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.

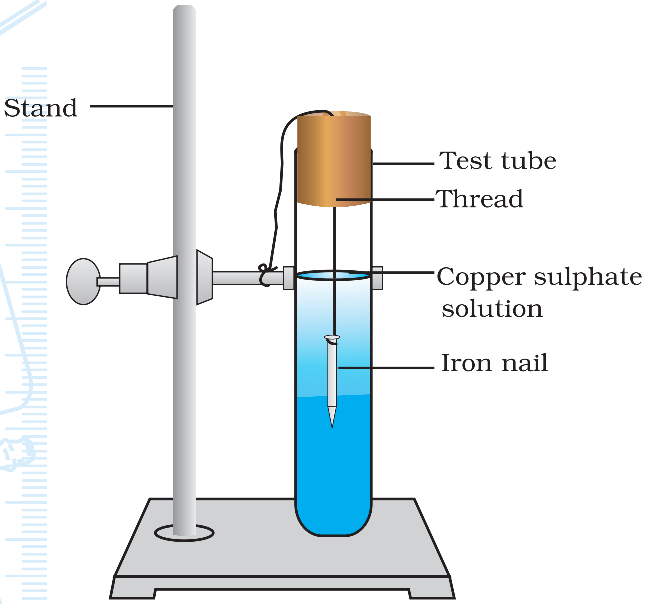
$$Fe\,(s) + Cu^{2+}\,(aq) \longrightarrow Fe^{2+}\,(aq) + Cu\,(s).$$

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,

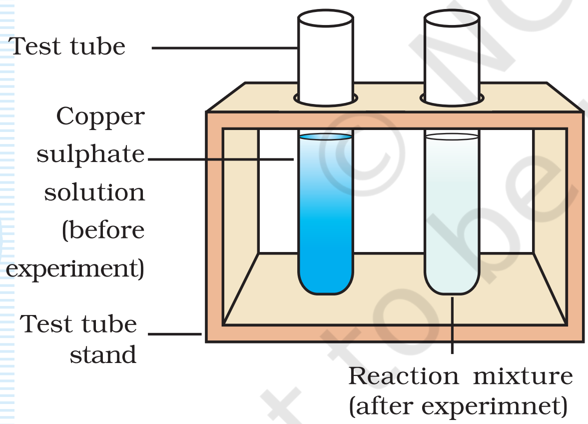
PROCEDURE

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

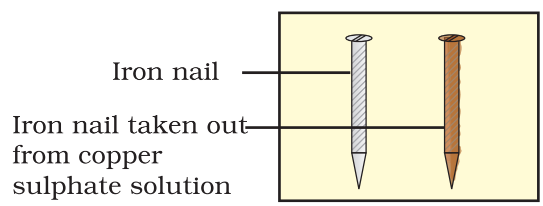


(a)

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?

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

(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.

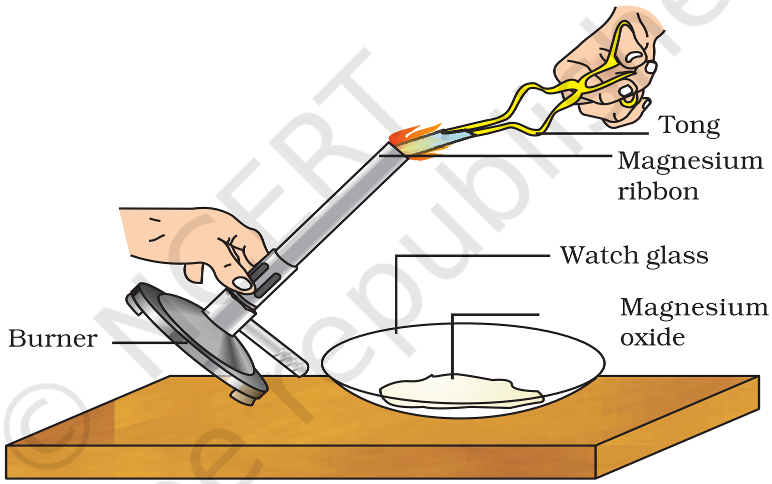


Fig. 1.2 : *Burning of magnesium ribbon and collection of magnesium oxide on a watch glass*

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.

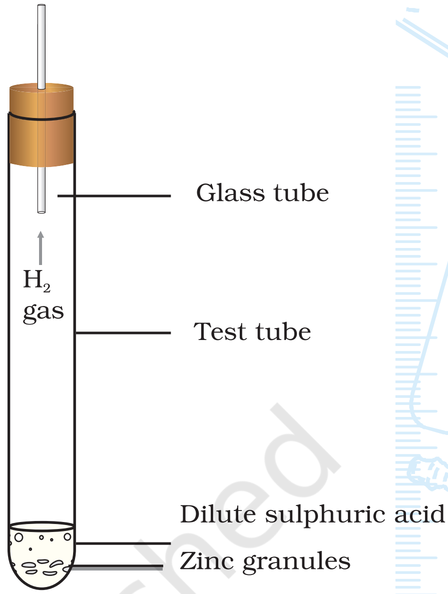


Fig. 2.1 : *Reaction of zinc granules with dil. sulphuric acid*

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.

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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.

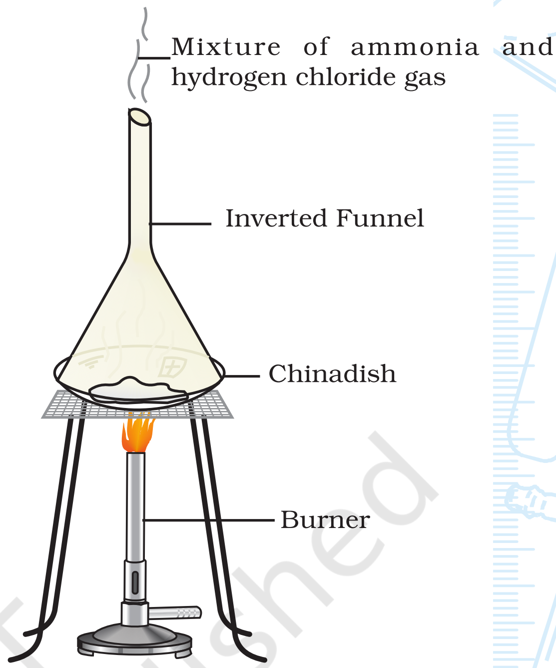


Fig. 2.2 : *Heating of ammonium chloride in an open container*

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 ºC to 110 ºC), 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	Inital temperature of the reaction mixture	Final temperature of the reaction mixture	Change in temperature
		¯N= (°C)	¯O= (°C)	¯0=~ ¯N= (°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) + 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.

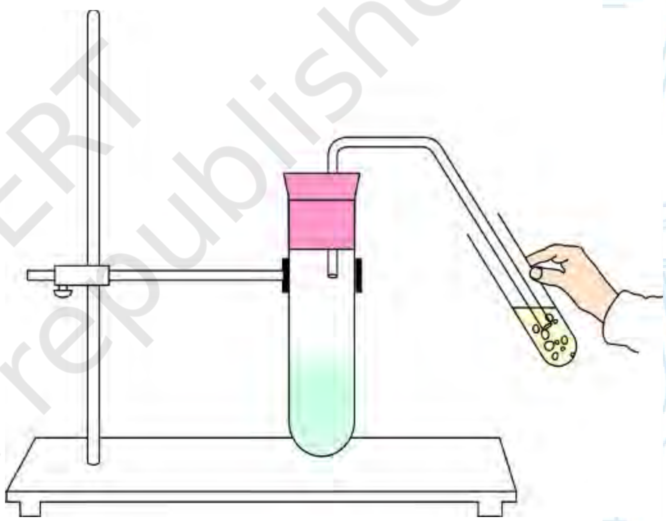


***Fig. 4.1 :** Passing of liberated gas through the freshly prepared lime water*

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 clectric 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.

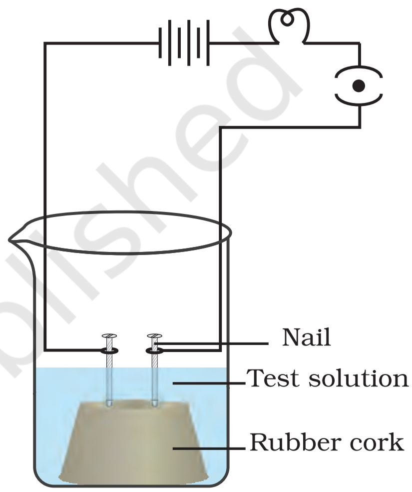


Fig. 6.1 : *Continuity test through an electrolyte*

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.
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.

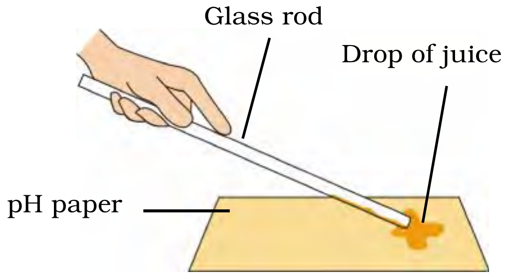


Fig. 7.1 : *(a) Testing the pH of a sample by putting a drop on pH paper by glass rod*

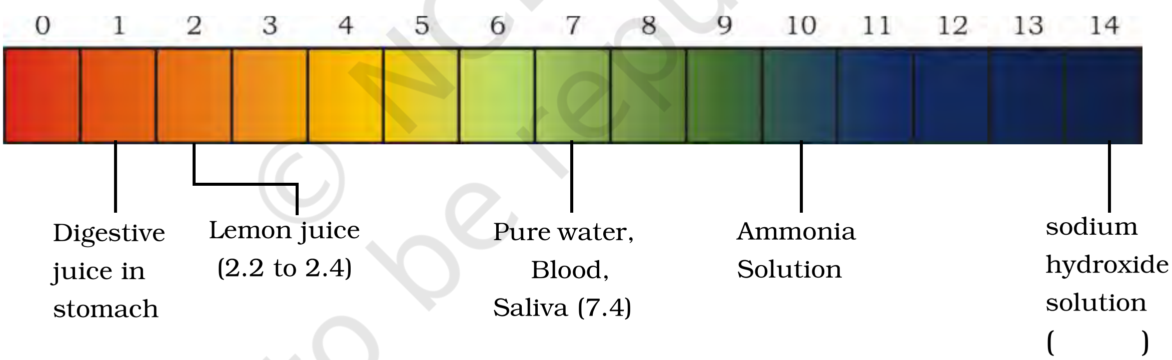


Fig. 7.1 : *(b) Colour of universal indicator at different pH*

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.

17-04-2018

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.

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

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

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

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

$$\underset{\text{Lime water}}{Ca(OH)_2(aq)} + CO_2(g) \longrightarrow \underset{\text{Milky}}{CaCO_3(s)} + H_2O(l)$$

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.

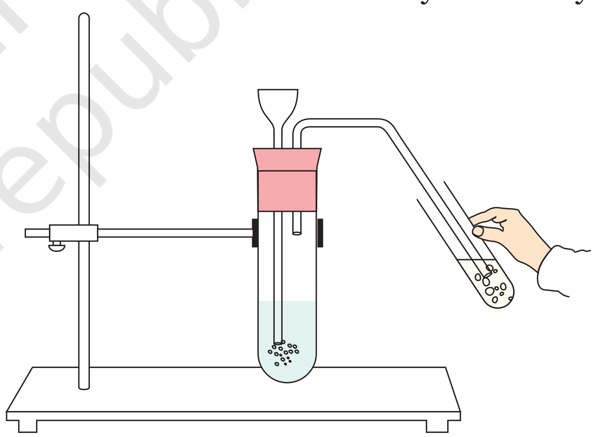


Fig. 9.1 : *Carbon dioxide gas formed by the reaction of dil/hydrochloric acid on washing soda or baking soda is being passed through lime water*

OBSERVATIONS

Sl. No.	Sample	Colour	Solubility in water (Soluble/ insoulble)	Action on red litmus paper (Changes to blue or not)	Action of dil. HCl acid (Effervescence observed or not?)	Lime water 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.

17-04-2018

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.

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

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)

Experiment 11

Aim

To study the interaction of metals such as magnesium, zinc, iron, tin, lead, copper, aluminum (any four) with their salt solutions and to arrange them according to their reactivity.

Theory

Different metals have different reactivities towards chemical reagents. Some metals are more reactive than others. The metals, which can lose electrons more readily to form positive ions are more reactive. Displacement reactions can be used to find out the relative reactivities of metals. A more reactive metal displaces a less reactive metal from its salt solution. For example, if a piece of zinc metal is dipped in a solution of copper sulphate, zinc will displace copper from copper sulphate. The blue colour of copper sulphate solution will gradually fade and finally, a colourless solution of zinc sulphate will be obtained.

$$\mathrm{Zn\ (s)} + \underset{\text{Blue}}{\mathrm{CuSO_4\ (aq)}} \longrightarrow \underset{\text{Colourless}}{\mathrm{ZnSO_4\ (aq)}} + \underset{\text{Reddish brown}}{\mathrm{Cu\ (s)}}$$

Materials Required

Pieces of metals such as zinc, copper, iron, and lead or other suitable metals (at least four strips of each metal), solutions like zinc sulphate; copper (II) sulphate; iron (II) sulphate; and lead nitrate, distilled water, four beakers (100 mL), four test tubes, a measuring cylinder (50 mL), sa test tube stand, and a piece of sand paper.

PROCEDURE

1. Take zinc, copper, iron, and lead metal pieces and clean their surfaces with a sand paper.
2. Prepare 50 mL solutions of 5% concentration (by volume) of zinc sulphate, copper (II) sulphate, iron (II) sulphate and lead nitrate in distilled in four different beakers. Label these beakers as W, X, Y, and Z. Note that these are the salt solutions of the four metals taken for studying the interaction.
3. Take 10 mL of each solution in four different test tubes and label them as tubes A, B, C, and D.
4. Put zinc metal strip in all the four test tubes, that is in tubes A, B, C, and D and observe the change that follows.
5. Repeat the above experiment with other metal strips by dipping them in fresh salt solutions of metals and observe for displacement reactions.

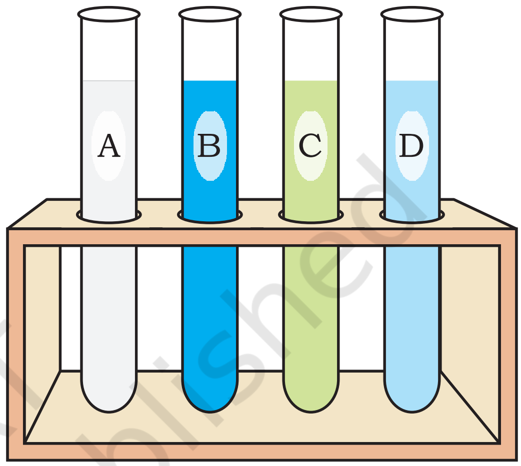


Fig 11.1 : *Zinc metal dipped in Zinc sulphate (A), copper sulphate (B), iron sulphate (C) and lead nitrate (D) solutions*

OBSERVATIONS

Sl.No.	Metal	Metal displacement and colour change of solution (Solution to which metal is added)			
		Zinc sulphate solution, A	Copper (II) sulphate solution, B	Iron (II) sulphate solution, C	Lead nitrate solution, D
1.	Zinc				
2.	Copper				
3.	Iron				
4.	Lead				

Results and Discussion

Infer from the observations and arrange the metals in the order of their decreasing reactivities.

Precautions

- Clean the metals by rubbing them with a piece of sand paper before dipping them in the salt solutions.
- Wash the test tubes after every set of observations of interaction of a particular metal with the four salt solutions.

Note for the Teacher

- One or two drops of conc. H_2SO_4 may be added during the preparation of salt solutions to avoid the hydrolysis of sulphate salts.
- For obtaining granuels of different metals, sheets of metals may be cut into smaller pieces.

Questions

- In the following reaction, A and B are metals. BX is a salt of metal B.

$$A + BX \longrightarrow AX + B$$

  Which one of the two metals is more reactive? Give reason.
- Name any two metals that are more reactive than iron?
- Why did the colour of copper (II) sulphate solution, change, when zinc metal was dipped in it?
- What is your observation when copper is added in iron (II) sulphate solution?
- Which is the most and the least reactive metal in the above experiment?
- Why can we safely preserve iron (II) sulphate in a copper vessel whereas the same can't be safely preserved in zinc vessel?

Experiment 12

AIM

To study the reaction of metals with water under different temperature conditions.

THEORY

Some metals react with water and produce metal hydroxides or oxides and liberate hydrogen gas. Metals like potassium and sodium react violently with cold water. The reaction is so violent that the liberated hydrogen immediately ignites. But all metals donot react with water.

$$2K\ (s) + 2H_2O\ (l) \longrightarrow 2\ KOH\ (aq) + H_2\ (g)$$
$$2Na\ (s) + 2H_2O\ (l) \longrightarrow 2\ NaOH\ (aq) + H_2\ (g)$$

However calcium reacts less violently with cold water. That is

$$Ca\ (s) + 2H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq) + H_2\ (g)$$

Magnesium does not react with cold water. It reacts with hot water to form magnesium hydroxide and hydrogen. Metals like aluminium, zinc and iron donot react either with cold or hot water, but they react with steam to form a metal oxide and hydrogen.

$$2Al\ (s) + 3H_2O\ (g) \longrightarrow Al_2O_3\ (s) + 3H_2\ (g)$$
$$3Fe\ (s) + 4H_2O\ (g) \longrightarrow Fe_3O_4\ (s) + 4H_2\ (g).$$

Metals like lead, copper, silver and gold donot react even with steam.

In this experiment we shall study reactions of some metals with water under different conditions.

Materials Required

Small pieces of same sizes of seven samples of metallic substances (sodium, magnesium, zinc, lead, iron, aluminium, and copper), distilled water, fourteen test tubes, two test tube stands, burner, a beaker (250 mL), a gas jar, a boiling tube, a laboratory stand, a delivery tube, a trough, a single bored cork, glass wool, and a piece of sand paper.

Procedure

1. Take seven test tubes each half filled with cold water. Place them in a test tube stand.
2. Put small samples of clean metallic substances in these seven test tubes.
3. Observe the test tubes to identify the metals that react with cold water. How fast these metals react with cold water? Do all metals react at the same rate? The order of reactivity in different test tubes can be compared by carefully observing the rate of formation of bubbles of liberated hydrogen gas in the test tubes.
4. Boil about 100 mL water in a beaker.

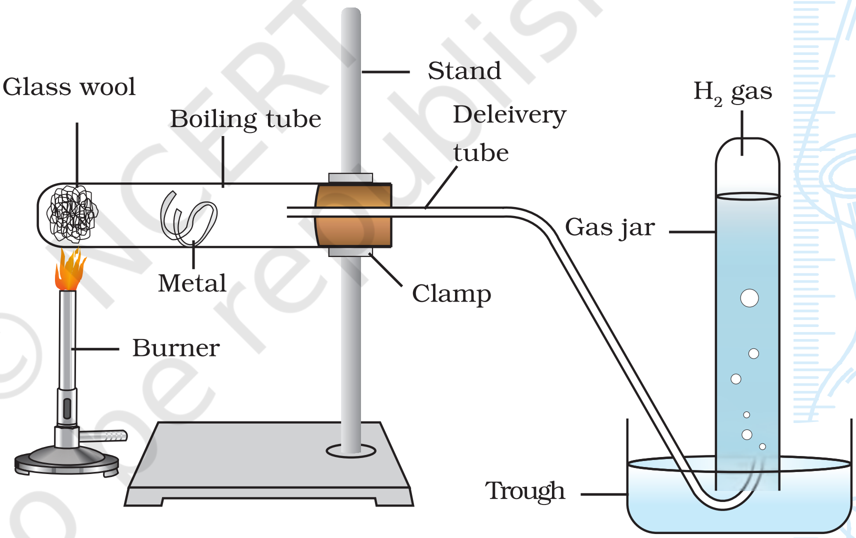


Fig. 12.1 : *Action of steam on a metal*

5. Take out the metallic pieces from the test tubes that did not react with cold water in steps 2 and 3.
6. Put these metallic pieces in test tubes half filled with hot water.
7. Observe the test tubes to identify the metals that react with hot water. Also observe that which metal reacts fast with hot water? Also compare their order of reactivity by observing the bubbles of liberated hydrogen gas in the test tubes.
8. Did you find any metallic substance reacting neither with cold water nor hot water? These may or may not react with steam. Take such samples out from the test tubes.

9. To see the reaction of metallic substances (as identified in step 8) with steam, arrange the apparatus as shown in Fig 12.1, and observe their reaction with steam.
10. Arrange the metals in the decreasing order of reactivity with water under different conditions.

Observations

Record your observation as vigrous, slow or no reaction with cold water or with hot water or with steam in the following table.

Sl.No.	Metals	Reaction conditions		
		cold water	hot water	steam
1.	Sodium			
2.	Magnesium			
3.	Zinc			
4.	Lead			
5.	Iron			
6.	Aluminium			
7.	Copper			

Results and Discussion

List the metals that react with cold water, with hot water, with steam separately. Also arrange the metals in the order of decreasing reactivity in each list. Also list the metals that do not react with water.

Precautions

- Always handle sodium metal carefully as it even react with moisture of skin.
- All metals except sodium should be cleaned by rubbing with a sand paper.
- The exposed surface area of all samples of metallic substances under observations should be same.

QUESTIONS

- Which gas is produced when an active metal reacts with water?
- Did any metal produce fire in water?
- Which metals did not react with cold water at all?
- Why should we use glass-wool soaked in water for action of sodium on cold water?
- Which metal did not react with water even in the form of steam?

Experiment 13

AIM

To study reaction of metals with dilute acids.

THEORY

Many metals react with dil. hydrochloric and dil. sulphuric acid to form a salt. In this reaction hydrogen gas is evolved. The metal replaces the hydrogen atoms from the acid to form a salt. That is

$$\text{Metal + dilute acid} \longrightarrow \text{Metal salt + Hydrogen}$$

$$Mg\ (s) + 2HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$$

$$2Al\ (s) + 6HCl\ (aq) \longrightarrow 2AlCl_3\ (aq) + 3H_2\ (g)$$

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

$$2Fe\ (s) + 6HCl\ (aq) \longrightarrow 2FeCl_3\ (aq) + 3H_2\ (g)$$

However all metals do not react with dil. HCl or dil. H_2SO_4. Hydrogen gas is rarely evolved when metal reacts with nitric acid. It is because HNO_3 is a strong oxidising agent. It oxidises the hydrogen produced to water and reduces to any of oxides of nitrogen (such as N_2O, NO_2, NO). However, magnesium and manganese metals react with dil. HNO_3 to liberate hydrogen gas.

MATERIALS REQUIRED

Five samples of metallic substances (such as magnesium, aluminium, zinc, iron, and copper), dil. hydrochloric acid, dil. sulphuric acid, dil. nitric acid, five test tubes, a test tube stand, a measuring cylinder (50 mL), and a piece of sand paper.

PROCEDURE

1. Take five test tubes in the test tube stand and label them as A, B, C, D, and E.
2. Take small pieces of sample metallic substances (magnesium, zinc, aluminium, iron, and copper metals). Clean their surfaces by rubbing with a sand paper.
3. Place these metals in test tubes A, B, C, D and E respectively.
4. Add about 10 mL dil. hydrochloric acid to each of these test tubes.
5. Observe carefully the rate of formation of bubbles in the test tubes. These bubbles are of the hydrogen gas, liberated in the reaction.
6. Arrange the metals in the decreasing order of reactivity with dil. hydrochloric acid.
7. Take out the metallic samples from the test tubes. Wash the test tubes with water. Put them in a test tube stand.
8. Repeat the experiment (steps 1 to 6) with dil. sulphuric acid and dil. nitric acid.
9. Record your observations as vigrous, slow or no reaction in the following table.

OBSERVATIONS

Sl.No.	Metal	Intensity of reaction with		
		dil. HCl	dil. H_2SO_4	dil. HNO_3
A	Magnesium			
B	Zinc			
C	Aluminium			
D	Iron			
E	Copper			

RESULTS AND DISCUSSION

List the metals that react with dil. HCl, dil. H_2SO_4 and dil. HNO_3 separately. Also arrange the metals A, B, C, D, and E in the order of decreasing reactivity in each case.

Precautions

- The exposed surface area of the metals should be approximately same.
- Clean the metal surface with sand paper, specially for Mg and Al.

Note for the Teacher

- The intensity of reaction depends on concentration of acid used and surface area of the metal exposed besides the nature of metal and the experimental temperature.

Applications

Such studies can help us to construct metal activity series.

Questions

- What was your observation when zinc was dipped in dil. hydrochloric acid?
- Which metals reacted vigorously with dil. hydrochloric acid?
- Which metal did not react with dil. hydrochloric acid?
- Metal reacts with dil. hydrochloric acid to give metal salt and hydrogen gas. Can you suggest any test to verify that evolved gas, if any, is hydrogen?
- Can we use dil. nitric acid in place of dil. HCl in this experiment?
- Why would iron dust reacts vigorously as compared to iron filings with dil. HCl?

17-04-2018

Experiment 14

AIM

To prepare sulphur dioxide gas and study its physical and chemical properties.

THEORY

Sulphur dioxide is prepared by the action of hot concentrated sulphuric acid on copper turnings.

$$Cu\ (s) + 2H_2SO_4\ (l) \xrightarrow{Heat} CuSO_4\ (aq) + SO_2\ (g) + 2H_2O$$

Sulphur dioxide is acidic in nature. It decolourises acidified potassium permanganate ($KMnO_4$) solution. Acidified potassium dichromate solution ($K_2Cr_2O_7$) is also turned green by SO_2. The reactions with $KMnO_4$ and $K_2Cr_2O_7$ are due to the reducing property of SO_2 and oxidising nature of acidified $KMnO_4$ and $K_2Cr_2O_7$.

Or $\quad 2KMnO_4 + 5SO_2 + 2H_2O \longrightarrow K_2SO_4 + 2MnSO_4 + 2H_2SO_4$.

Similarly,

$$K_2Cr_2O_7 + 4\,H_2SO_4 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + 3[O]$$
$$[SO_2 + H_2O + [O] \longrightarrow H_2SO_4] \times 3$$

Or $\quad K_2Cr_2O_7 + H_2SO_4 + 3SO_2 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O$.

MATERIALS REQUIRED

Copper turnings, conc. sulphuric acid, dil. sulphuric acid, potassium permanganate solution, potassium dichromate solution, red and blue litmus papers, a round bottom flask, a thistle funnel, a delivery tube, a double bored cork, a piece of card board as a lid, a laboratory stand, a burner, a wire gauze, a tripod stand, two test tubes, a trough, a measuring cylinder (50 mL), and a gas jar.

PROCEDURE

1. Place few pieces of copper turnings (about 5 g) in a round bottom flask and arrange the apparatus as shown in Fig. 14.1.
2. Add 15 - 20 mL of conc. sulphuric acid to it through a thistle funnel.
3. Place the cork in its position on the flask again and heat the contents gently. The gas formation starts after sometime.
4. Collect the gas in the gas jar and study its properties as per the steps given in the observation table.

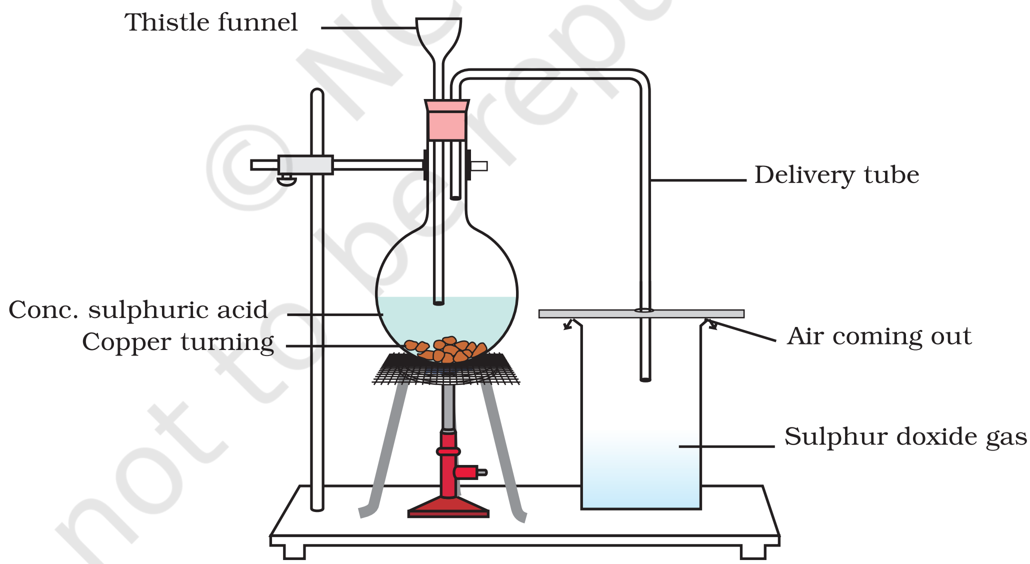


Fig. 14.1 : *Preparation of sulphur dioxide gas*

OBSERVATIONS

Sl.No.	Test	Experiment	Observation	Inference
1.	**Physical Properties**			
(a)	Colour	Look at the sulphur dioxide filled gas jar.		
(b)	Solublity in water	Take a gas jar filled with sulphur dioxide with its mouth closed with a lid. Invert it in water contained in a trough. Remove the lid carefully.		
2.	**Chemical Properties**			
(a)	Acidic or basic nature	Insert damp or wet litmus paper in the jar filled with sulphur dioxide.		
(b) (i)	Reaction with potassium permanganate ($KMnO_4$) solution	Take about 2 mL of potassium permaganate solution in a test tube, add about 1 mL of dilute H_2SO_4 and pass sulphur dioxide gas in this solution.		
(ii)	potassium dichromate ($K_2Cr_2O_7$) solution	Pass sulphur dioxide through another test tube containing acidified potassium dichromate solution		

RESULTS AND DISCUSSION

Infer the observations and note your inferences in observation table. On the basis of observations mention the physical and chemical properties of the liberated sulphur dioxide gas. Solubility of SO_2 in water is a chemical property on account of the following reaction.

$$SO_2\ (g) + H_2O\ (l) \longrightarrow H_2SO_3\ (l)$$

SO_2 is both oxidising and reducing in its behaviour as, it can take as well as supply oxygen.

PRECAUTIONS

- Keep the apparatus for preparation of gas airtight.
- Concentrated sulphuric acid should be handled carefully. It should not touch your skin.
- Avoid adding large quantity of acid at a time, otherwise a vigorous reaction may occur. Care should be taken while handling hydrochloric acid. It should not touch the skin.
- Do not inhale sulphur dioxide.

Note for the Teacher

- Sulphur dioxide can also be prepared by
  Na_2SO_3 (aq)+ dil. H_2SO_4 (aq) $\longrightarrow$ Na_2SO_4 (aq) + H_2O (l) + SO_2 (g)

Questions

- What type of reaction (oxidation or reduction) does sulphuric acid undergo during the laboratory preparation of sulphur dioxide?
- What happens when sulphur dioxide is passed through acidified potassium permanganate solution?
- How will you prove that sulphur dioxide is acidic in nature?
- Why is sulphur dioxide collected by upward displacement of air?
- What are the different roles of H_2SO_4 in chemical reactions? Justify your answer with an example of each?
- Identify the role (oxidant or reductant) of each gas in the following reaction:

  $2H_2S$ (g) + SO_2 (g) $\longrightarrow$ 3S (s) + $2H_2O$ (l)
- Compare the reactions:

  (i) Cu(s) + $2H_2SO_4$ (l) $\longrightarrow$ $CuSO_4$ (aq) + SO_2 (g) + $2H_2O$ (l)

  (ii) 4Zn (s) + $5H_2SO_4$ (l) $\longrightarrow$ $4ZnSO_4$ (aq) + H_2S + $4H_2O$ (l)

  What conclusion do you draw about the two metals here?

Experiment 15

AIM

To prepare carbon dioxide gas and study its physical and chemical properties.

THEORY

In a laboratory, carbon dioxide gas (CO_2) may be prepared by the action of dilute acids on calcium carbonate. Calcium carbonate is usually taken in the form of marble chips.

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O + CO_2\,(g)$$

Carbon dioxide gas is acidic in nature. It is also an oxidising agent. It turns a red litmus paper blue and turns lime water [$Ca(OH)_2$ (aq)] milky.

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

On passing excess of CO_2 through lime water, 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

Marble chips (10 g), dil. hydrochloric acid, red and blue litmus paper strips, freshly prepared lime water, a small piece of magnesium ribbon, Woulfe's bottle or a round bottom flask, a gas jar, a measuring cylinder (50 mL), a thistle funnel, a delivery tube, two single bored corks, a trough, a candle, and a piece of cardboard,

PROCEDURE

1. Take about 10 g of small pieces of marble chips in a woulfe's bottle and set up the apparatus as shown in Fig. 15.1.
2. Add, dropwise, 10 mL of dil. hydrochloric acid to the woulfe's bottle through the thistle funnel.
3. Do you see any reaction taking place in the woulfe's bottle? Do you see any gas formation?
4. Collect the liberated gas in gas jar to perform the colour, odour, solubility in water, combustibility and acidic tests. Record your observations in the observation table.
5. Pass the liberated gas through freshly prepared lime water in a test tube. Do you see any gas bubbles in the lime water? Does the colour of lime water turns milky?

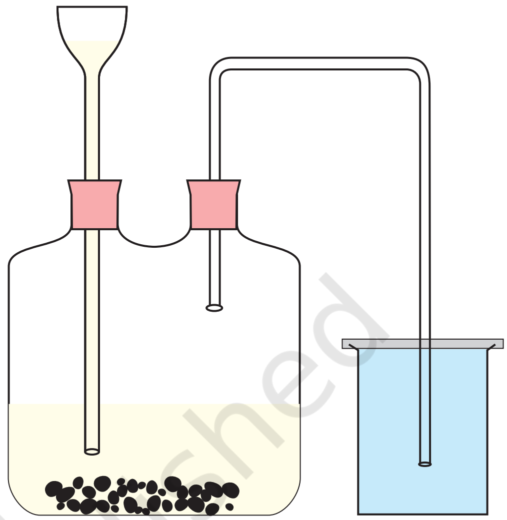


Fig. 15.1 : *Preparation of CO_2 gas in Woulfe's bottle*

OBSERVATIONS

Sl.No.	Test	Experiment	Observation	Inference
1.	Colour	Look at a gas jar filled with carbon dioxide gas.		
2.	Odour	With the help of your hand, fan the gas gently towards your nose and smell.		
3.	Solubility in water	Introduce a gas filled jar over a trough of water.		
4.	Combustibility	Introduce a lighted candle in the gas filled jar		
5.	Acidic nature	Insert a moist or wet red litmus paper in the gas filled jar.		
6.	Reaction with lime water	Take abour 5 mL of freshly prepared lime water in a test tube and pass the liberated gas from the Woulfe's bottle through lime water using a delivery tube. Excessively pass the liberated gas through the lime water.		

Results and Discussion

Infer the observations and note your inferences in observation table. On the basis of observations mention the properties of liberated gas.

Precautions

- The apparatus should be airtight.
- The lower end of the thistle funnel should be dipped in the acid taken in Woulfe's bottle otherwise carbon dioxide will escape through the thistle funnel.
- Avoid adding large quantity of acid at a time, otherwise a vigorous reaction may occur. Care should be taken while handling hydrochloric acid. It should not touch the skin.
- While collecting the gas in the jar, a piece of cardboard should be placed over the mouth of the gas jar.

Note for the Teacher

- *Preparation of lime water:* Shake about 5 g of calcium oxide, CaO, with 100 mL water. Allow it to stand for 24 hours. Decant the supernatant liquid and use it for the tests. Always use freshly prepared limewater.
- It is not necessary to set up the Woulfe's apparatus for preparing the carbon dioxide gas. This set up is needed for collecting the gas in a jar. The reaction may also be carried out in a test tube.
- In this experiment dil. H_2SO_4 may also be used in place of dil. HCl. It forms $CaSO_4$ on reacting with marble chips. In this reaction, initially CO_2 is liberated but with time a layer of $CaSO_4$ deposits on $CaCO_3$ and this will stop the reaction and no CO_2 will form.

Questions

- Why is carbon dioxide collected by upward displacement of air?
- Sulphuric acid is not used for preparing the carbon dioxide gas in the laboratory. Why?
- What happens to the burning magnesium strip when introduced in the jar filled with carbon dioxide?
- Why is lighted candle put off when inserted in the jar of carbon dioxide?

- What is the chemical name of the compound formed when carbon dioxide gas is passed through limewater?
- What is the effect of carbon dioxide gas on moist blue litmus paper?
- How can you prove that we exhale (breathe out) carbon dioxide?
- Can the reaction

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

be called acid-base reaction?
- What happens when excess of CO_2 is passed through lime water? Write the chemical equation for the reaction involved.
- Why can't you introduce the dil. HCl in Woulfe's bottle at a faster rate?

Experiment 16

AIM

To study the process of electrolysis.

THEORY

An electrolyte is a compound that, in solution or in the molten state, dissociate into ions and conduct an electric current. On passing an electric current from an external source, these ions migrate towards the oppositely charged electrodes. Positive ions migrate to the negative electrode and negative ions to the positive electrode and they discharge at respective electrodes. This phenomenon is called electrolysis and the container in which electrolysis occur is called electrolytic cell.

In this experiment, we take an aqueous solution of copper sulphate as an electrolyte; a strip of pure copper metal as cathode; and a strip of impure copper as an anode. When an electric current is passed through the aqueous solution of copper sulphate, copper ions (Cu^{2+}) migrate towards the pure copper strip cathode, get discharge and deposit over it. Sulphate (SO_4^{2-}) ions move towards the impure copper anode, dissolve its copper into the solution as copper ions and thus keep their concentration constant in the solution.

Materials Required

An impure thick copper metal strip, a thin pure copper metal strip, copper sulphate crystals, distilled water, dil. sulphuric acid, two dry cells with a cell holder (or a battery of 3 V or a battery eliminator), a plug key, a beaker (250 mL), a measuring cylinder (50 mL), physical balance with weight box, a small cardboard with two holes, connecting wires, and a piece of sand paper.

Procedure

1. Take a beaker to use it as an electrolyic cell.
2. Dissolve about 3 g of copper sulphate crystals in 100 mL distilled water and pour this solution in the electrolytic cell. Add to it about 1 mL of dil. sulphuric acid to make the solution acidified.
3. Clean the ends of the connecting wires using a sand paper.
4. Connect two wires with the two copper strips to be considered as positive electrode (impure copper strip) and negative electrode (pure copper strip). Pass these strips through the two holes of the cardboard(Fig. 16.1).
5. Connect the two copper strips with a combination of two dry cells (in a cell holder) through a plug key as shown in Fig. 16.1.

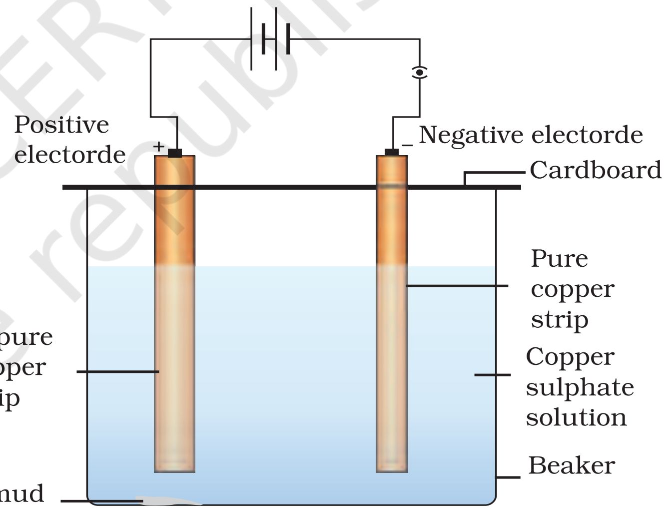


Fig. 16.1 : *An electrolysis process*

   [In place of cells a battery of 3 V or a battery eliminator may also be used.] Do not plug the key.
6. Immerse the two copper strips into the solution and cover the beaker with the cardboard.
7. Insert the key into the plug to allow the electric current to pass through the electrolytic solution.
8. Observe the electrolytic solution after sometime. Do you find any change in the thickness(es) of the two copper strips? Note your observations. If not, allow the current to flow through the electrolyte

for more time till you observe a change in the thickness of the two copper strips.

Observations

The thickeness of impure copper strip (positive electrode) decreases whereas the thickness of pure copper strip (negative electrode) increases on passing an electric current through the acidified copper sulphate for some time.

Results and Discussion

In this electrolysis process, copper ions are released from the impure copper strip (positive electrode). These ions move through the solution in the electrolytic cell towards the pure copper strip (negative electrode). Here they get discharged and deposited. That is how the thickness of impure copper strip keeps on decreasing and the thickness of pure copper strip keeps on increasing on passing the current through the electrolyte.

Discuss the role and movement of sulphate ions in the solution.

Precautions

- Copper sulphate is poisnous in nature. Handle it carefully,
- Thin strip of copper (negative electrode) should be of pure metal (why?).
- Never keep the two electrodes close to each other in the electrolytic cell. Similarly, the two electrodes should not touch the sides of the cell.

Note for the Teacher

- Instead of impure and pure copper strips, loops of copper wire can also used as two electrodes. The purpose of electrolysis is electroplating. In fact any conducting material can be used as cathode over which the copper can be deposited.
- Oxidation is also described as loss of electrons and takes place at anode for example in the reaction under discussion.

  $Cu\ (s) + SO_4^{2-}\ (aq) \longrightarrow CuSO_4\ (aq) + 2e^-$
- Reduction is gain of electrons and takes place at cathode and here it occurs in the following manner.

  $Cu^{2+} + 2e^- \longrightarrow Cu(s)$
- A little quantity of dil. sulphuric acid is added to acidify the copper sulphate solution to avoid its hydrolysis, failing which the precipitation of copper hydroxide will take place.

QUESTIONS

- What is electrolysis?
- How will you come to know that copper is deposited on cathode at the end of the experiment?
- What will happen to the impurities present in the impure copper strip (positive electrode)?
- How gold plated or silver plated articles are prepared?
- These days aluminium vessels by the name of 'anodised aluminium' are available in the market. How are these prepared?
- What is the role of electricity in the process of electrolysis.

Experiment 17

AIM

To study physical and chemical properties of acetic acid (ethanoic acid).

THEORY

Ethanoic acid (CH_3COOH) is an organic acid containing (-COOH) functional group. It has an odour of vinegar. It turns blue litmus paper red and reacts with

(a) sodium hydrogencarbonate and sodium carbonate to evolve carbon dioxide gas.

$$CH_3COOH\ (l) + NaHCO_3\ (s) \longrightarrow CH_3COONa\ (aq) + CO_2\ (g) + H_2O\ (l)$$

$$2CH_3COOH\ (l) + Na_2CO_3\ (s) \longrightarrow 2CH_3COONa\ (aq) + CO_2\ (g) + H_2O\ (l)$$

(b) sodium hydroxide to produce sodium ethanoate and water.

$$CH_3COOH + NaOH\ (aq) \longrightarrow CH_3COONa\ (aq) + H_2O.$$

MATERIALS REQUIRED

Sodium hydrogencarbonate, sodium carbonate, sodium hydroxide, phenolphthalein solution, conc. sulphuric acid, 5% ethanoic acid, blue litmus paper strips, two beakers (100 ml), four test tubes, measuring cylinder (10 mL), tripod stand, a burner, and wire gauge.

PROCEDURE

1. Study the physical and chemical properties of ethanoic acid according to the following table and record your observations.

17-04-2018

OBSERVATIONS

Sl.No.	Test	Experiment	Observations	Inference
1.	**Physical Property**			
	Smell	Smell the sample of ethanoic acid		
(b)	Solubility test	Add 1 mL of the given sample of acid in 2 mL water.		
2.	**Chemical Porperty**			
(a)	Litmus test	Put a drop of ethanoic acid over a (i) blue litmus paper and (ii) red litmus paper		
(b)	Reaction with sodium hydrogen-carbonate	Take 1 mL of the ethanoic acid and add to it a pinch of sodium hydrogen-carbonate.		
(c)	Reaction with sodium carbonate	Take 1 mL of the ethanoic acid and add to it a pinch of sodium carbonate.		
(d)	Reaction with aqueous sodium hydroxide solution.	Take 5 mL of the ethanoic acid and add 2-3 drops of phenolph-thalein solution to it. Then add sodium hydroxide solution to the mixture drop by drop. Shake the mixture gently. Count the number of drops of sodium hydroxide needed for appearance of pink colour in the reaction mixture.		

RESULTS AND DISCUSSION

Infer the physical properties and chemical properties on the basis of observations.

PRECAUTIONS

- Handle ethanoic acid carefully.
- Add only small amount (0.01 g) of $NaHCO_3$ or Na_2CO_3 to ethanoic acid to control the intensity of CO_2 evolved.

QUESTIONS

- Which gas is evolved when ethanoic acid reacts with sodium hydrogencarbonate?
- How will you test that the liberated gas is carbon dioxide?
- How will you show that ethanoic acid is acidic in nature?
- Where do you find the use of ethanoic acid in day-to-day food products?
- What is the common name of ethanoic acid as sold in the market in the form of its dilute solution?
- What type of reaction takes place between ethanoic acid and sodium hydroxide solution?

Experiment 18

AIM

To study esterification reaction between alcohol and carboxylic acid.

THEORY

Carboxylic acid react with alcohols in presence of conc. H_2SO_4 to form esters with a loss of water molecule. For example when ethanoic acid reacts with ethanol, the ethyl ethaonate ester is formed.

$$\underset{\text{Ethanoic acid}}{CH_3COOH} + \underset{\text{Ethanol}}{CH_3CH_2OH} \xrightarrow{\text{Conc. } H_2SO_4} \underset{\text{Ethyl ethnoate}}{CH_3COOCH_2CH_3} + H_2O$$

Ester has a fruity odour which is distinct from those of carboxylic acid and alcohol.

MATERIALS REQUIRED

Ethanoic acid (3 mL), ethanol (3 mL), few drops of conc. H_2SO_4, distilled water, sodium hydrogencarbonate (1 g), thermometer (– 10ºC to 110 ºC), a test tube, a cork, measuring cylinder (10 mL), a beaker (250 mL), burner, tripod stand, and a wire gauge.

PROCEDURE

1. In a clean test tube take 3 mL ethanoic acid.

2. Add about 3 mL ethanol to it. Also add four to five drops of conc. H_2SO_4 to the reaction mixture. Put a cork loosely over the mouth of the test tube.
3. Take about 150 mL water in the beaker. Heat it to about 60 °C.
4. Put the test tube in the warm water. (The reaction mixture would also get heated. This is the warming of reaction mixture on a water bath.)
5. Shake the reaction mixture occassionaly.
6. Pour the reaction mixture into a beaker containing aqueous solution of sodium hydrogencarbonate. This will remove the unreacted ethanoic acid from the reaction mixture. Do you see any effervesence coming out?
7. Fan the liberated vapours of ester formed with your hand gently towards your nose and smell.
8. Feel the difference in the odours of ethanoic acid, ethanol and ester.

Results and Discussion

Comment on the difference in odours of ethanoic acid, ethanol and ester.

Esters are formed when -OH of carboxylic acid are replaced by -OR (here R represent an alkyl group). In this reaction conc. H_2SO_4 is used as a catalyst.

Precautions

- Be careful while using conc. H_2SO_4.
- The organic compounds are extremely volatile and alcohol is combustible so never heat it directly on a flame. Always use water bath for heating the reaction mixture.

Note for the Teacher

- Ethanoic acid is often available as glacial acetic acid, which is 98% pure. The vinegar used at home is only approximately 2% ethanoic acid solution in water.
- The reaction mixture is poured into aqueous solution of sodium hydrogencarbonate to neutralise the unreacted acid. And Ethanol gets diluted in water. This is needed so that the smell of ester is not masked by smell of unreacted acid and alcohol.

QUESTIONS

- What will be the ester formed when propanoic acid reacts with propanol?
- What is the function of conc. sulphuric acid in this experiment?
- Will the ester formed, turn blue litmus to red?
- Name any subsatnce other than conc. H_2SO_4 that can be used as a catalyst.
- Why do you use a water bath in this experiment?

Experiment 19

Aim

To study some oxidation reactions of alcohol.

Theory

An oxidation process involves gain of oxygen by the elements or compounds. Alcohol can be oxidised to form various products under different conditions.

1. Complete combustion: On burning alcohol (ethanol)in an excess supply of oxygen (present in air) with a flame, gives carbon dioxide and water.

   $$CH_3CH_2OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

2. Oxidation using an oxidising agent: alcohol on oxidation with an oxidising agent (such as alkaline potassium permanganate solution) get oxidised to a carboxylic acid.

   $$CH_3CH_2OH \xrightarrow{\text{Heat with alkaline } KMnO_4} CH_3COOH + H_2O.$$

   The completion of reaction is characterised by the decolourisation of potassium permanganate solution.

Materials Required

Ethanol, 1% solution of alkaline $KMnO_4$, $NaHCO_3$, freshly prepared lime water, spirit lamp, two boiling tubes, a measuring cylinder (10 mL), a beaker (250 mL), rubber tubing, jet, a funnel, and filter paper.

PROCEDURE

1. Perform the reactions as mentioned in the Observation Table.

   For warming ethanol in a test tube (step 2), adopt the following:

   (i) Take about 150 mL water in a beaker. Heat it to about 60 °C.

   (ii) Put the test tube in the warm water. (The reaction mixture would also get heated. This is the warming of mixture on a water bath.)

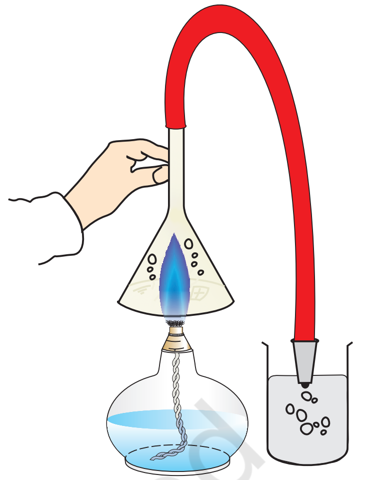


Fig. 19.1 : *Complete combustion of ethanol*

OBSERVATIONS

Sl. No.	Experiment	Observation	Inference
1.	*Complete combustion:* Take ethanol as fuel in a spirit lamp. Burn it as usual and cover the lamp with an inverted funnel (Fig. 19.1). Fix the stem of the inverted funnel with a rubber tubing which is attached to a jet on another end. Pass the evolved vapours through freshly prepared lime water. Do you observe any condensation of water vapours on the inner surface of the inverted funnel?		
2.	*Oxidation using an oxidising agent:* Take 3 mL of ethanol in a boiling tube confined to a water bath. To this add two or three drops of 1% alkaline $KMnO_4$ solution. Warm the tube till the reaction mixture decolourises. Filter it and then add a pinch of sodium hydrogen-carbonate ($NaHCO_3$).		

Results and Discussion

Infer the two oxidation reactions. Oxidation product of alcohol depends on conditions and nature of the process carried out.

Precautions

- Alcohols are extermely volatile and inflammable.
- The alkaline potassium permaganate solution should be very dilute and added dropwise only.
- Do not add potassium permangnate solution in the reaction excessively.

Note for the Teacher

- The alkaline potassium permanganate solution can be prepared by dissolving 1 pellet of KOH and 2-3 small crystals of $KMnO_4$ in 20 mL distilled water.
- If $KMnO_4$ solution is concentrated or is added in excess, the occurrence of reaction may not be observed, as it will not be decolourised.

Questions

- On adding diluted potassium permanganate solution to an alcohol, it decolourises initially and then on its excess addition, the colour of $KMnO_4$ persists. How?
- What are the species oxidised, reduced acting as an oxidising agent or as a reducing agent in the reaction.
- Why should the reaction of alcohol with potassium permaganate be considered as an oxidation reaction for an alcohol?

Experiment 20

AIM

To study saponification reaction for preparation of soap.

THEORY

Oil or fat when treated with sodium hydroxide solution, gets converted into sodium salt of fatty acid (soap) and glycerol. This reaction is known as saponification.

$$
\begin{array}{lll}
CH_2OCOC_{17}H_{35} & & CH_2OH \\
| & & | \\
CH\text{-}OCOC_{17}H_{35} + \underset{\text{Sodium hydroxide}}{3NaOH\ (aq)} & \longrightarrow & CHOH + \underset{\text{Soap}}{3\ C_{17}H_{35}COONa} \\
| & & | \\
CH_2OCOC_{17}H_{35} & & CH_2OH \\
\text{Triester or Triglyceride} & & \text{Glycerol} \\
\text{(oil or fat)} & &
\end{array}
$$

It is an exothermic reaction, that is heat is liberated during saponification.

MATERIALS REQUIRED

Sodium hydroxide, a sample of vegetable oil such as castor oil (25 mL), common salt (about 10 g), distilled water, red and blue litmus paper strips, two beakers (250 mL), two test tubes, a glass rod, a measuring cylinder (50 mL), and a knife.

PROCEDURE

1. Take about 20 mL of castor oil (triglyceride) in a beaker (250 mL).
2. Prepare about 50 mL 20% solution of sodium hydroxide in distilled water and add 30 mL of this solution in 25 mL castor oil.
3. Successively dip the red and blue litmus paper strips into this reaction mixture. Do you find any change in colour of any litmus paper strip. Note and record your observation.
4. Touch the beaker from outside. Is it hot or cold?
5. Add 5 g to 10 g of common salt to this mixture and using a glass rod to stir the mixture continuosly till the soap begins to set.
6. Leave it for a day till the mixture cools and becomes solid.
7. Remove the soap cake and cut it into desired shapes and sizes.

OBSERVATIONS

(i) The colour of red litmus paper (when dipped into mixture) turns __________, while the colour of blue litmus paper becomes _________

(ii) The temperature of reaction mixture on adding sodium hydroxide with oil _____________ (increases/decreases).

RESULTS AND DISCUSSION

On the basis of your observations with litmus paper, ascertain the medium of soap solution is (acidic/basic). Also comment whether the saponification reaction is exothermic or endothermic.

The saponification reaction suggests the formation of glycerol alongwith the soap which is present as a separate product.

Soap is salt of fatty acid and its precipitation is governed just like the precipitation of any other salt.

PRECAUTIONS

- Stir the soap solution carefully so that it does not spill out.

NOTE FOR THE TEACHER

- If castor oil is not available, any other edible oil may be taken.
- For commercial preparation of soap certain additives like colour, perfume, fillers are added. Fillers harden the soap and make the cutting of the soap easy.
- Common salt is used to favour precipitation of soap.

- Notice that in a saponification reaction, glycerol (commonly known gas glycerine) may be obtained as a by product.

QUESTIONS

- Why does a red litmus paper changes its colour when dipped in soap solution? Explain your observation.
- Why is it advised to add common salt while preparing the soap?
- Can we use Na_2CO_3 instead of NaOH? Explain.
- Was heat evolved or absorbed when sodium hydroxide was added to oil?
- What is the chemical reaction involved in the manufacture of soap?
- Can you devise a method to separate glycerine from the reaction mixture?

Experiment 21

Aim

To compare the foaming capacity of different samples of soap.

Theory

Foam is produced when soap is shaken with water. Foaming of a soap is due to the presence of hydrophilic and hydrophobic portions in its molecule $RCOO^-Na^+$. (Refer Chapter 4, *Science Textbook for Class X*, published by the NCERT). Foaming capacity of different soap samples can be compared by measuring the quantity of the foam produced by equal amount of various soap samples.

Materials Required

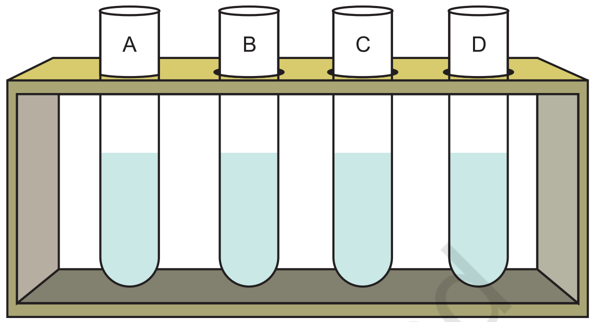
Four different samples of soaps, distilled water, physical balance and weight box, four test tubes, a test tube stand, four beakers (100 mL), a glass rod, burner, tripod stand, wire gauze, a measuring cylinder (50 mL) and a measuring scale.

Procedure

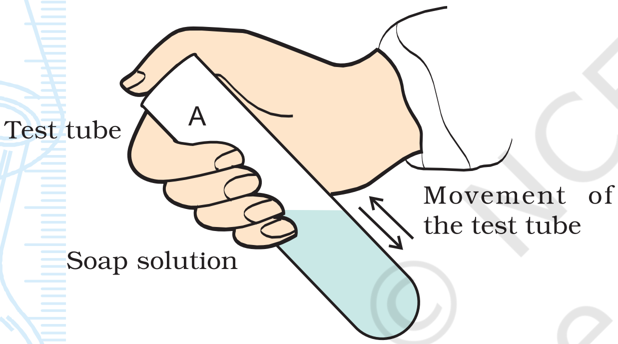
1. Take four 100 mL beakers and label them as beakers A, B, C, and D. Weigh 1 g each of the four different soap samples in a physical balance. Put them into four different beakers.
2. Add 20 mL of distilled water in all the beakers containing the soap

samples. Dissolve the soap in water by stirring the mixture with a glass rod. If a soap sample takes longer time to dissolve in distilled water, uniformly heat the contents of the beaker over a wire gauge.

3. Take a test tube stand and place the four test tubes in it and label them as tubes A, B, C, and D. Pour 1 mL each of the above-prepared soap solutions in corresponding test tubes. [Fig. 21.1(a)].
4. Add 5 mL of distilled water in each test tube.



(a)



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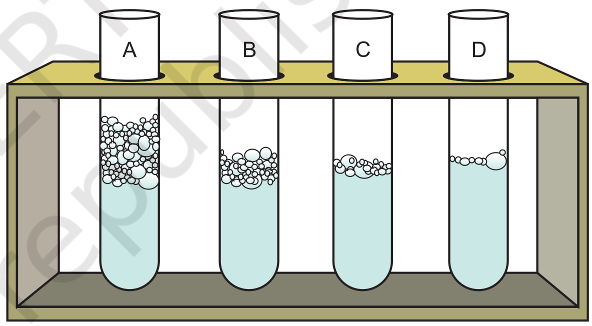

(c)

Fig. 21.1 : *(a) Test tubes containing different soap solutions*
(b) Showing the shaking of the test tube
(c) To compare the foaming capacity of different samples of soap

5. Take test tube labelled as tube A and shake it ten times by placing thumb on its mouth [see Fig. 21.1(b)].
6. By shaking the test tube, foam will be formed. Once the foam is formed, measure the length of the foam produced immediately with the help of a measuring scale [Fig. 21.1(c)].
7. Repeat steps 5 and 6 with the remaining three samples of soap solutions.

Observations and Calculations

(i) Mass of each soap sample taken in a beaker = ____ g

(ii) Volume of the distilled water added in each beaker = ____ mL

(iii) Volume of each soap sample taken in a test tube = ____ mL

(iv) Volume of distilled water added in each test tube = ____ mL

(v) Number of times each test tube shaken = _____

Soap solution	Test tube readings		Length of the foam produced
	Initial length	Final length	
	(cm)	(cm)	(cm)
1. 2. 3. 4.			

Results and Discussion

Infer from the observations that which soap sample produces the maximum length of foam (lather) in test tube.

Why does the different soap solutions have different foaming capacities? Is it due to the presence of different alkyl groups (R) in different soap solutions? An alkyl group in a soap is the hydrophobic part.

Precautions

- Use distilled water for each sample because foaming of a soap solution does not take place in hard water.
- Stir the mixture carefully while dissolving soap in water so as to avoid spilling of soap solution.
- The quantity of soap samples in all solutions must be same. The amount of distilled water added in every soap sample must be same. That is the concentration of all test solutions must be same.
- The mass of the soap samples must be determined very carefully using a physical balance. In case of any need, take help from your teacher.
- Shake every tube for equal number of times and in a similar manner.
- Measure the length of the foam produced immediately after its production.
- Use wire gauge to heat the beaker containing soap sample(s) uniformly.

Note for the Teacher

- Detergents should not be used in this experiment. However a similar experiment can separately be performed for comparing the foaming capacities of different detergents.
- Students may be guided to use a physical balance for weighing sample accurately.

Questions

- What is the name of the chemical reaction which takes place during the alkaline hydrolysis of oils and fats?
- Why is it necessary to shake every test tube for equal number of times and in a similar manner?
- Why does the concentration of every soap solutions be same?
- Was the length of foam formed same in each of the test tube containing soap solution?
- Which soap sample formed maximum foam?
- If distilled water had not been taken and salts of Mg^{2+} and Ca^{2+} were present in water sample. What would have been your observation?
- In hard water, which will form more foam - a soap or a detergent?
- In this experiment it is advised that the length of the foam produced should be measured immediately after its production. Why?

Experiment 22

AIM

To study the comparative cleansing capacity of a sample of soap in soft and hard water.

THEORY

Hardness of water is caused by the presence of the salts of calcium and magnesium (hydrogencarbonates, chlorides and sulphates) in water. These salts are soluble in water. When soap is added to hard water, it reacts with the salts to form a scum, which is insoluble and floats on top of the water surface. The scum is formed due to the formation of insoluble calcium or magnesium salts of the fatty acid used in the soap formation. The soap in solution then becomes ineffective.

$$Ca(HCO_3)_2\ (aq) + \underset{\substack{\text{Sodium stearate} \\ \text{(soap)}}}{2Na^{+}\text{-stearate}\ (aq)} \longrightarrow \underset{\text{Scum}}{Ca(\text{stearate})_2\ (s)} + 2NaHCO_3\ (aq)$$

$$CaSO_4\ (aq) + 2Na^{+}\text{-stearate}\ (aq) \longrightarrow \underset{\text{Scum}}{Ca(\text{stearate})_2\ (s)} + Na_2SO_4(aq)$$

The salts of calcium and magnesium show similar reactions. Therefore, the presence of calcium and magnesium salts in water precipitates the soap thereby reducing its cleansing power and foaming capacity.

MATERIALS REQUIRED

Underground water (well water), distilled water, calcium hydrogencarbonate or calcium sulphate, soap sample, a physical balance and weight box, three test tubes and a test tube stand, three beakers (100 mL), three glass rods, a measuring cylinder (50 mL), and a measuring scale,

PROCEDURE

1. Take three beakers and label them as A, B and C.
2. Take 20 mL of distilled water in beaker A. In beaker B, put 20 mL of underground water, and in beaker C add 2 g of calcium hydrogencarbonate (or calcium sulphate) to 20 mL of distilled water.
3. Stir the contents of beaker C so that calcium hydrogen carbonate (or calcium sulphate) dissolves in water.
4. Put 1 g of soap in each beaker A, B, and C (after weighing it using a physical balance).
5. Stir the contents of these beakers with separate glass rods.
6. Place three test tubes in a test tube stand and label them as tube A, B and C [Fig. 22.1(a)].
7. Pour 3 mL of the above-prepared soap solution in the corresponding test tubes.

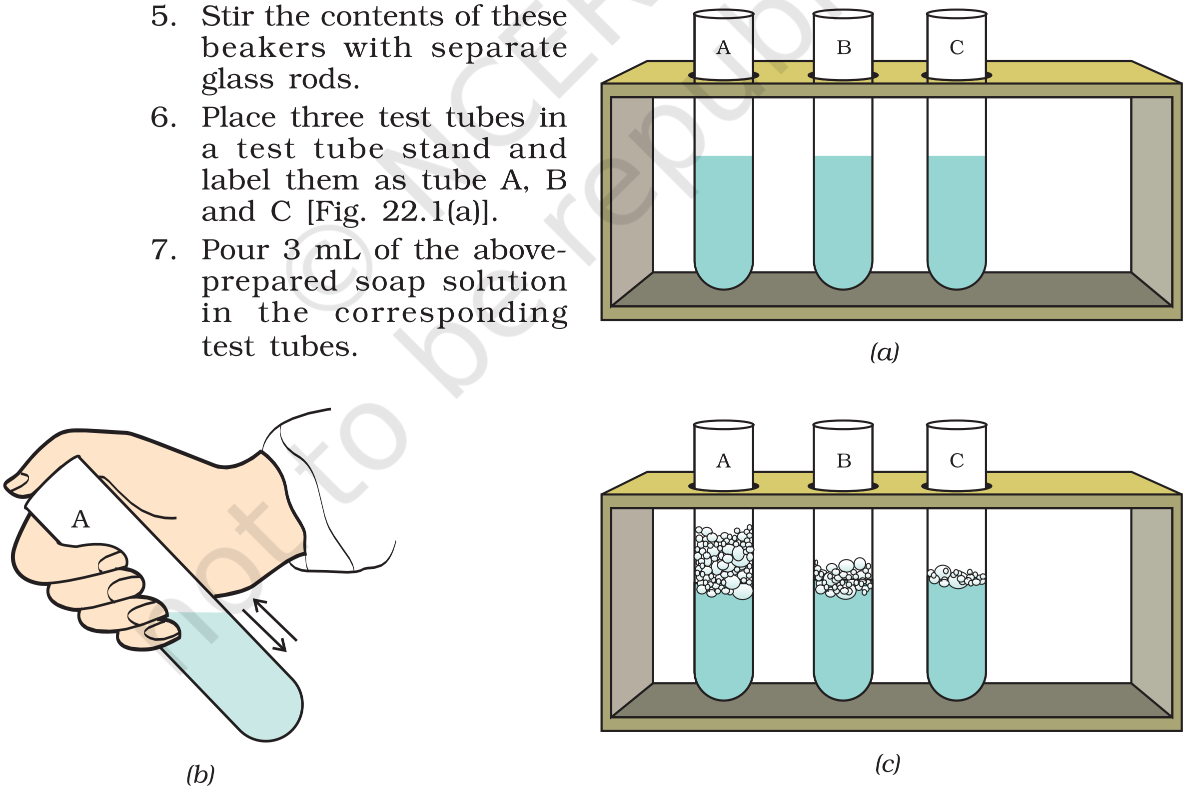

Fig. 22.1 : *(a) Test tubes containing different soap solutions*
(b) Showing the shaking of the test tube
(c) To compare the foaming capacity of different samples of soap

8. Take test tube A and shake it ten times by placing thumb on its mouth [Fig. 22.1(b)].
9. Foam or lather will be formed by shaking the test tube. Measure the length of the foam produced immediately with the help of a measuring scale [Fig. 22.1(c)].
10. Similarly, repeat steps 8 and 9 with the remaining two samples.

Observations and Calculations

(i) Mass of the soap sample taken in each beaker = _____ g

(ii) Volume of the distilled water and underground water added in each beaker = _____ mL

(iii) Volume of soap sample taken in each test tube = _____ mL

(iv) Number of times each test tube taken = _____

Sl. No.	Mixture (water + soap)	Test tube readings Initial length	Final length	Length of the foam produced
		(cm)	(cm)	(cm)
1.	Distilled water (soft water)			
2.	Well water or underground water (hard water)			
3.	Water containing $Ca(HCO_3)_2$ or $CaSO_4$ (hard water)			

Results and Discussion

Infer from the observations that which solution of the soap sample produces the maximum length of foam (lather).

For cleansing purpose, the foam needs to be produced which depends on free availability of hydrophobic portion of soaps (or alkyl groups). In hard water it is trapped due to scum or precipitation, this makes the hard water unsuitable for washing.

PRECAUTIONS

- Use same sample of soap for soft water and hard water.
- Stir the mixture carefully while dissolving soap in water so as to avoid spilling of soap solution.
- The quantity of soap sample in all solutions must be same. The amount of distilled water added in every soap sample must be same. That is the concentration of all test solutions must be same.
- The mass of the soap samples must be determined very carefully using a physical balance. In case of any need, take help from your teacher.
- Shake every tube for equal number of times and in a similar manner.
- Measure the length of the foam produced immediately after its production.

NOTE FOR THE TEACHER

- Students may be guided to use a physical balance for weighing sample accurately.

QUESTIONS

- Do both hard water and soft water produce foam with soap?
- Why is scum formed when hard water is treated with soap?
- Why did we add calcium hydrogencarbonate (or calcium sulphate) to beaker C?
- Was there any difference in the length of the foam formed in test tube C having water containing calcium hydrogencarbonate (or calcium sulphate) and test tube B containing well water or underground water?
- With their prolong use, white scales get deposited in the interior of boilers and electric kettles. What is the reason for this observation? How can these scales be removed?
- What do you understand by temporary and permanent hardness of water?
- What is the reaction between soap molecules and ions present in hard water?