3.2: Covalent Bonding

A second method by which atoms can achieve a filled valence shell is by *sharing* valence electrons with another atom. Thus fluorine, with one unpaired valence electron, can share that electron with an unshared electron on another fluorine to form the compound, F_2 in which the two shared electrons form a chemical bond holding the two fluorine atoms together. When you do this, each fluorine now has the equivalent of eight electrons in its valence shell; three unshared pairs and one pair that is *shared* between the two atoms. Note that when you are counting electrons, the electrons that are shared in the covalent bond are counted for each atom, individually. A chemical bond formed by *sharing* electrons between atoms is called a **covalent bond**. When two or more atoms are bonded together utilizing covalent bonds, the compound is referred to as a **molecule**.

There is a simple method, given below, that we can use to construct Lewis diagrams for diatomic and for polyatomic molecules:

- Begin by adding up all of the *valence electrons* in the molecule. For F_2, each fluorine has seven, giving a total of 14 valence electrons.
- Next, draw your *central atom*. For a diatomic molecule like F_2, both atoms are the same, but if several different atoms are present, the *central atom* will be to the *left* (or lower) in the periodic table.
- Next, draw the other atoms around the *central atom*, placing two electrons *between* the atoms to form a covalent bond.
- Distribute the remaining valence electrons, as pairs, around each of the outer atoms, so that they all are surrounded by *eight electrons.*
- Place any remaining electrons on the *central atom.*
- If the central atom is *not* surrounded by an octet of electrons, construct *multiple bonds* with the outer atoms until *all* atoms have a complete octet.
- If there are an *odd number* of valence electrons in the molecule, leave the remaining single electron on the *central atom.*

Let's apply these rules for the Lewis diagram for chlorine gas, Cl_2. There are 14 valence electrons in the molecule. Both atoms are the same, so we draw them next to each other and place two electrons between them to form the covalent bond. Of the twelve remain electrons, we now place six around one chlorine (to give an octet) and then place the other six around the other chlorine (our *central atom*). Checking, we see that each atom is surrounded by an octet of valence electrons, and so our structure is complete.

$$:\underset{\cdot\cdot}{\ddot{Cl}}\cdot\cdot\underset{\cdot\cdot}{\ddot{Cl}}:$$

All of the Group 7A elements (the halogens), have valence shells with seven electrons and all of the common halogens exist in nature as diatomic molecules; fluorine, F_2; chlorine, Cl_2; bromine, Br_2 and iodine I_2 (astatine, the halogen in the sixth period, is a short-lived radioactive element and its chemical properties are poorly understood). Nitrogen and oxygen, Group 5A and 6A elements, respectively, also exists in nature as diatomic molecules (N_2 and O_2). Let's consider **oxygen**; oxygen has six valence electrons (a Group 6A element). Following the logic that we used for chlorine, we draw the two atoms and place one pair of electrons between them, leaving 10 valence electrons. We place three pairs on one oxygen atom, and the remaining two pairs on the second (our *central atom*). Because we only have six valence electrons surrounding the second oxygen atom, we must move one pair from the other oxygen and form a second covalent bond (a **double bond**) between the two atoms. Doing this, each atom now has an octet of valence electrons.

$$:\underset{\cdot\cdot}{O}::\ddot{O}:$$

Nitrogen has five valence electrons. Sharing one on each atom gives the first intermediate where each nitrogen is surrounded by six electrons (not enough!). Sharing another pair, each nitrogen is surrounded by seven electrons, and finally, sharing the third, we get a structure where each nitrogen is surrounded by eight electrons; a noble gas configuration (or the "octet rule"). Nitrogen is a very stable molecule and relatively unreactive, being held together by a strong **triple covalent bond**.

$$:N\overset{\cdot\cdot}{::}N:$$

As we have constructed Lewis diagrams, thus far, we have strived to achieve an octet of electrons around every element. In nature, however, there are many exceptions to the "octet rule". Elements in the first row of the periodic table (hydrogen and helium) can only accommodate *two* valence electrons. Elements below the second row in the periodic table can accommodate, 10, 12 or even 14 valence electrons (we will see an example of this in the next section). Finally, in many cases molecules exist with single unpaired electrons. A classic example of this is oxygen gas (O_2). We have previously drawn the Lewis diagram for oxygen with an oxygen-oxygen double bond. Physical measurements on oxygen, however, suggest that this picture of bonding is not quite accurate. The magnetic properties of oxygen, O_2, are most consistent with a structure having *two unpaired* electrons in the configuration shown below:

$$: \ddot{\underset{\cdot}{O}} \cdot \cdot \ddot{\underset{\cdot}{O}} :$$

In this Lewis diagram, each oxygen atom is surrounded by *seven* electrons (not eight). This electronic configuration may explain why oxygen is such a *reactive* molecule (reacting with iron, for example, to form rust); the unpaired electrons on the oxygen molecule are readily available to interact with electrons on other elements to form new chemical compounds.

Another notable exception to the "octet rule" is the molecule NO (nitrogen monoxide). Combining one nitrogen (Group 5A) with one oxygen (Group 6A) gives a molecule with *eleven* valence electrons. There is *no way* to arrange eleven electrons without leaving one electron unpaired. Nitric oxide is an extremely reactive molecule (by virtue of its unshared electron) and has been found to play a central role is biochemistry as a reactive, short-lived molecule involved in cellular communication.

$$\cdot \ddot{\underset{\cdot}{N}} \cdot \cdot \ddot{\underset{\cdot}{O}} :$$

As useful as Lewis diagrams can be, chemists tire of drawing little dots and, for a shorthand representation of a covalent bond, a short line (a **line-bond**) is often drawn between the two elements. Whenever you see atoms connected by a line-bond, you are expected to understand that this represents two shared electrons in a covalent bond. Further, the *unshared pairs* of electrons on the bonded atoms are sometimes shown, and sometimes they are omitted. If unshared pairs ore omitted, the chemist reading the structure is assumed to understand that they are present.

$$\underset{\cdot\cdot}{\ddot{F}} \cdot \cdot \underset{\cdot\cdot}{\ddot{F}} : \;\; or \;\; : \underset{\cdot\cdot}{\ddot{F}} - \underset{\cdot\cdot}{\ddot{F}} : \;\; or \; F - F$$

$$: N \vdots\vdots N : \;\; or \;\; : N \equiv N : \;\; or \; N \equiv N$$