Chemical Bonding

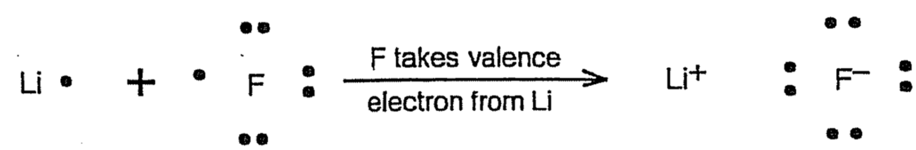
Ionic Bonding

– A bond formed usually between a metal and a non-metal. An electron from one atom (metal) is transferred to another atom (non-metal) forming a positive ion for the atom that lost the electron and a negative ion for the atom accepting the electron. The act of giving and receiving of the electron is what holds the different species together, which we call a "bond".

- Electronegativities can be used to find out if a bonding between atoms will be <u>ionic</u>, <u>polar covalent</u> or <u>covalent</u>. If the difference (subtraction) between the electronegativities is between $4.0 - 1.7$ then the bond is ionic. Since one atom is much more electronegative it pulls the electron(s) closer to its nucleus, resulting in one atom taking and the other losing. If the difference is between $1.7 - 0.2$ then the bond is polar covalent and if less than 0.2 then the bond is covalent. Polar covalent is an unequal sharing of electron(s). One atom is more electronegative so it hogs the electron(s). However, it is not "strong" enough to actually steal so an uneven sharing occurs. Covalent is when the atoms can both pull equally on the electrons resulting in an equal sharing.

- Bonding is usually depicted in chemistry using what are called "Lewis structures" or "electron dot diagrams". These are pictures used to show how the valence electrons are distributed around an atom, ion or molecule.

<u>Ex. 1</u> -



<u>Ex. 2</u> –

- Ionic bonds are occurring between positive and negative ions. If we think logically we can then solve for how ions are different then their corresponding neutral atoms.

- When non-metals become ions, they take electron(s) and become negative. If the same positive charge of the nucleus now has to attract extra electron(s) will it do this weaker or stronger? _________________________. If its attracting all of the electrons slightly weaker, will the overall atom be larger or smaller? ____________________. So, anions (negative charged ions) are **<u>larger</u>**.

- When metals become ions, they lose electron(s) and become positive. If the same positive charge of the nucleus now has to attract less electron(s) will it do this weaker or stronger? _____________________.

If its attracting all of the electrons slightly weaker, will the overall atom be larger or smaller?

____________________. So, cations (positively charged ions) are **<u>smaller</u>**.

- The strength of ionic bands is estimated by examining the melting temperatures of ionic compounds.

  $LiF = 845°C$ $NaF = 993°C$ $KCl = 770°C$ $LiCl = 605°C$

- Ionic bonds are relatively weak, so compounds held together by ionic bonds have a relatively low melting temperature

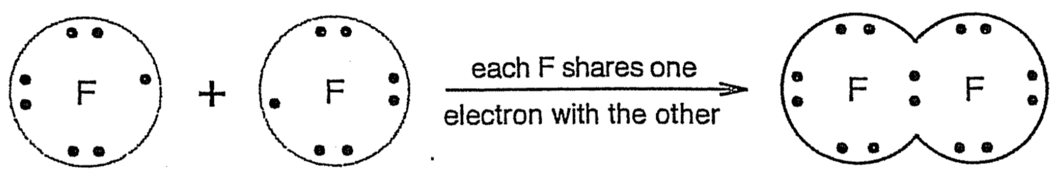
- Common charge trends on the periodic table follow the families or columns.

  1 = +1 2 = +2 13 = +3 15 = -3 16 = -2 17 = -1 18 = 0

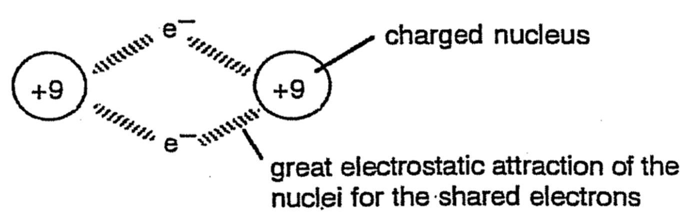
Covalent Bonding

- A bond formed between two non-metal atoms by an equal sharing of their valence electrons. The bond is formed between atoms having less than full shells of electrons by SHARING valence electrons so as to fill both atoms valence shell. The elements in families 14 to 17 tend to form covalent bonds to have eight electrons each in their valence shells.

<u>Ex. 1</u> -



The shared electrons in a F–F covalent bond can be visualized as follows.



- Again, the strength of the covalent bonds is estimated based on the melting temperatures.

  $BN = 3000°C$ $SiC = 2700°C$ C (diamond) = $3550°C$

- Covalent bonds have high melting temperatures, so are also very strong.

- BUT!!!! $CH_4 = -182°C$ $O^{2-} = -218°C$ $F_2 = -220°C$

- So, covalent compounds have a <u>wide range</u> of melting temperatures.

London's Forces and Dipoles

- So, what's going on? Which are strong and which are weak?

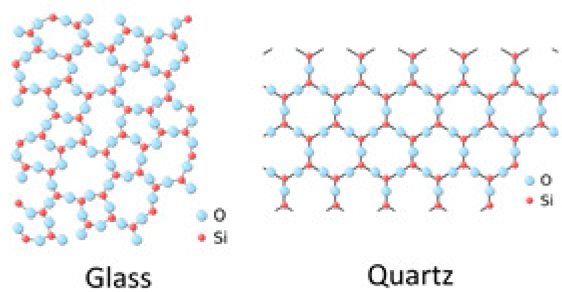
- When dealing with the bond holding different elements together to make a molecule we call these forces ________________________. The forces bonding individual molecules to other molecules we call those forces ________________________.

- The general rule for bonding is that;

	Ionic	**Covalent**
Intramolecular		
Intermolecular		

- **This differentiation becomes important in understanding bond strength**.

- An example of strong intramolecular bonding occurs in SiO_2 (quartz). The bonding is so strong that pure quartz has a hardness of 7 Mohs. Indigenous peoples often made arrowheads and knife blades using "chert" which is SiO_2.

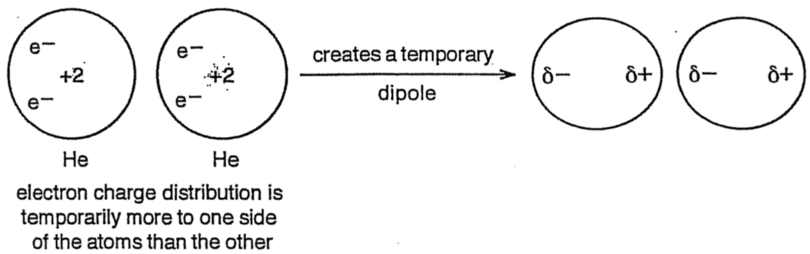


- Silicon dioxide also is an excellent example of weaker **intermolecular bonding** forces. The weaker bonds allow the quartz to chip conchoidally or "flake". This allowed the very fine, sharp edge to form that is needed for knives and cutting implements.

- When atoms come close to each other they often experience attractive and repulsive forces from the neighbouring atom. The electrons of each atom will repel the electrons of the neighbouring atom and the neighbouring nucleus will attract the neighbouring electrons. The repelling of the neighbouring atom's electrons is called "polarization". This polarization results in a very short-lived dipole. A dipole (δ) is a ***partial separation of charge*** which exists in an atom. In the diagram below two helium atoms have come close to each other and set up a slight excess of negative charge ($\delta-$) on one side (left side) and a slight excess of positive charge ($\delta+$) (right side) on the other side.

- These weak dipoles that result cause a weak force to pull different molecules together. This weak force caused by the dipoles is called London Forces.

Ex. 1 -



- **The rule for London Forces is that the more electrons you have the greater the London Forces.**

When predicting properties or traits of compounds, use the following order for deciding relative attractive force strengths:

1.) Bond Type (covalent stronger)

2.) Compound size (smaller stronger)

3.) Charge (larger charge stronger)