## Section 7: Extensions of the Lewis Structure Model

With these thoughts in mind, we turn to a set of molecules which challenge the limits of the Lewis model in describing molecular structures. First, we note that there are a variety of molecules for which atoms clearly must bond in such a way as to have more than eight valence electrons. A conspicuous example is , where the sulfur atom is bonded to six F atoms. As such, the S atom must have 12 valence shell electrons to form 6 covalent bonds. Similarly, the phosphorous atom in has 10 valence electrons in 5 covalent bonds, the Cl atom in has 10 valence electrons in 3 covalent bonds and two lone pairs. We also observe the interesting compounds of the noble gas atoms, e.g.

, where noble gas atom begins with eight valence electrons even before forming any bonds. In each of these cases, we note that the valence of the atoms S, P, Cl, and Xe are normally 2, 3, 1, and 0, yet more bonds than this are formed. In such cases, it is not possible to draw Lewis structures in which S, P, Cl, and Xe obey the octet rule. We refer to these molecules as "expanded valence" molecules, meaning that the valence of the central atom has expanded beyond the expected octet.

There are also a variety of molecules for which there are too few electrons to provide an octet for every atom. Most notably, Boron and Aluminum, from Group III, display bonding behavior somewhat different than we have seen and thus less predictable from the model we have developed so far. These atoms have three valence shell electrons, so we might predict a valence of 5 on the basis of the octet rule. However, compounds in which boron or aluminum atoms form five bonds are never observed, so we must conclude that simple predictions based on the octet rule are not reliable for Group III.

Consider first boron trifluoride,

. The bonding here is relatively simple to model with a Lewis structure if we allow each valence shell electron in the boron atom to be shared in a covalent bond with each fluorine atom.

Note that, in this structure, the boron atom has only six valence shell electrons, but the octet rule is obeyed by the fluorine atoms.

We might conclude from this one example that boron atoms obey a sextet rule. However, boron will form a stable ion with hydrogen, $BH_4^-$

, in which the boron atom does have a complete octet. In addition, $BF_3$ will react with ammonia $NH_3$ for form a stable compound, $NH_3BF_3$, for which a Lewis structure can be drawn in which boron has a complete octet, shown here.

Compounds of aluminum follow similar trends. Aluminum trichloride, $AlCl_3$, aluminum hydride, $AlH_3$, and aluminum hydroxide, $Al(OH)_3$, all indicate a valence of 3 for aluminum, with six valence electrons in the bonded molecule. However, the stability of aluminum hydride ions, $AlH_4^-$

, indicates that Al can also support an octet of valence shell electrons as well.

We conclude that, although the octet rule can still be of some utility in understanding the chemistry of Boron and Aluminum, the compounds of these elements are less predictable from the octet rule. This should not be disconcerting, however. The octet rule was developed in Section on the basis of the observation that, for elements in Groups IV through VIII, the number of valence electrons plus the most common valence is equal to eight. Elements in Groups I, II, and III do not follow this observation most commonly.

Section 7: Extensions of the Lewis Structure Model is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.