CHEMICAL REACTIONS & EQUATIONS

PHYSICAL AND CHEMICAL CHANGE

- In our daily life many processes occur around us. Some of them do not lead to formation of any new substance, while others may lead to formation of a new chemical substance.

- **Physical changes:** Such processes in which no new chemical substances are formed are called physical changes. e.g., melting of ice, evaporation of water, dissolving of sugar in water, etc.
- **Chemical changes:** Such processes in which the original substance loss their nature and identity to form a new chemical substance with different properties are called chemical changes. e.g., rusting of iron, sourcing of milk, freshly cut apple turning brown, cooking food, digestion of food, respiration, fermentation, burning of candle wax, burning of fuel like coal, petrol, LPG etc.

- **Chemical reactions:** The change of one or more substances into other substances having different composition and properties is called a chemical reaction.

Activity 1.1

CAUTION: This Activity needs the teacher's assistance. It would be better if students wear eye protection.

- Clean a magnesium ribbon about 2 cm long by rubbing it with sandpaper.
- Hold it with a pair of tongs. Burn it using a spirit lamp or burner and collect the ash so formed in a watch-glass as shown in Fig. 1.1. Burn the magnesium ribbon keeping it as far as possible from your eyes.
- What do you observe?

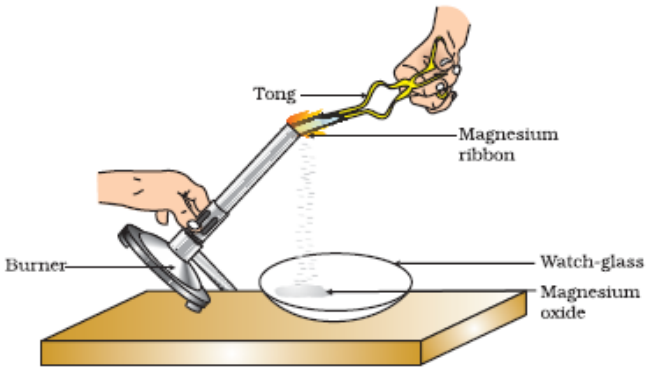

Figure 1.1
Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

- Magnesium ribbon burns with a dazzling white flame and changes into a white powder. This powder is magnesium oxide. It is formed due to the reaction between magnesium and oxygen present in the air.

Magnesium + *Oxygen* → *Magnesium oxide*
Reactants *Product*

REACTANTS AND PRODUCTS

- **Reactants:** The chemical substances taken originally in a chemical reaction is called reactants.
- **Products:** The new substances formed in a chemical reaction is called product.

Activity 1.2

- Take lead nitrate solution in a test tube.
- Add potassium iodide solution to this.
- What do you observe?

Activity 1.3

- Take a few zinc granules in a conical flask or a test tube.
- Add dilute hydrochloric acid or sulphuric acid to this (Fig. 1.2).
  CAUTION: Handle the acid with care.
- Do you observe anything happening around the zinc granules?
- Touch the conical flask or test tube. Is there any change in its temperature?

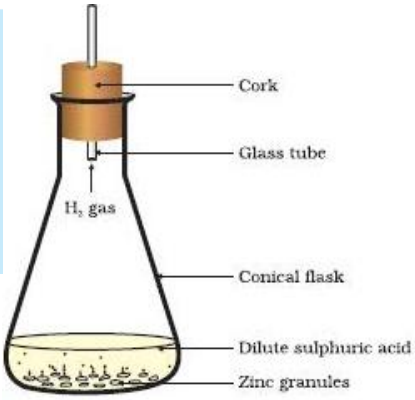

Figure 1.2 *Formation of hydrogen gas by the action of dilute sulphuric acid on zinc*

- During any chemical reaction the existing bonds present in the reactant molecules break and after rearrangement of atoms, new bonds are formed which give rise to product.
- A chemical reaction simply involves rearrangement of atoms and atoms of an element, it never undergo any change to form a new atom or element.

EXAMPLE OF SOME CHEMICAL REACTIONS

1. Burning of magnesium ribbon in air.
   Magnesium + *Oxygen* → *Magnesium oxide*
   White powder
2. Reaction between lead nitrate and potassium iodide.
   Lead nitrate + *Potassium iodide* → *Lead iodide* + *Potassium nitrate*
   colourless solution colourless solution yellow solid colourless solution
3. Reaction between zinc and dilute sulphuric acid.
   Zinc + Sulphuric acid → *Zinc sulphate* + *Hydrogen*
   colourless solution gas

CHARACTERISTICS OF CHEMICAL REACTIONS

- The easily observable changes that take place in a chemical reaction are called the characteristics of chemical reactions. These changes help to check a chemical reaction has taken place.
- Some important characteristics of chemical reaction are:
  1. Formation of precipitate
  2. Evolution of gas
  3. Change in colour
  4. Change in state
  5. Change in temperature

1. **Formation of precipitate:** Some chemical reactions are accompanied by formation of a precipitate. A precipitate is a solid substance formed on mixing two solutions in a chemical reaction. e.g., Mixing of sodium chloride solution with silver nitrate solution form a white precipitate of silver chloride. Mixing of sodium sulphate solution with barium chloride solution form white precipitate of barium sulphate.

2. **Evolution of gas:** Some chemical reactions are accompanied by evolution of gas. e.g., In the reaction between zinc and dilute sulphuric acid or hydrochloric acid, hydrogen gas is evolved. When calcium carbonate is heated or when calcium carbonate react with HCl, carbon dioxide gas is released.

3. **Change in colour:** Some chemical reactions are accompanied by change in colour. e.g., In iron rusting the rust is brown in colour whereas iron is grayish black in colour. Lead nitrate and potassium iodite solution are colourless but on mixing the two the precipitate of lead iodite is formed which is yellow in colour.

4. **Change in state:** Some chemical reactions undergo change in state from reactants to products. e.g., Solid wax in a candle burns to form water vapour and carbon dioxide which are gaseous. Petrol which is a liquid burns to form water vapour and carbon dioxide which are gaseous.

5. **Change in temperature:** In some chemical reaction there is either rise or fall in temperature. On the basis of this, chemical reactions are categorized in to two types:
    a. **Exothermic reactions:** Reactions in which we observe rise in temperature are called exothermic reactions. In this reaction, heat is evolved. e.g: When zinc reacts with sulphuric acid it is found to be warm. When water is added to quicklime (CaO) then in the beaker $Ca(OH)_2$ (slaked lime) is formed and the beaker is found to be hot.
    b. **Endothermic reactions:** Reactions in which there is a fall in temperature are called endothermic reactions. In this reaction the heat is absorbed from surroundings. e.g: When barium hydroxide is added to ammonium chloride in a test tube it is found to be cooler.

- There are many reactions where more than many changes are observed. When we burn coke, oxygen is used and a new black coloured substance is formed in which heat and carbon dioxide is also evlolved.

CHEMICAL EQUATIONS

- Chemical reactions are not always written in language as it takes lots of time and space. There are some chemical reactions which are very complex and it is difficult to write in language therefore we follow some short hand method to represent a chemical reaction.

- There are two ways of representing a chemical reaction.
    - **Word equation:** In this practice we write the reactants on left hand side and products on right hand side. We put an arrow mark between reactant and product.

Magnesium + Oxygen → *Magnesium oxide*

- **Chemical equation:** When we represent a chemical reaction in terms of symbols and formulae of different reactants and products it is called chemical equation.

UNBALANCED OR SKELETAL CHEMICAL EQUATIONS

- Such an equation in which the numbers of atoms of each element on the two sides of the equations are not equal.

$$Mg + O_2 \rightarrow MgO$$
$$Fe + H_2O \rightarrow Fe_3O_4 + H_2$$

BALANCED CHEMICAL EQUATION

- Such an equation in which the numbers of atoms of each element on the two sides of the equation are equal is called balanced chemical equation.

$$2Mg + O_2 \rightarrow 2MgO$$
$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

BALANCING OF CHEMICAL EQUATION

- Making the number of atoms of each element equal on both sides of the equation is called balancing of chemical equations.
- Steps involved in balancing chemicals equations are as follows:
  1. Write the chemical equation in the form of a word equation. Keep the reactants on LHS and products on the RHS, separating them by an arrowhead pointing from the reactants towards the products.
  2. Convert the word equation into the symbol equation by writing the symbol and formulae of the reactants and products.
  3. Start balancing the atoms of a compound which contains the maximum number of atoms whether a reactant or a product.
  4. Having selected the compound with the biggest formula as above, first balance the element of this compound which as the highest number of atoms. Than balance other elements one by one.
  5. Finally to check the correctness of balancing the equation, count atoms of each element on reactant and product sides.
  6. To make a chemical equation more informative, the physical states of the reactants and products are mentioned along with their chemical formulae. The gaseous, liquid, aqueous and solid states of reactants and products are represented by the notations (g), (l), (aq) and (s), respectively.
  7. Sometimes the reaction conditions, such as temperature, pressure, catalyst, etc., for the reaction are indicated above and/or below the arrow in the equation.

TYPES OF CHEMICAL REACTION

- Chemical reactions are of different types depending upon the nature of reaction.
    1. Combination reaction
    2. Decomposition reaction
    3. Displacement reaction
    4. Double displacement reaction
    5. Precipitation reaction
    6. Neutralization reaction
    7. Oxidation or Reduction reaction

COMBINATION REACTION

- Those reactions in which two or more elements or compounds combine to form a single substance are called combination reaction.

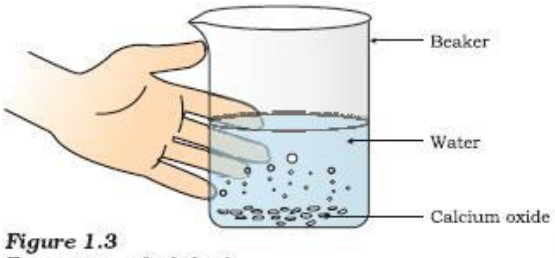

Figure 1.3 *Formation of slaked lime by the reaction of calcium oxide with water*

Activity 1.4

- Take a small amount of calcium oxide or quick lime in a beaker.
- Slowly add water to this.
- Touch the beaker as shown in Fig. 1.3.
- Do you feel any change in temperature?

EXAMPLE:

1. Calcium oxide reacts vigorously with water to produce slaked lime (calcium hydroxide) releasing a large amount of heat.

$$\underset{\text{(Quick lime)}}{CaO\ (s)} + H_2O\ (l) \rightarrow \underset{\text{(Slaked lime)}}{Ca(OH)_2(aq)} \quad (1.13)$$

In this reaction, calcium oxide and water combine to form a single product, calcium hydroxide. Such a reaction in which a single product is formed from two or more reactants is known as a combination reaction.

Do You Know?

A solution of slaked lime produced by the reaction 1.13 is used for white washing walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin layer of calcium carbonate on the walls. Calcium carbonate is formed after two to three days of white washing and gives a shiny finish to the walls. It is interesting to note that the chemical formula for marble is also $CaCO_3$.

$$\underset{\text{(Calcium hydroxide)}}{Ca(OH)_2(aq)} + CO_2(g) \rightarrow \underset{\text{(Calcium carbonate)}}{CaCO_3(s)} + H_2O(l) \quad (1.14)$$

2. Burning of coal (Exothermic reaction)

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

3. Formation of water from $H_2(g)$ and $O_2(g)$

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

4. Synthesis of hydrogen chloride

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

5. Synthesis of ammonia

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

Reactions in which heat is released along with the formation of products are called **exothermic chemical reactions.**

Examples of exothermic reactions are –

1. Burning of coal

$$C(s) + O_2(g) \rightarrow CO_2(g) + E$$

2. Burning of natural gas

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) + E$$

3. Respiration is an exothermic process

$$C_6H_{12}O_6(aq) + 6O_2(aq) \rightarrow 6CO_2(aq) + 6H_2O(l) + E$$

4. The decomposition of vegetable matter into compost is also an example of an exothermic reaction.

DECOMPOSITION REACTION

1. Those reactions in which a single compound breaks down to give two or more simpler substances are called decomposition reaction. Decomposition reaction is just the opposite of combination reaction.

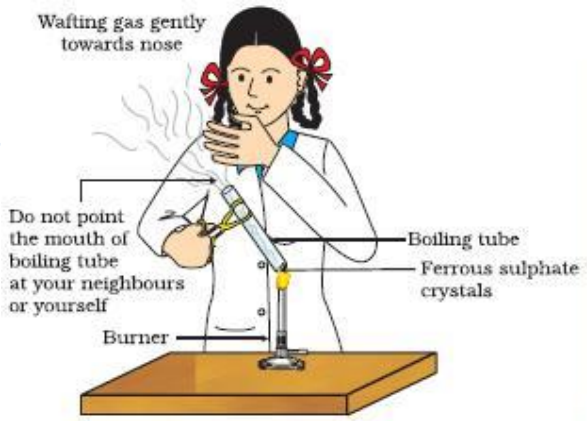

Figure 1.4
Correct way of heating the boiling tube containing crystals of ferrous sulphate and of smelling the odour

Activity 1.5

- Take about 2 g ferrous sulphate crystals in a dry boiling tube.
- Note the colour of the ferrous sulphate crystals.
- Heat the boiling tube over the flame of a burner or spirit lamp as shown in Fig. 1.4.
- Observe the colour of the crystals after heating.

THERMAL DECOMPOSITION REACTION

2. Those decomposition reactions which take place by absorption of heat are called thermal decomposition, or, when a decomposition reaction is carried out by heating, it is called thermal decomposition.

EXAMPLE:

1. In this reaction a single reactant breaks down to give simpler products. This is a decomposition reaction. Ferrous sulphate crystals ($FeSO_4$, $7H_2O$) lose water when heated and the colour of the crystals changes from green to white. It then decomposes to ferric oxide (Fe_2O_3), sulphur dioxide (SO_2) and sulphur trioxide (SO_3). We can also smell the characteristic odour of burning sulphur.

$$\underset{\text{(Ferrous sulphate)}}{2FeSO_4(s)} \xrightarrow{\text{(Heat)}} \underset{\text{(Ferric oxide)}}{Fe_2O_3(s)} + SO_2(g) + SO_3(g)$$

2. Decomposition of calcium carbonate to calcium oxide and carbon dioxide on heating is an important decomposition reaction used in various industries. Calcium oxide is called lime or quick lime; it is used in the manufacture of cement.

$$CaCO_3(s) \xrightarrow{\text{(Heat)}} CaO(s) + CO_2(g)$$

3. When lead nitrate is heated strongly, it breaks down into lead oxide, nitrogen dioxide (brown fume) and oxygen.

$$\underset{\text{(Lead nitrate)}}{2Pb(NO_3)_2(s)} \xrightarrow{\text{(Heat)}} \underset{\text{(Lead oxide)}}{2PbO(s)} + \underset{\text{(Nitrogen dioxide)}}{4NO_2(g)} + \underset{\text{(Oxygen)}}{O_2(g)}$$

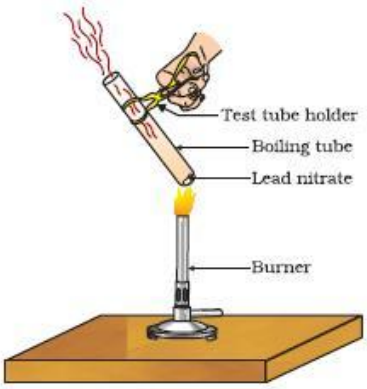

Figure 1.5
Heating of lead nitrate and emission of nitrogen dioxide

Activity 1.6

- Take about 2 g lead nitrate powder in a boiling tube.
- Hold the boiling tube with a pair of tongs and heat it over a flame, as shown in Fig. 1.5.
- What do you observe? Note down the change, if any.

ELECTROLYTIC DECOMPOSITION REACTION

3. Those decomposition reactions which take place when electric current is passed through the compound in the molten state or in aqueous solution.

EXAMPLE:

1. On supplying electrical energy, water decomposes into hydrogen and oxygen.

$$2H_2O \xrightarrow{\text{(Electric Current)}} 2H_2 + O_2$$

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

$$2NaCl \xrightarrow{\text{(Electric Current)}} 2Na + Cl_2$$

PHOTO DECOMPOSITION REACTION

4. Those decomposition reactions which take place by absorption of light are called photo decomposition reaction.

Activity 1.7

- Take a plastic mug. Drill two holes at its base and fit rubber stoppers in these holes. Insert carbon electrodes in these rubber stoppers as shown in Fig. 1.6.
- Connect these electrodes to a 6 volt battery.
- Fill the mug with water such that the electrodes are immersed. Add a few drops of dilute sulphuric acid to the water.
- Take two test tubes filled with water and invert them over the two carbon electrodes.
- Switch on the current and leave the apparatus undisturbed for some time.
- You will observe the formation of bubbles at both the electrodes. These bubbles displace water in the test tubes.
- Is the volume of the gas collected the same in both the test tubes?
- Once the test tubes are filled with the respective gases, remove them carefully.
- Test these gases one by one by bringing a burning candle close to the mouth of the test tubes.
  CAUTION: This step must be performed carefully by the teacher.
- What happens in each case?
- Which gas is present in each test tube?

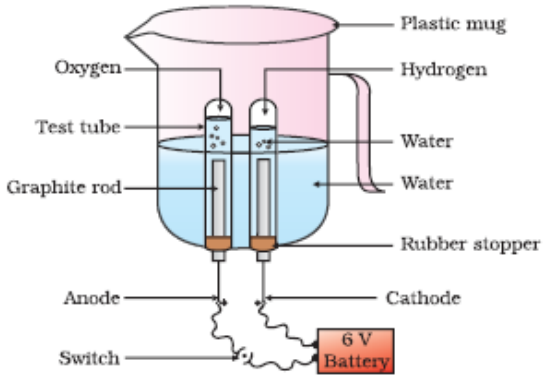

Figure 1.6
Electrolysis of water

Activity 1.8

- Take about 2 g silver chloride in a china dish.
- What is its colour?
- Place this china dish in sunlight for some time (Fig. 1.7).
- Observe the colour of the silver chloride after some time.

Figure 1.7
Silver chloride turns grey in sunlight to form silver metal

EXAMPLE:

1. White silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

$$2AgCl(s) \xrightarrow{\text{(Sunlight)}} 2Ag(s) + Cl_2(g)$$

2. Silver bromide also behaves in the same way, which is used in black and white photography.

$$2AgBr(s) \xrightarrow{\text{(Sunlight)}} 2Ag(s) + Br_2(g)$$

Decomposition reactions require energy either in the form of heat, light or electricity for breaking down the reactants. Reactions in which energy is absorbed are known as **endothermic reactions**.

DISPLACEMENT REACTION

- Those reactions in which one element takes the place of another element in a compound are known as displacement reaction. In these reactions the more reactive element displaces a less reactive element from the compound.

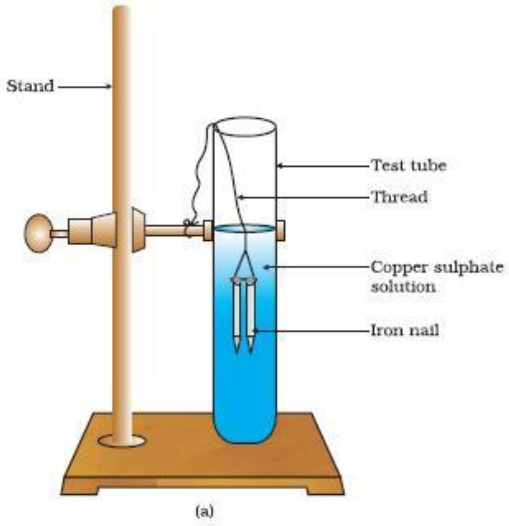

Figure 1.8
(a) Iron nails dipped in copper sulphate solution

Activity 1.9

- Take three iron nails and clean them by rubbing with sand paper.
- Take two test tubes marked as (A) and (B). In each test tube, take about 10 mL copper sulphate solution.
- Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [Fig. 1.8 (a)]. Keep one iron nail aside for comparison.
- After 20 minutes, take out the iron nails from the copper sulphate solution.
- Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B). [Fig. 1.8 (b)].
- Also, compare the colour of the iron nails dipped in the copper sulphate solution with the one kept aside [Fig. 1.8 (b)].

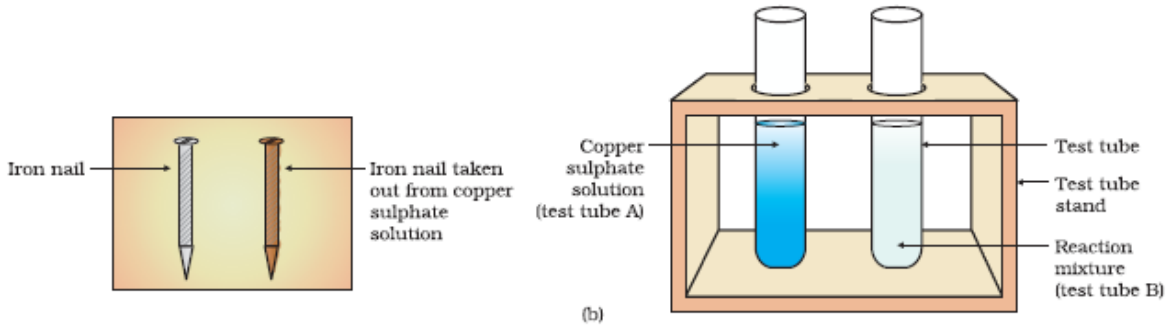

***Figure 1.8** (b) Iron nails and copper sulphate solutions compared before and after the experiment*

EXAMPLE:

1. The following chemical reaction takes place and iron nail become brownish in colour and the blue colour of copper sulphate solution fade off.

$$Fe(s) + \underset{\text{(Copper sulphate)}}{CuSO_4(aq)} \rightarrow \underset{\text{(Iron sulphate)}}{FeSO_4(aq)} + Cu(s)$$

In this reaction, iron has displaced or removed another element, copper, from copper sulphate solution. This reaction is known as displacement reaction.

2. When a strip of zinc metal is placed in copper sulphate solution then zinc sulphate solution and copper are obtained. The displacement of copper occurs because zinc is more reactive than copper.

$$Zn(s) + \underset{\text{(Copper sulphate)}}{CuSO_4(aq)} \rightarrow \underset{\text{(Zinc sulphate)}}{ZnSO_4(aq)} + Cu(s)$$

3. Lead is more reactive elements than copper. It displace copper from its compounds.

$$Pb(s) + \underset{\text{(Copper chloride)}}{CuCl_2(aq)} \rightarrow \underset{\text{(Lead chloride)}}{PbCl_2(aq)} + Cu(s)$$

DOUBLE DISPLACEMENT REACTION

- Those reactions in which two compounds react by exchange of ions to form new compounds are called double displacement reactions.

EXAMPLE

1. When sodium sulphate solution added to barium chloride solution a white substance, which is insoluble in water, is formed. This insoluble substance formed is known as a precipitate.

Activity 1.10

- Take about 3 mL of sodium sulphate solution in a test tube.
- In another test tube, take about 3 mL of barium chloride solution.
- Mix the two solutions (Fig. 1.9).
- What do you observe?

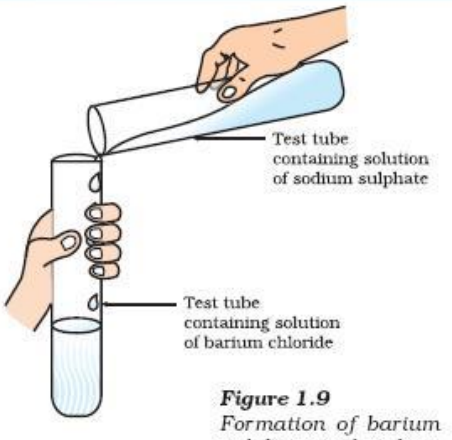

Figure 1.9 *Formation of barium sulphate and sodium chloride*

$$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

(Sodium sulphate) (Barium chloride) (Barium sulphate) (Sodium chloride)

Any reaction in which an insoluble solid or precipitate is formed is called precipitate reaction.

2. When silver nitrate solution is added to sodium chloride solution then a white precipitate of silver chloride is formed along with sodium nitrate solution.

$$AgNo_3(aq) + NaCl\,(aq) \rightarrow NaNo_3(aq) + AgCl(s)$$

OXIDATION AND REDUCTION

- **Oxidation:** It is defined as a process which involves:
  - Addition of oxygen
  - Removal of hydrogen.
- **Reduction**: It is defined as a process which involves:
  - Addition of hydrogen
  - Removal of oxygen.
- **Oxidizing Agent:** The substance which either gives oxygen or remove hydrogen is called oxidizing agent. It oxidizes the substance.
- **Reducing Agent:** The substance which either give hydrogen or removes oxygen is called reducing agent. It reduces the substances.
- Oxidation and reduction occurs together in any reaction and together they are known as "**Redox reaction**".

Activity 1.11

- Heat a china dish containing about 1 g copper powder (Fig. 1.10).
- What do you observe?

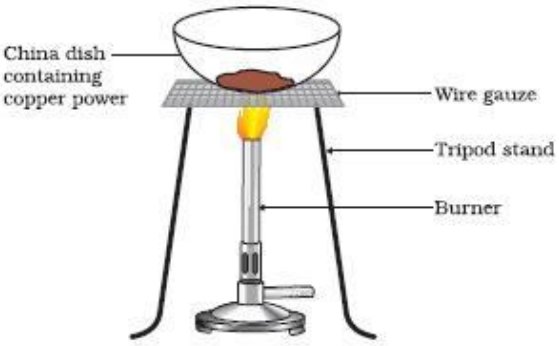

Figure 1.10
Oxidation of copper to copper oxide

EXAMPLE:

PRACTICE SESSION

SOON A VIDEO LECTURE WILL BE ADDED ON THIS PRACTICE SESSION.

EFFECTS OF OXIDATION REACTIONS ON EVERYDAY LIFE

- Oxidation damages metals and food by oxidizing them. The damaging effect of oxidation on metal is known as corrosion and the damaging effect on food is rancidity. Rancidity and corrosion are two common effects of oxidation that we observe in daily life.

CORROSION

- It is a process in which metals are catch up gradually by the action of air, moisture or a chemical on their surface. It is caused mainly by the oxidation of metal.
- Rusting of iron is most common form of corrosion. During the corrosion of iron or rusting of iron, iron metal is oxidized by the oxygen in air in the presence of water to form hydrated iron oxide which is a red brown substance called rust.
- The corrosion weakens the iron and steel objects and structures such as railing, car bodies, bridges, ships etc and it cuts short their life.

RANCIDITY

- When the fats and oils present in food material get oxidized by the oxygen of air they occur as unpleasant taste and smell and this is called rancidity.
- Rancidity can be prevented by adding anti-oxidant to food containing fats and oils. Anti-oxidants are actually reducing agent so when they are added to food it prevents the fats and oils from getting oxidized.
- Rancidity can be prevented by packaging fats and oil containing nitrogen gas. The unavailability of oxygen prevents oxidation.
- Rancidity can be retarded by keeping food in refrigerators as it slows down the oxidation due to low temperature.
- Rancidity can be retarded by storing food in air tight containers.
- Rancidity can be retarded by keeping food away from light.