Chapter 6 – Chemical Reactions

Chemical reactions occur often with one or more of the following indications of something happening:

a. 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)

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 and to write the equations that describe the reactions.

Section 6-1 – Chemical Equations

1) Sec 6-1.1 – Constructing an Equation

A chemical equation is a concise representation of a chemical reaction. A **chemical equation** *is the representation of a chemical reaction using the symbols of elements and the formulas of compounds.*

In a chemical equation the formulas of the **reactants** (the original reacting species) are on the left and are connected by an arrow with the formulas of the **products** (the results of the reaction) which are on the right. Chemical equations are always written with the products on the left and the products are on the right – flow is left to right. If there is more than one reactant and/or one product, these are separated by a plus (+) sign.

It is important to know that in a chemical reaction, atoms are neither created not destroyed. All atoms present in the reactants must be accounted for in the products.

Symbols Used in Chemical Equations

\+ Used to separate two reactants or two products; for reactants the words used are "combines with" or "reacts with"

-> Used to represent "produces" of "yields," separates reactants from products

(s) Designates a reactant or product in the solid state; placed after the formula

↓ Alternate to (s); used only for solid product (precipitate)

(l) Designates a reactant or product in the liquid state; placed after the formula

(aq) Designates an aqueous solution; the substance is dissolved in water

(g) Designates a reactant or product in the gaseous state; placed after the formula

↑ Alternate to (g); used only for a gaseous product

$\xrightarrow{\triangle}$ or $\xrightarrow{heat}$ Indicates that heat is supplied to the reaction

A formula written above or below the yield sign indicates its use as a catalyst (a substance that speeds up a reaction without being used up).

For chemical equations we need to use the correct formulas for the reactants and products. Say that we have oxygen reacting with hydrogen to make water. The chemical equation becomes:

$$O_2 \quad + \quad H_2 \quad \text{------>} \quad H_2O$$

Remember that oxygen and hydrogen are diatomic and the symbols for their formulas have to represent that, and that water consists of two hydrogens and one oxygen in the formula. This form of the chemical equation is referred to as a **skeleton equation**.

<u>Examples</u>:

1) Write a skeleton equation for the reaction that is described by the following statement: Liquid hydrogen peroxide decomposes into water and oxygen gas in the presence of the catalyst manganese(IV) oxide.

2) Write a sentence that describes the following reaction:

$$Mg_{(s)} \quad + \quad HCl_{(aq)} \quad \text{------->} \quad MgCl_{2\,(aq)} \quad + \quad H_{2\,(g)}$$

It is important in writing chemical equations that we remember to follow faithfully the Law of Conservation of Mass which simply states that *mass can neither be created nor destroyed* in a chemical reaction. If we look at the water reaction above we see that there are two oxygen atoms on the reactant side of the equation and only one on the product side. In order to follow the law of conservation of mass we need to balance the equation. *A* **balanced chemical equation** *has the same number and types of atoms on both sides of the equation*. The formulas for the compounds involved in the chemical reaction must **never** be changed in balancing a chemical equation. *An equation is balanced by introducing* **coefficients**. Coefficients are whole numbers that are placed in front of the symbols or formulas. Their value affects all the atoms in the symbol or formula. Many chemical equations can be balanced by trial and error.

Back to our water example: $O_2 \quad + \quad H_2 \quad \text{------>} \quad H_2O$

If we put a 2 in front of the H_2O, we then have an equal number of oxygen atoms on both sides:

$O_2 \quad + \quad H_2 \quad \text{------>} \quad 2\ H_2O$ But the number of hydrogens is not.

Now we have four hydrogens on the product side (2 x 2) and if we place a 2 in front of the hydrogen on the reactants side we get four hydrogens on the left (again, 2 x 2).

$$O_2 \quad + \quad 2\ H_2 \quad \text{------>} \quad 2\ H_2O$$

The equation is now balanced!

We can always check to make sure that the equation is balanced by checking when we are done to make sure there are the same numbers of each number and type of atom on both sides. The answer to whether an equation is balanced is either a "yes" or "no." If it is "yes" we can move forward. If it is "no" we need to go back and try some more.

2) Sec 6-1.2 – (Mr. Weinkauff's) Rules for Balancing Chemical Equations

1) Determine the correct formulas for the reactants and products in the reaction.

2) Write the formulas for the reactants on the left and the formulas for the products on the right with an arrow in between. If two or more reactants or products are involved, separate their formulas with plus signs.

3) Count the number of atoms of each element in the reactants and products. A polyatomic ion appearing unchanged on both sides of the equation can be counted as a single unit.

4) Balance the elements one at a time by using coefficients. A coefficient is a small whole number that appears in front of a formula in an equation. When no coefficient is written, it is assumed to be 1. It is usually best to begin balancing by assuming that the most complex substance has a coefficient of 1. (DO NOT change the subscripts in the chemical formula of a substance!)

   It is easier to start with all elements other than oxygen and hydrogen, except for any elements that are free elements (not in any compounds).

5) Keep working by trial and error.

6) Check each atom and/or polyatomic ion to be sure that the equation is balanced.

7) Finally, make sure that all the coefficients are in the lowest possible ratio.

Examples

1) Balance the following equations.
   a) ____ $Al_{(s)}$ + ____ $N_{2\,(g)}$ ----> ____ $AlN_{(s)}$

   b) ____ $Al_{(s)}$ + ____ $CuSO_{4\,(aq)}$ ----> ____ $Cu_{(s)}$ + $Al_2(SO_4)_3$

   c) ____ $NaCl_{(aq)}$ + ____ $H_2SO_{4\,(aq)}$ -----> ____ $Na_2SO_{4\,(aq)}$ + ____ $HCl_{(aq)}$

   d) ____ $C_2H_5OH_{(l)}$ + ____ $O_{2\,(g)}$ ----> ____ $CO_{2\,(g)}$ + ____ $H_2O_{(l)}$

2) Rewrite these word equations as balanced chemical equations. You do not have to indicate the physical state of the chemicals.
   a) sodium + water ----> sodium hydroxide + hydrogen

b) iron(III) chloride + calcium hydroxide ----> iron(III) hydroxide + calcium chloride

3) Write a balanced chemical equation for the following reaction (include symbols to represent the physical state of each substance): When solid mercury(II) sulfide is heated with oxygen gas, liquid mercury and gaseous sulfur dioxide are produced.

Section 6-2 – Combustion, Combination, and Decomposition Reactions

Types of Chemical Reactions – There are many ways to catagorize chemical reactions. Our classification scheme identifies five general reaction types: Combustion, Combination, Decomposition, Single Replacement, and Double Replacement. There are common factors that help us observe the connections within the types.

As we go forward we may use other classifications that will fit our purposes at those times.

3) Sec 6-2.1 – Combustion Reactions

In a **combustion reaction,** oxygen reacts with another substance, often producing energy in the form of heat and light. This reaction is generally referred to as “burning.” When combustion reactions occur they generally produce only one product - the oxide.

$C + O_2$ -----> CO_2

$Al + O_2$ -----> Al_2O_3

Combustion reactions involving hydrocarbons (compounds containing only carbon and hydrogen) and organic compounds containing only carbon, hydrogen and oxygen (for 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.] In these cases there is more than one product but notice that both products are the oxides of the elements.

$C_xH_y + O_2$ ----> $CO_2 + H_2O$ or

$C_xH_yO_z + O_2$ ----> $CO_2 + H_2O$

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

1) Methane, CH_4

2) Glucose, $C_6H_{12}O_6$

3) Al

4) Sec 6-2.2 – 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 simple 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. Since one of the reactants is sometimes elemental oxygen, these kinds of reaction can also be classified as combustion reactions.

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 ---->

5) Sec 6-2.3 – Decomposition Reactions

A chemical property of a compound may be its ability to decompose into simpler substances. This type of reaction is simply the reverse of combination reactions. That is, in a **decomposition reaction** one compound is decomposed into two or more elements or simpler compounds. Often these reactions occur when heat is supplied.

Some examples: $2\ HgO \xrightarrow{\Delta} 2\ Hg + O_2$

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

$$2\ KClO_3 \xrightarrow{\Delta} 2\ KCl + 3\ O_2$$

There are three decomposition reactions that you need to **memorize**. These will be important later on:

a) $H_2CO_3 \longrightarrow H_2O + CO_2$
b) $NH_4OH \longrightarrow H_2O + NH_3$
c) $H_2SO_3 \longrightarrow H_2O + SO_2$

Section 6-3 - The Formation of Ions in Water

6) Sec 6-3.1 – Salts in Aqueous Solution

When one adds sodium chloride to water, the salt disappears into the aqueous medium. We note that the salt will form a *solution* (a homogeneous mixture). However, if one adds chalk (calcium carbonate) to water, it does not dissolve and settles to the bottom.

For the salt we say that it is *soluble*. The chalk is said to be *insoluble*.

When the ionic compound sodium chloride is dissolved in water, the compound is separated into individual ions (Na^+ and Cl^-). The result can be represented as:

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

Likewise, we can describe the dissolving of calcium perchlorate in water, but we need to notice that the atoms in the perchlorate ion remain together in solution; it is only the ions that separate:

$$Ca(ClO_4)_{2\ (s)} \xrightarrow{H_2O} Ca^{+2}_{(aq)} + 2\ ClO_4^{-}{}_{(aq)}$$

If a compound is insoluble then none of this separation takes place!

There are some compounds that are *partially soluble*. This means that some of the material dissolves and the rest of it stays as the solid. We will get more into this shortly.

7) Sec 6-3.2 – Strong Acids in Aqueous Solution

We mentioned acids in Chapter 4. These are molecular compounds but are special in that they produce ions when dissolved in water (like ionic compounds). It is said that the neutral compound is "ionized" by the water. Acids are named because they form the H^+ ion when dissolved in water. **Strong acids** (such as HCl, HNO_3, and H_2SO_4) are completely ionized in water.

$$HCl_{(aq)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2SO_{4\,(aq)} \xrightarrow{H_2O} 2\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Like solids that are only partially soluble in water, some acids do not completely ionize when they are dissolved in water. These are generally referred to as **weak acids**.

Section 6-4 – Single-Replacement Reactions

8) 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

This type of reaction usually occurs in aqueous solutions. The usual explanation is that "like replaces like". That means that a metal will replace a metal ion in a compound, and a non-metal will replace a non-metal ion in a compound. The products are an element and a new compound.

9) Sec 6-4.1 – Types of Equations

A **molecular equation** shows all the reactants and products in a reaction as neutral compounds. The reaction between aqueous solutions of ionic compounds or strong acids in a 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 the cations and anions as free ions in solution is called a **total or complete ionic equation**.

(molecular equation) $Zn_{(s)} + CuCl_{2\,(aq)}$ -----> $ZnCl_{2\,(aq)} + Cu_{(s)}$

(total ionic equation) $Zn + Cu^{2+} + 2\,Cl^-$ ------> $Zn^{2+} + 2\,Cl^- + Cu$

The equation can be simplified by eliminating ions that do not participate in the reaction because they are in the same state on both sides of the total ionic equation. These ions are called **spectator ions**. When the spectator ions have been eliminated, the equation that results is called the **net ionic equation**. This equation shows us what actually happens in the reaction.

(net ionic equation) $Zn + Cu^{2+}$ ------> $Zn^{2+} + Cu$

This reaction occurs in one direction only. Copper will not react with zinc ion to go the other direction.

10) Sec 6-4.2 – The Activity Series

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 6-2, page 193) 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. Because hydrogen is not really a metal but is included in the list, the series is often referred to as the *Activity Series of Metals and Hydrogen*.

<u>Examples</u> – Complete and balance the following reactions:

1) zinc + hydrochloric acid ---->

3) potassium + water (HOH) ---->

3) zinc + stannous chloride ---->

4) tin + sodium nitrate ---->

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

<u>Examples</u> – Complete and balance the following reactions:

1) chlorine + sodium bromide ---->

2) iodine + calcium fluoride ---->

For an atom to be reactive means that it has a tendency to exist as an ion, not the free element. To say that an element is stable means that it is more likely to be found in its free state.

<u>Section 6-5 – Double-Replacement Reactions – Precipitation</u>

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.

AX + BY ------> AY + BX

11) <u>Sec 6-5.1 – Soluble and Insoluble Compounds</u>

There are all kinds of guidelines and rules for whether compounds are soluble or insoluble. Table 6-3 (page 195) give some of these.

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 that is given in the *Chemistry Helper*.)

12) <u>Sec 6-5.2 – Formation of a Precipitate</u>

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.*

This is one of the three kinds of double-replacement reactions.

Example:

(molecular equation) $2\ AgNO_3 + CuCl_2$ ------> $Cu(NO_3)_2 + 2\ AgCl$

(total ionic) $2\ Ag^+ + 2\ NO_3^{--} + Cu^{2+} + 2\ Cl^-$ ------> $Cu^{2+} + 2\ NO_3^{--} + 2\ AgCl$

(net ionic equation) $2\ Ag^+ + 2\ Cl^-$ ------> $2\ AgCl$

As mentioned before, the net ionic equation shows us the real reaction occurring in the system. As in the case of the single-replacement reactions, the net ionic equation focuses on the driving force for the reaction which is the formation of solid silver chloride.

For these reactions, if a solid is not formed there is no reaction.

13) Rules for Writing Precipitation Reactions

Predicting the results of precipitation reactions follows several simple steps:

1. Write the formulas for the reactants and the products (by switching partners) making sure that the compounds have the correct formulas based on the charges of the ions.

2. Use Table 6-3, the *Chemistry Helper*, or the following table to determine if one of the products is insoluble.

Ion	General Solubility Rule
NO_3^-	All nitrates are soluble
$C_2H_3O_2^-$	All acetates are soluble ($AgC_2H_3O_2$ only moderately)
Cl^-, Br^-, I^-	All chlorides, bromides and iodides are soluble except Ag^+, Pb^+ and Hg_2^{2+}. ($PbCl_2$ is slightly soluble in cold water and moderatel soluble in hot water.)
SO_4^{2-}	All sulfates are soluble except those of Ba^{2+}, Pb^{2+}, Ca^{2+} and Sr^{2+}
CO_3^{2-} and PO_4^{3-}	All carbonates and phosphates are insoluble except those of Na^+, K^+ and NH_4^+. (Many acid phosphates are soluble).
OH^-	All hydroxides are insoluble except those of Na^+ and K^+. Hydroxides of Ba^{2+} and Ca^{2+} are slightly soluble.
S^{2-}	All sulfides are insoluble except those of Na^+, K^+, NH_4^+ and those of the alkaline earths: Mg^{2+}, Ca^{2+}, Sr^{2+} and Ba^{2+}. (Sulfides of Al^{3+} and Cr^{3+} hydrolyze and precipiate as the corresponding hydroxides.
Na^+, K^+ and NH_4^+	All salts of sodium ion, potassium ion and ammonium ion are soluble except several uncommon ones.

3. If one of the two products is insoluble in water, then a precipitation reaction occurs and we can write the equation representing the reaction.

If no precipitate is formed from the combining of two solutions, then we need to also look at the next two types of double-replacement reactions to see if a reaction occurs.

Section 6-6 – Double-Replacement Reactions – Neutralization

Neutralization reactions are also referred to as **acid-base reactions** because acids and bases neutralize each other when they react.

14) Sec 6-6.1 – Strong Acids and Strong Bases

According to Bronsted Lowery:

An acid is any compound which can donate a proton (H^+).
A base is any compound which can accept a proton (H^+).

In Section 4-6 and Section 6-3 we described what strong acids are and how they behave in water. A second important group of compounds known as **strong bases** dissolve in water to form the hydroxide ion (OH^-). The bases are ionic compounds and they simply separate when dissolved in water.

$$NaOH_{(s)} \xrightarrow{H_2O} Na+_{(aq)} + OH^-_{(aq)}$$

$$Ba(OH)_{2\,(s)} \xrightarrow{H_2O} Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

15) Sec 6-6.2 – Neutralization Reactions

In a **neutralization reaction** with a strong acid and a strong base, the H^+ from an acid and the OH^- from a base will react to form a salt and water. During the reaction, a proton is donated by the acid to the base to yield water. The remaining ions form a salt. A **salt** is formed from the cation of the base and the anion from the acid.

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Here, hydrochloric acid (a proton donor) reacts with sodium hydroxide (a proton acceptor and a source of hydroxide ion) to form sodium chloride (a salt) and water.

The total ionic and net ionic equations are:

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

For these kinds of reactions the net ionic equation is **always** the same: H^+ and OH^- combine to make water, H_2O. You should notice that in the middle of the solubility table on the *Chemistry Helper* this also shows (with the water written as HOH).

[We will discuss neutralization reactions using weak acids and bases later in Chapter 13.]

16) Sec 6-6.3 – Gas-Forming Neutralization Reactions

Some reactions do not produce a solid product or water. These produce a molecular compound that is a gas (remember the physical properties of ionic compounds and molecular compounds). For example, calcium carbonate reacts with hydrochloric acid and produces a molecular compound other than water, called carbonic acid.

$$CaCO_{3\ (s)} + 2\ HCl_{(aq)} \text{ ------> } CaCl_{2\ (aq)} + H_2CO_{3\ (aq)}$$

At this point you might think that this reaction does not occur because both listed products are soluble and neither is water. The "trick" is that carbonic acid spontaneously decomposes (remember that this is one of the decomposition reactions I said you should learn) to form water and carbon dioxide.

$$H_2CO_{3\ (aq)} \text{ ----> } H_2O_{(l)} + CO_{2\ (g)}$$

The overall reaction then becomes:

$$CaCO_{3\ (s)} + 2\ HCl_{(aq)} \text{ ------> } CaCl_{2\ (aq)} + H_2O_{(l)} + CO_{2\ (g)}$$

And the net ionic equation is:

$$CO_3{}^{2-}{}_{(aq)} + 2\ H^+{}_{(aq)} \text{ ------> } H_2O_{(l)} + CO_{2\ (g)}$$

Salts containing the following neutralize strong acids by forming a gas:

- $HCO_3{}^-$ and $CO_3{}^{2-}$ salts produce H_2CO_3, which decomposes to H_2O and CO_2
- $HSO_3{}^-$ and $SO_3{}^{2-}$ salts produce H_2SO_3, which decomposes to H_2O and $SO_{2\ (g)}$
- HS^- and S^{2-} salts produce $H_2S_{(g)}$
- CN^- salts produce $HCN_{(g)}$

17. Complete and balance the following equations. Then give the net ionic equation. The compound in parentheses is the "product" of the reaction.

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

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

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

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

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

18. Copper dissolves in nitric acid forming copper(II) nitrate, nitrogen dioxide and water. Write out and balance this equation.