Class 10 (Science)

Chapter 1

Chemical Reactions and Equations

Handwritten Notes

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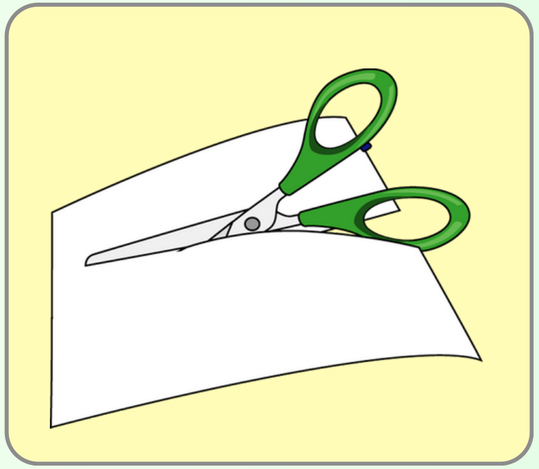
by Jay Rana

Types of changes around us

1) Physical change

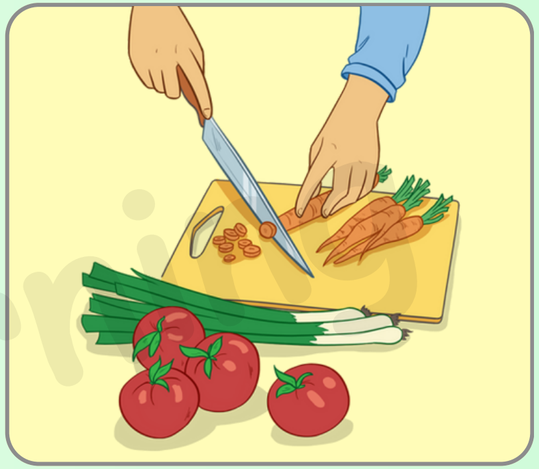
2) Chemical change

Physical change: Physical change do not involve any chemical reaction.



Cutting of paper

Crushing a can



Chopping vegetables

Melting of ice

Boiling of water

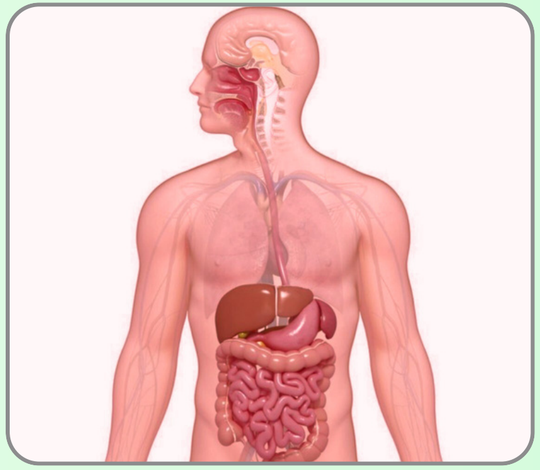
Chemical change: Chemical change always involves a chemical reaction.

Spoiling of milk

Cooking of food

Rusting of iron

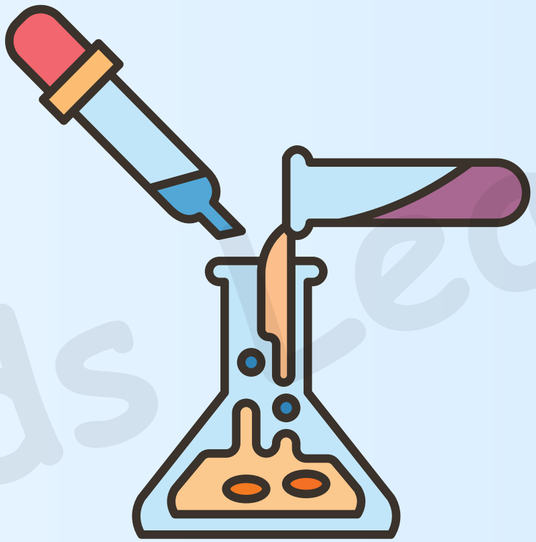
Digestion



Respiration

Chemical reaction

A chemical reaction is a process in which one or more substances are transformed into new substances through the breaking and forming of chemical bonds.



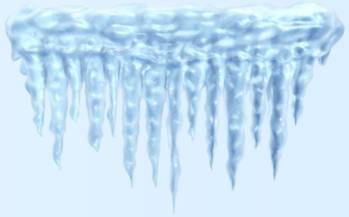
Examples of chemical reaction:

Hydrogen + Oxygen ⟶ Water

Magnesium + Oxygen ⟶ Magnesium oxide

What observation can we see during a chemical reaction?

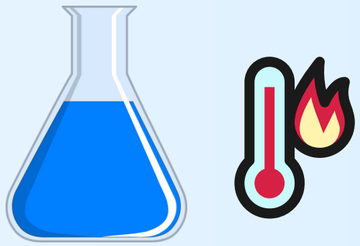
- Change in state



- Change in colour

- Evolution of a gas

- Change in temperature



Chemical equation

A chemical equation is a symbolic representation of a chemical reaction by using the symbols and formulas.

For example

Word equation: Magnesium + Oxygen ⟶ Magnesium oxide

Chemical equation: $Mg + O_2 \longrightarrow MgO$

Reactants: $Mg + O_2$

Products: MgO

Reactants: Reactants are the substances initially present before the reaction takes place.

Products: Products are the substances produced as a result of the chemical reaction.

Balanced chemical equation

An equation in which the number of atoms of each element in the reactant and product are same is called a balanced chemical equation.

Why do we need to balance the chemical equation?

- According to law of conservation of mass, mass can neither be created nor destroyed in a chemical reaction.
- That is, the total mass of the elements present in the products of a chemical reaction has to be equal to the total mass of the elements present in the reactants.
- In other words, the number of atoms of each element remains the same, before and after a chemical reaction.
- Hence, we need to balance a chemical equation.

Example of Balanced chemical equation: $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$ ✓

Example of Unbalanced chemical equation: $Fe + H_2O \longrightarrow Fe_3O_4 + H_2$ ✗

How to balance the chemical equation?

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More information/conditions shown in chemical equations

Physical states: $CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) + Heat$

Heat (Δ): $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

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

Pressure: $CO(g) + 2H_2(g) \xrightarrow{340\ atm} CH_3OH(l)$

Catalyst: $A + B \xrightarrow{X} C$

A catalyst is a substance that helps the reaction happen faster without being used up in the process

Types of Chemical Reactions

1. Combination reaction
2. Decomposition reaction
3. Displacement reaction
4. Double displacement reaction
5. Oxidation and reduction reaction
6. Endothermic and exothermic reaction

1) Combination reaction

The reaction in which two or more reactants combine to form a single product is called a combination reaction.

General form: $A + B \longrightarrow AB$

For example

Burning of coal: $C + O_2 \longrightarrow CO_2$

Formation of water: $2H_2 + O_2 \longrightarrow 2H_2O$

2) Decomposition reaction

A decomposition reaction is a type of chemical reaction where a single compound breaks down into two or more simpler substances.

This process is often triggered by heat, electricity, light, or the presence of other substances.

General form: $AB \longrightarrow A + B$

Three types of Decomposition reaction

1. Thermal Decomposition: It is a type of decomposition reaction triggered by heat.

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

2. Electrolytic Decomposition: It is a type of decomposition reaction that occurs through the application of an electric current.

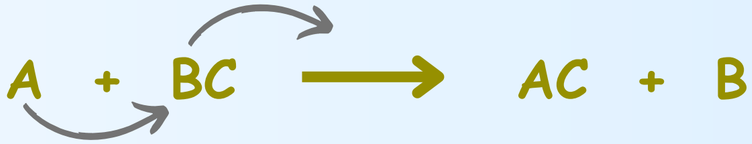
$$2H_2O \xrightarrow{\text{Electric current}} 2H_2 + O_2$$

3. Photolytic Decomposition: Photolytic decomposition is a type of decomposition reaction initiated by light energy.

$$2AgCl \xrightarrow{\text{Sunlight}} 2Ag + Cl_2$$

3) Displacement reaction

A displacement reaction is a type of chemical reaction in which a more reactive element displaces less reactive element from its compound or solution.

General form: 

Here, element A displaces element B, forming a new compound AC.

For example

$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

$Pb + CuCl_2 \longrightarrow PbCl_2 + Cu$

Iron, zinc and lead are more reactive than copper, so they displaces copper from its compound.

4) Double Displacement reaction

A double displacement reaction is a type of chemical reaction where two compounds exchange ions to form two new compounds.

In this process, the positive and negative ions of two different compounds switch places, resulting in the formation of two entirely new compounds.

General form: $AB + CD \longrightarrow AD + CB$

For example

$NaOH + HCl \longrightarrow NaCl + H_2O$

$Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2NaCl$

$AgNO_3 + NaBr \longrightarrow AgBr + NaNO_3$

5) Oxidation and Reduction reaction

Oxidation reaction

The process of

- addition of oxygen, or
- removal of hydrogen, or
- removal of electron,

is called oxidation reaction.

Examples:

$2Cu + O_2 \xrightarrow{\Delta} 2CuO$

$4HCl + O_2 \longrightarrow 2Cl_2 + 2H_2O$

$Zn \longrightarrow Zn^{2+} + 2e^-$

Reduction reaction

The process of

- removal of oxygen, or
- addition of hydrogen, or
- addition of electron,

is called reduction reaction.

Examples:

$2HgO \longrightarrow 2Hg + O_2$

$2Na + H_2 \longrightarrow 2NaH$

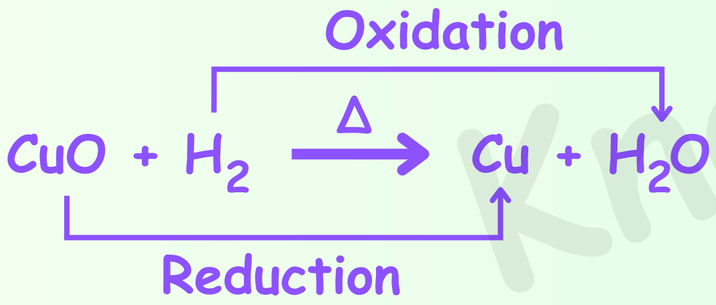
$Zn^{2+} + 2e^- \longrightarrow Zn$

Redox reaction

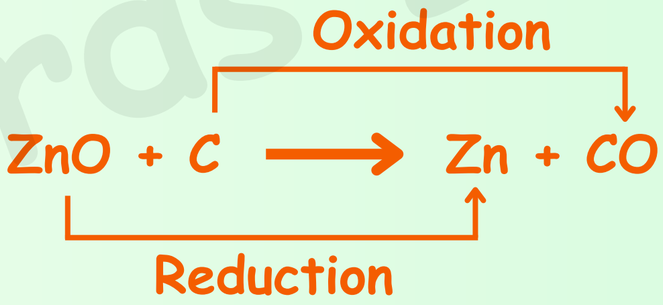
A redox reaction is a chemical reaction in which reduction and oxidation reaction occurs simultaneously.

Redox = Reduction + Oxidation

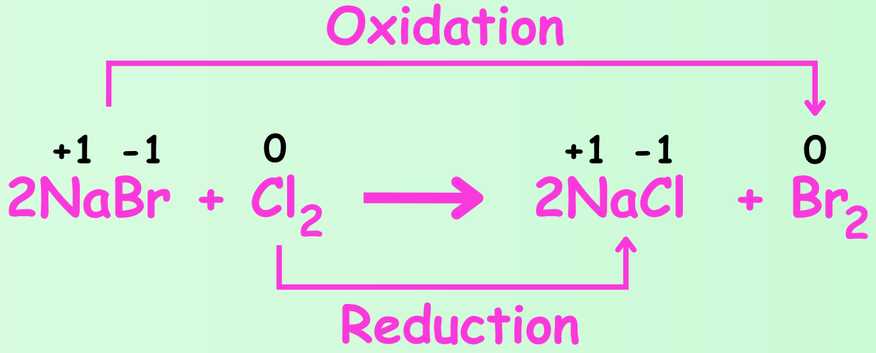
Example 1



Example 2



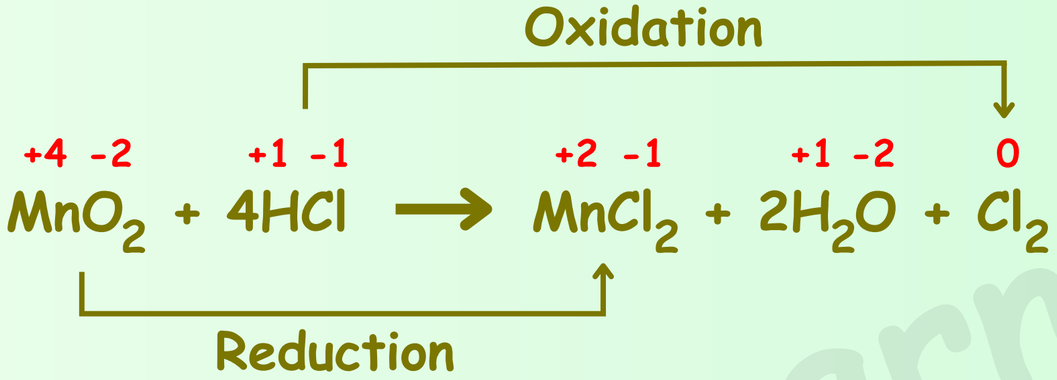
Example 3



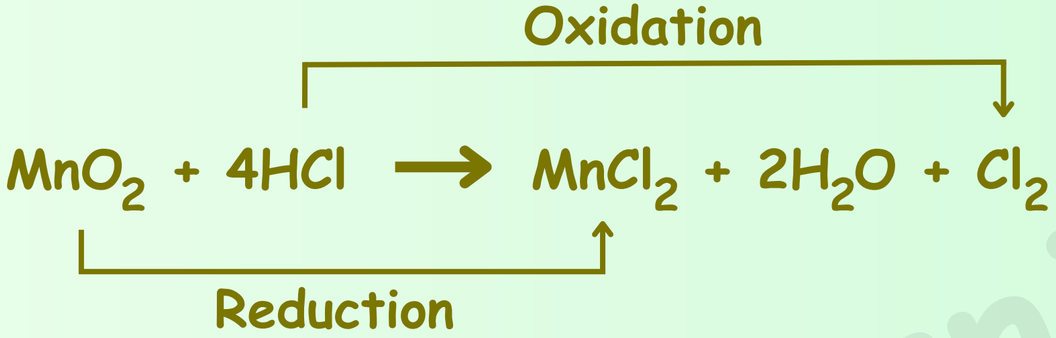
Remember:

- Charge Increase hua to Oxidation (e.g. 0 to +1)
- Charge Reduce hua to Reduction (e.g. 0 to -1)

Example 4



Some imp terms



Oxidised: **If a substance undergoes oxidation, it is said to be oxidised.** In above example, HCl is oxidised.

Reduced: **If a substance undergoes reduction, it is said to be reduced.** In above example, MnO2 is reduced.

Oxidizing agent: **The substance that causes oxidation is called oxidizing agent.**
In above example, MnO2 is oxidizing agent.

Reducing agent: **The substance that causes reduction is called reducing agent.**
In above example, HCl is reducing agent.

Home work:

Que: Identify the substance oxidized, substance reduced, oxidising agent and reducing agent in the following reactions.

$2H_2 + O_2 \longrightarrow 2H_2O$

$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$

$Mg + Br_2 \longrightarrow MgBr_2$

Answers

$\overset{0}{2H_2} + \overset{0}{O_2} \longrightarrow \overset{+1\ -2}{2H_2O}$

H2 is oxidized; O2 is oxidizing agent.
O2 is reduced; H2 is reducing agent.

$\overset{0}{Zn} + \overset{+1\ -1}{2HCl} \longrightarrow \overset{+2\ -1}{ZnCl_2} + \overset{0}{H_2}$

Zn is oxidized; HCl is oxidizing agent.
HCl is reduced; Zn is reducing agent.

$\overset{0}{Mg} + \overset{0}{Br_2} \longrightarrow \overset{+2\ -1}{MgBr_2}$

Mg is oxidized; Br2 is oxidizing agent.
Br2 is reduced; Mg is reducing agent.

6) Endothermic and Exothermic reaction

Endothermic reaction: Chemical reactions that absorb heat energy from their surroundings to proceed.

$$CaCO_3 + Heat \longrightarrow CaO + CO_2$$

Exothermic reaction: Chemical reactions that release heat energy to their surroundings as they proceed.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + Heat$$

Effects of Oxidation in everyday life

1) Corrosion

Corrosion: When a metal is attacked by substances around it such as **oxygen, moisture, acids**, etc., it is said to corrode and this process is called **corrosion**.

For example:

Rusting of iron

Black coating on silver

Green coating on copper

Effects of corrosion

Corrosion of car bodies

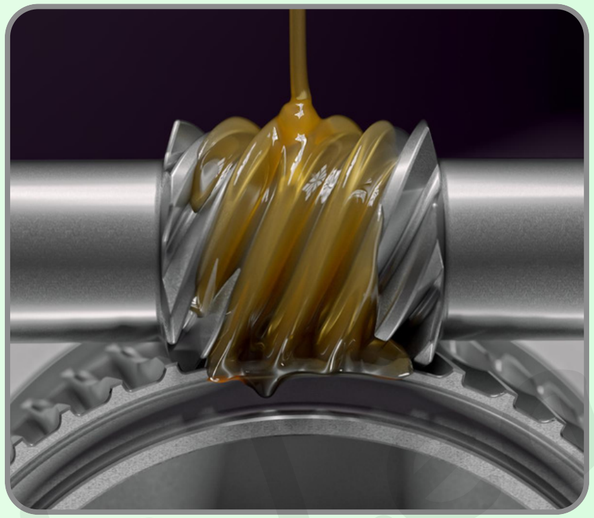
Corrosion on bridges

Corrosion of ships

Corrosion of iron railings

How to prevent corrosion?

Painting



Greasing

Oiling

Chrome plating

Galvanizing

Alloying

2) Rancidity

Rancidity: The oxidation of fats and oils when exposed to air is known as rancidity. It leads to bad smell and bad taste of food.

How to prevent rancidity?

Keeping food in airtight containers

Using flush bags with nitrogen gas

Refrigeration

Adding antioxidants

How to balance the chemical equation?

Note: Always start with the metal

Example 1:

$$Fe + H_2O \longrightarrow Fe_3O_4 + H_2$$

$$3\,Fe + 4\,H_2O \longrightarrow Fe_3O_4 + 4\,H_2$$

Example 2:

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

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

Example 3:

$$FeS_2 + O_2 \longrightarrow Fe_2O_3 + SO_2$$

$$2\,FeS_2 + \frac{11}{2}O_2 \longrightarrow Fe_2O_3 + 4\,SO_2$$

OR

$$4\,FeS_2 + 11\,O_2 \longrightarrow 2\,Fe_2O_3 + 8\,SO_2$$

Activity 1.1

Activity 1.1

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

- Clean a magnesium ribbon about 3-4 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 away as far as possible from your eyes.
- What do you observe?

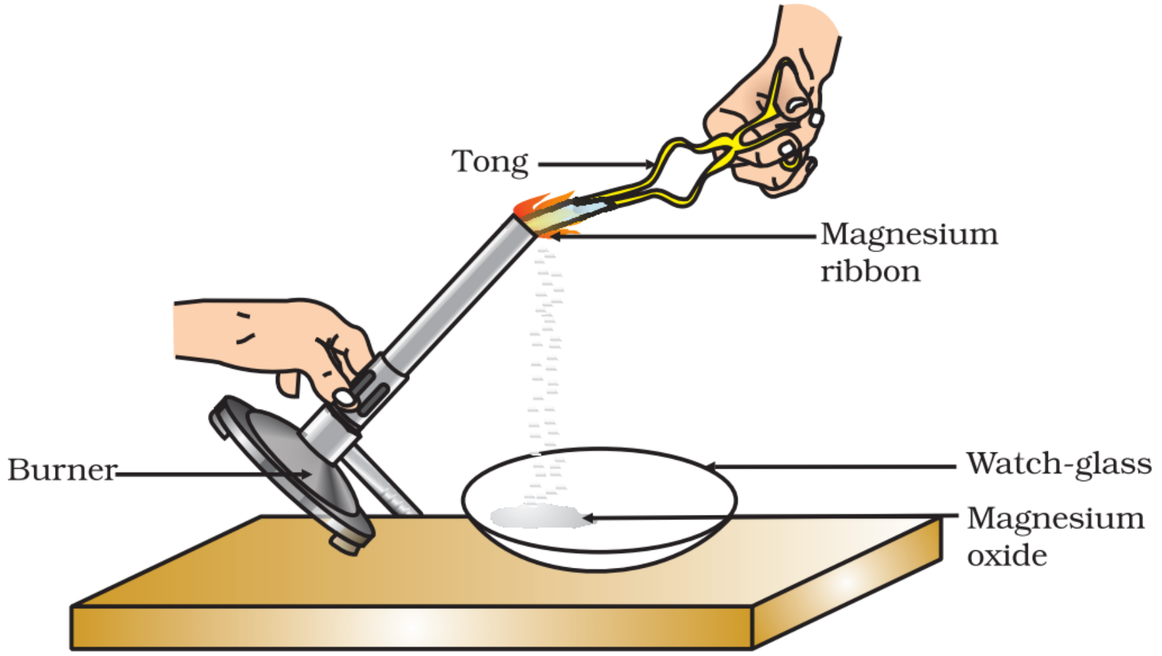


Figure 1.1
Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

Reaction involved:

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

Questions

Que: **Explain the chemical reaction involved in the formation of magnesium oxide.**

Ans: - The chemical reaction is;

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

- It represents the combination of magnesium (Mg) with oxygen (O2) to form magnesium oxide (MgO).

Que: **What happens when magnesium ribbon is burnt in the air?**

Ans: - When magnesium ribbon is burnt in the air, it undergoes a chemical reaction with oxygen present in the air, and produces bright white light.

- In this reaction, magnesium (Mg) reacts with oxygen (O2) to produce magnesium oxide (MgO).

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

Activity 1.2

Activity 1.2

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

Reaction involved:

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

Questions

Que: **What is a precipitation reaction? Give one example.**

Ans: - A precipitation reaction is a type of chemical reaction in which two soluble substances in a solution combine to form an insoluble product, known as a precipitate. The precipitate is a solid that separates from the solution and usually appears as a fine suspension of particles.

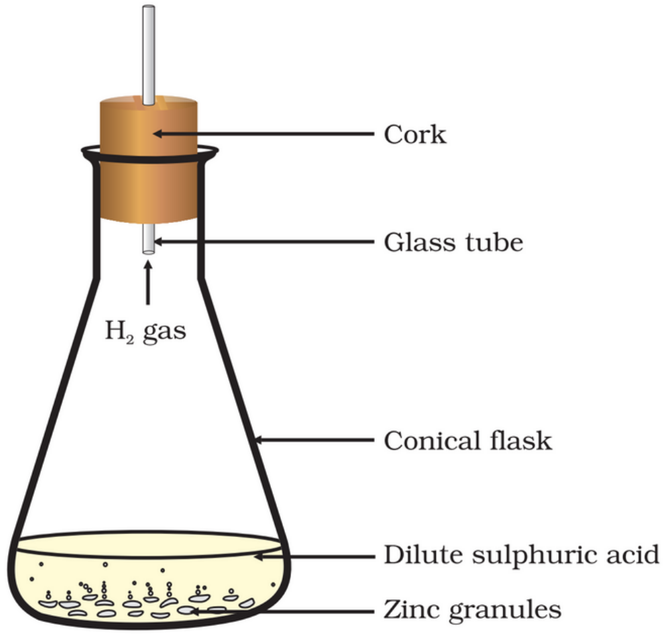
- Example of precipitation reaction: Lead nitrate (Pb(NO3)2) reacts with potassium iodide (KI) to form lead iodide (PbI2) as a yellow colored solid precipitate, while potassium nitrate (KNO3) remains in the solution.

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

Activity 1.3

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?



Reaction involved:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2 + Heat$$

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 + Heat$$

Questions

Que: **What will you observe when zinc granules reacts with dilute hydrochloric acid or dilute sulphuric acid?**

Ans: - When zinc granules react with dilute hydrochloric acid or dilute sulfuric acid, you will observe the evolution of hydrogen gas.

- Also these reactions are exothermic reactions, meaning heat will liberate during this reactions.

- The chemical equations for these reactions are as follows:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2 + Heat$$

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 + Heat$$

Que: **Why are dilute acids often used in chemical reactions instead of concentrated acids?**

Ans: - Dilute acids are often used in chemical reactions instead of concentrated acids because they **provide better control** over the reaction. The use of dilute acids allows for a more controlled and safer reaction environment.

- The other reason that dilute acids are commonly used in laboratory settings is to **avoid excessive heat generation** during the reaction.

- Additionally, **concentrated acids are risky in labs** because they are strong, corrosive, and can be dangerous.

Activity 1.4

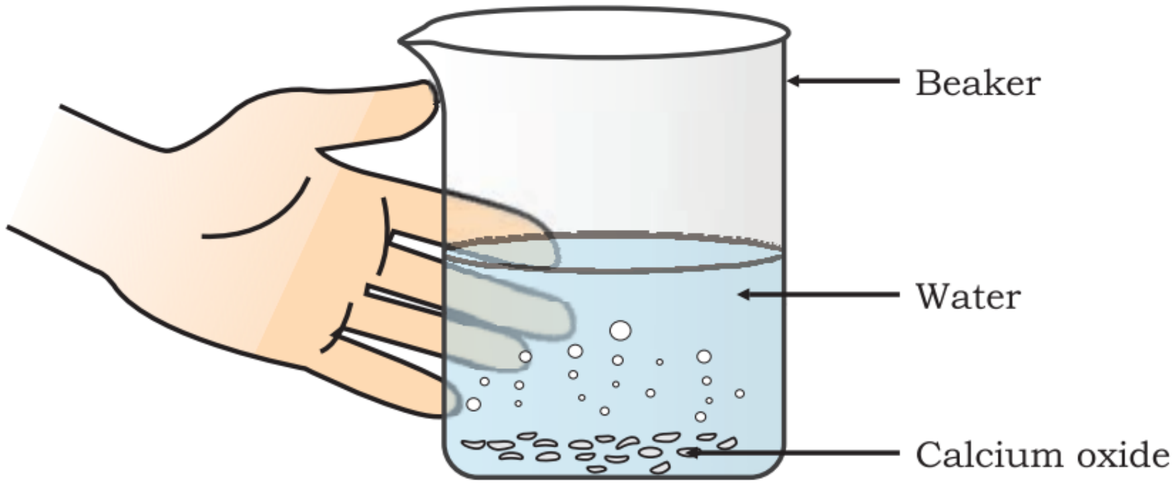


Figure 1.3
Formation of slaked lime by the reaction of calcium oxide with water

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?

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

Reaction involved:

$$CaO + H_2O \longrightarrow Ca(OH)_2 + \text{Heat}$$

(Quick lime) or (calcium oxide)

(Slaked lime) or (calcium hydroxide)

Questions

Que: **What happens when water is added to quick lime? What type of reaction it is?**

Or

How is slaked lime formed? Explain with a chemical reaction.

Ans: - When water is added to quick lime (calcium oxide), the quick lime reacts vigorously with water to produce slaked lime (calcium hydroxide) releasing a large amount of heat.

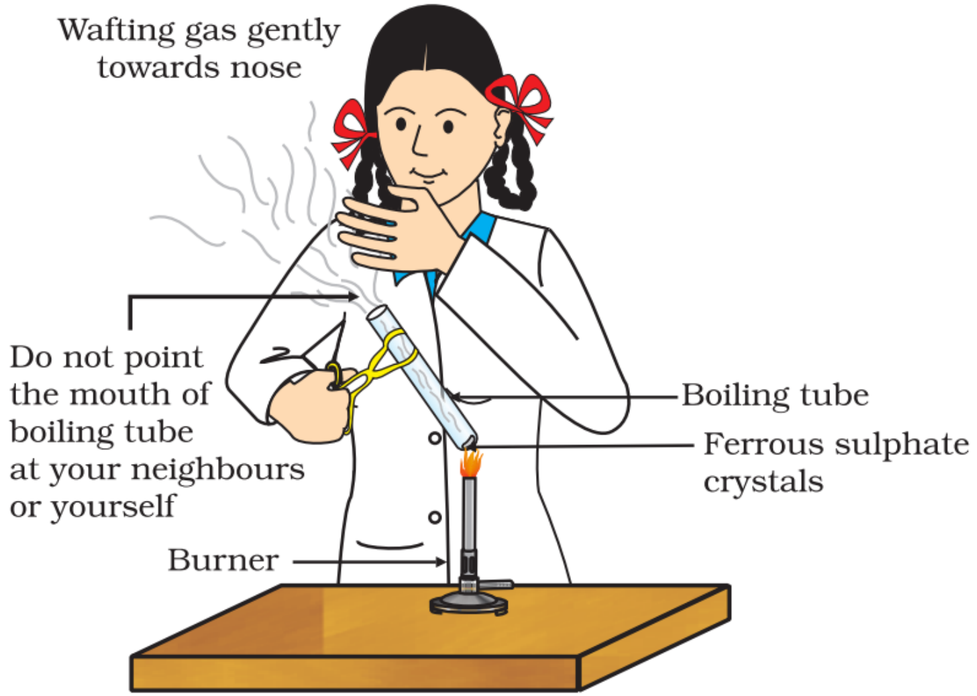
- This reaction is **exothermic reaction** as it releases large amount of heat.

- Additionally, it is also a **combination reaction** as two reactants combine to form a single product.

- The chemical equation for this reaction is as follows;

$$\underset{\text{(Quick lime)}}{CaO} + H_2O \longrightarrow \underset{\text{(Slaked lime)}}{Ca(OH)_2} + \text{Heat}$$

Activity 1.5



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.

Have you noticed that the green colour of the ferrous sulphate crystals has changed? You can also smell the characteristic odour of burning sulphur.

Reaction involved:

$$2\,FeSO_4 \xrightarrow{\Delta} Fe_2O_3 + SO_2 + SO_3$$

(Ferrous sulphate) (Ferric oxide)

Questions

Que: **What products are obtained when ferrous sulphate is heated in a test tube? Write the chemical reaction and also state the type of chemical reaction.**

Ans: - Ferrous sulphate crystals ($FeSO4 . 7H2O$) lose water when heated and the colour of the crystals changes. It then decomposes to ferric oxide ($Fe2O3$), sulphur dioxide ($SO2$) and sulphur trioxide ($SO3$).

- The chemical equation for this reaction is as follows;

$$2\,FeSO_4 \xrightarrow{\Delta} Fe_2O_3 + SO_2 + SO_3$$

(Ferrous sulphate) (Ferric oxide)

- The chemical reaction is a decomposition reaction (thermal decomposition reaction), as a single reactant breaks down to give simpler products due to the application of heat.

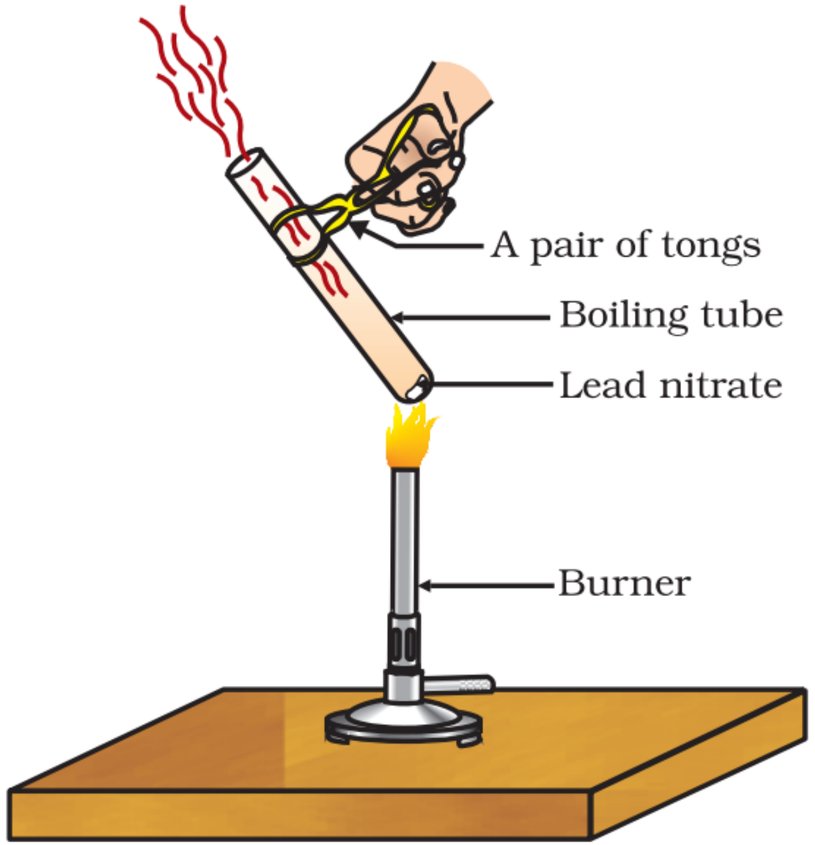
Que: **List down any two observations when ferrous sulphate is heated in a dry test tube.**

Ans: - Initially the ferrous sulphate crystals are green in color, but it changes to white colored crystals as it loses water on heating.

- After that, these crystals immediately turns to brown color due to the formation of ferric oxide.

Ferrous sulphate crystals

Activity 1.6



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.

Reaction involved:

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

Questions

Que: **What happens when lead nitrate powder is heated in a test tube?**

Ans: - When lead nitrate powder (Pb(NO3)2) is heated in a test tube, it decomposes into 3 products.
(i) Lead oxide (PbO)
(ii) Nitrogen dioxide (NO2)
(iii) Oxygen (O2)

- The brown fumes comes out of the test tube, which indicates the emission of nitrogen dioxide (NO2) gas.

- The chemical equation for this decomposition reaction is as follows;

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

Activity 1.7

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?

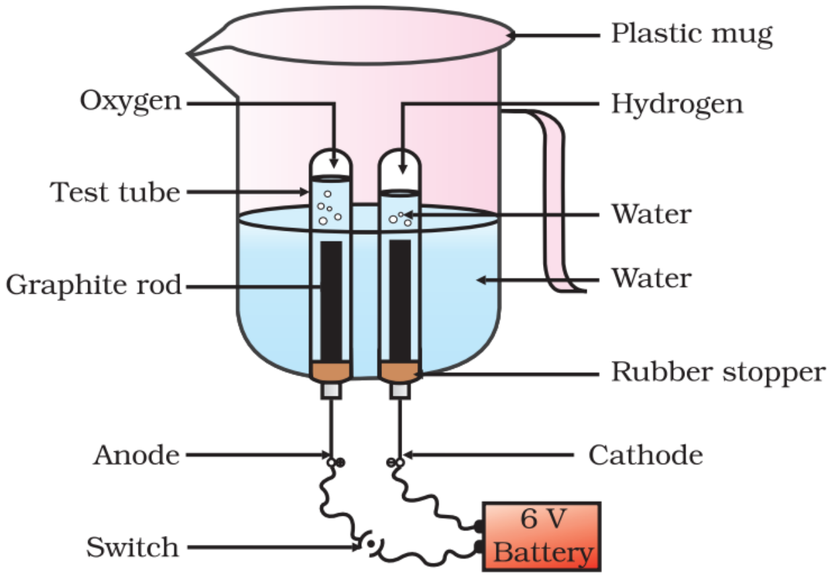


Figure 1.6
Electrolysis of water

Questions

Que: **What products are obtained by the electrolysis of water? What type of reaction is the electrolysis of water?**

Ans: - When electric current is passed through the water (H2O), it decomposes in to its two primary components, hydrogen (H2) and oxygen (O2).

- The chemical equation for this decomposition reaction is as follows;

$$H_2O \xrightarrow{\text{Electric current}} H_2 + O_2$$

- The electrolysis of water is a decomposition reaction, more specifically it is an electrolytic decomposition reaction.

Activity 1.8

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.

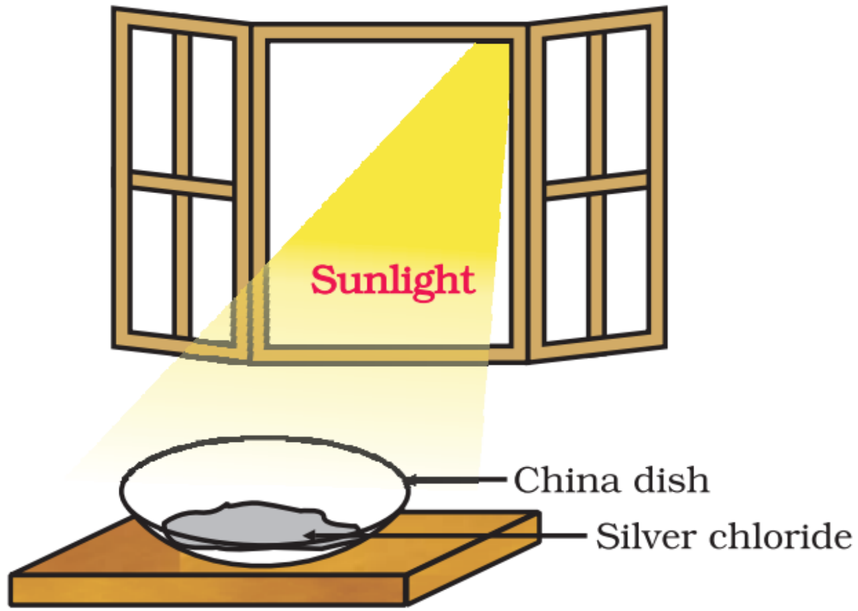


Figure 1.7
Silver chloride turns grey in sunlight to form silver metal

Reaction involved:

$$2AgCl \xrightarrow{\text{Sunlight}} 2Ag + Cl_2$$

Questions

Que: **What happens when silver chloride is kept open in the sunlight? Why?**

Ans: - When silver chloride is kept open in sunlight, the white silver chloride turns grey in color.

- This is because silver chloride (AgCl) undergoes a decomposition reaction, and forms silver (Ag) and chlorine (Cl2).

- The chemical equation for this decomposition reaction is as follows;

$$2AgCl \xrightarrow{\text{Sunlight}} 2Ag + Cl_2$$

Que: **Why is silver chloride always kept in black bottles?**

Ans: - Silver chloride is kept in black bottles because it's light-sensitive and breaks down when exposed to sunlight.

- Sunlight provides the energy for silver chloride to decompose into silver metal and chlorine gas, making it lose its properties and darken in color.

- Black bottles block out light, protecting the silver chloride and preserving its stability.

Activity 1.9

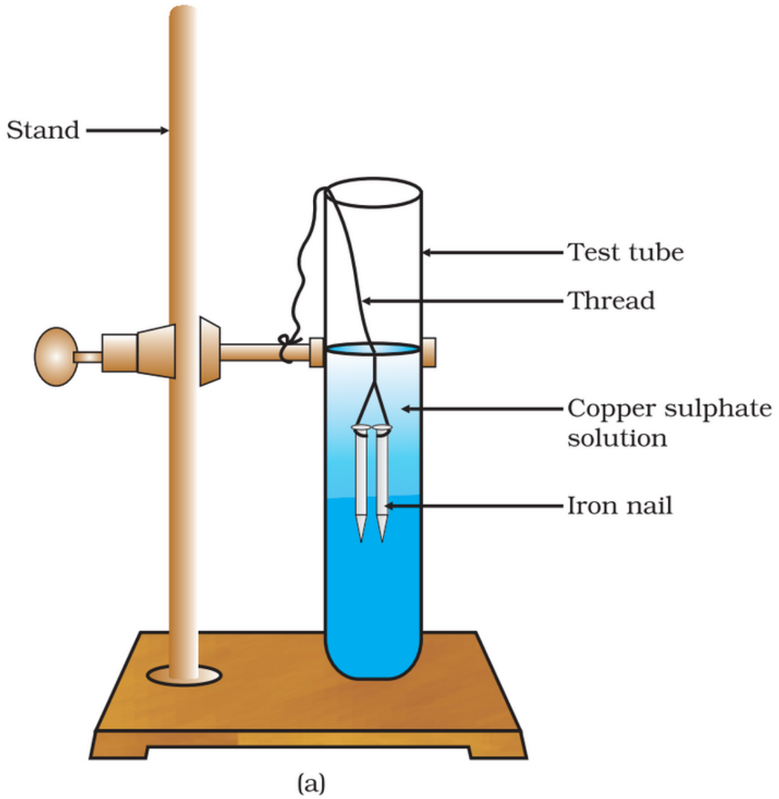


Figure 1.8
(a) Iron nails dipped in copper sulphate solution

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)].

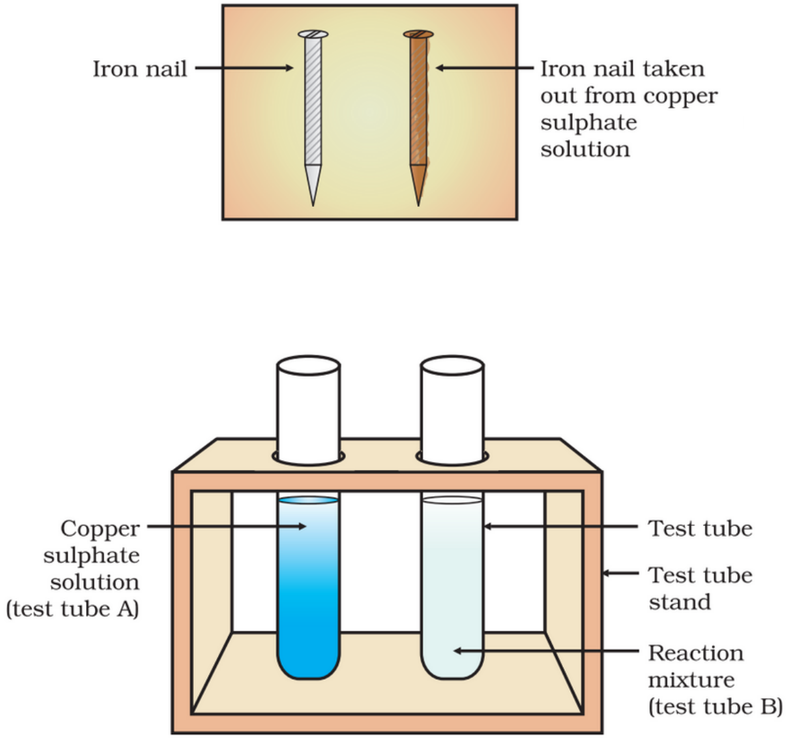


Figure 1.8 *(b) Iron nails and copper sulphate solutions compared before and after the experiment*

Reaction involved:

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$$

Questions

Que: **What happens when iron nail is dipped in the copper sulphate solution?**

Ans: - When an iron nail is dipped in a copper sulfate solution, a displacement reaction occurs, leading to the formation of iron sulfate (FeSO4) and metallic copper (Cu).

- The chemical equation for this displacement reaction is as follows;

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$$

Que: **When iron nail is dipped in the blue colored copper sulphate solution, it is observed that the color of the solution turns light green. Explain the reason for this color change.**

Ans: - The color change from blue to light green when an iron nail is dipped in blue-colored copper sulfate solution indicates a displacement reaction.

- Iron from the nail displaces copper from copper sulfate, forming iron sulfate in solution and depositing metallic copper.

- The light green color indicates the presence of ferrous sulfate, which gives the solution its green tint.

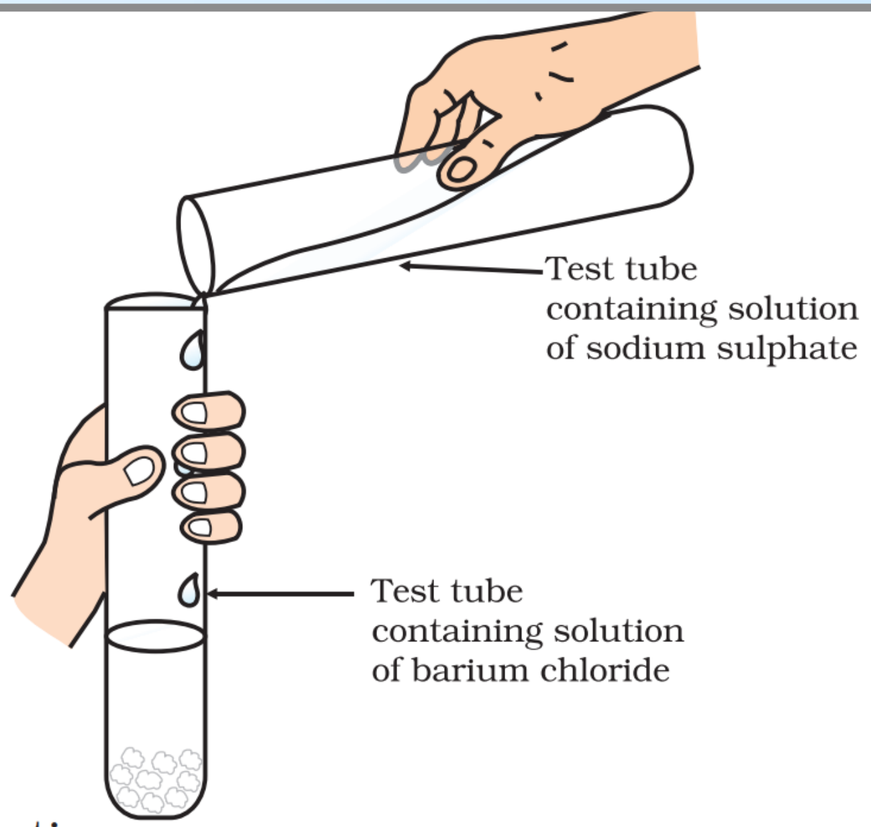
- The chemical equation for this displacement reaction is as follows;

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$$

Activity 1.10

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?



You will observe that a white substance, which is insoluble in water, is formed. This insoluble substance formed is known as a precipitate. Any reaction that produces a precipitate can be called a precipitation reaction.

Reaction involved:

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

Questions

Que: **Explain the double displacement reaction that occurs when mixing sodium sulphate and barium chloride solutions.**

Ans: - When the sodium sulphate and barium chloride solutions are mixed, a double displacement reaction occurs.

- The barium sulfate (BaSO4) is formed as a white precipitate and sodium chloride (NaCl) remains in the solution. This reaction is also knows as precipitation reaction.

- The chemical equation for this reaction is as follows;

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

Activity 1.11

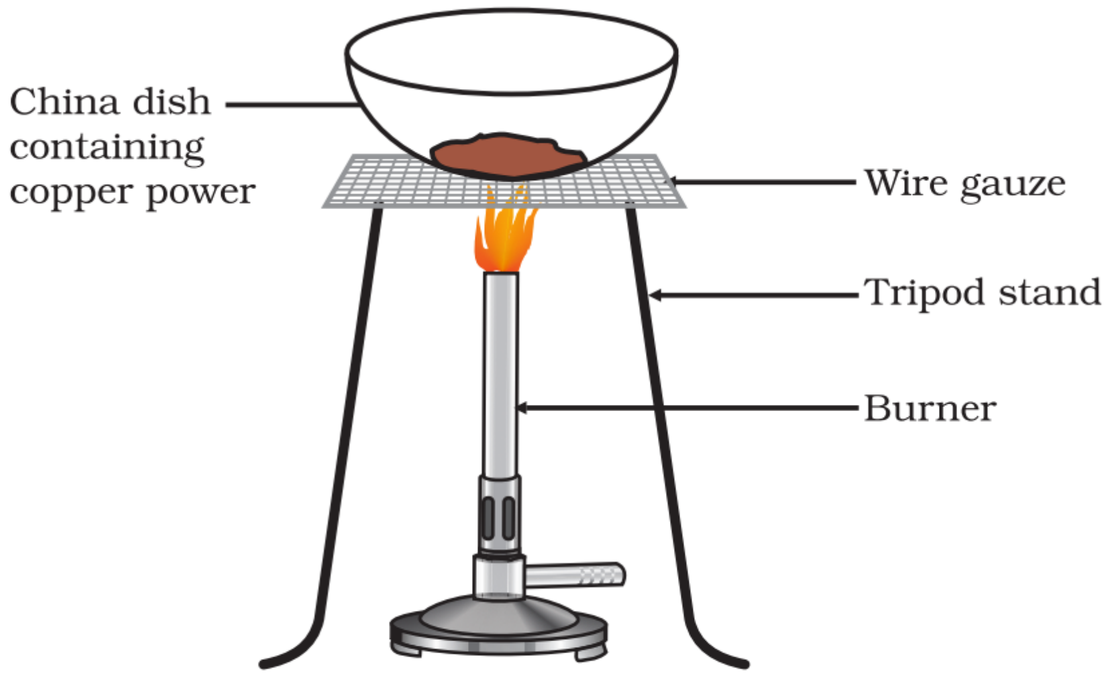


Figure 1.10
Oxidation of copper to copper oxide

Activity 1.11

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

Reaction involved:

$$2Cu + O_2 \xrightarrow{\Delta} 2CuO$$

Questions

Que: **What will you observe when a copper powder is heated in a china dish?**

Ans: - When a copper powder (Cu) is heated, it reacts with the oxygen (O2) of the air and produces copper oxide (CuO).

- The chemical equation for this reaction is as follows;

$$2Cu + O_2 \xrightarrow{\Delta} 2CuO$$

- During this reaction, the black coating is observed on the copper powder, which indicates the copper oxide (CuO).

Que: **You might have noted that when copper powder is heated in a china dish, the reddish brown color of the copper powder turns black. How can this black coating be turned reddish brown again?**

Ans: - The black coating formed on copper powder when heated is copper oxide (CuO).

- To turn it reddish-brown again, the copper oxide can be reduced back to copper by heating it and passing the hydrogen gas over this heated copper oxide.

- The chemical equation for this reaction is as follows;

$$CuO + H_2 \xrightarrow{\Delta} Cu + H_2O$$