Periodic Trends

Many properties of the elements change in a predictable way as you move through the periodic table. These systematic variations are called **periodic trends**. In this lesson, we will examine several of these periodic trends.

Atomic Radius

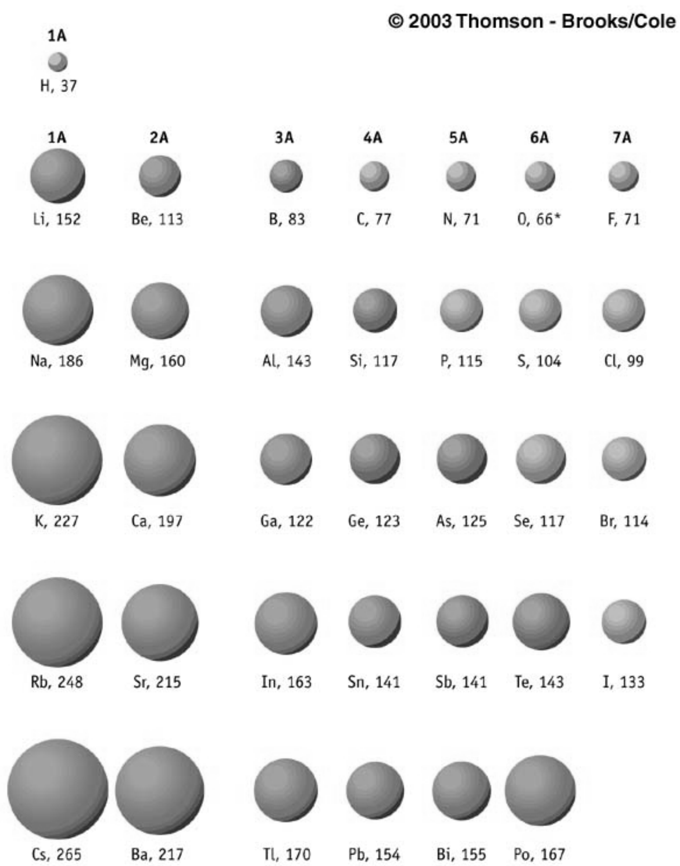
The **atomic radius** is the distance from the center of an atom's nucleus to its outermost electron. On the periodic table, two trends can be observed in regards to atomic radius.

1. Atomic radius increases as you move down a group.

2. Atomic radius decreases as you move from left to right across each period.

The first trend is easier to explain. As you move down a group, the principal quantum number of the outermost electrons increases. Electrons with a larger principal quantum number are found in orbitals that are farther away from the nucleus, which makes the atomic radius larger.

The second trend is a bit more difficult. In any period, the outer electrons of each element are in orbitals with the same principal quantum number. As you move from left to right across a period, the atoms' nuclei gain more protons. Atoms that have more protons in their nucleus exert a stronger pull on the electrons in a given principal quantum level. This stronger attractive force shrinks the electrons orbitals and makes the atom smaller.



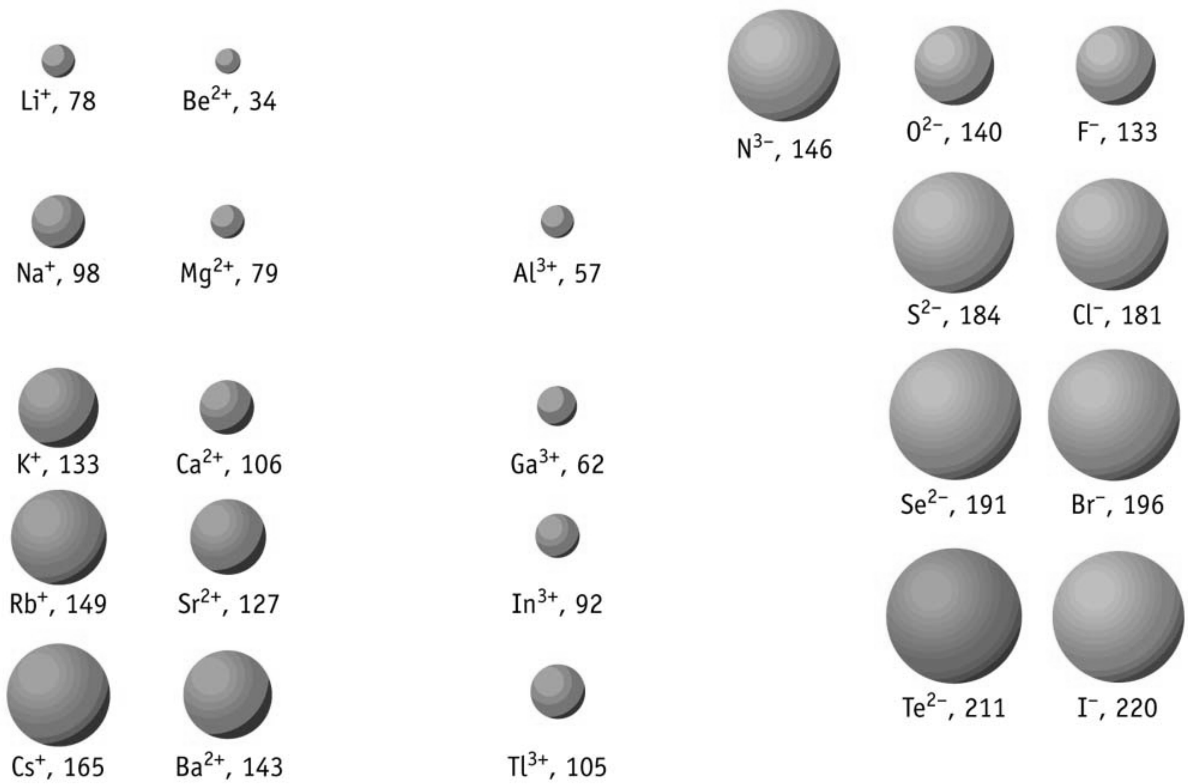
Ionic Size

As you have learned in past courses, an atom can gain or lose electrons to form an ion.

When an atom loses electrons (becoming a positive ion), it becomes smaller. There are two reasons for this. First, the loss of its valence electrons results in the atom's outermost orbital becoming empty. Second, the loss of one or more electrons reduces the repulsive forces between the remaining electrons, allowing them to pull closer together (and, therefore, closer to the nucleus).

When an atom gains electrons (becoming a negative ion), it becomes larger. This occurs because the addition of one or more electrons increases the repulsive forces between the remaining electrons. This causes them to spread farther apart, forcing them to move farther from the nucleus.

In addition to their different sizes, the diagram below illustrates other periodic trends for the common ions.



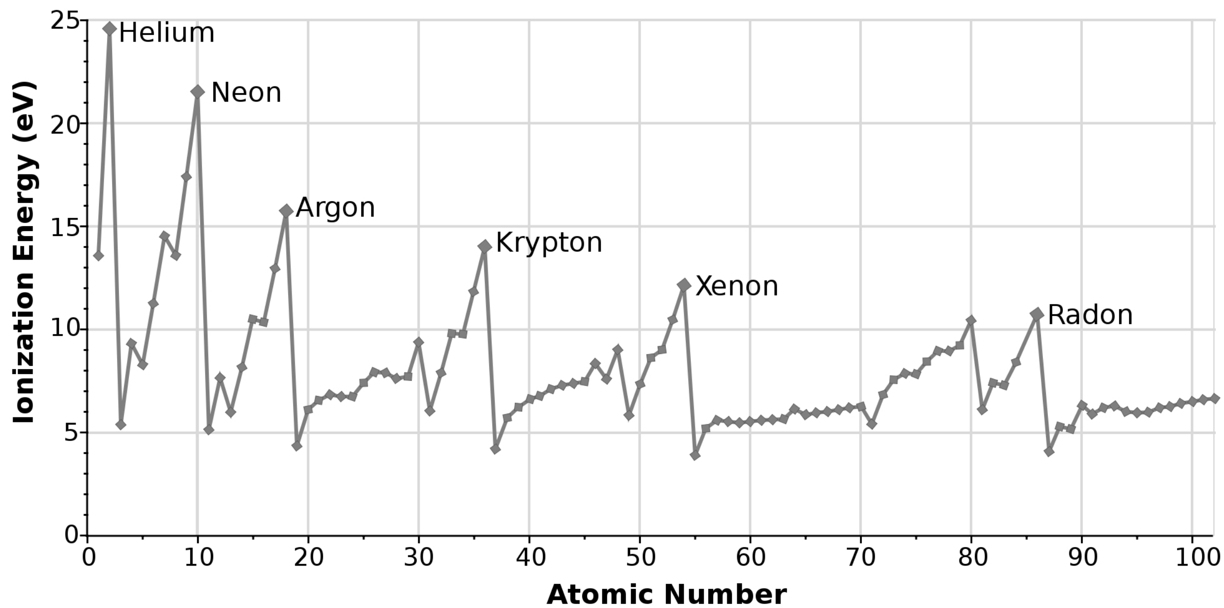
There are three trends that can be observed:

1. Ionic radii increase as you move down a group.
2. The ionic radii of positive ions decreases as you move from left to right across a period.
3. The ionic radii of negative ions decreases as you move from left to right across a period.

Ionization Energy

An atom's **ionization energy** is the energy needed to remove one of its electrons. The ionization energy is measured in the gaseous state. The ionization energy of an atom indicates how strongly the atom holds onto its valence electrons. High ionization energy values indicate the atom has a strong hold on its valence electrons. Low ionization energy values indicate the atom has a weak hold on its valence electrons. Atoms with high ionization energy values are unlikely to lose electrons and form positive ions.

The graph below shows how ionization energy varies with atomic number.



With respect to the periodic table, ionization energies show two important trends.

1. Ionization energies decrease as you move down a group.
2. Ionization energies increase as you move from left to right across a period.

Each time an electron is removed from an atom it requires more energy. Once all of the valence electrons have been removed from an atom, the amount of energy required to remove the next electron increases quite dramatically. This shows that an atom holds the electrons in its inner core much more strongly than it holds its valence electrons.

It is important to note that, while it is technically possible, atoms on the right side of the periodic table are far less likely to lose their valence electrons and form positive ions.

Electronegativity

An atom's **electronegativity** reflects its ability to attract electrons in a chemical bond. The periodic table below lists the electronegativities of the elements.

H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	Ac 1.1	Unq	Unp	Unh	Uns	Uno	Une									

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

Notice that electronegativity values have no units. Fluorine is the most electronegative element, with an electronegativity of 4.0. The least electronegative elements are in the bottom left corner of the periodic table. Both cesium and francium have electronegativities of 0.7.

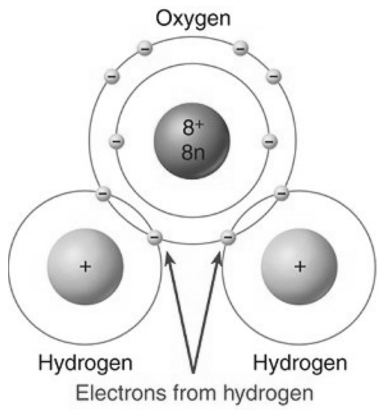
In general, electronegativity increases as you move from left to right across a period, and decreases as you move down a group.

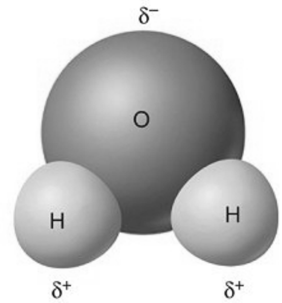
Electronegativity and Chemical Bonds

A covalent bond is a shared pair of electrons between two atoms. When atoms with different electronegativities form a covalent bond, the shared electrons are more strongly attracted to the atom that is more electronegative.

When one atom is significantly more electronegative than another, a covalent bond between them is said to be **polar**. In a polar covalent bond, the atom with greater electronegativity gains a slightly negative charge because the electrons in the bond are slightly closer to it. At the same time, the atom with lower electronegativity gains a slight positive charge.

The diagram on the next page illustrates a polar bond between a hydrogen atom and an oxygen atom in a molecule of water. Because oxygen is more electronegative than hydrogen, the electron pair that forms the bond between them is pulled closer to the oxygen atom. As a result, the oxygen atom in a water molecule has a slight negative charge, and the hydrogen atom has a slight positive charge.





When two atoms that have similar electronegativities form a covalent bond, the electron pair is shared equally. This kind of bond is called a **nonpolar** covalent bond. The best examples of nonpolar covalent bonds are found between atoms of the same element, such as the single bond between two fluorine atoms in molecular fluorine (F_2).

The three kinds of bonds that we know — polar covalent, nonpolar covalent, and ionic — form a spectrum of bond types. At one end of the spectrum are ionic bonds. At the other end are nonpolar covalent bonds. Polar covalent bonds are found between these two extremes.

You can use electronegativity to predict whether a bond will be nonpolar, polar, or ionic. This is done by determining the difference between the electronegativities of the atoms involved.

Electronegativity Difference	**Bond Type**
≤ 0.4	nonpolar covalent
between 0.4 and 2.0	polar covalent
≥ 2.0	ionic