Chemistry Honors – Lesson 2
Bonding

Bonding

The atoms of many elements can combine to form **compounds**. Individual, isolated unites of compounds are considered **molecules**. The atoms in most molecules are held together by strong attractive forces called **chemical bonds.** These bonds are formed through the interaction of valence electrons of the combining atoms. In addition to the very strong forces within a molecule, there are weaker forces acting between molecules. These **intermolecular forces**, although weaker than the intramolecular chemical bonds, are of considerable importance in understanding the properties of many substances.

Many molecules contain atoms bonded according to the **octet rule**, which states that an atom tends to bond with other atoms until it has eight electrons in its outermost shell, thereby forming stable configurations similar to that of Group VIII (noble gas) elements.

When classifying chemical bonds, it is helpful to introduce two types: ionic bonds and covalent bonds. **Ionic bonds** transfer an electron(s) and are held together by electrostatic force. **Covalent bonds** share electron pairs between atoms.

Ionic Bonds

When two atoms with large differences in **electronegativity** react, there is a transfer of electrons from the less electronegative atom to the more electronegative atom. The atom that loses electrons becomes a positively charged ion, or **cation**, and the atom that gains electrons becomes a negatively charged ion, or **anion**. In general, the elements of Groups I and II (low electronegativity) give up their electrons while Group VII (high electronegativity) gains the electrons to get noble gas configuration. The positive and negatives charges become a force of attraction between the anion and cation called **ionic bond**.

Covalent Bonds

Covalent bonds can be grouped into network covalent bonding and molecular covalent bonding. In network covalent, atoms were bonded in a intricate network, connecting each atom with all of the adjacent atoms. In molecular bonding, atoms are connected into small units called molecules. However, despite the difference in arrangement, the bonds work in the same way. When atoms with similar electronegativity interact, they share electrons to achieve noble gas electron configuration. The binding force between the atoms result from the attraction of each electron has for the positive nuclei.

A. Properties of Covalent Bonds

Atoms sharing electrons can share more than one pair. Each pair is considered a single bond, so if two atoms share one, two or three electron pairs, then they are said to be joined by a **single**, **double** or **triple** covalent bond. A covalent bond can be characterized by two features: **bond length** and **bond energy**.

a. **Bond length** – the average distance between two nuclei of the atoms in a bond. As the number of electron pairs increase, the two atoms are pulled closer together, leading to a decrease in bond length. Thus, for a given pair of atoms, a triple is shorter than a double bond, which is shorter than a single bond.
b. **Bond energy** – is the energy needed to separate the two bonded atoms, ie the energy needed to break a bond. For a given pair of atoms, the strength of a pond increases as more electron pairs are shared, requiring more energy to break them.

B. Covalent Bond Notation

a. **Lewis Structures** – or Lewis Dot Structure, is a model that shows an element surrounded by dots, each representing one of the valence electrons.

C	**Carbon**	**N**	**Nitrogen**
O	**Oxygen**	**F**	**Fluorine**
Li	**Lithium**	**Be**	**Beryllium**

C. Types of Covalent Bonds

a. **Polar Covalent Bond** – Polar covalent bonds occur with atoms of different electronegativity resulting in electron pairs not shared equally. The electrons are pulled more towards the atom/element with higher electronegativity. As a result, the more electronegative atom acquires a partial negative charge and the less electronegative one has a partial positive charge. A molecule that has such a separation of charges is called a polar molecule. A **dipole moment** is the direction of a positive and negative partial charge.
b. **Nonpolar Covalent Bond** – a nonpolar bond occurs with atoms that have the same electronegativities. The electron pair is shared equally and is commonly found in **diatomic molecules**.

D. Geometry and Polarity

a. **Valence Shell Electron-Pair Repulsion Theory (VSEPR)** – states that the three-dimensional arrangement of atoms surrounding a central atom is determined by the repulsions between the bonding and the nonbonding electron pairs in the valence shell of the central atom. The electron pairs arrange themselves as far apart as possible, minimizing repulsion.
b. **Polarity** – a molecule with a net dipole moment is called polar because it has a positive and negative pole. The polarity of a molecule dipole of each bond and the shape of the of the molecule.

Intermolecular Forces

The attractive forces that exist between separate and isolated molecules are known as **intermolecular forces**.

A. Dipole-Dipole Interaction

Polar molecules tend to orient themselves such that the positive region of one molecule is close to the negative region of another molecule. Dipole-dipole interactions are present in the solid and liquid phases but become negligible in the gas phase because the molecules are much father apart. Polar species tend to have higher boiling points than nonpolar species.

B. Hydrogen Bonding

Hydrogen bonding is a specific, unusually strong form of dipole-dipole interaction, which may be either intra- or intermolecular. When hydrogen is bound to a highly electronegative atom, the hydrogen atoms carries little of the electrons. The hydrogen becomes positively charged and interacts with partially negative regions. Substances with hydrogen bonding usually have high boiling points.

C. Dispersion Forces

The bonding electrons in covalent bonds may appear to be equally shared between two atoms, but at any one moment in time they will be located randomly throughout the orbital. This results in temporary short lived dipoles. These dipoles interact with the electron clouds of other atoms, inducing more dipole formation. These interactions of short lived dipoles are called **London Forces**.