Honors Pre-AP Chemistry Ch. 8 and 9 Bonding/Ionic and Covalent (Section 1)

Bonds—forces that hold groups of atoms together and make them function as a unit.

Three major types of bonds—ionic, covalent, and metallic. We will be focusing on Ionic and Covalent bonding.

Strength of a bond is dependent on its **bond energy**.
Bond energy—The energy needed to break a bond.

Ionic Bonding—The electrostatic attraction of oppositely charged ions. Ions form when electrons are transferred from on atom to another. Highly electronegative atoms attract and "pull off" electrons from atoms with less electronegativity. (Electronegativity difference between the atoms is usually ≥ 1.7). These bonds are usually very strong. Ionic bonds are between a metal and a non-metal or a metal and a polyatomic ion. Ionic compounds are all solid crystals at room temperature and the crystal structure is call a lattice. The energy holding the crystal ions together is call the lattice energy.
Lattice Energy—the change in energy that takes place when separated gaseous ions are packed together to form an ionic solid. (We will not be calculating lattice energy this semester. It will be taught next year during AP Chemistry.)

Whenever two oppositely charged particles come together, it lowers the potential energy of the particles. When the particles are at a distance where the potential energy is the lowest, they "bond". (This is also true for covalent bonding.) The particles move closer and farther apart until the potential energy is at its lowest.

Covalent Bonding—bonding that occurs when two atoms come close enough to "share" electrons by overlapping orbitals and lowering their potential energy. Neither atom "pulls" the electrons off the other atom, but atoms with higher electronegativity will pull the electrons in the overlapping orbitals closer to their nuclei (uneven sharing) than the other atoms' nuclei. They will set up negative (-) and positive (+) forces within the bond called **polar covalent bonds** (p. 355-257).
Whereas, if the atoms bond with equal or "close to" equal electronegativity, they will form **non-polar covalent bonds.**

Electronegativity—the ability of an atom in a molecule to attract shared electrons to itself.
(p. 357, table of electronegativities)
Electronegativities tend to increase across a period and decrease down a group.

Covalent vs Ionic Character

The greater the electronegativity differences between the two bonding atoms the greater the **polar → ionic** "character".
Metals have low electronegativity and non-metals have high electronegativity. Their electronegativity difference is usually a high number forming ionic compounds (diff ≥ 1.7).
The 1.7 is a "rule of thumb". There can be ionic substances with an electronegativity difference that is < 1.7 and polar covalent non-metal substances with and electronegativity difference that is > 1.7.
For example: Li_2S is an ionic cpd with an electronegativity difference of 1.5 (2.5-1.0).
HF is a polar covalent cpd with an electronegativity difference of 1.9 (4.0-2.1).

When electronegativity difference is ≤ 0.1 or 0.2, the molecule has a high non-polar or covalent character. (See diagram on bottom of p. 357)

Dipole Moment—Whenever a polar bond is set up, one side of the molecule is (+) (less electronegative), and the other side is (-) (more electronegative). It is represented by an arrow with at + on one end and an arrow point (-) on the other end **+-->.** This is the electrical field that is set up around the molecule (see p. 359).

H_2O	HF	CO_2
O< H H	H—F	O=C=O
<--+	**+-->**	**<--+ +-->**

***Sometimes, the dipole moments cancel (as you can see above with CO_2) resulting in a non-polar covalent molecule. The **molecule is non-polar even though the bonds are polar** with dipole moments. (This will be important to know later.)
If a molecule has "**symmetrical**" bonding, the dipole moments cancel and the molecule is non-polar. (See p.360.)

Ionic and covalent bond energies are important to know and calculate but, as I said, we will not be doing that this semester. Typical bond energies are listed on p. 374 for covalent compounds.

However, it is important that you know the different sizes and strengths of covalent bonds.
Single bond: longest and weakest
Double bond: shorter and stronger than single bonds
Triple bonds: shortest and stongest

Writing Lewis Structures

Atoms bond for the reason of **stability. Ionic bonding** occurs when ions of opposite charges attract each other and hold together in a bond and form a lattice structure giving the compound its lowest potential energy. **Covalent bonding** occurs in order that the atoms in the compound form a Noble Gas configuration "octet" in their valence shell.
There are exceptions to the "octet" rule (8 electrons in the valence shell) but they still form compounds that lower the potential energy of the atoms and become stable.

Ionic Lewis Structure

The ionic Lewis structure is the structure of a "formula unit" of an ionic compound. This is the smallest whole number ratio of elements in the compound that still have the properties of that compound. It is always in Empirical form.

Ionic Lewis Structures are very simple to represent. We represent the metal and non-metal ions with their charge and their octet of electrons.

Example: Lithium oxide. First we write the formula unit (Li_2O).

Then we write the ions with charges and valence electrons: $Li^+ + Li^+ + [:\overset{..}{\underset{..}{O}}:]^{-2}$

Writing this structure is difficult on the computer but you can see how it is done.
The positive metal ions have no valence electrons because they were lost with the positive charge for the electrons lost. The negative non- metal ions have 8 dots symmetrically around them inside a bracket with the charge on the outside representing the electrons gained to fulfill the octet.
All ionic Lewis structures are done this way. It is very simple and easy. Positive metals have no valence electrons and include their charges, negative non-metals have 8 electrons around them with brackets and their charges.
The only time this will be different is if there is a metal hydride (hydrogen with a metal).
Hydrogen with have a -1 charge and only two electrons around it (duet).
Example: Sodium hydride, NaH: $Na^+ + [.H.]^-$

Next section will include covalent Lewis structures with shape.
Shape is important so we will **not** be doing 2-dimentional symmetric Lewis Structures.
Most of the notes will be hand-written and posted as photos since it is too difficult to show all electrons using "word".