



CHAPTER

4

TYPES OF CHEMICAL REACTION

In previous classes, we have learnt about physical and chemical changes. Changes which occur in physical properties such as states of matter, shape, size, colour, texture, electrical and magnetic condition etc, of matter without bringing any change in composition are called physical changes. Evaporation of water, melting of ice, dissolving sugar in water, breaking of wood into smaller pieces, magnetization of iron, glowing of electric bulb are examples of physical changes.

There are also many changes which results in the formation of new substances whose composition and properties are entirely different from those of the original substances. Let us consider the following changes occurring around us.

- (i) Burning of paper producing ash, carbon dioxide and water vapour
- (ii) Burning of candle wax and fuels like kerosene oil, petrol etc.
- (iii) Rusting of iron rod, nail etc producing rust (hydrated iron oxide)
- (iv) Souring of milk (formation of curd)
- (v) Cooking of food
- (vi) Digestion of food in our body
- (vii) Taking of antacids when one gets stomach burn/ pain.

In all these changes, the original substances lose their nature and identity and form new substances. For example, when the paper is burnt, the products such as the ash, carbon dioxide, water vapour etc are entirely different from paper. Rust is a different substance from iron. These are examples of **chemical changes**. **Whenever such a chemical change occurs, we say that a chemical reaction has taken place.** In other words, a chemical change takes place as a result of chemical reaction. In a chemical reaction, the substances originally present and taking part in the reaction are called **reactants** and the new substances formed are called **products**. For example, in the burning of

paper, the paper (cellulose) and air (oxygen) are the reactants and the ash, carbon dioxide and water vapour are the **products**.

Depending upon the nature of the reaction and the chemical change taking place, we can classify chemical reactions into various types.

4.1. TYPES OF CHEMICAL REACTIONS

We have learnt in class IX, that in a chemical reaction, atoms can neither be created nor destroyed as given by the Law of Conservation of Mass. During a chemical reaction atoms of one element do not change into those of another element. Nor do atoms disappear from the mixture or appear from elsewhere. We have also learnt in the previous chapter that atoms in molecules are held together by a force of attraction called **bond**. **In chemical reactions, the bonds present in the reactants are broken and new bonds are formed to produce new substances or products.** That is, there take place rearrangements or regroupings of atoms in various ways to give products.

4.1.1. Combination reactions

Activity 4.1.

[This activity is better demonstrated by the teacher. Students should not see directly to the white dazzling light produced on burning magnesium ribbon as it is harmful to the eye]

- ☞ Take a strip of magnesium ribbon 2-3 cm long and clean it by rubbing with sand paper.
- ☞ Hold it with a pair of tongs in the flame of a spirit lamp or burner. (Do not observe directly the dazzling white light produced).
- ☞ Collect the ash formed in a watch glass. (Fig. 4.1.).
- ☞ What do you observe ?

It is observed that when heated strongly magnesium burns in air with dazzling white light. The white powder produced can be collected in a china dish. The white powder is magnesium oxide. It is formed due to the reaction between magnesium and oxygen present in the air.

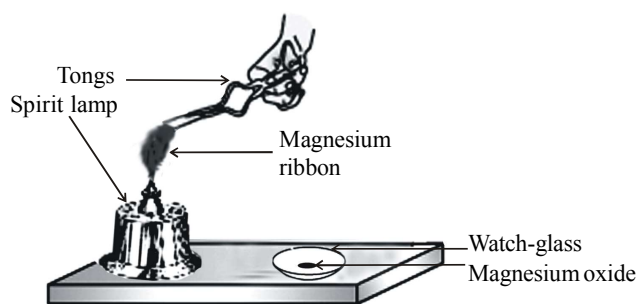
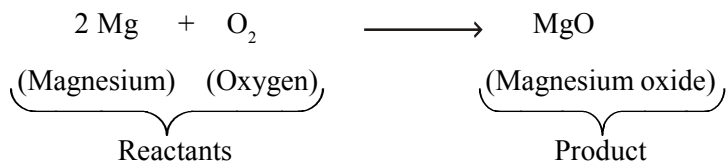
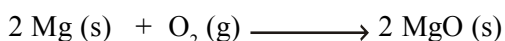


Fig.4.1. Burning of magnesium ribbon to produce magnesium oxide

The reaction can be represented by a chemical equation as follows :



Showing the physical states of the reactants and products, we may write the equation as follows :



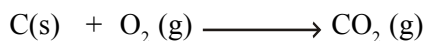
The gases, liquid, aqueous and solid states of reactants and products are represented by the notations (g), (l), (aq) and (s) respectively.

A reaction in which a single product is formed from two or more reactants is known as a combination reaction.

The reaction of magnesium and oxygen to form magnesium oxide is an example of combination reaction between two elements.

Let us discuss some more examples of this type.

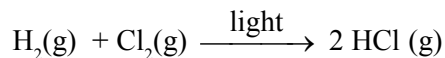
- (i) Carbon or coke burns in air to produce carbon dioxide gas.



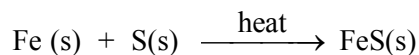
- (ii) Hydrogen burns in oxygen to form water.



- (iii) Hydrogen combines with chlorine in the presence of light to form hydrogen chloride.



- (iv) Iron and sulphur combine on heating to form iron sulphide.



- (v) Nitrogen combines with hydrogen to form ammonia gas.



Combination reactions also take place between two compounds to form a single product. Let us perform the following activity :

Activity 4.2.

- ☞ Take a small lump of calcium oxide (quick lime) in a beaker.
- ☞ Add water slowly to the lump of calcium oxide.
- ☞ Observe what happens.
- ☞ Touch the side of the beaker as shown in Fig 4.2.
- ☞ What do you feel ? Hot or cold ?

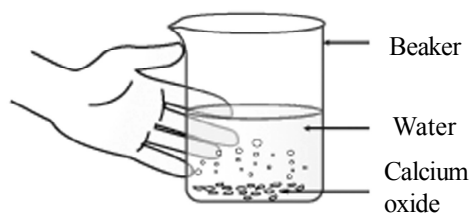


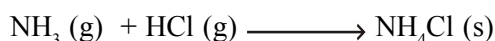
Fig. 4.2. Formation of slaked lime by the reaction of calcium oxide with water

When water is added to quick lime, they react vigorously to produce calcium hydroxide called slaked lime. A large amount of heat is evolved. The reaction can be represented by the following equation.

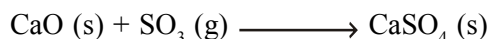


A few more examples are given below.

- (i) Ammonia combines with hydrogen chloride to form ammonium chloride.



- (ii) Calcium oxide reacts with sulphur trioxide to form calcium sulphate

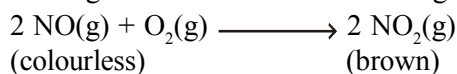


- (iii) Sodium oxide reacts with water to form sodium hydroxide

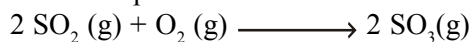


Combination reaction also takes place between an element and a compound to form a new compound.

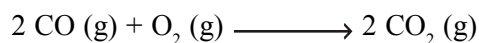
- (i) Nitric oxide (a colourless gas) combines with oxygen of the air to form nitrogen dioxide which is a brown gas.



- (ii) Sulphur dioxide reacts with oxygen (in the presence of catalyst) to form sulphur trioxide.



- (iii) Carbon monoxide gas combines with oxygen of air to form carbon dioxide gas.



4.1.2. Decomposition Reactions

The reactions in which a compound is broken down into two or more simpler substances are known as **decomposition reactions**. Thus, these reactions are just opposite of combination reactions. Let us perform the following activity.

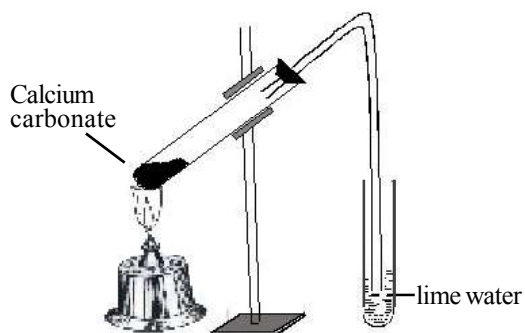


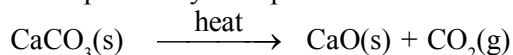
Fig.4.3. Heating of calcium carbonate and giving out carbon dioxide gas which turns lime water milky

Activity 4.3.

- ☞ Take about 4 g calcium carbonate in a hard glass test tube fitted with a delivery tube dipped in a test tube containing lime water as shown in Fig. 4.3.
- ☞ Heat the tube over the flame of a spirit lamp or burner.
- ☞ Observe the change in lime water

It is observed that lime water turns milky. When calcium carbonate is heated, it decomposes to carbon-dioxide gas.

The reaction that take place may be represented as follows.



Carbon dioxide which comes out turns lime water milky

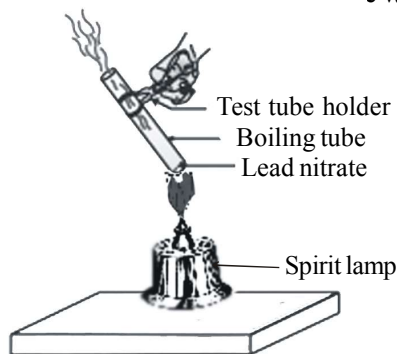
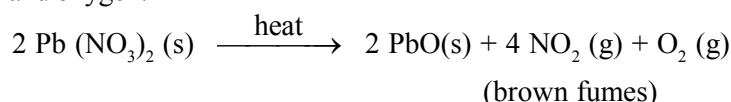


Fig.4.4. Decomposition of lead nitrate on heating

Activity 4.4.

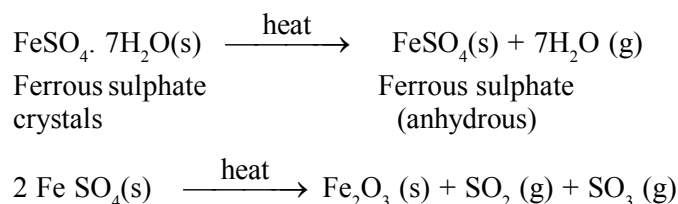
- ☞ Take about 2 g of powdered lead nitrate in a test tube.
- ☞ Hold it in the test tube holder.
- ☞ Heat it over a flame, as shown in Fig. 4.4.
- ☞ Observe the brown fumes evolved.
- ☞ Note down the changes observed.

When lead nitrate is heated, it breaks down into lead monoxide, nitrogen dioxide and oxygen.



The colourless lead nitrate becomes yellow due to the formation of lead monoxide (PbO). Brown fumes are observed due to the evolution of nitrogen dioxide (NO₂) gas.

Similarly, when we heat green coloured crystals of ferrous sulphate in a dry test tube, we observe that the green colour of ferrous sulphate becomes reddish brown due to the formation of iron (III) oxide. Sulphur dioxide and sulphur trioxide gases are also produced.



The reactions given above are examples of decomposition reactions caused by **heat**. They are called **thermal decomposition reactions**.

Decomposition reactions may also take place when electric current is passed through the compound in the molten state or in aqueous solution (decomposition by electricity). In certain cases, light also causes decomposition reactions. They are called *photochemical reactions*.

Decomposition by Electricity

Activity 4.5.

- ☞ Take a water voltameter with platinum electrodes.
- ☞ Pour sufficient water to the voltameter so that the platinum electrodes are well below the level of water.
- ☞ Fill two test tubes with water and close its mouth with your finger or thumb.
- ☞ Bring the test tubes with water upside down in the voltameter so that its mouth is under water. Then release your finger.
- ☞ Invert the test tubes with water over the two platinum electrodes so that the electrodes remain inside the tubes. (Fig 4.5).

- ☞ Now connect the electrodes to a 6 volt battery and see for bubbles of gas at both electrodes. If none appears, add a few mL of dilute sulphuric acid solution.
- ☞ Observe the formation of bubbles at both the electrodes. The gas in the bubbles displaces water in the test tubes.
- ☞ Compare the volume of the gases collected in both the test tubes.
- ☞ Test the gas over the anode with a glowing splinter.
- ☞ Try to light the gas collected over the cathode. Always bring out the gas collected over the cathode with mouth of the test tube downwards.
- ☞ What happens in each case ?
- ☞ Which gas is present in each test tube ?

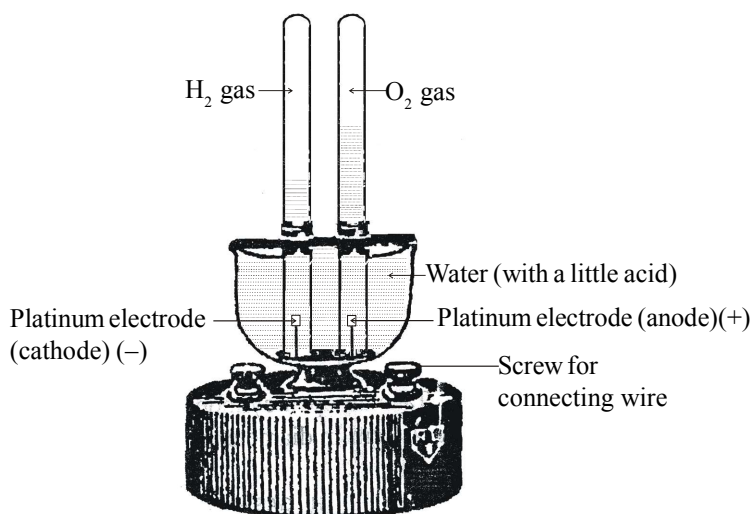
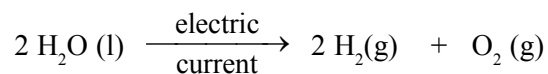


Fig. 4.5. Electrolysis of water

The following results are observed :

- (i) Volume of gas collected over the cathode is double that of the gas collected over the anode.
- (ii) The gas collected over the cathode burns with a 'pop' sound. It is hydrogen gas.
- (iii) The glowing splinter introduced into the gas collected over the anode burst into flame. The gas is oxygen.

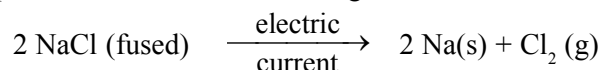
The above experiment shows that on passing electric current, water decomposes into hydrogen and oxygen according to the reaction :



Some more examples of decomposition reactions caused by passing electric current are given below.

(i) Electrolytic decomposition of molten sodium chloride

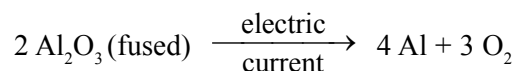
On passing electric current through molten sodium chloride, it decomposes to sodium and chlorine gas.



The reaction is used to obtain sodium metal on a large scale. Chlorine gas is obtained as a by-product.

(ii) Electrolytic decomposition of molten alumina (Al_2O_3)

When electric current is passed through molten alumina mixed with a little cryolite (it lowers melting point of alumina), alumina decomposes to give aluminium metal and oxygen gas.



Decomposition by Light (Photochemical decomposition)

There are some reactions which take place in the presence of light. These are called **photochemical reactions**.

Activity 4.6

- ☞ Take a pinch of silver chloride or silver bromide on a watch glass.
- ☞ Observe its colour.
- ☞ Keep the watch glass and its contents in sunlight for some time.
- ☞ Observe the change in the colour of the silver chloride in the sunlight. (Fig. 4.6).

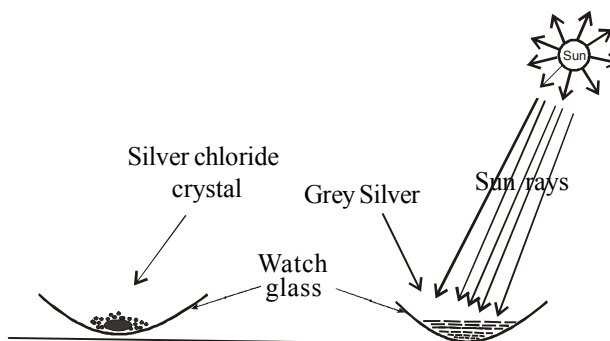
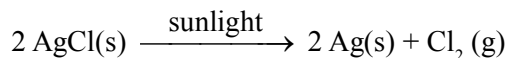


Fig. 4.6. Silver chloride turns grey in sunlight forming silver metal

It is observed that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.



Silver bromide behaves similarly,



Black and white photography is based upon these reactions. The photographic film is coated with these salts.

Exothermic and Endothermic Reactions

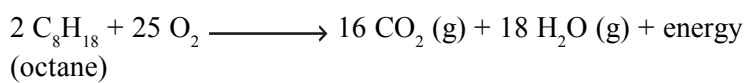
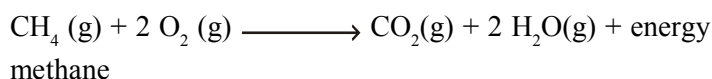
Every chemical reaction involves a change in energy. Either the reaction takes place and energy is liberated to the adjacent surroundings, or the reaction takes place with the absorption of energy i.e. energy is absorbed from the adjacent surroundings. The first type of reaction is said to be **exothermic** (heat comes out), and the second type **endothermic** (heat goes in). All chemical reactions are in one of these two categories.

As the root form **thermic** indicates, the original definition of these two terms was based only on heat energy. Today, however, *exothermic reactions refer to reactions which take place with the evolution of energy in any form — heat, light, electrical, mechanical, etc. Endothermic reactions refer to reactions which are accompanied by the absorption of any form of energy.*

In activity 4.2, we have observed that a large amount of heat is evolved in the reaction. This makes the reaction mixture warm. Hence, the reaction of calcium oxide with water to form calcium hydroxide or slaked lime is an example of exothermic reaction.

Other examples of exothermic reactions are :

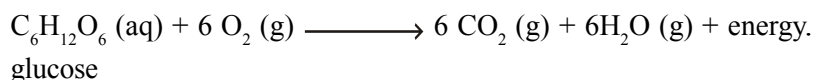
- (i) Burning of natural gas (methane) and, other fuels.



Respiration is an Exothermic Process

We get energy from the food we eat. The food that we eat mainly contains starch (carbohydrate) from rice, potato, bread etc and proteins from meat, egg, milk, pulses etc. During digestion, food materials are broken down into simpler substances. Starch decomposes into simple sugars such as glucose. Similarly, proteins decompose to form simpler molecules called amino acids. The glucose combines with oxygen in the cells of our body and provides energy. The process is called respiration.

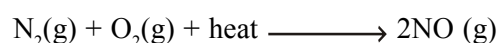
The reaction can be represented as follows :



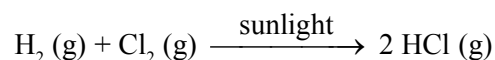
Decomposition Reactions are Endothermic Reactions

It is observed that decomposition reactions require energy in the form of **heat, light** or **electricity** for breaking down the reactants. As already learnt, those reactions in which energy is absorbed are called **endothermic reactions**.

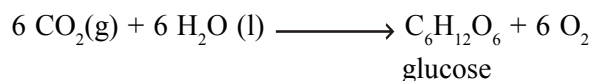
However, there are also certain combination reactions in which heat is absorbed. For example, combination of nitrogen and oxygen to form nitric oxide (NO) takes place with the absorption of heat.



A mixture of hydrogen and chlorine when exposed to sunlight forms hydrogen chloride.



Photosynthesis is also an endothermic reaction. Chlorophyll present in green plants absorbs light and helps in the conversion of carbon dioxide and water into glucose



Let us answer these.

1. A solution of an oxide 'X' is used for white washing.
(i) What could be 'X'? Write its formula.
(ii) Write the reaction of substance 'X' named in (i) above with water.
2. Why is the amount of gas collected in one of the test tubes in activity 4.5 double of the amount collected in the other ? Name this gas.

4.1.3. Displacement Reactions

Activity 4.7.

- ☞ Take two iron nails and clean them by rubbing with sand paper.
- ☞ Take two clean test tubes and mark A and B.
- ☞ Take about 10 mL of copper sulphate solution in each of them.
- ☞ Tie a thread on one iron nail and suspend it in test tube A. (Fig 4.7)

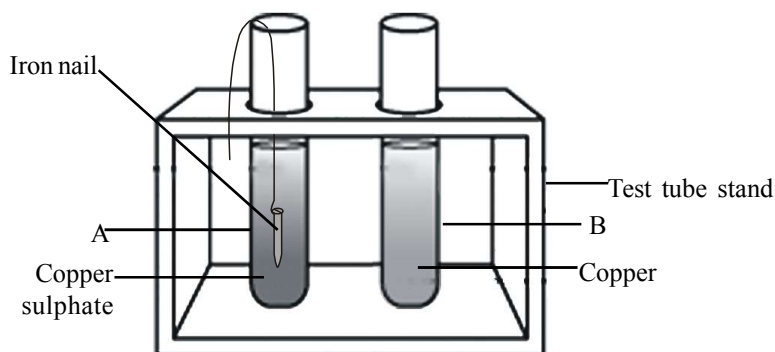
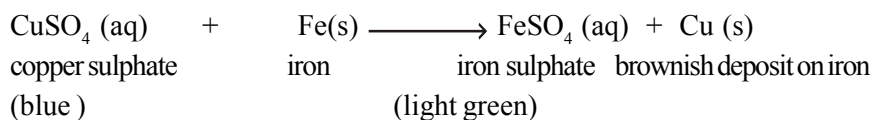


Fig. 4.7. Iron displaces copper from Copper sulphate solution

- ☞ Wait for about 30 minutes.
- ☞ Take out the nail from the solution.
- ☞ Observe the change in the colour of the solution in test tube A and compare it with solution in test tube B.
- ☞ Also, compare the colour of the iron nail that was suspended in the test tube with that of other nail.

It is observed that the blue colour of copper sulphate fades and iron nail becomes brownish in colour. These changes show that the following reaction takes place.



Some more examples of displacement reactions are :

- (i) When a piece of zinc is added to dilute sulphuric acid or dilute hydrochloric acid, hydrogen gas is evolved. The following reactions take place.



- (ii) Zinc and magnesium displace copper from copper sulphate solution.



- (iii) Copper displaces silver from silver nitrate solution



- (iv) When chlorine water is added to potassium iodide solution, iodine is liberated.



Reactivity Series Of Metals Or Activity Series Of Metals

We have learnt that iron can displace copper from its solution as follows.



However, it may be noted that copper cannot displace iron from its solution. This means that the following reaction does not take place.



This raises an important question, “**which metal can displace another metal from its salt solution**”?

We know that some metals are highly reactive while some others are less reactive. Based on the study of their reactions, metals can be arranged in order of their reactivities. **The arrangement of metals in the decreasing order of their reactivity is called reactivity series or activity series of metals. (Table 4.1).**

The metal which is higher in the reactivity series is more reactive than metals below it. Hence, in general, a more reactive metal (placed higher in the activity series) can displace the less reactive metal from its solution. Thus, iron displaces copper (less reactive) from its solution as



But copper cannot displace iron from FeSO_4 solution.



However, copper can displace silver (less reactive) from its solution.

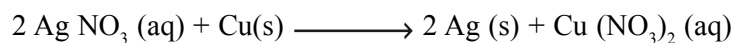
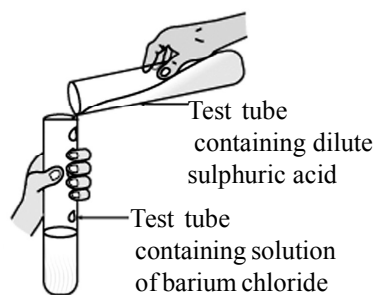


Table 4.1 Activity series of metals

<u>Element</u>	<u>Symbol</u>	
Potassium	K	(Most reactive)
Barium	Ba	
Calcium	Ca	
Sodium	Na	
Magnesium	Mg	
Aluminium	Al	
		Metals more reactive than hydrogen
Zinc	Zn	
Iron	Fe	
Cobalt	Co	
Nickel	Ni	
Tin	Sn	
Lead	Pb	
Hydrogen	H	
Copper	Cu	
Mercury	Hg	
Silver	Ag	
Gold	Au	
Platinum	Pt	(Least reactive)

4.1.4. Double Displacement Reactions**Activity 4.8.**

- ☞ Take about 3 mL of barium chloride solution in a test tube.
- ☞ In another test tube, take about 3 mL of dilute sulphuric acid.
- ☞ Mix the two solutions.
- ☞ Observe what happens.

**Fig 4.8. Formation of white substance**

We observe that a white substance which is insoluble in water is formed. This insoluble substance is known as a **precipitate**. The white precipitate is barium sulphate which is formed due to the following reaction :



The reactions which produce precipitates are also called **precipitation reactions**. What causes this ? The white precipitate of BaSO_4 is formed by the combination of SO_4^{2-} ion with Ba^{2+} ion. The other product formed is hydrogen chloride which is in solution.

Some more examples of double displacement reactions are given below.

(i) When silver nitrate solution is mixed with sodium chloride solution, a white precipitate of silver chloride is formed.



(ii) When dilute sulphuric acid is added to iron (ferrous) sulphide, a double displacement reaction takes place with the evolution of hydrogen sulphide gas ; ferrous sulphate formed remains in solution.



(iii) When hydrogen sulphide gas is passed through a blue solution of copper sulphate, a black precipitate of copper sulphide is formed. Sulphuric acid formed remains in solution.



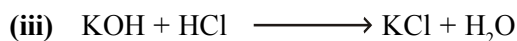
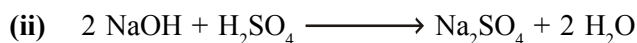
Neutralisation Reactions

A double displacement reaction also takes place when an acid reacts with a base to form salt and water



These reactions are called neutralisation reactions.

A few examples are given below:



Oxidation and Reduction Reactions (Redox reactions)**Activity 4.9.**

- ☞ Heat about 1g of copper powder in a China dish (Fig 4.9)
- ☞ Observe what happens.

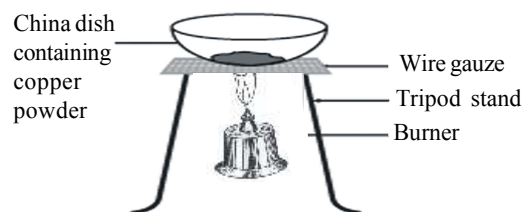
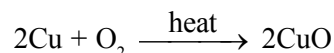


Fig 4.9. Oxidation of copper to copper oxide

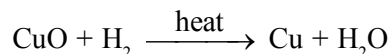
It is observed that reddish brown copper powder becomes coated with a black substance. How is this black substance formed ?

This is because when copper is heated, it combines with oxygen and black copper oxide is formed.



We say that copper has been oxidized to copper (II) oxide.

If hydrogen gas is passed over this heated material of copper (II) oxide, the black coating on the surface turns brown. The reverse reaction takes place and copper is obtained.

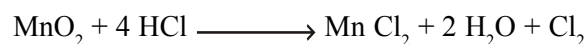


We say that copper oxide has been reduced to copper.

If a substance gains oxygen during a reaction, it is said to be oxidised. If a substance loses oxygen during a reaction, it is said to be reduced.

During the above reaction, the copper (II) oxide is losing oxygen and is being reduced. The hydrogen is gaining oxygen and is being oxidised. In other words, one reactant gets oxidised while the other gets reduced during the reaction. Thus oxidation and reduction occur simultaneously. Such reactions in which oxidation and reduction occur simultaneously are called **oxidation - reduction reactions** or **redox reactions**.

Let us consider the following reaction.



In this reaction, HCl is oxidised to Cl_2 whereas MnO_2 is reduced to

MnCl_2 . From this example, we can say that **if a substance gains oxygen or loses hydrogen during a reaction, it is said to be oxidised. If a substance loses oxygen or gains hydrogen during a reaction, it is said to be reduced.** The substance which oxidises other substances but itself undergoes reduction is called **oxidising agent**. The substance which reduces other substances but itself undergoes oxidation is called the **reducing agent**. In the above example, MnO_2 undergoes reduction and oxidises HCl . Hence MnO_2 is the oxidising agent, while HCl is the reducing agent.

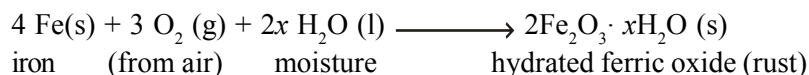
Some more examples of redox reactions are given below.



According to electronic concept, **oxidation is a process which involves loss of one or more electrons. Reduction is a process which involve gain of electrons.** You will learn about these in higher classes.

Oxidation-Reduction Reactions in Everyday Life

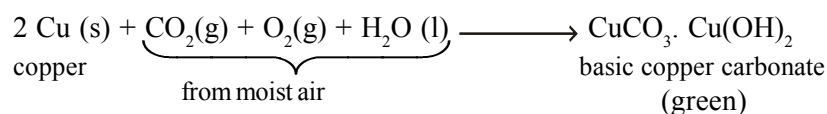
There are a number of redox reactions taking place around us which affect our everyday life. We have observed that iron articles are bright and shining when new. But they get coated with a brown flaky substance when left exposed to moist air for a long time. This process is commonly known as rusting of iron. Rust is mainly hydrated ferric oxide ($\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$). It is formed due to the action of oxygen, water vapour or moisture and other gases present in the air on the surface of iron. The reaction may be represented by the following equation :



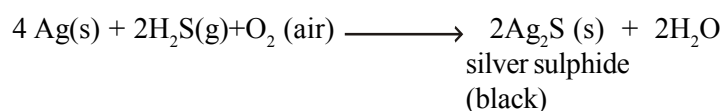
Rusting is an oxidation process. It causes damage to car bodies, bridges, iron railings, automobile parts and many other articles made of iron. The reason is that, the reddish brown crust of rust does not stick to the surface. It peels off exposing fresh surface for rusting. Thus, corrosion of iron is a continuous process which ultimately eats up the whole iron object.

Some other metals also lose their luster and shine after some time. The surface of articles made of copper and brass acquires a green coating of basic carbonate, CuCO_3 , Cu(OH)_2 when exposed to moist air. This is due to

the action of O_2 , CO_2 and water vapour present in the air on the surface of copper.



Likewise, the surface of silver metal gets tarnished on exposure to air. This is due to the formation of coating of black silver sulphide (Ag_2S) on the surface by the action of hydrogen sulphide gas present in the air.



When a metal is reacted with substances around it such as oxygen, moisture, acids etc, it is said to corrode and the process is called corrosion. Rusting of iron, the formation of black coating on silver and the green coating on copper are examples of corrosion.

We have seen that iron objects such as window grills, iron gates, steel furnitures, bodies of cars, buses etc are painted with colours and the painted iron does not rust. Why is it so ?

Corrosion of iron (rusting) can be prevented if iron objects are not allowed to come in contact with damp air. Painting not only makes the iron objects beautiful but also prevents from corrosion by forming a barrier between the iron surface and the damp air.

Iron articles such as mechanical tools, machine parts etc can be prevented from corrosion by **greasing** and **oiling**.

Rusting of iron can be effectively prevented by coating iron with zinc. The process is called **galvanization**

Rancidity

Have you ever tasted or smelt the food materials containing fat or oil left for a longer time ?

When fats and oils in foods are oxidised, they become rancid and their smell and taste change. **The oxidation of fats or oils in foods resulting into bad smell and taste is called rancidity.** A common example is the rancidity of butter on keeping for a longer time.

To prevent rancidity, usually substances which prevent oxidation are added to foods containing fats and oils. These substances are called **anti-oxidants**. Keeping food in air tight containers also helps to slow down oxidation. We have noticed that bags of chips are flushed with gas such as nitrogen to prevent them from oxidation. In a number of cases, while packing the foods, the container is evacuated before sealing so that no air or oxygen is available inside for the oxidation of food. This is called **vacuum packing**.

4.3. SLOW AND FAST REACTIONS

We have observed that a piece of iron rusts in the course of time. We have also observed that the colour of many clothes fades slowly even if they are not washed. If we thrust our hand into a heap of rotting grass or straw, we find it very hot inside. These are all due to chemical reactions and are evidently very slow.

Food is digested in the stomach and intestines very slowly and takes hours for the digestion to complete. Setting of cement and fermentation of sugar to form alcohol are also examples of slow reactions.

When ethyl alcohol and acetic acid is heated in the presence of a little concentrated sulphuric acid, the reaction takes place after a few minutes and an ester called ethyl acetate is formed.



It is a slow reaction. Generally reactions between covalent compounds are slow.

Fast reactions

Recall activity 4.6 where a white substance, barium sulphate is formed when a solution of barium chloride is mixed with dilute sulphuric acid solution. The precipitation is formed at once. So the reaction is very fast or instantaneous.

The reaction between silver nitrate solution and sodium chloride solution is also very fast. The white precipitate of silver chloride is formed at once. **Generally the reactions between ionic compounds in solution are fast.**

Burning of magnesium in air (oxygen) to form magnesium oxide and combination of nitric oxide (NO) with oxygen (O₂) to form nitrogen dioxide (NO₂) are also fast reactions.

In between very fast and very slow reactions, there are certain reactions which occur at moderate speed. These reactions may be called **moderately fast reactions**. The examples of such reactions are :

$$2 \text{H}_2\text{O}_2 (\text{l}) \longrightarrow 2 \text{H}_2\text{O} (\text{l}) + \text{O}_2 (\text{g})$$

Hydrogen peroxide

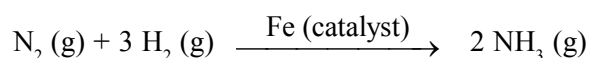
$$\text{C}_{12}\text{H}_{22}\text{O}_{11} + \text{H}_2\text{O} \xrightarrow{\text{acid}} \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + \text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$$

Sugar Glucose Fructose

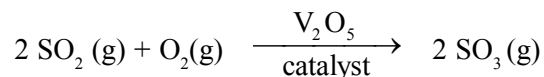
The progress of a reaction can be altered (increases or decreases) by the use of specific substances called catalysts. *A catalyst is a substance which increases the rate of a chemical reaction without itself being consumed and the phenomenon is called catalysis. On the other hand, if a catalyst decreases (retards) the rate of a reaction, it is called a negative catalyst and the phenomenon is called negative catalysis.*

- ☞ Take two beakers labelled A and B. In each of them take 10 mL of oxalic acid solution, 5 mL of dilute sulphuric acid solution and 10 mL of dilute potassium permanganate solution.
- ☞ Observe the colour of the solution in the two beakers.
- ☞ Now, in beaker A add one or two crystals of manganese (II) sulphate (MnSO_4) and observe carefully.
- ☞ Why is the colour of the solution in beaker A fades up faster and finally disappeared?
- ☞ Why is there slow change in the colour of the solution in the beaker B ?

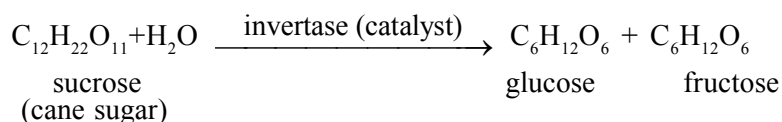
Catalysts are used in the preparation of a large number of industrially important compounds. For example, in the manufacture of ammonia from nitrogen and hydrogen, finely divided iron is used as a catalyst.



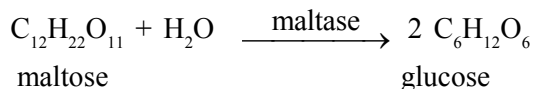
Similarly, vanadium pentoxide (V_2O_5) is used as a catalyst in the manufacture of sulphur trioxide.



A large number of reactions of industrial and biological importance takes place in the presence of catalyst, which would otherwise never occur in their absence. Most of the chemical reactions taking place in plants, animals and micro-organisms require their own specific catalyst called enzymes. Thus, enzymes are biocatalysts present in the living cells which catalyse reactions in living plants and animals. Enzymes are highly specific in their action. Each enzyme catalyses a particular reaction. For example, an enzyme called **invertase** can convert sucrose into glucose and fructose, but fails to in the case of very similar sugar maltose



Maltose requires another enzyme called **maltase** for this purpose.

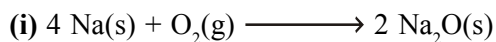


In most of the cases, catalysts are used to accelerate the reactions. In a few cases, however, catalysts are used to slow down the rate of reactions. Such catalysts are known as **negative catalysts**.

For example, glycerine is sometimes added to hydrogen peroxide as negative catalyst to slow down its rate of decomposition into hydrogen and oxygen.

Let us answer these.

1. Why does the colour of copper sulphate solution change when an iron nail is dipped in it ?
2. Give an example of a double displacement reaction other than the one given in activity 4.8.
3. Identify the substances that are oxidised and the substances that are reduced in the following reaction.



SUMMARY

- A chemical reaction is a process in which one or more substances called reactants are changed to one or more new substances called products.
- Combination reactions are those reactions in which two or more substances combine to form a new single substance (compound).
- Decomposition reactions are those reactions in which a single substance splits into two or more substances.
- Reactions in which energy is given out along with the products are called exothermic reactions.
- Reactions in which energy is absorbed are called endothermic reactions.
- A displacement reaction occurs when an element displaces another element from its compound.
- Two different atoms or group of atoms or ions are mutually exchanged in double displacement reactions.
- In a double displacement reaction, if one of the products is a precipitate, it is called precipitation reaction.
- When an acid reacts with a base to form salt and water, it is called neutralization reaction.
- Reactions may also involve the gain or loss of oxygen or hydrogen by substances. Oxidation is the process which involves gain of oxygen or loss of hydrogen. Reduction is loss of oxygen or gain of hydrogen.
- Depending upon the rate at which a chemical reaction proceeds, all chemical reactions may be classified into slow reactions, fast reactions and moderately fast reactions.
- The process of slowly eating away of the metal due to the action of the atmospheric gases on the surface of the metals resulting in the formation of compounds is called corrosion. Corrosion of iron is called rusting.
- Corrosion of metals can be prevented by providing protective layers with substances like grease, paint or metal coatings so that the metals do not come in contact with damp air.

- * The oxidation of fatty foods forming products which have bad taste and bad smell is called rancidity.

EXERCISES

1. Which of the statements about the reaction given below are NOT correct ?



- (a) Lead is reduced.
 - (b) Carbon dioxide is oxidised.
 - (c) Carbon is oxidised
 - (d) Lead oxide is reduced.
 - (i) (a) and (b)
 - (ii) (a) and (c)
 - (iii) (a), (b) and (c)
 - (iv) all
2. $\text{Fe}_2\text{O}_3 + 2 \text{Al} \longrightarrow \text{Al}_2\text{O}_3 + 2 \text{Fe}$
The above reaction is an example of a
- (a) combination reaction
 - (b) double displacement reaction.
 - (c) decomposition reaction.
 - (d) displacement reaction.
3. What are combination reactions ? Give one example.
4. What are decomposition reactions ? Give one experiment to demonstrate a decomposition reaction.
5. Why are decomposition reactions called the opposites of combination reactions ?
6. Write one equation each for decomposition reactions where energy is supplied in the form of heat, light and electricity
7. What is the difference between displacement and double displacement reactions. Write equations for these reactions.
8. What is a displacement reaction ? Give one example.
9. In the refining of silver, silver can be recovered from silver nitrate solution by adding zinc metal. Write down the reaction involved.

10. What is a precipitation reaction ? Explain it by giving one example.
11. What are exothermic and endothermic reactions ? Give two examples each.
12. Why is respiration considered an exothermic reaction ?
13. Explain the terms (a) oxidation and (b) reduction with two examples each.
14. Why does painted iron not rust ?
15. Fat and oil containing food items flushed with nitrogen, why ?

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