of 3. We found that the solubility limit lies between 25 and 925 g/L.

**Problem 32.** Consider \( \beta \)-sheets that are stacked in a protein. While the strength of interaction between individual sheet elements is small, the overall interaction between sheets is strong because of cooperative effects (i.e. the sum of a very large number of small interactions leads to a strong interaction). Estimate the total energy of interaction (bonding energy) between pairs of \( \beta \)-sheets.

**Solution.** Let us consider the case of silk fibroin. \( \beta \)-sheets are stabilized by hydrogen bonding between peptide strands. In a \( \beta \)-sheet, regions of the polypeptide backbone lie parallel to each other and are connected by hydrogen bonds. The hydrogen bonds are formed between the carbonyl oxygen and the amine hydrogen of amino acid (AA) in adjacent strands in a polypeptide (i.e., the hydrogen bonds are inter-stand). The distance between adjacent amino acids is 3.5 Å.

The protein fibroin in silk has a repeating sequence of amino acids: mostly Gly-Ser-Gly-Ala-Gly-Ala, extending in length to over 5000 residues.\(^{19}\) 5,000 residues correspond to a length of 3.5 Å* 5000 = 1,750 nm = 1.75 μm.

VDW interaction are weak (0.4-4.0 kJ/mol compared with 12-30 kJ/mol for a hydrogen bond).\(^{20}\) Thus, along the first dimension (peptide chain), there

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\(^{19}\)From: [http://www.biotopics.co.uk/jsmol/fibroin.html](http://www.biotopics.co.uk/jsmol/fibroin.html)

are 5000 residues; assume 1 hydrogen bond per residue (0.4–4.0 kJ/mol), and one VDW interaction (12–30 kJ/mol), the total interaction strength ranging from $5000 \times (0.4+12) = 62,000$ kJ/mol to $5000 \times (4+30) = 170,000$ kJ/mol. Along the second dimension, let’s assume it is 1.75 µm long, and that each neighboring AA chain are $5.5+3.7=11.2$ Å apart (see figure below).

The structure of silk fibroin is:

Thus, there are approximately $1,750 \text{ nm} / 11.2 \text{ Å} = 1562$ AA chains. For a pair of sheets, the total interaction energy is between

\[
1562 \times 62,000 \text{ kJ/mol} = 96,844,000 \text{ kJ/mol} \text{ to } 1562 \times 170,000 \text{ kJ/mol} = 265,540,000 \text{ kJ/mol}.
\]

And since 1 eV = 96.485 kJ/mol, this energy range is 1 to 2.7 MeV, a very large binding energy. If we assume only VDW bonding, the total energy is in the range

\[
1562 \times 5000 \times (0.4) = 3,124,000 \text{ kJ/mol} = 32 \text{ keV} \text{ to } 1562 \times 5000 \times (4) = 31,240,000 \text{ kJ/mol} = 323 \text{ keV}.
\]

This should be compared against the strength of covalent bonds (200–1,000 kJ/mol) and ionic interactions (3,000 kJ/mol; for NaCl, 7.97 eV). We see that VDW interaction between β-sheets is several orders of magnitude stronger. This is a very strong bond, which is a consequence of the cooperative effect of many small interactions. Breaking such a bond would require enormous forces (or a suitable enzyme).

**Problem 33.** Hydrogen bonding is an intermolecular force often denoted by $X - H \cdots X$ where $X=F,O$, or $N$ and the dotted line is the intermolecular attraction between molecules of a dimer. This attraction involves a strongly

---

negative partial charge (in fact, a lone pair) on X pulling the positively charged H atom from the neighboring molecule.

(a) Does \( \text{H}_2\text{O} \) or \( \text{NH}_3 \) form stronger hydrogen bonds with itself? State reasons to support your answer.

(b) Does \( \text{NH}_4^+ \) hydrogen bond with itself, why or why not?

**Solution.** (a) \( \text{H}_2\text{O} \cdots \text{H}_2\text{O} \) hydrogen bond is stronger than \( \text{NH}_3 \cdots \text{NH}_3 \) bond because (1) the electronegativity difference between O and H is larger than that between N and H and (2) the hydrogen bonding in \( \text{H}_2\text{O} \) is linear whereas the hydrogen bonding in \( \text{NH}_3 \) is bent.

(b) Draw the Lewis structure: \( \text{H} \ \text{N} \ \text{H} \). There are no lone electron pairs! Thus, \( \text{NH}_4^+ \) does not hydrogen-bond with itself because it does not have lone pairs to participate in hydrogen bonding.

**Problem 34.** Consider the compounds \( \text{CHCl}_3 \) and \( \text{CHBr}_3 \) and compare the strengths of their dipole-dipole force and the strengths of their VDW forces. The respective boiling points are \( 61^\circ\text{C} \) and \( 149^\circ\text{C} \). For these compounds, which type of intermolecular force (dipole-dipole or VDW) contributes more to the apparent difference in boiling points?

**Solution.** \( \text{CHCl}_3 \) has a stronger dipole-dipole force among its dimers because Cl is more electronegative than Br. However, the difference in electronegativities \( \Delta E_N = 3.0 - 2.5 = 0.5 \) for C-Cl versus \( \Delta E_N = 2.8 - 2.5 = 0.3 \) for C-Br is small. Thus, differences in the dipole-dipole interaction strength are small (0.3 vs 0.5). \( \text{CHBr}_3 \) has a stronger VDW force because Br is a heavier atom than Cl and has more electrons. But \( \text{CHBr}_3 \) has the higher boiling point, so differences in VDW interactions must be stronger than differences in the d-d interaction.

**Problem 35.** Calculate the mean-square displacement (msd) \( \overline{x^2} \), for a freely diffusing particle (no boundaries). The average of \( x^2 \) is defined as \( \overline{x^2} = \int_{-\infty}^{\infty} x^2 p_t(x)dx \), where \( p_t(x) \) is the PDF that solves the diffusion equation:

\[
p_t(x) = \frac{1}{\sqrt{4\pi D t}} \exp \left( -\frac{x^2}{4Dt} \right)
\]

The product \( p_t(x)dx \) is the probability that a particle will be found in the interval \([x, x + dx]\) at time \( t \) when it was initially at \( x = 0 \) at time \( t = 0 \). Using the formula for Gaussian integration I gave you in class you can easily check that this PDF is normalized, i.e. \( \int_{-\infty}^{\infty} p_t(x)dx = 1 \). Show by direct calculation that \( \overline{x^2} = 2Dt \). Note: the reason we don’t use \( \overline{x} \) to quantify the displacement of particles undergoing diffusion is that \( \overline{x} \) equals zero and is therefore trivial and non-informative. We instead use the second moment \( \overline{x^2} \),
2.17. Problems

which is nonzero because of the square. To get units of length we can take the square root, \( \sqrt{x^2} \), which is called the “root-mean-square” displacement.

**Solution.** By definition,

\[
\bar{x}^2 = \int_R \frac{1}{\sqrt{4\pi Dt}} x^2 \exp\left(-\frac{x^2}{4Dt}\right) \, dx.
\]

Let \( a = \frac{1}{4Dt} \), and observe that (using formula for Gaussian integral provided in class)

\[
\int_R x^2 e^{-ax^2} \, dx = -\frac{d}{da} \int_R e^{-ax^2} \, dx = -\frac{d}{da} \frac{\sqrt{\pi}}{a} = \frac{\sqrt{\pi}}{2} = \frac{\sqrt{\pi(4Dt)^3}}{2}.
\]

Thus,

\[
\bar{x}^2 = \frac{1}{\sqrt{4\pi Dt}} \frac{\sqrt{\pi(4Dt)^3}}{2} = 2Dt.
\]

**Problem 36.** The dipole moment \( \vec{p} = q\vec{d} \) is the product of the charge \( q \) and the displacement vector \( \vec{d} \) pointing from positive to negative charge. Use vector addition and the diagram below to calculate the dipole moment for a water molecule \( \vec{p}_{\text{H}_2\text{O}} \) given that the O-H\(^+\) bond moment \( |\vec{p}_{\text{O-H}^+}| = 1.51 \, \text{D} \).

\[
\begin{align*}
\text{O} & \quad \text{H} \\
\text{H} & \quad \text{O}
\end{align*}
\]

\[104.5^\circ\]

**Solution.**

\[
|\vec{p}_{\text{H}_2\text{O}}| = 2 \, |\vec{p}_{\text{O-H}^+}| \cos\left(\frac{1}{2} \theta\right) = 2 \times (1.51 \, \text{D}) \cos(52.25^\circ) = 1.85 \, \text{D}
\]

**Problem 37.** The spherical coordinate expression for the dipole-dipole interaction energy in the point dipole approximation is (Section 2.4.2):

\[
V(R, \theta_1, \theta_2, \phi) = -\frac{|\vec{p}_A||\vec{p}_B|}{4\pi \epsilon_0 \epsilon R^3} \left[2 \cos \theta_1 \cos \theta_2 - \sin \theta_1 \sin \theta_2 \cos \phi\right]
\]

where \( \vec{p}_A \) and \( \vec{p}_B \) are the dipole moments of dipoles \( A \) and \( B \), respectively. They are vectors with length and orientation. \( |\vec{p}_A| \) and \( |\vec{p}_B| \) are the corresponding magnitudes (strengths) of the dipole moments. \( \epsilon_0 = 8.854 \times 10^{-12} \, \text{F m}^{-1} \) is the vacuum permittivity, \( \epsilon \) is the dimensionless dielectric constant of the solvent (\( \epsilon = 1 \) in the case of vacuum) and \( R \) is the inter-dipole center-to-center separation length (assuming “point dipoles”,}
i.e. infinitely small dipoles). The angles $\theta_1$, $\theta_2$ and $\phi$, which describe the relative orientations of the dipoles, are illustrated in Fig. 2.19.

(a) Determine the expression for the maximum attraction energy.

(b) Consider the situation when the dipoles are aligned anti-parallel to one another (still undergoing attractive interaction). Compare this energy to your answer in part (a).

(c) Are dipole-dipole forces alone strong enough to hold pure polar liquids of dipole moment 1 D together without the contribution of any other intermolecular forces? Compare your answer to part (a) or (b) at typical liquid-phase inter-dipole separations of 0.29 - 0.36 nm to room temperature thermal energy $k_B T$ at 298 K.

Solution. (a) Maximum attraction occurs when $\theta_1 = \theta_2 = 0$. (Please note that attraction here is twice as strong as if we had $\theta_1 = 90^\circ$, $\theta_2 = 180^\circ$.)

$$V(R, 0, 0, \phi) = -\frac{|\vec{p}_A| |\vec{p}_B|}{(2\pi\epsilon_0 \epsilon) R^3}$$

The negative sign denotes attractive force.

(b) If the dipoles are anti-parallel, $\theta_1 = 90^\circ$ and $\theta_2 = 180^\circ$, the interaction energy is ($\phi = 0$):

$$V(R, 0, 0, 0) = -\frac{|\vec{p}_A| |\vec{p}_B|}{(4\pi\epsilon_0 \epsilon) R^3}$$

which is still attractive because of the negative sign. Except that it differs by a factor of 2 from the expression in a.

(c) Yes, for strongly polar liquids of high dielectric constant.

Problem 38. In Section 2.10 we defined the thermally averaged energy by angle-averaging the Boltzmann factor corresponding to some interaction energy $V(R, \Omega)$:

$$e^{-V(R)/k_B T} \equiv \left< e^{-V(R,\Omega)/k_B T} \right> = \frac{\int e^{-V(R,\Omega)/k_B T} d\Omega}{\int d\Omega}$$

In spherical coordinates the solid angle element is $d\Omega = \sin \theta \, d\theta \, d\phi$, where $\theta, \phi$ are polar and azimuthal angles, respectively. Suppose we are interested in the low-temperature limit. Let us write $V(R, \Omega_0)$ for the lowest energy configuration (ground state energy). We then write the remaining energies as differences from the lowest energy: $\Delta V(R, \Omega) = V(R, \Omega) - V(R, \Omega_0)$. The ground state energy corresponds to $\Delta V = 0$. Then,

$$e^{-V(R)/k_B T} = e^{-V_0(R,\Omega_0)/k_B T} \frac{\int e^{-\Delta V(R,\Omega)/k_B T} d\Omega}{\int d\Omega} \sim \text{const.} \times e^{-V(R,\Omega_0)/k_B T},$$
since when \( T \to 0 \), only the term \( e^{-\Delta V(R,\Omega_0)/k_BT} = 1 \) contributes to the integral. The value of the integral depends on how fast \( \Delta V \) falls to zero as function of \( \Omega \). Take log of both sides:

\[
-V(R)/k_BT = \log(\text{const}) - V(R,\Omega_0)/k_BT.
\]

In the low temperature limit, the term \( \log(\text{const}) \) is negligible compared to the \( 1/T \) terms and we obtain the important result that the thermally averaged intermolecular energy is simply the ground state energy:

\[
V(R) = V(R,\Omega_0) \text{ as } T \to 0.
\]

(a) Let us take the case of the charge-dipole interaction. The interaction energy is

\[
V(R, \Omega) = -\frac{Qp \cos \theta}{4\pi\epsilon_0 R^2},
\]

where \( Q \) is the electric charge (in Coulombs), \( p \) is the strength of the dipole moment, \( R \) is the charge-dipole separation, \( \theta \) is the angle of the dipole tilt with respect to the charge-dipole axis, as shown in Fig. 2.17. Using the procedure outlined above, calculate the thermally averaged charge-dipole interaction strength in the low temperature limit.

(b) The high temperature regime occurs when \( V(R,\Omega) < k_BT \). Calculate the charge-dipole interaction at high temperatures by expanding

\[
e^{-V(R)/k_BT} = \langle e^{-V(R,\Omega)/k_BT} \rangle = \frac{\int e^{-V(R,\Omega)/k_BT} d\Omega}{\int d\Omega}
\]

Explain what happens to \( V(R) \) in the limit \( T \to \infty \).

(c) The Boltzmann-averaged dipole-dipole interaction is known as the Keesom interaction and it is attractive and temperature dependent. Recall the dipole-dipole interaction potential for static dipoles can be written as (see Fig. 2.19):

\[
V(R, \theta_1, \theta_2, \phi) = -\frac{|\vec{p}_A||\vec{p}_B|}{4\pi\epsilon_0\epsilon R^3} [2 \cos \theta_1 \cos \theta_2 - \sin \theta_1 \sin \theta_2 \cos \phi]
\]

Explain the behavior of the interaction when \( k_BT \) is larger than the dipole-dipole interaction energy.

(d) Calculate the thermal energy \( k_BT \) at 298 K in eV and J.

(e) What is the inter-dipole separation at which orientational order breaks down for nitrobenzene in carbon tetrachloride? \( \epsilon_{\text{carbon tetrachloride}} = 2.24 \),
$|\vec{p}_{\text{nitrobenzene}}| = 4.2 \, \text{D}$ and at $T = 298 \, \text{K}$.

(f) What is the value of the Keesom interaction energy at the high temperature limit (i.e., as $T \to \infty$)?

(g) Calculate the dipole-dipole interaction energy at the low temperature limit.

(h) Consider the $1/R^2$ vs. $1/R^4$ dependence of the charge-dipole interaction energy at the low-$T$ vs. high-$T$ limits, respectively. Also consider the dipole-dipole interaction at its low-$T$ limit ($1/R^3$) and high-$T$ limit ($1/R^6$). What is the significance of these differences in the $R$-dependence when $T$ varies? Why is the interaction becoming more short-ranged in the high-$T$ limit?

**Solution.** (a) Since $V(R) = V(R, \Omega_0)$, we need to find $V(R, \Omega_0)$, the ground state energy. Suppose $Q > 0$. In that case, the minimum of $V(R)$ with respect to $\theta$ is found by inspection of the expression for $V(R, \Omega)$, which shows that the minimum value occurs when $\cos \theta = 1$, or when $\theta = n\pi$, $n = 0, 1, 2, \ldots$. Then, $V(R, \Omega_0) = -(Qp)/(4\pi\epsilon_0 R^2)$. If $Q < 0$, we would need $\cos \theta = -1$, which would occur when $\theta = \pi(2n + 1)$, $n = 0, 1, 2, 3, \ldots$.

b) Since we are in the high-$T$ limit, we may Taylor-expand both sides

$$e^{-V(R)/k_B T} = 1 - \frac{V(R)}{k_B T} + \cdots = \left< 1 - \frac{V(R, \Omega)}{k_B T} + \frac{1}{2} \left( \frac{V(R, \Omega)}{k_B T} \right)^2 - \cdots \right>$$

Cancelling out the 1’s on both sides and multiplying out by $k_B T$ leaves us with:

$$V(R) = \left< V(R, \Omega) - \frac{V(R, \Omega)^2}{2k_B T} + \cdots \right>$$

Substituting the charge-dipole interaction for $V(R, \Omega)$

$$V(R) = \left< -\frac{Qp \cos \theta}{4\pi\epsilon_0 R^2} - \frac{(Qp \cos \theta)^2}{4\pi\epsilon_0 R^2} \frac{\cos^2 \theta}{2k_B T} + \cdots \right>$$

The first term vanishes since the average $\langle \cos \theta \rangle = 0$. The second term survives:

$$V(R) \approx -\frac{Q^2 p^2}{6(4\pi\epsilon_0)^2 (k_B T) R^4}$$

where we used the fact that $\langle \cos^4 \theta \rangle$ is a nonzero constant (you can easily check that $\int_1^{-1} x^4 dx$ is nonzero by symmetry of $x^4$, where $x = \cos \theta$). When $T \to \infty$, the interaction energy $V(R)$ vanishes. As the temperature of the system increases, more molecules have sufficient energy to occupy the less favorable configurations. The higher, less favorable, configurations are those
that give less favorable interactions between the dipoles (i.e., higher potential energy configurations).

(c) At thermal energies $k_B T$ greater than the dipole-dipole interaction energy, the Keesom interaction averages over all angles and dipoles no longer have orientational preference to one another. The angle-averaged Keesom interaction potential and its condition is given below.

$$V(R)_{\text{Keesom}} = -\frac{\left|\vec{p}_A\right|^2 \left|\vec{p}_B\right|^2}{3(4\pi\epsilon_0\epsilon)^2 (k_B T) R^6} \quad \text{for} \quad k_B T > \frac{\left|\vec{p}_A\right|\left|\vec{p}_B\right|}{(4\pi\epsilon_0\epsilon) R^3}$$

(d) 

$$k_B T = (1.38 \times 10^{-23} \text{ J/K})(298 \text{ K}) = 4.11 \times 10^{-21} \text{ J}$$

$$4.11 \times 10^{-21} \text{ J} \left(\frac{1 \text{ eV}}{1.602 \times 10^{-19} \text{ J}}\right) = 0.0257 \text{ eV}$$

(e) 

$$R = \left(\frac{\left|\vec{p}\right|^2}{(4\pi\epsilon_0\epsilon)k_B T}\right)^{1/3}$$

$$= \left(\frac{(3.34 \times 10^{-30} \text{ C.m.D}^{-1})(4.2 \text{ D})^2}{4\pi(8.85 \times 10^{-12} \text{ F m}^{-1})(2.24)(1.38 \times 10^{-23} \text{ J K}^{-1})(298 \text{ K})}\right)^{1/3}$$

$$= 0.577 \text{ nm}$$

(f) As $T \to \infty$, $V(R)_{\text{Keesom}} \to 0$.

(g) In the low temperature limit, the energy is just the ground state energy. This is the result obtained in (a).

(h) As discussed in lecture, in the high-$T$ limit, there is an equal mix of ground and excited state configurations. When these average-out the interaction, it leads to partial cancellation and weakening of the interaction. This partial cancellation takes the long-range interaction $(1/R^3)$ and turns it into a short-range $(1/R^6)$ interaction.

**Problem 39.** Model the dissolution of NaCl in water by comparing the lattice energy of an NaCl crystal to the number of ion-dipole interactions of Na$^+$ and Cl$^-$ with the dipole of water that it would take to equal this energy according to the steps below.

(a) Calculate the ion-ion interaction energy between sodium cation and chloride anion by taking $R = 0.282$ nm as the distance between the ion centers (an experimental value determined by crystallography) in the NaCl crystal lattice and using the Coulomb interaction (ion-ion) below. Assume vacuum so $\epsilon = 1$. The variable $z$ refers to the charge valency of the ion (e.g., $z = 1$ for Cl$^-$, $z = 2$ for Ca$^{2+}$, $z = 3$ for Fe$^{3+}$, etc.). $z$ times the elementary
charge of the electron, $e$, gives the charge of the ion. The experimentally
determined value for the NaCl lattice energy is 787 kJ/mol. How close does
your estimate compare to the experimental lattice energy? Propose reasons
why your answer may deviate from the experimental value.

$$V(R)_{\text{ion-ion}} = \frac{z_1 z_2 e^2}{4\pi\epsilon_0\epsilon R}$$

(b) Calculate how many molecules of water it would take to equal the ex-
pensively determined lattice energy of 787 kJ/mol for an NaCl crystal
based entirely on ion-dipole interactions between water, Na$^+$ and Cl$^-$. Use
the value $R = 0.354$ nm, the average distance between the center of sodium
or chloride ions and the water dipole, in your calculation. Take the angle $\theta$
to be equal to $0^\circ$, the ion-dipole orientation that yields a maximum value for
the potential energy. Assume vacuum so $\epsilon = 1$ for a more direct comparison
to the lattice energy conditions. Note that in the diagram, $Q$ represents the
magnitude of the charge participating in the charge-dipole interaction. For
the case of monovalent ions such as Na$^+$ and Cl$^-$, $ze = Q$ where $z$
represents ion valency and $e$ represents the elementary charge $e = 1.602 \times 10^{-19}$ C. See
Fig. 2.17 for geometry.

$$V(R)_{\text{ion-dipole}} = -\frac{ze|\vec{p}|\cos \theta}{4\pi\epsilon_0\epsilon R^2}$$

(c) Convert the solubility limit of NaCl in water (36.0 g/100. mL) to a mole
fraction and compare your answer to the mole fraction of 1 mol NaCl per
mol H$_2$O from your answer to part (b). How close was your estimate to the
experimental value?

Solution. (a)

$$V(R)_{\text{ion-ion}} = \frac{(1)(1)(1.602 \times 10^{-19} \ C)^2(6.022 \times 10^{23} \ \text{mol}^{-1})}{4\pi(8.854 \times 10^{-12} \ \text{F m}^{-1})(1)(0.282 \times 10^{-9} \ \text{m})}$$

$$= 493 \ \text{kJ/mol}$$

This answer is not bad for a first order estimate but deviates significantly
due to the long-range nature of Coulombic interactions. Ions other than
nearest neighbors can interact with one another and must be considered. In
more rigorous models of lattice energies, the term that accounts for this long-
range interaction is called the Madelung constant. Furthermore, repulsions
between ions of the same charge in the lattice are also significant.

(b)

$$\frac{n_{\text{H}_2\text{O}}}{n_{\text{NaCl}}} \times V(R)_{\text{ion-dip.}} = V(R)_{\text{ion-ion}}$$
where $V(R)_{\text{ion-dip.}}$ evaluates to:

$$
= - \frac{ze|\vec{p}| \cos \theta}{4\pi\epsilon_0\epsilon R^2}
= - \frac{(1)(1.602 \times 10^{-19} \text{ C})(6.17 \times 10^{-30} \text{ C.m})(1)(6.022 \times 10^{23} \text{ mol}^{-1})}{4\pi(8.854 \times 10^{-12} \text{ F.m}^{-1})(1)(0.354 \times 10^{-9} \text{ m})^2}
= - 42.7 \text{ kJ/mol}
$$

Considering only the absolute magnitudes of the energies and solving for moles of water

$$\frac{n_{\text{H}_2\text{O}}}{n_{\text{NaCl}}} = \frac{787 \text{ kJ/mol}}{42.7 \text{ kJ/mol}} = 18.4$$

We find that it would take 18 molecules of water to dissolve 1 molecule of NaCl. In other words, 9 molecules of water per ion. Please note that in this problem and in the previous one, we have assumed that $\epsilon = 1$ (vacuum). We would get different results if we used the dielectric constant of each respective medium.

(c) \[
\frac{36.0 \text{ g NaCl}}{100. \text{ mL H}_2\text{O}} \approx \frac{36.0 \text{ g NaCl}}{100. \text{ g H}_2\text{O}}
\]
\[
\frac{36.0 \text{ g NaCl}}{100. \text{ g H}_2\text{O}} \left( \frac{18.02 \text{ g/mol H}_2\text{O}}{58.44 \text{ g/mol NaCl}} \right) = 0.111
\]

Taking the reciprocal of this mole fraction gives a number of 9 moles of water for every mole of NaCl at the solubility limit which is surprisingly close to the coarse estimate of 18.4 from part (b). Please note that this assumes ions are homogeneously and randomly distributed in solutions. This is not the case (ions of opposing charge associate loosely, and there is also an ionic atmosphere forming).

\[\blacksquare\]

**Problem 40.** At $90^\circ\text{C}$, the vapor pressure of methylbenzene is 400. Torr and that of 1,2-dimethylbenzene is 150. Torr. What is the composition of the liquid mixture that boils at $90^\circ\text{C}$ when the pressure is 0.50 atm? Calculate the composition of the vapor produced.

**Solution.** Combine Dalton’s law and Raoult’s law.

$$P_{\text{total}} = P_A + P_B = X_A P_A^\circ + (1 - X_A) P_B^\circ$$

$P_{\text{MB}}^\circ = 400$. Torr

$P_{1,2-\text{DMB}}^\circ = 150$. Torr

Solve for $X_{\text{MB}} = X_A$

$$X_A = \frac{P_{\text{total}} - P_B^\circ}{P_A^\circ - P_B^\circ}$$
Boiling occurs when

\[ P_{\text{total}} = 0.50 \text{ atm} \left( \frac{760. \text{ Torr}}{1 \text{ atm}} \right) = 380 \text{ Torr} \]

Substituting

\[ X_A = \frac{380 \text{ Torr} - 150 \text{ Torr}}{400 \text{ Torr} - 150 \text{ Torr}} = 0.920 \]

\[ X_{1,2-\text{DMB}} = X_B = 1 - 0.920 = 0.080 \]

Vapor phase composition is given by

\[ Y_{MB} = Y_A = \frac{X_A P_A^o}{P_B^o + (P_A^o - P_B^o) X_A} \]

\[ = \frac{(0.920)(400 \text{ atm})}{150 \text{ atm} + (400 \text{ atm} - 150 \text{ atm})(0.920)} = 0.968 \]

\[ Y_{1,2-\text{DMB}} = Y_B = 1 - 0.968 = 0.032 \]

**Problem 41.** Isothermal compressibility is a material property defined by:

\[ \kappa = -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T \]

(a) Calculate \( \kappa \) for an ideal gas.

(b) Calculate \( \kappa \) for a VDW gas.

**Solution.** (a)

\[ V = \frac{nRT}{P}, \quad \kappa = -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T = -\frac{P}{nRT} \left( -\frac{nRT}{P^2} \right) = \frac{1}{P} \]

(b)

\[ \kappa = -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T = -\frac{1}{V} \left( \frac{\partial P}{\partial V} \right)_T^{-1}, \quad P = \frac{nRT}{V - nb} - \frac{an^2}{V^2} \]

\[ \left( \frac{\partial P}{\partial V} \right)_T = -\frac{nRT}{(V - nb)^2} + \frac{2an^2}{V^3} \]

\[ \kappa = -\frac{1}{V} \left( \frac{1}{\frac{nRT}{(V - nb)^2} + \frac{2an^2}{V^3}} \right) = \frac{(V - nb)^2 \cdot V^{-1}}{nRT - 2an^2(V - nb) \cdot V^{-3}} \]

**Problem 42.** Calculate the thermal expansion coefficient, \( \alpha = \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_P \), for ideal and VDW gas.

**Solution.** Idea gas was done in class, \( \alpha = 1/T \). Regarding \( V \) as a function of temperature, \( V = V(T) \), we can solve for \( \left( \frac{\partial V}{\partial T} \right)_P \) by differentiation:

\[ \frac{\partial}{\partial T} \left[ (P + an^2/V^2)(V - nb) = nRT \right]_P \]
\[
\frac{\partial}{\partial T} \left[ PV - Pnb + a \frac{n^2}{V} - ab \frac{n^3}{V^2} = nRT \right]_p \\
\left( \frac{\partial V}{\partial T} \right)_p \left( P - a \frac{n^2}{V^2} + 2ab \frac{n^3}{V^3} \right) = nR \\
\left( \frac{\partial V}{\partial T} \right)_p = \frac{nR}{P - a \frac{n^2}{V^2} + 2ab \frac{n^3}{V^3}}
\]

and then divide by \( V \) to arrive at the result.

\[
\alpha = \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_p = \frac{nR}{PV - a \frac{n^2}{V} + 2ab \frac{n^3}{V^2}}
\]
Properties of Matter

The bulk properties of gases, liquids and solids — molar volume, density, compressibility, and thermal expansion, among others — differ widely, often by orders of magnitude. All of these properties depend on temperature and pressure. Gases, liquids and solids differ by their local structure. We have outlined the most important forces responsible for interactions between molecules. These forces help determine the structure of solids and liquids, solubility limits for different solutes, surface tension, boiling points, melting points, etc.

3.1. Structure of Liquids, Solids and Gases

Figure 3.1 illustrates the packing structure of solids (left), liquids (middle) and gases (right). As the Figure illustrates, gases are isolated molecules. Liquids are packed molecules, but the packing shows little ordering. Solids can be ordered.

![Figure 3.1. Structure of solids, liquids and gases.](image-url)
3.1.1. Density. Liquids and solids are dense: their density is about 1,000 times larger than gases. However, they differ in their ordering: solids are often crystalline whereas liquids are usually disordered. Solids can also be disordered (amorphous). Gases are dilute and disordered.

3.1.2. Compressibility. Gases are compressible whereas liquids and solids are said to be incompressible. (Liquids and solids can be compressed slightly, but much less than gases can be compressed.) The isothermal compressibility $\kappa$ of a substance is the change in its volume in response to a change in the applied pressure under conditions of constant temperature. The formal definition is:

$$
\kappa = -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T \approx -\frac{1}{V} \left( \frac{\Delta V}{\Delta P} \right)_T
$$

where the approximation is good in the limit of small $\Delta P$ and $\Delta V$. This says that $\kappa$ is the (negative) fractional change in volume in response to a small pressure change. Its unit are inverse pressure. The value of $\kappa$ for a real gas will depend on molecular properties such as molecular size and intermolecular interactions: for example, we expect that $\kappa$ decreases with molecular size but increases as intermolecular interaction strength increases.

On the other hand, let us take the ideal gas. Since $PV = nRT$, $V = nRT/P$. Thus, the derivative with respect to $P$ is $\partial V/\partial P = -nRT/P^2$. Then, $\kappa = -(1/V)(-nRT/P^2)$, and inserting $V = nRT/P$ gives:

$$
\kappa = \frac{1}{P}
$$

Thus, the lower the pressure, the higher the compressibility. The higher the pressure (in the container) the less compressible the gas is. Also, note that $\kappa$ is independent of the gas type, as expected for an ideal gas.

Homework problem: calculate the isothermal compressibility of a VDW gas and compare with the case of an ideal gas.

3.1.3. Mean Free Path. The mean free path is defined as the average distance traveled by a molecule (or atom/particle) between successive collisions. In Section 1.5.2 we derived an expression for the mean free path in gases using arguments from kinetic theory. The collision causes the particles/molecules to change direction or energy. If the collisions are elastic, total momentum and total energy are conserved quantities. At ambient pressure, there are approximately $10^{19}$ molecules/cc and the mean free path is on the order of 70 nm. The mean free path is inversely proportional to the gas density.
In a gas, the mean free path ($\lambda$) is much larger than the molecular diameter $d$: $\lambda \gg d$. In a liquid, the mean free path is less than $d$: $\lambda < d$. In a solid there exists no concept of mean free path.

### 3.1.4. Volume

In a gas, the volume is inversely proportional to the pressure: $V \propto P^{-1}$. In a liquid or solid, the volume remains mostly constant.

### 3.1.5. Thermal Expansion

When a material such as a solid is heated it will expand because of the increase in interatomic distances within the lattice as a result of increased thermal vibrations (phonons). This is quantified by the coefficient of thermal expansion:

$$\alpha = \left. \frac{1}{V} \frac{\partial V}{\partial T} \right|_P \approx \left. \frac{1}{V} \frac{\Delta V}{\Delta T} \right|_P,$$

where the approximation holds in the limit of small $\Delta T$ and $\Delta V$. The subscript $P$ indicates constant pressure. For solids and liquids, this coefficient is on the order of 0.02% per degrees Celsius, which is much smaller than in the case of a gas.

### 3.1.6. Diffusion Coefficient

In Section 1.5.2 we derived an expression for the diffusion coefficient of gases using arguments from kinetic theory. In a gas, molecular diffusion is fast: $D \sim 0.2 \text{ cm}^2/\text{s}$. In a liquid, the diffusion is moderate: $D \sim 2 \times 10^{-5} \text{ cm}^2/\text{s}$. In a solid, diffusion is slow: $D \sim 2 \times 10^{-9} \text{ cm}^2/\text{s}$ for hydrogen in iron or $D \sim 10^{-30} \text{ cm}^2/\text{s}$ for aluminum in copper.

Note: “aluminum in copper” refers to the diffusion of aluminum atoms in a copper lattice. When specifying diffusion coefficients one must indicate the diffusing substance and the surrounding matrix. If only one substance is specified, this refers to the “self-diffusion” coefficient, i.e. how fast atoms/molecules move within the same substance.

### 3.2. Effusion

Effusion is the flow of gas out of a container through a tiny hole (pinhole) as illustrated in Fig. 3.2. The effusion rate can be computed from the rate of collisions at the wall. The latter is proportional to:

$$\frac{N}{V} \propto \sqrt{\frac{8k_B T}{\pi m}} A.$$

Suppose that a container of volume $V$ is held at temperature $T$ and contains a mixture of two gases $A$ and $B$ ($N_A$ molecules of $A$ and $N_B$ molecules of $B$). This leads to:

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{N_A}{N_B} \sqrt{\frac{m_B}{m_A}}.$$
This is Graham’s law of effusion. A rigorous derivation requires arguments from the kinetic theory of gases. It was used in the past as a technique to separate isotopes by consecutive stages of enrichment. This technique was later replaced by the gas centrifuge method where the centrifugal force creates a gradient of molecules/atoms of different weights. Nowadays, the technique of laser isotope separation is the method of choice because of its much higher efficiency.

3.3. Phase Equilibrium

More than one phases (solid, liquid, gas) can co-exist. Suppose we evacuate a flask and put liquid water in it, with the temperature held at 25°C. We then monitor the pressure of water vapor inside the flask. At time $t = 0$, the pressure of water vapor begins to rise from zero. It increases with time and gradually levels off at a value of 0.03126 atm, which is the vapor pressure of water at 25°C. The contents of the flask have reached equilibrium, a condition in which no further changes in macroscopic properties occur as
long as the system remains isolated. If some of the water vapor that has formed is removed, additional water evaporates from the liquid to reestablish the same vapor pressure, $P_{vap}(H_2O)=0.03126$ atm.

![Figure 3.4. Schematic representation of the dynamics illustrated in Fig. 3.3, at three different points in time: (left) $t=0$, (middle) $t=\text{intermediate}$ and (right) $t=\infty$.](image)

What is happening on a microscopic scale to cause this spontaneous movement of the system toward equilibrium? According to kinetic theory, the molecules of water in the liquid are in a constant state of thermal motion. Some of those near the surface are moving fast enough to escape the attractive forces holding them in the liquid; this process of evaporation causes the pressure of the water vapor to increase. As the number of molecules in the vapor phase increases, the reverse process begins to occur: molecules in the vapor strike the surface of the liquid, and some are captured, leading to condensation. As the pressure of the gas increases, the rate of condensation increases until it balances the rate of evaporation from the surface (see Figs. 3.3 and 3.4). Once this occurs, there is no further net flow of matter from one phase to the other; the system has reached phase equilibrium, characterized by a particular value of the water vapor pressure. Water molecules continue to evaporate from the surface of the liquid, but other water molecules return to the liquid from the vapor at an equal rate. A similar phase equilibrium is established between an ice cube and liquid water at the freezing point.

Vapor pressure depends on temperature and type of substance, as shown in Fig. 3.5. For fixed volume and number of moles, the temperature at which the vapor pressure equals 1 atm defines the **normal boiling point of a liquid** and the **normal sublimation point of a solid**.

**3.4. Phase Transitions**

Suppose 1 mol of gaseous sulfur dioxide is compressed at fixed temperature of 30°C. The volume of the substance is measured at each pressure, and a
graph of volume against pressure is constructed (Fig. 3.6). At low pressures, the graph shows the inverse dependence \( V \propto 1/P \) predicted by the ideal gas law. As the pressure increases, deviations appear because the gas is not ideal.

For pressures up to 4.52 atm, this behavior is quite regular and can be described by the VDW equation. At 4.52 atm, something dramatic occurs: The volume decreases abruptly by a factor of 100 and remains small as the pressure is increased further. The gas underwent a phase transition from gas to liquid as a result of the applied pressure. If the compression of SO\(_2\) is continued, another abrupt (but small) change in volume will occur as the liquid freezes to form a solid.
3.5. Phase Diagrams

Condensed phases also arise when the temperature of a gas is reduced at constant pressure. If steam (water vapor) is cooled at 1 atm pressure, it condenses to liquid water at 100°C and freezes to solid ice at 0°C. Liquids and solids form at low temperatures once the attractive forces between molecules become strong enough to overcome the kinetic energy of random thermal motion.

Six phase transitions occur among the three states of matter (see Fig. 3.7).

![Phase Diagrams](image)

**Figure 3.7.** The most commonly encountered phase transitions in chemistry describe structural transitions between the liquid, solid and gas states.

3.5. Phase Diagrams

If the temperature of a substance is held constant and the applied pressure is changed, phase transitions between two phases will be observed at particular pressures. Making the same measurements at a number of different temperatures provides the data necessary to draw the phase diagram for that substance – a plot of pressure against temperature that shows the stable state for every pressure – temperature combination. Figure 3.8 shows a sketch of the phase diagram for water. For each substance there is a unique combination of pressure and temperature, called the triple point (marked “T”), at which the gas, liquid, and solid phases coexist in equilibrium. Extending from the triple point are three lines, each denoting the conditions for the coexistence of two phases at equilibrium. Along the line TA, solid and gas are in equilibrium; along TB, solid and liquid; and along TC, liquid and gas. The regions bounded by these lines represent conditions where only one phase exists.

The P-T diagram for H₂O at fixed volume is shown in Fig. 3.8.

Notice how the solid-liquid coexistence line tilts leftward. This is a result of the anomalous properties of water (owing to its abilities for hydrogen
bonding and directionality of the bonding); for H\textsubscript{2}O, the ice phase is less dense than the liquid phase. Other substances tilt towards the right, because the solid phase is denser than the liquid phase.

A P-V-T diagram may look like the one shown in Fig. 1.5. Note: This surface plot is definitely not for water, because the solid-liquid coexistence line does not tilt leftward. To learn more about the anomalous properties of water, see: http://www1.lsbu.ac.uk/water/water_anomalies.html

The gas-liquid coexistence curve extends upward in temperature and pressure from the triple point. This line (similar to a parabola) is the vapor pressure curve of the liquid substance. The gas-liquid coexistence curve does not continue indefinitely, but instead terminates at the critical point (labelled). Along this coexistence curve there is an abrupt, discontinuous change in the density and other properties from one side to the other. The differences between the properties of the liquid and the gas become smaller as the critical point is approached and disappear altogether at that point.

For pressures above the critical pressure (218 atm for water), it is no longer possible to identify a particular state as gas or liquid. A substance beyond its critical point is called a supercritical fluid because the term fluid includes both gases and liquids. The liquid-solid coexistence curve does not terminate as the gas-liquid curve does at the critical point, but continues to indefinitely high pressures. In practice, such a curve is almost vertical because large changes in pressure are necessary to change the freezing temperature of a liquid. For most substances, this curve inclines slightly to the
3.6. Solutions

Solutions are homogeneous systems that contain two or more substances. Usually, it consists of a solvent (liquid) in which we have dissolved a second substance (solid, gas or liquid). “Solid solutions” are also possible in which more than two different substances are mixed, such as bismuth and selenium.¹ The host substance is called the solvent and the added substance is the solute.²

3.6.1. Composition: Mole Fraction, Molarity and Molality. The mole fraction, $X_1$, of a substance, 1, in a mixture is the number of moles of that substance, $n_1$, divided by the total number of moles of all substances.

¹In solid solutions, it is more difficult to maintain homogeneity of the system, depending on the degree of miscibility of the constituent substances. We often observe a certain amount of phase separation.
²The usual convention is to label as “solvent” the most abundant substance.
present \( n_1 + \cdots + n_N \):

\[
X_1 = \frac{n_1}{n_1 + n_2 + \ldots + n_N}
\]

where \( N \) is the number of substances present (including the solvent) and \( n_i \) is the number of moles of the \( i \)-th substance.

The concentration of a substance is the number of moles per unit volume. The basic unit is molarity, defined as number of moles of solute per liter of solution:

\[
\text{molarity} = \frac{\text{moles solute}}{\text{liters solution}} = \text{mol L}^{-1} \equiv M.
\]

Because volume depends on temperature, some prefer to use molality, which is defined as the number of moles of solute per kilogram of solvent:

\[
\text{molality} = \frac{\text{moles solute}}{\text{kilograms solvent}} = \text{mol kg}^{-1}.
\]

We note that at 20° C the two concepts are identical for water, since its density is 1.00 g/cm³ (so that 1.00 L of water weighs 1.00 kg).

3.6.2. Preparing Solutions. Suppose that a solution of 1 M (molar) concentration is prepared. One possible method is to weigh 1 mol of solute, then add slightly less than 1 L of solvent (to avoid overshooting our target volume of 1 L). We then add solvent continuously until we reach a final volume of exactly 1 L.

In chemistry, we often prepare “stock solutions” which are highly concentrated, then dilute the stock solution to obtain the desired concentration. Suppose that the stock solution contains \( n_1 = c_1 V_1 \) moles of solute, where \( c_1 \) is the concentration of solute and \( V_1 \) is the volume of stock solution. To get a solution of concentration \( c_2 \) from this, all we need to do is realize that if we dilute the solution, it will still contain \( n_1 \) moles of solute. Thus, \( n_1 = n_2 \), where \( n_2 = c_2 V_2 \). So, we may write:

\[
c_2 = \frac{\text{moles of solute}}{\text{final solution volume}} = \frac{c_1 V_1}{V_2}.
\]

3.7. Non-Electrolyte Solutions

In organic and inorganic chemistries, different solvents than water are normally used. Here, we shall mainly discuss solutions of the aqueous type. When dissolving a solute, such as \( \text{C}_{12}\text{H}_{22}\text{O}_{11} \), in water, we write it as \( \text{C}_{12}\text{H}_{22}\text{O}_{11}(aq) \) to mean that the environment is aqueous. Solutions can be classified as electrolytes and non-electrolytes, due to the very different nature of the intermolecular interactions.

Non-polar molecules such as hydrocarbons do not dissolve significantly in water. Hydrocarbons are hydrophobic and tend to phase-separate. Molecules
that are polar are readily dissolved in water. A typical example is sugar, \( C_m(H_2O)_n \). Some examples of common sugars are shown in Fig. 3.10.

![Figure 3.10. Pentoses.](image)

As can be seen in Fig. 3.10 these molecules include OH (hydroxyl) groups bonded to carbon atoms, which provide sites for hydrogen bonding interactions with water molecules. Figure 3.11 shows the example of the fructose molecule whose hydroxyl groups readily form a hydrogen bonding network with surrounding water molecules. Also shown is a hydrogen bond between a water molecule and an oxygen atom in the fructose ring.

![Figure 3.11. Hydrogen bonding between fructose and surrounding water molecules.](image)

### 3.8. Electrolyte Solutions

3.8.0.1. Dissolution of Electrolytes. In electrolyte solutions the ion-dipole solute-solvent bonding can be much stronger than the case of dipole-dipole interactions (or hydrogen bonding) for non-electrolyte solutions. Consider the dissolution of potassium sulfate:

\[
K_2SO_4(s) \rightarrow 2K^+(aq) + SO_4^{2-}(aq).
\]
The dissolution occurs through the ion-dipole forces. Each positive ion is surrounded by water molecules with the negative end of their dipole moments toward the positive ion. For most ionic solids, several water molecules are needed to pull an ion from the lattice. In Problem 30 (p. 102) you are asked to calculate how many water molecules are needed to dissolve Na\(^+\) and Cl\(^-\) ions from the NaCl lattice.

In Fig. 3.12, each SO\(_4^{2-}\) anion in solution is surrounded by water molecules oriented with the positive end of their dipole moments toward the anion. Likewise, each K\(^+\) ion is surrounded by water molecules oriented with the negative end of their dipole moments toward the cation.

3.8.0.2. Hydration Shell of Ion. The ability of salts to dissolve in water depends strongly on their stabilization by the solvent molecules. This process is known generally as solvation, and in the case of water, as hydration. All ions dissolved in water are hydrated (see Fig. 3.13).
In the standard model for hydrated ions (Fig. 3.13), the water molecules (red) closest to the negative ion (green) are the most tightly bound, and constitute the primary hydration shell. H$_2$O molecules in the outer hydration shell experience local ordering by hydrogen bonding, but this is opposed by the electrostatic field of the central cation. As one moves farther out and the electrostatic field due to the central ion weakens, local hydrogen bonded-clusters enlarge and begin to dominate, finally becoming the sole source of local ordering within the bulk water region of the solution.

The attractive forces that are operational here [hydrogen bonding (dipole-dipole) and even the stronger ion-dipole forces that bind the waters in the primary shell to the central ion] are continually subject to disruption by thermal motions. This affects even the waters within the primary shell; residence times of these H$_2$O molecules are usually less than $10^{-4}$ s.

3.9. Ionic Conduction

The phenomenon of electrical conductivity is illustrated in Fig. 3.14. Electrical charges experiencing an electric field $\vec{E}$ accelerate due to the Lorentz force, $\vec{F} = Q\vec{E}$, but are slowed down because of scattering$^3$. The net effect

\[ J = \sigma E \]

\[ I = JA \]

$^3$While electrons in the conduction band of a conductor do not interact with each other (the “free electron gas” model), they do scatter at the positively charged lattice sites via the Coulomb interaction.
is to eliminate acceleration, leaving charges that are moving at constant velocity. This is the statement of Ohm’s law, $\vec{J} = \sigma \vec{E}$, where $\sigma$ is the electrical conductivity of the medium in which charges flow. Multiplication of $\vec{J}$ by the cross-sectional area of the conductor, $A$, gives the total current, $\vec{I}$.

Suppose that our conductor has a square cross-sectional area, $A = L^2$, where $L$ is the side length. Multiplication of $\vec{J} = \sigma \vec{E}$ by $A$ gives $\vec{I} = (\sigma L)(\vec{E}L)$. Denoting $\vec{E}L$ as the voltage $V$ and $(\sigma L)$ as $R^{-1}$ (the inverse resistance). This gives $I = R^{-1}V$ or $V = RI$, which is the familiar form of Ohm’s law. The units of $R$ are Ohm (Ω). The inverse of conductivity, $\sigma^{-1}$, is called the resistivity. The units of $\sigma^{-1}$ are Ω.m.

The units of $\vec{I}$ are C/s; thus charges are flowing at constant rate as long as $\vec{E}$ remains constant. The sign convention is that positive charges move along $\vec{E}$ (and therefore, $\vec{I}$) whereas negative charges move in the opposite direction. In a metallic conductor, negative charges are electrons and positive charges are holes. In ionic solutions, positive charges are the positively charged ions whereas negative charges are the negatively charged ions.

The conduction of electricity through an ionic solution differs from metallic conduction in two fundamental ways:

- The current is associated with the transport of relatively large and massive hydrated ions, rather than by nearly weightless electrons. Compared to ions, electrons move largely unimpeded through the metal. But ions, with their closely-held waters of hydration and more diffuse secondary hydration shell and oppositely-charged counterions, must disrupt the local hydrogen-bonded water structure as they move through the solution.
- Transfer of electric charge into and out of the solution occurs at electrodes, and is accompanied by chemical reactions at these interfaces.

Electrolytic conduction involves the transport of electric charge in the form of hydrated ions. Movement of these ions in response to an electric potential gradient is known as migration. Transport of electric charge is measured in Ampère (A). 1 A $\equiv$ 1 C/s. Current ($\vec{I}$) is the flow of charge per unit time: $I = dQ/dt$, where $Q(t)$ describes the electric charge in a given region as function of time.

An ion of charge $Q$ undergoes a drift from cathode to anode in solution, under the action of an electric field $\vec{E}$. In chemistry, the convention is to call the cathode the electrode which attracts positive ions and the anode the electrode which attracts negative ions.\(^4\) See Fig. 3.15.

When a voltage (potential difference), $V = V(\text{cathode}) - V(\text{anode})$, is applied an electric field is created between the two electrodes. The $E$-field

\(^4\)In physics, the convention is reversed: the anode is the negatively charged electrode and the cathode is the positively charged electrode.
arrows point from the positive electrode to the negative electrode. The \( E \)-field lines are shown in Fig. 3.16.

This electric field gives rise to a force \( \mathbf{F} = Q \mathbf{E} \). While \( \mathbf{F} = m \mathbf{a} \), the ions accelerate under the \( E \)-field, but also slow down because of random thermal motion and viscous drag (collision with fluid molecules). The overall motion is that of a constant velocity rather than acceleration.

Using this sign convention, in Fig. 3.16, the anode (positive electrode) is on the right, the cathode (negative electrode) is on the left, and the \( E \)-field arrows between the two electrodes point from right to left (with \( E \)-field lines always originating from the + charges at the electrodes). According to \( \mathbf{F} = Q \mathbf{E} \), the negative ions \( (Q < 0) \) drift from left to right (direction opposite to the \( E \)-field), i.e. from the cathode to the anode.
Force over distance leads to work:  
\[ W(\text{work}) = \int_a^b \mathbf{F} \cdot d\mathbf{l} = Q \int_a^b \mathbf{E} \cdot d\mathbf{l} \]
where the integral is carried out along a path (line or curve) connecting a to b. If \( \mathbf{E} \) is parallel to the path \( d\mathbf{l} \), \( \mathbf{E} \cdot d\mathbf{l} = E \, dl \) and we get
\[ W = Q \int_a^b E \, dl = Q[V(b) - V(a)]. \]
where \( V(b) - V(a) \) is the potential difference (voltage) between the two electrodes (one at a and one at b). Electric field has units of V/m; if we have two electrodes separated by a distance \( d \), then the \( E \)-field has magnitude \( V/d \), where \( V \) is the potential difference between the two electrodes, \( V = V(b) - V(a) \).

Electric charge is measured in units of Coulombs. One Coulomb (1 C) is one ampère-second (1 A.s); if a current of one ampère (1 A) flows for one minute (60 s), the quantity of charge transported will be 60 C.

When charges migrate in an electric field, thermodynamic work is done. One Coulomb (1 C) of charge moving through a potential difference of one volt results in the performance of one joule (1 J) of work.

**Ideal solutions** are solutions whose properties depend linearly on the concentration of a dissolved species. Water is not electrically conductive unless we add electrolytes. The dissolved ions conduct electricity. We would expect that the electric conductivity \( \sigma \) of the solution increases linearly with the concentration of ions [\( A^+ \)], i.e. \( \sigma \propto [A^+] \). This is true at low ion concentrations. However, the more ions are dissolved in solution, the more they interact among each other, leading to deviations from ideal behavior. In fact, for “strong” electrolytes, at high concentrations, the conductivity begins to decrease, as shown in Fig. 3.17.

In general, ionic solutions are not ideal, meaning that we can expect interaction among the ions leading to deviations from ideal behavior. The non-ideality character of a solution occurs at ion concentrations above 10-100 \( \mu \)m; the exact point of onset depends on the strength of intermolecular interactions.

Significant early developments in the theory of electrolyte solutions are due to Debye and Hückel (1923) as well as Lars Onsager. The decrease in the molar conductivity of a strong electrolyte was attributed to the mutual interference of the ions, which becomes more pronounced as the concentration increases. Because of the strong attractive forces between ions of opposite
3.9. Ionic Conduction

![Conductivity graph]

Figure 3.17. Conductivity of saline solution and comparison to theory (AEM, Advanced Electrolyte Model by Kevin Gering, R&D 100, FLC Far West, and Idaho Innovation awards, 2014). In the low concentration limit, conductivity increases with ion concentration. However, at higher ion concentrations, the conductivity decreases due to significant ion-ion interactions. The SI unit for conductivity is Siemens per meter (S/m). $1 \text{ S} = 1 \Omega^{-1} = 1 \text{ A/V}$, or $1 \text{ S} = 1 \text{ A}^2/(\text{kg} \cdot \text{m}^2 \cdot \text{s}^3)$.

The specific example of NaCl is shown in Fig. 3.18. As seen in Fig. 3.18(a), each sodium ion has six chloride ions as its nearest neighbor. When the sodium chloride is dissolved in water, this ordering is still preserved to a very slight extent [Fig. 3.18(b)]. The ions are much farther apart than in the solid; the electrical attractions are therefore much smaller and the thermal motions cause irregularity. The small amount of ordering that does exist, however, is sufficient to exert an important effect on the conductivity of the solution.

The effect of the ionic atmosphere is to exert a drag on the movement of a given ion. If the ion is stationary, the atmosphere is arranged symmetrically about it and does not tend to move it in either direction. See Fig 3.19(a). However, if a potential that tends to move the ion (indicated by the upward arrow) is applied, the atmosphere will decay to some extent under the ion and build up more above it. See Fig. 3.19(b). Since it takes time for these relaxation processes to occur, there will be an excess of ionic atmosphere...
Figure 3.18. Structure of NaCl in solid (crystalline) form (a) and in solution (b). The solution structure is time-averaged.

below the ion (i.e., behind it) and a deficit above (in front of it). This asymmetry of the atmosphere will have the effect of dragging the central ion back. This is the relaxation or asymmetry effect.

There is a second reason why the existence of the ionic atmosphere impedes the motion of an ion. Ions are attracted to solvent molecules mainly by ion-dipole forces; therefore, when they move, they drag solvent along with them. The ionic atmosphere, however, having a charge opposite to that of the central ion, moves in the opposite direction to it and therefore drags solvent in the opposite direction. This means that the central ion has to travel upstream, and it therefore travels more slowly than if there were no effect of this kind. This is the electrophoretic effect (Fig. 3.20).

The theory of Debye and Hückel supposes the ions travel through the solution in straight lines, neglecting zigzag Brownian motion brought about by
3.9. Ionic Conduction

The electrophoretic effect occurs due to the collisions of surrounding solvent molecules. Lars Onsager in 1926 improved the theory to account for these “thermal effects.”

3.9.1. Ionic Solutions Generally Exhibit Non-Ideal Behavior. The conceptual definition of an ideal solution requires that the interactions of solvent-solvent, solvent-solute, and solute-solute molecules are identical, something that is clearly not possible in ionic solutions. As a consequence, the operational criteria of solution ideality (linear dependence of colligative and other properties on nominal concentrations), does not apply. In general, we cannot realistically expect calculations based on equilibrium constants to reliably predict the compositions of solutions having significant concentrations of ions.

The actual quantity we use to express the summed concentrations of ions of all kinds in a solution is the ionic strength. See this Wikipedia article for a definition and examples: http://en.wikipedia.org/wiki/Ionic_strength

What “significant” means in this context depends on the charges of the ions, and this includes “spectator” ions that are not directly involved in a reaction. A very crude rule-of-thumb is that ion concentrations in excess of 0.0001 M are likely to exhibit non-ideal behavior.

3.9.2. Interactions Among Dissolved Ions. It is not possible using conventional synthetic chemistry techniques to introduce a single kind of ion into water. Bulk matter cannot possess a significant net electric charge. Thus, all ionic solutions contain both anions and cations in the proportions required to ensure electroneutrality. Coulombic interactions between charged particles fall off with separation distance far more slowly than do the weaker ion-dipole forces responsible for water ordering, and are thus operative even in solutions that we would normally regard as quite dilute, down to about $10^{-4}$ M.

3.9.2.1. Oppositely-Charged Ions Tend to Pair Up. Depending on the sizes and charges of the two kind of ions, they may form relatively stable ion
pairs in which the “bonding” is purely electrostatic and independent of the chemical properties of the ions (Fig. 3.21).

Figure 3.21. Oppositely-charged ions in solution tend to pair up.

Neglect of ion-pairing can lead to erroneous assumptions about the concentrations of “free” ions in solutions, and misinterpretation of calculations involving solubility. Here are a few examples:

- In addition to forming the sparingly-soluble solid CdI$_2$, part of the cadmium in a solution in equilibrium with the two ions will exist as the soluble species CdI$^+(aq)$ and CdI$_2^0(aq)$.
- Similarly, solutions in equilibrium with Ca(OH)$_2(s)$ will also contain species CaOH$^+(aq)$ and Ca(OH)$_2^0(aq)$.
- The ion-pair CaCO$_3^0(aq)$ is believed to tie up a significant fraction of the total carbonate when water is in equilibrium with solid calcium carbonate.

The ions in the more strongly-bound pairs will be in contact and within a common hydration shell, whereas others may have only their shells in contact.

3.9.2.2. Electrostatic Forces Lead to Local Charge Imbalances Near an Ion. Simple electrostatics predicts that oppositely-charged ions are more likely to be found close to a given ion, while like-charged ions tend to be farther away. Thus in addition to hydration shells, ions in solution are surrounded by a region containing an excess of counter-ions, known as the ionic atmosphere. The latter tends to move along with the ion as it diffuses or migrates through the solution.

Figure 3.22. Ionic atmosphere.

- In Fig. 3.22, the ionic atmosphere corresponds to the yellow region just outside the primary hydration shell [blue].
• Although only the positive counter-ions are explicitly represented in the ionic atmosphere in the diagram, it is important to understand that the atmosphere contains ions of both charges; the positive ions are just present in greater quantity. Near the outer boundary of the ionic atmosphere, the displaced negative ions are in excess. Only in the bulk solution are their concentrations identical.

• The atmosphere surrounding an cation will of course have the opposite charge. Because cations tend to be smaller than anions, the volume and charge density of the atmosphere is expected to be bigger.

3.9.3. Most Salts are Not Completely Dissociated in Water. The dissolution of cadmium iodide is water is commonly represented as

\[ \text{CdI}_2(s) \rightarrow \text{Cd}^{2+} + 2\text{I}^- \].

This is actually wrong.\(^6\) To most people, this would imply that a 0.1 M solution of this salt would contain 0.1 M of Cd\(^{2+}\)(aq)? This would be wrong because it fails to take into account that the two ions react with each other. Firstly, they combine to form neutral, largely-covalent molecular species:

\[ \text{Cd}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{CdI}_2(aq) \].

This non-ionic form accounts for 76% of the Cd present in the solution. In addition, they form a molecular ion CdI\(^+\)(aq) according to the following scheme (the K’s are equilibrium constants, see Section 4.1.0.1):

\[
\begin{align*}
\text{CdI}_2(s) &\rightarrow \text{Cd}^{2+} + 2\text{I}^- , \quad K_1 = 10^{-3.9} \\
\text{Cd}^{2+} + \text{I}^- &\rightarrow \text{CdI}^+ , \quad K_2 = 10^{+2.3} \\
\text{CdI}_2(s) &\rightarrow \text{CdI}^+ + \text{I}^- , \quad K = 10^{-1.6} = 0.023.
\end{align*}
\]

As a consequence, the concentration of “free” Cd\(^{2+}\)(aq) in an aqueous cadmium iodide solution is only about 2% of the value you would calculate by taking \(K_1\) as the solubility product. The principal component of such a solution is actually (covalently-bound) CdI\(_2\)(aq). It turns out that many salts, especially those of metals beyond Group 2, are similarly only partially ionized in aqueous solution. See Table 3.1.

\(\text{Table 3.1.}\) Species present in various ionic solutions. This is Table 2 from [13].

<table>
<thead>
<tr>
<th>Solute</th>
<th>Molarity</th>
<th>Ionic Strength</th>
<th>Fraction of Metal Present as Each Species</th>
</tr>
</thead>
<tbody>
<tr>
<td>KCl</td>
<td>0.52</td>
<td>0.5</td>
<td>K(^+) 0.95 KCl 0.05</td>
</tr>
<tr>
<td>CsCl</td>
<td>0.55</td>
<td>0.5</td>
<td>Cs(^+) 0.9 CsCl 0.1</td>
</tr>
<tr>
<td>KI</td>
<td>0.58</td>
<td>0.5</td>
<td>K(^+) 0.87 KI 0.13</td>
</tr>
<tr>
<td>CdI</td>
<td>0.61</td>
<td>0.5</td>
<td>Cd(^+) 0.82 Cd 0.18</td>
</tr>
<tr>
<td>Na(_2)SO(_4)</td>
<td>0.25</td>
<td>0.5</td>
<td>Na(^+) 0.76 NaSO(_4)(^-) 0.24</td>
</tr>
</tbody>
</table>

\(^6\) The data shown here and in the table are from Ref. [13]. This fact was stated by Arrhenius in 1887, but has been largely ignored and is almost never mentioned in standard textbooks.
3.9.4. Activities and Activity Coefficients. The quantitative treatment of ionic solutions is based on the Debye-Hückel theory, which was developed in the 1920s. This theory models the electrostatic interactions between ions and their ionic atmospheres, and can predict mean ionic activity coefficient for solutions whose ionic strengths are not very high. Many extensions to this model have been suggested; all are mathematically complicated and all fail for “highly concentrated” solutions.

The practical approach to the problem of non-ideality is to introduce a quantity known as the activity, which can be thought of as the thermodynamically-effective concentration. The relation between activity \( a \) and the “analytical” concentration \( c \) is given by the activity coefficient \( \gamma \) (gamma), where \( a = \gamma c \).

An ideal solution has \( \gamma = 1 \). As ionic concentrations increase, activity constants diminish. Activity coefficients approach unity in the limit of zero concentration. Because solutions containing ions of a single charge species cannot be prepared, all experimental measurements can only yield mean ionic activity coefficients (i.e. the coefficient describes both positive and negative ions simultaneously).

3.9.5. Reactions: Halide Compounds. An unknown solution that is suspected to contain halide compounds such as KCl, KBr and KI can be identified with silver nitrate solution, \( \text{AgNO}_3 \). The halogen will react with \( \text{Ag}^+ \) and form a precipitate, with varying color depending on the halogen:

- AgF: No Precipitate
- AgCl: White
- AgBr: Pale Yellow
- AgI: Bright Yellow

3.9.6. Reactions: Precipitation. Barium sulfate is not very soluble in water at room temperature. The following reaction is very unfavorable:

\[
\text{BaSO}_4(s) \rightarrow \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq).
\]
Instead, the reverse reaction is much more likely to happen. It is called a precipitation reaction

$$\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s).$$

One possible way to obtain this reaction is to mix barium chloride and potassium sulfate:

$$\text{BaCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{KCl}(aq).$$

What happens during this reaction is that the ions dissociate, followed by precipitation:

$$\text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq) + 2\text{K}^+(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) + 2\text{K}^+(aq) + 2\text{Cl}^-(aq).$$

The potassium and chloride ions appear on both sides of the equation. They are called spectator ions, as they do not take part in the reaction. We can omit the spectator ions and write the net ionic equation:

$$\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s).$$

### 3.9.7. Reactions: Ionic Exchange

Suppose that we dissolve magnesium sulfate and sodium hydroxide in aqueous solution. The ions dissolve to some extent:

$$\text{MgSO}_4 + 2\text{NaOH} \rightarrow \text{Mg}^{2+} + \text{SO}_4^{2-} + 2\text{Na}^+ + 2\text{OH}^-. $$

Of those ions that do dissolve there is a strong tendency for precipitation. In this particular case, the following precipitation reaction takes place:

$$\rightarrow \text{Mg(OH)}_2(s) + \text{SO}_4^{2-} + 2\text{Na}^+.$$
Another example is water softening by ion exchange. Water that contains divalent cations, such as Ca$^{2+}$, Mg$^{2+}$ and Fe$^{2+}$ causes problems. First, they combine with fatty-acid anions in soaps to produce hard deposits on clothes, washing machines and kitchen sinks:

$$\text{Ca}^2(aq) + 2\text{C}_{17}\text{H}_{35}\text{COONa}(aq) \rightarrow (\text{C}_{17}\text{H}_{35}\text{COO})_2\text{Ca}(s) + 2\text{Na}^+(aq).$$

Secondly, when bicarbonate (HCO$_3^-$) is present in water, the hard-water cations cause a buildup of deposits in boilers, pots and pipes, which not only looks dirty but can damage the plumbing:

$$\text{Ca}^{2+}(aq) + 2\text{HCO}_3^-(aq) \rightarrow \text{CaCO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(l).$$

The removal of hard-water ions (water softening) is accomplished by exchanging soft-water Na$^+$ ions for the hard-water ions. Ion-exchange resins are insoluble polymers with covalently bound anion groups, such as -SO$_3^-$ or -COO$^-$, with Na$^+$ ions attached to balance the charge. The divalent cations in hard water are attracted to the resin’s anionic groups and displace the Na$^+$ ions into the water. One type of ion is exchanged for another. The resin is replaced when all the resin sites are occupied, or it can be \textit{regenerated} by treating it with a very concentrated Na$^+$ solution, which exchanges Na$^+$ ions for the bound Ca$^{2+}$.

\section*{3.9.8. Importance of Matching the Number of Moles in Reactions.}
When solving problems, pay attention to the data given to you, such as masses of substances, volumes of solutions and concentrations. For example, a problem may ask you which volume or concentration of a solution is required to neutralize or precipitate a given substance. The problem frequently boils down to matching the number of moles and inspecting the stoichiometry of the chemical reaction to ensure the correct number of moles are used. For a given stoichiometry, the number of moles is calculated using the formula $n = c \cdot V$, where $n$ is the number of moles, $c$ is concentration and $V$ is volume. The example of the previous section is typical.

One more thing you will need to check is whether or not a substance has fully dissociated in the solvent or not. Knowing the extent of a reaction (how far it goes) is essential to knowing the correct number of moles.

Oxtoby [1] (p. 481-2) discusses the following two reactions:

Potassium dichromate with hydrochloric acid:

$$\text{K}_2\text{Cr}_2\text{O}_7(s) + 14\text{HCl}(aq) \rightarrow 2\text{K}^+(aq) + 2\text{Cr}^{3+}(aq) + 8\text{Cl}^-(aq) + 7\text{H}_2\text{O}(l) + 3\text{Cl}_2(g).$$
Preparation of elemental bromine from its salts in solution:

\[ 2\text{Br}^- (aq) + \text{Cl}_2 (aq) \rightarrow 2\text{Cl}^- (aq) + \text{Br}_2 (aq). \]

### 3.10. Ideal Solutions: Raoult’s and Henry’s Laws

Vapor pressure (Fig. 3.23) is the pressure of a substance as measured in the space above the liquid or solid substance. It is easiest to measure this pressure in closed container in the gas space above the substance that is initially (at time \( t = 0 \)) vacuum-pumped. Vapor pressure of a substance arises because of thermal energy \( (k_B T) \), which imparts some fraction\(^7\) of molecules with kinetic energy large enough to leave the liquid or solid\(^8\) phase and escape into the gas phase.

---

**Figure 3.23.** Vapor pressure.

A *solution* is a homogeneous phase that contains more than one component. Although if both components are in the same phase, e.g. liquids\(^9\), and the solution is homogeneous, there is no fundamental difference between components in a solution as far as labeling them solvent vs solute. It is then common practice to label the component that constitutes the larger proportion of the solution as the solvent; the component in lesser proportion is commonly called the solute.

\(^7\)For a given temperature and molecular mass, you should be able to calculate this fraction from the Maxwell-Boltzmann speed distribution. The energy should be larger than intermolecular forces that keep the liquid (or solid) together.

\(^8\)At sufficiently low temperature and pressure, a solid substance may sublimate to the gas phase without passing through the liquid phase (c.f. Fig. 1.4).

\(^9\)Solid solutions are also possible when there are one or more solutes in a solvent, and the chemical components remain in a single homogeneous phase.
Partial vapor pressures of the individual components of the solution yield information about the individual components in solution. This is because the partial vapor pressure measures the escaping tendency of a molecule from solution, which is in turn a measure of the cohesive forces present in solution. Thus, measurements of the vapor pressure of each component as a function of pressure, temperature, and mole fraction lead to some understanding of the system, namely, the intermolecular forces and the physical properties affected by them (e.g. boiling point, freezing point, etc).

*Ideal solutions* are homogeneous mixtures of substances that have physical properties linearly related to the amount of solute present. The classic statement of this condition is Raoult’s law, which is valid for many highly dilute solutions and for a limited class of concentrated solutions, namely, those in which the interactions between the molecules of solute and solvent are the same as those between the molecules of each substance by itself. Solutions of benzene and toluene, which have very similar molecular structures, are ideal solutions: any mixture of the two has a volume equal to the sum of the volumes of the separate components, and the mixing process occurs without absorption or evolution of heat. The vapor pressures of the solutions are mathematically represented by a linear function of the molecular composition. Ideal solutions are analogous to ideal gases ($P_i = n_iRT/V$, which expresses the linear dependence of the partial pressure, $P_i$, of the $i$-th component of a gas mixture on gas density $n_i/V$ of that component at fixed $T$).

When the components of nonideal solutions are mixed, the volume of the mixture ordinarily differs from the sum of the volumes of the pure components, and heat is evolved or absorbed. The properties of such solutions often are described in terms of their deviations from those of ideal solutions. Mixtures of acetone and chloroform, for example, are said to show negative deviations from ideality: their vapor pressures are lower than those calculated on the assumption of a linear relationship to the molecular composition. Solutions of acetone and carbon disulfide, on the other hand, have higher vapor pressures than those that would characterize an ideal solution. The cases of chloroform in acetone and chloroform in ethanol will be discussed later in the context of Fig. 3.28.

### 3.10.1. Raoult’s Law (François Marie Raoult, 1830-1916).

Raoult’s law says that the vapor pressure of a component (1), $P_1$, in solution is the product of the pure substance’s vapor pressure, $P_1^o$ (may depend on temperature), and its mole fraction, $X_1$:

$$P_1 = X_1P_1^o.$$
A solution which obeys this law is called an **ideal solution**. Solutions that deviate from this law are called non-ideal solutions. For non-ideal solution, this linear relationship only holds in the dilute limit ($X_2 \to 0, X_1 \to 1$, for a 2-component solution). Mathematically, this law is expressed as a limit:

\[
P_1^o = \lim_{X_1 \to 1} \frac{P_1}{X_1}.
\]

Real solutions may show positive deviations (with vapor pressures higher than those predicted by Raoult’s law) or negative deviations (with lower vapor pressures).

The total vapor pressure, of a solution consisting of $N$ components will be the sum of the vapor pressures:

\[
P_{tot} = P_1 + P_2 + \cdots + P_N,
\]

where 1, 2, $\ldots$ , $N$ include all components of the solution (all solutes plus solvent). Using Raoult’s law this can be written as:

\[
P_{tot} = X_1 P_1^o + X_2 P_2^o + \cdots + X_N P_N^o,
\]

where $X_1 + X_2 + \cdots + X_N = 1$. This relationship look similar to the Dalton’s law, but applies to ideal solutions.

For a 2-component system, where “1” is the solvent and “2” is the solute,

\[
X_1 = \frac{n_1}{n_1 + n_2}.
\]

Here, $n_1$ is the number of moles of solvent and $n_2$ is the number of moles of solute. $n_1 + n_2$ is the total number of moles in solution. $X_1 + X_2 = 1$.

Raoult’s law is a statement regarding the behavior of solvents when a solute is added. Raoul was known for his research on solutions. He published papers on solute-induced freezing point depression of solutions, depression of a solvent’s vapor pressure, and various experiments involving solvents, such as benzene and acetic acid, in addition to water. The laws he published concern the properties of the solvent, enable the determination of the molecular weights of dissolved substances, and are valid in the limit of dilute solutions.

Raoult’s law is typically only valid at low solute concentrations. The relationship $P_1 = X_1 P_1^o$ is analogous to the ideal gas law, $P_1 = n_1 RT/V$. Both laws express the linear dependence of $P_1$ on the amount of substance 1, namely $X_1$ (or $n_1$ for gases).

### 3.11. Deviations From Raoul’ts Law

Real solutions deviate from ideal behavior, as shown in Fig. 3.24.
3.11. Negative Deviations. On a molecular level, negative deviations arise when the solute attracts solvent molecules especially strongly, reducing their tendency to escape into the vapor phase. The substance wants to stay a liquid. (Recall our discussion of the VDW equation, where attractive forces lower the pressure.)

3.11.2. Positive Deviations. Positive deviations arise in the opposite case, when solvent and solute molecules are not strongly attracted to each other. The substance wants to be a gas, so its pressure is higher than it should be. Even non-ideal solutions with non-dissociating solutes approach Raoults law as \( X_1 \) approaches 1, just as all real gases obey the ideal gas law at sufficiently low densities.

No deviations: A number of pairs of liquids obey Raoults law over a wide range of compositions:

- benzene — toluene
- ethylene bromide — ethylene chloride
- carbon tetrachloride — trichloroethylene
- acetic acid — isobornyl acetate

It should be noted that air is generally present above the solution at a pressure that makes up the difference between the total vapor pressure and atmospheric pressure. The fact that air is present can generally be ignored in the vapor phase.

3.11.3. How Raoults Law Works. While studying colligative properties of solvents, Raoults discovered that adding a solute lowers the vapor pressure. This is because the additional solute particles occupy space (volume
fraction) in solution. Vapor pressure is due to surface molecules escaping into the gas phase. If some of these solvent molecules are replaced by solute molecules, there will be fewer solvent molecules able to escape into the gas phase (due to obstruction), lowering the vapor pressure of the solvent.

When a liquid is in a sealed container, evaporation proceeds until there are as many molecules returning to the liquid as there are escaping (per unit time). At this point the vapor is said to be saturated, and the pressure of that vapor is called the **saturated vapor pressure**. Figure 3.25 shows the case where we start from 100% alcohol and dilute it by adding water (the case of 70% alcohol, 30% water is shown). Figure 3.25 shows that there are fewer alcohol molecules in the gas phase after dilution with water.

![Figure 3.25.](image)

As mentioned earlier, some fraction of the solvent molecules have sufficient energy to escape from the surface and enter the gas phase. If the number of solvent molecules on the liquid’s surface is reduced by the addition of solute, there will be a smaller number of solvent molecules that will be able to escape during any given time interval. On the other hand, molecules in the gas phase can always return to the solution regardless of how much solute is present (assuming an ideal solution). The net effect of this is that when equilibrium is established, there will be fewer solvent molecules in the vapor phase, and the vapor pressure is lowered.
3.11.4. Limitations of Raoult’s Law. In practice, there’s no such thing as an ideal solution! However, features of one include:

- Ideal solutions satisfy Raoult’s law over a wide range of concentrations, meaning that some key macroscopic properties (such as vapor pressure, boiling point, freezing point, etc.) depend linearly on the amount of solute added and not on the nature of the solute itself.

- In an ideal solution (Fig. 3.26), it takes exactly the same amount of energy for a solvent molecule to break away from the surface of the solution as it did in the pure solvent. The forces of attraction between solvent and solute are exactly the same as between the original solvent molecules.

- There are very few cases of truly ideal solutions where Raoult’s law holds precisely over the entire range of concentrations. When solutions deviate from ideal solutions, the Raoult’s law will still hold, but in the dilute limit (see Eq. 3.1).

![Figure 3.26. Ideal solution. The intermolecular forces shown are average forces (averaged over all such interactions, and over time).](image)

3.11.5. Henry’s Law (William Henry, 1774-1836). While Raoult’s law applies to solvents, Henry’s law applies to the solute. Henry’s law is the statement that the vapor pressure of a volatile dissolved substance is proportional to the mole fraction (or concentration) of that substance in solution. Suppose we have a solute “2” with low mole fraction $X_2$. Its vapor pressure (even in non-ideal solution) is proportional to $X_2$:

$$P_2 = k_2 X_2,$$

where $k_2$ is a constant (the Henry’s law constant), which is usually denoted $k_H$. The SI units of the Henry’s law constant are Pa. Its value depends on both solute and solvent. A volatile substance is one for which the Henry’s law constant is non-zero. In mathematical terms, Henry’s law is the statement

$^{10}$Since $X_2 = 1 - X_1$, we can also express this as $P_2 = k_2 X_2 = k_2 (1 - X_1)$. 
that the following limit exists:

\[
(3.2) \quad k_2 = \lim_{\chi_2 \to 0} \frac{P_2}{\chi_2}
\]

We also find Henry’s law expressed in concentration units, \(P_2 = k_2 \cdot c_2\), where \(k_2\) has units of pressure per unit concentration (e.g. L.atm/mol) and \(c_2\) has concentration units (mol/L). Concentration units arise as follows: suppose we have a gas, such as CO\(_2\) dissolved in water. The chemical equilibrium is:

\[
\text{CO}_2(aq) \rightleftharpoons \text{CO}_2(g) + \text{H}_2\text{O}(l), \quad K_{eq} = \frac{[\text{CO}_2(g)][\text{H}_2\text{O}(l)]}{[\text{CO}_2(aq)]}.
\]

The concentration of water is a constant, so we can combine it with \(K_{eq}\) by dividing both sides by \([\text{H}_2\text{O}(l)]\):

\[
k_H = \frac{K_{eq}}{[\text{H}_2\text{O}(l)]} = \frac{P_{\text{CO}_2}}{[\text{CO}_2(aq)]}.
\]

Thus, \(k_H\) has units of atm/M if \(P_{\text{CO}_2}\) is in atm. Chemical equilibrium will be covered in Chapter 4.

### Henry’s law: (Applicability)

- Henry’s law applies in the limit of low concentrations of the solute.
- Henry’s law can be used to infer the concentration of dissolved gas in a solvent from the knowledge of the vapor pressure of this solute, provided of course, that Henry’s law constant is known.

Whenever Raoult’s law is valid for a solvent, Henry’s law is valid for a solute (see Fig. 3.27). Henry’s law may be applied to dilute solutions of a binary liquid system. It is found that in the limit of infinite dilution, most liquid solvents obey Raoult’s law but that under the same conditions the solute obeys Henry’s law.

One application of Henry’s law is the carbonation of beverages: by knowing the partial pressure of a given gas in the space above a solution, we can estimate how much of the substance is dissolved in solution. This is best understood by working an example.

3.11.5.1. Example. Suppose that we have a carbonated drink (soda). CO\(_2\) gas is dissolved and the can is pressurized. The partial pressure of CO\(_2\) above the soda (in the soda can) is 0.01 atm (this can be taken as the total gas pressure, since water is far less volatile than CO\(_2\)). Henry’s law constant for CO\(_2\) dissolved in water at 20\(^\circ\)C is 29.4 L.atm/mol. Find the amount of CO\(_2\) dissolved in solution.

**Solution:** If the Henry’s law constant is specified in units of L.atm/mol, this means the Henry’s law is in the form \(P = k_H \cdot c\), where \(P\) is the partial pressure of the solute above the solution and \(c\) is the concentration of the solute in solution. From \(c = P/k_H\), we find \(c = 0.01\ \text{atm}/29.4\ \text{L.atm/mol}=0.34\ \text{mM}.\)
Note: This problem is slightly more complicated when $k_H$ is specified in pressure units, e.g. $k_H = 0.163 \times 10^4$ atm. We instead use the form, $X_{CO_2}k_H = P_{CO_2}$. The mole fraction is:

$$X_{CO_2} = \frac{P_{CO_2}}{k_H} = \frac{0.01 \text{ atm}}{0.163 \times 10^4 \text{ atm}} = 0.061 \times 10^{-4}.$$ 

The mole fraction is defined as:

$$X_{CO_2} = \frac{n_{CO_2}}{n_{H_2O} + n_{CO_2}} = \frac{c_{CO_2}}{c_{H_2O} + c_{CO_2}} \approx \frac{c_{CO_2}}{c_{H_2O}} = \frac{c_{CO_2}}{55 \text{ M}}$$

Inserting the numerical value for $X_{CO_2}$ and solving for $c_{CO_2}$:

$$0.061 \times 10^{-4} = \frac{c_{CO_2}}{55 \text{ M}} \quad \rightarrow \quad c_{CO_2} = (0.061 \times 10^{-4})(55 \text{ M}) = 0.34 \text{ mM}.$$ 

This suggests another expression for the Henry’s law:

$$P_2(\text{atm}) = k_H(\text{atm})X_2 \frac{[H_2O]}{[H_2O]} = \frac{k_H(\text{atm})}{[H_2O]} \cdot c_2(\text{M}).$$

where $[H_2O]$=55 M and $c_2(\text{M}) = X_2 \cdot [H_2O]$.

3.11.6. **Henry’s Law Constant.** Henry’s law constant $k_H$ depends on the solvent, solute and temperature. You can find such constants tabulated under various conditions. See, for example, [http://www.henrys-law.org/henry.pdf](http://www.henrys-law.org/henry.pdf)

If a gas is dissolved in a solvent, the pressure of the gas in equilibrium with the solution is proportional to the amount of gas dissolved. (And
3.11. Deviations From Raoult’s Law

Likewise, if several gases from a mixture of gases dissolve in a solution, Henry’s law applies to each gas independently, regardless of the pressure of the other gases present in the mixture (meaning, for example, that air could be present as well, and we can neglect air).

Henry’s law constant is the vapor pressure of the liquified gas at the temperature of the solution. The dissolved gas may be viewed in terms of its solubility under the pressure $P_2$, or $P_2$ may be taken as the vapor pressure of the volatile solute.

Temperature dependence of Henry’s law constant: An expression for the temperature dependence of the Henry’s law constant is (see Problem 49 or Atkins [14] for derivation):

$$k_H(T) = k_H^\ominus \exp \left( -\frac{\Delta_{\text{soln}} H}{R} \left[ \frac{1}{T} - \frac{1}{T^\ominus} \right] \right)$$

where $\Delta_{\text{soln}} H$ is the enthalpy of solution (in molar units here, since the denominator uses $R$). We will introduce enthalpy later (see Section 6.9.5). $T^\ominus$ is a reference temperature and $k_H^\ominus$ is the value of $k_H$ measured at this reference temperature. Note: The value of $k_H^\ominus$ depends on both the solvent and the solute. For example $k_H^\ominus$ differs for oxygen in water compared to nitrogen in water. Likewise, $k_H^\ominus$ is different for oxygen in benzene versus oxygen in water.

3.11.7. Summary: Raoult’s and Henry’s Laws. The three plots in Fig. 3.28 summarize the behavior of Raoult’s and Henry’s laws.

Figure 3.28. Summary of Raoult’s and Henry’s laws. (left) Vapor pressure of two liquids obeying Raoult’s law. The total pressure at constant temperature is obtained from $P = X_1 P_1^0 + (1 - X_1) P_2^0$. (middle) Vapor pressure of the system acetone-chloroform at 328 K showing negative deviations from Raoult’s law. Dashed lines show Raoult’s law behavior. (right) Vapor pressure of the system ethanol-chloroform at 318 K showing positive deviations from Raoult’s law. Dashed lines show Raoult’s law behavior.
3.11.8. Henry’s Law Example. Dry air contains 78.084 mol % N\(_2\) and 20.946 mol % O\(_2\). Calculate the relative proportion of N\(_2\) and O\(_2\) dissolved in water under a total pressure of 1.000 bar. Henry’s law constants for N\(_2\) and O\(_2\) are 6.51 \times 10^7 \text{ Torr} and 3.30 \times 10^7 \text{ Torr}, respectively, at 25°C.

Solution: Application of Dalton’s law gives
\[
P_{N_2} = 0.78084 \times 750.06 \text{ Torr} = 585.7 \text{ Torr}
\]
\[
P_{O_2} = 0.20946 \times 750.06 \text{ Torr} = 157.1 \text{ Torr}
\]
We then apply Henry’s law to get the mole fractions
\[
X_{N_2} = \frac{P_{N_2}}{k_H(N_2)} = \frac{585.7 \text{ Torr}}{6.51 \times 10^7 \text{ Torr}} = 9.00 \times 10^{-6}
\]
\[
X_{O_2} = \frac{P_{O_2}}{k_H(O_2)} = \frac{157.1 \text{ Torr}}{3.30 \times 10^7 \text{ Torr}} = 4.76 \times 10^{-6}
\]
Since \(X_{N_2} = n_{N_2}/n_{tot}\) and \(X_{O_2} = n_{O_2}/n_{tot}\), the relative proportions are:

for N\(_2\)
\[
\frac{n_{N_2}}{n_{N_2} + n_{O_2}} = \frac{9.00 \times 10^{-6}}{9.00 \times 10^{-6} + 4.76 \times 10^{-6}} = \frac{9.00 \times 10^{-6}}{1.376 \times 10^{-6}} = 0.654
\]

for O\(_2\)
\[
\frac{n_{O_2}}{n_{N_2} + n_{O_2}} = \frac{4.76 \times 10^{-6}}{9.00 \times 10^{-6} + 4.76 \times 10^{-6}} = 0.346
\]

This is approximately a ratio of 2 N\(_2\) to 1 O\(_2\).


Raoult’s law forms the basis for the four colligative properties of dilute solutions. Colligative means that the properties depend on the number of dissolved particles, independently of the nature of the particles themselves. These properties are:

- Lowering of vapor pressure
- Boiling point elevation
- Freezing point depression
- Osmotic pressure

3.12.1. Lowering of Vapor Pressure (Raoult’s Law). Suppose we have a two-component solution where “1” is the solvent and “2” is the solute. Raoult’s law for the solvent component says that:

\[
\Delta P_1 = P_1 - P_1^o = X_1 P_1^o - P_1^o = (X_1 - 1) P_1^o = -X_2 P_1^o,
\]

where \(X_1 = 1 - X_2 = n_1/(n_1 + n_2)\) is the solvent mole fraction, \(P_1^o\) is the pure solvent vapor pressure and \(X_2 = n_2/(n_1 + n_2)\) is the solute’s mole fraction. Because the sign is negative, we have a lowering of the vapor pressure. Thus,
the vapor pressure is always lower when we add a solute, as compared to the pure solvent. Note that $\Delta P_1 = -X_2P_1^\circ$ depends only on $X_2$, the mole fraction of the solute. It does not depend at all on the nature of the solute. This is an example of a colligative property.

3.12.2. Boiling Point Elevation (Raoult’s Law). A consequence of the lowering of vapor pressure is the elevation of the boiling point. This fact has been used as a way to determine molar masses. In Fig. 3.29, $\Delta P_1$ is the decrease in vapor pressure at $T_b$ and $\Delta T_b$ is the change in temperature needed to keep the vapor pressure at 1 atm. The quantity

$$\Delta T_b = T_b' - T_b$$

is the increase in boiling point caused by the addition of solute to the pure solvent. Let us look at the slope of this graph near 1 atm, assuming that the two curves are parallel near that point:

$$-\frac{\Delta P_1}{\Delta T_b} = \text{slope of curve} = S \quad \rightarrow \quad \Delta T_b = -\frac{\Delta P_1}{S} = \frac{X_2P_1^\circ}{S} = \frac{1}{S} \left( \frac{n_2}{n_1 + n_2} \right)$$

from Raoult’s law with $P_1^\circ=1$ atm. $S$ is a property of the solvent (independent of the solute) because we assumed the two curves were parallel. For dilute solutions ($n_1 \gg n_2$) we have

$$\Delta T_b = \frac{1}{S} \frac{n_2}{n_1} = \frac{1}{S} \left( \frac{m_2/M_2}{m_1/M_1} \right),$$

where $m_1$ and $m_2$ are the masses of the solvent and solute (in grams) and $M_1$ and $M_2$ are their molar masses in grams per mole. Because $M_1$ is a property of the solvent, it can be lumped into $S$. We define a new constant:

$$K_b = \frac{M_1}{(1000 \text{ g.kg}^{-1})S}$$

**Boiling point elevation:**

$$\Delta T_b = K_b \left( \frac{m_2/M_2}{m_1/(1000 \text{ g.kg}^{-1})} \right) = K_b m.$$ If $m_1$ is in grams, $m_1/(1000 \text{ g.kg}^{-1})$ is the number of kilograms of solvent. $m_2/M_2$ is the number of moles of solute. Hence, this expression is equal to $K_b m$ where $m$ is the molality of the solution.

Note: if the substance *dissociates* in solution, we must count the total molality. For example, if we dissolve NaCl, we get two species: $\text{Na}^+$ and $\text{Cl}^-$. Thus, the molality should be doubled, as we now have 2 moles of ions for each mole of NaCl dissolved.
3. Properties of Matter

Temperature
Solvent vapor pressure
Pure solvent
Solvent in solution

![Figure 3.29](image)

Figure 3.29. The vapor pressure of the solvent above a dilute solution is lower than that of the pure solvent at all temperatures. As a result, for the solution to boil (that is, for the vapor pressure to reach 1 atm), a higher temperature is required for the solution than for the pure solvent. This amounts to an elevation of the boiling point, \( \Delta T_b = T'_b - T_b > 0 \).

3.12.3. Example w/ Multiple Ions. Lanthanum(III) chloride (LaCl₃) is a salt that completely dissociates into ions in dilute aqueous solution,

\[
\text{LaCl}_3(s) \rightarrow \text{La}^{3+}(aq) + 3\text{Cl}^-(aq)
\]
yielding 4 mol of ions per mole of LaCl₃. Suppose 0.2453 g of LaCl₃ is dissolved in 10.00 g of H₂O. What is the boiling point of the solution at atmospheric pressure, assuming ideal solution behavior?

**Solution:** The molar mass of LaCl₃ is 245.3 g/mol. The moles of LaCl₃ are \( \frac{0.2453 \text{ g}}{245.3 \text{ g/mol}} = 1.000 \times 10^{-3} \) mol. The molality, \( m \), is \( \frac{(4)(1.000 \times 10^{-3}) \text{ mol ions}}{0.0100 \text{ kg solvent}} = 0.400 \text{ mol/kg} \). We look up the value of \( K_b \) from tables and insert this molality into the equation for the boiling-point elevation:

\[
\Delta T_b = K_b m = (0.512 \text{ K.kg/mol})(0.400 \text{ mol/kg}) = 0.205 \text{ K}.
\]

This is the boiling point elevation. The new boiling point of the solution (water) is increased to 100.000°C + 0.205°C = 100.205°C. The actual boiling point is slightly lower than this because the solution is nonideal.

3.12.4. Freezing Point Depression (Raoult's Law). In a pure solvent, such as water, the formation of the solid phase usually involves the creation of a crystalline structure. The example of water is shown in Fig. 3.30.

The process of going from gas to liquid, or from liquid to solid, requires us to remove energy from the system (i.e. via cooling), as shown in Fig. 3.31.

\[\text{Not all solids are crystalline. Some are amorphous.}\]

Figure 3.30. Structure of water in the liquid and crystalline (solid) phases.

Figure 3.31. Phase transitions between different states of matter require the addition or removal of energy.

When a solute is added, however, the mere presence of the solute interferes with the formation of the solid crystalline structure (see Fig. 3.32). Indeed, the solute added (red) in the solvent (grey) prevents the rigid, ordered arrangement of a solid from being achieved. The particles of solute block the crystallization of the solvent, at least up until a certain point, causing a depression in the freezing point of the water. This is an entropic effect (see Problem 53).

Figure 3.32. Freezing point depression. (a) Pure solvent (b) Solvent plus solute.

Application 1: Why do we add ethylene glycol, \( \text{CH}_2\text{OHCH}_2\text{OH} \), to water in a car radiator in winter? Without any additive the water would freeze at 0°C and the resulting increase in volume would crack the radiator. Since ethylene glycol is very soluble in water, it can form a solution with a freezing point low enough to prevent freezing even on the coldest winter day.
Application 2: In many countries, salt is added to melt the ice on the sidewalk. The ice melts because the NaCl added lowers the freezing point of water. Of course, this only works at mild temperatures. At temperatures well below the freezing point, salt only has the effect of increasing friction.

Vapor pressure diagram. At the freezing point the vapor pressure of the solid solvent is equal to the vapor pressure of the solvent in solution. The freezing temperature of the solution is the temperature at which the vapor-pressure curve of the pure solid solvent intersects that of the solution (see Fig. 3.33). As solute is added to the solution, the vapor pressure of the solvent falls and the freezing point drops. \( \Delta T_f = T'_f - T_f \) is negative.

![Vapor Pressure Diagram](image)

**Figure 3.33.** Change in the temperature associated with the freezing point depression. \( \Delta T_f = T'_f - T_f < 0 \).

By analogy with the case of the boiling point elevation we write:

\[
\Delta T_f = T'_f - T_f = -K_fm.
\]

If the substance dissociates in solution, we must use the total molality, as the example below shows:

### 3.12.5. Example Problem

The number of moles of the major dissolved species in a 1.000-L sample of seawater are as follows:

- Na\(^+\): 0.458 mol.
- Mg\(^2+\): 0.052 mol.
- Ca\(^2+\): 0.010 mol.
- K\(^+\): 0.010 mol.
- Cl\(^-\): 0.533 mol.
- SO\(_4^{2-}\): 0.028 mol.
- HCO\(_3^-\): 0.002 mol.
- Br\(^-\): 0.001 mol.

Neutral species: 0.001 mol.

Estimate the freezing point of the seawater, assuming \( K_f = 1.86 \text{ K.kg/mol for water} \).

**Solution:** Because water has a density of 1.00 g/cm\(^3\), 1.00 L of water weighs 1.00 kg. For dilute aqueous solutions, the number of moles per kilogram of solvent (the molality, \( m \)) is therefore approximately equal to the number of moles per liter. The total molality, obtained by adding the individual
species molalities just given, is \( m = 1.095 \, \text{mol/kg} \). Then
\[
\Delta T_f = -K_f m = -(1.86 \, \text{K/kg/mol})(1.095 \, \text{mol/kg}) = -2.04 \, \text{K}.
\]
The seawater should freeze at approximately -2\(^\circ\)C. Nonideal solution effects make the actual freezing point slightly higher than this.

The graph of Fig. 3.34 shows freezing point of water vs molality of solute added. This graph appears to suggest that the freezing point of water depends not only on molality but also on the dissolved substance. This difference in slopes is actually due to the fact that different substances decompose into different numbers of ions (one: M, two: MX\(_1\), three: MX\(_2\) and four: MX\(_3\)). When using the formula \( \Delta T_f = -K_f m \), the molality \( m \) must account for the total number of moles of all solute species added to the solvent. For example, if 1 mol of NaCl added to water completely dissociates into Na\(^+\) and Cl\(^-\), then \( m \) should count 2 mol of ions. If the substance added does not dissociate, then \( m \) only counts the number of moles of undissociated substance. If the substance partially dissociates, then we must add all the moles of reactants and products according to the chemical equilibrium describing the partial dissociation of the solute added.

Table 3.2 lists freezing point depressions and boiling point elevations for different solvents.

<table>
<thead>
<tr>
<th>Solvent</th>
<th>Formula</th>
<th>( T_b ) ((^\circ)C)</th>
<th>( K_b ) (K/kg/mol)</th>
<th>( T_f ) ((^\circ)C)</th>
<th>( K_f ) (K/kg/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetic acid</td>
<td>CH(_3)COOH</td>
<td>118.1</td>
<td>3.07</td>
<td>17</td>
<td>3.9</td>
</tr>
<tr>
<td>Benzene</td>
<td>C(_6)H(_6)</td>
<td>80.1</td>
<td>2.53</td>
<td>5.5</td>
<td>4.0</td>
</tr>
<tr>
<td>Carbon tetrachloride</td>
<td>C(_4)Cl(_4)</td>
<td>76.7</td>
<td>5.03</td>
<td>-22.9</td>
<td>32</td>
</tr>
<tr>
<td>Diethyl ether</td>
<td>C(<em>4)H(</em>{10})O</td>
<td>34.7</td>
<td>2.02</td>
<td>-116.2</td>
<td>1.8</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C(_2)H(_5)OH</td>
<td>78.4</td>
<td>1.22</td>
<td>-114.7</td>
<td>1.9</td>
</tr>
<tr>
<td>Naphthalene</td>
<td>C(_10)H(_8)</td>
<td>—</td>
<td>—</td>
<td>80.5</td>
<td>6.8</td>
</tr>
<tr>
<td>Water</td>
<td>H(_2)O</td>
<td>100</td>
<td>0.512</td>
<td>0.0</td>
<td>1.86</td>
</tr>
</tbody>
</table>

### 3.13. Binary Mixtures of Liquids

Suppose that we have an “ideal binary mixture”, i.e. two liquids (A and B) that behave together as ideal solution. Both liquids are volatile and thus generate a non-zero vapor pressure. Figure 3.35 shows vapor pressure vs mole fraction of A. Traces shown are for the two partial vapor pressures of the mixture components (A, B) as well as the total vapor pressure, which is the sum of both partial pressures. For an ideal solution the Raoult’s and Henry’s laws have the same form, featuring the same slope. The only difference is that Raoult’s law refers to the solvent whereas Henry’s law refers to the solute. Otherwise, their form is identical (i.e., they both describe straight line behavior vs mole fraction).

Figure 3.36 shows the behavior of a real solution compared to an ideal solution. The real solution is the curved trace. The ideal solution is the
Figure 3.34. Freezing point vs molality. Molality as the number of moles of solute added (per kg solvent). The red colored lines give the observed depression of the freezing point of water by acetic acid, NaCl, and FeCl$_3$ as the molality of the solutions increases. Straight blue lines sketch the predicted ideal behavior for one through four moles of particles per mole of solute added to the solution. The experimental curve for NaCl (which gives two moles of dissolved particles per mole of NaCl added) stays close to the ideal straight line for MX; the experimental curve for FeCl$_3$ (which gives four moles of dissolved particles) stays fairly close to the ideal straight line for MX$_3$. The pattern suggests that acetic acid dissolves to give one mole of particles per mole of solute. As molality increases, the observed freezing-point depressions deviate in varying ways from the straight lines.

straight line. There are two straight lines: one for Raoult’s law and one for Henry’s law. The Raoult’s law is the straight line that is tangent to the real solution curve near mole fraction $X_B = 1$ (the limit of nearly pure solvent). It extrapolates to zero mole fraction. The Henry’s law is the case of a dilute solution, and refers to the solute. It is a straight line that is tangent to the curve at $X_B = 0$ and extrapolates at $X_B = 1$ to the pressure required to liquify the solute substance at the same temperature. Thus, the two slopes are different and carry different meanings. (The Henry’s law constant is a measure of the volatility of the solute in a particular solvent, whereas the Raoult’s law constant is just the vapor pressure of the pure solvent.)

Figure 3.37 illustrates the behavior of three different “real solutions”: carbon disulfide in acetone, benzene in methylbenzene and chloroform in acetone. For carbon disulfide in acetone, there is a positive deviation from the ideal
3.13. Binary Mixtures of Liquids

![Graph of vapor pressures and mole fractions](image)

**Figure 3.35.** The total vapor pressure and the two vapor pressures of an ideal binary mixture are proportional to the mole fractions of the components.

![Graph of ideal dilute solutions](image)

**Figure 3.36.** When a component (the solvent) is nearly pure, it has vapor pressure that is proportional to its mole fraction with a slope $P_B^0$ (Raoult’s law). When it is the minor component (the solute) its vapor pressure is still proportional to the mole fraction, but the constant of proportionality is now $K_B$ (Henry’s law).

solution. Neither substance likes each other, and intermolecular forces are weak, hence the high volatility.

For benzene in methylbenzene, we see behavior that is close to ideal solution. Both substances are very similar in terms of intermolecular bonding, yet the attraction forces are moderate. See Fig. 3.38.

For chloroform in acetone, the intermolecular forces between them are strong, as both substances are polar. Hence, the deviations are negative. See Fig. 3.39.
3. Properties of Matter

Figure 3.37. Strong deviations from ideality are shown by dissimilar liquids (in this case carbon disulfide and acetone, propanone).

Figure 3.38. Two similar liquids, in this case benzene and methylbenzene (toluene), behave almost ideally, and the variation of their vapor pressures with composition resembles that for an ideal solution.

3.13.1. Volatile Solutes: Distillation. Henry’s law applies to volatile solutes, because without volatility, the partial pressure of the solute in the gas space would be zero. Henry’s law enables us to calculate the amount of solute dissolved in a solution (mole fraction) based on the knowledge of its vapor pressure and the Henry’s law constant for that substance and solvent at a given temperature.

Using Henry’s law we can explain the phenomena of distillation, which is a method for separating component substances from liquid mixtures through vaporization and condensation. It is a separation process that is based on differences in volatility (vaporization point) of components in the mixture. The main applications of distillation include separation of crude oil into its components, desalination of sea water, purification of liquid chemicals (e.g.
solvents) and the production of distilled beverages with a higher alcohol content.

The equipment used for distillation is called a still (Fig. 3.40). The process is based on differences in volatility, as quantified by the vapor pressure of the substances at some given temperature.

3.13.2. Example. Consider a mixture of hexane and heptane. At ambient temperatures, the vapor pressure of pure hexane is 0.198 atm whereas that
of pure heptane is 0.060 atm. Suppose that a mixture consists of mole fractions $X_{\text{hexane}} = 0.4$ and $X_{\text{heptane}} = 0.6$, the partial pressures are:

$$P_{\text{hexane}} = X_{\text{hexane}}P^0_{\text{hexane}} = (0.4)(0.198 \text{ atm}) = 0.0792 \text{ atm}$$

$$P_{\text{heptane}} = X_{\text{heptane}}P^0_{\text{heptane}} = (0.6)(0.060 \text{ atm}) = 0.0360 \text{ atm}$$

From Dalton’s law the total pressure is $P = P_{\text{hexane}} + P_{\text{heptane}} = 0.1152$ atm. The mole fractions in the vapor are:

$$X'_{\text{hexane}} = \frac{0.0792 \text{ atm}}{0.1152 \text{ atm}} = 0.688$$

$$X'_{\text{heptane}} = 1 - X'_{\text{hexane}} = 0.312$$

Thus, the vapor is enriched in the more volatile component. In a temperature-mole fraction plot (see Fig. 3.41), the distillation process at constant total pressure is illustrated. The component with the lower vapor pressure (component 2) has the higher boiling point ($T^b_2$). If a solution starts at mole fraction 0.6 (see plot) and we raise its temperature until it touches the liquid line in the plot, the vapor in equilibrium with the solution is richer in the more volatile component 1. Its composition lies at the intersection of the horizontal constant-temperature line and the equilibrium vapor curve. Component 1 is enriched, as indicated by the component 2 mole fraction which is down to 0.312. The process is repeated until the desired enrichment is attained.

![Figure 3.41](image_url). The boiling point of an ideal solution varies with the composition of the solution. The upper curve is the boiling temperature referred to the vapor composition, and the lower curve is the boiling temperature referred to the liquid composition. The vapors boiling off a solution that has a 0.600-mol fraction of component 2 are enriched in the more volatile component 1 to the extent that their mole fraction of component 2 is only 0.312 (red arrow). The subsequent blue arrows show the further steps used in obtaining nearly pure component 1 by fractional distillation.
3.14. Osmotic Pressure

Osmosis describes the transport of small molecules through membranes. It is important in biology and reverse osmosis (a commercial process to desalinize water). Suppose a solution is contained in an inverted tube, the lower end of which is covered by a semipermeable membrane. This solution has a solute concentration of $c$ moles per liter. When the end of the tube is inserted in a beaker of pure solvent (see Fig. 3.42), solvent flows from the beaker into the tube. The volume of the solution increases, and the solvent rises in the tube until, at equilibrium, it reaches a height, $h$, above the solvent in the beaker.

![Figure 3.42. Osmotic pressure builds up when concentrations across a membrane are unequal.](image)

The pressure on the solution side of the membrane is greater than the atmospheric pressure on the surface of the pure solvent by an amount given by the osmotic pressure, $\pi$:

$$\pi = \rho gh$$

where $\rho$ is the density of the solution (water: 1.00 g/cm³) and $g$ is the gravitational acceleration constant (9.8 m/s²). For example, $h=0.17$ m in a dilute aqueous solution gives:

$$\pi = (1.00 \text{ g/cm}^3)(10^{-3} \text{ kg/g})(10^6 \text{ cm}^3\text{m}^{-3})(9.8 \text{ m/s}^2)(0.17 \text{ m})$$

$$\approx 1.7 \times 10^3 \text{ kg/m/s}^2 = 1.7 \times 10^3 \text{ Pa} = \frac{1.7 \times 10^3 \text{ Pa}}{1.013 \times 10^5 \text{ Pa/atm}} = 0.016 \text{ atm}.$$
van’t Hoff (1887) discovered a relationship between osmotic pressure $\pi$, concentration $c$ and absolute temperature $T$:

$$\pi = cRT$$

since $c = n/V$ we can write this in a form similar to the ideal gas law:

$$\pi V = nRT.$$

The following example illustrates the use of the van’t Hoff formula to find the molar mass of a chemical species, if an osmometer provides the osmotic pressure in units of atm.

**3.14.1. Example.** We dissolve 2.00 g of a protein in 0.100 L water. The osmotic pressure is 0.021 atm at 25°C. What is the approximate molar mass of the protein?

**Solution:** The concentration in moles per liter is

$$c = \frac{\pi}{RT} = \frac{0.021 \text{ atm}}{(0.08206 \text{ L.atm.mol}^{-1}.\text{K}^{-1})(298 \text{ K})} = 8.6 \times 10^{-4} \text{ mol/L}$$

The number of moles ($n = c \cdot V$, $V=0.100$ L) is:

$$n = c \cdot V = 8.6 \times 10^{-4} \text{ mol/L} \cdot 0.100 \text{ L} = 8.6 \times 10^{-5} \text{ mol}.$$

The molar mass, by definition, is:

$$M = \frac{m}{n} = \frac{2.00 \text{ g}}{8.6 \times 10^{-5} \text{ mol}} = 23,000 \text{ g/mol}.$$

**Remark:** If you are confused about the constant $R = 0.08206 \text{ L.atm/mol/K}$, when we earlier using $R = 8.31 \text{ J/mol/K}$, the relationship between the two is easily obtained by realizing that J (Joules) are a unit of energy, and energy is $PV$ (pressure $\times$ volume). In SI units, the units of pressure are Pa, whereas those of volume are m$^3$. Converting 1 J=$(1 \text{ Pa}) \cdot (1 \text{ m}^3)$ to L.atm:

$$(1 \text{ Pa}) \cdot (1 \text{ m}^3) \cdot \frac{1 \text{ atm}}{101325 \text{ Pa}} \cdot \frac{1000 \text{ L}}{1 \text{ m}^3} = \frac{1}{101.325} \text{ L.atm}$$

So that the gas constant becomes (in the new units):

$$8.31 \text{ J/mol} \cdot \frac{1}{101.325} \text{ L.atm} = 0.08206 \text{ L.atm/mol/K}.$$
3.15. Osmosis and Reverse Osmosis

Osmosis and Reverse Osmosis involve the movement of water molecules through a semi-permeable membrane. A semi-permeable membrane allows only water molecules to pass. Due to the difference in concentrations on either side of the membrane, water molecules move from a concentrated solution to a less concentrated solution, creating osmotic pressure. This pressure equalizes when the system reaches equilibrium.

**Figure 3.43.** Osmotic pressure builds up when concentrations across a membrane are unequal. This schematic shows the situation before (left) and after (right) equilibrium is reached.

Reverse osmosis (Fig. 3.44) can be described in terms of chemical potentials. A potential gives rise to a flux or force of some kind. In a way analogous to the gravitational potential which leads to objects falling from high places and the electric potential which leads to the flow of charges between electrodes, the chemical potential leads to chemical or physical transformations involving molecules. It is differences in the potential that lead to these forces. For example, we need a positively charged electrode and a negatively charged electrode for electric currents to flow through a conductive medium. Likewise, differences in the chemical potential of a substance between two compartments (or points in space) will lead to the flow of molecules.
In general, the chemical potential of a substance indicates its potential for physical or chemical transformation. In the absence of a chemical reaction, chemical potential differences drive physical transport such as diffusion. You can think of the chemical potential as indicative of the concentration of a given molecular species. Diffusion describes the random movement of chemical species. The consequence of this random motion is that it leads to uniform distributions of concentrations: gases diffuse away from regions of high concentrations toward regions of low concentrations\textsuperscript{12}. Equilibrium is reached when the concentration of each chemical species is uniform everywhere in space. Thus, we can think of each substance as having a chemical potential which indicates the local concentration of this particular substance.

In Fig. 3.44 the solute is found in high concentration in the right-most compartment and has zero concentration in the left-most compartment. Thus, the solute wants to move from the right compartment to the left compartment. However, the semi-permeable membrane in the middle does not allow this diffusion to take place. Therefore, the flux of solute particles toward the left is zero because of this membrane. On the other hand, there is a higher concentration of solvent in the left compartment compared to the right compartment. This is because of the presence of solute molecules in the right compartment, which leads to a lower concentration of solvent molecules. Thus, the solvent wants to move from the left compartment toward the

\textsuperscript{12}The flux of the \textit{i}th component, $J_i$, is driven by gradients in the chemical potentials according to relationships of the type $J_i = -\sum_k M_{ik} \nabla \mu_k$, where $\mu_k$ is the chemical potential of the \textit{k}th component and $M_{ik}$ is a coupling coefficient between fluxes in \textit{i} and gradients in \textit{k}. In the case where there is no coupling between the components ($M_{ik} = 0$ for $i \neq k$) we have $J_i = -M_{ii} \nabla \mu_i$. 

\textbf{Figure 3.45.} Experimental apparatus for reverse osmosis.
right compartment, in an attempt to equate the solvent concentrations everywhere. This flux of solvent from the left to the right leads to a higher pressure on the right compartment: the osmotic pressure. At equilibrium, the chemical potentials of the solvent are equal in both compartments. This causes the gradient (difference) in chemical potentials to be zero and the driving force vanishes.

**Reverse osmosis:** What happens if we apply an external pressure to overcome the osmotic pressure? In this case, the solvent will want to move from the top compartment to the bottom compartment. The solute will also want to move from the top to the bottom. However, it will not be able to do so because of the semi-permeable membrane. Thus, only the solvent will be transferred. This process leads to an accumulation of pure solvent in the bottom compartment. This process, called *reverse osmosis*, can be used to purify water and other solvents.

Reverse osmosis is a technique that can be used to purify water. When a sufficiently high pressure is applied to overcome the osmotic pressure across a membrane, water is forced back through the membrane, leaving ions and other dissolved substances behind. In a sense, this method filters out the ions and other unwanted substances at the molecular level.

Even after municipal water treatment, drinking water still contains heavy metal ions such as Pb\(^{2+}\), Cd\(^{2+}\) and Hg\(^{2+}\) that are toxic. Commercial reverse osmosis systems can be purchased for domestic use to remove these toxic ions. Reverse osmosis can also be used to desalinate (remove ions) from seawater. Seawater is pumped under high pressure into tubes containing hollow fiber membranes. Water molecules (but not ions) pass through the membranes into the fiber to be collected. This method can purify seawater to less than 400 ppm of total dissolved solids in one pass, making it suitable for drinking.

### 3.16. Problems

**Problem 43.** The average molar masses of unknown polymers can be determined through osmotic pressure experiments. Several aqueous solutions of non-volatile poly(ethylene) oxide of an unknown molar mass were prepared and the osmotic pressure expressed in terms of height of the solvent in a thin tube according to the experimental setup below. You may take the density of water to be 1.004 g cm\(^{-3}\), \(P = 1.00\) atm, and \(T = 298.15\) K.
3. Properties of Matter

<table>
<thead>
<tr>
<th>concentration (g L(^{-1}))</th>
<th>2.042</th>
<th>6.613</th>
<th>9.521</th>
<th>12.602</th>
</tr>
</thead>
<tbody>
<tr>
<td>height (cm)</td>
<td>0.592</td>
<td>1.910</td>
<td>2.750</td>
<td>3.600</td>
</tr>
</tbody>
</table>

Calculate the molar mass of the dissolved poly(ethylene) oxide sample. Hint: plotting the data will be helpful.

**Solution.** From the graph of height vs concentration, a straight line is fit through the data points. We should force the intercept to be zero, since \( \pi = cRT \) has no offset. Rearranging,

\[
\pi = cRT = \frac{mRT}{V\mathcal{M}} \quad \text{and} \quad \pi = \rho gh
\]

Therefore,

\[
h = \left( \frac{RT}{\rho g \mathcal{M}} \right) c
\]

The slope of the fitting curve is equal to the above quantity in parentheses. Rearranging, converting to mks units, and plugging in the slope,

\[
\mathcal{M} = \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298.15 \text{ K})}{(1.004 \times 10^3 \text{ kg m}^{-3})(9.81 \text{ m s}^{-2})(0.29 \times 10^{-2} \text{ m}^4 \text{ kg}^{-1})} = 87 \text{ kg mol}^{-1}
\]

**Problem 44.** You determine the osmotic pressure of an aqueous solution at 300 K to be 120 kPa. Calculate the freezing point of the solution based on this data.

**Solution.** The freezing point depression equation

\[
\Delta T_f = K_f m
\]

and osmotic pressure equation

\[
\pi = cRT
\]

can be combined based on the following relationships

\[
m = \frac{n_{\text{solute}}}{\text{kg H}_2\text{O}} \approx \frac{n_{\text{solute}}}{V_{\text{H}_2\text{O}}} = \frac{n}{V}, \quad n = \frac{\pi V}{RT}
\]
to form the following expression for temperature change

\[ \Delta T_f = \frac{K_f \pi}{RT\rho} \]

The freezing point depression constant and density for H\(_2\)O are \(K_f=1.86 \) K.kg.mol\(^{-1}\) and \(10^3 \) kg.m\(^{-3}\), respectively.

\[ \Delta T_f = \frac{(1.86 \text{ K.kg.mol}^{-1})(120 \times 10^3 \text{ Pa})}{(8.314 \text{ J.K}^{-1} \text{ mol}^{-1})(300. \text{ K})(10^3 \text{ kg.m}^{-3})} = 0.0895 \text{ K} \]

so solution will freeze at \(-0.0895^\circ\text{C}\). \[\blacksquare\]

**Problem 45.** Each increase in pressure of 100 atm decreases the melting point of ice by about 1.0°C.

(a) Estimate the temperature at which liquid water freezes under a pressure of 400 atm.

(b) One possible explanation of why a skate moves smoothly over ice is that the pressure exerted by the skater on the ice lowers its freezing point and causes it to melt. The pressure exerted by an object is the force (its mass \(\times\) the acceleration of gravity, 9.8 m.s\(^{-2}\)) divided by the area of contact.

Calculate the change in freezing point of ice when a skater with a mass of 75 kg stands on a blade of area \(8.0 \times 10^{-5}\) m\(^2\) in contact with the ice. Is this sufficient to explain the ease of skating at a temperature of, for example, \(-5^\circ\text{C}\) (23°F)?

**Solution.** (a)

\[ T_f = 0^\circ\text{C} + \Delta T_f = 0^\circ\text{C} - \left(\frac{400}{100}\right) \times 1^\circ\text{C} = -4^\circ\text{C} \]

(b)

\[ P = \frac{(75 \text{ kg})(9.8 \text{ m.s}^{-2})}{8.0 \times 10^{-5} \text{ m}^2} \times \frac{1 \text{ atm}}{101325 \text{ Pa}} = 90.67 \text{ atm} \]

No, this pressure is not high enough for this effect to explain the relative ease of skating at \(-5^\circ\text{C}\). \[\blacksquare\]

**Problem 46.** Glycerol was once used in antifreeze because it can lower the freezing point of water below 0°C. What mass percentage of glycerol in water must be used to reduce the freezing point of the mixture to \(-5.0^\circ\text{C}\), assuming ideal solution behavior? (\(K_f = 1.87 \) K.kg/mol.)

**Solution.** Using \(\Delta T_f = -K_f m\) with \(\Delta T_f = -5.0 \) K and the value of \(K_f\) provided, yields \(m = 2.7 \) mol/kg (moles of glycerol per kg of solvent). Multiplication of \(m\) by the molecular weight of glycerol (92.09 g/mol) gives 248.6 g glycerol per kg (1,000 g) solvent. (Let’s round this to 250 g to get round numbers; in practice, don’t round.) We have 250 g glycerol in 1,000 g water.
The total mass of the solution is 250 g + 1,000 g = 1,250 g. The mass fraction is 250 / 1,250 = 0.20, or 20% of total solution by mass. ■

Problem 47. Show that for one nonvolatile ionic molecule, that completely ionizes into x ions per molecule, the formula for the vapor pressure lowering is \( \Delta P = -xX_{salt}P^o_B \) as the limit of the solute concentration approaches zero. \( X_{salt} \) is the ratio of the number of moles of the ionic salt in solution to the total number of moles if the compound were to not ionize.

Solution. \( n_0 := \) moles of H\(_2\)O 
\( n_1 := \) moles of ionic compound 
\( X_B := \) mole fraction when the compound completely dissociates 
\( X_{salt} := \) mole fraction the compound does not dissociate 
\[
X_{salt} = \frac{n_1}{n_1 + n_0} \approx \frac{n_1}{n_0}
\]
\[
X_B = \frac{xn_1}{xn_1 + n_0} \approx \frac{xn_1}{n_0}
\]
\[
\Delta P_B = -X_B P^o_B = -x\frac{n_1}{n_0} P^o_B = -x X_{salt} P^o_B
\]
■

Problem 48. You have a non-volatile solvent Q, with a molar mass of 97.80 g mol\(^{-1}\), dissolved in 0.551 pounds of water (water is a solvent as well, but it is volatile compared to Q). The solution has a vapor pressure of 23.8 torr. Take the vapor pressure of water to be 0.056 atm at the temperature of the solution. Determine how many grams of Q are present in the solution.

Solution. The only volatile component is water. Thus, the vapor pressure of the solution, \( P_{vap} \), is due to the water solvent. The Raoult’s law is applied to the water solvent: 
\[
P_{vap} = X_{solvent} P^o_{solvent},
\]
\[
X_{solvent} = \frac{P_{vap}}{P^o_{solvent}} = \frac{0.031 \text{ atm}}{0.056 \text{ atm}} = 0.56.
\]
Since we have 0.551 lb (249.92 g) of water and its molar mass is 18 g/mol, this implies we have \( n_{H_2O} = 249.92/18 = 13.9 \text{ mol} \) of water. Then,
\[
0.56 = \frac{n_{H_2O}}{n_{H_2O} + n_Q} = \frac{13.9}{n_Q + 13.9},
\]
\( n_Q = 10.9 \text{ mol}, \)
\( n_Q \times 97.8 \text{ g.mol}^{-1} = 1066 \text{ g}. \) ■
Problem 49.  A total of 10 g of nonvolatile solute are added to five moles of a solvent whose total volume is 200 cm$^3$. The freezing point of the solvent is 280 K, while its molar enthalpy of fusion is 2,000 cal mol$^{-1}$. It’s found that the solvent in the solution freezes at 279.894 K. By assuming the solution is ideal, compute the osmotic pressure at 300 K.

Solution. This is a freezing point depression problem, a colligative property (Raoult’s law). Consider the heterogeneous equilibrium where we have a solid $A(s)$ melting into a liquid $A(l)$, but the liquid phase has some amount of solute $B$ dissolved. $A(s)$ is in equilibrium with $A(l)$; this means that the chemical potentials of $A$ in the two phases are equal:

\[ \mu^*_A(s) = \mu^*_A(l) + RT \log X_A \]

where $X_A$ is the mole fraction of $A$ and $*$ denotes the pure substance. This leads to the formula for freezing point depression $\Delta T_f = K' X_B$ presented in class, $K' = RT_f^2/(\Delta_{fus}H)$, because differentiation of:

\[ \log X_A = \frac{\mu^*_A(s) - \mu^*_A(l)}{RT} = \frac{\Delta_{fus}G}{RT} \]

with respect to $T$:

\[ \frac{d \log X_A}{dT} = \frac{d}{dT} \frac{\Delta_{fus}G}{RT} \]

Then we use the Gibbs-Helmholtz equation, $\partial(\Delta G/T)/\partial T|_P = -\Delta H/T^2$, to get $d \log X_A = -\Delta_{fus}H/(RT^2)$. Multiply both sides by $dT$ and integrate from $X_A = 1$ to $X_A$, and from $T = T_f$ to $T$:

\[ \int_0^{\log X_A} d \log X_A = -\frac{1}{R} \int_{T_f}^{T} \frac{\Delta_{fus}H}{T^2} dT \]

The left hand side integrates to $\log X_A$, which equals $\log(1 - X_B)$, but for small values of $X_B$, this is approximately equal to $-X_B$. Then, assuming that $\Delta_{fus}H$ is constant across this temperature range,

\[ X_B = \frac{\Delta_{fus}H}{R} \left( \frac{1}{T_f} - \frac{1}{T} \right) \]

where $X_B = X_{\text{solute}}$. Writing $\frac{1}{T_f} - \frac{1}{T} = \frac{T - T_f}{TT_f} \approx \frac{\Delta T_f}{T_f}$ (since $T \approx T_f$), $\Delta T_f = T - T_f$, yields the result $\Delta T_f = K' X_B$, where $K' = RT_f^2/(\Delta_{fus}H)$. (Note: this convention results in $\Delta T_f > 0$.)

---

13 This problem requires Chapter 6 as a pre-requisite; it is placed here because it deals with the topic of freezing point depression. Upon first reading, you can skip the derivations and merely use the boxed equation.
Finally, \[ X_{\text{solute}} = \frac{2000 \text{ cal mol}^{-1}}{1.987 \text{ cal K}^{-1} \text{ mol}^{-1}} \left( \frac{1}{279.894 \text{ K}} - \frac{1}{280 \text{ K}} \right) = 0.0013614 \]
and
\[ \pi = \frac{RTX_{\text{solute}}}{V_n} = \frac{(0.0821 \text{ atm L.K/mol})(300 \text{ K})(0.0013614)}{\frac{0.2 \text{ L}}{5 \text{ mol}}} = 0.8379 \text{ atm} \]

NOTE: We have employed two results. The first is the Gibbs-Helmholtz equation, \( \partial(\Delta G/T)/\partial T|_P = -\Delta H/T^2 \). This equation is derived as follows:

\[
\left( \frac{\partial(G/T)}{\partial T} \right)_P = \frac{1}{T} \left( \frac{\partial G}{\partial T} \right)_P + \frac{G}{T} \frac{d(1/T)}{dT} = \frac{1}{T} \left( \frac{\partial G}{\partial T} \right)_P - \frac{G}{T^2} = \frac{1}{T} \left[ \left( \frac{\partial G}{\partial T} \right)_P - \frac{G}{T} \right]
\]

Then, from \( dG(P,T) = VdP - SdT \), \( (\partial G/\partial T)_P = -S \), by definition \( G = H - TS \), or \( S = (H - G)/T \), \( (\partial G/\partial T)_P = (G - H)/T \), we get for the square bracket:

\[
\left( \frac{\partial(G/T)}{\partial T} \right)_P - \frac{G}{T} = \frac{G - H}{T} - \frac{G}{T} = -\frac{H}{T}
\]

Thus,
\[
\frac{\partial(G/T)}{\partial T}_P = -\frac{H}{T^2},
\]

which holds for each term in \( \Delta G = G_f - G_i \); hence we can replace \( G \) by \( \Delta G \) and \( H \) by \( \Delta H \). The second result is the Gibbs energy at constant temperature \( (dT = 0) \), \( dG = VdP \), integrated from initial to final state:

\[ G(P_f) = G(P_i) + \int_{P_i}^{P_f} VdP \]

which, for an ideal gas, gives \( G(P_f) = G(P_i) + RT \int_{P_i}^{P_f} (1/P)dP = G(P_i) + RT \log(P_f/P_i) \). The corresponding molar quantity is called the molar Gibbs energy (Gibbs energy per mole of substance), or chemical potential; hence \( \mu = \mu_\oplus + RT \log(P/P_\oplus) \), where \( \oplus \) denotes a reference state. For simplicity, we write \( P \equiv P/P_\oplus \). Denoting the quantities relating to pure substances by a *, so the chemical potential of pure \( A \) is written \( \mu^*_A \) or \( \mu^*_A(l) \) to denote a liquid. Because the vapor pressure of the pure liquid is \( P_A^* \), it follows that the chemical potential of \( A \) in the vapor (treated as perfect gas) is \( \mu_A^\ominus + RT \log P_A \) (with \( P_A \), a relative pressure \( P_A/P_\oplus \)). If another substance, a solute, is also present in the liquid, the chemical potential of \( A \) in the liquid is changed to \( \mu_A \) and its vapor pressure is \( P_A \). The vapor and solvent are in equilibrium, so we can equate the chemical potentials, \( \mu_A = \mu_A^\ominus + RT \log P_A \). Combining these two equations to eliminate \( \mu_A^\ominus \), we get \( \mu_A = \mu_A^* - RT \log(P_A^*) + RT \log P_A = \mu_A^* + RT \log(P_A/P_A^*) \). Invoking Raoult’s law, \( P_A = X_A P_A^* \), we get \( \mu_A = \mu_A^* + RT \log X_A \). This derivation is from Atkins [14].
Problem 50. Let’s estimate the rate of evaporation for water at RT. 1) Calculate the energy, $\Delta H$, that binds water molecules in a liquid. You can find more information about the structure of water at the link below:

http://www1.lsbu.ac.uk/water/water_hydrogen_bonding.html

Here are some figures showing tetrahedrally coordinated water molecules:

As always, such structures are understood in the sense of time-averages. The actual picture is very dynamic. You may also consult any other useful articles you may find. 2) Then calculate the fraction, $f$, of water molecules with kinetic energies sufficiently large to break the hydrogen bonding structure apart and escape to the gas phase. Let’s assume that water molecules in the liquid phase also obey Maxwell-Boltzmann statistics. Feel free to use Wolfram Alpha to compute the relevant integrals numerically. 3) Use the fraction calculated in #2 to estimate the rate of evaporation in kg per square meter per second. The approach we shall use is loosely inspired by the Langmuir theory of adsorption on surfaces. How many molecules are there per unit area on the water’s surface? Call this quantity, $\sigma$, for area density (units: # molecules/m$^2$). Consider the following formula:

$$E = \frac{1}{6} \cdot \sigma \cdot \tau_0^{-1} e^{-\Delta H/RT} \cdot f$$

The factor of $1/6$ is a geometric factor I’ve added to account for the fraction of water molecules moving upwards (imagine a cube with 6 faces, and only those moving towards the top face go towards the gas region; the rest stay in solution). The factor $f$ gives the “area coverage”, that is, the fraction of water molecules that are eligible to escape from the solution because their kinetic energies are large enough to enable a jump. $\tau_0$ is a “residence time”, which we take to be the average lifetime of a water molecule in a hydrogen-bonded cluster. The paper below cites hydrogen bond lifetimes in the range 1-20 ps:

https://arxiv.org/pdf/0706.1355

Let’s take $\tau_0=20$ ps, the worst case scenario (very long lived clusters). The Arrhenius factor $e^{-\Delta H/RT}$ gives the fraction of eligible water molecules that will actually jump (at a given temperature $T$); this fraction grows larger with $T$. $\Delta H$ is the energy barrier required to break the hydrogen bonded cluster (see part #1). You can check that the units of $E$ are molecules/m$^2$/s.
3. Properties of Matter

Solution. 1) Per articles
https://arxiv.org/pdf/0706.1355
http://www1.lsbu.ac.uk/water/water_hydrogen_bonding.html
the energy required to break the bonds is on the order of a few kJ/mol. One value cited was 9.80 kJ/mol. Let’s take 9.80 kJ/mol.

2) For water, \( M = 18 \text{ g/mol} = 0.018 \text{ kg/mol} \).

\[
\sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8(8.31)(293)}{(\pi \cdot 0.018)}} \approx 587 \text{ m/s}.
\]
This is the average speed of a water molecule. From part 1), the kinetic energy must exceed 9.80 kJ/mol. From \( K = \frac{1}{2}Mv^2 \), this corresponds to a speed of \( \sqrt{2K/M} = \sqrt{2(9800 \text{ J/mol})/(0.018 \text{ kg/mol})} = 1043 \text{ m/s} \).

We must integrate

\[
f(v) = \int_{1043}^{\infty} \frac{(M/2\pi RT)^3}{4\pi^2} v^2 e^{-Mv^2/(2RT)} dv
\]

from 1043 m/s to \( \infty \), using \( T = 293 \text{ K} \):

\[
f = \sqrt{\left(\frac{M}{2\pi RT}\right)^3} \int_{1043}^{\infty} v^2 e^{-Mv^2/(2RT)} dv
\]

Numerically, the prefactor \( \sqrt{\left(\frac{M}{2\pi RT}\right)^3} \) evaluates to \( 1.276 \times 10^{-9} \).

Also, \( M/(2RT) = 3.696 \times 10^{-6} \). In Wolfram Alpha, the integral can be evaluated by typing:

\[\text{integrate} \ (v^2)*\exp(-v^2*(3.696e-4)) \text{ from 1043 to infinity}\]

This gives the result: \( 2.81 \times 10^6 \). Multiplication by \( 4\pi \times 1.276 \times 10^{-9} \) gives:

\[f = 0.045\]

so about 4.5% of molecules have enough kinetic energy to escape the solution and go into the gas phase.

3) Water has mass density \( \rho_m = 1 \text{ g/cc} = 10^6 \text{ g/m}^3 = 10^3 \text{ kg/m}^3 \). Number density is \( (10^3 \text{ kg/m}^3) \cdot (6.022 \times 10^{23} \text{ molecules/mol})/(0.018 \text{ kg/mol}) \).

Elevating this to the power \( 2/3 \) gives an area density of \( \sigma = 1.038 \times 10^{19} \text{ molecules/m}^2 \).

\[
E = \frac{1}{6} \sigma \cdot \tau_0^{-1} \cdot e^{-\Delta H_{\text{vap}}/RT} \cdot f = 1.39 \times 10^{27} \text{ molecules/m}^2/\text{s}.
\]

where \( \Delta H_{\text{vap}} = 9.8 \text{ kJ/mol} \). \( RT = (8.31 \text{ J/mol/K})(293 \text{ K}) = 2.434 \text{ kJ/mol} \). Then, \( e^{-\Delta H_{\text{vap}}/RT} \approx 0.01784 \). To get units of \( \text{kg/m}^2/\text{s} \) we must convert molecules to kg, using molar mass (0.018 kg/mol) and Avogadro’s constant:

\[
E_{\text{mass}} = \frac{E}{N_A} \cdot (0.018 \text{ kg/mol}) = 2.07 \text{ kg/m}^2/\text{s}
\]
This value is close to the experimental value of 1.97 kg/m²/s cited here: http://likesundanese.blogspot.com/2013/09/water-evaporation-rate-per-surface-area.html

**Problem 51.** Use the evaporation rate, $E$, you found in Problem 50 to estimate the amount of water evaporated around the planet by all bodies of water during 1 day. (You will need to look up this information about the total surface area of the Earth that is water.) Assume that all water is at RT and that the temperature is uniform (held at RT) everywhere. (This assumption is probably OK, since the average temperature on Earth is 16°C, per https://www.space.com/17816-earth-temperature.html, so not too far off from RT.) Compare the amount of water evaporated in 1 day to the total rainfall in 1 day. (Look up the average annual rainfall, and divide by 365 days to get the daily rate.) These two numbers should balance each other, right? (If not, explain why not.)

**Solution.** About 70% of the Earth’s surface is water-covered (with the oceans accounting for 95.6% of all Earth’s water). The surface area of the Earth is 196.9 millions of square miles, or $5.1 \times 10^8$ km². So 70% of this works out to $361$ million km² (or 139.4 million mi²). This data is from: https://www.universetoday.com/25756/surface-area-of-the-earth/

Multiplication of $E$ by 361×10^{12} m² and by 1 day (86,400 seconds) gives:

$$(2.07 \text{ kg/m}^2/\text{s})(361 \times 10^{12} \text{ m}^2)(86,400 \text{ s}) = 6.45 \times 10^{19} \text{ kg}$$

The Earth’s average annual precipitation is approximately $5.36 \times 10^{14}$ m³. Source: https://hypertextbook.com/facts/2008/VernonWu.shtml

Divided by 365 days, this gives $1.47 \times 10^{12}$ m³ per day. Because water has density of 1000 kg/m³, this corresponds to $1.47 \times 10^{15}$ kg of water.

The estimate is off by 4 orders of magnitude ($1.47 \times 10^{15}$ kg vs. $6.45 \times 10^{19}$ kg). We neglected the rate of condensation. We can estimate the rate of condensation if we know the amount of water in the air (humidity level), [H₂O] (in moles/m³), the average (say, rms) speed $v_{rms}$ of the air molecules, and $\gamma$ the uptake coefficient. $\gamma$ is the probability that a molecule hitting the water surface will be reabsorbed into the solution. Then,

$$J = \frac{\gamma}{6} \cdot v_{rms} \cdot [\text{H}_2\text{O}]$$

where the factor 1/6 is introduced for the same reason as before. Let’s assume the air temperature is the same as the solution temperature and take $\gamma = 1$. Taking $v_{rms} = 1043$ m/s. According to https://www.cactus2000.de/uk/unit/masshum.shtml

if the relative humidity were 51.73%, which appears to be some ideal value “comfortable to humans”

https://en.wikipedia.org/wiki/Relative_humidity
the water concentration at normal pressure and temperature (25°C and 101.325 kPa) is $3.9798 \times 10^{17}$ molecules/cc, or $3.9798 \times 10^{23}$ molecules/m$^3$. Then, $J$ in units of kg/m$^2$/s is:

$$J = \frac{(0.18 \text{ kg/mol})}{N_A} \cdot (1043) \cdot (3.9798 \times 10^{23}) \cdot \frac{0.018}{6.022 \times 10^{23}} \approx 2.067 \text{ kg/m}^2/\text{s}$$

Then, condensation accounts for the difference. Note: the choice of 51.73% is semi-arbitrary here and the exact value chosen matters. The moral of the story here is that precipitations alone do not account for the difference, by a long shot. We must include the rate of condensation at the very least. ■

**Problem 52.** Suppose you want to build a swimming pool in your backyard (assume a surface area of 500 square foot) and you want to estimate how much water it will consume. How many gallons of water per day do you lose by evaporation? (Assume RT.) Note: this calculation (and the one in Problem 51) neglects the effects of direct sunlight.

**Solution.** 500 square feet = 46.5 square meters. Multiplication of $E$ by 46.5 m$^2$ and by 1 day (86,400 seconds) gives:

$$(2.07 \text{ kg/m}^2/\text{s})(46.5 \text{ m}^2)(86,400 \text{ s}) = 8.32 \times 10^6 \text{ kg}$$

Again, this number is much too high. If instead, we account for condensation and air humidity (51% r.h.), we get a much smaller loss

$$(2.07 - 2.067 \text{ kg/m}^2/\text{s})(46.5 \text{ m}^2)(86,400 \text{ s}) = 12,053 \text{ kg}$$

A pool of this size (say 5 m wide x 10 m long, 1.5-3.5 m deep) may have a volume of 125,000 liters (kg). This represents about 10% loss of water volume per day. Losses may be less if the pool temperature is lower. ■

**Problem 53.** Freezing point depression. The addition of a (non-volatile) solute to the solvent lowers the freezing point because the solute impedes the formation of the crystalline solid. This is mainly an entropic effect that describes the space (configurations) occupied by the solute, thereby forcing the solvent molecules to rearrange themselves around the solute particles to accommodate them. Recall that the freezing point depression is given by the formula $\Delta T_f = K_f \cdot m$, where $m$ is the molality of the solute and $K_f$ is a constant. (With this convention, $\Delta T_f > 0$.) This formula is a statement of the fact that $\Delta T_f$ a colligative property. (a) Explain why? $K_f$ has a name; it is called the cryoscopic constant. It can be shown that the cryoscopic constant can be written in the form:

$$K_f = \frac{MRT_f^2}{\Delta H_{fus}}$$
where $\Delta H_{fus}$ is the enthalpy of fusion, $M$ is the molar mass of the solvent, $R$ is the gas constant, $T_f$ is fusion temperature of the pure solvent. It can also be shown

www4.ncsu.edu/~franzen/public_html/CH433/lecture/

Colligative_Prop_h.pdf


that the enthalpy of fusion is related to the entropy of fusion, $\Delta S_{fus}$:

$$\Delta S_{fus} = \Delta H_{fus}/T_f,$$

because $\Delta G_{fus} = \Delta H_{fus} - T_f \Delta S_{fus} = 0$ at equilibrium. The entropy of fusion, $\Delta S_{fus} = S(\text{liquid}) - S(\text{solid})$, is the increase in entropy when melting a substance. For a pure substance, this is almost always positive [meaning that $S(\text{liquid}) > S(\text{solid})$] since the degree of disorder increases in the transition from an organized crystalline solid to the disorganized structure of a liquid; the only known exception is helium. Suppose for simplicity that $S(\text{liquid})$ does not change significantly with addition of the solute [i.e. it may change a bit, but not as much as $S(\text{solid})$, so that $S(\text{liquid})$ is always larger than $S(\text{solid})$; this is, of course, an approximation, but good enough to illustrate the main point].

Thus, $K_f$ (and thereby, $\Delta T_f$) depend on $\Delta S_{fus}$. Let us look at the origins of these entropic effects from a very simple picture. Our task here shall be to examine the effect of solute addition on $S(\text{solid})$. Imagine a 2D fluid, for example, a 3×3 matrix (square array). When the solvent is pure, this array looks like: $\circ \circ \circ \circ \circ \circ \circ \circ \circ$, where $\circ$ denotes a solvent particle. When a solute particle is added, this matrix may look like $\bullet \circ \circ \circ$, where $\bullet$ denotes a solute particle. In fact, there can be 9 different configurations, corresponding to all possible ways to insert a $\bullet$ in any of the 9 positions. The entropy of the solid, according to Boltzmann, is proportional to the logarithm of the number of configurations of fixed energy, $\Omega$: $S(\text{solid}) = k_B \log(\Omega)$. (You can take log to be the natural log.) For pure solvent, $\Omega = 1$, since there is only $\circ \circ \circ$ 1 possible configuration, namely, $\circ \circ \circ$. Since $\log(1) = 0$, $S_0 = 0$. For 1 solute particle, $\Omega = 9$ and $S_1 = k_B \log(9)$. Thus, $S_1 > S_0$, i.e. the entropy for 1 solute particle is greater than the entropy of the pure solvent.

(b) What do you conclude about the effect of solute on $S(\text{solid})$, and its impact on $\Delta T_f$, the freezing point depression? Now let’s go to 3D. Imagine a 3×3×3 array/cube (27 total lattice positions).
3. Properties of Matter

(c) What is the entropy, $S_0$, of the pure solvent. What is the entropy, $S_1$, of the solution containing 1 solute particle? What is the entropy, $S_2$ of the solution containing 2 particles? etc. (Continue this process up to $S_{27}$, the case of the pure solute.) What can you conclude about $S_n$, $n = 1, \ldots, 27$, and its effect on $\Delta T_f$, the freezing point depression?

(d) Here, we have considered a solid of size $3 \times 3 \times 3$ for simplicity; what happens in the limit of large volume $V$?

Solution. (a) Colligative property only depends on the mole fraction of the solute, not on its nature. (b) As solute is added, $S(solid)$ increases, up to the point where the amount of solute equals the amount of solvent (50-50 mixture). At the same time, $S(liquid) - S(solid)$ decreases. This causes $\Delta T_f$ to increase.

(c) There are $\Omega(27, 1) = \binom{27}{1} = 27$ ways of placing 1 particle in a $3 \times 3 \times 3$ lattice, $\Omega(27, 2) = \binom{27}{2} = 27 \cdot 26/2 = 351$ ways of placing 2 particles in a $3 \times 3 \times 3$ lattice, $\Omega(27, 3) = \binom{27}{3} = 27 \cdot 26/2 = 2925$ ways of placing 2 particles in a $3 \times 3 \times 3$ lattice, $\ldots$, $\Omega(27, 26) = 27$, $\Omega(27, 27) = 1$. So the entropy increases as we add more solute, up to the point there the amount of solute equals the amount of solvent (at which point the solute becomes the solvent), followed by a decrease. As far as the freezing point depression, $\Delta T_f$, is concerned, the effect becomes more pronounced as more solute is added.

(d) In the limit of large volume, the same idea applies to the freezing point depression. The discrete variables become continuous parameters.

Problem 54. Find out what conditions $(T, P)$ are used to carry out ammonia synthesis in industrial processes. Explain why the reaction is carried out under those conditions. How do these conditions affect the chemical equilibrium?

Solution. The conditions are approximately 200 atm and 450°C. Carrying out the reaction at high pressure shifts the equilibrium towards the product, per Le Châtelier’s principle. However, the reaction is exothermic, so lower temperatures are better. However, the reaction rate can be pretty slow at low temperatures. Why 450°C then? At high temperatures, the reaction rate increases. 450°C represents a compromise (between equilibrium and reaction rate). This compromise results in a significant amount of ammonia produced, in a very short time. More details can be found here: https://chemguide.co.uk/physical/equilibria/haber.html

Problem 55. Dialysis relies on a membrane to separate molecules in a solution. Consider a system in equilibrium consisting of a flexible dialysis vessel filled with a solution of a purified protein sitting in water. Describe
what happens to the system if:
\( a \) You increase the concentration of the protein in the vessel.
\( b \) You decrease the concentration of the protein in the vessel.
\( c \) Add a small, impure protein sample (containing salts and dyes) to the vessel.

**Solution.** Let us invoke Le Châtelier’s principle:
\( a \) The vessel expands as more water enters to balance the increased concentration.
\( b \) The vessel contracts as water leaves in response to the reduced concentration.
\( c \) The salt and dyes diffuse through the dialysis vessel, leaving a more pure sample of the protein. The vessel also likely expands some due to the increased concentration.
Suppose we have a chemical reaction which describes a cobalt-HCl system (see Fig. 4.1), where at time $t = 0$ we only have reactants and zero products. As time evolves, the reactants are depleted and the products are formed. A possible time course of this reaction is illustrated in the left hand plot (a) of Fig. 4.1, where the time courses of concentration of a reactant (A) and product (C) are shown. In the right hand plot (b) the roles are reversed: species A is produced whereas species C is depleted. This plot could correspond to the case where at time $t = 0$ we only have products present and zero reactants, in which case only the reverse reaction initially takes place. In both cases (a,b) some equilibrium concentrations are reached after sufficiently long time. At that point, species A and C do not “remember” the initial conditions anymore.

**Figure 4.1.** Chemical reactions reach equilibrium over time. The dynamics of the reaction are determined by its kinetics.
4. Chemical Reactions and Equilibrium

Suppose that we have a reversible chemical reaction:

\[ \alpha A + \beta B \rightleftharpoons \sigma S + \tau T \]

- The forward reaction rate is \( k_f[A]^\alpha[B]^\beta \),
- The backward reaction rate is \( k_r[S]^\sigma[T]^\tau \).

Here, \( k_r \) and \( k_f \) are called affinity constants.

At equilibrium the reaction rates for forward and backward reaction are equal:

\[ k_f[A]_{eq}^\alpha[B]_{eq}^\beta = k_r[S]_{eq}^\sigma[T]_{eq}^\tau, \]

where the subscript \( \text{eq} \) denotes equilibrium concentrations.

4.1.0.1. Law of Mass-Action. If we call \( K \) the equilibrium constant and set it equal to \( k_f/k_r \), we have:

\[ K \equiv \frac{k_f}{k_r} = \frac{[S]_{eq}^\sigma[T]_{eq}^\tau}{[A]_{eq}^\alpha[B]_{eq}^\beta} \]

This equation for \( K \) is called the law of mass-action. Sometimes the law of mass-action refers to the "chemical force" (affinity) between two reactants reacting together, in the above case, this force is proportional to \( k_f[A]^\alpha[B]^\beta \).

**Chemical equilibrium:** At equilibrium, the concentrations do not change in time:

\[ \frac{d[\text{reactants}]}{dt} = \frac{d[\text{products}]}{dt} = 0. \]

There are still chemical transformations taking place at equilibrium, but the net change is zero: this is where the rates of the forward and backward reaction are equal.

4.1.1. Comment on Reaction Rates. That the forward reaction rate is of the form

\[ k_f[A]^\alpha[B]^\beta \]

can easily be understood as follows. Suppose \( \alpha = \beta = 1 \) for simplicity. The reaction is of the form:

\[ A + B \rightleftharpoons \sigma S + \tau T \]

1 molecule of \( A \) reacts with 1 molecule of \( B \) to produce some quantities of \( S \) and \( T \). Thus, this event depends on the binary collision of \( A \) and \( B \). For a binary collision to happen in a given volume element, the rate of this reaction should be proportional to the product of probabilities of finding a molecule of type \( A \) in that volume times the probability of finding a molecule of type \( B \) in that same volume. (Assuming "statistical independence" of these individual
events, i.e. the presence of a molecule of type A within a volume element does not influence, and is not influenced by the presence of a molecule of type B within that same volume, which is the case for dilute concentrations), but these probabilities are proportional to the concentrations of each species. Namely, the probability of finding A in a volume element is proportional to \([A]\), the concentration of A. Same for the probability of finding a molecule of type B within that same volume, which is also proportional to \([B]\). Hence, the rate is proportional to the product \([A][B]\). When the stoichiometric coefficients \(\alpha\) and \(\beta\) are different from 1, this means we have three-body, four-body, etc. collisions. But in those instances, the same reasoning can be applied. In a three-body collision requiring the encounter of A, A and B, the rate of this event is proportional to \([A][A][B] = [A]^2[B]\). This idea is easily generalized to arbitrary \(\alpha\) and \(\beta\) to give \([A]^{\alpha}[B]^\beta\).

4.1.2. For Solution Phase Reactions.

4.1.3. Notation. \(K_C\) is the equilibrium constant in terms of concentrations. This applies to solutions. For example,

\[
aA + bB \rightleftharpoons cC + dD
\]

\[
K_C = \frac{[C]_eq^c[D]_eq^d}{[A]_eq^a[B]_eq^b}
\]

where the subscript eq denotes the concentrations at equilibrium (in steady state, where concentrations do not change with time anymore).

4.1.4. For Gas-Phase Reactions. For a gas-phase reaction, the law of mass-action is written in terms of gas pressures (all pressures are taken at equilibrium)

\[
K_P = \frac{(P_C)_eq^c(P_D)_eq^d}{(P_A)_eq^a(P_B)_eq^b}
\]

and likewise we can omit the subscript eq from this.

4.1.5. Gas-Phase Reactions. For reasons of convenience and compatibility with concepts from thermodynamics (which we will cover later in the course), all pressures are expressed in units of atm, relative to a reference pressure \(P_{\text{ref}}\) of 1 atm and we define a dimensionless quantity \(K\):

\[
K = \frac{(P_C/P_{\text{ref}})_eq^c(P_D/P_{\text{ref}})_eq^d}{(P_A/P_{\text{ref}})_eq^a(P_B/P_{\text{ref}})_eq^b}
\]

\[
K(P_{\text{ref}})^{(c+d-a-b)} = \frac{(P_C)_eq^c(P_D)_eq^d}{(P_A)_eq^a(P_B)_eq^b}
\]
where, with $P_{\text{ref}} = 1$ atm, the $K$ is identical to $K_P$ (i.e. they both have the same numerical value):

$$K = K_P = \frac{(P_C)_{\text{eq}}(P_D)_{\text{eq}}}{(P_A)_{\text{eq}}(P_B)_{\text{eq}}}.$$  

(The equality $K = K_P$ holds as long as $P_{\text{ref}} = 1$ atm and partial pressures are in units of atm.)

4.1.6. Liquid Phase Reactions. Likewise if we define a $K$ value in terms of concentrations referenced to a reference concentration $c_{\text{ref}}$

$$K = \frac{([C]_{\text{eq}}/c_{\text{ref}})^c([D]_{\text{eq}}/c_{\text{ref}})^d}{([A]_{\text{eq}}/c_{\text{ref}})^a([B]_{\text{eq}}/c_{\text{ref}})^b}$$

and express our concentrations in units of molar (M) and take $c_{\text{ref}} = 1$ M, $K$ has the same numerical value as $K_C$:

$$K = K_C = \frac{[C]_{\text{eq}}^c[D]_{\text{eq}}^d}{[A]_{\text{eq}}^a[B]_{\text{eq}}^b}.$$  

(The equality $K = K_C$ holds as long as $c_{\text{ref}} = 1$ M and concentrations are in units of M.)

4.2. Operations on Reaction

4.2.1. Inverse Reaction. Suppose we have a reaction

$$aA + bB \rightleftharpoons cC + dD, \quad K_1 = \frac{[C]_{\text{eq}}^c[D]_{\text{eq}}^d}{[A]_{\text{eq}}^a[B]_{\text{eq}}^b}$$

and another reaction, which is the inverse of the first one:

$$cC + dD \rightleftharpoons aA + bB.$$  

Then the rate constants are inverses of each other:

$$K_2 = \frac{[A]_{\text{eq}}^a[B]_{\text{eq}}^b}{[C]_{\text{eq}}^c[D]_{\text{eq}}^d} = \frac{1}{K_1}.$$  

4.2.2. Multiplication of the Stoichiometric Coefficients. Suppose that we multiply the reaction:

$$aA + bB \rightleftharpoons cC + dD, \quad K_1 = \frac{[C]_{\text{eq}}^c[D]_{\text{eq}}^d}{[A]_{\text{eq}}^a[B]_{\text{eq}}^b}$$

by a factor of 2,

$$2aA + 2bB \rightleftharpoons 2cC + 2dD.$$  

The effect is to raise the equilibrium constant to the power 2:

$$K_2 = \frac{[C]_{\text{eq}}^{2c}[D]_{\text{eq}}^{2d}}{[A]_{\text{eq}}^{2a}[B]_{\text{eq}}^{2b}} = (K_1)^2$$
4.2.3. Adding Two Reactions. If we add:

\[ 2\text{BrCl}(g) \rightleftharpoons \text{Cl}_2(g) + \text{Br}_2(g), \quad K_1 = \frac{(P_{\text{Cl}_2})_{eq} (P_{\text{Br}_2})_{eq}}{(P_{\text{BrCl}})_{eq}} \]

to the equation

\[ \text{Br}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{IBr}(g), \quad K_2 = \frac{(P_{\text{IBr}})_{eq}}{(P_{\text{Br}_2})_{eq} (P_{\text{I}_2})_{eq}} \]

we get an overall reaction:

\[ 2\text{BrCl}(g) + \text{Br}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{IBr}(g) + \text{Cl}_2(g) + \text{Br}_2(g) \]

and its corresponding rate constant:

\[ K_3 = \frac{(P_{\text{IBr}})_{eq} (P_{\text{Cl}_2})_{eq} (P_{\text{Br}_2})_{eq}}{(P_{\text{I}_2})_{eq} (P_{\text{Br}_2})_{eq} (P_{\text{BrCl}})_{eq}} = \frac{(P_{\text{IBr}})_{eq} (P_{\text{Cl}_2})_{eq}}{(P_{\text{I}_2})_{eq} (P_{\text{Br}_2})_{eq} (P_{\text{BrCl}})_{eq}} \]

Deleting \( \text{Br}_2(g) \) from both sides:

\[ 2\text{BrCl}(g) + \text{I}_2(g) \rightleftharpoons 2\text{IBr}(g) + \text{Cl}_2(g) \]

Thus, we see that the equilibrium constants multiply each other:

\[ K_3 = K_1 \cdot K_2 \]

4.2.4. Subtracting Two Reactions. By the same token subtracting two reactions leads to the division of the equilibrium constants. This follows from the fact that subtracting a reaction is equivalent to adding the reverse of the second one, which we saw leads to taking the inverse of its equilibrium constant.

4.3. Multi-Phase (Heterogeneous) Reactions

For reactions that involve some mixture of solid, liquid, and gas phases (also known as heterogeneous reactions), we have the following equilibrium constants

\[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g) \quad K = (P_{\text{H}_2\text{O}})_{eq} \]
\[ \text{I}_2(s) \rightleftharpoons \text{I}_2(aq) \quad K = [\text{I}_2]_{eq} \]
\[ \text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g) \quad K = (P_{\text{CO}_2})_{eq} \]

To write down the equilibrium constant for such a heterogeneous reaction we must follow certain conventions:

- Gases enter the equilibrium expression as partial pressures, measured in units of atm
- Dissolved species enter as concentrations, in units of mol/L (M)
- Pure solids and pure liquids do not appear in equilibrium expressions. Neither does a solvent that participates in a chemical reaction provided the solution is dilute.
4. Chemical Reactions

- Partial pressures and concentrations of products appear in the numerator, and those of reactants in the denominator. Each is raised to a power equal to its coefficient in the balanced chemical equation for the reaction.

**Example:** For

\[ 2\text{Cl}_2(g) + 2\text{HgO}(s) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HgO} \cdot \text{HgCl}_2(s) + 2\text{HOCl}(aq) \]

we write:

\[ K = \frac{[\text{HOCl}]_\text{eq}^2}{(P_{\text{Cl}_2})_\text{eq}} \]

The HgO and HgO-HgCl₂ do not appear because they are solids, and water does not appear because it is a pure liquid. Chlorine gas enters as its partial pressure in atm. The HOCl appears as its concentration in M. Both the concentration of HOCl and the partial pressure of Cl₂ are raised to the second power because their coefficients in the balanced chemical reaction are 2.

4.3.1. Dependence on Stoichiometry. It matters how you write the equation. For example,

\[ 2\text{H}_2 + \text{O}_2 \rightleftharpoons 2\text{H}_2\text{O}, \quad K_1 = \frac{[\text{H}_2\text{O}]_\text{eq}^2}{[\text{H}_2]_\text{eq}[\text{O}_2]_\text{eq}} \]

and

\[ \text{H}_2 + \frac{1}{2}\text{O}_2 \rightleftharpoons \text{H}_2\text{O}, \quad K_2 = \frac{[\text{H}_2\text{O}]_\text{eq}}{[\text{H}_2]_\text{eq}[\text{O}_2]_\text{eq}^{1/2}} \]

implies that \( K_1 \neq K_2 \). In fact, we have \( K_2 = \sqrt{K_1} \). However, both equations describe the same chemical reaction, so it would make sense that their equilibrium constants be the same constant. This ambiguity can be avoided by following the convention of using “stoichiometric coefficients that are the smallest possible whole numbers”.

The convention we learned so far is: 1) express partial pressures in the equilibrium constant in units of 1 atm, concentrations in units of 1 M, 2) omit pure solids and pure liquids from the equilibrium constant, 3) use smallest possible whole numbers as stoichiometric coefficients. This convention removes ambiguities in the many possible values and units of the equilibrium constant. In this course, we shall not necessarily follow this convention. The reason for this is because many textbooks do not; hence, you should be prepared to deal with different conventions. For example, in the first reaction, \( K_1 \) could have units of \( \text{atm}^{-1} \) or \( \text{M}^{-1} \), depending on whether the quantities are expressed in units of partial pressures or concentrations. Likewise, \( K_2 \) could have units of \( 1/\sqrt{\text{atm}} \) or \( 1/\sqrt{\text{M}} \).

4.3.2. Detailed Balance. If two reactions are in equilibrium:

\[ aA + bB \rightleftharpoons eE \quad \text{and} \quad eE \rightleftharpoons cC + dD \]
Then the whole thing is in equilibrium:

\[ aA + bB \rightleftharpoons cC + dD. \]

Thus, if each step is in equilibrium (balanced) then the whole thing is in equilibrium.

This is a consequence of the addition of chemical reactions. The equilibrium constant of the first two equations are:

\[ K_1 = \frac{[E]^e}{[A]^a[B]^b} \quad \text{and} \quad K_2 = \frac{[C]^c[D]^d}{[E]^e} \]

The sum of the two reactions is:

\[ aA + bB + eE \rightleftharpoons eE + cC + dD \quad (\ast) \]

Cancelling \( eE \) on both sides leads to the desired result:

\[ aA + bB \rightleftharpoons cC + dD \]

whose equilibrium constant is

\[ K_3 = \frac{[C]^c[D]^d}{[A]^a[B]^b}. \]

We note that this is also equal to the equilibrium constant of (\ast):


4.4. Direction of a Reaction

Suppose we have a reaction

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g), \quad K_C = \frac{[\text{HI}]_{eq}^2}{[\text{H}_2]_{eq}[\text{I}_2]_{eq}} \]

for which \( K_C = 46 \) at 783 K. Let’s say we mix \( \text{H}_2, \text{I}_2 \) and \( \text{HI} \) at each 0.0020 mol/L at 783 K. (We’ll use concentrations instead of pressures.) There is a way to predict whether or not HI will have a tendency to form.

**Strategy:** As you will recall \( K_C \) is products over reactants. If the quantity \( [\text{HI}]^2 / ([\text{H}_2][\text{I}_2]) \) at non-equilibrium is greater than \( K_C \) this means there is too much product, so reactants will form. Thus, our strategy is to calculate the reaction quotient

\[ Q_C = \frac{[\text{HI}]_{eq}^2}{[\text{H}_2]_{eq}[\text{I}_2]_{eq}} = \frac{(0.0020)^2}{0.0020 \times 0.0020} = 1.0 \]

and compare to \( K_C \). Here, we find \( Q_C < K_C \) and conclude that the reaction will tend to form more product and consume reactants.

\( Q_C \) is called the quotient of the reaction. It is calculated under the conditions given by the problem (not necessarily at equilibrium). Comparison of \( Q_C \)
to $K_C$ tells us which direction the reaction will proceed. This concept is illustrated in Fig. 4.2.

**Figure 4.2.** The value of $Q$ relative to $K_C$ (equilibrium constant) tells us if the reaction needs to favor the reactants or products. Red: reactants; blue: products.

**Exercise:** A mixture of $H_2$, $N_2$ and $NH_3$ with molar concentrations $3.0 \times 10^{-3}$ mol/L, $1.0 \times 10^{-3}$ mol/L and $2.0 \times 10^{-3}$ mol/L, respectively, was prepared and heated to 500 K, at which temperature $K_C = 61$ for the reaction

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Decide whether ammonia tends to form or decompose. [Answer: $Q_C = (2.0 \times 10^{-3})^2 / (3 \times 10^{-3})^3 (1.0 \times 10^{-3}) = 1.48 \times 10^5$, so it tends to decompose.]

Note: certain textbooks would express $K_C$ and $Q_C$ in units of $1/M^2$.

### 4.4.1. Solving for Equilibrium Concentrations.

Suppose that in equilibrium between $HCl$, $Cl_2$ and $H_2$,

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g), \quad K_C = \frac{[HCl]_{eq}^2}{[H_2]_{eq}[Cl_2]_{eq}}$$

the concentrations are:

$[H_2]_{eq} = 1.0 \times 10^{-17}$ mol $\cdot$ L$^{-1}$ and $[Cl_2]_{eq} = 2.0 \times 10^{-16}$ mol $\cdot$ L$^{-1}$.

What is the equilibrium molar concentration of $HCl$ given $K_C = 4.0 \times 10^{31}$ for the reaction?

**Strategy:** At equilibrium, the molar concentrations of the reactants and products satisfy the expression for $K_C$. Rearrange the expression to give the unknown concentration, and substitute the data.
From the expression $K_C = \frac{[\text{HCl}]^2_{eq}}{[\text{H}_2]_{eq}[\text{Cl}_2]_{eq}}$ we solve for $[\text{HCl}]_{eq}$:
\[
[\text{HCl}]_{eq} = \sqrt{K_C[\text{H}_2]_{eq}[\text{Cl}_2]_{eq}},
\]
and substitute the data given
\[
[\text{HCl}]_{eq} = \sqrt{(4.0 \times 10^{31})(1.0 \times 10^{-17})(2.0 \times 10^{-16})} = 0.28.
\]
That is, the molar concentration of HCl at equilibrium is 0.28 mol/L. This result means that, at equilibrium, the amount of product in the system is overwhelming compared to the amounts of reactants.

### 4.4.2. Effecting a Change From Initial to Equilibrium Pressures.

The reversible reaction between dinitrogen tetroxide and nitrogen dioxide is written as:
\[
\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g).
\]
1 L $\text{N}_2\text{O}_4$ (at 25°C, 1 atm) is allowed to come to equilibrium. $K=0.14$ atm at 25°C. What are $P_{\text{N}_2\text{O}_4}$, $P_{\text{NO}_2}$ and $P_{\text{tot}}$ at equilibrium?

**Strategy:** Chemical reaction gives relationship between the partial pressures. We setup $x$ as the progress variable.

\[
P_{\text{N}_2\text{O}_4} = 1 \text{ atm} - x \quad \text{and} \quad P_{\text{NO}_2} = 2x.
\]

From our knowledge of the equilibrium constant (its value is given to us)
\[
K = 0.14 \text{ atm} = \frac{(P_{\text{NO}_2})_{eq}}{(P_{\text{N}_2\text{O}_4})_{eq}}
\]
\[
0.14 \text{ atm} = \frac{(2x)^2}{1 \text{ atm} - x}
\]
\[
0.14 \text{ atm}^2 - 0.14x \text{ atm} = 4x^2
\]
\[
4x^2 + 0.14x \text{ atm} - 0.14 \text{ atm}^2 = 0.
\]
Thus we need to solve a quadratic equation of the form $Ax^2 + Bx + C = 0$ whose roots are:
\[
x^\pm = \frac{-B \pm \sqrt{B^2 - 4AC}}{2A}.
\]
We get two roots, only one of which makes physical sense (the other root, $-0.21$ is rejected because $x$ cannot be negative if product must be formed):
\[
x = 0.17 \text{ atm}.
\]
From this, we get:
\[
(P_{\text{NO}_2})_{eq} = 2x = 0.34 \text{ atm}, \quad (P_{\text{N}_2\text{O}_4})_{eq} = 1 - 0.17 = 0.83 \text{ atm},
\]
\[
(P_{\text{tot}})_{eq} = (0.34 + 0.83) \text{ atm} = 1.17 \text{ atm}.
\]
4.5. Le Châtelier’s Principle: Response to (small) Changes (from Equilibrium)

Henri Louis Le Châtelier (1850-1936) stated a general principle that lets us predict how the composition of a reaction mixture at equilibrium tends to change when the conditions are perturbed:

**Le Châtelier’s principle:** When a stress is applied to a system in dynamic equilibrium, the equilibrium tends to adjust to minimize the effect of the stress.

Other statements you may encounter include: 1) When any system at equilibrium is subjected to change in concentration, temperature, volume, or pressure, then the system readjusts itself to counteract (partially) the effect of the applied change and a new equilibrium is established. 2) Any change in status quo prompts an opposing reaction in the responding system.

The main points to remember are:

- System initially at equilibrium (“dynamic equilibrium” = steady state)
- It is perturbed *slightly* from equilibrium (big changes not allowed)
- DIRECTION: Response is such that the change is opposed

4.5.1. Adding and Removing Reagents. Suppose that the following reaction has reached equilibrium:

\[
N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)
\]

So at time \(t = 0\) we are at equilibrium. Suppose that we pump in more hydrogen gas. According to Le Châtelier’s principle, the reaction will tend to minimize the increase in the number of hydrogen molecules. Hydrogen will tend to react with nitrogen. As a result, ammonia will be formed. If instead of hydrogen, we were to add ammonia, then the reaction would tend to form reactants.

The explanation is found in the previous discussion of \(Q\) and \(K\). When reactants are added, \(Q\)

\[
Q = \frac{[NH_3]^2}{[H_2]^3[N_2]}
\]

falls below \(K\) momentarily, because the reactant concentrations appear in the denominator of \(Q\). As we have seen, when \(Q < K\), the reaction mixture responds by forming products. Likewise when products are added \(Q\) rises above \(K\). Because \(Q > K\) the reaction mixture responds by forming reactants at the expense of the products until \(Q = K\) again.

4.5.2. Compressing a Reaction Mixture. A gas-phase equilibrium responds to compression (a reduction in volume) of the reaction vessel. According to Le Châtelier’s principle, the composition will tend to change in a
way that minimizes the increase in pressure resulting from the compression. For instance, the formation of \( \text{NH}_3 \) from \( \text{N}_2 \) and \( \text{H}_2 \) decreases the number of gas-phase molecules in the container because 4 mol of reactant molecules produce 2 mol of product molecules. The forward reaction therefore decreases the pressure the mixture exerts. When the mixture is compressed, the equilibrium composition will tend to shift in favor of the products, for that minimizes the increase in pressure. The opposite response, a tendency for products to decompose, occurs in an expansion. In order to increase the yield of ammonia, the synthesis must be carried out with highly compressed gases. The industrial process uses pressures of 250 atm and more.

Predicting the Effect of Compression: Predict the effect of compression on the equilibrium composition of the reaction mixture in which the equilibria

\[
\begin{align*}
(a) & \quad \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \\
(b) & \quad \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)
\end{align*}
\]

have been established.

Solution: (a) In the reverse reaction, two \( \text{NO}_2 \) molecules combine to form one \( \text{N}_2\text{O}_4 \) molecule. Hence, compression favors the formation of \( \text{N}_2\text{O}_4 \). (b) Because neither direction corresponds to the reduction of gas-phase molecules, compressing the mixture should have little effect on the composition of the equilibrium mixture.

Comment: We can deduce the Le Châtelier principle by comparing the \( Q \) value to \( K \). For the first reaction we have \( Q = [\text{NO}_2]^2/[\text{N}_2\text{O}_4] \). Upon compression all concentrations increase by the same factor. However, the numerator will grow faster than the denominator because it is squared. Thus, \( Q \) will increase. We end up with \( Q > K \), which implies the reaction will want to go in the \( \leftarrow \) direction to form more reactants in order to lower the number of moles to counter the increase in pressure.

4.5.3. Temperature. The effect of temperature is most easily understood by writing the heat of the reaction as a reactant or product according to whether the reaction is endothermic or exothermic. The reaction of ammonia synthesis:

\[
\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)
\]

is exothermic with \( \Delta H = -92 \text{ kJ/mol} \) at some temperature. This heat is produced, meaning that we can write it on the product side:

\[
\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + \text{heat}.
\]

Thus, with the equation written as such we easily see that upon increasing the temperature the reaction will want to shift toward the reactants and
consume heat in the process to counter-act the increase in heat supplied to
the reaction (due to the elevated temperature).
Chapter 5

Acids and Bases

5.1. Definitions of Acids: Arrhenius, Lewis, Brønsted-Lowry

5.1.1. Arrhenius Acid. The definition of acid and base, according to Swedish chemist Svante Arrhenius (1859-1927), classifies acids as substances that increase the concentration of $[\text{H}_3\text{O}^+]$ ions, e.g.

$$\text{HCl}(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^-(aq),$$

and bases as substances that increase the concentration of $[\text{OH}^-]$ ions, e.g.

$$\text{NaOH}(s) \rightleftharpoons \text{Na}^+(aq) + \text{OH}^-(aq).$$

If we combine these two reactions, we get an acid-base neutralization reaction

$$\text{HCl} + \text{NaOH} \rightleftharpoons \text{H}_2\text{O} + \text{NaCl},$$

which is seen to be of the form:

$$\text{acid} + \text{base} \rightleftharpoons \text{water} + \text{salt}.$$  

This creates an electrolyte solution, since

$$\text{NaCl} \rightleftharpoons \text{Na}^+(aq) + \text{Cl}^-(aq).$$

The neutralization reaction involves the combination of hydrogen ions and hydroxide ions to form water: $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$.

This strong acid/base reaction produces a strong electrolyte solution in the sense that its dissociation constant is very large: $K \gg 1$, i.e. the equilibrium constant heavily favors the formation of products.

One problem with this Arrhenius definition is it makes only reference to water as the solvent. What if the solvent is different from water? Another problem arises with the weak base ammonia, $\text{NH}_3$, because Arrhenius theory suggests that all bases contain $\text{OH}^-$. Where is $\text{OH}^-$ in $\text{NH}_3$? Moreover,
the Arrhenius theory does not account for the behavior of amphiprotic substances.\textsuperscript{1}

5.1.2. Brønsted-Lowry Definition. A slightly more general definition of an \textbf{acid} is a proton donor and a \textbf{base} is a proton acceptor. To describe the behavior of ammonia:

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-. 
\]

In this reaction, H\textsubscript{2}O acts as an \textit{acid}, giving up a proton H\textsuperscript{+}, which is taken up by NH\textsubscript{3}, the \textit{base}. As a result of this, the ions OH\textsuperscript{−} and NH\textsubscript{4}\textsuperscript{+} are formed. The pair NH\textsubscript{3}/NH\textsubscript{4}\textsuperscript{+} is called a \textit{conjugate pair}. NH\textsubscript{4}\textsuperscript{+} is the conjugate acid of NH\textsubscript{3}, the conjugate base. Similarly, H\textsubscript{2}O/OH\textsuperscript{−} is a conjugate pair where H\textsubscript{2}O is the acid and OH\textsuperscript{−} is the conjugate base. Pairs (1) and (2) are explicitly identified below:

\[
\text{NH}_3 (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{NH}_4^+ (aq) + \text{OH}^- (aq).
\]

The equilibrium constant for this reaction is:

\[
K_b = \frac{[\text{NH}_4^+]_{eq}[\text{OH}^-]_{eq}}{[\text{NH}_3]_{eq}} = 1.8 \times 10^{-5}.
\]

Some authors prefer to call $K_b$ the \textit{ionization constant} rather than the equilibrium constant, where the latter is the ratio $\frac{[\text{NH}_4^+]_{eq}[\text{OH}^-]_{eq}}{[\text{NH}_3]_{eq}[\text{H}_2\text{O}]_{eq}}$. But according to our rules for writing down equilibrium constants [H\textsubscript{2}O\textsubscript{eq}] does not appear in the expression since it is a pure liquid. $K_b$ is normally called the \textit{base ionization constant}. See Fig. 5.1.

5.1.2.1. \textit{Example: ionization of acetic acid.}

\[
\text{HC}_2\text{H}_3\text{O}_2 (\text{acid}(1)) + \text{H}_2\text{O (base}(2)) \rightleftharpoons \text{C}_2\text{H}_5\text{O}_2^- (\text{acid}(1)) + \text{H}_3\text{O}^+ (\text{base}(2)).
\]

The acetate ion C\textsubscript{2}H\textsubscript{3}O\textsubscript{2}\textsuperscript{−} is the conjugate base of the acid HC\textsubscript{2}H\textsubscript{3}O\textsubscript{2}. Here, H\textsubscript{2}O acts as a base. Its conjugate acid is the hydronium ion H\textsubscript{3}O\textsuperscript{+}.

The ionization of acetic acid can be described in the following way:

\[
K_a = \frac{[\text{C}_2\text{H}_5\text{O}_2^-]_{eq}[\text{H}_3\text{O}^+]_{eq}}{[\text{HC}_2\text{H}_3\text{O}_2]_{eq}} = 1.8 \times 10^{-5}.
\]

\textsuperscript{1}An amphiprotic substance is one that can act as an acid or a base. The most common example is water, which acts as a base in the presence of an acid:

H\textsubscript{2}O + HA $\rightleftharpoons$ H\textsubscript{3}O\textsuperscript{+} + A\textsuperscript{−}.

It also acts as an acid in the presence of a base:

H\textsubscript{2}O + B $\rightleftharpoons$ OH\textsuperscript{−} + B\textsuperscript{+}. 
5.1. Definitions of Acids: Arrhenius, Lewis, Brønsted-Lowry

Both forward and backward “proton transfer” reactions proceed because \( \text{NH}_4 \) is a stronger acid than \( \text{H}_2\text{O} \) and \( \text{OH}^- \) is a stronger base than \( \text{NH}_3 \). \( \text{NH}_3 \) is only slightly ionized.

\( K_a \) is called the **acid ionization constant**. See Fig. 5.2.

5.1.2.2. **Example: ionization of HCl.** The reaction is shown in Fig. 5.3. Here \( K_a \) is so large (about \( 10^6 \)) so the ionization equation is written with a single arrow

\[
\text{HCl} + \text{H}_2\text{O} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+.
\]

5.1.2.3. **Identifying Brønsted-Lowry acids and bases and their conjugates.**

Chlorous acid reacts with water to give chlorine dioxide and hydronium ion:

(1) \( \text{HClO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{ClO}_2^- + \text{H}_3\text{O}^+ \).
HClO₂ gives up a proton, H⁺, to become ClO₂⁻, therefore, HClO₂ is an acid and ClO₂⁻ is its conjugate base. H₂O takes the proton from HClO₂ and becomes H₃O⁺. Thus H₂O is a base and H₃O⁺ is its conjugate acid.

(2) OCl⁻ + H₂O ⇌ HOCl + OH⁻
OCl⁻ is a base and gains a proton from water. OH⁻ produced in this reaction is the conjugate base of H₂O.

Consider the reactions below; can you identify the conjugate acids and bases?

\[
\begin{align*}
\text{NH}_3 + \text{H}_2\text{PO}_4^- & \equiv \text{NH}_4^+ + \text{HPO}_4^{2-} \\
\text{HCl} + \text{H}_2\text{PO}_4^- & \equiv \text{Cl}^- + \text{H}_3\text{PO}_4 \\
\text{HF} + \text{H}_2\text{O} & \equiv \text{F}^- + \text{H}_3\text{O}^+ \\
\text{HSO}_4^- + \text{NH}_3 & \equiv \text{SO}_4^{2-} + \text{NH}_4^+ \\
\text{C}_2\text{H}_3\text{O}_2^- + \text{HCl} & \equiv \text{HC}_2\text{H}_3\text{O}_2 + \text{Cl}^-
\end{align*}
\]

5.2. Acid-Base Titrations

According to the Arrhenius definition of acid and base, an acid is a substance that donates a proton:

\[
\text{HA} \rightarrow \text{H}^+ + \text{A}^-.
\]

whereas a base is a proton acceptor:

\[
\text{BOH} \rightarrow \text{B}^+ + \text{OH}^-.
\]

OH⁻ is a base because it accepts a proton:

\[
\text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O}.
\]

An acid-base titration reaction is one where acid is added to base (or vice-versa) to neutralize the original species and form H₂O. An example of an acid-base titration is the reaction of nitric acid with ammonium hydroxide:

\[
\text{HNO}_3 + \text{NH}_4\text{OH} \rightarrow \text{H}^+ + \text{NO}_3^- + \text{NH}_4^+ + \text{OH}^- \rightarrow \text{NO}_3^- + \text{NH}_4^+ + \text{H}_2\text{O}.
\]

In acid-base reactions we are often asked to neutralize an acid using a base or vice-versa. This is done by matching the number of moles.

5.2.0.1. Example Calculation (Acid-Base Titration). Suppose that we are asked to titrate a solution which contains 10 mL of 3.0 M sulfuric acid, H₂SO₄(aq), using a sodium hydroxide solution of unknown concentration. We find that 20 mL NaOH was required. What is the molarity c₇NaOH of the NaOH solution?

Solution: The reaction is:

\[
\text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)
\]
5.3. Redox Reactions

The units of molarity $c_{\text{NaOH}}$ are mol/L. The concentration of NaOH was (we need to obtain $n_{\text{NaOH}}$):

$$c_{\text{NaOH}} = \frac{n_{\text{NaOH}}}{20 \text{ mL}}$$

The number of moles of sulfuric acid is:

$$n_{\text{H}_2\text{SO}_4} = c_{\text{H}_2\text{SO}_4} \cdot 10 \text{ mL} = (3.0 \text{ M})(10 \text{ mL}) = (3.0 \text{ M})(0.010 \text{ L}) = 0.03 \text{ mol}.$$ 

The minimum number of moles needed, $n_{\text{NaOH}}$, is (according to the stoichiometry of the equation, we must have 2 moles of NaOH for each mole of sulfuric acid)

$$n_{\text{NaOH}} = 2n_{\text{H}_2\text{SO}_4} = 0.06 \text{ mol}.$$ 

So the concentration of NaOH is

$$c_{\text{NaOH}} = \frac{n_{\text{NaOH}}}{20 \text{ mL}} = \frac{0.06 \text{ mol}}{0.020 \text{ L}} = 3.0 \text{ M}.$$ 

Another example (Oxtoby [1], Example 11.6, p. 484) is the titration of acetic acid using NaOH:

$$\text{CH}_3\text{COOH}(aq) + \text{Na}^+(aq) + \text{OH}^-(aq) \rightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l) + \text{Na}^+(aq).$$

5.3. Redox Reactions

An **oxidized** substance is one that gives up its electrons. Its oxidation number increases, whereas a reduced substance is one which accepts electrons. Example:

$$2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$$

where magnesium is oxidized: it gives up two electrons. Its oxidation number increases from 0 (as in elemental Mg) to +2 (in MgO). Oxygen accepts these electrons and its oxidation number decreases from 0 to -2. The transfer of electrons can be emphasized using arrows (Fig. 5.4).

![Figure 5.4. Redox reaction.](image-url)
Another example is the reaction of iron(III) chloride with hydrogen iodide (hydroiodic acid),

$$2\text{FeCl}_3 + 2\text{HI} \rightarrow 2\text{FeCl}_2 + 2\text{HCl} + \text{I}_2$$

Which involves the following intermediate steps:

$$2\text{FeCl}_3 + 2\text{HI} \rightarrow 2\text{Fe}^{3+} + 6\text{Cl}^- + 2\text{H}^+ + 2\text{I}^- \rightarrow 2\text{Fe}^{2+} + 6\text{Cl}^- + 2\text{H}^+ + \text{I}_2$$

where the electron was transferred from I to Fe.

### 5.4. Lewis Acids & Bases

A Lewis acid (Gilbert N. Lewis, 1985-1946) is a chemical species that contains an empty orbital which is capable of accepting an electron pair from a Lewis base to form a Lewis adduct. A Lewis base, then, is any species that has a filled orbital containing an electron pair which is not involved in bonding but may form a dative bond with a Lewis acid to form a Lewis adduct. For example, NH$_3$ is a Lewis base, because it can donate its lone pair of electrons. Trimethylborane (Me$_3$B) is a Lewis acid as it is capable of accepting a lone pair. In a Lewis adduct, the Lewis acid and base share an electron pair furnished by the Lewis base, forming a dative bond. In the context of a specific chemical reaction between NH$_3$ and Me$_3$B, the lone pair from NH$_3$ will form a dative bond with the empty orbital of Me$_3$B to form an adduct NH$_3$·BMe$_3$ (the dot here simply denotes a bond, not an electron). Another example is shown in Fig. 5.5. See also Problem 58.

![Figure 5.5](image-url)

**Figure 5.5.** Major structural changes accompany binding of the Lewis base to the coordinatively unsaturated, planar Lewis acid BF$_3$.

### 5.4.1. For Acids & Bases, What is the Direction of the Reaction?

In an acid-base reaction, the favored direction of the reaction is from the stronger to the weaker member of a conjugate acid-base pair. For example:

$$\text{HCl} \overset{\text{acid(1) strong}}{\overset{\overset{\text{acid(2)} \text{ weak}}{\text{base(1) weak}}} \rightarrow \text{Cl}^- + \text{H}_2\text{O} \overset{\text{base(2) strong}}{\text{acid(2) weak}}}$$

---

\(^2\)A dative bond is also known as a coordinate covalent bond.
Likewise, the following reaction proceeds almost exclusively in reverse:

\[
\begin{align*}
\text{H}_2\text{O} + \text{I}^- & \underset{\text{weak base (1)}}{\overset{\text{weak acid (2)}}{\rightleftharpoons}} \text{OH}^- + \text{HI}^+. \\
\end{align*}
\]

In general, we need to use a table of acid and base strengths, such as the one in Fig. 5.6, which also incidentally shows that the stronger an acid, the weaker its conjugate base.

![Figure 5.6. Strengths of acids and bases.](image)

**5.4.2. Water and the pH Scale.** Pure water is nonconducting. This is because the concentration of ions is far too low. The reaction of self-ionization (auto-ionization, auto-protolytic)

\[
\text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-.
\]

The equilibrium constant for pure water is therefore,

\[
K_C = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}[\text{OH}^-]_{\text{eq}}}{[\text{H}_2\text{O}]_{\text{eq}}^2}.
\]

However, it is customary to absorb the factor $[\text{H}_2\text{O}]_{\text{eq}}^2$ into the definition of $K$ and call this the ionic product or water-dissociation equilibrium constant,

\[
K_w = [\text{H}_3\text{O}^+]_{\text{eq}}[\text{OH}^-]_{\text{eq}}.
\]

This is a consequence of our rules for writing down $K$ (pure liquids are not included in the expression). The numerical value of the ionic product is

\[
K_w = 1.0 \times 10^{-14}. \quad \text{(pure water, 25\textdegree C)}
\]

Thus, the equilibrium is heavily shifted to the left, i.e., there are very few ionized species. The stoichiometry of the equation indicates that $[\text{H}_3\text{O}^+]_{\text{eq}}$
and \([\text{OH}^-]_{\text{eq}}\) are equal in pure water. Therefore,
\[
[\text{H}_3\text{O}^+]_{\text{eq}} = [\text{OH}^-]_{\text{eq}} = 10^{-7} \text{ M}, \quad \text{(pure water, } 25^\circ\text{C)}
\]
which is exceedingly low: this means that only about 1 molecule in 200 million is ionized. This concentration is far too low for significant electrical conductivity. The pH of a solution is:

**pH stands for “potential of hydrogen ion” (Sørensen, 1909):**
The pH is defined as the negative logarithm of \([\text{H}^+]\):
\[
\text{pH} = -\log_{10} [\text{H}^+]_{\text{eq}} = -\log_{10} [\text{H}_3\text{O}^+]_{\text{eq}}.
\]
The concentration of hydronium ions is a result of the ionization of water plus any other species present that may contribute or remove \(\text{H}^+\) ions (e.g. acids or bases).

In a solution of 0.0025 M HCl,
\[
[\text{H}_3\text{O}^+]_{\text{eq}} = 2.5 \times 10^{-3} \text{ M} \quad \text{and} \quad \text{pH} = -\log_{10}(2.5 \times 10^{-3}) = 2.60.
\]
Conversely in a solution with pH=4.50 we have:
\[
[\text{H}_3\text{O}^+]_{\text{eq}} = 10^{-4.50} = 3.2 \times 10^{-5} \text{ M}.
\]

**pOH:** The pOH is defined as the negative of the logarithm of \([\text{OH}^-]_{\text{eq}}\):
\[
\text{pOH} = -\log_{10} [\text{OH}^-]_{\text{eq}}.
\]
Let us take the negative log of \(K_w = [\text{H}_3\text{O}^+]_{\text{eq}}[\text{OH}^-]_{\text{eq}} = 1.0 \times 10^{-14}\):
\[
\text{p}K_w = -\log_{10} [\text{H}_3\text{O}^+]_{\text{eq}}[\text{OH}^-]_{\text{eq}} = -\log_{10} [\text{H}_3\text{O}^+]_{\text{eq}} - \log_{10} [\text{OH}^-]_{\text{eq}} = \text{pH} + \text{pOH}.
\]
Thus,
\[
\text{p}K_w = \text{pH} + \text{pOH} = 14.00. \quad (25^\circ\text{C})
\]

An aqueous solution with \([\text{H}_3\text{O}^+]_{\text{eq}} = [\text{OH}^-]_{\text{eq}}\) is said to be neutral. In pure water at 25°C, \([\text{H}_3\text{O}^+]_{\text{eq}} = [\text{OH}^-]_{\text{eq}} = 1.0 \times 10^{-7} \text{ M} \quad \text{and} \quad \text{pH}=7.00. \quad \text{If the pH is less than 7.00 the solution is } \text{acidic} \quad \text{and if the pH is greater than 7.00 it is } \text{basic} \quad \text{or alkaline.}

The following remark reinforces the notion that an equilibrium constant is indeed a constant. However, note that it is temperature-dependent.
Because $K_w$ is an equilibrium constant, the product of the molarities of H$_3$O$^+$ and OH$^-$ ions is always equal to $K_w$. We cannot increase the molarities of both hydronium and hydroxide ions. We can increase the concentration of H$_3$O$^+$ ions by adding acid, in which case the concentration of OH$^-$ ions must decrease to preserve the value of $K_w$. Alternatively, we can increase the concentration of OH$^-$ ions by adding base, but then the concentration of H$_3$O$^+$ ions must decrease. The auto-protolysis equilibrium links the molarities of H$_3$O$^+$ and OH$^-$ ions so that when one goes up, the other goes down.

5.4.3. Strong Acid (Strong Base). A strong acid is one for which the dissociation constant is very large (i.e. $K \gg 1$). At equilibrium it is completely dissociated (deprotonated). Similarly for a base, a strong base is one which is entirely protonated at equilibrium. Let’s look at an example:

$$\text{H}_2\text{O}(l) + \text{HCl}(aq) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq).$$

In this case, the pH of this solution will be determined solely by the amount of acid dissolved. This applies to any strong acid. In other words, a 0.10 M solution of any strong acid that donates one proton per molecule is simply 0.10 M= $10^{-1}$ M, whose pH equals 1. The pOH is found from the equation for $K_w$:

$$[\text{OH}^-]_{eq} = \frac{K_w}{[\text{H}_3\text{O}^+]_{eq}} = \frac{1.0 \times 10^{-14}}{0.10} = 1.0 \times 10^{-13} \text{ M},$$

which gives pOH=13. Likewise a strong base means we get complete dissociation to yield OH$^-$ ions:

$$\text{NaOH}(s) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq).$$

If we have a 0.10 M NaOH solution, we have

$$[\text{OH}^-]_{eq} = 0.10 \text{ M} \quad [\text{H}_3\text{O}^+]_{eq} = \frac{1.0 \times 10^{-14}}{0.10} = 1.0 \times 10^{-13} \text{ M}.$$ 

Remark: We have neglected the autoionization of water in both cases on the basis that the contribution of H$_3$O$^+$ or OH$^-$ form this process is very small ($\sim 10^{-7}$ M). When very small amounts of a strong acid or base is added to pure water (for example, $\sim 10^{-7}$ M), we have to include the autoionization of water to describe the concentration of hydronium and hydroxide ions accurately.
5. The pH Scale

In experiments the concentrations of hydronium ions can range from $10 \text{ M}$ to $10^{-15} \text{ M}$ and even smaller. Thus, the range of concentrations extends over an enormous range. This is why we define the pH scale: the scale is compressed to a much smaller range as a result of taking the logarithm. The logarithm is indicative of the exponent of the concentration.

Can we have negative pH values?: It is certainly possible to have a negative pH value. For example, if the concentration of hydronium ions is $10 \text{ M} = 10^1 \text{ M}$, this corresponds to the negative value $\text{pH} = -1$.

5.5.1. Example: Calculating pH. What is the pH of (a) human blood, in which the molarity of $\text{H}_3\text{O}^+$ ions is $4.0 \times 10^{-8} \text{ mol/L}$; (b) 0.020 M HCl($aq$); (c) 0.040 M KOH($aq$)?

Strategy: Apply the definition of pH. For strong acids, the molarity of $\text{H}_3\text{O}^+$ is equal to the molarity of the acid. For strong bases we must first find the molarity of $\text{OH}^-$ then convert to molarity of $\text{H}_3\text{O}^+$ using the equation for $K_w$.

Solution: (a) For a solution in which the molarity of $\text{H}_3\text{O}^+$ ions is $4.0 \times 10^{-8} \text{ mol/L}$, we write $[\text{H}_3\text{O}^+]_{eq} = 4.0 \times 10^{-8}$ and obtain

$$\text{pH} = -\log_{10}(4.0 \times 10^{-8}) = 7.40$$

(b) Because HCl is a strong acid, the molarity of $\text{H}_3\text{O}^+$ is 0.020 mol/L. Hence,

$$\text{pH} = -\log_{10}0.020 = 1.70$$

(c) Each formula unit of KOH (a strong base) provides one $\text{OH}^-$ ion; therefore the molarity of $\text{OH}^-$ is 0.040 mol/L and

$$[\text{H}_3\text{O}^+]_{eq} = \frac{K_w}{[\text{OH}^-]_{eq}} = \frac{1.0 \times 10^{-14}}{0.040} = 2.5 \times 10^{-13}$$

Hence $\text{pH} = -\log_{10}(2.5 \times 10^{-13}) = 12.60$.

5.5.2. Example Problem: Strong Base. What are the molarities of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ in a 0.0030 M Ba(OH)$_2$ solution at 25°C? [Ba(OH)$_2$ is a hydroxide of an alkaline earth metal and is therefore a strong base.]

Strategy: Strong bases are almost entirely present as $\text{OH}^-$ in water. How many $\text{OH}^-$ ions per formula unit? Calculate the molarity of these ions in solution. Find molarity of $\text{H}_3\text{O}^+$ ions using the equation for $K_w$.

Solution: The equation

$$\text{Ba(OH)}_2(s) \rightarrow \text{Ba}^{2+}(aq) + 2\text{OH}^-(aq)$$
5.5. The pH Scale

It tells us that each \( \text{Ba(OH)}_2 \) gives \( 2\text{OH}^- \). Because the molarity of \( \text{Ba(OH)}_2(aq) \) is 0.0030 mol/L, it follows that the molarity of \( \text{OH}^- \) is twice that value, or 0.0060 mol/L. Then, for the molarity of \( \text{H}_3\text{O}^+ \) ions, we write

\[
[H_3O^+]_{eq} = \frac{K_w}{[OH^-]_{eq}} = \frac{1.0 \times 10^{-14}}{0.0060} = 1.7 \times 10^{-12}
\]

That is, the molarity of \( \text{H}_3\text{O}^+ \) ions in the solution is only \( 1.7 \times 10^{-12} \) mol/L.

5.5.3. \( pK_a \) and \( pK_b \). According to the Brønsted-Lowry theory of an acid (HA), its ionization in aqueous solution is written as

\[
\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)
\]

where \( \text{A}^- \) is the conjugate base of HA. The equilibrium expression for this chemical reaction is

\[
K_a = \frac{[H_3O^+]_{eq}[A^-]_{eq}}{[HA]_{eq}}
\]

where \( K_a \) is the acid ionization constant. The negative of its log is the \( pK_a \) value of this acid:

\[
pK_a = -\log_{10} K_a.
\]

This is a quantitative measure of the strength of the acid in a particular solvent (here, water). Weak acids do not react completely and the concentrations of product are low at equilibrium and so \( K_a \) is small. By contrast, a strong acid reacts nearly completely and \( K_a \) is large (\( K_a \gg 1 \)).

A specific example is:

\[
\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-
\]

\[
K_a = \frac{[H_3O^+]_{eq}[F^-]_{eq}}{[HF]_{eq}} = 6.7 \times 10^{-4} \text{ M}
\]

where HF is a weak acid and \( \text{F}^- \) is a weak base.

In general, the concept of \( pK_a \) applies to two types of reactions:

\[
\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-,
\]

\[
\text{HA}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}.
\]

There is a similar concept for bases. For example, consider the reaction involving the weak base \( \text{F}^- \):

\[
\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-
\]

\[
K_b = \frac{[\text{HF}]_{eq}[\text{OH}^-]_{eq}}{[\text{F}^-]_{eq}} = 1.5 \times 10^{-11} \text{ M}.
\]
Relationship between $K_a$ and $K_b$:

For the two reactions describing the equilibria in water of each conjugate acid-base pair considered above:

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-,$$  
$$K_a = \frac{[H_3O^+]_{eq}[F^-]_{eq}}{[HF]_{eq}}$$

and

$$F^- + H_2O \rightleftharpoons HF + OH^-,$$  
$$K_b = \frac{[HF]_{eq}[OH^-]_{eq}}{[F^-]_{eq}}$$

We have the product:

$$K_a \cdot K_b = \frac{[H_3O^+]_{eq}[F^-]_{eq}}{[HF]_{eq}} \cdot \frac{[HF]_{eq}[OH^-]_{eq}}{[F^-]_{eq}} = [H_3O^+]_{eq}[OH^-]_{eq} = K_w.$$  

This relationship $K_w = K_a \cdot K_b$ is general and indicates an inverse relationship between the strength of a base and its conjugate acid. Taking the logs of both sides of the equation we get:

$$pK_w = pK_a + pK_b.$$  

Note: this relationship only applies to conjugate pairs, such as HF and F\(^-\). It does not relate $pK_a$ and $pK_b$ of, for example, HI and F\(^-\), respectively, since they are not a conjugate pair.

This trend is illustrated in Fig. 5.7.

\[ \text{Acid strength vs conjugate base strength. } pK_a \text{ is obtained from } pK_b \text{ from the relationship } pK_w = pK_a + pK_b. \]

We see from Fig. 5.7 that when $H_2O$ is viewed as the acid, $pK_a$ is 14. What is the significance of this? $pK_a$ of 14 implies that $K_a = 10^{-14}$. This value of $K_a$ is the equilibrium constant for the auto-ionization reaction of water:

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-.$$  

Here,

$$K_a = [H_3O^+]_{eq}[OH^-]_{eq} = 10^{-14}.$$
This value indicates that H$_2$O has a poor tendency to ionize. Why? Because H$_2$O is a very stable molecule and in order to ionize it, we would need to break an O-H bond. The O-H bond dissociation energy in water is approximately 5.15 eV per bond, which is considered very strong. In order to break this bond, the ion-dipole interaction between H$^+$ and another H$_2$O molecule would need to be comparable or stronger. In Example 2.5 (see p. 71), we calculated this interaction energy for sodium and water; we can do a similar calculation here. The distance between nearest-neighbor water molecules in liquid water (1.85 Å) is approximately 2.8 Å. At time $t = 0$, this gives the interaction energy between the ion-dipole that could potentially compete against the O-H bond energy. This ion-dipole interaction energy is:

$$
(9 \times 10^9 \text{N.m}^2/\text{C}^2)(1.6 \times 10^{-19} \text{C})(1.85 \times 3.336 \times 10^{-30} \text{C.m})/\left(2.8 \times 10^{-10} \text{ m}\right)^2 \approx 8.9 \times 10^{-19} \text{ J},
$$

and since 1 eV = $1.6 \times 10^{-19}$ J, this energy is $\sim 0.7$ eV, which is much weaker than the O-H bond dissociation energy (5.15 eV). Thus, this energy is highly unlikely to break the O-H bond of water and lead to ionization. This is why the equilibrium favors reactants, i.e. water is an ineffective acid.

From the relationship $pK_w = pK_a + pK_b$, we find that $pK_b = 0$. The $pK_a$ should not be confused with the pH. The pH is $-\log_{10}[\text{H}_3\text{O}^+]$. For pure water at 25°C the pH is 7; however, hydronium ion concentration can be altered if other proton donors or acceptors are also present in solution. It is best to think of $pK_a$ as a property of a given acid (proton donor in a given solvent), whereas pH is a property of the solution (accounting for all species present). For example, the $pK_a$ of HF in water at 25°C will always equal 3.18. However, the pH of the solution depends on the amount of HF in solution (and any other combination of acids and/or bases that may be present).

Let us now look at the acid H$_3$O$^+$. The $pK_a$ of H$_3$O$^+$ is 0. What is the significance of $pK_a=0$? In the reaction:

$$\text{H}_3\text{O}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{O} + \text{H}_3\text{O}^+,$$

we have (by definition),

$$K_a = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}[\text{H}_2\text{O}]_{\text{eq}}}{[\text{H}_3\text{O}^+]_{\text{eq}}[\text{H}_2\text{O}]_{\text{eq}}} = 1,$$

which implies that $pK_a=0$ (strong acid). In other words, a proton is donated from one hydronium ion to a neighboring water molecule. But this is the same as proton transfer in the reverse direction. We therefore suspect that the energy cost of swapping the proton from one H$_2$O to the next is minimal, hence the reason why H$_3$O$^+$ is a strong acid. Can you quantitatively explain this behavior in terms of interaction energies?
5.5.4. Example. Suppose we have a reaction:

\[
\text{CH}_3\text{COOH}(aq) + \text{NH}_3(aq) \rightleftharpoons \text{CH}_3\text{COO}^-(aq) + \text{NH}_4^+(aq).
\]

Its equilibrium constant is

\[
K = \frac{[\text{CH}_3\text{COO}^-]_{\text{eq}}[\text{NH}_4^+]_{\text{eq}}}{[\text{CH}_3\text{COOH}]_{\text{eq}}[\text{NH}_3]_{\text{eq}}}.
\]

This reaction is the difference between two acid ionization reactions:

\[
\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+, \quad K_{a1} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}[\text{CH}_3\text{COO}^-]_{\text{eq}}}{[\text{CH}_3\text{COOH}]_{\text{eq}}},
\]

\[
\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3, \quad K_{a2} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}[\text{NH}_3]_{\text{eq}}}{[\text{NH}_4^+]_{\text{eq}}}.
\]

The equilibrium constant for the net reaction is the ratio of the equilibrium constants for the separate reactions:

\[
K = \frac{K_{a1}}{K_{a2}}.
\]

This method can be used to predict the outcome of the net reaction, provided it can be decomposed as the difference between two elementary reactions. Equilibrium constants for elementary acid ionization reactions can be found in Table 5.1. Looking up values from the Table for CH₃COOH and NH₄⁺, we find:

\[
K = \frac{1.76 \times 10^{-5}}{5.6 \times 10^{-10}} \approx 31,428.
\]

Table 5.1. Ionization constants of acids at 25°C. For a given conjugate pair, pKₐ is obtained from pKₐ from the relationship pKₐ = pKₐ + pKₐ.

<table>
<thead>
<tr>
<th>Acid</th>
<th>HA (acid)</th>
<th>A⁻ (conjugate base)</th>
<th>Kₐ</th>
<th>pKₐ</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydroiodic</td>
<td>HI</td>
<td>I⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Hydrobromic</td>
<td>HBr</td>
<td>Br⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Perchloric</td>
<td>HClO₄</td>
<td>Cl⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Hydrochloric</td>
<td>HClO₃</td>
<td>ClO₄⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Chloric</td>
<td>HClO₃</td>
<td>ClO₅⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Sulfuric (first ionization)</td>
<td>H₂SO₄</td>
<td>HSO₃⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Nitric</td>
<td>HNO₃</td>
<td>NO₃⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Hydronium ion</td>
<td>H₃O⁺</td>
<td>H₂O⁺</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Iodic</td>
<td>HIO₄</td>
<td>IO₅⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Oxalic (first ionization)</td>
<td>H₂C₂O₄</td>
<td>H₂C₂O₄⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Sulfurous (first ionization)</td>
<td>H₂SO₃</td>
<td>HSO₃⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Sulfuric (second ionization)</td>
<td>HSO₄⁻</td>
<td>SO₄²⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Chlorous</td>
<td>HClO₄</td>
<td>ClO₅⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Phosphoric (first ionization)</td>
<td>H₃PO₄</td>
<td>H₃PO₄⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Arsenic (first ionization)</td>
<td>H₃AsO₄</td>
<td>H₃AsO₄⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Chloroacetic</td>
<td>CH₂CICOOH</td>
<td>CH₂CICOO⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Hydrofluoric</td>
<td>HF</td>
<td>F⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Nitrous</td>
<td>HNO₂</td>
<td>NO₂⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Formic</td>
<td>HCOOH</td>
<td>HCOO⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Benzoic</td>
<td>C₆H₅COOH</td>
<td>C₆H₅COO⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Oxalic (second ionization)</td>
<td>H₂C₂O₄</td>
<td>H₂C₂O₄⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Hydrazoic</td>
<td>HN₃</td>
<td>N₅⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
<tr>
<td>Acetic</td>
<td>CH₃COOH</td>
<td>CH₃COO⁻</td>
<td>0.01</td>
<td>2.08</td>
</tr>
</tbody>
</table>
5.5. The pH Scale

<table>
<thead>
<tr>
<th>Acid Name</th>
<th>Chemical Formula</th>
<th>pK_a Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Propionic</td>
<td>CH₃CH₂COOH</td>
<td>4.75</td>
</tr>
<tr>
<td>Pyridinium ion</td>
<td>C₅H₅N⁺</td>
<td>5.25</td>
</tr>
<tr>
<td>Carbonic (first ionization)</td>
<td>H₂CO₃</td>
<td>6.37</td>
</tr>
<tr>
<td>Sulfurous (second ionization)</td>
<td>HSO₃⁻</td>
<td>6.91</td>
</tr>
<tr>
<td>Arsenic (second ionization)</td>
<td>H₃AsO₄⁻</td>
<td>7.03</td>
</tr>
<tr>
<td>Hydroxulfuric</td>
<td>H₂S</td>
<td>7.04</td>
</tr>
<tr>
<td>Phosphoric (second ionization)</td>
<td>H₃PO₄⁻</td>
<td>7.21</td>
</tr>
<tr>
<td>Hypochlorous</td>
<td>HClO</td>
<td>7.53</td>
</tr>
<tr>
<td>Hydrocyanic</td>
<td>HCN</td>
<td>9.21</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
<td>9.25</td>
</tr>
<tr>
<td>Carbonic (second ionization)</td>
<td>HCO₃⁻</td>
<td>10.32</td>
</tr>
<tr>
<td>Arsenic (third ionization)</td>
<td>H₃AsO₄⁻</td>
<td>11.53</td>
</tr>
<tr>
<td>Hydrogen peroxide</td>
<td>H₂O₂</td>
<td>11.62</td>
</tr>
<tr>
<td>Phosphoric (third ionization)</td>
<td>H₃PO₄⁻</td>
<td>12.67</td>
</tr>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>14.00</td>
</tr>
</tbody>
</table>

5.5.5. The Meaning of $K_a$ and Strength of the Acid. Let us clarify the meaning of $K_a$. $K_a$ is a property of an acid. Its numerical value depends on the type of acid and solvent. It quantifies the tendency (affinity) of an acid to give up a proton. Strong acids have large $K_a$ values whereas weak acids have small $K_a$ values. The precise value of $K_a$ reflects the strength of intermolecular interactions binding the proton to its acid relative to the ability (strength of interaction) of the solvent molecules to pull this proton away from the acid. For ionization to occur, the solvent must be able to overcome the bond dissociation energy that’s keeping this proton bound to its acid. In the dilute limit, the $K_a$ does not depend on acid concentration, and does not significantly depend on other species dissolved in that solution.

For a given solvent, and keeping the concentration of an acid fixed (say, 1 M), the pH of the solution depends on $K_a$. For example, the pH of a 1 M HCl aqueous solution at 25°C will be 0. According to Table 5.1, the $pK_a$ of HCl is -7. The $pK_a$ and pH are different: the pH will change if $[\text{HCl}]$ changes (the more acid we add, the lower the pH), but the $pK_a$ itself will not change.

In contrast, a 1 M solution of CH₃COOH ($pK_a$=4.75) has a pH which is determined from the equilibrium:

$$K_a = 1.76 \times 10^{-5} = [\text{CH}_3\text{COO}^-]_{\text{eq}}[\text{H}^+]_{\text{eq}}/[\text{CH}_3\text{COOH}]_{\text{eq}}.$$ 

where the concentrations are equilibrium concentrations. If we start at time $t = 0$ with 1 M of CH₃COOH. At equilibrium, $x$ amount of CH₃COOH will dissociate. At the same time $x$ amount of CH₃COO⁻ and H⁺ will be produced. Thus, $1.76 \times 10^{-5} = x^2/(1-x) \approx x^2$ since $x \ll 1$. This leads to $x = \sqrt{1.76 \times 10^{-5}} \approx 0.0042$ M. The pH is therefore,

$$-\log_{10}(\sqrt{1.76 \times 10^{-5}}) \approx 2.5.$$ 

Here, we had the acid concentration fixed (1 M). But the two acids (HCl vs CH₃COOH) had a different $pK_a$. The acid with the lowest $pK_a$ (HCl at -7)
had the lowest pH (0). The acid with the highest $pK_a$ ($\text{CH}_3\text{COOH}$ at 4.75) had the highest pH (2.4).

5.6. “ICE” Tables

In subsequent sections, we will use the concept of an “ICE table”. The ICE table is a convenient way to organize your equilibrium calculations. They are used in cases where initial and final conditions are to be related to one another. ICE stands for (I)initial, (C)change and (E)quilibrium. First, we have the initial (I) conditions of a reaction. Then, these reactants react to form products. This “change” (C) leads to a new equilibrium (E). So I, C and E form the rows of an ICE table. The columns are the reactants and products of a reaction. The entries of the table are the amounts (moles) or concentrations (molar or molal), of reactants and products at a given time (initial or final conditions). We sometimes don’t know the “extent” of the reaction; however, it can be inferred algebraically by analyzing the ICE table and relating initial and final conditions using known equilibrium constants. ICE tables are frequently used in problems of acid-base titration.

5.6.1. Example 1. Suppose that we start with a mixture of 0.500 mol/L of $\text{N}_2$ and 0.800 mol/L $\text{H}_2$ and allow it to reach equilibrium with the product, ammonia. It is found that, at equilibrium at a certain temperature, the molar concentration of $\text{NH}_3$ is 0.150 mol/L. Calculate the equilibrium constant for the reaction at that temperature.

Strategy: Write down an equation for the equilibrium constant. Identify the change in concentration of one substance, use reaction stoichiometry to calculate the implied changes in the other substances. Solve for new concentrations and plug into equation for $K_C$.

The reaction is

$$\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g), \quad K_C = \frac{[\text{NH}_3]^2}{[\text{N}_2]_\text{eq}[\text{H}_2]^3_\text{eq}}.$$  

where $K_C$ has units of $1/(\text{concentration})^2$, due to the particular stoichiometry of this reaction (2 mol on products side, 4 mol on reactants side).

The reaction implies that 1 mol $\text{N}_2 = 2$ mol $\text{NH}_3$ and 3 mol $\text{H}_2 = 2$ mol $\text{NH}_3$. Therefore, because the molar concentration of $\text{NH}_3$ increases by 0.150 mol/L to reach equilibrium, the concentration of $\text{N}_2$ decreases by half that much, 0.075 mol/L, and the concentration of $\text{H}_2$ by 1.5 times, 0.225 mol/L. We have:
5.6. “ICE” Tables

<table>
<thead>
<tr>
<th></th>
<th>N₂</th>
<th>H₂</th>
<th>NH₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Initial molar concentration</td>
<td>0.500</td>
<td>0.800</td>
<td>0</td>
</tr>
<tr>
<td>2. Change in molar concentration</td>
<td>-0.075</td>
<td>-0.225</td>
<td>+0.150</td>
</tr>
<tr>
<td>3. Equilibrium molar concentration</td>
<td>0.425</td>
<td>0.575</td>
<td>0.150</td>
</tr>
</tbody>
</table>

It follows that:

\[ K_C = \frac{[NH_3]_{eq}^2}{[N_2]_{eq}[H_2]_{eq}^3} = \frac{(0.150)^2}{0.425 \times (0.575)^3} = 0.278 \]

where the units are L²/mol². If you are given the value of an equilibrium constant with no units, it will either have no units (due to normalization to 1 atm or 1 mol/L) or its units will be in terms of the concentration (or pressure) units given in the problem.

**Exercise:** A sample of gaseous N₂O₄ with an initial partial pressure of 3.0 atm is prepared at a certain temperature and allowed to react as follows:

\[ N_2O_4(g) \rightleftharpoons 2NO_2(g) \]

Once equilibrium is reached, it is found that the partial pressure of N₂O₄ has dropped to 1.0 atm. Find the equilibrium partial pressure of NO₂, then calculate \( K_P \) [Answer: \( P_{NO_2} = 4.0 \) atm and \( K_P = 16 \).]

5.6.2. **Example 2.** Suppose that 1.50 mol PCl₅ (phosphorous pentachloride) is placed in a reaction vessel of volume 500 mL and allowed to reach equilibrium with its decomposition products phosphorous trichloride and chlorine at 250°C, when \( K_C = 1.80 \). What is the composition of the equilibrium mixture? All three substances are gases at 250°C.

**Strategy:** Write chemical equation and expression for equilibrium constant. Denote by \( x \) the change in molar concentration of the decomposing substance. Use reaction stoichiometry to express the molar concentrations of the decomposition products in terms of \( x \). Calculate the molar concentrations by dividing the number of moles of each substance by the volume of the container.

The chemical equation is

\[ PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g), \quad K_C = \frac{[PCl_3]_{eq}[Cl_2]_{eq}}{[PCl_5]_{eq}}. \]

The initial molar concentration of PCl₅ is

Molar concentration of PCl₅ = \( \frac{1.50}{0.500} = 3.00, \)

and we suppose that the change in its molar concentration is \(-x\) mol/L. Using the chemical equation above, we see that (with all concentrations in mol/L):
Step 3: these values are the sums of the initial concentrations, step 1, and
the changes in concentration brought about by reaction, step 2.
Step 2: the stoichiometry of the reaction implies that if the molar concen-
tration of PCl\textsubscript{5} decreases by \( x \), then the molar concentrations of PCl\textsubscript{3} and
Cl\textsubscript{2} both increase by \( x \).
Step 4: Substitution of these equilibrium values into the expression for the
equilibrium constant gives:
\[
K_C = \frac{[\text{PCl}_3]_{eq}[\text{Cl}_2]_{eq}}{[\text{PCl}_5]_{eq}} = \frac{x \times x}{3.00 - x},
\]
and because we are told that \( K_C = 1.80 \), the equation we have to solve is
\[
1.80 = \frac{x^2}{3.00 - x}.
\]
This expression rearranges to the quadratic equation:
\[
x^2 + 1.80x - 5.40 = 0.
\]
The solutions of this equation are:
\[
x = \frac{-1.80 \pm \sqrt{(1.80)^2 - 4(1)(-5.40)}}{2} = 1.59 \text{ and } -3.39.
\]
Because the concentrations must be positive and because (from step 3) \( x \)
is the molar concentration of each of the products, we select 1.59 as the
solution. It follows that at equilibrium
\[
[\text{PCl}_5]_{eq} = 3.00 - x = 3.00 - 1.59 = 1.41,
[\text{PCl}_3]_{eq} = x = 1.59,
[\text{Cl}_2]_{eq} = x = 1.59,
\]
that is, the equilibrium concentrations of PCl\textsubscript{5}, PCl\textsubscript{3} and Cl\textsubscript{2} are 1.41 mol/L,
1.59 mol/L and 1.59 mol/L, respectively.

**Exercise:** Bromine monochloride, BrCl, decomposes into bromine and
chlorine and reaches the equilibrium \( 2\text{BrCl}(g) \rightleftharpoons \text{Br}_2(g) + \text{Cl}_2(g) \), for
which \( K_C = 32 \) at 500 K. If initially pure BrCl is present at a concentra-
tion of \( 3.30 \times 10^{-3} \) mol/L, what is its molar concentration in the mixture
at equilibrium? [Answer: \( 3 \times 10^{-4} \) mol/L.]

5.7. Weak Acids and Bases
5.7. Weak Acids and Bases

5.7.1. Weak Acids. A weak acid has a $K_a$ smaller than 1. We write the reactions as bidirectional reactions:

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

Example (Oxtoby [1], Example 15.6):

**Question:** Acetic acid (CH$_3$COOH) has $K_a$ of $1.76 \times 10^{-5}$ at 25°C. Suppose 1.000 mol is dissolved in enough water to give 1.000 L of solution. Calculate the pH and the fraction of acetic acid ionized at equilibrium.

**Solution:** The initial concentration is 1.000 M. Let’s denote our “progress variable” as $y$. i.e. $y$ can be taken to be the amount of acetic acid that ionizes. Then our “ICE table” looks like this:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

<table>
<thead>
<tr>
<th>1. Initial concentration (M)</th>
<th>$[CH_3COOH]$</th>
<th>$[H_3O^+]$</th>
<th>$[CH_3COO^-]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.000</td>
<td>1.000</td>
<td>$\approx 0$</td>
<td>0</td>
</tr>
<tr>
<td>2. Change in molar concentration</td>
<td>$-y$</td>
<td>$+y$</td>
<td>$+y$</td>
</tr>
<tr>
<td>3. Equilibrium molar concentration</td>
<td>$1 - y$</td>
<td>$y$</td>
<td>$y$</td>
</tr>
</tbody>
</table>

Writing down the equation for the equilibrium constant:

$$K_a = 1.76 \times 10^{-5} = \frac{[H_3O^+]_{eq}[CH_3COO^-]_{eq}}{[CH_3COOH]_{eq}} = \frac{y^2}{1 - y}$$

We can solve the quadratic equation or proceed by approximating $1 - y \approx 1$ because we know $y$ should be small (acetic acid is a weak acid, so only a small fraction should ionize at equilibrium).

$$1.76 \times 10^{-5} = \frac{y^2}{1 - y} \approx \frac{y^2}{1}$$

$$y = \sqrt{1.76 \times 10^{-5}} = 4.20 \times 10^{-3} \text{ M},$$

which is large compared to self-ionization of water ($10^{-7}$ M). The pH is $-\log_{10}(4.2 \times 10^{-3}) = 2.38$ and fraction ionized is $y/1.000 = y = 4.2 \times 10^{-3}$. 
Without approximation: Suppose in the previous problem that instead of dissolving 1.000 mol we dissolve 0.001 mol. Then, we can no longer make this approximation in the denominator. Our equation is:

\[
y^2 - (0.001 - y) = 1.76 \times 10^{-5}
\]

Because \( y \) is not small compared to 0.001 we should instead solve the quadratic equation:

\[
y^2 + (1.76 \times 10^{-5})y - (1.76 \times 10^{-8}) = 0,
\]

\[
y = 1.24 \times 10^{-4} \text{ M} = [H_3O^+] = [CH_3COO^-],
\]

which leads to \( pH = -\log_{10}(1.24 \times 10^{-4}) = 3.91 \) and the percentage of acid that is ionized is \( (100\%) \times 1.24 \times 10^{-4}/0.001 \text{ M} = 12.4\% \).

5.7.2. Weak Bases. A weak base is only partially protonated in water and its \( K_b \) is less than 1:

\[
\text{H}_2\text{O}(l) + \text{NH}_3(aq) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq), \quad K_b = \frac{[\text{NH}_4^+]_{eq}[\text{OH}^-]_{eq}}{[\text{NH}_3]_{eq}} = 1.8 \times 10^{-5}
\]

Example (Oxtoby [1], Example 15.8):

\[\text{Question:} \text{ Calculate the pH of a solution made by dissolving 0.0100 mol of NH}_3 \text{ in enough water to give 1.000 L of solution at 25}^\circ \text{C. The} \quad K_b \text{ for ammonia is } 1.8 \times 10^{-5}.\]

\[\text{Solution: Setup an ICE table:}\]

<table>
<thead>
<tr>
<th></th>
<th>( \text{H}_2\text{O}(l) + \text{NH}_3(aq) )</th>
<th>( \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq) )</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Initial (M)</td>
<td>0.0100</td>
<td>( \approx 0 )</td>
</tr>
<tr>
<td>2. Change</td>
<td>(-y)</td>
<td>(+y)</td>
</tr>
<tr>
<td>3. Equilibrium</td>
<td>(0.0100 - y)</td>
<td>(y)</td>
</tr>
</tbody>
</table>

We get a quadratic equation, which we can solve to get:

\[
K_b = 1.8 \times 10^{-5} = \frac{y^2}{0.0100 - y}, \quad y = 4.15 \times 10^{-4} \text{ M} = [\text{OH}^-]_{eq}.
\]

\[
[H_3O^+]_{eq} = \frac{K_w}{[\text{OH}^-]_{eq}} = \frac{1.0 \times 10^{-14}}{4.15 \times 10^{-4}} = 2.4 \times 10^{-11},
\]

\[
pH = -\log_{10}(2.4 \times 10^{-11}) = 10.64
\]
5.8. Buffer Solutions

The goal of a buffer solution is to maintain the pH approximately constant with respect to small fluctuations of acids or bases. Biological systems require maintaining the pH within very narrow ranges. A buffer is made from a weak acid or base. Consider a weak acid, formic acid (HCOOH), and its conjugate base, formate ion (HCOO\(^{-}\)). Equilibrium is established:

\[
\text{HCOOH}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCOO}^{-}(aq),
\]

\[
K_a = 1.77 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+]_{eq}[\text{HCOO}^{-}]_{eq}}{[\text{HCOOH}]_{eq}}, \quad pK_a = 3.75.
\]

The formate ion can also be obtained by dissolving a salt such as sodium formate (NaHCOO). To form a buffer, we also add salt for purposes of stabilizing the pH. This is the method used to create a buffer.

5.8.1. Calculating the pH of a Buffer Solution. Example (Oxtoby [1], Example 15.8):

**Question:** Suppose 1.00 mol of HCOOH and 0.500 mol of NaHCOO are added to water and diluted to 1.00 L. Calculate the pH of the solution.

**Solution:** As usual, we set up the ICE table:

<table>
<thead>
<tr>
<th></th>
<th>HCOOH(aq) + H(_2)O(l) ⇌ H(_3)O(^{+})(aq) + HCOO(^{-})(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Initial (M)</td>
<td>1.00</td>
</tr>
<tr>
<td>2. Change</td>
<td>−y</td>
</tr>
<tr>
<td>3. Equilibrium</td>
<td>1.00−y</td>
</tr>
</tbody>
</table>

\[
y(0.500+y) = K_a = 1.77 \times 10^{-4},
\]

\[
y(0.500) \approx 1.77 \times 10^{-4}, \quad y \approx 3.54 \times 10^{-4} \text{ M} = [\text{H}_3\text{O}^+] , \quad \text{pH} = 3.45.
\]

5.8.1.1. *How a Buffer Works.* Consider the reaction for a weak acid HA:

\[
\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^{-}(aq)
\]

\[
K_a = \frac{[\text{H}_3\text{O}^+]_{eq}[\text{A}^{-}]_{eq}}{[\text{HA}]_{eq}} , \quad [\text{H}_3\text{O}^+]_{eq} = K_a \frac{[\text{HA}]_{eq}}{[\text{A}^{-}]_{eq}} \approx K_a \frac{[\text{HA}]_{eq} + \delta_1}{[\text{A}^{-}]_{eq} + \delta_2}.
\]

This equation holds regardless of the concentration values involved.

---

\(^3\)Lysosome pH is in the range 4.5-5.0; blood plasma, 7.4; average pH in typical muscle cells, 6.1; average pH in typical liver cells, 6.9; blood plasma, 7.4; mitochondria, 7.8; Golgi apparatus, 6.5; endoplasmic reticulum, 7.0; cytosol, 7.4.
If we make both concentrations $[HA]$ and $[A^-]$ very large, then small perturbations in each quantity ($\delta_1, \delta_2$) produce little change in their ratio\(^4\). Thus, $[H_3O^+]$ is relatively unaffected by these perturbations.

But if $[H_3O^+]$ is unaffected, then the pH is unaffected. Thus, a buffer stabilizes the pH of the solution against perturbations.

If both $[HA]$ and $[A^-]$ very large, then their ratio approaches a constant, say 1. In that case, the pH $\approx pK_a$, according to the above equation. The buffer therefore attempts to keep the pH near the $pK_a$ value. But other values of the pH are possible, by selecting different concentrations for $[HA]$ and $[A^-]$. From a practical standpoint, it is easiest to prepare solutions for which $[HA] \approx [A^-]$, in which case, a buffer is prepared by selecting an acid whose $pK_a$ value is close to the desired pH value and the concentrations $[HA] \approx [A^-]$ are selected to “tweak” the pH to the desired value.

To make $[HA]$ very large we simply dissolve a large amount of acid. To make $[A^-]$ very large, we simply add salt that dissociates to yield the conjugate base. In the first example, we used sodium formate as the salt.

### 5.8.2. pH Stability

Adding Strong Acid to Buffer Solution: in previous problem suppose that we add 0.10 mole of a strong acid (e.g. HCl).

**Method 1:** HCl is a strong acid and will completely dissociate. The $H^+$ ions can be assumed to react with formate ions to produce formic acid, i.e.,

\[
HCl(aq) \rightarrow H^+(aq) + Cl^−(aq)
\]

\[
HCOOH(aq) \leftarrow HCOO^−(aq) + H^+(aq) \quad (*)
\]

For now, the reactions are assumed to proceed to completion (because HCl is a strong acid and HCOO$^-$ is a “strong” base). Thus, the initial conditions are such that the amount of HCOOH is 1.0 mol (from the formic acid) plus 0.1 mol (from the HCl, which leads to the formation of HCOOH in the same amount). Thus, the initial amount of HCOOH is 1.1 mol.

Similarly, while we had 0.5 mol of HCOO$^-$ from the sodium formate, now we have 0.4 mol of it, because 0.1 mol of it was used to form HCOOH due to the added HCl.

That being said, only the HCl reaction truly proceeds to completion since its $K_a$ value is $\sim 10^7$ (see Table 5.1). For the reaction (*) the $K_a$ value is $1.77 \times 10^{-4}$. Thus, the reaction (*) is not complete (for the reaction to be complete, its value should be closer to $\sim 10^{-7}$). Thus, our progress variable $y > 0$ will describe a decrease in the amount of HCOOH from 1.10 mol in

\[^4\]If $[HA]$ and $[A^-]$ are very large compared to $\delta_1$ and $\delta_2$, respectively, then adding some perturbation $\delta_i$ ($i=1$ or 2) will have negligible effect on the ratio $[HA]/[A^-] \approx ([HA] + \delta_1)/([A^-] + \delta_2)$. $\delta_1$ and $\delta_2$ can be thought of as a perturbation to the system introduced by the addition or removal of substances that would normally alter the pH.
5.8. Buffer Solutions

the ICE table below (since 1.10 mol is really an overestimate of the true amount of HCOOH that we have at equilibrium).

<table>
<thead>
<tr>
<th></th>
<th>HCOOH(aq) + H₂O(l) ⇌ H₃O⁺(aq) + HCOO⁻(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.10</td>
</tr>
<tr>
<td>Change</td>
<td>−y</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.10−y</td>
</tr>
</tbody>
</table>

\[
y(0.40 + y) = K_a = 1.77 \times 10^{-4},
\]

\[
y \approx 0.40 + y \
\]

\[
y \approx 1.77 \times 10^{-4}, \quad y \approx 4.9 \times 10^{-4} \ M = [\text{H}_3\text{O}^+], \quad \text{pH} = 3.31.
\]

If these approximations are confusing, do not worry; just solve the full quadratic equation instead.

As you can see, the pH did not change by much (from 3.45 down to 3.31) even though 0.1 mol of a strong acid was added. Had this amount of acid been added to pure water, the pH would have dropped from 7 to 1.

Method 2: There is nothing special about assuming that the H⁺ ions react with formate ions to produce formic acid. We could have assumed instead that HCl increases the number of H₃O⁺ ions and allowed y to describe the decrease in the number of such ions:

<table>
<thead>
<tr>
<th></th>
<th>HCOOH(aq) + H₂O(l) ⇌ H₃O⁺(aq) + HCOO⁻(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.00</td>
</tr>
<tr>
<td>Change</td>
<td>+y</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.00+y</td>
</tr>
</tbody>
</table>

\[
\frac{(0.1 - y)(0.5 - y)}{1 + y} = K_a = 1.77 \times 10^{-4},
\]

Instead of making approximations, we solve the quadratic equation:

\[
y^2 - (0.6 + 1.77 \times 10^{-4})y + 0.05 - 1.77 \times 10^{-4} = 0.
\]

This yields the same end result as Method 1 (as it should):

\[
y = 0.099514, \quad [\text{H}_3\text{O}^+] = 0.1 - y = 4.9 \times 10^{-4} \ M, \quad \text{pH} = 3.31.
\]

(Recall that the pKₐ of the buffer was 3.45.)

5.8.3. Design of a Buffer. The normal procedure is to choose an acid with pKₐ as close as possible to the desired pH. Then we adjust the concentrations of acid and conjugate base to give exactly the desired pH. To see, this, note that since equilibrium concentrations of acid and base in the
Buffers are close to the initial concentrations:

\[ K_a = \frac{[H_3O^+]}{[HA]} \approx \frac{[H_3O^+]}{[HA]}_0, \quad [H_3O^+]_0 \approx [HA]_0 \cdot K_a. \]

Taking the logs of both sides of this equation yields:

\[ pH \approx pK_a - \log_{10} \frac{[HA]}{[A^-]}_0. \]

This is called the Henderson-Hasselbach equation. It relates the desired pH to the pKa of the weak acid (chosen as close a possible to the desired pH) and the ratio of concentrations of weak acid to conjugate base. This ratio is used to “tweak” the effects of the pKa to bring it closer to the desired pH.

We also see from this equation that the buffer will be robust to changes in pH provided that the concentrations of weak acid to conjugate base are chosen large enough, so that small perturbation in their values does not affect the pH appreciably.

Let’s use this equation to design a buffer. We will need a table of pKa values to enable us to pick the acid whose pKa is close to the desired pH (see Table 5.1).

**Design a buffer with pH 4.60:**

\[ pH \approx pK_a - \log_{10} \frac{[HA]}{[A^-]}_0. \]

The pKa of acetic acid is 4.75, which is reasonably close to the desired value of 4.60, so we choose CH₃COOH/CH₃COO⁻ as the buffer system. We can fine-tune the pH from 4.75 down to 4.60 by adjusting the ratio of weak acid to conjugate base. First we write,

\[ pH = 4.60 = pK_a - \log_{10} \frac{[CH_3COOH]}{[CH_3COO^-]}_0 \]

\[ \log_{10} \frac{[CH_3COOH]}{[CH_3COO^-]}_0 = pK_a - pH = 4.75 - 4.60 = 0.15 \]

\[ \frac{[CH_3COOH]}{[CH_3COO^-]}_0 = 10^{0.15} = 1.4. \]

This can be obtained by dissolving 0.100 mol of sodium acetate (NaCH₃COO) and 0.140 mol of acetic acid in water and diluting to 1.00 L. Other solutions are possible.

**Note:** While the ratio \( \frac{[CH_3COOH]}{[CH_3COO^-]}_0 \) is important, the concentrations must be high enough. The higher the initial concentrations, the higher the ability of the buffer to resist changes in the pH.
5.8.4. Titration of Strong Acid with Strong Base. The simplest problem involves titrating a strong acid with a strong base. It simply involves counting number of moles to achieve neutralization:

\[ \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \rightarrow 2\text{H}_2\text{O}(l) \]

**Question:** Suppose a solution of 100.0 mL (0.1000 L) of 0.1000 M HCl is titrated with 0.1000 M NaOH. What does the titration curve look like?

Figure 5.7 shows the titration curve for titration of strong acid by a strong base.

![Titration curve (strong acid by a strong base).](image)

This curve should be analyzed into 3 regions:

- **Below the equivalence point** we have acidic conditions. Adding base will neutralize the acid, up the point where \( n_{\text{base}} = n_{\text{acid}} \), the equivalence point. Thus, below the equivalence point we have:

\[
\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}\left(\frac{n_{\text{H}_3\text{O}^+}}{V_{\text{tot}}}\right) = -\log_{10}\left(\frac{n_{\text{acid}} - n_{\text{base}}}{V_{\text{acid}} + V_{\text{base}}}\right)
\]

\[
= -\log_{10}\left(\frac{c_{\text{acid}}V_{\text{acid}} - c_{\text{base}}V_{\text{base}}}{V_{\text{acid}} + V_{\text{base}}}\right)
\]

which equals pH=1 in the limit \( V_{\text{base}} \rightarrow 0 \) since \( c_{\text{acid}}=0.1000 \text{ M HCl} \).

**Note:** This function is of the form \( f(x) = -\log_{10}( (1-x)/(1+x)) \). With a graphing calculator or a computer you can try plotting this function over the range \( x \in [0,1] \) and compare with the shape of the titration curve before the equivalence point.

- **At the equivalence point** the number of moles of base added equals the number of moles of acid originally present. The concentrations of OH⁻
and \( H_3O^+ \) must be equal and the pH is due solely to the auto-ionization of water. Thus, the pH is precisely 7.00.

- **Above the equivalence point** we have basic conditions. Adding more base makes it more basic. Thus, the only thing that matters is the moles of excess \( OH^- \) ions. The pH is obtained from the relation

\[
pK_w = pH + pOH.
\]

Because the base is a strong base, it is fully protonated. Thus, if we denote the concentration of excess \( OH^- \) ions by \([OH^-]\),

\[
[OH^-] = \frac{\text{moles of excess } OH^-}{\text{total volume}} = \frac{n_{\text{base}} - n_{\text{acid}}}{V_{\text{acid}} + V_{\text{base}}} = \frac{c_{\text{base}}V_{\text{base}} - c_{\text{acid}}V_{\text{acid}}}{V_{\text{acid}} + V_{\text{base}}}.
\]

So,

\[
pH = pK_w - pOH = 14.00 + \log_{10} \frac{c_{\text{base}}V_{\text{base}} - c_{\text{acid}}V_{\text{acid}}}{V_{\text{acid}} + V_{\text{base}}}.
\]

**Note:** This function is of the form \( \log_{10}((x-1)/(1+x)) \). With a graphing calculator or a computer you can try plotting this function over the range \( x \in [1, 4] \) and compare with the shape of the titration curve after the equivalence point.

### 5.8.5. Titration of Weak Acid with Strong Base.

Consider the titration of 100.0 mL of a 0.1000 M solution of acetic acid (\( CH_3COOH \)) with 0.1000 M NaOH. So the concentrations are the same as in the previous problem, except that the acid is a weak acid instead of a strong acid. Because we have a weak acid, the \( pK_a \) should enter the picture. In what follows, we denote the acid as \([HA]\) and its conjugate base as \([A^-]\).

We note the following main differences:

- The titration curve (see Fig.5.9) begins at higher pH because the acid is not completely dissociated, leading to lower concentrations of \( H_3O^+ \) ions compared to the case of HCl.
- A gradually rising portion of the curve, called the buffer region, appears before the steep rise to the equivalence point. At the midpoint, half of the original acid has reacted.
- The pH at equivalence point is greater than 7.00. When \( c_{\text{acid}}V_{\text{acid}} = c_{\text{base}}V_{\text{base}} \), the solution will be more basic (there will be excess \( OH^- \)) because we over-neutralize the acid, since the acid was not completely dissociated (weak acid).

Let us examine the main elements of this titration curve:

- **Initial point.** Before the base is added the weak acid is partially dissociated. The pH is obtained by setting up an ICE table for the dissociation
reaction

\[ \text{HA}(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq), \]

(let \( x = [\text{H}_3\text{O}^+] \))

\[ K_a = \frac{[\text{H}_3\text{O}^+]_{eq}[\text{A}^-]_{eq}}{[\text{HA}]_{eq}} = \frac{x^2}{[\text{HA}]_0 - x} \approx \frac{x^2}{[\text{HA}]_0}, \quad x \approx \sqrt{K_a \cdot [\text{HA}]_0}. \]

- **Buffer region.** As soon as we add NaOH, it reacts with HA to form \( \text{A}^- \). This means that up to the equivalence point we have a buffer over much of that interval. The pH is found from:

\[ [\text{H}_3\text{O}^+]_{eq} = K_a \frac{[\text{HA}]_{eq}}{[\text{A}^-]_{eq}} \quad \text{or} \quad \text{pH} = pK_a - \log_{10} \frac{[\text{HA}]_{eq}}{[\text{A}^-]_{eq}}. \]

We note that at the midpoint of the buffer region, half the original HA has reacted, so \([\text{HA}] = [\text{A}^-]\) or \([\text{A}^-]/[\text{HA}] = 1\) and \( \text{pH} = pK_a \). For the reaction at hand, \( \text{OH}^- \) is a stronger base than acetate ion, so it reacts almost completely with the acid originally present:

\[ \text{CH}_3\text{COOH} + \text{OH}^- \rightleftharpoons \text{H}_2\text{O} + \text{CH}_3\text{COO}^-, \quad \text{pH} \approx pK_a - \log_{10} \frac{[\text{CH}_3\text{COOH}]_0}{[\text{CH}_3\text{COO}^-]_0}. \]

- **At equivalence point** the solution is more basic. You can setup an ICE table for the reverse reaction, \( \text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^- \), to show that

\[ [\text{OH}^-] \approx \sqrt{K_b \cdot [\text{A}^-]}, \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}, \quad [\text{H}_3\text{O}^+] \approx \frac{K_w}{\sqrt{K_b \cdot [\text{A}^-]}}. \]

- **After the equivalence point** we have a basic solution. Its properties are calculated based on the quantity of excess \( \text{OH}^- \) ion.

See Fig. 5.9.

![Buffering region](image-url)
5.8.6. Phosphate Buffered Saline (PBS). A commonly used buffer in biology is PBS (phosphate buffered saline). Let’s look at ways to prepare a phosphate buffer.

**Question:** Prepare a “phosphate buffer” with pH of about 7.40.

**Solution:** From $pK_a$ tables we find the following stages of ionization of phosphoric acid:

$$\text{H}_3\text{PO}_4(aq) \rightleftharpoons \text{H}^+(aq) + \text{H}_2\text{PO}_4^-(aq) \quad K_{a_1} = 7.5 \times 10^{-3} \quad pK_{a_1} = 2.12$$

$$\text{H}_2\text{PO}_4^-(aq) \rightleftharpoons \text{H}^+(aq) + \text{HPO}_2^-(aq) \quad K_{a_2} = 6.2 \times 10^{-8} \quad pK_{a_2} = 7.21$$

$$\text{HPO}_2^-(aq) \rightleftharpoons \text{H}^+(aq) + \text{PO}_3^-(aq) \quad K_{a_3} = 4.8 \times 10^{-13} \quad pK_{a_3} = 12.32$$

The most suitable of the three systems is $\text{HPO}_2^-/\text{H}_2\text{PO}_4^-$, because the $pK_a$ of the acid $\text{H}_2\text{PO}_4^-$ is closest to the desired pH. From the Henderson-Hasselbach equation we write

$$\text{pH} = pK_a + \log_{10} \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$7.40 = 7.21 + \log_{10} \frac{[\text{HPO}_2^-]}{[\text{H}_2\text{PO}_4^-]}$$

$$\log_{10} \frac{[\text{HPO}_2^-]}{[\text{H}_2\text{PO}_4^-]} = 0.19, \quad \frac{[\text{HPO}_2^-]}{[\text{H}_2\text{PO}_4^-]} = 1.5$$

One way to prepare this buffer with pH 7.40 is to dissolve disodium hydrogen phosphate ($\text{Na}_2\text{HPO}_4$) and sodium dihydrogen phosphate ($\text{NaH}_2\text{PO}_4$) in a mole ratio of 1.5:1.0 in water. For example, we could dissolve 1.5 moles of $\text{Na}_2\text{HPO}_4$ and 1.0 mole of $\text{NaH}_2\text{PO}_4$ in enough water to make up a 1 L solution.

An actual PBS buffer is slightly different in composition. Instead of $\text{NaH}_2\text{PO}_4$, one dissolves $\text{KH}_2\text{PO}_4$. It also contains additional salts to provide physiological saline conditions.

5.8.7. Polyprotic Acids. Polyprotic acids can donate 2 or more protons. As an example, carbonic acid dissociates to form bicarbonate:

$$\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)$$

$$K_{a_1} = 4.3 \times 10^{-7} = \frac{[\text{H}_3\text{O}^+]_{eq}[\text{HCO}_3^-]_{eq}}{[\text{H}_2\text{CO}_3]_{eq}}$$

and

$$\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CO}_3^{2-}(aq)$$

$$K_{a_2} = 4.8 \times 10^{-11} = \frac{[\text{H}_3\text{O}^+]_{eq}[\text{CO}_3^{2-}]_{eq}}{[\text{HCO}_3^-]_{eq}}$$

Two comments:
• $[\text{H}_3\text{O}^+]$ in the two ionization equilibria are the same. This follows from the equivalence relation ($A \rightleftharpoons E$, $E \rightleftharpoons B$ implies $A \rightleftharpoons B$) at equilibrium, which follows from the principle of detailed balance. Here, $E = [\text{H}_3\text{O}^+]$ and the equations are rearranged as needed.

• $K_{a_2}$ is less than $K_{a_1}$ because the negative charge left behind by the loss of a hydrogen ion in the first ionization causes the second hydrogen ion to be more tightly bound.

The behavior of polyprotic acids is best described using a plot (Fig. 5.10). Take for example, the case of phosphoric acid ($\text{H}_3\text{PO}_4$), which dissociates into the ions $\text{H}_2\text{PO}_4^-$, $\text{HPO}_4^{2-}$, $\text{PO}_4^{3-}$ and $\text{H}_3\text{O}^+$ with acid dissociation equilibrium constants $K_{a_1} = 7.1 \times 10^{-3}$, $K_{a_2} = 6.3 \times 10^{-8}$ and $K_{a_3} = 4.2 \times 10^{-13}$.

**Figure 5.10.** Titration curve for a polyprotic acid.

5.9. Problems

**Problem 56.** Ammonia ($\text{NH}_3$) aqueous solution has a $pK_a$ (9.25 at $25^\circ\text{C}$) that is convenient for the control of pH. A buffer is prepared by mixing 0.10 mol of ammonia with 0.05 mol of $\text{HCl}$ in a volume of 1.00 L aqueous solution. Compute the pH of the solution.

**Solution.** The addition of $\text{HCl}$ converts some $\text{NH}_3$ to its conjugate acid

$$\text{NH}_3(aq) + \text{HCl}(aq) \rightleftharpoons \text{NH}_4^+(aq) + \text{Cl}^-(aq)$$

The resulting solution is a mixture of a weak acid ($\text{NH}_4^+$) and its conjugate base ($\text{NH}_3$). It is a buffer by virtue of the reaction

$$\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq)K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$
Compute the concentrations of the NH$_3$ and NH$_4^+$ after complete reaction with the HCl but before the preceding equilibrium is established

\[
[NH_3]_0 = \frac{(0.10 - 0.05 \text{ mol})}{1.00 \text{ L}} = 0.05 \text{ M} \quad \text{and} \quad [NH_4^+]_0 = \frac{0.05 \text{ mol}}{1.00 \text{ L}} = 0.05 \text{ M}
\]

The equilibrium now reduces the concentration of the NH$_4^+$ as it forms H$_3$O$^+$ and NH$_3$ in equal amounts. If $x$ is the equilibrium concentration of H$_3$O$^+$$(aq)$ ion, then

\[
\text{Init. Conc. (M)} \quad \begin{array}{c}
0.05 \\
- \end{array} \quad \begin{array}{c}
0.05 \\
\text{small}
\end{array} \\
\text{Change in Conc. (M)} \quad \begin{array}{c}
-x \\
- \\
+x \quad +x
\end{array} \\
\text{Equil. Conc. (M)} \quad \begin{array}{c}
0.05 - x \\
- \\
0.05 + x \quad x
\end{array}
\]

\[
K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = \frac{(0.05 + x)x}{(0.05 - x)}
\]

Assume that $x$ is small compared to 0.05. Then 0.05 in the numerator cancels out with that in the denominator

\[
[H_3O^+] = K_a \quad \text{so that} \quad pH = pK_a = 9.25
\]

Clearly $x$ is less than $10^{-9}$, so the assumption was justified.

Tip: The pH equals the $pK_a$ of the weak acid. This is a general result in buffer solutions in which the acid and conjugate base concentrations are equal (and not extremely low).

**Problem 57.** A buffer solution is prepared from 100 mL of 0.1 M acetic acid and 25 mL of 0.1 M sodium hydroxide. The $pK_a$ of acetic acid is 4.75 while the $pK_b$ of sodium hydroxide is 0.2. What is the pH of the resulting solution? Is the acid-base reaction of NaOH a competing equilibrium that we must include in our calculation? Explain why it is or why it is not.

**Solution.** Total volume is 125 mL. We have 0.01 mol acetic acid (HA) and 0.0025 mol NaOH. NaOH reacts with acetic acid to yield acetate (A$^-$), so the initial conditions are 0.0075 mol acetic acid and 0.0025 mol acetate

<table>
<thead>
<tr>
<th></th>
<th>HA</th>
<th>A$^-$</th>
<th>H$_3$O$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.060</td>
<td>0.020</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>E</td>
<td>0.060 - x</td>
<td>0.020 + x</td>
<td>+x</td>
</tr>
</tbody>
</table>

which gives:

\[
\frac{(0.020 + x)(x)}{0.060 - x} = K_a = 10^{-4.75}.
\]

The quadratic equation, $0 = x^2 + 0.02x - 1.1 \times 10^{-6}$, is solved and we keep $x = 5.3 \times 10^{-5}$. Therefore, $pH = -\log_{10}(5.3 \times 10^{-5}) = 4.3$. The association
of NaOH does not compete with the dissociation of acetic acid, as NaOH is a strong base with a low $pK_b$ of 0.2. This means that NaOH effectively dissociates entirely in our solution, so we do not need to consider the dissociation/association reaction in our calculation of the buffer solution pH.

**Problem 58.** For the following Lewis acid-base reactions below draw the Lewis structures (of each reactant and product), relevant chemical bonds, and indicate with arrows which bonds are formed or broken and which unpaired electrons are shared to form the adduct. (Note: when forming the adduct, bonds can be broken and formed apart from the sharing of electrons.) Identify the acid and the base:

(a) $\text{B(OH)}_3 + \text{OH}^- \rightarrow [\text{B(OH)}_4^-]

(b) $\text{N}_2\text{H}_4 + \text{H}_3\text{O}^+ \rightarrow \text{N}_2\text{H}_5^+ + \text{H}_2\text{O}$

(c) $(\text{C}_2\text{H}_5)_2\text{O} + \text{BF}_3 \rightarrow (\text{C}_2\text{H}_5)_2\text{OBF}_3$

(d) $\text{H}_2\text{O} + \text{SO}_2 \rightarrow \text{H}_2\text{SO}_3$

**Solution.** The chemical bonding changes as follows:

**Problem 59.** Consider the Lewis acid-base reaction:

$\text{Ag}^+(aq) + 2 \text{CN}^-(aq) \rightarrow [\text{Ag(CN)}_2^-](aq)$

(a) Identify the acid and base among reactants.

(b) Many Lewis acid-base “titration reactions”, such as this one, result in the formation of complex (polyatomic) ions. The equilibrium constant for this reaction is called a formation constant of complexes ($K_f$) (sometimes called a stability constant). Define the formation constant of $[\text{Ag(CN)}_2^-]$.

(c) Calculate the silver ion concentration, $[\text{Ag}^+]$, of a solution that initially is 0.10 M with respect to $[\text{Ag(CN)}_2^-]$. The formation constant of $[\text{Ag(CN)}_2^-]$ is $5.3 \times 10^{18}$ at room temperature.

**Solution.** (a) Lewis acid: $\text{Ag}^+$. Lewis base: $\text{CN}^-$
(b) \( K_f = \frac{[\text{Ag(CN)}_2^-]}{[\text{Ag}^+][\text{CN}^-]^2} \)

(c) The initial concentration of \([\text{Ag(CN)}_2^-]\) is 0.10 mol/L

\[
\begin{array}{c|c|c|c}
& \text{Ag}^+ (aq) & 2 \text{CN}^- (aq) & \Rightarrow [\text{Ag(CN)}_2^-] (aq) \\
\text{I (M)} & 0 & 0 & 0.10 \\
\text{C (M)} & x & 2x & -x \\
\text{E (M)} & x & 2x & 0.10-x \\
\end{array}
\]

\( K_f = 5.3 \times 10^{18} = \frac{[\text{Ag(CN)}_2^-]}{[\text{Ag}^+][\text{CN}^-]^2} = \frac{0.10-x}{x \times (2x)^2} \)

Both \( Q \) and \( K_f \) are much larger than 1, so let us assume that the changes in concentrations needed to reach equilibrium are small. Thus \( 0.10-x \) is approximated as 0.10:

\[
x^3 = \frac{0.10}{4 \times 5.3 \times 10^{18}} = 4.717 \times 10^{-21}
\]

\[x = [\text{Ag}^+] = 1.68 \times 10^{-7} \text{ mol/L} \]

**Problem 60.** Consider the following reaction with an equilibrium constant of 322:

\[\text{Br}_2(g) + \text{Cl}_2(g) \rightleftharpoons \text{BrCl}(g)\]

Determine the final partial pressure of each compound if a vessel of 0.0500 atm \( \text{Br}_2 \) and 0.0400 atm \( \text{Cl}_2 \) is left to achieve equilibrium.

**Solution.**

\[
\frac{(P_{\text{BrCl}})^2}{P_{\text{Br}_2}P_{\text{Cl}_2}} = K \\
(0.0500 \text{ atm} - x \text{ atm})(0.0400 \text{ atm} - x \text{ atm}) = 322 \\
318x^2 - 28.98x + 0.644 = 0 \\
0.0768 \text{ atm BrCl} \\
0.0116 \text{ atm Br}_2 \\
0.0016 \text{ atm Cl}_2
\]

**Problem 61.** Take the solubility product of \( \text{CaF}_2 \) to be \( 4.0 \times 10^{-11} \) and use it to find the solubility of the compound. What can be said about trying to dissolve \( \text{CaF}_2 \) in water?

**Solution.**

\( \text{CaF}_2 \rightleftharpoons \text{Ca}^{2+} + 2 \text{F}^- \)

\[
K_{sp} = [\text{Ca}^{2+}] [\text{F}^-]^2 = [\text{Ca}^{2+}] [2\text{Ca}^{2+}]^2 = 4.0 \times 10^{-1} \\
[\text{Ca}^{2+}] = [\text{CaF}_2] = 0.000215 \text{ M} \\
[\text{F}^-] = 0.000430 \text{ M} \\
\frac{0.000215 \text{ mol}}{L} \times \frac{78.07 \text{ g}}{\text{mol}} = 0.0169 \text{ g.L}^{-1}
\]

Based on the value computed above, we would expect \( \text{CaF}_2 \) to not be very soluble in water.
Chapter 6

Thermodynamics

Thermodynamics provides a framework to account for exchanges (transfers) of energy among different systems or subsets of a given system. Energy transfers can occur in the form of heat exchanges or different types of work. Thermodynamics is a topic that crosses many disciplines (physics, chemistry, engineering, life sciences, etc.). This chapter only provides a brief introduction to this vast topic. To learn more, the reader should consult the more authoritative treatises by Callen [15], Kittel [16], Moran & Shapiro [17], Atkins [14], McQuarrie & Simon [5], Ben-Amotz [18], Sonntag & Borgnakke [19] and Honig [20].

In thermodynamics the universe is divided into two parts: the system and its surroundings. We must then specify whether or not particles and heat are allowed to exchange between the system and its surroundings (Fig. 6.1).

6.1. Heat vs Work

Energy is the capacity to do work. A change in energy can be thought of as the sum of heat and work (done, say, by the system on its surroundings, or vice-versa). Heating is the transfer of energy that makes use of disorderly molecular motion. See Fig. 6.2.

Work is the transfer of energy that makes use of organized motion (organized, as opposed to random molecular motions). See Fig. 6.3.

An exothermic process is one that releases energy as heat to the surroundings.\(^1\) All combustion reactions are exothermic. An endothermic process

\(^{1}\)Heat transfer is a process. Heat is not viewed as an entity. It is instead a process: the transfer of energy as a result of temperature difference. Similarly to a diffusion process which is a flux \(J_t\)
Systems

Open

Matter

Energy

Closed

Matter

Energy

Isolated

Matter

Energy

Figure 6.1. If matter can be transferred to the system, it is called **open**, otherwise it is **closed**. An **isolated system** cannot exchange mass or heat with its surroundings. Walls of the system that allow transfer of heat are called **diathermic** whereas those that don’t are called **adiabatic**. **Energy** is the capacity of a system to do work. Energy can also be transferred as **heat**.

![Figure 6.2.](image)

Figure 6.2. When energy is transferred to the surroundings as heat, the transfer stimulates random motion of the atoms in the surroundings. Transfer of energy from the surroundings to the system makes use of random motion (thermal motion) in the surroundings.

in the $i$-th chemical species driven by gradients in the chemical potential of that species,

$$J_i(\text{particle flux}) = -D \frac{d}{dx} \mu_i(x)$$

Fick’s law of diffusion

($D$ is a diffusion coefficient), heat transfer is a flux driven by gradients in temperature

$$q(\text{heat flux}) = -k \frac{d}{dx} T(x)$$

Fourier’s law
6.2. Heat Capacity

Suppose we transfer some amount of heat $\Delta Q$ to a system, its temperature will increase by an amount $\Delta T$. If you transfer double the heat $2\Delta Q$, the temperature will increase by double the amount, $2\Delta T$. The proportionality “constant” is the heat capacity, $C(T)$, which itself depends on temperature:

$$q = \Delta Q = C(T)\Delta T.$$

where $k$ is the coefficient of thermal conductivity (in general, $k = k(T)$ is temperature dependent). At the interface between a hot and a cold body, the temperature gradient is large, causing a transfer of heat from the hot to the cold body. The direction hot $\rightarrow$ cold is enforced by the minus sign.
Heat capacity is thus a kind of “resistance” or reluctance that a material shows towards raising its temperature in response to receiving heat from its surroundings. Suppose that we transfer 1 J of heat to the material. The smaller the heat capacity, the larger the temperature rise. The larger the heat capacity, the smaller the temperature rise. A closely related quantity is thermal conductivity; the latter is proportional to the heat capacity (see Section 6.2.1).

In thermodynamics we make the distinction between the heat capacity at constant pressure \( (C_P) \) and at constant volume \( (C_V) \) using a subscript \( P \) or \( V \), respectively. This refers to the kind of experiment used to measure the heat capacity. As we will see later, those two values are different.

The specific heat capacity (not to be confused with concentration), denoted by a lowercase \( c \), is the heat capacity per unit material.

- **Per unit mass**: \( c = C/m \), where \( m \) is the mass of the body. SI Units: [J/kg.K]
- **Per mole**: \( c = C/n \), where \( n \) is the number of moles. This is called the molar heat capacity. SI Units: [J/mol.K]

Why should we care about specific heat capacity? Because without it, heat capacity is an extensive property, i.e. it depends on the amount of material used. \( q = C\Delta T \) says that for a 1°C increase, \( q \) will double if the amount of material doubles (because a body twice as large requires twice the amount of heat transfer to cause this 1°C increase), therefore \( C \) should double. Because of this dependence, it is impossible to tabulate values of \( C \). Dividing by the amount of material \( (n \) or \( m \) renders the heat capacity independent of the amount of material. It is then an intensive property which is intrinsic to the material type, not its quantity.

Another notation you may see for the definition of heat capacity is in terms of infinitesimal quantities:

\[
C(T) = \frac{\delta q}{dT},
\]

or the corresponding molar heat capacity

\[
e(T) \equiv \frac{1}{n} C(T) = \frac{1}{n} \frac{\delta q}{dT}.
\]

While this notation appears awkward, it is sometimes useful conceptually.

This notation appears to suggest that heat capacity is some type of derivative \( (\delta q/dT) \) of an extensive parameter (heat flow) with respect to an intensive property (temperature). These are, however, formal definitions. In practice, measured heat capacities are evaluated as derivatives with respect to temperature while holding some other variables constant \( (P \) or \( V \).
This is done as such, because the measured heat flow will depend on which parameters are held constant. \( C_V \) is the most convenient quantity for gases and theoreticians. \( C_P \) is the most common tabulated form.

### 6.2.1. Thermal Conductivity

If the temperature is non-uniform there will be a transfer of heat until a steady-state is reached. Similar to the case of particle diffusion, this transfer of heat acts to make the temperature more spatially uniform. The rate of heat loss per unit area \( (A) \) is related to the temperature gradient according to:

\[
\frac{1}{A} \frac{\Delta Q}{\Delta t} = -\kappa \frac{\Delta T}{\Delta x},
\]

where \( \kappa \) is the thermal conductivity.

For an ideal gas the heat transfer rate is proportional to the average molecular velocity \( (\bar{v}) \), the mean free path \( (\lambda) \) of the energy carrier (average distance traveled between consecutive scattering events, on the order of a few nm to \( \mu m \)), and the specific (volume) heat capacity of the gas \([c(T)]\).

\[
\kappa = \frac{1}{3} \bar{v} \lambda c(T).
\]

A related quantity is the thermal diffusivity, \( \alpha \) (units: \( m^2/s \)):

\[
\alpha = \frac{\kappa}{c(T)}.
\]

The units of \( c(T) \) are \( J/m^3.K \). Those of \( v \) are \( m/s \). This implies that \( \kappa \) has units of \( J/m.s.K \) or \( W/m.K \). The unit \( J/s \) is also known as the Watt \( (W) \). Inspection of Eq. (6.1) reveals that the units on the right hand side must be \( (J/m.s.K) \times (K/m) \) or \( W/m^2 \). For equality to hold, these must also be the units of the left hand side. Indeed, \( Q \) has units of \( J \), and therefore, \( \Delta Q/(A\Delta t) \) has units of \( W/m^2 \), as it should. This equation holds for each energy carrier (e.g. electron, phonon). For electrons, \( \bar{v} \) is on the order of \( 10^6 \) m/s; for phonons it is approximately \( 10^3 \) m/s.

For solids, there are two contributions to the heat capacity: phonons and conduction elections. (For non-metallic solids, only phonons contribute.) Atoms vibrating more energetically at one part of a solid transfer that energy to less energetic neighboring atoms. Phonons are propagating lattice waves that are quantized. For metals, the thermal conductivity is quite high,

\[\text{Using derivatives, this is:}\]

\[
\frac{1}{A} \frac{dQ}{dt} = -\kappa \frac{dT}{dx},
\]

where the derivative on the left, of a function \( Q(x, t) \), with respect to \( t \) holds \( x \) constant. Likewise, the derivative on the right hand side is with respect to \( x \) while holding \( t \) constant. This is normally denoted using partial derivatives:

\[
\frac{1}{A} \frac{\partial Q}{\partial t} = -\kappa \frac{\partial T}{\partial x}.
\]
and those metals which are the best electrical conductors are also the best thermal conductors.

6.2.2. Heat Capacity of a Solid: Dulong-Petit Law. In 1819, Dulong and Petit measured the specific heats\(^3\) of several solids and found that the values for the various materials differed considerably. However, their heat capacity was close to 6 cal/K/mol (Table 6.1) or 25 J/mol.K.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic weight (g/mol)</th>
<th>Specific heat (cal/K/g)</th>
<th>Heat capacity (cal/K/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>6.9</td>
<td>0.92</td>
<td>6.3</td>
</tr>
<tr>
<td>Beryllium</td>
<td>9.0</td>
<td>0.39</td>
<td>3.5</td>
</tr>
<tr>
<td>Magnesium</td>
<td>24.3</td>
<td>0.25</td>
<td>6.1</td>
</tr>
<tr>
<td>Carbon (diamond)</td>
<td>12.0</td>
<td>0.12</td>
<td>1.4</td>
</tr>
<tr>
<td>Aluminum</td>
<td>27.0</td>
<td>0.21</td>
<td>5.7</td>
</tr>
<tr>
<td>Iron</td>
<td>55.8</td>
<td>0.11</td>
<td>6.1</td>
</tr>
<tr>
<td>Silver</td>
<td>107.9</td>
<td>0.056</td>
<td>6.0</td>
</tr>
<tr>
<td>Lead</td>
<td>207.2</td>
<td>0.031</td>
<td>6.4</td>
</tr>
<tr>
<td>Mercury</td>
<td>200.6</td>
<td>0.033</td>
<td>6.6</td>
</tr>
</tbody>
</table>

Table 6.1. Heat capacities of some solid elements. Dulong and Petit stated their law as: specific heat (cal/K.g) \(\times\) atomic weight (g/mol) \(\approx\) 6 (cal/K.mol). 1 calorie equals 4.18 Joules. Source: [21]

Heat capacity arises because of the many different ways (degrees of freedom) in which molecules can absorb energy. The Law of Dulong and Petit is:

**Law of Dulong and Petit.** The molar specific heats of most solids at room temperature and above are nearly the same value. (This value is \(3R \approx 25\) J/mol.K.)

We note that the Dulong-Petit heat capacity is independent of temperature (in the high-\(T\) limit only). See Fig. 6.4.

\(^3\)Here, specific heat is the number of calories (or Joules) required to raise one gram (or kilogram) of material one degree celsius. The term heat capacity often refers to the heat required to raise one mole of material one degree celsius. In older chemistry books (e.g. [21]), \“heat capacity\” refers to \“molar heat capacity\”, whereas \“specific heat capacity\”, is also called \“specific heat\”, and that is the heat capacity per unit mass of a material. In modern chemistry textbooks (e.g. [1]), these two notions are called \“specific heat capacity\”, whereas \“heat capacity\” refers to the total heat capacity (of the entire volume of material). The latter, which is an extensive property, is less useful because such values cannot be tabulated for different materials, as the numerical value depends on the amount of material. In this section only, we use the old terminology.
6.2. Heat Capacity

Figure 6.4. The dimensionless heat capacity divided by three, as a function of temperature as predicted by the Debye and Einstein models. The horizontal axis is the temperature divided by the Debye temperature. The dimensionless heat capacity is zero at absolute zero, and rises to a value of three as the temperature becomes much larger than the Debye temperature. The red line corresponds to the classical limit of the Dulong-Petit law.

Units of heat transfer. The connection between heat transfer (SI units: Joule) and temperature change (units: Kelvin) should not be so unfamiliar to you at this point. We have seen in the case of gases that the average kinetic energy per molecule, \( K \equiv \frac{1}{2}mv^2 = \frac{3}{2}k_B T \), established a connection between temperature and energy. Thus, a change in the kinetic energy \( \Delta K \) is proportional to a change in temperature \( \Delta T \). The two quantities (energy and temperature) were related by the Boltzmann constant. \( k_B \) and \( C \) have the same SI units (J/K).

6.2.3. Heat Capacity of Ideal Gas. Consider an ideal monatomic gas held at constant volume \( V \). The internal energy of an ideal monatomic gas is purely kinetic. There is no potential energy. From

\[
K \equiv \frac{1}{2}mv^2 = \frac{3}{2}k_B T
\]

or \((3/2)RT\) for 1 mole, we can easily derive the heat capacity of an ideal gas by considering the change in temperature that accompanies an infinitesimal change in internal energy, which for an ideal gas is just the kinetic energy \( (U = K) \).

The heat capacity is obtained by taking the derivative of energy with respect to temperature. For 1 mole of substance, the heat capacity is,

\[
c_V = \frac{3}{2}R \text{ ideal gas}
\]
Here, \( c_V \) denotes the specific (molar) heat capacity at constant volume. (The volume is constant because the gas is contained in a volume \( V \).)

It can be shown that at constant pressure, the heat capacity of the ideal gas is:

\[
c_P = \frac{5}{2} R. \quad \text{ideal gas}
\]

6.2.4. Internal Energy. The internal energy of a system, \( U \), is the sum of total kinetic and potential energy of the molecules in the system. Potential energy includes the energy stored in the chemical bonds. The internal energy is a state function, meaning that it depends only on the state of the system, not on the history or path taken to reach this state.

State function: To describe a system completely, we must indicate its temperature, pressure and amounts of substances present. When we have done this, we have specified the state of the system. Any property that has a unique value for a specified state of a system is said to be a state function. Note: Any function that depends on the history of the system — rather than its current state — is not a state function.

For a state function, we can express changes in the state function in terms of the final state minus the initial state, \( \Delta U = U_f - U_i \), without any mention of the intermediate states taken to reach the final state from the initial state.

Change in internal energy: Because \( U \) is a state function, so is \( \Delta U = U_f - U_i \). Suppose that a system goes from state 1 to state 2 and back to state 1. Its change of internal energy is \( U_2 - U_1 \) followed by \( U_1 - U_2 \):

\[
\Delta U = \Delta U_{1\to2} + \Delta U_{2\to1} = (U_2 - U_1) + (U_1 - U_2) = 0.
\]

This says that the internal energy returns to its initial value, which it must do, since it is a function of state.

Internal energy is an extensive property of the system (i.e. if we double the size of the system, its internal energy doubles). The SI units of \( U \) are Joule (J). However, the internal energy per mole, \( U/n \), is an intensive property of the system. The SI units of \( U/n \) are J/mol.

The change in internal energy \( U \) will contain terms which can be classified as heat transfer or work:

\[
\Delta U \equiv \text{change in } U = \text{(heat)} + \text{(work)} \equiv q + w.
\]

In general, we do not need to enumerate all contributions to the energy because thermodynamics is only concerned about changes in the energy.
6.2. Heat Capacity

(e.g. $\Delta U$ or $dU$), in which case, only the terms that are involved in the change are needed.

In thermodynamics we are interested in $dU$, the total differential of the function $U$. $dU$ is the sum of “infinitesimal” heat transfer, $\delta Q$, and “infinitesimal” work done on the system $\delta W$:

$\textbf{dU is the sum of infinitesimal heat transfer and work}$

The change in internal energy is:

\[ dU = \delta Q + \delta W \]

The total change in heat and work is obtained by integrating the infinitesimal expressions, $q \equiv \Delta Q = \int \delta Q$ and $w \equiv \Delta W = \int \delta W$. For $\Delta U$, we have:

\[ \Delta U = \int dU = \int \delta Q + \int \delta W = \Delta Q + \Delta W \equiv q + w. \]

Some books write $q$ for heat transfer. We can write $\Delta Q$ to emphasize that it is a change. We will see later that the first law of thermodynamics states that $U$ is a state function, or equivalently, that $dU$ is an exact differential. This implies that $\Delta U = \int_{f}^{i} dU = U_{f} - U_{i}$ only depends on the state of the system (hence the name “state function”), not on the history of changes it undergoes between $i$ and $f$.

6.2.5. Compression Work (Hydrostatic Pressure Work). Suppose that we compress a gas using a piston by applying an external force (pressure), as shown in Fig. 6.5.

![Figure 6.5. Work accomplished when compressing a gas using a piston.](image)
The *system* is the gas inside the chamber. The *surroundings* is everything else external to it. The external pressure $P_{\text{ext}}$ originates from the surroundings. The pressure is force per unit area:

$$P_{\text{ext}} = \frac{F_{\text{ext}}}{A}.$$ 

### 6.2.6. Work Done by the Surroundings to the System

We also know that mechanical work, $\Delta W$, is obtained by integrating the force over distance (let $z$ denote the vertical direction which corresponds to the direction of the displacement $\Delta h$):

$$\Delta W = \int_{z_i}^{z_f} P_{\text{ext}} A \, dz = \int_{V_i}^{V_f} P_{\text{ext}} \, dV.$$ 

As a matter of convention in chemistry we put a negative sign to denote work done on the system by the surroundings. Engineers use the opposite convention (work done by the system). For example, if a gas is compressed (work done on the system), $\Delta V < 0$, leading to positive values of $W$ in the definition below:

<table>
<thead>
<tr>
<th>Compression work</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\Delta W = -\int_{V_i}^{V_f} P_{\text{ext}} , dV$ or $\delta W = -P_{\text{ext}} , dV$</td>
</tr>
</tbody>
</table>

If $P_{\text{ext}}$ remains constant during this change (as would be the case if the surroundings were atmospheric pressure, which remains essentially constant), we can pull it out of the integral and the work done on the system is the product of the external pressure and the change in the volume of the gas:

$$\Delta W = -P_{\text{ext}} \Delta V.$$ 

In general, depending on what you consider to be the surroundings, $P_{\text{ext}}$ may not necessarily be constant, in which case you need to use $\delta W = -P_{\text{ext}} \, dV$ and integrate $\Delta W = \int \delta W$ along the path of the change.

### 6.2.7. RECALL: Work = Force $\times$ Distance

In classical mechanics, we learn that force times distance is the mechanical work, i.e.

$$w = \int_{i}^{f} \vec{F} \cdot d\vec{x},$$ 

where $i$ and $f$ denote the initial and final points, respectively. If the force is in the $x$ direction, $\vec{F} = F\hat{x}$, we have

$$w = \int_{i}^{f} F \, dx.$$
6.3. Expansion ($P - V$) Work

The integral is taken along some path with initial point $i$ and final point $f$. The result is generally path-dependent, unless $F dx$ is the derivative of some function (say) $g$, i.e. $F dx = dg$, in which case it only depends on the initial and final points since $\int_i^f dg = g(f) - g(i)$ (fundamental theorem of calculus).

If $F$ is constant along the path, then the work is simply $F(x_f - x_i) = F \Delta x$.

Now let’s return to the example from the previous section, where a gas is compressed by a piston. Starting from $P_{ext} = F_{ext}/A$, substituting $A = \Delta V/\Delta h$ yields $P_{ext} = F_{ext}\Delta h/\Delta V$. Thus, $P_{ext}\Delta V = F_{ext}\Delta h$. For infinitesimal changes, $P_{ext}dV = F_{ext}dh$. Since the work done by the surroundings on the system is $-P_{ext}dV$,

$$-dw = P_{ext}dV = F_{ext}dh \quad P_{ext} \frac{dV}{dh} = F_{ext}.$$  

The external force (force exerted on the system by its surroundings) is seen to be related to the spatial derivative of the work $F_{ext} = -\frac{dw}{dx}$. If $h$ is the $x$ direction:

$$F_{ext} = -\frac{dw}{dx}.$$  

This is a general principle. The force is generally obtained by taking minus the gradient (spatial derivative) of the (potential) energy:

$$F_{ext} = -\frac{dU}{dx}.$$  

(The kinetic energy does not appear because it rarely depends on position.)

6.3. Expansion ($P - V$) Work

In chemistry, $P - V$ work is important because it describes the behavior of gases and the work arising from a change in volume. This includes the work done by a gas as it expands and pushes against the atmosphere. Chemical reactions generate gases. The $P - V$ work here is the work that must be done to make room for the gas it has produced. $P - V$ work also describes compression (negative changes in volume).

6.3.1. Free Expansion. Suppose that a gas is allowed to expand in a vacuum. Since there is zero opposing force, $P_{ext} = 0$ and the work is zero $\Delta W = 0$.

6.3.2. Expansion Against Constant Pressure. Suppose that a gas inside a piston expands, and the piston is pressing against the atmosphere. The latter corresponds to a constant pressure $P_{ext}$. Thus, it is constant and
6. Thermodynamics

we can take it outside the integral:

\[ \Delta W = - \int_{V_i}^{V_f} P_{ext} dV = -P_{ext} \int_{V_i}^{V_f} dV = -P_{ext} (V_f - V_i) = -P_{ext} \Delta V. \]

6.3.3. Reversible Process. Reversible processes are important because they represent the most efficient processes. They are idealizations (no process in nature can be perfectly reversible) which enable us to do useful calculations. A process is fully reversible if both system and surroundings will return exactly to their initial states if the process is run in reverse. Less efficient processes do not have this property, and so are referred to as irreversible. A process is said to be reversible if it remains arbitrarily close to equilibrium with its surroundings at all times.

When a system is arbitrarily close to equilibrium with its surroundings, we mean close in the sense of infinitesimals. If \( T, P, \mu \) are the temperature, pressure and chemical potential of the system and \( T_{ext}, P_{ext}, \mu_{ext} \) are those of the surroundings, then arbitrarily close means that they are close in the sense of infinitesimals:

\[
T_{ext} = T + dT \\
P_{ext} = P + dP \\
\mu_{ext} = \mu + d\mu
\]

According to the rules of calculus, we can handle expressions containing infinitesimals such as \( dT, dP \) and \( d\mu \). However, second-order quantities such as \( (dT)^2 \) or \( (dT)(dP) \) should be neglected because the square of an infinitesimally small quantity is so small that it is effectively zero.\(^4\) If you don’t want to deal with infinitesimals, it is usually safe to make the following substitutions (replace external pressure by internal pressure, etc.):

\[ T_{ext} \approx T, \quad P_{ext} \approx P, \quad \mu_{ext} \approx \mu. \]

\(^4\)From calculus, we know that

\[ df = \frac{df}{dx} dx. \quad (1) \]

However, it is also the case that:

\[ df(x) = f(x + dx) - f(x). \]

Taylor-expanding the first term,

\[ df = \left[ f(x) + (dx) \frac{df}{dx} + \frac{(dx)^2}{2} \frac{d^2f}{dx^2} + O((dx)^3) \right] - f(x) \quad (2) \]

But for (1) and (2) to be equal we are “forced” to drop terms of order \((dx)^2\) or higher. Thus, when handling differential expressions we keep terms of order \(O(dx)\) only. The higher order infinitesimal terms are negligible (too small).
6.3. Expansion ($P - V$) Work

6.3.4. Reversible Compression. If the mechanical compression work done by the surroundings on a system

\[ \delta W = -P_{\text{ext}} \, dV \]

is reversible, this means we can write $P_{\text{ext}} = P + dP$,

\[ \delta W = -(P + dP) \, dV. \]

Noting that the product of infinitesimals $dP \, dV$ is much smaller than the term $P \, dV$, we neglect $dP \, dV$:

\[ \delta W = -P \, dV. \]

The reversible work done on the system is then:

\[ \Delta W^{\text{rev}} = \int \delta W = -\int_{V_i}^{V_f} (P + dP) \, dV = -\int_{V_i}^{V_f} P \, dV. \]

What did we accomplish by changing $\delta W = -P_{\text{ext}} \, dV$ to $\delta W = -P \, dV$? The difference is significant because this allows us to replace $P_{\text{ext}}$ by $P$.

In many experiments $P$ is known, because the gas mixture is known, and obtained from the equation of state, whereas $P_{\text{ext}}$ is often unknown (e.g., the surroundings may be out of your control).

Indeed, when computing the work done on the gas, $-\int F \, dx = -\int P_{\text{ext}} \, dV$, the force $F$ must be the gravitational force acting downward on all of the mass supported by the gas at pressure $P$. This mass includes that of the piston, the piston rod, the pan, any weight placed on the piston (to compress the gas), and the atmosphere above the piston. In the case of a reversible process this force $F$ is never more than minutely out of balance with the force exerted upward on the piston face by the gas and given by the product of pressure and piston area. Thus, for all practical purposes, $F = PA$ for the reversible process. The volume change of the gas (system) is $dV = A \, dx$; thus $dx = dV/A$ and the work done on the gas is

\[ \Delta W = -\int F \, dx = -\int PA \, \frac{dV}{A} = -\int P \, dV. \]

For a reversible process, we replace $P_{\text{ext}}$ by $P$, which in this equation means we set $F = PA$. This allows us to calculate the work using the properties of the system rather than its surroundings. It is generally difficult to know the exact properties of the surroundings in such calculations, especially when the gas expansion is rapid. In a reversible process, the forces are never more than differentially out of balance (e.g., $P_{\text{ext}} = P + dP$).

For irreversible processes this substitution is not possible because we may not have thermal equilibrium. For example, suppose that gas expansion takes place: when a finite weight is removed from the piston in the processes described, the forces of gravity acting downward is overbalanced by the gas
pressure acting upward by a finite amount (i.e. $P_{ext} = P - \Delta P$, where $\Delta P$ is not an infinitesimal), and $F$ does not equal $PA$ again until a new equilibrium position of the piston is reached. (Similarly, for gas compression we have $P_{ext} = P + \Delta P$, and the internal gas pressure is overcome by the external pressure from the weight pressing downward.) Thus $PA$ cannot be substituted in the integral $\int F \, dx$, and it is not possible to calculate the work from a knowledge of the properties of the system (the gas). The integral $\int F \, dx$ is given by $\int P \, dV$ only when the process is reversible; that is,

$$\Delta W^{rev} = -\int_{V_1}^{V_2} P \, dV$$

This work for an expansion process is the maximum work which the system can produce.

**6.3.5. Why are Reversible Processes Important?** The reversible process is unique, and important in thermodynamics, because it represents the limit of what is possible in the real world. We cannot imagine anything better than a reversible process. Such processes lend themselves to exact mathematical analysis, which is not the case for other processes. In thermodynamics, our choice is often to do calculations for reversible processes, or to do no calculations at all. The reasons for this is that reversible processes are those for which the forces causing change are almost exactly in balance. Thus the states through which the system passes during a reversible process are for all practical purposes equilibrium states, or more precisely are never removed more than differentially from equilibrium states (e.g. $P_{ext} = P + dP$).

**6.3.6. Example: (Reversible Compression of an Ideal Gas).** To illustrate the application of this result, let’s look at a specific example. The easiest case is an isothermal compression ($T = const$) for an ideal gas ($P = nRT/V$):

$$\Delta W^{rev} = -\int_{V_1}^{V_2} P \, dV = -\int_{V_1}^{V_2} \frac{nRT}{V} \, dV = -nRT \int_{V_1}^{V_2} \frac{dV}{V} = nRT \log(V_1/V_2).$$

**6.3.7. Example: (Reversible Compression of VDW Gas).** A slightly more complicated example is the van der Waals gas. Suppose that $n$ moles of gas are expanded from $V_1$ to $V_2$ reversibly and isothermally. Calculate $\Delta W$ for a Van der Waals gas.

**Solution:**

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$
6.3. Expansion \((P - V)\) Work

\[
\Delta W^{rev} = - \int_{V_1}^{V_2} P \, dV = - \int_{V_1}^{V_2} \frac{nRT}{V - nb} \, dV + \int_{V_1}^{V_2} \frac{n^2a}{V^2} \, dV \\
= - nRT \int_{V_1}^{V_2} \frac{dV}{V - nb} + n^2a \int_{V_1}^{V_2} \frac{dV}{V^2} \\
= - nRT \log \frac{V_2 - nb}{V_1 - nb} + n^2a \left( \frac{1}{V_1} - \frac{1}{V_2} \right).
\]

6.3.8. Example: (w/ numbers) Calculate the work done when 50 g of iron reacts with hydrochloric acid to produce \(\text{FeCl}_2(aq)\) and hydrogen in (a) a closed vessel of fixed volume, (b) an open beaker at 25°C.

Solution: For (a) the volume cannot change, so the expansion work done is zero \((\Delta W = 0)\). For (b) the gas pushes against the atmosphere, \(\Delta W = -P_{ext}\Delta V\). After production of gas, the volume changed by \(\Delta V = V_f - V_i \approx V_f = nRT/P_{ext}\) (the initial volume is so small compared to \(V_f\) that it can be neglected), where \(n\) is the amount of \(\text{H}_2\) produced. Therefore,

\[
\Delta W = -P_{ext}\Delta V = -P_{ext} \frac{nRT}{P_{ext}} = -nRT.
\]

Because the reaction is \(\text{Fe}(s) + 2\text{HCl}(aq) \rightarrow \text{FeCl}_2(aq) + \text{H}_2(g)\), we know that 1 mol \(\text{H}_2\) is generated when 1 mol \(\text{Fe}\) is consumed, and \(n\) can be taken as the amount of \(\text{Fe}\) atoms that react. The molar mass of \(\text{Fe}\) is 55.85 g/mol. Then,

\[
\Delta W \approx - \frac{50 \text{ g}}{55.85 \text{ g/mol}} \times (8.3145 \text{ J/K.mol}) \times (298 \text{ K}) \approx -2.2 \text{ kJ}.
\]

6.3.9. Path-Dependence of Work. Work is not a state function, meaning that its numerical value depends on the path taken to go from one state to another. As an example, take a gas in a cylinder with piston. Suppose that state 1 correspond to \(P_{ext} = 2\text{ atm}\) and state 2 is \(P_{ext} = 1\text{ atm}\). Both states are at room temperature. We allow this gas to expand from state 1 to state 2 under two different circumstances (paths). In both cases, we expect the volume of an ideal gas to double (say, from 1 L to 2 L). For this to work, we must have (state 1):

\[
n = \frac{PV}{RT} = \frac{(2 \text{ atm})(1 \text{ L})}{(0.082 \text{ L.atm/K/mol})(298 \text{ K})} = 0.08185 \text{ mol}
\]

of gas. The first path allows the volume to expand reversibly:

\[
\Delta W^{rev} = nRT \log(V_1/V_2) = (0.08185 \text{ mol})(8.31 \text{ J/K/mol})(298 \text{ K}) \log(1 \text{ L}/2 \text{ L})
= - 140.5 \text{ J}.
\]

In the second path, we carry out the expansion at constant pressure, in two discrete steps: during step 1, the volume goes from 1 L to 1.5 L at 2 atm;
during step 2 the volume expands from 1.5 L to 2 L under a pressure of 1 atm:

\[ \Delta W = \Delta W_{1\rightarrow i} + \Delta W_{i\rightarrow 2} \]
\[ = -(2 \text{ atm})(0.5 \text{ L}) - (1 \text{ atm})(0.5 \text{ L}) \]
\[ = -1.5 \text{ L.atm} = -152.0 \text{ J}, \]

since 101.33 J converts to 1 L.atm. Here, \( i \) denotes the intermediate state. As you can see, the work is different in each case. Thus, the work depends on path. A negative value of \( \Delta W \) means that the work is done by the system on its surroundings. This work done by the system is lowest in the case of reversible expansion. This behavior is typical of reversible processes: they are the most efficient paths from state 1 to state 2, spending the least amounts of energy to accomplish any given change.

6.3.10. Why the Infinitesimals? In the previous sections we have been writing down the work in terms of infinitesimal changes \( \delta W \). The total work is obtained by integrating:

\[ \Delta W = \int \delta W, \]

where \( \delta W \) represents infinitesimal work. For example, the \( P - V \) work term is

\[ \delta W = -P_{\text{ext}} \, dV, \]

or

\[ \delta W = -P \, dV, \]

in the case of reversible \( P - V \) work. The total work is

\[ \Delta W = \int \delta W = - \int P(V) dV, \]

where \( P(V) \) denotes pressure expressed in term of volume, since the variable of integration is \( V \). This is done using an equation of state.

You may ask, why are we working with infinitesimal expressions? This is because the intensive quantity (e.g., the pressure) may change along the path of integration. To get the total work \( \Delta W \), we may need to express the intensive quantity in terms of the extensive quantity. This can be done using an equation of state, as we did in the above examples for the reversible compression of an ideal and VDW gas. The exact dependence of this intensive quantity is problem-dependent.

Working with infinitesimal expressions reminds us that integration is the process where infinitesimals are added along the path of integration. These infinitesimals are not necessarily constant along the path.
6.4. Other Types of Work

Apart from P-V work (compressive mechanical work), \( \delta W = -P_{\text{ext}} \, dV \), there exists several different other types of work corresponding to the various laws of physics available. Let us look at some examples.

6.4.1. Work Done by Gravity. Suppose we lift an object of mass \( m \) from the surface of the earth \( (z = 0) \) to a height \( (z = h) \). The gravitational force \( F = mg \), where \( m \) is mass and \( g \) is the gravitational acceleration, leads to a term of the form \( \Delta W = \int_0^h F \, dz = mgh \). The numerical value of \( g \) is approximately 9.8 m/s\(^2\). The force \( F \) points downward (toward the Earth). Thus, elevating a body of mass \( m \) to a height \( h \) from the ground requires an amount of work \( mgh \). The SI units of \( g \) are m/s\(^2\).

6.4.2. Work Done by a Spring. The linear restoring force \( F = kx \) (Hooke’s law), where \( k \) is the spring constant and \( x \) is the extension of the spring from its rest position, leads to a term of the form \( \Delta W = \int F \, dx = \int kx \, dx = \frac{1}{2}kx^2 \).

6.4.3. Electrostatic Work. Suppose that we add an electric charge to an existing (static) system of charged particles \( \{ (q_i, \mathbf{r}_i) \}_{i=1}^N \). The charged particles distribution is described by an electric potential \( \phi \) or electromotive force, \( \mathcal{E} \), which is additive (sum over all charges):

\[
\phi(\mathbf{r}) = \frac{1}{4\pi \epsilon_0} \sum_{i=1}^{N} \frac{q_i}{|\mathbf{r} - \mathbf{r}_i|}.
\]

This charge distribution gives rise to an electric field, which is obtained from the spatial derivative of \( \phi(\mathbf{r}) \):

\[
\mathbf{E}(\mathbf{r}) = -\nabla \phi(\mathbf{r}) = \frac{1}{4\pi \epsilon_0} \sum_{i=1}^{N} \frac{q_i (\mathbf{r} - \mathbf{r}_i)}{|\mathbf{r} - \mathbf{r}_i|^3}.
\]

Suppose that we place a charge \( q \) at \( \mathbf{r} \). It will experience a force:

\[
\mathbf{F} = q\mathbf{E}(\mathbf{r}).
\]

The work is force over distance. The work done in moving a charge \( q \) from \( A \) to \( B \) is:

\[
W = - \int_A^B \mathbf{F} \cdot d\mathbf{x} = -q \int_A^B \mathbf{E} \cdot d\mathbf{x}.
\]

Here, the minus sign denotes the work done on the charge against the action of the field. Then, using \( \mathbf{E}(\mathbf{r}) = -\nabla \phi(\mathbf{r}) \),

\[
W = q \int_A^B \nabla \phi \cdot d\mathbf{x} = q \int_A^B d\phi = q(\phi_B - \phi_A).
\]
We normally assume that a point charge is brought from infinity \((A \to \infty)\) to a point \(r\) in a region of localized electric fields described by the potential \(\phi\) (which vanishes at infinity). Then, \(q\phi\) can be interpreted as the potential energy of the test charge \(q\) in the electrostatic field. The integral is a line integral along a curve \(C\) from \(A\) to \(B\). \(dx\) is an infinitesimal line element along the curve \(C\).

Denote an infinitesimal charge added as \(dq\) and the associated work as \(\delta W\). The work is \(\delta W = E\, dq\) or \(\delta W = \phi\, dq\), depending on the notation used (\(E\) vs \(\phi\)). \(E\) vs \(\phi\) are the same thing; some books will use \(E\) whereas others will use \(\phi\). The units of \(E\) and \(\phi\) are volt (V), whereas those of \(dq\) are Coulomb (C). The total work is obtained by integrating \(\delta W\). Note: the meaning of \(dq\) is clear if we write \(\rho(\mathbf{r})\, d^3r\), where \(\rho(\mathbf{r})\) is the charge density and \(d^3r\) is a volume element centered on \(r\). Since \(\delta W\) is to be integrated, this integral is over all space:

\[
\Delta W = \int \delta W = \int_{\Omega} \phi(\mathbf{r}) \rho(\mathbf{r}) \, d^3r.
\]

where \(\Omega \in \mathbb{R}^3\) is a region of space. Suppose that a charge \(q\) is centered at the point \(r_q\). This means that \(\rho(\mathbf{r}) = q\delta(\mathbf{r} - \mathbf{r}_q)\), where \(\delta\) is the Dirac delta function (see Section A.23). For the above-mentioned electrostatic charge distribution (Eq. 6.2), then:

\[
\Delta W = \int_{\Omega} \phi(\mathbf{r}) q \delta(\mathbf{r} - \mathbf{r}_q) \, d^3r = \frac{q}{4\pi\epsilon_0} \sum_{i=1}^{N} \frac{q_i}{|r_q - r_i|}.
\]

If the force depends on a single dimension \((x)\), we may take:

\[
\phi(x) = \frac{1}{4\pi\epsilon_0} \sum_{i=1}^{N} \frac{q_i}{|x - x_i|}, \quad E(x) = -\frac{d}{dx} \phi(x) = \frac{1}{4\pi\epsilon_0} \sum_{i=1}^{N} q_i \frac{x - x_i}{|x - x_i|^3},
\]

\[
\delta W = \phi(x) \, dq, \quad \Delta W = \frac{q}{4\pi\epsilon_0} \sum_{i=1}^{N} \frac{q_i}{|x_q - x_i|},
\]

where \(|x - x_i| = \sqrt{(x - x_i)^2}\) and \(\frac{d}{dx} \frac{1}{|x - x_i|} = -\frac{(x - x_i)}{|x - x_i|^3}\).

6.4.4. Work Done by an Electric Field. Suppose that we have an electric field \(\mathbf{E}\) and an object with an electric dipole moment \(\mathbf{p}\). If this object is re-oriented in the presence of the electric field, work is required according to \(\delta W = \mathbf{E} \cdot d\mathbf{p}\). The units of \(\mathbf{E}\) are V/m whereas those of \(\mathbf{p}\) are C-m, since a dipole \(\mathbf{p} = q\mathbf{d}\) consists of two point charges \(+q\) and \(-q\), separated by a displacement vector \(\mathbf{d}\). We recall from electrostatics that \(\mathbf{E} = -\nabla \phi\), so you
can see how these units also agree with the previous work \( \delta W = \phi dq \) (the units \( \nabla \) are \( \text{m}^{-1} \)).

As an example, consider an induced dipole moment. Recall from Section 2.1.2 that \( p = \epsilon_0 \alpha E \). In other words, \( p \) is parallel to \( E \) and is proportional to it. The infinitesimal work is then

$$ \delta W = E \cdot dp = \alpha dE = \frac{1}{2} dE^2. $$

Suppose we turn on the electric field (ramp it up from 0 to \( E \)), the total work is then:

$$ \Delta W = \int \delta W = \frac{1}{2} \int_0^E dE^2 = \frac{1}{2} E^2. $$

Suppose we have argon, whose polarizability is 1.63 in units of \( 1.11 \times 10^{-40} \text{C}^2/\text{m}^2/\text{J} \), according to Table 2.5.

$$ \Delta W = \frac{(1.63 \times 1.11 \times 10^{-40} \text{C}^2/\text{m}^2/\text{J})}{2} E^2. $$

To get \( \Delta W = 1 \text{ J} \), we need \( E = 1.05 \times 10^{20} \text{ V/m} \). For \( \Delta W = 1 \text{ J} \), we need \( E = 1.05 \times 10^{20} \text{ V/m} \), which is enormous.

### 6.4.5. Work Done by a Magnetic Field.

Suppose that we have a magnetic field \( B \) and an object with a magnetic dipole moment \( m \). If this object is re-oriented in the presence of the magnetic field, work is required according to \( \delta W = B \cdot dm \). The units of \( B \) are Tesla (T) whereas those of \( m \) are \( \text{J/T} = \text{A.m}^2 \) (Ampère-meter square), since a magnetic dipole \( m = IA \) can be described as a current \( I \) flowing around a loop whose area is \( A \) (area is a vector pointing normal to the surface). This is shown in Fig. 6.6.

![Model for magnetic moment (\( \vec{m} \)) as a current loop. The current, \( I \), is the charge per unit time (\( dq/dt \)). The loop has area \( A \). \( \vec{A} \) is a vector pointing perpendicular to the loop surface area, in the direction given by the right-hand rule with respect to the current direction.](image)

Physically, magnetic moments in materials can arise from nuclear spins, unpaired electrons, or orbital moments. Let us write \( \mu \) instead of \( m \) to emphasize the atomic or molecular origin of the moment. Certain nuclei can have non-zero moments, depending on the configuration of its constituent
protons and neutrons. For example, protons have spin \( I = 1/2 \); hence, hydrogen nuclei have spin 1/2. The spin angular momentum is \( \hbar I \), where \( |I| = I \). The corresponding magnetic moment is

\[
\mu = g\mu_N I,
\]

where \( g \) is the proton \( g \)-factor (\( g = 5.5856947 \)),

\[
\mu_N = \frac{e\hbar}{2m_p},
\]

is the nuclear magneton. \( m_p \) is the mass of the proton, \( e \) is the elementary charge and \( \hbar \) is the Planck’s constant divided by \( 2\pi \) (\( \hbar = h/2\pi = 1.054 \times 10^{-34} \) J.s/rad). The numerical value of \( \mu \) (for the proton) is:

\[
\mu = g\mu_N I = \frac{e\hbar}{2m_p}I = (5.585)(1.602 \times 10^{-19})(1.054 \times 10^{-34}) \frac{1}{2} \frac{1}{2(1.672 \times 10^{-27})} = 1.41 \times 10^{-26} \text{ A.m}^2
\]

Note: since these formulas are in SI units, we do not need to explicitly work out or convert units as long as all the variables are stated in SI units, as was done here.

Magnetic moments can come from electrons, in the form of spin or orbital angular momentum. The proton magnetic moment is 1.521\( \times \)10\(-3\) times that of the Bohr magneton. The Bohr magneton is a natural unit for expressing the magnetic moment of an electron:

\[
\mu_B = \frac{e\hbar}{2m_e},
\]

where \( m_e \) is the electron mass. Its numerical value is 9.27401\( \times \)10\(-24\) J/T. Electrons bound to atoms that occupy filled atomic or molecular orbitals pair up, according to the Pauli exclusion principle, as spin-up and spin-down (\( \uparrow \downarrow \)), leaving zero net moment. However, unpaired electrons (spin-up or spin-down), have non-zero moment, leading to magnetic phenomena (paramagnetism, ferromagnetism, ferrimagnetism). An example is dioxygen, which is paramagnetic. The unpaired electrons in dioxygen are explained in terms of molecular orbital theory:

https://en.wikipedia.org/wiki/Molecular_orbital_diagram#Dioxygen

Non-zero orbital angular momentum of electrons gives rise to a magnetic moment:

\[
\mu = g\hbar \frac{(-e)}{2m_e} \mathbf{L} = -g\mu_B \mathbf{L},
\]

where \( \hbar \mathbf{L} \) is the orbital angular momentum; \( |L| = \sqrt{l(l+1)} \), where \( l = 0, 1, 2, \ldots, n-1 \) (corresponding to \( s, p, d, f, \ldots \) orbitals). \( n \) is the principal quantum number. Electrons in \( s \) orbitals have \( L = 0 \) and \( \mu = 0 \).
6.4.6. Adding Molecules to a System. It requires work to add a molecule to a system while keeping it at thermal equilibrium because at equilibrium the particle must have energy comparable to the average energy. Recall (see Section 1.3.3) that for a monatomic gas the average kinetic energy per molecule is related to temperature:

\[ \frac{1}{2} m v^2 = \frac{3}{2} k_B T. \]

The new molecule or atom that we add must be imparted with this amount of energy otherwise the total energy (and consequently, the temperature \( T \)) of the system will change. The energetic cost is given by the chemical potential:

\[ \mu_i \approx \left( \frac{\Delta U}{\Delta N_i} \right)_{S,V,N_j \neq i} \]

where the subscript \( S,V,N_j \neq i \) indicates that other variables are being held constant. This definition holds in the limit of small \( \Delta N_i \). For those familiar with partial derivatives (see Section A.5), the formal definition is:

\[ \mu_i \equiv \left( \frac{\partial U}{\partial N_i} \right)_{S,V,N_j \neq i}. \]

\( \mu_i \) is the energy required to add 1 particle to the system. You can think of it as the resistance to adding particles. The associated work is:

\[ \delta W_i = \mu_i dN_i. \]

If there are \( N \) chemical species, the infinitesimal work is the sum:

\[ \delta W = \sum_{i=1}^{N} \mu_i dN_i. \]

In chemistry, we normally work with molar quantities:

\[ \mu_i = \left( \frac{\partial U}{\partial n_i} \right)_{S,V,n_j \neq i}, \quad \delta W = \sum_{i=1}^{N} \mu_i dn_i. \]

The units of \( \mu_i \) are the same (Joule) except that the amount of energy involved is that which involves adding a mole of substance to the system.

6.4.7. The Infinitesimal Work Terms Follow a General Pattern. We note that all types of works are of the form:

\[ \delta W = (\text{intensive quantity}) \times d(\text{extensive quantity}). \]

In the above examples, our intensive and extensive quantities were:
### 6. Thermodynamics

#### 6.5. Laws of Thermodynamics

**6.5.1. Classical thermodynamics describes systems close to equilibrium.** The laws of thermodynamics describe the behavior of systems which are in thermal equilibrium or approaching thermal equilibrium. *An equilibrium system is defined as one whose macroscopically observable properties do not change with time.*

Equilibrium implies the macroscopic properties are spatially homogeneous, i.e. there are no gradients (spatial derivatives) in the properties. For, if gradients were present, forces (fluxes) would be set into motion to restore the system to equilibrium conditions, eliminating the gradients.

<table>
<thead>
<tr>
<th>Intensive variable</th>
<th>Corresp. Extensive variable</th>
<th>Type of work</th>
</tr>
</thead>
<tbody>
<tr>
<td>$P$ (pressure)</td>
<td>$V$ (volume)</td>
<td>mechanical (e.g. $P - V$) work: when a system goes from high pressure $P_{ext}$ to lower pressure while changing its volume $V$. The energy changes by $-P_{ext}dV$.</td>
</tr>
<tr>
<td>$\phi$ (electrostatic potential) or $E$ (electromotive force)</td>
<td>$q$ (electric charge)</td>
<td>electrical (e.g. electromotive force, emf): when a system goes from high emf $E$ to low emf by transferring electric charges $q$ between the two electrodes. The energy changes by $E dq$.</td>
</tr>
<tr>
<td>$E$ (electric field)</td>
<td>$\vec{p}$ (electric dipole moment)</td>
<td>electrical polarization work: the force is defined by the surroundings’ electric field strength and the “displacement” is change of the polarization of the medium (the sum of the electric dipole moments of the molecules).</td>
</tr>
<tr>
<td>$B$ (magnetic field)</td>
<td>$\vec{m}$ (magnetic dipole moment)</td>
<td>magnetic work: here the force is defined by the surroundings’ magnetic field strength and the “displacement” is the change of total magnetic dipole moment.</td>
</tr>
<tr>
<td>$mg$ (gravitational force)</td>
<td>$h$ (displacement)</td>
<td>gravitational work: the force is defined by the surroundings’ gravitational field and the generalized displacement is change of the spatial distribution of the matter within the system. For example, the work involved when a point particle of mass $m$ is dropped from height $h_2$ to $h_1$ is $mg(h_2 - h_1)$.</td>
</tr>
<tr>
<td>$kx$ (spring force)</td>
<td>$x$ (displacement)</td>
<td>(same as above)</td>
</tr>
<tr>
<td>$\mu$ (chemical potential)</td>
<td>$N$ (number of molecules) or $n$ (number of moles)</td>
<td>chemical: a system can go from high chemical potential $\mu_i$ to low potential by transferring particles $n_i$. The energy change by $\mu_i dn_i$.</td>
</tr>
<tr>
<td>$T$</td>
<td>$S$</td>
<td>thermal: a system equilibrates by going from high $T$ to low $T$ while transferring a quantity known as entropy $S$. If $T$ depends on position, heat will flow from regions of high $T$ to regions of low $T$. The energy of the system changes by $TdS$.</td>
</tr>
</tbody>
</table>
6.5. Laws of Thermodynamics

Thermodynamics is capable of describing changes away from equilibrium, but only provided we are not too far from equilibrium. For example, it can predict the direction of a chemical reaction, heat flow or mass flux.

However, if the system is far from equilibrium, equilibrium thermodynamics will not predict its detailed behavior except to indicate the direction of change. What it does not do is provide information about the microscopic properties of the system. For that, we need statistical mechanics.

6.5.2. The Four Classical Laws. There are four classical laws of thermodynamics. The first and second laws introduce state functions, \( U \) and \( S \). The zeroth law introduces the existence of temperature \( T \), whereas the second law establishes the absolute temperature scale. The third law is statement regarding the properties of systems in equilibrium at absolute zero temperature.

| Zeroth Law (Maxwell 1871, Carathéodory 1909) | If two systems are in thermal equilibrium with a third system, then they are in equilibrium with each other. In fact, there is a single system property (called temperature) that serves to indicate whether systems are in thermal equilibrium. Note: this law postulates the existence of a thermometer. |
| First Law (Hess 1841, Clausius & Rankine 1850) | There is a system property (i.e. state function) called energy, \( U \), that is conserved, but can take several different forms that can interconvert. Note: this law postulates the existence of a state function, \( U \). |
| Second Law (Clausius 1854, Kelvin 1851, Carathéodory 1909, Carnot 1824) | There is a system property (i.e. state function), \( S \), called entropy that, if the system is isolated from its environment, either increases or (in principle) remains constant during thermodynamic processes. Mathematically, \( dS_{\text{univ}} \geq 0 \), where \( S_{\text{univ}} \) is the entropy of the isolated system. Note: \( S_{\text{univ}} \) can be taken to be the whole universe (system plus surroundings) because the universe is an isolated system. Note: this law postulates the existence of a state function, the entropy. |
| Third Law (Nernst 1912) | The entropy of a system is a universal constant (set to zero, for a perfect crystal) at the absolute zero of temperature.\(^5\) Note: \( S(T \to 0) = 0 \) for a perfect crystal. The third law was first conjectured by Walther Nernst in 1912 and proved in 2017 by Lluis Masanes and Jonathan Oppenheim.\(^6\) |

\(^5\)The entropy of degenerate systems tends to a nonzero value when \( T \to 0 \). This is why the Third Law is formulated in terms of “the entropy of a perfect crystal”. For perfect crystals, the ground state is non-degenerate.

\(^6\)Masanes and Oppenheim, A general derivation and quantification of the third law of thermodynamics, Nature Comm. 8, 14538 (2017) https://www.nature.com/articles/ncomms14538
6.6. Zeroth Law

The Zeroth Law of thermodynamics states that thermal equilibrium is an equivalence relation: If \( A \) is in equilibrium with \( B \) and \( B \) is in equilibrium with \( C \) then \( A \) is in equilibrium with \( C \). What is an equivalence relation? A binary relation \( \sim \) on the set \( X = \{A, B, C, \ldots\} \) is an equivalence relation if and only if it is reflexive, symmetric and transitive, i.e. for all \( A, B, C \) in \( X \):

- \( A \sim A \) (reflexivity)
- if \( A \sim B \) then \( B \sim A \) (symmetry)
- if \( A \sim B \) and \( B \sim C \) then \( A \sim C \) (transitivity)

Here, \( A \sim B \) denotes “\( A \) is in thermodynamic equilibrium with \( B \)”. \( X \) denotes the set of all possible thermodynamical systems. In spite of its apparent simplicity, the zeroth law has the important consequence that it establishes any number of possible empirical temperature scales. Each thermodynamic system can be labeled with a temperature, similar to a real number system (systems with the same temperature are considered equivalent). The existence of an absolute temperature scale is then a consequence of the Second Law, which establishes the existence of the entropy. Other physical meanings of the zeroth law include: “All heat is of the same kind” (Maxwell) and “All diathermal walls are equivalent”.

6.7. First Law

The importance of changes is embedded in the statement of the First Law of thermodynamics,

<table>
<thead>
<tr>
<th>First Law of Thermodynamics</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Statement 1</strong>: There is a system property (i.e. state function) called energy that is conserved, but can take several different forms that can interconvert.</td>
</tr>
<tr>
<td><strong>Statement 2</strong>: [ dU = \delta Q + \delta W ] is an exact differential.</td>
</tr>
<tr>
<td>( dU ) is the sum of “infinitesimal” heat transfer (from the surroundings to the system), ( \delta Q ), and “infinitesimal” work done on the system ( \delta W ) by the surroundings. Here, ( q \equiv \Delta Q = \int \delta Q ) and ( w \equiv \Delta W = \int \delta W ). Many textbooks write ( q ) for heat transfer.</td>
</tr>
</tbody>
</table>

6.7.1. \( U \) is a state function, meaning it is path independent. If \( dU \) is an exact differential, its integral is independent of path, \( \Delta U = \int_{f}^{f} dU = U_f - U_i \) (according to the fundamental theorem of calculus), and only depends on the initial and final points. This is the definition of a state function.
6.7. First Law

First Law of Thermodynamics (in integral form)
\[ \Delta U = \Delta Q + \Delta W \] is independent of path.

A function \( U \) which is independent of path is called a state function in the language of thermodynamics. This is because it only depends on the state of the system, not on the path taken to arrive at the final state.

Other books (e.g. Oxtoby [1], p.529) use a slightly different notation (\( q \equiv \Delta Q, \ w \equiv \Delta W \)) which writes \( \Delta U = q + w \). These definitions are equivalent. A state function implies that the change \( \Delta U \) can be written as the difference of final and initial states, \( U_f - U_i \). Thus, internal energy is a state function. This is the same as saying that \( dU \) is an exact differential, which implies that \( \Delta U = \int_{U_i}^{U_f} dU = U_f - U_i \). Obviously, this is independent of path (it only depends on initial and final states).

6.7.2. First Law is a Statement on Conservation of Energy. In this section we show that the First Law can be interpreted as a statement of conservation of energy.

Conservation of energy and energy transfer to the system or its surroundings. In any process, the heat added to the system is removed from the surroundings (\( q = -q_{ext} \)). The work done on the system is the work done by the surroundings (\( w = -w_{ext} \)). Adding these two and invoking the first law,
\[ \Delta U = -\Delta U_{ext}. \]

The total energy changes in the universe (system plus surroundings) is always zero in any process:
\[ \Delta U_{univ} = \Delta U + \Delta U_{ext} = 0. \]
Thus, energy is conserved (even in the presence of heat exchange and dissipation). It can only be exchanged between the system and its surroundings.

The statement \( \Delta U_{univ} = 0 \) (“The energy of the universe is constant”) is another formulation of the First Law. In thermodynamics, the universe is another word for isolated system, this is equivalent to:

The energy of an isolated system is constant. (1st Law)

6.7.3. Summary of First Law: \( dU = \delta(\text{heat}) + \delta(\text{work}) \). The change in the internal energy for a closed system, by any process, is defined by the
change in external variables:

\[ dU = \delta Q + \delta W. \]

Since heat, \( Q \), and work, \( W \), are not state functions, many alternative paths can result in a given change \( dU \).

As we will see in Section 6.8 below, there exists a state function \( S \) called entropy. In the case of a reversible process, the heat transfer, \( \delta Q \) is of the form \( TdS \). Suppose one chooses a reversible path. Then, \( \delta Q = TdS \) and \( \delta W = -PdV \) and thus \( dU \) can be expressed by internal variables:

\[ dU =TdS - PdV. \]

Here, the change in the internal energy is expressed by state functions only. The change in a state function is independent of the path and this last equation is valid, regardless of which path the change in internal energy was brought about. For a closed system, \( U \) is a function of \( S \) and \( V \); thus \( dU \) is a total differential of internal variables.

For an open system, the internal energy is, in addition, a function of the amount of matter, i.e. \( U \) is a function of \( S,V,\{n_i\} \), where \( n_i \) is the amount (moles) of component \( i \). The change in internal energy, expressed as the total differential of internal variables, will then be

\[ dU (S,V,\{n_i\}) = TdS - PdV + \sum \mu_i dn_i, \]

where

\[ \mu_i \approx \left( \frac{\Delta U}{\Delta n_i} \right)_{S,V,n_j \neq i} \quad \text{or} \quad \mu_i = \left( \frac{\partial U}{\partial n_i} \right)_{S,V,n_j \neq i} \]

is the chemical potential of component \( i \). The summation is over all components. This latter expression for the differential \( dU \) contains only state functions. \( dU \) is a total differential of internal variables and the equation is generally valid for an open system. Chemical potentials have no absolute values, and the value of \( dU \) for an open system can only be given with respect to a reference state.

6.7.4. Choice of System vs Surroundings. The differential \( dU \) used in the statement of the first law, \( dU = \delta(\text{heat}) + \delta(\text{work}) \), can refer to the system, its surroundings or the whole universe. It’s up to us to decide what we call the system and its surroundings. For example, when writing \( dU = \delta Q - P_{\text{ext}}dV \), we could be referring to the circular region below, or its surroundings. See Fig. 6.7.

On the left, we have a system and its surroundings. The union of these two is the universe. On the right, we have the inverted situation: the system what used to be the surroundings, and the surroundings is what used to be the system. This inverted configuration is also possible. On the left, \( dU \)
Figure 6.7. The universe is divided into a system and its surroundings. This division is largely arbitrary and a matter of convenience.

refers to the circular region, \( \delta Q \) is the heat transferred into it, \( dV \) refers to its volume change, \( P_{\text{ext}} \) refers to the external pressure acting on it from the surroundings. On the right, \( dU \) refers to the change in internal energy of the grey region, \( dV \) refers to the change in volume of the grey region and \( P_{\text{ext}} \) refers to the pressure exerted on the grey region by the colored region.

6.8. Second Law

There is an extensive thermodynamic state function called entropy, \( S \), which can be used to indicate the direction in which systems may spontaneously evolve. \( S \) also dictates the maximum amount of work that can be obtained from a given chemical transformation or maximum efficiency for extracting work out of a heat engine.

To operate at maximum efficiency, a process must be fully reversible. Clausius defined change in entropy of a system\(^7\) \( dS \) in terms of the quantity of heat \( \delta Q^{\text{rev}} \) that is reversibly added to the system at a given temperature \( T \):

\[ dS = \frac{\delta Q^{\text{rev}}}{T} \]

\(^7\)In the definition of entropy, \( dS = \frac{\delta Q^{\text{rev}}}{T} \), the \( 1/T \) factor is an integrating factor. This integrating factor is analogous to integrating factors in the theory of differential equations, which are used to transform the differential equation into an exact differential equation. The reversible heat transfer, \( \delta Q^{\text{rev}} \), by itself is not an exact differential. However, when multiplied by \( 1/T \), it is. Another way to say this is that while \( dS \) is exact differential, the product \( TdS \) is no longer exact. An example discussed in class is the exact differential \( df = ydx + xdy \) (which we verified was exact using Euler’s test). When multiplying by another function, \( df \) may no longer be exact. For example, \( \omega = ydf = y^2dx + xydy \) is no longer an exact differential even though \( df \) was exact.
Thermodynamic definition of entropy: (of a system)

\[ dS \equiv \frac{\delta Q^{\text{rev}}}{T}. \]

\( dS \): infinitesimal entropy change of the system.
\( \delta Q^{\text{rev}} \): heat added reversibly to the system.

Note: A process for which \( dS = 0 \) is called isentropic. A process for which \( \delta Q = 0 \) is called adiabatic. The two notions are equivalent when the process is reversible \((\delta Q = \delta Q^{\text{rev}} = TdS)\).

This implies that to find the entropy change in the system, we need to find a reversible path and integrate. (In practice, a reversible path is one that involves slow changes.) It will yield a state function which depends on initial and final point:

\[ \Delta S = \int_{i}^{f} dS = S_f - S_i = \int_{i}^{f} \frac{\delta Q^{\text{rev}}}{T}. \]

The equality on the right hand side applies for a reversible process, in which case, \( dS = \frac{\delta Q^{\text{rev}}}{T} \).

The problem is further simplified if the process is isothermal, in which case temperature is a constant and can be taken out of the integral:

\[ \Delta S = \frac{1}{T} \int_{i}^{f} \delta Q^{\text{rev}} = \frac{\Delta Q^{\text{rev}}}{T}, \text{ where } \Delta Q^{\text{rev}} = \int_{i}^{f} \delta Q^{\text{rev}}. \]

The total entropy is \( S_{\text{univ}} = S + S_{\text{ext}} \). Its change is:

\[ dS_{\text{univ}} = -S_{\text{ent}} + dS_{\text{ext}}. \]

The second law states that the total entropy change (of the universe) must be nonnegative:
Second Law of Thermodynamics (Clausius inequality):
The entropy of the universe is a nondecreasing function of time:

\[ \frac{\delta S_{\text{univ}}}{\delta t} = \frac{\delta Q_{\text{rev}}}{T} \leq 0, \]

where

\[ S_{\text{univ}} = S + S_{\text{ext}}. \]

The equality holds at thermal equilibrium (when processes become reversible). This statement applies to an isolated system ("the entropy of an isolated system is nondecreasing"), not only the universe. The universe, of course, is an isolated system. We often use the terms "universe" and "isolated" (system) interchangeably.

6.8.1. Classical Statement of the Second Law. Clausius defined change in entropy of a system \( S \) in terms of the quantity of heat \( \delta Q_{\text{rev}} \) that is reversibly added to the system at a given temperature \( T \):

**Thermodynamic definition of entropy:**
The addition of heat to a system changes its entropy by an amount:

\[ dS \equiv \frac{\delta Q_{\text{rev}}}{T}. \]

Note: \( S \) refers to the entropy of the system, not its surroundings. \( dS \) is a total differential which can be positive, zero or negative (since heat can be added or removed from the system). \( \delta Q_{\text{rev}} \) by itself is not exact, but when multiplied by \( 1/T \) it becomes exact.

Consider a system with entropy \( S \) and its surroundings at entropy \( S_{\text{ext}} \). Let \( S_{\text{univ}} = S + S_{\text{ext}}. \)

Second Law of Thermodynamics (Clausius inequality):

\[ dS_{\text{univ}} = dS + dS_{\text{ext}} \geq 0. \]

This is the statement that entropy of the universe is a nondecreasing function of time. The equality holds at thermal equilibrium (when processes become reversible). This statement applies to an isolated system ("the entropy of an isolated system is nondecreasing"). The universe, of course, is an isolated system.

The equality on the right hand side applies for a reversible process, in which case, \( dS = \delta Q_{\text{rev}} / T \).
Heat added to the surroundings is opposite in sign to the heat added to the system

\[ \delta Q_{\text{ext}}^{\text{rev}} = -\delta Q^{\text{rev}}. \]

Thus, for a reversible process \( (dS = \delta Q^{\text{rev}}/T) \),

\[ dS_{\text{univ}} = dS + dS_{\text{ext}} = \frac{\delta Q^{\text{rev}}}{T} + \frac{\delta Q_{\text{ext}}^{\text{rev}}}{T} = \frac{\delta Q^{\text{rev}}}{T} - \frac{\delta Q^{\text{rev}}}{T} = 0. \]

This is the case \( dS_{\text{univ}} = 0 \) (reversible process). The remaining case is \( dS_{\text{univ}} > 0 \) (irreversible process). In a general process, the heat exchanged between the system and surroundings may be irreversible. The difference (excess) between heat transferred irreversibly and reversibly gives the increase in entropy:

\[ dS_{\text{univ}} = dS + dS_{\text{ext}} = \frac{\delta Q^{\text{rev}}}{T} - \frac{\delta Q}{T} \geq 0, \]

which leads to the following form of the Clausius inequality:

\[ \delta Q^{\text{rev}} \geq \delta Q, \]

where \( \delta Q \) is the amount of heat absorbed by the system from its surroundings (in a process that can be irreversible), whereas \( \delta Q^{\text{rev}} \) is the heat added to the system in a reversible process. The equality holds in the reversible case.

Note: there are many different formulations of the Clausius inequality. See, for example:

http://www.eoht.info/page/Clausius+inequality

For a general irreversible process, we then have the change in entropy of the system:

\[ \Delta S = \int_A^B dS = S_B - S_A = \int_A^B \frac{\delta Q^{\text{rev}}}{T} \geq \int_A^B \frac{\delta Q}{T}. \]

For any cyclic process (process which begins and ends at the same state)

\[ S_B - S_A = 0 \]

so that

\[ \int_A^B \frac{\delta Q^{\text{rev}}}{T} \equiv \oint \frac{\delta Q^{\text{rev}}}{T} = 0. \]

Therefore,

\[ \oint \frac{\delta Q}{T} \leq 0. \] Clausius’ theorem (1854)
The symbol $\oint$ denotes that the process is carried out along a path that is cyclic (end point = start point).

It is helpful to learn more about the Carnot engine (see, e.g., Oxtoby [1], Section 13.4, “Carnot cycles, efficiency and entropy”), where this integral is carried out along paths of constant $T$ and $S$.

### 6.8.2. Isentropic vs Adiabatic Process

From the thermodynamic definition of entropy $dS \equiv \frac{\delta Q_{\text{rev}}}{T}$, we have that the heat transferred to the system in a reversible process is $\delta Q_{\text{rev}} = TdS$. A process for which $dS = 0$ is called isentropic. A process for which $\delta Q = 0$ is called adiabatic. The two notions are equivalent when the process is reversible ($\delta Q = \delta Q_{\text{rev}} = TdS$).

### 6.8.3. Reversible Process, Fundamental Equation of Thermodynamics

For a reversible process the heat transferred is $\delta Q_{\text{rev}} \equiv TdS$. The infinitesimal change in energy is then

$$dU = TdS - PdV + \sum \mu_idN_i$$

This equation is so important that it is called the fundamental equation of thermodynamics. It is valid for a reversible process. For the case of $P - V$ work only,

$$dU \approx \left(\frac{\Delta U}{\Delta S}\right)_V dS + \left(\frac{\Delta U}{\Delta V}\right)_S dV$$

or, in the language of calculus,

$$dU \equiv \left(\frac{\partial U}{\partial S}\right)_V dS + \left(\frac{\partial U}{\partial V}\right)_S dV = TdS - PdV$$

is plotted in Fig. 6.8.

### 6.8.4. Inequality for First Law, Irreversible Process

For a general process we have $\delta Q \leq TdS$, according to the Clausius inequality

$$\delta Q_{\text{rev}} \geq \delta Q.$$ 

Therefore, the Clausius inequality implies that

$$dU \leq TdS - PdV + \sum \mu_idN_i.$$ (First Law)

On the left hand side, $dU$ refers to a general (not necessarily reversible) process. This inequality is true for a general process.

Note: we have not changed the First Law, which is still “$dU = \delta Q + \delta W$ is an exact differential”. All we have done is express $dU$ as an inequality because $\delta Q \leq \delta Q_{\text{rev}} = TdS$. 


The equality $\delta Q^{rev} = TdS$, of course, follows from the thermodynamic definition of entropy $dS = \delta Q^{rev}/T$.

6.8.5. Proof of Second Law. A proof of the Second Law can be found in the articles [22, 23]. By examining the second law in the context of a finite system evolving over a short time period, the authors show from first principles that random fluctuations in the behavior of a non-equilibrium system can appear to violate the second law of thermodynamics over isolated periods of time, and that these local violations become more probable exponentially as we decrease either the size of the system or the time over which we measure the entropy. The mathematical form of this result is known as the fluctuation theorem. Remarkable, at the same time that the fluctuation theorem shows that the total entropy can decrease over very short periods of time in small systems, it also shows that in the limit of long times and macroscopic systems the entropy must increase, providing a mathematical proof of the second law.

6.9. Third Law

According to the Third Law of thermodynamics (Nernst), we can find the absolute entropy of a perfect crystal:

$$\lim_{T \to 0} S = 0.$$  

Combining $dS = \frac{\delta Q}{T}$, with $\delta Q = C_VdT$ (heating at constant volume),

$$S(T) = S_0 + \int_0^T \frac{C_V(T)}{T}dT.$$
6.9. Third Law

$S_0$ is the entropy at absolute zero. An analogous formula exists at constant pressure:

$$S(T) = S_0 + \int_0^T \frac{C_P(T)}{T} \, dT.$$ 

Note that since the heat capacity is always positive, the entropy increases with increasing temperature.

If $S_0$ is known, the absolute entropy $S(T)$ can be obtained at any temperature $T$ through measurements of the heat capacity versus temperature from 0 up to $T$.

Invoking the third law, $S_0 = 0$, and

$$S(T) = \int_0^T \frac{C_V(T)}{T} \, dT.$$ 

The plausibility of the Third Law can be appreciated from the Boltzmann entropy:

$$S = k_B \log \Omega,$$

which is zero ($S = 0$), because at $T = 0$ K the system is in its unique ground state, i.e. there is only 1 ground state configuration.\(^8\)

6.9.1. Absolute Zero. The third law is equivalent to the statement that “It is impossible by any procedure, no matter how idealized, to reduce the temperature of any system to zero temperature in a finite number of finite operations” (Guggenheim, 1967).

6.9.2. Proof of Third Law. A proof of the Third Law can be found in [24].

6.9.3. Common Thermodynamic Potentials. In classical thermodynamics, we often use thermodynamics potentials, such as $U$, $H$, $A$ and $G$, depending on the experimental situation. They are all energies, and serve the same purpose of calculating changes in the energy of the system. Except that which one we should use depends on the experimental conditions. $U$

\(^8\)(Optional) For those familiar with the Gibbs-Shannon entropy,

$$S = -k_B \sum_i p_i \log p_i,$$

this situation corresponds to $p_0 = 1$ (where 0 denotes the ground state) and $p_{i\neq 0} = 0$. But this implies that $S = 0$. Even if the ground state is degenerate, their probability would be $p_i = 1/N$ for each of these eigenstates of the ground state. Then, $S = -k_B \sum_{i=1}^N \frac{1}{N} \log \frac{1}{N} = k_B \log N$, where $N$ denotes the number of states (not the number of particles). Suppose that the degeneracy is as large as Avogadro’s number ($6 \times 10^{23}$), $k_B = 1.38 \times 10^{-23}$, so $S = (1.38 \times 10^{-23}) \log(6 \times 10^{23})$ $\approx (1.38 \times 10^{-23}) \cdot 55 \approx 10^{-21}$ J/K/mol, which is vanishingly small (well below a measurable value of $S$). Note: The Third Law is formulated in terms of “the entropy of a perfect crystal”. This is because in a perfect crystal the ground state is unique and non-degenerate. When the ground state is degenerate, the entropy tends to a nonzero value when $T \to 0$.\)
depends on $S, V, \{N_i\}$. This may not be a convenient representation to use in many situations, because it may be difficult to keep $S$ and $V$ constant. It may be easier, for example, to keep $T$ and $P$ constant instead, as is the case for many chemistry experiments. In which case, we prefer to use $G$, the Gibbs free energy.

6.9.4. **Internal Energy.** The change in internal energy is

$$dU(S, V, \{N_i\}) = TdS - PdV + \sum_i \mu_i dN_i$$

if there is more than one chemical species. The differential form $dU$ is exact, according to the First Law.

6.9.5. **Enthalpy.** The enthalpy is defined as:

$$H(S, P, \{N_i\}) = U + PV.$$ Differentiating:

$$dH(S, P, \{N_i\}) = dU + (PdV + VdP) = (TdS - PdV + \sum_i \mu_i dN_i) + (PdV + VdP).$$

This gives:

$$dH(S, P, \{N_i\}) = TdS + VdP + \sum_i \mu_i dN_i.$$ The enthalpy is useful for processes carried out at constant entropy and pressure, such as chemical reactions. Constant pressure because the reaction is often exposed to atmospheric pressure. The entropy is constant if there is no heat transferred to the system, since $\delta Q^{rev} = TdS = 0$ (for a reversible process).
A calorimeter is a device that measures the heat flow during a process. They are the chief diagnostic tool in thermodynamics. Many calorimeters operate with the sample at fixed volume. We can design a calorimeter to maintain a constant temperature (using a suitable temperature control system, such as water bath or heating mantle) and constant pressure (by exposure to the atmosphere). Keeping certain variables fixed is important because it isolates the ones of interest. During a process, the enthalpy change is:

$$\Delta H = \int dH = \int (TdS + VdP).$$

If the pressure is fixed during a process (dP = 0), then the enthalpy change is equal to the heat transferred:

$$\langle \Delta H \rangle_P = \int TdS = \int \delta Q = \Delta Q.$$

The enthalpy was invented to make this relationship true.

6.9.6. Helmholtz Free Energy. The Helmholtz free energy is denoted as \(A\) or \(F\), depending on the textbook. It is defined as:

$$A(T, V, \{N_i\}) = F = U - TS.$$

Applying the same ideas as above, you can verify that

$$dA(T, V, \{N_i\}) = -SdT - PdV + \sum_i \mu_i dN_i.$$

The differential form \(dA\) is also exact because it was obtained by differentiating \(A\). The Helmholtz free energy is useful for processes carried out at constant temperature and volume (e.g., such as solids, whose volumes don’t change).

Note: here, \(F\) denotes Helmholtz free energy, not force. In the section below, we will use \(F\) to denote force. While \(F\) is widely used in the literature to denote Helmholtz free energy, we will use \(A\) in these notes to avoid confusing the Helmholtz potential with the force \(F\).

The definition \(A = U - TS\) shows that \(A\) is determined by a balance between internal energy and entropy. At low temperatures, the energy dominates. At high temperatures, the entropy dominates.

6.9.7. Gibbs Free Energy. The Gibbs Free energy is defined as:

$$G(T, P, \{N_i\}) = U + PV - TS.$$

Similarly,

$$dG(T, P, \{N_i\}) = -SdT + VdP + \sum_i \mu_i dN_i.$$
The differential form $dG$ is exact because it was obtained by differentiating $G$. The Gibbs free energy is useful in chemistry for processes carried out at constant temperature and pressure. Notice that $H = U + PV$ and $G = U + PV - TS$ are related by $G = H - TS$.

By integrating $dG$ at constant temperature ($dT = 0$) we have the famous formula: $\Delta G = \Delta H - T \Delta S$.

Notice that at constant $T, P$ ($dT = 0, dP = 0$), the Gibbs free energy is simply a measure of the energy exchanged through chemical transformation:

$$dG(T, P, \{N_i\}) = \sum_i \mu_i dN_i. \quad \text{(constant } T, P)$$

The Gibbs energy was *invented* with this experimental situation in mind. Many chemistry experiments are carried out at constant $T, P$.

By its construction, $G = H - TS$, the Gibbs free energy describes the balance between enthalpy and entropy (enthalpy dominates at low temperatures; entropy dominates at high temperatures). In a system held at constant $P, T$, the Gibbs energy will be at a minimum. The system can exchange energy with the surroundings by volume changes and heat transfer. Equilibrium is the state at which the entropy of the system plus surroundings is at a maximum. However, for the system itself, at constant $P, T, N$, equilibrium occurs when $G$ is at a minimum.

### 6.9.8. Landau Potential (Grand potential)

The grand potential is a quantity used in statistical mechanics, especially for irreversible processes in open systems. The grand potential is the characteristic state function for the grand canonical ensemble.

$$\Omega(T, V, \{\mu_i\}) = U - TS - \sum_i \mu_i N_i,$$

$$d\Omega(T, V, \{\mu_i\}) = -PdV - SdT - \sum_i N_idx_i.$$  

### 6.9.9. Meaning of Free Energy

The term “free energy” refers to any thermodynamic potential that has $T$ as its independent variable instead of $S$. We know that the entropy of an isolated system can never decrease, however, an isolated system is of little experimental interest. We consider instead a system connected to a much larger system (heat bath) whose properties do not change significantly. For example, a bath may be chosen

---

9Take $G = H - TS$ and differentiate to get $dG = dH - TdS - SdT = dH - TdS$ (since $dT = 0$). Integrating gives $\Delta G = \Delta H - T \Delta S$, where $\Delta G = \int dG$, $\Delta H = \int dH$ and $\Delta S = \int dS$.  

with $T, P$ constant. Chemical reactions are rarely studied under conditions of constant entropy or constant energy. Usually, the chemist carries out reactions at constant temperature and pressure ($T, P$), or constant volume and temperature ($V, T$).

Examples of “free energies” include the Gibbs free energy, $G \equiv G(T, P, \{N_i\})$, and the Helmholtz free energy, $A \equiv A(T, V, \{N_i\})$. Per this definition, the Landau potential is also a type of free energy, since $\Omega \equiv \Omega(T, V, \{\mu_i\})$. On the other hand, $H$ and $U$ are not free energies, since they depend on $S$ not $T$.

The free energies include:

\[
A(T, V, \{N_i\}) = U - TS, \\
G(T, P, \{N_i\}) = H - TS, \\
\Omega(T, V, \{\mu_i\}) = A - \sum_i \mu_i N_i.
\]

The thermodynamic potentials that are not free energies include: $U(S, V, \{N_i\})$, $S(U, V, \{N_i\})$, $H(S, P, \{N_i\}) = U + PV$.

6.9.9.1. Meaning of Free Energy. First law, $dU = \delta Q + \delta W$ and second law, $dS \geq \delta Q/T$, combined: $TdS \geq dU - \delta W$. (When changes are slow\(^{10}\), and the process is reversible, the equality holds, $TdS = dU - \delta W$.) Thus, $\delta W \geq dU - TdS$. Since $A = U - TS$, at constant temperature, $dA = dU - TdS$, we have that $\delta W \geq dA$ or $\Delta W \geq \Delta A$. In a reversible process, the equality holds and the total work done is equal to the free energy change. For an irreversible process, the work done on the system by its surroundings is always greater than the increase in free energy. An equivalent formulation is obtained: the work done by the system to the surroundings is always less than or equal to the decrease in its free energy. Thus, the free energy of a system determines the maximum amount of work that the system can do.

6.9.10. Any (“Create-Your-Own”) Potential. There are situations where other choices of thermodynamic variables may be more convenient than the ones given above. For example, consider the following potential:

\[
\Phi(S, P, \{\mu_i\}) = U + PV - \sum_i \mu_i N_i
\]

\[
d\Phi(S, P, \{\mu_i\}) = TdS + VdP - \sum_i N_id\mu_i
\]

d\Phi has the units of work. In classical mechanics infinitesimal work is of the form $Fdx$.\(^{11}\)

\(^{10}\)This relation holds for closed systems (closed to particle exchange). When particles can transfer, $TdS = dU - \delta W - \mu dN$.

\(^{11}\)(Optional) $\vec{F} \cdot d\vec{x}$ in 3D.
Suppose that the variables $S$, $P$ and $\mu_i$ are all spatially-dependent, i.e. $S = S(x)$ [or $S = S(\vec{x})$ in 3D]. Dividing through by $dx$ gives the force:

$$\begin{align*}
F &= \frac{d\Phi}{dx} = T \frac{dS}{dx} + V \frac{dP}{dx} - \sum_i N_i \frac{d\mu_i}{dx} \\
&= \frac{dS}{dx} (\text{entropic}) + \frac{dP}{dx} (\text{mechanical}) - \sum_i N_i \frac{d\mu_i}{dx} (\text{chemical})
\end{align*}$$

where $T \frac{dS}{dx}$ is interpreted as an entropic force.

The three terms are:

- $T \frac{dS}{dx}$ is the entropic force. Entropic force is an emergent force that reflects a system’s tendency to reach equilibrium by increasing entropy (in accordance to the Second Law).
- $V \frac{dP}{dx}$ is the force of $P - V$ work originating from a pressure gradient (spatially non-uniform pressure). Pressure gradients drive flow.
- $N_i \frac{d\mu_i}{dx}$ is the force corresponding to the chemical work.

### 6.9.11. Entropy Representation, Fundamental Equation.

It is sometimes convenient to treat $S$ (rather than $U$) as the dependent variable, so that $S(U,V,N_i)$ leads to the following entropy representation of the fundamental equation

$$dS(U,V,N_i) = \left( \frac{\partial S}{\partial U} \right)_{V,N_i} dU + \left( \frac{\partial S}{\partial V} \right)_{U,N_i} dV + \sum_i \left( \frac{\partial S}{\partial N_i} \right)_{U,V,N_j \neq N_i} dN_i.$$

Which can be read out from the Fundamental equation by isolating $dS$:

**Entropy Representation of the Fundamental Equation:**

$$dS(U,V,N_i) = \left( \frac{1}{T} \right) dU + \left( \frac{P}{T} \right) dV - \sum_i \left( \frac{\mu_i}{T} \right) dN_i.$$ 
(reversible process, $P = P_{ext}$)

---

12. We are not really dividing through by $dx$ but rather invoking the chain rule, i.e. since $S = S(x)$, $dS(x) = S'(x)dx = \frac{dS}{dx}dx$.

13. (Optional) In 3D, the analogous expression is:

$$\vec{F} \cdot d\vec{x} = \nabla \Phi \cdot d\vec{x} = T \nabla S \cdot d\vec{x} + V \nabla P \cdot d\vec{x} - \sum_i N_i \nabla \mu_i \cdot d\vec{x}.$$

where $\nabla S = \frac{\partial S}{\partial x} \hat{x} + \frac{\partial S}{\partial y} \hat{y} + \frac{\partial S}{\partial z} \hat{z}$ is the gradient of $S$, $\nabla S \cdot d\vec{x}$ is the dot product of $\nabla S$ and $d\vec{x} = dx \hat{x} + dy \hat{y} + dz \hat{z}$, i.e. $\nabla S \cdot d\vec{x} = \frac{\partial S}{\partial x} dx + \frac{\partial S}{\partial y} dy + \frac{\partial S}{\partial z} dz$, etc.
Matching the coefficients we get:

\[ \frac{1}{T} = \left( \frac{\partial S}{\partial U} \right)_V, \quad \frac{P}{T} = \left( \frac{\partial S}{\partial V} \right)_U, \quad \mu_i = -\left( \frac{\partial S}{\partial N_i} \right)_{U,V,N_j \neq N_i}. \]

The entropy representation gives a natural decomposition of the entropy as the sum of two terms: heat \((dU/T)\) + work (here, \(PV\) + chemical term). When \(dV = 0\), we are left with heat transfer \(dU/T\) plus chemical terms. The chemical term can describe transport of chemical species in and out of the system, or it can describe spontaneous (irreversible) chemical transformations within the system. The heat term can be positive, zero or negative. It can describe reversible or irreversible heat transfer. Spontaneous chemical reactions are irreversible and lead to “entropy production”. The latter is nonnegative \((\geq 0)\).

6.9.12. Gibbs Free Energy, Criteria for Spontaneity. The criterion for spontaneity is given by the Second Law \((dS_{univ} \geq 0)\):

<table>
<thead>
<tr>
<th>Spontaneity Criterion (Entropy, (dS_{univ} \geq 0)):</th>
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<tbody>
<tr>
<td>(dS_{univ} &gt; 0) spontaneous (irreversible)</td>
</tr>
<tr>
<td>(dS_{univ} = 0) reversible</td>
</tr>
<tr>
<td>(dS_{univ} &lt; 0) not spontaneous (“forbidden”)</td>
</tr>
</tbody>
</table>

In terms of the Gibbs energy \(G = H - TS\), this becomes (fixed \(T\) and \(P\)):

<table>
<thead>
<tr>
<th>Spontaneity Criterion (Gibbs energy, (dG_{univ} \leq 0)):</th>
</tr>
</thead>
<tbody>
<tr>
<td>(dG_{univ} &lt; 0) spontaneous (irreversible)</td>
</tr>
<tr>
<td>(dG_{univ} = 0) reversible</td>
</tr>
<tr>
<td>(dG_{univ} &gt; 0) not spontaneous (“forbidden”)</td>
</tr>
</tbody>
</table>

Thus, maximizing the entropy of the universe is equivalent to minimizing the Gibbs energy. Note that spontaneity does not imply that the reaction will necessarily occur at a detectable rate.

6.9.13. Euler Relation. Since \(U\) is extensive and is also a function of extensive variables \(S, V, N_i\), if these variables are all multiplied by \(\lambda\), then \(U\) will also change by the same factor:

\[ U(\lambda S, \lambda V, \{\lambda N_i\}) = \lambda U(S, V, N_i). \]

Differentiating both sides with respect to \(\lambda\) gives:

\[ \frac{d}{d\lambda} \lambda U(S, V, N_i) = U(S, V, N_i). \]
and
\[
\left. \frac{d}{d\lambda} \right|_{\lambda=1} U(\lambda S, \lambda V, \{\lambda N_i\}) = \frac{\partial U}{\partial S} \frac{d(\lambda S)}{d\lambda} + \frac{\partial U}{\partial V} \frac{d(\lambda V)}{d\lambda} + \sum_i \frac{\partial U}{\partial N_i} \frac{d(\lambda N_i)}{d\lambda}
\]
\[
= \frac{\partial U}{\partial S} S + \frac{\partial U}{\partial V} V + \sum_i \frac{\partial U}{\partial N_i} N_i
\]
by the chain rule. Here, \( \left. \frac{d}{d\lambda} \right|_{\lambda=1} \) denotes the derivative with respect to \( \lambda \) evaluated at the point \( \lambda = 1 \).

Thus, \( U \) obeys
\[
U = \left( \frac{\partial U}{\partial S} \right)_T S + \left( \frac{\partial U}{\partial V} \right)_P V + \sum_i \left( \frac{\partial U}{\partial N_i} \right)_{\mu_i} N_i.
\]

Recall\(^{14}\) that \( T \equiv (\partial U/\partial S), \quad P \equiv -(\partial U/\partial V), \quad \mu_i \equiv (\partial U/\partial N_i) \). Then, we get the Euler relation:
\[
U = TS - PV + \sum_i \mu_i N_i,
\]
which can also be expressed in the entropy representation as:
\[
S = \frac{U}{T} + \frac{PV}{T} - \sum_i \frac{\mu_i N_i}{T}.
\]

6.9.14. Gibbs-Duhem Relation. Differentiating the Euler relation (reversible process, \( P = P_{ext} \))\(^{15}\),
\[
U = TS - PV + \sum_i \mu_i N_i,
\]
we obtain
\[
dU = TdS + SdT - PdV - VdP + \sum_i \mu_i dN_i + \sum_i N_i d\mu_i.
\]
Comparing with the First Law:
\[
dU = TdS - PdV + \sum_i \mu_i dN_i,
\]
we see that for both to be correct we must have:
\[
SdT - VdP + \sum_i N_i d\mu_i = 0. \quad \text{ (Gibbs-Duhem)}
\]

\(^{14}\)From \( dU = TdS - PdV + \mu dN \) we get \( T = \partial_S U, \quad P = -\partial_V U \) and \( \mu = \partial_N U \).

\(^{15}\)Reversibility is assumed because the expression for \( dU \) (First Law) makes use of \( \delta Q = TdS \), which is true for a reversible process. For an irreversible process we have \( \delta Q \leq TdS \). (See Clausius inequality below.) Incidentally, you can see that this Clausius inequality also implies that \( dU \leq TdS - PdV + \sum_i \mu_i dN_i \).
Note: Writing the heat term as $TdS$ implicitly assumes reversibility of the process. Therefore, we wrote $P$ instead of $P_{ext}$. The Gibbs-Duhem equation shows that three intensive variables are not independent; if we know two of them, the value of the third can be determined from the Gibbs-Duhem equation.

6.9.15. Application: (Chemical Potential of Ideal Gas). Write down Gibbs-Duhem for a single component system:

$$d\mu = \left(\frac{V}{N}\right) dP - \left(\frac{S}{N}\right) dT.$$  

If we assume that the temperature is constant ($dT = 0$) and integrate:

$$\int d\mu = \mu(T, P) - \mu^\circ(T, P) = \int_{P_0}^{P} \left(\frac{V}{N}\right) dP.$$  

Substituting $V/N = k_B T/P$ from the ideal gas law:

$$\mu(T, P) = \mu^\circ(T, P) + k_B T \log\left(\frac{P}{P_0}\right) = \mu^\circ + k_B T \log\left(\frac{P}{P_0}\right)$$  

(6.3)

$$\mu^* = \mu^\circ + k_B T \log\left(\frac{[c]}{[c]_0}\right)$$

where $\rho = N/V$ and $[c] = n/V = \rho/N_A$.

6.10. Multi-Component Gas

Suppose we have a mixture of gases. Such an expression holds for each component:

$$\mu_i(T, P) = \mu_i^\circ(T, P) + k_B T \log\left(\frac{P_i}{P_0}\right)$$

For ideal gases, Dalton’s law holds, $P_i = X_i P$, where $P$ is the total pressure and $X_i$ is the mole fraction of component $i$. $P_0$ is a reference pressure, often taken to be 1 atm. Substituting Dalton’s law and collecting the terms independent of composition into the square bracket:

$$\mu_i(T, P) = \left[\mu_i^\circ(T, P) + k_B T \log(P/P_0)\right] + k_B T \log X_i$$

Some textbooks call the square bracket $\mu_i^*$. Regardless of notation, this offset/baseline term corresponds to the chemical potential of the pure substance under some reference conditions $(P, T)$. $P_0$ can be regarded as the unit of pressure used. When $X_i = 1$ (pure substance of type $i$), the log

\[\text{An easy trick to remember this equation is to start from the First Law for a reversible chemical process, } dU = TdS - PdV + \sum_i \mu_i dN_i, \text{ and reverse the roles of the intensive and extensive variables, then set } dU = 0.\]

\[\text{Gibbs-Duhem assumes reversibility because its derivation involved a } TdS \text{ term; hence we may write } P = P_{ext}.\]
term is zero and only the square bracket remains. What should this square bracket represent? It is the chemical potential of the pure substance. If our pure substance were an ideal gas, \( U = \frac{3}{2} N k_B T \). Taking the derivative with respect to \( N \) gives \( \frac{3}{2} k_B T \), which is the energy cost of adding a molecule to a gas at temperature \( T \).


Definition of Gibbs energy. The Gibbs energy is defined as

\[
G(P, T, \{N_i\}) = U + P V - T S,
\]

where the transformation from \( U \) swaps the roles of \( V \) in favor of \( P \) and \( S \) in favor of \( T \). In chemistry it is common to work at constant \( T \) and \( P \).

\[
dU = T dS - P dV + \sum_i \mu_i dN_i \quad \text{for a reversible process},
\]

\[
dG = -SdT + V dP + \sum_i \mu_i dN_i \quad \text{.}
\]

Constant \( T, P \). At constant \( T, P \) (\( dT = 0, dP = 0 \))

\[
dG = \sum_i \mu_i dN_i \quad .
\]

6.10.1.1. Integrated Form. Combining \( G = U - TS + PV \) with the Euler formula\(^\text{19} \)

\[
U = T S - PV + \sum_i \mu_i N_i
\]

\[
G = \sum_i \mu_i N_i
\]

Thus, the **Gibbs free energy is equal to the sum of all the chemical potentials of the molecules in any system**, regardless of any assumptions of constant \( T \) or \( P \). (Here we did not impose any restrictions such as constant temperature or pressure\(^\text{20} \).)

6.10.1.2. Single-Component, Ideal Gas. For a single component\(^\text{21} \), \( G = \mu N \) and substituting Eq. (6.3) yields:

\[
G = G^\circ + k_B T N \log \left( \frac{P}{P_0} \right) = G^\circ + k_B T N \log \left( \frac{\rho}{\rho_0} \right) = G^\circ + k_B T N \log \left( \frac{[c]}{[c]_0} \right) .
\]

---

\(^\text{18}\)Note: we have assumed reversibility because of the presence of the \( T dS \) term. In this case, we write \( P \) in place of \( P_{\text{ext}} \).

\(^\text{19}\)An easy way to remember the Euler formula is to start with the First Law for a reversible chemical process, \( dU = T dS - P dV + \sum_i \mu_i dN_i \), and remove the ’\( d \)’ in the differentials.

\(^\text{20}\)The only two elements that were used to derive this formula were the definition of the Gibbs energy and the Euler relation. The latter was derived using the extensively property of the thermodynamic variables. Thus, the result was obtained without assuming constant \( P \) or \( T \).

\(^\text{21}\)We often write \( P \), \( \rho \) and \( [c] \) instead of \( P/P_0 \), \( \rho/\rho_0 \) and \( [c]/[c]_0 \), where \( P \), \( \rho \) and \( [c] \) are expressed in units of \( P_0 \), \( \rho_0 \) and \( [c]_0 \), respectively.
or in terms of molar Gibbs energy $G/n$,

$$
\frac{G}{n} = \mu^° + RT \log \left( \frac{P}{P₀} \right) = \mu^° + RT \log \left( \frac{ρ}{ρ₀} \right) = \mu^° + RT \log \left( \frac{[c]}{[c]₀} \right).
$$

### 6.10.2. At Equilibrium, $dG = 0$. For a chemical reaction (usually, $T, P$ are constant), $dG = \sum_i \mu_i dN_i$. The second law states that $dG = 0$ at equilibrium. Thus, $\sum_i \mu_i dN_i = 0$. Or, in integrated form,

$$
\Delta G = \int_{\text{react.}}^{\text{prod.}} dG = G(\text{prod.}) - G(\text{react.}) = \sum_i \mu_i \int dN_i = \sum_i \mu_i \Delta N_i.
$$

The equilibrium condition is therefore equivalent to $\sum_i \mu_i \Delta N_i = 0$. However, $\Delta N_i$ (and also $dN_i$) is proportional to the stoichiometric coefficient $\nu_i$. Thus, the equilibrium condition is $\sum_i \mu_i \nu_i = 0$.

There is a sign convention that we must follow for the stoichiometric coefficients: the products are positive, the reactants are negative. Let’s look at a simple example:

$$
aA + bB \rightleftharpoons cC + dD
$$

is written in the more general form

$$
0 = \sum_i \nu_i A_i,
$$

where $A_i$ denotes the $i$-th chemical species and $\nu_i$ is the associated stoichiometric coefficient (taken negative for a reactant species or positive for a product species) in the balanced chemical equation.

For example, with the reaction $aA + bB \rightleftharpoons cC + dD$, we have that $\nu_1 = -a$, $\nu_2 = -b$, $\nu_3 = c$, $\nu_4 = d$ and $A_1 = A$, $A_2 = B$, $A_3 = C$ and $A_4 = D$.

### 6.10.3. Chemical Potential, Activity. Activities are used as a way to account for non-idealities of solutions. Chemical potentials can be written in terms of activities ($a_i$), as follows:

$$
\mu_i = \mu^°_i + RT \log a_i.
$$

This can be taken as the definition of activity:

$$
a_i \equiv \exp \left( \frac{\mu_i - \mu^°_i}{RT} \right).
$$

The equilibrium condition $\sum_i \mu_i \nu_i = 0$ becomes

$$
0 = \sum_i \mu^°_i \nu_i + RT \sum_i \nu_i \log a_i.
$$

The values $\mu^°_i$ can be found tabulated for different (pure) chemical species $i$ and some standard conditions.
6.10.4. Gibbs Free Energy of Reaction. To recap what we have done, we just calculated $\Delta G$ of the reaction, by integrating $dG$ from reactants to products. Some textbooks use a subscript “rxn” to emphasize this:

$$\Delta G_{\text{rxn}} = \int_{\text{react.}}^{\text{prod.}} dG = G(\text{prod.}) - G(\text{react.})$$

$$= \sum_i \mu_i \nu_i = \sum_i \mu_i^0 \nu_i + RT \sum_i \nu_i \log a_i.$$

At equilibrium $\Delta G_{\text{rxn}} = 0$. (Obviously, this is only valid at constant $T, P$.)

6.10.5. Standard Gibbs Free Energy of Reaction. The first term is a collection of constants describing an energy change going from reactants to products:

$$\sum_i \mu_i^0 \nu_i = G_{\text{prod.}}^0 - G_{\text{react.}}^0 = \Delta G_{\text{rxn}}^0$$

which depends only on the properties of the unmixed species. The quantity $\Delta G_{\text{rxn}}^0$ is known as the standard Gibbs free energy of the reaction. Its value can be obtained from tables.

The second term is rewritten as

$$RT \sum_i \nu_i \log a_i = RT \sum_i \log a_i^{\nu_i} = RT \log \left( \prod_i a_i^{\nu_i} \right).$$

Then we have, at equilibrium,

$$\Delta G_{\text{rxn}}^0 + RT \log \left( \prod_i a_i^{\nu_i} \right) = 0.$$

The quantity in the brackets can be identified as the reactant quotient, $Q$, the ratio of product and reactant activities, each raised to its stoichiometric power. For example, for the symbolic reaction:

$$aA + bB \rightleftharpoons cC + dD$$

$$Q = \prod_i a_i^{\nu_i} = \frac{\text{products}}{\text{reactants}} = \frac{a_C^d a_D^d}{a_A^a a_B^b}.$$  

6.10.6. Reaction Constant. At equilibrium $\Delta G_{\text{rxn}}^0$ and $T$ are constants, $Q = K$ and we have the central equation of chemical equilibrium theory:

$$\Delta G_{\text{rxn}}^0 = -RT \log K.$$

where $K = \frac{a_C^d a_D^d}{a_A^a a_B^b}$ (activities are equilibrium values). Your book uses the notation $K_C = \frac{[C]_{eq}[D]_{eq}}{[A]_{eq}[B]_{eq}}$ when dealing with solutions. Here, the concentrations
are equilibrium values. Note that \( a_i \propto \frac{[A_i]}{n_i/V} \) and the proportionality constants have been absorbed into the definition of \( K_C \). Similarly, for gaseous mixtures, \( a_i \propto P_i \) and \( K_P = \frac{P_c^a P_d^b P_b^c}{P_a^d} \).

6.10.7. What is the Meaning of the Standard Gibbs Free Energy?

We saw in the previous lecture that chemical reactions can be written in the general form, \( \sum \nu_i A_i = 0 \), where \( A_i \) denotes the \( i \)-th chemical species and \( \nu_i \) is the associated stoichiometric coefficient (taken negative for a reactant species or positive for a product species) in the balanced chemical equation.

For example, with the reaction \( aA + bB \rightleftharpoons cC + dD \), we have that \( \nu_1 = -a, \nu_2 = -b, \nu_3 = c, \nu_4 = d \) and \( A_1 = A, A_2 = B, A_3 = C \) and \( A_4 = D \). At chemical equilibrium we have that \( dG = 0 \) (constant \( T, P \)), where \( dG = \sum \mu_i d n_i \). Integrating from reactants to products gives

\[
\Delta G_{\text{rxn}} = G(\text{prod.}) - G(\text{react.}) = \sum \mu_i \nu_i,
\]

since

\[
\int_{\text{react.}}^{\text{prod.}} d n_i = \Delta n_i = \nu_i
\]

(with the above sign convention). Thus, \( \sum \nu_i = 0 \). Finally, we note that the combination of \( G = U - TS + PV \) and the Euler formula, \( U = TS - PV + \sum \mu_i n_i \), leads to \( G = \sum \mu_i n_i \), which, for a single component, implies that \( G_i/n_i = \mu_i \) (\( \mu_i \) is the chemical potential, also the molar Gibbs energy).

We have seen that substitution of the ideal gas expression for \( \mu_i \) into \( \sum \mu_i \nu_i = 0 \) leads to \( \Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^0 + R T \log Q \), where \( \Delta G_{\text{rxn}}^0 = \sum \nu_i \mu_i \nu_i \). It is important to note that \( \Delta G_{\text{rxn}} \) and \( \Delta G_{\text{rxn}}^0 \) are two different things.

\( \Delta G_{\text{rxn}}^0 \) is computed as follows:

\[
\Delta G_{\text{rxn}}^0 = \sum \nu_i (\Delta G_f^0)_i,
\]

where \( \nu_i \) are the stoichiometric coefficients of species \( i \) in the reaction (products are positive, reactants are negative) and \( \Delta G_f^0 \) are standard Gibbs energies of formation (\( i \) denotes the \( i \)-th chemical species in the reaction). These values are found in tables. The symbol ° attached to \( \Delta G_{\text{rxn}} \) means “standard” Gibbs energy, which has a single value for a particular reaction at a given temperature and pressure. We can obtain the value of \( \Delta G_{\text{rxn}}^0 \) from tables (such as Appendix D of Oxtony [1]). It corresponds to the free energy change for a process that never really happens: the complete transformation of pure reactants into pure products at a constant pressure of 1 atm (or whatever the reference conditions happen to be; but usually 1 atm). In the case of the reaction:

\[
C(s, \text{graphite}) + 2O_2(g) \rightarrow CO_2(g),
\]
\( \Delta G_{\text{rxn}}^0 \) is the free energy change for the complete transformation of pure graphite and \( \text{O}_2 \) into pure \( \text{CO}_2 \) at a constant pressure of 1 atm.

In contrast to \( \Delta G^\circ_{\text{rxn}} \), which is a constant for a given reaction, \( \Delta G_{\text{rxn}} \) varies continuously as the composition changes, finally reaching zero at equilibrium. \( \Delta G_{\text{rxn}} \) is the “distance” (in free energy) from the equilibrium state of a given reaction. The two are related by:

\[
\Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^0 + RT \log Q,
\]

where \( Q \) is the reaction quotient, which would be, for a reaction

\[
aA + bB \rightarrow cD + dD
\]

\[
Q = \frac{P_C^c P_D^d}{(P_A^a P_B^b)} \quad \text{or} \quad Q = \frac{[C]^c[D]^d}{([A]^a[B]^b)}.\]

At equilibrium we have \( \Delta G_{\text{rxn}} = 0 \) and the above equation gives \( \Delta G_{\text{rxn}}^0 = -RT \log K \) (because \( Q = K \) at equilibrium).

The change in standard Gibbs free energy for the above reaction is the sum of standard Gibbs energies of formation:

\[
\Delta G_{\text{rxn}}^0 = c(\Delta G_f^\circ)_C + d(\Delta G_f^\circ)_D - a(\Delta G_f^\circ)_A - b(\Delta G_f^\circ)_B.
\]

The equation, \( \Delta G_{\text{rxn}}^0 = -RT \log K \), is important because it relates the equilibrium composition of a chemical reaction system to measurable physical properties of the reactants and products. If you know the entropies (\( S^\circ \)) and the enthalpies (\( \Delta H_f^\circ \)) of formation of a set of substances, you can predict (since \( \Delta G_{\text{rxn}}^0 = \Delta H_{\text{rxn}}^\circ - T \Delta S_{\text{rxn}}^\circ \) at constant temperature) the equilibrium constant of any reaction involving these substances without the need to know anything about the mechanism of the reaction. We may find tables of \( \Delta G_f^\circ \), \( S^\circ \) and \( \Delta H_f^\circ \) for different substances in chemistry handbooks (e.g., [12]) or textbooks (e.g., [1]).

The standard enthalpy of formation or standard heat of formation of a compound is the change of enthalpy during the formation of 1 mole of the substance from its constituent elements, with all substances in their standard states, and at a pressure of 1 atm. There is no standard temperature. It is usually denoted \( \Delta H_f^\circ \). The superscript \( ^\circ \) on this symbol indicates that the process has occurred under standard conditions at the specified temperature (usually 25°C).

The standard Gibbs free energy of formation of a compound is the change of Gibbs free energy that accompanies the formation of 1 mole of a substance in its standard state from its constituent elements in their standard states (the most stable form of the element at 1 atm of pressure and the specified temperature, usually 25°C).
6.10.8. Where to Find Values for $\Delta H^\circ_f$, $\Delta G^\circ_f$ and $S^\circ$. Values of $\Delta H^\circ_f$, $S^\circ$ and $\Delta G^\circ_f$ can be found tabulated in most chemistry textbooks (usually, in appendices), and on the internet. In chemistry books, we do not find $\Delta G^\circ_{rxn}$ values for all possible chemical reactions, because it would be impossible to list all possible chemical reactions (there are too many of them). Instead, we will find tables of the energy of formation of various compounds, which may be involved in your reaction. Textbooks often denote these energies of formation as $\Delta G^\circ_f$.

Standard enthalpies of formation can be found at https://en.wikipedia.org/wiki/Standard_enthalpy_of_forma


Appendix D of Oxtoby [1], entitled “Standard Chemical Thermodynamic Properties”, contains tabulated values for enthalpies, entropies and Gibbs energies of formation. Please note that Oxtoby [1] writes $S^\circ$ and not $\Delta S^\circ_f$. This is because the Third Law guarantees the existence of an absolute entropy scale. So what is being tabulated here is the absolute entropy $S^\circ$ for each substance. On the other hand, there exists no absolute scale for $\Delta H^\circ_f$ and $\Delta G^\circ_f$, hence the $\Delta$ notation.

Table 6.2 is an excerpt from page 1 of the Appendix D of Oxtoby [1]. Oxtoby [1] prefaces the table as follows: “This table lists standard enthalpies of formation $\Delta H^\circ_f$, standard third-law entropies $S^\circ$, standard free energies of formation $\Delta G^\circ_f$, and molar heat capacities at constant pressure, $c_P$, for a variety of substances, all at 25°C (298.15 K) and 1 atm. [...] Note that the solution-phase entropies are not absolute entropies but are measured relative to the arbitrary standard $S^\circ(H^+(aq)) = 0$. Consequently, some of them (as well as some of the heat capacities) are negative.”

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H^\circ_f(25^\circ C)$</th>
<th>$S^\circ(25^\circ C)$</th>
<th>$\Delta G^\circ_f(25^\circ C)$</th>
<th>$c_P(25^\circ C)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>units</td>
<td>kJ/mol</td>
<td>J/K/mol</td>
<td>kJ/mol</td>
<td>J/K/mol</td>
</tr>
<tr>
<td>H(g)</td>
<td>217.96</td>
<td>114.60</td>
<td>203.26</td>
<td>20.78</td>
</tr>
<tr>
<td>H$_2$(g)</td>
<td>0</td>
<td>130.57</td>
<td>0</td>
<td>28.82</td>
</tr>
<tr>
<td>H$^+$ (aq)</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>H$_3$O$^+$ (aq)</td>
<td>-285.83</td>
<td>69.91</td>
<td>-237.18</td>
<td>75.29</td>
</tr>
<tr>
<td>Li(s)</td>
<td>0</td>
<td>29.12</td>
<td>0</td>
<td>24.77</td>
</tr>
<tr>
<td>Li(g)</td>
<td>159.37</td>
<td>138.66</td>
<td>126.69</td>
<td>20.79</td>
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<tr>
<td>Li$^+$ (aq)</td>
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<td>13.4</td>
<td>-293.31</td>
<td>68.6</td>
</tr>
<tr>
<td>LiH(s)</td>
<td>-90.54</td>
<td>20.01</td>
<td>-68.37</td>
<td>27.87</td>
</tr>
</tbody>
</table>
6. Thermodynamics

<table>
<thead>
<tr>
<th>Compound</th>
<th>$\Delta G^\circ$</th>
<th>$\Delta S$</th>
<th>$\Delta H^\circ$</th>
<th>$\Delta C_p$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li₂O(s)</td>
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<td>37.57</td>
<td>-561.20</td>
<td>54.10</td>
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<tr>
<td>LiF(s)</td>
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<td>35.65</td>
<td>-587.73</td>
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</tr>
<tr>
<td>LiCl(s)</td>
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<td>59.33</td>
<td>-384.39</td>
<td>47.99</td>
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<tr>
<td>LiBr(s)</td>
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<td>-342.00</td>
<td>—</td>
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<tr>
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</tr>
<tr>
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<td>0</td>
<td>28.24</td>
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<tr>
<td>Na(g)</td>
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<td>153.60</td>
<td>76.79</td>
<td>20.79</td>
</tr>
<tr>
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<td>59.0</td>
<td>-261.90</td>
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</tr>
<tr>
<td>Na₂O(s)</td>
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<td>-375.48</td>
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</tr>
<tr>
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<td>-425.61</td>
<td>64.46</td>
<td>-379.53</td>
<td>59.54</td>
</tr>
<tr>
<td>NaF(s)</td>
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<td>51.46</td>
<td>-543.51</td>
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<tr>
<td>NaCl(s)</td>
<td>-411.15</td>
<td>72.13</td>
<td>-384.15</td>
<td>50.50</td>
</tr>
</tbody>
</table>

6.10.9. The Standard States. When writing $\Delta G^\circ$, the $^\circ$ indicates that all components (reactants and products) are in their standard states. This concept of standard states is especially important in the case of the free energy. For most practical purposes, the following definitions are used (see https://en.wikipedia.org/wiki/Standard_enthalpy_of_formation)

- **Gases:** 1 atm partial pressure (i.e., the hypothetical state it would have assuming it obeyed the ideal gas equation at a pressure of 1 atm).
- **Pure liquids:** for a solute present in an ideal solution, a concentration of exactly 1 M at a pressure of 1 atm.
- **Solutes:** an effective concentration of 1 mol/L. (“Effective” concentrations approach real concentrations as the latter approach zero; for practical purposes, these can be considered identical at real concentrations smaller than about $10^{-4}$ molar.)
- **Solids:** for a pure substance or a solvent in a condensed state (a liquid or a solid), the standard state is the pure liquid or solid under a pressure of 1 atm.
- **For an element:** the form in which the element is most stable under 1 atm of pressure. One exception is phosphorus, for which the most stable form at 1 atm is black phosphorus, but white phosphorus is chosen as the standard reference state for zero enthalpy of formation.
- **All elements in their standard states** (oxygen gas, hydrogen gas, nitrogen gas, solid carbon in the form of graphite, etc.) have a standard enthalpy of formation ($\Delta H^\circ$) of zero, as there is no change involved in their formation.
- **There is no “standard temperature” in thermodynamics, but one often uses 25°.** (Thermodynamic standard state is different from STP used in gas law calculations.)
- **Some texts will use 1 atm as the reference pressure; others may use 1 bar.**
6.10. **Example Calculation.** Find the standard Gibbs energy change for the reaction

\[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]

where the following values of standard Gibbs energies of formation are found from tables: \( \text{CaCO}_3(s): -1128 \text{ kJ/mol} \), \( \text{CaO}(s): -603.5 \text{ kJ/mol} \), \( \text{CO}_2(g): -137.2 \text{ kJ/mol} \).

**Solution:** Substitute these values into the equation for \( \Delta G^\circ_{\text{rxn}} \), we find:

\[ \Delta G^\circ_{\text{rxn}} = (-603.5 - 137.2) - (-1128) \text{ kJ/mol} = +387.3 \text{ kJ/mol}. \]

A positive value for \( \Delta G^\circ_{\text{rxn}} \) means that the process is not spontaneous (under standard conditions), i.e. solid calcium carbonate will not form solid calcium oxide and \( \text{CO}_2 \) at 1 atm at 25\(^\circ\). Note: This reaction is carried out on a large scale to manufacture cement, so obviously the process can be spontaneous under different conditions.

Recall the relationship between \( K \) and \( \Delta G^\circ_{\text{rxn}} \):

\[ K = \exp \left( -\frac{\Delta G^\circ_{\text{rxn}}}{RT} \right). \]

Here, since \( \Delta G^\circ_{\text{rxn}} \) is positive, \( K < 1 \), meaning the equilibrium favors the reactants, meaning that \( \text{CaCO}_3 \) wants to remain in its solid form.

The practical importance of the Gibbs energy is that it allows us to make predictions based on the properties (\( \Delta G^\circ_{\text{rxn}} \) values) of the reactants and products themselves, eliminating the need for experiment.

6.10.11. **Heat Transfer at Constant Volume and Pressure.** For a closed system \( (dN_i = 0) \) under constant volume,

\[ dU = \delta Q - P_{\text{ext}} dV = \delta Q_V = C_V(T) dT, \]

which implies that

\[ C_V(T) = \left( \frac{\partial U}{\partial T} \right)_V. \]

If the temperature dependence of \( C_V \) over some temperature range of interest is negligible, \( C_V(T) = C_V \), the expression for \( dU \) can be integrated to give:

\[ \Delta U = C_V \Delta T. \]

For a closed system at constant pressure,

\[ dH = \delta Q + V \delta P_{\text{ext}} = \delta Q_P = C_P(T) dT, \]

\( dP_{\text{ext}} = 0 \)
which implies that

$$C_P = \left( \frac{\partial H}{\partial T} \right)_P.$$ 

If $C_P$ is independent of temperature we can integrate $dH$ to get:

$$\Delta H = C_P \Delta T.$$ 

### 6.10.12. Relationship Between Specific Heats

Take the derivative of

$$H = U + PV \quad \text{(definition of enthalpy)}$$

with respect to temperature:

$$\frac{\partial H}{\partial T} = \frac{\partial (U + PV)}{\partial T} = \frac{\partial U}{\partial T} + \frac{\partial (PV)}{\partial T}.$$ 

$$C_P = C_V + \frac{\partial (PV)}{\partial T}.$$ 

For an ideal gas, $PV = nRT$, so the second term equals $nR$, and we obtain:

$$C_P = C_V + nR.$$ 

Dividing through by $n$ we get a relationship between the specific heats:

$$c_P = c_V + R. \quad \text{(Mayer’s relation)}$$ 

A look at the term $\frac{\partial (PV)}{\partial T}$ reveals that it is always positive because as temperature increases, $PV$ also increases. Thus, $C_P > C_V$.

The situation is illustrated in Fig. 6.8.

Thus,

$$\left( \frac{\partial U}{\partial T} \right)_V \equiv C_V < C_P \equiv \left( \frac{\partial H}{\partial T} \right)_P.$$ 

The enthalpy accounts for energy transfer to the environment through the expansion of the system.

### 6.10.13. Latent Heat

Latent heats are associated with phase transitions, such as the melting of solids or vaporization of liquids and solids. It is the heat absorbed or released by a thermodynamic system as it undergoes a phase transition.

At the microscopic level, this heat serves to break or form bonds. Such heats are also called heat of fusion (melting), heat of vaporization, etc. There are entropy changes associated with such phase transitions, namely, the entropy of fusion, entropy of vaporization, etc. Entropy changes because there is a spatial rearrangement of the molecules in the system (e.g. from ordered crystal to disorder in a liquid). See Fig. 6.10.
For a reversible process,

\[ dS = \frac{\delta Q^{rev}}{T} \]

Consider a small temperature interval \([T_C - \epsilon, T_C + \epsilon]\) centered on \(T_C\) (critical temperature), with \(\epsilon\) so small that the temperature is essentially constant. The system undergoes a phase transition at \(T_C\). The latent heat associated with this transition is

\[ \delta Q_{lat} = T_C \, ds_{lat}, \quad \text{or} \quad \Delta Q_{lat} = T_C \, \Delta S_{lat}, \]

where \(\Delta S_{lat}\) is called the entropy of fusion or vaporization, if the process is one of melting or evaporation, respectively.
We can also express it in terms of the enthalpy change $dH = TdS + VdP$ at constant pressure ($dP = 0$) gives

$$\Delta S = \int_{T_C-\epsilon}^{T_C+\epsilon} \frac{dH}{T}$$

across the phase transition (see Fig. 6.11).

![Figure 6.11. Latent heat.](image)

and

$$\Delta S_{\text{lat}} = \frac{\Delta H_{\text{lat}}}{T_C},$$

where $T_C$ is the phase transition temperature (critical temperature) and $\Delta H_{\text{lat}}$ is the latent heat. Since we assumed constant $P$, $\Delta H_{\text{lat}}$ is the same as $\Delta Q_{\text{lat}}$. This formula gives the entropy change $\Delta S_{\text{lat}}$ corresponding to latent heat $\Delta H_{\text{lat}}$ of a phase transition at constant $P$.

Remarks:

- $\Delta S_{\text{lat}}$ and $\Delta H_{\text{lat}}$ represent discontinuous changes between two phases of the same substance occurring at the transition temperature $T_C$.
- Latent heats are tabulated as specific latent heats, $L$, which is an intensive property: $L \equiv \Delta Q_{\text{lat}}/m$, where $m$ is the mass of the substance. The common units are kJ/kg.

6.10.14. Application of First and Second Laws: Equilibrium Establishes a Common Temperature. Suppose that $A$ denotes the system and $B$ denotes the surroundings (see Fig. 6.12) and that $A$ and $B$ are separated by a thermally conductive wall. Assume that the volumes of $A$ and $B$ do not change. The combined system is the universe and it is isolated. Because entropy is an extensive property, $S_A$ and $S_B$ added yields the entropy of the universe:

$$S = S_A + S_B.$$ 

First, we rewrite the fundamental equation

$$dU = TdS - PdV + \sum \mu_i d n_i$$
The second law requires that at thermal equilibrium the temperatures of the two subsystems will necessarily become equal ($T_A = T_B$).

The entropy of the isolated system ("universe") is the sum of entropies of $A$ and $B$, $S = S_A + S_B$ (entropy is additive). Then, we invoke the fundamental equation in the entropy representation (with $dV = 0$ and $dN_i = 0$),

$$dS = dS_A + dS_B = \left(\frac{1}{T_A}\right) dU_A + \left(\frac{1}{T_B}\right) dU_B$$

where we have invoked the First Law twice, once for $S_A$ and once for $S_B$. Since energy is conserved ($A + B$ is an isolated system) $dU_A = -dU_B$:

$$dS = \left[\left(\frac{1}{T_A}\right) - \left(\frac{1}{T_B}\right)\right] dU_A.$$

Since $V_A, V_B, N_A, N_B$ are all constant, there can be no mechanical or chemical work exchanged between the system and its surroundings and so $dU_A = \delta Q$.

$$dS = \left[\left(\frac{1}{T_A}\right) - \left(\frac{1}{T_B}\right)\right] \delta Q$$

where $\delta Q$ is the heat transferred to the system (or to the surroundings if $\delta Q < 0$). At thermal equilibrium, entropy of the system is maximized $dS = 0$. So a nonzero infinitesimal heat exchange $\delta Q \neq 0$ implies that the temperatures of $A$ and $B$ must be equal:

$$[T_A = T_B] \quad (A \text{ and } B \text{ in thermal equilibrium})$$

**Note:** Suppose that $T_A$ (the system) is higher than $T_B$. Then we expect heat to flow from the system into the surroundings. The second law $dS \geq 0$ and Eq. (6.4) imply that $dU_A = \delta Q < 0$ and indeed heat flows out of the system. So the Second Law correctly predicts the direction of the heat exchange.
6.10.15. When is Heat Transfer Reversible? We have previously seen that reversibility is defined as a process for which the surroundings have intensive variables that are arbitrarily close to those of the system, i.e. \( P = P_{ext} + dT \), \( P = P_{ext} + dP \), \( \mu = \mu_{ext} + d\mu \). We’ve also seen that \( dS = 0 \) when a process is reversible. Let us check that these two conditions imply each other. Using the notation from the previous section, we would expect that if \( T_A \) is vastly different than \( T_B \), the process would not be reversible. Let \( T_B = T \) and \( T_A = T - dT \) so that the process is truly reversible. Then,

\[
dS = \left[ \frac{1}{T - dT} - \frac{1}{T} \right] \delta Q
\]

Factoring out \( 1/T \) and invoking the geometric series, \( 1/(1 - r) = 1 + r + r^2 + \ldots \) with \( r = dT/T \), we have:

\[
dS = \left[ 1 + \frac{dT}{T} + \frac{dT^2}{T^2} + \ldots \right] - 1 \right] \frac{\delta Q}{T} \approx \frac{\delta Q dT}{T^2} \approx 0,
\]

since \( \delta Q dT \) is an infinitesimal of the second order. Thus, the entropy change \( dS \) is zero when the process is reversible.

6.10.16. Can we Have Isothermal Heat Transfer? Isothermal means \( dT = 0 \) or \( T = constant \). From \( dS = \delta Q^{rev}/T \), a heat transfer \( \delta Q^{rev} \) to the system corresponds to a change in its energy in the amount of \( TdS \). Thus, it describes a change in entropy, i.e. given \( \Delta Q^{rev} > 0 \), we must have \( \Delta S > 0 \). This can happen, for example, during the course of a phase transition. Some amount of heat (latent heat) is transferred to the system at constant temperature while the entropy changes.

6.10.17. Application of First and Second Laws: Chemical Equilibria. In a similar way, we can use the Second Law to predict that at thermal equilibrium the chemical potentials of the reactant and product species should be equal to each other. Consider a chemical a system in which one chemical component (A) may react to form another component (B). For example, A and B could be the gauche and trans conformational isomers of n-butane, or they could be the native and denatured states of a protein. See Fig. 6.13.

We will also assume that the container which holds the system is insulated and sealed (so neither heat nor matter can enter or leave), and that the volume of the container is constant. Thus, no work or heat exchange can take place between the system and its surroundings (so it is isolated).

Now imagine that we initially put some amounts of A and B into the system which are not equal to their equilibrium concentrations. The entropy representation of the fundamental equation implies that the total entropy
6.10. Multi-Component Gas

Figure 6.13. A closed, insulated, constant-volume, system contains molecules that can inter-convert between two different chemical states (A and B). The second law requires that the chemical potentials of the two compounds must become equal (∆A = ∆B) at equilibrium.

\[
\frac{dS(U, V, N_i)}{T} = \left(\frac{1}{T}\right) dU + \left(\frac{P}{T}\right) dV - \sum_i \left(\frac{\mu_i}{T}\right) dN_i,
\]

and since energy and volume are constant \(dU = dV = 0\),

\[
dS = -\left(\frac{\mu_A}{T_A}\right) dN_A - \left(\frac{\mu_B}{T_B}\right) dN_B.
\]

At thermal equilibrium \(T_A = T_B = T\). Moreover, stoichiometric balance requires that \(dN_B = -dN_A\),

\[
dS = -\left[\frac{\mu_A}{T} - \frac{\mu_B}{T}\right] dN_A.
\]

At equilibrium \(dS = 0\) and (provided \(dN_A \neq 0\)),

\[
\mu_A = \mu_B.
\]

Away from equilibrium \(dS > 0\), so if \(\mu_A > \mu_B\) then we must have \(dN_A < 0\) (reactants are consumed), as would be expected. Thus, the Second Law correctly predicts the direction of change in a chemical reaction.

6.10.18. Generalization to Arbitrary Reactions. The above result may be extended to chemical reactions involving more complex stoichiometry,

\[
\sum_{i=1}^{n_R} a_i A_i \rightleftharpoons \sum_{i=1}^{n_P} b_i B_i
\]

where \(a_i\) and \(b_i\) are stoichiometric coefficients for the reactant \(A_i\) and product \(B_i\) species, respectively. \(n_R\) is the number of reactants. \(n_P\) is the number of products.

The chemical reaction can be written as an equation, where \(\{A_i\}\) denote the set of all reactants and products, \(\{\nu_i\}\), the set of stoichiometric coefficients
with the convention that \( \nu_i \) are positive numbers for the reactants \((\nu_i = a_i)\) and negative numbers for the products \((\nu_i = -b_i)\),

\[
\sum_{i=1}^{n_S} \nu_i A_i = 0,
\]

with \( n_S = n_R + n_P \).

For any such reaction one may use the following equation to express the relationship between the chemical potentials of the reacting species at equilibrium:

<table>
<thead>
<tr>
<th>Chemical Equilibrium:</th>
</tr>
</thead>
</table>
| \[
\sum_{i=1}^{n_S} \nu_i \mu_i = 0 \text{ at equilibrium.}
\] |

This result follows from the same analysis as in the previous section. There, we would also have that at thermal equilibrium, all temperatures are the same, \( T_i = T \) for all \( i \).

The \( dN_i \) are parametrized by a new variable \( t \), i.e. so that \( dN_i(t) = N'_i(t) dt \), where \( t \in [0, 1] \) is a variable that denotes the “extent of the reaction”. Then \( t = 0 \) the reaction begins and when \( t = 1 \), it ends. \( N'_i(t) \) is the speed at which species \( i \) is produced or depleted; it is proportional to the stoichiometric coefficient \( \nu_i \), so we write \( N'_i(t) = \alpha \nu_i \), where \( \alpha \) is a constant.

Under those conditions, the chemical term in \( dS \) reads:

\[
dS = - \left[ \sum_{i=1}^{n_S} \left( \frac{\mu_i}{T_i} \right) \alpha \nu_i \right] dt = - \frac{\alpha}{T} \left[ \sum_{i=1}^{n_S} \mu_i \nu_i \right] dt.
\]

At equilibrium, \( dS = 0 \) (provided \( dt \neq 0 \)), from which it follows that the coefficient of \( dt \) must vanish, i.e. \( \sum_{i=1}^{n_S} \nu_i \mu_i = 0 \).

6.10.19. Application of First and Second Laws: Phase Equilibrium. Consider a system composed of two sub-systems whose boundary allows the exchange of molecules, as well as heat and mechanical work. For example, one sub-system could be a vapor phase A and the other a liquid phase B, as shown in Fig. 6.14. In such a situation, we may again invoke the Second Law to obtain general relations between all of the intensive variables of the two sub-systems at equilibrium.

For simplicity we will assume there is only one chemical component in the system (such as water). We assume that the number of water molecules in the vapor phase is \( N_A \) and in the liquid phase is \( N_B \).

---

\(^{22}\) If \( t \) is time, then \( dN_i \) denotes the change in number of molecules of species \( i \) whereas \( N'_i(t) dt \) denotes the rate of change in \( N_i \) times \( dt \).
6.10. Multi-Component Gas

Figure 6.14. A closed, insulated, constant-volume, system contains two phases (A and B). The second law requires that at equilibrium the temperature, pressure and chemical potentials of the two phases will necessarily become equal \((T_A = T_B, \ P_A = P_B \text{ and } \mu_A = \mu_B)\).

Note that when molecules are exchanged between the two phases the volumes \(V_A\) and \(V_B\), as well as the energies \(U_A\) and \(U_B\), of each phase will also change. We further assume that both phases are enclosed by a sealed and insulated container of fixed total volume (so the entire system is isolated from its surroundings).

When we apply the fundamental equation to each phase we obtain the following expressions.

\[
\begin{align*}
\text{d}S_A &= \left(\frac{1}{T_A}\right) \text{d}U_A + \left(\frac{P_A}{T_A}\right) \text{d}V_A - \left(\frac{\mu_A}{T_A}\right) \text{d}N_A \\
\text{d}S_B &= \left(\frac{1}{T_B}\right) \text{d}U_B + \left(\frac{P_B}{T_B}\right) \text{d}V_B - \left(\frac{\mu_B}{T_B}\right) \text{d}N_B.
\end{align*}
\]

The entropy change of the entire system is again \(\text{d}S = \text{d}S_A + \text{d}S_B\). Since the total energy, volume and number of molecules are all conserved, we may equate \(\text{d}U_A = -\text{d}U_B = \text{d}U\), \(\text{d}V_A = -\text{d}V_B = \text{d}V\), and \(\text{d}N_A = -\text{d}N_B = \text{d}N\). Thus, an infinitesimal entropy change of the whole system, resulting from the transfer of some molecules between phase A and phase B, may be expressed as follows.

\[
\text{d}S = \left(\frac{1}{T_A} - \frac{1}{T_B}\right) \text{d}U + \left(\frac{P_A}{T_A} - \frac{P_B}{T_B}\right) \text{d}V - \left(\frac{\mu_A}{T_A} - \frac{\mu_B}{T_B}\right) \text{d}N.
\]

At equilibrium \(\text{d}S = 0\) when \(\text{d}U \neq 0\) and \(\text{d}V \neq 0\) and \(\text{d}N \neq 0\), so all three quantities in parentheses must equal to zero. Thus implies that all three of the following intensive variables of the two phases must be equal at equilibrium:

\[
T_A = T_B, \quad P_A = P_B, \quad \mu_A = \mu_B.
\]
In fact, you can check that

\[ T_A = T_B + dT, \quad P_A = P_B + dP, \quad \mu_A = \mu_B + d\mu, \]

leads to equilibrium \((dS = 0)\).

In any two phase equilibrium all of the intensive variables of the two phases – temperatures, pressures and chemical potentials – must be in perfect balance with each other.

If the system contained more than one chemical species then similar expressions would hold for each chemical species. Moreover, if the system contained molecules that can chemically react with each other, then \(\sum_i \nu_i \mu_i = 0\) would impose an additional equilibrium condition which must hold for any such reactive species.


Gibbs energy of mixing. Suppose that we have two perfect gases, \(A\) and \(B\), in two containers, in the amounts \(n_A\) and \(n_B\), respectively. Both are at temperature \(T\) and pressure \(P\). The chemical potentials of the two gases have their “pure” values,

\[ \mu = \mu^0 + RT \log(P/P_0) \]

where \(\mu^0\) is the standard chemical potential, the chemical potential of the pure gas at 1 bar. See Fig. 6.15.

![Figure 6.15. Arrangement for calculating the thermodynamic functions of mixing of two perfect gases.](image)

Taking \(P_0\) to be 1 bar and writing \(P\) instead of \(P/P_0\), the Gibbs energy of the total system is (before mixing):

\[ G_i = n_A \mu_A + n_B \mu_B = n_A (\mu_A^0 + RT \log P) + n_B (\mu_B^0 + RT \log P). \]
After mixing, the partial pressures are $P_A$ and $P_B$ ($P_A + P_B = P$). The total Gibbs energy changes to:

$$G_f = n_A(\mu_A^0 + RT \log P_A) + n_B(\mu_B^0 + RT \log P_B)$$

The difference, $G_f - G_i$, the Gibbs energy of mixing, $\Delta G_{mix}$, is therefore

$$\Delta G = G_f - G_i = n_A RT \log(P_A/P) + n_B RT \log(P_B/P).$$

Replacing $n_A$ by $X_A n$, where $n$ is the total number of moles of $A$ and $B$, and using Dalton’s law, $P_A = X_A \cdot P$, we have:

$$\Delta G_{mix} = nRT(X_A \log X_A + X_B \log X_B).$$

Because $0 \leq X_A, X_B \leq 1$, the logs are negative and $\Delta G_{mix} < 0$. Thus, perfect gases mix spontaneously in all proportions. And from $dG = -SdT + VdP$, $S = -(\partial G/\partial T)$, the entropy of mixing is:

$$\Delta S_{mix} = -nR\{X_A \log X_A + X_B \log X_B\}.$$

Note: since $\Delta G = \Delta H - T \Delta S$, it follows that $\Delta H_{mix} = 0$. This is expected for a system in which there are no interactions between the molecules forming the gaseous mixture. It follows that the driving force for mixing comes from the increase in entropy of the system because the entropy of the surroundings is unchanged.

6.10.20.1. **Example: (Gibbs energy of mixing).** A container is divided into two equal compartments (Fig. 6.16). One contains 3.0 mol H$_2$(g) at 25°C. The other contains 1.0 mol N$_2$(g) at 25°C. Calculate the Gibbs energy of mixing when the partition is removed. Assume perfect behavior.

![Figure 6.16](image-url) The initial and final states considered in the calculation of the Gibbs energy of mixing of gases at different initial pressures.

*Solution:* The formula from the previous section cannot be used directly because the initial gas pressures are different. Given that the pressure of
nitrogen is $P$ and the pressure of hydrogen is $3P$, the initial Gibbs energy is

\[
G_i = (3.0 \text{ mol})\{\mu(\text{H}_2) + RT \log 3P\} + (1.0 \text{ mol})\{\mu(\text{N}_2) + RT \log P\}
\]

When the partition is removed and each gas occupies twice the original volume, the partial pressure of nitrogen falls to $\frac{1}{2}P$ and that of hydrogen falls to $\frac{3}{2}P$. Therefore, the Gibbs energy changes to

\[
G_f = (3.0 \text{ mol})\{\mu(\text{H}_2) + RT \log \frac{3}{2}P\} + (1.0 \text{ mol})\{\mu(\text{N}_2) + RT \log \frac{1}{2}P\}
\]

Taking the difference gives the Gibbs energy of mixing:

\[
\Delta G_{\text{mix}} = G_f - G_i = (3.0 \text{ mol})RT \log \left(\frac{\frac{3}{2}P}{3P}\right) + (1.0 \text{ mol})RT \log \left(\frac{\frac{1}{2}P}{P}\right)
\]

\[
= - (3.0 \text{ mol})RT \log 2 - (1.0 \text{ mol})RT \log 2
\]

\[
= - (4.0 \text{ mol})RT \log 2 = -6.9 \text{ kJ}
\]

Here, $\Delta G_{\text{mix}} < 0$, however, the negative sign does not necessarily indicate spontaneity. $dG < 0$ indicates spontaneity only at constant temperature and pressure. (Here, the pressure dropped from $4P$ to $2P$ when opening the partition.)

Problem. Suppose that 2.0 mol H\textsubscript{2} at 2.0 atm and 25°C and 4.0 mol N\textsubscript{2} at 3.0 atm and 25°C are mixed at constant volume. Calculate $\Delta G_{\text{mix}}$. What would be the value of $\Delta G_{\text{mix}}$ had the pressures been identical initially? [-9.7 kJ, -9.5 kJ]

---

**Figure 6.17.** At equilibrium, the chemical potential of the gaseous form of a substance $A$ is equal to the chemical potential of its condensed phase. The equality is preserved if a solute is also present. Because the chemical potential of $A$ in the vapor depends on its partial vapor pressure, it follows that the chemical potential of liquid $A$ can be related to its partial vapor pressure.

---

**6.10.21. Ideal Solutions & Raoult’s Law.** Suppose we have an ideal solution made up of two components, $A$ and $B$ (Fig. 6.17). Let us denote
with a $^*$ the quantities relating to pure substances. Let $\mu^*_A(l)$ denote the chemical potential of pure $A$ (liquid). Vapor pressure of the pure liquid is $P^*_A$. Chemical potential in the vapor is $\mu^*_A + RT \log P^*_A$. These two chemical potentials are equal at equilibrium:

$$\mu^*_A(l) = \mu^*_A + RT \log P^*_A.$$  \hspace{1cm} \text{(pure A)}

If another substance, a solute, is also present in the liquid, the chemical potential of $A$ in the liquid is changed to $\mu_A(l)$ and its vapor pressure is changed to $P_A$. The vapor and solvent are still in equilibrium, so we can write

$$\mu_A(l) = \mu^*_A + RT \log P_A.$$ \hspace{1cm} \text{(add some B)}

Next, we combine these two equations to eliminate the standard chemical potential of the gas. To do so, we solve for $\mu^*_A$ in both equations an equate the two results to obtain

$$\mu_A(l) = \mu^*_A(l) - RT \log P_A^* + RT \log P_A = \mu^*_A(l) + RT \log(P_A/P_A^*).$$

Substituting Raoult’s law, $P_A = X_A \cdot P^*_A$, we get:

$$\mu_A(l) = \mu^*_A(l) + RT \log X_A.$$  

This important equation can be used as the definition of an ideal solution (so that it implies Raoult’s law rather than stemming from it) and is better, because it does not assume that the vapor is a perfect gas.

### 6.10.22. Liquid Mixtures: Ideal Solutions.

Suppose that we mix two liquids $A$ and $B$. Before mixing, we have:

$$G_i = n_A \mu^*_A + n_B \mu^*_B.$$  

When they are mixed, the individual chemical potentials are modified, and the total Gibbs energy is,

$$G_f = n_A \{ \mu^*_A + RT \log X_A \} + n_B \{ \mu^*_B + RT \log X_B \}.$$  

Consequently, the Gibbs energy of mixing, $G_f - G_i$, is

$$\Delta G_{\text{mix}} = nRT \{ X_A \log X_A + X_B \log X_B \}.$$  

And from $dG = -SdT + VdP$, $S = -(\partial G/\partial T)$, the entropy of mixing is:

$$\Delta S_{\text{mix}} = -nR \{ X_A \log X_A + X_B \log X_B \}.$$  

### 6.10.23. Elevation of Boiling Point.

The vapor pressure of a pure liquid represents a balance between the increase in disorder arising from vaporization and the decrease in disorder of the surroundings. See Figs. 6.18 and 6.19.
PROOF of Gibbs-Helmholtz: on one hand, 
\[ - \Delta_{vap} T \] 
For constant enthalpy of vaporization, \( \log \left( \frac{1}{X} \right) \) integrates from \( T_1 \) to \( T \), and therefore, 
\[ \int_{T_1}^{T} d\log X_A = \frac{1}{T} \int_{T_1}^{T} \Delta_{vap} H \, dT. \] 
The left hand side integrates to \( \log X_A = \log(1 - X_B). \)

For constant enthalpy of vaporization, \( \log(1 - X_B) = -\frac{\Delta_{vap} H}{R} \int_{T_1}^{T} \frac{1}{T^2} dT, \) and therefore, 
\[ \log(1 - X_B) = -\Delta_{vap} H \left( \frac{1}{T^2} - \frac{1}{T_1^2} \right). \] 
We now suppose that the amount of solute present is so small that \( X_B \ll 1 \). We can write \( \log(1 - X_B) \approx -X_B \) and obtain \( X_B = \frac{\Delta_{vap} H}{R} \left( \frac{1}{T^2} - \frac{1}{T_1^2} \right). \)

Finally, because \( T \approx T^* \), it follows that \( \frac{1}{T^2} - \frac{1}{T} = \frac{T^* - T}{T^2} \approx \frac{\Delta T}{T^2} \), with \( \Delta T = T - T^*. \) 

PROOF of Gibbs-Helmholtz: on one hand, \( -S = (\partial G/\partial T)_P = (G - H)/T \), on the other hand, 
\[ \left( \frac{\partial G}{\partial T} \right)_P = \frac{1}{T} \left( \frac{\partial G}{\partial T} \right)_P + G \left( \frac{\partial G}{\partial T} \right)_P = \frac{1}{T} \left( \frac{\partial G}{\partial T} \right)_P - T \frac{G}{T} = \frac{1}{T} \left( \left( \frac{\partial G}{\partial T} \right)_P - T \frac{G}{T} \right). \] 
In the curly bracket, we substitute \( T \frac{G}{T} = G - H \) and \( G = -\frac{H}{T} \). Gibbs-Helmholtz follows.

The heterogeneous equilibrium of interest when considering boiling is between the solvent vapor and the solvent in solution at 1 atm. We denote the solvent by \( A \) and the solute by \( B \). The equilibrium is established at a temperature for which 
\[ \mu_A^*(g) = \mu_A^*(l) + RT \log X_A. \] 
(The pressure of 1 atm is the same throughout and is not written explicitly.) 
This equation implies that the presence of a solute at a mole fraction \( X_B \)
causes an increase in normal boiling point from $T^*$ to $T^* + \Delta T$, where

$$\Delta T = K \cdot X_B, \quad K = \frac{RT^*^2}{\Delta_{vap}H}.$$  

**Figure 6.20.** Heterogeneous equilibrium involved in the calculation of the lowering of freezing point is between $A$ in the pure solid and $A$ in the mixture, $A$ being the solvent and $B$ a solute that is insoluble in solid $A$.

### 6.10.24. Depression of Freezing Point.

The heterogeneous equilibrium now of interest is between pure solid solvent $A$ and the solution with solute present at a mole fraction $X_B$ (Fig. 6.20). At the freezing point, the chemical potentials of $A$ in the two phases are equal:

$$\mu_A^*(s) = \mu_A^*(l) + RT\log X_A$$

The calculation is the same as in the previous section, and the result is

$$\Delta T = K' \cdot X_B, \quad K' = \frac{RT^*^2}{\Delta_{fus}H},$$

where $\Delta T$ is the freezing point depression, $T^* - T$, and $\Delta_{fus}H$ is the enthalpy of fusion of the solvent. Larger depressions are observed in solvents with low enthalpies of fusion at high melting points. When the solution is dilute, the mole fraction is proportional to the molality of the solute, $b$, and it is common to write the last equation as

$$\Delta T = K_f \cdot b,$$

where $K_f$ is the empirical freezing-point constant.

### 6.10.25. Solubility.

When a solid solute is left in contact with a solvent, it dissolves until the solution is saturated. Saturation is a state of equilibrium, with the undissolved solute in equilibrium with the dissolved solute. Therefore, in a saturated solution the chemical potential of the pure solid solute, $\mu_B^*(s)$, and the chemical potential of $B$ in solution, $\mu_B$, are equal (Fig. 6.21).

Because the latter is $\mu_B = \mu_B^*(l) + RT\log X_B$, we can write

$$\mu_B^*(s) = \mu_B^*(l) + RT\log X_B.$$
This expression is the same as the starting expression in the last section, except that the quantities refer to the solute \( B \), not the solvent \( A \). It leads to:

\[
\log X_B = \frac{\Delta f_{us} H}{R} \left( \frac{1}{T_f} - \frac{1}{T} \right).
\]

6.10.26. Example: (Ion transport Across Cell Membrane). There are ions (e.g. \( \text{Na}^+ \), \( \text{Ca}^{2+} \), \( \text{K}^+ \), etc.) found in the intra and extra-cellular spaces at different concentrations. Cell membrane proteins participate in the transport of such ions. As an example, let us analyze the chemical equilibrium of \( \text{K}^+ \) ions using the first two laws of thermodynamics. Recall that for a chemical reaction:

\[
\Delta G_{rxn} = \Delta G_{rxn}^0 + RT \log Q
\]

where \( \Delta G_{rxn}^0 = -RT \log K_{eq} \). This is valid when a chemical reaction is the only “work” done. However, there may be other terms in the expression for \( \Delta G \) that need to be accounted for. Consider the “chemical reaction” where a potassium ion is shuttled in and out of the cell cytoplasm:

\[
\text{K}_0^+ \rightleftharpoons \text{K}_i^+
\]

Chemical reaction equilibrium alone would state that

\[
\mu_i - \mu_0 = RT \log \left( \frac{[\text{K}^+]_i}{[\text{K}^+]_o} \right)
\]

Where \([\text{K}^+]_i\) is the potassium ion concentration inside the cell and \([\text{K}^+]_o\) is its concentration found outside (in the extracellular matrix). The difference \( \mu_i - \mu_0 \) represents gradient (difference) in the chemical potential across the cell membrane. This chemical potential gradient acts as a driving force to shuttle ions in and out of the cell (depending on whether \([\text{K}^+]_i > [\text{K}^+]_o\) or \([\text{K}^+]_i < [\text{K}^+]_o\)). See Fig. 6.22.

However, this cannot be the only force because if we shuttle all ions out of the cell (for instance), the electrostatic repulsion among potassium ions in
the extracellular matrix will be large enough to drive some of these ions to re-enter the cytoplasm. Thus, we must include electrostatic repulsions.

Electrostatic repulsions can be accounted for by adding electrostatic work. Recall that the electrostatic work is of the form $\delta W_{el.} = \phi dq$ where $dq$ is the charge added or removed, and $\phi$ is the local electrostatic potential. Here we write $\phi_o$ for the electrostatic potential in the extracellular matrix and $\phi_i$ for the potential inside the cytoplasm. The total work is obtained by integrating from reactants to products. As far as the integral $\Delta W_{el.} = \int_{\text{reactants}}^{\text{products}} \phi dq$ is concerned, going from reactants to products means we subtract a charge $Q$ from the extracellular space (work = $-Q\phi_0$) and add it to the intracellular space (work = $+Q\phi_i$).

The total integral along this path yields the sum of these two terms: one from the extracellular matrix, and one for the cytoplasm:

$$\Delta W_{el.} = Q(\phi_i - \phi_o),$$

where $\phi_0$ acquires a negative sign because we are removing an ion of charge $Q$ from the extracellular matrix (and add it to the cytoplasm). The difference $\phi_i - \phi_o$ is the potential difference across the cell membrane (units: volts). When the electrostatic force balances out the chemical force, $\Delta G = 0$ and these two terms are equal:

$$\Delta \phi \equiv \phi_i - \phi_o = \frac{RT}{Q} \log \left( \frac{[K^+]_i}{[K^+]_o} \right).$$

This potential difference, $\Delta \phi \equiv \phi_i - \phi_o$ (in volts), is called the Nernst potential. Inspection of this equation shows that $Q$ must be specified in units of Coulomb per mole; $Q$ in these units is often called the Faraday’s constant, $F \equiv 9.65 \times 10^4$ C/mol. (Otherwise, replace $RT$ by $k_B T$ and $Q$ is then specified in units of Coulomb.)
Please note that the above analysis is valid in thermal equilibrium (steady state) and does not describe dynamics, such as those observed in cardiac tissues or neuronal networks.

There are ways to directly measure ion concentrations and voltages. For a discussion of Ca\(^{2+}\) and voltage sensing via specially-designed fluorophores, see the paper [25].

### 6.10.27. Affinity of Reaction and Reaction Enthalpy

General chemical reactions are represented by an equation, \(\sum_i \nu_i A_i = 0\), where the stoichiometric coefficients \(\nu_i\) are negative for reactants and positive for products. For example:

\[
0 = -N_2 - 3H_2 + 2NH_3.
\]

The *extent of reaction* \(\xi\) is defined by:

\[
\xi = \frac{n_i(\xi) - n_i(0)}{\nu_i},
\]

where \(n_i(\xi)\) is the amount of substance \(A_i\) present when the extent of reaction is \(\xi\). \(n_i(0)\) is the amount of \(A_i\) present when \(\xi = 0\), corresponding to the specified initial conditions. The units of \(\xi\) are moles. For example, if 1 mol of \(N_2\) and 3 mol of \(H_2\) are converted completely into 2 mol of \(NH_3\)

\[
\xi = \frac{n_{NH_3}(\xi) - n_{NH_3}(0)}{\nu_{NH_3}} = \frac{2 \text{ mol} - 0 \text{ mol}}{2} = 1 \text{ mol}
\]

\[
\xi = \frac{n_{N_2}(\xi) - n_{N_2}(0)}{\nu_{N_2}} = \frac{0 \text{ mol} - 1 \text{ mol}}{-1} = 1 \text{ mol}
\]

\[
\xi = \frac{n_{H_2}(\xi) - n_{H_2}(0)}{\nu_{H_2}} = \frac{0 \text{ mol} - 3 \text{ mol}}{-3} = 1 \text{ mol}
\]

\(\xi\) is independent of which chemical is used to compute it. Note also that

\[
d\xi = \frac{dn_i}{\nu_i}.
\]

Using this notation, the fundamental equation

\[
dG = -SdT + VdP + \sum_i \mu_i dn_i,
\]

becomes:

\[
dG = -SdT + VdP + \sum_i \mu_i \nu_i d\xi.
\]

If we define the *reaction affinity* by:

\[
A = - \sum_i \mu_i \nu_i.
\]
Then \( dG \) is written:
\[
dG = -SdT + VdP - Ad\xi.
\]

From this, we see that
\[
A = -\left( \frac{\partial G}{\partial \xi} \right)_{T,P}.
\]

For the special case of constant \( T, P \), we see that \( dG = -Ad\xi \). Since \( G \) is a state function, integrating from \( \xi = 0 \) to \( \xi = 1 \) mol yields:
\[
A = -\Delta G_{rxn} = -\sum_i \nu_i \mu_i \geq 0.
\]

(This assumes that \( A \) is independent of \( \xi \).)

In a similar manner, the fundamental equations for the thermodynamic potentials \( U, S, H, A \):
\[
dU = TdS - PdV + \sum_i \mu_idn_i \quad \to \quad dU = TdS - PdV - Ad\xi
\]
\[
dS = \frac{1}{T}dU + \frac{P}{T}dV - \frac{1}{T}\sum_i \mu_idn_i \quad \to \quad dS = \frac{1}{T}dU + \frac{P}{T}dV + \frac{A}{T}d\xi
\]
\[
dH = TdS + VdP + \sum_i \mu_idn_i \quad \to \quad dH = TdS + VdP - Ad\xi
\]
\[
dA = -SdT - PdV + \sum_i \mu_idn_i \quad \to \quad dA = -SdT - PdV - Ad\xi
\]
\[
dG = -SdT + VdP + \sum_i \mu_idn_i \quad \to \quad dG = -SdT + VdP - Ad\xi
\]

**6.10.28. Enthalpy is the “Heat Content”**. Recall the second law, \( dS_{univ} = dS + dS_{ext} \geq 0 \), where \( S \) is the system and \( S_{ext} \) is the surroundings. Now consider the system \( S \); the total entropy production for this system can be decomposed into two terms:
\[
dS = d_eS + d_iS.
\]

The first term, \( d_eS \), is the entropy produced by the interaction of the system with its surroundings, e.g.
\[
d_eS = \frac{\delta Q}{T}. \quad \text{(if process is reversible)}
\]

The second term, \( d_iS \), is the entropy produced by irreversible processes going on inside the system (for example, chemical reactions). **When the only interaction with a system’s surroundings is the absorption of heat \( \delta Q \), and the only irreversible process going on within the system is a chemical reaction, we can write explicit expressions for \( d_eS \) and \( d_iS \):**
The fundamental equation for \( dS \) reduces to:

\[
dS = \frac{\delta Q}{T} + \frac{A}{T} d\xi.
\]

Thus,

\[
d_{c} S = \frac{\delta Q}{T}, \quad d_{i} S = \frac{A d\xi}{T}.
\]

These two contributions describe the measurable heat transfer and the irreversible heat (absorbed or released) of the reaction. *We note that the term \( d_{i} S \) can describe a bona fide chemical reaction or the transfer of matter to another subsystem* (essentially, any process that can be described using chemical potentials). When matter is transported, this causes a transfer of heat.

Consider the complete transformation of stoichiometric amounts of reactants at \( T, P \) into products at \( T, P \), so that \( \Delta \xi = 1 \) mol. Because \( H \) is a function of the state of the system, the enthalpy change due to the reaction

\[
\Delta H_{\text{rxn}} = (\Delta Q)_{P}
\]

is equal to the heat absorbed at constant pressure. This quantity was formerly called the *heat of reaction at constant pressure*, but is now called the *reaction enthalpy*. Another historic name for enthalpy is *heat content*. The subscript \( \text{rxn} \) tells us that this is the enthalpy change during the reaction. \( \Delta H_{\text{rxn}} \) is the change in an extensive property, but numerical data always refer to \( \Delta \xi = 1 \) mol so that the units of \( \Delta H_{\text{rxn}} \) are J/mol, or more commonly kJ/mol.

The idea that enthalpy is the heat transfer at constant pressure is illustrated by the fundamental equation for \( H \):

\[
(dH)_{P} = T dS + V dP - A d\xi = T dS - A d\xi
\]

which, at constant pressure, features two terms, \( T dS \), the measurable heat transfer and \( -A d\xi \), the enthalpy associated with mass transport or irreversible chemical reaction. It is because of this property that enthalpy is called *heat content*.

**6.10.29. Forces and Fluxes.** The fundamental equation in the entropy representation is particularly important because of its special structure:

\[
dS = \left( \frac{1}{T} \right) dU + \left( \frac{P}{T} \right) dV - \sum_{i} \left( \frac{\mu_{i}}{T} \right) dn_{i}.
\]

It has the structure of a sum of terms, each of which represent a force times a flux. The forces are:

- \( \frac{1}{T} \): tendency for heat flow
6.10. Multi-Component Gas

- $\frac{P}{T}$: tendency for volume change
- $\frac{\mu}{T}$: tendency for particle exchange

Figure 6.23. A closed, insulated, constant-volume, system contains two phases (A and B). The second law requires that at equilibrium the temperature, pressure and chemical potentials of the two phases will necessarily become equal ($T_A = T_B$, $P_A = P_B$ and $\mu_A = \mu_B$).

Recall our previous analysis of phase equilibrium (Fig. 6.23), which begins by writing down the First Law for each subsystem $dS = dS_A + dS_B$ in the entropy representation ($A + B$ is a closed system, so that $dU = dU_A = -dU_B$, $dV = dV_A = -dV_B$, $dn = dn_A = -dn_B$):

$$dS = \left( \frac{1}{T_A} - \frac{1}{T_B} \right) dU + \left( \frac{P_A}{T_A} - \frac{P_B}{T_B} \right) dV - \left( \frac{\mu_A}{T_A} - \frac{\mu_B}{T_B} \right) dn.$$

dS is called the dissipation function because $dS > 0$ indicates an irreversible process where energy is dissipated (cannot be converted into useful work).

Notice that all the terms in the expression for $dS$ are of the form of a gradient in intensive property ($1/T, P/T, \mu/T$) times the differential of an extensive property ($U, V, n$). The former is called a force; the latter is called a flux:

$$\text{dissipation function} = \text{force} \times \text{flux} = X \cdot J$$

6.10.30. Internal vs External Variables: Heat Transfer Only. A system with only heat transfer can be exemplified as consisting of two pieces of metal at different absolute temperatures, $T_1$ and $T_2$, where $T_1 > T_2$. A small quantity of heat, $\delta Q$, is allowed to pass from subsystem 1 to subsystem 2 in course of a short time interval, $dt$ (Fig. 6.24).

The quantity of heat is so small that changes in the temperatures of the subsystems can be disregarded: Since there is no transfer of matter, each
subsystem behaves as a closed system. The transfer of $\delta Q$ leads to the following changes in the subsystems:

$$dU_1 = T_1 dS_1 - P_1 dV_1 \quad dU_2 = T_2 dS_2 - P_2 dV_2,$$

or with the equations solved with respect to entropy change:

$$dS_1 = \frac{dU_1}{T_1} + \left(\frac{P_1}{T_1}\right) dV_1 \quad dS_2 = \frac{dU_2}{T_2} + \left(\frac{P_2}{T_2}\right) dV_2.$$

When $P-V$ work is the only work carried out, $dU_1$ and $dU_2$ can be expressed by the external variables, which are measurable quantities,

$$dU_1 = \delta Q_1 - P_1 dV_1 \quad dU_2 = \delta Q_2 - P_2 dV_2.$$

The total system is adiabatic,

$$\delta Q = -\delta Q_1 = \delta Q_2,$$

so that

$$dU_1 + dU_2 = -P_1 dV_1 - P_2 dV_2.$$

Total entropy production is:

$$dS = dS_1 + dS_2 = \frac{\delta Q}{T_1} + \frac{\delta Q}{T_2} = \left(\frac{1}{T_2} - \frac{1}{T_1}\right) \delta Q = \Delta(1/T)\delta Q,$$

or

$$\frac{dS}{dt} = \Delta(1/T) \frac{\delta Q}{dt}.$$

Multiplying by $T_2$ gives the dissipated energy

$$T_2 \frac{dS}{dt} = T_2 \Delta(1/T) \frac{\delta Q}{dt},$$

where $\delta Q/dt$ is the rate of transport of heat. The units of $T_2 \frac{dS}{dt}$ and $\delta Q/dt$ are Watts (1 W=1 J/s).

#### 6.10.31. Internal vs External Variables: Heat Transfer + Matter Transport

A system with transfer of heat and matter can be exemplified by two gas containers at different absolute temperatures, $T_1$ and $T_2$, where $T_1 > T_2$ (Fig. 6.25).
A small quantity of heat, $\delta\Phi$, and small quantities of the different gas components, $dn_i$, are allowed to pass from subsystem 1 to subsystem 2 in course of the time $dt$. The total heat transferred, $\delta\Phi$, is composed of a measurable heat and the enthalpy of the transferred matter. Since the total system is adiabatic, the total heat removed from subsystem 1 is equal to the total heat received by subsystem 2. The enthalpy of the transferred matter, however, may be different in the two subsystems, and thus the measurable heat change will be different in the two subsystems.

The changes in the subsystems are:

$$\delta\Phi = \delta\Phi_2 = -\delta\Phi_1.$$  

In terms of internal variables,

$$dU_1 = T_1 dS_1 - P_1 dV_1 + \sum_i \mu_{i,1} dn_{i,1} \quad dU_2 = T_2 dS_2 - P_2 dV_2 + \sum_i \mu_{i,2} dn_{i,2}.$$  

Solving for $dS$,

$$dS_1 = \frac{dU_1}{T_1} + \frac{P_1}{T_1} dV_1 - \frac{1}{T_1} \sum_i \mu_{i,1} dn_{i,1}, \quad dS_2 = \frac{dU_2}{T_2} + \frac{P_2}{T_2} dV_2 - \frac{1}{T_2} \sum_i \mu_{i,2} dn_{i,2}.$$  

In terms of external variables, when we only have $P - V$ work:

$$dU_1 = \delta\Phi_1 - P_1 dV_1 \quad dU_2 = \delta\Phi_2 - P_2 dV_2,$$

$$dU_1 + dU_2 = -P_1 dV_1 - P_2 dV_2.$$  

Since $\delta\Phi = \delta\Phi_2 = -\delta\Phi_1$. Then, plugging Eq. (6.6) into (6.5),

$$dS_1 = -\frac{\delta\Phi}{T_1} + \frac{1}{T_1} \sum_i \mu_{i,1} dn_i, \quad dS_2 = \frac{\delta\Phi}{T_2} - \frac{1}{T_2} \sum_i \mu_{i,2} dn_i.$$  

The production of entropy in the total system is $dS = dS_1 + dS_2$ and therefore

$$dS = \left(\frac{1}{T_2} - \frac{1}{T_1}\right) d\Phi + \sum_i \left(\frac{\mu_{i,2}}{T_2} - \frac{\mu_{i,1}}{T_1}\right) dn_i,$$

$$dS = \Delta(1/T)d\Phi - \sum_i \Delta(\mu/T)dn_i.$$  

**Figure 6.25.** Both heat and matter can transfer.
Multiplying by \( T_2 \) and dividing by time \((dt)\), we have the dissipated energy \((J/s)\)

\[
T_2 \frac{dS}{dt} = T_2 \Delta(1/T) \frac{d\Phi}{dt} - T_2 \sum_i \Delta(\mu_i/T) \frac{dn_i}{dt}.
\]

It is possible to develop Eq. (6.8) further and put it in a more useful form (see Section 6.10.32):

\[
dS = \Delta(1/T) \delta Q_1 - \frac{1}{T} \sum_i \Delta \mu_{i,T} dn_i.
\]

### 6.10.32. Algebraic Details

Starting from Eq. (6.7),

\[
dS_1 = -\frac{\delta \Phi}{T_1} + \frac{1}{T_1} \sum_i \mu_{i,1} dn_i, \quad dS_2 = \frac{\delta \Phi}{T_2} - \frac{1}{T_2} \sum_i \mu_{i,2} dn_i,
\]

and solving for \( \delta \Phi \), we have:

\[
\delta \Phi = -T_1 dS_1 + \sum_i \mu_{i,1} dn_{i,1} = T_2 dS_2 + \sum_i \mu_{i,2} dn_{i,2}.
\]

The entropy of an open system can be expressed as a function of temperature, pressure and amounts of substances, \( S = S(T, P, \{n_i\}) \). The experiment can be arranged such that the pressure within each subsystem is kept constant. The change in entropy with temperature and amounts of substances for subsystem 1 can be written,

\[
dS_1 = \frac{\partial S_1}{\partial T} \frac{dT}{P,n_i} + \sum_i \left( \frac{\partial S_1}{\partial n_{i,1}} \right)_{P,T,n_{i,j}\neq i} dn_{i,1}
\]

\[
=(C_{P,1}/T_1)dT + \sum_i S_{i,1} dn_{i,1},
\]

and in a similar way for subsystem 2. The heat capacity of the system multiplied by the change in temperature is equal to the measurable heat absorbed by the system, \( C_{P,1}dT = -\delta Q_1 \) where \( \delta Q_1 \) is the heat removed from subsystem 1. Here it is assumed that \( dT \) is so small that changes in \( C_{P,1} \) can be neglected. When the amounts of substances transferred from subsystem 1 to subsystem 2 are small, the change in partial molar entropy, \( S_{i,1} \) can be neglected. Thus, for subsystem 1,

\[
dS_1 = -\delta Q_1/T_1 - \sum_i S_{i,1} dn_i.
\]

This equation combines with Eq. (6.7) for \( dS_1 \) to give

\[
\delta \Phi = \delta Q + \sum_i (\mu_{i,1} + T_1 S_{i,1}) dn_i,
\]
or, since $G = H - TS$, the partial molar quantity at constant $T$ is:

$$G_i \equiv \mu_i = \left( \frac{\partial G}{\partial n_i} \right)_T = \left( \frac{\partial H}{\partial n_i} \right)_T - T \left( \frac{\partial S}{\partial n_i} \right)_T \equiv H_i - TS_i.$$  

(6.9)  \hspace{1cm} \delta \Phi = \delta Q_1 + \sum_i H_{i,1}dn_i.

In a similar way, we can express $d\Phi$ by the heat absorbed by subsystem 2, $\delta Q_2$, and the enthalpy added to subsystem 2 by the transfer of $dn_i$ moles of substance:

(6.10)  \hspace{1cm} d\Phi = \delta Q_2 + \sum_i H_{i,2}dn_i.

Comparing Eq. (6.9) and (6.10) we see that the difference in $\delta Q$ values corresponds to a difference in the $H$ values:

$$\delta Q_2 - \delta Q_1 = \sum_i (H_{i,1} - H_{i,2})dn_i.$$  

For gases, $H_{i,1} - H_{i,2} \approx 0$.

The enthalpies, $H_i$, do not have absolute values, and therefore $\delta \Phi$ does not have an absolute value. We shall choose subsystem 1 as the reference state, and replace $\delta \Phi$ in Eq. (6.8) by the expression given in Eq. (6.9). The second term in Eq. (6.8), $\sum_i \Delta(\mu_i/T)dn_i$, can also be referred to subsystem 1 as the reference state. The chemical potentials are functions of temperature, pressure and composition, and for small differences in these parameters between the two subsystems, we have by the rules of derivation

$$\Delta(\mu_i/T) = \Delta(1/T)\mu_{i,1} + \frac{1}{T_1} (\partial \mu_{i,1}/\partial T)_{P,n_i} \Delta T + \frac{1}{T_1} \Delta \mu_{i,T},$$

where $\Delta \mu_{i,T}$ gives the variation in $\mu_i$ with changes in composition and pressure.

Since (from $H_i = \mu_i + TS_i$)

$$(\partial \mu_{i,1}/\partial T)_{P,n_i} = -S_{i,1},$$

and

$$\Delta(1/T) = \left( \frac{1}{T + \Delta T} - \frac{1}{T} \right) \approx -\frac{\Delta T}{T^2},$$

and

$$T_1 \approx T.$$

Therefore,

(6.11)  \hspace{1cm} \sum_i \Delta(\mu_i/T)dn_i = \Delta(1/T) \sum_i (\mu_{i,1} + TS_{i,1})dn_i + \frac{1}{T} \sum_i \Delta \mu_{i,T}dn_i.$
6. Thermodynamics

Introducing Eq. (6.9) and (6.11) into (6.8) we obtain
\[ dS = \Delta (1/T) \delta Q_1 - \frac{1}{T} \sum_i \Delta \mu_{i,T} d n_i. \]

This form may be better because it explicitly separates the two contributions to the entropy production originating from heat transfer and particle transport.

6.10.33. Gibbs Free Energy: Ideal Gas Equilibria. We have seen that the progress of a chemical reaction \( \sum_i \nu_i A_i = 0 \) can be described using the extent of the reaction, \( \xi = \frac{n_i(\xi) - n_i(0)}{\nu_i} \). From this, \( d \xi = d n_i / \nu_i \) or \( d \xi = d n_i / d t = (1/\nu_i) d n_i / d t \). The fundamental equation, \( dG = -S dT + V dP + \sum_i \mu_i d n_i \), can be rewritten as \( dG = -S dT + V dP + \sum_i \mu_i d \xi \), or \( dG = -S dT + V dP - A d \xi \), where \( A = -\Delta G_{\text{rxn}} = -\sum_i \nu_i \mu_i \).

Consider a gas-phase reaction \( A \rightleftharpoons B \). If the pressures are low enough, the gases can be considered ideal gases. Recall that for ideal gases, the chemical potentials are:
\[ \mu_A(T,P) = \mu_A^\circ(T) + RT \log \left( \frac{P_A}{1 \text{ atm}} \right) \]
\[ \mu_B(T,P) = \mu_B^\circ(T) + RT \log \left( \frac{P_B}{1 \text{ atm}} \right) \]
where \( P_A \) and \( P_B \) are partial pressures in units of 1 atm.

Inserting these expressions into the expression for \( \left( \frac{\partial G}{\partial \xi} \right) \) gives:
\[ \left( \frac{\partial G}{\partial \xi} \right) = \mu_B - \mu_A = (\mu_B^\circ + RT \log P_B) - (\mu_A^\circ + RT \log P_A) = \Delta G_{\text{rxn}}^\circ + RT \log \left( \frac{P_B}{P_A} \right), \]
where \( \Delta G_{\text{rxn}}^\circ = \mu_B^\circ - \mu_A^\circ \). Writing \( Q = P_B/P_A \), this is
\[ \Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^\circ + RT \log Q. \]

At equilibrium \( \Delta G_{\text{rxn}} = 0 \), \( Q \) is denoted by \( K \), and we have:
\[ \Delta G_{\text{rxn}} = -RT \log K + RT \log Q = RT \log (Q/K). \]

This expression, together with the second law \( \Delta G_{\text{rxn}} \leq 0 \), allows us to study the direction of a reaction.

For a general chemical reaction of the form \( \sum_i \nu_i A_i = 0 \), the quantity \( \Delta G_{\text{rxn}} \equiv \left( \frac{\partial G}{\partial \xi} \right) = \sum_i \mu_i \nu_i \), is the Gibbs free energy change of the reaction. Then,
\[ \left( \frac{\partial G}{\partial \xi} \right) = \Delta G_{\text{rxn}} = \sum_i \left[ \mu_i^0 + RT \log \left( \frac{P_i}{P_i^0} \right) \right] \nu_i \]
\[ = \Delta G_{\text{rxn}}^0 + RT \sum_i \nu_i \log \left( \frac{P_i}{P_i^0} \right), \]
where \( P^o \) is a reference pressure (usually 1 atm or 1 bar). Then,
\[
\Delta G_{rxn} = RT \log(Q/K),
\]
where
\[
Q = \prod_i (P_i/P^o)^{\nu_i}, \quad K = \prod_i (P_{eq}/P^o)^{\mu_i}, \quad \Delta G_{rxn}^0 = \sum_i \nu_i \mu_i^0 = -RT \log(K).
\]
Recall that although \( Q \) has the form of an equilibrium constant, the pressures are not necessarily at equilibrium.

The quantity \( \Delta G_{rxn}^0(T) \) is the change in standard Gibbs energy for the reaction between unmixed reactants in their standard states at temperature \( T \) and a pressure of one bar to form unmixed products in their standard states at the same temperature \( T \) and a pressure of \( P_0 \) (usually, taken to be 1 atm or 1 bar). If all pressures are referenced to 1 bar (or 1 atm), \( Q \) is unitless.

From the previous equation,
\[
\Delta G_{rxn}(T) = -RT \log(K) + RT \log(Q) = RT \log(Q/K).
\]
We see that:
- At equilibrium, \( Q = K \).
- If \( Q < K \) then \( Q \) must increase as the system moves toward equilibrium. Partial pressures of the products must increase and those of reactants must decrease. The reaction proceeds from left to right. In terms of \( \Delta G_{rxn} \), if \( Q < K \), then \( \Delta G_{rxn} < 0 \), indicating that the reaction is spontaneous from left to right as written.
- If \( Q > K \), then \( Q \) must decrease as the reaction proceeds to equilibrium. Pressures of products must decrease and those of reactants must increase. Equivalently, if \( Q > K \) then \( \Delta G_{rxn} > 0 \), indicating that the reaction is spontaneous from right to left as written.

### 6.10.34. Enthalpy and Entropy Contribute to \( \Delta G_{rxn}^0 \)

We know that at constant temperature we have:
\[
\Delta G_{rxn}^0 = \Delta H_{rxn}^0 - T\Delta S_{rxn}^0.
\]
Consequently, we may express the equilibrium constant as a product of two contributions, one from enthalpy and one from entropy:
\[
K = \exp \left( -\frac{\Delta G_{rxn}^0}{RT} \right) = \exp \left( -\frac{\Delta H_{rxn}^0}{RT} \right) \exp \left( \frac{\Delta S_{rxn}^0}{R} \right).
\]

### 6.10.35. Example 1: Three Components Reaction

The above formula states that the change in standard Gibbs energy for a reaction is related to the equilibrium constant. But the latter depends on equilibrium partial
pressures. For example,

\[ 3\text{NO}(g) \rightleftharpoons \text{N}_2\text{O}(g) + \text{NO}_2(g) \]

has

\[ K = \frac{(P_{\text{N}_2\text{O}})_{eq} \cdot (P_{\text{NO}_2})_{eq}}{(P_{\text{NO}})_{eq}^3} \]

so (taking \( P_0 = 1 \) atm and expressing pressures in units of \( P_0 \)),

\[ \Delta G_{\text{rxn}}^\circ = -RT \log \frac{(P_{\text{N}_2\text{O}})_{eq} \cdot (P_{\text{NO}_2})_{eq}}{(P_{\text{NO}})_{eq}^3}. \]

6.10.36. Example 2: Four Components Reaction. For a reaction of the type:

\[ \nu_A \text{A}(g) + \nu_B \text{B}(g) \rightleftharpoons \nu_Y \text{Y}(g) + \nu_Z \text{Z}(g), \]

we have:

\[ \Delta G_{\text{rxn}} = \nu_Y \mu_Y + \nu_Z \mu_Z - \nu_A \mu_A - \nu_B \mu_B \]

\[ = \nu_Y \mu_Y^\circ(T) + \nu_Z \mu_Z^\circ(T) - \nu_A \mu_A^\circ(T) - \nu_B \mu_B^\circ(T) \]

\[ + RT \left( \nu_Y \log \frac{P_Y}{P_0} + \nu_Z \log \frac{P_Z}{P_0} - \nu_A \log \frac{P_A}{P_0} - \nu_B \log \frac{P_B}{P_0} \right), \]

which can be written in the form:

\[ \Delta G_{\text{rxn}}(T) = \Delta G_{\text{rxn}}^\circ(T) + RT \log Q, \]

where

\[ \Delta G_{\text{rxn}}^\circ(T) = \nu_Y \mu_Y^\circ(T) + \nu_Z \mu_Z^\circ(T) - \nu_A \mu_A^\circ(T) - \nu_B \mu_B^\circ(T), \]

and

\[ Q = \frac{(P_Y/P_0)^{\nu_Y} \cdot (P_Z/P_0)^{\nu_Z}}{(P_A/P_0)^{\nu_A} \cdot (P_B/P_0)^{\nu_B}}. \]

6.10.37. Standard Gibbs Energies of Reactions (Using Tabulated Values). Standard entropies \( (\Delta S_{\text{rxn}}^\circ) \) and enthalpies \( (\Delta H_{\text{rxn}}^\circ) \) of reactions can be combined to obtain the standard Gibbs energy of a reaction, \( \Delta G_{\text{rxn}}^\circ = \Delta H_{\text{rxn}}^\circ - T \Delta S_{\text{rxn}}^\circ \) (valid for constant temperature). The standard Gibbs energy of a reaction is the difference in standard molar Gibbs energies of the products and reactants in their standard states at the temperature specified for the reaction as written.

In Oxtoby [1], what we denote here as \( \mu_i^0 \) is called standard Gibbs energy of formation and denoted \( \Delta G_i^f \). Appendix D of your book can be used to calculate equilibrium constants; e.g.

\[ \nu_A \text{A} + \nu_B \text{B} \rightarrow \nu_Y \text{Y} + \nu_Z \text{Z} \]

has

\[ \Delta G_{\text{rxn}}^\circ = \nu_Y \Delta G_i^f[Y] + \nu_Z \Delta G_i^f[Z] - \nu_A \Delta G_i^f[A] - \nu_B \Delta G_i^f[B]. \]
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where the values for the $\Delta G^o_f[Y]$ can be looked up in tables.

The **standard Gibbs energy of formation** of a compound is the change of Gibbs free energy that accompanies the formation of 1 mole of a substance in its standard state from its constituent elements in their standard states (the most stable form of the element at 1 bar of pressure and the specified temperature, usually 298.15 K or 25°C). Standard Gibbs energies of formation are tabulated. See Appendix D of Oxtoby [1] for such a list.

In terms of the **standard Gibbs energies of formation**, $\Delta G^o_f$:

$$
\Delta G^o_{\text{rxn}} = \sum_{I \in \text{prod.}} \nu_I (\Delta G^o_f)_I - \sum_{J \in \text{react.}} \nu_J (\Delta G^o_f)_J.
$$

**Example:** To calculate the standard Gibbs energy of the reaction

$$
\text{CO}(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{CO}_2(g)
$$

at 25°C. Looking up tabulated values for each substance in Appendix D, we write:

$$
\Delta G^o_{\text{rxn}} = \Delta G^o_f(\text{CO}_2,g) - \left\{ \Delta G^o_f(\text{CO},g) + \frac{1}{2} \Delta G^o_f(\text{O}_2,g) \right\}
$$

$$
= -394.4 \text{ kJ/mol} - \left\{ (-137.2) + \frac{1}{2}(0) \right\} \text{ kJ/mol}
$$

$$
= -257.2 \text{ kJ/mol}
$$

6.10.38. **Adiabatic Decompression.** The process of adiabatic decompression is used for cooling gases. The principle is illustrated in Fig. 6.26. In the equations below, ‘1’ refers to the ‘initial’ state and ‘2’ refers to the ‘final’ state. Step 1 uses the law for isothermal compression (see Section 6.3.6):

$$
\Delta W = -nRT \log \left( \frac{V_2}{V_1} \right)
$$

with $\Delta U = 0$ for an ideal gas (isothermal process), so that $\Delta Q = \Delta W$. The second step uses the law for adiabatic expansion:

$$
T_1V_1^{\gamma-1} = T_2V_2^{\gamma-1}.
$$

**Problem:** Derive the law for adiabatic expansion.

**Problem:** Explain how this leads to cooling.

6.11. **Electric and Magnetic Work**

Can electric and magnetic fields generate work? In this section we look at the case of electrical charges subjected to the influence of electric and magnetic fields. This material is standard and can be found in most textbooks of electrodynamics, e.g. see Griffith [26] or Jackson [27]. In electrodynamics,
the electric field is denoted $\mathbf{E}$. There is an associated quantity called the *electric displacements*, denoted $\mathbf{D}$. An applied electric field gives rise to a polarization $\mathbf{P}$ (molecules develop an induced dipole moment). These three fields are related by:

$$\mathbf{D} = \varepsilon_0 \mathbf{E} + \mathbf{P}.$$ 

In linear media, $\mathbf{D} = \varepsilon \mathbf{E}$, where $\varepsilon$ is the dielectric constant of the linear medium. The corresponding magnetic quantities are $\mathbf{B}$ (magnetic induction), $\mathbf{H}$ (magnetic field) and $\mathbf{M}$ magnetic polarization or magnetization. Magnetic polarization is magnetic moment ($\vec{m}$) per unit volume: $\vec{m} = \text{vol} \cdot \mathbf{M}$. Similarly, $\vec{p} = \text{vol} \cdot \mathbf{P}$ for the electric dipole moment. The magnetic moments align when an external magnetic field is applied. These three fields are related by:

$$\mathbf{H} = \frac{1}{\mu_0} \mathbf{B} - \mathbf{M}.$$ 

In linear media, we have $\mathbf{H} = \frac{1}{\mu} \mathbf{B}$, where $\mu$ is the magnetic permeability of the medium. $\mathbf{E}$ and $\mathbf{B}$ are the fundamental fields because they determine the force on electric charges according to the Lorentz force law:

$$\mathbf{F} = q(\mathbf{E} + \mathbf{v} \times \mathbf{B}).$$
6.11. Electric and Magnetic Work

H and D are called auxiliary fields.

6.11.1. Derivation. Upon first reading, you may skip this section, which is provided for completeness and whose details are not essential to the development of thermodynamical relations. We start with the Lorentz force,

\[ \mathbf{F} = q(\mathbf{E} + \mathbf{v} \times \mathbf{B}) \]

Take the dot product\(^{24}\) with \(d\mathbf{l}\), an element of length:

\[ \mathbf{F} \cdot d\mathbf{l} = q(\mathbf{E} + \mathbf{v} \times \mathbf{B}) \cdot d\mathbf{l} \]

We replace \(d\mathbf{l}\) by \(\mathbf{v} dt\), where \(\mathbf{v}\) is the velocity of electric charges, then divide by \(dt\) to get \(dW/dt\), the rate of change in the work (why we use the rate of change will be clear later, as the Poynting vector emerges). Also, because electric charge is possibly distributed elsewhere in space, we really should replace \(q\) by a charge density, \(\rho\), and integrate over volume to get the total work (and use \(J = \rho \mathbf{v}\)):

\[ (6.13) \]

\[ \frac{dW}{dt} = - \int_{\text{vol}} \rho(\mathbf{E} + \mathbf{v} \times \mathbf{B}) \cdot \mathbf{v} \, d^3r = - \int_{\text{vol}} \rho \mathbf{E} \cdot \mathbf{v} \, d^3r = - \int_{\text{vol}} \mathbf{E} \cdot \mathbf{J} \, d^3r. \]

The negative sign appears because we are calculating the work done on the charge against the action of the field (see Jackson \[27\], Chapter 1).

The magnetic force term dropped out because the triple product\(^{25}\) \((\mathbf{v} \times \mathbf{B}) \cdot \mathbf{v}\) vanishes, as two of its vectors are parallel. However, this does not mean that the magnetic field does not play a role. Let us look at the current density, \(\mathbf{J}\). We can obtain an expression for \(\mathbf{E} \cdot \mathbf{J}\) by invoking Ampère’s law:

\[ \nabla \times \mathbf{H} = \mathbf{J} + \frac{\partial \mathbf{D}}{\partial t}. \]

Take the dot product with \(\mathbf{E}\):

\[ (6.14) \]

\[ \mathbf{E} \cdot \mathbf{J} = \mathbf{E} \cdot (\nabla \times \mathbf{H}) - \mathbf{E} \cdot \frac{\partial \mathbf{D}}{\partial t}. \]

Next, we use the identity,\(^{26}\)

\[ \mathbf{H} \cdot (\nabla \times \mathbf{E}) - \mathbf{E} \cdot (\nabla \times \mathbf{H}) = \nabla \cdot (\mathbf{E} \times \mathbf{H}), \]

in which we substitute Faraday’s law

\[ \nabla \times \mathbf{E} = -\frac{\partial \mathbf{B}}{\partial t}, \]

\(^{24}\)The dot product is defined in Section A.3.

\(^{25}\)A triple product, \((\mathbf{A} \times \mathbf{B}) \cdot \mathbf{C}\), is a scalar (number) which equals to the volume generated by the three vectors \(\mathbf{A, B, C}\). When any two vectors are parallel, the volume is zero.

\(^{26}\)This can be verified as follows:

\[ \nabla \cdot (\mathbf{E} \times \mathbf{H}) = \partial_j (\epsilon_{ilm} E_l H_m) = \epsilon_{ilm} (\partial_j E_l) H_m + \epsilon_{ilm} E_l (\partial_j H_m) = \mathbf{H} \cdot (\nabla \times \mathbf{E}) - \mathbf{E} \cdot (\nabla \times \mathbf{H}), \]

where in the second term we have used \(\epsilon_{ilm} = \epsilon_{nml} = -\epsilon_{nim}\).
to get

\[ (6.15) \quad -\mathbf{H} \cdot \frac{\partial \mathbf{B}}{\partial t} - \mathbf{E} \cdot (\nabla \times \mathbf{H}) = \nabla \cdot (\mathbf{E} \times \mathbf{H}). \]

In Eq. (6.14), we replace \( \mathbf{E} \cdot (\nabla \times \mathbf{H}) \) by Eq. (6.15), to get:

\[ \mathbf{E} \cdot \mathbf{J} = -\mathbf{H} \cdot \frac{\partial \mathbf{B}}{\partial t} - \nabla \cdot (\mathbf{E} \times \mathbf{H}) - \mathbf{E} \cdot \frac{\partial \mathbf{D}}{\partial t}. \]

Substituting this expression into Eq. (6.13), we get:

\[ (6.16) \quad \frac{dW}{dt} = -\int_{\text{vol}} \left( -\mathbf{H} \cdot \frac{\partial \mathbf{B}}{\partial t} - \nabla \cdot (\mathbf{E} \times \mathbf{H}) - \mathbf{E} \cdot \frac{\partial \mathbf{D}}{\partial t} \right) \, d^3r. \]

The middle term can be converted to a surface integral using the Gauss divergence theorem:

\[ (6.17) \quad \frac{dW}{dt} = \int_{\text{vol}} \left( \mathbf{H} \cdot \frac{\partial \mathbf{B}}{\partial t} + \mathbf{E} \cdot \frac{\partial \mathbf{D}}{\partial t} \right) \, d^3r + \int_{\partial \text{vol}} (\mathbf{E} \times \mathbf{H}) \cdot d\mathbf{s}. \]

In linear media, \( \mathbf{B} = \mu \mathbf{H} \) and \( \mathbf{D} = \epsilon \mathbf{E} \), and we can write this as:

\[ (6.18) \quad \frac{dW}{dt} = \frac{d}{dt} \int_{\text{vol}} \frac{1}{2} (\mathbf{H} \cdot \mathbf{B} + \mathbf{E} \cdot \mathbf{D}) \, d^3r + \int_{\partial \text{vol}} (\mathbf{E} \times \mathbf{H}) \cdot d\mathbf{s}. \]

The first integral is the total energy stored in the electromagnetic fields. The first term represents the rate of change in the total energy stored. The quantity

\[ W_{EH} = \frac{1}{2} (\mathbf{H} \cdot \mathbf{B} + \mathbf{E} \cdot \mathbf{D}) \]

is the energy density of the electromagnetic field. The second term describes the transport of energy in and out of the volume (vol), through the surface \( \partial \text{vol} \). The vector \( \mathbf{S} \equiv \mathbf{E} \times \mathbf{H} \) is called the Poynting vector. The Poynting vector is relevant in cases where electromagnetic radiation is present.

The first term in Eq. (6.18) can also be written as (for linear media):

\[ (6.19) \quad dW = \int_{\text{vol}} \frac{d}{dt} \frac{1}{2} \left( \mathbf{B}^2 / \mu + \epsilon \mathbf{E}^2 \right) \, d^3r = \int_{\text{vol}} \frac{1}{2} \left( \mu^{-1} \mathbf{B} \cdot d\mathbf{B} + \epsilon \mathbf{E} \cdot d\mathbf{E} \right) \, d^3r. \]

If the fields \( \mathbf{E} \) and \( \mathbf{B} \) are spatially uniform (over the volume \( \text{vol} \)), integration over the volume yields a factor \( \text{vol} \). Then, the total energy differential is:

\[ dW = \text{vol} \left( \mu^{-1} \mathbf{B} \cdot d\mathbf{B} + \epsilon \mathbf{E} \cdot d\mathbf{E} \right). \]

This is often written as:

\[ (6.20) \quad dW = \text{vol} (\mathbf{H} \cdot d\mathbf{B} + \mathbf{E} \cdot d\mathbf{D}). \]

Note: some authors go directly from Eq. (6.17) to

\[ dW = \int_{\text{vol}} (\mathbf{H} \cdot d\mathbf{B} + \mathbf{E} \cdot d\mathbf{D}) \, d^3r + \int_{\partial \text{vol}} (\mathbf{E} \times \mathbf{H}) \cdot d\mathbf{s}. \]
This implies elimination of the time parametrization
\[ \frac{\partial B}{\partial t} dt \rightarrow dB, \quad \frac{\partial D}{\partial t} dt \rightarrow dD. \]

Since \( B \) and \( D \) are functions of position and time, their total derivatives are:
\[ dB = \frac{\partial B}{\partial r} \cdot dr + \frac{\partial B}{\partial t} dt, \quad dD = \frac{\partial D}{\partial r} \cdot dr + \frac{\partial D}{\partial t} dt. \]

Thus, the substitution (6.21) can only be justified if spatial variations of \( B \) and \( D \) are small (i.e. \( \frac{\partial B}{\partial r} \approx 0 \) and \( \frac{\partial D}{\partial r} \approx 0 \)) or if \( \text{vol} \) is sufficiently large so that edge effects (spatial non-uniformities) contribute negligibly to the volume integral.

6.11.2. Magnetic and Electric Polarizations. In Eq. (6.20), namely
\[ dW = \text{vol} (H \cdot dB + E \cdot dD), \]

only the fields \( E, D, B, H \) appear. When we derived Eq. (6.20), we could have just as easily written instead (reader should check this):\(^{27}\)
\[ dW = \text{vol} (B \cdot dH + D \cdot dE). \]

It would be desirable to decompose the contributions of the external and internal fields. Since \( B = \mu_0 H + M \) and \( D = \varepsilon_0 E + P \), we have:
\[ dW = \text{vol} \left[ (\mu_0 H \cdot dH + \varepsilon_0 E \cdot dE) + M \cdot dM + P \cdot dP \right]. \]

A useful approximation can be obtained by decomposing the fields \( E, H \) as the sum of externally applied fields (subscript 0) and any additional fields (subscript \( dm \))\(^{28}\) arising after introduction of the sample in the external field:
\[ H = H_0 + H_{dm}, \quad E = E_0 + E_{dm}. \]

The \( dm \) fields are usually very small compared to the externally applied fields: \( ||H_{dm}|| \ll ||H_0|| \) and \( ||E_{dm}|| \ll ||E_0|| \). In this approximation the

\(^{27}\)Check also that substituting, \( B = \mu_0 H + M \) and \( D = \varepsilon_0 E + P \) into \( dW = \text{vol} (H \cdot dB + E \cdot dD) \) yields:
\[ dW = \text{vol} \left[ \mu_0 H \cdot dH + \varepsilon_0 E \cdot dE + M \cdot dM + P \cdot dP \right]. \]

It’s a matter of convenience which set of independent variables one uses. However, \( \vec{m} = \text{vol} \cdot M \) and \( \vec{p} = \text{vol} \cdot P \) are extensive variables whereas \( H \) and \( E \) are intensive. We recall from Section 6.4.7 that work differentials are of the form \( \delta W = \text{(intensive quantity)} \times \text{d(extensive quantity)}. \) Thus, the correct form is:
\[ dW = \text{vol} \left[ \frac{1}{2} \mu_0 H^2 + \varepsilon_0 E^2 \right] + H \cdot d\vec{m} + E \cdot d\vec{p}. \]

\(^{28}\)An example of field perturbation is the demagnetizing field associated with magnets.
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expression for work reduces to:

\[
(6.22) \quad dW = \text{vol} \left[ \frac{1}{2} d \left( \mu_0 H_0^2 + \epsilon_0 E_0^2 \right) + M \cdot dH_0 + P \cdot dE_0 \right].
\]

The advantages of this expression are two-fold:

1. The terms are broken down into two parts: the energy required to create the external field (the \(\mu_0 H_0^2 + \epsilon_0 E_0^2\) term), and the energy required to realign the external fields (the \(M \cdot dH_0 + P \cdot dE_0\) term).
2. For the \(M \cdot dH_0 + P \cdot dE_0\), we see that the external fields (\(H_0\) and \(E_0\)) are the control parameters. This is better representative of real experimental conditions, where the experimentalist controls the external fields rather than the polarization fields inside the material.

6.11.3. Linear Media, Curie’s Law. The internal energy differential with electromagnetic work, Eq. (6.22), reads:

\[
\frac{dU}{dQ} = \text{vol} \left[ \frac{1}{2} d \left( \mu_0 H_0^2 + \epsilon_0 E_0^2 \right) + M \cdot dH_0 + P \cdot dE_0 \right].
\]

In paramagnetic materials, the Curie-Weiss law holds,

\[M = \chi H, \quad \chi = \frac{C}{T},\]

where \(C\) is a material-dependent constant. In dielectric materials a similar law often holds,

\[P = \chi_\epsilon \epsilon_0 E = \alpha E,\]

where \(\alpha\) is a material-dependent constant called the polarizability (see Section 2.1.2 and Table 2.5). Under these circumstances, we have:

\[
dU = \delta Q + \text{vol} \left[ \frac{1}{2} d \left( \mu_0 H_0^2 + \epsilon_0 E_0^2 \right) + \frac{1}{2} d \left( \frac{C}{T} H_0^2 + \alpha E_0^2 \right) \right]
\]

\[
= \delta Q + \text{vol} \left[ \frac{1}{2} d \left( \mu_0 + \frac{C}{T} \right) H_0^2 + \left( \epsilon_0 + \alpha \right) E_0^2 \right]
\]

The term \(C/T\) is much larger than \(\mu_0\); thus, we neglect the \(\mu_0\) term.

---

29The factor of \(\frac{1}{2}\) arises because we used \(HdH = \frac{1}{2}dH^2\).
30Energy is required to bring the charges and currents required to create the fields.
31The conditions include: Curie law, linear media, and approximating \(H \approx H_0\) and \(E \approx E_0\).
32For those interested in a justification of this approximation: The Curie constant is given by \(\frac{\mu_0^2 g^2}{3k_B} N g^2 J(J+1)\), where \(J\) is the total angular momentum quantum number of the paramagnetic atoms, \(g\) is the Landé factor, \(g = 1 + \frac{J(J+1)+S(S+1)-L(L+1)}{2J(J+1)}\) (\(S\): spin, \(L\): orbital, \(J\): total angular momenta). For an electron spin, \(g \approx 2\). \(N\) is the density of magnetic atoms (10-100 \(\times 10^{27}\) atoms/m\(^3\) for solids). \(\mu_B\) is the Bohr magneton (9.274 \(\times 10^{-24}\) J/T). Plugging these typical values at room temperature gives \(C/T \approx 10^3\mu_0\), which is much larger than \(\mu_0\). In adiabatic demagnetization experiments, the temperature is typically much lower (liquid helium), so the approximation is well justified.
6.11.4. Adiabatic Demagnetization. Adiabatic demagnetization (W.F. Giauque) can be used for cooling (Fig. 6.27) in a process that is analogous to adiabatic decompression (Section 6.10.38) but instead involves the spin degrees of freedom of paramagnetic salts. Adiabatic demagnetization of a paramagnetic salt is a two-steps sequence that enables cooling below 1 K. Temperatures around 1 mK have been reached. The steps are: (1) isothermal magnetization followed by (2) adiabatic demagnetization (Fig. 6.27):

(1) **Isothermal magnetization** (thermal contact with cold He gas). At time zero, spins in the paramagnetic salt are randomly oriented (disordered state). The system undergoes isothermal demagnetization in the presence of a strong magnetic field. The entropy of the spin system decreases, as the spins align with the magnetic field (ordered state). While the final temperature is unchanged, heat is lost to the surroundings in order to dissipate the entropy. (Lowering the temperature would also lead to alignment of the spins.)

(2) **Adiabatic demagnetization** (He gas pumped away; no thermal contact with He). Next, heat transfer is stopped (paramagnetic salt is insulated) and the external field is turned off, leading the spins to randomly reorient. However, the temperature has decreased. While heat transfer outside the system is not allowed, heat can be transferred between spin and lattice. However, the lattice entropy is much lower than the spin entropy. The total entropy of the isolated system (paramagnetic salt) has not changed much, since heat was not allowed to transfer ($\delta S = \delta Q / T \approx 0$).

Let us see how this two-step process could possibly work (Fig. 6.28). Our starting point is Eq. (6.22). Let us take the “system” as the paramagnetic salt plus electromagnetic coils as well as the surrounding helium (liquid helium bath plus helium gas in contact with the salt). This forms an isolated system. For isolated systems, the change in internal energy is zero ($\delta U = 0$). During isothermal magnetization, heat is allowed to exchange between the helium and the salt in order to maintain the salt temperature constant:

$$\delta Q_{\text{He}} = -\delta Q_{\text{salt}}.$$
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**Figure 6.27.** Adiabatic demagnetization. The first step involves turning on a magnetic field at constant temperature. The spins align with the magnetic field, reducing the entropy. In the second step, the entropy is held constant while the field is turned off. Since the entropy remains constant, the only possible outcome left to compensate for the loss of magnetic field is for the temperature to decrease.

If the process is reversible and isothermal,

\[ T_{\text{He}} dS_{\text{He}} = -T_{\text{salt}} dS_{\text{salt}}, \]

we have \( T_{\text{He}} = T_{\text{salt}} \equiv T \) and \( dS_{\text{He}} = -dS_{\text{salt}} \). The entropy in the context of heat transfer is the lattice entropy (random atomic motions). The paramagnetic salt also has spin entropy, which does not exchange with the helium in any way because no exchange mechanism exists\(^{35}\).

The total salt entropy should be viewed as the sum of lattice \((L)\) and spin \((S)\) contributions:

\[ dS_{\text{salt}} = dS_{\text{salt}}^L + dS_{\text{salt}}^S. \]

Our previous statement \( \delta Q_{\text{He}} = -\delta Q_{\text{salt}} \) refers to the lattice contributions because heat can only be transferred via lattice motions. We are left with a non-vanishing \( T dS_{\text{salt}}^S \) term describing changes in the spin entropy. (\( T dS_{\text{He}}^S = 0 \) because helium is not magnetic.) The physical origin of \( dS_{\text{salt}}^S \) is the alignment of the magnetic moments of the paramagnetic salt due to the ramping up of the magnetic field.

\(^{35}\)For spin entropy to exchange, we would need a spin-spin coupling mechanism, which does not exist here.
The first law then reads (with $E_0 = 0$ and neglecting the $\mu_0$ term):

$$dU = 0 = TdS_{salt} + \frac{vol}{2} d\left[\frac{C}{T}H_0^2\right].$$

During isothermal magnetization, the field is ramped up from 0 to some value $H$. The spin entropy decreases by an amount proportional to $(H/T)^2$:

$$\Delta S_{salt}^S = \int dS_{salt}^S = S_{salt}^S(H) - S_{salt}^S(0) = -\frac{vol}{2}C\left(\frac{H}{T}\right)^2.$$

During adiabatic demagnetization, the entropy remains constant since $\delta Q = TdS = 0$. The field is ramped down from $H$ to $H_{int}$, where $H_{int}$ is the internal magnetic field due to coupling to neighboring magnetic moments. If the entropy is proportional to $(H/T)^2$, the condition of the initial entropy to equal the final entropy:

$$S_{salt}^S(initial) = const \times \left(\frac{H}{T_i}\right)^2 = S_{salt}^S(final) = const \times \left(\frac{H_{int}}{T_f}\right)^2,$$

leads to:

$$T_f = T_i \left(\frac{H_{int}}{H}\right).$$

The internal field is typically on the order of a few Gauss ($10^{-4}$ T). The external field ($H$) is on the order of a Tesla. Therefore, cooling by 3-4 orders of magnitude is possible.

6.12. Problems

Problem 62. We have seen that the standard Gibbs free energy of a reaction, $\Delta G_{rxn}^0$, is given as a sum of standard Gibbs energy of formation, for
each reactant and products (with reactants taken as negative):
\[ \Delta G^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta G^\circ_f)_i \]

where the sum \( i \) runs over all reactants and products. (a) Prove that this formula follows from the fact that \( dG \) is an exact differential. (b) Using only the formula from (a) and the relationship between \( G \), \( H \) and \( S \), prove also that:

\[ \Delta H^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta H^\circ_f)_i, \quad \Delta S^\circ_{\text{rxn}} = \sum_i \nu_i \delta^\circ_i \]

where \( S^\circ_i \) and \((\Delta H^\circ_f)_i\) are the entropies (absolute) and enthalpies of formation for species \( i \), respectively. \( \Delta H^\circ_{\text{rxn}} \) and \( \Delta S^\circ_{\text{rxn}} \) are the standard enthalpy and entropy changes for the reaction.

**Solution.** (a) \( \Delta G^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta G^\circ_f)_i \) follows from the fact that \( G \) is a state function, i.e. \( \Delta G = \int_{\text{products}}^{\text{reactants}} dG \). Integrate along the following path: Destroy all reactants one by one (order is unimportant) and Form all the products one by one (order is unimportant). Each molecule destroyed and formed contributes a term in the summation (with appropriate sign).

(b) First we have the basic relationship from thermodynamics (constant \( T \)), \( \Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T \Delta S^\circ_{\text{rxn}} \). Each term in the summation \( \Delta G^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta G^\circ_f)_i \), we view \((\Delta G^\circ_f)_i \) as a small increment \( dG_i \) that contributes to the integral, for which we apply the following decomposition \( dG_i = dH_i - T dS_i \). Let’s call each term \( dH_i \equiv (\Delta H^\circ_f)_i \) and \( dS_i \equiv \delta^\circ_i \), respectively. This gives the decomposition

\[ \Delta G^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta G^\circ_f)_i = \sum_i \nu_i (\Delta H^\circ_f)_i - T \sum_i \nu_i S^\circ_i. \]

And since \( \Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T \Delta S^\circ_{\text{rxn}} \), in the limit \( T \to 0 \), the second term vanishes, we have \( \Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} \) and \( \Delta G^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta H^\circ_f)_i \). Therefore,

\[ \Delta H^\circ_{\text{rxn}} = \sum_i \nu_i (\Delta H^\circ_f)_i. \]

Taking the limit \( T \to \infty \), the first term is negligible compared to the second term, leaving only \( \Delta G^\circ_{\text{rxn}} = -T \Delta S^\circ_{\text{rxn}} \) and \( \Delta G^\circ_{\text{rxn}} = -T \sum_i \nu_i \delta^\circ_i \). It follows that:

\[ \Delta S^\circ_{\text{rxn}} = \sum_i \nu_i \delta^\circ_i. \]

**Problem 63.** Hess’s law states that the total enthalpy change during the complete course of a reaction is the same whether the reaction is made in one step or in several steps. Mathematically, the statement of Hess’s law,
as pertains to the net or overall reaction, is:

\[ \Delta H^\circ_{\text{rxn}} \equiv \sum_{I \in \text{products}} \nu_I (\Delta H_f^\circ)_I - \sum_{J \in \text{reactants}} \nu_J (\Delta H_f^\circ)_J, \]

where the reaction is the net reaction. The summation is over all reactants (products) in the net equation. \((\Delta H_f^\circ)_I\) is called enthalpy of formation (of species \(I\)) and \(\nu_I\) is the stoichiometric coefficient of species \(I\). (Here, \(\nu_I\) and \(\nu_J\) are all positive.) When the conditions are standard, the enthalpy of formation is denoted as \((\Delta H_f^\circ)^{\text{standard}}\) (standard enthalpy of formation).

(a) Prove that Hess’s law is a consequence of the fact that \(H\) is a state function (\(dH\) is an exact differential), i.e. it does not depend on the path taken to get from the reactants to the products. Prove that the above statement for the net reaction is entirely equivalent to summing up reaction enthalpies from individual reactions:

\[ \Delta H^\circ_{\text{rxn}} = \sum_{i \in \text{reactions}} \Delta H^\circ_{\text{rxn}}(i), \]

where \(\Delta H^\circ_{\text{rxn}}(i)\) denotes the net (overall) reaction enthalpy of the \(i\)-th reaction. The sum runs over all reactions that make up the net reaction.

Suppose that we have two concurrent reactions taking place:

\[
\begin{align*}
\text{C(s,gr)} + \frac{1}{2} \text{O}_2(g) & \rightarrow \text{CO}_2(g), \quad \Delta H^\circ_{\text{rxn}}(1) = -393.5 \text{ kJ} \\
\text{CO}_2(g) & \rightarrow \text{CO}(g) + \frac{1}{2} \text{O}_2(g) \quad \Delta H^\circ_{\text{rxn}}(2) = +283.0 \text{ kJ}.
\end{align*}
\]

(b) Apply Hess’s law to obtain \(\Delta H^\circ_{\text{rxn}}\) for the net reaction.

**Solution.** (a) Hess’s law follows from \(\Delta H^\circ_{\text{rxn}} = \int_{\text{reactants}}^{\text{products}} \Delta H^\circ_{\text{rxn}} \, dH\) and choosing a path in the net reaction that destroys all reactant molecules regardless of order (destruction implies a negative sign) and creates all product molecules (without regards to order). This is equivalent to summing over all reactions: because \(H\) is a state function, the value of \(\Delta H^\circ_{\text{rxn}}\) is the same, regardless of the choice of path. Here, we choose a different path, which takes us through all reactions one by one. This is entirely equivalent, because the chemical reaction is treated as an equation; multiple reactions are added to give the net reaction. The integral is a linear operator, hence, it is applied to each term in the summation (and the summation is identical whether we look at the net reaction or the sum of individual reactions).

(b) The net reaction (and net enthalpy change) is the sum:

\[
\text{C(s,gr) +} \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO}(g) \quad \Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}}(1)+\Delta H^\circ_{\text{rxn}}(2) = -110.5 \text{ kJ}.
\]
**Problem 64.** METHOD OF BOND ENTHALPIES. In chemical reactions, bonds are broken and new ones are formed. Think about reactants: one way to consume reactants is to destroy all the chemical bonds of all reactants, until atoms are separate. On the products side: to form the product molecules, we assemble the atoms together and create a number of chemical bonds. There is an energy associated with breaking and forming these bonds. It is possible to use this information to estimate the enthalpy change \( \Delta H^\circ_{rxn} \) of a reaction, which can be calculated from:

\[
\Delta H^\circ_{rxn} = \sum_{I \in \text{bonds broken}} \Delta H^\circ_I - \sum_{J \in \text{bonds formed}} \Delta H^\circ_J,
\]

where \( \Delta H^\circ_I \) is the energy absorbed when a particular bond is broken and \( \Delta H^\circ_J \) is the energy released when a bond is formed. (a) Prove that this restatement of Hess’s law is entirely equivalent to the one in Problems 1 and 2. (b) Prove also that this new statement follows from the fact that \( H \) is a state function (and \( dH \) is an exact differential). (c) Consider the following example on cracking of propane (thermal decomposition, \( \Delta H^\circ_{rxn} > 0 \) endothermic) to make ethene (Fig. 6.29).

![Figure 6.29. Method of bond enthalpy illustrated.](image)

Here is some data on average bond enthalpy (units are kJ/mol)

- C – H bond: +415
- C – C bond: +345
- C = C bond: +611

Add up all the energies of the broken bonds; add up all the energies of the bonds that are reformed and subtract one from the other. Show that \( \Delta H^\circ_{rxn} = 4.010 - 3.931 = 79 \text{ kJ/mol} \). (Hint: 10 bonds must be broken on the reactants side to consume the reactants, whereas 9 bonds must be formed on the products side to synthesize the needed products.)

**Solution.** (a) This reformulation of Hess’s law is justified by the choice of path, which here involves destroying molecules or forming them, one chemical bond at a time. (This is longer, and involves more steps, but the end result is the same, i.e. the value of \( \Delta H^\circ_{rxn} \) is identical.)
(b) This follows from the fact that \( dH \) is an exact differential because the original Hess’s law is also based on this fact. And since the two are equivalent, the proof follows.

(c) We have, for the left hand side (bonds broken):

\[
\begin{align*}
2 \text{ C} - \text{ C} & \quad 2 \times 345 = 690 \\
8 \text{ C} - \text{ H} & \quad 8 \times 415 = 3320
\end{align*}
\]

These energies add up to a total of 4,010 kJ/mol.

For the right hand side (bonds formed),

\[
\begin{align*}
1 \text{ C} = \text{ C} & \quad 1 \times 611 = 611 \\
8 \text{ C} - \text{ H} & \quad 8 \times 415 = 3320
\end{align*}
\]

for a total of 3,931 kJ/mol. The net difference between reactant bonds destroyed and product bonds formed is:

\[
\Delta H^\circ_{\text{rxn}} = 4010 - 3931 = 79 \text{ kJ/mol}.
\]

Problem 65. (e) Is it really necessary to completely break down all bonds and reform them or can you find a short-cut to computing \( \Delta H \) of the reaction without having to completely take apart all molecules and reform them (including reforming the same bonds that were broken!)? Reformulate the statement for the method of bond enthalpies with this short-cut in mind.

Solution. No, it is not necessary to break ALL the bonds and reform them. When we break a bond and reform it, there is no net change in this particular bond, so why bother breaking and reforming it? Such pairs of events can be discounted because they cancel each other out. In other words, for a given reaction (reactants + products), the initial conditions (reactants) and final conditions (products) are fixed regardless of the path (from reactants to products) we choose. Instead of the ‘long path’ (that involves breaking ALL bonds), we can instead choose a shorter path that does not involve destroying ALL bonds. Instead we may limit the breaking and forming of bonds to the bare minimum necessary. In real reactions, not all bonds in all reactants are broken; only a small number of bonds are broken.

Problem 66. Prove the following formula

\[
d(\Delta G) = (\Delta V) dP - (\Delta S) dT,
\]

which allows us to calculate how \( \Delta G \) varies with temperature and pressure. \( \Delta G = \int dG \) denotes a change in \( G \), for example, \( \Delta G = G(\text{products}) - G(\text{reactants}) \) being one possible application. (Note: we have not derived
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this formula in class; the fundamental equation of thermodynamics (1st law) was given in terms of \( G \), not \( \Delta G \).

**Solution.** Apply the fundamental equation, \( dG = VdP - SDT \), twice, once to all the reaction products collectively, \( G(\text{products}) \), and once to all the reactants collectively, \( G(\text{reactants}) \):

\[
dG_{\text{prod}} = V_{\text{prod}}dP - S_{\text{prod}}dT, \quad dG_{\text{react}} = V_{\text{react}}dP - S_{\text{react}}dT
\]

\[
dG_{\text{prod}} - dG_{\text{react}} = d(G_{\text{prod}} - G_{\text{react}}) = d(\Delta G)
\]

\[
=(V_{\text{prod}} - V_{\text{react}})dP - (S_{\text{prod}} - S_{\text{react}})dT = (\Delta V)dP - (\Delta S)dT.
\]

**Problem 67.** Consider the reaction \( \text{CuBr}_2(s) \rightleftharpoons \text{CuBr}(s) + \frac{1}{2} \text{Br}_2(g, 1 \text{ atm}) \).

(a) In what direction does this reaction proceed at 300 K and 1 atm pressure? At 300 K, 1 atm, the following data is available from tables:

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H^\circ ) kcal/mol</th>
<th>( S^\circ_{300} ) cal/mol/K</th>
</tr>
</thead>
<tbody>
<tr>
<td>CuBr_2(s)</td>
<td>-33.2</td>
<td>30.1</td>
</tr>
<tr>
<td>CuBr(s)</td>
<td>-25.1</td>
<td>21.9</td>
</tr>
<tr>
<td>Br_2(g, 1 atm)</td>
<td>+7.34</td>
<td>58.64</td>
</tr>
</tbody>
</table>

**Solution.** The favored direction of reaction will be indicated by the sign of \( \Delta G^\circ_{300} \); to calculate which we need only determine \( \Delta H^\circ_{300} \) and \( \Delta S^\circ_{300} \) for the reaction.

\[
\Delta H^\circ_{300} = (\Delta H^\circ_f)_{\text{CuBr}} + \frac{1}{2}(\Delta H^\circ_f)_{\text{Br}_2} - (\Delta H^\circ_f)_{\text{CuBr}_2} = -25.1 + 3.67 - (-33.2) = +11.8 \text{ kcal/mol}.
\]

\[
\Delta S^\circ_{300} = (S^\circ)_{\text{CuBr}} + \frac{1}{2}(S^\circ)_{\text{Br}_2} - (S^\circ)_{\text{CuBr}_2} = +21.9 + 29.32 - 30.1 = +21.1 \text{ kcal/mol}.
\]

\[
\Delta G^\circ_{300} = \Delta H^\circ_{300} - T\Delta S^\circ_{300} = 11,800 - (298)(21.1) = 5500 \text{ cal/mol} = 5.5 \text{ kcal/mol}.
\]

Not rightward, but rather, leftward progress of the reaction is thus strongly favored at 300 K, and will reduce the Br_2 pressure far below 1 atm.

**Problem 68.** (b) For the previous reaction, at what temperature will the three substances coexist at equilibrium under a pressure of 1 atm?

**Solution.** At whatever temperature \( T \) the three substances coexist at 1 atm pressure, we will have \( \Delta H^\circ_T - T\Delta S^\circ_T = \Delta G^\circ_T = 0 \). Noting that \( \Delta H^\circ \) and \( \Delta S^\circ \) are the same in sign, we essay the approximation in which the two terms are treated as substantially constant over the temperature range concerned. In that case,

\[
\Delta H^\circ_T \approx \Delta H^\circ_{300} = 11.8 \text{ kcal/mol}
\]
\[ \Delta S_T^0 \approx \Delta S_{300}^0 = 21.1 \text{ kcal/mol.K} \]
and then
\[ 11,800 - T(21.1) = \Delta G_T^0 = 0, \quad T = 11,800/21.1 = 560^\circ \text{ K}. \]

Problem 69. (a) Which is the more stable form of carbon at 300 K and 1 atm pressure, diamond or graphite? (b) At 300 K, what pressure would be required to form diamond (density, 3.5 g/ml) from graphite (density, 2.25 g/ml)? (c) How can you be confident of the soundness of this entire mode of analysis? Some relevant data are as follows:

Heat capacity data yield for diamond (300 K), \( S_{300}^0 = 0.58 \text{ cal/mol.K} \), where \( S = -(dG/dT)_P \) and \( G \) is the molar Gibbs free energy.

Heat capacity data yield for graphite (300 K), \( S_{300}^0 = 1.37 \text{ cal/mol.K} \)

For \( \text{C(graphite)} \rightarrow \text{C(diamond)} \), \( \Delta S_{300}^0 = -0.79 \text{ cal/mol.K} \)

On combustion: \( \text{C(graphite)} + \text{O}_2 \rightleftharpoons \text{CO}_2 \), \( \Delta H_{300}^0 = -94.03 \text{ kcal.} \)

On combustion: \( \text{C(diamond)} + \text{O}_2 \rightleftharpoons \text{CO}_2 \), \( \Delta H_{300}^0 = -94.48 \text{ kcal.} \)

For \( \text{C(graphite)} \rightarrow \text{C(diamond)} \): \( \Delta H_{300}^0 = +450 \text{ cal/mol.} \)

And the given densities imply:

For diamond: gram-atomic volume = \( 12/3.5 = 3.4 \text{ ml/gr-atom.} \) (Note: Gram atomic mass is another term for the mass, in grams, of one mole of atoms of that element. “Gram atom” is a former term for a mole of substance (expressed in grams). Example: gram atomic mass of Hydrogen is mass of 1 mole atoms =1.008 grams.)

For graphite: gram-atomic volume = \( 12/2.25 = 5.3 \text{ ml/gr-atom.} \)

For \( \text{C(graphite)} \rightarrow \text{C(diamond)} \): \( \Delta V_{300}^0 = -1.9 \text{ ml/gr-atom, where } V = (dG/dt)_T. \)

Solution. (a) For the reaction \( \text{C(graphite)} \rightarrow \text{C(diamond)}, \)
\[ \Delta G_{300}^0 = \Delta H_{300}^0 - T\Delta S_{300}^0 = 450 - 300(-0.79) = +685 \text{ cal/mol} \]

At room temperature and atmospheric pressure, the favored direction of reaction is thus the inverse change of diamond into graphite. Under ordinary conditions diamond is therefore a thermodynamically unstable species, which exists only because of the extreme slowness of its conversion into the more stable graphite.

(b) At atmospheric pressure the reaction \( \text{C(graphite)} \rightarrow \text{C(diamond)} \) is characterized by \( \Delta G_{300}^0 > 0 \). But with \( \Delta V < 0 \) in this reaction, a sufficient rise of pressure should change the sign of \( \Delta G \) and, hence, the favored direction of reaction at 300 K. For consider that the effect of a change of pressure at
constant temperature will be given by the following reduced form
\[ d(\Delta \overline{G}) = \Delta V dP. \]

If we approximate by treating \( \Delta V \) as constant (= -0.0019 lit) over the entire pressure range involved, the last equation can easily be integrated. For the upper limit we choose that pressure \( (\overline{P}^*) \) at which \( \Delta \overline{G}^\circ_{300} = 0 \), i.e. the pressure under which diamond and graphite stand in equilibrium with each other at 300 K. For the lower limit we use the standard pressure of 1 atm, at which we have just found \( \Delta \overline{G}^\circ_{300} = +685 \text{ cal/mol} \). However, with \( \Delta V \) in liters and pressure in atmospheres, the appropriate unit for \( \Delta \overline{G} \) is not calories but lit-atm (1 lit-atm = 101.325 J). Multiplication by the factor 0.0413 converts a-figure in calories to one in lit-atm, so that \( \Delta \overline{G}^\circ_{300} = 0.0413(685) = 28.3 \text{ lit atm} \). We have then,
\[ \int_{28.3}^0 d(\Delta \overline{G}) = -0.0019 \int_1^{\overline{P}^*} dP, \]
\[ 0 - 28.3 = -0.0019(\overline{P}^* - 1), \]
\[ \overline{P}^* \approx 15,000 \text{ atm}. \]

At 25°C diamond and graphite would stand in equilibrium under a pressure of 15,000 atm. At still higher pressures graphite becomes thermodynamically unstable, and its conversion into diamond is then possible in principle, though so slow in practice as to be wholly undetectable.

(c) Since at room temperature no equilibrium of graphite with diamond ever is attained in practice, there remains room for skepticism that we have correctly calculated what would be the equilibrium condition. But, given expressions for \( \Delta \overline{H} \) and \( \Delta V \) as functions of temperature and pressure, by integration of equation \( d(\Delta \overline{G}) = \Delta V dP \), we can calculate that the equilibrium pressure is of the order of 75,000 atm at 1500°C. And here the soundness of our calculation is attested by an unmistakable production of diamond from graphite at pressures that exceed the equilibrium pressure.

\[ \square \]

**Problem 70.** Iron has a heat capacity of 25.1 J K\(^{-1}\) mol\(^{-1}\), approximately independent of temperature between 0°C and 100°C.

(a) Calculate the enthalpy and entropy change of 1.00 mol iron as it is cooled at atmospheric pressure from 100°C to 0°C.

(b) A piece of iron weighing 55.85 g and at 100°C is placed in a large reservoir of water held at 0°C. It cools irreversibly until its temperature equals that of the water. Assuming the water reservoir is large enough that its temperature remains close to 0°C, calculate the entropy changes for the iron and the water and the total entropy change in this process.
6.12. Problems

Solution. a)\[
\Delta H_{\text{Fe}} = n c_p \Delta T = (1.00 \text{ mol})(25.1 \text{ J K}^{-1} \text{ mol}^{-1})(273.15 - 373.15 \text{ K})
\]
\[= -2510 \text{ J} = -2.51 \times 10^3 \text{ J} \]
\[
\Delta S_{\text{Fe}} = n c_p \log \frac{T_f}{T_i} = (1.00 \text{ mol})(25.1 \text{ J K}^{-1} \text{ mol}^{-1})\log\frac{273.15}{373.15} = -7.83 \text{ J K}^{-1}
\]
b) The entropy $S$ is a function of state, and the initial and final states of the piece of iron are the same as in part a). Therefore, $\Delta S_{\text{Fe}} = -7.83 \text{ J K}^{-1}$. The reservoir of water gains the 2510 J of heat from the piece of iron at a constant temperature of 273.15 K. Therefore $\Delta S_{\text{water}} = 2510 \text{ J}/273.15 \text{ K} = +9.19 \text{ J K}^{-1}$.
\[
\Delta S_{\text{total}} = \Delta S_{\text{Fe}} + \Delta S_{\text{water}} = 1.36 \text{ J K}^{-1}
\]

Problem 71. (a) Use data from Appendix D from Oxtoby [1] to calculate $\Delta H^\circ$ and $\Delta S^\circ$ at 25 °C for the reaction
\[2\text{CuCl}_2(g) \rightleftharpoons 2\text{CuCl}(s) + \text{Cl}_2(g)\]
(b) Calculate $\Delta G^\circ$ at 590 K, assuming $\Delta H^\circ$ and $\Delta S^\circ$ are independent of temperature.
(c) Careful high-temperature measurements show that when this reaction is performed at 590 K, $\Delta H^\circ_{590}$ is 158.36 kJ and $\Delta S^\circ_{590}$ is 177.74 J K$^{-1}$. Use these facts to compute an improved value of $\Delta G^\circ_{590}$ for this reaction. Determine the percentage error in $\Delta H^\circ_{590}$ that comes from using the 298-K values in place of the 590-K values in this case.

Solution. a) The reaction of interest is
\[2\text{CuCl}_2(s) \rightarrow 2\text{CuCl}(s) + \text{Cl}_2(g)\]
Text Appendix D supplies $\Delta H^\circ_{298}$ and $S^\circ$ values at 298 K for the computation of $\Delta H^\circ_{298}$ and $\Delta S^\circ_{298}$
\[
\Delta H^\circ_{298} = 2(-137.2) + 1(0) - 2(-220.1) = 165.8 \text{ kJ}
\]
\[
\Delta S^\circ_{298} = 2(86.2) + 1(222.96) - 2(108.07) = 179.2 \text{ J K}^{-1}
\]
b)
$\Delta G^\circ_{590} \approx \Delta H^\circ_{590} - T \Delta S^\circ_{590} = 165.8 \text{ kJ} - (590 \text{ K})(0.1792 \text{ kJ K}^{-1}) = 60.1 \text{ kJ}$
(c) Use the experimental values at 590 K instead of the values at 298.15 K
\[
\Delta G^\circ_{590} = \Delta H^\circ_{590} - T \Delta S^\circ_{590} = 158.36 \text{ kJ} - (590 \text{ K})(0.17774 \text{ kJ K}^{-1}) = 53.5 \text{ kJ}
\]
The answer using $\Delta H^\circ_{298}$ and $\Delta S^\circ_{298}$ is about 12 % larger than the actual $\Delta G^\circ_{590}$. Tip. The temperature dependence of $\Delta H^\circ$ and $\Delta S^\circ$ should not always be neglected. Taking it into consideration becomes important when the temperature differs a lot from 298.15 K.
Problem 72. Find the Gibbs free energy and the entropy associated with mixing 10 g of liquid water, 5 g of liquid ethanol, and 1 g of liquid methanol at 298 K.

Solution.

\[ G_i = n_{H_2O} \mu^*_{H_2O} + n_{EtOH} \mu^*_{EtOH} + n_{MeOH} \mu^*_{MeOH} \]

\[ G_f = n_{H_2O}(\mu^*_{H_2O} + RT \log(X_{H_2O})) + n_{EtOH}(\mu^*_{EtOH} + RT \log(X_{EtOH})) + n_{MeOH}(\mu^*_{MeOH} + RT \log(X_{MeOH})) \]

\[ G_{mix} = G_f - G_i = nRT(X_{H_2O} \log(X_{H_2O}) + X_{EtOH} \log(X_{EtOH}) + X_{MeOH} \log(X_{MeOH})) \]

\[ S_{mix} = -nR(X_{H_2O} \log(X_{H_2O}) + X_{EtOH} \log(X_{EtOH}) + X_{MeOH} \log(X_{MeOH})) \]

\[ n_{H_2O} = 0.555 \text{ mol}, n_{EtOH} = 0.108 \text{ mol}, n_{MeOH} = 0.0312 \text{ mol}, n = 0.694 \text{ mol} \]

\[ X_{H_2O} = 0.800, X_{EtOH} = 0.156, X_{MeOH} = 0.0450 \]

\[ G_{mix} = -454 \text{ J} \]

\[ S_{mix} = 1.52 \text{ J/K} \]

Problem 73. Find the reaction affinity at constant pressure and temperature for the following reaction going to completion:

\[ \text{CH}_3\text{COOH}(g) + 2\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]

Data: \( G^\circ_f(\text{CO}_2(g)) = -394.36 \text{ kJ/mol}, G^\circ_f(\text{H}_2\text{O}(g)) = -228.59 \text{ kJ/mol}, G^\circ_f(\text{CH}_3\text{COOH}(g)) = -374.1 \text{ kJ/mol} \)

Solution. \( \Delta G_{rxn} = G_f(\text{products}) - G_f(\text{reactants}) = 2 \times G^\circ_f(\text{CO}_2(g)) + 2 \times G^\circ_f(\text{H}_2\text{O}(g)) - G^\circ_f(\text{CH}_3\text{COOH}(g)) = -871.8 \text{ kJ/mol} \)

Affinity = \(-\left(\frac{\partial G}{\partial \xi}\right)_{T,P}\)

Affinity = \(-\Delta G_{rxn} = -871.8 \text{ kJ/mol} \)

Problem 74. For the reaction \( 2 \text{Ca}(s) + \text{O}_2(g) \rightarrow 2 \text{CaO}(s) \) find the equilibrium constant at 300 K given that \( \mu^\circ_{\text{CaO}} = -604.17 \text{ kJ mol}^{-1} \) at that temperature. What can you conclude about the reaction based on the magnitude of the equilibrium constant?

Solution. \( \Delta \mu^\circ = \mu^\circ_{\text{CaO}}(2) - \mu^\circ_{\text{O}_2}(1) - \mu^\circ_{\text{Ca}}(2) = -1208.34 \text{ kJ mol}^{-1} \)

\( \Delta \mu^\circ = -RT \log K \)

\[ K = \exp\left(-\frac{\Delta \mu^\circ}{RT}\right) \]

\[ K = 2.5 \times 10^{210} \]

Problem 75. Assume the reaction \( \text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g) \) has taken place in a 2 L container and has reached equilibrium. There are 0.1 mole of
H₂(g), 0.15 mole of N₂(g), and 1.5 moles of NH₃(g) present at a temperature of 500 K. Find the change in chemical potential for the reaction if the equilibrium constant is 0.1744 under the stated conditions.

**Solution.**

\[ \mu_i = \mu_i^\circ + RT \log(P_i) \]

\[ \Delta \mu = \mu_{\text{NH}_3}^\circ(2) - \mu_{\text{H}_2}^\circ(3) - \mu_{\text{N}_2}^\circ(1) + RT \log \left( \frac{P_{\text{NH}_3}^2}{P_{\text{H}_2}^3 P_{\text{N}_2}} \right) \]

Let \( \Delta \mu^\circ = \mu_{\text{NH}_3}^\circ(2) - \mu_{\text{H}_2}^\circ(3) - \mu_{\text{N}_2}^\circ \)

\[ \Delta \mu^\circ = -RT \log (K) = 7260 \text{ J mol}^{-1} \]

If we treat the gases as ideal then we have:

\[ P_{\text{H}_2} = nRT/V = (0.1 \text{ mol})(0.08312 \text{ L bar mol}^{-1} \text{ K}^{-1})(500 \text{ K})/(2 \text{ L}) = 2.0786 \text{ bar} \]

\[ P_{\text{N}_2} = 3.1179 \text{ bar} \]

\[ P_{\text{NH}_3} = 31.179 \text{ bar} \]

\[ RT \log \left( \frac{P_{\text{NH}_3}^2}{P_{\text{H}_2}^3 P_{\text{N}_2}} \right) = (4157.24 \text{ J mol}^{-1})(3.55) = 14758 \text{ J mol}^{-1} \]

\[ \Delta \mu = 22018 \text{ J mol}^{-1} \]

**Problem 76.** Suppose we have an ideal gas. We double the gas volume by an isothermal expansion that proceeds (1) reversibly and (2) irreversibly into a vacuum. What is the change in entropy during this process (1 and 2)?

**Solution.**

(1) For the reversible expansion

\[ \Delta S = nR \log(V_2/V_1) = (2.30)(1)(1.99) \log 2 = +1.38 \text{ cal/mol.K} \]

(2) For the irreversible expansion to the same final state, the entropy change in the gas must again by +1.38 cal/mol.K.

**Problem 77.** What is the change in entropy of the surroundings when the volume of an ideal gas is doubled by an isothermal expansion that proceeds (1) reversibly or (2) irreversibly?

**Problem 78.** When the volume of an ideal gas is reduced two-fold by a reversible isothermal compression, what is the change of entropy of (1) the gas and (2) the surroundings?

**Problem 79.** Under 1 atm pressure at 279 K, the reversible crystallization of benzene from its melt is accompanied by an entropy change of -8.53 cal/mol.K. What is the molar heat of fusion of benzene at its melting point?

**Problem 80.** The heat capacity at constant pressure \( (c_P) \) for carbon monoxide is 7.0 cal/mol.K (ignore temperature dependences). What is the entropy change when a mole of CO is (1) heated from 100 K to 200 K, or (2) cooled from 1500 K to 750 K?

**Problem 81.** At 1000 K, \( 4.49 \times 10^{-2} \) as many molecules in a sample of carbon monoxide (CO) are in the first excited vibrational-energy state as
the ground state. How far above the ground state is the first excited state, in kJ/mol?

**Solution.** Given $T = 1000$ K, molecule=CO (carbon monoxide) and $N_1/N_0 = 4.49 \times 10^{-2}$ (vibrational energy). Let $x$=number of molecules in the ground state (vibrational) and $4.49 \times 10^{-2}x$ = number of molecules in the 1st excited vibrational state. Using the equation

$$\frac{N_i}{N_j} e^{-(\epsilon_i-\epsilon_j)/k_BT} = e^{-\Delta\epsilon/k_BT}$$

$$\frac{N_1}{N_0} = \frac{4.49 \times 10^{-2}x}{x} = e^{-\Delta\epsilon/(1.38 \times 10^{-23} \text{ J/K})(1000 \text{ K})}$$

$$\log(4.49 \times 10^{-2}) = \frac{-\Delta\epsilon}{1.38 \times 10^{-20} \text{ J}}$$

$$(-3.10)(1.38 \times 10^{-20} \text{ J}) = -\Delta\epsilon$$

$$\Delta\epsilon = 4.28 \times 10^{-20} \text{ J}$$

Now, to convert to the final units of kJ/mol.

$$(4.28 \times 10^{-20} \text{ J}) \left( \frac{1 \text{ kJ}}{1000 \text{ J}} \right) \left( \frac{6.022 \times 10^{23}}{1 \text{ mol}} \right) = 25.8 \text{ kJ/mol}$$

**Problem 82.** Molecular nitrogen has a vibrational energy $h\nu = 2230$ cm$^{-1}$. A sample of nitrogen is heated in an electric arc. Spectroscopic methods are used to determine the relative populations of excited vibrational levels. The results are presented in the table below

<table>
<thead>
<tr>
<th>Vibrational level (v)</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>...</th>
</tr>
</thead>
<tbody>
<tr>
<td>$N_v/N_0$</td>
<td>1.000</td>
<td>0.200</td>
<td>0.040</td>
<td>0.008</td>
<td>0.002</td>
<td>...</td>
</tr>
</tbody>
</table>

Use these results in combination with the information about the spacing of vibrational energy levels to determine whether the nitrogen is in thermodynamic equilibrium with respect to vibrational energy. In other words, does the vibrational population obey the Maxwell-Boltzmann distribution law? If so, what is the vibrational temperature of the gas? Is this value necessarily the same as the translational temperature of the gas? Why or why not? Hint: try to rearrange the Maxwell-Boltzmann equation appropriate for this problem so that there is a linear relationship between $v$ and some function of $N_v/N_0$. Then plot the data and calculate the slope.

**Solution.** Given: molecule is $N_2(g)$ and $h\nu = 2230$ cm$^{-1}$. Relative populations of the first four excited state to the ground state. One way to approach this problem is to rearrange the Maxwell-Boltzmann distribution (equation) to look like a linear function, graph the function and calculate $T$ from the slope. We are looking at vibrational energy, so we use:
\[ \epsilon_v = (v + 1/2)h\nu \] and \[ N_v/N_0 = e^{-(\epsilon_v-\epsilon_0)/k_BT} = e^{-(v+1/2)h\nu-(1/2)h\nu)/k_BT} , \]
where the second term in the exponent is from \( \epsilon_0 = (0 + 1/2)h\nu \). Therefore, \( N_V/N_0 = e^{-\epsilon v/k_BT} \). Linearize this equation by taking the log of both sides:

\[ \log(N_v/N_0) = -n k\nu/k_BT \]

or \( \log(N_v/N_0) = -v(h\nu/k_BT) \) which is of the form \( y = mx \).

The slope of this line is \( m = -1.5649 = -k\nu/k_BT \). Note: depending on your program and graphing the data, the slope value may differ from mine.

Solving for \( T \), we get: \( T = 2050K \), which becomes \( T = 2000 \) K (1 sig. fig.).

The \( T_v \) is usually the same as \( T_{\text{translational}} \) but doesn’t have to be under certain conditions.

**Problem 83.** A tank is filled with 1000 g of nitrogen at 0.00°C and 16.0 atm pressure. The tank is then heated to 50.00°C and the valve is opened. What is the total mass (in grams) of the nitrogen that escapes if the external pressure is 1.00 atm. and the temperature is maintained at 50.00°C?

**Solution.** Using \( PV = nRT \) we can calculate the volume of the tank from the initial data. \( PV = nRT, V = nRT/P, n =? \).

\[ n = (1000. \text{ g N}_2(g)) \left( \frac{1 \text{ mol N}_2(g)}{28.00 \text{ g N}_2(g)} \right) = 35.714 \text{ mol N}_2(g) \]

\[ V = \frac{(35.714 \text{ mol})(0.08206 \text{ L.atm/mol/K})(273.15 \text{ K})}{16.0 \text{ atm}} = 50.03 \text{ L} \rightarrow 50.0 \text{ L} \]

This gas will continue to escape until the internal pressure equals the external pressure of 1.00 atm. The amount of gas remaining in the tank:

\[ n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(50.03 \text{ L})}{(0.08206 \text{ L.atm/mol/K})(323.15 \text{ K})} = 1.8867 \text{ mol N}_2(g) \]

\[ (1.8867 \text{ mol N}_2(g)) \left( \frac{28.00 \text{ g N}_2(g)}{1 \text{ mol N}_2(g)} \right) = 52.827 \text{ g N}_2(g) \]

The amount of gas that escaped = 1000. g-52.827 g=947.172 g → 927 g \( \text{N}_2(g) \).

**Problem 84.** The Lennard-Jones potential is as follows: \( V(r) = 4\epsilon[(\sigma/r)^{12}-(\sigma/r)^6] \), where \( \epsilon \) and \( \sigma \) are constants. The \( \epsilon \) corresponds to the potential well depth at the equilibrium intermolecular (interatomic) distance. (i.e. where the potential energy is a minimum). Calculate the intermolecular
(interatomic) distance in Å that corresponds to $\epsilon$ for a He..He interaction. Derive an expression in terms of $\sigma$ for this distance for any atomic/molecular interaction.

**Solution.** You might start out by sketching a potential energy curve. From such a graph we see that the distance we need to calculate occurs at the minimum of the function, so one technique to calculate this $r$ is to take the derivative of the function and set it equal to zero and solve for an $r$ value that makes the derivative equal to zero.

You may also want to find the constants for He.

$$
\epsilon = 1.41 \times 10^{-22} \text{ J}, \quad \sigma = 2.56 \times 10^{-10} m = 2.56 \text{ Å}.
$$

$$
V(r) = 4\epsilon \left[ \left( \frac{\sigma}{r} \right)^{12} - \left( \frac{\sigma}{r} \right)^{6} \right] - 4\epsilon \sigma^{6} r^{-6}.
$$

$$
\frac{dV(r)}{dr} = (-12)(4\epsilon)\sigma^{12} r^{-11} - (-6)(4\epsilon)(\sigma^{6})(r^{-5})
$$

$$
= (-6)(4\epsilon)(\sigma) \left[ 2 \left( \frac{\sigma}{r} \right)^{11} - \left( \frac{\sigma}{r} \right)^{5} \right] = 0
$$

When the square bracket equals zero, $dV/dr = 0$. Thus, $2(\sigma^{11}/r^{11}) - (\sigma^{5}/r^{5}) = 0$ and $2(\sigma^{11}/r^{11}) = (\sigma^{5}/r^{5})$ which gives $2(\sigma)^{6} = (r)^{6}$ and $r = (2\sigma^{6})^{1/6} = \sigma(2)^{1/6} = \sigma(1.122)$. This is the general expression in terms of $\sigma$ for the distance $r$. Now, plug in $\sigma$ for He to obtain the specific answer

$$
\epsilon = (2.56 \text{ Å})(1.122) = 2.87 \text{ Å}
$$

for the He...He interaction.

**Problem 85.** The vapor pressures of each component in a mixture of propanone (acetone, A) and trichloromethane (chloroform, C), were measured at 35°C with the following results:

<table>
<thead>
<tr>
<th>$\chi_c$</th>
<th>0</th>
<th>0.2</th>
<th>0.4</th>
<th>0.6</th>
<th>0.8</th>
<th>1.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>$P_C$/torr</td>
<td>347</td>
<td>270</td>
<td>185</td>
<td>102</td>
<td>37</td>
<td>0</td>
</tr>
<tr>
<td>$P_A$/torr</td>
<td>35</td>
<td>82</td>
<td>142</td>
<td>219</td>
<td>293</td>
<td></td>
</tr>
</tbody>
</table>

Confirm that the mixture "conforms" to Raoult’s law for the component in large excess and to Henry’s law for the component that is minor. Estimate the Henry’s law constants, $K_A$ and $K_C$.

**Solution.** You should graph this data and realize that it shows a general conformity to Raoult’s law by both components. Both components exhibit a negative deviation to Raoult’s law. So, at large mole fractions the component obeys Raoult’s law (experimental data will approach the line) and at small mole fractions the component obeys Henry’s law.
Problem 86. Ethanol (CH$_3$CH$_2$OH) has a normal boiling point, $T_b$ of 78.4°C. When 46.58 g of Na$_2$SO$_4$ (sodium sulfate) is dissolved in the ethanol (mass = 1000 g) the boiling point becomes 79.6°C. Calculate the $K_B$ for ethanol (including appropriate units). Assume the salt completely dissolves in the ethanol and behaves as a strong electrolyte.

Solution. We will use: $\Delta T_b = K_b \cdot m$. The solute, Na$_2$SO$_4$ has a molar mass of 142.04 g/mol. There is 1000 g of solvent (1.000 kg solvent). Therefore, the molality can be calculated. Strong electrolyte:

$$\text{Na}_2\text{SO}_4 \rightarrow 2\text{Na}^+ + \text{SO}_4^{2-}$$

$$\frac{46.58 \text{ g Na}_2\text{SO}_4}{142.04 \text{ g Na}_2\text{SO}_4} = 0.3279 \text{ mol Na}_2\text{SO}_4 \times 3 = 0.9837 \text{ mol particles}$$

$$m = \frac{0.9837 \text{ mol particles}}{1.000 \text{ kg solvent}} = 0.9837 \text{ m}$$

So, $\Delta T = T_{\text{new}} - T_b = K_b \cdot m = 79.6^\circ\text{C} - 78.4^\circ\text{C} = 1.2^\circ\text{C}$. So, $1.2^\circ\text{C} = (K_b)/(0.9837 \text{ mol/kg})$ and $K_b = \frac{1.2^\circ\text{C}}{0.9837 \text{ mol/kg}} = 1.22^\circ\text{C} \text{ kg/mol}$ which gives $1.2^\circ\text{C} \text{ kg/mol}$.

Problem 87. In thermodynamics, the partial derivatives of physical quantities are related to one another by application of the commutativity property of partial derivatives, $\partial_x \partial_y f(x, y) = \partial_y \partial_x f(x, y)$, which follows from the equality of mixed partial derivatives for sufficiently smooth functions. For example, if the internal energy is expressed as

$$dU = TdS + \mu dN$$

it follows that

$$T = \frac{\partial U}{\partial S} \bigg|_N \quad \text{and} \quad \mu = \frac{\partial U}{\partial N} \bigg|_S$$

The joint second derivative of $U$ is then given by

$$\frac{\partial^2 U}{\partial S \partial N} = \frac{\partial^2 U}{\partial N \partial S} = \frac{\partial T}{\partial N} \bigg|_S = \frac{\partial \mu}{\partial S} \bigg|_N$$

Since $(\partial y/\partial x) = (\partial x/\partial y)^{-1}$, the above equation can be inverted to give

$$\frac{\partial S}{\partial \mu} \bigg|_N = \frac{\partial N}{\partial T} \bigg|_S.$$
Similar identities can be obtained from the variations of other state functions. Show that:

\[
\left( \frac{\partial T}{\partial V} \right)_S = - \left( \frac{\partial P}{\partial S} \right)_V = \frac{\partial^2 U}{\partial S \partial V}
\]
\[
\left( \frac{\partial T}{\partial P} \right)_S = + \left( \frac{\partial V}{\partial S} \right)_P = \frac{\partial^2 H}{\partial S \partial P}
\]
\[
\left( \frac{\partial S}{\partial V} \right)_T = + \left( \frac{\partial P}{\partial T} \right)_V = - \frac{\partial^2 A}{\partial T \partial V}
\]
\[
\left( \frac{\partial S}{\partial P} \right)_T = + \left( \frac{\partial V}{\partial T} \right)_P = \frac{\partial^2 G}{\partial T \partial P}
\]

Recall that \(U = U(S, V), \ H = H(S, P), \ A = A(T, V)\) and \(G = G(T, P)\).

**Solution.** These “Maxwell relations” are all derived the same way. Take the third one, for example, \(A(T, V) = U - TS\) and \(dA = -SdT - pdV\) lead to

\[
S = - \left( \frac{\partial A}{\partial T} \right)_V \quad \text{and} \quad P = - \left( \frac{\partial A}{\partial V} \right)_T
\]

Equating the second partial derivatives \((\partial^2 A / \partial T \partial V)\) we get

\[
\left( \frac{\partial S}{\partial V} \right)_T = \left( \frac{\partial V}{\partial T} \right)_P
\]

similarly for all other relations.

**Problem 88.** The reaction

\[
\frac{1}{2} O_2(g) + H_2(g) \rightarrow H_2O(l)
\]

is used to produce an electrical current. The details of the production of electrical current are not important for solving this problem; we only need to assume that the free energy can be converted into current (electrical work) through some process.

From tables of thermodynamic data, we find the following information in J/mol.K for the standard entropies of formation (\(\Delta S^{\circ}_f\)): \(O_2(g): 205.0, \ H_2(g): 130.6, \ H_2O(l): 70.0\). Standard enthalpy of formation for water, \(\Delta H^{\circ}_f = -285.9\) kJ/mol; those for \(H_2\) and \(O_2\) are zero.

Use this information to find:

a) The amount of heat released if the reaction were to take place by direct combustion (of oxygen and hydrogen)

b) The amount of electrical work the same reaction can perform when carried out in a fuel cell at 298K under reversible conditions

c) The amount of heat released under the same conditions (when electrical current is produced)
6.12. Problems

Solution. First, we need to find $\Delta H^\circ_{\text{rxn}}$ and $\Delta S^\circ_{\text{rxn}}$ for the process. Recalling that the standard enthalpy of formation of the elements is zero, $\Delta H^\circ_{\text{rxn}} = H_f^\circ(\text{products}) - H_f^\circ(\text{reactants}) = -285.9 \text{ kJ mol}^{-1}$. Similarly, $\Delta S^\circ_{\text{rxn}} = S_f^\circ(\text{products}) - S_f^\circ(\text{reactants}) = (70.0) - \left(\frac{1}{2} \times 205.0 + 130.6\right) = -163 \text{ JK}^{-1}\text{mol}^{-1}$.

a) When the hydrogen and oxygen are combined directly, the heat released will be $\Delta H^\circ_{\text{rxn}} = -285.9 \text{ kJ mol}^{-1}$.

b) The maximum electrical work the fuel cell can perform is given by $\Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T \Delta S^\circ_{\text{rxn}} = -285.9 \text{ kJ mol}^{-1} - (298 \text{ K})(-163 \text{ JK}^{-1}\text{mol}^{-1}) = -237.2 \text{ kJ mol}^{-1}$.

c) The heat released in the reaction is the difference between the enthalpy change (the total energy available) and the reversible work that was expended: $\Delta H^\circ_{\text{rxn}} - \Delta G^\circ_{\text{rxn}} = T \Delta S^\circ_{\text{rxn}} = (298 \text{ K})(-163 \text{ J.K}^{-1}\text{mol}^{-1}) = -48800 \text{ J mol}^{-1} = -48.8 \text{ kJ mol}^{-1}$.

Problem 89. Two ideal gases at constant temperature and pressure are separated by a partition. There are 0.07 mol of gas A and 0.12 mol of gas B. Find the entropy change of the system after the partition is released and the two gases are allowed to mix. It is useful to know that entropy is a state function and that the entropy of a system is given by $S = k_B \log \Omega$ where $\Omega$ is the number of accessible microstates to the system and $\log$ is the natural logarithm. For an ideal gas, $\Omega \propto V^N$ where $V$ is the volume of the gas and $N$ is the number of gas particles. Since $S$ depends on $N$ through $\Omega$, $S$ is an extensive property. To find the entropy change, you will want to consider the system as two subsystems, one consisting of gas A and the other consisting of gas B, and then calculate the entropy of the system before and after the gases mix. The change is equal to the difference between pre- and post-mixing.

Solution. Since entropy is a state function, the entropy change is equal to the entropy of the mixture subtracted by the entropy of the two gases prior to mixing.

$$\Delta S = S_{\text{mixture}} - S_A - S_B$$

$$S_{\text{mixture}} = k_B \log(C(V_A + V_B)^{N_A+N_B})$$

where $C$ is some constant of proportionality.

$$S_{\text{mixture}} - S_A - S_B = k_B(N_A+N_B)\log(C(V_A + V_B)) - k_B N_A \log(CV_A) - k_B N_B \log(CV_B)$$

$$= k_B N_A \log \left(\frac{V_A + V_B}{V_A}\right) + k_B N_B \log \left(\frac{V_A + V_B}{V_B}\right)$$
Since we are dealing with ideal gases, we can relate the volumes to moles.

\[ \frac{V_A + V_B}{V_A} = \frac{n_A + n_B}{n_B} \]

Next we insert \( k_B = \frac{R}{N_{\text{Avogadro}}} \) and \( X_A = \frac{n_A + n_B}{n_A} \).

\[ \Delta S = R(n_A \log X_A + n_B \log X_B) \]

\[ = 8.314 \text{ J K}^{-1} \left[ 0.07 \log \left( \frac{0.19}{0.07} \right) + 0.12 \log \left( \frac{0.19}{0.12} \right) \right] = 1.04 \text{ J K}^{-1} \]

**Problem 90.** A non-uniform chemical potential \( (d\mu/dx \neq 0) \) leads to a mass flux from regions of high concentration to regions of low concentration. Let’s see how we can relate this gradient (slope, \( d\mu/dx \)) to the diffusion of particles. We know that the chemical potential for an ideal gas is of the form

\[ \mu = \mu_0 + k_B T \log(c/c_0) \]

where \( \mu_0 \) is a reference chemical potential (only depends on \( T \), but otherwise a constant), \( c(x,t) \) is the concentration ([M]) of a chemical species of interest at position \( x \) and time \( t \), and \( c_0 \) is a reference concentration (say, 1 M).

Let’s first establish the principle of “conservation of mass”. Let us consider a 1D flow of particles along the \( x \) direction and a “volume element” of length \( dx \) centered at \( x + \frac{dx}{2} \). Since we are working in 1D we may take the concentration \( c(x,t) \) to have units of particles per unit length. The flow of particles can be described in terms of the particle flux, \( J = c(x,t)v \), where \( v \) is the velocity of particles. \( J \) has units of particles per unit time. Of course, \( J \) represents a net flux, meaning that \( J = 0 \) describes equal number of particles moving to the left as to the right. Thus, when \( J = 0 \), there is no change in the number of particles in the volume element \( dx \). Therefore, its rate of change is zero: \( \partial c/\partial t = 0 \). Moreover, even when \( J = \text{constant} \), we still have \( \partial c/\partial t = 0 \) because even though there may be a net flux of particles, the number of particles that enter \( dx \) equals the number that leaves \( dx \) during some time interval \( dt \). The only way that \( \partial c/\partial t \) can be nonzero is if \( J \) is not spatially uniform. If there is a slope in \( J \), there will be a change in \( c \) over time.

In a time interval \( dt \), the increase in concentration with time, \( \partial c/\partial t \), will equal the excess of molecules diffusing into the region at position \( x \) over those diffusing out at position \( x + dx \), divided by the volume \( (dx) \):

\[ \frac{\partial c}{\partial t} = \frac{1}{dx} [J(x) - J(x + dx)] \]
(a) Show that $\frac{\partial c}{\partial t} = -\frac{\partial J}{\partial x}$. This is a statement of the conservation of mass. Explain why/how this equation amounts to mass conservation.

Next we must deal with the particle flux $J = c(x, t)v$. What should $v$ be? It turns out that according to linear response theory, systems near equilibrium have currents proportional to the first derivative (gradient) of their properties. The examples given in class are Fourier’s law, $J = k(\partial T/\partial x)$ (heat flow proportional to temperature gradient), and Fick’s law, $J = -D(\partial \mu/\partial x)$ (particle flux proportional to the chemical potential gradient). Other examples include Ohm’s law (electrical current proportional to the gradient of the electromagnetic potential) and in viscous fluids, where the shear rate is proportional to the shear stress.

Thus, we set $v = -\gamma(\partial \mu/\partial x)$, where $\gamma$ is a coefficient of mobility. This gives $J \propto \partial \mu/\partial x$. Namely, $J = -\gamma c(x, t)\partial \mu/\partial x$.

b) Insert the ideal gas expression for $\mu$ into $J = -\gamma c(x, t)\partial \mu/\partial x$ to find $J = -\gamma k_B T \partial c/\partial x$. Then use the expression found in part (a) to arrive at a diffusion equation where $D = \gamma k_B T$. Confirm that this diffusion equation is identical to the one derived in class using Einstein’s method.

c) Consider two solutes, $A$ and $B$, dissolved in solution. Write an expression to describe the diffusion of the sum of $A$ and $B$ and call $c_{AB}(x, t) = c_A(x, t) + c_B(x, t)$. (Hint: no work is needed here, just borrow the result you derived in the previous question.)

**Solution.** a) Taylor expansion of $J(x + dx)$ gives

$$J(x + dx) = J(x) + dx \left( \frac{\partial J}{\partial x} \right) + O(dx^2)$$

where $O(dx^2)$ are terms of order equal to or higher than $dx^2$. Hence, in the limit $dx \to 0$:

$$\frac{\partial c}{\partial t} = \frac{1}{dx} \left[ J(x') - J(x) - dx \left( \frac{\partial J}{\partial x} \right) + O(dx^2) \right] = -\frac{\partial J}{\partial x}$$

b) Inserting the expression for $\mu$ for the ideal gas, we find:

$$J = -\gamma c(x, t)\frac{\partial \mu}{\partial x} = -\gamma c(x, t) \frac{\partial (k_B T \log(c(x, t)/c_0))}{\partial x}$$

$$= -\gamma c(x, t) \frac{k_B T}{c} \frac{\partial c}{\partial x} = -\gamma k_B T \frac{\partial c}{\partial x}$$

---

36Hint: Taylor expand $J(x + dx)$. Recall from calculus that the Taylor expansion of $f(x)$ at $x$ is the series $\sum_{n=0}^{\infty} \frac{f^{(n)}(a)}{n!}(x - a)^n$, where $f^{(n)}$ denotes the n-th derivative of $f$ evaluated at the point $a$. In practice (i.e. in the physical sciences), only the lowest order terms are needed when $x$ is close to $a$. In that case, we only need Taylor’s theorem, $f(x) = f(a) + f'(a)(x-a) + O((x-a)^2)$, where $O((x-a)^2)$ is a term that tends to zero as $(x-a)$ goes to zero.
and thus the rate of change of $c(x, t)$ is given by the diffusion equation

$$\frac{\partial c}{\partial t} = - \frac{\partial J}{\partial x} = \gamma k_B T \frac{\partial^2 c}{\partial x^2}$$

where $D = \gamma k_B T$.

c) The above treatment is for ideal solutions by choice of $\mu = \mu_0 + k_B T \log(c/c_0)$. Thus, solutes $A$ and $B$ are non-interacting and the diffusion of the sum, $c_{AB}(x, t)$, is the sum of the parts, $c_A(x, t)$ and $c_B(x, t)$.

$$c_{AB}(x, t) = c_A(x, t) + c_B(x, t)$$

$$\frac{\partial c_{AB}}{\partial t} = \frac{\partial c_A}{\partial t} + \frac{\partial c_B}{\partial t} = \gamma_A k_B T \frac{\partial^2 c_A}{\partial x^2} + \gamma_B k_B T \frac{\partial^2 c_B}{\partial x^2}$$

Problem 91. For the given equation of states (a) and (b) below, find the 2nd virial coefficient, $B_2(T)$, by both methods (1) and (2) which are explained below. Recall that the virial expansion is a polynomial expansion of $P/k_B T$ with respect to the number density $\rho = N/V$. To find $B_2(T)$, you will want to re-write the given equation of state in the proper form on the left hand side, and then do a polynomial expansion with respect to $\rho$ on the remaining terms on the right hand side. Recall that a Taylor expansion of $f(\rho)$ at $\rho = 0$ will yield a polynomial expansion. Note: You may want to rationalize why it is a sound idea to Taylor expand $P/k_B T$ at $\rho = 0$. Once you have the virial expansion, you can find $B_2(T)$ by two methods.

(1) You can group terms by powers of $\rho$ and then “select” the coefficient that goes with $\rho^2$.

(2) You can divide your virial expansion by $\rho$, subtract by one, divide by $\rho$ again, then take the limit as $\rho \to 0$ for the remaining terms. You may recognize the term $Z = \frac{PV}{Nk_B T}$.

(a) $P = \frac{RT}{V_{m-b} - a \sqrt{TV_m(V_m+b)}}$

(b) $P = \frac{RT}{V_{m-b}} \frac{a}{V_m (1 - \exp(b/V_m))}$

Solution. The general strategy is to write the given equation of state in the form of the virial expansion $\frac{P}{k_BT} = \rho + B_2(T)\rho^2 + B_3(T)\rho^3 + \ldots$ and then compare orders of $\rho$ to determine the second virial coefficient, $B_2(T)$. Alternatively, we can manipulate the expression then take the limit as $\rho \to 0$. Both methods are demonstrated below.

(a) First we rewrite the expression to isolate constants and then do Taylor expansions with respect to $\rho$ to write the equation of state in powers of $\rho$.

$$P = \frac{Nk_B T}{V} \frac{1}{1 - \frac{N_b}{V}} = \frac{a N^2}{\sqrt{TV^2 N_A}} \frac{1}{1 - \frac{N_b}{V N_A}}$$
\[
\rho k_B T \frac{1 - \frac{a \rho^2}{\sqrt{T N_A^2}}}{1 - \frac{b \rho}{N_A}}
\]

By Taylor expansion,
\[
\frac{1}{1 - b \rho} = 1 + b \rho + (b \rho)^2 + \ldots
\]
\[
\frac{1}{1 - \frac{b \rho}{N_A}} = 1 + \frac{b \rho}{N_A} + \left(\frac{b \rho}{N_A}\right)^2 + \ldots
\]

Now utilizing the Taylor expansions and fitting to the form of the virial expansion,
\[
\frac{P}{k_B T} = \rho + \rho^2 \left(b - \frac{a}{\sqrt{T N_A^2}}\right) + \mathcal{O}(\rho^3)
\]

where \(\mathcal{O}(\rho^3)\) means terms of order equal to or higher than \(\rho^3\). By method (1), clearly \(B_2(T) = \left(b - \frac{a}{\sqrt{T N_A^2}}\right)\). Alternatively by method (2), we can manipulate the expression then take limit as \(\rho \to 0\). Let \(Z = \frac{P}{k_B T \rho}\).
\[
Z = 1 + \rho \left(b - \frac{a}{\sqrt{T N_A^2}}\right) + \mathcal{O}(\rho^2)
\]
\[
\lim_{\rho \to 0} \frac{Z - 1}{\rho} = \lim_{\rho \to 0} \left(b - \frac{a}{\sqrt{T N_A^2}} + \mathcal{O}(\rho)\right) = b - \frac{a}{\sqrt{T N_A^2}} = B_2(T)
\]
Since \(\mathcal{O}(\rho)\) depends on \(\rho\) or higher orders of \(\rho\), it will vanish as \(\rho \to 0\).

(b) We use the same strategy as in part (a)
\[
P = k_B T \rho \left(\frac{1}{1 - b \rho}\right) - k_B T \rho (1 - \exp(b \rho / N_A))
\]

By Taylor expansion,
\[
\frac{1}{1 - b \rho} = 1 + b \rho + (b \rho)^2 + \ldots
\]
\[
1 - \exp(b \rho / N_A) = \frac{b \rho}{N_A} + \frac{1}{2!} \left(\frac{b \rho}{N_A}\right)^2 + \ldots
\]

Now plugging in those Taylor expansions and fitting to the form of virial expansion,
\[
\frac{P}{k_B T} = \rho + \rho^2 \left(b + \frac{b}{N_A}\right) + \mathcal{O}(\rho^3)
\]
Since we have grouped terms by order of \(\rho\) it is clear that
\[
B_2(T) = b + \frac{b}{N_A}
\]
Alternatively,
\[
Z = 1 + \rho \left(b + \frac{b}{N_A}\right) + \mathcal{O}(\rho^2)
\]
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\[ \lim_{\rho \to 0} \frac{Z - 1}{\rho} = \lim_{\rho \to 0} \left( b + \frac{b}{N_A} + \mathcal{O}(\rho) \right) = b + \frac{b}{N_A} = B_2(T) \]

**Problem 92.** Thermodynamic relationships can be used to accurately describe the slopes of phase diagram coexistence curves. Specifically, the Clapeyron equation can be used to characterize the discontinuous phase transition of a single material.

a) Show that at constant temperature \( T \) and pressure \( P \), the infinitesimal change in Gibbs free energy can be written as

\[ dG = (\mu^g - \mu^l)dn^g \]

Note that for a liquid-gas phase change, \( dn^l = -dn^g \), the Gibbs free energy of a liquid-gas mixture is given by \( G = G^l + G^g \), and the chemical potential \( \mu \) can be defined as

\[ \mu = \left( \frac{\partial G}{\partial n} \right)_{P,T} \]

b) Derive the Clapeyron equation (subscript \( t = \) phase transition),

\[ \frac{dP}{dT} = \frac{\Delta_t \overline{H}}{T \Delta_t \overline{V}} \]

by taking the total derivative* of both sides of the following expression

\[ \mu^\alpha(T, P) = \mu^\beta(T, P) \]

where \( \alpha \) and \( \beta \) represent two states of a phase transition at equilibrium. You may find the following relationships useful, where the overline indicates the intensive molar quantity

\[ \left( \frac{\partial \mu}{\partial P} \right)_T = \left( \frac{\partial \overline{G}}{\partial P} \right)_T = \overline{V} \quad \text{and} \quad \left( \frac{\partial \mu}{\partial T} \right)_P = \left( \frac{\partial \overline{G}}{\partial T} \right)_P = -\overline{S} \]

\[ \Delta_t \overline{S} = \frac{\Delta_t \overline{H}}{T} \]

*The total derivative \( df \) is best explained by an example: consider a function \( f(x, y) \). Its total derivative is

\[ df = \left( \frac{\partial f}{\partial y} \right)_x dy + \left( \frac{\partial f}{\partial x} \right)_y dx. \]

c) How much does the melting point of benzene increase per atmosphere of pressure (within the vicinity of 1 atm)? Take \( \Delta_{\text{fus}} \overline{H} \) (at 278.7 K) to be 9.95 kJ mol\(^{-1}\) and \( \Delta_{\text{fus}} \overline{V} \) (at 278.7 K) to be 10.3 cm\(^3\) mol\(^{-1}\).
Solution. a) Take the derivative of $G = G^l + G^g$ with respect to $n$ at constant $T$ and $P$.

$$dG = \left( \frac{\partial G^g}{\partial n^g} \right)_{P,T} dn^g + \left( \frac{\partial G^l}{\partial n^l} \right)_{P,T} dn^l$$

Since $dn^l = -dn^g$,

$$dG = \left[ \left( \frac{\partial G^g}{\partial n^g} \right)_{P,T} - \left( \frac{\partial G^l}{\partial n^l} \right)_{P,T} \right] dn^g$$

$$dG = (\mu^g - \mu^l)dn^g \quad \text{(constant } T \text{ and } P)$$

b) Take the total derivative

$$\left( \frac{\partial \mu^e}{\partial P} \right)_T dP + \left( \frac{\partial \mu^e}{\partial T} \right)_P dT = \left( \frac{\partial \mu^\beta}{\partial P} \right)_T dP + \left( \frac{\partial \mu^\beta}{\partial T} \right)_P dT$$

Evoking the relationships

$$\left( \frac{\partial \mu}{\partial P} \right)_T = \left( \frac{\partial G}{\partial P} \right)_T = \nabla \quad \text{and} \quad \left( \frac{\partial \mu}{\partial T} \right)_P = \left( \frac{\partial G}{\partial T} \right)_P = -\mathcal{S}$$

results in

$$\nabla^\alpha dP - \mathcal{S}^\alpha dT = \nabla^\beta dP - \mathcal{S}^\beta dT$$

Since this expression is in equilibrium, it is appropriate to evoke

$$\Delta_t \mathcal{S} = \frac{\Delta_t \mathcal{H}}{T}$$

where it follows that

$$\frac{dP}{dT} = \frac{\Delta_t \mathcal{H}}{T \Delta_t \nabla}$$

c) $\frac{dP}{dT}$ evaluates to:

$$\frac{9950 \text{ J mol}^{-1}}{(278.68 \text{ K})(10.3 \text{ cm}^3 \text{ mol}^{-1})} \left( \frac{10 \text{ cm}}{1 \text{ dm}} \right)^3 \left( \frac{0.08206 \text{ dm}^3 \text{ atm} \text{ mol}^{-1} \text{ K}^{-1}}{8.314 \text{ J mol}^{-1} \text{ K}^{-1}} \right) = 34.2 \text{ atm K}^{-1}$$

Taking the reciprocal

$$\frac{dT}{dP} = 0.0292 \text{ K atm}^{-1}$$

Problem 93. A chemical reaction is at equilibrium, according to its equilibrium constant, $K$. It is possible to estimate the equilibrium constant at temperatures other than standard conditions by using the enthalpy of reaction $\Delta H_{rxn}^\circ$ and the van’t Hoff equation.
a) First, show that the expression
\[ \Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^\circ + RT \log(Q) \]
follows from the definition
\[ dG = \sum J \mu_J dn_J \]
Recall that \( \mu_J = \mu_J^\circ + RT \log(a_J) \) where \( \mu_J \) and \( a_J \) are the chemical potential and activity of species \( J \), respectively. Furthermore, the extent of reaction \( \xi \) can be defined by the expression \( dn_J = v_J d\xi \) where \( v_J \) is the stoichiometric number of species \( J \).

b) Derive the van’t Hoff equation
\[ \frac{d(\log K)}{d(1/T)} = -\frac{\Delta H_{\text{rxn}}^\circ}{R} \]
by taking the derivative with respect to temperature of \( \Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^\circ + RT \log Q \) at equilibrium standard conditions. Use the Gibbs-Helmholtz equation
\[ \frac{d(\Delta G_{\text{rxn}}^\circ/T)}{dT} = -\frac{\Delta H_{\text{rxn}}^\circ}{T^2} \]
in your derivation.

c) Consider the synthesis of ammonia at 298 K
\[ N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad K = 6.1 \times 10^5 \]
and using the van’t Hoff equation, estimate the equilibrium constant at 500 K. Take \( \Delta H_{\text{rxn}}^\circ = -92.2 \text{ kJ mol}^{-1} \) and assume it is constant over this temperature range.

**Solution.** a) Express the infinitesimal change in Gibbs free energy in terms of extent of reaction \( \xi \)
\[ dG = \sum J \mu_J dn_J = \left( \sum J v_J \mu_J \right) d\xi \]
It follows that
\[ \Delta G_{\text{rxn}} = \left( \frac{\partial G}{\partial \xi} \right)_{P,T} = \sum J v_J \mu_J \]
Substituting the chemical potential of species \( J \)
\[ \Delta G_{\text{rxn}} = \sum J v_J \mu_J^\circ + RT \sum J v_J \log(a_J) = \Delta G_{\text{rxn}}^\circ + RT \sum J \log(a_J^{v_J}) \]
where the identity \( a \log x = \log x^a \) has been used. Finally, evoking the identity \( \log x + \log y + \ldots = \log(xy\ldots) \) results in

\[
\Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^\circ + RT \prod_j a_j^{v_j} = \Delta G_{\text{rxn}}^\circ + RT \log(Q)
\]

where

\[ Q = \prod_j a_j^{v_j} \]

b) Taking the derivative with respect to \( T \) of the following

\[
\log K = -\frac{\Delta G_{\text{rxn}}^\circ}{RT}
\]

which results in

\[
\frac{d \log K}{dT} = -\frac{1}{R} \frac{d(\Delta G_{\text{rxn}}^\circ/T)}{dT}
\]

Substitute the Gibbs-Helmholtz equation

\[
\frac{d \log K}{dT} = \frac{\Delta H_{\text{rxn}}^\circ}{RT^2}
\]

The required form of the van't Hoff equation can be obtained by substituting the following identity

\[
\frac{d(1/T)}{dT} = -\frac{1}{T^2} \quad \text{so} \quad dT = -T^2 d(1/T)
\]

c) Integrate the result of part (b) assuming constant standard enthalpy of reaction

\[
\log K_2 - \log K_1 = -\frac{1}{R} \int_{1/T_1}^{1/T_2} \Delta H_{\text{rxn}}^\circ d(1/T) = -\frac{\Delta H_{\text{rxn}}^\circ}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)
\]

\[
\log K_2 = \log(6.1 \times 10^5) - \frac{(-92.2 \times 10^2 \text{ J mol}^{-1})}{8.3145 \text{ J K}^{-1} \text{ mol}^{-1}} \left( \frac{1}{500 \text{ K}} - \frac{1}{298 \text{ K}} \right) = -1.71
\]

Thus

\[ K_2 = 0.18 \]

**Problem 94.** Consider the chemical equilibrium reaction under standard conditions

\[ \text{H}_2(g) \rightleftharpoons \text{H}(g) + \text{H}(g) \]

By what factor is the rate of the reverse reaction greater than that of the forward? Take the chemical potential of monatomic hydrogen gas \( \mu_\text{H}^\circ \) to be 203.26 kJ/mol.

**Solution.**

\[
\Delta G_{\text{rxn}}(T, P) = \sum_j v_j \mu_j
\]

\[
2(203.26 \text{ kJ/mol}) - 1(0) = 406.52 \text{ kJ/mol}
\]
\[ \Delta G_{rxn}^\circ(T, P) = -RT \log K_{eq} = 406.52 \times 10^3 \text{ kJ/mol} \]

\[ K_{eq} = \frac{K_+}{K_-} = 5.5 \times 10^{-72} \]

The reverse reaction occurs at a rate on the order of \(10^{72}\) times faster (under standard conditions).

**Problem 95.** Consider the following situation: (a) Let \( K = \exp \left( \frac{-\Delta G_{rxn}^\circ}{k_B T} \right) \).

If \( K_w = 10^{-14} \) at RT, will \( K_w \) be greater or less than \( 10^{-14} \) at \( T = 0^\circ\text{C} \)?

(b) At a different temperature \( K_w = 6.3 \times 10^{-13} \). Find pH of neutral water at this temperature. (c) Relate pH and pOH using the value of \( K_w \) given in the previous step.

**Solution.**

(a) \( K \propto \frac{1}{e^x} \), so it will be lower since the argument of the exponential will be larger.

(b) \( K_w = [H^+][OH^-] = [H^+]^2 \)

\[ \text{pH} = -\log(\sqrt{6.3 \times 10^{-13}}) = 6.10 \]

(c) \( \text{pH} + \text{pOH} = 12.2 \)

**Problem 96.** Liquid helium and liquid argon are both cryogenic fluids; \( \text{He}(l) \) boils at 4.21 K, and \( \text{Ar}(l) \) boils at 90.2 K. The specific heat of liquid helium near its boiling point is 4.25 J.K\(^{-1}\).g\(^{-1}\), and the specific heat of liquid argon near its boiling point is 1.12 J.K\(^{-1}\) g\(^{-1}\). The enthalpy of vaporization of \( \text{He}(l) \) is 25.1 J.g\(^{-1}\), and the enthalpy of vaporization of \( \text{Ar}(l) \) is 161.1 J.g\(^{-1}\). Discuss which is a better coolant near its boiling point and which is better at its boiling point.

**Solution.**

Near the boiling point, the larger the specific heat the better the coolant. It takes more energy from whatever is cooled in order to raise its temperature, thus better cooling the other substance. Thus, \( \text{He} \) is the better coolant near its boiling point.

At the boiling point, the larger the enthalpy of vaporization the better the coolant as it takes more energy from whatever is cooled for it to vaporize. Therefore, \( \text{Ar} \) is the better coolant at its boiling point.

**Problem 97.** Find the \( \Delta H_f^\circ \) and \( \Delta U_f^\circ \) for \( \text{SiH}_4(g) \) from its elements at 298 K, if 0.17 L of the compound at 298 K and \( P=0.981 \) bar is burned in a constant-volume calorimeter in an excess of oxygen and causes the evolution of 9.757 kJ of heat. The reaction is:

\[ \text{SiH}_4(g) + 2\text{O}_2(g) \rightarrow \text{SiO}_2(s, \text{quartz}) + 2\text{H}_2\text{O}(l) \]

and the formation of silane from its elements is:

\[ \text{Si}(s) + 2\text{H}_2(g) \rightarrow \text{SiH}_4(g) \]
6.12. Problems

Solution.

\[ n = \frac{PV}{RT} = \frac{(0.968 \text{ atm}) \cdot (0.17 \text{ L})}{(0.082057 \text{ L.atm/mol.K})(298 \text{ K})} = 0.00673 \text{ mol} \]

9.757 kJ / 0.00673 mol = 1450 kJ/mol = \( \Delta H^\circ_C \)

\[ \Delta H^\circ_C = 2\Delta H_f^\circ(\text{H}_2\text{O}(l)) + \Delta H_f^\circ(\text{SiO}_2(s, \text{quartz})) - \Delta H_f^\circ(\text{SiH}_4(g)) \]

\[ \Delta H_f^\circ(\text{SiH}_4(g)) = (2 \times -285.83 \text{ kJ/mol}) + (-910.94 \text{ kJ/mol}) - (-1450 \text{ kJ/mol}) \]

\[ \Delta H_f^\circ(\text{SiH}_4(g)) = -32.8 \text{ kJ/mol} \]

\[ \Delta H_f^\circ = \Delta U_f^\circ + \Delta n_fRT, \Delta n_f = 1 - 2 = -1 \]

\[ \Delta U_f^\circ = \Delta H_f^\circ - \Delta n_fRT = -32.8 \text{ kJ/mol} - (-1 \times 8.314 \text{ J/mol.K} \times 298 \text{ K} \times (1 \text{ kJ}/1000 \text{ J})) = -30.3 \text{ kJ/mol} \]

Problem 98. A container with 3 moles of ideal gas initially has a pressure and temperature of 5 atm and 300 K, respectively. Through an isothermal irreversible expansion against a constant pressure of 1 atm its final volume becomes 62 L. (a) Find the work done in the process. (b) Find the work done if the expansion had been done reversibly. (c) Why are (a) and (b) different?

Solution.

\[ dw = -P_{\text{ext}}dV \]

For a reversible process, \( P_{\text{ext}} = P \)

(a)

\[ w_{\text{irrev}} = -P_{\text{ext}}\Delta V = -(1 \text{ atm})(62 \text{ L} - 14.8 \text{ L}) = -47.2 \text{ atm L} \]

(b)

\[ w_{\text{rev}} = -nRT \log(V_2/V_1) = -105.8 \text{ atm.L} \]

(c) Work is not a state function. The reversible work will always be greater.

Problem 99. (a) If an ideal gas is reversibly heated in such a way that its volume is exponentially dependent on the temperature (\( V = [1.3 \text{ L}]e^{(0.01 K^{-1})T} \)), what is the relationship between pressure and temperature? Find the work done by the system during an isothermal expansion from \( V_1 \) to \( V_2 \). (b) What is the change in the internal energy of a system with 3 moles that is heated from 100 K to 250 K? Would it be different or equal if the heating had been done through a non-reversible process?

Solution. (a)

\[ P = \frac{nRT}{V} = nRT \frac{1}{[1.3 \text{ L}]e^{(0.01 K^{-1})T}} = \frac{nRT}{1.3} e^{-(0.01 K^{-1})T} \]

In this case, \( dT = 0 \) implies \( dV = 0 \) so the expansion work is zero.

(b)

\[ dU = C_vdT = (3/2)nRTdT \]
\[ \Delta U = \int_{100}^{250} \frac{3}{2} \frac{3}{2} RT \, dT = \frac{9}{2} R (250 \text{ K} - 100 \text{ K}) = 5612.3 \text{ J}. \]

The change in \( U \) would be the same for both paths.
Chapter 7

Chemical Kinetics

Chemical reactions proceed at different rates, depending on the chemical species reacting together and on the conditions of the reaction. The study of chemical kinetics involves the construction of models of chemical reactions and the validation of such models against experiments. A model for a chemical reaction consists of the following parts:

- A mechanism. This is a set of elementary chemical reactions to describe how reactants form intermediates, intermediates combine with one another and reactants, and ultimately products are produced.

- A set of rate equations. These are differential equations corresponding to the reaction mechanism and giving the rates of change of all reactants, intermediates, and products.

- A set of integrated rate equations. These show the concentrations as functions of time for reactants, intermediates, and products. They are obtained by integrating the rate (differential) equations.

The criterion for an acceptable theoretical model is that it agrees with experimental observations of measured time varying concentrations. When a theoretical chemist finds an acceptable model he says he “understands” the reaction.

If we know the reaction mechanism (i.e. the elementary steps of a reaction) we can write down the rate law. Otherwise, we may not have the reaction mechanism but instead only a net reaction; in this case, the rate law must be determined experimentally.
7. Chemical Kinetics

7.1. Order of Reaction

Many chemical reactions you will encounter are net reactions. For example,

$$5\text{Fe}^{2+}(aq) + \text{MnO}_4^{-}(aq) + 8\text{H}_3\text{O}^+(aq) \rightarrow 5\text{Fe}^{3+}(aq) + 12\text{H}_2\text{O}(l) + \text{Mn}^{2+}(aq)$$

is a net reaction because the left hand side appears to imply that 5 ions \(\text{Fe}^{2+}\) collide with 1 ion \(\text{MnO}_4^{-}\) simultaneously to form the stated products. This is not possible. Usually, collisions are binary. Many-body collisions are rarely observed due to their unlikely occurrence\(^1\); at most, 3-body collisions are observed experimentally. Reactions involving only 1 body are possible in the case of a decomposition reaction or radioactive decay.

From knowledge of a net reaction we cannot determine the order of the reaction by inspection of the balanced equation. Instead, experiments must be carried out to determine its order. The rate of the reaction will be proportional to

$$\text{rate of forward reaction} \propto [\text{Fe}^{2+}]^\alpha [\text{MnO}_4^-]^\beta [\text{H}_3\text{O}^+]^\delta,$$

$$\text{rate of backward reaction} \propto [\text{Fe}^{3+}]^\gamma [\text{Mn}^{2+}]^\zeta [\text{H}_2\text{O}]^\xi,$$

where \(\alpha, \beta, \delta, \gamma, \zeta, \xi\) are non-negative constants (not necessarily integers) which are determined from experimental data. The order of the reaction is the sum \(\alpha + \beta + \delta\) or \(\gamma + \zeta + \xi\). Sometimes we speak of the order with respect to a particular species. For example, the reaction is of order \(\alpha\) with respect to \([\text{Fe}^{2+}]\).

The above reaction instead proceeds through a series of elementary steps (path) involving binary collisions (sometimes elementary steps can involve 3-body collisions). Some of these steps include (this list is not complete):

$$\text{Fe}^{2+}(aq) + \text{MnO}_4^{-}(aq) \rightarrow \text{Fe}^{3+}(aq) + \text{MnO}_4^{2-}(aq)$$

$$\text{MnO}_4^{2-}(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{HMnO}_4^{-}(aq) + \text{H}_2\text{O}(l)$$

$$\text{HMnO}_4^{-}(aq) + \text{Fe}^{2+}(aq) \rightarrow \text{HMnO}_4^{2-}(aq) + \text{Fe}^{3+}(aq).$$

A series of elementary steps is called the reaction mechanism. It is generally quite challenging to determine the nature of the reaction mechanism because these involve reaction intermediates that are produced and consumed immediately. In the above reaction mechanism the ions \(\text{MnO}_4^{2-}\) and \(\text{HMnO}_4^{-}\) are reaction intermediates. The short-lived nature of reaction intermediates poses a problem from the standpoint of experimental techniques. Currently, chemists employ a combination of experimental techniques together with results from ab initio calculations to provide insights into reaction mechanisms.

\(^1\)Such reactions are unlikely because they required the simultaneous collision of many different molecules, which is much less likely to occur than binary collisions.
7.2. Reaction Order from Experiments

The order of the reaction can be determined from experiments by plotting reaction rate vs concentration of reactants or products. If a reaction is known to be an elementary step, then its order can be read off from the stoichiometric coefficients by writing down these coefficients as exponentials (similarly to the procedure we use for writing down equilibrium constants).

Some decomposition reactions are first order:

\[ \text{N}_2\text{O}_5(g) \rightarrow 2\text{NO}_2(g) + \frac{1}{2}\text{O}_2(g). \]

It is found from experiments that the rate of the forward reaction is \( k[\text{N}_2\text{O}_5] \), where \( k \) (the rate constant) is determined from the slope of a reaction rate vs concentration plot, so it is first order. The rate constant, \( k \), is a quantity that depends on temperature, as we will see in the next lecture.

In general, a reaction that depends on a single reactant, \( aA \rightarrow \text{products} \), has a rate proportional to the reactant concentration raised to a power \( n \):

\[ \text{rate} = k[A]^n, \]

where \( n \) may be different from \( a \). \( n \) must be determined from experiments. For example, the rate of \( \text{C}_2\text{H}_6(g) \rightarrow 2\text{CH}_3(g) \) is found to be \( k[\text{C}_2\text{H}_6]^2 \) from experiments. This is a second order reaction. Some reactions are independent of reactant concentrations (rate=\( k \)). Fractional order is also possible. An example is the decomposition of CH₃CHO which is found from experiments to be of order 3/2, i.e., \( \text{rate}=k[\text{CH}_3\text{CHO}]^{3/2} \).

**Reaction Order is Derived from Experimental Data:** In the simplest cases (e.g., reaction rate depends on only 1 reactant), it is easy to determine the reaction order from a small number of data points. See, for example, Oxtoby [1], Example 18.1. In the general case, you need to fit the reaction rate expression as function of concentration for each reactant (or product) concentration. The measurement of reaction rates involves measuring reactants as function of time. The “rate” \( \frac{d[A]}{dT} \) is the time derivative of the concentration of \( [A] \) and is obtained from a plot of \([A]\) vs time.

If there is more than 1 chemical species involved, such as \( aA + bB \rightarrow \text{products} \), the rate depends on more than one exponent, namely \( \text{rate} = -(1/a)d[A]/dt = k[A]^m[B]^n \), where \( m,n \) are exponents to be determined experimentally. The overall order of the reaction is \( m + n \).
7.3. The Reaction Rate

Suppose we have a reaction of the type \( \sum_i \nu_i A_i = 0 \), where \( \nu_i \) are the stoichiometric coefficients (negative on the reactant side, and positive on the products side) and \( A_i \) are the chemical species. IUPAC defines the “rate of reaction” as \( \dot{\xi} \equiv \frac{d\xi}{dt} = \frac{1}{\nu_i} \frac{dn_i}{dt} \), where \( n_i \) is the number of moles of species \( A_i \) and \( \xi \) is called the “extent of the reaction”. For example,

\[
N_2 + 3H_2 \rightarrow 2NH_3
\]

The reaction rate is:

\[
\dot{\xi} = \frac{1}{2} \frac{d[NH_3]}{dt} = -\frac{d[N_2]}{dt} = -\frac{1}{3} \frac{d[H_2]}{dt}.
\]

7.3.1. First order reaction: example. A first order reaction is a reaction whose rate law is of the form:

\[
-\frac{d[A]}{dt} = k[A]
\]

where \( k \) is a rate coefficient.

For example, the following reaction

\[
2N_2O_5 \rightarrow 4NO_2 + O_2
\]

has a rate law that is found to be:

\[
-\frac{1}{2} \frac{d[N_2O_5]}{dt} = k[N_2O_5]
\]

The rate law can be integrated if we rewrite it in the form (constants such as ‘2’ can be absorbed in the definition of \( k \)):

\[
\frac{d[A]}{[A]} = -k dt
\]

to get

\[
\int_{[A]_0}^{[A](t)} \frac{d[A]}{[A]} = -k \int_0^t dt \Rightarrow \log[A](t) - \log[A]_0 = -kt
\]

\[
\Rightarrow \log \left( \frac{[A](t)}{[A]_0} \right) = -kt \Rightarrow [A](t) = [A]_0 e^{-kt}
\]

In Fig. 7.1, the concentration as function of time (left) and log concentration vs time (right) are shown. On a semilog plot,

\[
\log[A](t) = \log[A]_0 - kt
\]

the graph of log\([A](t)\) vs \(t\) is a straight line with slope \(-k\) and intercept \(\log[A]_0\).
Thus, an experimental measurement of \([A](t)\) vs time will yield the initial concentration \([A]_0\) by extrapolating to \(t = 0\) and the rate constant \(k\) from the slope. The rate constant is a basic property of every chemical reaction.

The *half-life* is defined by the time \(t\) for which \([A](t)/[A]_0 = \frac{1}{2}\). That is, 
\[
1/2 = e^{-kt}, \text{ or } t_{1/2} = \frac{\log(2)}{k}.
\]

### 7.3.2. Second Order Reaction: Examples.

An example of a second-order reaction is

\[
2\text{ClO}^- \rightarrow 2\text{Cl}^- + \text{O}_2(g)
\]

which is found experimentally to obey the rate law:

\[
-\frac{d[\text{ClO}^-]}{dt} = k[\text{ClO}^-]^2.
\]

This is an equation of the form:

\[
-\frac{d[A]}{dt} = k[A]^2, \text{ or } -\frac{d[A]}{[A]^2} = kd\text{.}
\]

Integrating,

\[
-\int_{[A]_0}^{[A](t)} \frac{d[A]}{[A]^2} = k \int_0^t dt.
\]

which gives:

\[
\frac{1}{[A](t)} = \frac{1}{[A]_0} + kt.
\]

Plotting \(1/[A](t)\) vs time gives a straight line with slope \(k\) and intercept \(1/[A]_0\), as shown in Fig. 7.2.

Another type of second-order reaction is a reaction

\[
aA + bB \rightarrow \text{products}.
\]
For second order reactions, the inverse of $[A](t)$ decays linearly with time from some initial concentration $[A]_0$. The rate is of the form

$$\text{rate} = k[A][B]$$

i.e.

$$-\frac{1}{a} \frac{d[A]}{dt} = -\frac{1}{b} \frac{d[B]}{dt} = k[A][B].$$

For example,

$$S_2O_8^{2-} + 2I^- \rightarrow 2SO_4^{2-} + I_2$$

whose rate law is found experimentally to be:

$$\frac{d[S_2O_8^{2-}]}{dt} = -k[S_2O_8^{2-}][I^-]$$

### 7.4. Measurement Methods

In order to study reaction kinetics we need analytical techniques that will report on concentrations of different chemical species as function of time. In general, we may not know what the stoichiometry of the overall reaction is (which is the sum of all the steps in the reaction mechanism). When the stoichiometry is known, we can write down the equilibrium constant. Sometimes, we don’t know what all the products are.

Chemical analysis can be done using gravimetric or volumetric techniques, which can be made specific to one or more reactants or products. Chemical analysis is generally slow and can perturb the reaction itself, as it involves withdrawing a physical sample from the system. Separation of chemicals can be done via chromatography, gel electrophoresis, or capillary electrophoresis, followed by analysis of the separated species.
Physical methods, such as pressure or volume measurements (without sample extraction and analysis), can be much more rapid than chemical methods. Suppose that a reaction is carried out in a vessel of fixed volume:

\[ 2\text{NO}_2 \rightarrow 2\text{NO} + \text{O}_2. \]

The reaction (in the forward direction) produces two moles of product for every mole of reactant, leading to an increase in the pressure inside the vessel. If the gas mixture is ideal, we may invoke the ideal gas law to relate the concentration of chemical species to the pressure inside the vessel:

\[
\frac{d[\text{O}_2]}{dt} = \frac{1}{2} \frac{d[\text{NO}]}{dt} = -\frac{1}{2} \frac{d[\text{NO}_2]}{dt} = \frac{1}{V} \frac{dn_{\text{total}}}{dt} = \frac{1}{RT} \frac{dP}{dt}.
\]

The last equality follows from the ideal gas law. The previous equality follows from:

\[ n_{\text{total}} = n_{\text{O}_2} + n_{\text{NO}} + n_{\text{NO}_2}, \]

stoichiometry

\[ \Delta n_{\text{NO}_2} = -\frac{1}{2} \Delta n_{\text{O}_2}, \quad \Delta n_{\text{NO}} = \frac{1}{2} \Delta n_{\text{O}_2}, \]

and

\[
\frac{dn_{\text{total}}}{dt} = \frac{dn_{\text{O}_2}}{dt} + \frac{dn_{\text{NO}}}{dt} + \frac{dn_{\text{NO}_2}}{dt} = \frac{dn_{\text{O}_2}}{dt} (1 + \frac{1}{2} - \frac{1}{2}) = \frac{dn_{\text{O}_2}}{dt}.
\]

Pressure measurements (e.g., with strain gauges) take on the order of a millisecond, so while the method is faster than chemical methods, it is slower than optical methods. We can also work at constant pressure by allowed the gas in a vessel expand through a piston and monitor volume changes. The total volume is the sum of the partial volumes of the species present.

Electrochemical methods can also be used to monitor the course of a reaction. For example, the electrical conductivity of the following solution will vary as the reaction proceeds:

\[ \text{C}_2\text{H}_5\text{Cl} + 2\text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH} + \text{H}_3\text{O}^+ + \text{Cl}^- . \]

Finally, spectroscopic methods are commonly used to monitor reactions: electron spin resonance (ESR), nuclear magnetic resonance (NMR), mass spectrometry, and optical spectroscopy (ultraviolet, IR and visible). Other methods such as light scattering, optical rotation, flash photolysis, pulse radiolysis and relaxation have been used. Spectroscopic methods are chemically selective and report on the various chemical species. ESR, NMR and optical methods rely on the absorption of radiation (microwave, radiofrequency, optical photons) at different wavelengths to probe various chemical groups. Optical methods are generally the fastest ones (down to picoseconds and femtoseconds).
7.4.1. Steady State, Elementary Reactions of the Type \( aA + bB \rightleftharpoons cC + dD \). In general, “steady state” and “equilibrium” are not equivalent. The term steady state means that time derivatives of observable quantities are zero. Steady state often implies equilibrium. However, some steady states are not equilibrium states. For example, a fluid flowing at constant velocity is by no means an equilibrium state. However, it is a steady state because the velocity, an observable quantity, is constant. Equilibrium refers to thermal equilibrium, which is the point where the differential of \( G \) is zero.

Let’s look at an example of a steady state. Suppose we have the elementary reaction

\[
\begin{align*}
    aA + bB & \rightleftharpoons cC + dD \\
\end{align*}
\]

Its rate is the sum of two terms. The first term describes the rate of depletion of \( A \) while the second term describes the rate of production of \( A \):

\[
-\dot{\xi} = \frac{1}{a} \frac{d[A]}{dt} + \frac{1}{b} \frac{d[B]}{dt} = -\frac{1}{c} \frac{d[C]}{dt} - \frac{1}{d} \frac{d[D]}{dt} = -k_f [A]^a [B]^b + k_r [C]^c [D]^d.
\]

At steady-state (\( \dot{\xi}=0 \)), the rates of the forward and backward reactions are equal:

\[
    k_f [A]_{eq}^a [B]_{eq}^b = k_r [C]_{eq}^c [D]_{eq}^d.
\]

This is also known as the principle of detailed balance. And since \( K \equiv k_f / k_r \),

\[
    K = \frac{k_f}{k_r} = \frac{[C]_{eq}^c [D]_{eq}^d}{[A]_{eq}^a [B]_{eq}^b}.
\]

An example of such an elementary reaction is:

\[
    \text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}.
\]

Its rate is:

\[
-\dot{\xi} = \frac{d[I_2]}{dt} = -k_f [\text{H}_2][\text{I}_2] + k_r [\text{HI}]^2.
\]

At equilibrium, \( d/dt=0 \), the steady-state gives the equilibrium constant:

\[
    K = \frac{k_f}{k_r} = \left( \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \right)_{eq}.
\]

7.4.2. Reaction Order, From Elementary Reactions. If the elementary steps are known, then the reaction order is known. For example, the elementary step

\[
    k \quad \text{N}_2\text{O}_5^*(g) \rightarrow \text{NO}_2(g) + \text{NO}_3(g)
\]
is first order (asterisk denotes an excited state) and its rate is $k[N_2O_5^*]_1$. Likewise, the elementary step
\[ NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g) \]
yields a second order reaction whose rate is $k[NO]_1[O_3]_1$. It is second order because the exponents add up to $1 + 1 = 2$.

### 7.4.3. Net Reactions vs Reaction Mechanism: Example 1.
As a reminder, the reactions we usually write are net reaction, e.g.,
\[ NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g), \quad (*) \]
whose order is unknown without data from experiments. However, if the elementary steps are known, we can determine the order. In this particular case, the reaction mechanism is known. It involves the following steps:
\[ NO_2(g) + NO_2(g) \rightarrow NO_3(g) + NO(g) \quad \text{(slow)} \]
\[ NO_3(g) + CO(g) \rightarrow NO_2(g) + CO_2(g) \quad \text{(fast)} \]
You can check that adding these two reactions yields the net reaction (*). NO$_3$ is a reaction intermediate because it is produced then consumed in the subsequent step and does not show up in the net reaction.

### 7.4.4. Net Reactions vs Reaction Mechanism: Example 2.
Suppose we have an overall reaction:
\[ 2NO(g) + 2H_2(g) \rightleftharpoons N_2(g) + 2H_2O(g) \]
whose elementary steps describe a series of binary collisions shown below. You can check that adding these 3 equations together yields the overall reaction (we indicate which terms cancel out - those are the reaction intermediates).
\[ NO(g) + NO(g) \overset{k_1}{\rightleftharpoons} N_2O_2(g) \]
\[ N_2O_2(g) + H_2(g) \overset{k_2}{\rightleftharpoons} N_2O(g) + H_2O(g) \]
\[ N_2O(g) + H_2(g) \overset{k_3}{\rightleftharpoons} N_2(g) + H_2O(g) \]
At equilibrium, we take the rate laws and set $\dot{\xi} = 0$. This results in detailed balance:
\[ k_1[NO]_\text{eq}^2 = k_{-1}[N_2O_2]_\text{eq} \]
\[ k_2[N_2O_2]_\text{eq}[H_2]_\text{eq} = k_{-2}[N_2O]_\text{eq}[H_2O]_\text{eq} \]
\[ k_3[N_2O]_\text{eq}[H_2]_\text{eq} = k_{-3}[N_2]_\text{eq}[H_2O]_\text{eq} \]
We also know that the equilibrium constants are equal to the ratio of the forward and reverse reaction rate constants:

\[ K_1 = \frac{[N_2O_2]_{eq}^2}{[NO]_{eq}^2} = \frac{k_1}{k_1}, \quad K_2 = \frac{[N_2O]_{eq}[H_2O]_{eq}}{[N_2O_2]_{eq}[H_2]_{eq}} = \frac{k_2}{k_2}, \]

\[ K_3 = \frac{[N_2]_{eq}[H_2O]_{eq}}{[N_2O]_{eq}[H_2]_{eq}} = \frac{k_3}{k_3}. \]

We also know that adding reactions together leads to the multiplication of its equilibrium constants. Let’s verify that this gives the correct equilibrium constant for the overall reaction:

\[ K = K_1K_2K_3 = \frac{k_1k_2k_3}{k_{-1}k_{-2}k_{-3}} = \frac{[N_2O_2]_{eq}^2[N_2O]_{eq}[H_2O]_{eq}[N_2]_{eq}[H_2O]_{eq}}{[NO]_{eq}^2[N_2O_2]_{eq}[H_2]_{eq}[N_2O]_{eq}[H_2]_{eq}} = \frac{[H_2O]_{eq}^2[N_2]_{eq}}{[NO]_{eq}^2[H_2]_{eq}} \]

as it should.

This suggests that the rate of the forward net reaction is the product (or at least, proportional to the product) of forward rates of the elementary reactions, \( k_1k_2k_3 \), whereas the rate of the backward net reaction is the product \( k_{-1}k_{-2}k_{-3} \).

### 7.4.5. Elementary Reactions: Single Reactions.

When writing down rates of reactions for a chemical species, we first need to identify which reactions this species is involved in, and add a term describing the production or depletion of this species according to the rate constant. In other words, every arrow that points towards or away from the species will correspond a term in the rate equation. This is best explained by looking at many examples.

The rate of the reaction

\[ A \overset{k}{\rightarrow} B, \]

which describes the unidirectional depletion of \( A \) is:

\[ \dot{\xi} = \frac{d[A]}{dt} = -k[A]. \]

There is only one term because the reaction is unidirectional. There is only 1 arrow that points away from \( A \) and no arrows that point toward \( A \). Thus, only 1 term is needed. This term, \(-k[A]\), describes the depletion of \( A \) (conversion of \( A \) to \( B \)). It is proportional to the rate constant \( k \), and to \([A]\) (the concentration of \( A \)), because this elementary reaction describes a “1-body collision” (i.e. this could describe the decay of a molecule from an excited state to a ground state, or the folding of a protein). Hence, the negative sign.
If the reaction were two-sided,

\[ A \xrightleftharpoons[k_1]{k_2} B, \]

we would need 2 terms in the rate equation. One term describing the depletion of \( A \), with a negative sign and rate constant \( k_1 \), and one term describing the production of \( A \), with a positive sign and rate constant \( k_2 \):

\[ -\dot{\xi} = \frac{d[A]}{dt} = -k_1[A] + k_2[B]. \]

The first term is proportional to the concentration of \( A \) because this elementary reaction describes a “one-body collision” involving only 1 molecule of \( A \). Also, if there were no amount of \( A \), we wouldn’t be producing any \( B \); similarly, the second term is proportional to \( B \), as we need a nonzero concentration of \( B \) in order to produce \( A \) in the reverse reaction.

A unidirectional reaction

\[ A + B \rightarrow^k C + D, \]

only describes the depletion of \( A \). In order to deplete \( A \), we need a “two-body collision” involving 1 molecule of \( A \) colliding with 1 molecule of \( B \). Thus, the rate should be proportional to the product \([A][B]\), since the concentrations describe the probabilities of finding a molecule of \( A \) (or \( B \)) within a given volume element, and the product of \([A]\) and \([B]\) describes the joint probability of finding 1 molecule of \( A \) and 1 molecule of \( B \) in the same volume.

The rate contains only 1 term. This term has a negative sign, because it describes the depletion of \( A \):

\[ \frac{d[A]}{dt} = -k[A][B]. \]

A bidirectional reaction

\[ A + B \xrightleftharpoons[k_f]{k_r} C + D, \]

now describes both the depletion of \( A \) (in the forward reaction) and the production of \( A \) (in the reverse reaction). The first event proceeds at a rate \( k_f[A][B] \) while the second event happens at the rate \( k_r[C][D] \). The rate law is:

\[ \frac{d[A]}{dt} = -k_f[A][B] + k_r[C][D]. \]

### 7.4.6. Elementary Reactions: Simultaneous Reactions

Suppose we have two simultaneous reactions:

\[ A \rightarrow^k_1 B \quad \text{and} \quad A \rightarrow^k_2 C, \]
Both reactions describe the depletion of $A$. The rate of the reaction is:

$$-\dot{\xi} = \frac{d[A]}{dt} = -k_1[A] - k_2[A].$$

### 7.4.7. Elementary Reactions: Reversible Reactions.

A reversible reaction

$$A + B \rightleftharpoons C + D,$$

can be viewed as two simultaneous unidirectional forward reactions:

$$A + B \rightarrow C + D,$$

$$C + D \rightarrow A + B.$$

We can write down the rate for each reaction. The first reaction describes the depletion of $A$ at a rate $k_f[A][B]$ while the second reaction describes the production of $A$ at a rate $k_r[C][D]$. Summing up these two rates gives:

$$\frac{d[A]}{dt} = -k_f[A][B] + k_r[C][D].$$

### 7.4.7.1. Elementary Reactions: Consecutive Reactions.

Consider the consecutive elementary reactions:

$$\begin{align*}
A & \rightarrow B \\
A & \rightarrow C,
\end{align*}$$

A rate analysis yields:

$$\begin{align*}
\frac{d[A]}{dt} & = -k_1[A], \\
\frac{d[B]}{dt} & = k_1[A] - k_2[B] \quad \text{and} \quad \frac{d[C]}{dt} = k_2[B].
\end{align*}$$

Consider the following variant:

$$\begin{align*}
A & \rightarrow B \\
A & \rightleftharpoons C,
\end{align*}$$

The rate laws would be:

$$\begin{align*}
\frac{d[A]}{dt} & = -k_1[A], \\
\frac{d[B]}{dt} & = k_1[A] - k_2[B] + k_{-2}[C] \quad \text{and} \quad \frac{d[C]}{dt} = k_2[B] - k_{-2}[C].
\end{align*}$$

Another variant of the above reaction could be:

$$\begin{align*}
A & \rightleftharpoons B \\
A & \rightarrow C,
\end{align*}$$
The rate laws would be:
\[
\frac{d[A]}{dt} = -k_1[A] + k_{-1}[B],
\]
\[
\frac{d[B]}{dt} = k_1[A] - k_2[B] - k_{-1}[B] + k_{-2}[C],
\]
\[
\frac{d[C]}{dt} = k_2[B] - k_{-2}[C].
\]

7.4.8. Series of Unidirectional Steps. We have to be slightly careful when the same chemical species appears on both sides of the same reaction. Suppose we have the following reaction mechanism:

\[
\begin{align*}
    A + X &\rightarrow 2X & \text{(step 1)} \\
    X + Y &\rightarrow 2Y & \text{(step 2)} \\
    Y &\rightarrow Z & \text{(step 3)}
\end{align*}
\]

The rate laws are:
\[
\frac{d[A]}{dt} = -k_1[A][X],
\]
\[
\frac{d[X]}{dt} = k_1[A][X] - k_2[X][Y],
\]
\[
\frac{d[Y]}{dt} = k_2[X][Y] - k_3[Y],
\]
\[
\frac{d[Z]}{dt} = k_3[Y].
\]

In the rate equation for \([X]\), there are two terms. The first term follows from step 1 (forward direction) and describes the production\(^2\) of \(X\) from the reactants of step 1. The production of \(X\) enters with a positive sign. The second term describes the depletion of \(X\) as a result of step 2 (forward direction). The depletion of \(X\) enters with a negative sign. Notice that there is no term of the form \(-k_{-1}[X]\)\(^2\) because step 1 is unidirectional.

\(^2\)This mechanism normally assumes that the concentration of reactant \(A\) is much larger than its equilibrium value. Another assumption is that the forward rates are much larger than the reverse rates. Thus, the reverse rates are negligible.

\(^3\)It should be clear that step 1, \(A + X \rightarrow 2X\), describes the production of \(X\). This reaction describes the transformation of \(A\) into \(X\), so \(X\) is indeed being produced. Another way to think about it is to manipulate the chemical reaction as an equation: subtracting \(2X\) on both sides gives, \(A - X \rightarrow 0\), whose negative sign on the reactants side reminds us that \(A\) isn’t being consumed in the forward direction, but produced instead.
In the rate equation for \([Y]\), there are two terms, \(k_2[X][Y]\) and \(-k_3[Y]\). These follow by inspection of steps 2 and 3: step 2 describes the production of \(Y\), which depends on the forward reaction. This leads to the term \(k_2[X][Y]\). Note that there is no term of the form \(-k_{-2}[Y]^2\) because the reaction is in the forward direction only. The second term \(-k_3[Y]\) describes the depletion of \(Y\) as a result of its transformation into \(Y\) and follows from step 3. Notice that there is no term of the form \(+k_{-3}[Z]\) because step 3 is unidirectional (forward direction only).

The lesson here is to be careful when you see a species appear on both sides of a reaction. Make sure you get the sign right, depending on whether the species is depleted or produced. Another example is:

\[
k A + 2X \rightarrow 3X
\]

Again, this reaction describes the production of \(X\) in a 3-body collision. This sign can be seen by subtracting \(3X\) on both sides (so that \(X\) shows up only on 1 side of the equation, but not on the other);\(^4\) the \(X\) appears with a negative sign on the reactants side, implying that \(X\) is produced not depleted. The reactants side, \(A + 2X\), shows that this is a 3-body collision involving 1 molecules of \(A\) and 2 molecules of \(X\); thus, its rate is proportional to \(+k[A][X]^2\), not \(-k[A][X]^2\).

7.5. Tricks for Writing Down Rate Laws

Let’s look at some techniques that help us write down rate laws when given a reaction mechanism. Examples are given below. This is far from a complete survey of all techniques. In general, we must look at the reaction mechanism (if it is available to us) and analyze it on a case-by-case basis.

7.5.1. Rate-Determining Steps. In many reactions, there is often one step that is significantly slower than all other steps. In this case, the rate of the reaction can be taken to be equal to the rate of the slowest step. Suppose that an overall reaction is:

\[
2\text{NO}_2(g) + F_2(g) \rightarrow 2\text{NO}_2F(g)
\]

and suppose that the reaction mechanism is:

\[
k_1 \text{NO}_2(g) + F_2(g) \rightarrow \text{NO}_2F(g) + F(g) \quad \text{(slow)}
\]

\[
k_2 \text{NO}_2(g) + F(g) \rightarrow \text{NO}_2F(g) \quad \text{(fast)}
\]

\(^4\)The purpose of this is simply to see if \(X\) is produced or depleted. Alternatively, you can also subtract \(2X\) on both sides to get \(A \rightarrow X\), which shows that \(X\) is produced. This does not imply that the rate is \(k[A]\) instead of \(k[A][X]^2\).
The first step is the slow step (suppose that we know this from experiments). It is called the rate-limiting step. The rate of the overall reaction is determined by this slow step, $k_1[\text{NO}_2][\text{F}_2]$. Note: the rate is not obtained from the net reaction as $\propto [\text{NO}_2]^2[\text{F}_2]$ because the net reaction does not allow us to determine the rate or order.

### 7.5.2. Fast Equilibrium.

In the above section, the elementary steps were unidirectional. Suppose we have bidirectional steps. If one of the reactions happens to be in “fast equilibrium”, we can use this information to solve for concentrations. For example, the overall reaction

$$2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$$

proceeds according to the mechanism:

$$\begin{align*}
\text{NO}(g) + \text{NO}(g) & \overset{k_1}{\rightleftharpoons} \text{N}_2\text{O}_2(g), \quad \text{(fast equilibrium)} \\
\text{N}_2\text{O}_2(g) + \text{O}_2(g) & \overset{k_2}{\rightleftharpoons} 2\text{NO}_2(g). \quad \text{(slow)}
\end{align*}$$

Since the slow step determines the overall rate we have:

$$\text{rate} = k_2[\text{N}_2\text{O}_2][\text{O}_2].$$

However, $\text{N}_2\text{O}_2$ doesn’t appear in the net reaction. We can express this intermediate’s concentration in terms of NO concentration using the condition for fast equilibrium. The fast equilibrium condition is:

$$K_1 = \frac{k_1}{k_{-1}} = \frac{[\text{N}_2\text{O}_2]}{[\text{NO}]^2}, \quad [\text{N}_2\text{O}_2] = K_1[\text{NO}]^2, \quad \text{rate} = k_2K_1[\text{NO}]^2[\text{O}_2].$$

We note that the order here is the same as that of the net reaction. This is a coincidence and may not be the case in general.

### 7.5.3. Steady-State Approximation.

If there is no step that is much slower than the others, another trick we can use is the steady-state approximation. This requires a reaction intermediate and we assume that this intermediate’s concentration remains constant through the reaction. This assumption is reasonable because intermediates are rapidly produced and consumed. Therefore, they rapidly reach a steady-state concentration, which enables us to assume their concentration is constant.

Suppose that the net reaction

$$\text{N}_2\text{O}_5 \rightarrow 2\text{NO}_2 + \frac{1}{2}\text{O}_2$$

...
Chemical Kinetics proceeds according to the mechanism:

\[ \begin{align*}
N_2O_5(g) + M(g) & \xrightleftharpoons[k_{-1}][k_1] N_2O_5^*(g) + M(g) \quad \text{(step 1)} \\
N_2O_5^*(g) & \xrightarrow[k_2]{\text{fast}} NO_3(g) + NO_2(g) \quad \text{(step 2)} \\
NO_3(g) + NO_2(g) & \xrightarrow[k_3]{\text{fast}} NO(g) + NO_2(g) + O_2(g) \quad \text{(step 3 - fast)} \\
NO_3(g) + NO(g) & \xrightarrow[k_4]{\text{fast}} 2NO_2(g) \quad \text{(step 4 - fast)}
\end{align*} \]

A reaction intermediate is \( N_2O_5^* \). Shortly after the start of the reaction, the concentration of the intermediate reaches a steady state. (Since they don’t appear in the net reaction, intermediates can’t be easily detected, mainly because they are short-lived, as they are rapidly consumed after being produced.) This means we can set its time derivative equal to zero:

\[
\frac{d[N_2O_5^*]}{dt} = k_1 [N_2O_5][M] - k_{-1} [N_2O_5^*][M] - k_2 [N_2O_5^*] = 0.
\]

This gives an algebraic equation which enables us to solve for the steady state concentration:

\[
[N_2O_5^*] = \frac{k_1 [N_2O_5][M]}{k_2 + k_{-1}[M]}.
\]

Using the rate law for step 2, we can write the rate of the overall reaction as:

\[
\text{rate} = \frac{1}{2} \frac{d[NO_2]}{dt} = k_2 [N_2O_5^*] = \frac{k_1 k_2 [N_2O_5][M]}{k_2 + k_{-1}[M]}.
\]

We note that the factor of \(1/2\) here has been absorbed in the definition of \(k_2\) (this is how the textbook of Oxtoby [1] does it).

7.5.4. Steady State, Elementary Reactions of the Type \( A \rightleftharpoons B \).

Consider the simple reaction,

\[ A \xrightleftharpoons[k_r][k_f] B, \]

where \(k_f\) is the rate constant of the forward reaction and \(k_r\) is the rate constant of the backward reaction.

Suppose that we start with only species \( A \) present at time \( t = 0 \). Thus, there is a relationship between the concentrations of \( A \) and \( B \) at all times:

\[ [B](t) + [A](t) = [A]_0. \]

If this is an elementary reaction, the rate of the reaction is the sum of two terms. The first term describes the rate of depletion of \( A \) while the second
term describes the rate of production of $A$:
\[
-\dot{\xi} = \frac{d[A]}{dt} = -k_f[A] + k_r[B] = -k_f[A] + k_r([A]_0 - [A]).
\]
At equilibrium, the time derivative is zero, as the reaction has reached a steady state,
\[
0 = k_f[A]_{eq} - k_r ([A]_0 - [A]_{eq}).
\]
This is the same as saying that at steady-state ($d/dt=0$), the rates of the forward and backward reactions are equal:
\[
k_f[A]_{eq} = k_r ([A]_0 - [A]_{eq}).
\]
We have a relationship between equilibrium concentration, initial concentration and the rate constants:
\[
[A]_{eq} = \frac{k_r[A]_0}{k_f + k_r}.
\]

7.6. Reaction Rates From Kinetic Theory

Suppose we have an elementary reaction
\[
A + B \rightarrow P,
\]
whose rate is
\[
rate = k[A][B],
\]
so far we have not said anything about the value of $k$ and how it depends on the parameters of the molecules involved in the reaction. Simple arguments from kinetic theory can be invoked to explain the behavior of $k$.

$k$ should depend on the frequency of molecular collisions, and therefore the mean speed of the molecules, $\propto \sqrt{T/M}$, where $T$ is temperature and $M$ is molar mass of the molecules. We also expect it to be proportional to the collision cross-section, $\sigma$ and to the number densities of $A$ and $B$ (and therefore, their concentrations):

\[
rate \propto \sigma \sqrt{\frac{T}{M}} [A][B].
\]
However, not all collisions will lead to reaction. Only those whose kinetic energy exceeds a minimum value $E_a$ (activation energy) will succeed. Thus, the rate should be proportional to a Boltzmann factor of the form $e^{-E_a/RT}$:

\[
rate \propto \sigma \sqrt{\frac{T}{M}} e^{-E_a/RT} [A][B]; \quad k \propto \sigma \sqrt{\frac{T}{M}} e^{-E_a/RT}.
\]
Not every collision will lead to reaction even if the energy requirement is satisfied because the reactants may need to collide in certain relative orientations, as illustrated in Fig. 7.3. This “steric requirement” can be accounted for using another factor $P$,

$$k \propto P\sigma \sqrt{\frac{T}{M} e^{-E_a/RT}}.$$  
(Note: $M$ is a type of effective or “reduced mass”, i.e. $m_1 m_2/(m_1 + m_2)$, when molecular collisions between pairs of molecules are involved.)

7.7. Arrhenius Law and Activation Energy Barriers

The activation energy, $E_a$, may be different if you go in the forward or reverse direction. Thus when stating $k$ we must also state whether this refers to the forward or reverse direction.

The expression bears close resemblance to the Arrhenius law, which is usually expressed in the form:

$$k(T) = A(T) e^{-E_a/RT},$$

where $A(T)$ is the pre-exponential factor and $E_a$ is the activation energy. The activation energy is the height of the barrier required to reach the “transition state” (‡ symbol in Fig. 7.4). Of course, the value of $E_a$ depends on whether we are referring to the forward or reverse (backward) reaction. We denote these two values as $E_{a,f}$ and $E_{a,r}$. These two values can be very different, as seen in Fig. 7.4.

The standard Gibbs free energy change of a reaction is related to the equilibrium constant $K$ through

$$K = e^{-\Delta G_{\text{rxn}}^{0}/RT},$$

The rate of the forward reaction, $k_f$, is

$$k_f = A(T) e^{-E_{a,f}/RT},$$

where $E_{a,f,\text{act}}$ is the difference between Gibbs free energy of the activated complex and that of the reactants (i.e. the barrier height of the forward
In the previous section we have seen that the equilibrium constant of an uncatalyzed reaction is:

\[ K = \frac{k_f}{k_r} = e^{-(E_{a,f} - E_{a,r})/RT}, \]

In the above diagram, we see that \( \Delta G_{rxn}^0 < 0 \), so the reaction is spontaneous in the forward direction. Its speed in the forward direction, however, is determined by \( E_{a,f} \), not \( \Delta G_{rxn}^0 \).

These expressions for \( k_f \) and \( k_r \) are called Arrhenius laws. They give the temperature dependence of the reaction rates. It also states that the rates depend on the activation energy. This activation energy is the energy required to reach the transition state (top of the energy barrier), either from the reactant side or the product side.

The standard free energy change of the reaction is the energy difference between reactants and products. Thus, we see how the free energy does not tell the whole story: it says nothing about the energy barrier, hence, it is not expected to predict the reaction rates. The reaction rates are predicted instead by the activation energy barrier.

**7.8. Effect of Catalyst**

In the previous section we have seen that the equilibrium constant of an uncatalyzed reaction is:

\[ K = \frac{k_f}{k_r} = e^{-(E_{a,f} - E_{a,r})/RT}. \]
Suppose that we add a catalyst. The effect of a catalyst is to lower the “activation energy barrier” by some amount $C$, as illustrated in the Fig. 7.5.

\[ k'_f = A(T)e^{-(E_{a,f} - C)/RT}, \quad k'_r = A(T)e^{-(E_{a,r} - C)/RT}, \]

and

\[ K' = \frac{k'_f}{k'_r} = e^{-(E_{a,f} - C - E_{a,r} + C)/RT} = e^{-(E_{a,f} - E_{a,r})/RT} = e^{-\Delta G^{\circ}_{\text{rxn}}/RT} = K, \]

which is the same equilibrium constant as the uncatalyzed reaction ($K = K'$). The equilibrium constant depends only on the difference between the energies of the reactants and the products. Thus, while the introduction of a catalyst speeds up both forward and reverse reactions by the same amount, it does not change the equilibrium.

### 7.9. Origin of the Boltzmann Factor, $e^{-\epsilon/RT}$

The Boltzmann distribution gives the statistical distributions of the energy levels of atoms/molecules in a gas at thermal equilibrium at some temperature $T$. It is a purely “classical” distribution, and does not describe quantum (boson, fermion) statistics. Our derivation is based on the First Law of thermodynamics and the Boltzmann definition of entropy.
Suppose that a system and a reservoir, are together isolated from their surroundings, i.e. think of the system and reservoir together as the universe. The system is in contact with the reservoir as shown in Fig. 7.6.

The system can be anything we want, but here we shall take the system to be a single atom or molecule from the gas. The energy of the reservoir is denoted $U_{\text{ext}}$ and the energy of the system, $U$.

Consider two states (e.g. eigenstates) of our atom/molecule: state $s_1$ and state $s_2$. They are assigned energies $U(s_1)$ and $U(s_2)$, respectively. To these, correspond probabilities $P(s_1)$ and $P(s_2)$.

For an isolated system (e.g., here, the system plus reservoir), all accessible microstates are equally probable. This assumption is reasonable because in thermodynamics we do not have any knowledge about the microscopic state of the system: any state that is degenerate is (from our point of view) no different than any other eigenstate of the same energy. The system can freely hop between eigenstates of the same energy, since there is no energy cost to doing so.

Their probabilities are then proportional to the degeneracy (multiplicity) of the reservoir: $\Omega_{\text{ext}}(s_1)$ and $\Omega_{\text{ext}}(s_2)$. $\Omega_{\text{ext}}(s_i)$ is the number of microstates of the reservoir that are consistent with the system being in state $s_i$, and the total energy (system + reservoir) fixed. We need not be concerned about the actual value of $\Omega_{\text{ext}}$ because we will invoke the First Law to relate it to the macroscopic properties of the reservoir temperature ($T$) and entropy ($S_{\text{ext}}$).\footnote{This is the usefulness of the Boltzmann definition of entropy, which relates the microscopic world ($\Omega$) to the macroscopic world ($S$).}

### 7.9.1. Ratio of Probabilities

In this section we will show that the ratio of probabilities $P(s_2)/P(s_1)$ is a function of the energy difference $\Delta = U(s_2) - U(s_1)$ and identify the precise functional form for this dependence on $\Delta$. We shall assume no particle exchange between the system and the reservoir.
reservoir. We will also assume that the volume of the reservoir does not change significantly when the atom/molecule goes into an excited state.

Invoking the Boltzmann definition of entropy $S = k_B \log \Omega$,

$$
P(s_2) = \frac{\Omega_{ext}(s_2)}{\Omega_{ext}(s_1)} = e^{S_{ext}(s_2)/k_B} = e^{S_{ext}(s_1)/k_B} = e^{[S_{ext}(s_2) - S_{ext}(s_1)]/k_B}
$$

In the exponent, we have the change in entropy of the reservoir as the system goes from state 1 to state 2. This is an infinitesimal change from the reservoir’s point of view. Thus we invoke the First Law and apply it to the reservoir (subscript $ext$):

$$
S_{ext}(s_2) - S_{ext}(s_1) \approx dS_{ext} = \frac{1}{T} [dU_{ext} + PdV_{ext} - \mu dN_{ext}] \approx \frac{1}{T} [U_{ext}(s_2) - U_{ext}(s_1)].
$$

First Law

with $dV_{ext} = 0$ (the $P - V$ term is negligible compared to the $dU$ term, when the atom goes into an excited state) and $dN_{ext} = 0$ (no particle exchange).

Then, since the energy lost by the reservoir is the energy gained by the system ($\Delta U_{ext} = - \Delta U$):

$$
S_{ext}(s_2) - S_{ext}(s_1) = \frac{1}{T} [U_{ext}(s_2) - U_{ext}(s_1)] = -\frac{1}{T} [U(s_2) - U(s_1)],
$$

where $U$ is the energy of the system (atom/molecule). Then,

$$
\frac{P(s_2)}{P(s_1)} = e^{\left[-U(s_2) - U(s_1)/k_BT\right]},
$$

Thus, we have derived the famous result $\frac{P(s_2)}{P(s_1)} = e^{-\Delta E/k_BT}$, where $\Delta E$ is the energy separation between states 1 and 2.

### 7.9.2. Canonical Distribution, No Particles Exchanged

In this section we will derive an exact expression for $P(s_2)$ [not the ratio $P(s_2)/P(s_1)$] including the proper normalization factor.

Let us rewrite the previous equation as:

$$
\frac{P(s_2)}{P(s_1)} = e^{-[U(s_2) - U(s_1)]/k_BT} = \frac{e^{-U(s_2)/k_BT}}{e^{-U(s_1)/k_BT}}.
$$

This can be rearranged as follows: $\frac{P(s_2)}{e^{-U(s_2)/k_BT}} = \frac{P(s_1)}{e^{-U(s_1)/k_BT}}$, where the left hand side depends on $s_2$ and the right hand side depends on $s_1$. The only way this can happen is if both sides are equal to a constant. With proper normalization, we obtain the Boltzmann factor:

$$
P(s) = \frac{1}{Z} e^{-U(s)/k_BT}.
7.9. Origin of the Boltzmann Factor, \( e^{-\epsilon/RT} \)

where \( Z = \sum_{s'} e^{-U(s')/k_B T} \) and \( \sum_{s'} \) is a sum over all states. This Boltzmann distribution is also called the canonical distribution.

7.9.2.1. Sum Over Energies. Because many states may correspond to the same energy, we can also express this in terms of a sum over energies, rather than a sum over states. Let \( g(\epsilon) \) denote the degeneracy of energy level \( \epsilon \) and \( P(\epsilon) \) (degeneracy = number of quantum states corresponding to the same energy \( \epsilon \)) the probability of the system being found in energy \( \epsilon \). Then,

\[
P(\epsilon) = \frac{1}{Z} g(\epsilon) e^{-\epsilon/k_B T},
\]

where \( Z = \sum_{\epsilon'} g(\epsilon') e^{-\epsilon'/k_B T} \) (summation over energy levels, not over quantum states).

7.9.3. Degeneracy. In case you forgot what quantum state degeneracy means, you can refresh your memory at:

https://en.wikipedia.org/wiki/Degenerate_energy_levels

You may have seen the case of the “particle in a box” or the “hydrogen atom” in Chem 20A. Both are famous examples illustrating quantum state degeneracy. For example, hydrogen atom energy levels only depend on \( n \). The degeneracy of level \( n \) is determined by the quantum number \( l \) (angular momentum), \( m_l \) (z-projection of angular momentum) and \( m_s \) (spin). \( l \) is an integer ranging from 0 to \( n-1 \) (so \( n \) different values of \( l \)). Similarly, for given values of \( n \) and \( l \), \( m_l \) is an integer ranging from \(-l\) to \(+l\) (so \( 2l+1 \) different values of \( m_l \)). Thus, the degeneracy of level \( n \) is \( \sum_{l=0}^{n-1} (2l+1) = n^2 \). If we add spin \( (m_s = -1/2, 1/2) \), the degeneracy is multiplied by 2, i.e. \( g(n) = 2n^2 \).

7.9.4. Particles Allowed to Exchange: the Grand Canonical distribution. Now let’s repeat the steps from the previous sections but allow for exchange of particles between the system and the reservoir. Suppose now that particles are allowed to be exchanged between the system and reservoir. Our starting point is the same as before:

\[
\frac{P(s_2)}{P(s_1)} = \frac{\Omega_{ext}(s_2)}{\Omega_{ext}(s_1)} = e^{S_{ext}(s_2)/k_B} e^{-S_{ext}(s_1)/k_B} = e^{[S_{ext}(s_2)-S_{ext}(s_1)]/k_B}
\]

In the exponent, we have the change in entropy of the reservoir as the system goes from state 1 to state 2. This is an infinitesimal change from the reservoir’s point of view. Thus, we invoke the identity

\[
dS_{ext} = \frac{1}{T}(dU_{ext} + PdV_{ext} - \mu dN_{ext}).
\]
We throw away the $PdV$ term because it is so small, but keep the chemical term:\footnote{We do not bother to write $\mu_{ext}$ because at equilibrium, $\mu = \mu_{ext}$; just like temperature, $T = T_{ext}$.}

$$S_{ext}(s_2) - S_{ext}(s_1) = \frac{1}{T}[U_{ext}(s_2) - U_{ext}(s_1) - \mu_{N_{ext}(s_2)} + \mu_{N_{ext}(s_1)}]$$

$$= -\frac{1}{T}[U(s_2) - U(s_1) - \mu_{N(s_2)} + \mu_{N(s_1)}]$$

Here, $U$ and $N$ refer to the small system, hence the overall minus sign. The sign change is a consequence of conservation of energy, $\Delta U_{ext} = -\Delta U$ (energy lost by reservoir is gained by the system) and conservation of mass $\Delta N_{ext} = -\Delta N$ (number of particles lost by reservoir is gained by the system). Then,

$$P(s_2) = \frac{e^{-\frac{U(s_2) - \mu_{N(s_2)}}{k_BT}}}{e^{-\frac{U(s_1) - \mu_{N(s_1)}}{k_BT}}},$$

The Gibbs factor is

$$P(s) = \Xi e^{-\frac{U(s) - \mu_{N(s)}}{k_BT}},$$

where the normalization factor, $\Xi$, viewed as a function of temperature $T$,

$$\Xi(T) = \sum_{s'} e^{-\frac{U(s') - \mu_{N(s')}}{k_BT}}$$

is the grand partition function or Gibbs sum.

### 7.10. Problems

**Problem 100.** The rate constant of the elementary reaction

$$BH_4^-(aq) + NH_4^+(aq) \rightarrow BH_3NH_3(aq) + H_2(g)$$

is $k = 1.94 \times 10^{-4}$ L mol\(^{-1}\) s\(^{-1}\) at 30.0\(^\circ\)C, and the reaction has an activation energy of 161 kJ mol\(^{-1}\).

(a) Compute the rate constant of the reaction at a temperature of 40.0\(^\circ\)C.

(b) After equal concentrations of $BH_4^-(aq)$ and $NH_4^+(aq)$ are mixed at 30.0\(^\circ\)C, 1.00 \times 10^4 s is required for half of them to be consumed. How long will it take to consume half of the reactants if an identical experiment is performed at 40.0\(^\circ\)C?

**Solution.** (a) Recall the Arrhenius equation

$$k = A \exp(-E_a/RT)$$
Given \( k = 1.94 \times 10^{-4} \text{ L mol}^{-1} \text{ s}^{-1} \) at 30.0°C and \( E_a = 161 \text{ kJ mol}^{-1} \),

\[
A = \frac{1.94 \times 10^{-4} \text{ L mol}^{-1} \text{ s}^{-1}}{\exp(-161 \text{ kJ mol}^{-1}/(8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1})(303 \text{ K}))} = 1.106 \times 10^{24} \text{ L mol}^{-1} \text{ s}^{-1}
\]

At 40.0°C

\[
k = A \exp(-161 \text{ kJ mol}^{-1}/(8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1})(313 \text{ K})) = 1.49 \times 10^{-3} \text{ L mol}^{-1} \text{ s}^{-1}
\]

b) This reaction is second order, but the following reasoning applies to reactions of an order. The larger the rate constant, the more rapid is the reaction. Fast reactions require less time to reach any designated point in their progress. Increasing the temperature of this reaction from 30.0°C to 40.0°C increases \( k \) from \( 1.94 \times 10^{-4} \text{ L mol}^{-1} \text{ s}^{-1} \) to \( 1.49 \times 10^{-3} \text{ L mol}^{-1} \text{ s}^{-1} \), which is a factor of 7.68. The 50% conversion requires only \( 1.30 \times 10^3 \text{ s} \) at 40.0°C instead of 10000 s it requires at 30.0°C.

**Problem 101.** The first-order reaction \( X \rightarrow Y \) has a half-life of 10 minutes. Determine the percentage of the reactant remaining after one hour has elapsed.

**Solution.**

\[
k = \log(2)/t_{1/2} = \log(2)/(10 \times 60 \text{ s}) = 0.0011552 \text{ s}^{-1}
\]

\[
\frac{[X]}{[X]_0} = \exp[-kt] = \exp[-(0.0011552 \text{ s}^{-1})(3600 \text{ s})] = 0.0156275
\]

**Problem 102.** For a first-order reaction the time needed for the concentration of the reactant to be \( [X] = (1/e)[X]_0 \) is called the relaxation time and is equal to the inverse of the rate constant. Show that this relaxation time is the average lifetime of the reactant population.

**Solution.** The concentration at time \( t \) is given by \( [X] = [X]_0 e^{-kt} \)

The fraction of molecules present at a time between \( t \) and \( t + dt \) is \( E(t) = \frac{[X]}{\int_0^\infty [X]dt} \)

\[
\int_0^\infty [X]dt = \int_0^\infty [X]_0 e^{-kt}dt = [X]_0/k
\]

\( E(t) \) can be used as a PDF, so: \( t = \int_0^\infty tE(t)dt \)

\[
E(t) = ke^{-kt}
\]

\[
\bar{t} = \int_0^\infty tE(t)dt = k\int_0^\infty te^{-kt}dt = k/k^2 = 1/k.
\]
Problem 103. DTBP [(CH$_3$)$_3$COOC(CH$_3$)$_3$] decomposes into acetone and ethane by the following reaction:

$$(\text{CH}_3)_3\text{COOC(CH}_3)_3(\text{g}) \rightarrow 2(\text{CH}_3)_2\text{CO(}\text{g}) + \text{C}_2\text{H}_6(\text{g})$$

Given the table of data below which begins with a sample of pure DTBP, what is the approximate order of the reaction with respect to DTBP?

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>$P_{\text{tot}}$ (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.2362</td>
</tr>
<tr>
<td>2</td>
<td>0.2466</td>
</tr>
<tr>
<td>6</td>
<td>0.2613</td>
</tr>
<tr>
<td>10</td>
<td>0.2770</td>
</tr>
<tr>
<td>14</td>
<td>0.2911</td>
</tr>
<tr>
<td>18</td>
<td>0.3051</td>
</tr>
<tr>
<td>20</td>
<td>0.3122</td>
</tr>
<tr>
<td>22</td>
<td>0.3188</td>
</tr>
<tr>
<td>26</td>
<td>0.3322</td>
</tr>
<tr>
<td>30</td>
<td>0.3449</td>
</tr>
<tr>
<td>34</td>
<td>0.3570</td>
</tr>
<tr>
<td>38</td>
<td>0.3687</td>
</tr>
<tr>
<td>40</td>
<td>0.3749</td>
</tr>
<tr>
<td>42</td>
<td>0.3801</td>
</tr>
<tr>
<td>46</td>
<td>0.3909</td>
</tr>
</tbody>
</table>

Solution. $P_{\text{DTBP}}(t = 0) = P_{\text{tot}}(t = 0) = P_{\text{Ethane}}(t = 0) = 0 = P_{\text{Acetone}}(t = 0) = 0$

$P_{\text{tot}}(t) = P_{\text{DTBP}}(t) + P_{\text{Ethane}}(t) + P_{\text{Acetone}}(t)$

Because of stoichiometric constraints:

$P_{\text{tot}}(t) = P_{\text{DTBP}}(0) - x + P_{\text{Ethane}}(0) + x + P_{\text{Acetone}}(0) + 2x$

$P_{\text{tot}}(t) = P_{\text{DTBP}}(0) - x + x + 2x$

$x = (P_{\text{tot}}(t) - P_{\text{DTBP}}(0))/2$

$P_{\text{DTBP}}(t) = P_{\text{DTBP}}(0) - x = P_{\text{tot}}(0) - x = \frac{3}{2}P_{\text{tot}}(0) - \frac{1}{2}P_{\text{tot}}(t)$

So finding the partial pressures of DTBP gives us the following table:
Since the partial pressure changes the reaction can’t be 0th order by DTBP. The following table tests first order with $\log(P_{DTBP})$:

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>$P_{DTBP}$ (atm)</th>
<th>$\log(P_{DTBP})$</th>
<th>log(atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.2362</td>
<td>-1.443</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>0.2310</td>
<td>-1.465</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>0.2237</td>
<td>-1.497</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>0.2158</td>
<td>-1.533</td>
<td></td>
</tr>
<tr>
<td>14</td>
<td>0.2088</td>
<td>-1.566</td>
<td></td>
</tr>
<tr>
<td>18</td>
<td>0.2018</td>
<td>-1.600</td>
<td></td>
</tr>
<tr>
<td>20</td>
<td>0.1982</td>
<td>-1.618</td>
<td></td>
</tr>
<tr>
<td>22</td>
<td>0.1949</td>
<td>-1.635</td>
<td></td>
</tr>
<tr>
<td>26</td>
<td>0.1882</td>
<td>-1.670</td>
<td></td>
</tr>
<tr>
<td>30</td>
<td>0.1819</td>
<td>-1.704</td>
<td></td>
</tr>
<tr>
<td>34</td>
<td>0.1758</td>
<td>-1.738</td>
<td></td>
</tr>
<tr>
<td>38</td>
<td>0.1700</td>
<td>-1.772</td>
<td></td>
</tr>
<tr>
<td>40</td>
<td>0.1669</td>
<td>-1.806</td>
<td></td>
</tr>
<tr>
<td>42</td>
<td>0.1643</td>
<td>-1.839</td>
<td></td>
</tr>
<tr>
<td>46</td>
<td>0.1589</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

This yields a linear relation with a slope of about 0.089, so the reaction is first order relative to DTBP.
Problem 104. Short answer problems:
(a) What activation energy would result in a doubling of the rate of a reaction when the temperature is increased by 10°C near room temperature? Compare your answer with typical bond energies.

Solution. From the Arrhenius equation \( k = Ae^{-E_a/RT} \) we obtain the van't Hoff equation:

\[
\log \frac{k_2}{k_1} = -\frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)
\]

Since the rate constant is independent of concentration we can say:

\( k_2 = 2k_1 \)

when a reaction rate doubles (and we assume the concentration of reactants do not change). So,

\[
\log \frac{2k_1}{k_1} = -\frac{E_a}{R} \left( \frac{\Delta T}{T_1T_2} \right)
\]

where \( \Delta T = T_2 - T_1 \) (in this form of equation).

\[
E_a = \frac{R \log(2)(T_1T_2)}{\Delta T} = \frac{(8.315 \text{ J/mol/K})(0.693)(308 \text{ K})(298 \text{ K})}{10 \text{ K}} - 52,885 \text{ J/mol} = 52.9 \text{ kJ/mol}
\]

Typical bond energies are 300-800 kJ/mol. So, this ”\( E_a \)” is small compared to typical bond energies. Remember that \( E_a \) is the energy required to get to the transition state of a reaction and bonds are generally not completely broken in this transition state.

Problem 105. (b) The rate constant for the disappearance of chlorine in the third-order reaction of NO with Cl\(_2\) to form NOCl is 4.5 M\(^{-2}\)s\(^{-1}\) at 0°C and 8.0 M\(^{-2}\)s\(^{-1}\) at 22°C. What is the activation energy for this reaction?

Solution. Write the reaction

\[
2\text{NO}(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{NOCl}(g)
\]

We are told the reaction is third order so we can infer the rate law:

\[- \frac{d[\text{Cl}_2(g)]}{dt} = k[\text{NO}(g)]^2[\text{Cl}_2(g)]\]

Use the equation:

\[
\log \frac{k_2}{k_1} = \frac{E_a}{R} \left( \frac{\Delta T}{T_1T_2} \right)
\]

where \( \Delta T = T_2 - T_1 \).

\[
\log \frac{8.0 \text{ M}^{-2}\text{s}^{-1}}{4.5 \text{ M}^{-2}\text{s}^{-1}} = \frac{E_a}{8.3145 \text{ J/mol/K}} \left( \frac{22 \text{ K}}{(273.15 \text{ K})(293.15 \text{ K})} \right)
\]

solve for \( E_a \) to get \( E_a = 18 \text{ kJ/mol} \).
**Problem 106.** (c) At 627°C and 1 atm, SO$_3$ is partly dissociated into SO$_2$ and O$_2$: \( \text{SO}_3(g) = \text{SO}_2(g) + \frac{1}{2} \text{O}_2(g) \). The density of the equilibrium mixture is 0.925 g/liter. What is the degree of dissociation of SO$_3$ under these circumstances?

**Solution.** Let \( y = \) equilibrium \( P_{\text{dimer}} \). Let \( x = \) equilibrium \( P_{\text{monomer}} \). \( P_{\text{total}} = 0.725 \text{ atm} \) (at equilibrium).

\[
P_{\text{total}} = P_{\text{dimer}} + P_{\text{monomer}} \quad \text{(Dalton’s law)} \\
\therefore P_{\text{dimer}} + P_{\text{monomer}} = 0.725 \text{ atm} \\
y + x = 0.725 \text{ atm}
\]

We also know that:

\[
K_p = 3.72 = \frac{P_{\text{dimer}}}{(P_{\text{monomer}})^2}
\]

So:

\[
3.72 = \frac{(0.725 - x)}{x^2}
\]

Use the quadratic equation to solve for \( x \). \( x = 0.327 \). \( y = 0.398 \). \( P_{\text{monomer}} = 0.327 \text{ atm} \). \( P_{\text{dimer}} = 0.398 \text{ atm} \).

If none of the monomer is dimerized the \( P_{\text{monomer}} \) at equilibrium would be:

\[
0.327 \text{ atm} + (2 \times 0.398 \text{ atm}) = 1.123 \text{ atm}
\]

There is actually 0.327 atm of monomer present at equilibrium:

\[
\frac{0.327 \text{ atm}}{1.123 \text{ atm}} \times 100 = 29.1\%
\]

\[\therefore 100\% - 29.1\% = 90.9\% \text{ dimerized} \]

**Problem 107.** (d) Write an expression for the equilibrium constant for the endothermic reaction

\[ \text{AgCl} \cdot \text{NH}_3(s) \rightleftharpoons \text{AgCl}(s) + \text{NH}_3(g) \]

What is the effect on \( P_{\text{NH}_3} \) at equilibrium if additional AgCl(s) is added or if additional NH$_3$ is pumped into or out of the system, provided that neither of two solids phases shown in the chemical equation are completely used up? What is the effect on the \( P_{\text{NH}_3} \) of lowering the temperature? Why?

**Solution.**

\[ K = [\text{NH}_3(g)] \quad K_p = P_{\text{NH}_3} \]

Addition of AgCl(s) has no effect on \( P_{\text{NH}_3} \).

If \( P_{\text{NH}_3} \) increases, it will react with solid AlCl(s) to relieve the stress on the system. The reaction shifts to the left (from equilibrium) to achieve a new equilibrium condition. The opposite will occur if \( P_{\text{NH}_3} \) is decreased.
The reaction is endothermic and can be written as
\[ \text{AgCl} \cdot \text{NH}_3(s) + \text{(heat)} \leftrightharpoons \text{AgCl}(s) + \text{NH}_3(g) \]
In this context, lowering the temperature means removing heat. The reaction shifts to the left and \( P_{\text{NH}_3} \) is reduced.

**Problem 108.** Reaction rate problems:
(a) Consider the following reactions:

(I) \( \text{COCl}_2(g) \leftrightharpoons \text{CO}(g) + \text{Cl}_2(g) \)

(II) \( \text{FeCl}_2\text{O}_4(s) \leftrightharpoons \text{FeO}(s) + \text{CO}(g) + \text{CO}_2(g) \)

Write an expression for the equilibrium constant for each reaction.

**Solution.** For reaction I

\[ K_p = \frac{P_{\text{CO}}P_{\text{Cl}_2}}{P_{\text{COCl}_2}} \quad K = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]} \]

for reaction II

\[ K_p = P_{\text{CO}}P_{\text{CO}_2} \quad K = [\text{CO}][\text{CO}_2] \]

**Problem 109.** Reaction (I) from the previous problem is endothermic. Suppose that an equilibrium mixture of the three gases is heated. How will the degree of dissociation of \( \text{COCl}_2 \) change?

**Solution.** We can rewrite the reaction as follows:

\( \text{(heat)} + \text{COCl}_2(g) \leftrightharpoons \text{CO}(g) + \text{Cl}_2(g) \)

to denote an endothermic reaction. A simple way to view the question is to look at the revised chemical reaction and realize that "heat" is a reaction reactant (as opposed to a product). If we add heat, the reaction will be out of equilibrium and shift to the right toward more products. The best way to approach this question is to find the dependence of \( K \) on \( T \)

\[ K = \frac{A_f}{A_b} e^{-\Delta H^\circ/RT} \]

We assume \( A_f/A_b \) is temperature independent. \( \Delta H^\circ > 0 \) (endothermic). Therefore, as \( T \) goes up, \( e^{-\Delta H^\circ/RT} \) goes up, and \( K \) goes up. (As \( T \uparrow, K \uparrow \)). The degree of dissociation of the reactant \( \text{COCl}_2(g) \) increases.

**Problem 110.** Predict the effects on the amount of CO of a decrease in volume for reaction (I) and of a decrease in amount of FeO(s) in reaction (II).
**Solution.** Decrease in volume for reaction (I)

\[
\text{COCl}_2(g) \rightleftharpoons \text{CO}(g) + \text{Cl}_2(g)
\]

1 mole gas \quad 2 moles of gas

The decrease in volume will increase the total pressure and the system will relieve this stress by reducing the number of moles of gaseous species. Since there are 2 moles of gas on the product side for every 1 mole of gas on the reactant side, the reaction will shift to the left and the amount of CO(g) will decrease.

Decrease in amount of FeO(s) for reaction (II)

Since FeO(s) is a pure substance, it is not in the equilibrium equation and therefore any change in the amount of FeO(s) does not affect the equilibrium constant, \(K\) (or \(K_p\)) and the amount of CO(g) would not change.

**Problem 111.** (b) At high temperature, nitrogen dioxide decomposes into NO and \(O_2\): \(2\text{NO}_2 \rightarrow 2\text{NO} + \text{O}_2\) by the second order rate law \(d[\text{NO}_2]/dt = -k[\text{NO}_2]^2\). At 592 K, the rate constant is \(5.0 \times 10^{-1}\) liter/mol/s and at 656 K it is 4.7 liter/mol/s. Calculate the activation energy for this process.

**Solution.**

\[
\log \frac{K(T_2)}{K(T_1)} = -\frac{E_a}{R} \left( \frac{T_1 - T_2}{T_1 T_2} \right)
\]

Therefore

\[
E_a = -\frac{RT_1 T_2}{\Delta T} \log \frac{K(T_2)}{K(T_1)} = -(8.314)(592)(656) \frac{1}{-64} \log \frac{4.7 \times 10}{5.0 \times 10^{-1}} \text{kJ/mol} = 113 \text{kJ/mol}
\]

**Problem 112.** The equilibrium constant for the reaction \(\text{NO}_2 \rightleftharpoons \text{NO} + (1/2)\text{O}_2\) is 0.0604 atm\(^{1/2}\) at 592 K and 0.358 atm\(^{1/2}\) at 700 K. What is the rate for the reverse reaction, \(2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2\) at 592 K?

**Solution.** First reaction has \(K_p\). Second reaction has \(K'_p = (K_p)^2\). When you multiply a chemical reaction equation by \(n\) you must raise its equilibrium constant to the power of \(n\). The rate law is given in terms of molarity so convert the \(K'_p\) to \(K'\) as follows:

\[
K'_p = K'(RT)^{\Delta n_g} \quad \Delta n_g = 3 \text{ mol(gas)} - 2 \text{ mol(gas)}
\]

\[
K' = \frac{K'_p}{RT} = \frac{(K_p)^2}{RT} = \frac{(0.0604 \text{ atm}^{1/2})^2}{(0.08206 \text{ L.atm/mol}/K)(592 \text{ K})} = 7.51 \times 10^{-5} \text{ M}
\]
7. Chemical Kinetics

\[ K' = \frac{k_f}{k_b} \rightarrow k_b = \frac{k_f}{K'} \]

\[ k_b = \frac{5.0 \times 10^{-1} \text{ L/mol/s}}{7.51 \times 10^{-5} \text{ mol/L}} = 6.658 \times 10^3 \text{ L}^2/\text{mol}^2/\text{s} = 6.7 \times 10^3 \text{ L}^2/\text{mol}^2/\text{s} \]

**Problem 113.** What is the activation energy for the reverse reaction? Make a sketch of energy versus reaction coordinate illustrating your answer. [Note that the rate equation and the equilibrium equation differ by a factor of 2.]

**Solution.** 1st calculate \( \Delta H^\circ \) for \( \text{NO}_2(g) \rightarrow \text{NO}(g) + (1/2)\text{O}_2(g) \)

\[
\log \frac{K(T_2)}{K(T_1)} = -\frac{\Delta H^\circ}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)
\]

\[
\Delta H^\circ = -\frac{RT_1T_2}{\Delta T} \log \frac{K(T_2)}{K(T_1)}
\]

\[
= -\frac{(8.314 \text{ J/mol/K})(592 \text{ K})(700 \text{ K})}{-108 \text{ K}} \log \frac{0.358 \text{ atm}^{1/2}}{0.0604 \text{ atm}^{1/2}} = 56.7 \text{ kJ/mol}
\]

Now, \( \Delta H^\circ = 2(\Delta H^\circ') = 2(56.7 \text{ kJ/mol}) = 113 \text{ kJ/mol} \)

\[ E_{a,b} = E_{a,f} - \Delta E_a \]

and \( \Delta E_a \approx \Delta H^\circ \). Therefore,

\[ E_{a,b} \approx 113 \text{ kJ/mol} - 113 \text{ kJ/mol} = 0 \]

**Problem 114.** (c) What volume of 1.0 M KOH must be added to 10 ml of 1.0 M acetic acid (\( \text{CH}_3\text{COOH} \)) to give pH of 5.0? (\( K_a \) for acetic acid is \( 1.8 \times 10^{-5} \) M). Is this mixture a good buffer? Why or why not?

**Solution.** The reaction is

\[ \text{CH}_3\text{COOH}(aq) + \text{OH}^- (aq) \rightarrow \text{CH}_3\text{COO}^- (aq) + \text{H}_2\text{O}(l) \]

\[ K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5} \]

\[ \text{pH} = 5.0 \quad [\text{H}_3\text{O}^+] = 1.0 \times 10^{-5} \text{ M} \]

Let \( x=\text{KOH} \) added in "liters"

\[ [\text{CH}_3\text{COO}^-] = \frac{(1.0 \text{ M})(x)}{0.010 \text{ L} + x} \]

\[ [\text{CH}_3\text{COOH}] = \frac{(1.0 \text{ M})(0.010 \text{ L}) - (1.0 \text{ M})(x)}{0.010 \text{ L} + x} \]

\[ K_a = 1.8 \times 10^{-5} = \frac{(1.0 \text{ M})(x)}{(0.010 \text{ L}+x)} \cdot (1 \times 10^{-5} \text{ M}) \]

\[ (1.0 \text{ M})(0.010 \text{ L})-(1.0 \text{ M})(x) \]

\[ (0.010 \text{ L}+x) \]
7.10. Problems

\[ x = 6.4 \text{ mL of KOH} \]

\[ \therefore [\text{CH}_3\text{COO}^-] = \frac{(1.0 \text{ M})(0.0064 \text{ L})}{(0.010 \text{ L} + 0.0064 \text{ L})} = 0.0064 \text{ mol} = 0.39 \text{ M} \]

\[ [\text{CH}_3\text{COOH}] = \frac{(1.0 \text{ M})(0.010 \text{ L}) - (1.0 \text{ M})(0.0064 \text{ L})}{(0.010 \text{ L} + 0.0064 \text{ L})} = 0.22 \text{ M} \]

This is a good buffer system since \( [\text{CH}_3\text{COO}^-] \approx [\text{CH}_3\text{COOH}] \)

Problem 115. (d) Suppose you have a buffer at pH 5.0 that is 1.00 M in HX and 1.00 M in X\(^-\), where HX is a weak acid with \( K_a = 1.0 \times 10^{-5} \) M. Calculate the effects of adding 50 ml of 1.0 M HCl to a 1-liter portion of this buffer and of adding 50 ml of 1.0 M NaOH to a separate 1-liter portion. What will the final pH be in each case?

Solution. 1st addition of the 50 mL of 1.0 M HCl

\[ [X^-] = \frac{(1.0 \text{ M})(1 \text{ L}) - (1.0 \text{ M})(0.050 \text{ L})}{1.050 \text{ L}} = 0.90 \text{ M} \]

\[ [\text{HX}] = \frac{(1.0 \text{ M})(1 \text{ L}) + (1.0 \text{ M})(0.050 \text{ L})}{1.050 \text{ L}} = 1.0 \text{ M} \]

\[ p\text{H} = pK_a + \log_{10} \frac{[X^-]}{[\text{HA}]} = 5.00 + \log_{10} \frac{0.90}{1.0} = 4.95 \]

2nd addition of the 50 mL of 1.0 M NaOH

\[ [X^-] = \frac{(1.0 \text{ M})(1 \text{ L}) + (1.0 \text{ M})(0.050 \text{ L})}{1.050 \text{ L}} = 1.0 \text{ M} \]

\[ [\text{HX}] = \frac{(1.0 \text{ M})(1 \text{ L}) - (1.0 \text{ M})(0.050 \text{ L})}{1.050 \text{ L}} = 0.90 \text{ M} \]

\[ p\text{H} = pK_a + \log_{10} \frac{[X^-]}{[\text{HA}]} = 5.00 + \log_{10} \frac{1.0}{0.90} = 5.05 \]

Problem 116. (e) Describe how you would prepare 1 liter of a buffer at pH=9.0, using 1.0 moles NH\(_3\) and as much of any strong acid as needed.

Solution. We must look up the \( K_a \) for NH\(_4^+\)

\[ K_A = 5.67 \times 10^{-10} \quad pK_a = 9.25 \]

\[ p\text{H} = pK_a + \log_{10} \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \]
\[
\text{NH}_3 + \text{H}^+ \rightleftharpoons \text{NH}_4^+
\]

| Initial (M) | 1.0 | 0 |
| Change (M) | \(-x\) | \(+x\) |
| Equilibrium (M) | \((1.0 - x)\) | \(+x\) |

\[
\therefore 9.00 = 9.25 + \log_{10} \left( \frac{1.0 - x}{x} \right)
\]

Solve for \(x = 0.64\) moles HCl added to a liter (M)

“moles HCl added = moles NH\text{$_4^+$}”

**Problem 117.** (e) A 40-ml sample of a 0.10 M solution of nitric acid is added to 20 ml of 0.30 M aqueous ammonia. What is the pH of the resulting solution?

**Solution.**

\[
pH = pK_a + \log_{10} \left( \frac{[\text{NH}_3]}{[\text{NH}_4^+] \cdot [\text{H}^+]} \right)
\]

\[
= 9.25 + \log_{10} \left( \frac{(0.020 \text{ L})(0.30 \text{ M}) - (0.040 \text{ L})(0.10 \text{ M})}{(0.040 \text{ L})(0.10 \text{ M})} \right)
\]

\[
pH = 9.25 + \log_{10}(0.50) = 9.25 - 0.30 = 8.95
\]

**Question:** why did we not calculate the molarity of the conjugate base and weak acid in the \(\log_{10}(x)\) in the above equation?

**Problem 118.** (f) The molar enthalpy of fusion of ice at 0\(^\circ\)C is 6.02 kJ/mol. The molar heat capacity of supercooled water is 75.3 J/mol/K. One mole of supercooled water at -10\(^\circ\)C is induced to crystallize in a heat-insulated vessel. The result is a mixture of ice and water at 0\(^\circ\)C. What fraction of this mixture is ice? What is \(\Delta S\) for the system?

**Solution.** The freezing of metastable supercooled water is certainly a spontaneous and irreversible process. The transformation can be thought of as two separate reversible processes.

1. \(l(-10^\circ \text{ C}) \rightarrow l(0^\circ \text{ C}) \) “heat is required”
2. \(l(0^\circ \text{ C}) \rightarrow s(0^\circ \text{ C}) \) “heat is evolved”

Because the system is thermally isolated from the surroundings, \(Q = 0\). Therefore, the heat required to raise the temperature of the liquid from -10\(^\circ\) C to 0\(^\circ\) C equals the heat evolved from the fusion of the ice. i.e. \(q_1 = -q_2\). Therefore,

\[
q_1 = n\bar{C}_p\Delta T = (1.00 \text{ mol})(75.3 \text{ J/K.mol})(10 \text{ K}) = 733 \text{ J}
\]

\[
q_2 = n\Delta\tilde{H}_{fus}
\]
but the magnitude of \( q_1 \) and \( q_2 \) are equal. So,

\[
n = \frac{q_2}{\Delta H_{\text{fus}}} = \frac{753 \text{ J}}{6.02 \times 10^3 \text{ J/mol}} = 1.25 \times 10^{-1} \text{ mol}
\]

(amount of ice formed)

\[
\%\text{ICE} = \frac{1.25 \times 10^{-1} \text{ mol frozen}}{1.00 \text{ mol total}} \times 100 = 13\%
\]

\[
\Delta S_1 = n \tilde{C}_p \log \frac{T_f}{T_i} = (1.00 \text{ mol})(75.3 \text{ J/mol}) \log \left( \frac{273}{263} \right) = 2.81 \text{ J/K}
\]

\[
\Delta S_2 = \frac{-(n)\Delta \tilde{H}_{\text{fus}}}{T_{\text{fus}}} = \frac{-(1.25 \times 10^{-1} \text{ mol})(6.02 \times 10^3 \text{ J/K.mol})}{273 \text{ K}} = -2.76 \text{ J/K}
\]

\[
\Delta S_{\text{total}} = 2.81 \text{ J/K} + (-2.76 \text{ J/K}) = 5 \times 10^{-2} \text{ J/K}
\]

Small, but still greater than zero. ■

**Problem 119.** Calculate \( \Delta H^\circ \) and \( \Delta G^\circ \) at 25°C for the reaction:

\[
3\text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}_3\text{O}_4(s) + (1/2)\text{O}_2(g)
\]

Which of the two oxides is more stable at 25°C and \( P_{\text{O}_2} = 1 \text{ atm} \)?

**Solution.**

\[
K_p = (P_{\text{O}_2})^{1/2}
\]

\[
\Delta G^\circ = \left( [\frac{1}{2} \text{ mol}](0) + (2 \text{ mol})(-1014.2 \text{ kJ/mol}) \right) - [(3 \text{ mol})(-741.0 \text{ kJ/mol})] = 195 \text{ kJ} > 0
\]

The process is not spontaneous as written. Therefore, \( \text{Fe}_2\text{O}_3(s) \) is the more "stable" of the two iron oxides. ■

**Problem 120.** At 1200 K and in the presence of solid carbon, an equilibrium mixture of CO and \( \text{CO}_2 \) contains 98.3 mol percent CO and 1.69 mol percent \( \text{CO}_2 \) with the total pressure at 1 atm. What are \( P_{\text{CO}} \) and \( P_{\text{CO}_2} \), what is the equilibrium constant, and what is \( \Delta G^\circ \) associated with the reaction \( \text{CO}_2(g) + \text{C(graphite)} \rightleftharpoons 2\text{CO}(g) \) at 1200 K?

**Solution.**

\[
K_p = \frac{P_{\text{CO}}^2}{P_{\text{CO}_2}} = \frac{(0.983 \text{ atm})^2}{(1.69 \times 10^{-2} \text{ atm})} = 57.2 \text{ atm}
\]

\[
\Delta G^\circ = -RT \log K_p = -(8.315)(1200) \log(57.2) \text{ J/mol} = 40,376 \text{ J/mol}
\]

= 40.4 kJ/mol ■
Problem 121. (i) Part (1): Calculate the standard free-energy change and the equilibrium constant for dimerization of NO\textsubscript{2} to N\textsubscript{2}O\textsubscript{4} at 25° C.

Part (2): Calculate \( \Delta G \) for this reaction at 25° C when the pressure of NO\textsubscript{2} and N\textsubscript{2}O\textsubscript{4} are each held at 0.01 atm. Which way will the reaction tend to proceed?

Solution. Part (1):
\[
\Delta G^\circ = [(1 \text{ mol})(98.3 \text{ kJ/mol})] - [(2.00 \text{ mol})(51.8 \text{ kJ/mol})] = -5.3 \text{ kJ/mol of } \text{N}_2\text{O}_4
\]

\[K_p = e^{-\Delta G^\circ / RT} = e^{-(5.3 \text{ kJ/mol})(1000 \text{ J/kJ})/(8.315 \text{ J/K mol})/(298 \text{ K})} = 8.50\]

Part (2):
\[
Q = \frac{P_{\text{N}_2\text{O}_4}}{P_{\text{N}_2\text{O}_4}^2} = \frac{(0.01 \text{ atm})}{(0.01 \text{ atm})^2} = 100 \text{ atm}^{-1}
\]
\[
\Delta G = \Delta G^\circ + RT \log \frac{Q}{1 \text{ atm}} = -5.3 \times 10^3 \text{ J/mol} + (8.315)(298) \log(100) = 6.1 \text{ kJ/mol} > 0
\]

not spontaneous! Reaction shifts back toward reactants. ■

Problem 122. The gas phase decomposition of ozone, O\textsubscript{3} (i.e. 2 O\textsubscript{3} \rightarrow 3 O\textsubscript{2}) is second order with respect to ozone:

\[-\frac{d[O_3]}{dt} = k[O_3]^2\]

This “constant” \( k \) can be a function of the concentration of other species, and in fact, at 100° C, \( k \) is found to depend on the partial pressure of molecular oxygen according to the following table:

<table>
<thead>
<tr>
<th>( P_{O_3}/\text{torr} )</th>
<th>( P_{O_2}/\text{torr} )</th>
<th>( k/\text{M}^{-1}\text{s}^{-1} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>6.0</td>
<td>800</td>
<td>0.16</td>
</tr>
<tr>
<td>6.0</td>
<td>600</td>
<td>0.22</td>
</tr>
<tr>
<td>6.0</td>
<td>400</td>
<td>0.35</td>
</tr>
</tbody>
</table>

(Note that the data presented above involves a low percentage of O\textsubscript{3} in O\textsubscript{2}.)

(a) What is the rate law for the O\textsubscript{2} concentration under these conditions?

(b) What is the value (with units) of the rate constant in the rate law that includes the dependence on [O\textsubscript{2}]?
(c) If the table presented above was obtained by measuring \([\text{O}_3]\) as a function of time, what graphs might you construct to obtain the \(k\) values?

(d) Why does the \(\text{O}_2\) produced by the decomposition of \(\text{O}_3\) not seriously affect the accuracy of the graphs?

(e) The half-life of a reaction is the time required for the concentration to fall to half its initial value. What would be the half-life of the ozone in each of the above experiments, i.e., for each of the initial conditions in the table presented above?

**Solution.** (a) In general, \(k = 2k'[\text{O}_2(g)]^\alpha\), where \(k'\) is a constant. Now, from the experimental data, choose two experiments to compare. \(k_1 = 2k'[\text{O}_2(g)]_1^\alpha\) and \(k_2 = 2k'[\text{O}_2(g)]_2^\alpha\)

\[
\frac{k_1}{k_2} = \frac{2k'[\text{O}_2(g)]_1^\alpha}{2k'[\text{O}_2(g)]_2^\alpha} = \left(\frac{[\text{O}_2(g)]_1}{[\text{O}_2(g)]_2}\right)^\alpha
\]

taking logs

\[
\log(k_1/k_2) = \alpha \log([\text{O}_2(g)]_1/[\text{O}_2(g)]_2)
\]

we can solve for \(\alpha\)

\[
\alpha = \frac{\log(k_1/k_2)}{\log([\text{O}_2(g)]_1/[\text{O}_2(g)]_2)} = \frac{\log(\frac{0.16 \text{ M}^{-1} \text{s}^{-1}}{0.22 \text{ M}^{-1} \text{s}^{-1}})}{\log(\frac{0.034 \text{ M}}{0.026 \text{ M}})}
\]

\[
= \frac{\log(0.727)}{\log(1.308)} = -0.319
\]

The rate law can be expressed as

\[
-\frac{d[\text{O}_3(g)]}{dt} = 2k'[\text{O}_3(g)]^2[\text{O}_2(g)]^{-1.1}.\]

(b) Since \([\text{O}_2(g)] \gg [\text{O}_3(g)]\), we can use rate = \(2k'[\text{O}_2(g)]^{-1.1}\) only to calculate the rate constant \(k'\).

\[
0.16 \text{ M}^{-1} \text{s}^{-1} = 2k'(0.034 \text{ M})^{-1.1}
\]

\[
k' = 1.9 \times 10^{-3} \text{ s}^{-1}
\]
(c) Graph $1/[O_3(g)]$ vs. time for the three experiments.

(d) The $[O_2(g)]$ is so much larger than the $[O_3(g)]$ that any $O_2(g)$ produced from $O_3(g)$ would not appear to change the overall $[O_2(g)]$. In other words, since $[O_2(g)]$ is so large, we can treat it as a constant compared to $[O_3(g)]$.

(e) Using $[A]/[A]_0 = e^{-kt_{1/2}}$ and appropriate initial concentration and rate, calculate the half-life for each experiment. Setting $[A] = [A]_0/2$ per the definition of half-life we obtain $\log(1/2) = -kt_{1/2}$, $[A]_0$ goes away and $t_{1/2} = 0.693/k$. Plug in "k" and solve for $t_{1/2}$.

Experiment 1 $t_{1/2} = 0.693/0.16 \text{ s}^{-1} = 4.3 \text{ s}$

Experiment 2 $t_{1/2} = 0.693/0.22 \text{ s}^{-1} = 3.2 \text{ s}$

Experiment 3 $t_{1/2} = 0.693/0.35 \text{ s}^{-1} = 2.0 \text{ s}$

Why did we drop the M$^{-1}$ part from the units of k designated in the table?

Problem 123. The following data have been obtained for the instantaneous rate of the following reaction occurring in solution: $3A + 2B \rightarrow A_3B_2$.

<table>
<thead>
<tr>
<th>[A]/M</th>
<th>[B]/M</th>
<th>$(d[A]/dt)/(\text{M.s}^{-1})$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.10</td>
<td>0.10</td>
<td>$6.0 \times 10^{-4}$</td>
</tr>
<tr>
<td>0.30</td>
<td>0.30</td>
<td>$5.4 \times 10^{-3}$</td>
</tr>
<tr>
<td>0.50</td>
<td>0.10</td>
<td>$1.5 \times 10^{-2}$</td>
</tr>
<tr>
<td>0.50</td>
<td>0.50</td>
<td>$1.5 \times 10^{-2}$</td>
</tr>
</tbody>
</table>

(a) Calculate the order of the reaction with respect to A and with respect to B, and calculate the value of k (including its units).

(b) Suppose that you have 2.0 liters of a solution that is 0.50 M in A and 0.50 M in B. How many moles of $A_3B_2$ are formed in the first second of
reaction?

(c) How long will it take for the concentration of A to fall to 0.05 M from the initial conditions described in (b)?

(d) A reaction proceeds by the following mechanism: \( mA + nB + pC \rightarrow \) products, with \( m, n \) and \( p \) all positive integers. Doubling the concentration of A, B and C increases the overall rate of reaction by a factor of 16. Tripling the concentration of C has the same effect as tripling the concentration of A. An increase in the concentration of B has a larger effect than an increase in the concentration of A. What are \( m, n \) and \( p \)? What is unrealistic about this mechanism?

**Solution.** (a) In general the rate is

\[ k[A]^α[B]^β \]

Increasing \([A]\) by 5\( \times \) (0.10 M \(\rightarrow\) 0.50 M) while holding the \([B]\) constant at 0.10 M, the rate increases by 25\( \times \) \((6.0 \times 10^{-4} \text{ M/s} \rightarrow 1.5 \times 10^{-2} \text{ M/s})\)

\[ \therefore \alpha = 2. \]

Increasing \([B]\) by 5\( \times \) (0.10 M \(\rightarrow\) 0.50 M) while holding the \([A]\) constant at 0.50 M, the rate remains constant at 1.5 \(\times\) 10\(^{-2}\) M/s.

\[ \therefore \beta = 0. \]

Thus,

\[ \text{rate} = k[A]^2. \]

Now, to calculate \( k \) we will use the rate law determined with one set of experimental data (your choice). However, for the calculation of \( k \) and for the number of moles of \( A_3B_2 \) formed (part b) there are a number of possible approaches (two answers). It is unclear that the “rate” data refers to the “reaction rate” and includes the 1/3 factor for “A” or if the “rate” data should be corrected by multiplication. i.e.,

\[ \text{rate} = -\frac{1}{3} \frac{d[A]}{dt} \]

The table gives us \( d[A]/dt \). So, the most correct way to view the data is that a reaction rate must be calculated from the change in concentration of A with time data. If you assumed the data included the 1/3 factor you will receive full credit. Thus,

\[ k = \frac{(1/3)(6.0 \times 10^{-4} \text{ M/s})}{(0.10 \text{ M})^2} = 0.020 \text{ M}^{-1}\text{s}^{-1}. \]
(b) $[A] = [B] = 0.5$ M. rate = $1.5 \times 10^{-2}$ M/s for $-d[A]/dt$. But $-(1/3)d[A]/dt = +d[A_3B_2]/dt$ and
\[
\frac{d[A_3B_2]}{dt} = \frac{1}{3}(1.5 \times 10^{-2} \text{ M/s}) = 5.0 \times 10^{-3} \text{ M/s}.
\]
So, $(5.0 \times 10^{-3} \text{ M/s})(1 \text{ s}) = 5.0 \times 10^{-3} \text{ mol/L}$ and $(5.0 \times 10^{-3} \text{ mol/L})(2.0 \text{ L}) = 1.0 \times 10^{-2} \text{ mol A}_3\text{B}_2$ produced.

(c) From $[A] = [A]_0 e^{-kt}$ we have $\log([A]/[A]_0) = -kt$ and $t = -(1/k) \log[A]/[A]_0$. Plug in the values given
\[
t = -\left(\frac{1}{0.020 \text{ M}^{-1}\text{s}^{-1}}\right) \log \left(\frac{0.05 \text{ M}}{0.50 \text{ M}}\right) = 115 \text{ s} \approx 120 \text{ s}
\]
(d) 1st step: setup the general rate law
\[
k[A]^m[B]^n[C]^p
\]
Why can we represent the rate law by using the reaction coefficients?

1st step: determine $m, n, p$

Doubling all concentrations results in a 16x increase in the rate
\[
(2)^x = 16, \quad x = 4
\]
Thus, the overall order of this reaction is 4th order. Tripling $[A]$ or $[C]$ results in the same rate increase
\[
m = p
\]
Then,
\[
m + n + p = 4, \quad \text{with } m = p
\]
\[
m = p = 1, \quad n = 2
\]
Thus, we have that the rate is
\[
k[A][B]^2[C]
\]
Why is this mechanism unrealistic: Remember, this is a mechanism (elementary step) and the collision of four molecules is very unlikely.
Entropy Production, Chemical Kinetics and Irreversibility

In the context of chemical reactions we have stated that the reaction rate, $\dot{\xi}$, can be written down by inspection of the reaction mechanism (elementary reactions) as a sum of terms describing various molecular collision events times a reaction rate. We also previously stated, in the context of the Second Law, that differences in the chemical potential, $\Delta(\mu/T)$ give rise to a driving force for chemical transformations (or mass transport, if the properties are spatially inhomogeneous). The method of molecular collisions for determining $\dot{\xi}$ agrees with experiments and molecular dynamics simulations. On the other hand, $\dot{\xi}$ should also depend on $\Delta(\mu/T)$. But exactly how? An arbitrary function of $\Delta(\mu/T)$ could be expressed as a power series in $\Delta(\mu/T)$. For small deviations from equilibrium, the affinity is very low and only the first power matters. Thus, $\dot{\xi}$ should depend linearly on $\Delta(\mu/T)$ in this near-equilibrium regime.

This chapter is a very short introduction to a vast subject. To learn more about non-equilibrium phenomena in the context of thermodynamics, see \cite{5, 7, 6}. Our presentation follows Chapter 14 of McQuarrie \cite{5}. We use both upper and lower case letters for the state variables, with the latter referring to “per unit volume” quantities, i.e. $u \equiv U/V$ is the internal energy per unit volume.
8. Entropy Production, Chemical Kinetics and Irreversibility

About notation: we consider the case of two variables (time, t, and space, x) and write d/dx for the spatial derivative, where it is understood that when differentiating we keep t constant (i.e. it means partial differentiation, ∂/∂x). In several spatial dimensions, ∂/∂x is replaced by the gradient operator, ∇. (And of course, an equation that involves d/dx and d/dt implies that both are partial derivatives: ∂/∂x and ∂/∂t, respectively.)

8.1. Entropy Production

The change in the entropy of a system, dS, can be decomposed into two terms,

\[ dS = d_eS + d_iS, \]

where \( d_eS \) is the entropy supplied to the system by its surroundings (e, external), and \( d_iS \) is the entropy produced inside the system (i, internal). The second law states that \( d_iS = 0 \) for a reversible process and positive for irreversible transformations of the system, i.e.

\[ d_iS \geq 0. \]

This statement is true regardless of the nature of the system (provided that the e and i contributions are correctly identified).

The term \( d_iS \) is called the entropy production, and describes the various irreversible phenomena which may occur inside the system. It will be convenient to divide \( S \) by \( V \) to get the specific entropy \( s \), i.e. the entropy per unit volume. Division of \( ds \) by \( dt \) will be called the rate of entropy production:

\[ \dot{s} = \frac{d_iS}{dt} = \frac{d_iS}{dV}. \]

Furthermore, we shall assume that the fundamental equation of thermodynamics (also known as the Gibbs equation) holds at every point in space, and even away from equilibrium:

\[ dU = TdS - PdV + \sum_i \mu_i dn_i. \]

This equation was previously assumed to hold for a reversible process. Here, we assume that it holds for an irreversible process as well. This assumption may be justified in a volume element that is sufficiently small. We will write it in terms of specific quantities, \( u = U/V, \), \( s = s/V, \) \( [X_i] = n_i/V \). For constant volume (\( dV = 0 \)),

\[ du = Tds + \sum_i \mu_i d[X_i]. \]
Dividing by $dt$, and solving for $ds/dt$, we get the rate of change in the entropy:

\[[8.1] \dot{s} \equiv \frac{ds}{dt} = \frac{1}{T} \frac{du}{dt} - \sum_i \frac{\mu_i}{T} \frac{d[X_i]}{dt} \]

The first term is the heat transfer (can be positive or negative). The second term, which always non-negative $(\geq 0)$, is entropy production, an irreversible process. Entropy production describes the thermodynamic lost work. It quantifies irreversibility, and is therefore a measure of inefficiency. $\frac{d[X_i]}{dt}$ is discussed in the next section.

### 8.2. Chemical Reaction Away From Equilibrium

Suppose that we have a reversible first-order elementary chemical reaction

\[X \xrightleftharpoons[k_{XY}]^{k_{YX}} Y.\]

The reaction rate is:

\[[8.2] \dot{\xi} = - \frac{d[X]}{dt} = \frac{d[Y]}{dt} = k_{XY} [X] - k_{YX} [Y].\]

Close to equilibrium, $\dot{\xi}$ should be a linear function of $A$ (lowest order term in a power series in $A$). Indeed, according to Eq. (8.2), $\dot{\xi}$ should be a function of $A$ because it depends on the differences in concentrations of $X$ and $Y$. Close to equilibrium, where $A$ is small, we expect $\dot{\xi}$ to depend linearly on $A$:

\[[8.3] \dot{\xi} = L \left( \frac{A}{T} \right) = L \left( \frac{\mu_X - \mu_Y}{T} \right),\]

where $L$ is a coefficient to be determined.

At equilibrium, $\dot{\xi} = 0$, and we have $k_{XY}[X]_{eq} = k_{YX}[Y]_{eq}$, or $k_{XY} = k_{YX}[Y]_{eq} / [X]_{eq}$. In the first term, let’s substitute this relationship for $k_{XY}$; in the second term, let’s multiply it by $1 = [Y]_{eq} / [Y]_{eq}$:

\[\dot{\xi} = k_{YX} \cdot \frac{[Y]_{eq}}{[X]_{eq}} [X] - k_{YX} \cdot [Y] \cdot \frac{[Y]_{eq}}{[Y]_{eq}} = k_{YX} \cdot [Y]_{eq} \left( \frac{[X]}{[X]_{eq}} - \frac{[Y]}{[Y]_{eq}} \right).\]

On the other hand, consider the affinity, $\mu_X - \mu_Y$, in the case where we have ideal solutions. Choosing the reference conditions to be the chemical

---

1This term is also equal to $\dot{\xi}(-\Delta G_{rxn}/T)$. If a reaction proceeds from left to right, $\Delta G_{rxn} < 0$, and $\dot{\xi} > 0$. (And conversely, if the reaction proceeds from right to left, $\Delta G_{rxn} > 0$, $\dot{\xi} < 0$.) At equilibrium, of course, $\Delta G_{rxn} = 0$. Thus, the term can never be negative.
equilibrium concentrations:
\[ \mu_X = \mu_X^0 + RT \log \frac{[X]}{[X]_{eq}}, \quad \mu_Y = \mu_Y^0 + RT \log \frac{[Y]}{[Y]_{eq}}. \]

Then,
\[ \frac{\mu_X - \mu_Y}{T} = R \left( \log \frac{[X]}{[X]_{eq}} - \log \frac{[Y]}{[Y]_{eq}} \right), \]

Adding and subtracting the same quantity is like adding zero; a given term does not change. Thus, we rewrite
\[ \frac{[X]}{[X]_{eq}} = \left( \frac{[X] - [X]_{eq}}{[X]_{eq}} + 1 \right), \]

and similarly for \( Y \). The term \( \frac{[X] - [X]_{eq}}{[X]_{eq}} \) now is a small parameter if we are close to equilibrium. Thus, the logs are now of the form \( \log(1 + x) \), where \( x \) is small. Taylor expansion for small \( x \) gives \( \log(1 + x) \approx x \)

Thus,
\[ \dot{\xi} = k_{XY} \cdot [Y]_{eq} \left( \frac{[X]}{[X]_{eq}} - \frac{[Y]}{[Y]_{eq}} \right) = \frac{k_{XY}}{R} \cdot \frac{\mu_X - \mu_Y}{T} = L \left( \frac{A}{T} \right). \]

This is an example of a flux-force term. \( L \) is the flux and \( A/T \) is the force. Therefore, we have established the following equivalence close to equilibrium:

8.2.1. Entropy Production. For the previous reaction (consider the system to be isolated), at constant \( u \) the Gibbs equation is
\[ \dot{s} = - \sum_i \frac{\mu_i}{T} \frac{d[X_i]}{dt} = - \frac{\mu_X}{T} \frac{d[X]}{dt} - \frac{\mu_Y}{T} \frac{d[Y]}{dt} = - \frac{d[X]}{dt} \left( \frac{\mu_X}{T} - \frac{\mu_Y}{T} \right) = \dot{\xi} \left( \frac{A}{T} \right), \]

where \( A \) is called the affinity. Recall that \( A = -\Delta G_{rxn} \), where \( \Delta G_{rxn} \equiv \left( \frac{\partial G}{\partial \xi} \right)_{T,P} = \sum_i \nu_i \mu_i \), with the convention that \( \nu_i \) is negative for reactants and positive for products. Thus, differences in chemical potential drive entropy production. Since entropy production describes the evolution of the reaction toward equilibrium, affinity is a driving force of the chemical reaction.
8.2.2. Example with Three Components. Consider the reaction:

\[ X + Y \rightleftharpoons Z. \]

Let’s check that the affinity is:

\[ \mathcal{A} = \frac{\mu_X + \mu_Y - \mu_Z}{T}. \]

The flow of reaction from left to right is:

\[ \dot{\xi} = -\frac{d[X]}{dt} = -\frac{d[Y]}{dt} = +\frac{d[Z]}{dt} \]

so the equation for \( \dot{s} \) is

\[ \dot{s} = -\frac{\mu_X}{T} \frac{d[X]}{dt} - \frac{\mu_Y}{T} \frac{d[Y]}{dt} - \frac{\mu_Z}{T} \frac{d[Z]}{dt} = -\frac{d[X]}{dt} \left( \frac{\mu_X}{T} + \frac{\mu_Y}{T} - \frac{\mu_Z}{T} \right) = \dot{\xi} \left( \frac{\mathcal{A}}{T} \right), \]

where \( \mathcal{A}/T = (\frac{\mu_X}{T} + \frac{\mu_Y}{T} - \frac{\mu_Z}{T}) \).

8.3. External vs Internal Variables

As mentioned previously, the change in the entropy of a system, \( dS \), can be decomposed into two terms,

\[ dS = d_eS + d_iS, \]

where \( d_eS \) is the entropy supplied to the system by its surroundings (\( e=\text{external} \)), and \( d_iS \) is the entropy produced inside the system (\( i=\text{internal} \)). The second law states that \( d_iS = 0 \) for a reversible process and positive for irreversible transformations of the system, i.e.

\[ d_iS \geq 0. \]

The entropy supplied, \( d_eS \), however, may be positive, zero or negative, depending on the interaction of the system with its surroundings. Thus, for an adiabatically insulated system (i.e. a system which can exchange neither heat nor matter with its surroundings), \( d_eS \) is equal to zero, and it follows that

\[ dS \geq 0. \] (for an adiabatically insulated system)

For a so-called closed system, which may only exchange heat with its surroundings, we have

\[ d_eS = \frac{\delta Q}{T}, \quad \text{(Carnot’s theorem)} \]

where \( \delta Q \) is the heat supplied to the system by its surroundings and \( T \) is the absolute temperature at which heat is received by the system. It then follows that

\[ dS = d_eS + d_iS = \frac{\delta Q}{T} + d_iS \geq \frac{\delta Q}{T} \] (for a closed system)
For open systems (heat+matter exchange allowed w/surroundings), \(d_e S\) contains also a term connected with the transfer of matter. The theorem of Carnot-Clausius, \(d_e S = \frac{\delta Q}{T}\), does not apply to such systems (we will see later how to correct it). However, \(dS = d_e S + d_i S\) and \(d_i S \geq 0\) remain valid.

### 8.3.1. General Form of Entropy Production.

Previously, we used the first law (fundamental equation for \(dS\)) to write down an equation for the time-derivative of the entropy:

\[
\dot{s} = \frac{ds}{dt} = \frac{1}{T} \frac{du}{dt} - \sum_i \frac{\mu_i}{T} \frac{d[X_i]}{dt}.
\]

The first term, \(\frac{1}{T} \frac{du}{dt}\), is associated with heat transfer (can be positive or negative). The last term can be rewritten as \(\dot{\xi}(A/T)\) or \(\dot{\xi}(\Delta G_{rxn}/T)\), using \(\dot{\xi} = \frac{1}{T} \frac{d[X_i]}{dt}\), since \(A = -\sum \mu_i \nu_i\) and \(A = -\Delta G_{rxn}\). This term is also equal to \(\dot{\xi}(-\Delta G_{rxn}/T)\). If a reaction proceeds from left to right, \(\Delta G_{rxn} < 0\), and \(\dot{\xi} > 0\). And conversely, if the reaction proceeds from right to left, \(\Delta G_{rxn} > 0\), \(\dot{\xi} < 0\). At equilibrium, \(\Delta G_{rxn} = 0\). Thus, the term can never be negative. This nonnegative (\(\geq 0\)) term, is called entropy production, and describes an irreversible process. We obtain a restatement of the Clausius theorem,

\[
\dot{s} \geq \frac{1}{T} \frac{du}{dt}.
\]

We have seen in the specific case of a reaction \(X \rightleftharpoons Y\) that the reaction term is equal to\(^2\)

\[
\dot{\xi}(A/T) = -\frac{d[X]}{dt} \left(\frac{\mu_X - \mu_Y}{T}\right).
\]

We have also found that for that particular reaction \((X \rightleftharpoons Y)\),

\[
\dot{\xi} = \frac{k_{YX}[Y]_{eq}}{R} \left(\frac{\mu_X - \mu_Y}{T}\right) = L \left(\frac{A}{T}\right)
\]

where \(L = k_{YX}[Y]_{eq}/R = k_{XY}[X]_{eq}/R\). Thus, \(\dot{s}\) has the form

\[
\dot{s} = L \left(\frac{A}{T}\right) = LXX,
\]

where \(X = \left(\frac{A}{T}\right)^2\) is a force. \(LX\) is a flux. The flux, often denoted \(J = LX\), is proportional to the force (no forces, no fluxes).

\(^2\)The form \(\dot{\xi}(A/T)\) of the chemical term is generally true for any reaction. \(A/T\) is a driving force for the chemical reaction (i.e. differences in chemical potential lead to transformation, whereas equilibrium is reached when the chemical potentials are equal). \(\xi\) is a flux.
This turns out to be a general principle. In the general case there may be several forces (and fluxes) acting on the system. Then, \( \dot{s} \) will be of the form:

\[
\dot{s} = \sum_{i=1}^{N} \sum_{j=1}^{N} L_{ij} X_i X_j,
\]

where the \( L_{ij} \) are called \textit{Onsager coefficients} and the \( X_i \)'s are the thermodynamic forces. The quantities

\[
J_j = L_{1j} X_1 + L_{2j} X_2 + L_{3j} X_3 + \cdots + L_{Nj} X_N
\]

are called \textit{fluxes}. We can think of the fluxes as the leading terms of a Taylor expansion of \( J_j \) as a power series in \( X_1, X_2, \ldots, X_N \). Close to equilibrium, only the first order term is needed. Also, there are no constant terms because the fluxes all vanish if all the forces vanish. The Onsager coefficients are reciprocal, \( L_{ij} = L_{ji} \).

Some examples of commonly encountered fluxes and forces can be found in Table 8.1.

<table>
<thead>
<tr>
<th>Flux, J</th>
<th>Force, X</th>
</tr>
</thead>
<tbody>
<tr>
<td>Heat</td>
<td>( J_u ) = \frac{d(1/T)}{dx} = -\frac{1}{T^2} \frac{dT}{dx}</td>
</tr>
<tr>
<td>Matter</td>
<td>( J_i ) = -T \frac{d(\mu_i/T)}{dx} (voltage/T)</td>
</tr>
<tr>
<td>Electric current density</td>
<td>( J_I = I/A ) = -\frac{1}{T} \frac{d\phi}{dx}</td>
</tr>
<tr>
<td>Chemical reaction</td>
<td>( \dot{\xi} ) = -\sum_i \nu_i \mu_i</td>
</tr>
</tbody>
</table>

**8.4. Particle Flux and Heat Flux**

So far we have looked at chemical reactions localized in the same volume element. These chemical reactions are driven by differences in the chemical potential. Now let’s look at spatially dependent processes. For simplicity, we consider at processes that vary along a single dimension (x). The system is subdivided into slices of thickness \( \Delta x = x_2 - x_1 \), as shown in Fig. 8.1.

\( \Delta x \) is small enough so that \( T, P, \mu \) are constant over the volume element and so that local thermodynamic equilibrium can be assumed.

Recall (see section 6.10.19) that in the case of two volume elements A and B adjacent to each other (setting \( dV = 0 \)), we found that

\[
dS = (T_A^{-1} - T_B^{-1}) dU_A - \left( \frac{\mu_A}{T_A} - \frac{\mu_B}{T_B} \right) dn_A,
\]
where \( \mu_{A/B} \) are in molar units. (This equation was obtained by applying the fundamental equation of thermodynamics in the entropy representation.)

Now take \( A \) and \( B \) to be thin slices centered around \( x_1 \) and \( x_2 \) in the above diagram. Dividing by \( dt \) gives the equation for entropy production:

\[
\dot{S} = \frac{dU_1}{dt} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) + \frac{dn_1}{dt} \left( \frac{\mu_2}{T_2} - \frac{\mu_1}{T_1} \right) \geq 0,
\]

which can be written in the form,

\[
\dot{S} = J_U X_U + J_n X_n,
\]

where

\[
J_U = \frac{dU_1}{dt} = -\frac{dU_2}{dt}, \quad J_n = \frac{dn_1}{dt} = -\frac{dn_2}{dt}, \quad X_U = \frac{1}{T_1} - \frac{1}{T_2}, \quad X_n = \frac{\mu_2}{T_2} - \frac{\mu_1}{T_1}.
\]

\( X_U \) is the thermal force and \( X_n \) is the chemical force. Writing \( \Delta(1/T) = 1/T_2 - 1/T_1 \), \( \Delta(\mu/T) = \mu_2/T_2 - \mu_1/T_1 \), and dividing by the volume \( V = A \Delta x \) gives:

\[
\dot{s} = -\frac{1}{A} \frac{dU_1}{dt} \frac{\Delta(1/T)}{\Delta x} + \frac{1}{A} \frac{dn_1}{dt} \frac{\Delta(\mu/T)}{\Delta x}.
\]

Taking the limit of small \( \Delta x \) gives:

\[
\dot{s} = -\frac{1}{A} \frac{dU_1}{dx} \frac{d(1/T)}{dx} + \frac{1}{A} \frac{dn_1}{dt} \frac{d(-\mu/T)}{dx},
\]

which can be written as

\[ 8.4 \]

\[
\dot{s} = J_U \frac{d(1/T)}{dx} + J_n \frac{d(-\mu/T)}{dx},
\]

where the derivatives are called gradients of \( 1/T \) and \( -\mu/T \), respectively, and

\[
J_U = -\frac{1}{A} \frac{dU_1}{dt}, \quad J_n = -\frac{1}{A} \frac{dn_1}{dt}
\]

are fluxes.
8.5. Conservation of Mass

In the case of diffusion and reaction taking place simultaneously, the flow of the reaction from reactant to product is:

\[
\frac{d}{dt}[X_j] = -\frac{dJ_j}{dx} + \nu_j \dot{\xi}, \quad j = 1, \ldots, n
\]

where \([X_j] = n_j/V, J_j\) are the component fluxes \((n: \text{number of components})\), all directed along the \(x\) axis, i.e. \(\nu_j\) are the stoichiometric constants in a chemical reaction, and \(\dot{\xi}\) is its rate in the volume element. This is merely the addition of two terms: the component flux term \(-dJ_j/dx\) due to mass transport, plus the chemical reaction term which we know from chemical kinetics, by writing down the rate law, \(\xi = (1/\nu_j)d[X_j]/dt\), for the reaction.

Our convention was to take \(\nu_j\) as negative for the reactants and positive for the products. Since \(\nu_j\) are dimensionless, \(\dot{\xi}\) has units of mol/m\(^3\)/s.

This is the total rate of change, which is the sum of two terms. The second term \((\dot{\xi})\) we have seen already. But where does the first term come from?

Consider a volume \(V = A\cdot dx\) of length \(dx\) and transport along the \(x\) direction. The volume has cross-sectional area \(A\) and length beginning at \(x\) and ending at \(x+dx\). The change in the number of moles of a component, \(n_j\), in a small volume \(V\), is due to the difference between the flux of the components in and out of the volume. We have

\[
\frac{dn_j}{dt} = -A \left[ J_j(x + dx) - J_j(x) \right]
\]

where \(A\) is the cross sectional area of the volume orthogonal to the flux direction. Positive fluxes \(J_j\) indicate particles flowing toward positive \(x\).

The flux \(-J_j(x + dx)\) denotes a decrease in \(n_j\), as particles leave the volume \(dx\) at \(x+dx\). \(+J_j(x)\) denotes particles entering \(dx\) at \(x\).

The fluxes have units of moles per area per unit time. The area is equal to the volume divided by \(dx\). In the limit of small \(dx\), we therefore have

\[
\frac{dn_j}{dt} = -V \left[ J_j(x + dx) - J_j(x) \right] dx = -V \frac{dJ_j(x)}{dx}.
\]

By dividing this equation left and right by the (constant) volume, one obtains the desired result for \(\dot{\xi} = 0\)

\[
\frac{d[X_j]}{dt} = -\frac{dJ_j(x)}{dx}
\]

where \([X_j] = n_j/V\). This equation describes the conservation of moles (mass) within the volume.
The total rate of change, $d_{\text{tot}}[X_j]/dt$, is then formed by adding this mass transport term, $-\frac{dJ_j}{dx}$, to the chemical reaction term, $\nu_j \dot{\xi}$.

The mass transport term, $-\frac{dJ_j}{dx}$, is called a surface term because it describes particle flux entering and leaving the volume element (by crossing its surface). The chemical reaction term, $\nu_j \dot{\xi}$ is called a volume term because the chemical reaction takes place everywhere inside the volume (with zero influence from the surroundings).

### 8.6. Conservation of Energy

Similarly, we have a conservation law

$$\frac{du}{dt} = -\frac{d}{dx}J_u.$$ 

Here $u = U/V$ is the internal energy density, $J_u$ is the energy flux (e.g. heat flow). Here, there is no “volume term” because energy is not created within the volume unless a “heat source” is added. This equation expresses the fact that energy can only be transferred, never created.

On the other hand, if there is a “heat source” is placed within the volume, then a volumetric term is added:

$$\frac{du}{dt} = -\frac{d}{dx}J_u + \dot{q}_V,$$

where $\dot{q}_V$ is a volumetric heat source (units: W/m$^3$).

### 8.7. Entropy Balance

The entropy balance equation is

$$\frac{dS}{dt} = -A [J_s(x + dx) - J_s(x)] + V\sigma$$

where $\sigma$ is the entropy production per unit volume. Since $A = V/dx$, we have, in the limit of small $dx$

$$\frac{dS}{dt} = -V \frac{[J_s(x + dx) - J_s(x)]}{dx} + V\sigma = -V \frac{dJ_s(x)}{dx} + V\sigma.$$

Dividing by volume $V$:

$$\frac{ds}{dt} = -\frac{dJ_s(x)}{dx} + \sigma.$$ (8.5)

Substituting $\frac{d[X_j]}{dt} = -\frac{d}{dx}J_j + \nu_j \dot{\xi}$ and $\frac{du}{dt} = -\frac{d}{dx}J_u$ into

$$\dot{s} \equiv \frac{ds}{dt} = \frac{1}{T} \frac{du}{dt} - \sum_i \frac{\mu_i}{T} \frac{d[X_i]}{dt},$$
and comparison with Eq. (8.5) enables us to identify \( J_s \) and \( \sigma \):

\[
\frac{ds}{dt} = \frac{1}{T} \frac{du}{dt} - \frac{1}{T} \sum_j \mu_j \frac{d[X_j]}{dt} = \frac{1}{T} \left( -\frac{dJ_u}{dx} \right) - \frac{1}{T} \sum_j \mu_j \left( -\frac{d}{dx} J_j + \nu_j \dot{\xi} \right)
\]

\[
= -\frac{d}{dx} \left[ \frac{1}{T} \left( J_u - \sum_{j=1}^n \mu_j J_j \right) \right] + J_u \frac{d}{dx} \left( \frac{1}{T} \right) + \sum_{j=1}^n J_j \frac{d}{dx} \left( -\frac{\mu_j}{T} \right) + \dot{\xi} \left( -\frac{\Delta G_{rxn}}{T} \right)
\]

where \( \Delta G_{rxn} = \sum_i \nu_i \mu_i \), with the convention that \( \nu_i \) is negative for reactants and positive for products. In the second line we have twice used the product rule for differentiation: \( d(fg)/dx = f \cdot \frac{dg}{dx} + g \cdot \frac{df}{dx} \).

The surface term, \(-dJ_s/dx\), is given in terms of the entropy flux,

\[
J_s = \left[ \frac{1}{T} \left( J_u - \sum_{j=1}^n \mu_j J_j \right) \right].
\]

It describes two phenomena: the measurable heat flux (“reduced” heat flow \( J_u/T \)) and the enthalpy of matter transfer (heating due to flows of matter). The enthalpy flux is carried by the component fluxes, \( J_j \). The entropy flux term is a “surface term” which describes changes in entropy of the system due to inputs from the surroundings.

The volume term, which is also called “entropy production”,

\[
\sigma = \frac{J_u}{T} \frac{d}{dx} \left( \frac{1}{T} \right) + \sum_{j=1}^n J_j \frac{d}{dx} \left( -\frac{\mu_j}{T} \right) + \dot{\xi} \left( -\frac{\Delta G_{rxn}}{T} \right)
\]

has 3 contributions. The first term arises from heat conduction (temperature gradients), since \( d(1/T)/dx = (-1/T^2)dT/dx \). The second term is connected to diffusion, which is driven by gradients in the chemical potential; here, \((d/dx)(\mu_j/T)\). The third term is due to chemical reactions.

The entropy flux term describes external contributions from the surroundings. It can be positive, negative or zero. This is the \( deS \) term. The entropy production term is non-negative, \( \sigma \geq 0 \). Entropy production is the \( diS \) term.

8.8. Entropy Production vs External Entropy Flow

We recall that \( dS \geq 0 \) for an isolated system \( S \). If the system is not isolated, we have the decomposition \( dS = deS + diS \). Some books write \( dS_{exch} \) in lieu of \( deS \) and \( dS_{prod} = diS \). Thus, \( dS \) consists of two parts. One part is the entropy created in the system by any spontaneous processes occurring within it, and the other part is the change in entropy resulting from the
If the exchange is carried out reversibly, \( \delta Q = \delta Q^\text{rev} \); if it is irreversible, \( \delta Q = \delta Q^\text{irr} \). For any process (reversible or irreversible),
\[
dS = d_iS + d_eS = d_iS + \frac{\delta Q}{T}.
\]
For a reversible process, \( d_iS = 0 \) and \( \delta Q = \delta Q^\text{rev} \), so that we have \( dS = (\delta Q^\text{rev}/T) \). For an irreversible process, \( d_i > 0 \) and \( \delta Q = \delta Q^\text{irr} \), so that \( dS > (\delta Q^\text{irr}/T) \). Combining these two relations together we obtain the Clausius statement of the second law, \( dS \geq (\delta Q/T) \).

### 8.8.1. Fluxes Depend Linearly on Forces.
Consider two systems, 1 and 2 in contact with each other. Particles and heat can be exchanged between them. The two combined systems are isolated from their surroundings, thus explaining the term \( d_eS = 0 \). We are left with the entropy production, \( d_iS \), as the only contribution to the entropy:
\[
\frac{d_iS}{dt} = \frac{dU_1}{dt} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) + \frac{dn_1}{dt} \left( \frac{\mu_2}{T_2} - \frac{\mu_1}{T_1} \right) = J_U X_U + J_n X_n \geq 0.
\]
where \( X_U = 1/T_1 - 1/T_2 \), \( X_n = \mu_2/T_2 - \mu_1/T_1 \), \( J_U = dU_1/dt = -dU_2/dt \) and \( J_n = dn_1/dt = -dn_2/dt \). The fluxes and forces are related to each other. Experimentally, the relation is often linear. For example, Fourier’s law of heat conduction says that the flux of energy as heat is directly proportional to the temperature difference,
\[
J_U \propto T_2 - T_1.
\]
But \( T_2 - T_1 \propto (T_2 - T_1)/T_1T_2 = 1/T_1 - 1/T_2 = X_U \), so we write \( J_U \propto 1/T_1 - 1/T_2 = X_U \). Similarly, Fick’s law of (isothermal) diffusion says that the flux of matter is directly proportional to the concentration difference,
\[
J_n \propto c_2 - c_1
\]
which can be expressed\(^4\) in the form \( J_n \propto \mu_2 - \mu_1 \) (constant \( T \)). Finally, we have Ohm’s law, \( V = RI \), which states that the electrical current, \( I \), is proportional to the potential difference, \( V = \phi_2 - \phi_1 \). Thus, in all cases, \( J \propto X_i \) (fluxes linearly depend on the forces).

---

\(^3\)The right hand side was explained previously. It is obtained from the fundamental equation in the entropy representation.

\(^4\)The proof is similar to our previous discussion. \( c_1 \) and \( c_2 \) are different concentrations of the same substance. Write \( \mu_1 = \mu^c + RT \log(c_1/c^o) \) and \( \mu_2 = \mu^c + RT \log(c_2/c^o) \), so that \( \mu_2 - \mu_1 = RT \log(c_2/c^o - \log(c_1/c^o)) \). Write \( c_2/c^o = [(c_2 - c^o) + c^o]/c^o = (c_2 - c^o)/c^o + 1 \) and similarly for \( c_1 \), and Taylor expand in the small parameter \( (c_2 - c^o)/c^o \). Then, \( \mu_2 - \mu_1 \approx \frac{RT}{c^o}(c_2 - c_1) \).
8.8. From Discrete to Continuous Case. In the continuum case, it is gradients of thermodynamic quantities that determine the forces (not differences). We have seen that \[ \frac{dS}{dt} = \frac{dU_1}{dt} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) + \frac{dn_1}{dt} \left( \frac{\mu_2}{T_2} - \frac{\mu_1}{T_1} \right) \geq 0. \]

We take the continuum limit by dividing by \( V = A \cdot \Delta x \) and taking the limit \( \Delta x \to 0 \):

\[
\sigma \equiv \frac{1}{V} \frac{dS}{dt} = \left( -\frac{1}{A} \frac{dU_1}{dt} \right) \frac{\Delta(1/T)}{\Delta x} + \frac{1}{A} \left( \frac{dn_1}{dt} \right) \frac{\Delta(\mu/T)}{\Delta x},
\]

where \( \Delta(1/T) = 1/T_2 - 1/T_1, \Delta(\mu/T) = \mu_2/T_2 - \mu_1/T_1 \). Thus,

\[
\sigma = J_U \frac{d(1/T)}{dx} + J_n \frac{d(-\mu/T)}{dx} \geq 0,
\]

where \( J_U = -\frac{1}{A} \frac{dU_1}{dt} \) and \( J_n = -\frac{1}{A} \frac{dn_1}{dt} \).

8.8.3. Application to Isothermal Diffusion. As an example of the application to continuous systems, we will consider diffusion in an isothermal system (isothermal diffusion). Let \( T \) be constant, so it can be factored out of each term. The entropy production is:

\[
\sigma = \frac{1}{T} \sum_{i=1}^{n} J_i \left( -\frac{d\mu_i}{dx} \right)_{P,T},
\]

where \( n \) is the number of components in the system. The chemical potentials are not all independent of each other because of the Gibbs-Duhem equation (\( \sum_{i=1}^{n} n_i d\mu_i + S dT - V dP = 0 \)), which for constant \( T, P \) reads, \( \sum_{i=1}^{n} n_i d\mu_i = 0 \). Dividing by \( V \) and \( dx \) gives, \( \sum_{i=1}^{n} c_i \left( -\frac{d\mu_i}{dx} \right) = 0 \). Solving for \( d\mu_i/dx \),

\[
\frac{d\mu_i}{dx} = -\frac{1}{c_1} \sum_{i=2}^{n} c_i \left( \frac{d\mu_i}{dx} \right)_{P,T}.
\]

Then,

\[
T \sigma = J_1 \left( -\frac{d\mu_1}{dx} \right) + \sum_{i=2}^{n} J_i \left( -\frac{d\mu_i}{dx} \right) = J_1 \sum_{i=2}^{n} c_i \left( \frac{d\mu_i}{dx} \right) + \sum_{i=2}^{n} J_i \left( -\frac{d\mu_i}{dx} \right)
\]

\[
= \sum_{i=2}^{n} \left( J_i - \frac{c_i}{c_1} J_1 \right) \left( -\frac{d\mu_i}{dx} \right),
\]

where \( J_i^d = \left( J_i - \frac{c_i}{c_1} J_1 \right) = c_i \left( \frac{J_i}{c_i} - \frac{J_1}{c_1} \right) = c(v_i - v_1) \), since particle flux equals concentration times velocity \( J_i = c_i v_i \). (If \( v_i = v_1 \), then entropy production is zero.) Thus, \( J_i^d \) is the flux of solute \( i \) relative to solvent (1).

8.8.4. Example of Binary Solution. Let’s consider a binary solution, such as sucrose in water or sodium chloride in water. In such cases, there is
only 1 independent flow:

\[ T\sigma = J^d_2 \left( -\frac{d\mu^2}{dx} \right)_{T,P}. \]

The force is \( X = \frac{d\mu^2}{dx} \). There is just one flux force equation \( J = LX \):

\[ J^d_2 = L_{22} \left( -\frac{d\mu^2}{dx} \right)_{T,P} = -L_{22} \left( -\frac{d\mu^2}{dx} \right). \]

Using \( \mu_2 = \mu_2(T,P) + RT \log c_2 \) (\( c_2 \) in units of mol/L). Then,

\[ \frac{d\mu_2}{dx} = \frac{d\mu_2}{dc_2} \cdot \frac{dc_2}{dx} = \frac{RT}{c_2} \left( \frac{dc_2}{dx} \right). \quad \text{(chain rule)} \]

So that

\[ J^d_2 = -L_{22} \frac{RT}{c_2} \frac{dc_2}{dx}. \]

We get Fick’s law of diffusion, which gives the concentration flux as \( J_c = -D \frac{dc}{dx} \), where \( D = L_{22}RT/c \).

From the conservation of mass,

\[ \frac{dc}{dt} = -\frac{1}{A} J_c, \]

where \( J_c \) is the concentration flux (particle flux, \( J_n \), divided by \( V \)), substitution of \( J_c = -D \frac{dc}{dx} \) gives the diffusion equation

\[ \frac{dc}{dt} = D \frac{d^2c}{dx^2}, \]

which we derived in Section 1.5.1 using Einstein’s method.

### 8.9. Diffusion Leads to Entropy Production

Diffusion is an irreversible process. This is obvious from experiments: releasing a gas in the corner of the room leads to its spreading across the room until the concentration is uniform. The gas will never return to its original position. This can be seen also from the diffusion equation, which is not time-reversal invariant: the transformation \( t \to -t \) leads to a sign change. That diffusion leads to entropy production can be seen by computing the flow of entropy.

Recall from Section 1.5.1 that the solution to the diffusion equation for initial conditions, \( c(x,0) = n\delta(x) \), is:

\[ c(x,t) = \frac{n}{\sqrt{4\pi Dt}} e^{-x^2/4Dt}. \]
8.9. Diffusion Leads to Entropy Production

8.9.1. Entropy Production and the Gibbs-Shannon Entropy. It can be shown that Boltzmann entropy, \( S = k_B \log \Omega \) is a special case of the more general Gibbs-Shannon entropy: for a discrete probability distribution \( p \) on the countable set \( \{x_1, x_2, \ldots \} \), with \( p_i = p(x_i) \), the Gibbs-Shannon entropy is defined as

\[
S = -k_B \sum_{i \geq 1} p_i \log p_i.
\]

Indeed, taking \( p_i = 1/\Omega \) (the uniform distribution, \( \sum_{i=1}^{\Omega} 1/\Omega = 1 \)), we get:

\[
S = k_B \sum_{i=1}^{\Omega} 1/\Omega \log \Omega = k_B \log \Omega.
\]

For a continuous probability density function \( p(x) \), the Shannon entropy is

\[
S = -k_B \int_{\mathbb{R}} p(x) \log p(x) dx.
\]

The entropy of the Gaussian density is seen to be:

\[
S = -k_B \int_{\mathbb{R}} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-x^2/2\sigma^2} \left( -\log(\sqrt{2\pi\sigma^2}) - \frac{x^2}{2\sigma^2} \right) dx = \frac{k_B}{2} (1 + \log(2\pi\sigma^2))
\]

where \( \sigma^2 = 2Dt \) for the solution to the diffusion equation. Differentiation of \( S \) with respect to \( t \) gives:

\[
\frac{dS}{dt} = \frac{d}{dt} \left( \frac{k_B}{2} (1 + \log(4\piDt)) \right) = \frac{k_B}{2} \frac{1}{4\pi D t} \cdot 4\pi D = \frac{k_B}{2t} > 0,
\]

which is positive for all times \( t > 0 \), meaning that diffusion leads to entropy production. Note that \( dS/dt \) decreases with time as the steady state is approached, eventually reaching 0 asymptotically.

Now, a slightly more general proof. We can still show that \( dS/dt > 0 \) without assuming a Gaussian distribution, but instead only invoking the form of the diffusion equation (and not its solution). Take the time derivative of \( S \), then substitute the right hand side of the diffusion equation, \( \frac{d\rho}{dt} = D \frac{d^2\rho}{dx^2} \), for each instance of \( d\rho/dt \):

\[
\frac{dS}{dt} = -k_B \int \left[ \frac{d\rho}{dt} \log \rho(x) + \rho(x) \frac{1}{\rho(x)} \frac{d\rho}{dt} \right] dx
\]

\[
= -k_B D \int \left[ \frac{d^2\rho}{dx^2} \log \rho(x) + \frac{d^2\rho}{dx^2} \right] dx.
\]

Integration by parts eliminates the second term. Integration by parts of the first term leads to:

\[
\frac{dS}{dt} = -k_B D \int \frac{d\rho}{dx} (-1) \frac{1}{\rho} \frac{d\rho}{dx} dx,
\]
which is a non-negative quantity:

\[
\frac{dS}{dt} = k_B D \int \left( \frac{d\rho}{dx} \right)^2 \frac{1}{\rho} \, dx \geq 0,
\]

because the integrand is the product of two positive quantities: \((d\rho/dx)^2 \geq 0\) and \(\rho^{-1} > 0\).

### 8.10. Heat Conduction Leads to Entropy Production

We have seen that the heat conduction term

\[
\sigma = J_u \frac{d}{dx} \left( \frac{1}{T} \right) = -J_u \cdot \frac{1}{T^2} \frac{dT}{dx}
\]

leads to entropy production. From this equation, we see that the force \(X = dT/dx\). There is a flux-force relation, \(J = LX\), of the form:

\[
J_u = -L_{qq} \frac{dT}{dx},
\]

where \(L_{qq}\) is a “constant” (Onsager coefficient; the subscript \(qq\) refers to heat), which we will call \(\alpha\) in a moment. Starting from the equation for energy conservation (with no heat source term),

\[
\frac{du}{dt} = -\frac{d}{dx} J_u.
\]

Inserting the flux-force relation, invoking \(du = d(U/V) = \delta Q/V\) (\(dV = 0\), so no work, change in internal energy is due to heat transfer) and \(\delta Q/V = \rho c_V dT\) (\(\rho\): density, \(c_V\): specific heat capacity, \(C/\rho V\)), we obtain the famous Fourier’s heat conduction equation

\[
\frac{dT}{dt} = \alpha \frac{d^2 T}{dx^2}
\]

where \(\alpha = k/\rho c_V\) is the thermal diffusivity of the medium where heat flows and \(k\) is the thermal conductivity (units: W/m.K). It is identical to the diffusion equation. Therefore, heat conduction also leads to entropy production (as the diffusion equation does).

### 8.11. Summary: Onsager’s Formulation of the Second Law

The Second Law can be formulated in terms of the entropy production \(\sigma\), namely, \(\sigma \geq 0\). In Onsager’s formulation, the entropy production is given by the product sum of so-called conjugate fluxes, \(J_i\), and forces \(X_i\), in the system. The second law becomes

\[
\sigma = \sum_i J_i X_i \geq 0
\]
where $\sigma$ is larger than or equal to zero. Each flux is taken to be a linear combination of all forces,

$$J_i = \sum_j L_{ij} X_j$$

and the reciprocal relations

$$L_{ij} = L_{ji}$$

apply. They are called Onsager’s reciprocal relations. In order to use the theory, one first has to identify a complete set of extensive independent variables, $\alpha_i$. The resulting conjugate fluxes and forces are

$$J_i = \frac{d\alpha_i}{dt}, \quad \text{and} \quad X_i = \left( \frac{\partial S}{\partial \alpha_i} \right)_{\alpha_j \neq i},$$

respectively. Here $t$ is the time and $S$ is the entropy of the system. The three equations above contain then all the information on the behavior of the system. Accurate expressions for the fluxes are required to model real systems. The simplest descriptions of heat, mass, charge, and volume transport are the equations of Fourier, Fick, Ohm, Darcy and Newton. Fourier’s law expresses the measurable heat flux in terms of the temperature gradient by:

$$J_q' = -\lambda \frac{dT}{dx}$$

where $\lambda$ is the thermal conductivity, $T$ is the absolute temperature, and $x$ is the direction of transport. Fick’s law gives the mass flux of one of the components in terms of the gradient of its molar concentration $c$:

$$J = -D \frac{dc}{dx}$$

where $D$ is the diffusion coefficient. Ohm’s law gives the electric current in terms of the gradient of the electric potential:

$$J_I = -\kappa \frac{d\phi}{dx}$$

where $\kappa$ is the electrical conductivity, and $\phi$ is the electric potential. Darcy’s law says that the volume flow $J_v$ in a tube is proportional to the pressure gradient $dP/dx$ via the coefficient $L_p$:

$$J_v = -L_p \frac{dP}{dx}.$$

The fluxes are all caused by a gradient, or driving force. Fick’s law, for instance, says that there is no mass flux if there is no concentration gradient. We know from experiments that a temperature gradient and an electric potential gradient also can give rise to a mass flux. To neglect such effects can have severe consequences. In general situations, we must take all driving forces into account. Equations (8.7) and (8.8) form the second law of
thermodynamics

\[
\sigma \equiv \frac{dS}{dt} = \sum_i J_i X_i = \sum_i \sum_j L_{ij} X_j X_i \geq 0.
\]

In a stationary state there is no accumulation of internal energy, mass or charge. This means that the heat, molar and electric fluxes are independent of position. The derivative of the above equations with respect to \(x\) are then zero. For the first equation, we have:

\[-\frac{d}{dx} J'_a = \frac{d}{dx} \lambda \frac{dT}{dx} = 0.\]

Equations like these can be used to calculate the temperature, concentration, electric potential and pressure as a function of the position, when their values on the boundaries of the system and \(\lambda, D, \kappa, L_p\) and \(\eta\) are known.

8.12. Problems

**Problem 124.** Suppose that we have a metal rod whose two end points are held at different temperatures. This is a non-equilibrium situation because thermodynamic equilibrium would imply that the entire rod is at the same temperature. Instead, the temperature is a function of position along the length of the rod. It is a steady state situation because the temperature does not vary with time (i.e. each point along the rod is at a fixed temperature). Use the concepts of non-equilibrium thermodynamics (namely, the heat conduction equation) to calculate the temperature as a function of position between the two end points of the rod. The rod is 10 cm long. The walls are kept at constant temperature, 4 and 25 °C, respectively. Assume that the thermal conductivity is constant.

**Solution.** The heat conduction equation is Eq. 8.6. In the steady state, we have \(d^2T/dx^2 = 0\). Thus, \(T\) is of the form, \(T(x) = a + bx\). The constants \(a\) and \(b\) follow from the boundary conditions. We have \(T(0) = 278\) K and \(T(10) = 298\) K. It follows that \(T(x) = (278 + 2x/\text{cm})\) K.

**Problem 125.** Indoor ice rinks employ a cooling system placed underneath a concrete slab (similar to radiant floor cooling systems, but lower temperatures). On top of the concrete slab is the ice. Suppose that we place a temperature sensor \(d = 8\) cm under the concrete slab. Calculate the entropy production per surface area due to the heat transport through the concrete slab. The temperature sensor reads 243 K. This is the temperature of the radiant cooling system. The concrete is in contact with ice (273 K). The thermal conductivity of the pavement is 0.7 W/K.m (\(\lambda\) parameter in the Fourier’s law). Note: this is a 1D problem with depth as the relevant spatial variable.
Solution. Fourier’s law for heat conduction is \( J_u = -\lambda (dT/dx) \), where \( x \) is the depth coordinate. The entropy production per surface area is (let \( \Delta T = T_{top} - T_{below} \)):

\[
\left| \int_0^d \sigma \, dx \right| = \left| \int_0^d J_u \frac{d}{dx} \left( \frac{1}{T} \right) \, dx \right| = \left| \lambda \frac{\Delta T}{d} \left( \frac{1}{T_{top}} - \frac{1}{T_{below}} \right) \right| \\
= \left| 0.7 \, (30) \, \left( \frac{1}{273} - \frac{1}{243} \right) \right| = 0.12 \, \text{W/K.m}^2.
\]

Problem 126. For the previous problem, calculate the lost work per surface area near the top of the concrete slab (at the interface of ice and concrete). Lost work is the work lost due to entropy production (a type of \( T\Delta S \) term, where \( \Delta S \) is due to entropy production).

Solution. The lost work (\( A \): area)

\[
\Delta W_{\text{lost}} = T_{top} \cdot A \int_0^d \sigma \, dx
\]

per surface area \( A \) is \( \Delta W_{\text{lost}} / A = 273 \, \text{K} \cdot 0.12 \, \text{W/K.m}^2 = 33 \, \text{W/m}^2 \). It is typical for heat conduction around room temperature that the entropy production is large.

Problem 127. Suppose that a closed 1D system (ideal gas in a pipe) does not undergo any irreversible changes (\( \sigma = 0 \)) and its temperature is spatially uniform [\( d(1/T)/dx = 0 \)] and equal to 300 K everywhere, but that there is an enthalpy flux driven by particle flow. Moreover, the entropy of the system does not change \( ds/dt = 0 \) (isentropic process). There is flow in the pipe which we measure to have an average local velocity of 10 m/s at \( x = 0 \) and 12 m/s at \( x = 1 \) m. (Such an acceleration is possible if the pipe diameter narrows slightly; you can assume that the velocity increases uniformly from 10 to 12 m/s over this distance.) The concentration of the substance in the pipe is 1 mol/m\(^3\) and uniform everywhere. Then, because of the nature of this enthalpy flux, we expect the heat flux to be spatially non-uniform. Calculate the gradient in the measurable heat flux, \( dJ_u/dx \).

Solution. Since the particle flux \( J = cv \), at \( x = 0 \) we have

\[
J(0) = (1 \, \text{mol/m}^3)(10 \, \text{m/s})
\]

and

\[
J(1) = (1 \, \text{mol/m}^3)(12 \, \text{m/s}).
\]

This gives a gradient \( dJ/dx \) of \( (1 \, \text{mol/m}^3)(2 \, \text{m/s})/(1 \, \text{m}) = 2 \, \text{mol/m}^3/\text{s} \). For ideal gas, the chemical potential is \( \mu = \partial U / \partial n \), where \( U = \frac{3}{2} nRT \). Thus,
\[ \mu = \frac{3}{2}RT, \text{ where } T = 300 \text{ K.} \]

Then invoke:

\[ \frac{ds}{dt} = -\frac{d}{dx} \left[ \frac{1}{T} (J_u - \mu J_j) \right] = 0 \]

Thus, \( \frac{d}{dx} J_u = \mu \frac{d}{dx} J_j \), where \( \mu = \frac{3}{2}RT \). Plug in the numbers.

**Problem 128.** If two systems \( A \) and \( B \) are in thermal contact, the assembly \((A + B)\) being isolated, exchange a heat quantity 1 kJ every second (i.e., 1 kW), calculate the entropy production if \( T_A = 200 \text{ K} \) and \( T_B = 300 \text{ K} \). Is the entropy production positive or negative (explain)?

**Solution.** Heat flows from \( B \) to \( A \) (from hot body to the cold body), so

\[ \delta Q = \delta Q_B = -\delta Q_A = -1 \text{ kJ.} \]

\[ \sigma = \frac{\delta Q}{dt} \left( \frac{1}{T_B} - \frac{1}{T_A} \right) > 0 = -(1 \text{ kW}) \left( \frac{1}{300 \text{ K}} - \frac{1}{200 \text{ K}} \right) = \frac{5}{3} \text{ kW K}. \]

Entropy production is always positive.

**Problem 129.** Explain what is the driving force, \( X \), of a chemical reaction. What is the corresponding flux, \( J \)? What are the units of the flux and force?

**Problem 130.** Does a chemical reaction produce entropy? Why (or why not)?

**Solution.** Yes, spontaneous irreversible reactions produce entropy, the entropy production is \( \dot{\xi}(-\Delta G_{rxn}/T) \).
Appendices

A.1. Vector

For our purposes, it will be sufficient to define a vector $\mathbf{x}$ as a $n$-tuple:

$$\mathbf{x} = (x_1, x_2, \ldots, x_n),$$

where $x_i \in \mathbb{R}$ are all real numbers. These numbers $x_1, \ldots, x_n$ are called the coordinates of the vector. The coordinates are normally taken to be Cartesian, although other coordinate systems are possible. The vector $\mathbf{x}$ lives in $\mathbb{R}^n$.

A.2. Euclidean Length

A vector $\mathbf{r}$ with components $\mathbf{r} = (x, y, z)$ has Euclidean length:

$$|\mathbf{r}| \equiv r \equiv \sqrt{x^2 + y^2 + z^2}.$$  

A.3. Dot Product of Two Vectors

The notation $\mathbf{x} \cdot \mathbf{y}$ denotes a dot product. In 3D ($\mathbf{x}, \mathbf{y} \in \mathbb{R}^3$) the dot product of these vectors whose components are $\mathbf{x} = (x_1, x_2, x_3)$ and $\mathbf{y} = (y_1, y_2, y_3)$ can be expressed as:

$$\mathbf{x} \cdot \mathbf{y} = x_1 y_1 + x_2 y_2 + x_3 y_3.$$  

The geometric definition of the dot product is the projection of $\mathbf{x}$ along the vector $\mathbf{y}$:

$$\mathbf{x} \cdot \mathbf{y} = |\mathbf{x}| |\mathbf{y}| \cos \theta.$$
where $\theta$ is the angle between $\mathbf{x}$ and $\mathbf{y}$. $|\mathbf{x}|$ and $|\mathbf{y}|$ denote the Euclidean length of the vectors $\mathbf{x}$ and $\mathbf{y}$, respectively. One often writes $x$ and $y$ for the length (omitting the norm symbol) for simplicity.

A.4. Derivative of a Function of One Variable

Let $f(x)$ be a function from $U$ to $V$, where $U$ and $V$ are subsets on the real line ($U,V \subset \mathbb{R}$). Let $x$ be a point in the interior of $U$.\footnote{It is easier to take interior points rather than points on the boundary because calculating the derivative of $f$ at $x$ requires looking at the neighborhood of $x$. At the boundary, there is one side with no neighborhood.} If $f$ is differentiable, then\footnote{You can think of $h$ as a “small parameter”.}

$$f(x + h) = f(x) + h \cdot Df(x) + o(|h|),$$

where $o(|h|)$ is a function that tends to 0 as $|h| \to 0$. $Df(x)$ is called the derivative of $f$ at the point $x$. This definition is equivalent to the statement:

$$Df(x) = \lim_{|h| \to 0} \frac{f(x + h) - f(x)}{h}.$$  

The derivative $Df(x)$ is also denoted as $df(x)/dx$ or $f'(x)$.

A.5. More Than One Variable: Partial Derivative

When $f$ is a function of more than one variable, we must specify which variable we differentiate with respect to. Hence, the notion of partial derivative. Partial derivatives are required in thermodynamics because the thermodynamic potentials depend on several variables.

Suppose that $f$ depends on two variables, $f(x,y)$. We have the option of differentiating with respect to $x$ or $y$. (In fact, we must specify which variable we differentiate with respect to.) These derivatives are called partial derivatives and are defined by differentiating with respect to the variable of choice while holding the other constant:

$$\frac{\partial f}{\partial x} = \lim_{h \to 0} \frac{f(x+h,y) - f(x,y)}{h}, \quad \frac{\partial f}{\partial y} = \lim_{h \to 0} \frac{f(x,y+h) - f(x,y)}{h}.$$  

As stated, these derivatives are evaluated at the point $(x,y)$. For example, partial differentation of $f(xy) = xy$ with respect to $x$ yields $\partial f/\partial x = y$. (If you don’t see why, take $y$ to be constant, any constant, say, $y = 4$, i.e. $f = 4x$, whose derivative with respect to $x$ is 4.) Likewise $\partial f/\partial y = x$. Another example is $f(x,y) = y \sin(x)$. Its derivative with respect to $x$ is:

$$\frac{\partial f}{\partial x} = y \cos(x)$$
because this is what we get if we treat $y$ as if it were a constant. Similarly, the derivative with respect to $y$ is:

$$\frac{\partial f}{\partial y} = \sin(x).$$

In thermodynamics, the practice is to explicitly state which variables are held constant. For example, the internal energy $U(S, V, N)$ is a function of three variables ($S$: entropy, $V$: volume, $N$: number of particles) and its total derivative is denoted:

$$dU(S, V, N) = \left(\frac{\partial U}{\partial S}\right)_{V,N} dS + \left(\frac{\partial U}{\partial V}\right)_{S,N} dV + \left(\frac{\partial U}{\partial N}\right)_{S,V} dN.$$

The notation $\left(\frac{\partial U}{\partial S}\right)_{V,N}$ means “derivative of $U$ with respect to $S$ while holding $V$ and $N$ constant”. In general, however, it is not necessary to specify which variables are held constant because the definition of partial derivative makes it clear that all variables except the one of interest are to be held constant.

**A.6. Taylor’s Theorem**

Let $f$ be a function that is $n$ times differentiable at a point $x_0$. Then,

$$f(x_0 + \epsilon) = f(x_0) + \epsilon \frac{df}{dx}(x_0) + \frac{1}{2!} \epsilon^2 \frac{d^2f}{dx^2}(x_0) + \cdots + \frac{1}{n!} \epsilon^n \frac{d^n f}{dx^n}(x_0) + o(|\epsilon|^n),$$

where $o(|\epsilon|^n)$ is a quantity such that

$$\lim_{|\epsilon| \to 0} \frac{o(|\epsilon|^n)}{|\epsilon|^n} = 0.$$

In plain English, Taylor’s theorem says that if $\epsilon$ is small enough, and $f$ is smooth enough (sufficiently differentiable), we can expand $f$ around the neighborhood of $x_0$ using $n$ of its derivatives. These $n$ derivatives are evaluated at $x_0$ and thus are $n$ numbers. Together with $f(x_0)$, we only need $n+1$ numbers to approximate the function in this neighborhood.

When $n \to \infty$, this expansion is called Taylor’s series. When $x_0 = 0$ the Taylor series is called a Maclaurin series. To obtain the Taylor series from an analytical formula for $f$, we start by evaluating the derivatives of $f$ at the point $x = x_0$.

**A.6.1. Examples.** The Taylor (Maclaurin) series for $f(x) = e^x$ about $x_0 = 0$ is

$$e^x = \sum_{n=0}^{\infty} \frac{x^n}{n!}.$$
The Taylor series of $f(x) = \cos(x)$ about $x_0 = 0$ is

$$\cos(x) = \sum_{n=0}^{\infty} \frac{(-1)^n x^{2n}}{(2n)!}. $$

The case of $e^x$ is trivial since all derivatives of this function are $e^x$ (which, when evaluated at $x = 0$ give 1). For $\cos(x)$, the derivatives alternate between $\sin$ and $\cos$ together with some sign changes:

$$
\begin{align*}
\frac{df}{dx}(x) &= -\sin(x) & \frac{df}{dx}(0) &= 0, \\
\frac{d^2f}{dx^2}(x) &= -\cos(x) & \frac{d^2f}{dx^2}(0) &= -1, \\
\frac{d^3f}{dx^3}(x) &= \sin(x) & \frac{d^3f}{dx^3}(0) &= 0, \\
\frac{d^4f}{dx^4}(x) &= \cos(x) & \frac{d^4f}{dx^4}(0) &= 1,
\end{align*}
$$

which leads to

$$\cos(x) = 1 - \frac{1}{2!} x^2 + \frac{1}{4!} x^4 - \frac{1}{6!} x^6 + \ldots = \sum_{n=0}^{\infty} \frac{(-1)^n x^{2n}}{(2n)!}.$$  

A.7. Chain Rule

The chain rule is needed when we differentiate a function with respect to a variable that is found within the argument of the function. Example:

$$\frac{d}{dy} f(\lambda y) = \frac{df}{d\lambda} \cdot \frac{d(\lambda y)}{dy} = \frac{df}{d\lambda}.$$ 

A function $f$ expressed in the form $f(\cdot)$ is a function of some argument $\cdot$. The argument is the slot in which we insert a number or variable in order to evaluate the function’s value at that point. Here, the argument of $f$ is $\lambda y$. However, the variable we differentiate with respect to is $y$. $y$ is found in the argument of $f$, i.e. $\lambda y$ is a function of $y$. In the first equality, the term $\frac{df}{d\lambda}$ means “derivative of $f$ with respect to its argument”.

In two variables, the idea is the same except that we use partial derivatives to stress which argument of the function is targeted. For example, if we must differentiate $f(x, \lambda y)$ with respect to $y$, the chain rule gives:

$$\frac{\partial}{\partial y} f(x, \lambda y) = \frac{\partial f}{\partial y} \cdot \frac{\partial (\lambda y)}{\partial y} = \frac{\partial f}{\partial y} \lambda.$$
A.8. Total Differential

Here, \( \frac{\partial f}{\partial y} \) means derivative of \( f \) with respect to its second argument. A better notation for \( \frac{\partial f}{\partial y} \) would be \( D_2f \) to emphasize the second argument, rather than \( y \) (since \( y \) is different from \( \lambda y \)).

A.8. Total Differential

For a function of one variable, \( f(x) \), the total differential is defined as:

\[
df(x) = f'(x)dx = \frac{df}{dx} \cdot dx.
\]

It is also defined as

\[
df(x) = f(x + dx) - f(x).
\]

Taylor expansion of \( f(x + dx) = f(x) + f'(x)dx + O(dx^2) \) leads to:

\[
df(x) = f(\bar{x}) + f'(x)dx + O(dx^2) - f(\bar{x}) = f'(x)dx + O(dx^2),
\]

where \( O(dx^2) \) represents higher-order terms \((dx^2) \) and higher). And for this to agree with \( df(x) = f'(x)dx = \frac{df}{dx} \cdot dx \), then we must have that \( O(dx^2) = 0 \).

For a function of two variables,

\[
df(x,y) = f(x + dx, y + dy) - f(x,y) = f(x, y) + \frac{\partial f}{\partial x} \cdot dx
\]

\[
+ \frac{\partial f}{\partial y} \cdot dy + O(|dx, dy|^2) - f(x, y),
\]

where \( O(|dx, dy|^2) \) denotes higher order terms. For the rules of calculus to hold, we must have that \( O(|dx, dy|^2) = 0 \). Then,

\[
\boxed{df(x,y) = \frac{\partial f}{\partial x} \cdot dx + \frac{\partial f}{\partial y} \cdot dy.}
\]

This can be generalized to any number of variables. For example, a function of 3 variables, \( U \equiv U(S,V,N) \), has:

\[
dU(S,V,N) = \frac{\partial U}{\partial S} dS + \frac{\partial U}{\partial V} dV + \frac{\partial U}{\partial N} dN.
\]

And for the case \( U \equiv U(S,V,\{N_i\}) \)

\[
dU(S,V,\{N_i\}) = \frac{\partial U}{\partial S} dS + \frac{\partial U}{\partial V} dV + \sum_i \frac{\partial U}{\partial N_i} dN_i.
\]

Comparing this to the fundamental equation of thermodynamics, \( dU = TdS - PdV + \sum_i \mu_i dN_i \), we get thermodynamic definitions for \( T,P,\mu\): \( T \equiv \frac{\partial U}{\partial S}, \ P \equiv -\frac{\partial U}{\partial V}, \ \mu_i \equiv \frac{\partial U}{\partial N_i} \).
A.9. Exact Differentials

In the statement of the first law ($dU$ is an exact differential), we invoked the concept of exact differentials or exact forms. Exact forms are those that are obtained as the differential of some function $f(x, y)$:

$$\omega(\text{exact}) \equiv df = f_x(x, y)dx + f_y(x, y)dy = \nabla f \cdot d\vec{r},$$

where here $f_x$ denotes the partial of $f$ with respect to $x$ and $d\vec{r} = (dx, dy)$. The fundamental theorem of calculus says that integration of an exact 1-form $df$ on a closed interval $[a, b]$ is equal to the difference of $f$ at the boundaries of the interval, $\partial[a, b] = \{a, b\}$:

$$\int_{[a, b]} \omega = \int_a^b \omega = \int_a^b df = f(b) - f(a).$$

Another equivalent statement is:

$$\int_a^b \nabla f \cdot d\vec{r} = f(b) - f(a),$$

where $\nabla f = (\partial_x f, \partial_y f)$ and $d\vec{r} = (dx, dy)$. Thus, for exact differentials, their integrals only depend on the end points and not on the path taken to reach the end point.

On the other hand, more general expressions of the form

$$\omega = f_1(x, y)dx + f_2(x, y)dy$$

are called differential forms (1-forms). While all exact differentials are 1-forms, not all 1-forms happen to be exact.

It is helpful at this point to introduce Euler’s test to help us determine whether an expression containing first-order differentials is exact. Take, for example, the following differential 1-form:

$$\omega = (xy)dx + (xy)dy.$$ 

If this were an exact differential we would write $\omega = df$ for some $f$. If this were the case then $df$, according to the definition of differential, must also be equal to (subscript denotes which variable is held constant):

$$df = \left(\frac{\partial f}{\partial x}\right)_y dx + \left(\frac{\partial f}{\partial y}\right)_x dy$$

But since mixed partials should be equal:

$$\frac{\partial}{\partial y} \left(\frac{\partial f}{\partial x}\right) = \frac{\partial}{\partial x} \left(\frac{\partial f}{\partial y}\right).$$

This “Euler test” provides a necessary condition for exact differential. Let’s check to see if this is the case: $\partial_y(xy) = x$ whereas $\partial_x(xy) = y$, which are not equal for all $x, y$. So we cannot say that $\omega$ is an exact differential.
On the other hand,
\[ \omega = ydx + xdy \]
meets the necessary condition since \( \partial_y(y) = 1 \) and \( \partial_x(x) = 1 \) are equal. In fact, we see that \( \omega = d(xy) \). Thus, it is an exact differential because it is the derivative \( d \) of some function \( f(x, y) = xy \).

Any expression \( \omega \) which can be written in the form \( df \) for some \( f \) is an exact differential. In fact, this is the correct definition of \textbf{exact differential}. Euler’s test provides a necessary condition. The sufficient condition is provided by identifying the function \( f \) for which \( \omega = df \) over some specified domain.\(^3\) The latter task of finding \( f \) requires more work.

\textbf{A.10. Einstein Summation Convention}

Whenever two indices are repeated in the same term, a summation is implied. For example, in the dot product of \( \mathbf{u} = (u_x, u_y, u_z) \) and \( \mathbf{v} = (v_x, v_y, v_z) \) we have:
\[ \mathbf{u} \cdot \mathbf{v} = u_\alpha v_\alpha \equiv \sum_{i=1}^{3} u_i v_i = u_xv_x + u_yv_y + u_zv_z. \]
It is simpler to write \( u_\alpha v_\alpha \) than the entire summation.

\textbf{A.11. Multivariate Taylor Expansion}

In 1D the Taylor expansion of \( f(x + h) \) at \( x \) is:
\[ f(x + h) = f(x) + f'(x)h + \frac{1}{2}f''(x)h^2 + \frac{1}{3!}f^{(3)}(x)h^3 + \ldots \]
In \( n \)-D, a scalar-valued function \( f(\mathbf{x}) \), where \( \mathbf{x} \in \mathbb{R}^n \), is expanded as:
\[ f(\mathbf{x} + \mathbf{h}) = f(\mathbf{x}) + \sum_{i=1}^{n} \frac{\partial f(\mathbf{x})}{\partial x_i} h_i + \frac{1}{2} \sum_{i=1}^{n} \sum_{j=1}^{n} \frac{\partial^2 f(\mathbf{x})}{\partial x_i \partial x_j} h_i h_j + \frac{1}{3!} \sum_{i=1}^{n} \sum_{j=1}^{n} \sum_{k=1}^{n} \frac{\partial^3 f(\mathbf{x})}{\partial x_i \partial x_j \partial x_k} h_i h_j h_k + \ldots \]

\textbf{A.11.1. Abbreviation for Partial Derivatives.} We will often use the shorthand notation to abbreviate the notation for partial differentiation:
\[ \partial_\alpha \equiv \frac{\partial}{\partial x_\alpha} \]

\(^3\)The choice of domain is important because some forms may be not be exact over all space, but may be exact over a restricted domain.
In this notation, and using the summation convention, the multivariate Taylor expansion looks particularly neat:

\[ f(x + h) = f(x) + h_i \partial_i f(x) + \frac{1}{2} h_i h_j \partial_i \partial_j f(x) + \frac{1}{3!} h_i h_j h_k \partial_i \partial_j \partial_k f(x) \ldots \]

(summation convention) The components of \( h \) have been moved to the left-hand-side of all derivative operators to avoid any possible confusion about which quantity is differentiated.

**A.12. Example: Derivative of \( 1/r \)**

This example is often encountered when dealing with potentials (gravitational, electromagnetic). The first order partial derivative of

\[ \frac{1}{r} \equiv \frac{1}{|r|} = \frac{1}{\sqrt{x^2 + y^2 + z^2}} \]

with respect to \( x \) is:

\[ \frac{\partial}{\partial x} \left( \frac{1}{r} \right) \equiv \partial_x \left( \frac{1}{r} \right) = -\frac{1}{2} \frac{(2x)}{(x^2 + y^2 + z^2)^{3/2}} = -\frac{x}{r^3} \]

Similar expressions are found for differentiation with respect to \( y \) or \( z \). Thus, for any component \( \alpha = x, y, z \) we have:

\[ \frac{\partial}{\partial r_\alpha} \left( \frac{1}{r} \right) \equiv \partial_\alpha \left( \frac{1}{r} \right) = -\frac{r_\alpha}{r^3} \]

where \( r_\alpha \) can be any of \( x, y, z \) depending on the choice of index \( \alpha = x, y, z \).

**A.13. Probability Concepts**

Molecular velocities are an inherently probabilistic concept. This is because we are usually dealing with a large number of molecules and we do not have perfect knowledge of every single molecule’s position and velocity. The best we can do is describe their behavior statistically. Their description requires us to introduce tools to manipulate random variables. Random variables can be continuous or discrete, depending on whether they take continuous or discrete values, respectively. An example of a continuous random variable is the length of an object. Another example of continuous random variable is the velocity component of a molecule chosen at random. The velocity component, for example \( v_x \), takes values in the interval \((-\infty, \infty)\). The speed of a molecule, \( v = \sqrt{v_x^2 + v_y^2 + v_z^2} \), is another example of a continuous random variable, taking values in the interval \([0, \infty)\). Length is a random variable which can take positive real values, a continuous interval, \([0, \infty)\). An example of a discrete random variable is the number of counts within a time interval. Counts can only take discrete values \((1, 2, 3, \ldots)\), in this
case, the positive integers. Similarly, the roll of a (normal) die is a discrete random variable, which can only take values in the set \{1, 2, 3, 4, 5, 6\}.

A.14. Continuous Random Variables

A continuous random variable \(X\) cannot be a simple variable because it describes the value of a measurement made in a given experiment and all statistical properties of this measurement.\(^4\) Instead it is better to think of \(X\) as a function that encodes the statistical properties of the measurement. To this end, let us associate with \(X\) a probability density function (PDF), \(p(x)\).

A.15. Probability Distribution Functions

A PDF, \(p(x)\), describes the likelihood that a random variable will take a certain value. For example, the probability, \(P\), that a random variable \(X\) will take a value in the interval \([a, b]\) is given by:

\[
P(a < X < b) = \int_a^b p(x) dx
\]

where \(p(x)\) is the probability density function for the random variable \(X\). The requirements are that \(p(x)\) should be non-negative (\(\geq 0\)) and normalized to 1:

\[
p(x) \geq 0, \forall x, \quad \int_{-\infty}^{\infty} p(x) dx = 1,
\]

where the limits of integration (denoted here by \([\pm \infty]\)) range over the domain of definition of the random variable, i.e. if this domain is \([0, 1]\) then the limits of integration are from 0 to 1. Its main purpose is to compute average values. For example, the average (mean) value of \(X\) is:

\[
\mathbb{E}X \equiv \langle X \rangle = \int_{-\infty}^{\infty} xp(x) dx.
\]

The average of a function of \(X\) is obtained by replacing the random variable \(X\) by a dummy integration variable \(x\), multiplying by \(p(x)\) and integrating over \(x\):

\[
\mathbb{E}f(X) \equiv \langle f(X) \rangle = \int_{-\infty}^{\infty} f(x)p(x) dx.
\]

\(^4\)This experiment could consist, for example, of choosing a molecule at random within an ensemble and recording its velocity (assuming this is even possible!). Every time a molecule is chosen, it will usually be a different molecule than in previous experiments, especially if the ensemble is very large. Even if the same molecule is chosen twice, this will be done at different times, and its velocity will be different. In each case, the result (velocity measurement) will be very different. Hence, we need a statistical description that will tell us what is the average velocity, what is the variance (spread of values), etc.
A famous example of a PDF is the normal (Gaussian) distribution, whose pdf is

\[ p(x) = \frac{1}{\sqrt{2\pi}} e^{-x^2/2}. \]

**Meaning of the PDF:** The PDF \( p(x) \) has the following meaning. The quantity \( p(x)dx \) is the probability \( \mathbb{P} \) that the random variable \( X \) lies in the interval \( [x, x + dx] \):

\[ p(x)dx = \mathbb{P}(x \leq X \leq x + dx), \]

where \( dx \) is an infinitesimally small quantity and \( \mathbb{P}(\cdot) \) denotes the probability \( \mathbb{P} \) of the event \( \cdot \) occurring. The quantity \( p(x)dx \) in itself is rarely used. Instead, one integrates this expression to find the probability that \( X \) will take some value in a finite interval \( [a, b] \):

\[ \mathbb{P}(a \leq X \leq b) = \int_a^b p(x)dx. \]

\[ ^a \text{Since } \mathbb{P}(x \leq X \leq x + dx) = \mathbb{P}(X \leq x + dx) - \mathbb{P}(X \leq x) = d\mathbb{P}(X \leq x), \text{ integrating } \mathbb{P}(x \leq X \leq x + dx) = d\mathbb{P}(X \leq x) \text{ from } a \text{ to } b \text{ yields } \int_a^b d\mathbb{P}(X \leq x) = \mathbb{P}(X \leq b) - \mathbb{P}(X \leq a) = \mathbb{P}(a \leq X \leq b) \text{ whereas integrating } p(x)dx \text{ yields } \int_a^b p(x)dx. \text{ Since the two are equal, we have that } \mathbb{P}(a \leq X \leq b) = \int_a^b p(x)dx. \]

### A.16. Mean Value

The mean value of a random variable \( X \), denoted \( \mathbb{E}X \), is defined as:

\[ \mathbb{E}X \equiv \int_{-\infty}^{\infty} xp(x)dx. \]

That is, to obtain the mean value of \( X \), we replace the random variable \( X \) by a regular variable \( x \) that represents its value, then multiply by \( p(x) \) and integrate over all \( x \).

We note that this expression differs from the sample mean, \( \frac{1}{n} \sum_{i=1}^{n} x_i \). The sample mean is only an estimate of the mean.

### A.17. Variance

The variance of \( X \) is defined as the sum of square differences between \( X \) and its mean, weighted by the pdf:

\[ \sigma^2 = \int_{-\infty}^{\infty} p(x)(x - \mathbb{E}X)^2dx. \]
A.18. Average of a Function of a Random Variable

The square can be expanded to give $\int_{-\infty}^{\infty} p(x)(x^2 + (EX)^2 - 2xEX)dx$ and thus

$$\sigma^2 = E(X^2) - (EX)^2.$$  

The square root of the variance is called the standard deviation and is denoted $\sigma$.

A.18. Average of a Function of a Random Variable

The average (or mean, or expectation value) of a function $f$ of a random variable $X$ is defined as:

$$E[f(X)] = \int_{-\infty}^{\infty} f(x)p(x)dx.$$  

As an example, take $f(x) = x^n$. This is called the $n$-th moment of $X$:

$$E(X^n) = \int_{-\infty}^{\infty} p(x)x^ndx.$$  

The case $n = 1$ (first moment) is called the “mean value of $X$”:

$$EX = \int_{-\infty}^{\infty} p(x)x dx$$  

“mean value of $X$”

The variance of $X$ is defined as $\sigma^2 = \int_{-\infty}^{\infty} p(x)(x - EX)^2dx$, which also equals $\sigma^2 = E(X^2) - (EX)^2$. Thus, the variance is the second moment of $X$ minus the square of the average of $X$.

A.18.1. Example 1. Suppose that the kinetic energy, $K(v) = \frac{1}{2}mv^2$, of an object of mass $m$ is to be calculated using experimentally measured values of the velocity $v$. Thus, $v$ is a random variable $V$. Since $v$ is a random variable, $K(v)$ is also a random variable. You determine from experiments that the velocities, $v$, are Gaussian-distributed around 100 m/s, with a standard deviation of 1 m/s, i.e. $p(v) = \frac{1}{\sqrt{2\pi}}e^{-(v-100)^2/2}$. What is the average kinetic energy, $EK(V)$? You expect that $EK(V)$ should be close to $K(100) = \frac{1}{2}m(100)^2$. However, the exact value of $EK(V)$ will depend on the distribution $p(v)$. We need to calculate:

$$EK(V) = \int_{-\infty}^{\infty} \frac{1}{2}mv^2 \frac{1}{\sqrt{2\pi}}e^{-(v-100)^2/2}dv$$  

$$= \frac{m}{2\sqrt{2\pi}} \int_{-\infty}^{\infty} v^2e^{-(v-100)^2/2}dv \approx \frac{m}{2}1.0001 \cdot 10^4.$$  

So fairly close to $K(100)$ but slightly higher.

A.18.2. Example 2. Suppose that $f(\theta) = \cos(\theta)$ and $p(\theta) = 1$ for $\theta \in [-\frac{1}{2}, \frac{1}{2}]$ and $p = 0$ elsewhere (uniform distribution). Denote the random variable as $\Theta$. Using $p(\theta)$ you can easily check that $E\Theta = \int_{-1/2}^{1/2} \theta d\theta = 0$.  


A. Appendices

The average of \( \cos(\Theta) \) is:

\[
\mathbb{E}\cos(\Theta) = \int_{-1/2}^{1/2} \cos(\theta) d\theta \approx 0.9589.
\]

Note: it is not equal to 1 even though the average of \( \Theta \) is 0.

A.19. More than One (Continuous) Random Variable

If we are to compute the average of an expression that is a function of more than one random variable, we need to use the joint probability density \( p_{XY}(x, y) \), which is everywhere non-negative (\( p_{XY}(x, y) \geq 0 \)) and integrates to 1:

\[
\int_{-\infty}^{\infty} \int_{-\infty}^{\infty} p_{XY}(x, y) dx dy = 1.
\]

The average of a function \( g(X, Y) \) would be:

\[
\mathbb{E}g(X, Y) = \int_{-\infty}^{\infty} \int_{-\infty}^{\infty} p_{XY}(x, y) g(x, y) dx dy.
\]

Given a joint PDF, \( p_{XY}(x, y) \), the marginal density function for \( X \) is obtained by integrating over \( y \):

\[
p_X(x) = \int_{-\infty}^{\infty} p_{XY}(x, y) dy.
\]

Similarly, the marginal density for \( Y \) is obtained by integrating over all \( x \):

\[
p_Y(y) = \int_{-\infty}^{\infty} p_{XY}(x, y) dx.
\]

Note: you can easily check that both marginals \( p_X(x) \) and \( p_Y(y) \) are bona fide densities, i.e. nonnegative and normalized to 1.

A.20. Statistical Independence

The marginal density is a useful concept if you are asked to check whether or not two random variables are statistically independent. Two random variables \( X \) and \( Y \) are statistically independent if the joint probability density is written as the product of densities of each variable:

\[
p_{XY}(x, y) = p_X(x)p_Y(y),
\]

where \( p_X(x) \) and \( p_Y(y) \) are the marginal densities of \( X \) and \( Y \), respectively. They can be computed from \( p_{XY}(x, y) \) by integrating.

There are at least two consequences of statistical independence that we can immediately point out. First, one concerns expectation values. Consider
the average of a function \( g(X, Y) \) of two r.v. \( X \) and \( Y \):

\[
E[g(X, Y)] = \int_{-\infty}^{\infty} \int_{-\infty}^{\infty} p_{XY}(x, y)g(x, y)\,dx\,dy
\]

If \( g \) factors as a product of a function of \( X \) times a function of \( Y \), for example \( g(X, Y) = XY \) then the expectation value of \( XY \) is equal to the product of expectation values of \( X \) and that of \( Y \):

\[
E(XY) = \int_{-\infty}^{\infty} \int_{-\infty}^{\infty} p_X(x)p_Y(y)xy\,dx\,dy = \int_{-\infty}^{\infty} p_X(x)x\,dx \cdot \int_{-\infty}^{\infty} p_Y(y)y\,dy = E(X) \cdot E(Y).
\]

The second consequence pertains to the calculation of probabilities in general, which also factors as a product. For example, the joint probability:

\[
P(X \in A, Y \in B) = \int_{\{(x,y)|x\in A, y \in B\}} p_{XY}(x, y)\,dx\,dy
\]

\[
= \int_{\{(x,y)|x\in A, y \in B\}} p_X(x)p_Y(y)\,dx\,dy
\]

\[
= \int_{\{x|x \in A\}} p_X(x)\,dx \cdot \int_{\{y|y \in B\}} p_Y(y)\,dy
\]

\[
= P(X \in A)P(Y \in B).
\]


The normal distribution \( N(\mu, \sigma^2) \) with mean \( \mu \) and variance \( \sigma^2 \) is defined by the PDF:

\[
p(x) = \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{(x-\mu)^2}{2\sigma^2}}
\]

The prefactor \( \frac{1}{\sqrt{2\pi\sigma^2}} \) is such that \( p(x) \) adds up to 1:

\[
\int_{-\infty}^{\infty} p(x)\,dx = 1
\]

This can be verified using the well-known result for a Gaussian integral

\[
\int_{-\infty}^{\infty} e^{-ax^2}\,dx = \sqrt{\pi/a}, \text{ where } a > 0.
\]

This PDF is plotted in Fig. A.1. The width is proportional\(^5\) to \( \sigma \) and mean value is \( \mathbb{E}X \).

If a random variable \( X \) follows a Gaussian distribution with mean \( \mu = \mathbb{E}X \) and variance \( \sigma^2 \) we write

\[
X \sim N(\mu, \sigma^2).
\]

\(^5\)In fact, the full width at half maximum of the Gaussian is \( 2\sqrt{2\log 2}\sigma \approx 2.355\sigma \). You can check this by finding the values of \( x \) for which \( \frac{1}{2} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{x^2}{2\sigma^2}} = \frac{1}{2} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{x^2}{2\sigma^2}} = e^{-\frac{x^2}{2\sigma^2}}, \) since the maximum of the function is \( \frac{1}{\sqrt{2\pi\sigma^2}} \) (set \( x = 0 \)). Taking logs of both sides gives \( x = \pm \sqrt{2\sigma^2 \log 2} \).
A.20.1.1. The Gaussian Probability Density. The Gaussian probability density (in 1D) with mean zero is:

\[ p(x) = \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{x^2}{2\sigma^2}}. \]

Its integral is 1:

\[ \int_{\mathbb{R}} p(x) dx = 1 \]

This can be easily checked using the standard result from integration tables:

\[ \int_{-\infty}^{\infty} e^{-ax^2} dx = \sqrt{\frac{\pi}{a}}. \]

\( \sigma \) is called the standard deviation. \( \sigma \) is related to the width of the bell curve. \( \sigma^2 \) is called the variance. The function \( p(x) \) has a bell shape centered at 0 and a full-width at half-maximum of approximately \( \approx 2.355\sigma \).

A.20.2. Other Examples of Probability Distributions. The Gibbs distribution for the probability of the system \( X \) being in state \( x \) (random variable \( X \) taking the value \( x \)) is:

\[ \mathbb{P}(X = x) = \frac{1}{Z(\beta)} \exp(-\beta E(x)) \]

where \( Z(\beta) \) is a normalization factor called the partition function:

\[ Z(\beta) = \int dx \exp(-\beta E(x)). \]

\( \beta \) is a free parameter called the inverse temperature (\( \beta = 1/k_B T \)). When integrated over \( x \), this distribution is normalized:

\[ \int \mathbb{P}(X = x) dx = \frac{1}{Z(\beta)} \int dx \exp(-\beta E(x)) = \frac{Z(\beta)}{Z(\beta)} = 1. \]
The Rayleigh distribution has density
\[ p(x) = \frac{x}{\sigma^2} \exp \left( -\frac{x^2}{2\sigma^2} \right), \quad x \geq 0, \]
where \( \sigma \) is a parameter of the distribution.

The Gaussian distribution is
\[ p(x) = \frac{1}{\sqrt{2\pi\sigma^2}} \exp \left( -\frac{x^2}{2\sigma^2} \right), \]
where \( \sigma \) is a parameter of the distribution.

**A.20.3. Boltzmann Factor.** Another example of probability density is the Boltzmann distribution, or the Boltzmann factor. The latter gives the probability of a system to be in the state \( i \), with energy \( E_i \):
\[ p(\text{state } i) \propto \exp \left( -\frac{E_i}{k_B T} \right) \]

A state is a set of quantum numbers. For example, if we specify the values of \( n, l, m_l, m_s \) for an electron in an atom, we have specified a state. Note: state is different from energy level. More than 1 state can correspond to the same energy. That is the case for the hydrogenic atom, where only the principal quantum number \( n \) determines the energy. For a given \( n \), all values of \( l, m_l, m_s \) give the same energy.

The meaning of the Boltzmann factor is the probability of finding the system in a given state when the temperature is \( T \), assuming classical particles (not fermions or bosons). In the limit \( T \to 0 \) only the ground state is occupied (the reader should check this!). In the high-temperature limit \( (T \to \infty) \) all levels are equally likely to be populated; the distribution behaves like a uniform distribution.

**A.20.4. Derivation of the Maxwellian Distribution.** We seek to determine the constants \( a \) and \( c \) in:
\[ f(v_i) = c \exp(-av_i^2), \quad (i = x, y, z) \]

**A.20.4.1. Finding \( c \).** In Section 1.3.3, the constant \( c \) was obtained from the normalization condition:
\[ \int_{-\infty}^{\infty} dv_i f(v_i) = c \int_{-\infty}^{\infty} dv_i \exp(-av_i^2), \quad a > 0 \]
which led to \( c = \sqrt{a/\pi} \).

**A.20.4.2. Finding \( a \).** \( a \) can be found by invoking the two previous results:
\[ \langle \epsilon_{\text{kin}} \rangle = \frac{3}{2} k_B T, \quad \langle \epsilon_{\text{kin}} \rangle = \frac{1}{2} m \langle v^2 \rangle = \frac{3}{2} m \langle v_x^2 \rangle. \]
But $\langle v_z^2 \rangle$ is an average which can be computed using the probability density for $\vec{v}$.\(^6\)

\[
k_B T = m \langle v_z^2 \rangle = m \int_{-\infty}^{\infty} dv_z v_z^2 f(v_z) = m \sqrt{\frac{a}{\pi}} \int_{-\infty}^{\infty} dv_z v_z^2 \exp(-av_z^2)
\]

\[
= -m \sqrt{\frac{a}{\pi}} \frac{d}{da} \left( \int_{-\infty}^{\infty} e^{-av_z^2} dv_z \right) = -m \sqrt{\frac{a}{\pi}} \frac{d}{da} \left( \sqrt{\frac{\pi}{a}} \right) = m \sqrt{\frac{a}{2} \frac{1}{\sqrt{a^{3/2}}} = \frac{m}{2a}.
\]

Therefore, $a = \frac{m}{2k_B T}$ and $c = \sqrt{\frac{a}{\pi}} = \sqrt{\frac{m}{2\pi k_B T}}$, so that:

\[
f(v_i) = \sqrt{\frac{m}{2\pi k_B T}} \exp \left( -\frac{mv_i^2}{2k_B T} \right), \quad f(\vec{v}) = \left( \frac{m}{2\pi k_B T} \right)^{3/2} \exp \left( -\frac{mv^2}{2k_B T} \right).
\]

A.20.4.3. Mean Square Displacement. The mean square displacement, $\langle x(t)^2 \rangle$, or also denoted $\langle X_t^2 \rangle$ where $X_t$ is the random variable for position at time $t$, for the 1D diffusion equals $2Dt$. This is proved as follows. Apart from the factor $n$, Eq. (A.4) is merely the probability density for finding a particle at $x$ at time $t$:

\[
p_t(x) = \frac{c(x,t)}{n} = \frac{1}{\sqrt{4\pi Dt}} e^{-x^2/4Dt}.
\]

So using the definition of the average (let $a = 1/4Dt$):

\[
\langle X_t^2 \rangle = \int_{-\infty}^{\infty} x^2 p_t(x) dx = \frac{1}{\sqrt{4\pi Dt}} \int_{-\infty}^{\infty} x^2 e^{-x^2}/4Dt dx = \sqrt{\frac{a}{\pi}} \int_{-\infty}^{\infty} x^2 e^{-ax^2} dx
\]

\[
= -\sqrt{\frac{a}{\pi}} \frac{d}{da} \left( \int_{-\infty}^{\infty} e^{-ax^2} dx \right) = -\sqrt{\frac{a}{\pi}} \frac{d}{da} \left( \sqrt{\frac{\pi}{a}} \right) = \frac{1}{2} \sqrt{\frac{a}{\pi}} \left( \frac{\sqrt{\pi}}{a^{3/2}} \right) = \frac{1}{2a}
\]

which is equal to $2Dt$.

A.20.4.4. Transition Probabilities. Equation (A.4), $\left( \frac{\partial}{\partial \eta} \right)_z = \frac{1}{\left( \frac{\partial^2}{\partial \eta^2} \right)_z}$, asks the question: what is the concentration at point $x$ at time $t + \tau$. This concentration consists of a sum of contributions from all possible spatial positions located some distance $\Delta$ away from $x$ weighted by the probability of undergoing a jump of size $\Delta$ to reach the point $x$. This weighting factor is $\phi(\Delta)$. This makes intuitive sense and can be taken as the definition of the function $\phi(\Delta)$.

\(^6\)Alternatively, we could have started by using the 3D probability density for $\vec{V}$ by writing, $\langle v_z^2 \rangle = \int f(\vec{v}) v_z^2 d^3\vec{v}$, because this is equal, by normalization of the two functions $f(v_x)$ and $f(v_y)$, to

\[
\int_{-\infty}^{\infty} dv_x f(v_x) \int_{-\infty}^{\infty} dv_y f(v_y) \int_{-\infty}^{\infty} dv_z v_z^2 f(v_z) = \int_{-\infty}^{\infty} dv_z v_z^2 f(v_z)
\]

But this step can be skipped in favor of simply using $f(v_z)$ the probability density for the $V_z$ component.
Those familiar with probability theory may recognize Eq. (A.4) as the definition of transition probability density for a Markov process. The probability density $p_t(x)$ of finding the particle at position $x$, time $t$, obeys the condition

$$p_s(x) = \int_{-\infty}^{\infty} p_t(y) p(x, s \mid y, t) dy, \quad s > t$$

where $p(x, s \mid y, t)$ a conditional probability density describing the probability of the random process to be found at the point $x$ at time $s$ ($s > t$) under the condition that it was at the point $y$ at an earlier time $t$. It is known as the transition probability density. Taking $p_s(x)$ to be proportional to the concentration $c(x, s)$, where $s = t + \tau$ yields Eq. (A.4).

### A.21. Angular Averaging

A PDF can also depend on several variables. For example, $p(r, \theta, \phi)$. It is called a joint probability density. It is possible to integrate over some of these variables and obtain a reduced probability density, also known as marginal probability density. Suppose that $p_{XY}(x, y)$ is a joint probability density for the random variables $X$ and $Y$, i.e.

$$\int \int p_{XY}(x, y) dx dy = 1 \text{ (normalization)}, \quad p_{XY}(x, y) \geq 0.$$ 

Integrating with respect to $y$ eliminates the random variable $Y$ and yields a probability density for the random variable $X$ only:

$$p_X(x) = \int p_{XY}(x, y) dy$$

An example of marginal distribution is an angle-averaged distribution. Suppose that we have a (joint) probability density function, $p(r, \theta, \phi)$, for the three random variables $r, \theta, \phi$ and that we want to average over $\theta$ and $\phi$. This is done as follows:

$$p(r) \equiv \frac{\int_{-\frac{1}{2}}^{\frac{1}{2}} \int_{0}^{\pi} p(r, \theta, \phi) d\cos\theta d\phi}{\int_{-\frac{1}{2}}^{\frac{1}{2}} \int_{0}^{\pi} d\cos\theta d\phi} = \frac{1}{4\pi} \int_{-\frac{1}{2}}^{\frac{1}{2}} \int_{0}^{\pi} p(r, \theta, \phi) d\cos\theta d\phi$$

where the denominator is added for normalization purposes. The specific case shown in the above equation is the example of averaging over the unit sphere.

The need for marginal distribution arises when physical quantities are being averaged away by some process, such as molecular tumbling motion. The end result, $p(r)$, contains no information about the angles except the effect of averaging over those angles.

In the next section, we will perform angle-averaging of the intermolecular interactions. This is done because molecular rotations (tumbling) causes the angles to be random functions of time. In experiments, we almost always
measure time-averaged quantities because the measurement times (milliseconds to seconds) are much longer than the molecular rotations (picoseconds or less).

Note: it is the PDF that is averaged over the angles and not the interaction energy itself. In the high-temperature limit, those two operations are identical. At finite temperatures, there is a difference. In the last section below, we will show how the Keesom interaction can be derived by averaging the interaction energy over the angles, but weighted by the Boltzmann factor. Nearly identical results are obtained.

A.22. Spherical Coordinates

The spherical coordinates are defined by

\[ x = r \sin \theta \cos \phi, \quad y = r \sin \theta \sin \phi, \quad z = r \cos \theta \]

and the unit vectors are shown in Fig. A.2.

![Orthogonal frame in spherical coordinates.](image)

The inverse transformation is

\[ r = \sqrt{x^2 + y^2 + z^2}, \quad \theta = \cos^{-1} \left( \frac{z}{\sqrt{x^2 + y^2 + z^2}} \right), \quad \phi = \tan^{-1} \left( \frac{y}{x} \right) \]

A.23. Dirac Delta Function

The Dirac delta can be loosely thought of as a function on the real line which is zero everywhere except at the origin, where it is infinite,

\[ \delta(x) = \begin{cases} +\infty, & x = 0 \\ 0, & x \neq 0 \end{cases} \]
and which is also constrained to satisfy the identity
\[ \int_{-\infty}^{\infty} \delta(x) \, dx = 1. \]

This is not a rigorous mathematical definition, but it will be enough for our purpose later in Section A.24.2, where it will be used to express the initial condition to the diffusion equation whereby at time \( t = 0 \) all molecules are at \( x = 0 \) and nowhere else.

The Dirac delta function has another purpose, namely, to extract the value of a function at some point \( a \) under the integral sign:
\[ \int_{-\infty}^{\infty} f(x) \delta(x - a) \, dx = f(a). \]

This is called the “sifting property” of the Dirac delta function.

You can also think of it as a sharply peaked Gaussian density:
\[ \delta(x) = \lim_{\sigma \to 0} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-x^2 / 2\sigma^2}. \]

There are many other representations of the Dirac delta function.

In 3D the Dirac delta function has the meaning \( \delta(x) = \delta(x)\delta(y)\delta(z) \).

A.24. Integration Over a Sphere

Suppose that you are asked to average a function over the unit sphere. We simply integrate over the solid angle \( d\Omega = \sin \theta \, d\theta \, d\phi \). This is equivalent to setting \( r = 1 \) in the volume integral in spherical coordinates and omitting the \( dr \) integration (which is equivalent to adding a Dirac delta function \( \delta(r - 1) \) in the integrand). The area element in spherical coordinates is shown in Fig. A.3.
A.24.1. Change of Variables Formula. Suppose that \( x, y \) and \( z \) are expressed in terms of new coordinates \( u, v, w \). The volume integral

\[
\int_R f(x, y, z) \, dx \, dy \, dz
\]

of the function \( f \) over the region \( R \) is also equal to

\[
\int_R f(x(u, v, w), y(u, v, w), z(u, v, w)) \left| \frac{\partial(x, y, z)}{\partial(u, v, w)} \right| \, du \, dv \, dw
\]

where \( u, v, w \) is a parametrization of the region \( R \). The integral is still computed over the region \( R \), except that the limits of integration are specified in the \( u, v, w \) coordinates. \( \left| \frac{\partial(x, y, z)}{\partial(u, v, w)} \right| \) is the Jacobian of the transformation, which is computed from the determinant:

\[
\begin{vmatrix}
  \frac{\partial x}{\partial u} & \frac{\partial x}{\partial v} & \frac{\partial x}{\partial w} \\
  \frac{\partial y}{\partial u} & \frac{\partial y}{\partial v} & \frac{\partial y}{\partial w} \\
  \frac{\partial z}{\partial u} & \frac{\partial z}{\partial v} & \frac{\partial z}{\partial w}
\end{vmatrix}
\]

where \( \frac{\partial x}{\partial u} \) denotes the partial of \( x \) with respect to \( u \), i.e. \( \frac{\partial x}{\partial u} = \frac{\partial x}{\partial u} \), etc.

You can check that in spherical coordinates the Jacobian is \( r^2 \sin \theta \). Thus, the volume element

\[
dV = dx \, dy \, dz = r^2 \sin \theta \, dr \, d\theta \, d\phi.
\]

A.24.2. Exercise: the 3D case. Using Taylor expansions in multiple variables, we can extend the analysis from the previous section and show that the diffusion equation in 3D is of the form:

\[
\frac{\partial c}{\partial t} = D\nabla^2 c(x, t),
\]

where \( c = c(x, t) \), \( x = (x, y, z) \) is a vector and \( \nabla^2 \) is the Laplacian, i.e. \( \nabla^2 = \frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2} \). Find the solution to this diffusion equation in 3D for similar initial conditions, i.e. \( c(x, 0) = n\delta(x) \). Show that the mean-square displacement in 3D is \( 6Dt \) (c.f. \( 2Dt \) in 1D).

A.24.3. Exercise: Mean-Square Displacement. The mean-square displacement is defined as \( \mathbb{E}(X^2) \). Using the definition of average this means solve \( \mathbb{E}(X^2) = \int \! x^2 \, p(x, t) \, dx \). Calculate this average using the probability density we found when solving the diffusion equation:

\[
p(x, t) = \frac{1}{\sqrt{4\pi Dt}} e^{-x^2/4Dt}
\]
where \( \int_{\mathbb{R}} p(x, t) \, dx = 1 \). Show that \( \mathbb{E}(x^2) = 2Dt \). Note: the reason we don’t use \( \mathbb{E}X = 0 \) is that it’s equal to zero and therefore trivial.

A.24.4. Exercise: Solution for Arbitrary Initial Conditions. When solving the diffusion equation we assumed that the concentration at time zero, was proportional to \( \delta(x) \). This led to the following result:

\[
p(x, t) = \frac{1}{\sqrt{4\pi Dt}} e^{-x^2/4Dt}.
\]

Suppose instead that we have the initial condition \( c(x, 0) = g(x) \). Show that the solution would be a convolution:

\[
c(x, t) = \int_{\mathbb{R}} p(x - y, t) g(y) \, dy.
\]

In 3D this would be of the form:

\[
c(x, t) = \int_{\mathbb{R}^3} p(x - y, t) g(y) \, d^3y,
\]

where \( d^3y = dy_1 dy_2 dy_3 \) is a volume element.

A.25. Gradient

The gradient of a function of three variable \( f(x, y, z) \) is defined in Cartesian coordinates as the vector

\[
\nabla f = \frac{\partial f}{\partial x} \hat{x} + \frac{\partial f}{\partial y} \hat{y} + \frac{\partial f}{\partial z} \hat{z}.
\]

In spherical coordinates (using the “physics convention” for the angles \( \theta, \phi \) shown in Fig. A.4), the gradient is

\[
\nabla f = \frac{\partial f}{\partial r} \hat{r} + \frac{1}{r} \frac{\partial f}{\partial \theta} \hat{\theta} + \frac{1}{r \sin \theta} \frac{\partial f}{\partial \phi} \hat{\phi}.
\]

For a function, such as \( V(r) \), that depends only on \( r \), then only the first term matters:

\[
\nabla V(r) = \frac{\partial V}{\partial r} \hat{r}.
\]

A.26. Legendre Transformation

The Legendre transformation is central to the structure of thermodynamics and classical mechanics. In classical mechanics it establishes the correspondence between Lagrangian and Hamiltonian mechanics. Here we follow the presentation of Arnold’s outstanding book [28]. In thermodynamics it is used for two purposes. The first is to change the “natural variables” to a more convenient set of variables. The natural variables are the independent variables that can be controlled in the laboratory. The precise set of natural variables one should use is a choice that depends on the application.
The second purpose is the statement of thermodynamic equilibrium. The second law says entropy should be maximized. However, this can also be shown to mean that energy is minimized, as we will explain below. Legendre transforms can be used to transform nonlinear partial differential equations (PDEs) into linear ones, for which known methods can be used to solve them.

The Legendre transformation of \( f(x) \) is the function \( f^*(p) \) defined by

\[
 f^*(p) = \max_x \left\{ px - f(x) \right\} .
\]

In the case of multiple variables, \( \vec{x} \) and \( \vec{p} \) are vectors and the definition is:

\[
 f^*(\vec{p}) = \max_{\vec{x}} \left\{ \vec{p} \cdot \vec{x} - f(\vec{x}) \right\} .
\]

A.26.1. How to Obtain the Legendre Transform. (Definition) Let

\( y = f(x) \) be a convex function, \( f''(x) > 0 \). The Legendre transformation of the function \( f \) is a new function \( g \) of a new variable \( p \), which is constructed in the following way (Fig. A.5).

- We draw the graph of \( f \) in the \( x, y \) plane.
- Let \( p \) be a given number.
- Consider the straight line \( y = px \).
- We take the point \( x = x(p) \) at which the curve is farthest from the straight line in the vertical direction: for each \( p \) the function \( px - f(x) = F(p, x) \) has a maximum with respect to \( x \) at the point \( x(p) \).
- Now we define the function \( g(p) = F(p, x(p)) \) and call it the Legendre transformation of \( f \), which is basically \( F \) but in which we have substituted \( x = x(p) \) (\( x \) as a function of \( p \)).
The point \( x(p) \) is obtained by solving for \( x \) in the expression obtained from the extremal condition \( \partial F/\partial x = 0 \), i.e. \( f'(x) = p \). Since \( f \) is convex, the point \( x(p) \) is unique (if it exists).

**Example.** Let \( f(x) = x^2 \). Then \( F(p,x) = px - x^2 \), \( x(p) = (1/2)p \), \( g(p) = (1/4)p^2 \).

Example. Let \( f(\dot{x}) = m\dot{x}^2/2 \). Then \( g(p) = p^2/2m \). This can be seen as follows:

\[
F(p,\dot{x}) = \dot{x}p - f(\dot{x}) \quad \Rightarrow \quad p = f'(\dot{x}) = m\dot{x} \quad \Rightarrow \quad \dot{x}(p) = \frac{p}{m}
\]

\[
g(p) = F(p,\dot{x}(p)) = \frac{p^2}{m} - \frac{p^2}{2m} \quad \Rightarrow \quad g(p) = \frac{p^2}{2m}.
\]

Example. Let \( f(x) = x^\alpha/\alpha \). Then \( g(p) = p^\beta/\beta \), where \((1/\alpha) + (1/\beta) = 1 \) (\( \alpha > 1, \beta > 1 \)).

This can be seen as follows:

\[
F(x,p) = xp - \frac{x^\alpha}{\alpha}
\]

\[
F_x = 0 \quad \Rightarrow \quad p = \frac{\alpha x^{\alpha-1}}{\alpha} = x^{\alpha-1} \quad \Rightarrow \quad x(p) = p^{\frac{1}{\alpha-1}}
\]

\[
g(p) = p \cdot p^{\frac{1}{\alpha-1}} - \frac{p^{\frac{\alpha}{\alpha-1}}}{\alpha} = p^{\frac{\alpha}{\alpha-1}} \left( \frac{\alpha - 1}{\alpha} \right) = p^\beta
\]

where \( \beta = \frac{\alpha}{\alpha-1} \) or \((1/\beta) + (1/\alpha) = 1 \).

Example. Let \( f(x) \) be a convex polygon. Then \( g(p) \) is also a convex polygon, in which the vertices of \( f(x) \) correspond to the edges of \( g(p) \), and the edges of \( f(x) \) to the vertices of \( g(p) \). For example, the corner depicted in Fig. A.6 is transformed to a line segment under the Legendre transformation.
A.26.2. Cyclic Property of Partial Derivatives. In order to establish the energy minimization principle (see Section A.27) we also need to introduce the cyclic property of partial derivatives. Suppose that we have a constraint of the form \( F(x, y, z) = 0 \) then \( x, y \) and \( z \) depend on each other. Then we can write:

\[
dx = \left( \frac{\partial x}{\partial y} \right)_z dy + \left( \frac{\partial x}{\partial z} \right)_y dz \quad \text{and} \quad dz = \left( \frac{\partial z}{\partial x} \right)_y dx + \left( \frac{\partial z}{\partial y} \right)_x dy.
\]

Substitution of the first equation into the second one gives:

\[
dz = \left( \frac{\partial z}{\partial x} \right)_y \left[ \left( \frac{\partial x}{\partial y} \right)_z dy + \left( \frac{\partial x}{\partial z} \right)_y dz \right] + \left( \frac{\partial z}{\partial y} \right)_x dy
\]

\[
= \left[ \left( \frac{\partial z}{\partial x} \right)_y \left( \frac{\partial x}{\partial y} \right)_z + \left( \frac{\partial z}{\partial y} \right)_x \right] dy + \left( \frac{\partial z}{\partial x} \right)_y \left( \frac{\partial x}{\partial y} \right)_z dz.
\]

Rearranging gives:

\[
\left[ 1 - \left( \frac{\partial z}{\partial x} \right)_y \left( \frac{\partial x}{\partial z} \right)_y \right] dz = \left[ \left( \frac{\partial z}{\partial x} \right)_y \left( \frac{\partial x}{\partial y} \right)_z + \left( \frac{\partial z}{\partial y} \right)_x \right] dy.
\]

Because \( dz \) and \( dy \) are arbitrary variations, their coefficients must vanish for this equation to hold in the general case. Setting the first term equal to zero gives the reciprocal relation of partial derivatives:

\[
\left( \frac{\partial z}{\partial x} \right)_y = \frac{1}{\left( \frac{\partial x}{\partial z} \right)_y}.
\]

---

\(^7\)For example, the constraint \( x^2 + y^2 + z^2 - 1 = 0 \) defines a sphere.
A.27. Maximizing Entropy vs Minimizing Energy

Setting the second term equal to zero, and applying the reciprocal relation in order to replace \( \frac{\partial x}{\partial y} \) by \( \frac{\partial y}{\partial x} \), gives

\[
\left( \frac{\partial y}{\partial x} \right)_z = -\left( \frac{\partial x}{\partial y} \right)_y.
\]

### A.27. Maximizing Entropy vs Minimizing Energy

The second law says that for isolated systems the entropy is a nondecreasing function that reaches a maximum value at equilibrium. The total energy of the system is \( U(S,X) \), where \( X \) denotes any set of extensive variables (e.g. \( V, N_i \), etc.). Using the inverse function theorem of calculus, we can solve for \( S \) in terms of \( U \) and get \( S(U,X) \). “Maximum” entropy (\( dS \geq 0 \)) means that:

\[
\left( \frac{\partial S}{\partial X} \right)_U = 0 \quad \text{and} \quad \left( \frac{\partial^2 S}{\partial X^2} \right)_U < 0 \quad \text{at equilibrium}
\]

From calculus we have that

\[
\left( \frac{\partial U}{\partial X} \right)_S = -\left( \frac{\partial S}{\partial X} \right)_U = -T \left( \frac{\partial S}{\partial U} \right)_X = 0.
\]

Thus, the energy is at an extremum at equilibrium. We can also show that

\[
\left( \frac{\partial^2 U}{\partial X^2} \right)_S = -T \left( \frac{\partial^2 S}{\partial U^2} \right)_U.
\]

In terms of the Helmholtz potential, \( A(T,V,N_i) = U - TS \), we start from \( U \) and replace \( S \) by \( T \). The transformation is:

\[
\tilde{A}(T,V,N_i) = \max_S \{ U(S,V,N_i) - TS \},
\]

where max is the maximum over all possible \( S \). We write \( \tilde{A} \) to indicate that \( A \) was obtained by maximization. This maximum will be reached whenever \( T \) is equal to the temperature \( \partial S U \), i.e., \( T = \left( \frac{\partial U}{\partial S} \right)_{V,N_i} \). The particular choice of \( S \) that fulfills the condition \( T = \left( \frac{\partial U}{\partial S} \right)_{V,N_i} \) is denoted \( S(T) \). The transform is written as:

\[
\tilde{A}(T,V,N_i) = U - TS
\]

where \( S \) is understood to stand for \( S(T) \) obtained from the condition \( T = \left( \frac{\partial U}{\partial S} \right)_{V,N_i} \). We will furthermore denote \( S \) by \( \tilde{S} \), to indicate that \( S \) is the value of the entropy at equilibrium. If we combine this with the fact that
at equilibrium, the internal energy is minimized

$$U(\tilde{S}) = \min_X (U(\tilde{S}, X)),$$

we may write the Helmholtz potential as:

$$\tilde{A} = \max_S \left\{ \min_X (U(S, X)) - TS \right\} = \min_X \left\{ \max_S (U(S, X) - TS) \right\},$$

where in the second equality we have interchanged the order of the min and max. But the bracketed term is simply $\tilde{A}(T, X)$. Thus, we conclude that

$$\tilde{A} = \min_X (\tilde{A}(T, X))$$

and the Helmholtz free energy is minimized at equilibrium.

Likewise, for the enthalpy

$$H(S, P) = U + PV$$

we have

$$\tilde{H} = \max_V \left\{ \min_X (U(V, X)) + PV \right\} = \min_X \left\{ \max_V (U(V, X) + PV) \right\},$$

so that

$$\tilde{H} = \min_X (\tilde{H}(P, X)).$$

For the Gibbs free energy we have a double transformation, or equivalently a transformation in two variables where $V$ is replaced by $P$ and $S$ is replaced by $T$:

$$G(P, T) = U + PV - TS.$$  

Using the formula transformation in two variables:

$$\tilde{G} = \max_{S,V} \left\{ \min_X (U(S, V, X)) + PV - TS \right\}$$

$$= \min_X \left\{ \max_{S,V} (U(S, V, X) + PV - TS) \right\},$$

we find

$$\tilde{G} = \min_X (\tilde{G}(P, T, X)).$$

Thus, at thermal equilibrium the Gibbs free energy is minimized, the Helmholtz free energy is minimized and the enthalpy is also minimized.

A.27.1. Legendre Transformation in Classical Mechanics. In classical mechanics, we begin with the Lagrangian, $L$, which is a function of particle position and velocities, $L \equiv L(q, \dot{q}, t)$, where $q$ denotes particle generalized positions and $\dot{q}$ denotes the generalized velocities:

$$q = (q_1^1, q_1^2, q_1^3, \ldots, q_N^1, q_N^2, q_N^3).$$
where \( N \) is the number of particles. Similarly for \( \dot{q} \), we have a \( 3N \)-dimensional vector of generalized velocities. The Lagrange equations of motion are the Euler-Lagrange equations:\(^8\)

\[
\frac{d}{dt} \left( \frac{\partial L}{\partial \dot{q}_j} \right) - \frac{\partial L}{\partial q_j} = 0.
\]

In general, \( L \) is the difference between kinetic (\( T \)) and potential (\( V \)) energies: \( L = T - V \). For example, a particle with kinetic energy \( T = \frac{1}{2}m\dot{z}^2 \) and gravitational potential energy \( V = -mgz \) has Lagrangian \( L = T - V = \frac{1}{2}m\dot{z}^2 + mgz \). The Euler-Lagrange equations of motion become

\[
0 = \frac{d}{dt} \left( \frac{\partial L}{\partial \dot{q}_j} \right) - \frac{\partial L}{\partial q_j} = \frac{d}{dt}(m\dot{z}) - mg = m\ddot{z} - mg,
\]

which yields Newton’s law (\( F = ma \), or more succinctly, \( \ddot{z} = g \). This equation is second-order in time. Sometimes, equations that are first-order in time may offer certain advantages. This is possible if we define the Hamiltonian as the Legendre transform of the Lagrangian:

\[
(A.1) \quad H(p, q, t) = p \cdot \dot{q}(p) - L(q, \dot{q}(p), t),
\]

where \( p \cdot \dot{q} \) denotes the dot product of

\[
p = (p^1_x, p^1_y, p^1_z, \ldots, p^N_x, p^N_y, p^N_z)
\]

with

\[
\dot{q} = (\dot{q}^1_x, \dot{q}^1_y, \dot{q}^1_z, \ldots, \dot{q}^N_x, \dot{q}^N_y, \dot{q}^N_z),
\]

---

\(^8\)These equations are obtained from the principle of least action, \( \delta S = 0 \), where \( S[q(t)] = \int_{t_1}^{t_2} L(q(t), \dot{q}(t), t) dt \). \( \delta S \) is obtained by perturbing its argument \( q \): \( \delta S[q(t)] = S[q(t) + \delta q(t)] - S[q(t)] \) and Taylor-expanding the first term. The exact nature of the variation \( \delta q(t) \) is important; here, we consider variations in the path \( \delta q(t) \) where the initial and final points are unchanged: \( \delta q(t_1) = \delta q(t_2) = 0 \). Taylor expanding the functional to first order and integrating by parts,

\[
\delta S = \int_{t_1}^{t_2} \delta L(q(t), \dot{q}(t), t) dt = \int_{t_1}^{t_2} \{ L((q + \delta q)(t), (\dot{q} + \delta \dot{q})(t), t) \} dt - \int_{t_1}^{t_2} L(q(t), \dot{q}(t), t) dt
\]

\[
= \int_{t_1}^{t_2} \left\{ L(q(t), \dot{q}(t), t) + \delta q \frac{\partial L}{\partial \dot{q}} + \delta \dot{q} \frac{\partial L}{\partial q} + \ldots \right\} dt - \int_{t_1}^{t_2} L(q(t), \dot{q}(t), t) dt
\]

\[
= \int_{t_1}^{t_2} \left\{ \frac{\partial L}{\partial \dot{q}} - \frac{d}{dt} \frac{\partial L}{\partial q} \right\} \delta q dt = 0
\]

(The surface term from integration by parts vanished due to the boundary condition \( \delta q(t_1) = \delta q(t_2) = 0 \).) Since \( \delta S = 0 \) should hold for arbitrary \( \delta q(t) \), the integrand must vanish, yielding the Euler-Lagrange equation \( \frac{d}{dt} \frac{\partial L}{\partial \dot{q}} - \frac{\partial L}{\partial q} = 0 \).
and \( \dot{q}(p) \) denotes the velocity expressed in terms of the momentum. The quantity \( \dot{q}(p) \) is obtained from the equation\(^9\)

\[
p = \frac{\partial L}{\partial \dot{q}},
\]

which, for the case of a kinetic energy term \( \frac{1}{2}m\dot{z}^2 \) yields \( p = m\dot{z} \), or \( \dot{z}(p) = p/m \), the usual concept of linear momentum. The equations of motion are obtained by taking the total differential of (A.1):

\[
dH = \sum \left[ \left( \frac{\partial H}{\partial q_i} \right) dq_i + \left( \frac{\partial H}{\partial p_i} \right) dp_i \right] + \left( \frac{\partial H}{\partial t} \right) dt
\]

\[
= \sum \left[ \left( -\frac{\partial L}{\partial q_i} \right) dq_i + \left( \dot{q}_i + p_i \frac{\partial \dot{q}_i}{\partial p_i} - \frac{\partial L}{\partial q_i} \frac{\partial \dot{q}_i}{\partial p_i} \right) dp_i \right] - \left( \frac{\partial L}{\partial t} \right) dt
\]

Substituting the definition of conjugate momenta, \( p_i = \frac{\partial L}{\partial \dot{q}_i} \),

\[
dH = \sum \left[ -\frac{\partial L}{\partial q_i} dq_i + \dot{q}_i dp_i + p_i dq_i - \frac{\partial L}{\partial q_i} \frac{\partial \dot{q}_i}{\partial p_i} \right] - \left( \frac{\partial L}{\partial t} \right) dt.
\]

and equating the coefficients of like differentials between (A.3) and (A.2), we get the Hamilton equations of motion which are first-order in time\(^{10}\):

\[
\dot{p}_i = -\frac{\partial H}{\partial q_i}, \quad \dot{q}_i = \frac{\partial H}{\partial p_i}, \quad \frac{\partial H}{\partial t} = -\frac{\partial L}{\partial t}.
\]

You can check, for example, that the Lagrangian

\[
L(q, \dot{q}) = \sum_i \frac{1}{2} m_i \dot{q}_i^2 - U(q)
\]

transforms to the Hamiltonian:

\[
H(q, p) = \sum_i \frac{p_i^2}{2m_i} + U(q).
\]

The Hamiltonian is a convex function of the momenta, just like the Lagrangian is a convex function of the velocities.

---

\(^9\)With this definition of the conjugate momenta, \( p_j = \frac{\partial L}{\partial \dot{q}_j} \), the Euler-Lagrange equation,

\[
\frac{d}{dt} \left( \frac{\partial L}{\partial \dot{q}_j} \right) - \frac{\partial L}{\partial q_j} \text{ reads } \dot{p}_j = \frac{\partial L}{\partial \dot{q}_j}.
\]

\(^{10}\)These Hamilton equations, \( \dot{p}_i = -\frac{\partial H}{\partial q_i}, \dot{q}_i = \frac{\partial H}{\partial p_i}, \frac{\partial H}{\partial t} = -\frac{\partial L}{\partial t} \), are to be contrasted with the Lagrangian equations, \( p_j = \frac{\partial L}{\partial \dot{q}_j} \) and \( \dot{p}_j = \frac{\partial L}{\partial \dot{q}_j} \).
A.27.2. Legendre Transformation in Thermodynamics. We have seen how the Legendre transform preserves convexity and therefore plays a role in energy minimization of thermodynamic potentials at equilibrium. Chemists also use the Legendre transform because it enables us to change the independent variables to a set of variables that is more convenient. Those are called the “natural variables”.

The differential of the internal energy $U(S,V,N)$ is

$$dU = \left( \frac{\partial U}{\partial S} \right)_{V,N} dS + \left( \frac{\partial U}{\partial V} \right)_{S,N} dV + \left( \frac{\partial U}{\partial N} \right)_{S,V} dN = T dS - P dV + \mu dN,$$

where

$$T \equiv \left( \frac{\partial U}{\partial S} \right)_{V,N}, \quad P \equiv -\left( \frac{\partial U}{\partial V} \right)_{S,N}, \quad \mu \equiv \left( \frac{\partial U}{\partial N} \right)_{S,V}$$

to the Helmholtz free energy

$$A(T,V,N) = U - TS,$$

$$dA = dU - S dT - T dS = -S dT - P dV + \mu dN,$$

Thus, we have changed the independent variables from $(S,V,N)$ to $(T,V,N)$. This is the case because the expression for $dA$ contains the differentials $dT$, $dV$ and $dN$ whereas $dU$ contained the differentials $dS$, $dV$ and $dN$. Please note that since we have swapped the variables $S$ and $T$ by way of the Legendre transform, the new variable $T$ is required to be equal to the energy $\partial_S U$:

$$T \equiv \left( \frac{\partial U}{\partial S} \right)_{V,N}.$$

(This was the condition for finding the function $x(p)$ in the formal definition of the Legendre transformation, see pages 1 and 2 of these lecture notes.) Finally, the convexity of the function has been preserved, i.e. the new thermodynamic potential is convex.

There is also the enthalpy $H(S,P,N) = U + PV$, whose differential becomes

$$dH = T dS + V dP + \mu dN$$

or to the free enthalpy (Gibbs’ potential) $G(T,P,N) = U - TS + PV$, whose differential is

$$dG = dU - T dS - S dT + P dV + V dP = -S dT + V dP + \mu dN.$$

A thermodynamic potential is chosen according to which variables are the most convenient to use from an experimental standpoint. For example, if pressure is held constant in an experiment, then $G$ is the potential of choice. If volume is held constant, then $A$ is convenient to use.

\[\text{Notes:}\]

11 We note that $A = U - TS$ differs from the above-defined Legendre transformation by a minus sign. The origins of this sign difference is historical. In spite of this sign difference, the technique is the same.
A.27.3. Involutivity of the Legendre transformation. The Legendre transform to be involutive, meaning that applying twice recovers the original function without loss of information. Consider a function $f$ which is differentiable as many times as necessary, with $f''(x) > 0$. It is easy to verify that a Legendre transformation takes convex functions to convex functions. Therefore, we can apply it twice.

**Theorem.** The Legendre transformation is involutive, i.e., its square is the identity: if under the Legendre transformation $f$ is taken to $g$, then the Legendre transform of $g$ will again be $f$.

**Proof.** In order to apply the Legendre transform to $g$, with variable $p$, we must by definition look at a new independent variable (which we call $x$), construct the function

$$G(x, p) = xp - g(p),$$

and find the point $p(x)$ at which $G$ attains its maximum: $\partial G/\partial p = 0$, i.e., $g'(p) = x$. Then the Legendre transform of $g(p)$ will be the function of $x$ equal to $G(x, p(x))$.

We will show that $G(x, p(x)) = f(x)$. To this end we notice that $G(x, p) = xp - g(p)$ has a simple geometric interpretation: it is the ordinate of the point with abscissa $x$ on the line tangent to the graph of $f(x)$ with slope $p$ (Fig. A.7). For fixed $p$, the function $G(x, p)$ is a linear function of $x$, with $\partial G/\partial x = p$, and for $x = x(p)$ we have $G(x, p) = xp - g(p) = f(x)$ by the definition of $g(p)$.

Let us now fix $x = x_0$ and vary $p$. Then the values of $G(x, p)$ will be the ordinates of the points of intersection of the line $x = x_0$ with the line tangent to the graph of $f(x)$ with various slopes $p$. By the convexity of the graph it follows that all these tangents lie below the curve, and therefore the maximum of $G(x, p)$ for a fixed $x(p_0)$ is equal to $f(x)$ (and is achieved for $p = p(x_0) = f'(x_0)$).

A.27.4. Inequalities. By definition of the Legendre transform, $F(x, p) = px - f(x)$ is less than or equal to $g(p)$ for any $x$ and $p$. From this we have Young’s inequality:

$$px \leq f(x) + g(p).$$

**Example.** If $f(x) = \frac{1}{2}x^2$, then $g(p) = \frac{1}{2}p^2$ and we obtain the well-known inequality $px \leq \frac{1}{2}x^2 + \frac{1}{2}p^2$ for all $x$ and $p$.

**Example.** If $f(x) = x^\alpha/\alpha$, $g(p) = p^\beta/\beta$, where $(1/\alpha) + (1/\beta) = 1$, and we obtain Young’s inequality $px \leq (x^\alpha/\alpha) + (p^\beta/\beta)$ for all $x > 0$, $p > 0$, $\alpha > 1$, $\beta > 1$. 
\( \beta > 1 \) and \( (1/\alpha) + (1/\beta) = 1 \).

**Example.** Let \( f \) by a quadratic form \( f(x) = \sum f_{ij} x_i x_j \). The reader can show that its Legendre transform is again a quadratic form \( g(p) = \sum g_{ij} p_i p_j \), and that the values of both forms at corresponding points coincide:

\[
\begin{align*}
g(p(x)) &= f(x) \\
g(p(x)) &= f(x).
\end{align*}
\]

A.28. Problems

**Problem 131.** You are a member of a club. (a) The club consists of only 50 members. In how many ways can a president, vice president, secretary, and treasurer be chosen (amongst the 50 members)? (b) In how many ways can a committee of 4 members be chosen? (Hint: in the first case, order matters; in the second case, order does not matter.)

**Solution.** (A) \( \frac{50!}{(50-4)!} \), (B) \( \frac{50!}{4!} \)

**Problem 132.** Two students are working separately on the same homework problem. If the first student has probability \( \frac{1}{2} \) of solving it and the second student has probability \( \frac{3}{4} \) of solving it (due to higher IQ), what is the probability that at least one of them solves it given that the probability of both of them solving it simultaneously is \( \frac{3}{8} \)?

**Solution.** For overlapping events,

\[
P(A \text{ or } B) = P(A) + P(B) - P(A \text{ and } B),
\]

\[
P(A \text{ or } B) = \frac{1}{2} + \frac{3}{4} - \frac{3}{8}
\]
Problem 133. Find the Taylor series for \( f(x) = \frac{1}{1 - x} \) about \( x = 0 \), where \(|x| < 1\). Explicitly compute all the terms needed to arrive at a general expression for the Taylor series of \( f \) at \( x \).

Solution.
- First term \( \rightarrow f(x = 0) = \frac{1}{1 - 0} = 1 \)
- Second term \( \rightarrow \left( \frac{df}{dx} \right)_{x=0} x = (-1)(1 - 0)^{-2}(-1)x = x \)
- Third term \( \rightarrow \left( \frac{1}{2!} \left( \frac{d^2f}{dx^2} \right)_{x=0} \right) x^2 = \left( \frac{1}{2} \right)2(1 - 0)^{-3}x^2 = x^2 \)

\[
\frac{1}{1 - x} = 1 + x + x^2 + x^4 + \cdots = \sum_{n=0}^{\infty} x^n
\]

Problem 134. The osmotic pressure \( \Pi \) can be described by the function:

\[
\Pi(x_2) = -RT \log(1 - x_2)
\]

where \( 0 < x_2 < 1 \) is the mole fraction of the solute, \( R > 0 \) is the universal gas constant and \( T > 0 \) is the thermodynamic temperature also held at a constant value.

(a) Write down the Taylor expansion of the function that defines \( \Pi \) at \( x_2 = a \).

(b) Use Taylor series

\[
\log(1 + x) = x - \frac{x^2}{2} + \frac{x^3}{4} - \frac{x^4}{5} + \cdots
\]

to prove that when \( x_2 \ll 1 \), \( \Pi \) is proportional to \( x_2 \) at a certain temperature.

Solution. (a) Taylor expansion of the function is

\[
\Pi(x_2) = -RT \log(1 - x_2)
\]

\[
= RT \left[ \log(1 - a) - \frac{x - a}{1 - a} - \frac{1}{2} \frac{(x - a)^2}{(1 - a)^2} - \frac{1}{3} \frac{(x - a)^3}{(1 - a)^3} - \frac{1}{4} \frac{(x - a)^4}{(1 - a)^4} - \cdots \right]
\]

\[
= RT \left[ \log(1 - a) - \sum_{n=1}^{\infty} \frac{1}{n} \frac{(x - a)^n}{(1 - a)^n} \right]
\]

(b) With \(-1 < x \leq 1\), when \( x_2 \ll 1 \), we can neglect the higher-order terms in the Taylor expansion and using

\[
\log(1 + x) \approx x
\]

we obtain

\[
\log(1 - x_2) \approx -x_2
\]
and
\[ \Pi(x_2) \approx x_2 RT \]

**Problem 135.** In the ground state of the hydrogen atom, the PDF describing the likelihood of finding the electron at \( r \) is:
\[
p(r) = \frac{1}{\pi a_0^3} e^{-2r/a_0}
\]
where \( a_0 \) is Bohr radius. Compute the average potential energy
\[
V(r) = -\frac{e^2}{4\pi \epsilon_0 r}
\]
Express the result finally in the units of eV. (\( \epsilon_0 = 8.8542 \times 10^{-12} \text{ F/m} \), \( a_0 = 5.2918 \times 10^{-11} \text{ m} \), \( e = 1.6022 \times 10^{-19} \text{ C} \))

**Solution.**
\[
EV = -\frac{e^2}{4\pi \epsilon_0 a_0^3} \int_{r=0}^{\infty} e^{-2r/a_0} \frac{1}{r} 4\pi r^2 dr = -\frac{e^2}{\pi \epsilon_0 a_0^3} \int_{r=0}^{\infty} e^{-2r/a_0} r dr
\]
\[
= -\frac{e^2}{\pi \epsilon_0 a_0^3} \times \frac{a_0^2}{4} = -\frac{e^2}{4\pi \epsilon_0 a_0} = -27.2 \text{ eV}
\]

**Problem 136.** The mean value of a random variable \( X \) taking values in discrete space can be calculated from the weighted sum:
\[
\mathbb{E} X = \sum_{i=1}^{N} P(X = x_i) x_i
\]
where \( x_i \) ranges over all the allowed values of \( X \), e.g. 1, 2, ..., \( N \), whereas \( P(X = x) \) is the probability of that \( X \) takes the value \( x \). Calculate the average of an eight-sided die roll.

**Solution.**
\[
\mathbb{E} X = \sum_{i=1}^{8} (1/8) x_i = (1/8) \times \sum_{i=1}^{8} x_i = 36/8 = 4.5
\]

**Problem 137.** Let \( X \) be a random variable with PDF \( f(x) \). The mean value of \( X \), when \( X \) takes continuous (not discrete) values in the range \( (-\infty, \infty) \), can be calculated by an integral:
\[
\mathbb{E} X = \int_{-\infty}^{\infty} f(x) x dx
\]
where \( x \) is the value of the random variable \( X \). Calculate the average position \( X \) of a particle with the following PDF (if it is not normalized, normalize it):
\[
f(x) = \begin{cases} 
  x^2 + 4 & -4 < x < 0 \\
  x + 4 & 0 < x < 4
\end{cases}
\]

**Solution.** First, normalize the probability distribution function (normalization factor of \( A \)):
1 = A(\int_{-4}^{0} x^2 + 4dx + \int_{0}^{4} x + 4dx)
1 = A(\frac{x^3}{3} + 4x|_{-4}^{0} + \frac{x^2}{2} + 4x|_{0}^{4})

1 = A(64/3 + 16 + 16/2 + 16), \quad 1 = \frac{184A}{3}, \quad A = \frac{3}{184}

Now, the average:

\begin{align*}
\mathbb{E}X &= \frac{3}{184}(\int_{-4}^{0} x(x^2 + 4)dx + \int_{0}^{4} x(x + 4)dx) \\
\mathbb{E}X &= \frac{3}{184}(\int_{-4}^{0} x^3 + 4xdx + \int_{0}^{4} x^2 + 4xdx) \\
\mathbb{E}X &= \frac{3}{184}(\frac{x^4}{4} + 2x^2|_{-4}^{0} + \frac{x^3}{3} + 2x^2|_{0}^{4}) \\
\mathbb{E}X &= \frac{3}{184}(-64 - 32 + 64/3 + 32) \\
\mathbb{E}X &= \frac{3}{184}(-64 - 32 + 64/3 + 32) \\
\mathbb{E}X &= \frac{-16}{23}
\end{align*}

Problem 138. Taylor-expand to low order (just a few terms needed) the Coulomb interaction energy about the point \( R = r_0 \), where \( r_0 \) is the molecular radius. In the Coulomb interaction, the fall-off with distance is described by the \( 1/R \) dependence (which gets smaller as \( R \) increases). In the Taylor expansion, however, only positive powers of \( R \) occur. Show how the Taylor series describes the fall-off of the interaction energy with distance.

Solution. From \( V(R) = \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{R} \), we get, from Taylor expansion about \( r_0 \) using the “small parameter” \( R - r_0 \):

\[
V(R) = \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{r_0} - \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{r_0^3} (R - r_0) + \frac{2}{4\pi\varepsilon_0} \frac{q_1 q_2}{r_0^5} (R - r_0)^2 + \ldots
\]

The zeroth order term, \( \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{r_0} \) is constant (doesn’t change with \( R \). As \( R \) increases beyond \( r_0 \), \( R - r_0 \) is positive, but the first order term has an overall negative sign, and this reflects the decrease in interaction energy as the intermolecular distance increases. The second-order term is positive, but its magnitude is smaller than the first-order term.

The problems below are related to probability. CDF stands for cumulative distribution function, \( \mathbb{P}(X < x) \), where \( X \) is a random variable. The CDF related to the PDF, \( p(x) \), according to \( \mathbb{P}(X < x) = \int_{-\infty}^{x} p(x)dx \). (The lower limit may be different from \(-\infty\); it is the lowest value that \( X \) is allowed to
The CDF can be used to compute probabilities between two limits, e.g., \( \mathbb{P}(a < X < b) = \mathbb{P}(X < b) - \mathbb{P}(X < a) \).

**Problem 139.** Let \( X \) be a continuous random variable with PDF

\[
p(x) = \begin{cases} 
  kx^2 & 0 < x < 1 \\
  0 & \text{otherwise}
\end{cases}
\]

(i) Determine the value of \( k \) and sketch \( p(x) \)

(ii) Derive a formula for the corresponding CDF. (You can assume that the lower limit of integration for the CDF is 0.) Sketch the CDF.

(iii) Find \( \mathbb{P}(\frac{1}{4} < X \leq 2) \).

(iv) Compute the mean and variance of the random variable \( X \) associated with this PDF.

**Solution.** (i) We must have \( k > 0 \). The normalization condition yields:

\[
\int_0^1 kx^2 \, dx = k \frac{x^3}{3} \bigg|_0^1 = \frac{k}{3} = 1
\]

which gives \( k = 3 \). Thus,

\[
p(x) = \begin{cases} 
  3x^2 & 0 < x < 1 \\
  0 & \text{otherwise}
\end{cases}
\]

(ii) The CDF is

\[
\mathbb{P}(X < x) = \begin{cases} 
  0 & x < 0 \\
  \int_0^x 3\xi^2 \, d\xi = x^3 & 0 \leq x < 1 \\
  \int_0^1 3\xi^2 \, d\xi = 1 & 1 \leq x
\end{cases}
\]

These functions are sketched below:
(iii) \( P(\frac{1}{4} < X \leq 2) = P(X < 2) - P(X < \frac{1}{4}) = 1 - (\frac{1}{4})^3 = \frac{63}{64}. \) (Note: \( P(X < 2) = P(X \leq 2) \) because the distribution function is continuous; so we need not worry about the equal sign.)

(iv) We need to compute:

\[
\mathbb{E}X = \int x p(x) \, dx = \int_0^1 x (3x^2) \, dx = \left. \frac{3x^4}{4} \right|_0^1 = 3/4.
\]

\[
\mathbb{E}(X^2) = \int x^2 p(x) \, dx = \int_0^1 x^2 (3x^2) \, dx = \left. \frac{3x^5}{5} \right|_0^1 = 3/5.
\]

Variance is

\[
\mathbb{E}(X^2) - (\mathbb{E}X)^2 = 3/5 - (3/4)^2 = 3/80 = 0.0375.
\]

\[\blacksquare\]

**Problem 140.** The Stefan-Boltzman law says that the power irradiated by a black body is proportional to the fourth power of its temperature:

\[ p = cT^4 \]

where \( c \) is a constant. This law has been used to measure the temperature of stars in the galaxy. By measuring the light output (power), we can infer the value of \( T \). Suppose that \( p \) is Gaussian-distributed with mean \( \mu \) and variance \( \sigma^2 \), find the distribution of \( T \). Obtain both the CDF and PDF. In particular, show that you can obtain the PDF from the CDF and vice-versa.

**Solution.** We need to find an expression for:

\[ P(T < t) = P((p/c)^{1/4} < t) = P(p < ct^4) \]

and since \( p \) follows a normal law \( N(\mu, \sigma) \),

\[ P(T < t) = P(p < ct^4) = \frac{1}{\sqrt{2\pi\sigma^2}} \int_{-\infty}^{ct^4} e^{-(\xi - \mu)^2/2\sigma^2} \, d\xi = \text{Erf}(ct^4\mu, \sigma). \]

This is the CDF. The PDF is obtained by differentiating with respect to \( t \):

\[ p_T(t) = \frac{dP(T < t)}{dt} = \frac{4c\tau^3}{\sqrt{2\pi\sigma^2}} e^{-(c\tau^4 - \mu)^2/2\sigma^2} \]

We can check that integrating the PDF from \(-\infty\) to \( t \) gives the CDF

\[ \int_{-\infty}^{t} p_T(\tau) \, d\tau = \int_{-\infty}^{ct^4} \frac{4c\tau^3}{\sqrt{2\pi\sigma^2}} e^{-(c\tau^4 - \mu)^2/2\sigma^2} \, d\tau \]

\[ = \int_{-\infty}^{ct^4} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-(\xi - \mu)^2/2\sigma^2} \, d\xi = P(p < ct^4) = P(T < t). \]

In the second equality we made a change of variables \( \xi = c\tau^4, \, d\xi = 4c\tau^3 \, d\tau. \) \[\blacksquare\]
Problem 141. Suppose that we have

\[ F(t) = A \cos(\omega t) \]

where \( \omega \) is a normal random variable with mean \( \omega_0 \) and variance \( \Delta \omega^2 \). Find the properties of the random variable \( F \), i.e. derive the PDF, CDF, its mean and standard deviation (or variance).

Solution. We note that

\[ \omega = t^{-1} \cos^{-1}(F/A) \]

Then,

\[
P(F < f) = P(\omega < t^{-1} \cos^{-1}(f/A)) = \int_{-\infty}^{t^{-1} \cos^{-1}(f/A)} \frac{1}{\sqrt{2\pi \Delta \omega^2}} e^{-\frac{(w-\omega_0)^2}{2\Delta \omega^2}} dw.
\]

The PDF is obtained by differentiating the CDF with respect to \( f \):

\[
p_F(f) = \frac{dP(F < f)}{df} = \frac{1}{\sqrt{2\pi \Delta \omega^2}} e^{-\frac{(t^{-1} \cos^{-1}(f/A)-\omega_0)^2}{2\Delta \omega^2}} \frac{d}{df} t^{-1} \cos^{-1}(f/A)
\]

\[
= -\frac{t^{-1}}{\sqrt{2\pi \Delta \omega^2}} e^{-\frac{(t^{-1} \cos^{-1}(f/A)-\omega_0)^2}{2\Delta \omega^2}} \frac{1}{\sqrt{1-\frac{f^2}{A^2}}} \cdot \frac{1}{A}.
\]

Problem 142. Let the PDF of a continuous r.v. \( X \) be given by:

\[
p(x) = \frac{1}{\sqrt{2\pi \sigma^2}} e^{-\frac{(x-\mu)^2}{2\sigma^2}}
\]

Prove, via direct calculation of the integrals, that the mean and variance of \( X \) are: \( EX = \mu \) and \( var(X) = \sigma^2 \).

Solution. The mean is

\[
EX = \frac{1}{\sqrt{2\pi \sigma^2}} \int_{-\infty}^{\infty} xe^{-\frac{(x-\mu)^2}{2\sigma^2}} dx
\]

Writing \( x \) as \((x - \mu) + \mu\), we have

\[
EX = \frac{1}{\sqrt{2\pi \sigma^2}} \int_{-\infty}^{\infty} (x - \mu)e^{-\frac{(x-\mu)^2}{2\sigma^2}} dx + \mu \frac{1}{\sqrt{2\pi \sigma^2}} \int_{-\infty}^{\infty} e^{-\frac{(x-\mu)^2}{2\sigma^2}} dx
\]

Letting \( y = x - \mu \) in the first integral we obtain

\[
EX = \frac{1}{\sqrt{2\pi \sigma^2}} \int_{-\infty}^{\infty} ye^{-\frac{y^2}{2\sigma^2}} dy + \mu \int_{-\infty}^{\infty} p(x) dx
\]

The first integral is zero since the integrand is an odd function. In the second integral, we used the normalization condition (\( \int p(x) dx = 1 \)). Thus,

\[
EX = \mu
\]

Now for the variance:

\[
var(X) = (X - \mu)^2 = \frac{1}{\sqrt{2\pi \sigma^2}} \int_{-\infty}^{\infty} (x - \mu)^2 e^{-\frac{(x-\mu)^2}{2\sigma^2}} dx
\]
We note that
\[ \int_{-\infty}^{\infty} e^{-(x-\mu)^2/(2\sigma^2)} \, dx = \sigma \sqrt{2\pi}. \]
Differentiating with respect to \( \sigma \), we obtain:
\[ \int_{-\infty}^{\infty} \frac{(x-\mu)^2}{\sigma^3} e^{-(x-\mu)^2/(2\sigma^2)} \, dx = \sqrt{2\pi}. \]
Multiplying both sides by \( \sigma^2 / \sqrt{2\pi} \), we have
\[ \frac{1}{\sqrt{2\pi}} \int_{-\infty}^{\infty} \frac{(x-\mu)^2}{\sigma^3} e^{-(x-\mu)^2/(2\sigma^2)} \, dx = \sigma \]
Thus,
\[ \text{var}(X) = \sigma^2 \]

**Problem 143.** Let \( Y = Ae^X + B \). (a) Find the distribution of \( Y \) when \( X \) is a uniform r.v. over the interval \((0,1)\), i.e. its density takes the value 1 over the interval and 0 elsewhere.
(b) Find the distribution of \( Y \) when \( X \) follows a normal law with mean \( \mu \) and variance \( \sigma^2 \).

**Solution.** The PDF of \( X \), as stated in the problem, is:
\[ p_X(x) = \begin{cases} 
1 & 0 < x < 1 \\
0 & \text{otherwise}
\end{cases} \]
From the definition of a CDF, \( \mathbb{P}(Y < y) = \int_{-\infty}^{y} p(y) \, dy \), and differentiating it with respect to \( y \), we get that \( p(y) = (d/dy)\mathbb{P}(Y < y) \). Next, we need to know \( \mathbb{P}(Y < y) \). The CDF of \( Y \) is (for \( A + B < y < Ae + B \))
\[ \mathbb{P}(Y < y) = \mathbb{P}(Ae^X + B < y) = \mathbb{P}(X < \log A^{-1}(y-B)) = \int_{-\infty}^{\log A^{-1}(y-B)} p_X(x) \, dx \]
\[ = \int_{0}^{\log A^{-1}(y-B)} dx = \log A^{-1}(y-B) \]
To get the PDF of \( Y \) we differentiate the CDF with respect to \( y \):
\[ p_Y(y) = \frac{d\mathbb{P}(Y < y)}{dy} = \frac{d}{dy} \log A^{-1}(y-B) = \frac{A^{-1}}{A^{-1}(y-B)} = \frac{1}{y-B} \]
for \( A + B < y < Ae + B \) and 0 otherwise.
From the definition of a CDF, \( \mathbb{P}(Y < y) = \int_{-\infty}^{y} p(y) \, dy \), and differentiating it with respect to \( y \), we get that \( p(y) = (d/dy)\mathbb{P}(Y < y) \). Next, we need to know \( \mathbb{P}(Y < y) \). Since \( Y = Ae^X + B \), the probability
\[ \mathbb{P}(Y < y) = \mathbb{P}(Ae^X + B < y) = \mathbb{P}(Ae^X < y - B) = \mathbb{P}(e^X < (y - B)/A) \]
\[ = \mathbb{P}(X < \log[(y - B)/A]) \]
And because $X$ follows a normal law with mean $\mu$ and variance $\sigma^2$,

$$P(Y < y) = \int_{-\infty}^{\log[(y-B)/A]} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{(x-\mu)^2}{2\sigma^2}} \, dx$$

From this, we have:

$$\frac{dP(Y < y)}{dy} = \frac{d}{dy} \left[ \int_{-\infty}^{\log[(y-B)/A]} \frac{1}{\sqrt{2\pi\sigma^2}} e^{-\frac{(x-\mu)^2}{2\sigma^2}} \, dx \right]$$

where we made use of the formula for differentiation of an integral (only term 2 applies, the other two are zero):

$$\frac{d}{dy} \left( \int_{a(y)}^{b(y)} f(x, y) \, dx \right) = \int_{a(y)}^{b(y)} \frac{\partial}{\partial y} f(x, y) \, dx + \frac{f(b(y), y) \cdot b'(y)}{2} - \frac{f(a(y), y) \cdot a'(y)}{3}$$

\[\square\]

**Problem 144.** Normalize the probability function given by $f(x) = \cos(x)$ so that it becomes a probability density function $p(x)$ (shown below) over the interval $[-\pi/2, \pi/2]$. What is the normalization constant? Calculate the mean (expectation value) and standard deviation $\sigma$ of a random variable $X$ defined in the interval $[-\pi/2, \pi/2]$, i.e. calculate $E_X = \int_{-\pi/2}^{\pi/2} xp(x) \, dx$ and the square root of the variance, $\sigma^2 = \text{var}(X) = \int_{-\pi/2}^{\pi/2} (x - E_X)^2 p(x) \, dx$.

(Note: I write $X$ for the random variable, and $x$ for its numerical value.)

![Graph of cos(x) from -\pi/2 to \pi/2]

\[\text{Solution.}\] A PDF is said to be normalized if it can be described by the following relationship

$$\int f(x) \, dx = 1.$$  

The integral is calculated over all space where $X$ takes its values. For example, $X$ could be defined over the interval $[-\infty, \infty]$, or $[-\pi/2, \pi/2]$. Solving for the normalization constant, $C$

$$\int_{-\pi/2}^{\pi/2} C \cos(x) \, dx = 1, \quad C = 1/2.$$
The mean value, $E_X$ is

$$E_X = \int_{\mathbb{R}} xp(x) dx$$

If $p(x) = (1/2) \cos(x)$ then $E_X$ is zero, since $x$ is odd and $\cos(x)$ is even:

$$E_X = \int_{-\pi/2}^{\pi/2} x(1/2) \cos(x) dx = \cos(x) + x \sin(x)]_{-\pi/2}^{\pi/2} (1/2) = 0$$

The standard deviation $\sigma$ is the square root of the variance $\sigma^2$

$$\sigma^2 = \int_{\mathbb{R}} p(x)(x - E_X)^2 dx = \int_{-\pi/2}^{\pi/2} \frac{1}{2} \cos(x) x^2 dx$$

$$= \frac{1}{2} (2x \cos(x) + (-2 + x^2) \sin(x)]_{-\pi/2}^{\pi/2} = 0.4674$$

$$\sigma = \sqrt{0.4674} = 0.6837$$

**Problem 145.** This problem may be helpful to those who need practice doing Taylor expansions. Consider a function, $f(x)$, that is infinitely differentiable at a point $x_0$ (i.e. a “nice” function). The Taylor expansion of this function for points $x$ centered around the point $x_0$ is given by

$$\sum_{n=0}^{\infty} f^{(n)}(x_0) \frac{(x - x_0)^n}{n!}$$

where $f^{(n)}(x_0)$ is the $n$th derivative of $f(x)$ at $x = x_0$. Consider the function $f(x) = \exp(x)$. The Taylor expansion, $T(x)$, of $\exp(x)$ for $x$ near $x_0 = 2$ is

$$T(x) = f(x_0) + f^{(1)}(x_0)(x - x_0) + f^{(2)}(x_0) \frac{(x - x_0)^2}{2!} + \ldots$$

$$= \exp(2) + \exp(2)(x - 2) + \exp(2) \frac{(x - 2)^2}{2!} + \ldots$$

Let $T_n(x)$ denote the Taylor expansion truncated at the $n$-th term:

$$T_n(x) = f(x_0) + f^{(1)}(x_0)(x - x_0) + \ldots + f^{(n)}(x_0) \frac{(x - x_0)^n}{n!}.$$ 

Notice that when $x = x_0$, all terms after the first go to zero because $x - x_0 = 0$ which then gives $T(x) = f(x_0)$. When $x \neq x_0$ and near $x_0$, the second term $f^{(1)}(x_0)(x - x_0)$ acts as a linear approximation. Subsequent terms act as correction terms that bring $T_n(x)$ progressively close to $f(x)$ as $n$ increases.

(a) Check graphically that the second term in $T(x)$ acts as a linear approximation of $f(x)$ for $x \neq x_0$ and near to $x_0$. Plot $f(x) = \exp(x)$ for $x = [2, 2.5]$ then plot $T(x = 2)$ and the first two terms of $T(x = 2.1)$ for $x_0 = 2$. 
(b) Let’s add more terms to \( T_n(x = 2.1) \) and look at its limiting behavior. Plot \( T_n(x = 2.1) \) keeping the first three terms (i.e. \( n = 2 \)), then repeat for the first four terms (\( n = 3 \)). You should see that adding more terms brings \( T_n(x = 2.1) \) closer to \( f(x = 2.1) = \exp(2.1) \).

For any \( \epsilon > 0 \) and \( x \) we can show that \( |T_n(x) - f(x)| < \epsilon \) for large enough \( n \) (this is called point-wise convergence, i.e. convergence at any \( x \)). One can also show that the Taylor expansion is unique.

Solution. (a) The arrow for “1st correction” is \( T_1(x = 2.1) \)

(b) Here the graph is zoomed in such that \( T_2(2.1) \) and \( T_3(2.1) \) do not overlap

Problem 146. Taylor expansions are often used as approximations to \( f(x) \) when \( x \) is near the point of expansion \( x_0 \). The farther \( x \) is from \( x_0 \) the more terms we need for a good approximation. Here is some practice for you; these functions are found in chemistry:

(a) Taylor expand \( f(T) = \log(T/\tau) \) as \( T \to \tau \) where \( \tau \) is a reference temperature to make \( T/\tau \) dimensionless.

(b) Taylor expand \( f(x) = (1 - x^2)^{-1} \) as \( x \to 0 \).

(c) Taylor expand \( f(T) = \exp(-T^{-1}) \) as \( T \to \infty \).

(d) Taylor expand \( f(x) = \sqrt{1 - x} \) as \( x \to 0 \).

Solution. (a) Let’s do a change of variables where \( c = \frac{T}{\tau} - 1 \). As \( T \to \tau \), \( c \to 0 \). Then

\[
f(T) = f(c) = \log(c + 1)
\]
\[ f(0) = \log(1) = 0 \]
\[ f^{(1)}(0) = 1 \]

By taking successive derivatives, we arrive at
\[ f(c) = c - \frac{c^2}{2} + \frac{c^3}{3} + \ldots \]
\[ f(T) = \left( \frac{T}{\tau} - 1 \right) - \frac{(T - 1)^2}{2} + \frac{(T - 1)^3}{3} + \ldots \]

(b) Recall the sum of a geometric series for \(|r| < 1\)
\[ \sum_{n=0}^{\infty} r^n = \frac{1}{1-r} \]

We can fit the given function, \( f(x) = (1 - x^2)^{-1} \) into the above equation by change of variables where \( c = x^2 \). For \(|c| < 1\), the geometric series
\[ \frac{1}{1-c} = \sum_{n=0}^{\infty} c^n \]
\[ \frac{1}{1-c} = 1 + c + c^2 + \ldots \]

with \( c = x^2 \) leads to:
\[ f(x) = 1 + x^2 + x^4 + \ldots \]

A Taylor expansion is a polynomial expansion and is unique for a given function. Therefore, \( f(x) = 1 + x^2 + x^4 + \ldots \) is the Taylor expansion of \( \frac{1}{1-x^2} \).

(c) Let \(-T^{-1} = c\)
\[ f(T) = f(c) = \exp(c) \]
\[ \exp(c) = 1 + c + \frac{c^2}{2!} + \frac{c^3}{3!} + \ldots \]
\[ f(T) = 1 - \frac{1}{T} + \frac{1}{2!T^2} - \frac{1}{3!T^3} + \ldots \]

(d) Let’s take the first few derivatives
\[ f^{(1)}(x) = -\frac{1}{2}(1-x)^{-1/2} \]
\[ f^{(2)}(x) = -\left(\frac{1}{2}\right)^2 (1-x)^{-3/2} \]
\[ f^{(3)}(x) = -3\left(\frac{1}{2}\right)^3 (1-x)^{-5/2} \]
\[ f^{(4)}(x) = -5 \times 3\left(\frac{1}{2}\right)^4 (1-x)^{-7/2} \]
\[ f^{(n)}(x) = -|2n - 3|!!\left(\frac{1}{2}\right)^n (1-x)^{-(2n-3)/2} \]
where \( m!! = m \times (m-2) \times (m-4) \times \cdots \times 1 \) for \( m \) odd.

\[
f(x) = 1 - \frac{x}{2} - \frac{x^2}{8} - \frac{x^3}{16} - \cdots
\]

**Problem 147.** The triple product rule, also known as the cyclic relation, is a formula commonly found in thermodynamics. If a function of three variables can be expressed as \( f(x, y, z) = 0 \), then each variable can be expressed as a function of the remaining two variables. Show that,

\[
\left( \frac{\partial x}{\partial y} \right)_z \left( \frac{\partial y}{\partial z} \right)_x \left( \frac{\partial z}{\partial x} \right)_y = -1
\]

where \( \left( \frac{\partial x}{\partial y} \right)_z \) denotes the derivative of the function \( x(y, z) \) with respect to \( y \) (holding \( z \) constant), etc. Does the triple product rule hold true for the ideal gas law? Show your work.

**Solution.** Since we have \( f(x, y, z) = 0 \) we may solve for \( x \) as a function of \( y \) and \( z \), and write the total differential of \( x(y, z) \):

\[
dx(y, z) = \left( \frac{\partial x}{\partial y} \right)_z dy + \left( \frac{\partial x}{\partial z} \right)_y dz
\]

Likewise, we may solve for \( y \) as a function of \( x \) and \( z \). The total differential of \( y(x, z) \) is:

\[
dy(x, z) = \left( \frac{\partial y}{\partial x} \right)_z dx + \left( \frac{\partial y}{\partial z} \right)_x dz
\]

Substitute \( dy \) into \( dx \)

\[
dx = \left( \frac{\partial x}{\partial y} \right)_z \left[ \left( \frac{\partial y}{\partial x} \right)_z dx + \left( \frac{\partial y}{\partial z} \right)_x dz \right] + \left( \frac{\partial x}{\partial z} \right)_y dz
\]

Because \( dx \) and \( dz \) are arbitrary displacements, their coefficients must vanish. Setting the coefficient of \( dx \) to zero gives the reciprocal relationship:

\[
\left( \frac{\partial x}{\partial y} \right)_z = \frac{1}{\left( \frac{\partial y}{\partial x} \right)_z}
\]

(A.4)

Thus, the inverse of a partial derivative is the same as inverting a fraction. Setting the coefficient of \( dz \) to zero:

\[
\left( \frac{\partial x}{\partial y} \right)_z \left( \frac{\partial y}{\partial z} \right)_x + \left( \frac{\partial x}{\partial z} \right)_y = 0
\]

Using a reciprocal relation such as Eq. (A.4) for \( \left( \frac{\partial x}{\partial z} \right)_y \) gives:

\[
\left( \frac{\partial x}{\partial y} \right)_z \left( \frac{\partial y}{\partial z} \right)_x \left( \frac{\partial z}{\partial x} \right)_y = -1
\]
Take the ideal gas law to be \( pv = kT \) where we take \( v \) used here to be the molar volume and \( k \) is Boltzmann’s constant.

\[
\left( \frac{\partial p}{\partial v} \right)_T : p + v \frac{\partial p}{\partial v} = 0 \implies \frac{\partial p}{\partial v} = -\frac{p}{v}
\]

\[
\left( \frac{\partial v}{\partial T} \right)_p : p \frac{\partial v}{\partial T} = k \implies \frac{\partial v}{\partial T} = \frac{k}{p}
\]

\[
\left( \frac{\partial T}{\partial p} \right)_v : v \frac{\partial T}{\partial p} = -\frac{v}{k}
\]

Now multiply the expressions together to confirm they equal -1

\[
\left( \frac{\partial p}{\partial v} \right)_T \left( \frac{\partial v}{\partial T} \right)_p \left( \frac{\partial T}{\partial p} \right)_v = -\frac{p k}{v p k} = -1
\]

**Problem 148.** Partial derivatives of a multivariable function are those which are taken for each individual variable while treating the remaining variables as constants. This is more rigorously achieved by taking limits. The definitions of the derivatives of a function \( f \) of three variables \( (x, y, z) \) with respect to each argument are:

\[
f_x(x, y, z) = \frac{\partial f(x, y, z)}{\partial x} = \lim_{h \to 0} \frac{f(x + h, y, z) - f(x, y, z)}{h}
\]

\[
f_y(x, y, z) = \frac{\partial f(x, y, z)}{\partial y} = \lim_{h \to 0} \frac{f(x, y + h, z) - f(x, y, z)}{h}
\]

\[
f_z(x, y, z) = \frac{\partial f(x, y, z)}{\partial z} = \lim_{h \to 0} \frac{f(x, y, z + h) - f(x, y, z)}{h}
\]

Find the partial derivatives \( f_x(x, y, z) \), \( f_y(x, y, z) \), and \( f_z(x, y, z) \) by computing directly these limits in the case of \( f(x, y, z) = xy^2z \).

**Solution.** Each of the three limits given above are applied to the function \( f(x, y, z) = xy^2z \).

\[
f_x(x, y, z) = \frac{\partial(xy^2z)}{\partial x} = \lim_{h \to 0} \frac{(x + h)y^2z - xy^2z}{h} = \lim_{h \to 0} y^2z = y^2z
\]

\[
f_y(x, y, z) = \frac{\partial(xy^2z)}{\partial y} = \lim_{h \to 0} \frac{x(y + h)^2z - xy^2z}{h} = \lim_{h \to 0} \frac{xz(2yh + h^2)}{h} = \lim_{h \to 0} xz(2y + h)
\]

As \( h \) approaches zero, it is much smaller than \( 2y \). Thus, only the first term is nonzero: \( 2xyz \).

\[
f_z(x, y, z) = \frac{\partial(xy^2z)}{\partial z} = \lim_{h \to 0} \frac{xy^2(z + h) - xy^2z}{h} = xy^2
\]
Problem 149. Use your knowledge of partial derivatives and the ideal gas law to show that
\[ V \frac{\partial P}{\partial V} = -P \quad \text{and} \quad V \frac{\partial P}{\partial T} + T \frac{\partial P}{\partial T} = 0 \]

Solution.
\[ V \frac{\partial P}{\partial V} = V \left( -\frac{k_B T}{V^2} \right) = -\frac{k_B T}{V} = -P \]
\[ V \frac{\partial P}{\partial V} + T \frac{\partial P}{\partial T} = -\frac{k_B T}{V} + T \left( \frac{k_B}{V} \right) = 0 \]

Problem 150. Show that the second-order quasilinear partial PDE with independent variables \((x,t)\)
\[ a(u_x, u_t)u_{xx} + 2b(u_x, u_t)u_{xt} + c(u_x, u_t)u_{tt} = 0, \]
can be transformed into the linear PDE
\[ a(u_x, u_t)\frac{\partial^2 w}{\partial u_x^2} - 2b(u_x, u_t)\frac{\partial^2 w}{\partial u_x \partial u_t} + c(u_x, u_t)\frac{\partial^2 w}{\partial u_t^2} = 0 \]
by way of a Legendre transformation.

Solution. We transform the old variables \(u(x,t)\) to new ones \(w(\xi, \eta)\) Setting,
\[ \xi = u_x, \quad \eta = u_t, \quad w_\xi = x, \quad w_\eta = t, \]
so that the Legendre transform is:
\[ w(\xi, \eta) = x\xi + t\eta - u(x,t) \]
The Jacobian of the transformation is
\[ J = \begin{vmatrix} u_{xx} & u_{xt} \\ u_{xt} & u_{tt} \end{vmatrix} = u_{xx}u_{tt} \neq 0. \]
Using the same approach as Problem 151 solution, we obtain the following transformations:
\[ u_{xx} = Jw_{\eta\eta}, \quad u_{xt} = -Jw_{\xi\eta}, \quad \text{and} \quad u_{tt} = Jw_{\xi\xi}. \]
Substitution into the original PDE gives:
\[ a(u_x, u_t)Jw_{\eta\eta} - 2b(u_x, u_t)(-Jw_{\xi\eta}) + c(u_x, u_t)Jw_{\xi\xi} = 0. \]
Dividing through by \(J\) and then using \(\xi = u_x\) ad \(\eta = u_t\) gives the desired result.

Problem 151. Consider the minimal surface equation:
\[ \nabla \cdot \left( \frac{\nabla u}{\left(1 + |\nabla u|^2\right)^{1/2}} \right) = 0, \]
(a) Explain what is the minimal surface equation.
(b) Show that for \( n = 2 \), the minimal surface equation may be rewritten as
\[
(1 + u_x^2)u_{x_1}u_{x_2} - 2u_{x_1}u_{x_2}u_{x_{1}x_2} + (1 + u_{x_1}^2)u_{x_1x_2} = 0.
\]

(c) Show that this nonlinear equation can be transformed, by means of a Legendre transformation (whose new variables are denoted \( p_1, p_2 \), to a linear equation
\[
(1 + p_2^2)v_{p_2p_2} + 2p_1p_2v_{p_1p_2} + (1 + p_1^2)v_{p_1p_2} = 0.
\]

**Solution.** (b) For \( n = 2 \), we have that (using \( x_1 = x \) and \( x_2 = y \))
\[
\nabla u = \left( \frac{\partial u}{\partial x}, \frac{\partial u}{\partial y} \right) = (u_x, u_y)
\]
and
\[
|\nabla u|^2 = (\nabla u) \cdot (\nabla u) = u_x^2 + u_y^2.
\]
Thus,
\[
\nabla \cdot \left( \frac{\nabla u}{(1 + |\nabla u|^2)^{1/2}} \right) = (u_x, u_y) \cdot \left( \frac{u_x}{(1 + u_x^2 + u_y^2)^{1/2}}, \frac{u_y}{(1 + u_x^2 + u_y^2)^{1/2}} \right)
\]
\[
= \frac{\partial}{\partial x} \left( \frac{u_x}{(1 + u_x^2 + u_y^2)^{1/2}} \right) + \frac{\partial}{\partial y} \left( \frac{u_y}{(1 + u_x^2 + u_y^2)^{1/2}} \right)
\]
\[
= \frac{u_{xx} + u_{xx} u_y^2 - u_{x} u_y u_{xy}}{(1 + u_x^2 + u_y^2)^{3/2}} + \frac{u_{yy} + u_{yy} u_x^2 - u_{x} u_y u_{xy}}{(1 + u_x^2 + u_y^2)^{3/2}}
\]
\[
= \frac{(1 + u_y^2)u_{xx} - 2u_x u_y u_{xy} + (1 + u_x^2)u_{yy}}{(1 + u_x^2 + u_y^2)^{3/2}} = 0.
\]
For the equation to hold, we must have the numerator equal to zero. That is,
\[
(1 + u_y^2)u_{xx} - 2u_x u_y u_{xy} + (1 + u_x^2)u_{yy} = 0.
\]

(c) The idea is to transform from \( u(x_1, x_2) \) to \( v(p_1, p_2) \). The Legendre transform is:
\[
v(p_1, p_2) = \tilde{p} \cdot \tilde{x} - u(x_1, x_2) = p_1 x_1 + p_2 x_2 - u(x_1(p_1, p_2), x_2(p_1, p_2))
\]
where \( p_1 = u_{x_1} \) and \( p_2 = u_{x_2} \). We assume the mapping from \( (x_1, x_2) \) to \( (p_1, p_2) \) is one-to-one and invertible so the condition (here we view \( u_{x_1} \) and \( u_{x_2} \) as independent variables)
\[
J = \begin{vmatrix}
  u_{x_1x_1} & u_{x_1x_2} \\
  u_{x_2x_1} & u_{x_2x_2}
\end{vmatrix} = u_{x_1x_1} u_{x_2x_2} - u_{x_1x_2}^2 \neq 0
\]
holds. Now we find a PDE for \( v \) with independent variables \( p_1, p_2 \). The inverse transformation has the following Jacobian:
\[
J' = \begin{vmatrix}
  v_{p_1p_1} & v_{p_1p_2} \\
  v_{p_2p_1} & v_{p_2p_2}
\end{vmatrix} = v_{p_1p_1} v_{p_2p_2} - v_{p_1p_2}^2 \neq 0.
\]
By the involutivity of the Legendre transformation, we just have $JJ' = 1$, or in matrix form,

$$
\begin{bmatrix}
  u_{x_1x_1} & u_{x_1x_2} \\
  u_{x_2x_1} & u_{x_2x_2}
\end{bmatrix}
\begin{bmatrix}
  v_{p_1p_1} \\
  v_{p_2p_1}
\end{bmatrix}
= 
\begin{bmatrix}
  1 & 0 \\
  0 & 1
\end{bmatrix}.
$$

From this, we obtain the following set of equations:

$$
\begin{align*}
  u_{x_1x_2}v_{p_1p_1} + u_{x_1x_2}v_{p_2p_1} &= 1 \\
  u_{x_1x_1}v_{p_1p_2} + u_{x_1x_2}v_{p_2p_2} &= 0 \\
  u_{x_2x_1}v_{p_1p_1} + u_{x_2x_2}v_{p_2p_1} &= 0 \\
  u_{x_2x_1}v_{p_1p_2} + u_{x_2x_2}v_{p_2p_2} &= 1.
\end{align*}
$$

Note that $u_{x_1x_2} = u_{x_2x_1}$ and $v_{p_1p_2} = v_{p_2p_1}$, and observe that:

$$
\begin{align*}
  u_{x_1x_1} &= u_{x_1x_1} - u_{x_1x_1} \times 0 \\
  &= u_{x_1x_1} - u_{x_1x_2}(u_{x_1x_1}v_{p_1p_2} + u_{x_1x_2}v_{p_2p_2}) \\
  &= u_{x_1x_1} - u_{x_1x_2}u_{x_1x_1}v_{p_1p_2} - u_{x_1x_2}^2v_{p_2p_2} \\
  &= u_{x_1x_1}(1 - u_{x_1x_2}v_{p_1p_2}) - u_{x_1x_2}^2v_{p_2p_2} \\
  &= u_{x_1x_1}(u_{x_2x_2}v_{p_2p_2}) - u_{x_1x_2}^2v_{p_2p_2} \\
  &= (u_{x_1x_1}u_{x_2x_2} - u_{x_1x_2}^2)v_{p_2p_2} \\
  &= Jv_{p_2p_2}.
\end{align*}
$$

Similarly, we find that $u_{x_1x_2} = -Jv_{p_1p_2}$ and $u_{x_2x_2} = Jv_{p_1p_1}$. Substituting these transformed variables into the original differential equation gives:

$$(1 + p_2^2)Jv_{p_2p_2} + 2Jp_1p_2v_{p_1p_2} + (1 + p_1^2)Jv_{p_1p_1} = 0.$$

Dividing through by $J$ gives the desired result. ■
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