Chemistry XL 14A

Saturday, June 25, 2011

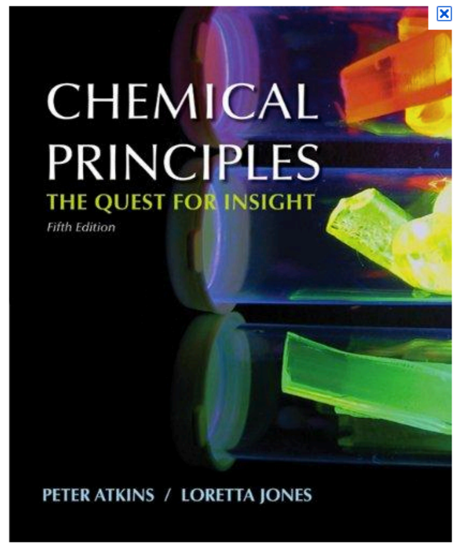
Syllabus

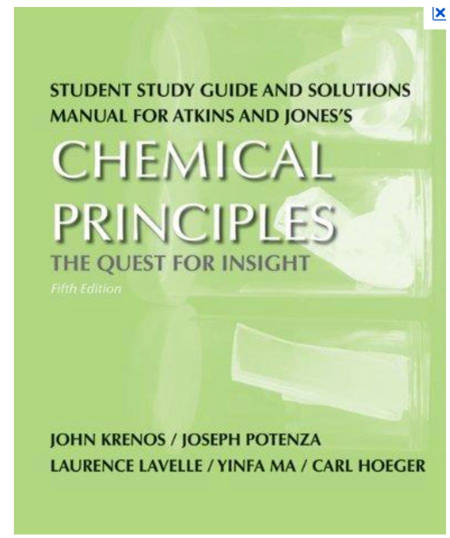
Review of the Basics:

1. Math of Chemistry/Science
2. Fundamentals of Chemistry

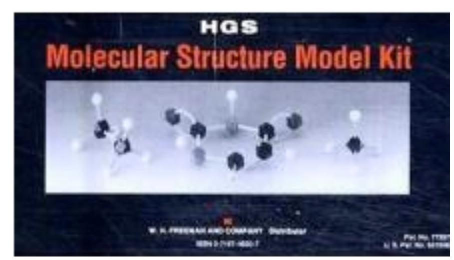
Chem XL 14A

- General Chemistry for Life Science Majors
  - Atomic and Molecular Structure
  - Quantum Mechanics
  - Gas Laws
  - Chemical Equilibrium
  - Acids and Bases





- Textbook:
  - Chemical Principles, 5th Ed. Atkins and Jones
  - Hardcover ISBN-13: 978-1-4292-1955-6
  - Model kit not necessary



Important Information

- Dr. Robert Iafe
  - Email: rgiafe@chem.ucla.edu
  - Office: MS-B 3234

Success and Academic Integrity

- Success
  - Good working knowledge and understanding of high school chemistry
  - Minimum of 7 hrs/wk outside of class
  - Attend Lecture
  - Participate in lecture – respond, bring calculator
  - Read required text PRIOR to class
  - Get help if falling behind – OH/tutoring
  - Understand basic concepts in addition to doing problems
  - Web enhanced course – Blackboard
- Academic Honesty
  - It is not worth cheating in this class
  - http://www.studentgroups.ucla.edu/dos/students/integrity/

Lecture Schedule

Date	Lecture Topic	Readings	Quizzes/Exams and Assignment Due Dates
Week 1	06/25	Fundamentals B, C, D, E, F, G, H, L1-3, M; Appendix 1B, 1C, 1D	
Week 2	No Class – Independence Day		
Week 3	07/09	Chapter 1. Atoms: The Quantum World	
Week 4	07/16	Chapter 2. Chemical Bonds	
Week 5	07/23	Chapter 3. Molecular Shapes and Structures	Exam I: Fundamentals, Chapter 1, 2
Week 6	07/30	Chapter 3. Molecular Shapes and Structures Chapter 16. Coordination Cmpds	
Week 7	08/06	Chapter 4. Gasses	
Week 8	08/13	Chapter 9. Chemical Equilibrium	Exam II: Chapter 3, 4
Week 9	08/20	Fundamentals J. Chapter 10. Acids and Bases	
Week 10	08/27	Chapter 11. Aqueous Equilibrium	
Week 11	No Class		
Week 12	09/10	Review session followed by	Final Exam: Cumulative

Assessment and Grading

Course Grade Distribution

• 2 Midterm Exams: (Friday, July 23rd and Aug 13th)	25% each	50%
• Final Exam (Saturday, Sept 10th)		50%

- Quizzes
- Homework
  - Suggested problems assigned on the syllabus
  - Not collected/checked
- Calculators:
  - Graphing calculators are allowed on exams
  - Go buy a cheap (<$20) scientific calculator today!
    - Recommendation: TI-30X IIS

Math for Chem 14A

1. Metric System
2. Scientific Notation
3. Significant Figures

The Metric System

- Every measurement consists of:
  - A numerical value
  - A reference unit
- All SI units can be derived from the 7 base units:

• Length:	meter (m)
• Mass:	kilogram (kg)
• Time:	second (s)
• Temperature:	kelvin (K)
• Chemical amount:	mole (mol)
• Luminescence	candela (cd)
• Current	ampere (A)

The Metric System

- Derived Units:

• Coulomb	C	A·s
• Newton	N	$kg \cdot m/s^2$
• Joule	J	$N \cdot m$ or $kg \cdot m^2/s^2$
• Volt	V	J/C
• Hertz	Hz	1/s or s^{-1}

SI Prefixes

Modifiy any unit using a prefix:

Prefix	*Prefix Symbol*	*Number*	*Word*	*Exponential Notation*
tera	T	1,000,000,000,000	trillion	10^{12}
giga	G	1,000,000,000	billion	10^{9}
mega	M	1,000,000	million	10^{6}
kilo	k	1,000	thousand	10^{3}
hecto	h	100	hundred	10^{2}
deka	da	10	ten	10^{1}
-----	----	1	one	10^{0}
deci	d	0.1	tenth	10^{-1}
centi	c	0.01	hundredth	10^{-2}
milli	m	0.001	thousandth	10^{-3}
micro	μ	0.000001	millionth	10^{-6}
nano	n	0.000000001	billionth	10^{-9}
pico	p	0.000000000001	trillionth	10^{-12}
femto	f	0.000000000000001	quadrillionth	10^{-15}

Converting

- Important Relationships
  - 1 in = 2.54 cm
  - 1 min = 60 s
  - 1 cal = 4.184 J
  - 1 atm = 760 Torr = 760 mm Hg = 101.3 kPa
  - K = ^{o}C + 273.15
- Dimensional Analysis
  - 155 mm → m
  - 0.31 km → mm
  - 3200 J → kcal
  - 30 cm^3 → m^3

Scientific Notation

Shorthand for writing either very large or very small numbers….

$0.000000000333 = 3.33 \times 10^{-10}$

$3330000000000 = 3.33 \times 10^{13}$

Numbers are written as:

$A \times 10^{a}$

A is a decimal number

a is a whole number

Significant Figures

The number of digits in a reported measurement

T = 85.3 ºF → 3 sig figs

What about zeros?

<u>Leading zeros</u> are not significant

0.0035 → 2 sig figs

<u>Trailing zeros</u> are significant if a decimal point is present

3.560 → 4 sig figs

3560 → 3 sig figs

3560. → 4 sig figs

Significant Figures

While a measured value has some uncertainty,

A number that is counted is an exact value

There are exactly 122 students enrolled in this class

Rounding off:

1. Round up if the last digit is above 5: 4.567 → 4.57
2. Round down if the last digit is below 5: 4.563 → 4.56
3. If the last digit is 5, round to the nearest even number

4.565 → 4.56   4.575 → 4.58

Math with Significant Figures

Addition/Subtraction:

Round to smallest number of decimal places

Multiplication/Division:

Round to smallest number of significant figures

Exact numbers and Counted numbers:

Treat as if have infinite significant figures

Logarithms and exponentials

Significant figures of the mantissa

$$10^{3.56} = 10^{0.56+3} = 10^{0.56} \times 10^{3} = 3.6 \times 10^{3}$$

Practice

$$\frac{6.342 + 0.94}{602} \times \frac{39.6 - 1}{592 \times 0.054} =$$

$$\frac{7.282}{602} \times \frac{38.6}{31.968} = \frac{7.28}{602} \times \frac{39}{32} = 0.0147384$$

$$= 0.015$$

$$= 1.5 \times 10^{-2}$$

Fundamentals of Chemistry

1. Properties
2. Elements and the atom
3. Compounds
4. The Mole and molar mass
5. Percent Composition and determining empirical Formulas
6. Molarity
7. Chemical Equations
8. Stoichiometry

Properties

Physical properties – can be observed without changing the identity of the substance

Ex. Melting pt Color State of Matter

Chemical properties – refer to the chemical reactivity of a substance

Ex. Zinc reacts with acids to produce H_2 gas

Intensive Property – independent of the mass of a substance

Ex. Temperature Color Hardness

Extensive Property – depends on the mass of the substance

Ex. Mass Volume Energy

Measurements

All measured values have some error

<u>Systematic Error</u> – a repeated error present in many measurements

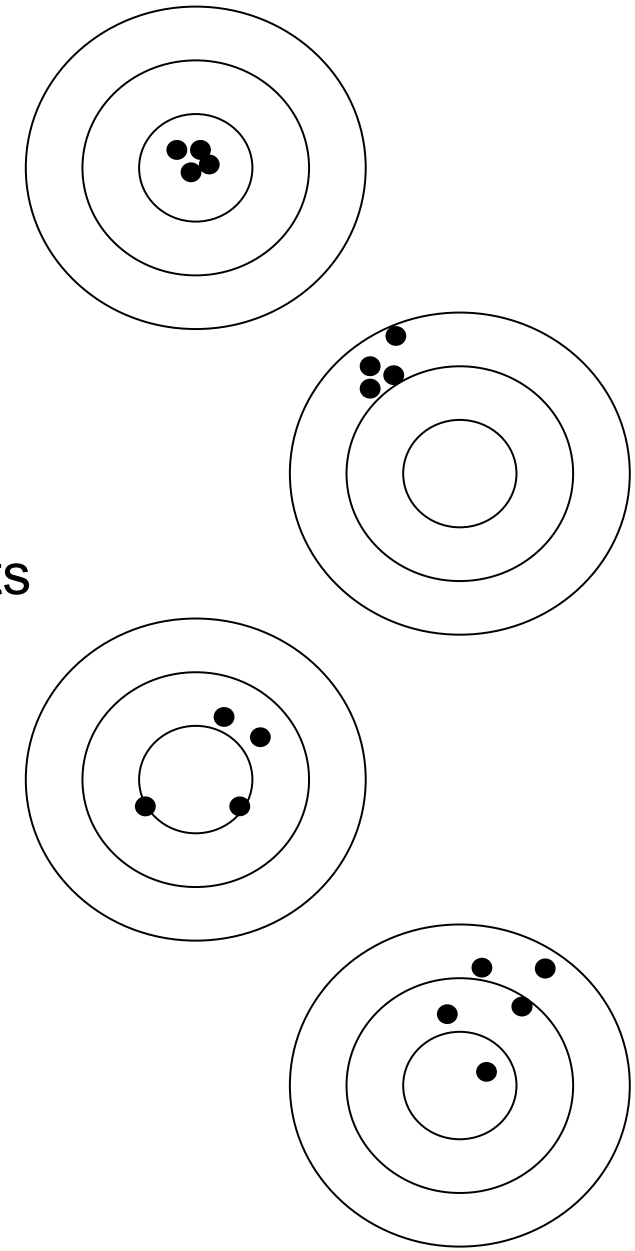
Ex. Instrument is incorrectly calibrated

<u>Random Error</u> – varies at random and can average to zero over many measurements

Ex. Random air currents affect measurement

<u>Precise Measurements</u> – repeated measurements are similar to each other

<u>Accurate Measurements</u> – measurements are close to average value



The Atom and Dalton

<u>Dalton's Atomic Hypothesis</u>

1. All atoms of a given element are identical
2. Atoms of different elements have different masses
3. A compound is a specific combination of more than one element
4. In a chemical reaction, atoms are neither created nor destroyed; they exchange partners to produce new substances.

The Atom

Atom – the smallest possible unit of an element

- small, positively charged nucleus surrounded by negatively charged electrons
- the nucleus is made up of positively charged protons and neutral neutrons
- mass of proton/neutron = ~ 2000 mass of electron
- a neutral atom has equal numbers of protons and electrons

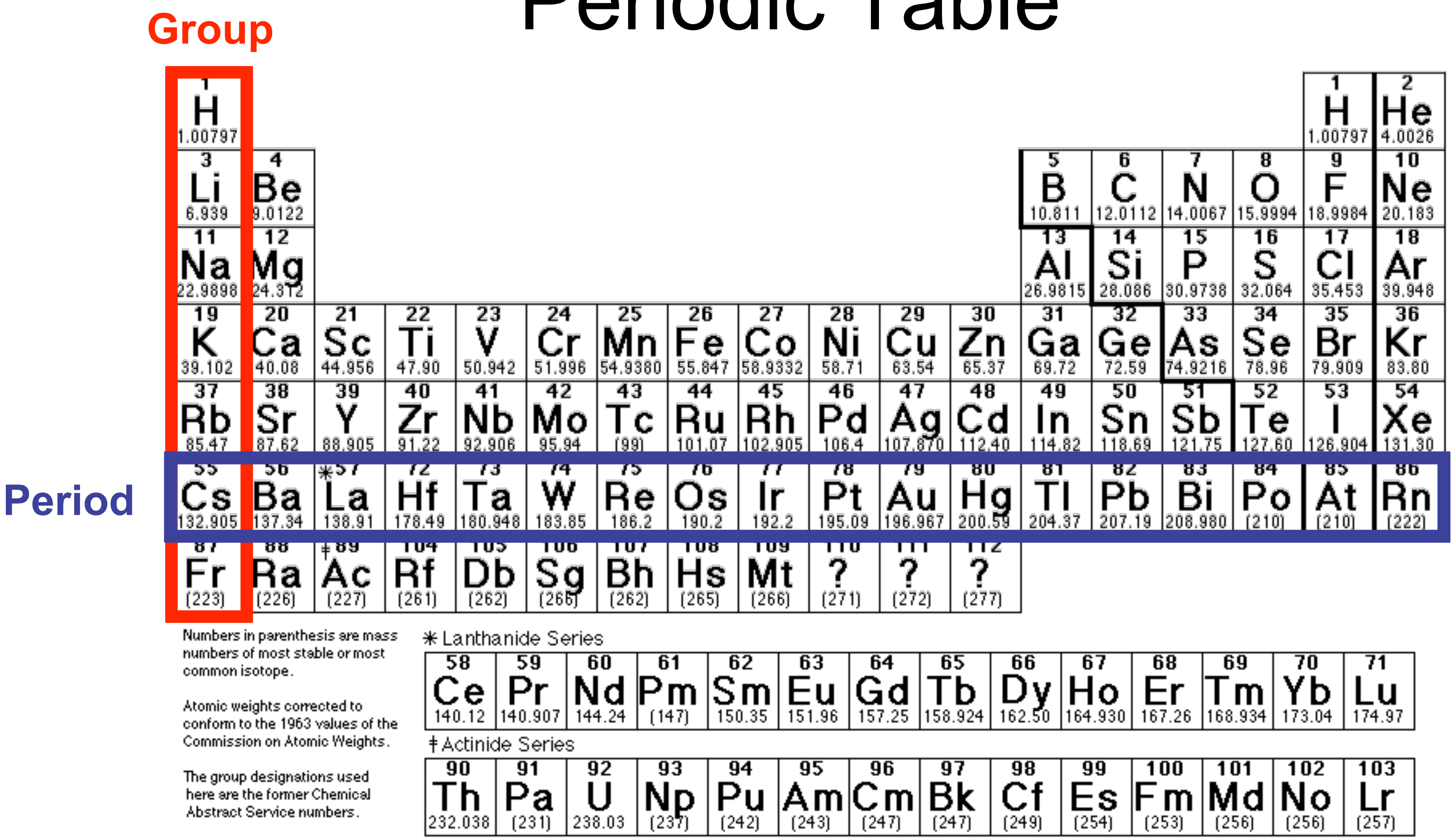
Atomic Number (Z) – the # of protons in an atom

- each element has a different Z
- each atom of an element has the same Z

Hydrogen → $Z = 1$ Chlorine → $Z = 17$

The Atom

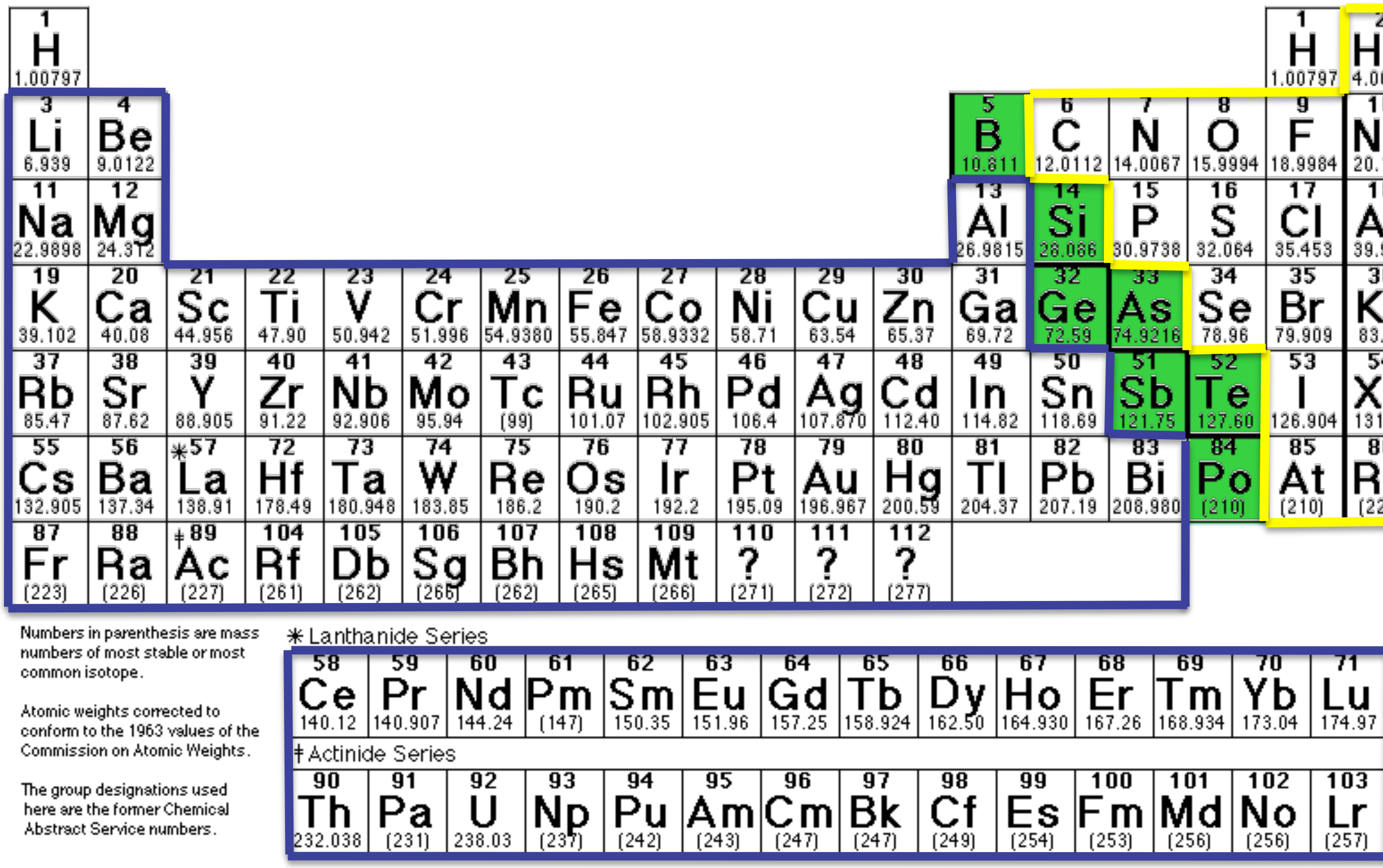
Periodic Table



Organization of the Elements based on periodic properties

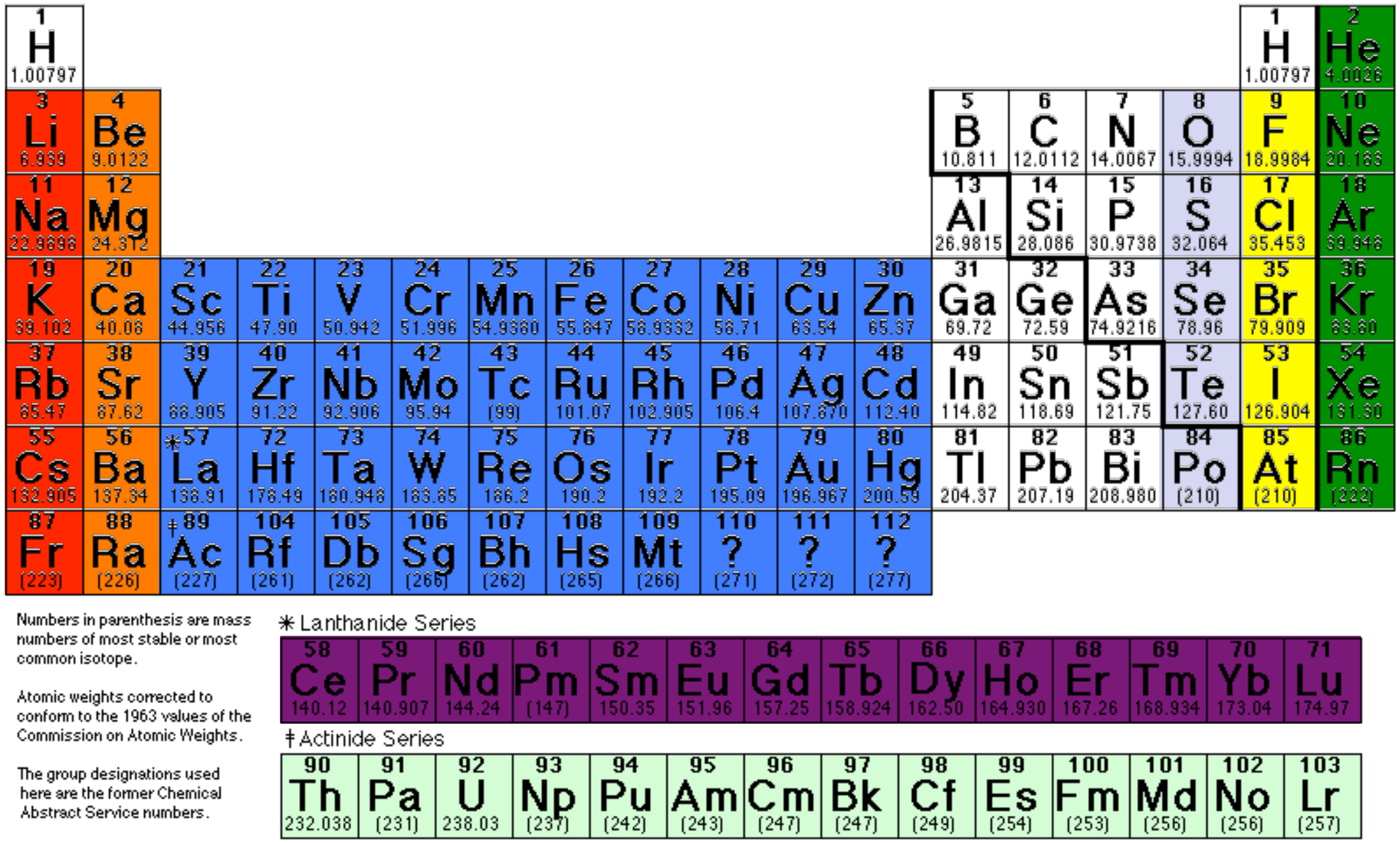
- Atomic number Z
- Molar mass
- Valence Electrons and Electron configurations

Periodic Table - Organization



- Metals: lustrous, malleable, ductile, conduct electricity
- Non-metals: brittle, dull, doesn't conduct electricity
- Metalloids: characteristics of both metals and non-metals

Periodic Table - Organization



Alkali Metals

Alkaline Earth Metals

Transition Metals

Lanthanoids

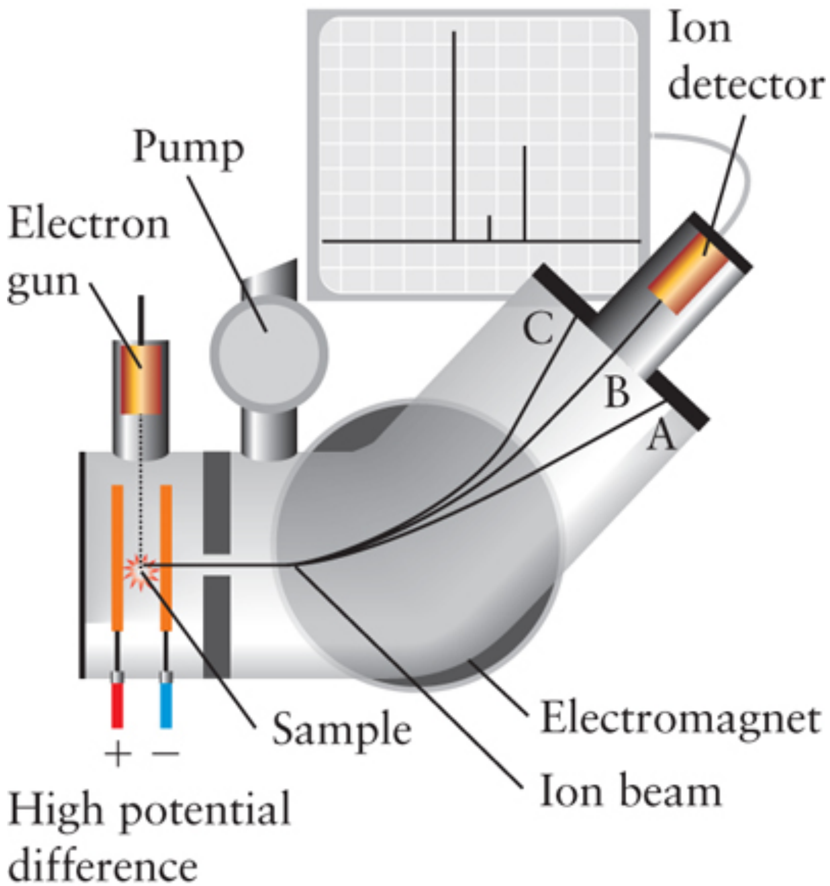
Actinoids

Chalcogens

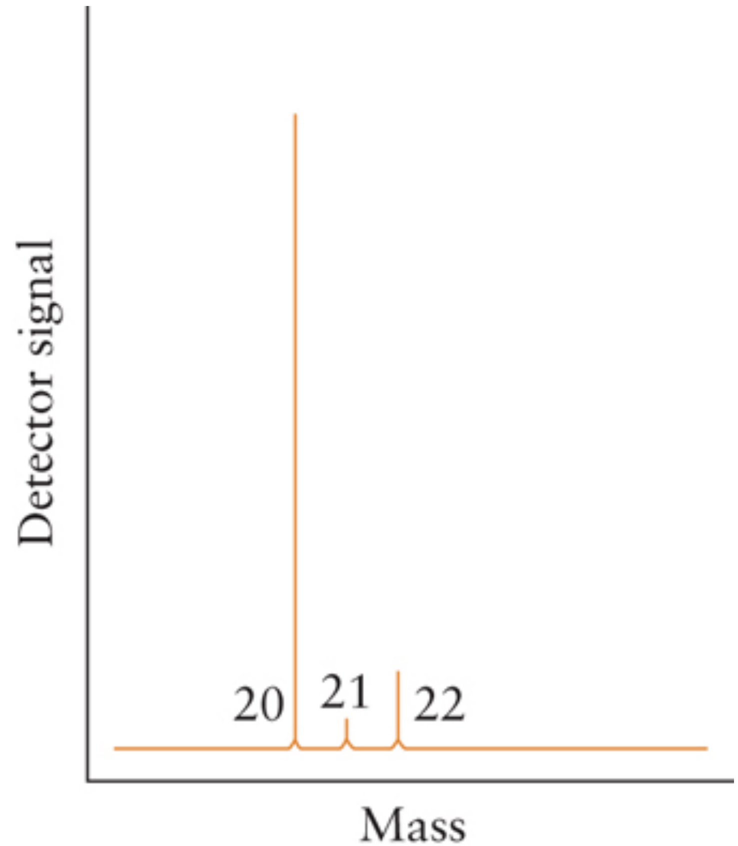
Halogens

Nobel Gases

Isotopes – Neon and MS



Mass Spectrometer



Isotopes

Mass Number (A) – the total # of protons and neutrons of an atom

Isotopes – Atoms with the same Z and different #'s of neutrons

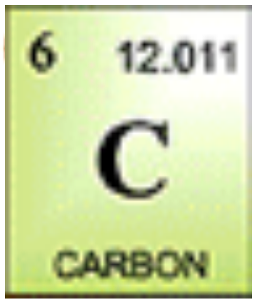
Neon-20 Neon-21 Neon-22

$^{A}_{Z}Sy$ $\quad$ $^{20}_{10}\mathrm{Ne}$ $\quad$ $^{21}_{10}\mathrm{Ne}$ $\quad$ $^{22}_{10}\mathrm{Ne}$

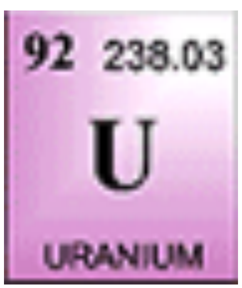
Isotopes have the same # of protons and electrons and

have essentially the same chemical and physical properties

Isotopes – Natural Abundance



^{12}C	12.000000	98.93
^{13}C	13.003355	1.07
^{14}C	14.003242	*



^{234}U	234.0409521	0.0054%
^{235}U	235.0439299	0.7204%
^{238}U	238.0507882	99.2742%

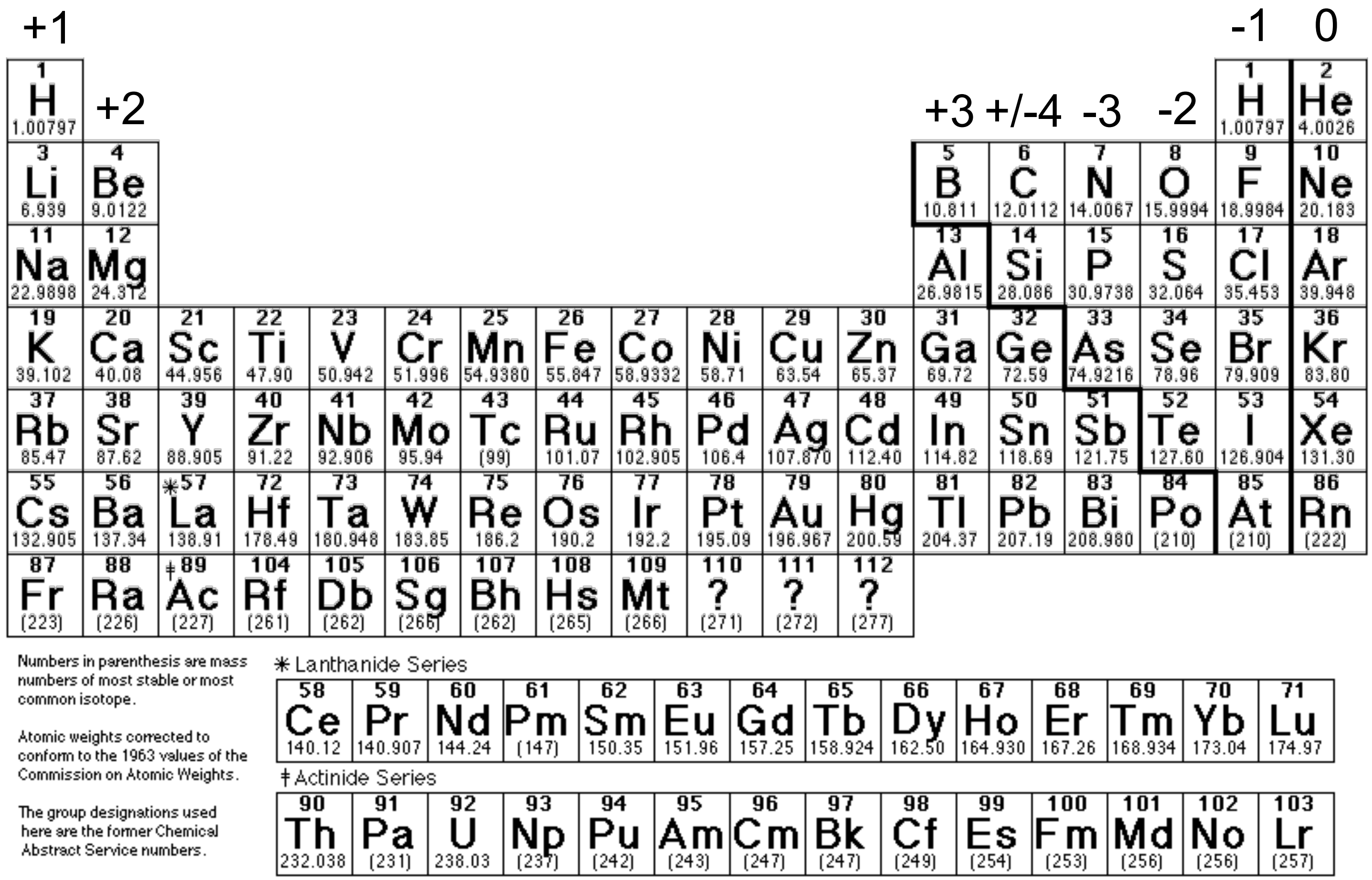
Ions

Ion – positively or negatively charged atom or molecule

- Ions of the same element have the same Z, different # electrons
- Cation – positively charged atom, an atom that has lost electrons
- Ex. Sodium loses 1 electron $Na \rightarrow Na^{+} + 1\ e^{-}$
- Anion – negatively charged atom, an atom that has gained electrons
- Ex. Oxygen gains 2 electrons $O + 2\ e^{-} \rightarrow O^{2-}$

Atoms tend to form specific ions based on their electron configurations

How to Find Ion Charge



- For Main Group Elements, look at the group number
- For Transition Metals and Metalloids, look at the anion

Compounds

An **Atom** is the smallest unit of an element.

H (hydrogen) O (oxygen) F (fluorine) Sn (tin)

A **compound** is a substance that consists of atoms of two or more different elements in a definite ratio

H_2O (water) Fe_3O_4 (magnetite) $C_6H_{12}O_6$ (glucose)

A binary compound consists of only two elements

Organic Compounds – always contain C (carbon) and usually H as well.

Inorganic Compounds – everything else. Some carbon compounds are treated as inorganic (CO_2, $CaCO_3$)

Ionic vs Molecular Compounds

Ionic compound – consists of ions in a ratio that results in a substance that is neutral

Ions are held together by electrostatic interactions:
the attraction between cations (+) and anions (-)

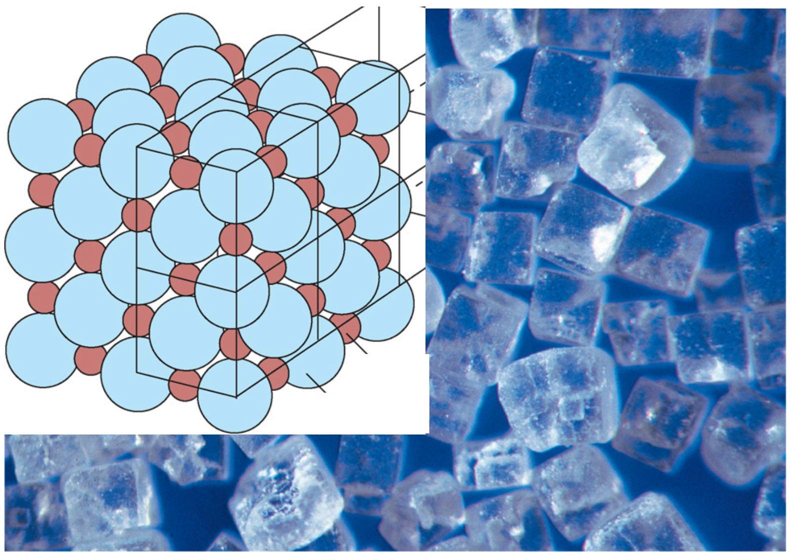
Ionic bond - Transfer of electrons from cation to the anion

Molecular compound – consists of electrically neutral molecules

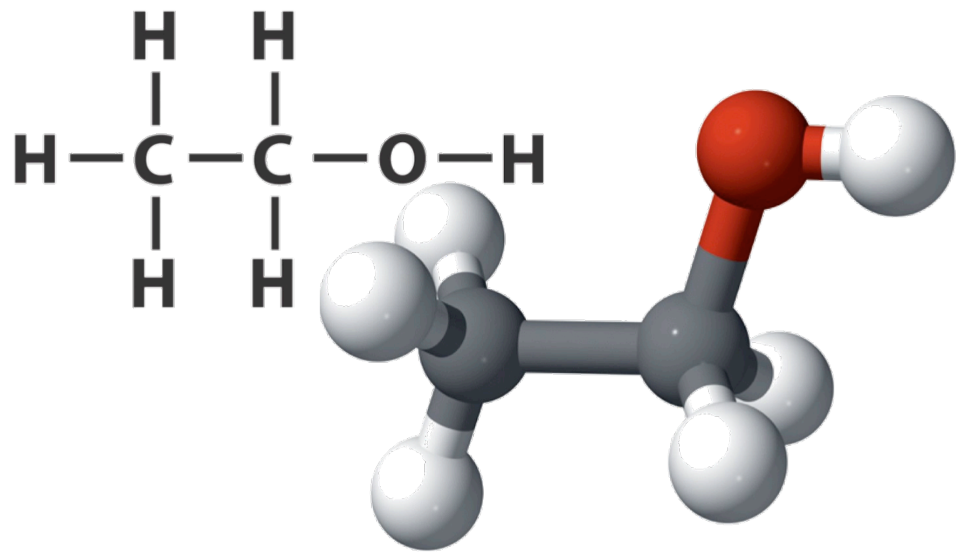
Molecules are groups of atoms covalently bonded together in a specific arrangement

Covalent bond – Sharing of electrons between atoms

Compounds and Formulas



NaCl, Table Salt

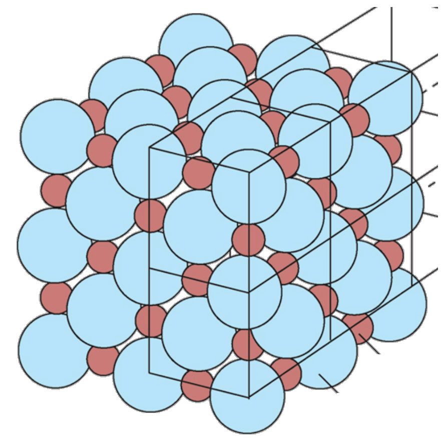


C_2H_5OH, Ethanol

<u>Chemical Formula</u> – represents the composition of a compund in terms of chemical symbols. Subscripts identify how many atoms of each element.

<u>Molecular Formula</u> – a chemical formula that shows how many atoms of each element are present in 1 molecule.

More about Ions!



NaCl, Table Salt

<u>Formula Unit</u> – Smallest whole # ratio of ions present in ionic compound

The ratio of ions is determined by charges of cations and anions

Atoms gain or lose electrons until they have the same # of electrons as the nearest noble gas atom

Na forms Na^+ ion

Chlorine forms Cl^- ion

1:1 ratio of Na^+ to Cl^-

Formula Unit is NaCl

What about Aluminum and Oxygen?

Polyatomic Ions

So far, we've just talked about the ions of different elements…

Polyatomic ions : 2 or more atoms that are covalently bonded together. A polyatomic ion has an overall charge

Cations	Anions
NH_4^+	CN^-
H_3O^+	SO_4^{2-}

H_2SO_4 NH_4Cl $NaCN$ $(NH_4)_2SO_4$

Naming Compounds - Nomenclature

Many compounds have common names

Water, ammonia, sugar, testosterone

Systematic names identify the ratio of elements present. In organic chemistry the systematic name also reveals the structure of the compounds

Different Rules for naming different types of compounds

Ionic Compounds

Inorganic Molecular Compounds

Organic Compounds – you will begin this in 14B

Naming Ionic Compounds

For elements which <u>predictably form specific ions</u>:

- No indication of ratio of ions is needed. The charges will identify the ratios of the elements
- Cation named first, followed by the anion.
- Monatomic anions – add "-ide" to the stem of the element name

Ex. $MgCl_2$ → magnesium chloride

Ex. sodium sulfide → Na_2S

Ex. $NH_4C_2H_3O_2$ → ammonium acetate

For elements which <u>can have more than one oxidation state</u>:

- Indicate the charge of the ion with a Roman Numeral

Ex. Fe_2O_3 → Iron(III) oxide

Polyatomic Ion Nomenclature

Charge number	Chemical formula	Name	Oxidation number of central element	Charge number	Chemical formula	Name	Oxidation number of central element
+2	Hg_2^{2+}	mercury(I)	+1		O_3^-	ozonide	$-\frac{1}{3}$
	UO_2^{2+}	uranyl	+6		OH^-	hydroxide	−2(O)
	VO^{2+}	vanadyl	+4		SCN^-	thiocyanate	—
+1	NH_4^+	ammonium	−3	−2	C_2^{2-}	carbide (acetylide)	−1
	PH_4^+	phosphonium	−3		CO_3^{2-}	carbonate	+4
−1	$CH_3CO_2^-$	acetate (ethanoate)	0(C)		$C_2O_4^{2-}$	oxalate	+3
	HCO_2^-	formate (methanoate)	+2(C)		CrO_4^{2-}	chromate	+6
					$Cr_2O_7^{2-}$	dichromate	+6
	CN^-	cyanide	+2(C), −3(N)		O_2^{2-}	peroxide	−1
	ClO_4^-	perchlorate*	+7		S_2^{2-}	disulfide	−1
	ClO_3^-	chlorate*	+5		SiO_3^{2-}	metasilicate	+4
	ClO_2^-	chlorite*	+3		SO_4^{2-}	sulfate	+6
	ClO^-	hypochlorite*	+1(Cl)		SO_3^{2-}	sulfite	+4
	MnO_4^-	permanganate	+7		$S_2O_3^{2-}$	thiosulfate	+2
	NO_3^-	nitrate	+5	−3	AsO_4^{3-}	arsenate	+5
	NO_2^-	nitrite	+3		BO_3^{3-}	borate	+3
	N_3^-	azide	$-\frac{1}{3}$		PO_4^{3-}	phosphate	+5

*These names are representative of the halogen oxoanions.

Naming Inorganic Molecular Compounds

Naming Binary molecular compounds:

- Must use greek prefixes to indicate the ratios of the elements
- Name the element which is further to the right in the periodic table second with the suffix “–ide”

Mono	1	Penta	5	Nona	9
Di	2	Hexa	6	Deca	10
Tri	3	Hepta	7	Undeca	11
Tetra	4	Octa	8	Dodeca	12

Ex. PCl_5 → Phosphorous pentachloride

Ex. Trinitrogen disulfide → N_3S_2

Acid Nomenclature

Binary Acids: hydro-(element)-ic acid

HF	hydrofluoric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid

Oxyacids: –ate becomes –ic acid

–ite becomes –ous acid

H_2SO_4	sulfuric acid	(from SO_4^{2-} sulfate)
HNO_2	nitrous acid	(from NO_2^{1-} nitrite)
HClO	hypochlorous acid	(from ClO^{1-} hypochlorite)

“Chemical Free” Products

Active Ingredients: Titanium Dioxide (5%) (Sunscreen), Zinc Oxide (10%) (Sunscreen)

Inactive Ingredients: Water, Ethylhexyl Palmitate, C12 15 Alkyl Benzoate, Ethylhexyl Stearate, Polyglyceryl 4 Isostearate, Cetyl PEG/PPG 10/1 Dimethicone, Hexyl Laurate, Propylene Glycol, Cetyl Dimethicone, Trimethylated Silica/Dimethicone, Octyldodecyl Neopentanoate, VP/Hexadecene Copolymer, Methyl Glucose Dioleate, PEG 7 Hydrogenated Castor Oil, Sorbitol Oleate, Hydrogenated Castor Oil, Beeswax (Apis Mellifera), Stearic Acid, Methylparaben, Propylparaben, Ethylparaben, Disodium EDTA, Diazolidinyl Urea, Tocopheryl Acetate (Vitamin E)

The Mole

A **<u>mole</u>** is <u>6.02214×10^{23}</u> of anything

In Chemistry, we need to compare ratios of molecules (atoms).

- Different elements have different masses:

  1 g of H_2 does not have the same # of molecules as 1 g of O_2

- Just like bakers count in “dozens,”

  Chemists count in “moles”

- 1 mole is defined as the # of carbon atoms in 12 g of ^{12}C

  Avogadro’s Number

  $N_A = 6.02214 \times 10^{23}\ mol^{-1}$

The Mole

1 mole of H atoms has $N_A = 6.02214 \times 10^{23}\ mol^{-1}$ H atoms

n = moles N = Number

We use N_A to convert between <u>particles</u> and <u>moles</u>

objects = (moles of objects) x (# objects/mol)

$$N = n \times N_A$$

- How many atoms of Fe are in 2.56×10^{-5} moles of Fe?
- 3.80×10^{24} molecules of Cl_2 is how many moles of Cl_2?

Molar Mass

Its just a little impractical to count atoms or molecules

All scales invented to date measure the mass of a substance, not the moles of a substance…

We use the **Molar Mass** of a substance to convert between moles (n) and mass (m)

- The **Molar Mass (M)** is the mass of 1 mole of a substance
- Molar mass has units of g/mol or kg/mol
- M of any element can be found on the Periodic Table
- Calculate the molar mass of a compound by multiplying the molar mass of each element by the number of atoms of that element:

$$M(H_2O) = 2 \times M(H) + 1 \times M(O) = ?$$

$$M(H_2O) = 18.02 \text{ g/mol}$$

Molar Mass

We use the **<u>Molar Mass (MM)</u>** of a substance to convert between moles (n) and mass (m)

mass of sample = amount in moles x molar mass

$$m = n \times M$$

- What is the mass of 1.45×10^{-2} moles of gold (Au)?
- 3.50 g of Fe_2O_3 contains how many moles of Fe_2O_3?
- 3.50 g of Fe_2O_3 contains how many molecules of Fe_2O_3?

More about Molar Mass

All Molar Masses in the Periodic Table are average values based on the percent of each isotope of an element present in nature

Carbon in nature is present as ~99% ^{12}C and ~1% ^{13}C

The average molar mass of Carbon is 12.01 g/mol

Molar mass vs Atomic/Molecular/Formula weight?

<u>Atomic 'weight'</u> – Molar mass of an element

<u>Molecular 'weight'</u> – Molar mass of a molecular compound

<u>Formula 'weight'</u> – Molar mass of an ionic compound

Review on Mass and Moles

Concept <u>NEEDS</u> to be <u>SECOND NATURE</u>

$$\text{molecules} \xrightleftharpoons[\times N_A]{\div N_A} \text{moles} \xrightleftharpoons[\div MM]{\times MM} \text{grams}$$

Mass Percent Composition

Mass % composition - the mass of each element in a compound expressed as a percentage of the total mass

$$\text{Mass \% of element} = \frac{\text{mass of element in a sample}}{\text{total mass of sample}} \times 100\%$$

What is the mass % composition of acetic acid ($C_2H_4O_2$)?

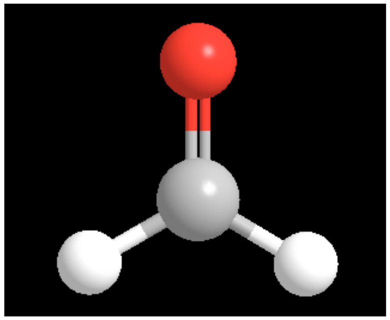
Determination of Chemical Formulas

Empirical formula - the smallest whole number ratio of atoms in a molecule

Molecular formula – gives the actual #'s of each atom of an element present in 1 molecule

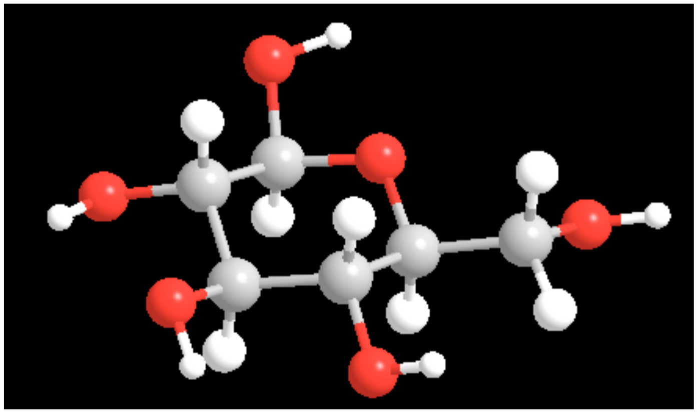
A substance has an empirical formula of CH_2O. What is it?

CH_2O



Formaldehyde

$C_6H_{12}O_6$



a-D-Glucose

To get the molecular formula, we need the Molar Mass of the substance

Determining Empirical Formulas

Through combustion analysis, a scientist can determine the percent composition of the elements in a sample

To Determine the Empirical formula of a compound:

1. Convert the mass percentages into moles
2. Determine the whole number ratio of moles of the elements

Tips:

1. Assume a sample of exactly 100 g
2. When you have the moles of each element, divide each by the smallest # of moles
3. Recognize decimal values of simple fractions:

$^1/_2 = 0.50$ $^1/_3 = 0.33$ $^1/_4 = 0.25$ $^1/_5 = 0.20$

Determining Empirical Formulas

From combustion analysis, a sample of benzene has a mass composition of:

92.26 % Carbon   7.74 % Hydrogen

What is the empirical formula of benzene?

1. Convert the mass percentages into moles
   - Assume a sample of exactly 100 g
2. Determine the whole number ratio of moles of the elements
   - When you have the moles of each element, divide each by the smallest # of moles
   - Recognize decimal values of simple fractions:

$^1/_2 = 0.50$ $^1/_3 = 0.33$ $^1/_4 = 0.25$ $^1/_5 = 0.20$

Determining Empirical Formulas

From combustion analysis, a sample of Salicylic Acid has a mass composition of:

60.87 % Carbon   4.38 % Hydrogen   34.75 % Oxygen

What is the empirical formula of Salicylic Acid?

1. Convert the mass percentages into moles
   - Assume a sample of exactly 100 g
2. Determine the whole number ratio of moles of the elements
   - When you have the moles of each element, divide each by the smallest # of moles
   - Recognize decimal values of simple fractions:

$^1/_2 = 0.50$ $^1/_3 = 0.33$ $^1/_4 = 0.25$ $^1/_5 = 0.20$

Determining Molecular Formulas

Once you have determined the empirical formula, you can determine the molecular formula if you know the molar mass of the substance:

1. Find the molar mass of the empirical formula
2. Divide by the molar mass of the substance
3. Multiply the # of atoms of each element by the answer

The empirical formula of Benzene is CH. If the molar mass of benzene is 78.11 g/mol, what is the molecular formula of Benzene?

Solutions

$CuSO_4$ in HCl and H_2O

Solute — Solvent

Solution

A solution can consist of multiple solvents

Units of Concentration

Solution = Solute + Solvent

[0.24] = 0.24M

$O_{2(g)}$ in $H_2O_{(l)}$

$K_3[Fe(CN)_6]_{(s)}$ in $H_2O_{(l)}$

$CoSO_{4(s)}$ in 1 M HCl

Molarity (M) – mol/L solution

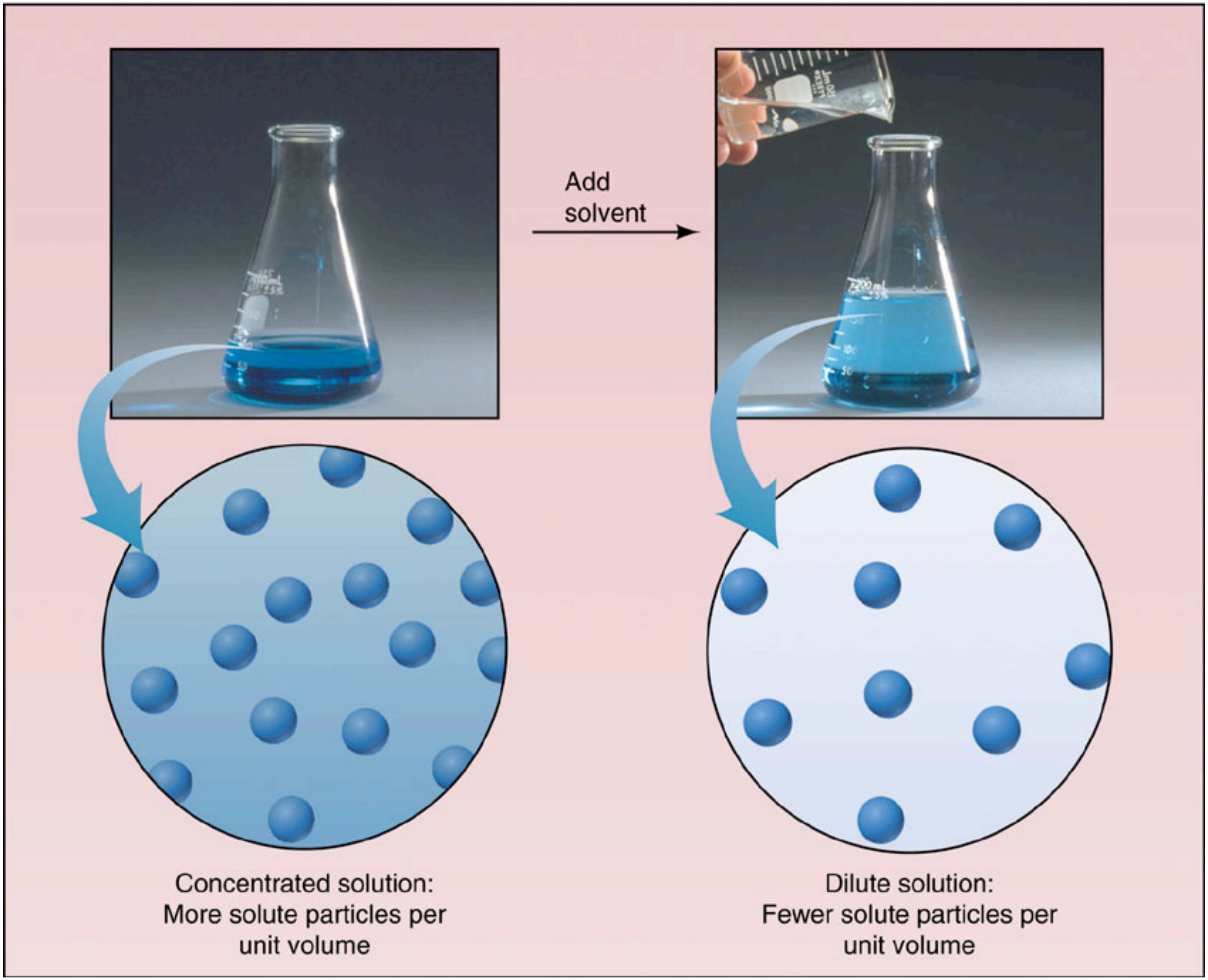
Molality – mol/kg solvent

w/v % - g solute / 100mL of solution

w/w % - g solute / 100g of solution

ppm – mg/L

ppb – µg/L or mcg/L


Add
solvent
Concentrated solution:
More solute particles per
unit volume
Dilute solution:
Fewer solute particles per
unit volume

Sample Dilution

Basic Skill

Number of molecules is invariant in a dilution

$$Concentration_i * Volume_i = Concentration_f * Volume_f$$

$$M_1V_1 = M_2V_2$$

How would you prepare a 100 mL sample of 2.0 M HCl solution from a stock 12.1 M HCl solution?

Chemical Equations

Chemical Reaction – a process in which one or more substances are chemically changed

- The starting materials are called the **reactants**
- The substances present after a reaction are called **products**

$$\text{Reactants} \rightarrow \text{Products}$$

A skeletal equation shows the reactants and products:

$$Na + H_2O \rightarrow NaOH + H_2$$

Law of Conservation of Mass – matter is neither created nor destroyed in a chemical or physical process

There are more H atoms in the products than in the reactants

This equation is unbalanced!

Chemical Equations - Types

• Synthesis	$A + B \rightarrow C$
• Decomposition	$A \rightarrow B + C$
• Single Displacement	$A + BC \rightarrow B + AC$
• Double Displacement	$AB + CD \rightarrow AD + BC$
• Combustion	$A + O_2 \rightarrow CO_2 + H_2O$
• Acid/Base	$HA + BOH \rightarrow H_2O + BA$

Chemical Equations

$$Na + H_2O \rightarrow NaOH + H_2$$

Must balance the equation

$$2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$$

There are now an equal # of atoms of each element in the reactants and the products

State symbols show the physical state of each substance:

(s) solid (l) liquid (g) gas (aq) aqueous

$$2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2(g)}$$

A 'Δ' symbol over the reaction arrow indicated high T is required

If a catalyst is used, it is indicated over the reaction arrow

Stoichiometry

Stoichiometry – quantitative aspect of chemical reactions

Chemical reaction Math!

Chemicals react in ratios of moles

The molar coefficients in a chemical reaction tell us these ratios:

$$\underline{\mathbf{2}}\ Na_{(s)} + \underline{\mathbf{2}}\ H_2O_{(l)} \rightarrow \underline{\mathbf{2}}\ NaOH_{(aq)} + H_{2(g)}$$

2 moles of $Na_{(s)}$ will react with 2 moles of $H_2O_{(l)}$
to produce 2 moles of $NaOH_{(aq)}$ and 1 mol of $H_{2(g)}$

Can make mole to mole predictions using the molar coefficients!

7.80 moles of Na reacts to produce how many moles of H_2 gas?

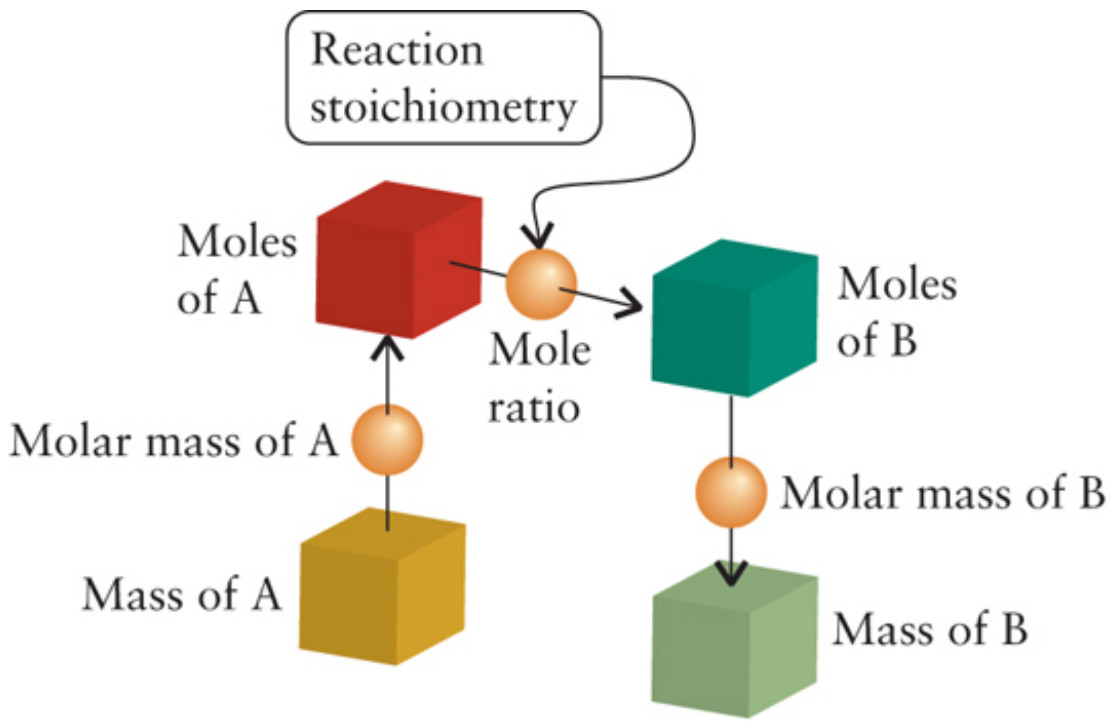
Stoichiometry

Mass to Mass Predictions require a little more work…

$\underline{2}\ Na_{(s)} + \underline{2}\ H_2O_{(l)} \rightarrow \underline{2}\ NaOH_{(aq)} + H_{2(g)}$

Convert Mass → Moles → Moles → Mass

7.80 g of Na reacts to produce what mass of H_2 gas?



Reaction Yield

Theoretical Yield – the maximum quantity of product that can be obtained from a given amount of reactants

% Yield – the percent of the theoretical yield actually produced:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Limiting Reactant – the reactant that determines the maximum yield of product

$$N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$$

If we have 2 moles of $N_{2(g)}$ and 3 moles of $H_{2(g)}$, we can only produce 2 moles of $NH_{3(g)}$

Reaction Yield

A mixture of 10.325 g of FeO and 5.734 g of Al react to produce 3.053 g Al_2O_3:

$$3\ FeO_{(s)} + 2\ Al_{(l)} \rightarrow 3\ Fe_{(l)} + Al_2O_{3(s)}$$

- What is the limiting reactant?
- What is the theoretical yield?
- What is the % yield?
- How much of the excess reactant is left over?

For Next Time

- Do the assigned questions for the fundamentals
- Prepare for quiz on fundamentals
- Read chapter 1
- Have a great 4th of July weekend!!