

**Illustrate the periodic law with graphs of selected properties of the elements as a function of atomic number.**

The periodic law states that the physical and chemical properties of the elements vary periodically with atomic number. This is shown in the graphs of atomic volume, atomic radius, ionization energy and melting point against atomic number. The same general pattern of peaks and valleys occurs in each graph. Although peaks or valleys do not occur for elements in the same family on each graph, there is one peak and one valley within each period of the periodic table for each property that is plotted. Similar periodic variation is evident in a periodic table of atomic properties.

**Use the terms periods, groups, families, representative elements and transition elements to describe individual elements and groupings of elements in the periodic table.**

A period is a horizontal row in the periodic table. There are seven periods. A group or a family is a column in the periodic table. Elements of the same family are called cogeners. For example, He, Ne, Kr, Xe, and Rn are cogeners. Several families have special names. Family 1A is the alkali metals. Family 2A is the alkaline earth metals. Family 7A is the halogens (salt formers). Family 8A is the noble gases. Family 6A is the chalcogens (chalcogen former). Family 1B is the coinage metals. Fe, Co and Ni form the iron triad [ferromagnetic elements]. Ru, Rh, Pd, Os, Ir, and Pt are the noble metals. The “B” families (the waist of the periodic table) is the transition metals. Ce through Lu are the lanthanide metals. Th through Lr (and beyond) are the actinides. The lanthanides and actinides together are the inner transition elements. The “A” families are the representative [main group] elements.

**Use the periodic table to describe the Aufbau process and explain the basic features of the electron configurations of the representative and transition elements, especially the number of valence electrons in each representative group.**

We used the periodic table to help remember the order of sub-shell filling. Period 1 fills 1s; period 2 fills 2s and 2p; period 3 fills 3s and 3p. Period 4 fills 4s followed by 3d then 4p and so on. The electron configuration of an element determines its physical and chemical properties. An atom displays mainly its outer electrons—those in the shell of highest principal quantum number. These outer shell electrons are called valence electrons and are said to be in the valence shell. Therefore, there is a major difference in the chemical and physical properties between elements in different representative families, since each family differs in the number of outer shell electrons. There is not so large a difference between transition families, because these families differ by inner-shell electrons. The number of outer shell (that is **s** and **p**) electrons of a representative element equals its family number.

**Describe metals, nonmetals, metalloids and noble gases in several ways and locate them in the periodic table.**

Noble gases are the elements in group 8A. Nonmetals are in the triangular block in the upper right corner of the A groups: H, C, N, O, F, P, S, Cl, Se, Br, I. Metalloids are in the two diagonal rows just below and to the left of the nonmetals: B, Si, Ge, As, Sb, Te, Po and At. Metals are the remaining elements.

Physically, pure metals generally are malleable and ductile, melt at moderate temperatures and conduct heat and electricity well. These properties may change substantially when two metals are alloyed or melted together.

Nonmetals tend to be insulators and form small molecules that condense poorly into solids; hence the solids often melt easily (except carbon with its high melting point). Chemically, metals

lose electrons and form cations while nonmetal gain electrons and form anions. Metallic oxides form basic solutions with water while nonmetallic oxides form acidic aqueous solution. Metals rarely combine chemically with each other. Nonmetals combine chemically to form definite compounds with nonmetallic properties. When metals combine with nonmetals, ionic compounds are formed. These compounds have high melting points and the melt (liquid state) conducts electricity while the solid does not.

Metalloids have physical and chemical properties that are between those of metals and those of nonmetals. For example, the aqueous solutions of their oxides often are amphoteric, meaning that they can react as acids or as bases.

**State the factors that influence atomic size, ionic size; distinguish among covalent, ionic, metallic bonds and describe the general trends in atomic size that occur within families and groups.**

Atoms are not small hard balls but indefinite spheres. Defining where the electron cloud of an atom ends is like trying to measure the extent of the metropolitan area of a city. (Such a non-bonded interaction defines the Van der Waals radius.) But with both cities and atoms we can determine the center. For atoms, one half of the distance between two centers is called the atomic radius. This idea is complicated by the fact that not all atoms are normally bound together in the same way. Some are bound by covalent bonds in molecules, some are attracted to each other in ionic crystals, and some are held in metallic crystals by the force of a sea of electrons acting as glue between cations. (The distance between nuclei in such a metallic crystal equals the metallic radius) Fortunately, it is possible to form molecules of nearly every element (except the smaller noble gases) in which two like atoms are held together by a covalent bond. The covalent radius of these molecules often is called the atomic radius.

Atomic size regularly increases from top to bottom of a family. More electrons are present with the same outer configuration are present around the nucleus and the electron of interest is in a higher energy level. (increases in the orbital size in successive principal quantum levels). There is a lower attraction for the outer most electrons even though  $Z_{\text{eff}}$  is increasing slightly down a family.

Atomic size gradually decreases from left to right in a period of representative elements. The explanation is that within a family all electrons are being added to the same shell. It is more populated. At the same time, protons are being added to the nucleus, making it more positively charged. This means the nucleus attractions are stronger, and even though the electron–electron repulsions increase, they are not large enough to counter act the increased nuclear charge ( $Z$ ).  $Z_{\text{eff}}$  essentially is increasing; the shell is pulled closer to the nucleus.

When electrons are removed from the atom, the electron–electron repulsions go down. The nuclear attraction goes up :  $Z_{\text{effective}}$  increases. This means when electrons are added, the electron–electron repulsions go up and the nuclear attractions decrease:  $Z_{\text{effective}}$  decreases.

Across a transition series, size gradually but somewhat irregularly decreases then increases at the end of the series. Within a transition series electrons are added to an inner **d** shell. This inner shell gradually will decrease in size, but since these **d** electrons are between the nucleus and the outer electrons, they “shield” each other from the nuclear charge. Thus the outer shell should increase in size. The two trends oppose each other with the shielding effect initially being less important, but predominating at the end of the transition series. A consequence of the atomic radius being the smallest in family 8B—the iron triad (Fe, Co, Ni) and the noble metals (Ru, Rh, Os, Ir, Pt) is that these metals are denser than those in the rest of the periodic table. In fact, Os and Ir are the densest substances known.

Across the lanthanides, atomic radii gradually but irregularly decreases and the radii of the +3 cations steadily decrease. This steady decrease in size, called the lanthanide contraction, is due

to the addition of electrons to a deep inner sub-shell (4f). For most lanthanides the relative sizes of ions are very much smaller than the atoms. The lanthanide contraction helps to explain the similarity of chemical and physical properties between the transition elements of the fifth (Y to Cd) and sixth (La to Hg) periods. Because of the lanthanide contraction, atomic sizes are almost the same in these two transition periods. Since the outer electron configurations are the same, the properties are very similar.

The sizes of iso-electronic ions decrease as the positive charge on the ions increase. The explanation is that the larger nuclear charge exerts more pull on each electron and thus pulls these electrons closer to the nucleus.

### **Explaining shielding, $Z$ , $Z_{\text{effective}}$ and penetration of electrons to the nucleus.**

In the poly electron atom, there are three energy contributions that must be considered: the kinetic energy of the electrons as they move around the nucleus, the potential energy of attraction between the nucleus and the electrons, and the potential energy of repulsion between the two electrons. If we think of this problem in terms of the quantum mechanical model, the Schrödinger equation cannot be solved exactly. The difficulty arises with the repulsions of the electrons. We, therefore, make an approximation: treat each electron as if it were moving in a field of charge that is the net result of the nuclear attraction ( $Z$ ) and the average repulsions of all the other electrons. We single out the outermost electron and consider the forces that electron feels. Take sodium: 11 protons, 10 electrons + 1 outermost electron. The electron should be attracted to the positive 11 charge, but it also feels the repulsions of the other 10 electrons. The net effect is that the electron is not bound nearly as tightly as it would be if the 10 electrons were gone. The outermost electron is shielded or screened from the full nuclear charge by the other electrons.

The picture of the atom we use is the hydrogen atom. In other words, we are thinking of spheres. They have the same shape as hydrogen but with different energies. The difference occurs because of the effect of the nuclear attraction versus the electron repulsions.

### **Define first, second, ... ionization energies; describe the factors that affect the magnitude of these ionization energies ; and relate ionization energies to the location of elements in the periodic table.**

Ionization energy (I. E. or IP) is the energy required by the process  $M_{(g)} \rightarrow M^+_{(g)} + 1 e^-$ . The ease of losing an electron is a measure of the ability of an element to act as a metal. Ionization energies are measured in kJ/mol or in electron volts per atom, where 1 eV/atom = 96.49 kJ/mol. Ionization energy decreases down a family of representative elements. As atoms get larger, the outermost electrons are farther from the nucleus and hence are less strongly held. Ionization energy gradually but irregularly increases across a representative period. As atomic size decreases, electrons are more strongly attracted to the nucleus. The trend is irregular in that it is harder than expected to remove electrons from  $s^2$  and  $p^3$  configurations and relatively easier to ionize  $p^1$  and  $p^4$  configurations. This demonstrates the unusual stability of full and half full sub-shells.

As electrons are removed from the valence shell, the electron–electron repulsions are decreased. The nuclear charge is stronger ( $Z_{\text{effective}}$ ). It is harder to remove further electrons, especially once the noble gas core is reached.

### **Define electron affinity; cite the factors influencing its magnitude; and describe the variation of electron affinity with in periods and groups of the periodic table.**

Electron affinity is the energy required by the process  $X_{(g)} + e^- \rightarrow X^-_{(g)}$ . This is an indication of an element's degree of being nonmetallic. Two trends are evident. Electron affinity increases

(becomes more positive, indicating a less exothermic reaction) down a group and decreases gradually across a period (becomes more negative, indicating a more exothermic reaction).

The increase in electron affinity down group is explained by each atom being larger than the atom above it in the group. This means that the added electron is further away from the nucleus than in the smaller (and lighter) atom. With a larger distance between the negatively charged electron and the positively charged nucleus, the force of attraction is relatively smaller. Thus electron affinity increases—becomes more positive, as one proceeds down a group.

The trend toward smaller (more exothermic) electron affinities across a period also is size related. Recall that atoms become smaller across a period. The irregularities are due to adding an electron to a full shell or a half filled sub-shell. ( $s^2$  and  $p^3$ ) because these two electron configurations are relatively stable, adding an electron to them is not very favorable energetically. The electron–electron repulsions have increased. The added electron feels a lower nuclear charge. In addition, adding an electron to an  $s^1$  or a  $p^2$  or  $p^5$  electron configuration will result in the creation of a full or half filled sub-shell. Creation of such a stable electron configuration is energetically favored, an exothermic process.

Although the first electron affinity often is negative (an exothermic process), the second EA is usually positive. Adding an electron to an ion that has a charge of  $-1$  is a process that is expected to require energy.

### **Relate the metallic character of an element to its position in the periodic table.**

The ease of losing an electron is a measure of an element's metallic character. This ease is measured by ionization potential, which decreases from right to left in a period and from top to bottom in a family. The eagerness to gain electrons is a measure of an element's nonmetallic character. This is the electron affinity. It becomes more negative from left to right and bottom to top. A good metal is a poor nonmetal. Metallic character decreases from left to right and bottom to top.