ACTIVITY 7– HOW ELECTRONS DETERMINE CHEMICAL BEHAVIOR

Background Information

The ionization potential of each atom was a great aid in understanding how the electrons of each atom are arranged. From this knowledge we were able to develop the electron configuration and found that each family had identifiable electron arrangements. For example, the alkali metals' outermost electron was always in the highest quantum energy level and was located in the *s* orbital. The halides on the other hand, always have their final electron placed in a *p* orbital, giving them a total of 5 electrons in their *p* orbitals.

If we take this one more step, we can look at the property of electron affinity. Electron affinity can be defined as the amount of energy released, or in some cases gained, when an electron is added to an atom. When an electron is added to a halide it then has the same electron configuration (isoelectronic) as its neighboring noble gas. Since noble gases are very stable, we find that when an electron is added to a halide it releases a large amount of energy. This implies that the halide anion is more stable than the gaseous halide atom would be. If we try to add an electron to a noble gas the process would require addition of energy, and this implies that the noble gas is more stable than the resulting anion would be.

Metals have low electron affinities and this supports our understanding of why they also have low ionization energies. So we can conclude that metals want to lose electrons and nonmetals want to gain electrons to achieve a more stable configuration. Alkali and alkaline earth metals want to lose electrons quite easily to form cations and when they come in contact with halides we see the formation of ionic compounds.

If we look at the periodic table and speak in general terms we can say that the elements in the upper right corner have the highest ionization energies and as you go left and down the periodic table, the ionization energies will decrease. Also, the electron affinity would follow the same pattern if we exclude the noble gases. With a little more information, we will soon be able to develop the electronegativity scale.

Goals and Assessment

Clarify that the goals indicate what the students should be able to do as a result of the activity. Make sure that students understand that the **Chapter Challenge** is based on these goals.

Goal	Location in Activity	Assessment Opportunity
Investigate more patterns in the electron arrangements of atoms.	**Investigate** Step 1	Students' answers match those given in Teacher's Edition.
Relate the positions of elements on the periodic table, their electron arrangements, and their distances from the nearest noble gas, to chemical properties of the elements.	**Investigate** Step 2 **Chemistry to Go** Questions 1-2	Students can describe and give explanations for the chemical behaviors of elements. Students' answers match those given in Teacher's Edition.
Relate electron arrangements to ionization energies.	**Chemistry to Go** Question 3 **Inquiring Further** Question 2	Students are able to explain which element of a pair has the more stable electron arrangement. Students can explain relative sizes of first, second, and third ionization energies of an element.
Assign valence numbers to elements and organize the periodic table according to valence numbers.	**Investigate** Steps 3-4 **Chemistry to Go** Questions 4-5 **Inquiring Further** Question 1	Students' answers match those given in Teacher's Edition.