2 CHEMICAL CHANGES AND REACTIONS

SYLLABUS

(i) Types of chemical changes.
• *Direct combination* • *decomposition* • *displacement;* • *double decomposition*
(The above to be taught with suitable chemical equations as examples).
(ii) Energy changes in a chemical change.
Exothermic and endothermic reactions with examples – evolution/absorption of heat, light and electricity.

2.1 CHEMICAL REACTION

A chemical reaction is the process of breaking the *chemical bonds* of the reacting substances (reactants) and making new bonds to form new substances (products).

A chemical bond is the force that holds the atoms of a molecule together, as in a compound.

A chemical change or chemical reaction occurs when particles collide. Collisions occur when reactants are in close contact or by supply of energy. Thus, one or more of the following conditions are necessary for a chemical change :

(i) Mixing (close contact) : In some cases, a chemical reaction occurs when two substances are mixed in their solid states.

Example 1 : Iodine and phosphorus react explosively when brought into close contact.

2 : Lead nitrate (white) and potassium iodide (white) react to make lead iodide (yellow).

$$Pb(NO_3)_2\ (s) + 2KI\ (s) \rightarrow 2KNO_3\ (s) + PbI_2\ (s)$$

(ii) Solution : In some cases, a chemical reaction occurs when substances are mixed in, *i.e.* molten or aqueous state.

Example 1 : Oxalic acid crystals and sodium carbonate react in water solution only.

2 : Sodium chloride and silver nitrate also react in a solution state to form the precipitate of silver chloride and sodium nitrate.

$$NaCl\ (aq) + AgNO_3\ (aq) \rightarrow \underset{\text{white ppt.}}{AgCl\downarrow} + NaNO_3\ (aq)$$

(iii) Heat : Some chemical reactions occur only on heating.

Example 1 : Copper carbonate decomposes on heating (Δ symbol for heating) into copper oxide and carbon dioxide.

$$CuCO_3(s) \xrightarrow{\Delta} CuO\ (s) + CO_2(g)$$

2 : Lead nitrate decomposes on heating leaving yellow residue lead monoxide, brown gas nitrogen dioxide and colourless gas oxygen.

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

(iv) Light : Some chemical reactions take place by the action of light. These are called *photochemical reactions* or *photolysis*. Molecules of the reactants absorb light energy to get activated, and then react rapidly.

Example 1 : Plants form glucose from carbon dioxide and water in the presence of light.

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

2 : Hydrogen and chlorine react in the presence of sunlight.

$$H_2 + Cl_2 \xrightarrow{\Delta} 2HCl$$

3 : If chlorine water is exposed to sunlight, oxygen is evolved.

$$Cl_2 + H_2O \rightarrow HCl + HClO$$

$$2HClO \xrightarrow[\text{UV radiation}]{\text{Sunlight}} 2HCl + O_2$$

Note : Solutions of silver nitrate, hydrogen per oxide are kept in brown bottles in the laboratory because they decompose in the presence of light.

4 : Silver nitrate decomposes in the presence of light.

$$2AgNO_3 \xrightarrow{\text{sun light}} 2Ag + 2NO_2 + O_2$$

(v) Electricity : Certain chemical reactions, like the decomposition of certain compounds, occur only when electricity is passed through the substance.

Example 1 : Electrolysis of acidulated water occurs only in the presence of electricity, to give hydrogen and oxygen.

$$2H_2O \xrightarrow{\text{electricity}} 2H_2 + O_2$$

2 : On passing current through molten sodium chloride, sodium and chlorine are separately obtained.

$$2NaCl \xrightarrow{\text{electricity}} 2Na + Cl_2$$

(vi) Pressure : Some chemical reactions take place only when the involved substances are subjected to high pressure.

Example 1 : Mercuric chloride and potassium iodide when **rubbed** in a mortar, give a scarlet-coloured substance called mercuric iodide.

2 : Nitrogen and hydrogen, when subjected to high pressure, produce ammonia (in the presence of iron as catalyst)

$$N_2 + 3H_2 \underset{\text{above 200 atm}}{\rightleftharpoons} 2NH_3$$

(vii) Catalyst : *Some chemical reactions need a catalyst to accelerate or decelerate the rate at which they ocour.* The catalysts themselves do not take part in the reaction.

Example 1 : Potassium chlorate decomposes only at 700°C, and even then the rate of release of oxygen is very slow. But when potassium chlorate is heated in the presence of **manganese dioxide**, decomposition begins at a much lower temperature, 300°C, and manganese dioxide remains unaffected. Thus, in this reaction, manganese dioxide acts as a catalyst.

$$2KClO_3 \xrightarrow{MnO_2} 2KCl + 3O_2$$

2 : Ammonia reacts with oxygen to produce nitric oxide and water vapour in the presence of **(platinum)** a catalyst.

$$4NH_3 + 5O_2 \xrightarrow[800°C]{Pt} 4NO + 6H_2O$$

(a) Positive catalyst

When a catalyst *accelerates* a reaction, it is known as a *positive* catalyst.

For example, the rate of decomposition of hydrogen peroxide gets increased in the presence of ***manganese dioxide***.

$$2H_2O_2 \xrightarrow{MnO_2} 2H_2O + O_2$$

Promoters : Substances that influence the rate of a chemical reaction by improving the efficiency of the catalyst are called promoters.

For example, in the manufacture of ammonia, iron acts as a catalyst and molybdenum as a promoter. Molybdenum increases the efficiency of the catalyst iron.

$$N_2 + 3H_2 \underset{\text{molybdenum}}{\overset{\text{iron +}}{\rightleftharpoons}} 2NH_3$$

(b) Negative catalyst

A catalyst employed to *retard* a reaction is known as a *negative* catalyst.

Example 1 : ***Phosphoric acid*** retards the rate of decomposition of hydrogen peroxide.

2 : The rate of oxidation of chloroform decreases in the presence of ***alcohol***.

Note : Certain chemical reactions proceed by absorption of sound energy.

For example : Acetylene breaks up into carbon and hydrogen by absorbing sound energy

$$C_2H_2 \xrightarrow[\text{energy}]{\text{Sound}} 2C + H_2.$$

2.2 CHARACTERISTICS OF CHEMICAL REACTIONS

Certain chemical reactions are characterized by changes that are quite easily observed. Some of these typical changes are :

(i) Evolution of gas : In many chemical reactions, one of the products is a gas.

Examples :

(a) When zinc reacts with dilute sulphuric acid, hydrogen gas is evolved, with an effervescence.

$$\underset{\text{[zinc]}}{Zn} + \underset{\text{[dil. sulphuric acid]}}{H_2SO_4} \rightarrow \underset{\text{[zinc sulphate]}}{ZnSO_4} + \underset{\text{[hydrogen]}}{H_2}$$

Effervescence : The formation of gas bubbles in a liquid during a reaction is called *effervescence.* In the reaction given above, one of the reactants is a liquid. In such cases, *i.e.* when one of the reactants is a liquid, the gas produced forms bubbles in the liquid, *i.e.* effervescence takes place.

(ii) Change of colour : Certain chemical reactions are characterized by a change in the colour of the reactants.

Example

When a few pieces of iron are dropped into a blue coloured copper sulphate solution, the blue colour of the solution fades and eventually turns into light green due to the formation of ferrous sulphate.

Fe	+	$CuSO_4(aq)$	$\rightarrow$	$FeSO_4$	+	Cu
[iron]		[blue solution]		[green solution]		[copper] (red deposit)

(iii) Formation of precipitates : Certain chemical reactions are characterized by the formation of *insoluble solid substances called* ***precipitates***.

Examples

(a) When a solution of silver nitrate is added to a solution of sodium chloride, a white insoluble substance (precipitate), silver chloride, is formed.

$AgNO_3(aq)$	+	$NaCl(aq)$	$\rightarrow$	AgCl(ppt)	+	$NaNO_3(aq)$
[silver nitrate solution]		[sodium chloride solution]		[silver chloride] (white ppt)		[sodium nitrate solution]

(b) When ferrous sulphate solution is added to sodium hydroxide solution, a dirty green precipitate of ferrous hydroxide is formed.

$FeSO_4(aq)$	+	$2NaOH(aq)$	$\rightarrow$	$Fe(OH)_2\downarrow$	+	$Na_2SO_4(aq)$
[ferrous sulphate]		[sodium hydroxide]		[ferrous hydroxide] (dirty green ppt)		[sodium sulphate]

(iv) Change of state : In many chemical reactions, a change of state is observed. The reaction might start with gaseous or liquid reactants and end up with solid products, and vice versa.

Example

Ammonia gas reacts with hydrogen chloride gas to produce solid ammonium chloride.

$$NH_3(g) + HCl(g) \rightleftharpoons NH_4Cl(s)$$

EXCERCISE 2(A)

1. (a) What is a chemical reaction ?
   (b) State the conditions necessary for a chemical change or reaction.
2. Define the following terms
   (a) Chemical bond
   (b) Effervescence
   (c) Precipitate
3. Give an example of a reaction where the following are involved
   (a) Heat (b) Light
   (c) Electricity (d) Close contact
   (e) Solution (f) Pressure
   (g) Catalyst
4. Define :
   (a) Photochemical reaction
   (b) Electrochemical reaction.
   Give one example in each case.
5. Give an example of each of the following chemical changes.
   (a) A photochemical reaction involving
   (i) silver salt (ii) water
   (b) A reaction involving
   (i) blue solution
   (ii) formation of dirty green precipitate
   (c) Two gases combine to form white solid.
   (d) A reaction where colour change is noticed.
6. Write the chemical reaction where the following changes are observed.
   (a) Gas is evolved
   (b) Colour change is noticed
   (c) Precipitate is formed
   (d) Physical state is changed
7. Give reason for the following :
   (a) Silver nitrate solution is kept in coloured bottles.
   (b) Molybdenum is used in the manfacture of ammonia.
   (c) Blue solution of copper sulphate changes to green when a piece of iron is added to this solution.

2.3 TYPES OF CHEMICAL CHANGE OR CHEMICAL REACTION

1. Direct combination (or synthesis)
2. Decomposition
3. Displacement
4. Double decomposition

1. Direct combination or synthesis

A reaction in which two or more substances combine together to form a single substance is called a *combination reaction.*

$$A + B \rightarrow AB$$

In the above reaction, a combination of substances A and B takes place to give a molecule of a new substance, AB.

In combination reactions:

(i) *two elements combine to form a compound* [this reaction is also called **synthesis**].

- Carbon burns in oxygen to form a gaseous compound, carbon dioxide.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2} + \text{Heat}$$

- When magnesium is burnt, it combines directly with the oxygen of air to form magnesium oxide.

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

(ii) *an element and a compound combine to give a new compound.*

Example : Carbon monoxide, a compound, burns in the presence of oxygen, an element, to form a single product, carbon dioxide.

$$\underset{\text{[carbon monoxide]}}{2CO\,(g)} + \underset{\text{[oxygen]}}{O_2\,(g)} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{2CO_2\,(g)}$$

- Sulphur dioxide and oxygen combines under certain conditions to form sulphur trioxide.

$$2SO_2 + O_2 \rightarrow 2SO_3$$

(iii) *two or more compounds combine to form a single product.*

Example : Ammonia and hydrogen chloride, both compounds, combine to form a new compound, ammonium chloride.

$$\underset{\text{[ammonia]}}{NH_3\,(g)} + \underset{\text{[hydrogen chloride]}}{HCl\,(g)} \rightarrow \underset{\text{[ammonium chloride]}}{NH_4Cl\,(s)}$$

Experiments :

How to perform direct combination reactions.

(i) Take some black lead sulphide in a test tube and heat it. The black sulphide reacts with oxygen to form white lead sulphate.

$$PbS + 2O_2 \rightarrow PbSO_4$$

(combination)

(ii) Hold a piece of magnesium ribbon over a flame. It burns with a **dazzling light**, forming magnesium oxide.

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

2. Decomposition

It is the breaking up of a compound either into elements or simpler compounds, such that these products do not recombine to form the original compound.

Decomposition may occur in the presence of **heat or light**, or by the passage of an **electric current**.

> *A decomposition reaction that is brought about by heat is known as* **thermal decomposition.**

In a **decomposition reaction** :

(i) a compound breaks up into two or more elements.

Examples :

(a) The compound mercuric oxide, when heated, decomposes to form two elements, mercury and oxygen.

$$\underset{\text{[mercuric oxide]}}{2HgO\,(s)} \xrightarrow{\Delta} \underset{\text{[mercury]}}{2Hg\,(l)} + \underset{\text{[oxygen]}}{O_2\,(g)}$$

Note : The symbol Δ (delta) is used to signify that heat has caused the reaction.

(b) When **electric current** is passed through acidulated water, the latter decomposes into hydrogen and oxygen.

$$2H_2O\,(l) \xrightarrow[\text{current}]{\text{electric}} 2H_2\,(g) + O_2\,(g)$$

(ii) a compound can break up to form both elements and compounds.

Example : The compound potassium nitrate on heating decomposes to produce a compound, potassium nitrite, and an element, oxygen.

$$2KNO_3(s) \xrightarrow{heat} 2KNO_2(s) + O_2(g)$$

[potassium nitrate] [potassium nitrite] [oxygen]

(iii) a compound can break up to form two or more new compounds.

Example : The compound calcium carbonate on strong heating decomposes to form two compounds, calcium oxide and carbon dioxide.

$$CaCO_3(s) \xrightarrow[1000°C]{heat} CaO(s) + CO_2(g)$$

[calcium carbonate] [calcium oxide] [carbon dioxide]

Thermal decomposition of metal compound

(i) Metal hydroxide

Hydroxides

K, Na	Metal hydroxides are stable to heat.
Ca, Mg, Al, Zn, Fe, Pb, Cu	Decompose on heating to form metal oxide and water vapour. $Ca(OH)_2 \rightarrow CaO + H_2O$; $Zn(OH)_2 \rightarrow ZnO + H_2O$; $Pb(OH)_2 \rightarrow PbO + H_2O$; $Cu(OH)_2 \rightarrow CuO + H_2O$
Hg, Ag	Yields metal, oxygen and water vapour. $4AgOH \rightarrow 4Ag + O_2 + 2H_2O$

(ii) Metal carbonates

Carbonates

K, Na	Stable to heat and soluble in water.
Ca, Mg, Al, Zn, Fe, Pb, Cu	Decompose on heating with decreasing vigour to form the metal oxide and carbon dioxide. $MgCO_3 \rightarrow MgO + CO_2$; $ZnCO_3 \rightarrow ZnO + CO_2$; $CuCO_3 \rightarrow CuO + CO_2$
Hg, Ag	Forms the metal, oxygen and carbon dioxide. $2Ag_2CO_3 \rightarrow 4Ag + O_2 + 2CO_2$

(iii) Metal bicarbonates (metal hydrogen carbonate)

Metal bicarbonates or metal hydrogen carbonates decomposes to give metal carbonate, water vapour and carbon dioxide.

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

sodium hydrogen carbonate — sodium carbonate

$$Ca(HCO_3)_2 \xrightarrow{\Delta} CaCO_3 + H_2O + CO_2$$

calcium hydrogen carbonate — calcium carbonate

$$Mg(HCO_3)_2 \xrightarrow{\Delta} MgCO_3 + H_2O + CO_2$$

magnesium bicarbonate — magnesium carbonate

(iv) Metal nitrates

Nitrates

K, Na	On heating they melt and decompose to give metal nitrite and oxygen. $2KNO_3 \rightarrow 2KNO_2 + O_2$
Ca, Mg, Al, Zn, Fe, Pb, Cu	Decompose on heating to form the metal oxide, nitrogen dioxide and oxygen. $2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$; $2Zn(NO_3)_2 \rightarrow 2ZnO + 4NO_2 + O_2$; $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$; $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$
Hg, Ag	Forms metal, nitrogen dioxide and oxygen. $2AgNO_3 \xrightarrow{\Delta} 2Ag + 2NO_2 + O_2$

Decomposition reactions in our body

Digestion of food by our body is an example of a decomposition reaction.

The starch present in the food we eat decomposes into glucose and sugar. Proteins undergo decomposition to form amino acids. Fats and oils are decomposed to fatty acids and finally oxidized by respiration into carbon dioxide and water.

$$\text{Starch} \xrightarrow{\text{Enzymes}} \text{Glucose} \xrightarrow{[O]} CO_2 + H_2O$$

Experiments : *How to perform decomposition reactions.*

(i) Take some lead nitrate crystals in a test tube and heat them. The crystals first melt and, on further heating, give out both nitrogen dioxide, a **reddish brown gas**, and oxygen. A **yellow solid** (lead monoxide) is left behind in the test tube.

$$2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 \uparrow + O_2 \uparrow$$

(ii) Put some zinc carbonate in a test tube fitted with a cork and a bent glass tube. On heating, carbon dioxide is given out, which will turn lime water milky. *The residue, i.e. zinc oxide, is yellow when hot, but it turns white on cooling.*

$$ZnCO_3 \rightarrow ZnO + CO_2 \uparrow$$

(iii) Heat orange-coloured ammonium dichromate in a test tube. Upon heating, it swells and decomposes, evolving nitrogen and water vapours and a green solid, chromium oxide is left behind.

$$\underset{\text{ammonium dichromate}}{(NH_4)_2Cr_2O_7} \rightarrow \underset{\text{chromium oxide}}{Cr_2O_3} + 4H_2O \uparrow + N_2 \uparrow$$

(iv) When hydrated copper(II) sulphate is heated in a test tube, the blue-coloured crystals change into white anhydrous salt. The change may be represented by the following equation :

$$\underset{\substack{\text{hydrated salt}\\\text{(blue)}}}{CuSO_4 \cdot 5H_2O(s)} \rightarrow \underset{\substack{\text{anhydrous salt}\\\text{(colourless)}}}{CuSO_4(s)} + 5H_2O(g)$$

However, anhydrous copper(II) sulphate may be changed to the blue hydrated form by taking a sample of the anhydrous salt and adding water to it (this is the test to detect the presence of water).

$$\underset{\substack{\text{anhydrous salt}\\\text{(colourless)}}}{CuSO_4(s)} + 5H_2O(l) \rightarrow \underset{\substack{\text{hydrated salt}\\\text{(blue)}}}{CuSO_4 \cdot 5H_2O(g)}$$

So far, most of the reactions that we have considered proceed quite definitely in a certain direction, and it is possible to identify the *reactants* and the *products*. However, there also exists a group of reactions in which the direction of chemical change can be reversed by changing the conditions under which the reaction is taking place. Such reactions are called **reversible reactions**.

Thus, this is a reversible reaction, and the equation for the reaction is :

$$CuSO_4 \cdot 5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(g)$$

To show that the reaction is reversible, we put the sign "$\rightleftharpoons$" in between the reactants and the products.

If steam is passed over red hot iron, the latter is partially converted into magnetic oxide, and hydrogen is released. If, on the other hand, hydrogen is passed over the heated oxide, it partially changes back to steam.

$$3Fe + 4H_2O \rightleftharpoons Fe_3O_4 + 4H_2$$

In either case, the reaction ends with a mixture of reactants and products.

Experiments : *To show thermal dissociation*

(v) Heat some solid ammonium chloride in a test tube. Two colourless gases ammonia and hydrogen chloride are produced. As these gases move up at the upper part of the test tube whic is cooler, they combine to form ammonium chloride, which appears as a white sublimate on the cooler upper side of the test tube.

$$NH_4Cl \rightleftharpoons NH_3 \uparrow + HCl \uparrow$$

A simultaneous reversible decomposition reaction brought about only by heat is **thermal dissociation**.

Example of thermal dissociation. On heating, nitrogen tetraoxide changes to nitrogen dioxide, a reddish brown gas. On cooling, nitrogen dioxide changes into the original compound, nitrogen tetraoxide.

$$N_2O_4 \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} 2NO_2$$

3. Displacement

It is a chemical change in which a more active element displaces a less active element from its salt solution.

Experiments : *To show displacement reactions.*

(i) Take a solution of copper sulphate in a beaker, add a few pieces of zinc, and stir with a glass rod. The blue colour of the solution gradually fades, and soon the solution becomes colourless. At the same time, reddish brown particles of copper settle down in the beaker.

$$CuSO_4 + Zn \rightarrow ZnSO_4 + Cu \downarrow$$

(ii) In a test tube, take some dilute sulphuric acid and drop a small piece of magnesium ribbon. Brisk effervescence takes place, and hydrogen is evolved, which burns with a pop sound on bringing a burning match stick near the mouth of the test tube.

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$$

(iii) Pass chlorine gas through a solution of potassium iodide. The colourless solution turns yellow brown as iodine appears.

$$2KI + Cl_2 \rightarrow 2KCl + I_2$$

From the above examples, it can be noticed that in metal, zinc is more active (reactive) than copper, and magnesium is more active compared to hydrogen. In non-metals chlorine is more active as compared to iodine.

By taking similar examples, the following activity series can be prepared.

Activity (reactivity) of elements

Metals			Non metals		
Potassium	K	*Most active metal*			
Sodium	Na				
Calcium	Ca	*Reactivity decreases downwards*	Fluorine	F	*Most active*
Magnesium	Mg		Chlorine	Cl	
Aluminium	Al		Bromine	Br	
Zinc	Zn		Iodine	I	*Least active*
Iron	Fe				
Lead	Pb				
[Hydrogen]	[H]				
Copper	Cu				
Mercury	Hg				
Silver	Ag				
Gold	Au				
Platinum	Pt	*Least active (reactive) metal*			

Note : The more reactive element displaces the less reactive element from its salt solution.

4. Double decomposition

This is a type of chemical change in which two compounds in a solution react to form two new compounds by mutual exchange of radicals. Double decomposition reaction is also called double displacement reaction.

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

These reactions are of two types: **(a) precipitation** reactions and **(b) neutralization** reactions.

(a) *Precipitation reaction :* *A chemical reaction in which two compounds in their aqueous state react to form an insoluble salt (a precipitate) as one of the products is known as a precipitation reaction.*

For example :

$$BaCl_2\,(aq) + Na_2SO_4\,(aq) \rightarrow \underset{\text{white ppt.}}{BaSO_4(s)} + NaCl\,(aq)$$

$$CuSO_4\,(aq) + H_2S\,(g) \rightarrow \underset{\text{black ppt}}{CuS(s)} + H_2SO_4\,(aq)$$

Experiments : *To show double decomposition reactions.*

(i) Take a solution of silver nitrate in a test tube and add dilute hydrochloric acid or a solution of sodium chloride. A white, curdy precipitate is formed.

$$AgNO_3 + HCl \rightarrow AgCl\downarrow + HNO_3$$

$$AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_3$$

(ii) Fill one-third of a test tube with a solution of dilute sulphuric acid, and add to it a solution of barium chloride. A thick white precipitate is immediately formed.

$$BaCl_2 + H_2SO_4 \rightarrow BaSO_4\downarrow + 2HCl$$

(b) *Neutralization*: *The reaction between an acid and a base that forms salt and water only is referred to as reaction of neutralisation.*

The reaction takes place because the hydrogen ion(H^+) from the acid combines with the hydroxyl ion(OH^-) from the base, to form water.

$$NaOH + HCl \rightarrow NaCl + H_2O$$

OR

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

Ionic form

Cancelling the common ions, Na^+ and Cl^-, the only change is the combination of H^+ and OH^- ions to form un-ionized water,

i.e. $H^+ + OH^- \rightarrow H_2O$

Note : Double decomposition reactions may also occur with evolution of gas.

Example :

$$FeS(s) + H_2SO_4\,(aq) \rightarrow FeSO_4 + H_2S\uparrow$$

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

As a chemical process, neutralization has many uses

(a) When someone is stung by a bee, formic acid enters the skin and gives pain, which can be relieved by rubbing the spot with slaked lime or baking soda, both of which are bases.

(b) If our stomach glands secrete excess HCl, we experience pain, which is relieved by taking milk of magnesia or sodium hydrogen carbonate, both of which are bases. On the other hand, if there is a deficiency of HCl, some or the other suitable acid is taken in dilute form to make up for it.

(c) Acid that is accidentally spilled on to our clothes or body can be neutralized with ammonia solution.

(d) If the soil is somewhat acidic, and thus unfavourable for growing of certain crops, slaked lime is added to neutralize the excess acid.

Hydrolysis : *It is the process in which a salt and water react to form an acidic or a basic solution.*

In the process of hydrolysis, only those salts hydrolyse that are formed by the reaction of (i) strong base and weak acid and (ii) strong acid and weak base.

This happens because a salt formed due to a strong base and a weak acid, on dissolving in water, will make a basic solution. **A basic solution turns red litmus blue.**

$$Na_2CO_3 + 2H_2O \rightarrow \underset{\text{Strong base}}{2NaOH} + \underset{\text{Weak acid}}{H_2CO_3}$$

$$K_3PO_4 + 3H_2O \rightarrow \underset{\text{Strong base}}{3KOH} + \underset{\text{Weak acid}}{H_3PO_4}$$

But the salt formed due to a strong acid and a weak base, on dissolving in water, will make an acidic solution. **Acidic solutions turn blue litmus red.**

$$FeCl_3 + 3H_2O \rightarrow \underset{\text{Weak base}}{Fe(OH)_3} + \underset{\text{Strong acid}}{3HCl}$$

$$ZnSO_4 + 2H_2O \rightarrow \underset{\text{Weak base}}{Zn(OH)_2} + \underset{\text{Strong acid}}{H_2SO_4}$$

Note : Strong acids and strong bases ionize to a large extent in aqueous solution. Weak acids and weak bases ionize to only a small extent in aqueous solution, and the salts formed by them usually do not hydrolyse.

EXCERCISE 2(B)

1. Complete the following statements.
   (a) The chemical change involving iron and hydrochloric acid illustrates a .... reaction.
   (b) In the type of reaction called ...., two compounds exchange their positive and negative radicals.
   (c) A catalyst either .... or .... the rate of a chemical change but itself remains .... at the end of the reaction.
2. When hydrogen burns in oxygen, water is formed; when electricity is passed through water, hydrogen and oxygen are given out. Name the type of chemical change involved in the two cases.
3. Explain, giving one example for each of the following chemical changes:
   (a) Double decomposition (b) Thermal dissociation
   (c) Reversible reaction (d) Displacement
4. What is synthesis ? What kind of chemical reaction is synthesis ? Support your answer by an example.
5. Decomposition brought about by heat is known as thermal decomposition. What is the difference between thermal dissociation and thermal decomposition.
6. Define neutralization reaction. Give three applications of neutralization reactions.
7. What do you understand by hydrolysis ? Explain giving examples.
8. Iron (III) chloride is acidic while sodium carbonate is basic. Explain.
9. What is decomposition ? Support your answer by an example.
10. State the type of reactions each of the following represent and balance the ones that are not balanced.
    (a) $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$
    (b) $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$
    (c) $2HgO \rightarrow 2Hg + O_2$
    (d) $PbO_2 + SO_2 \rightarrow PbSO_4$
    (e) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
    (f) $2KClO_3 \rightarrow 2KCl + 3O_2$
    (g) $2H_2O_2 \rightarrow 2H_2O + O_2$
    (h) $KNO_3 + H_2SO_4 \rightarrow HNO_3 + KHSO_4$
    (i) $CuO + H_2 \rightarrow Cu + H_2O$.
    (j) $CaCO_3 \rightarrow CaO + CO_2$
    (k) $NH_4Cl \rightarrow NH_3 + HCl$

2.4 ENERGY CHANGE IN CHEMICAL REACTIONS

In every chemical change or chemical reaction, change in energy is involved, *i.e.* there is a difference between the chemical energies of the reactants and the products. This energy can be in the form of heat, light, sound and electricity.

> *Every substance has a fixed amount of stored energy, which is in the form of potential energy, and is referred to as its chemical energy.*

A chemical reaction involves the breaking up of chemical bonds between atoms resulting in absorption of energy in the form of heat, and simultaneous formation of bonds with release of energy. These two types of energy are different from each other, *i.e.* there is either a surplus or a deficit of energy during the reaction. Therefore, in a chemical reaction, energy is either absorbed or released.

Depending upon the energy released or absorbed, chemical changes or chemical reactions are of two types :

1. Exothermic
2. Endothermic

1. *Exothermic reaction* : A chemical reaction in which heat (a form of energy) is given out is called *exothermic reaction.* It causes a rise in temperature.

Examples :

(a) When carbon burns in oxygen to form carbon dioxide, a lot of heat is produced.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \rightarrow \underset{\text{[carbon dioxide]}}{CO_2} + \text{Heat}$$

(b) When water is added to quicklime, a lot of heat energy is produced [along with alkaline calcium hydroxide (slaked lime)].

$$\underset{\text{[quicklime]}}{CaO} + H_2O \rightarrow \underset{\text{[slaked lime]}}{Ca(OH)_2} + \text{Heat}$$

(c) Formation of water : When hydrogen is burnt in oxygen, water is formed and heat is released.

$$2H_2 + O_2 \xrightarrow{\Delta} 2H_2O + \text{Heat}$$

(d) Formation of ammonia : Nitrogen reacts with hydrogen in the presence of catalyst (finely divided) iron at 450°C to 500°C, and above 200 atmospheres of pressure to form ammonia.

$$N_2 + 3H_2 \xrightarrow{\Delta} 2NH_3 + \text{Heat}$$

Note : The sign Δ indicates heat.

Respiration, rusting, and burning of coal, petrol, kerosene, etc., are some common exothermic reactions.

2. *Endothermic reaction* : A chemical reaction in which heat is absorbed is called *endothermic reaction*. It causes a fall in temperature.

This reaction cannot be sustained without supply of energy from an external source.

Examples :

(a) Formation of carbon disulphide : When carbon is heated with sulphur at high temperature, liquid carbon disulphide is formed.

$$C + 2S + \xrightarrow{\Delta} CS_2$$

(b) When nitrogen and oxygen are heated together to a temperature of about 3000°C, nitric oxide gas is formed. This is an endothermic reaction.

$$N_2 + O_2 + \xrightarrow[(3000°C)]{\text{Heat}} 2NO\ (g)$$

(c) Calcium carbonate decomposes into carbon dioxide and calcium oxide when it is heated to a temperature of about 1000°C; this is also an endothermic reaction.

$$\underset{\text{[calcium carbonate]}}{CaCO_3\ (s)} + \underset{(1000°C)}{\text{heat}} \rightarrow \underset{\text{[calcium oxide]}}{CaO\ (s)} + \underset{\text{[carbon-dioxide]}}{CO_2\ (g)}$$

PHOTOCHEMICAL REACTION

> *It is a reaction that occurs with absorption of light energy.*

Examples :

1. Photosynthesis

$$6CO_2 + 12\ H_2O \xrightarrow[\text{sunlight}]{\text{chlorophyll}} C_6H_{12}O_6 + 6\ H_2O + 6O_2$$

2. $2AgNO_3 \xrightarrow{\text{sunlight}} 2Ag + 2NO_2 + O_2$

ELECTROCHEMICAL REACTION

> *It is a reaction that occurs with absorption of electrical energy.*

Examples :

1. Fused potassium chloride, on passing current through it, breaks into charged particles (ions) of potassium and chloride..

$$KCl \xrightarrow[\text{current}]{\text{electric}} K^+ + Cl^-$$

2. Acidulated water breaks into hydrogen and oxygen.

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

CHAPTER AT A GLANCE

- Whenever energy is applied to matter, changes occurs. Changes are mainly of two types : physical and chemical.
- A change in which only the physical properties of a substance get changed, not its chemical composition, is called a physical change, viz. breaking of glass.
- A physical change is a temporary change and is reversible in nature.
- During a physical change, only the physical properties, such as state, colour, shape, etc. undergo a change.
- A change in which chemical composition and chemical properties of reacting substances undergo a change is called a chemical change, viz. burning of paper.
- A chemical change is a permanent change, *i.e.* it is irreversible.
- Changes occur by (i) mixing the substances in solution state (ii) by heat (iii) light (iv) electricity (v) pressure (vi) in the presence of a catalyst.
- A chemical reaction is the process of breaking the chemical bonds of reactants and formation of new bonds to form new products.
- Chemical reaction is confirmed by (i) evolution of gas (ii) change of colour (iii) formation of precipitate (iv) change of state.
- A reaction in which two or more substances combine to form a new substance is called a **combination reaction** : $Fe + S \rightarrow FeS$.
- A reaction in which a substance is broken down into two or more substances is called a **decomposition reaction** : $2HgO \rightarrow 2Hg + O_2$.
- A reaction in which one part of a molecule is replaced by another is called a **displacement reaction** : $CuSO_4 + Fe \rightarrow FeSO_4 + Cu$.
- A reaction in which two reacting molecules exchange their corresponding ions, is called a **double displacement reaction** : $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$.
- A reaction in which one of the products formed is an insoluble substance that is thrown out of the solution as solid (precipitate) is called **precipitation reaction** : $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$.
- A reaction in which heat is liberated (or given out) is called an **exothermic reaction**.
- A reaction in which heat is absorbed is called an **endothermic reaction**.

EXCERCISE 2(C)

1. State the main characteristics of chemical reactions. Give at least one example in each case.
2. Define exothermic and endothermic changes. Give two examples in each case.
3. State the effects of endothermic and exothermic reactions on the surroundings.
4. Give an example of a reaction where the following are involved
   (a) Evolution of heat
   (b) Absorption of heat
   (c) High pressure is required
5. Define :
   (a) Photochemical reaction
   (b) Electrochemical reaction.
   Give one example in each case.
6. Give an example of each of the following chemical changes.
   (a) A reaction involving
      (i) change of state
      (ii) formation of precipitate
   (b) An exothermic and an endothermic reaction involving carbon as one of the reactants.
   (c) A reaction where colour change is noticed.
7. What do you understand by 'chemical reaction.' ?
8. Complete and balance the following reactions :
   (a) $NaCl\ (aq) + AgNO_3\ (aq) \rightarrow$
   (b) $Pb(NO_3)_2 + KI \rightarrow$
   (c) $CuCO_3 \xrightarrow{\Delta}$
   (d) $Pb(NO_3)_2 \xrightarrow{\Delta}$
   (e) $NH_3 + O_2 \xrightarrow{Pt}$
9. What do you observe. When
   (a) Lead nitrate is heated.
   (b) Chlorine water is exposed to sunlight.
   (c) Hydrogen peroxide is exposed sunlight.
   (d) H_2S gas is passed through copper sulphate solution.
   (e) Barium chloride is added to sodium sulphate solution.
10. Name :
   (a) a carbonate which do not decompose on heating.
   (b) a nitrate which produces oxygen as the only gas.
   (c) a compound which produces carbon dioxide on heating
   (d) a nitrate which produces brown gas on heating.