Ionic Bonds

<u>Valence electrons</u>
- the electrons in the highest occupied energy level
- always electrons in the "s" and "p" orbitals
- maximum of 8 valence electrons
- elements in the same group have the same number of valence electrons (He only has 2, while other noble gases have 8, because the 1st energy level can only hold 2 e^-)
- valence electrons are the electrons involved in forming chemical bonds

<u>Electron dot symbols</u>
- the inner electrons and the nucleus are represented by the symbol of the element
- the valence electrons are shown as dots

H· He:

		·	·	·	·	·	··
Li·	·Be·	·B·	·C ·	:N·	:O ·	:F:	:Ne:
			·	·	··	··	··

<u>Octet Rule</u> - atoms react by sharing or transferring electrons in order to obtain a stable noble gas configuration. Noble gases (except He) have 8 valence electrons.

Metals lose electrons and form positive ions called "cations".

Ex. Na $1s^2 2s^2 2p^6 3s^1$

Sodium has one valence electron, which it loses to become:

Na^+ $1s^2 2s^2 2p^6$

The sodium ion has eight valence electrons, the same as the noble gas, neon.

Nonmetals gain electrons and form negative ions called "anions".

Ex. Cl $1s^2 2s^2 2p^6 3s^2 3p^5$

Chlorine has seven valence electrons, it gains one to become:

Cl^- $1s^2 2s^2 2p^6 3s^2 3p^6$

The chlorine ion (chloride) has eight valence electrons, the same as the noble gas argon.

Ionic Compounds

Ionic bonds are formed when one or more electrons are transferred from one atom to another atom. This produces positively and negatively charged particles called ions. The forces of attraction that bind oppositely charged ions together are called ionic bonds. Ionic bonds generally occur between metals and nonmetals.

Properties of Ionic Compounds

- crystalline solids at room temperature
- the ions are arranged in a repetitive three dimensional pattern called a lattice structure
- conduct electricity when molten or aqueous, but do not conduct electricity when solid
- many are soluble in water
- high melting and boiling points
- hard but brittle
- if enough force is applied, the crystal will break apart along smooth, flat surfaces

Covalent Bonds

Covalent bonds occur when atoms share electrons in order to obtain a stable noble gas configuration.

<u>Single Covalent Bonds</u> – Two atoms share one electron each in order to obtain eight valence electrons.

Ex. Cl_2 Each chlorine atom has seven valence electrons, if they share one electron each, they will have eight valence electrons.

:Cl· + **:**Cl· ➔ **:**Cl**:**Cl**:** or **:**Cl**-**Cl**:**

<u>Double Covalent Bonds</u> - Two atoms share two electrons each in order to obtain eight valence electrons.

Ex. O_2 Each oxygen atom has six valence electrons, if they share two electrons each, they will have eight valence electrons.

:O· + **:**O· ➔ **:**O**::**O**:** or **:**O**=**O**:**

<u>Triple Covalent Bonds</u> - Two atoms share three electrons each in order to obtain eight valence electrons.

Ex. N_2 Each nitrogen atom has five valence electrons, if they share three electrons each, they will have eight valence electrons.

:N· + **:**N· ➔ **:**N**:::**N**:** or **:**N≡N**:**

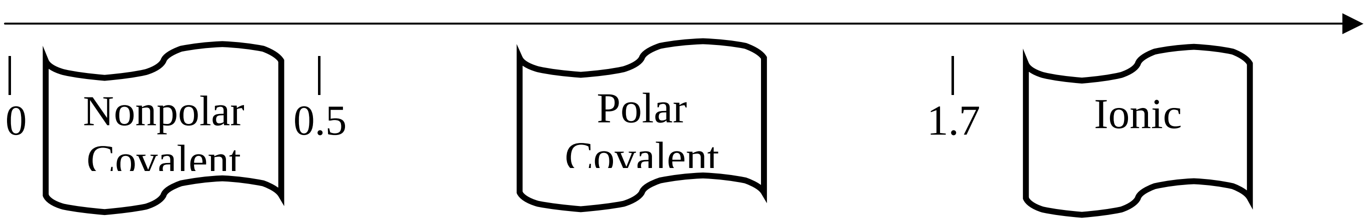
Polar Covalent Bonds – Two atoms share electrons, but they do not share them equally. One atom is more electronegative than the other atom. The shared electrons spend more time closer to the more electronegative atom, causing it to become slightly negative (δ^-) and the less electronegative atom becomes slightly positive (δ^+).

Nonpolar Covalent Bonds – The electrons are shared equally.

Ionic Bonds – The difference in electronegativities is so great that the electron(s) is no longer shared. The electron(s) is transferred from the less electronegative atom to the more electronegative atom. This produces positive and negative ions.

There is no sharp distinction between ionic and covalent bonding. The polarity of the bond depends on the nature of the elements involved, and their ability to attract electrons. The bonding continuum ranges from equal sharing, to unequal sharing, to complete electron transfer. The type of bonding obtained depends on the difference in electronegativities in the atoms involved in the bond.

Difference in Electronegativities



Properties of Covalent Compounds

Covalent compounds exist as gases, liquids, and solids at room temperature. Solid covalent compounds are usually soft. Compared with ionic compounds, covalent compounds usually evaporate readily and have low melting points and boiling points. Many covalent compounds are not soluble in polar substances such as water, but they are soluble in nonpolar substances such as gasoline. They do not conduct electricity in either the liquid or the solid state.

Metallic Bonds

In nonmetal atoms the valence electronic shells generally contain more electrons than they do vacant or partially filled orbitals, the opposite is true for metal atoms – they have more valence shell orbitals than electrons. A bonding scheme for metals must also account for the distinctive properties that metals share.

1. Ability to conduct electricity.
2. Ability to conduct heat.
3. Ease of deformation [i.e., ability to be flattened into sheets (malleability) and to be drawn into wires (ductility)].
4. Lustrous appearance.

One simple model that can account for these properties is the **electron-sea model**. The metal is pictured as a network of positively charged ions immersed in a "sea of electrons".

Because the valence electrons are delocalized, and not bound to any particular nucleus, they are easily displaced by an externally applied electric field. This delocalization of electrons is what gives metals their characteristic properties. For example, if the ends of a bar of metal are connected to a source of electric current, electrons from the external source enter the bar at one end. Free electrons pass through the metal and leave the other end at the same rate. In thermal conductivity no electrons enter or leave the metal, but those in the region that is being heated gain kinetic energy and transfer this to other electrons.

Delocalized electrons also explain the ease of deformation of metals. In the metal all the ions are positively charged, but the sea of electrons that flows around them offsets their mutual repulsions. If one layer of metal ions is forced across another (ex. by hammering) the internal structure remains essentially unchanged because the sea of electrons adjusts rapidly to the new situation. However, in an ionic crystal, displacement of one layer of ions with respect to another brings like charged ions into proximity. The strong repulsive forces set up between them can cause the ionic crystal to cleave or shatter.

The number of outer electrons available determines the properties of metals. Group 1 metals have only one valence electron and are soft. Group 2 metals are harder than group 1 metals. In the transition elements, electrons from the partially filled d orbitals may take part in the metallic bond. Many of these metals are very hard.

The strong metallic bond of the structural metals, such as iron, chromium, and nickel, makes them hard and strong. It is possible to strengthen some of the elements with fewer delocalized electrons by combining them with other metals to form alloys. (See examples – page 311, chemistry textbook)

Intermolecular Attractions

Intramolecular forces – are those holding atoms together in molecules, i.e., covalent bonds.

Intermolecular forces – are those holding molecules together, i.e., van der Waals forces.

Van der Waals forces include dipole-dipole forces and dispersion forces.

Dipole-dipole attractions occur between polar molecules. The slightly negative end of one molecule is attracted to the slightly negative end of another molecule. Dipole interactions are similar to, but much weaker than, ionic bonds.

Dispersion forces (London forces) – When electrons move around in an atom, their motion is somewhat random. In any atom, there is a chance that at a given instant there will be more electrons on one side of the atom than the other. Thus, one side of the atom will be slightly negative, and the other will be slightly positive. The atom will contain an instantaneous dipole. This dipole can induce a dipole in a neighboring atom. Then, since the negative end of one dipole is close to the positive end of the other dipole, there is an attraction between them.

Dispersion forces are quite weak because the instantaneous dipoles exist only for a moment and then only when the particles are very close to each other. Dispersion forces increase as the number of electrons and the size of the atom (or molecule) increases

Hydrogen bonds occur only when hydrogen atoms are bonded to extremely electronegative elements such as fluorine, oxygen, or nitrogen. These electronegative elements form covalent bonds with hydrogen atoms, which are of such an extremely polar nature that the attractive forces between neighbouring molecules are too large to be considered normal van der Waals forces. A hydrogen bond is a force of attraction, the attraction between the hydrogen atom in one molecule to a highly electronegative atom in another molecule.