Chemical Bonding

Chemical compounds are formed by the joining of two or more atoms. A stable compound occurs when the total energy of the combination has lower energy than the separated atoms. The bound state implies a net attractive force between the atoms ... a chemical bond. The two extreme cases of chemical bonds are:

Covalent bond: bond in which one or more pairs of electrons are shared by two atoms.

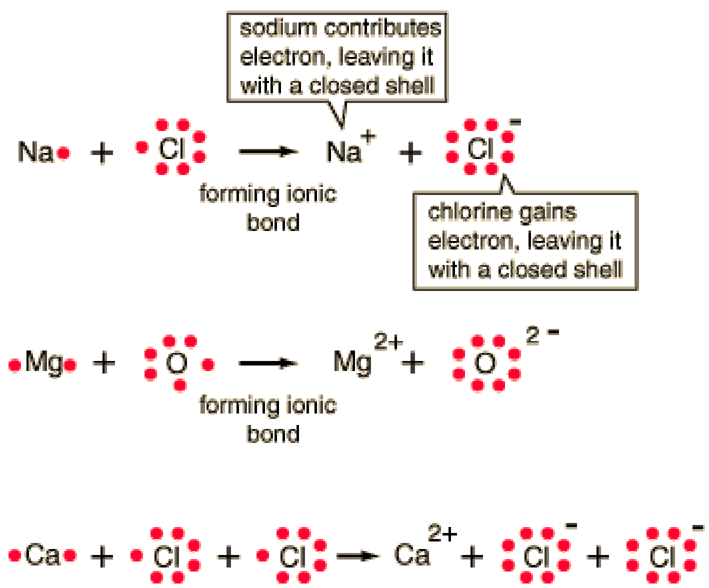
Ionic bond: bond in which one or more electrons from one atom are removed and attached to another atom, resulting in positive and negative ions which attract each other.

Other types of bonds include metallic bonds and hydrogen bonding.

Lewis Diagrams for Compound Formation

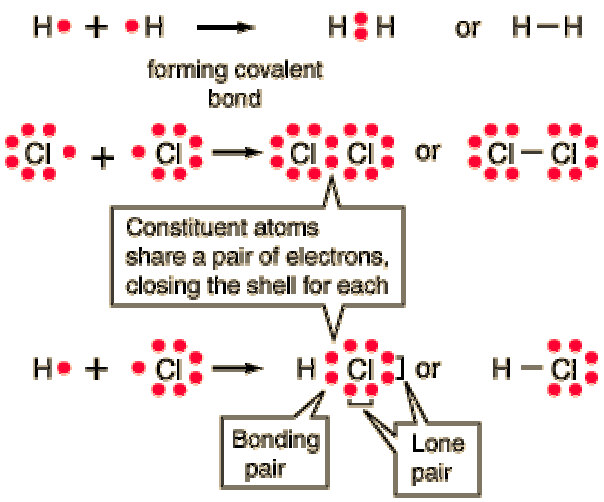
The formation of many common compounds can be visualized with the use of Lewis symbols and Lewis diagrams. In a Lewis symbol, the inner closed shells of electrons can be considered as included in chemical symbol for the element, and the outer shell or valence electrons are represented by dots. The dots are placed in four groups of one or two electrons each, with 8 electrons representing a closed shell or noble gas configuration. Lewis diagrams are useful for visualizing both ionic and covalent bonds.

In the idealized ionic bond, one atom gives up an electron to the other, forming positive and negative ions.



The conditions for bonds are that the total charge is zero and that each atom must have a noble gas electron configuration.

In the idealized covalent bond, two atoms share a pair of electrons, closing the shell for each of them.



The atoms share a pair of electrons, and that pair is referred to as a bonding pair. The pairs of electrons which do not participate in the bond have traditionally been called "lone pairs". A single bond can be represented by the two dots of the bonding pair, or by a single line which represents that pair. The single line representation for a bond is commonly used in drawing Lewis structures for molecules.

Comparison of Properties of Ionic and Covalent Compounds

Because of the nature of ionic and covalent bonds, the materials produced by those bonds tend to have quite different macroscopic properties. The atoms of covalent materials are bound tightly to each other in stable molecules, but those molecules are generally not very strongly attracted to other molecules in the material. On the other hand, the atoms (ions) in ionic materials show strong attractions to other ions in their vicinity. This generally leads to low melting points for covalent solids, and high melting points for ionic solids. For example, the molecule carbon tetrachloride is a non-polar covalent molecule, CCl_4. It's melting point is -23°C. By contrast, the ionic solid NaCl has a melting point of 800°C.

Ionic Compounds

1. Crystalline solids (made of ions)
2. High melting and boiling points
3. Conduct electricity when melted
4. Many soluble in water but not in non-polar liquid

Covalent Compounds

1. Gases, liquids, or solids (made of molecules)
2. Low melting and boiling points
3. Poor electrical conductors in all phases
4. Many soluble in non-polar liquids but not in water

Ionization Energy and Electron Affinity

The ionization energy or ionization potential is the energy necessary to remove an electron from the neutral atom. It is a minimum for the alkali metals which have a single electron outside a closed shell. It generally increases across a row on the periodic maximum for the noble gases which have closed shells. For example, sodium requires only 496 kJ/mol or 5.14 eV/atom to ionize it while neon, the noble gas immediately preceding it in the periodic table, requires 2081 kJ/mol or 21.56 eV/atom. The ionization energy can be thought of as a kind of counter property to electronegativity in the sense that a low ionization energy implies that an element readily <u>gives</u> electrons to a reaction, while a high electronegativity implies that an element strongly seeks to <u>take</u> electrons in a reaction.

The **electron affinity** is a measure of the energy change when an electron is added to a neutral atom to form a negative ion. For example, when a neutral chlorine atom in the gaseous form picks up an electron to form a Cl- ion, it releases an energy of 349 kJ/mol or 3.6 eV/atom. It is said to have an electron affinity of -349 kJ/mol and this large number indicates that it forms a stable negative ion. Small numbers indicate that a less stable negative ion is formed. Groups VIA and VIIA in the periodic table have the largest electron affinities.

Electronegativity

Electronegativity is a measure of the ability of an atom in a molecule to draw bonding electrons to itself. The most commonly used scale of electronegativity is that developed by Linus Pauling in which the value 4.0 is assigned to fluorine, the most electronegative element. Lithium, at the other end of the same period on the periodic table, is assigned a value of 1. Electronegativity generally increases from left to right on the periodic table and decreases from top to bottom. Metals are the least electronegative of the elements. The Pauling electronegativities for the elements are often included as a part of the chart of the elements.

An important application of electronegativity is in the prediction of the polarity of a chemical bond. Because hydrogen has an electro negativity of 2.1 and chlorine has an electronegativity of 3.0, they would be expected to form a polar molecule with the chlorine being the negative side of the dipole. The difference between the electronegativities of Na(0.9) and Cl(3.0) are so great that they form an ionic bond. The hydrogen molecule on the other hand, with zero electro negativity difference, becomes the classic example of a covalent bond.

Metals and Non-metals

As shown on the periodic table of the elements , the majority of the chemical elements in pure form are classified as metals. It seems appropriate to describe what is meant by "metal" in general terms. This general description is adapted from Shipman, et al.

Chemical Properties	
Metals	Non-metals
• Usually have 1-3 electrons in their outer shell. • Lose their valence electrons easily. • Form oxides that are basic. • Are good reducing agents. • Have lower electronegativities.	• Usually have 4-8 electrons in their outer shell. • Gain or share valence electrons easily. • Form oxides that are acidic. • Are good oxidizing agents. • Have higher electronegativities.
Physical Properties	
Metals	Non-metals
• Good electrical conductors and heat conductors. • Malleable - can be beaten into thin sheets. • Ductile - can be stretched into wire. • Possess metallic luster. • Opaque as thin sheet. • Solid at room temperature (except Hg).	• Poor conductors of heat and electricity. • Brittle - if a solid. • Non-ductile. • Do not possess metallic luster. • Transparent as a thin sheet. • Solids, liquids or gases at room temperature.

http://hyperphysics.phy-astr.gsu.edu/hbase/pertab/metal.html