Chapter 8 – Chemical Reactions

1) Sec 8.1 – Predicting Whether a Reaction Will Occur
Chemical reactions occur often with one or more of the following indications of something happening:
   a. color change
   b. a solid forms
   c. a gas (bubbles) forms
   d. heat and/or flame is produced (exothermic) or heat is absorbed (endothermic)

   Chemists generally use the term **precipitate** to describe the solid that is formed in a chemical reaction and **precipitation** to describe the process of the solid being formed.

2) Types of Chemical Reactions – There are many ways to catagorize chemical reactions. One classification scheme identifies five general reaction types: Combination, Decomposition, Single Replacement, Double Replacement, and Combustion. Occasionally a reaction may fit equally well into two of these categories. Recognizing the type of reaction will make it easier to **predict** the products of reactions.

   This scheme is a different approach than the one that the authors of the book use. This is a personal preference on my part. As we go through the chapter we will be talking about how these different approaches relate. I think that these "reaction types" are easier ways to look at possible reactions and understand/predict what might occur.

   The types of reactions we are going to be using are: Combination Reactions, Decomposition Reactions, Single Replacement Reactions, Double Replacement Reactions, and Combustion Reactions.

   Your book talks about some of these reaction types in Section 8.7 – Other Ways to Classify Reactions.

3) Combination Reactions
In a combination reaction, two or more substances react to form a single substance. The product of the reaction must be a compound. The reactants may be elements or compounds. In this class we will only deal with combination reactions involving metals with oxygen or one of the halogens (fluorine, chlorine, bromine, or iodine). In addition, we will discuss reactions of hydrogen, carbon, and sulfur with oxygen.

   a) Almost all metals will react with oxygen and the halogens. (The only exceptions are that Pt, Ag and Au DO NOT react with oxygen – in short, they do not rust.) For those metals that form more than one ion, the product will usually contain the higher charged ion.

      Examples – Complete and balance the following reactions:
      1) sodium + oxygen

      2) magnesium + bromine

      3) iron + oxygen

      4) platinum + iodine

b) The products of the reaction between hydrogen, carbon, and sulfur with oxygen cannot be determined using ionic charges because the reaction does not produce ions, but the products are easy to remember.

Examples – Complete and balance the following equations:
1) hydrogen + oxygen ---->

2) carbon + oxygen ---->

3) sulfur + oxygen ---->

4) Decomposition Reactions

In a decomposition reaction a single compound is broken down into two or more simpler products. These products can be any combination of elements or compounds. Most decomposition reactions require energy in the form of heat, light, or electricity.

Examples – Write balanced equations for these decomposition reactions:
1) The decomposition of water by the use of electrical energy into its elements.

2) The decomposition of calcium carbonate by the use of heat into calcium oxide and carbon dioxide. (Most carbonates undergo this type of reaction when heated.)

There are three decomposition reactions that you need to memorize:

a) H_2CO_3 ----> H_2O + CO_2
b) NH_4OH ----> H_2O + NH_3
c) H_2SO_3 ----> H_2O + SO_2

These will be especially important in dealing with Double Replacement Reactions because the reason that the double replacement reaction occurs is that one of the possible products then decomposes. If that did not happen, then we would say that there is no reaction.

5) Single Replacement Reactions

In a single replacement reaction, atoms of an element replace the atoms of a second element in a compound. That is, A + BX ----> AX + B or
X + AY ----> AX + Y

a) Whether one metal will replace another metal from a compound can be determined by the relative activity of the two metals. The activity series (see table) of metals lists the metals in order of decreasing reactivity. A reactive metal will replace any metal found below it in the activity series; however, only the top 4 metals are able to replace hydrogen from water.

Examples – Complete and balance the following reactions:

1) zinc + hydrochloric acid ----> 3) potassium + water (HOH) ---->

3) zinc + stannous chloride ----> 4) tin + sodium nitrate ---->

b) A nonmetal can also replace another nonmetal from a compound. The replacement is usually limited to the halogens. The activity of the halogens decreases as you go down the group VIIA column on the Periodic Table.

Examples – Complete and balance the following reactions:

1) chlorine + sodium bromide ---->

2) iodine + calcium fluoride ---->

6) Double Replacement Reactions

Double replacement reactions involve an exchange of positive ions between two compounds. These reactions generally take place between two ionic compounds in an aqueous solution. For a double replacement reaction to occur, one of the following statements is true concerning at least one of the products of the reaction.

AX + BY ----> AY + BX

a) One product is only slightly soluble or is insoluble and precipitates from the solution. (I = insoluble on the solubility table, s/I = slightly soluble on the solubility table.)

The book just calls these "reactions in which a solid forms - a precipitation reaction."

b) One product is a gas that bubbles out of the mixture. In this class you will be expected to predict the correct products of any reaction that produces:

H_2S (produced directly)
CO_2 (from the decomposition of H_2CO_3)
SO_2 (from the decomposition of H_2SO_3)
NH_3 (from the decomposition of NH_4OH)

(When one of these latter three decomposition reactions accompanies a double replacement reaction you will note that water is also formed which is a further indication that a reaction is taking place!)

c) One product is water (H_2O)

The book calls these "reactions that form water: acids and bases."

If you can recognize two ionic compounds possibly reacting in a double replacement

reaction, then you KNOW to use the solubility table to see if a precipitate forms and the table also reminds you that H^+ and OH^- react to give HOH (water).

7) Combustion Reactions

In a combustion reaction oxygen reacts with another substance, often producing energy in the form of heat and light. Combustion reactions involving hydrocarbons (compounds containing only carbon and hydrogen) and organic compounds containing only carbon, hydrogen and oxygen (fro example, alcohols) produce the compounds **carbon dioxide** and **water** if the combustion is complete (plenty of oxygen is available). Incomplete combustion (lack of enough oxygen) may produce unburned carbon and/or carbon monoxide.

$CxHy + O_2 \text{ ----> } CO_2 + H_2O$ or

$CxHyOz + O_2 \text{ ----> } CO_2 + H_2O$

Examples – Write a balanced equation for the complete combustion of these compounds.

1) Methane, CH_4

2) glucose, $C_6H_{12}O_6$

8) Sec 8.3 – Complete and Net Ionic Equations

The reaction between aqueous solutions of ionic compounds in a double displacement reaction can be written more realistically if you recognize that most ionic compounds dissociate, or separate, into cations and anions when they dissolve in water. An equation that shows dissolved ionic compounds as free ions is called a **complete ionic equation**. The equation can be simplified by eliminating ions that do not participate in the reaction (these are called **spectator ions**). When the spectator ions have been eliminated, the equation that results is called the **net ionic equation**.

Examples – Give the net ionic equation for the following reactions. The product that is not in solution is given.

1) potassium hydroxide + phosphoric acid ----> (H_2O)

2) barium chloride + potassium carbonate ----> ($BaCO_3$)

3) ammonium chloride + sodium hydroxide ----> (NH_3)

Others

sulfuric acid + potassium hydroxide ----> (H_2O)

silver nitrate + potassium phosphate ----> (Ag_3PO_4)

9) Sec 8.5 – Reactions of Metals with Non-Metals (Oxidation–Reduction)
Oxidation–Reduction reactions basically are reactions that involve the exchange of electrons. We have been discussing these kinds of reactions since Chapter 3 when we began talking about ions and electrons being either "gained" or "lost" by atoms in forming ions. Not much has changed we are just now putting these ideas into a larger context.

10) Further Thoughts

At the beginning of these Chapter Notes I mentioned that there were a number of ways that chemists use to classify reactions. No one of them can satisfy everything or everyone. The authors present this chapter with only three kinds of reaction – acid-base, precipitation, and oxidation-reduction. I think, however, that using the five types of reactions that I have mentioned goes farther into being able (a) to describe more things and (b) to help **predict** what is taking place by identify what is happening.

Consider one of the reactions I have told you to learn -- H_2CO_3 ----> H_2O + CO_2, for example. In Chapter 11 we will talk about "oxidation states" which is a term that is associated with the idea of electron transfers, etc. This reaction involves a rearrangement of atoms but does not, strictly, involve any change in the electronic structure (Chapter 11) about any of the atoms. Chapter 18 is devoted to reactions that do meet the requirements for electron transfers, etc. We may or not do that chapter in our study this year, but it is there for you to look at if you want.

Confusing? Somewhat, but this is why I prefer what we are doing with reactions compared to the presentation in your book because I want you to see relationships, and I think your book's approach is "stand alone" - just more facts to learn.