CHEM101: GENERAL CHEMISTRY
Lecture 5 – Chemical Reactions

I. Introduction

A. Chemical reactions are processes involving chemical change

II. Chemical Equations

A. In a chemical reaction, one or more pure substances are changed to one or more other chemical substances.
   1. The substances that are changed in the reaction are called **reactants**.
   2. The substances that they are changed into are called **products**.

B. Chemical equations are used to describe chemical reactions.
   1. A chemical equation has the following elements:
      a. The chemical symbols for the reactants are placed on the left-land side of the equation and the chemical symbols for the products are placed on the right-hand side.
      b. Either an arrow ($\rightarrow$) or and equal sign (=) are used to indicate the reactants converting to products.
      c. A plus sign (+) is used to separate individual reactants and products.
   2. For the reaction: "hydrogen and oxygen react to form water" the chemical equation is

$$2H_2(g) \; + \; O_2(g) \; \rightarrow \; 2H_2O(l)$$

      a. The coefficients indicate that two molecules of H_2 are needed to react with on molecule of O_2 to produce two molecules of H_2O.
      b. The letters in parentheses are used to indicate that state of the reactant or the product
         i. (*s*) for solid
         ii. (*l*) for liquid
         iii. (*g*) for gas
         iv. (*aq*) for aqueous solution (dissolved in water)
   3. Chemical equations embody a fundamental law of nature called the **law of conservation of matter**.
      a. The law states, that in a chemical reaction atoms are neither created or destroyed, only rearranged.
      b. All of the matter present in the reactants is also present in the products of the reaction.
      c. This allows us to treat chemical equation like mathematical equations
   4. The coefficients are use to **balance** the equation.
      a. For example, in the chemical equation describing the formation of liquid water from hydrogen gas and oxygen

gas there are 4 hydrogens on both the left and right hand sides of the equation and there are 2 oxygens on both the left and right hand sides of the equation.

Exercise 5.3

5.3 Identify which of the following are consistent with the law of conservation of matter. For those that are not, explain why they are not.

a. $4\ Al(s) \ + \ 3O_2(g) \ \rightarrow \ 2Al_2O_3(s)$

b. $P_4(s) \ + \ O_2(g) \ \rightarrow \ P_4O_{10}(s)$

c. 3.2*g* G oxygen + 3.21 g sulfur $\rightarrow$ 6.41g sulfur dioxide

d. $CH_4(g) \ + \ 2\ O_2(g) \ \rightarrow \ CO_2(g) \ + \ 2\ H_2O(g)$

Exercise 5.5

5.5 Determine the number of atoms of each element on each side of the following equations and decide which equations are balanced:

b. $H_2S(aq) \ + \ I_2(aq) \ \rightarrow \ 2HI(aq) \ + S(s)$

c. $KClO_3(s) \ + \ KCl(s) \ + \ O_2(g)$

d. SO2(g) + H2O(l) _ H2SO3(aq)

e. $Ba(ClO_3)_2(aq) \ + \ H_2SO_4(aq) \ \rightarrow \ 2HClO_3(aq) \ + \ BaSO_4(s)$

III. Types of Reactions

A. There are many types of reactions, of which we will only focus a few.

B. There are different ways of classifying reactions

C. Your text chooses to first classify all reactions as either redox or nonredox reactions.

Figure 5.2 – Classification of chemical reactions

1. This classification is made on the basis of whether the reactants and products exchange electrons during the reaction.

D. Under both of these heading are the combination and the decomposition reactions.

1. Combination reactions involve the combining of two or more reactants to produce a single product.

2. Decomposition reactions involve the breakdown of a single reactant to produce two or more products.

IV. Redox Reactions

A. Redox reactions is an abbreviation for oxidation/reduction reactions.

B. **Oxidation** reactions originally described reactions involving combining oxygen with the various elements to form oxides.

1. Because oxygen is a very electronegative element, when it combines with other elements, the other elements are viewed lose electrons
   a. This is strictly true only if an ionic compound is formed.

$$4Fe(s) \quad + \quad 3O_2(g) \quad \rightarrow \quad 2Fe_2O_3(s)$$

   i. In this reaction the oxygen is taking away 3 electrons from each iron to produce Fe^{3+} ions.

   b. However, even in covalent compounds, where the valence electrons are shared, the sharing is not equal when oxygen is involved.
   i. It is useful to consider that the oxygen is taking the electrons away from the other element, even though technically they are sharing the electrons.

$$2H_2(g) \quad + \quad O_2(g) \quad \rightarrow \quad 2H_2O(l)$$

2. A broader definition of *oxidation* is a chemical process that results in the loss of electrons by a substance.
   a. In this definition, oxygen does not necessarily need to be involved.
   i. Other reactants can be responsible for the removal of the electrons.

C. The opposite of oxidation is reduction.

1. **Reduction** reaction originally described reactions in which ionic forms of metals were converted to elemental forms by supplying them with electrons:

$$2Fe_2O_3(s) \quad + \quad 3C(s) \quad \rightarrow \quad 4Fe(s) \quad + \quad 3CO_2(g)$$

   a. As with oxidation, the electrons do not have to be complete removed from the element; if in the course of a chemical reaction and element becomes bonded to a less electronegative element, it is considered to be reduced.
   i. For example, in the reaction

$$2H_2(g) \quad + \quad O_2(g) \quad \rightarrow \quad 2H_2O(l)$$

the oxygen is reduced.

D. There is a systematic set of rules that can be used to determine if an element has been oxidized or reduced in a chemical reaction.

1. These rules result in the assignment of an oxidation number for each element participating in a reaction:
    a. *Rule 1*: The oxidation number (O.N.) of any uncombined element is 0.
    b. *Rule 2*: The O.N. of a simple ion is equal to the charge on the ion.
    c. *Rule 3*: The O.N.'s of group IA(1) and IIA(2) elements are +1 and +2 respectively.
    d. *Rule 4*: The O.N. of hydrogen is +1.
    e. *Rule 5*: The O.N. of oxygen is -2 except in peroxides (-O-O-), whre it is -1.
    f. *Rule 6*: The algebraic sum of the O.N.'s of all atoms in a complete compound formula equals zero.
    g. *Rule 7*: The algebraic sum of the O.N.'s of all atoms in a poly atomic ion is equal to the charge on the ion.
2. If in a reaction the oxidation number for an element increases, it is oxidized; conversely, if its oxidation number decreases it is reduced.
3. Applying these rules to a reaction:

$$S(s) \quad + \quad O_2(g) \quad \rightarrow \quad SO_2(g)$$

    a. As reactants both S and O_2 have oxidation numbers of 0 (Rule 1)
    b. The oxidation number for the oxygen in SO_2 is -2 (Rule 5)
    c. The oxidation number for the sulfur in SO_2 is +4 (Rule 6)
    d. In this reaction the sulfur is oxidized while the oxygen is reduced.

E. The various definitions used for oxidation and reduction are shown in Table 5.1

Table 5.1 - Common uses of the terms exidation and reduction

1. Oxidation
    a. To combine with oxygen
    b. To lose hydrogen
    c. To lose electrons
    d. To increase in oxidation number
2. Reduction
    a. To lose oxygen
    b. To combine with hydrogen
    c. To gain electrons
    d. To decrease in oxidation number

F. Oxidizing and reducing agents
1. A molecule that is capable of removing electrons from another molecule is called an **oxidizing agent**.

a. In the process the oxidizing agent becomes reduced because it has gained electrons

2. A molecule that is capable of donating electrons to another molecule is called a **reducing agent**.

   a. In the process the reducing agent becomes oxidized because it has lost electrons.

V. Decomposition Reactions

A. **Decomposition reactions** have the form $A \rightarrow B + C$

   1. They can be recognized because they have only one reactant, which breaks down to produce two or more products.

Figure 5.4 - Decomposition reaction

B. Examples:

   1. $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

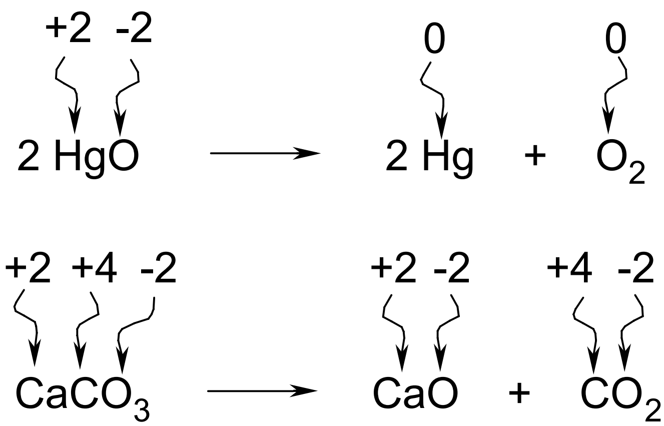
      a. This is the reaction that Joseph Priestley used to discover oxygen in 1774.

   2. $CaCO_3(s) \rightarrow CaO(s) + CO_2(s)$

      a. This reaction is called slaking and is the one used to produce lime (CaO) from limestone ($CaCO_3$).

C. Decomposition reactions can be either redox reactions or non redox reactions.

   1. Which of these a particular decomposition reaction is can be determined by determining the oxidation numbers of the each element in both the reactants and products of the reaction.



      a. The first reactions is a redox reaction because the mercury is reduced while the oxygen is oxidized.

      b. The second reaction is a nonredox reaction because there is no change in the oxidation state of any of the elements in the reaction.

VI. Combination Reactions

A. **Combination reactions** have the form $A + B \rightarrow C$
  1. They can be recognized because they have only one product, which is formed from two or more products.
  2. It is also called an addition reaction

Figure 5.6 - Combination reaction

B. Examples:
  1. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
  2. $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

C. Combination reactions can also be either redox reactions or non redox reactions.
  1. Which of these a particular decomposition reaction is can be determined by determining the oxidation numbers of the each element in both the reactants and products of the reaction.

VII. Replacement Reactions

A. There are two types of replacement reactions
  1. **Single-replacement**, where one of the reactants and one of the products is an element
  2. **Double-replacement**, where all of the products and reactants are compounds.

B. Single-replacement reactions
  1. Are also called substitution reactions
  2. They are always redox reactions
  3. The have the form $A + BX \rightarrow B + AX$, where A and B are elements and BX and AX are compounds.

Figure 5.8 - Single-replacement reaction

  4. Example
     a. $3C(s) + 2Fe_2O_3 \rightarrow 4Fe(s) + 3CO_2$

C. Double-replacement reactions
  1. Are also called metathesis reactions
  2. They are always nonredox reactions
  3. The have the form $AX + BY \rightarrow BX + AY$, where A and B are elements and BX and AX are compounds.

Figure 5.10 - Double-replacement reaction

  4. Often these reactions involve substances dissolved in water.
     a. Precipitation reactions
     b. Acid-base reactions
  5. Example, acid
     a. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

VIII. Ionic Equations

A. Ionic compounds that dissolve in water, along and some highly polar molecular compounds, such as strong acids, that dissolve in water, dissociate into ions

B. Total ionic equation

1. It useful to identify which species are actually present in solution to determine which are involved in the reaction.
2. To do this
    a. Any ionic compound (salt) that is soluble in water is written as dissociated ions
    b. Any strong acid is written as disscociated ions
    c. Any weak acid is written as undissociated.
3. **Spectator ions**
    a. Any ions that appear on both sides of the total ionic equation, they can be cancelled out.
        i. This produces the **Net Ionic Equation**.

C. Net ionic equation

1. The species, ions and molecules, that remain are the ones involved in the reaction
    a. If nothing remains, then there is no reaction

D. Examples

1. $HCl + NaOH \rightarrow NaCl + H_2O$
    a. Total Ionic Equation:

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$$

    b. Net Ionic Equation:

$$H^+ + OH^- \rightarrow H_2O$$

2. $2KBr + CoCl_2 \rightarrow 2KCl + CoBr_2$
    a. Total Ionic Equation:

$$2K^+ + Br^- + Co^{2+} + 2Cl^+ \rightarrow 2K^- + Cl^- + Co^{2+} + 2Br^-$$

    b. Net Ionic Equation
        i. No reaction (nothing remains)

IX. Energy and Reactions

A. Energy changes accompany all reactions

B. Can be in the form of

1. Heat
2. Light
3. Sound
4. High energy chemical bonds

C. Expressed in units of calories or Joules

1. $2\,H_2(g) + O_2(g) \rightarrow 2H_2O(g) + 115.6\text{ kcal }(483.7\text{ kJ})$

D. When heat is given off it is **exothermic**

E. When heat is absorbed it is **endothermic**.

X. The Mole and Chemical Equations

A. Stoichiometry - calculating the mass relationships in a chemical reactions
B. Typically quantitate substances by their mass
C. However, chemical reactions relate numbers of reactants to numbers of products:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

1. Statements consistent with this equation include:
   α. 1 CH_4 molecule + 2 O_2 molecules $\rightarrow$ 1 CO_2 molecule + 2 H_2O molecules.
   b. 10 CH_4 moleucles + 20 O_2 molecules $\rightarrow$ 10 CO_2 molecules + 20 H_2O molecules.
   c. 100 CH_4 moleucles + 200 O_2 molecules $\rightarrow$ 100 CO_2 molecules + 200 H_2O molecules.
   d. 6.02×10^{23} CH_4 moleucles + 12.0×10^{23} O_2 molecules $\rightarrow$ 6.02×10^{23} CO_2 molecules + 12.0×10^{23} H_2O molecules.
   e. 1 mol CH_4 + 2 mol O_2 $\rightarrow$ 1 mol CO_2 + 2 mol H_2O.
   f. 16 g CH_4 + 64.0 g O_2 $\rightarrow$ 44.0 g CO_2 + 36.0 g H_2O

D. To make use of the chemical equation we usually need to convert from mass to moles, then use the coefficients in the balanced chemical equation to relate the moles of one substance in the reaction with moles of another.
   1. In the end we usually need to convert back to mass

Figure 5.12 - Relationships for problem solving based on balanced equations

Exercise 5.49

5.49 - An important metabolic process of the body is the oxidation of glucose to water and carbon dioxide. The equation for the reaction is

$$C_6H_{12}O_6(aq) + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

a. What mass of water in grams is produced when the body oxidizes 1.00 mol of glucose?

b. How many grams of oxygen are needed to oxidize 1.00 mol of glucose?

XI. The Limiting Reactant

A. Reactions can proceed only as long as all of the reactants are present.
   1. In most cases one of the reactants will be used up before the others

a. This reactant is called the **limiting reactant** because it determines how much product can be made.

b. When the limiting reactant is used up the reaction comes to a stop.

Figure 5.13 – Limiting reactant is used up in a reaction.

B. The quantity of product that is produced when all of the limiting reactant is used up is called the **theortical yield**.

Exercise 5.51

Exercise 5.51: A sample of 4.00 *g* of methane (CH_4) is mixed with 15.0 *g* of chlorine (Cl_2).

a. Determine which is the limiting reactant according to the following equation:

$$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$$

b. What is the maximum mass (theoretical yield) of CCl_4 that can be formed?

XII. Reaction Yields

A. When carried out in the lab, a reaction rarely produces the theoretical yield

1. Loss of reactants and products, along with side reactions that produce different products reduce the yield of the desired product.
2. The experimental yield is called the **actual yield**.

B. The actual yield is compared to the theoretical yield by determining the **percent yield**:

$$PercentYield = \frac{Actual\,Yield}{Theoretical\,Yield} x 100\%$$

Exercise 5.59

Exercise 5.59: A sample of calcium metal with a mass of 2.00 *g* was reacted with excess oxygen. The following equation represents the reaction that took place:

$$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$$

The isolated product (CaO) weighed 2.26 *g*. What is the percent yield of the reaction?
