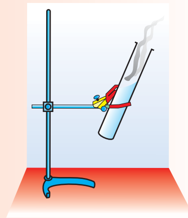
UNIT 4



Chemical Reactions and Stoichiometry

Unit Outcomes

After completing this unit, you will be able to:

- *understand the fundamental laws of chemical reactions and how they are applied*;
- *develop skills in writing and balancing chemical equations*;
- *understand energy changes in chemical reactions*;
- *know types of chemical reactions*;
- *develop skills in solving problems based on chemical equations* (*mass-mass, volume-volume and mass-volume problems);*
- *develop skills in determining limiting reactant, theoretical yield, actual yield, and percentage yield*;
- *understand oxidation-reduction reactions and analyze redox reactions by specifying the oxidizing agents and reducing agents, the substance reduced or oxidized;*
- *understand the rate of chemical reaction, the state of a chemical equilibrium and factors affecting them; and*
- *demonstrate scientific enquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, communicating, measuring, asking questions, designing experiments, interpreting data, drawing conclusions, applying concepts, relating cause and effect, and problem-solving.*

MAIN CONTENTS

Start-up Activity

A chemical reaction enables a space shuttle to be launched, which is powered by a chemical reaction between pure liquid hydrogen (serving as a fuel) and oxygen. Assume that the fuel tank contains 32,000 litres of H_2 and the oxidizer tank contains 40,000 litres of O_2;

Analysis

1. What type of reaction takes place?
2. Write the balanced chemical equation for the reaction.
3. What volume of product is formed in the reaction?
4. What mass of product is formed in the reaction?
   (Assume that the pressure remains constant in this process).
   Submit your findings to the teacher.

4.1 INTRODUCTION

Competencies

By the end of this unit, you will be able to:

- define chemical reaction; and
- give some examples of chemical reactions.

Activity 4.1

Form a group and discuss the following phenomenon:

1. When a space shuttle leaves the ground on its way into orbit, what does the brightness and warmth of the flame indicate?
2. What are the notations that indicate a chemical change might be taking place?

Present your conclusion to the class.

Chemical reactions are the basis of chemistry. Chemical reactions occur around us all the time. For example, the burning of fuel, the souring of milk, metabolic processes of our body and the decay of plants are some familiar chemical reactions in daily life.

A chemical reaction is the process in which reacting substances, called reactants, are converted to new substances, called products. The characteristics of the products are completely different from those of the reactants. The conversion process is a chemical change.

$$\text{Reactants} \longrightarrow \text{Products}$$

For example, if you burn magnesium with oxygen, the magnesium and oxygen are completely converted to magnesium oxide. Magnesium oxide is a soft, white, crumbling powder. These characteristics of magnesium oxide are completely different from the characteristics of the original substances, magnesium and oxygen. Magnesium and oxygen are no longer present in the elemental form.

In summary, a chemical reaction has occurred in which the reactants, magnesium and oxygen, underwent a complete chemical change, giving the product magnesium oxide.

All chemical reactions include three types of changes in the original substances. These are changes in composition, properties and energy.

Activity 4.2

Form a group and perform the following task.

List some chemical processes that occur in your daily life. Identify the reactants and products in each of these chemical processes.

Present your findings to the class.

Note that, in daily life, we use different terms for the same process of chemical change. For example "the souring of milk" occurs due to the process of fermentation. In scientific discussion we generally have a single term for each process.

4.2 FUNDAMENTAL LAWS OF CHEMICAL REACTIONS

Competencies

By the end of this unit, you will be able to:

- state the law of conservation of mass and illustrate the law, using examples;
- demonstrate the law of conservation of mass, using simple experiments;
- state the law of definite proportion and illustrate it, using examples;
- demonstrate the law of definite proportion, using a simple experiment; and
- State the law of multiple proportion and illustrate it, using examples.

Activity 4.3

Form a group and discuss the following phenomenon. When wood burns, the ash weighs much less than the original wood. Where did the "lost mass" go? How can you estimate the mass of the wood that is no longer present?

Present your conclusion to the class.

While investigating the quantitative relations between substances in chemical reactions, scientists formulated the three basic laws of chemical combination. These are:

i) The law of conservation of mass

ii) The law of definite proportions

iii) The law of multiple proportions

i) *The Law of Conservation of Mass*

In 1774, the French chemist Antoine Lavoisier performed an experiment in which he heated a sealed glass container that held a sample of tin in air. He measured the mass of the substances before and after heating and found them to be the same. This and other similar experimental observations became the basis of the law of conservation of mass.

Historical Note

Antoine Lavoisier

French chemist Antoine Lavoisier is considered the founder of modern chemistry. He found that the amount of matter before a chemical reaction is equal to the amount of matter afterwards, even though the matter may change in its form. Lavoisier also experimented with the role of oxygen in combustion and respiration in both plants and animals.

The law of conservation of mass states that matter is neither created nor destroyed in a chemical reaction. In other words, the mass of the reactants is exactly equal to the mass of the products, within the limits of experimental error. This law is also known as the law of indestructibility of matter.

Mass of reactants = Mass of products

There is no loss or gain of substances during a chemical reaction, and mass is conserved.

For example, consider the decomposition of mercury (II) oxide. When 100 g of mercury (II) oxide decomposes by heat, 92.6 g of mercury and 7.4 g of oxygen are formed. Note that the total mass of mercury and oxygen after decomposition is 100 g:

$$\begin{array}{ccc} 2HgO & \rightarrow & 2Hg \; + \; O_2 \\ \underbrace{100g}_{100g} & & \underbrace{92.6g \quad 7.4g}_{100g} \\ \text{Reactant} & & \text{Products} \end{array}$$

Experiment 4.1

Investigation of the Law of Conservation of Mass

Objective: To determine the mass of substances before and after a reaction.

Apparatus: Flask, test tube, thread, rubber stopper, balance.

Chemicals: Sodium chloride, silver nitrate.

Procedure:

1. Take 50 mL of silver nitrate solution in a conical flask.
2. Tie a thread around the top of a test tube. Fill the test tube with a saturated solution of sodium chloride. Suspend the test tube in the flask by means of a thread held by a rubber stopper, as shown in Figure 4.1.
3. Weigh the flask (*and its contents*). Record the result as m_1.
4. Mix the liquids by tilting the conical flask so that the sodium chloride pours into the silver nitrate solution.
5. Reweigh the conical flask and contents and record as m_2. Compare m_1 and m_2.

Observations and analysis

1. What was the colour of the solution after the reaction?
2. Is there any difference in mass between m_1 and m_2?
3. What is your conclusion from this experiment?
4. Write the balanced chemical equation for the reaction.

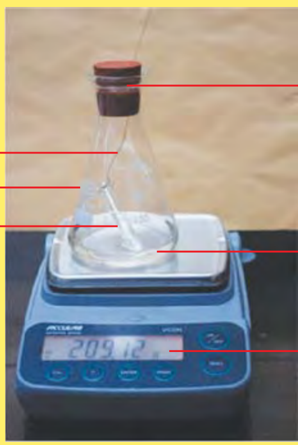


Figure 4.1 Conservation of mass.

ii) The Law of Definite Proportions

The law of *definite proportions states* that a compound always contains the same elements in the same proportion by mass. This means that all pure samples of a compound have the same composition regardless of the source of sample. This law is also known as the law of constant composition. For example, sample of water could be obtained from different sources, such as from a river, the ground, or the ocean. But whatever the original source, all forms of pure water contains 11.2% hydrogen and 88.8% oxygen by mass. These percentages represent a ratio of 1.0 to 8.0 (1:8), by mass, of hydrogen to oxygen.

This ratio is constant (fixed) for water. In other words, a compound with a different ratio of hydrogen and oxygen is not water.

Similarly, in forming the compound ZnO, 65.0 g of zinc combines with 16.0 g of oxygen. This is 80.2% zinc and 19.8% oxygen, by mass. As is the case for water, the composition of ZnO is constant. In forming ZnO, zinc combines with oxygen in a definite proportion.

Experiment 4.2

Investigation of the Law of definite proportions

Objective: To determine the mass of copper from copper (II) oxide.

Apparatus: Burner, stand, combustion tube, two glass test tubes, two watch glasses.

Chemicals: Copper powder, copper (II) carbonate, hydrogen gas

Procedure:

1. Prepare samples of copper (II) oxide using the following two methods:
   i) Make copper (II) oxide by heating copper powder in one of the test tubes.
   ii) Make copper (II) oxide by heating copper (II) carbonate in the second test tube.
   (In this case, the heating process produces a chemical change through thermal decomposition)
2. Take 1 g from each of the samples of copper (II) oxide (*from i and ii*). Place each of these samples in a watch glass.
3. Reduce each of these samples: use the combustion tube to heat the samples in a stream of hydrogen as shown in Figure 4.2.
4. Weigh the copper metal that remains in each case. Compare the measurements.

Observations and analysis

1. What is the mass of copper produced in each case?
2. Why is copper metal produced in each case?
3. What can you conclude from the experiment? Write a short report on your observations.

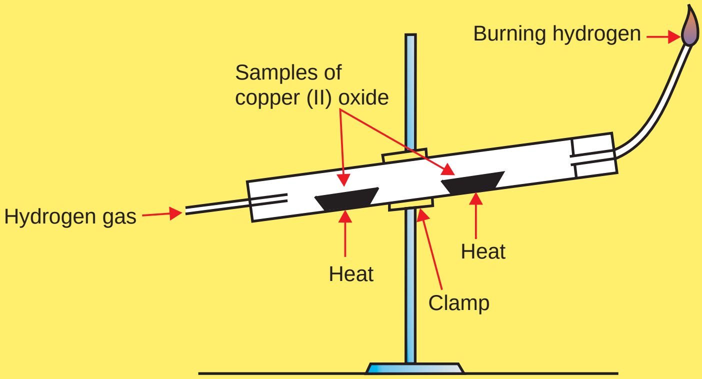


Figure 4.2 Reduction of copper (II) oxide by hydrogen.

iii) The Law of Multiple Proportions

The law of multiple proportions states that when two elements combine to form more than one compound, the masses of one element combined with a fixed mass of the second element are in the ratio of small whole numbers. This law can be illustrated by the two oxides of carbon. The two oxides of carbon are carbon monoxide (CO) and carbon dioxide (CO_2). In CO_2, 1.0 g of carbon is combined with 2.67 g of oxygen; whereas in CO, 1.0 g of carbon is combined with 1.33 g of oxygen. By comparing 2.67 g of oxygen with 1.33 g of oxygen, it is found that the masses of oxygen in the two compounds that combine with the same mass of carbon are in the simple whole number ratio, 2:1.

$$\frac{2.67\text{ g of oxygen in } CO_2}{1.33\text{ g of oxygen in } CO} = \frac{2}{1} = 2:1$$

Activity 4.4

Form a group and perform the following task:

The following table illustrates the law of multiple proportions using five oxides of nitrogen. In the table, fill the mass ratio of nitrogen to oxygen and determine the mass of oxygen in each compound that combine with a fixed mass (1g) of nitrogen.

Present your conclusion to the class.

Compound	Molecular formula	Mass ratio of N to O (N:O)	Ratio of Oxygen in the oxides
Dinitrogen monoxide	N_2O		
Nitrogen monoxide	NO		
Dinitrogen trioxide	N_2O_3		
Nitrogen dioxide	NO_2		
Dinitrogen pentoxide	N_2O_5		

Exercise 4.1

Give appropriate answers for the following questions.

1. Classify the following as chemical or physical changes:
   - a the souring of tella
   - b freezing ice cream
   - c plant growth
   - d boiling of an egg
   - e heating sugar
   - f fermentation
   - g the magnetization of iron
   - h the fading of dye in cloth
2. Iron and chlorine form two compounds, A and B. Compound A contains 1.27 g of chlorine for each 1 g of iron whereas compound B contains 1.9 g of chlorine for each 1 g of iron. Show that the masses of chlorine are in the ratio 2:3. Do they obey the law of multiple proportions? Explain.
3. Consider the following two chemical changes:

   i) When a material made of iron rusts, its mass increases.

   ii) When a match stick burns, its mass decreases.

   Do you think that these two observations violate the law of conservation of mass? Explain.

Critical Thinking

4. Discuss how the law of conservation of matter is explained by Dalton's atomic theory.

4.3 CHEMICAL EQUATIONS

Competencies

By the end of this unit, you will be able to:

- describe the conventions used to write chemical equations;
- balance chemical equations, using the inspection method;

- balance chemical equations, using the Least-Common-Multiple (LCM) method.

Activity 4.5

Form a group and discuss each of the following:

1. What is the difference between a chemical equation and a chemical reaction?
2. Which law is satisfied when a chemical equation is balanced? Take a simple chemical reaction to illustrate this law.

Present your conclusion to the class.

A chemical equation is a shorthand representation of a chemical reaction in terms of chemical symbols and formulas. In a chemical equation the starting substances are called reactants; and the new substances produced are known as products.

Reactants are written on the left side and products on the right side of the equation. An arrow ($\rightarrow$) is placed between the two sides to indicate transformation of reactants into products.

$$\text{Reactants} \rightarrow \text{products}$$

4.3.1 Writing Chemical Equation

In writing chemical equation, instead of using words, chemical symbols and formulas are used to represent the reaction.

Steps to Write a Chemical Equation

1. Write a word equation: A word equation is stated in words. For example, the word equation for the reaction between sodium and chlorine to produce sodium chloride is written as:

   Sodium + Chlorine $\rightarrow$ Sodium chloride (*word equation*)

   Note that we read the ' + ' sign as 'reacts with' and the arrow can be read as 'to produce', 'to form', 'to give' or 'to yield'.

2. Write the symbols and formulas for the reactants and products in the word equation.

   $Na + Cl_2 \rightarrow NaCl$ (*Chemical equation*)

3. Balance the equation.

   $$2Na + Cl_2 \rightarrow 2NaCl$$

Generally, any chemical equation must fulfil the following conditions:

i) The equation must represent a true and possible chemical reaction.

ii) The symbols and formulas must be written correctly. The elements– hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine exist as diatomic molecules. These elements should be written as molecules in the equation.

iii) The equation must be balanced.

A chemical equation has both *qualitative* and *quantitative meanings*.

Qualitatively, a chemical equation indicates the types of the reactants and products in the reaction.

Quantitatively, a chemical equation expresses the relative number (*amount)* of moles, molecules or masses of the reactants and products.

4.3.2 Balancing Chemical Equation

Which should be adjusted in balancing a chemical equation, the subscripts or the coefficients?

According to the law of conservation of mass, atoms are neither created nor destroyed during a chemical reaction. As a result, the number of atoms of each element should remain the same before and after the reaction. Therefore, the main reason why all chemical equations must be balanced is just to obey the law of conservation of mass.

To balance a chemical equation means to equalize the number of atoms on both sides of the equation by putting appropriate coefficients in front of the formulas.

Only two methods of balancing chemical equations will be discussed under this topic. These are the inspection and the Least Common Multiple (LCM) method.

1. *The Inspection Method*

Most simple chemical equations can be balanced using this method. Balancing an equation by inspection means to adjust coefficients by trial and error until the equation is balanced. Follow the following four steps to balance the chemical equation.

Step 1: Write the word equation.

Step 2: Write the correct symbols or formulas for the reactants and products.

Step 3: Place the smallest whole number coefficients in front of the symbols or formulas until the number of atoms of each element is the same on both sides of the equation.

Step 4: Checking: By counting the number of atoms on both sides of the equation, make sure that the atoms of all elements are balanced and also the coefficients are expressed as the smallest whole number ratio.

Note:

When you balance an equation, do not change any symbol or formula of any compound. If you change a symbol or formula, it no longer represents the element or compound required by the equation.

Example 1

Balance the equation for the reaction between magnesium and oxygen to produces magnesium oxide.

Solution:

Step 1: Magnesium + Oxygen $\rightarrow$ Magnesium oxide

Step 2: $Mg + O_2 \rightarrow MgO$ (*unbalanced*)

Step 3: Put coefficients to balance the equation

- Oxygen is not balanced. There are two oxygen atoms on the left side and one on the right side. Hence, place the coefficient 2 in front of MgO.

$$Mg + O_2 \rightarrow 2MgO \text{ (\textit{unbalanced})}$$

- Now Mg is not balanced. There is one Mg on the left side and two on the right side. Thus, place the coefficient 2 in front of Mg.

$$2Mg + O_2 \rightarrow 2MgO \text{ (\textit{balanced})}$$

Step 4: **Checking:** There are two Mg and two O atoms on each side of the equation. Therefore, the equation is correctly balanced.

$$2Mg + O_2 \rightarrow 2MgO$$

Exercise 4.2

Balance the following chemical equation, using the inspection method:

1. $Na + H_2O \rightarrow NaOH + H_2$
2. $CaCO_3 \rightarrow CaO + CO_2$
3. $H_2O_2 \rightarrow H_2O + O_2$
4. $Al + H_3PO_4 \rightarrow AlPO_4 + H_2$
5. $HNO_3 + H_2S \rightarrow NO + S + H_2O$

2. *The LCM Method*

In the LCM method, the coefficients for the balanced chemical equation are obtained by taking the LCM of the total valency of reactants and products and then dividing it by total valency of reactants and products. All the necessary steps to balance a chemical equation by the LCM method, are shown by the following examples.

Example 2

When aluminium reacts with oxygen, aluminium oxide is formed. Write the balanced chemical equation for the reaction.

Solution:

Step 1: Represent the reaction by a word equation.

$$\text{Aluminium} + \text{Oxygen} \rightarrow \text{Aluminium oxide}$$

Step 2: Change the words to symbols and formulas for the reactants and products.

$$Al + O_2 \rightarrow Al_2O_3$$

Step 3: Place the total valency of each atom above it.

$$\overset{3}{Al} + \overset{4}{O_2} \rightarrow \overset{6}{Al_2}\,\overset{6}{O_3}$$

Now the equation shows

- The valency of aluminium as 3.
- The total valency of oxygen is 2 × 2 = 4.
- The total valency of aluminium in Al_2O_3 is 3 × 2 = 6.
- The total valency of oxygen in Al_2O_3 is 2 × 3 = 6.

Step 4: Find the LCM of each total valency and place it above the arrow.

$$\overset{3}{Al} + \overset{4}{O_2} \xrightarrow[\text{L.C.M}]{12} \overset{6}{Al_2}\,\overset{6}{O_3}$$

Step 5: Divide the LCM by each total valency number to obtain the coefficients for each of the reactants and products. Place the obtained coefficients in front of the respective formulas.

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \text{ (\textit{balanced})}$$

Checking: There are 4 aluminium and 6 oxygen atoms on both sides of the equation. Hence, the chemical equation is correctly balanced.

Example 3

When iron reacts with water, iron (III) oxide and hydrogen are produced. Write the balanced equation.

Solution:

Step 1: Iron + water $\rightarrow$ Iron (III) oxide + hydrogen.

Step 2: $Fe + H_2O \rightarrow Fe_2O_3 + H_2$

Step 3: $\overset{3}{Fe} + \overset{2}{H_2}\overset{2}{O} \rightarrow \overset{6}{Fe_2}\overset{6}{O_3} + \overset{2}{H_2}$

Step 4: $\overset{3}{Fe} + \overset{2}{H_2}\overset{2}{O} \xrightarrow[L.C.M]{6} \overset{6}{Fe_2}\overset{6}{O_3} + \overset{2}{H_2}$

Step 5: $2Fe + 3H_2O \rightarrow Fe_2O_3 + 3H_2$ (*balanced*)

Checking: There are 2 iron, 6 hydrogen, and 3 oxygen atoms on each side of the equation. Thus, the equation is balanced.

Example 4

The reaction of ammonium sulphate with aluminium nitrate would form aluminium sulphate and ammonium nitrate.

Solution:

Step 1: Ammonium sulphate + Aluminium nitrate $\rightarrow$ Aluminium sulphate + Ammonium nitrate

Step 2: $(NH_4)_2SO_4 + Al(NO_3)_3 \longrightarrow Al_2(SO_4)_3 + NH_4NO_3$

Step 3: $\overset{2}{(NH_4)_2}\overset{2}{SO_4} + \overset{3}{Al}\overset{3}{(NO_3)_3} \longrightarrow \overset{6}{Al_2}\overset{6}{(SO_4)_3} + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 4: $\overset{2}{(NH_4)_2}\overset{2}{SO_4} + \overset{3}{Al}\overset{3}{(NO_3)_3} \xrightarrow[L.C.M]{6} \overset{6}{Al_2}\overset{6}{(SO_4)_3} + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 5: $3(NH_4)_2SO_4 + 2Al(NO_3)_3 \longrightarrow Al_2(SO_4)_3 + 6NH_4NO_3$ (balanced)

Checking: There are 12 nitrogen, 24 hydrogen, 3 sulphur, 30 oxygen and 2 aluminium atoms on both sides of the equation. Thus, the equation is correctly balanced.

Exercise 4.3

1. Write the balanced chemical equation to represent the following reactions.

   a Sulphur dioxide reacts with oxygen to produce sulphur trioxide.

   b Potassium chlorate when heated produces potassium chloride and oxygen.

c Sodium carbonate reacts with hydrochloric acid to form water, carbon dioxide and sodium chloride.

d Silver oxide decomposes to silver and oxygen gas.

2. Balance the following equations by the LCM method.

a $PCl_5 + H_2O \longrightarrow H_3PO_4 + HCl$

b $Mg + H_2O \longrightarrow Mg(OH)_2 + H_2$

c $Zn(NO_3)_2 \longrightarrow ZnO + NO_2 + O_2$

d $H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O$

e $NH_3 + O_2 \longrightarrow NO + H_2O$

4.4 ENERGY CHANGES IN CHEMICAL REACTIONS

Competencies

***By the end of this section, you will be able to*:**

- explain energy changes in chemical reactions;
- define endothermic and exothermic reactions;
- describe endothermic and exothermic reactions;
- illustrate endothermic and exothermic reactions using diagrams;
- conduct simple experiment to demonstrate exothermic and endothermic reactions;
- describe the importance of chemical changes in production of new substances and energy.

Activity 4.6

Form a group and discuss each of the following phenomena:

When the bread baked, does the bread absorb or release heat energy? Justify your answer.

Present your conclusion to the class.

Almost all chemical reactions are accompanied by energy changes. These energy changes could be in the form of heat energy, light energy, electrical energy, and so on.

On the basis of energy changes, chemical reactions can be divided into exothermic and endothermic reactions.

4.4.1 Exothermic and Endothermic Reactions

Can heat energy be considered as a reactant or product?

Exothermic Reaction

A chemical reaction that releases heat energy to the surroundings is known as an exothermic reaction. During an exothermic process, heat is given out from the system to its surroundings and this heat energy is written on the right side of the equation as shown below.

$$\text{Reactants} \rightarrow \text{Products} + \text{Heat}$$

For example, the burning of carbon with oxygen produces carbon dioxide and heat is released during the reaction. Thus, the reaction is exothermic and written as:

$$C + O_2 \rightarrow CO_2 + \text{Heat}$$

Endothermic Reaction

A chemical reaction which absorbs heat energy from the surroundings is known as an endothermic reaction. During an endothermic process, heat flows into the system from its surroundings and the heat is written on the left side of the equation.

$$\text{Reactants} + \text{Heat} \rightarrow \text{products}$$

For example, the reaction between carbon and sulphur to form carbon disulphide is an endothermic reaction because heat is absorbed in the reaction.

$$C + 2S + \text{Heat} \rightarrow CS_2$$

The amount of heat energy liberated or absorbed by a chemical reaction is called heat of reaction or change in enthalpy for the reaction. It is symbolized as ΔH. Its unit is expressed in kilojoules per mol $\left(\text{kJ}/\text{mol}\right)$. The change in enthalpy (ΔH) is the difference between the energy of the products and the energy of the reactants.

$\Delta H = H_p - H_r$; where H_p is the heat content (*energy*) of the product, H_r is the heat content (*energy*) of the reactant.

4.4.2 Energy Diagrams

For endothermic reactions, ΔH is positive because the energy of the product is higher than the energy of the reactant. As a result, the enthalpy of the system increases as shown in Figure 4.3.

since, $H_p > H_r$, ΔH = positive ($\Delta H > 0$)

For example, when nitrogen reacts with oxygen to form nitrogen dioxide, 66.4 kJ of heat energy is absorbed (*i.e.*, ΔH = + 66.4 kJ/mol) and thus the reaction is endothermic.

$$N_2\ (g) + 2O_2\ (g) \rightarrow 2NO_2\ (g);\ \Delta H = +\ 66.4\ \text{kJ/mol}$$

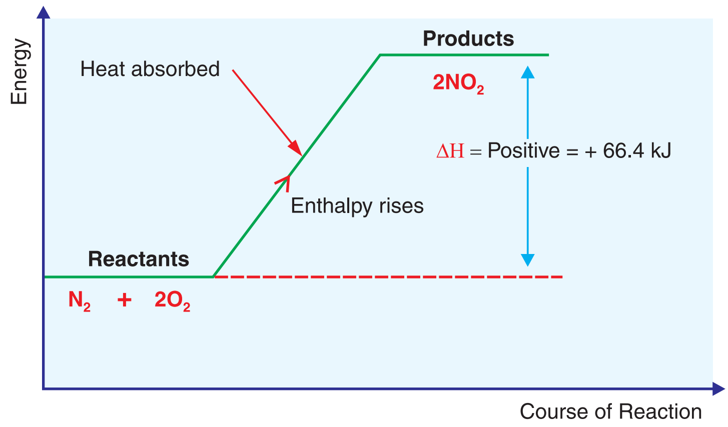


Figure 4.3 Energy diagram for an endothermic reaction.

For exothermic reactions, ΔH is negative because the energy of the reactants is greater than the energy of the products. Thus, the enthalpy of the system decreases, as shown in Figure 4.4.

$$H_p < H_r \Rightarrow \Delta H = \text{negative } (\Delta H < 0)$$

For example, when carbon burns in oxygen to produce carbon dioxide, 393.5 kJ of heat energy is liberated and hence the reaction is exothermic (ΔH = –393.5 kJ/mol).

$$C(s) + O_2(g) \rightarrow CO_2(g):\ \Delta H = -393.5\ \text{kJ/mol}$$

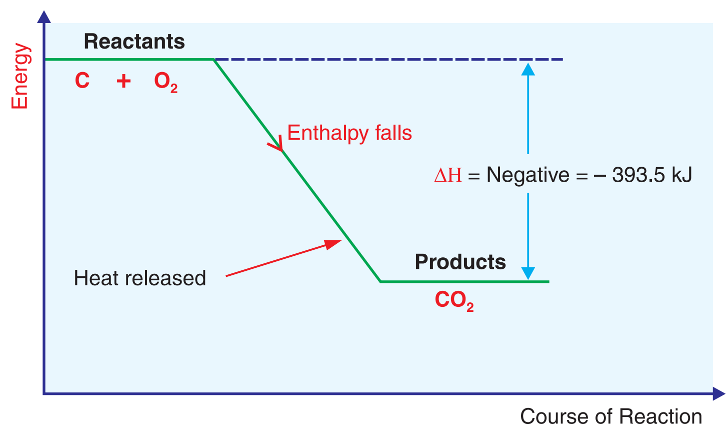


Figure 4.4 Energy diagram for an exothermic reaction.

Experiment 4.3

Investigating the Heat Involved in a Chemical Reaction

Objective: To determine the exothermic/endothermic nature of the reaction between sulphuric acid and sugar.

Apparatus: beaker and reagent bottle.

Chemicals: Concentrated H_2SO_4 and sugar.

Procedure:

1. Take small amount of sugar in a beaker.
2. Add a little concentrated sulphuric acid to the sugar.
3. Touch the outer surface of the beaker and record your observation.

Figure 4.5 Reaction between sulphuric acid and sugar.

Observations and analysis

1. Does the beaker feel hot or cold when you touch it?
2. Did you see any steam in the beaker?
3. What is the colour of the product formed?
4. Write a balanced chemical equation.
5. What can you conclude from the experiment?

   [Caution-When mixing concentrated acid and water, always add the acid to the water; never add water to concentrated acid.]

Experiment 4.4

Investigating the Heat Involved in a Reaction

Objective: To investigate the exothermic/endothermic nature of the process when ammonium nitrate is dissolved in water.

Apparatus: Beaker, thermometer, stirrer.

Chemicals: Ammonium nitrate and water.

***Procedure*:**

1. Take 100 mL of water in a beaker and record its temperature.
2. Dissolve 15 g of solid ammonium nitrate (NH_4NO_3) in the 100 mL of water.
3. Touch the outer surface of the beaker and record the temperature of the solution with the help of a thermometer.

Observations and analysis:

1. Does the beaker feel hot or cold when you touch it?
2. Is the temperature increased or decreased after the addition of NH_4NO_3?
3. What do you conclude from this experiment?

Figure 4.6 Dissolution of ammonium nitrate in water.

Activity 4.7

Form a group and perform the following task. In your daily life you encounter with many chemical changes involving energy. List some of such changes and discuss their importance.

Share your findings with the rest of the class.

Chemical reactions bring about chemical changes. All chemical changes are accompanied by energy changes. This energy is usually in the form of heat, light, or electricity.

Energy changes produced by chemical reactions have many practical applications (*uses*). For example, energy lifts rockets, runs cars, and extracts metal from compounds.

Many applications involve the energy produced by fuel combustion, which liberates large amounts of heat. The energy can be converted from one form to another. For example, the energy that fuel combustion produces can convert water to steam. The steam can run a turbine that creates electricity.

Respiration (breathing) creates energy for our bodies. Breathing releases the energy our living cells produce by oxidizing glucose. This energy helps to maintain our body temperature and body exercises.

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{Energy}$$

Exercise 4.4

In each of the following cases, determine the sign of ΔH. State whether the reaction is exothermic or endothermic, and draw an enthalpy diagram.

a $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 285.8\text{ kJ}$

b $H_2O(I) + 40.7\text{ kJ} \rightarrow H_2O(g)$

4.5 TYPES OF CHEMICAL REACTIONS

Competencies

***By the end of this section, you will be able to*:**

- list the four types of chemical reactions;
- define combination reaction and give examples;
- conduct some experiments on combination reactions in groups;

- define decomposition reaction and give examples;
- conduct some experiments on decomposition reactions in group;
- define single displacement reactions and give examples;
- conduct some experiments on simple displacement reactions in groups;
- define double displacement reaction and give examples; and
- conduct some experiments on double displacement reactions in groups.

Activity 4.8

Form a group and discuss the following chemical reactions that occur during the:

a digestion of food in our body.

b fermentation of 'tej'.

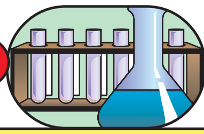
c burning of kerosene in a stove.

Share your discussion with the rest of the class.

Chemical reactions are classified into four categories. These are combination, decomposition, single displacement and double displacement reactions.

i) Combination Reactions

Experiment 4.5



Investigation of Combination Reaction

Objective: To investigate the reaction between sulphur and iron.

Apparatus: Test tube, stand, burner, watch glass

Chemicals: Sulphur powder, iron filings

Procedure:

1. Mix about 3 g of iron filings and 2 g of powdered sulphur in a watch glass.
2. Transfer the mixture in a glass test tube.
3. Mount the test tube in a sloping position on a stand as shown in Figure 4.7.
4. Heat the test tube until the mixture in the glass glows red hot.
5. Remove the test tube from the flame and observe the result.

Observations and analysis:

1. What were the colours of iron filings and sulphur before the reaction?
2. What was the colour of the resulting compound after the reaction?
3. Write a balanced chemical equation for the reaction.
4. Identify the type of reaction.

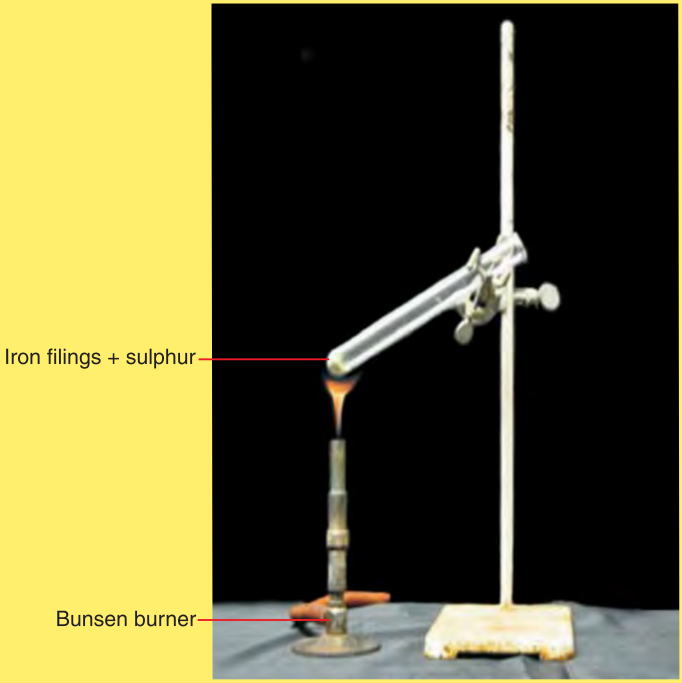


Figure 4.7 The reaction between iron and sulphur.

A reaction in which two or more substances combine to form a single substance is called a combination reaction. In a combination reaction, two elements, two compounds, or an element and a compound react to form a single compound. Combination reactions can be represented by the following general form of equation.

$$A + B \rightarrow AB;$$

where the reactants A and B are elements or compounds, the product AB is a compound.

Such type of reaction is also known as synthesis or composition reaction.

Examples

- Magnesium burns in oxygen to form magnesium oxide.

$$\underset{\text{element}}{2Mg} + \underset{\text{element}}{O_2} \rightarrow 2MgO$$

- Water and carbon dioxide combine to form carbonic acid.

$$H_2O + CO_2 \rightarrow H_2CO_3$$

Exercise 4.5

Complete and balance the following combination reactions.

1. $CO + O_2 \rightarrow$
2. $H_2O + SO_2 \rightarrow$
3. $Na_2O + H_2O \rightarrow$
4. $CaO + CO_2 \rightarrow$

ii) Decomposition Reactions

Experiment 4.6

Investigation of Decomposition Reaction

Objective: To investigate the decomposition of copper (II) carbonate.

Apparatus: Test tube, stand, burner, cork, delivery tube.

Chemicals: Copper (II) carbonate and lime water.

Procedure:

Put copper (II) carbonate powder in a glass test-tube. Mount the test tube in a sloping position on a stand as shown in Figure 4.8. Fit a cork and a delivery tube to the test tube. Put another test tube containing lime water at the end of the delivery tube. Heat the copper (II) carbonate with a burner.

Observations and analysis:

1. What was the colour of copper (II) carbonate before heating?
2. What was the colour during heating and after cooling?
3. What change did you observe in the lime water?
4. Write a balanced chemical equation for the reaction.

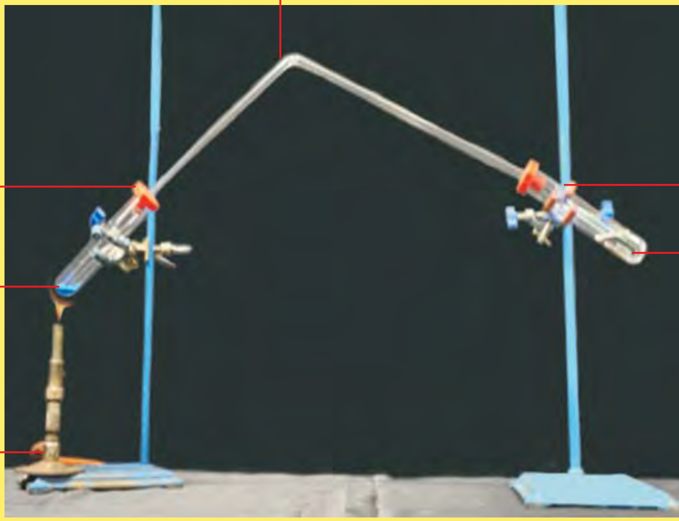


Figure 4.8 Decomposition of copper (II) carbonate.

A decomposition reaction is a reaction that involves the breaking down of a single compound into two or more elements or simpler compounds. A decomposition reaction can be carried out using heat, light, electricity or a catalyst. But most decomposition reactions are carried out when heat is supplied and this heat energy is indicated by a 'delta' (Δ) symbol above the arrow. The general form of equation for a decomposition reaction is:

$$AB \rightarrow A + B$$

where the reactant AB must be a compound and the products A and B could be elements or compounds.

Examples

- Water is decomposed to hydrogen and oxygen gases when electricity is passed through it.

$$2H_2O \xrightarrow[\text{current}]{\text{electric}} 2H_2 + O_2$$

- When sodium bicarbonate is heated, it decomposes to give sodium carbonate, carbon dioxide, and water.

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

Let us consider the decompositions of nitrates and carbonates:

1. *Decomposition of Metallic Nitrates*

a) Decomposition of group IA nitrates produces nitrites and oxygen.

Examples

$$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$$

$$2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$$

b) Decomposition of all metal nitrates, except group IA metals, gives nitrogen dioxide, metal oxide and oxygen gas.

Examples

$$2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$$

$$2Zn(NO_3)_2 \xrightarrow{\Delta} 2ZnO + 4NO_2 + O_2$$

$$2Pb(NO_3)_2 \xrightarrow{\Delta} 2PbO + 4NO_2 + O_2$$

2. *Decomposition of Metallic Carbonates*

All metal carbonates, except sodium and potassium, decompose when heated to form the metal oxide and carbon dioxide.

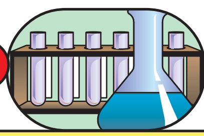
Examples

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2$$

$$CuCO_3 \xrightarrow{\Delta} CuO + CO_2$$

iii) Single Displacement Reactions

Experiment 4.7



Investigation of Single Displacement Reaction

Objective: To investigate the displacement reaction between iron and copper (II) sulphate.

Apparatus: Iron rod and beaker.

Chemicals: Copper (II) sulphate.

Procedure:

1. Clean a piece of iron rod or iron knife with emery paper to remove any rust.
2. Take copper sulphate solution in a beaker.
3. Dip the iron rod into the copper (II) sulphate solution as shown in Figure 4.9 and wait for a few minutes. What did you observe on the iron rod?
4. Allow the reactants to stand for one day and observe any change on the iron rod.

Observations and analysis:

1. What did you observe on the iron rod after one day?
2. Write a balanced chemical equation for the reaction.
3. Write the conclusion for the experiment.

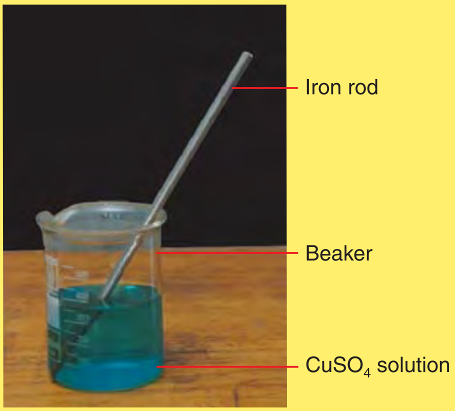


Figure 4.9 Reaction between iron and copper (II) sulphate.

A reaction in which one element displaces another element from its compound is known as single displacement or replacement reaction. Such a reaction is represented by the following two general forms.

$$A + BC \rightarrow B + AC$$

If A is a metal, it will displace B to form AC, provided A is a more active metal than B.

$$A + BC \rightarrow BA + C$$

If A is a non-metal, it will displace C to form BA, provided A is a more active non-metal than C.

In general, a more reactive element displaces a less reactive element from a compound.

Examples of single-displacement reactions

a Active metals displace hydrogen from acids.

Reactive metals such as potassium, calcium, sodium, and zinc displace hydrogen gas from dilute acids.

For example, zinc is an active metal, and it displaces hydrogen from hydrochloric acid; but copper metal cannot do so.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
$$Au + HCl \rightarrow \textit{No reaction}$$

b Reactive metals, such as potassium, calcium, and sodium react vigorously with water to displace hydrogen:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$
$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

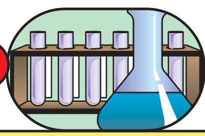
Exercise 4.6

Complete and balance the following single displacement reactions:

1. $Zn + CuSO_4 \longrightarrow$
2. $Cu + Zn(NO_3)_2 \longrightarrow$
3. $F_2 + CaCl_2 \longrightarrow$
4. $Br_2 + NaCl \longrightarrow$

iv) Double Displacement Reactions

Experiment 4.8



Investigation of Double Displacement Reaction

Objective: To observe the displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.

Apparatus: Beaker, stirrer, filter paper, filter funnel.

Chemicals: Na_2SO_4 and $Ba(NO_3)_2$.

Procedure:

1. Take solution of $Ba(NO_3)_2$ into a beaker and add dropwise Na_2SO_4 solution. Then stir it continuously.
2. Filter the precipitate using a filter paper and funnel. Collect the filtrate or the solution in a clean beaker.

Observations and analysis:

1. Write the names of the compounds that are formed as a precipitate and as solution at the end of the reaction.
2. What was the colour of the precipitate.
3. Write the balanced chemical equation for the reaction.

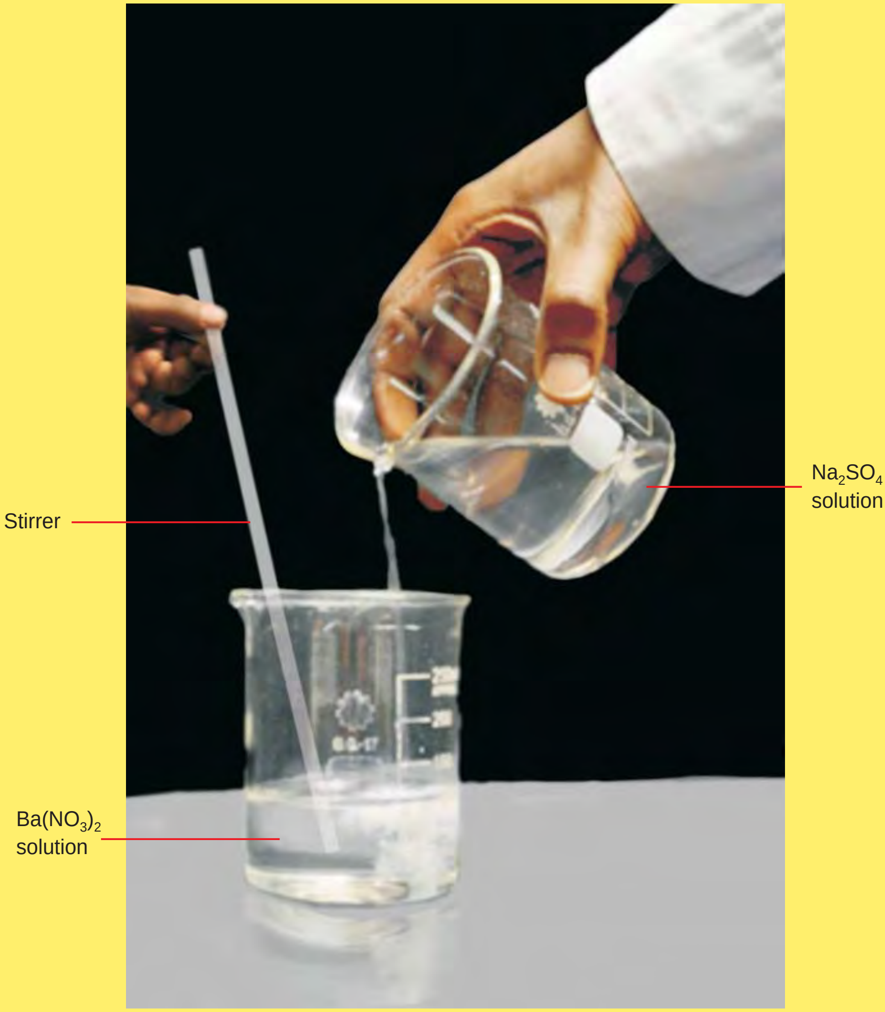


Figure 4.10 The double displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.

Double displacement reaction is a reaction in which two compounds react together to form two new compounds by exchange of the positive and negative ions of each reactant. Such a reaction is also known as double replacement reaction or metathesis.

This type of reaction can be written in the following general form of equation.

$$AB + CD \rightarrow AD + CB$$

Examples

- The two soluble compounds $AgNO_3$ and NaCl react to produce an insoluble precipitate of AgCl and a soluble $NaNO_3$ solution.

$$\underset{}{AgNO_3} + NaCl \rightarrow \underset{\text{Insoluble}}{AgCl\downarrow} + \underset{\text{Soluble}}{NaNO_3}$$

- When aqueous solutions of $BaCl_2$ and Na_2SO_4 react, a precipitate of $BaSO_4$ is formed.

$$BaCl_2 + Na_2SO_4 \rightarrow \underset{\text{Insoluble}}{BaSO_4\downarrow} + \underset{\text{Soluble}}{2NaCl}$$

Exercise 4.7

Give appropriate answers for the following questions.

1. What type of reaction does usually take place in each of the following reactions?
   a a metal reacting with water.
   b a metal reacting with a non-metal.
   c an acid reacting with a metal hydroxide.
   d heating of a metal hydrogen carbonate.
2. Classify the following reactions as combination, decomposition, single or double displacement reactions.
   a $FeO + C \rightarrow Fe + CO$
   b $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
   c $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
   d $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$
   e $2Na_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6NaOH$
   f $CuSO_4.5H_2O \rightarrow CuSO_4 + 5H_2O$

3. Complete and balance the following equations. If the reaction does not take place, write "No Reaction"

   a $Mg + N_2 \xrightarrow{\Delta}$

   b $Na_2CO_3 \xrightarrow{\Delta}$

   c $BaCO_3 + HNO_3 \rightarrow$

   d $Zn + H_2SO_4 \rightarrow$

   e $FeCO_3 \xrightarrow{\Delta}$

   f $H_2CO_3 + NaOH \rightarrow$

4.6 STOICHIOMETRY

Competencies

By the end of this section, you will be able to:

- deduce mole ratios from balanced chemical equations;
- solve mass-mass problems based on the given chemical equation;
- define molar volume;
- state Avogadro's principle;
- solve volume-volume problems based on the given chemical equation;
- solve mass-volume problems based on the given chemical equation;
- define limiting and excess reactants;
- determine limiting and excess reactants of a given chemical reaction;
- show that the amount of product formed in a chemical reaction is based on the limiting reactant;
- define the term theoretical yield, actual yield and percentage yield; and
- calculate the percentage yield of a chemical reaction from given information.

Activity 4.9

Form a group and discuss the following concepts:

a A bicycle mechanic has 10 frames (body parts) and 16 wheels in the shop. How many complete bicycles can he assemble using these parts? Which parts of the bicycle are left over?

b Based on your concULsion in (a), do you think that the masses of reactants are always completely converted to products in a chemical reaction?

Present your conclusion to the class.

The quantitative relationship between reactants and products in a balanced chemical equation is known as stoichiometry. In other words, stoichiometry is the study of the amount or ratio of moles, mass, energy and volumes (for gases) of reactants and products. Stoichiometric calculations are based on the following two major principles.

i) The composition of any substance in the chemical equation should be expressed by a definite formula.

ii) The law of conservation of mass must be obeyed (the mass of reactants equals the mass of products).

4.6.1 Molar Ratios in Balanced Chemical Equation

From a balanced chemical equation, it is possible to determine the:

- number of moles of each reactant and product; and
- relative mass of each of the reactants and products

For example, in the reaction of hydrogen with oxygen to produce water, 2 moles of H_2 combines with 1 mole of O_2 to yield 2 moles of H_2O. The equation also tells us 4 g of hydrogen reacts with 32 g of oxygen to produce 36 g of water. This can be further interpreted as follows:

$$\begin{array}{rclcll}
 & & 2H_2(g) & + & O_2(g) & \rightarrow \quad 2H_2O(l) \\
\text{Mole} & \rightarrow & 2 \text{ mole} & & 1 \text{ mole} & \quad\;\; 2 \text{ mole} \\
\text{Molecule} & \rightarrow & 2 \text{ molecule} & & 1 \text{ molecule} & \quad\;\; 2 \text{ molecule} \\
\text{Mass} & \rightarrow & \multicolumn{3}{l}{\underbrace{4 \text{ g} \qquad\qquad\quad 32 \text{ g}}_{36 \text{ g reactant}}} & \quad\;\; \underbrace{2 \times 18 \text{ g}}_{36 \text{ g product}}
\end{array}$$

Calculations based on chemical equations (*stoichiometric problems*) are classified into mass-mass problem, volume-volume problems and mass-volume problems.

4.6.2 Mass–Mass Relationships

In mass-mass problems, the mass of one substance is given, and the mass of the second substance is determined from the same reaction. There are two methods for solving such types of problems:

i) Mass-ratio method

ii) Mole-ratio method

Let us see each method by using the necessary steps.

i) The mass-ratio method

In this type of stoichiometric calculation, the mass of one substance is determined from the given mass of the other substance using the following steps.

Step 1: Write the balanced chemical equation.

Step 2: Place the given mass above the corresponding formula, and x above the formula of the substance whose mass is to be determined.

Step 3: Write the total molar mass of the substances below the formula of each substance. (*Total molar mass is the molar mass of the substance multiplied by its coefficient*).

Step 4: Set up the proportion.

Step 5: Solve for the unknown mass, x.

Example 1

How many grams of calcium chloride are formed when 15 g of calcium metal reacts with hydrochloric acid?

Solution:

Step 1: $Ca + 2HCl \rightarrow CaCl_2 + H_2$

Step 2: $\overset{15\ g}{Ca} + 2HCl \rightarrow \overset{x}{CaCl_2} + H_2$

Step 3: $\underset{40\ g}{\overset{15\ g}{Ca}} + 2HCl \rightarrow \underset{111 g}{\overset{x}{CaCl_2}} + H_2$

Step 4: $\dfrac{15\ g}{40\ g} = \dfrac{x}{111\ g}$

Step 5: $x = 41.63$ g

Therefore, 41.63 g of $CaCl_2$ is produced.

Example 2

How many grams of oxygen are produced by the decomposition of 145 grams of potassium chlorate?

Solution:

Step 1: $2KClO_3 \rightarrow 2KCl + 3O_2$

Step 2: $\overset{145\,g}{2\,KClO_3} \rightarrow 2KCl + \overset{x}{3\,O_2}$

Step 3: $\underset{(2\times122.5)\,g}{\overset{145\,g}{2\,KClO_3}} \rightarrow 2KCl + \underset{(3\times32)\,g}{\overset{x}{3\,O_2}}$

Step 4: $\dfrac{145\,g}{245\,g} = \dfrac{x}{96\,g}$

Step 5: $x = 56.8\ g$

ii) The mole-ratio method

The mole ratio is the ratio between the numbers of moles of any two substances in a given reaction. In this method, the given mass is converted into moles, and the number of moles for the required substance is calculated. If needed, convert the obtained moles back to mass.

Follow the steps given below to solve problems of mass-mass relationships by the mole ratio method:

Step 1: Write the balanced chemical equation.

Step 2: Convert the given mass to moles and write the obtained moles and the required quantity, *x*, above the formulas of the respective substances.

Step 3: Place the coefficients as the number of moles under the formula of each substance involved.

Step 4: Set up the proportion.

Step 5: Solve for the unknown value, *x*; and convert the moles obtained into mass.

Example 3

How many grams of sodium metal are needed to react with 10.0 g of water?

Solution:

Step 1: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Step 2: moles of $H_2O = \dfrac{\text{given mass}}{\text{molar mass}} = \dfrac{10.0\,g}{18\,g/mol} = 0.56\ mol$

$$\overset{x}{2\,Na} + \overset{0.56\,mol}{2H_2O} \rightarrow 2NaOH + H_2$$

Step 3: $\underset{2\,mol}{\overset{x}{2Na}} + \underset{2\,mol}{\overset{0.56\,mol}{2H_2O}} \rightarrow 2NaOH + H_2$

Step 4: $\dfrac{x}{2\,mol} = \dfrac{0.56\,mol}{2\,mol}$

Step 5: $x = 0.56$ mol of Na

Now, convert 0.56 mole of Na to grams

mass of Na = mole × molar mass

= 0.56 mol × 23 g/mol

= 12.88 g

Therefore, 12.88 g of sodium metal is needed to react with 10 g of water.

Example 4

What mass of nitrogen dioxide is produced by the decomposition of 182 g of magnesium nitrate?

Solution:

Step 1: $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

Step 2: moles of $Mg(NO_3)_2 = \dfrac{182\,g}{148\,g/mol} = 1.23$ mol

$$\overset{1.23\,mol}{2Mg(NO_3)_2} \rightarrow 2MgO + \overset{x}{4NO_2} + O_2$$

Step 3: $\underset{2\,mol}{\overset{1.23\,mol}{2Mg(NO_3)_2}} \rightarrow 2MgO + \underset{4\,mol}{\overset{x}{4NO_2}} + O_2$

Step 4: $\dfrac{1.23\,mol}{2\,mol} = \dfrac{x}{4\,mol}$

Step 5: $x = 2.46$ moles of NO_2

Mass of NO_2 = 2.46 mol × 46 g/mol = 113.2 g

Therefore, 113.2 g of NO_2 is produced.

Exercise 4.8

1. How many grams of $CaCO_3$ are needed to react with 15.2 g of HCl in according to the following equation?

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

2. How many grams of NaOH are needed to neutralize 50 grams of H_2SO_4?
3. Calculate the mass of $CaCl_2$ formed when 5 moles of chlorine reacts with calcium metal.
4. How many moles of H_2O are required to produce 4.5 moles of HNO_3 according to the following reaction:

$$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

5. In the decomposition of $KClO_3$, how many moles of KCl are formed in the reaction that produces 0.05 moles of O_2?
6. How many moles of CaO are needed to react with excess water to produce 370 g of calcium hydroxide?

4.6.3 Volume-Volume Relationships

In reactions involving gases, the volume of gases can be determined on the principle that 1 mole of any gas occupies a volume of 22.4 litres at STP (*standard temperature* and *pressure*, STP, the *temperature is* 0°C and *the pressure is 1 atm*). It is also known that 22.4 L of any gas weighs exactly its molecular mass at STP. This volume, 22.4 litres, of a gas is known as the molar volume.

At STP, 1 mole of any gas = 22.4 L = gram volume mass of the gas

The relationship between the volume of a gas and its number of molecules was explained by Avogadro. Avogadro's law states that equal volumes of different gases, under the same conditions of temperature and pressure, contain equal number of molecules. This law can also be stated as the volume of a gas is proportional to the number of molecules (moles) of the gas at STP.

Mathematically, $V \propto n$; where V is the volume and n is the number of moles.

In volume-volume problems, the volume of one substance is given and the volume of the other substance is calculated. All the steps to solve volume-volume problems are shown by the following example.

Example 5

What volume of oxygen will react with carbon monoxide to produce 20 litres of carbon dioxide at STP?

Solution:

Step 1: Write the balanced chemical equation

$$2CO + O_2 \longrightarrow 2CO_2$$

Step 2: Place the given volume and the required volume, x, above the corresponding formulas.

$$2CO + \overset{x}{O_2} \longrightarrow \overset{20\text{ L}}{2CO_2}$$

Step 3: Write the total molar volume (22.4 L multiplied by any coefficient) below the formulas.

$$2CO + \underset{22.4\text{ L}}{\overset{x}{O_2}} \longrightarrow \underset{2(22.4\text{ L})}{\overset{20\text{ L}}{2CO_2}}$$

Step 4: Set up the proportion.

$$\frac{x}{22.4\text{ L}} = \frac{20\text{ L}}{44.8\text{ L}}$$

Step 5: Solve for the unknown volume, x.

x = 10 L of O_2 are needed.

Exercise 4.9

1. What volume of nitrogen reacts with 33.6 litres of oxygen to produce nitrogen dioxide?
2. How many litres of sulphur trioxide are formed when 4800 cm^3 of sulphur dioxide is burned in air?
3. How many litres of ammonia are required to react with 145 litres of oxygen according to the following reaction?

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

4. Calculate the volume of oxygen produced in the decomposition of 5 moles of $KClO_3$ at STP?
5. How many moles of water vapour are formed when 10 litres of butane gas, C_4H_{10} is burned in oxygen at STP?

4.6.4 Mass–Volume Relationships

In mass-volume problems, either the mass of one substance is given and the volume of the other is required or the volume of one substance is given and the mass of the other one is required. The steps to solve such type of problems are the same as the previous steps except putting the masses on one side and the volumes on the other side of the equality sign.

Example 6

How many grams of calcium carbonate are decomposed to produce 11.2 L of carbon dioxide at STP?

Solution:

Step 1: $CaCO_3 \rightarrow CaO + CO_2$

Step 2: $\overset{x}{CaCO_3} \rightarrow CaO + \overset{11.2\,L}{CO_2}$

Step 3: $\underset{100\,g}{\overset{x}{CaCO_3}} \rightarrow CaO + \underset{22.4\,L}{\overset{11.2\,L}{CO_2}}$

Step 4: $\dfrac{x}{100\,g} = \dfrac{11.2\,L}{22.4\,L}$

Step 5: $x = 50$ g of $CaCO_3$ is decomposed.

Example 7

How many litres of oxygen at STP react with 72 g of aluminum to produce aluminum oxide?

Solution:

Step 1: $4Al + 3O_2 \rightarrow 2Al_2O_3$

Step 2: $\overset{72\ g}{4Al} + \overset{x}{3O_2} \rightarrow 2Al_2O_3$

Step 3: $\underset{(4\times27)\ g}{\overset{72\ g}{4Al}} + \underset{(3\times22.4)\ L}{\overset{x}{3O_2}} \rightarrow 2Al_2O_3$

Step 4: $\dfrac{72\ g}{108\ g} = \dfrac{x}{67.2\ L}$

Step 5: $x = 44.8$ L of O_2

Hence, 44.8 litres of oxygen is required at STP to react with 72 g of aluminium.

Exercise 4.10

1. How many litres of oxygen are required to react with 23 g of methane according to the following equation?

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

2. What mass of aluminium would be completely oxidized by 44.8 L of oxygen to produce Al_2O_3 at STP?
3. Calculate the mass of calcium carbide that is needed to produce 100 cm^3 of acetylene according to the following equation.

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

4. How many millilitres of sulphur dioxide are formed when 12.5 g of iron sulphide ore (pyrite) reacts with oxygen according to the equation at STP?

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

4.6.5 Limiting and Excess Reactants

When all the reactants are completely consumed in a chemical reaction, then such reactants are said to be in stoichiometric proportions. But, practically these types of chemical reactions do not always occur. In many cases, an excess of one or more reactants is encountered in the reaction and the other reactant is completely converted into products. Thus, the reactant that is completely consumed in the reaction is known as the limiting reactant, because it limits or determines the amount of products that can be formed.

For example, consider the following reaction:

$$\underset{1\ \text{mole}}{H_2} + \underset{1\ \text{mole}}{Cl_2} \rightarrow \underset{2\ \text{moles}}{2HCl}$$

According to the equation, 1 mole of H_2 reacts with 1 mole of chlorine to produce 2 mole of HCl. Thus, all the reactants are completely consumed and only products appear. However if 1 mole of H_2 reacts with 1.5 mole of Cl_2, there is insufficient H_2 to react with all of the Cl_2. Therefore, Cl_2 will be in excess and H_2 will be the limiting reactant. Only 2 moles of HCl are formed and at the end of the reaction 0.5 mole of Cl_2 remains unreacted.

Example 8

How much ammonia is produced if 10 g of hydrogen reacts with 18 g of nitrogen?

Solution:

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

First determine the number of moles;

$$\text{Moles of } H_2 = \frac{10\text{ g}}{2\text{ g/mol}} = 5\text{ mol}$$

$$\text{Moles of } N_2 = \frac{18\text{ g}}{28\text{ g/mol}} = 0.64\text{ mol}$$

Now, calculate the number of moles or masses of the product that would be formed by each reactant.

The reactant that gives the smallest amount of product is the limiting reactant.

i. **Using the quantity of H_2**

$$\underset{3\text{ mol}}{\overset{5\text{ mol}}{3H_2}} + N_2 \rightarrow \underset{2\text{ mol}}{\overset{x}{2NH_3}}$$

$$\frac{5\text{ mol}}{3\text{ mol}} = \frac{x}{2\text{ mol}}$$

$x = 3.33$ mol NH_3

$$\text{Mass of } NH_3 = 3.33\text{ mol} \times 17\text{ g/mol}$$

$$= 56.6\text{ g}$$

ii. **Using the quantity of N_2**

$$3H_2 + \underset{1\text{ mol}}{\overset{0.64\text{ mol}}{N_2}} \rightarrow \underset{2\text{ mol}}{\overset{x}{2NH_3}}$$

$$\frac{0.64\text{ mol}}{1\text{ mol}} = \frac{x}{2\text{ mol}}$$

$x = 1.28$ mol NH_3

$$\text{Mass of } NH_3 = 1.28\text{ mol} \times 17\text{ g/mol}$$

$$= 21.8\text{ g}$$

Therefore, the limiting reactant is nitrogen, because it gives less amount of NH_3, *i.e.*, 21.8 g NH_3. In the reaction, 0.64 mole (18 g) of N_2 is consumed. Hydrogen is in excess. The amount of hydrogen consumed will be:

$$\underset{6g}{\overset{x}{3H_2}} + \underset{28\,g}{\overset{18g}{N_2}} \rightarrow 2NH_3 \qquad \text{or} \qquad \underset{3\,\text{moles}}{\overset{x}{3H_2}} + \underset{1\,\text{mole}}{\overset{0.64\,\text{mol}}{N_2}} \rightarrow 2NH_3$$

$$\frac{x}{6} = \frac{18}{28} \qquad\qquad \frac{x}{3\,\text{mol}} = \frac{0.64\,\text{mol}}{1\,\text{mol}}$$

x = 3.86 g of H_2 $\qquad$ x = 1.92 mol of H_2

Therefore, 3.86 g or 1.92 moles of H_2 is used in the reaction, and 6.14 g or 3.08 moles of H_2 is left unreacted.

Example 9

In the chemistry laboratory, a student performed a displacement reaction by adding 9.5 g of zinc into 9.5 g of HCl in a beaker. What weight of $ZnCl_2$ will be produced?

Solution:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

$$\text{Moles of Zn} = \frac{9.5\,g}{65\,g/mol} = 0.15\ \text{mol}$$

$$\text{Moles of HCl} = \frac{9.5\,g}{36.5\,g/mol} = 0.26\ \text{mol}$$

Even though the given masses of the two reactants are the same, they are not mixed in equimolar ratio as shown above. Thus, the limiting reactant must be determined first.

i. **Using the quantity of Zn**

$$\underset{1\,\text{mol}}{\overset{0.15\,\text{mol}}{Zn}} + 2HCl \rightarrow \underset{1\,\text{mol}}{\overset{x}{ZnCl_2}} + H_2$$

$$\frac{0.15\,\text{mol}}{1\,\text{mol}} = \frac{x}{1\,\text{mol}}$$

x = 0.15 mol $ZnCl_2$

ii. **Using the quantity of HCl**

$$Zn + \underset{2\,\text{mol}}{\overset{0.26\,\text{mol}}{2HCl}} \rightarrow \underset{1\,\text{mol}}{\overset{x}{ZnCl_2}} + H_2$$

$$\frac{0.26\,\text{mol}}{2\,\text{mol}} = \frac{x}{1\,\text{mol}}$$

x = 0.13 mol $ZnCl_2$

Hence, the limiting reactant is HCl.

$$\text{Mass of } ZnCl_2 = 0.13\ \text{mol} \times 136\ g/mol = 17.68\ g\ ZnCl_2$$

Exercise 4.11

1. If 6.5 g of zinc reacts with 5.0 g of HCl, according to the following reaction.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

   a Which substance is the limiting reactant?

   b How many grams of the reactant remains unreacted?

   c How many grams of hydrogen would be produced?

2. What mass of Na_2SO_4 is produced if 49 g of H_2SO_4 reacts with 80 g of NaOH?

3. If 20 g of $CaCO_3$ and 25 g of HCl are mixed, what mass of CO_2 is produced?

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

4. If 3 moles of calcium reacts with 3 moles of oxygen, then

   a Which substance is the limiting reactant?

   b How many moles of calcium oxide are formed?

5. For the reaction:

$$2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$$

   How many grams of hydrogen are produced if 0.8 mole of aluminium reacts with 1.0 mole of sulphuric acid?

4.6.6 Theoretical, Actual and Percentage Yields

The measured amount of product obtained in any chemical reaction is known as the actual yield. The theoretical yield is the calculated amount of product that would be obtained if the reaction proceeds completely. The actual yield (*experimentally determined yield*) of a product is usually less than the theoretical yield (*calculated yield*).

The percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100.

$$\text{Percentage yield} = \frac{\text{Acutal yield}}{\text{Theoretical yield}} \times 100$$

Example 10

25 grams of methane gas (CH_4) burns in oxygen according to the following reaction:

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

What is the percentage yield if 60.3 grams of carbon dioxide is produced?

Solution:

The actual yield is 60.3 g of CO_2.

Determine the theoretical yield using mass-mass relationship

$$\underset{16\,g}{\overset{25\,g}{CH_4}} + 2O_2 \rightarrow \underset{44\,g}{\overset{x}{CO_2}} + 2H_2O$$

$$\frac{25\,g}{16\,g} = \frac{x}{44}$$

x = 68.75 g of CO_2 (*theoretical yield*)

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

$$= \frac{60.3\,g}{68.75\,g} \times 100\% = 87.7\ \%$$

Exercise 4.12

1. When 20 g of sulphur dioxide reacts with oxygen, 23 g of sulphur trioxide is formed. What is the percentage yield?
2. When 14.5 g of SO_2 reacts with 21 g of O_2, what will be the theoretical yield and percentage yield of the reaction if the actual yield is 12 g?
3. In the reaction:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

When 52.7 g of octane (C_8H_{18}) burns in oxygen, the percentage yield of carbon dioxide is 82.5%. What is the actual yield in grams?

4.7 OXIDATION AND REDUCTION REACTIONS

Competencies

By the end of this section, you will be able to:

- define redox reactions;
- define the terms oxidation and reduction in terms of electron transfer;
- define oxidation number (oxidation state),
- state oxidation number rules,
- determine the oxidation number of an element in a given formula;
- describe the oxidizing and reducing agents;
- analyze a given redox reaction by specifying the substance reduced and the substance oxidized, and also the oxidizing and reducing agents; and
- Distinguish between redox and non-redox reactions.

Activity 4.10

Form a group and discuss the following phenomenon:

When you dry your meal dishes with a towel, the towel can be termed as a drying agent and the dish as wetting agent. What happens to the towel and the dish, in terms of getting wet and dry, after cleaning? Relate this idea with oxidizing agent and reducing agent, oxidation and reduction of substances. Present your conclusion to the class.

In our day to day activity, we are familiar with the chemical processes like rusting of iron, burning of substances, breathing of air, digestion of food and so on. All such types of processes or reactions are known as oxidation and reduction or redox reactions.

4.7.1 Oxidation-Reduction

Can oxidation take place without reduction?

Oxidation: is the process in which a substance loses electrons in a chemical reaction.

For example, in the reaction

$$2Na + Cl_2 \rightarrow 2NaCl$$

Each sodium atom has lost one electron and has turned to a sodium ion. Hence, sodium is oxidized.

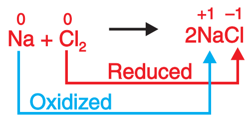
$$Na + 1e^- \rightarrow Na^+$$

Reduction: is the process in which a substance gains electrons in a chemical reaction.

For example, in the above reaction each chlorine atom has gained an electron and has changed to chloride ion. Thus, chlorine is reduced;

$$Cl_2 + 2e^- \rightarrow Cl^-$$

The processes of oxidation and reduction always occur simultaneously because if one substance loses electrons, the other substance must gain these electrons. Since the process of oxidation and reduction involves the transfer of electrons, it also results in the changes of oxidation number. Thus, oxidation and reduction can also be defined in terms of oxidation number. Oxidation is an increase in the oxidation number of an element and reduction is a decrease in the oxidation number. For example, in the reaction,



The oxidation number of sodium is increased from 0 to +1 and thus sodium is oxidized. The oxidation number of chlorine is decreased from 0 to –1, and therefore chlorine is reduced.

4.7.2 Oxidation Number or Oxidation State

Oxidation number or oxidation state is the number of electrons that an atom appears to have gained or lost when it is combined with other atoms.

Rules for Assigning Oxidation Numbers

Rule 1: The oxidation number of all elements in free state is zero. This rule is also applied for diatomic or polyatomic elements.

Example: The oxidation number of Na = 0, Cu = 0, Cl in Cl_2 = 0, O in O_3 = 0, S in S_8 = 0.

Rule 2: The oxidation number of a monatomic ion is equal to the charge on the ion.

Example: $Na^+ = +1$, $Mg^{2+} = +2$, $S^{2-} = -2$.

Rule 3: The oxidation number of oxygen in a compound is usually – 2 except in the following cases:

Exceptions

The oxidation number of oxygen in:

i) peroxides is –1. **Example:** Na_2O_2

ii) superoxides is –1/2. **Example:** KO_2

iii) oxygen diflouride is +2. **Example:** OF_2

Rule 4: The oxidation number of hydrogen in its entire compounds is +1 except in metal hydrides, (like NaH, CaH_2 and AlH_3), where its oxidation number is –1.

Rule 5: The sum of the oxidation number of all the atoms in a neutral compound is zero.

Example: $\overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4}$ ((+2) + (+6) + (–8) = 0)

Rule 6: In a polyatomic ion, the sum of the oxidation numbers of the constituent atoms equals the charge on the ion.

Example: $(\overset{+6}{S}\overset{-2}{O_4})^{-2}$ ((+6) + (–8) = –2)

Rule 7: Elements of group IA have +1 and group IIA have +2 oxidation states in all of their compounds.

Rule 8: In a compound, the more electronegative element is assigned a negative oxidation number, and the less electronegative element is assigned a positive oxidation number.

Example: $\overset{+3}{N}\overset{-1}{Cl_3}$ (*chlorine is more electronegative than nitrogen*)

Example 1

What is the oxidation number of chromium in $Na_2Cr_2O_7$?

Solution:

The oxidation number of O is –2 (*Rule 3*)

The oxidation number of Na is +1 (*Rule 7*)

Let the oxidation number of Cr be *x*.

$\overset{+1}{Na_2}\,\overset{x}{Cr_2}\,\overset{-2}{O_7}$

Since the sum of the oxidation numbers of Na, Cr, and O in $Na_2Cr_2O_7$ is 0 (*Rule 5*)

Then, $\overset{+1}{Na_2}\ \overset{x}{Cr_2}\ \overset{-2}{O_7}$

$$(1 \times 2) + (x \times 2) + (-2 \times 7) = 0$$

$$2 + 2x - 14 = 0$$

$$x = +6$$

Therefore, the oxidation number of Cr in $Na_2Cr_2O_7$ is +6.

Example 2

What is the oxidation number of manganese in MnO_4^-?

Solution :

Let the oxidation number of Mn be x.

$$\left(\overset{x}{Mn}\ \overset{-2}{O_4}\right)^-$$

The sum of the oxidation numbers of Mn and O in MnO_4^- is –1 (*Rule 6*)

$$x + (-2 \times 4) = -1$$

$$x - 8 = -1$$

$$x = +7$$

Therefore, the oxidation number of Mn in MnO_4^- is +7.

Example 3

Determine the oxidation number of phosphorus in $Ca(H_2PO_4)_2$.

Solution:

The oxidation number of Ca is +2 (*Rule 7*).

Let, the oxidation number of P be x.

$$\overset{+2}{Ca}\left(\overset{+1}{H_2}\ \overset{x\ -2}{PO_4}\right)_2$$

$$+2 + (4 \times (+1)) + (2 \times x) + (8 \times (-2)) = 0$$

$$2 + 4 + 2x - 16 = 0$$

$$2x - 10 = 0 \quad \text{or} \quad x = +5$$

Hence, the oxidation number of P in $Ca(H_2PO_4)_2$ is +5.

Exercise 4.13

1. Determine the oxidation number of the specified element in each of the following:

   a C in $H_2C_2O_4$
   b N in NH_4F
   c S in $Na_2S_4O_6$
   d P in $Ca_3(PO_4)_2$
   e H in AlH_3
   f N in NH_4HCO_3
   g Fe in $K_4[Fe(CN)_6]$

2. Determine the oxidation number of the specified element in each of the following.

   a S in S^{-2}
   b Cl in ClO_3^-
   c N in NH_4^+
   d P in PO_4^{-3}
   e Cr in $Cr_2O_7^{-2}$
   f S in $S_2O_8^{-2}$

3. Determine whether the following processes are oxidation or reduction reactions:

   a $Cu^2 + 2e^- \rightarrow Cu$
   b $K \rightarrow K^+ + e^-$
   c $O + 2e^- \rightarrow O^{2-}$
   d $S^{2-} \rightarrow S + 2e^-$
   e $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   f $N + 3e^- \rightarrow N^{3-}$

4.7.3 Oxidizing and Reducing Agents

In a redox reaction, the substance that causes another substance to get oxidized, but itself gets reduced, is known as an oxidizing agent, or oxidant. In the same manner, the substance that causes another substance to get reduced, but itself oxidized, is referred to as a reducing agent or reductant.

- Oxidizing agents are substances that:
  - are reduced (*gain electrons*)
  - contain elements whose oxidation number decreases
- Reducing agents are substances that:
  - are oxidized (*lose electrons*)
  - contain elements whose oxidation number increases

Tests for an oxidizing agent are accomplished by mixing it with a substance which is easily oxidized to give a visible colour change when the reaction takes place.

For example,

i) Permanganate ion (MnO_4^-) in acidic solution changes colour from purple to colourless.

$$MnO_4^- \rightarrow Mn^{2+}$$

ii) Dichromate in acidic solution changes colour from orange to green.

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

Other common oxidizing agents are chlorine, potassium chromate, sodium chlorate and manganese (IV) oxide.

Similarly, certain reducing agents undergo a visible colour change with a substance which is easily reduced.

For example,

i) A moist starch solution changes potassium iodide paper to blue-black to show that iodine is formed, $2I^- \rightarrow I_2$. That is potassium iodide is a reducing agent.

ii) Hydrogen sulphide bubbled through a solution of an oxidizing agent forms a yellow precipitate, $S^{2-} \rightarrow S$. That is H_2S is a reducing agent.

Other common reducing agents are carbon, carbon monoxide, sodium thiosulphate, sodium sulphite and iron (II) salts.

The oxidizing or reducing ability of substances depend on many factors. Some of these are:

- **Electronegativity:** Elements with high electronegativity such as F_2, O_2, N_2 and Cl_2 are good oxidizing agents. Elements with low electronegativity for example, metallic elements like Na, K, Mg and Al are good reducing agents.
- **Oxidation states:** In a compound or ion, if one of its elements is in a higher oxidation state, then it is an oxidizing agent. Similarly, if an element of a compound or ion is in a lower oxidation state, then it is a reducing agent..

Examples

$K\overset{+7}{Mn}O_4, Na\overset{+7}{Cl}O_4, K_2\overset{+6}{Cr_2}O_7$ ... are oxidizing agents

$\overset{+2}{Fe}S, \overset{+2}{C}O, Na_2\overset{+4}{S}O_3$ ... are reducing agents

4.7.4 Analysing Redox Reactions

Oxidation and reduction reactions are called Redox reactions. Oxidation and reduction reaction take place simultaneously in a given reaction.

Activity 4.11

Form a group and discuss the following:

1. Why a reducing agent undergoes oxidation?
2. Why must every redox reaction involve both an oxidizing agent and a reducing agent?

Present your discussion to the class.

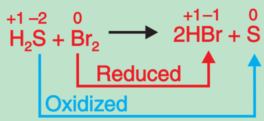
Example 4

Identify the oxidizing agent, reducing agent, the substance oxidized and reduced in the following reaction.

$$H_2S + Br_2 \rightarrow 2HBr + S$$

Solution:

Let us assign oxidation number to all the elements of the reactants and products.



In the reaction, the S atom in H_2S increases its oxidation number from –2 to 0. Hence, S is oxidized and H_2S is a reducing agent. The oxidation number of Br is decreased from 0 to –1. Thus, Br_2 is reduced and is an oxidizing agent.

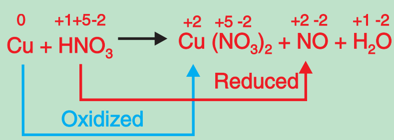
Example 5

Write the oxidizing and reducing agents for the reaction given below:

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$$

Solution:

Write the oxidation numbers of each element and identify the substances which undergo a change in oxidization number.



Therefore, copper is a reducing agent, and HNO_3 is an oxidizing agent.

Non-redox Reactions

So far we have discussed oxidation and reduction reaction or redox reactions. However, there are also reactions in which oxidation and reduction do not occur and such types of reactions are known as non-redox reactions.

In non-redox reactions, no electrons are exchanged between the reacting substances. Therefore, the oxidation numbers of the atoms do not change in the reaction. Usually such types of reactions involve the exchange of positive and negative ions. Most of the double displacement reactions and acid-base reactions are not oxidation-reduction reactions.

Examples

$$\overset{+2\ -1}{BaCl_2} + \overset{+1\ +6\ -2}{Na_2SO_4} \rightarrow \overset{+2\ +6\ -2}{BaSO_4}\downarrow + \overset{+1\ -1}{2NaCl}$$

$$\overset{+1\ -1}{HCl} + \overset{+1\ -2\ +1}{NaOH} \rightarrow \overset{+1\ -1}{NaCl} + \overset{+1\ -2}{H_2O}$$

Exercise 4.14

1. In each of the following equations, identify the substance oxidized, the substance reduced, the oxidizing agent and reducing agent.
   a $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
   b $CuO + H_2 \rightarrow Cu + H_2O$
   c $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$
   d $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
   e $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   f $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$
   g $SO_4^{2-} + I^- + H^+ \rightarrow H_2S + I_2 + H_2O$
   h $MnO_4^- + H^+ \rightarrow Mn^{2+} + H_2O$
2. In each of the following, is the change indicated oxidation, reduction or no change at all?
   a $Cl_2 \rightarrow Cl^-$
   b $MnO_4^- \rightarrow MnO_2$
   c $CrO_2 \rightarrow CrO_4^{2-}$
   d $SO_4^{2-} \rightarrow SO_3^{2-}$
   e $Cr_2O_7^{2-} \rightarrow CrO_4^{2-}$

4.8 RATES OF CHEMICAL REACTIONS AND CHEMICAL EQUILIBRIUM

Competencies

By the end of this section, you will be able to:

- define rate of reaction;
- describe rate of reaction using a graph;
- carry out an experiment to illustrate the relative rate of reactions;
- list the pre-conditions for a chemical reaction to occur;
- explain how collision, activation energy and proper orientation of reactants cause a chemical reaction to occur;
- list factors that affect rate of chemical reaction;
- explain the effects of changes in temperature, concentration or pressure and surface area on the rates of a chemical reaction;
- explain the effect of catalysts on the rates of chemical reaction;
- carry out an activity on how the factors affect the rate of chemical reaction;
- define the terms reversible reaction and irreversible reaction;
- define chemical equilibrium;
- describe the characteristics of chemical equilibrium;
- write the expression for equilibrium constant of a reversible reaction;
- state Le Châtelier's principle; and
- use Le Châtelier's principle to explain the effect of changes in temperature, pressure and concentration of reactants at equilibrium.

Activity 4.12

Discuss each of the following phenomena in groups and present your findings to the class:

1. Why do some reactions take place rapidly and others slowly? Give examples of fast and slow reactions.
2. Does sugar dissolve faster in hot or in cold tea? Why?

4.8.1 Reaction Rate

Experiment 4.9

Measuring Rate of a Reaction

Objective: To measure the rate of the reaction between HCl and $CaCO_3$.

Apparatus: Balance, conical flask, cotton wool, stop watch.

Chemicals: HCl, $CaCO_3$

Procedure:

1. Take 50 mL of dilute hydrochloric acid into a conical flask.
2. Place the conical flask on a laboratory balance.
3. Add 25 g of calcium carbonate (*marble chips*) into the flask. Plug the cotton wool immediately as shown in Figure 4.11. This will help to prevent escape of the acid spray.
4. Read and record the mass of the flask and its content at one minute intervals until the reaction is over. (*Use a stop watch*).

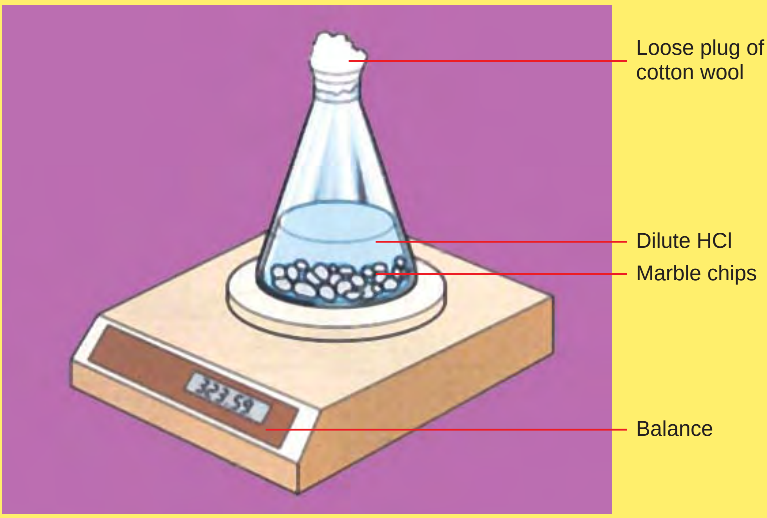


Figure 4.11 Measuring the rate of the reaction between HCl and $CaCO_3$.

Observations and analysis:

1. What happens to the mass during the reaction?
2. Write the balanced chemical equation.

Use the following table to record your results.

Time (min)	0	1	2	3	4	5	6	7
Mass (g)								
Loss in mass (g)								

3. Plot a graph between time (*x*-axis) and loss in mass (*y*-axis), and draw a smooth curve through maximum points.
4. Why the graph is steep in the beginning but horizontal at the end of the reaction?
5. At what time does the reaction stop?

Every chemical reaction proceeds at different rates or speed. Some reactions proceed very slowly and may take a number of days to complete; while others are very rapid, requiring only a few seconds.

The rate of a chemical reaction measures the decrease in concentration of a reactant or the increase in concentration of a product per unit time. This means that the rate of a reaction determines how fast the concentration of a reactant or product changes with time. The rate of a reaction is obtained by determining the concentration of reactants or products during the reaction. Methods for determining the concentration of reactants or products depend on the type of reactions. Some of the methods are:

a. Colour (*changes in colour*)
b. Pressure (*increase or decrease in pressure, particularly in gases*)
c. Volume (*increase or decrease in size, particularly in gases*)
d. Mass (*gain or loss in weight*)
e. Amount of precipitate formed

Generally, the rate of a reaction can be obtained by measuring either one of the above changes in properties of substances and consequently relating to changes in their concentrations during the course of the reaction.

$$\text{Rate of reaction} = \frac{\text{Change in concentration of substance}}{\text{Change in time}} = \frac{\Delta C}{\Delta t}$$

From this expression, it follows that the rate of a reaction is inversely proportional to the time taken by the reaction.

$$\text{Rate} \propto \frac{1}{\text{Time}}$$

Figure 4.12 illustrates the changes of the rate of a chemical reaction with time. A reaction becomes slower as reactants are consumed. The reaction rate curve becomes less steep until it becomes a horizontal straight line. No more reactant is used up at this point.

Note that the rate of a reaction is the slope of the tangent to the curve at any particular time.

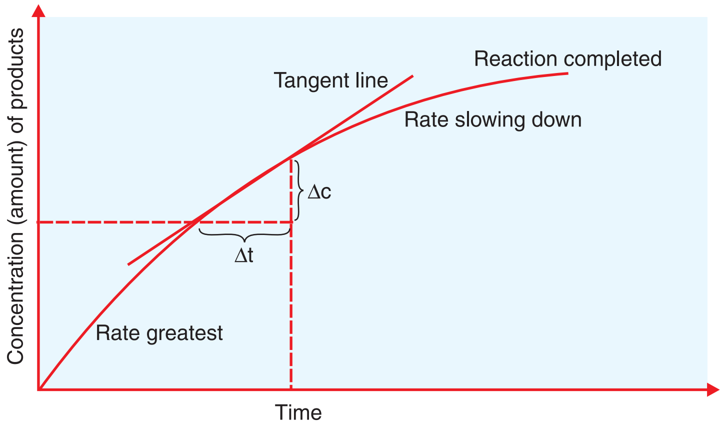


Figure 4.12 The change in concentration of product with time.

Reading Check

When a clean piece of magnesium ribbon is added to excess dilute hydrochloric acid, hydrogen gas is evolved. When a graph of volume versus time is drawn, show that the total volume of the gas evolved can be measured at fixed intervals.

Pre-conditions for a Chemical Reaction

Chemical reactions are usually explained by the collision theory. The assumption of the collision theory is that chemical reactions take place due to the collision between molecules.

1. *Collisions between reactants*

The first precondition for a reaction to occur is the direct contact of the reacting substance with each other. However, all collisions between molecules are not necessarily effective in bringing a reaction.

2. *Activation energy*

If the collisions between the reactant molecules do not have sufficient energy, then no reaction will occur. Therefore, for the reaction to take place collision must always occur with sufficient energy to break the bonds in the reactants and form new bonds in the product. Thus, minimum amount of energy needed for the reaction is known as activation energy.

3. *Proper Orientation*

Collision of molecules with sufficient activation energy will not bring a reaction if the reacting molecules are poorly oriented. Thus, the collision between molecules should have the proper orientation.

Consider the reaction between H_2 and Cl_2 molecules:

$$H_2 + Cl_2 \rightarrow 2HCl$$

For the reaction to occur, first H_2 and Cl_2 molecules must collide with each other. For the collision to be effective, these colliding molecules (H_2 and Cl_2) must have sufficient energy to break the H – H and Cl – Cl bonds and consequently to form new H – Cl bonds.

Unless, the H_2 and Cl_2 molecules are oriented in proper positions, there is no product formed. Therefore, as shown in Figure 4.13(c) the H_2 and Cl_2 molecules rearrange themselves so as to form a new H – Cl molecule.

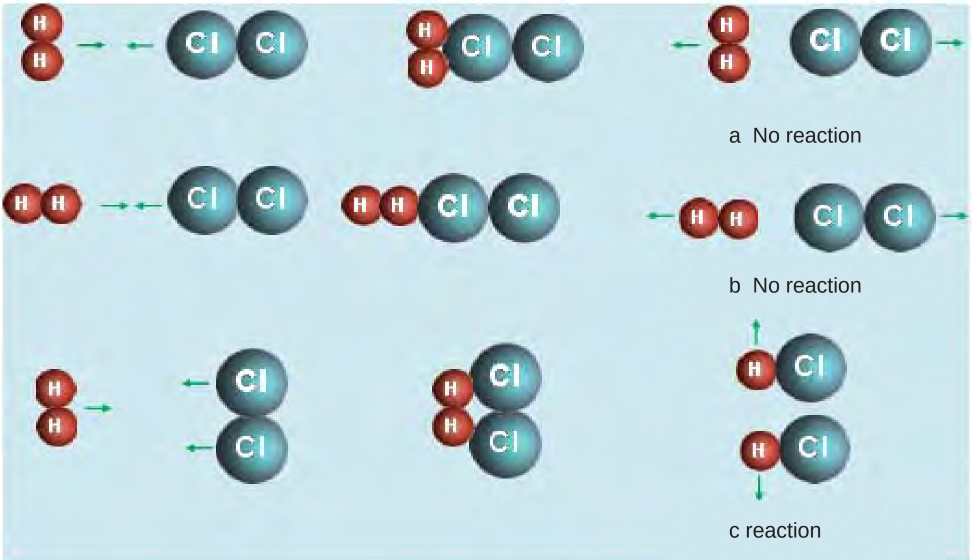


Figure 4.13 Molecular collisions and chemical reactions.

Factors Affecting the Rates of Chemical Reaction

Activity 4.13

Form a group and discuss each of the following:

1. How is the burning of charcoal affected by:
   a increasing the amount of air used
   b adding more charcoal.
2. How can you increase the rate of combustion of a given block of wood ?

Present your findings to the class.

Activity 4.14

Form a group and compare the rate of combustion of the following substances:

a paper

b kerosene

c wood charcoal

d copper

Present your findings to the class.

The rate of a chemical reaction depends on nature of reactants, temperature, and concentration of reactants, surface area and catalysts.

The collision theory assumes that the rate of a reaction depends on the number (*frequency*) of collisions of particles and those factors which affect the frequency of collisions will also affect the rate of a reaction. Now let us discuss each factor.

1. *Nature of the reactants*

The rate of a reaction is influenced by the type and nature of the reacting substances. For example, the following reactions have different rates due to the nature of the reactants, Mg, Fe and Cu.

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \quad \text{... (\textit{very fast reaction})}$$

$$Fe + 2HCl \rightarrow FeCl_2 + H_2 \quad \text{... (\textit{slow reaction})}$$

$$Cu + HCl \rightarrow \textit{No reaction}$$

2. *Temperature*

An increase in temperature increases the rate of a reaction. This is because as the temperature increases, the average kinetic energy of the particles increases which in turn increases the number of effective collisions.

In general, for many chemical reactions, the rate of a reaction doubles for every 10°C rise in temperature.

3. *Concentration of reactants*

The number of collisions is proportional to the concentration of reactants. The higher the concentration of the reactants, the more collisions between the reacting particles and thus the higher the rate of the reaction.

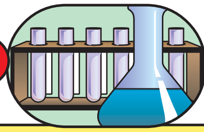
For example, if you heat a piece of steel wool in air (21% oxygen by volume) it burns slowly. But in pure oxygen (100% oxygen by volume) it bursts in to a dazzling white flame. This indicates that the rate of burning increases as the concentration of oxygen is higher.

Activity 4.15

Form a group and discuss the following idea. Present your discussion to the class.

If a certain reaction is carried out with water as a solvent, what will be the rate of the reaction if more water is added to the reaction vessel? Explain.

Experiment 4.10



Investigating the effect of concentration on reaction rate

Objective: To determine the rate of reaction of magnesium with 0.1M and 5M of sulphuric acid.

Apparatus: Beakers.

Chemicals: H_2SO_4.

Procedure:

a
1. Take 20 mL of 0.1M H_2SO_4 into the first beaker.
2. Add 1 cm long magnesium ribbon into the beaker.
3. Note how fast the reaction occurs.

b
1. Take 20 mL of 5M H_2SO_4 into the second beaker.
2. Add 1cm long magnesium ribbon into the beaker as shown in Figure 4.14.
3. Observe how fast the reaction occurs.

Observations and Analysis:

1. In which of the reactions does the evolution of gas bubble faster? (*a*) or (*b*).
2. Write the balanced chemical equation.
3. What do you conclude from the experiment?

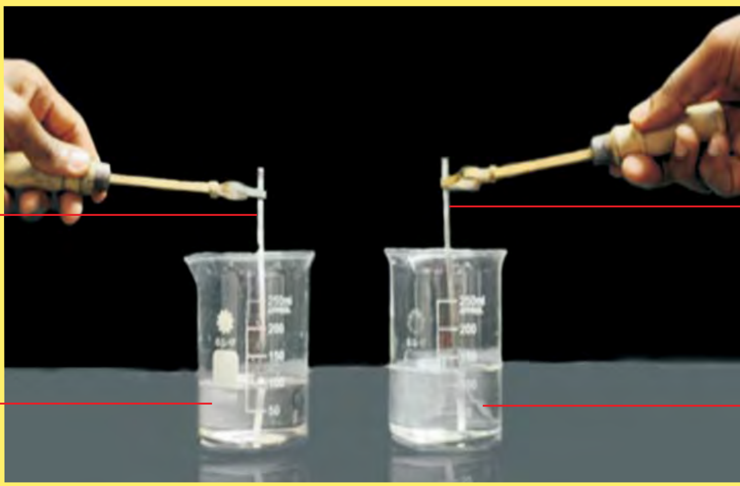


Figure 4.14 The effect of concentration on the rate of a reaction.

4. *Surface area*

When the reactants are in different phases, be it solid, liquid or gas, then the surface area of the substances affect the rate of the reaction. The higher the surface area of reactants, the faster is the rate of the reaction. This is because more contact results in more collisions between each small particle of reactants.

Experiment 4.11

Investigating the effect of surface area on reaction rate

Objective: To determine the rate of reaction of a lump and powdered calcium carbonate with hydrochloric acid.

Apparatus: Beaker, dish, and grinder.

Chemicals: Calcium carbonate and dilute hydrochloric acid.

Procedure:

a
1. Take a 5 g of calcium carbonate and put it in a beaker.
2. Add 100 mL dilute hydrochloric acid into the beaker carefully.
3. Observe how fast the reaction occurs.

b
1. Add 5 g of calcium carbonate into a dish and grind until it becomes powder.
2. Put this powdered calcium carbonate into the beaker as shown in Fig. 4.15.
3. Add 100 mL dilute hydrochloric acid carefully.
4. Observe how fast the reaction occurs.

Observations and Analysis:

1. Which of the reaction is faster? (*a*) or (*b*).
2. What do you conclude from the experiment?

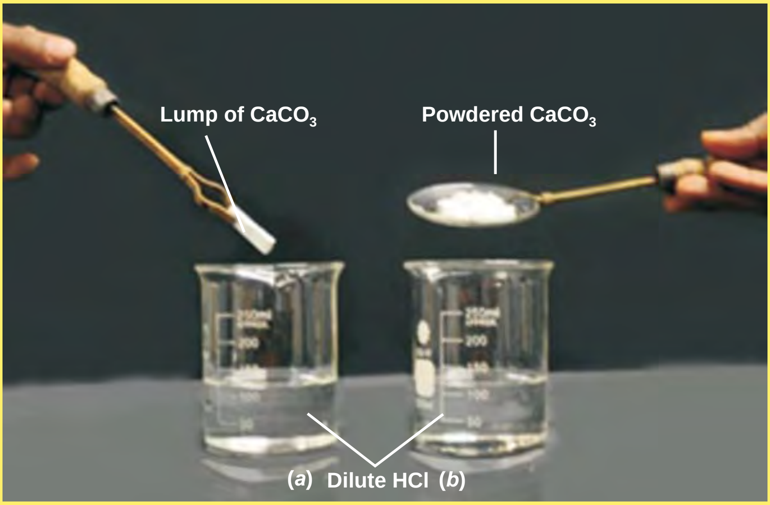


Figure 4.15 Effect of surface area on the rate of a reaction.

5. *Catalysts*

A catalyst is a substance that changes the rate of a chemical reaction without itself being consumed in the reaction. For example, the decomposition of potassium chlorate, $KClO_3$ into KCl and O_2 is made faster in the presence of MnO_2 catalyst.

$$2KClO_3 \xrightarrow[\text{Heat}]{MnO_2} 2KCl + 3O_2$$

A catalyst speeds up the rate of a reaction by providing an alternative reaction path with lower activation energy. Lower activation energy for a reaction corresponds to the higher reaction rate. Figure 4.16 illustrates how the presence of catalysts speeds up a reaction.

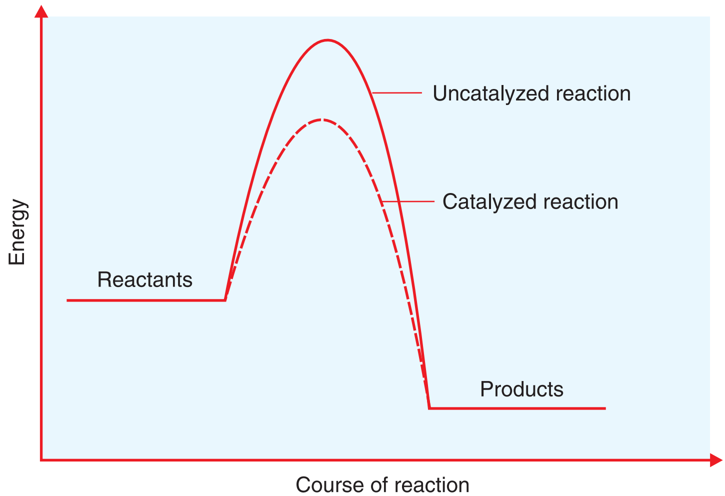


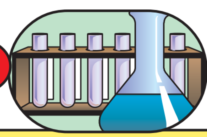
Figure 4.16 Activation energy and catalysts.

Activity 4.16

Form a group and discuss the following concept. Present your discussion to the class.

Do the factors that affect the rate of a chemical reaction influence a physical change in the same manner? Explain, by giving appropriate example.

Experiment 4.12



Investigating the effect of a catalyst on the Rate of a Chemical Reaction

Objective: To determine the effect of MnO_2 catalyst on the rate of decomposition of H_2O_2.

Apparatus: Test tube, test-tube rack.

Chemicals: MnO_2, H_2O_2.

Procedure:

1. Take 5 mL of hydrogen peroxide in each of the two test tubes.
2. Add a small amount of coarse manganese (IV) oxide to the first test tube (*a*). What do you observe?
3. Leave the second test tube (*b*) without adding any chemical.
4. What change do you observe in each test tube?
5. Introduce a glowing splint in the test tube and test for the evolution of a gas.

Observations and Analysis:

1. In which of the test tubes does the reaction occur at a fast rate? Why?
2. Which gas is evolved in the test tube?
3. What do you conclude from the experiment?

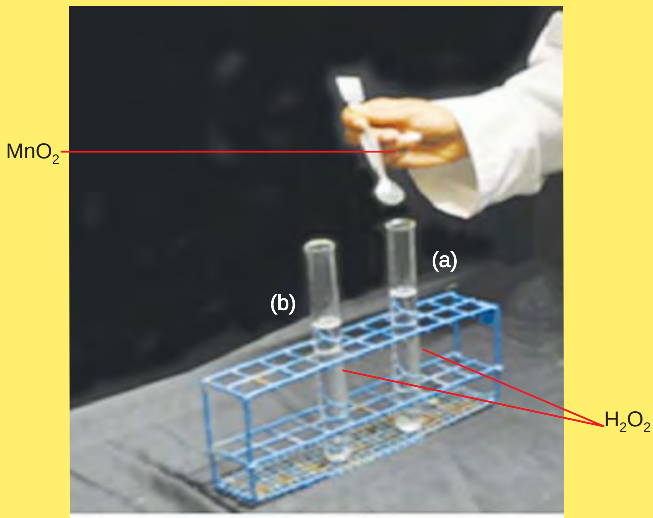


Figure 4.17 Effect of a catalyst on the decomposition of H_2O_2.

Reading Check

What is the difference between positive and negative catalysts? Explain the importance of negative catalysts by giving examples.

4.8.2 Chemical Equilibrium

In the previous chapters, we came across chemical reactions in which all the reactants are completely converted to products. Such types of reactions are known as irreversible reactions. Irreversible reactions proceed only in one direction (forward direction) and expressed by a single arrow ($\rightarrow$).

Examples

$$2Na + Cl_2 \rightarrow 2NaCl$$

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

However, there are many chemical reactions that do not proceed to completion. The products at the same time react to give (produce) the reactants. These are called reversible reactions.

Reversible reactions take place in both the forward and backward directions under the same conditions. A double arrow ($\rightleftarrows$) or ($\rightleftharpoons$) pointing in opposite directions is used in such reaction equations.

Example

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The forward reaction proceeds from left to right and the reaction that goes from right to left is the reverse reaction.

Does a reaction stop if it attains equilibrium?

Chemical equilibrium is the state of a chemical system in which the rates of the forward and reverse reactions are equal. At the state of chemical equilibrium, there is no net change in the concentrations of reactants and products because the system is in dynamic equilibrium. Dynamic equilibrium means the reaction does not stop and both the forward and the backward reactions continue at equal rates.

At equilibrium, Rate of forward reaction = Rate of reverse reaction

The law of chemical equilibrium can be expressed mathematically using the molar concentrations of reactants and products at equilibrium. The concentration of species is denoted by enclosing the formula in square bracket [ ].

Thus, for the reversible reaction:

$$aA + bB \rightleftharpoons cC + dD$$

Rate of forward reaction = $K_f\,[A]^a\,[B]^b$

Rate of reverse reaction = $K_r\,[C]^c\,[D]^d$

where K_f and K_r are rate constants for the forward and reverse reactions respectively.

Since at equilibrium the rate of the forward reaction equals the rate of the reverse reaction, it follows:

$$K_f\,[A]^a\,[B]^b = K_r\,[C]^c\,[D]^d$$

$$\frac{K_f}{K_r} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Solving for the constants, K_f/K_r, gives a new constant, termed as the equilibrium constant, K_{eq}.

Therefore, $$K_{eq} = \frac{K_f}{K_r} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Example

For the reaction,

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

Rate of forward reaction $= K_f\,[N_2][H_2]^3$

Rate of reverse reaction $= K_r\,[NH_3]^2$.

$$K_{eq} = \frac{K_f}{K_r} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

The rates of the forward and reverse reactions are also illustrated by the following graph.

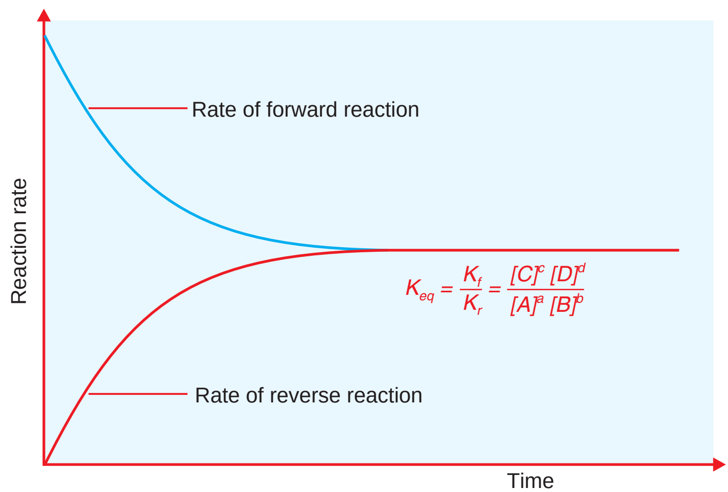


Figure 4.18 Change of the rate of the forward and reverse reactions with time.

As it is noted in the figure the rate of the forward reaction decreases with time as the concentrations of the reactants, A and B decrease with time. The reverse reaction rate starts at zero and increases as more of the products, C and D are produced. However, at equilibrium the forward and the reverse reaction rates are equal.

Example

The following equilibrium has been studied at 230°C.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

In one experiment, the concentrations of the reacting species at equilibrium are found to be [NO] = 0.0542 M, $[O_2]$ = 0.127 M and $[NO_2]$ = 15.5 M. Calculate the equilibrium constant of the reaction at this temperature.

Solution :

The equilibrium constant is given by

$$K = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

Substituting the concentration, we find that

$$K = \frac{(15.5)^2}{(0.0542)^2(0.127)} = 6.44 \times 10^5 \text{ M}^{-1}$$

Exercise 4.15

Give appropriate answers to the following questions.

1. What is the reason for all collisions between reactant molecules not to lead to products? Explain on the basis of the collision theory.
2. Explain why the rates of a reaction change with time.
3. Write the equilibrium constant expression for each of the following reactions.

   a $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$

   b $CH_4(g) + Cl_2(g) \rightleftharpoons CH_3Cl(g) + HCl(g)$

   c $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

   d $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$

4. At the start of a reaction there are 0.0249 mol of N_2, 3.21×10^{-2} mol H_2 and 6.42×10^{-4} mol of NH_3 in a 3.50 L reaction vessel at 375°C.

   If the equilibrium constant, *K*, for the reaction:

   $$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

   is 1.2 at this temperature, decide whether the system is at equilibrium or not. If it is not, predict in which direction, the net reaction will proceed.

Activity 4.17

Form a group and discuss the importance of equilibrium in the study of chemical reactions.

Present your discussion to the class.

Factors that Affect Chemical Equilibrium

Activity 4.18

Form a group and try to explore at least two properties that can be utilize to determine the state of chemical equilibrium in system.

Present your findings to the class.

A system remains at a state of chemical equilibrium if there is no change in the external conditions that disturb the equilibrium. But the point of equilibrium could be affected due to any external factors like temperature, pressure, concentration and so on. How a system at equilibrium adjusts itself to any of these changes is stated by the French chemist Henri Le Chatelier in 1888.

Le Chatelier states that if *a stress is applied to a system in equilibrium, the system will respond in such a way to counteract the stress*. The stress could be change in temperature, concentration or pressure.

The factors affecting chemical equilibrium and their effects:

1. *Effect of temperature*

The effect of temperature changes on equilibrium depends on whether the reaction is exothermic or endothermic. An increase in the temperature of a system will favour an endothermic reaction and a decrease in temperature favors an exothermic reaction. For example, consider the following reaction:

$$H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g);\ \Delta H = -41\ kJ$$

Since the reaction is exothermic,

i) if temperature is increased, the system will shift to the left.

ii) if temperature is decreased, the system will shift to the right and a high yield of products (H_2 and CO_2) is obtained.

2. *Effect of Pressure or Volume*

Pressure changes only affect equilibrium reaction involving gaseous reactants and products. The effect of pressure on liquids and solids is negligible. An increase in pressure (or decrease in volume) on a gaseous system shifts the equilibrium in the direction of forming smaller number of moles of gas. On the contrary, decreasing the pressure shifts the equilibrium in the direction of forming more number of moles of the gas. For example, in the reaction,

$$\underbrace{2CO(g) + O_2(g)}_{\text{3 moles of gases}} \rightleftharpoons \underbrace{2CO_2(g)}_{\text{2 moles of gases}}$$

There are 3 moles of reactants and 2 moles of product.

Therefore, increasing the pressure of the system shifts the equilibrium to the forward direction. This results in a higher yield of CO_2. Decreasing the pressure of the gaseous mixture shifts the equilibrium in the reverse direction.

When the number of moles of reactants and products are equal, pressure has no effect on the equilibrium. For example, in the following reaction;

$$\underbrace{H_2(g) + I_2(g)}_{\text{2 moles of gases}} \rightleftharpoons \underbrace{2HI(g)}_{\text{2 moles of gases}}$$

the number of moles of reactants and products are equal (2 mol each) and hence no effect of pressure on the equilibrium.

Exercise 4.16

For the following equilibrium system, how would the position of the equilibrium be changed if:

a the temperature is increased;

b the pressure is decreased?

c the temperature is increased, and

d the pressure is decreased

1. $N_2(g) + O_2(g) \rightleftharpoons 2NO(g) + \text{heat}$
2. $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g), \Delta H = +241.7 \text{ kJ}$
3. $N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$
4. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{heat}$
5. $N_2O_4(g) + \text{heat} \rightleftharpoons 2NO_2(g)$
6. $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) + 21 \text{ kJ}$

3. *Effect of Concentration*

According to Le Chatelier's principle, any change in the concentration of a reactant or product will lead to a change in the concentration of the substances on the other side of the equation.

If one of the reactants is added to the equilibrium mixture, the system shifts to the forward direction and a high yield of product is obtained. To the contrary, if more product is added to the system, the equilibrium shifts to the reverse direction. For example, in the reaction,

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

If the concentration of N_2 or H_2 is increased (i.e., if more N_2 or H_2 is added), the equilibrium will shift to the right direction and more NH_3 will be produced.

4. *Effect of catalysts*

Catalysts change the speed of both the forward and reverse reactions equally. However, catalysts do not affect the state of chemical equilibrium of a reaction. This means that the position of equilibrium will not be shifted due to the presence of a catalyst.

How can Le-Chatelier's principle help in maximizing the yields of products in industrial production?

Many industrial reactions are reversible reactions. The Haber and contact processes provide excellent illustrations of the effects of temperature, pressure and catalyst on the equilibrium systems.

A Haber process (the *industrial production of ammonia*). In the Haber process, ammonia (NH_3) is industrially manufactured using gaseous nitrogen and hydrogen.

$$N_2(g) + 3H_2 \underset{\text{Catalyst}}{\overset{\text{Fe}}{\rightleftharpoons}} 2NH_3(g);\ \Delta H = -92\ \text{kJ/mol}$$

High yield of ammonia is produced by decreasing the temperature and increasing the pressure of the system.

However, the yield per unit time is highly increased by using an iron catalyst.

B Contact process (*the industrial production of sulphuric acid*). Similarly, in the contact process, sulphuric acid (H_2SO_4) is commercially manufactured by the reaction of SO_2 and O_2. The resulting SO_3 is then converted to H_2SO_4.

$$SO_2(g) + O_2 \overset{V_2O_5}{\rightleftharpoons} 2SO_3(g) + \text{heat}$$

To get high yield of SO_3, how would you adjust the temperature, pressure and concentration in the above reaction?

Check list

Key terms of the unit

- *Activation energy*
- *Catalysts*
- *Chemical equilibrium*
- *Chemical reaction*
- *Collision theory*
- *Combination reaction*
- *Decomposition reaction*
- *Double displacement reaction*
- *Dynamic equilibrium*
- *Endothermic reaction*
- *Energy changes*
- *Energy diagrams*
- *Enthalpy*
- *Equilibrium constant*
- *Exothermic reaction*
- *Heat of reaction*
- *Irreversible reaction*
- *Le Chatelier's principle*
- *Limiting reactants*
- *Oxidation*
- *Oxidizing agents*
- *Percentage yield*
- *Products*
- *Reactants*
- *Reaction rates*
- *Redox reaction*
- *Reducing agents*
- *Reduction*
- *Reversible reaction*
- *Single displacement reaction*
- *Stoichiometry*
- *Theoretical and actual yield*

Unit Summary

- *Chemical reactions are represented by chemical equations.*
- *The three basic laws of chemical reactions are: the law of conservation of mass, the law of definite proportion and the law of multiple proportions.*
- *A balanced chemical equation is an equation in which all the number of atoms of reactants and products are equal.*
- *Most of the chemical reactions are accompanied by energy changes.*
- *Exothermic reactions release heat energy to the surrounding.*
- *Endothermic reactions absorb heat energy from the surrounding.*
- *Heat of reaction or change in enthalpy is the amount of heat energy liberated or absorbed by a chemical reaction.*
- *Chemical reactions are classified into combination, decomposition, single displacement and double displacement reactions.*

- *Stoichiometry is the quantitative relationship between reactants and products.*
- *Mass-mass problems, mass-volume and volume-volume problems are the main types of stoichiometric calculations.*
- *Oxidation is the process where a substance loses electron (s) or increases its oxidation number.*
- *Reduction is the process where a substance gains electron (s) or decreases its oxidation number.*
- *Oxidizing agents are the substances reduced, and reducing agents are the substances oxidized.*
- *Rate of reaction is the change in concentration of reactants or products with time.*
- *Reversible reactions take place in both the forward and backward directions.*
- *Irreversible reactions proceed only in the forward direction.*
- *Chemical equilibrium is the state of a chemical reaction in which the rate of the forward and reverse reactions is equal.*
- *Le Chatlier's principle states that when an equilibrium is subjected to a stress, the system will shift in a direction so as to relieve the stress.*
- *The chemical equilibrium is affected by temperature, pressure, and concentrations.*

REVIEW EXERCISE ON UNIT 4

Part I: True-False Type Questions

1. In any chemical reaction, each type of atoms is conserved.
2. In a balanced chemical equation, both sides of the equation have the same number of moles.
3. The oxidizing agent is oxidized by the reducing agent.
4. Most metallic elements are strong reducing agents, whereas most non-metallic elements are strong oxidizing agents.
5. The higher the percentage yield of a chemical reaction, the more successful the reaction.
6. The rate of a reaction is directly proportional to time.
7. A catalyst increases the reaction rates by changing the reaction mechanism.
8. The concentrations of reactants and products are equal at the state of chemical equilibrium.

9. A chemical equilibrium involves two opposite reactions, where one is endothermic and the other is exothermic.

10. The amount of product obtained at equilibrium directly indicates how fast equilibrium is attained.

Part II: Write the missing words in your exercise book

11. The oxidation number of sulphur in $CuSO_4.5H_2O$ is ____________ .

12. The value of the equilibrium constant (K) is changed or affected by ____________ .

13. Out of all factors, the rate of a heterogeneous reaction is highly influenced by ____________ .   .

Part III: Problems to Solve

14. How many grams of oxygen can be prepared by the decomposition of 12 grams of mercury (II) oxide?

15. 25 g of NH_3 is mixed with 4 moles of O_2 in the given reaction:

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(l)$$

   a Which is the limiting reactant?

   b What mass of NO is formed?

   c What mass of H_2O is formed?

16. Consider the following reaction:

$$2N_2O_5(g) \rightleftharpoons 4NO_2(g) + O_2(g)$$

When 40 g of N_2O_5 decomposes, 4.5 g of O_2 is formed. What is the percent yield?

17. Consider the following equilibrium:

$$N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$$

How would the equilibrium of the system be affected by the following changes?

   a Increasing the temperature.

   b Increasing the concentration of N_2.

   c Removing all NO and H_2.

   d Compressing the reaction vessel.

   e Decreasing the volume of the reaction vessel.

   f Decreasing the amount of N_2 or H_2O.