NTI Day 6 Lesson
Content Standard: Bonding
Class: Chemistry Teacher: K. Kelly

Chemical Bonds

A ***chemical bond*** *is the attraction between two or more atoms that results in the formation of compounds or molecules.*

There are three main types of chemical bonds: ionic bonds, covalent bonds, and metallic bonds.

Why Atoms Combine

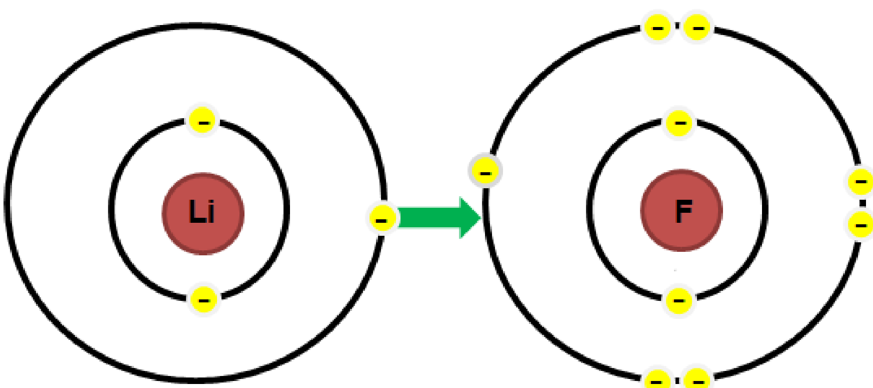
Atoms interact with one another in ways that result in improved chemical stability. To obtain maximum stability, atoms need to have the same electron configuration as a noble gas. That is, they need to have eight outer-shell electrons (or two outer-shell electrons for the elements in periods 1 and 2). The electrons of an atom that are not present in the noble gas of the preceding period (excepting electrons found in full *d* or *f* shells) are **valence electrons**. Only valence electrons participate in chemical bonds. Protons, neutrons, and other electrons are not involved in chemical reactions.

Chemical bonds form so atoms can donate, accept, or share the appropriate number of electrons to achieve a more stable electron configuration.

Ionic Bonds

Ionic bonds tend to form between metals and nonmetals. This type of bonding occurs when a metal donates electrons to form a positively-charged ion (cation) and a nonmetal accepts electrons to form a negatively-charged ion (anion).

For example, the compound lithium fluoride (LiF) is formed when a lithium atom gives one electron to a fluorine atom. By donating one electron to fluorine, lithium is able to achieve the electron configuration of the noble gas helium. By accepting one electron from lithium, fluorine is able to achieve the electron configuration of the noble gas neon. So, the ionic bond formed between lithium and fluorine increases the chemical stability of the atoms.

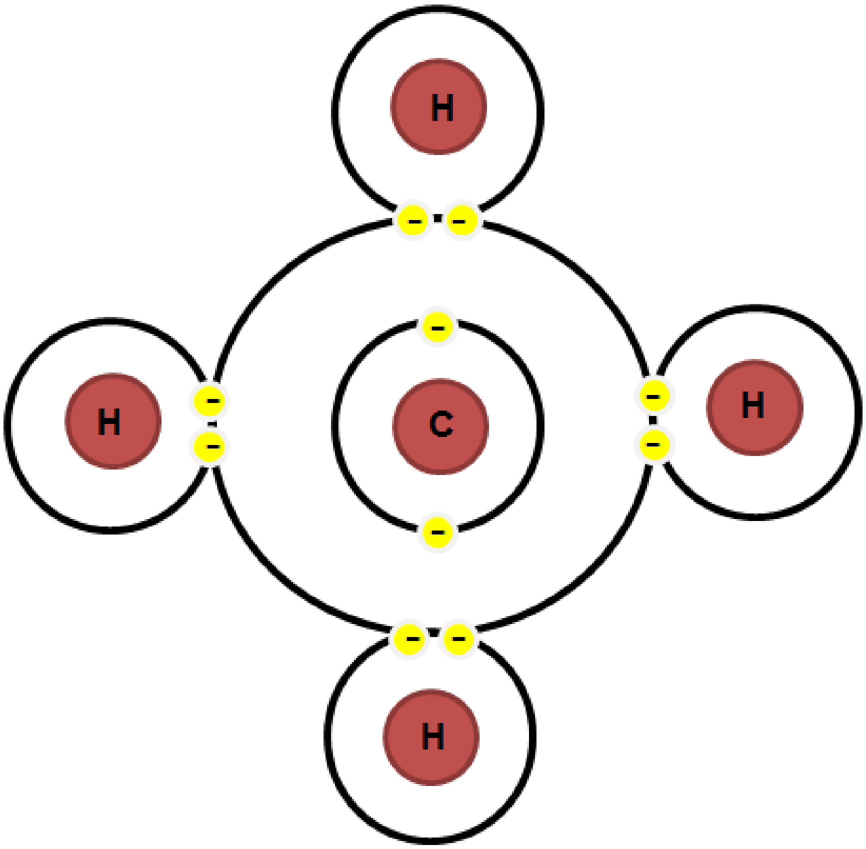


Ionic compounds tend to be brittle and have high melting points and high boiling points. When in the molten state or in aqueous solutions, ionic compounds have high electrical conductivities.

Covalent Bonds

Covalent bonds tend to form between nonmetals. This type of bonding occurs when valence electrons are shared between two or more atoms. For this to happen, the atoms involved should have high electron affinities and high ionization energies so that they both attract electrons, but neither atom is sufficiently electronegative to remove the other atom's electron.

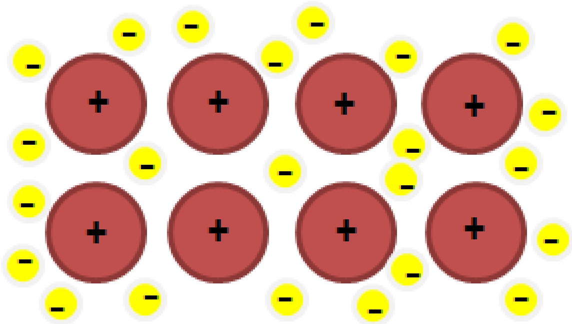
For example, the compound carbon tetrahydride (CH_4), or methane, is formed when four hydrogen atoms share their electrons with a carbon atom. By sharing electrons, each hydrogen atom is able to achieve the electron configuration of the noble gas helium, and the carbon atom is able to achieve the electron configuration of the noble gas neon.



Covalent compounds tend to have low melting points, low boiling points, and low electrical conductivities.

Metallic Bonds

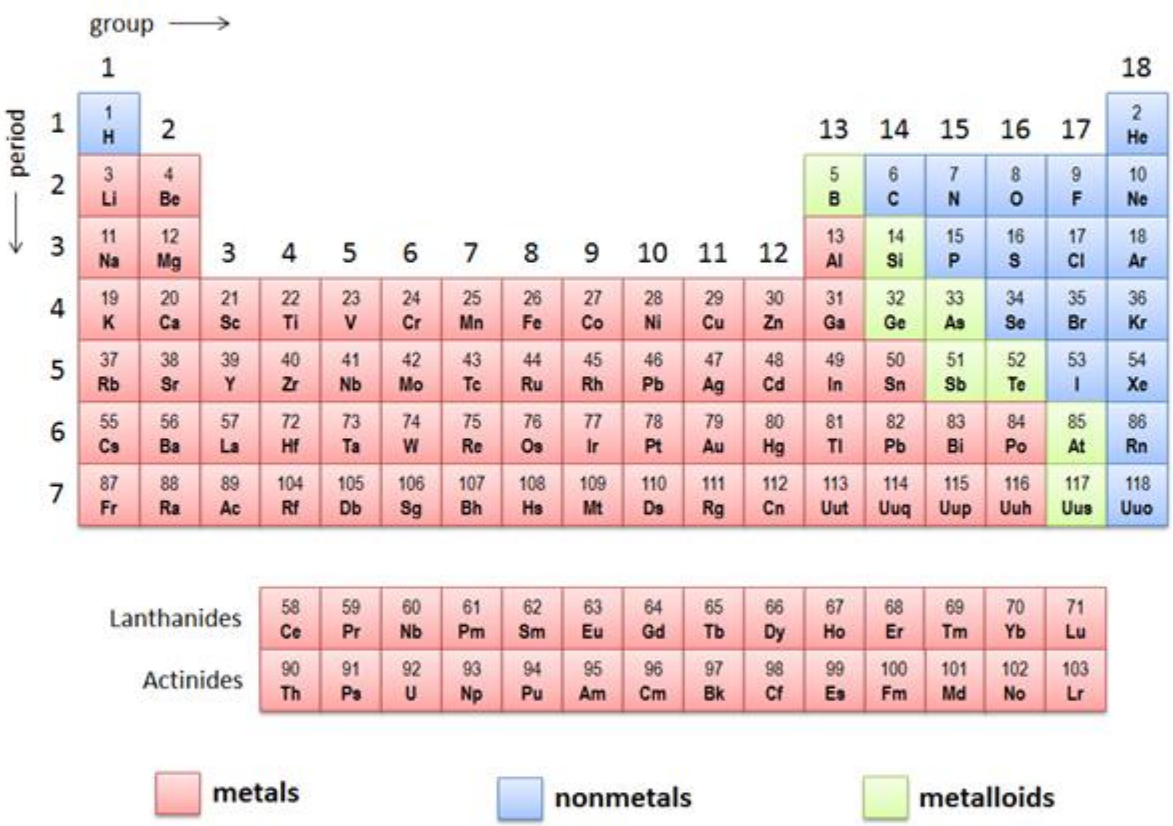
Metallic bonds are forces of attraction that exist between metal atoms. These bonds can be described as metal ions in a sea of mobile electrons.



The strong attraction between the cations and electrons give metals high melting points and high boiling points. The mobility of the electrons allows metals to have high conductivities. It also allows metals to be malleable (able to be hammered into shapes without breaking), ductile (able to be stretched into wires), and to have a shiny luster.

Bonding & the Periodic Table

The periodic table may be used to predict the type of bond that is formed between elements.



For example, since potassium (K) is a metal and chlorine (Cl) is a nonmetal, if these two elements combine, they will form an ionic bond. If carbon (C), a nonmetal, and oxygen (O), a nonmetal, combine, they will form a covalent bond.

Bonding & Electronegativity

Another way to predict the type of bond that is formed is by calculating differences in electronegativity.

Electronegativity refers to an atom's ability to attract electrons to itself. Elements with high electronegativities have stronger tendencies to attract electrons than elements with lower electronegativities.

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	Lu 1.2	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	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cp						

La 1.1	Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.2	Gd 1.2	Tb 1.2	Dy 1.2	Ho 1.2	Er 1.2	Tm 1.2	Yb 1.1
Ac 1.1	Th 1.3	Pa 1.5	U 1.4	Np 1.4	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

This table shows the approximate electronegativities of the elements. These numbers may vary slightly depending on the scale that is used. Some numbers are missing because those values are unknown.

If the electronegativity difference between elements is greater than 1.7, an ionic bond will likely be formed. For example, sodium (Na) has an electronegativity of 0.9, and bromine (Br) has an electronegativity of 2.8. The difference between these numbers is 1.9. Since 1.9 is greater than 1.7, an ionic bond forms between sodium and bromine.

If the electronegativity difference between elements is less than 1.7, a covalent bond will likely be formed. For example, phosphorus (P) has an electronegativity of 2.1, and chlorine (Cl) has an electronegativity of 3.0. The difference between these numbers is 0.9. Since 0.9 is less than 1.7, a covalent bond forms between phosphorus and chlorine.

If the electronegativity difference between elements is greater than 0.5 but less than 1.7, a **polar** covalent bond will likely be formed. In a polar bond, the electrons are not shared equally between the atoms. One end—or *pole*—of a polar molecule may end up slightly negative while the other becomes slightly positive. For example, water molecules are polar. The (positive) hydrogens of one molecule are attracted by the (negative) oxygen of another molecule, creating an attractive intermolecular force that gives water its high surface tension.

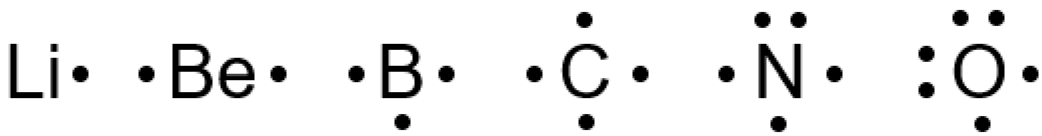
The strength of a chemical bond can also be related to electronegativity differences. In general, the greater the difference is in electronegativities, the stronger the bond. So, a bond between hydrogen and fluorine, for example, is stronger than a bond between hydrogen and iodine, since the electronegativity difference between hydrogen and fluorine is 1.9, and the electronegativity difference between hydrogen and iodine is only 0.4.

Electron-Dot Diagrams

The arrangement of electron pairs in a molecule can be represented by an electron-dot diagram.

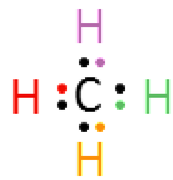
Electron-dot diagrams represent the valence electron arrangement in elements, molecules, and ions.

For an element, the electron-dot diagram simply shows how many valence electrons the atom has. The diagram below shows electron-dot diagrams for six elements in the second period of the periodic table.



Electron-dot Diagrams & Molecules

Electron-dot diagrams are often used to show the arrangement of atoms within a covalent molecule. The picture below shows the electron-dot diagram for methane, CH_4.



In a correctly drawn electron-dot diagram, each atom will have a full outer shell of electrons. To count an atom's number of valence electrons, count all of the dots that surround the atom. Methane is a covalent molecule, so two electrons are shared between its carbon atom and each hydrogen atom. Therefore, the two electrons that make up a single bond between carbon and each hydrogen are counted as part of *both atoms'* outer shells. Carbon needs eight electrons to fill its outer shell, and there are eight electrons surrounding the carbon atom in the structure of methane. In this case, carbon is said to have a **full octet**. Each hydrogen in the structure has two dots around it, and hydrogen only needs two electrons to have a full outer shell. Remember, *all bonded electrons should be counted as part of an atom's outer shell*.