

ACTIVITY 6– ATOMS WITH MORE THAN ONE ELECTRON

Background Information

The importance of understanding electronic configurations for the elements should be stressed. Electronic configurations provide a theoretical foundation for the arrangement of the periodic table. Trends in ionization potential and other periodic relationships (such as atomic and ionic radii, magnetic properties, and electron affinity) may be predicted with knowledge of electronic configurations. The chemical reactivity of elements may be understood based on the arrangement of their electrons. We know that metals tend to lose electrons in chemical reactions and nonmetals tend to gain electrons. Why is this so?

Electronic Configurations

The fact that the elements were arranged in the periodic table years before electrons were even discovered is strong evidence for the assignment of electrons to orbitals and subshells. The layout of the table can be explained by electronic configurations.

Bohr's atomic model does not fit elements with more than one electron because these elements have the complex situation of having several electrons that are attracted to the nucleus while simultaneously repelling one another. We can describe these elements as having hydrogen-like orbitals, but the energy differences between these orbitals is greater than those of the hydrogen atom.

Electrons can be grouped into **shells** or main energy levels. A shell may contain electrons that are approximately the same distance from the nucleus and thus possess about the same energy. Electron shells are identified by a number, n , which can have the values, 1, 2, 3, and so on. These shells are also referred to as the principal quantum levels. The higher the number n , the higher the energy of the shell, and the greater the number of electrons that can occupy that shell. The principal shells are broken into sublevels, or **subshells** that are identified by the letters s , p , d , or f . The increasing order of subshell energies is $s < p < d < f$. Subshells are identified with both a number and a lowercase letter. The number indicates the shell to which the subshell belongs. A superscript number following the subshell letter indicates the number of electrons in the subshell. For example, the notation $1s^2 2s^2 2p^1$ indicates that this atom has two electrons in the $1s$ subshell, two electrons in the $2s$ subshell, and one electron in the $2p$ subshell.

The lowest energy shell ($n=1$) has only one subshell, $1s$. The second shell ($n=2$) has two subshells, $2s$ and $2p$. Elements with electrons only in the $2s$ and $2p$ subshells are found in the 2nd period. As you continue down the periodic table you will find that in the 3rd period you will have three sublevels (s , p , and d) and the fourth period will have four sublevels (s , p , d , and f). It is important to note that the s and p energies are found in the outermost quantum of each element and the d and f energies are found in the inner levels.

The subshells contain **orbitals** that may accommodate no more than two electrons (of opposite spin, according to **Pauli's Principle**). An s subshell can accommodate

up to 2 electrons in its orbital. A p subshell can accommodate 6 electrons in its three equal energy (**degenerate**) orbitals. A d subshell can be occupied by up to 10 electrons in its five degenerate orbitals. An f subshell can accommodate 14 electrons in its seven degenerate orbitals. The notation used to designate orbitals is the same as that used for subshells. We call the three orbitals in the $3p$ subshell the $3p$ orbitals.

The maximum number of electrons in any given shell can be calculated by the following equation: $2n^2$.

Ionization Energies

As we study the graph of the ionization energies of the elements, we find that every time we reach a noble gas, it is at the highest point for that period of the periodic table. Recall that the **ionization energy** (I.E., also called ionization potential) is defined as the amount of energy required to remove the most loosely bound electron from a gaseous atom. If we look at the electronic configuration of the noble gases, we see that their highest occupied shell is completely filled. This imparts a special degree of stability to these atoms. Remember that atoms do not spontaneously eject electrons; energy must be supplied to remove one. The stability of the noble gases is reflected in their high ionization energies.

The particular stability of the noble gases is a state that other elements would like to achieve. For example, if we look at the electron configuration of sodium, we find that it has one more electron than neon. If the noble gas electron arrangement is considered to be very stable, then it should

seem logical that sodium would want to give up an electron quite easily. When we look at the ionization energy graph we find that the alkali metals do give up their valence electron quite easily. However, when we look at an element like chlorine, we find that it needs to gain only one electron in order to have the same number of electrons (to be **isoelectronic**) as argon. So we would assume that it does not want give up an electron very easily and that is confirmed by the ionization energy graph.

Looking across the 2nd period we see an increase in ionization energy going from Li to Be. This makes sense as the nuclear charge has increased. Going from Be to B, however, we notice a slight decrease in I.E. This can be explained by the fact that removing a $2p$ electron from boron is more easily accomplished than removing a $2s$ electron from beryllium. Going from B to C to N, we see an expected increase in I.E., again due to increased nuclear charge. The irregularity seen upon going from B to O can be explained by considering repulsions between electrons in an orbital. One of the paired electrons in oxygen's $2p$ orbital is more easily removed than unpaired electron in nitrogen's half-filled $2p$ orbital. There is the expected increase going from O to F.

We see that for every element, the 2nd ionization energy is larger than the first. This is due simply to the fact that once an atom has given up one electron it becomes a positively charged ion (a cation). It is more difficult to remove the second electron from the ion, as the electrons are more highly attracted to the nucleus.