Chapter 8 - Bonding: General Concepts

Section 8.1 - Types of Chemical Bonds

1. **Bond Energy** - the energy required to break the bond.
2. **Ionic Bond** - an electrostatic attraction between oppositely charged particles. A metal reacts with a nonmetal.
3. **Bond Length** - distance between atoms where the energy is minimum.
4. **Covalent Bonding** - type of bond in which a pairs of electrons are shared by nuclei.
5. **Polar Covalent Bond** - unequal sharing of a pair of electrons.
6. **Coulomb's Law** - calculates the energy of interaction between a pair of ions.

$$E = 2.31 \times 10^{-19} \; \text{J} \cdot \text{nm} \left(\frac{Q_1 Q_2}{r} \right)$$

Where **r** is the distance between the ion centers in nanometers (nm), and **Q1** and **Q2** are the numerical charges. A negative value indicates an attractive force. That is , the ion pair has lower energy than the separated ions. The repulsive energy can also be calculated by brining together two like ions. The value will be positive in this case.

Section 8.2 - Electronegativity

Electronegativity - the ability of an atom in a molecule to attract shared electrons to itself.

> **Periodic Trend** - Electronegativity increases as you go from left to right within a period on the periodic table. As the number of protons increases within the nucleus across a period, the elements ability to attract electrons increases.
>
> **Periodic Trend** - Electronegativity decreases as you go form top to bottom within a group (or family) on the periodic table. Do to increased **shielding** and **electron repulsion**, as you move down through a group (or family), the nucleus will have less of an ability to attract electrons.

Bond Type - The three major type of intramolecular bonds can be described by the electronegativity difference.

> **Covalent (nonpolar covalent)** - bonds which occur between atoms with little or no electronegativity difference.
> **Polar Covalent** - bonds which occur between atoms with a definite electronegativity difference.
> **Ionic** - bonds which occur between atoms with a large electronegativity difference where electron transfer can occur.

Section 8.3 - Bond Polarity and Dipole Moments

Dipole Moment (Dipolar) - a diagram to illustrate the unequal sharing of electrons. The dipole character of a molecule is often represented by an arrow pointing to the negative charge center with the tail of the arrow indicating the positive center of charge:

δ^+ δ^-
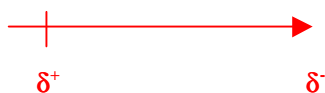
> Any diatomic molecule that has a polar bond also will show a molecular dipole moment. **Polyatomic** molecules that have polar bonds may or may not show a molecular dipole moment, that is why we must consider the molecular geometry of the molecule. The lower case delta (δ) is used to indicate a fractional charge.

Section 8.4 - Ions: Electron Configurations

When talking about **ionic compounds**, we are usually referring to them in their solid state. In this solid state, ions are close together and are arranged to minimize repulsion and maximize attraction between the ions. In a gaseous state, the ions are no longer very close to each other, thus interaction between ions is at a minimum.

Ionic size - the size of an ion plays an important role in determining the structure and stability of ionic solids.

> **Periodic Trend** - Metals tend to lose electrons. The size of the atom becomes smaller with the loss of each consecutive electron. The nuclear charge can have a greater pull on the electrons with the loss of each electron.

> **Periodic Trend** - Nonmetals tend to gain electrons. The size of the atom becomes larger with the addition of each consecutive electron. The nuclear charge has less of a pull on the electrons with the gain of each electron due to greater electron repulsion. **See page 356 - Figure 8.7**

Isoelectronic ions - ions containing the same number of electrons, thus having the same electron configuration.

# Section 8.5 - Formation of Binary ionic Compounds

Lattice Energy - the change in energy that takes place when separated gaseous ions are packed together to form an ionic solid:

$$M^+(g) + X^-(g) \rightarrow MX(s)$$

> Lattice energy can also be defined as the energy released when an ionic solid forms from its ions.

To see the energy terms associated with this process, we take advantage of the fact that energy is a **state function** and break this reaction into steps, the sum of which gives the overall reaction.

Step 1: **Sublimation** - energy involved in taking a substance from the solid state to the gaseous state:
$$M(s) \rightarrow M(g)$$
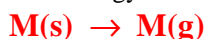
Step 2: **Ionization energy** - energy involved in removing electrons from the gaseous metal state:
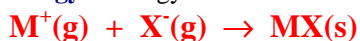$$M(g) \rightarrow M^+(g) + e^-$$

Step 3: **Dissociation of molecules** - energy involved in breaking the bonds that hold together diatomic molecules:
$$\tfrac{1}{2}X_2(g) \rightarrow X(g)$$

Step 4: **Electron Affinity** - energy involved in adding an electron to the gaseous nonmetal state:
$$X(g) + e^- \rightarrow X^-(g)$$

Step 5: **Lattice Energy** - energy involved in the formation of the solid ionic compound from the gaseous states:
$$M^+(g) + X^-(g) \rightarrow MX(s)$$

> The sum of these five processes yields the desired energy for the overall reaction. **See page 359 - figure 8.8**

Section 8.6 - Partial Ionic Character
any compound that conducts an electrical current when melted will be classified as ionic.

Section 8.7 - The Covalent Chemical Bond
Chemical Bond - forces that cause a group of atoms to behave as a unit.
Why do chemical bonds occur? - bonds result from the tendency of a system to seek its lowest possible energy. (becomes more stable)
Bonding - is a model proposed to explain molecular stability.
read purple box on page 366

Section 8.8 - Covalent Bond Energies
Single Bond - bonds in which one pair of electrons is shared.
Double Bond - bonds in which two pairs of electrons are shared.
Triple Bond - bonds in which three pairs of electrons are shared.

ΔH (Enthalpy) = sum of the energies required to break old bonds (positive signs) plus the sum of the energies released in the formation of new bonds (negative signs).

ΔH = ΣD (bonds broken) – ΣD (bonds formed)
                energy required   energy released

Section 8.9 - The Localized Electron Bonding Model

Localized electron (LE) model - assumes that a molecule is composed of atoms that are bonded together by sharing pairs of electrons using the atomic orbitals of the bound atoms. It is divided into 3 parts; Lewis Structures, VESPR models, and a description of the types of atomic orbitals used to share electrons or hold lone pairs (chapter 9).
Lone pairs - valence electrons which are not bonded to other atoms. (or unshared pairs of electrons)
Bonding pairs - valence electrons which are involved in the bonding of atoms.

Section 8.10 - Lewis Structures

Lewis structure of a model shows how the valence electrons are arranged among the atoms in the molecule.
        * The most important requirement for the formation of a stable compound is that the atoms achieve noble gas electron configurations.
        * Only the valence electrons are included in Lewis structures.
Duet rule - Hydrogen will form stable molecules where it shares two electrons.
Octet rule - eight valence electrons are required to form a noble gas configuration, so elements should be surrounded by eight electrons.
        Each set of dots represents an orbital in which the valence electrons are located.

Section 8.11 - Exceptions to the Octet Rule
See purple table on page 376

Section 8.12 - Resonance
Resonance - occurs when more than one valid Lewis structure can be written for a particular molecule.
Formal Charge - the difference between the number of valence electrons on the free atom and the number of valence electrons assigned to the atom in the molecule.

(Valence electrons)$_{assigned}$ = (number of lone pair electrons) + ½(number of shared electrons)

**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.