3.1: Compounds, Lewis Diagrams and Ionic Bonds

If we take two or more atoms and bond them together chemically so that they now behave as a single substance, we have made a chemical compound. We will see that the process of bonding actually involves either the *sharing*, or the net *transfer*, of electrons from one atom to another. The two types of bonding are **covalent**, for the sharing of electrons between atoms, and **ionic**, for the net transfer of electrons between atoms. Covalent or ionic bonding will determine the *type* of compound that will be formed.

In Chapter 1, we used atomic theory to describe the structure of the fluorine atom. We said that neutral fluorine has nine protons in its nucleus (an atomic number of 9), nine electrons surrounding the nucleus (to make it neutral), and the most common isotope has ten neutrons in its nucleus, for a mass number of 19. Further, we said that the nine electrons exist in two energy levels; the first energy level contains two electrons and is written $1s^2$. The second energy level contains seven electrons, distributed as $2s^2\ 2p^5$. The *outermost* electron level in any atom is referred to as the **valence shell**. For the representative elements (remember, this includes all of the elements except for the transition metals), the number of electrons in the valence shell corresponds to the Group number of the element in the periodic table. Group 1A elements will have one valence electron, Group 6A elements will have six valence electrons, and so on. Fluorine is a Group 7A element and has seven valence electrons. We can show the electron configuration for fluorine using a **Lewis diagram** (or *electron-dot structure*), named after the American chemist G. N. Lewis, who proposed the concepts of electron shells and valence electrons. In a Lewis diagram, the electrons in the valence shell are shown as small "dots" surrounding the atomic symbol for the element.

$$\cdot\ddot{\underset{\cdot\cdot}{F}}\cdot$$

When more than four electrons are present in the valence shell, they are shown as pairs when writing the Lewis diagram (but never more than pairs). Lewis diagrams for the atoms in the second period are shown below:

$$\cdot Li \qquad \cdot Be\cdot \qquad \cdot\dot{B}\cdot \qquad \cdot\dot{\underset{\cdot}{C}}\cdot$$

$$\cdot\ddot{\underset{\cdot}{N}}\cdot \qquad :\ddot{\underset{\cdot}{O}}\cdot \qquad \cdot\ddot{\underset{\cdot\cdot}{F}}\cdot \qquad :\ddot{\underset{\cdot\cdot}{Ne}}:$$

As you look at the dot-structures please understand that it makes *no difference* where you place the electrons, or the electron pairs, around the symbol, as long as pairs are shown whenever there are four or more valence electrons.

If you examine the Lewis diagram for neon (Ne) above, you will see that the valence shell is *filled*; that is, there are eight electrons in the valence shell. Elements in Group 8A of the periodic table are called *noble gasses*; they are very stable and do not routinely combine with other elements to form compounds (although today, many compounds containing noble gasses are known). Modern bonding theory tells us that this stability arises because the valence shell in the noble gasses is completely filled. When the valence shell is *not* full, theory suggests that atoms will transfer or share electrons with other atoms in order to achieve a filled valence shell… that is, the electron configuration of the noble gasses. Chemical bonding can then be viewed as a quest by atoms to acquire (or lose) enough electrons so that their valence shells are filled, that is, to achieve a "noble gas configuration". This is often referred to as the "**octet rule**"; the desire for elements to obtain eight electrons in the valence shell (except of course for helium where the noble gas configuration is *two* valence electrons).

Atoms can achieve a noble gas configuration by two methods; the transfer of electrons from their valence shells to another atom, or by sharing electrons with another atom. If you examine the Lewis diagram for lithium (Li), you will see that it has only one valence electron. If lithium was to *transfer* this electron to another atom, it would be left with two electrons in the 1*s*-orbital (denoted as $1s^2$). This is the same electron configuration as helium (He), and so by losing this electron, lithium has achieved a *noble gas configuration*. Because electrons carry a negative charge, the loss of this electron leaves lithium with a single positive charge. This is the **lithium cation** and it is shown as Li^+.

Returning to fluorine (F), in order to achieve the $2s^2\ 2p^6$ configuration of neon (Ne), fluorine needs to *gain* one valence electron. Because fluorine has *gained* one electron, it now has one negative charge. This is the fluoride anion and it is shown as F^-. The transfer of electrons in order to achieve a noble gas configuration is the process known as **ionic bonding**, and this will be covered in more detail later in this chapter.

Exercise 3.1.1

a. Sodium and chlorine are both third-period elements. Draw Lewis diagrams for each of these elements.
b. What number of electrons would chlorine have to gain in order to achieve a “noble gas configuration”? What would be the charge on chlorine?
c. What number of electrons would Na have to lose to obtain the noble gas configuration of Ne with *eight* valence electrons? What charge would Na have?