Changes in a material or system are called reactions, and they are divided into chemical and physical reactions. Energy is the driving force of all changes, both physical and chemical reactions. Energy is always involved in these reactions. If a system is more stable by losing some energy, a reaction takes place, releasing energy. Such a reaction is said to be exothermic. Supplying energy to a system also causes a reaction. Energy absorbing reactions are called endothermic reactions. Sometimes, the amount of energy involved in a reaction may be so small that the change in energy is not readily noticeable.

An equation can be used to describe a physical reaction, which involves a change of states. For example, melting, sublimation, evaporation, and condensation can be represented as follows. In these equations, (s) stands for solid, (l) for liquid (l), and (g) for gas.

- melting: $\text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(l)$
- sublimation: $\text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(g)$
- evaporation: $\text{C}_2\text{H}_5\text{OH}(l) \rightarrow \text{C}_2\text{H}_5\text{OH}(g)$
- condensation: $\text{NH}_3(g) \rightarrow \text{NH}_3(l)$

In these changes, no chemical bonds are broken or formed, and the molecular identities of the substances have not changed.

Is the phase transition between graphite and diamond a chemical or physical reaction?

$\text{C(graphite)} \rightarrow \text{C(diamond)}$.

The crystal structures of diamond and graphite are very different, and bonding between the carbon atoms is also different in the two solid states. Because chemical bonds are broken and new bonds are formed, the phase transition of diamond and graphite is a chemical reaction.

Chemicals or substances change converting to one or more other substances, and these changes are called chemical reactions. At the molecular level, atoms or groups of atoms rearrange resulting in breaking and forming some chemical bonds in a chemical reaction. The substances undergoing changes are called reactants, whereas substances newly
formed are called **products**. Physical appearances of products are often different from reactants. Chemical reactions are often accompanied by the appearance of gas, fire, precipitate, color, light, sound, or odor. These phenomena are related to energy and properties of the reactants and products. For example, the oxidation of propane releases heat and light, and a rapid reaction is an explosion,

\[
\text{C}_3\text{H}_8 + 5 \text{O}_2 \rightarrow 3 \text{CO}_2 + 4 \text{H}_2\text{O}
\]

A balanced equation also shows a macroscopic quantitative relationship. This balanced reaction equation shows that five moles of oxygen reacts with one mole of propane generating three moles of carbon dioxide and four moles of water, a total of 7 moles of products in the combustion reaction.

At the molecular level, this equation shows that for each propane molecule, 5 oxygen molecules are required. The three carbon atoms are converted to three molecules of carbon dioxide, whereas the 8 hydrogen atoms in propane are oxidized to 4 water molecules. The numbers of \(\text{H}\), \(\text{C}\), and \(\text{O}\) atoms are the same on both sides of the equation.

We study properties of substances so that we know how to make use of them. Tendencies of a substance to react, either by itself or with others, are important chemical properties. Via properties, we understand chemical reactions, which are best studied by experimentation and observation. After you have performed many experiments, you may generalize certain rules and facts. Knowing these rules and facts enables you to solve problems that you have not yet encountered.

The most important aspect of a chemical reaction is to know what are the reactants and what are the products. For this, the best description of a reaction is to write an equation for the reaction. A **chemical reaction equation** gives the reactants and products, and a **balanced chemical reaction equation** shows the mole relationships of reactants and products. Often, the amount of energy involved in the reaction is given. Dealing with the quantitative aspect of chemical reactions is called **reaction stoichiometry**.

For example, when clamshells, \(\text{CaCO}_3\), are heated, a gas \(\text{CO}_2\) will be released, leaving a white powder (solid \(\text{CaO}\)) behind. The equation of the reaction is written as:

\[
\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2
\]

The equation indicates that one mole of \(\text{CaCO}_3\) gives one mole each of \(\text{CaO}\) and \(\text{CO}_2\). Amounts of substances represented by chemical formulas have been introduced on the two previous pages, and these concepts should help to figure out the stoichiometry of reactions when a reaction equation is given.

**Example 1**

When 10.0 g pure calcium carbonate is heated and converted to solid calcium oxide \(\text{CaO}\), how much calcium oxide should be obtained? If only 5.0 grams \(\text{CaO}\) is obtained, what is the actual yield?

**HINT**

Under ideal conditions, amounts of substance in the reaction equation is as indicated below:

\[
\begin{alignat}{2}
\end{alignat}
\]
\[ \ce{CaCO_3 \rightarrow CaO + CO_2} \]
\[ 100.0 \text{ g} \times \frac{1 \text{ mol} \ce{CaCO_3}}{100 \text{ g} \ce{CaCO_3}} \times \frac{1 \text{ mol} \ce{CaO}}{1 \text{ mol} \ce{CaCO_3}} \times \frac{56 \text{ g} \ce{CaO}}{1 \text{ mol} \ce{CaO}} = 5.6 \text{ g} \ce{CaO} \]

**DISCUSSION**

An inefficient conversion is given here, but the method shows the details of consideration. If the amount of \(\ce{CaO}\) obtained is not 5.6 g, one can conclude that the sample may not be pure.

**Example 2**

When 10.0 g pure calcium carbonate is heated and converted to solid calcium oxide \(\ce{CaO}\), how much \(\ce{CO_2}\) at standard condition is released?

**HINT**

\[ \ce{CaCO_3 \rightarrow CaO + CO_2} \]
\[ 10.0 \text{ g} \ce{CaCO_3} \times \frac{1 \text{ mol} \ce{CO_2}}{100 \text{ g} \ce{CaCO_3}} \times \frac{22.4 \text{ L} \ce{CO_2}}{1 \text{ mol} \ce{CO_2}} = 2.24 \text{ L} \ce{CO_2} \]

**DISCUSSION**

We have taken a short cut in this formulation compared to Example 1. Examples 1 and 2 illustrate the evaluation of quantities in g and in L.

**Writing Equations for Chemical Reactions**

**Chemical reaction equations** truly represent changes of materials. For many reactions, we may only be able to write equations for the overall reactions. For example, common sense tells us that when sugar is fully oxidized, carbon dioxide and water are the final products. The oxidation reaction is the same as the combustion reaction. Thus we write

\[ \ce{C12H22O11 + 12 O2 \rightarrow 12 CO2 + 11 H2O} \]

This illustrates the methods used for writing balanced reaction equations:

1. **Determine the reactants and products**: In this case, the products are \(\ce{CO2}\) and \(\ce{H2O}\), determined by common sense. We know that.

2. Apply the fundamental principle of **conservation of atoms**. Numbers of atoms of each kind must be the same before and after the reactions.

3. **Balance one type of atoms at a time**: We may use \(\ce{H}\) or \(\ce{C}\) to begin. Since there are 12 \(\ce{C}\) atoms on the left, the coefficient is 12 for \(\ce{CO2}\). Similarly, 22 \(\ce{H}\) atoms produce 11 \(\ce{H2O}\) molecules.

4. **Balance the oxygen atoms on both sides**: There are a total of 35 \(\ce{O}\) atoms on the right hand, and the
coefficient for $\ce{O2}$ should be 11.

Example 3

The compound $\ce{N2O5}$ is unstable at room temperature. It decomposes yielding a brown gas $\ce{NO2}$ and oxygen. Write a balanced chemical reaction equation for its decomposition.

HINT

The first step is to write an unbalanced equation indicating only the reactant and products:

$$\ce{N2O5 -> NO2 + O2}$$

A $\ce{N2O5}$ molecule decomposes into two $\ce{NO2}$ molecule, and half of $\ce{O2}$.

$$\ce{N2O5 -> 2 NO2 + \frac{1}{2}O2}$$

In order to give whole number **stoichiometric coefficients** to the equation, we multiply all the stoichiometric coefficients by 2.

$$\ce{2 N2O5 -> 4 NO2 + O2}$$

DISCUSSION

This example illustrates the steps used in writing a balanced equation for a chemical reaction. This balanced equation does not tell us how a $\ce{N2O5}$ molecule decomposes, it only illustrates the overall reaction.

Example 4

When solutions of $\ce{CaCl2}$ and $\ce{AgNO3}$ are mixed, a white precipitate is formed. The same precipitate is also observed when $\ce{NaCl}$ solution is mixed with $\ce{AgCH3CO2}$ solution. Write a balanced equation for the reaction between $\ce{CaCl2}$ and $\ce{AgNO3}$.

HINT

The common ions between $\ce{NaCl}$ and $\ce{CaCl2}$ are $\ce{Cl-}$ ions, and $\ce{Ag+}$ ions are common between the two silver containing compounds. The question illustrates a scientific deduction used in the determination of products. The product is $\ce{AgCl}$, and the balanced reaction is

$$\ce{CaCl2 + 2 AgNO3 -> 2 AgCl + Ca(NO3)2}$$

DISCUSSION

In reality, solutions of salts contain ions. In this case, the solutions contain $\ce{Ca^{2+}}$, $\ce{Cl-}$, $\ce{Ag+}$, and $\ce{NO3-}$ ions. The $\ce{Cl-}$ and $\ce{Ag+}$ ions form an insoluble solid, and a precipitate is formed,

$$\ce{Cl- + Ag+ -> AgCl(s)}$$

$\ce{Ca^{2+}}$ and $\ce{NO3-}$ ions are **spectator ions**.
Chemical Reactions

One of the most important topics in chemistry is chemical reaction. In this page, we only concentrate on the stoichiometry conveyed by reaction equations. Other topics related to chemical reactions are:

- Excess and Limiting Reagents or reactants left over or used up
- Features of chemical reactions or classification of reactions
- Chemical kinetics or reaction rates
- Reaction mechanism or how actually reaction proceed

Balancing Redox Reactions

Balancing oxidation and reduction reaction equations is a little more complicated than what we discussed here. You have to have the skills to assign oxidation states, explain oxidation and reduction in terms of oxidation-state change, and write half reaction equations. Then you will be able to balance redox reactions. All these are given in the next module on Chemical Reactions.

Skill Developing Problems

1. What is/are the product(s) containing carbon when methane, $\ce{CH4}$, is burned in the air?
   
   Hint: $\ce{CO2}$

   Generalization:
   Combustion of $\ce{C}$-containing compounds converts all $\ce{C}$ to $\ce{CO2}$.

2. Use the common sense method to find the molecular formula for hydrogen sulfide, whose molecular weight is 34.1. (Atomic weight, $\ce{H}$, 1.008; $\ce{S}$, 32.066)
   
   Hint: $\ce{H2S}$

   Generalization:
   Sulfur and oxygen are group 6 elements, and they form $\ce{H2O}$ and $\ce{H2S}$.

3. When 30.0 g of $\ce{Al}$ (atomic weight 27.0) is heated in oxygen (atomic mass 16.0), an aluminum oxide, $\ce{Al2O3}$, is formed. How much oxide should be obtained?
   
   Hint: 56.7 g

   A Variation:
   How much (in g) oxygen is required?

4. When $\ce{KClO3}$ is heated, it decomposes to give solid $\ce{KCl}$ and oxygen gas. If 0.500 mol $\ce{O2}$ is collected, how many grams of $\ce{KCl}$ should be obtained? (Atomic wt: $\ce{K}$, 39.098; $\ce{Cl}$, 35.453)
   
   Hint: 24.9 g
Method suggestion:
For the reaction: \(\ce{2 KClO3 \rightarrow 2 KCl + 3 O2}\)

the formulation suggestion is:

\[
\text{0.50 mol O}_2 \times \frac{2 \text{ mole KCl}}{3 \text{ mol O}_2} \times \frac{74.6 \text{ g KCl}}{1 \text{ mol KCl}} = \text{?? g KCl}
\]

5. A solution containing pure \(\ce{BaCl2}\) is treated with excess amounts of \(\ce{H2SO4}\), and the precipitate \(\ce{BaSO4}\) is collected and dried. If 13.2 g of \(\ce{BaSO4}\) are collected, how many moles of \(\ce{Cl^{-}}\) ions are left in the solution?

Atomic wt: \(\text{H}\), 1.008; \(\text{O}\), 16.00; \(\text{S}\), 32.06; \(\text{Cl}\), 35.45; \(\text{Ba}\), 137.33.

Hint: 0.113 mol

Variations:
- How much (in g) \(\ce{BaCl2}\) is present in the solution?
- How much silver nitrate is required to precipitate all the chloride ions?

The reaction is: \(\text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2 \text{ H}^+ + 2 \text{ Cl}^-\).

0.0566 mole of \(\text{Ba}\) correspond to 0.113 mol of \(\text{Cl}^-\) in \(\text{BaCl}_2\).

Method Suggestion:

\[
\text{13.2 g} \times \frac{1 \text{ mol BaSO}_4}{233.39 \text{ g}} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol BaSO}_4} = 0.113 \text{ mol}
\]

6. A power plant burns coal, and this process is equivalent to burning 999 kg of sulfur a day. How many kg of \(\ce{SO2}\) are emitted per day if the power plant does not have pollution control devices to recover the sulfur? Atomic wt: \(\text{C}\), 12.00; \(\text{O}\), 16.00; \(\text{S}\), 32.06.

Hint: 1998 kg

Further consideration:
The molecular weight of \(\ce{SO2}\) is about twice the atomic weight of \(\ce{S}\). Thus the weight of \(\ce{SO2}\) is twice that of \(\ce{S}\).

Variations: How much (in mole and L) \(\ce{SO2}\) is generated per day?
If all \(\ce{SO2}\) is converted to \(\ce{H2SO4}\), how much (in mol and kg) sulfuric acid is produced? (3055 kg)

7. How many moles of water will be formed when one mole of propane \(\ce{C3H8}\) is burned in an excess amount of air?

Hint: 4 moles; \(\text{C3H8 + 5 O2 \rightarrow 3 CO2 + 4 H2O}\)

Skill:
Work out a balanced reaction equation.

Variations:
- How many grams of water will be produced?
- How many moles of \(\ce{CO2}\) will be produced?
8. A mixture containing $\text{Na}_2\text{SO}_4$, but no other sulfate, is analyzed by precipitation with $\text{BaCl}_2$. A 2.37 g mixture sample gave a 2.57 g $\text{BaSO}_4$ precipitate. What is the percentage of $\text{Na}_2\text{SO}_4$ in the mixture?

Hint: 66.0%

Skill:
The problem illustrates a strategy for chemical analysis.

9. Suppose 2.33 g of $\text{CaCl}_2$ and $\text{Ca(NO}_3)_2$ mixture gives 2.22 g of $\text{AgCl}$ when $\text{Ag(NO}_3)$ is used as a reagent to precipitate the chloride $\text{Cl}^-$ ions. What is the percentage of $\text{CaCl}_2$ in the mixture?

Hint: 36.9%

**Atomic wt:** $\text{N}$, 14.0; $\text{O}$, 16.0; $\text{Cl}$, 35.5; $\text{Ca}$, 40.1; $\text{Ag}$, 107.9.

Skill:
This problem also illustrates a strategy for chemical analysis.

**Contributors**

- **Chung (Peter) Chieh** (Professor Emeritus, Chemistry @ University of Waterloo)