Chapter 8 Chemistry Notes; Chemical Equations and Reactions

Chemical Equations

- The quantitative nature of chemical formulas and reactions is called **stoichiometry.**
- Lavoisier observed that mass is conserved in a chemical reaction.
  - This observation is know as the **law of conservation of mass.**
- **Chemical equations** give a description of a chemical reaction.
- There are two parts to any equation:
  - **Reactants** (written to the left of the arrow) and
  - **Products** (written to the right of the arrow): $2H_2 + O_2 \rightarrow 2H_2O$
- There are two sets of numbers in a chemical equation:
- Numbers in front of the chemical formulas (called stoichiometric *coefficients*) and
- Numbers in the formulas (they appear as subscripts).
- Stoichiometric coefficients give the *ratio* in which the reactants and products exist.
- The subscripts give the ratio in which the atoms are found in the molecule.
  - Example:
    - H_2O means there are <u>two H atoms</u> for each one molecule of water.
    - $2H_2O$ means that there are <u>two water molecules</u> present.
- Note in $2H_2O$ there are *four* hydrogen atoms present (two for each water molecule).
- Matter cannot be lost in any chemical reaction.
  - Therefore, the products of a chemical reaction have to account for all the atoms present in the reactants—we must *balance* the chemical equation.
  - When balancing the chemical equation we adjust the stoichiometric coefficients in front of the chemical formulas.
    - Subscripts in a formula are *never* changed when balancing an equation.
- Example: The reaction of methane with oxygen: $CH_4 + O_2 \rightarrow CO_2 + H_2O$
  - Counting *atoms* in the reactants:
    - 1C;
    - 4H; and
    - 2O.
  - In the products:
    - 1C;
    - 2H; and
    - 3O.
  - It appears as though H has been lost and O has been created.
  - To balance the equation, we adjust the stoichiometric coefficients:

$$CH_4 + \mathbf{2}O_2 \rightarrow CO_2 + \mathbf{2}H_2O$$

- The physical state of each reactant and product may be added to the equation:

$$CH_4(g) + \mathbf{2}O_2(g) \rightarrow CO_2(g) + \mathbf{2}H_2O(g)$$

- Reaction conditions occasionally appear above or below the reaction arrow(e.g., "$\triangle$" often is used to indicate the addition of heat).

Some Simple Patterns of Chemical Reactivity

Combination and Decomposition Reactions

- In **combination reactions (synthesis reactions)** two or more substances react to form one product.
- Combination reactions have more reactants than products.
  - Consider the reaction: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

- Since there are fewer products than reactants, the Mg has combined with O_2 to form MgO.
- Note that the structure of the reactants has changed:
- Mg consists of closely packed atoms and O_2 consists of dispersed molecules.
- MgO consists of a lattice of Mg2+ and O2- ions.

- In **decomposition reactions** one substance undergoes a reaction to produce two or more other substances.
- Decomposition reactions have more products than reactants.
  - Consider the reaction that occurs in an automobile air bag: $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$
  - Since there are more products than reactants, the sodium azide has decomposed into sodium metal and nitrogen gas.

Combustion in Air

- **Combustion reactions** are rapid reactions that produce a flame.
  - Most combustion reactions involve reaction of $O_2(g)$ from air.
  - Example: Combustion of a hydrocarbon (propane) to produce carbon dioxide and water.

$$C_3H_3(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

Oxidation of Metals by Acids and Salts

- The reaction of a metal with either an acid or a metal is called a **displacement reaction.**
- General pattern: $A + BX \rightarrow AX + B$ (Single Replacement)
- Example: It is common for metals to produce hydrogen gas when they react with acids. Consider the reaction between Mg and HCl: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
  - In the process the metal is oxidized and the H+ is reduced.
- Example: It is possible for metals to be oxidized in the presence of a salt:

$$Fe(s) + Ni(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Ni(s)$$

  - The net ionic equation shows the redox chemistry well:Fe(s) + Ni2+(aq) → Fe2-(aq) + Ni(s)
  - In this reaction iron has been oxidized to Fe2+ while the Ni2+ has been reduced to Ni.
  - Always keep in mind that whenever one substance is oxidized, some other substance *must* be reduced.

Exchange (Metathesis) Reactions – Double Replacement Reactions

- **Double replacement reactions** involve swapping ions in solution: $AX + BY \rightarrow AY + BX$
- Many precipitation and acid-base reactions exhibit this pattern.
- Double-replacement reactions can be called precipitation reactions.
- Reactions that result in the formation of an insoluble product are know as precipitation reactions.
- A precipitate is an insoluble solid formed by a reaction in solution.
  - Example: $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

The Activity Series

- We can list metals in order of decreasing ease of oxidation.
  - This list is an **activity series.**
- The metals at the top of the activity series are called *active metals.*
- The metals at the bottom of the activity series are called *noble metals.*
- A metal in the activity series can only be oxidized by a metal ion below it.
- If we place Cu into a solution of Ag+ ions, the Cu2+ ions can be formed because Cu is above Ag in the activity series: $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$