UNIT I

Materials

Experiment 1

Aim

To study the chemical reaction of an iron nail with aqueous copper sulphate solution; and to study the burning of magnesium ribbon in air.

(a) Chemical reaction of iron nail with copper sulphate solution in water.

Theory

Iron displaces copper ions from an aqueous solution of copper sulphate. It is a single displacement reaction of one metal by another metal. Iron is placed above copper in the activity series. Elements placed above in this series are more reactive than those placed below them. Thus iron is more reactive than copper. In this reaction, metallic iron is converted into ferrous ion (Fe^{2+}) and cupric ion (Cu^{2+}) is converted into metallic copper.

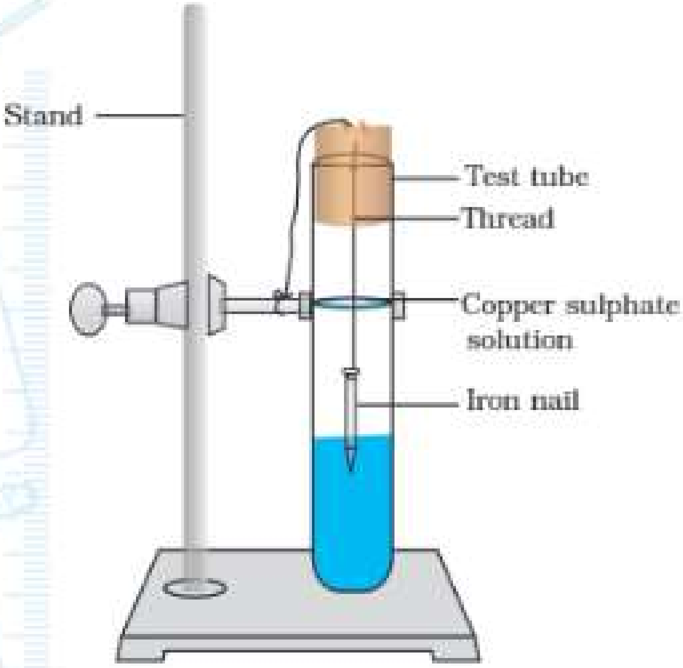
$$Fe\ (s) + Cu^{2+}\ (aq) \longrightarrow Fe^{2+}\ (aq) + Cu\ (s).$$

Materials Required

Two test tubes, two iron nails, measuring cylinder (50 mL), laboratory stand with clamp, test tube stand, thread, a piece of sand paper, single bored cork, copper sulphate, distilled water, and dil. sulphuric acid.

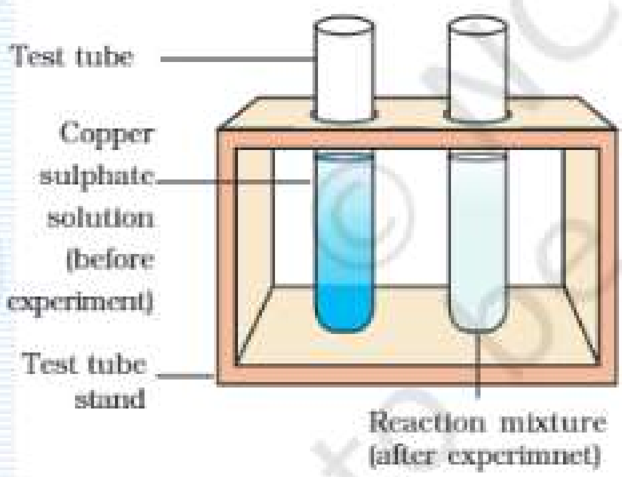
Procedure

1. Take two iron nails and clean them with a sand paper.

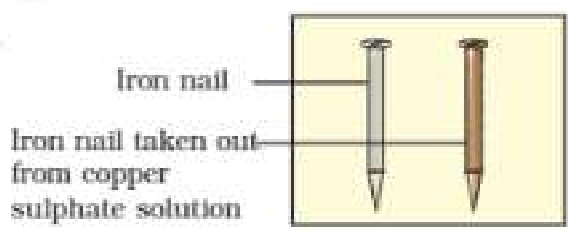


(a)

2. Take 20 mL of distilled water in a clean test tube and dissolve 1.0 g of copper sulphate in it. Add 2 or 3 drops of dil. sulphuric acid to it to check hydrolysis of $CuSO_4$ in water. Label this test tube as A.
3. Transfer about 10 mL of copper sulphate solution from tube A to another clean test tube. Label this test tube as B.
4. Tie one iron nail with a thread and immerse it carefully in copper sulphate solution in test tube B through a bored cork [as shown in the Fig 1.1(a)] for about 15 minutes [Fig. 1.1(a)]. Keep the another iron nail separately for comparison afterwards.
5. After 15 minutes take out the iron nail from the copper sulphate solution.



(b)



(c)

Fig. 1.1 : *(a) Iron nail dipped in copper sulphate solution; and (b) Iron nails and copper sulphate solutions are compared*

6. Compare the intensity of blue colour of copper sulphate solution before and after the experiment in tubes A and B, and also compare the colour of iron nail dipped in copper sulphate solution with the one kept separately [Fig. 1.1(b) and (c)]. Record your observations.

OBSERVATIONS

Sl.No.	Property	Before experiment	After experiment
1.	Colour of copper sulphate solution		
2.	Colour of iron nail		

RESULTS AND DISCUSSION

Infer from your observations about the changes in colours of copper sulphate solution and iron nail. Discuss the reason(s).

PRECAUTIONS

- The iron nails must be cleaned properly by using sand paper before dipping them in copper sulphate solution.

QUESTIONS

- Why does the colour of copper sulphate solution change, when an iron nail is dipped in it?
- How would you devise the procedure to show that Mg > Fe > Cu in reactivity series?
- What is the basic principle involved in this experiment?
- Why does the following reaction takes place?

$$2I^-(aq) + Cl_2(aq) \xrightarrow{CCl_4} 2Cl^-(aq) + I_2 \text{ (solvated)}.$$

(b) Chemical reaction of burning of magnesium ribbon in air.

THEORY

Magnesium forms magnesium oxide on burning in presence of air. It is a combination reaction between two elements. Magnesium oxide is basic in nature and thus its aqueous solution turns red litmus blue.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Materials Required

Magnesium ribbon (2 to 3 cm long), a pair of tongs, burner, a pair of dark coloured goggles, watch glass, red and blue litmus papers, distilled water, beaker, and a piece of sand paper.

Procedure

1. Take a magnesium ribbon (2 to 3 cm long) and clean it with a sand paper. This will remove the oxide layer deposited over the magnesium ribbon, which makes it passive.
2. Hold the magnesium ribbon with a pair of tongs over a watch glass and burn it in air with a burner (Fig. 1.2). Watch the burning of magnesium ribbon using a pair of dark coloured goggles.

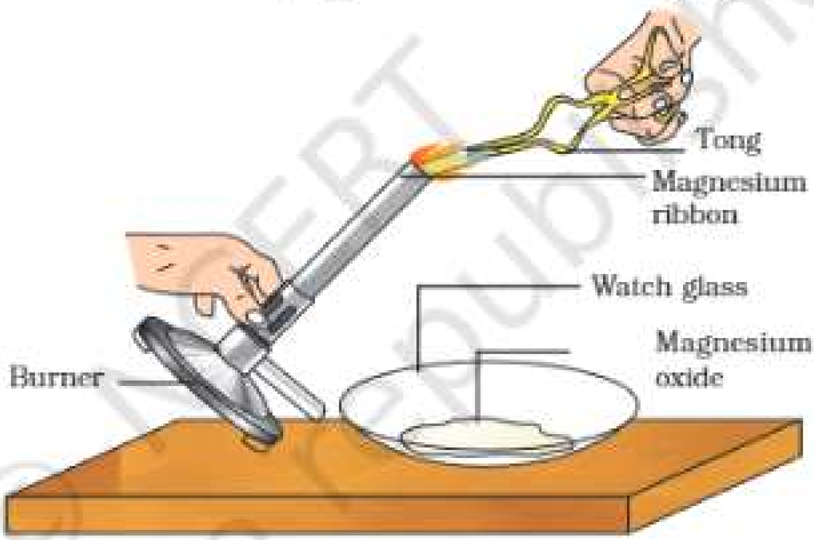


Fig. 1.2 : *Burning of magnesium ribbon and collection of magnesium oxide on a watch glass*

3. Collect the white powder obtained on a watch glass.
4. Transfer and mix the white powder in a beaker containing a little amount of distilled water.
5. Put drops of this mixture over the red and blue litmus papers and record your observations.

Observations

On putting a drop of mixture over the red litmus paper, colour of red litmus paper changes into ____ .

On putting a drop of mixture over the blue litmus paper, colour of blue litmus paper changes into ____ .

Results and Discussion

The change in the colour of ____ litmus paper into _____ suggests that the aqueous solution of magnesium oxide is _______ in nature.

Preacautions

- Clean the magnesium ribbon carefully to remove the deposited oxide layer on it.
- Burn the magnesium ribbon keeping it away from your eyes as far as possible and use dark coloured goggles to see dazzling light emitted during burning of magnesium. (Why?)
- Collect magnesium oxide powder carefully so that it does not touch your skin.

Note for the Teacher

- Oxides on account of their interacting capability with water are classified as acidic, basic and neutral oxides.
- Magnesium oxide (MgO) dissolves in water to form magnesium hydroxide $Mg(OH)_2$ (aq) which is a strong base.

  $MgO\ (s) + H_2O\ (l) \longrightarrow Mg\ (OH)_2\ (aq)$

  Here the reaction is:

  $O^{2-}\ (s) + H_2O\ (l) \longrightarrow 2OH^-\ (aq).$
- It is advised to tilt the burner in order to collect the magnesium oxide (product).

Questions

- Why should magnesium ribbon be cleaned before burning it in air?
- Which reaction that takes place when magnesium burns in air? Why is it called a combination reaction?
- Why does the red litmus paper turn blue when touched with aqueous solution of magnesium oxide?
- What is the total electron content of the species Mg^{2+} and O^{2-} ? Name 5 more such species?
- Is there a possibility of a compound other than MgO formed in the above reaction?
- Is there any similarity between compounds LiH, MgO, and K_2S?
- Why is it suggested to wear dark coloured goggles while watching the burning of magnesium ribbon in air?

Experiment 2

AIM

To study the following chemical reactions: (a) zinc with sulphuric acid; (b) precipitation reaction between aqueous solution of barium chloride and aqueous solution of sodium sulphate; and (c) thermal decomposition of ammonium chloride in an open container.

(a) Chemical reaction of zinc with sulphuric acid.

THEORY

Zinc metal reacts with dil. sulphuric acid and produces hydrogen gas.

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

This is an example of a single displacement reaction of a non-metal by a metal.

MATERIALS REQUIRED

Zinc metal granules, dil. sulphuric acid, red and blue litmus papers, test tube, and candle.

PROCEDURE

1. Take few zinc granules in a test tube.
2. Add about 10 mL of dil. sulphuric acid to zinc granule. Effervescence

comes out from the reaction mixture, as shown in Fig. 2.1.

3. Perform the tests as given in the observation table and record your observations.

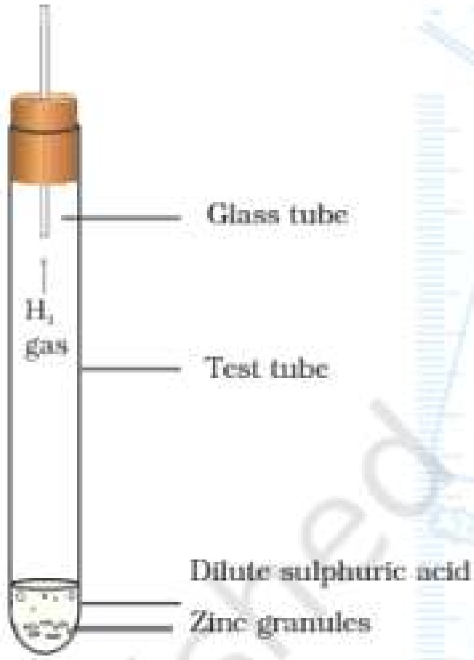


Fig. 2.1 : *Reaction of zinc granules with dil. sulphuric acid.*

Observations

Sl. No.	Test	Activity	Observations
1.	Colour	Look at the colour of the gas liberated	
2.	Smell	Fan the gas gently towards your nose with your hand	
3.	Litmus test	Bring moist blue and red litmus papers near to the mouth of the test tube	
4.	Combustion test	Bring a lighted candle near to the mouth of the test tube	

Results and Discussion

Infer from the observations about the nature of the gas liberated. Is it acidic or basic or neutral? Does it burn in air (or ignites exothermically) to produce water?

Precautions

- Clean zinc granules should be used.
- Care should be taken while pouring the dil. sulphuric acid in the test tube and performing combustion test.

NOTE FOR THE TEACHER

- The combustion test must be performed very carefully. It is advised that this test may first be demonstrated in the laboratory.

QUESTIONS

- Write the chemical reaction of zinc with dil. sulphuric acid.
- How does the combustion of hydrogen gas produce water?
- How will you show that the hydrogen gas is neutral in behaviour?
- What are the others metals among the species Mg, Al, Fe, Sn, Pb, Cu, Ag metals which react with dil. sulphuric acid to produce hydrogen gas?
- Which of the above metal(s) would not evolve hydrogen gas from dilute hydrochloric acid?

(b) Precipitation reation between aqueous solution of barium chloride with aqueous solution of sodium sulphate.

THEORY

When a solution of sodium sulphate is mixed with a solution of barium chloride, the following double displacement reaction takes place:

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

In this reaction, sulphate ions (SO_4^{2-}) from sodium sulphate are displaced by chloride ions (Cl-) and chloride ions in barium chloride are displaced by sulphate ions. As a result, a white precipitate of barium sulphate is formed and sodium chloride remains in the solution.

MATERIALS REQUIRED

Two test tubes, a small measuring cylinder (50 mL), aqueous solution of sodium sulphate, aqueous solution of barium chloride.

PROCEDURE

1. Take 3 mL of sodium sulphate solution in a test tube and label it as A.
2. In another test tube, take 3 mL of barium chloride and label it as B.

3. Transfer the solution from test tube A to the test tube B.
4. Mix the two solutions with gentle shaking.
5. Observe the changes in colours of the solutions as per the steps given in observation table below.

OBSERVATIONS

Sl.No.	Experiment	Observations
1.	Observe the colour of the two solutions in test tubes A and B before mixing them.	
2.	Mix the two solutions and leave the mixture undisturbed for some time. Does anything precipitates in the test tube? If so, what is the colour of it?	

RESULTS AND DISCUSSION

Confirm whether you have obtained a white precipitate of barium sulphate in the test tube. Does it suggests that the substances which produce ions in water result into precipitation reaction under favourable condition?

NOTE FOR THE TEACHER

- The aqueous solutions of barium chloride and sodium sulphate can be prepared by dissolving 6.1 g $BaCl_2.2H_2O$ and 3.2 g of $Na_2SO_4.10H_2O$ in water and then diluting them to 100 mL separately.

QUESTIONS

- Fill in the blanks:
  (a) Sodium sulphate and barium chloride are ______ (ionic/ covalent) compounds.

(b) As the white precipitate of barium sulphate is formed __________ (immediately/sometime after mixing the two solutions), the reaction between __________ (ionic/ covalent) compounds is ____________ (instantaneous/ slow).

- What may happen on mixing $Pb(NO_3)_2$ and KCl solutions? Predict (you may try to experimentally verify).
- What are the industrial applications of the type of the reaction being studied?
- Why do the persons suffering from the ailment of stone formation advised not to take too much milk and tomato juice?

(c) Thermal decomposition of ammonium chloride in an open container.

THEORY

Ammonium chloride on heating in an open container is decomposedinto hydrogen chloride and produces ammonia gas. This is an example of decomposition reaction.

$$NH_4Cl\,(s) \xrightarrow{\text{Heat}} NH_3\,(g) + HCl\,(g)$$

MATERIALS REQUIRED

Ammonium chloride, Nessler's reagent $K_2[HgI_4]$, blue litmus paper, laboratory stand with clamp, tripod stand, burner, china dish, wire gauge, and a funnel.

PROCEDURE

1. Take about 5 g of ammonium chloride in a clean and dry china dish.
2. Place the china dish on a wire gauge and keep it on a tripod stand.
3. Place an inverted clean and dry funnel over the china dish containing the sample.
4. Heat the china dish containing the sample of ammonium chloride (Fig. 2.2).
5. Vapours are formed that come out from the stem of the funnel. Check whether any liquid is produced in the china dish?

6. Bring a strip of filter paper dipped in Nessler's reagent $K_2[HgI_4]$ near the tip of funnel. Observe the change in the colour of the filter paper.
7. Bring a wet blue litmus paper near the tip of the funnel. Observe the change in its colour.

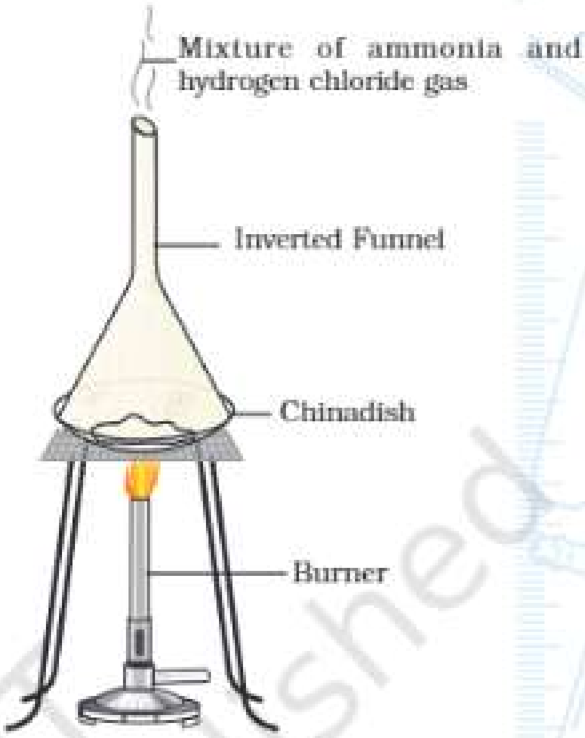


Fig. 2.2 : *Heating of ammonium chloride in an open container*

Observations

Sl. No.	Experiment	Observations	Inference
1.	Nessler's reagent test		
2.	Litmus paper test		

Results and Discussion

Infer from the observations from the Nessler's reagent test and litmus paper test on the vapours evolving from the funnel for the presence of ammonia and hydrogen chloride gases respectively. Now conclude that ammonium chloride when heated in open system, decomposes to give ammonia and hydrogen chloride gases.

Precautions

- The heating must be stopped when most of the ammonium chloride is decomposed.

Note for the Teacher

- If this reaction takes place in a closed container, the hydrogen chloride and ammonia gases cannot escape. (This reaction can be performed by tightly plugging the top of the stem of the funnel by cotton.) These gases then recombine to form ammonium chloride (NH_4Cl).

  $$NH_3\ (g) + HCl\ (g) \xrightarrow{\text{Cool}} NH_4Cl\ (s)$$

  Thus, an equilibrium exists between ammonium chloride, ammonia and hydrogen chloride in a closed container.

  In this reaction, solid ammonium chloride is converted directly into gaseous state without changing into liquid. It is thus sublimation reaction.

- *Preparation of Nessler's Reagent* : Dissolve 10 g of potassium iodide in 10 mL water (solution A). Dissolve 6 g mercury(II) chloride in 100 mL water (solution B). Dissolve 45 g potassium hydroxide in water and dilute to 80 mL (solution C). Add solution B to solution A dropwise, until a slight permanent precipitate is formed. Add solution C to it, mix and dilute with water to 200 mL. Allow it to stand overnight and decant the clear solution.

Questions

- What gases are liberated on heating ammonium chloride?
- How will you distinguish between hydrogen chloride and ammonia gases in a laboratory?
- Can you think of decomposing water into its elemental components H_2 and O_2, using this method?
- How does the decomposition reaction

  $$2Pb(NO_3)_2 \xrightarrow{\text{Heat}} 2PbO\ (s) + 4\ NO_2\ (g) + O_2\ (g)$$

  differ from the one being discussed in this experiment?
- Limestone decomposes thermally into quick lime. What is the industrial importance of this chemical reaction?
- On thermal decomposition, ammonium chloride produces a mixture of ammonia gas (basic) and hydrogen chloride gas (acidic). This gas mixture does not show neutral behaviour in litmus test. Why?

Experiment 3

AIM

To measure the change in temperature during chemical reactions and to conclude whether the reaction is exothermic or endothermic.

THEORY

Most of the chemical reactions are accompanied by energy changes. In some reactions, energy is absorbed while in some energy is released in the form of heat. The chemical reactions in which energy is absorbed are called endothermic reactions and those in which energy is released are known as exothermic reactions. The reaction can be identified as exothermic or endothermic by measuring the change in temperature of the reaction mixture.

In this experiment the following chemical reactions can be carried out:

(i) $NaOH\ (aq) + HCl\ (g) \longrightarrow NaCl\ (g) + H_2O\ (l)$; and

(ii) $Ba(OH)_2.8H_2O\ (s) + 2NH_4Cl\ (s) \longrightarrow BaCl_2\ (aq) + 10H_2O\ (l) + 2NH_3\ (aq)$

MATERIALS REQUIRED

Sodium hydroxide solution, hydrochloric acid, ammonium chloride (solid) and barium hydroxide (solid), weighing balance, watch glass, four beakers (100 mL), a thermometer (–10 °C to 110 °C), and a glass rod.

PROCEDURE

1. Mark all the four clean beakers as 1, 2, 3, and 4.
2. Take 20 mL of sodium hydroxide solution in beaker no. 1; 20 mL of hydrochloric acid in beaker no. 2; 15.75 g of barium hydroxide in beaker no. 3, and 5.35 g of ammonium chloride in beaker no. 4.
3. Successively insert a thermometer in each beaker for some time and record their temperatures. Also record the room temperature.
4. To see the reaction of sodium hydroxide solution with hydrochloric acid, pour the contents of beaker no. 1 in beaker no. 2. Quickly insert the thermometer in the reaction mixture. Note and record its initial temperature reading. Stir well the reaction mixture gently using a glass rod. Note and record the final temperature reading of the thermometer. Wash the thermometer and glass rod after noting the readings.
5. Similarly, to see the reaction of barium hydroxide solution with ammonium chloride, pour the contents of beaker no. 3 in beaker no. 4. Quickly insert the thermometer in this reaction mixture. Note and record the initial temperature. Stir well the reaction mixture gently using the glass rod. Note and record the final temperature readings of the thermometer.

OBSERVATIONS AND CALCULATIONS

(i) Temperature of the sodium hydroxide solution = ___ °C = ___ K

(ii) Temperature of the hydrochloric acid = ___ °C = ___ K

(iii) Temperature of the barium hydroxide solution = ___ °C = ___ K

(iv) Temperature of the ammonium chloride = ___ °C = ___ K

(v) Room temperature = ___ °C = ___ K

Sl. No.	Reactants of the reaction	Initial temperature of the reaction mixture	Final temperature of the reaction mixture	Change in temperature
		θ_1 (°C)	θ_2 (°C)	$\theta_2 - \theta_1$ (°C)
1.	$NaOH + HCl$			
2.	$Ba(OH)_2.8H_2O + 2NH_4Cl$			

Results and Discussion

Based on your observations for the change in tempearture in two reactions, infer about the nature of the two chemical reactions (exothermic or endothermic).

The reaction between sodium hydroxide solution and hydrochloride acid is ____________________ ; and the reaction between barium hydroxide solution and ammonium chloride is ____________________ (exothermic or endothermic).

Precautions

- Stir the reaction mixture very gently so that there is no heat loss during stirring.
- Wash the thermometer and glass rod with water before inserting it in another reactant or reaction mixture.

Questions

- The reaction between HCl and NaOH in its simplified version is

  $H^+ (aq) \quad + \quad OH^- (aq) \longrightarrow H_2O (l)$

  (from acid) (from base)

  Can you assign a plausible explanation as to why the reaction should be exothermic?
- Consider the changes;

  $2HCl (g) \longrightarrow H_2 (g) + Cl_2 (g)$;

  $2Mg (s) + O_2 (g) \longrightarrow 2MgO (s)$.

  Which according to you is exothermic change?
- What precautions did you take while measuring the temperature of a reaction mixture?

Experiment 4

AIM

To study the reactions of hydrochloric acid with zinc metal, sodium carbonate, and sodium hydroxide.

THEORY

An acid (HCl) reacts (i) with zinc metal to produce hydrogen gas, (ii) with carbonates and hydrogen carbonates to form carbon dioxide gas, and (iii) with sodium hydroxide (base) to neutralise and to produce sodium chloride (salt) and water.

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

$$Na_2CO_3\ (s) + 2HCl\ (aq) \longrightarrow 2NaCl\ (aq) + H_2O\ (l) + CO_2\ (g)$$

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

MATERIALS REQUIRED

Zinc metal granules, dil. hydrochloric acid, sodium carbonate, sodium hydroxide solution, freshly prepared lime water, red and blue litmus papers, distilled water, four test tubes, a delivery tube, single bore cork to be fixed on a test tube, and a piece of sand paper.

PROCEDURE

(i) *Reaction with Zinc Metal*

1. Take a clean zinc granule in a clean and dry test tube.

2. Put about 5 mL of dil. hydrochloric acid into it.
3. Effervescence will come out from the reaction mixture.
4. Successively bring wet blue and red litmus papers to the mouth of the test tube. Note and record the observation.

OBSERVATIONS

Sl.No.	Experiment	Observations	Inference
1.	Litmus test: Action on red litmus Action on blue litmus		

(ii) *Reaction with Sodium Carbonate*

1. Take about 1 g of sodium carbonate in a clean and dry test tube.
2. Add about 2 mL of dil. hydrochloric acid to it.
3. Effervescence will start coming from the reaction mixture.
4. Fix a delivery tube through a cork to the mouth of the test tube and pass the liberated gas through the freshly prepared lime water (Fig. 4.1). Observe what happens? Do you see bubbles of it in lime water? Does it turn milky? If yes, it shows the presence of carbon dioxide.

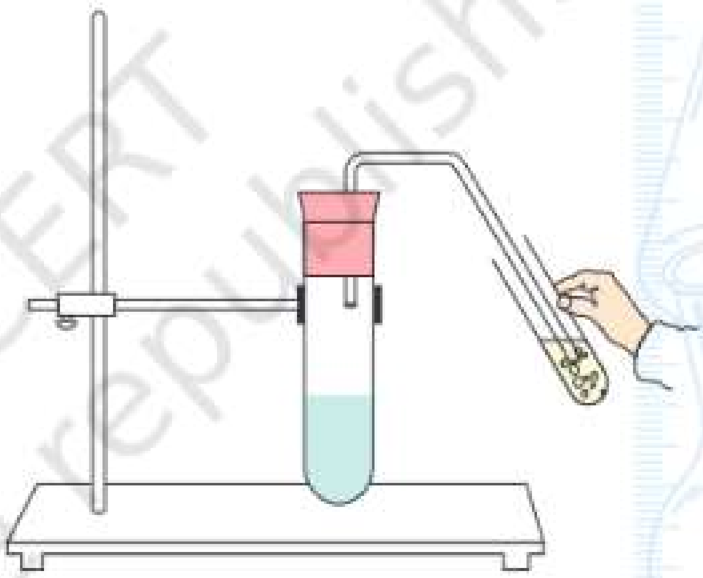


Fig. 4.1 : *Passing of liberated gas through the freshly prepared lime water*

OBSERVATIONS

Sl. No.	Experiment	Observations	Inference
1.	Lime water test		

(iii) *Reaction with Sodium Hydroxide*

1. Take about 5 mL dil. hydrochloric acid in a test tube and label it as A.
2. Similarly, take 5 mL of 10% sodium hydroxide solution in another test tube and label it as B.

3. Dip a blue litmus paper in test tube A containing dil. HCl. What do you see? Do you find the the blue litmus paper turns red.
4. Similarly, dip a red litmus paper in test tube B. Does it turn blue now?
5. Add dil. HCl from test tube A dropwise to dil. NaOH contained in test tube B.
6. Shake the mixture slowly but continously and observe the change by dipping litmus paper in the test tube B. (Which litmus paper will you use for this purpose?)
7. Keep on adding the dil. HCl from test tube A to 10% NaOH in test tube B dropwise till the reaction mixture in test tube B becomes neutral to litmus paper. Ascertain the neutrality of this mixture by successive dipping red and blue litmus papers.
8. Touch the test tube and feel the temperature. Do you find it warm or cold? What does that mean?

OBSERVATIONS

Sl. No.	Activity	Observations	Inference
1.	**Litmus Paper Test** *In the beginning of experiment:* Dip blue litmus paper in test tube A Dip red litmus paper in test tube B After adding *n* drops of dil. HCl from test tube A in dil. NaOH in test tube B *(i)* n = ___ : *action on red litmus paper* *action on blue litmus paper* *(ii)* n = ___ : *action on red litmus paper* *action on blue litmus paper* *(iii)* n = ___ : *action on red litmus paper* *action on blue litmus paper* *(iv)* n = ___ : *action on red litmus paper* *action on blue litmus paper* ...___ n = ___ : *action on red litmus paper* *action on blue litmus paper*	 *No change* *No change*	 *The solution in test tube B is neutralised*
2.	**Thermal Change** After the completion of litmus test, touch the test tube B from outside Heat absorbed/evolved during the reaction	Cold/Warm	Reaction is endothermic/ exothermic

Results and Discussion

State and discuss the performance of each test in all reactions performed in this experiment.

Precautions

- Always carry out the test for hydrogen with a very small volume of gas.
- Handle hydrochloric acid and sodium hydroxide solutions very carefully.
- Shake the solutions and reaction mixtures carefully without spilling.
- Care must be taken while performing the combustion test.

Note for the Teacher

- *Preparation of lime water*: Shake 5 g calcium hydroxide $Ca(OH)_2$, with 100 mL water. Allow it to stand for about 24 hours. Decant the supernatant liquid and use it for the tests. It is suggested to always use freshly prepared limewater.

Questions

- What will be the colour of a blue litmus paper on bringing it in contact with a drop of dil. hydrochloric acid?
- Explain why hydrogen gas is not collected by the downward displacement of air?
- What will happen to a lighted candle if it is brought near the mouth of a gas jar containing hydrogen gas?
- Which gas is produced when zinc metal reacts with hydrochloric acid?
- Which gas is liberated when sodium carbonate reacts with hydrochloric acid?
- Hydrogen gas is neutral to litmus paper. Explain how?
- What is the utility of the reaction between $NaHCO_3$ and HCl in daily life situation?
- How can the deposits of carbonates and hydrogencarbonates on the metal surface be cleaned?

Experiment 5

AIM

To study the reactions of sodium hydroxide with aluminum metal and hydrochloric acid.

THEORY

Sodium hydroxide is a base. It reacts with aluminium metal to produce hydrogen gas. It also neutalises the hydrochloric acid to produce sodium chloride salt and water.

$$2NaOH\ (aq) + 2Al\ (s) + 2H_2O\ (l) \longrightarrow 2NaAlO_2\ (aq) + H_2\ (g)$$

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

MATERIALS REQUIRED

Dil. hydrochloric acid, sodium hydroxide solution, some pieces of aluminium metal, red and blue litmus papers, a small measuring cylinder (100 mL), three test tube, and a candle.

PROCEDURE

(i) *Reaction with Aluminium Metal*

1. Take a small piece of aluminium metal and place it in a clean and dry test tube.

2. Add about 5 mL sodium hydroxide solution in it.
3. Observe the effervescence coming out from the reaction mixture. Look at the colour of the gas liberated.
4. Perform the smell test on the gas liberated by fanning the gas gently towards your nose.
5. Bring moist blue and red litmus papers to the mouth of the test tube.
6. Perform combustion test by bringing a lighted candle near to the mouth of the test tube. Does the liberated gas ignites exothermically to produce water?

Observations

(i) *Reaction with Aluminium Metal*

Sl. No.	Test	Experiment	Observations	Inference
1.	Colour	Look at the colour of the gas liberated		
2.	Smell	Fan the gas gently towards your nose with your hand		
3.	Litmus test	Bring moist blue and red litmus papers near to the mouth of the test tube		
4.	Combustion test	Bring a lighted candle near to the mouth of the test tube		

(ii) *Reaction with hydrochloric acid*

The experiment should be carried out as done in Experiment 4.

Results and Discussion

State and discuss the performance of each test in all reactions performed in this experiment.

Precautions

- Always carry out the test for hydrogen with a very small volume of gas.
- Handle hydrochloric acid and sodium hydroxide solutions very carefully.
- Shake the solutions and reaction mixtures carefully without spilling.
- Care must be taken while performing the combustion test.

QUESTIONS

- What will be the colour of a blue litmus paper on bringing it in contact with a drop of dil. NaOH?
- Explain why hydrogen gas is not collected by the downward displacement of air?
- What will happen to a lighted candle if it is brought near the mouth of a gas jar containing hydrogen gas?
- Which gas is produced when aluminium metal reacts with sodium hydroxide?
- Hydrogen gas is neutral to litmus paper. Explain how?
- What are the metals (other than Al) which react with alkalies to produce hydrogen gas? What are these metals called?

Experiment 6

Aim

To show that acids, bases, and salts are electrolytes.

Theory

An electrolyte is a compound that, in solution or in the molten state, conducts an electric current and is simultaneously decomposed by it. The current in electrolytes is carried by the ions and not by the electrons as in metals. Electrolytes may be acid, bases, or salts. In this experiment we shall observe it by means of continuity test in an electric circuit that contains either an acid or a base or a salt solution as a part of it.

Materials Required

Hydrochloric acid (about 5 mL), sodium hydroxide flakes (about 100 mg), sodium chloride (about 5 g), distilled water, four beakers (250 mL), four dry cells of 1.5 V each with a cell holder (or a battery of 6 V or a battery eliminator), a torch bulb of 6 V with a torch bulb holder, a rubber cork, two iron nails, a plug key, connecting wires, and a piece of sand paper.

Procedure

1. Using a sand paper, clean the insulation layers from the ends of connecting wires.

2. Take a dry rubber cork and fix two iron nails in it at a distance. The two nails will work as two electrodes. Also connect these two nails, separately, with connecting wires.
3. Draw a circuit diagram for performing a continuity test in an electric circuit that contains either an acid or a base or a salt solution as a part of it (see Fig. 6.1). Observe how different components like the dry cells (or battery or battery eliminator), torch bulb, a plug key, and the solution are connected in the circuit.
4. Take nearly 100 mL distilled water in each of the four beakers (250 mL). Label them as beakers A, B, C, and D respectively.
5. Add about five drops of hydrochloric acid in distilled water in beaker A to get an acidic solution; add about 100 mg flakes of sodium hydroxide in beaker B to get a basic solution; and add about 2 - 3 g of sodium chloride salt (about half a teaspoon) in water in beaker C to get a sodium chloride salt solution. Do not add anything in the distilled water in beaker D.

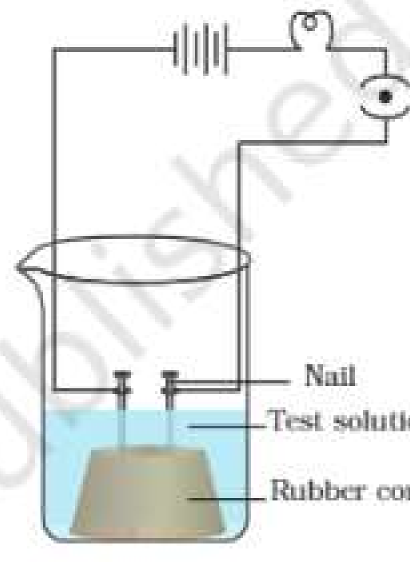


Fig. 6.1 : *Continuity test through an electrolyte*

6. Set up the electric circuit by connecting different components with the help of connecting wires. Do not dip the rubber cork (in which two iron nails are fixed and connected in the circuit) in any beaker. Insert the key into the plug. Check whether the torch bulb glows. It does not. Does it mean that the electric circuit is yet not complete or the dry rubber cork does not conduct electricity? Remove the key from the plug.
7. For observing the continuity test through the dil. hydrochloric acid (say), place the rubber cork in the beaker A such that the two iron nails are partially dipped in the solution.
8. Insert the key in the plug and allow the current to flow in the circuit containing dil. hydrochloric acid solution as a component. Does the bulb glow now? Yes, it glows. It means that the electric circuit is now complete and that the hydrochloric acid conducts electricity. Thus it is an electrolyte. Record your observation.

9. Remove the key and take out the rubber cork from the beaker A. Wash the rubber cork and make it dry using a clean cloth.
10. Repeat the experiment for the continuity test through the dil. sodium hydroxide solution, sodium chloride solution, and distilled water by successively dipping the rubber cork in beakers B, C, and D respectively.

Observations

Sl.No.	Experiment	Observations	Inference
	Electric continuity test through		
1.	Beaker A: dil. hydrochloric acid solution	Bulb glows or not?	
2.	Beaker B: dil. sodium hydroxide solution	Bulb glows or not?	
3.	Beaker C: sodium chloride solution	Bulb glows or not?	
4.	Beaker D: distilled water	Bulb glows or not?	

Results and Discussion

Infer from the observations that acids, bases and salts are electrolytes. Discuss the following dissociation reactions:

$$HCl\ (aq) \longrightarrow H^+\ (aq) + Cl^-\ (aq)$$
$$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$$
$$NaCl\ (aq) \longrightarrow Na^+\ (aq) + Cl^-\ (aq).$$

Precautions

- The ends of the connecting wires must be cleaned and connected tightly with the other components of the circuit.
- The acidic concentration in the distilled water must be highly dilute otherwise the nails will start reacting with the acid.
- The nails must be partially dipped inside the liquid while performing electric continuity test.
- The rubber cork must be washed and dried after every test and before dipping it in another liquid solution.

NOTE FOR THE TEACHER

- In place of four cells of 1.5 V each, a 6 V battery or a battery eliminator may also be used. Please make sure that if a 6 V source is used in the circuit, a torch bulb of 6 V must be used. This experiment can also be performed with 3 V source preferably with a 3 V torch bulb.
- In place of a torch bulb, a galvanometer or an ammeter (0 - 3 A) may also be used to perform the continuity test. Please also connect a resistor of about 1 or 2 W resistance in series with the galvanometer or ammeter.
- Experiment Nos. 48 to 51, involve observations with electric circuits. It is advised that students may be suggested to perform any of these experiments before performing this experiment.
- In place of hydrochloric acid solution, sulphuric acid solution may also be used.

QUESTIONS

- Though sodium chloride and potassium chloride crystals are composed of ions. Why do they not conduct electricity?
- How does an alcoholic solution of potassium hydroxide conduct electricity?
- How does the hydrochloric acid solution proves to be a better conductor of electricity than acetic acid solution (CH_3COOH)?
- Which substance is used as an electrolyte in lead storage battery and which one in dry cells.
- What are the current carriers in electrolytes?

Experiment 7

AIM

To find the pH of the given samples of solutions of solids or fruit juices using pH paper.

THEORY

The pH is the measure of the acidic (or basic) power of a solution. It is a scale for measuring hydrogen ion concertration in a solution. The pH scale varies from 0 to 14. At 25 °C (298 K), a neutral solution has pH equal to 7. A value less than 7 on the pH scale represents an acidic solution. Where as pH value more than 7 represents basic solution. Generally a paper impregnated with the universal indicator is used for finding the approximate pH value. It shows different colour at different pH [Fig. 7.1(b)].

MATERIALS REQUIRED

Test solutions of samples (a) a dilute acid (HCl or H_2SO_4 or CH_3COOH etc.); (b) a dilute base (NaOH or KOH); (c) a salt (such as NaCl, Na_2CO_3, NH_4Cl etc.;1 g salt in 10 mL distilled water); (d) soil water extract (dissolve 1 g of soil sample in 10 mL distilled water and filter to get a soil water extract); and (e) a fruit juice, five test tubes and a test tube stand, a measuring cylinder (10 mL), pH papers, and a glass rod.

PROCEDURE

1. Place five clean test tubes in a test tube stand.
2. Take the solutions of a dilute acid (say HCl), dilute base (say NaOH), salt (say NaCl), soil, and a fruit juice separately in five test tubes and label them.
3. Put one or two drops of each test solution on different strips of pH papers, using a glass rod [Fig. 7.1(a)]. Glass rod used for one sample must be washed with water before used for the other sample.

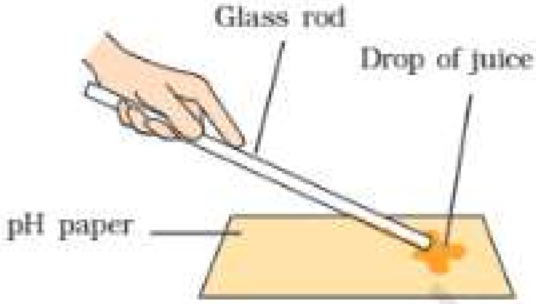


Fig. 7.1 : *(a) Testing the pH of a sample by putting a drop on pH paper by glass rod*

4. Note the pH by comparing the colour appeared on the pH paper with those on colour chart for pH paper [Fig. 7.1(b)].
5. For determining the pH of a fruit juice, squeeze the fruit and place 1 or 2 drop of the juice on the pH paper.

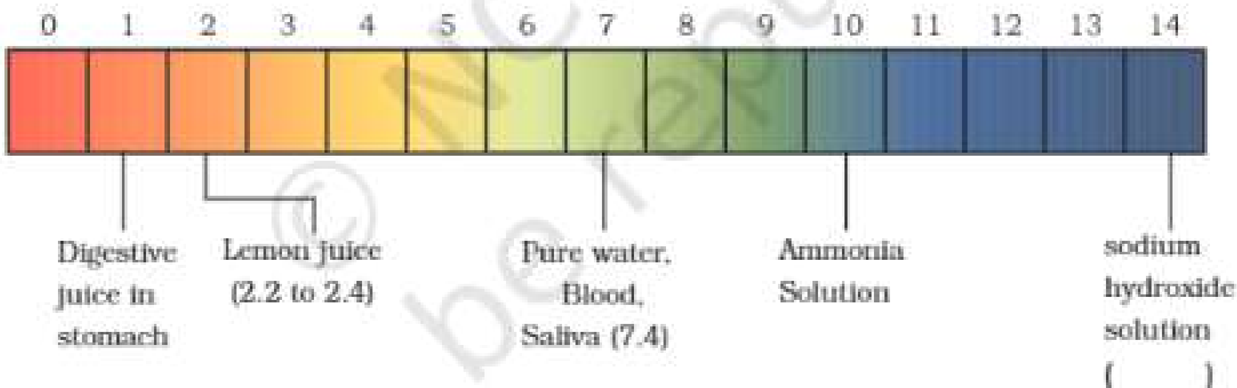


Fig. 7.1 : *(b) Colour of universal indicator at different pH*

OBSERVATIONS

Sl. No.	Sample	Approximate pH
1.	Dilute acid (HCl)	
2.	Dilute base (NaOH)	
3.	Salt solution (NaCl)	
4.	Soil water extract	
5.	A fruit (__________) juice	

Results and Discussion

As pH depends upon H^+ concentration and in an aqueous solution H^+ and OH^- ion concentrations are correlated, therefore, every acidic and basic solution shows different colour at different pH.

Precautions

- The test sample solutions should be freshly prepared and the fruit juice samples should also be fresh.
- Glass rod used for one sample should be used for the other sample only after washing it with water.

Note for the Teacher

- It is advised to explain the pH value of salt solutions and differentiate between acid and acidic compounds, bases and basic compounds.
- Teachers may take a solid chemical like oxalic acid and juices of citrus fruits, carrot, grapes etc., for making solutions for determining their pH values. Students may be suggested to compare the pH values of juices of unripe and ripe fruits and note the change in pH during ripening.

Questions

- What do you mean by pH?
- What is the pH of pure water at 25 °C (298 K)?
- What according to you should be the pH of dil. HCl and dil. NaOH solutions? Observe and explain your findings.
- On opening the soda water bottle the dissolved CO_2 comes out, would the pH of the solution increase or decrease as the gas comes out? Explain your answer either way.

Experiment 8

AIM

To identify bleaching powder among given samples of chemicals.

THEORY

Bleaching powder is calcium oxochloride ($CaOCl_2$). On treatment with small quantity of dilute acid, it liberates hypochlorous acid which can easily furnish oxygen (called nascent oxygen) and thus acts as an oxidising and bleaching agent.

$$2CaOCl_2\ (s) + H_2SO_4\ (aq) \longrightarrow CaSO_4\ (aq) + CaCl_2\ (aq) + 2HClO\ (aq)$$

$$HClO \rightarrow HCl + \underset{\text{(Nasent oxygen)}}{[O]}$$

In this experiment, we shall make use of this bleaching reaction to identify the bleaching powder from the given samples of chemicals (four, say).

MATERIALS REQUIRED

Given four samples: bleaching powder; sodium chloride; calcium chloride; and ammonium chloride (or alternate salts), dil. sulphuric acid, flowers petals or small pieces of coloured cotton cloth, eight beakers (100 mL), a measuring cylinder (100 mL), and a glass rod.

PROCEDURE

1. Prepare about 50 mL, 5% solution (by volume) of each of the four given samples of chemicals in four beakers. Label these beakers as A, B, C, and D.
2. Take about 20 mL of dilute sulphuric acid in each of remaining four beakers. Label them as E, F, G, and H.
3. Dip a small piece of coloured cloth or flower petal in beaker A.
4. Take out the cloth or flower petal from the beaker A and dip it in dil. sulphuric acid in beaker E and stir it gently with the help of a glass rod. Does the cotton cloth or flower petal decolourise? Record your observation.
5. Repeat steps 3 and 4 with other three samples of given chemicals and dil. sulphuric acid and record your observations.

OBSERVATIONS

Sl. No.	Sample	Colour of the cloth or flower petal dipped in the solution of sample chemical	Colour of cloth or flower petal (dipped in sample solution and then in dil. sulphuric acid)
1.	A		
2.	B		
3.	C		
4.	D		

RESULTS AND DISCUSSION

Infer from the observations that which chemical solution decolourises the cotton cloth or flower petal. The chemical which decolourises, exhibits bleaching action. Thus the bleaching powder can be identified.

In this case, the solution in beaker ___ shows bleaching reaction and therefore the chemical in solution in that beaker is bleaching powder. The nascent oxygen produced by the decomposition of hypochlorous acid (HClO) is the cause for bleaching action.

PRECAUTIONS

- Handle the sample solutions and sulphuric acid carefully. These must not touch your skin.
- Glass rod used for one sample solution should be used for the other sample solution only after washing it with water.

Note for the Teacher

- In the samples of chemicals, sulphites (SO_3^{2-}) and hydrogen sulphites (HSO_3^-) should not be given because these chemicals react with dil. sulphuric acid and produce sulphur dioxide gas which also acts as a temporary bleaching agent.

Questions

- Name the substance which on treating with chlorine yields bleaching powder.
- Why does the bleaching powder known as a mixture?
- What happens when bleaching powder is exposed to air?
- How does the bleaching powder help in the purification of water?
- What is the chemical name of bleaching powder?

Experiment 9

AIM

To identify washing soda or baking soda among given samples of chemicals.

THEORY

Washing soda ($Na_2CO_3.10H_2O$) and baking soda (mainly $NaHCO_3$) are white solids. Their aqueous solutions are alkaline and turn red litmus blue. Carbonates and hydrogencarbonates react with dilute acids and produce carbon dioxide gas which turns lime water milky.

$$CO_3^{2-}\,(aq) + H_2O\,(l) \rightarrow H_2CO_3\,(aq) + 2OH^-\,(aq)$$

$$HCO_3^-\,(aq) + H_2O\,(l) \longrightarrow H_2CO_3\,(aq) + OH^-\,(aq)$$

$$Na_2CO_3\,(s) + 2HCl\,(aq) \longrightarrow 2NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$$

$$NaHCO_3\,(s) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$$

$$Ca(OH)_2\,(aq) + CO_2\,(g) \longrightarrow CaCO_3\,(s) + H_2O\,(l)$$

Lime water Milky

On passing excess of CO_2 through limewater, calcium hydrogencarbonate is formed. It is soluble in water and forms a colourless solution.

$$CaCO_3\,(s) + H_2O\,(l) + CO_2\,(g) \longrightarrow Ca(HCO_3)_2\,(aq).$$

Materials Required

Samples: sodium carbonate (washing soda), sodium hydrogencarbonate (baking soda), ammonium chloride, sodium chloride etc., red litmus paper strips, freshly prepared lime water, dil. hydrochloric acid, five test tubes, a test tube stand, a boiling tube, a thistle funnel, a double bored cork, a delivery tube, and a glass rod.

Procedure

1. Take about 1 g each of the four given samples separately in four boiling tubes and label them as A, B, C, and D.
2. Add about 5 mL distilled water in each boiling tube. Gently shake the contents of the tubes.
3. Put a drop of every salt's solution on separate red litmus paper strips, using a glass rod. (Wash the glass rod used for one sample before using it for other sample.) Note the change in colour of the litmus paper, if any in each case.
4. Add 1 mL of dil. hydrochloric acid in each test tube. Do you see any effervescence from any test tube? If yes, perform the lime water test as detailed below.
5. For performing lime water test, take the solution of test tube A in a boiling tube and set up the apparatus (delivery tube, thistle funnel etc.) as shown in Fig. 9.1.
6. Add dil. hydrochloric acid drop by drop to the solution through the thistle funnel.
7. Pass the liberated gas evolved through the lime water in a test tube. Does the lime water turns milky? If yes, then it shows the presence of CO_2 gas.
8. Continue passing the liberated gas through the lime water. Does it again become colourless? This reconfirms that the liberated gas is CO_2.
9. Repeat the lime water test on all samples that give effervescence in step 4. Do not forget to wash the boiling tube when you change the sample in it for performing the lime water test.

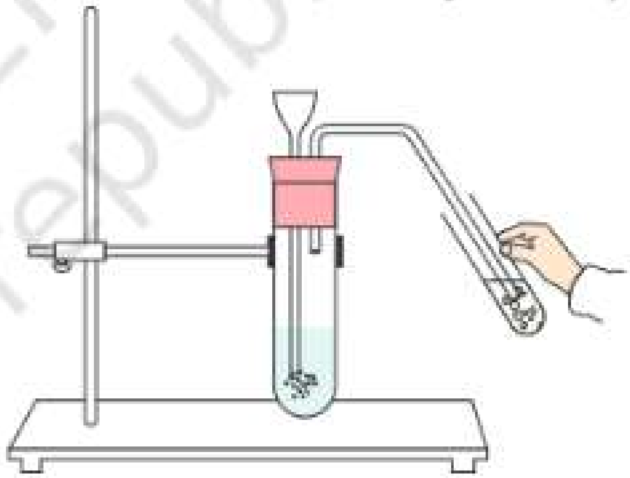


Fig. 9.1 : *Carbon dioxide gas formed by the reaction of dil/hydrochloric acid on washing soda or baking soda is being passed through lime water*

Observations

Sl. No.	Sample	Colour	Solubility in water	Action on red litmus paper	Action of dil. HCl acid	Lime water
			(Soluble/ insoulble)	(Changes to blue or not)	(Effervescence observed or not?)	Turns milky or not?
1.	A					
2.	B					
3.	C					
4.	D					

Results and Discussion

Infer from the observations about the identification of washing soda or baking soda out of the samples given for testing. Discuss about the litmus paper and lime water tests performed.

Sample in test tube ___ is washing soda/baking soda.

Precautions

- Add dil. hydrochloric acid to the salt solution drop by drop. If the addition of dil. hydrochloric acid is not slow, a vigorous reaction may occur and the reaction mixture may come out of the reaction tube and pass into lime water.
- Handle hydrochloric acid and washing soda carefully. These should not touch your skin.
- Freshly prepared lime water should be used for performing lime water test

Note for the Teacher

- Students may be given three or four samples of salts, of which one of the salts is washing soda or baking soda. The remaining samples of salts should not be carbonates, hydrogencarbonates, sulphites or hydrogensulphites. These salts liberate either CO_2 or SO_2. Sulphur dioxide also turn lime water milky.
- *Preparation of lime water*: Shake 5 g calcium oxide CaO, with 100 mL water. Allow it to stand for about 24 hours. Decant the supernatant liquid and use it for the tests. It is suggested to always use freshly prepared limewater.

QUESTIONS

- Explain why should dil. hydrochloric acid be added dropwise to the salt solution while performing lime water test?
- What will happen if crystalline washing soda is left open in air?
- CO_2 and SO_2 both turn lime water milky and their aqueous solutions turn blue litmus paper red. How can you then distinguish between these?
- Why should carbon dioxide be soluble in aqueous solution of potassium carbonate?

Experiment 10

AIM

To show that crystals of copper sulphate contain water of crystallisation.

THEORY

Blue crystals of copper sulphate contains water of crystallisation. These crystals dehydrate on heating to lose water of crystallisation at a particular temperature and also change their colour.

$$\underset{\text{Blue}}{CuSO_4.5H_2O} \xrightarrow{100^\circ C} \underset{\text{Bluish white}}{CuSO_4.H_2O} \xrightarrow{250^\circ C} \underset{\text{White}}{CuSO_4}$$

If the dehydrated copper sulphate solid material is allowed to cool in air, then it regains blue colour after gaining water molecules from the atmosphere.

$$\underset{\text{White}}{CuSO_4} + \underset{\substack{\text{From}\\\text{atmosphere}}}{5H_2O} \longrightarrow \underset{\text{Blue}}{CuSO_4.5H_2O}$$

MATERIALS REQUIRED

Spatula, watch glass, copper sulphate, and a burner.

PROCEDURE

1. Take some crystals of copper sulphate ($CuSO_4.5H_2O$) in a spatula.

2. Heat these crystals on a burner by keeping the spatula directly over the flame of the burner.
3. Note the change in colour of the copper sulphate crystals during the heating. Does it show a bluish white colour? If yes, keep on heating the crystals for some more time. After some time as temperature reaches around 250 °C, the copper sulphate crystals starts appearing white.
4. Stop heating when it becomes complete white .
5. Transfer the content (white powder) to a watch glass.
6. Keep the watch glass in open atmosphere for some time and allow it to cool. Do you find a change in the colour of copper sulphate crystals.

OBSERVATIONS

(i) On heating, the blue colour of copper sulphate crystals first changes into __________ and then to ________.

(ii) On cooling, the colour of copper sulphate again turns ________ .

RESULTS AND DISCUSSION

Infer from your observations that the hydrated sample of copper sulphate loses water of crystallisation on heating and becomes dehydrated whose colour is white. This dehydrated copper sulphate regains water of crystallisation on cooling and it again becomes blue. Thus the hydration and dehydration is the precise cause of colour change.

PRECAUTION

- Hold the spatula containing copper sulphate crystals very carefully. Do not bring your face near to hot spatula, as it may hurt.

QUESTIONS

- How can you test that a given sample contains water or not?
- What shall be the total action of heat on copper sulphate?
- It is regarded that each molecule of copper suphate crystals at room temperature contains five water molecules as water of crystallisation. Do you see any difference in them? (Hint: Look at the dehydration reaction of copper sulphate)