6

Chemical Arithmetic and Reactions

Total number of reactions we study in chemistry is very large. They are of numerous types. In lesson 2, you have learnt how to write and balance chemical equations. In this lesson you will learn, how chemical equations can be classified into various categories on the basis of some of their features. You will also learn about the information that can be obtained from a balanced chemical equation and how we can use this information for making calculations. You have learnt about acids, bases and salts in earlier classes. In this lesson you will learn more about them.

OBJECTIVES

After completing this lesson, you will be able to:

- list various types of reactions;
- distinguish between various types of reactions;
- classify the reactions according to their rates and energy changes;
- work out simple problems based on stoichiometry;
- define acids, bases and salts and give their examples;
- explain the acid-base equilibrium in aqueous systems;
- define pH and solve simple problems based on pH.

6.1 TYPES OF CHEMICAL REACTIONS

Chemical reactions can be classified on the basis of some of their features. One classification is based on the nature of chemical change that occurs in the reaction. On this basis reactions can be classified into five types. These are:

(i) Combination reactions
(ii) Decomposition reactions
(iii) Displacement reactions
(iv) Double-displacement reactions
(v) Oxidation-reduction reactions

Let us now learn about these reactions.

6.1.1 Combination reactions

A reaction in which two or more substances react to form a new substance is called a **combination reaction.**

A special category of combination reactions is the one in which a compound is formed by combination of its constituent elements. Such a reaction is known as **synthesis reaction.**

Following are some examples of combination reactions:

1. Carbon (charcoal, coke) burns in presence of oxygen (or air) to form carbon dioxide (*synthesis reaction*).

$$\underset{\text{carbon}}{C(s)} + \underset{\text{oxygen}}{O_2(g)} \longrightarrow \underset{\text{carbon dioxide}}{CO_2(g)}$$

2. Hydrogen burns in presence of oxygen (or air) to form water (*synthesis reaction*).

$$\underset{\text{hydrogen}}{2H_2(g)} + \underset{\text{oxygen}}{O_2(g)} \longrightarrow \underset{\text{water}}{2H_2O(l)}$$

3. Phosphorus combines with chlorine to form phosphorus pentachloride (*synthesis reaction*).

$$\underset{\text{phosphorus}}{P_4(s)} + \underset{\text{chlorine}}{10Cl_2(g)} \longrightarrow \underset{\text{phosphorus pentachloride}}{4PCl_5\,(s)}$$

4. Ammonia combines with hydrogen chloride to form ammonium chloride.

$$\underset{\text{ammonia}}{NH_3\,(g)} + \underset{\text{hydrogen chloride}}{HCl(g)} \longrightarrow \underset{\text{ammonium chloride}}{NH_4Cl_{(s)}}$$

6.1.2 Decomposition reactions

A reaction in which one substance breaks down into two or more simpler substances is known as **decomposition reaction.**

A decomposition reaction always involves breaking of one or more chemical bonds and therefore occurs only when the required amount of energy is supplied. The energy may be supplied in any of the following forms:

(i) **Heat:** Such decomposition reactions are called **thermal decomposition reactions.**

(ii) **Electricity:** Such decomposition reactions are called **electro-decomposition** reactions and the process is known as **electrolysis.**

(iii) **Light:** Such decomposition reactions are called **photo-decomposition** reactions and the process is known as **photolysis.**

Following are some examples of decomposition reactions:

1. Potassium chlorate decomposes on heating into potassium chloride and oxygen.

$$\underset{\text{potassium chlorate}}{2KClO_3\,(s)} \longrightarrow \underset{\text{potassium chloride}}{2KCl(s)} + \underset{\text{oxygen}}{3O_2(g)}$$

2. When calcium carbonate (*limestone*) is heated strongly it decomposes into calcium oxide (*quicklime*) and carbon dioxide.

$$\underset{\text{calcium carbonate}}{CaCO_3(s)} \longrightarrow \underset{\text{calcium oxide}}{CaO(s)} + \underset{\text{carbon dioxide}}{CO_2(g)}$$

3. Hydrogen peroxide decomposes into water and oxygen on heating.

$$\underset{\text{hydrogen peroxide}}{2H_2O_2(l)} \longrightarrow \underset{\text{water}}{2H_2O(l)} + \underset{\text{oxygen}}{O_2(g)}$$

4. Water decomposes into hydrogen and oxygen on passing electricity through it (*electrolysis*).

$$\underset{\text{water}}{2H_2O(l)} \longrightarrow \underset{\text{hydrogen}}{2H_2(g)} + \underset{\text{oxygen}}{O_2(g)}$$

5. Lead nitrate decomposes on heating into lead monoxide, nitrogen dioxide and oxygen.

$$\underset{\text{lead nitrate}}{2Pb(NO_3)_2(s)} \longrightarrow \underset{\text{lead monoxide}}{2PbO(s)} + \underset{\text{nitrogen dioxide}}{4NO_2(g)} + \underset{\text{oxygen}}{O_2(g)}$$

6.1.3 Displacement reactions

A reaction in which one element present in a compound is displaced by another element is known as **displacement reaction.**

Following are examples of displacement reactions:

1. Displacement of a metal by a more reactive metal.
   a. Zinc displaces copper from a solution of copper sulphate.

$$\underset{\text{zinc}}{Zn(s)} + \underset{\text{copper sulphate}}{CuSO_4(aq)} \longrightarrow \underset{\text{zinc sulphate}}{ZnSO_4(aq)} + \underset{\text{copper}}{Cu(s)}$$

   b. Magnesium displaces copper from a solution of copper sulphate.

$$\underset{\text{magnesium}}{Mg(s)} + \underset{\text{copper sulphate}}{CuSO_4(aq)} \longrightarrow \underset{\text{magnesium sulphate}}{MgSO_4(aq)} + \underset{\text{copper}}{Cu(s)}$$

2. Displacement of hydrogen from solutions of acids by more reactive metals.
   a. Zinc displaces hydrogen from dilute sulphuric acid.

$$\underset{\text{zinc}}{Zn(s)} + \underset{\text{dil. sulphuric acid}}{H_2SO_4(aq)} \longrightarrow \underset{\text{zinc sulphate}}{ZnSO_4(aq)} + \underset{\text{hydrogen}}{H_2(g)}$$

   b. Magnesium displaces hydrogen from dilute hydrochloric acid

$$\underset{\text{magnesium}}{Mg(s)} + \underset{\text{dil hydrochloric acid}}{2HCl(aq)} \longrightarrow \underset{\text{magnesium chloride}}{MgCl_2(aq)} + \underset{\text{hydrogen}}{H_2(g)}$$

2. Displacement of a halogen by a more reactive halogen.

   Chlorine displaces bromine from a solution of potassium bromide.

$$\underset{\text{chlorine}}{Cl_2(g)} + \underset{\text{potassium bromide}}{2KBr(aq)} \longrightarrow \underset{\text{potassium chloride}}{2KCl(aq)} + \underset{\text{bromine}}{Br_2(aq)}$$

6.1.4 Double-displacement reactions

A reaction in which two ionic compounds exchange their ions is known as **double displacement reaction.** The following are the examples of double displacement reactions:

a. Reaction between sodium chloride and silver nitrate.

$$\underset{\text{sodium chloride}}{NaCl(aq)} + \underset{\text{silver nitrate}}{AgNO_3(aq)} \longrightarrow \underset{\text{silver chloride}}{AgCl(s)} + \underset{\text{sodium nitrate}}{NaNO_3(aq)}$$

b. Neutralization of hydrochloric acid by sodium hydroxide.

$$\underset{\text{hydrochloric acid}}{HCl(aq)} + \underset{\text{sodium hydroxide}}{NaOH(aq)} \longrightarrow \underset{\text{sodium chloride}}{NaCl(aq)} + \underset{\text{water}}{H_2O(l)}$$

6.1.5 Oxidation–reduction or redox reactions

These are the reaction in which oxidation and reduction processes occur. Let us first learn what these processes are.

a) **Oxidation:** It is a process which involves loss of electrons. Earlier it was defined as a process involving addition of oxygen or loss of hydrogen.

b) **Reduction:** Reduction is a process which involves gain of electrons. Earlier it was defined as a process involving removal of oxygen or addition of hydrogen.

c) **Redox reactions:** From the above definitions, you must have noticed above that oxidation and reduction processes are just opposite to each other. None of these processes can occur alone. During a reaction if one substance gets oxidized the other gets reduced. Thus, both the processes occur simultaneously. That is why the reactions in which oxidation and reduction processes occur are called redox reactions or oxidation-reduction reactions. Now let us understand these processes with the help of some examples.

(i) Consider burning of coke (carbon) in presence of oxygen:

$$\underset{\text{carbon}}{C(s)} + \underset{\text{oxygen}}{O_2(g)} \longrightarrow \underset{\text{carbon dioxide}}{CO_2(g)}$$

In this reaction carbon is getting *oxidized* as oxygen is added to it and oxygen is *reduced*.

(ii) When hydrogen sulphide reacts with sulphur dioxide the products are sulphur and water.

$$\underset{\text{hydrogen sulphide}}{2H_2S(g)} + \underset{\text{sulphur dioxide}}{SO_2(g)} \longrightarrow \underset{\text{sulphur}}{3S(s)} + \underset{\text{water}}{2H_2O(l)}$$

Here, hydrogen sulphide is *oxidized* to sulphur due to loss of hydrogen while sulphur dioxide is *reduced* to sulphur due to loss of oxygen.

(iii) When copper (II) oxide is treated with hydrogen, copper and water are produced.

$$\underset{\text{cupric oxide}}{CuO(s)} + \underset{\text{hydrogen}}{H_2(g)} \longrightarrow \underset{\text{copper}}{Cu(s)} + \underset{\text{water}}{H_2O(l)}$$

Here cupric oxide is *reduced* to copper due to loss of oxygen while hydrogen is *oxidized* to water due to addition of oxygen.

(iv) When sodium metal reacts with chlorine it forms sodium chloride.

$$\underset{\text{sodium}}{2Na(s)} + \underset{\text{chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{sodium chloride}}{2NaCl(s)}$$

Sodium chloride is an ionic compound. Sodium is present in it as sodium ion (Na^+) and chlorine as chloride ion (Cl^-). This reaction can be considered to occur in the following steps:

- Each sodium atom loses one electron and forms sodium ion. Since two sodium atoms are involved in the reaction, the process is:

  $$\underset{\text{sodium}}{2Na} \longrightarrow \underset{\text{sodium ion}}{2Na^+} + 2e^-$$

  Thus, sodium is *oxidized* due to loss of electron.
- Each chlorine atom gains one electron and forms chloride ion. Since one chlorine molecule has two atoms of chlorine the process is:

  $$\underset{\text{chlorine}}{Cl_2} + 2e^- \longrightarrow \underset{\text{chloride ion}}{2Cl^-}$$

  Thus, chlorine is *reduced* due to gain of electrons.

(v) When zinc is added to an aqueous solution of copper sulphate, it displaces copper.

$$\underset{\text{zinc}}{Zn(s)} + \underset{\text{copper sulphate}}{CuSO_4(aq)} \longrightarrow \underset{\text{zinc sulphate}}{ZnSO_4(aq)} + \underset{\text{copper}}{Cu(s)}$$

Here zinc is *oxidized* to zinc ions and copper ions are *reduced* to copper. This reaction is displacement reaction as well as a redox reaction.

(d) Oxidizing and reducing agents : Consider the reaction between zinc and copper sulphate:

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

In this reaction zinc reduces cupric ions to copper. Such a substance which reduces another substance is called a **reducing agent.** Here, zinc is the *reducing agent*.

Also, in this reaction cupric ions oxidize zinc to zinc ions. Such a substance which oxidizes another substance is called an **oxidizing agent.** Here, cupric ions are the *oxidizing agent*.

CHECK YOUR PROGRESS 6.1

Match the type of reaction given in column I with the reactions given in column II.

I	II
1. Displacement reaction	A. $2H_2S(g) + SO_2(g) \longrightarrow 3S(g) + 2H_2O(1)$
2. Double displacement reaction	B. $NH_3 + HCl \longrightarrow NH_4Cl$
3. Combination reaction	C. $3CaCl_2 + 2K_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6KCl$
4. Redox reaction	D. $Mg(s) + CuSO_4(aq) \longrightarrow MgSO_4(aq) + Cu(s)$
5. Decomposition reaction	E. $2H_2O_2 \longrightarrow 2H_2O + O_2$

6.2 NATURE OF CHEMICAL REACTIONS

In the last section, we have learnt how chemical reactions have been classified into various types on the basis of the nature of chemical change that occurs in them. In this section we shall learn about some other features of chemical reactions. These features have been discussed below.

6.2.1 Homogeneous–heterogeneous reactions

Chemical reactions can be classified on the basis of physical states of reactants and products as *homogeneous* and *heterogeneous* reactions.

a) Homogeneous reactions

The reactions in which all the reactants and products are present in the same *phase* are called **homogeneous reactions**. Such reactions can occur in gas phase or solution phase only.

A. Gas phase homogeneous reactions

These are the reactions in which all reactants and products are gases.

(i) $\underset{\text{hydrogen}}{H_2(g)} + \underset{\text{chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{hydrogen chloride}}{2HCl(g)}$

(ii) $\underset{\text{sulphur dioxide}}{2SO_2(g)} + \underset{\text{oxygen}}{O_2(g)} \xrightarrow{H+} \underset{\text{sulphur trioxide}}{2SO_3(g)}$

(iii) $N_2(g)$ (nitrogen) + $3H_2(g)$ (hydrogen) $\longrightarrow$ $2NH_3(g)$ (ammonia)

B. Solution phase homogeneous reactions

These are the reactions in which all reactants and products are present in a solution.

(i) $HCl(aq)$ (hydrochloric acid) + $NaOH(aq)$ (sodium hydroxide) $\longrightarrow$ $NaCl(aq)$ (sodium chloride) + $H_2O(l)$ (water)

(ii) $CH_3COOC_2H_5(l) + H_2O(l) \longrightarrow CH_3COOH(l) + C_2H_5OH(l)$

b) Heterogeneous reactions

The reactions in which reactants and products are present in more than one *phase* are called heterogeneous reactions. Such reactions involve at least one solid substance along with one or more substances in solid, solution or gaseous phase. The following are the examples of heterogeneous reactions.

(i) $CaCO_3(s)$ (calcium carbonate) $\longrightarrow$ $CaO(s)$ (calcium oxide) + $CO_2(g)$ (carbon dioxide)

(ii) $2Mg(s)$ (magnesium) + $O_2(g)$ (oxygen) $\longrightarrow$ $2MgO(s)$ (magnesium oxide)

(iii) $BaCl_2(aq)$ (barium chloride) + $Na_2SO_4(aq)$ (sodium sulphate) $\longrightarrow$ $BaSO_4(s)$ (barium sulphate) + $2NaCl(aq)$ (sodium chloride)

6.2.2 Slow and fast reactions

Different reactions occur at different rates. Rusting of iron is a slow process and requires few days time. On the other hand burning of cooking gas is a fast reaction. On the basis of their rates chemical reactions can be classified as slow and fast reactions. Rusting of iron, curdling of milk, hydrolysis of esters at room temperature (e.g. reaction between ethyl acetate and water), fading of colours of clothes, burning of coal, etc. are some examples of slow reactions. On the other hand, neutralization reaction (e.g. reaction between hydrochloric acid and sodium hydroxide), explosion reactions (e.g. in a fire cracker bomb), action of acids or bases on litmus, and burning of cooking gas are some examples of fast reactions.

A large number of reactions are neither *slow* nor *fast*. They may be termed as moderate reactions. Burning of candle, thermal decomposition of potassium chlorate, and reaction of zinc with dilute sulphuric acid are some examples of moderate reactions.

6.2.3 Exothermic and endothermic reactions

All chemical reactions are accompanied by some energy changes. Energy is either evolved or absorbed during the reaction usually in the form of heat. Depending upon this, the reactions are classified as *exothermic* and *endothermic* reactions.

a) Exothermic reactions

The reactions in which heat is *liberated or evolved* are called exothermic reactions. In such reactions *heat* is shown as one of the products. If exact amount of heat evolved is known then this amount is written otherwise simply the word *heat* is written. Following are the examples of exothermic reactions.

(i) $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) + \text{heat}$

or $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) + 571.5\ \text{kJ}$

(ii) $C(s) + O_2(g) \longrightarrow CO_2(g) + 393.5\ kJ$

(iii) $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) + 57.3\ kJ$

b) Endothermic reactions

The reactions in which heat is *absorbed* are called endothermic reactions. In such reactions *heat* is shown as one of the reactants. If exact amount of heat absorbed is known then this amount is written otherwise simply the word *heat* is written. Following are the examples of endothermic reactions

(i) $N_2(g) + 2O_2(g) + \text{heat} \longrightarrow \underset{\text{nitric oxide}}{2NO(g)}$

or $N_2(g) + O_2(g) + 180.7\ kJ \longrightarrow 2NO(g)$

(ii) $\underset{\text{potassium chlorate}}{2KClO_3(s)} + \text{heat} \longrightarrow \underset{\text{potassium chloride}}{2KCl(s)} + 3O_2(g)$

(iii) $\underset{\text{lead nitrate}}{2Pb(NO_3)_2(s)} + \text{heat} \longrightarrow 2PbO(s) + \underset{\text{nitrogen dioxide}}{4NO_2(g)} + O_2(g)$

6.2.4 Reversible and irreversible reactions

Chemical reactions can also be classified on the basis whether they can occur only in the forward direction or in forward as well as backward directions.

a) Irreversible reactions

Most of the reactions would occur till the reactants (or atleast one reactant) have been completely converted into products. For example, if a small piece of zinc metal is put in a test tube containing excess of dilute hydrochloric acid, it completely reacts with it.

$$Zn(s) + 2HCl(aq) \longrightarrow H_2(g) + ZnCl_2(aq)$$

Such reactions occur in forward direction only. *The reactions which occur in forward direction only are called* irreversible reactions.

The following are some more examples of irreversible reactions:

(i) $\underset{\text{magnesium}}{2Mg(s)} + O_2(g) \longrightarrow \underset{\text{magnesium oxide}}{2MgO(s)}$

(ii) $\underset{\text{mercuric oxide}}{2HgO(s)} \longrightarrow \underset{\text{mercury}}{2Hg(l)} + O_2(g)$

(iii) $NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

b) Reversible reactions

On the other hand consider the reaction:

$$H_2(g) + I_2(g) \longrightarrow 2HI(g)$$

In this reaction hydrogen and iodine are not completely converted into hydrogen iodide. The reason for this is that the moment some HI is formed it starts decomposing back into H_2 and I_2.

$$2HI(g) \longrightarrow H_2(g) + I_2(g)$$

The reactions that can occur in forward and reverse directions, simultaneously under same set of conditions are called reversible reactions. Reversible nature of a reaction is indicated by writing two arrows (or two-half arrows) in opposite directions between reactants and products as shown below;

$$H_2(g) \quad + \quad I_2(g) \rightleftarrows 2HI(g)$$

or $$H_2(g) \quad + \quad I_2(g) \rightleftharpoons 2HI(g)$$

Some more examples of reversible reactions are :

(i) Synthesis of ammonia

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

(ii) Oxidation of sulphur dioxide to sulphur trioxide

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

6.2.5 Equilibrium in reversible reactions

In the last section we have learned that a reversible reaction can occur in forward as well as reverse directions simultaneously. Consider the following reaction:

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

When the reaction is started by taking a mixture of sulphur dioxide (SO_2) and oxygen (O_2) it would initially occur only in the forward direction and formation of sulphur trioxide (SO_3) would begin. Initially the rate of this reaction is fast. As it progresses its rate decreases. This happens because as reactants are consumed their concentrations decrease.

Concentration

Concentration is a measure of the amount of a substance contained per unit volume. In chemistry it is commonly measured in terms of *molarity*. Molarity is the number of moles of a substance present in one litre volume. It has the unit of mol L^{-1}. In case of *gases* it is their number of moles present in one litre volume. And in case of *solutions* it is the number of moles of solute present in one litre volume of solution. The molar concentration of a substance X is denoted by writing its formula/symbol within a square bracket [X].

As soon as SO_3 is formed, it starts decomposing and the backward reaction also starts. Initially its rate is very slow but as the reaction progresses the concentration of SO_3 (which is *reactant* for the reverse reaction) increases and the rate of reverse reaction also increases.

Thus, with the progress of reaction, the rate of forward reaction decreases and that of the reverse reaction increases with time. These changes are depicted in the figure 6.1.

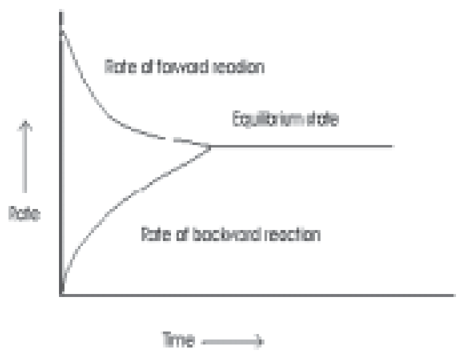

Fig. 6.1 Changes in rates of forward and backward (reverse) reactions in a reversible reaction. When the two become equal, the reaction attains equilibrium.

After some time, the rate of the forward reaction becomes equal to the rate of the reverse reaction and the reaction reaches equilibrium state (Fig. 6.1). Under these conditions, there is no change in concentration of any reactant or product. *A system is said to be in a state of equilibrium if none of its properties change with time.* In other words, when a system is in a state of equilibrium, all its properties remain constant.

At equilibrium the concentrations of reactants $[SO_2]$ and $[O_2]$ and product $[SO_3]$ are related by the following expression known as the law of equilibrium:

$$K_c = \frac{[SO_3]^2}{[SO_2][O_2]}$$

How to write the expression of the law of equilibrium ?

To understand how to write the expression of law of equilibrium for any reaction let us take a general reaction:

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

For this reaction, the law of equilibrium is given by the following expression :

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

In this expression, K_c is the equilibrium constant for the reaction.

The numerator is obtained by multiplying the concentration terms for *all products* after each term has been raised to the power which is equal to the stoichiometric coefficient of *that* product. Here **C** and **D** are the two products and **c** and **d** are their respective **stoichiometric coefficients**. Therefore, numerator would be the obtained by multiplying $[C]^c$ and $[D]^d$ terms. Also, conventionally if any *pure solid or liquid is taking part in the equilibrium, its concentration is taken as 1.*

Similarly, the denominator is obtained by multiplying the concentration terms of *all reactants* after each term has been raised to the power which is equal to the stoichiometric coefficient of *that* reactant.

Static and dynamic equilibrium

The type of **e**quilibrium attained by reversible reactions is called **dynamic equilibrium**. Such an equilibrium state is attained as a result of *two equal but opposite changes occurring simultaneously so that no net change occurs* in the system. Therefore, all the properties of the system acquire constant values. You can encounter a similar situation when a person is walking on a treadmill. His speed of walking is exactly matched by the speed of the treadmill which moves in the backward direction. The net result is that position of the person does not change and he stays there only. Another similar situation in encountered when a person using an escalator for climbing starts moving down on it and matches his speed with that of the escalator.

Another type of equilibrium is attained when a system is acted upon by a set of forces that cancel out each other. Such an equilibrium state is attained when *no change occurs in it.* This type of equilibrium is called **static equilibrium**. A book lying on a table is in state of static equilibrium because the downward acting gravitational force is balanced and cancelled by the upward acting force of reaction from the table (*Newton's third law of motion*). Another similar situation is encountered in the game *tug of war* when the efforts of the two opponent teams (*forces by which they pull the rope*) exactly match and they remain where they are.

K_c of a reaction is its characteristic property at a given temperature and it characterizes the equilibrium state. Its value changes only when temperature is changed. The same equilibrium state (characterized by the value of K_c) is reached finally whether the reaction is started from the reactant side or from the product side or all reactants and products are mixed in arbitrary amounts.

CHECK YOUR PROGRESS 6.2

Select the correct choice about the nature of each reaction out of the two options mentioned against it.

1. Burning of petrol in a car (*homogeneous / heterogeneous*).
2. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ (*exothermic / endothermic*)
3. $2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$ (*reversible / irreversible*)
4. Bursting of crackers (*slow / fast)*
5. A reversible reaction at a stage when concentration of reactants and products is changing (*equilibrium state /non-equilibrium state*)

6.3 CHEMICAL CALCULATIONS AND STOICHIOMETRY

In lesson 2 you have learnt how to write and balance chemical equations. Stoichiometry *deals with the proportions in which elements or compounds react with one another.* In this section we shall learn how to use the stoichiometric information in a balanced chemical equation for making some calculations.

6.3.1 Significance of balanced chemical equation

Balanced chemical equation carries the following information:

a) Qualitative information carried by a balanced chemical equation

- *Reactants taking part* in the reaction
- *Products formed* in the reaction
- *Physical states* of different reactants and products (*if given*)

b) Quantitative information carried by a balanced chemical equation

- *Number of molecules* of different reactants and products taking part in the reaction
- *Number of moles* of different reactants and products taking part in the reaction
- *Masses* of different reactants and products taking part in the reaction
- *Relationship between moles* of different reactants and products taking part in the reaction
- *Relationship between masses* of different reactants and products taking part in the reaction
- *Relationship between volumes* of different *gaseous reactants and products* taking part in the reaction

Let us understand how to get this information from a chemical equation with the help of an example.

Information carried by a chemical equation

	2Na(s)	+	$2H_2O(l)$	$\longrightarrow$	2NaOH(aq)	+	$H_2(g)$
Names	sodium		water		sodium hydroxide		hydrogen
Physical states	solid		liquid		aqueous solution		gas
Moles	2 moles		2 moles		2 moles		1mole
(Molar masses)	(Na= 23)		(H_2O = 2 +16 = 18)		(NaOH = 23 + 16 + 1 = 40)		(H_2 = 2)
Masses	2 x 23 = 46g			2 x 18 = 36g			2 x 40 = 80g
	1 x 2 = 2g						
Volume* of							1 x 22.7
L=22.7L							
gaseous substance							

From the information listed above we can conclude that:

(i) Sodium metal (solid) reacts with water (liquid) and produces sodium hydroxide (aqueous solution) and hydrogen (gas).

(ii) 2 moles of sodium react with 2 moles of water and produce 2 moles of sodium hydroxide and 1 mole of hydrogen. Thus the ratio of number of moles of these substances is 2:2:2:1.

(iii) 46 g sodium reacts with 36 g water and produces 80 g of NaOH and 2 g of hydrogen.

(iv) 2 moles or 46 g sodium produces 22.7 L of hydrogen gas when it reacts with water.

(v) 2 moles or 36 g water produces 22.7 L of hydrogen gas when it reacts with sodium.

c) Limitations or information not carried by a chemical equation

- Conditions under which the reaction takes place
- Rate of the reaction whether it is fast, slow or moderate
- The extent up to which the reaction takes place before equilibrium state is reached in case of a reversible reaction

6.3.2 Calculations based on chemical equations

The information that can be obtained from a chemical equation can be used to make several types of calculations. Let us carry out few such calculations.

a) Mole-mole relationship

Example 6.1: In the reaction

$$2KClO_3(s) \longrightarrow 2KCl(s) \quad + \quad 3O_2(g)$$

calculate the following:

(i) How many moles of oxygen will be produced if 10 moles of $KClO_3$ are decomposed?

(ii) How many moles of KCl would be produced with 0.6 moles of O_2?

Solution: The given reaction is

$$\underset{\text{2 moles}}{2KClO_3(s)} \longrightarrow \underset{\text{2 moles}}{2KCl(s)} \quad + \quad \underset{\text{3 moles}}{3O_2(g)}$$

*Volume of a gaseous substance can be calculated by making use of the fact that one mole of a gas occupies a volume of 22.7 L at STP (standard temperature and pressure) i.e. at 273 K temperature and 1 bar pressure.

(i) 2 moles of $KClO_3$ produce 3 moles of oxygen.

Therefore, 10 moles of $KClO_3$ would produce

$= \dfrac{3 \times 10}{2} = 15$ moles of oxygen.

(ii) With 3 moles of oxygen the number of moles of KCl produced = 2 moles

With 0.6 moles of oxygen the number of moles of KCl produced

$= \dfrac{2 \times 0.6}{3} = 0.4$ moles

b) Mass-mass relationship

Example 6.2: For the reaction

$$N_2(g) \quad + \quad 3H_2(g) \longrightarrow 2NH_3(g)$$

Calculate the masses of nitrogen and hydrogen required to produce 680 g of ammonia?

Solution: The given reaction is:

$N_2(g)$ +	$3H_2(g)$ $\longrightarrow$	$2NH_3(g)$
1mole	3 moles	2 moles
1 x 28	3 x 2	2 x (14+3)
28 g	6 g	34 g

Thus to produce 34 g ammonia the mass of nitrogen required = 28 g

Therefore to produce 680 g ammonia the mass of nitrogen required

$= \dfrac{28 \times 680}{34} = 560$ g

Similarly, to produce 34 g ammonia the mass of hydrogen required = 6 g

Therefore to produce 680 g ammonia the mass of hydrogen required

$= \dfrac{6 \times 680}{34} = 120$ g

c) Volume-volume relationship

Example 6.3 : The following reaction is used industrially for manufacture of sulphuric acid.

$$2SO_2(g) \quad + \quad O_2(g) \longrightarrow 2SO_3(g)$$

How much volume of oxygen at STP (Standard Temperature and Pressure) would be required for producing 100 L of SO_3 (at STP)?

Solution: In the reaction

$2SO_2(g)$ +	$O_2(g)$ $\longrightarrow$	$2SO_3(g)$
2 mole	1 mole	2 mole
2 x 22.7 L	22.7 L	2 x 22.7 L
2 volumes	1 volume	2 volumes

To produce 2 volumes or 2 L of SO_3 the oxygen required is 1 volume or 1 L.

To produce 1L of SO_3 the oxygen required is 0.5 L

Therefore to produce 100 L of SO_3 the volume of oxygen required is 0.5 x 100 = 50 L

d) Mixed calculations

Example 6.4: Calculate the mass of hydrochloric acid required for neutralizing 1 kg of NaOH

Solution: The neutralization reaction involved between hydrochloric acid and sodium hydroxide is as follows :

HCl(aq)	+	NaOH(aq)	$\longrightarrow$	NaCl(aq)	+	H_2O (l)
1mole		1 mole				
1 + 35.5		23 + 16 + 1				
= 36.5 g		= 40 g				

Thus, for neutralizing 40 g of NaOH the mass of HCl required is 36.5 g.

For neutralizing 1 kg or 1000 g of NaOH the mass of HCl required is

$$\frac{36.5 \times 1000}{40} = 912.5 \text{ g}$$

Example 6.5: In the reaction

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

calculate the following:

(i) The maximum number of moles of sodium that can react with 4 moles of water.

(ii) The mass of sodium hydroxide that would be produced when 4.6 g of sodium reacts with excess of water.

(iii) The mass and volume at STP of hydrogen gas that would be produced when 1.8 g of water reacts completely with sodium metal.

Solution:

2Na(s)	+	$2H_2O$(l)	$\longrightarrow$	2NaOH(aq)	+	H_2(g)
2 moles		2 moles		2 moles		1 mole
2 x 23 = 46 g		2 x 18 = 36 g		2 x 40 = 80 g		1 x 2 = 2 g
						22.7 L at STP

(i) From the equation it can be seen that

2 moles of water react with 2 moles of sodium

4 moles of water can react with a maximum of 4 moles of sodium.

(ii) 46 g sodium reacts to produce 80 g sodium hydroxide

4.6 g sodium would produce $\frac{80 \times 4.6}{46}$ = 8.0 g sodium hydroxide.

(iii) 6 g of water produces 2 g or 22.7 L of hydrogen at STP

1.8 g of water would produce $\frac{2 \times 1.8}{36}$ = 0.1 g of hydrogen and $\frac{22.7 \times 1.8}{36}$ = 1.135 L of hydrogen at STP.

CHECK YOUR PROGRESS 6.3

Consider the equation for combustion of benzene (C_6H_6):

$$2C_6H_6(l) \quad + \quad 15O_2(g) \longrightarrow 12CO_2(g) \quad + \quad 6H_2O(g) \quad + \quad \text{heat}$$

Some statements about this reaction are given below. Read them carefully and indicate against each statement whether it is true (T) or false (F).

1. It is an exothermic reaction.
2. 0.1 mole of benzene would require 7.5 moles of oxygen for its combustion.
3. 1 mole of benzene would produce 134.4 L of CO_2 at STP.
4. 10.8 g water would be produced by combustion of 15.6 g benzene.
5. 200 g of O_2 is sufficient to convert 1 mole of benzene completely into CO_2 and H_2O.

6.4 ACIDS, BASES AND SALTS

You have learnt in your earlier classes about three types of substances–acids, bases and salts. They are vital to many life processes and are valuable to industry. Let us do a quick revision about them.

6.4.1 Acids

An **acid** *is defined as a substance that furnishes hydrogen ions (H^+) in its solution.* Actually, the hydrogen ion, H^+ does not exist in the aqueous solution as such. Instead, it attaches itself to a water molecule to form the hydronium ion (H_3O^+). It is customary, however, to simplify equations by using the symbol for the hydrogen ion (H^+).

The strongest acids are the mineral or inorganic acids. These include sulphuric acid, nitric acid, and hydrochloric acid. More important to life are hundreds of weaker organic acids. These include acetic acid (in vinegar), citric acid (in lemons), lactic acid (in sour milk), and the amino acids (in proteins).

Acids have sour taste and turn blue litmus red. They react with metals (which are more reactive than hydrogen) to liberate hydrogen.

$$Zn(s) \quad + \quad H_2SO_4(aq) \longrightarrow ZnSO_4(aq) \quad + \quad H_2(g)$$

6.4.2 Bases

Bases are the substances which furnish hydroxyl ions OH^- in their solutions. The hydroxides of metals are the compounds that have the hydroxyl group. They are called *bases*. Hydroxides of alkali metals–lithium, sodium, potassium, rubidium, and caesium have the special name of alkalies. A basic solution is also called an *alkaline* solution. Bases have bitter taste and turn red litmus blue.

Taste of acids and bases

Although you will find mention of taste of acids being sour and that of bases being bitter in books, never attempt to taste them yourself. Many of them can cause serious damage if swallowed or even on their contact with tongue.

6.4.3 Salts

A **salt** *is a substance produced by the reaction of an acid with a base.* It consists of the cation (positive ion) of a base and the anion (negative ion) of an acid. *The reaction between an acid and a base is called a* **neutralization reaction**. In solution or in the molten state, most salts are completely dissociated into cation and anion and are good conductors of electricity.

$$\underset{\text{sodium hydroxide}}{2NaOH(s)} + \underset{\text{sulphuric acid}}{H_2SO_4(l)} \longrightarrow \underset{\text{sodium sulphate}}{Na_2SO_4(aq)} + 2H_2O(l)$$

Another typical acid-base reaction is between calcium hydroxide and phosphoric acid to produce calcium phosphate and water:

$$\underset{\text{calcium hydroxide}}{3Ca(OH)_2(s)} + \underset{\text{phosphoric acid}}{2H_3PO_4(l)} \longrightarrow \underset{\text{calcium phosphate}}{Ca_3(PO_4)_2(aq)} + 6H_2O(l)$$

CHECK YOUR PROGRESS 6.4

A substance AB is formed by reaction between an acid X and a base Y along with water. The cation and anion of the compound AB are monovalent.

1. What type of substance is AB?
2. Which one out of AB, X and Y would turn red litmus blue?
3. Which one out of AB, X and Y would have sour taste?

6.5 ACID-BASE EQUILIBRIA IN AQUEOUS SYSTEMS

In the last section we discussed the nature of three important types of substances–acids, bases and salts. They show their typical properties in aqueous solutions. In this section we shall learn about their behaviour in such solutions.

6.5.1 Electrolytes and non-electrolytes

An **electrolyte** *is a substance that conducts electric current through it in the molten state or through its solution*. The most familiar electrolytes are acids, bases, and salts, which dissociate in their molten state when dissolved in such solvents as water or alcohol. When common salt (sodium chloride, NaCl) is dissolved in water, it forms an electrolytic solution, dissociating into positive sodium ions (Na^+) and negative chloride ions (Cl^-).

A **non-electrolyte** *is a substance that does not conduct electric current through it in the molten state or through its solution*. Non-electrolytes consist of molecules that bear no net electric charge and they do not dissociate in their molten state or in their solutions. Sugar dissolved in water maintains its molecular integrity and does not dissociate and it is a non-electrolyte.

6.5.2 Strong and weak electrolytes

In the last section we learned that electrolytes dissociate into ions in their solutions. *Some electrolytes are completely dissociated into ions*. They are called **strong electrolytes**. Sodium chloride, potassium hydroxide and hydrochloric acid are strong electrolytes. On the other hand *some other electrolytes are dissociated only partially into ions*. They are called **weak electrolytes.** Acetic acid and ammonium hydroxide are weak electrolytes.

6.5.3 Dissociation of acids and bases in water

In the last section we learned that some electrolytes are strong while others are weak. In this section we shall study more about dissociation processes that occur in aqueous solutions of acids and bases.

6.5.3a Dissociation of acids

(i) Dissociation of strong acids

Strong acids are completely dissociated into ions in their aqueous solutions. Consider dissociation of hydrochloric acid:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

From the above equation it can be seen that

- HCl is completely converted into its ions and no amount of it remains in the undissociated form.
- One mole of HCl forms one mole each of hydrogen ions and chloride ions. Thus, concentration (molarity) of H^+ ions is same as that of HCl in the solution.

(ii) Dissociation of weak acids

Weak acids are only partially dissociated into ions in their aqueous solutions. Consider dissociation of acetic acid.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

From the process depicted above it can be seen that:

- CH_3COOH is only partially dissociated into ions.
- The process of dissociation is reversible and an equilibrium is established between dissociated and undissociated CH_3COOH.
- The amount of hydrogen ions and acetate ions formed is less than the total amount of acetic acid taken initially. Thus, if one mole of acetic acid was dissolved in one litre of solution (concentration = 1 mol L^{-1}) the concentration of hydrogen ions H^+ formed in the solution would be less than 1 mol L^{-1}. In fact acetic acid is such a weak electrolyte that less than 1% of it would dissociate in this solution.
- We can write expression of the law of equilibrium for the above equilibrium as

$$Ka = \frac{[H^+][CH3COO^-]}{[CH_3COOH]}$$

Here the symbol used for equilibrium constant is K_a in place of K_c. Here K_a is **dissociation constant** of acetic acid.

6.5.3b Dissociation of bases

(i) Dissociation of strong bases

Strong bases like sodium hydroxide are completely dissociated in their solutions.

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

From the above equation it can be seen that

- NaOH is completely converted into its ions and no amount of it remains in the undissociated form.
- One mole of NaOH forms one mole each of sodium ions and hydroxyl ions. Thus concentration (molarity) of OH^- ions is same as that NaOH of in the solution.

(ii) Dssociation of weak bases

Weak bases like ammonium hydroxide are only partially dissociated in their solutions.

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$NH_4OH(aq) \rightleftharpoons NH_4+(aq) + OH^-(aq)$$

From the process shown above it can be seen that

- NH_4OH is only partially dissociated.
- The dissociation process is a reversible process and in the solution equilibrium is established between dissociated and undissociated NH_4OH.
- The amount of OH^- ions and NH_4^+ ions formed is less than the total amount of ammonium hydroxide taken initially.
- We can write expression of the law of equilibrium as

$$Ka = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$$

Here the symbol used for equilibrium constant is K_b in place of K_c. K_b is the *dissociation constant of* ammonium hydroxide.

6.5.4 Self-dissociation of water

Pure water is neutral in nature. It ionizes to a small extent and releases an equal number of hydrogen and hydroxide ions.

$$H_2O(1) \rightleftharpoons H^+(aq) + OH^-(aq)$$

It can be seen from the above equation that in pure water

$$[H^+] = [OH-]$$

Also, for this equilibrium

$$K_w = [H^+].\ [OH^-]$$

where K_w is known as **ionic product** of water. This is in fact the equilibrium constant for self dissociation process of water. The term in the denominator is $[H_2O]$ which by convention is taken as 1 for any pure solid or liquid (*see section 6.2.5)*.

The concentration of H^+ and OH^- ions in water has been measured and found to be 1×10^{-7} mol L^{-1} each at 25 ^{0}C. Instead of saying that the hydrogen ion concentration in pure water is 1×10^{-7} mol L^{-1}, it is customary to say that the pH of water is 7.0 . *The* pH *is the logarithm* (see box) *of the reciprocal of the hydrogen ion concentration.* It is written:

$$pH = \log \frac{1}{[H^+]}$$

Alternately, the pH *is the negative logarithm of the hydrogen ion concentration* i.e.

$$pH = -\log [H^+]$$

Because of the negative sign in the expression, if $[H^+]$ *increases* pH would *decrease* and if it *decreases* the pH would *increase*.

LOGARITHM

Logarithm is a mathematical function.

If, $x = 10^y$

Then $y = \log x$

You will study more about logarithm in your higher classes.

Similarly, we may define pOH and pK_w as:

$pOH = -\log [OH^-]$

and $pK_w = -\log K_w$

Since the concentration of OH^- ions, $[OH^-]$ is

1×10^{-7} mol L^{-1} ; pOH = 7

The relationship between $pK_{w,}$ pH and pOH is

$$pK_w = pH + pOH$$
$$= 7+7$$
$$= 14$$

The following points should be noted regarding self-dissociation of water:

(i) Water produces H^+ and OH- ions in equal amounts therefore:

$[H^+] = [OH^-]$

(ii) Water is a neutral liquid.

(iii) pH of water is 7.0 at 25 ^{0}C temperature.

(iv) The sum of pH and pOH of any aqueous solution is always 14 at 25 ^{0}C.

6.5.5 Neutral, acidic and basic solutions and their pH

In the light of discussion on self-dissociation of water in the last section, we can now discuss the characteristics of neutral, acidic and basic aqueous solutions.

6.5.5a Neutral aqueous solutions

Neutral solutions would be similar to water, which is also neutral in nature. Therefore, the following are the characteristics of neutral aqueous solutions:

(i) $[H^+] = [OH^-]$

(ii) pH = 7.0 at 25 ^{0}C

6.5.5b Acidic aqueous solutions

Acidic solutions would have more $[H^+]$ than in water. Therefore the following are the characteristics of acidic aqueous solutions:

(i) $[H^+] > [OH^-]$

(ii) Since hydrogen ion concentration in acidic solutions is more than in water their pH would be *less* than that of water i.e.

pH < 7.0 at 25 ^{0}C

(iii) For calculation of pH of acidic solutions first the concentration of H^+ ions i.e. $[H^+]$ is calculated. From it the pH is calculated by the relation $pH = -\log [H^+]$

Such calculations have been shown in the next section.

6.5.5c Basic aqueous solutions

Basic solutions would have more $[OH^-]$ than in water. Therefore they would have less $[H^+]$ than water. The following are the characteristics of basic aqueous solutions:

(i) $[H^+] < [OH^-]$

(ii) Since hydrogen ion concentration in basic solutions is less than in water their pH would be more than that of water i.e.

$$pH > 7.0 \text{ at } 25\ ^0C$$

The pH of such solutions can be calculated indirectly. First pOH is calculated from the concentration of OH^- ions using the relation (*see next section*).

$$pOH = -\log [OH^-]$$

Then pH is calculated by the relation

$$pH = 14 - pOH$$

Thus in brief we may conclude that *at 25 ^0C*:

(i) Water has a pH of 7 and is neutral.

(ii) Solutions with pH 7 are neutral.

(iii) Solutions with pH less than 7 are acidic.

(iv) Solutions with pH more than 7 are basic.

6.5.7 Calculations based on pH concept

In the last section we learned the concept of pH and its relationship with hydrogen ion or hydroxyl ion concentration. In this section we shall use these relations to perform some calculations.

It may be noted that the methods of calculation of pH used in this lesson are valid for solutions of *strong* acids and bases only. The method is *not valid* for solutions, which are extremely dilute. The concentration of H^+ or OH^- *should not be less than 10^{-6}molar.*

Example 6.6 : Calculate the pH of 0.001 molar solution of HCl.

Solution: HCl is a strong acid and is fully dissociated in its solutions according to the process:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

From the above process it is clear that one mole of HCl will give one mole of H^+ ions. Therefore the concentration of H^+ ion would also be 0.001 molar or 1×10^{-3} mol L^{-1}.

Thus

$$[H^+] = 1 \times 10^{-3} \text{ mol } L^{-1}$$

$$pH = -\log [H^+] = -(-3) = 3$$

Thus pH= 3

Example 6.7 : What would be the pH of an aqueous solution of sulphuric acid which is 5×10^{-5} molar in concentration?

Solution : Sulphuric acid dissociates in water as:

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

Thus each mole of sulphuric acid gives two moles of H^+ ions in solutions. One litre of 5 x 10^{-5} molar solution contains 5 x 10^{-5} moles of H_2SO_4, which would give 2 x 5 x 10^{-5} = 10 x 10^{-5} = 10^{-4} mol of H^+ , therefore

$$[H^+] = 10^{-4}\ \text{mol L}^{-1}$$

Therefore, $pH = -\log [H^+] = -\log 10^{-4} = -(-4)$

$= 4$

Example 6.8 : Calculate the pH of $1\text{x}10^{-4}$ molar solution of NaOH.

Solution: NaOH is a strong base and dissociates in its solutions as:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

One mole of NaOH would give one mole of OH- ions. Therefore

$[OH^-] = 1\text{x}10^{-4}$ molar

$pOH = -\log [OH^-] = -(\log 10^{-4}) = -(-4)$

$pOH = 4$

Since $pH = 14 - pOH$

$= 14 - 4$

$= 10$

CHECK YOUR PROGRESS 6.5

1. Aqueous solution of a substance does not conduct electricity through it. What type of substance is it?
2. A substance completely dissociates in to ions when dissolved in water. What type of substance is it?
3. X is a strong acid while Y is a weak acid. In whose aqueous solution a dynamic equilibrium will be established?
4. In an aqueous solution

   $$[H^+] = [OH^-]$$

   What type of solution is it, acidic, basic or neutral?
5. pH of a solution is 4. What is the hydrogen ion concentration in it?

LET US REVISE

- Based on the nature of chemical changes, reactions can be classified into five types (i) combination reactions, (ii) decomposition reactions, (iii) displacement reactions, (iv) double-displacement reactions, and (v) oxidation-reduction reactions.
- The reactions in which all the reactants and products are present in the same phase are called homogeneous reactions and the reactions in which reactants and products are present in different phases are called heterogeneous reactions.
- The reactions in which heat is evolved are called exothermic reactions and the reactions in which heat is absorbed are called endothermic reactions.
- The reactions that can occur in forward and reverse directions simultaneously under same set of conditions are called reversible reactions.
- A system is said to be in a state of equilibrium if none of its properties changes with time.

- One mole of a gas occupies a volume of 22.7 L at STP (standard temperature and pressure) i.e. at 273 K temperature and 1 bar pressure.
- An acid is a substance that furnishes hydrogen ions, H^+; a base is a substance that furnishes hydroxyl ions, OH^- in its solutions and a salt is produced when an acid and a base react with each other.
- An electrolyte conducts electric current through itself in the molten state or through its solution. If it dissociates completely it is known as a strong electrolyte and if it dissociates only partially it is known as a weak electrolyte.
- pH of a neutral solution is 7, that of an acidic solution is less than 7 and that of a basic solution is more than 7 at 25^0C

TERMINAL EXERCISES

A. Multiple choice type questions.

1. The reaction given below is:

   $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$

   (a) Combination reaction
   (b) Displacement reaction
   (c) Redox reaction
   (d) Displacement and redox reaction.

2. The reaction given below is *not* a:

   $C(s) + O_2(g) \longrightarrow CO_2(g)$

   (a) Heterogeneous reaction
   (b) Displacement reaction
   (c) Exothermic reaction
   (d) Redox reaction.

3. In the reaction

   $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$

   (a) 1 mole of $KClO_3$ produces 1.5 mole of O_2
   (b) 1 mole of $KClO_3$ produces 3 moles of O_2
   (c) 2 moles of $KClO_3$ produce 1 mole of KCl
   (d) when 1 mole of KCl is produced 3 moles of O_2 are produced

4. Which of the following statements about chemical equilibrium is not correct ?

   (a) It is dynamic equilibrium.
   (b) It can be established by a reversible reaction only.
   (c) It is established in any aqueous solution of a strong acid or a strong base.
   (d) On changing the temperature the equilibrium constant's value would also change.

5. pH of a solution is equal to

   (a) log $[H^+]$
   (b) - log $[H^+]$
   (c) log $[OH^-]$
   (d) - log $[OH^-]$

6. In which of the following reactions H_2O_2 acts as a reducing agent?
   (a) $H_2O_2 + 2KI \longrightarrow 2KOH + I_2$
   (b) $H_2O_2 + SO_2 \longrightarrow H_2SO_4$
   (c) $H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2$
   (d) $4H_2O_2 + PbS \longrightarrow PbSO_4 + 4H_2O$

B. Descriptive type questions.

1. Write electronic definitions of oxidation and reduction.
2. Give one example each of *slow* and *fast* reactions.
3. Give *any two* examples of quantitative information carried by a chemical equation.
4. What is an acid?
5. What is pH?
6. What is an exothermic reaction? Give one example.
7. Differentiate between displacement reactions and double displacement reactions.
8. What are weak electrolytes? Give one example.
9. In the reaction
   $$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$
   How much mass of Cl_2 is required to produce 1.5 moles of Br_2?
10. What is the pH of $5\text{x}10^{-4}$ molar solution of H_2SO_4?
11. In the reaction:
    $$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$
    Identify the species that is getting (i) reduced (ii) oxidized.
12. What is the difference between dynamic and static equilibrium? Give example of each.
13. NH_4OH is a weak base. Write down the equilibrium established in its aqueous solution and the expression of its dissociation constant K_b.
14. Given the following reaction
    $$2Al(s) + Fe_2O_3(s) \longrightarrow 2Fe + Al_2O_3(s)$$
    calculate the mass of Fe_2O_3 in grams required to produce 20.0 g of Fe. (Relative atomic masses: Fe = 55.8; O = 16).
15. Calculate the pH of (i) 10^{-5} mol L^{-1} HCl and (ii) 10^{-4} mol L^{-1} NaOH.
16. What are oxidation and reduction? Give one example with equation of a redox reaction. Identify the oxidizing agent and the reducing agent in it.
17. (i) What is a homogeneous reaction? Give one example each of gas phase and solution phase homogeneous reactions.
    (ii) What is a reversible reaction? Give one example.
18. In the reaction
    $$3C_3H_6 + 2KMnO_4 + 4H_2O \longrightarrow 3C_3H_8O_2 + 2KOH + 2MnO_2$$
    Calculate,
    (i) the number of moles of MnO_2 produced by 12 moles of C_3H_6.
    (ii) the number of moles of $KMnO_4$ needed to react with 0.006 moles of C_3H_6.
    (iii) the number of moles of $KMnO_4$ needed to produce 0.15 moles of $C_3H_8O_2$.
    (iv) the mass of C_3H_6 required to produce 5.6 grams of KOH. (Atomic mass of K = 39)

19. What is a neutralization reaction? A titration was started by taking 20 mL of 10^{-2} molar HCl. Then a solution of NaOH was gradually added from the burette. By mistake the student missed the end point and added excess of NaOH. When he finished the titration, the solution was 10^{-4} molar in NaOH. What was the pH of the solution present in the titration flask?
    (i) In the beginning of the titration
    (ii) at the end point when NaOH had just neutralized the HCl and
    (iii) at the end of the titration.
20. Sodium metal reacts with excess of water according to the reaction:
    $$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$
    (i) Calculate the mass of sodium required to produce 1 kg of NaOH.
    (ii) Find out the volume of H_2 evolved at STP when 1.012 kg of sodium reacts with excess of water.

ANSWERS TO CHECK YOUR PROGRESS

6.1
1. D
2. C
3. B
4. A
5. E

6.2
1. Heterogeneous
2. Endothermic
3. Irreversible
4. Fast
5. Non-equilibrium

6.3
1. T
2. F
3. T
4. T
5. F

6.4
1. Salt
2. Y
3. X

6.5
1. Non-electrolyte
2. Strong electrolyte
3. In solution of Y
4. Neutral
5. 10^{-4} mol L^{-1}

GLOSSARY

Acid: A substance containing hydrogen that furnishes hydrogen ions (H^+) in its solutions.

Base: A substance that furnishes hydroxyl ions, OH^- in its solutions.

Combination reaction: A reaction in which two or more substances react to form a new substance.

Decomposition reaction: A reaction in which one substance breaks down into two or more substances.

Displacement reaction: A reaction in which an ion present in a compound is displaced by another ion.

Double displacement reactions: The reactions in which two ionic compounds exchange their ions.

Electrolyte: A substance that conducts electric current through it in the molten state or through its solution.

Endothermic reactions: The reactions in which heat is absorbed

Equilibrium state: A state in which no property of system changes with time.

Exothermic reactions: The reactions in which heat is evolved

Heterogeneous reactions: Reactions in which reactants and products are present in more than one phase.

Homogeneous reaction: Reactions in which all the reactants and products are present in the same phase.

Molarity: It is the number of moles of a substance present in one litre volume.

Neutralization: The reaction between an acid and a base to produce salt and water.

Non-electrolyte: A substance that does not conduct electric current through it in the molten state or through its solution.

Oxidation: A process which involves loss of electrons.

pH: The negative logarithm of the hydrogen ion concentration.

Reduction: A process which involves gain of electrons.

Reversible reactions: The reactions that can occur in forward and reverse directions simultaneously under same set of conditions.

Salt: A substance produced by the reaction of an acid with a base along with water.

STP: Standard temperature and pressure i.e. when temperature is 273 K temperature and pressure is 1 bar.

Strong electrolytes: The electrolytes that dissociate completely in their solutions.

Synthesis reaction: The reaction in which a compound is formed by combination of its constituent elements.

Weak electrolytes: The electrolytes that dissociate only partially into ions in their aqueous solutions.