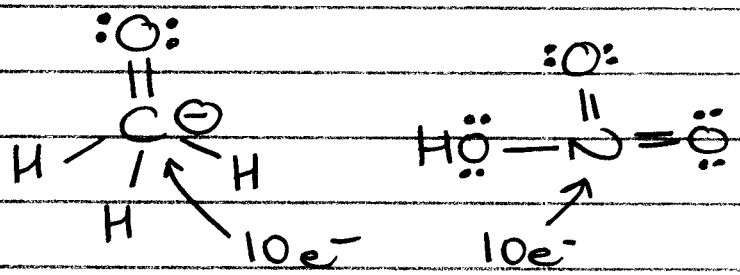
# CLARIFICATION OF THE OCTET RULE

Atoms in the second row of the periodic table only have four orbitals in their valence shell, i.e., 2s and 2p (x, y and z). Each orbital can hold a maximum of 2 electrons, hence only 8 electrons can be accommodated in the valence shell.

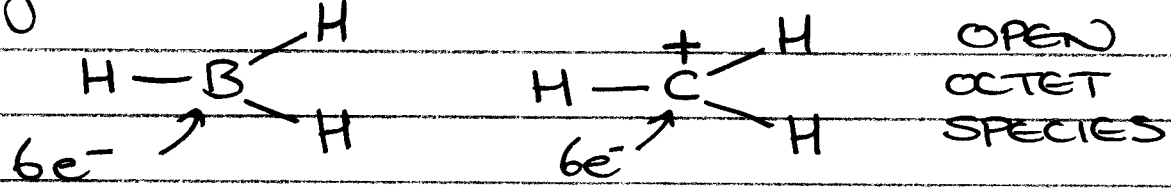
To 'violate' the OCTET rule, would be to draw a structure in which there are <u>more</u> than 8 electrons in such a valence shell, e.g.




In each case here, the OCTET rule is violated.

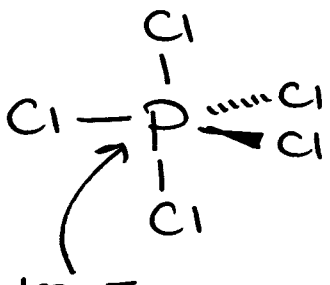
It is possible to draw structures in which there are <u>less</u> than 8 electrons in a valence shell, and this is <u>NOT</u> a violation of the octet rule, rather it is just an EXCEPTION, and generally such species are quite reactive (because they will want to try and attain a full valence shell)

e.g.

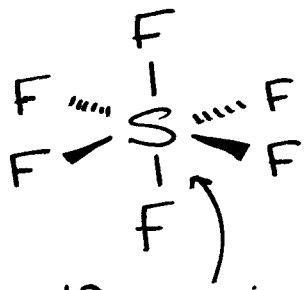



OPEN OCTET SPECIES

In the third row of the periodic table, d orbitals come into play, and so it is possible to 'EXPAND' the octet – there are now more orbitals available to accommodate electrons and so structures such as $PCl_5$ and $SF_6$ are OK.




$10e^-$ in valence shell




$12e^-$ in valence shell

Chemistry is not about blindly following a set of 'rules', e.g. the octet rule. You MUST understand the origins of these rules that you encounter, and thus understand the exceptions that are possible, and the violations that are not.

So, in summary, the octet rule in its strictest sense can never be violated, but having less than $8e^-$ in a valence shell is allowed, and we can 'EXPAND an octet' to more than $8e^-$ if there are d orbitals available to be filled.