(a) Name
   (i) compound P.
   (ii) solution Q.
   (iii) colourless gas R.
(b) Write a word equation for the reaction that took place.
(c) Comment on the observation that would be made if a similar compound of calcium was used in place of that of magnesium.

CHAPTER FOUR

Air and Combustion

Composition of Air

Air is a mixture of gases such as oxygen, nitrogen, carbon (IV) oxide, noble gases and water vapour. The approximate percentage composition by (Carbon dioxide) volume of these gases in air is; Nitrogen 78.1%, Oxygen 20.9%, Carbon (IV) oxide 0.03%, while water vapour varies from place to place. Noble gases comprise of about 1% by volume of air. They include Argon, Neon, Krypton and Xenon. Argon is the most abundant noble gas and forms about 0.9% while the other noble gases form 0.07% by volume of air.

Experiment 4.1(a): What percentage of air supports combustion?

Put dilute sodium hydroxide solutio in a trough. Place a small candle on a cork and float it on the solution as shown in figure 4.1. Cover it with a gas jar. Mark on the gas jar, the level of the solution. Remove the jar and light the candle. Gently, cover the burning candle with the gas jar. Leave the apparatus to cool to room temperature. Mark on the gas jar, the final level of the solution. Remove the gas jar and measure the change in height. Record all your observations.

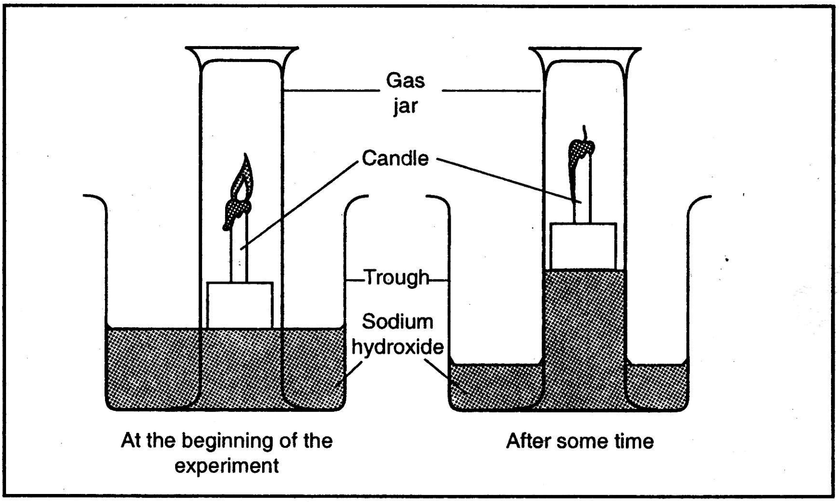

Fig. 4.1: Measurement of amount of air that supports combustion

Answer the following questions

1. Why does the candle go off after burning for sometime?
2. Explain why the level of dilute sodium hydroxide rises in the gas jar.
3. What was the length of the air column in the gas jar before and after burning?
4. Determine the percentage of air used up in burning.
5. Why is sodium hydroxide preferably used in the experiment instead of water?
6. State the sources of errors in the experiment.
7. Why is it necessary to leave the apparatus cool before taking the final reading?

Discussion

The candle burns for a while in a fixed volume of air, then it goes off. The level of the solution in the jar rises.

The following are sample results for a similar experiment

Volume of air before burning = 90 cm^3
Volume of air after burning = 76 cm^3
Volume of air used during burning = 14 cm^3
Percentage of air by volume used up:

$$= \frac{\text{Volume of oxygen used}}{\text{Initial volume of air}} \times 100$$

$$= \frac{14}{90} \times 100 = 15.5\%$$

Therefore, the solution in the jar rises to occupy the volume of air used up during combustion.

Combustion is a process of burning in which a substance combines with oxygen with the production of heat.

The part of air that supports combustion is active air. The 'active' part is oxygen which forms about 20% of dry air by volume. The part of air that remains in the gas jar does not support combustion. This component of air is 'inactive' and is mainly nitrogen.

Dilute sodium hydroxide is preferably used to absorb carbon (IV) oxide as that was initially in the gas jar and that produced during combustion.

Heating cause expansion of gases therefore the apparatus should be allowed to cool before the final reading is taken.
In this experiment, the answers obtained may not be accurate due to the following reasons:

- Dilute sodium hydroxide may not absorb all the carbon (IV) oxide gas.
- The candle may go off before all the oxygen is used up due to build-up of carbon (IV) oxide levels.

Experiment 4.1(b): What proportion of air is used up when copper is heated in a fixed volume of air?

Pack copper turnings completely in a six-centimetre long hard glass tube. Connect the tube to two glass syringes as shown in figure 4.2 with one syringe containing a specific volume of air while the other is empty.
CAUTION: Always handle glass wool with a pair of tongs.

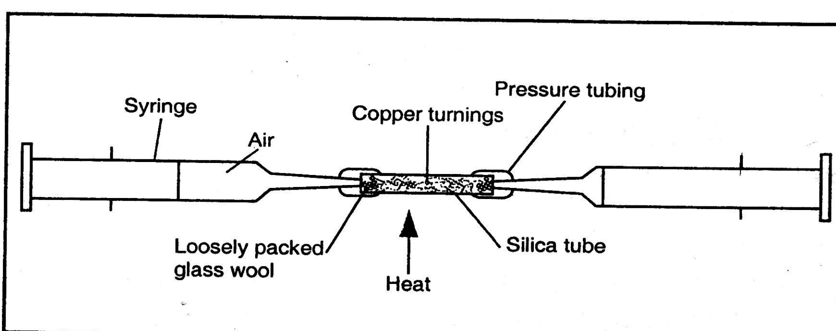

Fig 4.2: Measurement of amount of air used when copper is used

DANIEL MURAGE

Heat the copper turnings until they are red hot. Slowly pass the air from syringe A through the hot turnings to syringe B and back. Repeat this process four times while heating the copper turnings until the new volume of air in syringe B is constant. Allow the glass tube to cool and record the volume of the gas in syringe B.

Answer the following questions

1. What is the use of the glass wool plug in the experiment?
2. Why is the air passed through copper turnings
   (a) slowly?
   (b) repeatedly until there is no further change in volume?
3. (a) What is the volume of the air in the syringe
      (i) before heating?
      (ii) after heating?
   (b) Use the data obtained to calculate the
      (i) volume of air used up during the experiment.
      (ii) percentage by volume of air used during the experiment.
4. Why does the volume of air in the syringe decrease?
5. State the sources of error in this experiment.

Discussion

Copper is a reddish brown metal. When it is heated in air, it turns black. This is due to copper combining with oxygen to form copper (II) oxide which is black. Below is a word equation for the reaction.

Copper +	Oxygen	$\longrightarrow$	Copper (II) oxide
(red brown solid)	(colourless gas)		(black solid)

The following are sample results for the experiment:

Volume of air in the syringe before heating = $75cm^3$

Volume of air in the syringe after heating = $60cm^3$

Volume of air used up during heating = $15cm^3$

Percentage of air by volume used up is:

$$\frac{\text{Volume used during heating}}{\text{Original volume of air}} \times 100 = \frac{15 \times 100}{75} = 20\%$$

When air is passed over heated copper, a black solid is formed. About 20% by volume of air is used in this combustion and the 80% of air left does not react with heated copper.

In this experiment, the glass wool plug is used to stop the copper turnings from being blown into the syringe. The air is passed repeatedly over heated copper to ensure that all oxygen in the syringes, tubes and sample is used up. The air is passed slowly to allow enough time of contact between the reactants. The gas left over in the syringe does not react with copper. It is mainly nitrogen.

Experiment 4.1 (c): What percentage of air is used by rusting iron filings?

Wet a measuring cylinder and sprinkle some iron filings on the wet surface. Remove the excess iron filings. Invert the measuring cylinder in a trough of water. Leave the set up for 48 hours. Record all the observations.

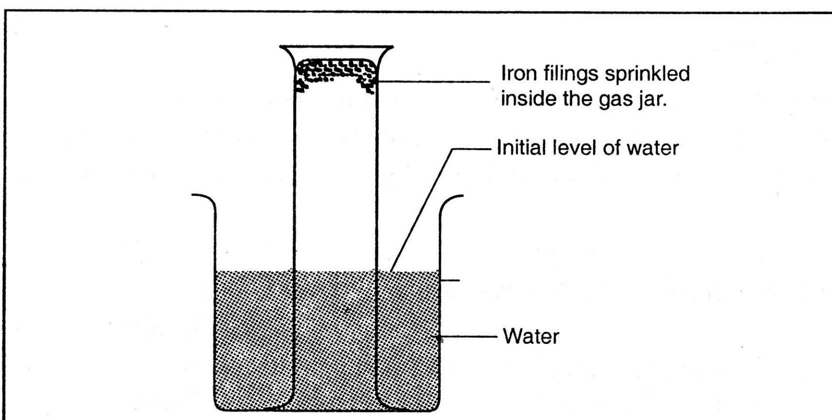

Fig. 4.3: Measurement of the amount of air used up during rusting of iron

Answer the following questions

1. What is the importance of wetting the measuring cylinder before sprinkling the iron filings?
2. What percentage volume of air used up in the experiment?

Discussion

When iron filings were left for 48 hours in a measuring cylinder, a brown coat formed on the filings. The brown coat is rust. Which is a compound of iron and oxygen. During rusting oxygen was used and therefore water rose up in the measuring cylinder to replace the volume of air used during rusting.

Experiment 4.1 (d): What percentage of air is used by smouldering white phosphorus?

Invert an empty measuring cylinder in a trough of water and note the level.of water in the cylinder. Attach a small piece of white phosphorus to the end of a piece of copper wire and place it in the measuring cylinder as shown in figure 4.4.

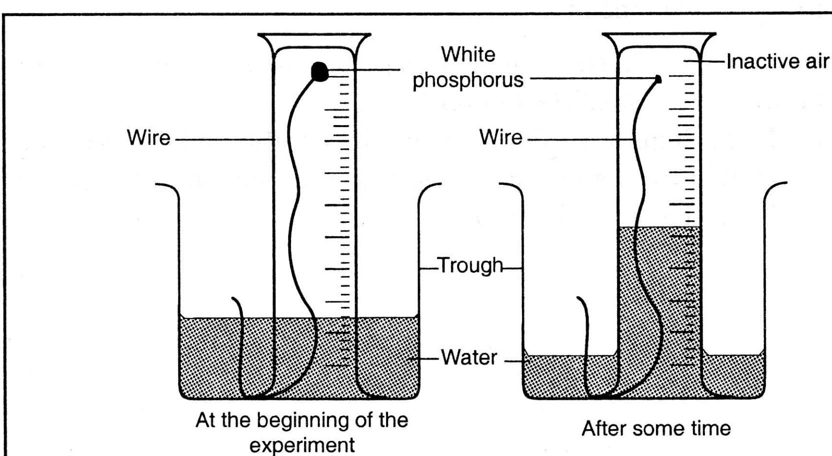

Fig. 4.4: Measuring of the amount of air used by smouldering white phosphorus

Note the level of water in the measuring cylinder. Leave the set up over-night and record your observations.

CAUTION: Avoid getting in contact with the phosphorus or inhaling the fumes.

Answer the following questions

1. What is observed when phosphorus is exposed to air?
2. Calculate the percentage volume of air used up in this experiment.
3. Why is phosphorus stored under water?

Discussion

Yellow or white phosphorus smoulders in air. This is because phosphorus reacts with oxygen to form oxides. This is why it is stored under water.

Word equation

Phosphorus + Oxygen $\longrightarrow$ Phosphorus (III) oxide

Phosphorus + Oxygen $\longrightarrow$ Phosphorus (V) oxide

After 24 hours, the water level inside the measuring cylinder will have risen to occupy the volume of oxygen used up. The difference in volume can be used to calculate the percentage of oxygen by volume in air.

Experiment 4.2: How can the presence of Carbon (IV) oxide and water be established in air?

(a) Place $2cm^3$ of fresh lime water in a boiling tube. Pass water slowly from a tap into an aspirator as shown in figure 4.5. Record your observations.

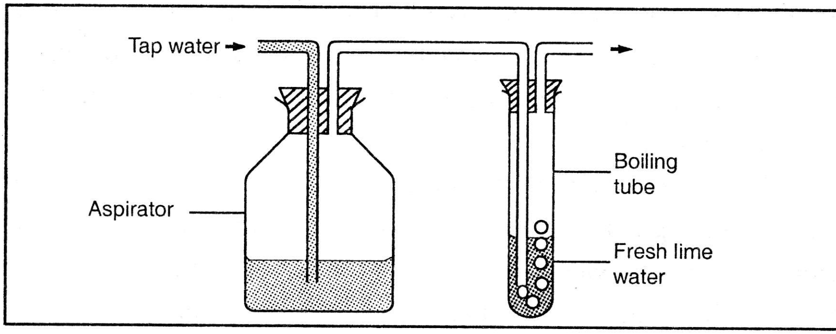

Fig. 4.5: Demonstration of presence of carbon (IV) oxide in air

(b) Pack the bottom of a U-tube with anhydrous calcium chloride as shown in figure 4.6. Pass air through the U-tube by means of a suction pump. Record your observations.

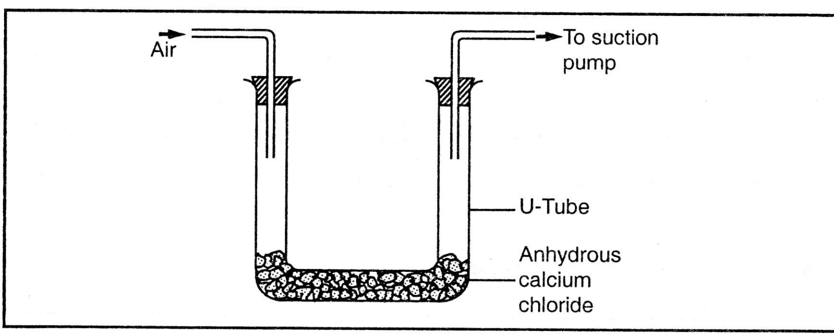

Fig. 4.6: Anhydrous calcium chloride set-up for demonstration of presence of water vapour in air

Answer the following questions

1. Why is it necessary to draw water into the aspirator in experiment 4.2 (a)?
2. Explain what was observed when air was bubbled through lime water?
3. What happens when air is passed through anhydrous calcium chloride?

Discussion

When a stream of air is passed through lime water a white precipitate is formed. This indicates that carbon (IV) oxide gas is present in air. Since carbon (IV) oxide forms a white precipitate with lime water.

Anhydrous calcium chloride is a white solid. When air is passed through the anhydrous calcium chloride, the calcium chloride absorbs water vapour from the air to form a colourless solution. Substances which absorb moisture from the air to form a solution are called **deliquescent substances**. These substances are used as drying agents in the preparation of dry gases. Other deliquescent substances are; anhydrous iron (III) chloride, magnesium chloride and zinc chloride.

Word equation

Anhydrous calcium chloride (white)	+ Water $\longrightarrow$	Calcium chloride solution (colourless)

The experiment above shows that air contains water vapour. A summary of the average composition of air is given in the table 4.1.

Table 4.1: Average composition of air

Gas	*Percentage Composition of volume*
Nitrogen	78.1
Oxygen	20.9
Carbon (IV) oxide	0.03
Argon	0.9
Other noble gases	0.07
Humidity, dust and impurities	variable

Fractional Distillation of Liquefied Air

As already established, air is a mixture of many gases. It can be separated into its constituent gases by fractional distillation of liquefied air.

The air is first purified by passing it through filters. The dust free air is passed through a solution of sodium hydroxide to remove carbon (IV) oxide gas. The air is then cooled to –25°C to remove water vapour which solidifies out as ice. The remaining air is then compressed to a pressure of 200 atmospheres. Repeated expansion and contraction of the air cools it to a liquid at–200°C. The liquid air now consists of oxygen, nitrogen and noble gases. Since these gases have different boiling points, they can be separated by fractional distillation. Liquid oxygen boils at –183°C and nitrogen at –196°C. Nitrogen distills out first because it has a lower boiling point. The other gases, made of mainly argon, boil at –186°C. The argon can be separated from oxygen by further distillation.

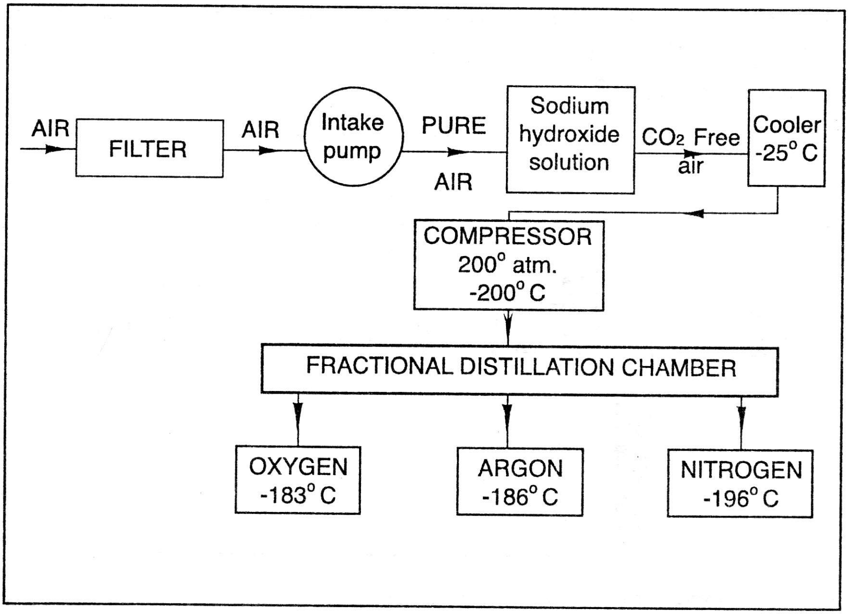

Fig. 4.7: Summary of fractional distillation of air

Rusting

Have you ever known why old iron nails or sheets look different from new ones? Old iron sheets are brown and dull while new ones are shiny. Many other metals change after sometime due to some chemical reactions which take place on their surfaces. These reactions are referred to as corrosion. Corrosion of iron due to reaction of iron with atmospheric oxygen is rusting. Rusting is the formation of a brown solid coating on the surface of iron or steel. The coating weakens their structure and hence eventually destroy them.

Experiment 4.3: What conditions are necessary for rusting?

Take five boiling tubes and label them 1, 2, 3, 4 and 5 respectively. Put two clean nails in each of the boiling tubes. To the first tube, add 10 cm^3 of tap water. To the second tube, add 10 cm^3 of boiled water followed by about 3 cm^3 of oil. To the third tube, push a piece of cotton wool half-way and place some anhydrous calcium chloride on it and cork the tube. Leave the fourth tube as set up initially. To the fifth tube, add salted water. The whole set-up is shown in figure 4.8.

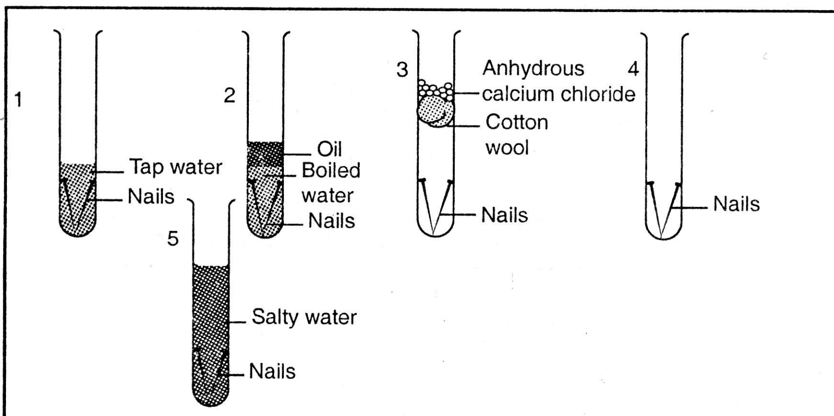

Fig. 4.8: Finding out what conditions favour rusting

Examine the nails after two days and record your observations.

Answer the following questions:

(i) What is observed in each of the tubes after two days?

(ii) Why is the water in the second tube

(a) boiled?

(b) covered with oil?

(iii) What is the purpose of anhydrous calcium chloride?

(iv) Why is it necessary to cork the third tube?

(v) List the conditions necessary for rusting to occur.

(vi) Suggest methods by which rusting can be prevented.

Discussion

It is observed that the nails in boiling tube 1 would have rusted after two days. The nails in tube 4 would have rusted to a smaller extent. The rusting is even intense in tube 5. No rust is observed on the nails in tubes 2 and 3.

Tap water contains dissolved oxygen. The iron nails combine with oxygen in the presence of water to form hydrated iron (III) oxide.

Iron + Oxygen $\longrightarrow$ Iron (III) oxide

Iron (III) oxide + water $\longrightarrow$ hydrated Iron (III) oxide (brown)

The brown and porous substance is what is referred to as rust.

Rusting occurred in tube 1 because both water and oxygen were present. Some rusting occurred in tube 4 since there was some moisture in the air that facilitated rusting of nails. The rusting was even faster when salt was added to the water was seen in tube 5. When water was boiled, all the gases dissolved in it were expelled. Further entry of air into the water was prevented by covering the boiled water with oil. There was no oxygen in tube 2 and therefore the nails did not rust. Anhydrous calcium chloride absorbs moisture from the air. It was necessary to cork tube 3 to prevent further entry of water vapour from the atmosphere. The nails in tube 3 do not rust because there was no water in it. The presence of water and oxygen are thus necessary for iron to rust. Rusting occurs faster in salty or acidic surroundings. For example, cars rust faster in Mombasa than in Nairobi.

Preventing Rusting

Rusting destroys machinery, equipment and roofs made of iron. Rusting can be very expensive. Prevention of rusting is therefore of great importance. The basis of prevention of rusting is to keep iron or steel out of directl contact with water and oxygen.

The following methods are widely used to prevent rusting of iron:

(i) Painting e.g. cars, roofs, ship etc.
(ii) Coating with other metals. This can be done through galvanisation or electroplating.
iii) Alloying: This involves the mixing of iron with one or more metals to produce a substance which does not rust.
(iv) Oiling and greasing: This method is used in moving engine parts where other methods cannot be used due to friction.

Oxygen

Oxygen exists freely in the atmosphere as a gas. Its chemical symbol is O. Two atoms of oxygen combine to form a molecule with a symbol of O_2. Oxygen may also be found in combined forms such as in water and metal oxides. It is the most active component of air.

Experiment 4.4 (a): How is oxygen prepared in the laboratory?

Half-fill a trough with water. Fill a gas jar with tap water and cover it with a gas jar cover slip. Invert the gas jar under water and remove the cover slip. Put about 2 g of Manganese (IV) oxide into a flat-bottomed flask. Add hydrogen peroxide from a thistle funnel drop-wise into the flask. Let the few bubbles escape then collect the gas as shown in figure 4.9.

Note the colour and smell of the gas. Lower a glowing splint into a gas jar of oxygen. Record all your observations.

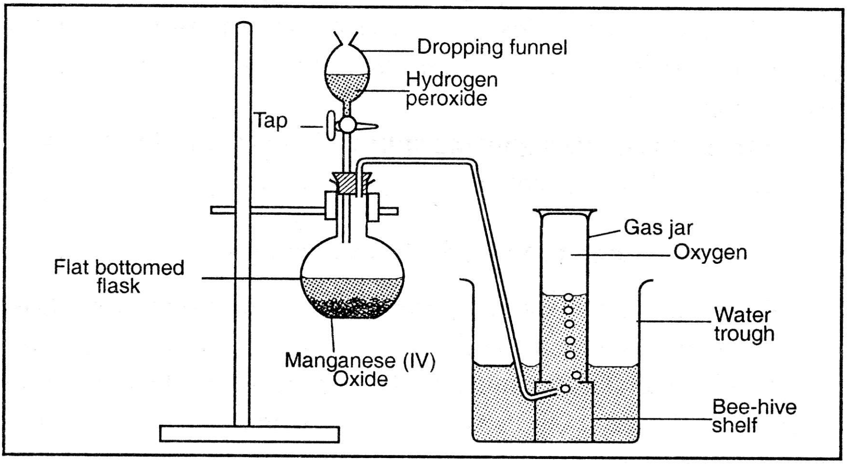

Fig. 4.9: Laboratory preparation of oxygen

Answer the following questions

1. Why are the first few bubbles of oxygen produced not collected?
2. What is the colour and smell of oxygen?
3. Why is it possible to collect oxygen over water?
4. What is the function of manganese (IV) oxide in the experiment?
5. Describe how oxygen is tested.

Discussion

Hydrogen peroxide decomposes slowly to oxygen and water under normal conditions. On adding manganese (IV) oxide decomposition is speeded up. A substance that alters the rate of reaction is called a catalyst.

Manganese (IV) oxide is a catalyst in the decomposition of hydrogen peroxide.

$$\text{Hydrogen peroxide} \xrightarrow[\textit{(IV) oxide}]{\textit{Manganese}} \underset{\text{(gas)}}{\text{Oxygen}} + \text{Water}$$

The first few bubbles of gas are not collected because the gas is mixed with air which was originally in the flask.

Oxygen is a colourless, odourless gas with a low boiling point of –183°C. It is almost insoluble in water and so it is collected over water.

Oxygen relights a glowing splint. This is the test for oxygen. Oxygen can also be prepared in the laboratory by:

(i) Adding Water to Solid Sodium Peroxide

Put about 2 g of solid sodium peroxide in a round-bottomed flask. Fill a gas jar with water and invert the jar under a trough half-filled with water. Let the few bubbles escape and collect the gas as shown in figure 4.10. Test the gas collected. Record your observations.

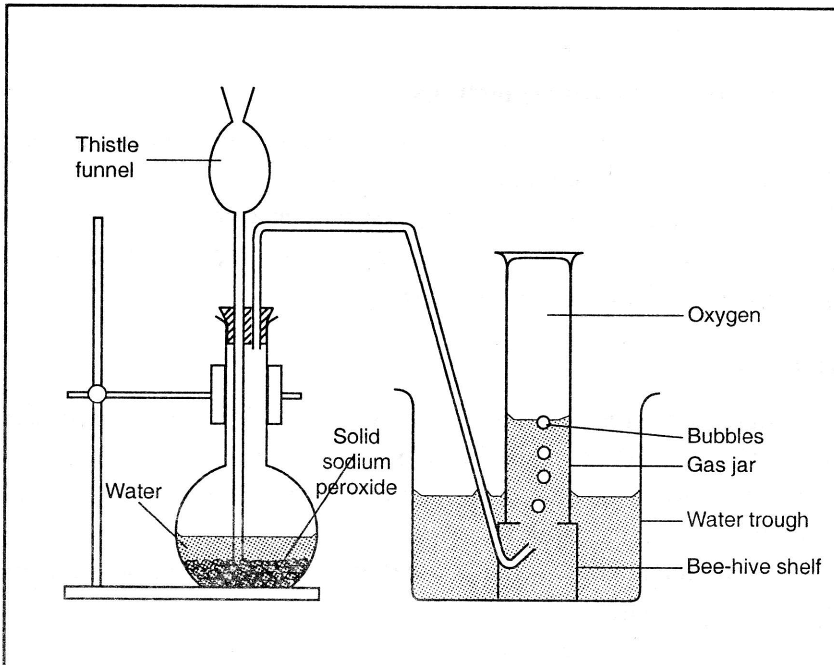

Fig. 4.10: Laboratory preparation of oxygen using sodium peroxide

(ii) Heating Potassium Manganate (VII) (Potassium Permanganate)

Oxygen is liberated and a black residue is formed.

Burning of Substances in Air

The most familiar chemical reaction of air is burning. There are many substances which burn in air. It has been established that heat causes change in the nature of a substance. The effect of burning simple elements, which are metals or non-metals, can be investigated.

Experiment 4.5: Is there change in mass when a substance burns in air?

Weigh about 1 g of clean magnesium ribbon in a crucible. Set up the apparatus as show in figure 4.11.

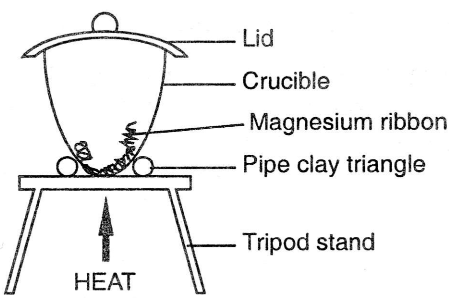

Fig. 4.11: Burning of substances in air

Heat the crucible, occasionally lifting the lid to allow air in. Do not allow any contents to escape from the crucible. When all the magnesium has burned, let the crucible to cool. Weigh the cooled crucible and its contents again. Record your observations as follows:

Mass of crucible + Magnesium before burning	= Xg
Mass of crucible + contents after burning	= Yg
Change in Mass	= Y–Xg

Answer the following questions

(i) Why is it necessary to occasionally allow air into the crucible?
(ii) Does the mass increase or decrease after the burning? Explain.

Discussion

When magnesium is burned in a closed crucible, most of the air inside is consumed. It is therefore necessary to allow in air so that the burning can continue. The mass of the product of burned magnesium is more than the original magnesium.

This shows that during burning, the magnesium combines with air to form a new product. Generally, when metals burn in air, there is an increase in mass. If the product of the burning substance is gaseous, there is a decrease in mass due to the escapes. For example, when phosphorus burns in air, it forms phosphorus (V) oxide gas resulting in the reduction in the mass.

Experiment 4.6: How do metals burn in air and in oxygen?

(i) Warm a piece of sodium or a deflagrating spoon until it begins to burn. Lower it into a gas jar of air as shown in figure 4.12. Record your observations.

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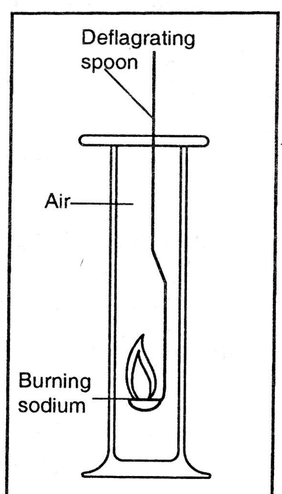

Fig. 4.12: Burning metals in oxygen

Allow the gas jar to cool and add some water to the product and shake well. Check for any gas given out and test it with damp red litmus paper. Test the products in the gas jar with litmus paper and record your observations. Repeat the experiment by lowering the sodium into a gas jar of oxygen. Now repeat the whole procedure using calcium in place of sodium.

(ii) Lower a piece of burning magnesium ribbon into a gas jar of air by using a pair of tongs. Let the jar to cool and add some water to the product. Shake well and test the solution formed with litmus paper. Note the smell of any gas produced and test it with damp red litmus paper. Record your observations. Repeat the procedure by lowering the burning magnesium into a gas jar of oxygen.

(iii) Attach a small piece of iron wire or steel wool to the end of a deflagrating spoon. Heat until the wire or steel wool glows. Lower it into gas jar of air. Record any observations. Repeat the experiment in a gas jar of oxygen.
Repeat the procedure using a piece of copper in place of iron. Record your observation as shown in table 4.2

Table 4.2: Burning Metals in Oxygen

M

Metal	*How it burns in air*	*How it burns in oxygen*	*Appearance of product*	*Name of product*	*Solubility of the product in water*	*Effect on litimus paper*
Sodium						
Magnesium						
Calcium						
Iron						
Copper						

Answer the following questions

(i) How do you compare the speed at which metals burn in air to that in oxygen?
(ii) What is the general name given to the products?
iii) Write a word equation for the reaction taking place for each of the metals.
(iv) Arrange the metals in order of their reactivity.
(v) Suggest the name of the gas given out when sodium and magnesium burned in air and water added to the products.

Discussion

Many metals burn in air and in oxygen at different speeds. They burn faster in oxygen than in air. In air, there are other constituents such as nitrogen and other gases which slow down burning. When metals burn in oxygen they form metal oxides. For example, Magnesium burns in oxygen with a bright white flame to form magnesium oxide.

Magnesium + Oxygen $\longrightarrow$ Magnesium oxide

Reactive metals such as sodium and magnesium react with nitrogen in the air to form nitrides.

Sodium + Nitrogen $\longrightarrow$ Sodium nitride

Magnesium + Nitrogen $\longrightarrow$ Magnesium nitride

When the nitrides react with water, alkaline ammonia gas is given out.

Sodium reacts most vigorously with oxygen while copper is the least reactive. A summary of the observations made when metals burn in air and in oxygen is given in table 4.3.