4 Air and Its Constituents

There is a thick blanket of air around the earth. Life on earth would not have been possible without air, which is actually a mixture of gases. In this chapter we will study about air and its constituents.

The Composition of Air

Air contains about 78% nitrogen and 21% oxygen by volume. The remaining 1% consists of argon, carbon dioxide, water vapour and other gases. Normally, air also contains dust and smoke. The average composition of dry air is given in Table 4.1.

Table 4.1 The average composition of dry air

Constituents	Formula	Percentage by Volume
Nitrogen	N_2	78.08
Oxygen	O_2	20.95
Argon	Ar	0.93
Carbon dioxide	CO_2	0.03
Other gases		0.01

Air is a Mixture, not a Compound

That air is a mixture and not a compound is clear from the following facts.

1. As you know, the composition of a compound is fixed, but that of a mixture is not. The composition of air is not fixed. It varies with location and season. For example, the proportion of carbon dioxide in air is greater in cities than in villages. The moisture content of air is greater in the rainy season than in other seasons.
2. The constituents of a compound retain their properties. So do the constituents of air. For example, oxygen present in air continues to support combustion.
3. The constituents of air can be separated by the fractional distillation of liquefied air.
4. An artificially prepared mixture of oxygen, nitrogen, carbon dioxide, etc., in the same proportion as in air, behaves in the same manner as air.

Pollution of Air

In urban and industrial areas, the air contains substances that are harmful to us and our environment. Such air is said to be polluted, and the harmful substances present in air are known as air pollutants. The common pollutants of air are:

- Carbon monoxide (CO)
- Oxides of sulphur (SO_2 and SO_3, collectively represented as SO_x)
- Oxides of nitrogen (NO and NO_2, collectively represented as NO_x)
- Chemicals called chlorofluorocarbons (CFCs) used as coolants and in aerosol sprays
- Dust and soot (collectively known as suspended particulate matter—SPM). Soot is unburnt particles from fuels, which give smoke its colour.

Air pollutants come from various sources, as mentioned in Table 4.2.

Table 4.2 Sources of the chief pollutants of air

Pollutant	Source
1. Carbon monoxide	(i) The incomplete combustion* of fuel in vehicles (ii) The incomplete combustion of coal in thermal power plants
2. Oxides of sulphur (Mainly SO_2. SO_3 is formed by the slow reaction of SO_2 with oxygen in the presence of soot.)	(i) Volcanic eruptions (ii) The burning of sulphur (iii) The extraction of metals from minerals containing sulphides
3. Oxides of nitrogen	The combustion of fuels like petrol, diesel, kerosene or coal
4. CFCs	(i) Leaking refrigerators and air conditioners (ii) Spray cans
5. Soot	The combustion of fuels
6. Dust particles	(i) Wind-blown particles of soil, plant spores, etc. (ii) The processing of minerals

*The complete combustion of a fuel containing carbon gives CO_2, which is not poisonous.

$2C + O_2 \rightarrow 2CO$ Incomplete combustion

$C + O_2 \rightarrow CO_2$ Complete combustion

Effects of Air Pollution

Effect on health

Air pollution causes many health problems.

- The oxides of sulphur and nitrogen cause breathing problems and respiratory infections. They also trigger asthma attacks.
- Carbon monoxide hampers oxygen transport in our body. In the lungs, oxygen gets attached to the haemoglobin present in the blood. From there it gets carried to the other parts of the body with blood. When carbon monoxide is inhaled, it combines with haemoglobin to form carboxyhaemoglobin. As a result, less haemoglobin is available for transporting oxygen. This causes weakness, headaches and, in extreme cases, death.
- The chlorine present in CFCs converts ozone to oxygen. CFC pollution is, therefore, reducing the amount of ozone in the ozone layer of the atmosphere. This layer shields us from the harmful effects of the ultraviolet (UV) rays of the sun. The decrease in ozone has led to an increase in skin cancer and eye problems.
- Very small particles that pollute the air get inhaled by us. These particles irritate and damage the lungs. They cause frequent attacks of asthma and bronchitis.

Acid rain

The oxides of sulphur and nitrogen present as pollutants in the air react with the water present in the atmosphere to form acids. These acids come down with rain, making the rainwater significantly acidic. Such rain is called acid rain. The acids found in acid rain are formed as follows.

$$\underset{\text{sulphur dioxide}}{SO_2} + \underset{\text{water}}{H_2O} \longrightarrow \underset{\text{sulphurous acid}}{H_2SO_3}$$

$$\underset{\text{sulphurous acid}}{2H_2SO_3} + \underset{\text{oxygen}}{O_2} \longrightarrow \underset{\text{sulphuric acid}}{2H_2SO_4}$$

$$\underset{\text{nitrogen dioxide}}{3NO_2} + \underset{\text{water}}{H_2O} \longrightarrow \underset{\text{nitric acid}}{2HNO_3} + NO$$

In some industrial areas, hydrogen chloride gas is present in the air as a pollutant. It dissolves in water to form hydrochloric acid, which comes down with the rain.

Effects of acid rain

1. The water of lakes and rivers becomes acidic and, thus, unsuitable for aquatic plants and animals.
2. The soil becomes acidic and, therefore, unsuitable for cultivation.
3. Sculptures, monuments and buildings are eroded.

Methods of reducing air pollution

The following methods are commonly used to reduce air pollution.

1. Modifying automobile engines Automobile engines are so modified that complete combustion of fuel takes place. This reduces the proportion of carbon monoxide and particulate matter in air.

2. Using a catalytic converter A catalytic converter, used in automobiles, converts unburnt fuel, CO and NO_x into carbon dioxide, water vapour and nitrogen. It uses platinum as a catalyst to help these conversions.

> A catalyst is a substance that is generally used to speed up a reaction, without being changed itself.

3. Using compressed natural gas (CNG) Nowadays, CNG is preferred to petrol or diesel in trucks and buses. It contains methane (CH_4) which on burning produces much less pollutants than coal, petrol, etc.

4. Using unleaded petrol In the past, a lead compound was mixed with petrol to help automobile engines run smoothly. This mixing produced poisonous lead compounds in the exhaust, which polluted the air. So unleaded (or lead-free) petrol is now used.

5. Removing pollutants from industrial waste gases You may have seen waste gases coming out of the chimneys of factories. But, before being discharged, the gases must be freed from most pollutants. This is usually done in two ways.

(i) They are passed through electrostatic precipitators, where particulate matter (like dust and soot) get arrested.

(ii) They are chemically treated to remove pollutants like the oxides of nitrogen and sulphur. Being acidic in nature, these oxides are absorbed in an alkaline solution (e.g., that of lime).

6. Using cleaner sources of energy A major part of the electricity in our country is produced by thermal power plants. Coal is burnt in these plants to get heat. The burning of coal produces gases that pollute our atmosphere. Hence, we must utilise cleaner sources of energy. For example, we can use hydroelectricity, solar energy, wind energy and hydrogen energy (the burning of hydrogen produces water, which is not a pollutant).

Oxygen

As you know, oxygen is essential for us. Yet chemists discovered it only in the eighteenth century. In the 1770s, C W Scheele and Joseph Priestly, working independently, discovered a gas that supported combustion. Soon after, Antoine Lavoisier recognised the gas as an element. He named it oxygen.

Occurrence

Oxygen gas constitutes one-fifth of air. In nature, it is also found in a large number of compounds such as water (H_2O), carbon dioxide (CO_2) and sand (SiO_2).

Preparation of Oxygen

1. From hydrogen peroxide

Principle Hydrogen peroxide decomposes slowly at room temperature to form water and oxygen. Manganese dioxide acts as a catalyst in this reaction.

$$\underset{\text{hydrogen peroxide}}{2H_2O_2} \xrightarrow[\text{catalyst}]{MnO_2} \underset{\text{water}}{2H_2O} + \underset{\text{oxygen}}{O_2}$$

Procedure

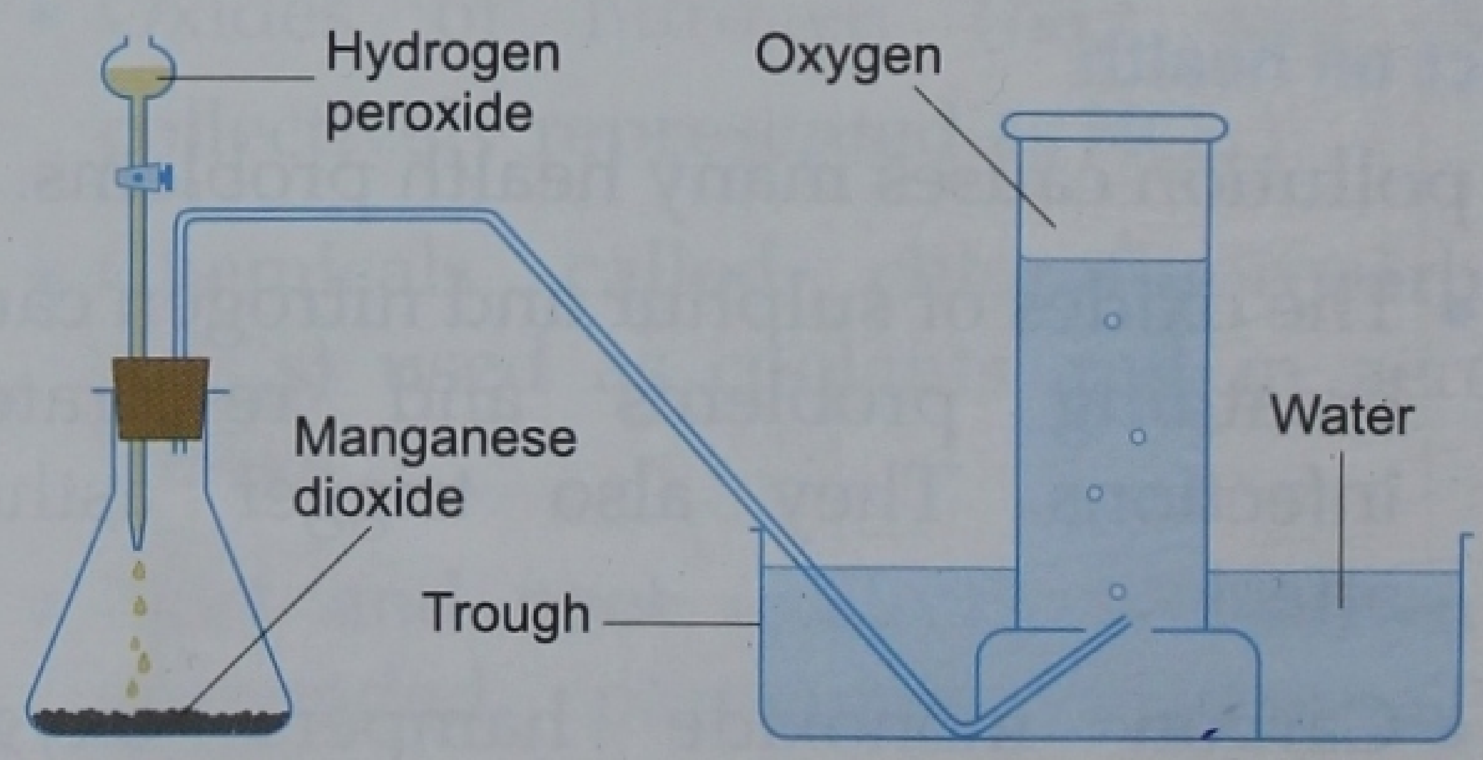

Fig. 4.1 The preparation of oxygen

The apparatus is fitted as shown in Figure 4.1. Some manganese dioxide, a black solid, is placed in the conical flask. Hydrogen peroxide is added drop by drop with the help of the dropping funnel. A brisk evolution of oxygen takes place.

Collection The gas pushes water downwards, and collects at the top of the inverted jar. This is known as collection by the downward displacement of water.

2. From some oxides of metals

Some oxides of metals decompose on being heated to give oxygen.

$$\underset{\text{mercury(II) oxide}}{2HgO} \xrightarrow{\text{heat}} \underset{\text{mercury}}{2Hg} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{silver oxide}}{2Ag_2O} \xrightarrow{\text{heat}} \underset{\text{silver}}{4Ag} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{lead dioxide}}{2PbO_2} \xrightarrow{\text{heat}} \underset{\text{lead oxide}}{2PbO} + \underset{\text{oxygen}}{O_2\uparrow}$$

3. From sodium/potassium nitrate

On being heated, sodium nitrate and potassium nitrate first melt and then decompose to give oxygen.

$$\underset{\text{sodium nitrate}}{2NaNO_3} \rightarrow \underset{\text{sodium nitrite}}{2NaNO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{potassium nitrate}}{2KNO_3} \rightarrow \underset{\text{potassium nitrite}}{2KNO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

4. From potassium chlorate

When heated, potassium chlorate crystals produce oxygen.

$$\underset{\text{potassium chlorate}}{2KClO_3} \xrightarrow[\text{[MnO}_2\text{]}]{\text{heat}} \underset{\text{potassium chloride}}{2KCl} + \underset{\text{oxygen}}{3O_2\uparrow}$$

Manganese dioxide is used as a catalyst in this reaction. If the manganese dioxide contains any carbon particles (both of them are black), the mixture will explode.

Physical Properties of Oxygen

1. Oxygen is a colourless and odourless gas.
2. It is slightly heavier than air.
3. Oxygen is only slightly soluble in water.
4. Solid oxygen, light blue in colour, melts at –218.4°C.
5. Liquid oxygen, pale blue in colour, boils at –183°C.

Chemical Properties of Oxygen

Oxygen is a reactive element and adds on to many substances.

> A reaction in which a substance adds on oxygen is called an oxidation reaction.

An oxidation reaction may be fast (as in burning) or slow (as in rusting).

Burning—Fast oxidation

> When a substance combines with oxygen, giving out heat and light, the phenomenon is called combustion.

Thus, burning is a combustion process. But, generally, to enable a substance to burn, it has to be heated first. For example, a strip of paper held above a flame catches fire when it becomes sufficiently hot. To enable coal to start burning, i.e., to ignite it, we use a flame to heat it.

> The temperature to which a substance must be heated before combustion takes place is known as its ignition temperature.

The combustion of nonmetals Many nonmetals, like hydrogen, carbon, sulphur and phosphorus, when ignited, burn in oxygen to form their oxides.

$$\underset{\text{hydrogen}}{2H_2} + O_2 \xrightarrow{\text{heat}} \underset{\text{water vapour}}{2H_2O\uparrow}$$

$$\underset{\text{carbon}}{C} + O_2 \xrightarrow{\text{heat}} \underset{\text{carbon dioxide}}{CO_2\uparrow}$$

$$\underset{\text{sulphur}}{S} + O_2 \xrightarrow{\text{heat}} \underset{\text{sulphur dioxide}}{SO_2\uparrow}$$

$$\underset{\text{phosphorus}}{P_4} + 5O_2 \xrightarrow{\text{heat}} \underset{\text{phosphorus pentoxide}}{2P_2O_5}$$

The combustion of metals Active metals, like calcium, sodium, magnesium and iron, when ignited, burn brightly in oxygen to form their oxides.

$$\underset{\text{calcium}}{2Ca} + O_2 \xrightarrow{\text{heat}} \underset{\text{calcium oxide}}{2CaO}$$

$$\underset{\text{sodium}}{2Na} + O_2 \xrightarrow{\text{heat}} \underset{\text{sodium peroxide}}{Na_2O_2}$$

$$\underset{\text{magnesium}}{2Mg} + O_2 \xrightarrow{\text{heat}} \underset{\text{magnesium oxide}}{2MgO}$$

$$\underset{\text{iron}}{3Fe} + 2O_2 \xrightarrow{\text{heat}} \underset{\text{tri-iron tetroxide, or ferrosoferric oxide}}{Fe_3O_4}$$

The combustion of compounds Many compounds are combustible, e.g., carbon monoxide and methane.

$$\underset{\text{carbon monoxide}}{2CO} + O_2 \xrightarrow{\text{heat}} 2CO_2 \uparrow$$

$$\underset{\text{methane}}{CH_4} + 2O_2 \xrightarrow{\text{heat}} CO_2 + 2H_2O \uparrow$$

Rusting—Slow oxidation

Iron is not affected by dry air, but it rusts in moist air.

> The slow conversion of iron into reddish brown hydrated iron(III) oxide by the action of moist air is known as rusting. The reddish brown product ($Fe_2O_3 \cdot xH_2O$) is known as rust.

In rust, the number of water molecules is not fixed.

Activity You can perform the following experiment to determine the conditions in which rusting occurs.

1. Take a few unrusted iron nails in a test tube A, and pour some tap water over them. Leave the test tube on a stand.
2. Boil some distilled water in test tube B in order to expel the dissolved air. Drop a few unrusted nails into it. Put some oil over the water surface to cut it off from the air above. Cork the test tube and allow it to stand.
3. Take some pieces of anhydrous calcium chloride ($CaCl_2$) in test tube C. Slip in a perforated cardboard disc (cut to size) into the test tube. Place a few unrusted nails over the disc, cork the test tube and allow it to stand. Remember that anhydrous calcium chloride absorbs moisture and so it will keep the air inside the test tube moisture-free.

Examine the nails after a few days.

Observation Rusting occurs only in the nails placed in test tube A, and not in the others.

Conclusion

(i) Only *moist air* causes rusting.

(ii) Neither air-free water nor moisture-free air can cause rusting.

How to prevent rusting

Rusting causes great damage. When rust is formed, cracks develop in the iron body. As a result, the metal loses strength and needs to be replaced. So rusting must be prevented as far as

Fig. 4.2 Rust on a nail

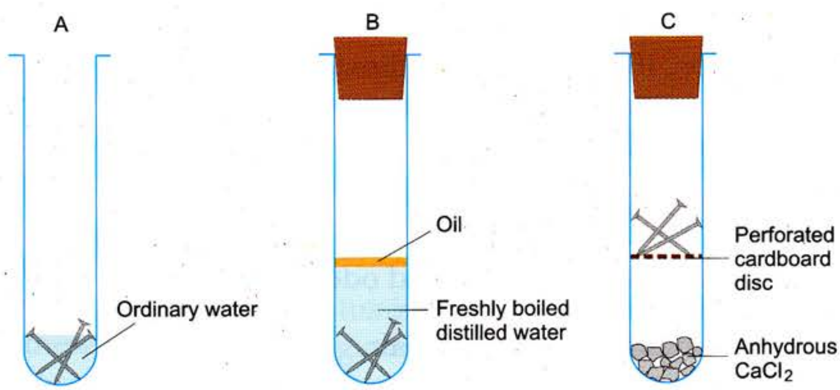

Fig. 4.3 Only moist air causes rusting.

possible. Some of the methods to do so are given below.

1. Painting Applying a paint over the metal surface protects it from air and moisture. This prevents it from rusting.

2. Greasing Applying grease over the surface also prevents rusting.

3. Galvanising Things made of iron are galvanised by dipping them in molten zinc. A coating of zinc forms on iron, which protects it from rusting. GI (galvanised iron) sheets and pipes are in common use.

4. Electroplating A fine deposit of tin, nickel or chromium over iron protects the metal from rusting. Such a deposit is made by a process known as electroplating. Tin cans are made of tin-plated iron sheets. Some cycle parts and bathroom fittings are made of iron plated with nickel or chromium.

Fig. 4.4 An electroplated tap

5. Alloying

> An alloy is a material consisting of a metal mixed homogeneously with one or more metals or nonmetals.

Generally, alloying prevents rusting. For example, stainless steel, which is an alloy of iron with carbon, manganese, nickel and chromium, does not rust.

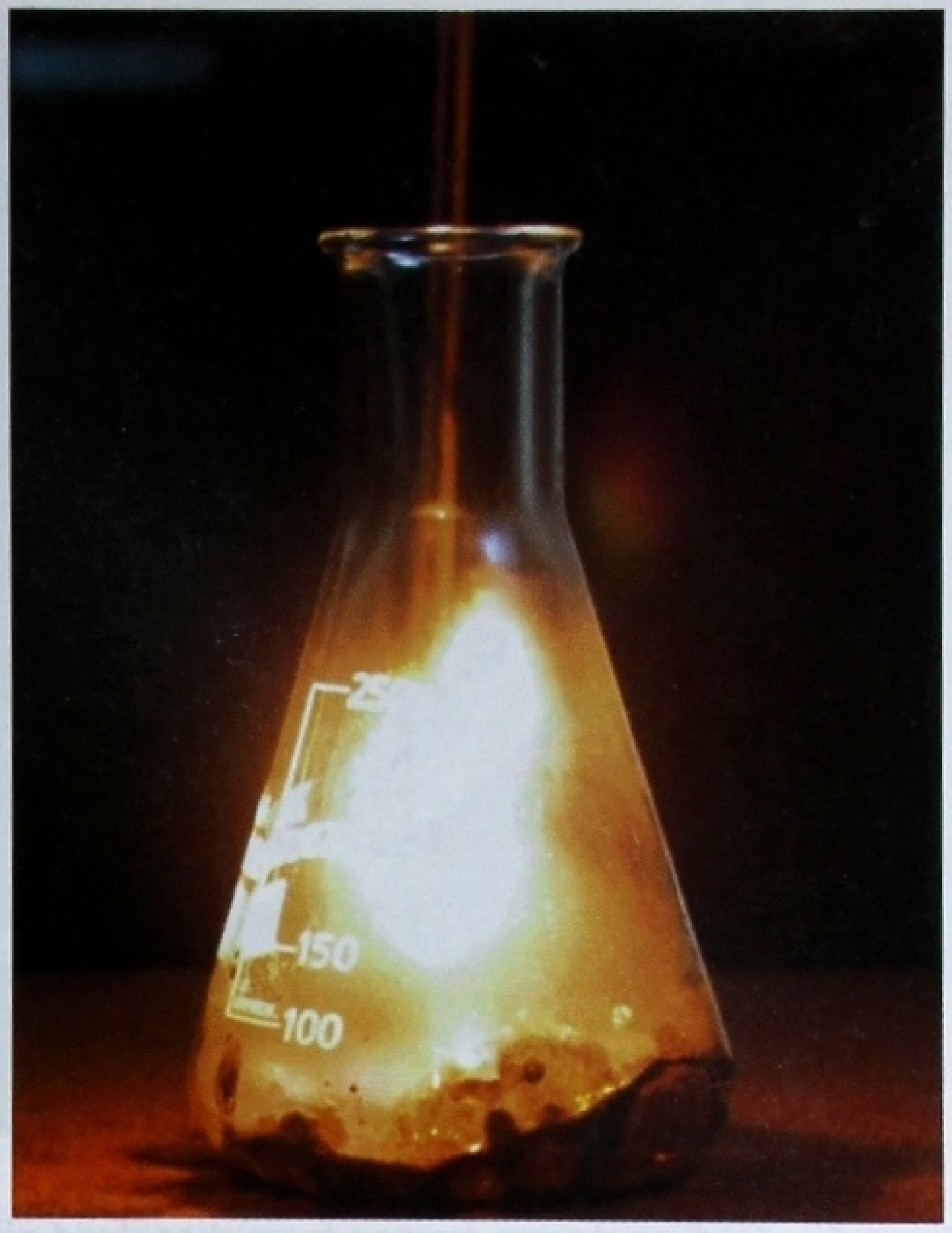

Fig. 4.5 A glowing splinter burns brightly in oxygen.

Test for Oxygen

A glowing splinter burns brightly in oxygen. Insert a glowing splinter into a conical flask containing MnO_2 and H_2O_2. Oxygen produced by their reaction will cause the splinter to burn brightly.

Uses of Oxygen

1. All combustion processes require oxygen, which is usually derived from air.
2. It is required for respiration. Oxygen burns food in our body to liberate energy.

$$\underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{oxygen}}{6O_2} \rightarrow \underset{\text{carbon dioxide}}{6CO_2} + \underset{\text{water}}{6H_2O} + \text{energy}$$

3. Oxygen is given to patients suffering from respiratory problems (due to heart attacks, asthma, pneumonia, smoke inhalation, etc.).
4. Deep-sea divers carry a mixture of oxygen and helium for respiration. If they carry air instead, the nitrogen present in it would dissolve in the blood under the high pressure in the depths of the sea. On returning to the surface, the nitrogen would come out of solution, forming bubbles. The nitrogen bubbles can form in the bones, muscles, brain, etc., causing pain, paralysis or even death.

(a)

(b)

Fig. 4.6 (a) Fish derive oxygen from dissolved air. (b) A deep-sea diver carries a mixture of oxygen and helium for respiration.

5. An oxyacetylene torch is used for cutting and welding metals. In such a torch, a mixture of oxygen and acetylene is burnt to produce a high-temperature flame.
6. Large amounts of oxygen are used in the iron and steel industry for various purposes.
7. Oxygen is used in large amounts in the manufacture of sulphuric and nitric acid.
8. Rockets carry liquid oxygen or some other source of oxygen for the combustion of fuels.

Fig. 4.7 Some rockets carry liquid oxygen for combustion.

How is the Oxygen of Air Renewed?

Although oxygen is used up in various processes, the proportion of this gas in air remains almost the same. This is because oxygen is released into the atmosphere during photosynthesis in green plants. In photosynthesis, the green pigment chlorophyll, present in leaves, makes carbon dioxide of the air react with water from the soil in the presence of sunlight. Glucose is formed, and oxygen is liberated into the atmosphere (Figure 4.8).

Carbon Dioxide

Carbon dioxide was first prepared in the laboratory around 1630. In 1783, Lavoisier recognised it as an oxide of carbon.

Preparation of Carbon Dioxide

Carbon dioxide may be prepared by various methods.

1. By burning carbon

When burnt in a sufficient supply of air, carbon forms carbon dioxide.

$$\underset{\text{carbon}}{C} + \underset{\text{oxygen}}{O_2} \xrightarrow{\text{heat}} \underset{\text{carbon dioxide}}{CO_2} \uparrow$$

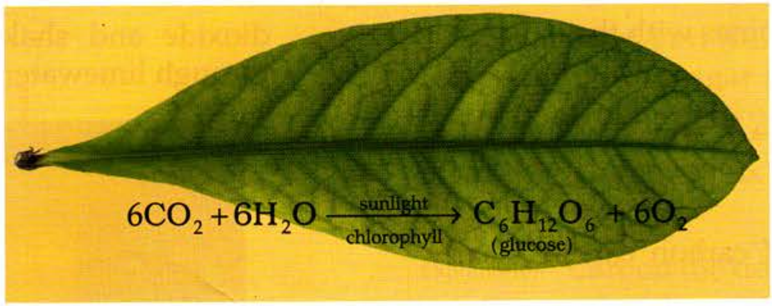

Fig. 4.8 The liberation of oxygen in photosynthesis

In an insufficient supply of air or oxygen, carbon monoxide is formed.

$$\underset{\text{carbon}}{2C} + \underset{\text{oxygen}}{O_2} \xrightarrow{\text{heat}} \underset{\text{carbon monoxide}}{2CO} \uparrow$$

2. By heating some metal carbonates

On being strongly heated, some metal carbonates decompose to give carbon dioxide.

$$\underset{\substack{\text{calcium carbonate}\\ \text{(limestone)}}}{CaCO_3} \xrightarrow{\text{heat}} \underset{\substack{\text{calcium oxide}\\ \text{(quicklime)}}}{CaO} + \underset{\text{carbon dioxide}}{CO_2} \uparrow$$

$$\underset{\substack{\text{magnesium}\\ \text{carbonate}}}{MgCO_3} \xrightarrow{\text{heat}} \underset{\substack{\text{magnesium}\\ \text{oxide}}}{MgO} + CO_2 \uparrow$$

3. By the action of an acid on a carbonate

Principle In the laboratory, carbon dioxide is prepared by the action of dilute hydrochloric acid on marble chips, which contain mostly calcium carbonate.

$$\underset{\substack{\text{calcium}\\ \text{carbonate}}}{CaCO_3} + \underset{\substack{\text{hydrochloric}\\ \text{acid (dilute)}}}{2HCl} \rightarrow \underset{\substack{\text{calcium}\\ \text{chloride}}}{CaCl_2} + \underset{\text{water}}{H_2O} + \underset{\substack{\text{carbon}\\ \text{dioxide}}}{CO_2} \uparrow$$

Procedure Some marble chips are placed in a conical flask fitted with a thistle funnel and a delivery tube. Dilute hydrochloric acid is added through the thistle funnel till the stem of the funnel dips in the liquid. Carbon dioxide is evolved with effervescence.

Collection of the gas The gas is heavier than air and is, therefore, collected by the upward displacement of air.

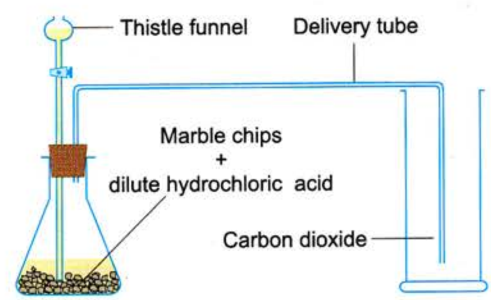

Fig. 4.9 Preparation of carbon dioxide

Physical Properties

1. Carbon dioxide is a colourless and odourless gas.
2. It is about 1.5 times as heavy as air.
3. It is only moderately soluble in water at ordinary pressure, but highly soluble under high pressure.
4. It can be liquefied by compression.
5. Carbon dioxide solidifies at a low temperature. Solid carbon dioxide sublimes, i.e., it gets converted into gas without melting. It is used to keep materials cold without wetting them. So it is called dry ice.

Chemical Properties

1. Carbon dioxide does not burn. It does not support combustion either. So, a burning splinter is extinguished in this gas.
2. A burning piece of magnesium continues to burn in carbon dioxide, though carbon dioxide is a nonsupporter of combustion. This is because magnesium is highly

reactive and combines with the oxygen of carbon dioxide.

$$\underset{\text{magnesium (burning)}}{2Mg} + CO_2 \rightarrow \underset{\text{magnesium oxide (white)}}{2MgO} + \underset{\text{carbon (black)}}{C}$$

Black particles of carbon can be clearly seen in the gas jar.

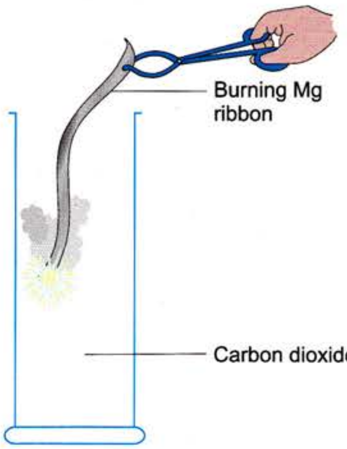

Fig. 4.10 A burning piece of magnesium continues to burn in carbon dioxide.

3. A solution of carbon dioxide in water is acidic due to the formation of carbonic acid.

$$\underset{\text{carbon dioxide}}{CO_2} + \underset{\text{water}}{H_2O} \longrightarrow \underset{\text{carbonic acid}}{H_2CO_3}$$

So, the gas turns moist blue litmus paper wine-red.

4. Carbon dioxide turns limewater milky due to the formation of white, insoluble calcium carbonate.

$$\underset{\text{calcium hydroxide (limewater)}}{Ca(OH)_2} + \underset{\text{carbon dioxide}}{CO_2} \rightarrow \underset{\text{calcium carbonate}}{CaCO_3} + \underset{\text{water}}{H_2O}$$

Tests for Carbon Dioxide

1. It extinguishes a burning splinter.
2. It turns moist blue litmus wine-red.
3. It turns limewater milky. To test, pour limewater into a gas jar containing carbon dioxide and shake it, or pass the gas through limewater in a test tube.

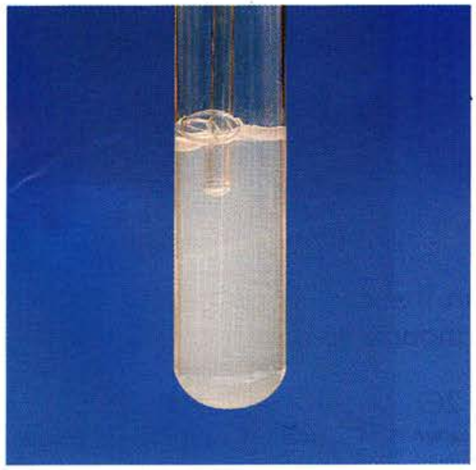

Fig. 4.11 Carbon dioxide turns limewater milky.

Uses of Carbon Dioxide

1. Carbon dioxide is used in photosynthesis. The food produced in photosynthesis is used by all plants and animals, either directly or indirectly.
2. It dissolves in blood—to form carbonic acid (H_2CO_3). This helps in maintaining the acid–base balance of the blood.
3. It is used in large amounts in the manufacture of urea, which is an important fertiliser.
4. It is used in preparing carbonated drinks like soda.
5. Solid carbon dioxide (dry ice) is used as a refrigerant to preserve ice cream, meat, etc. It is also used to preserve dead bodies.
6. As dry ice sublimes, it is used to produce the smokelike effect in movies, TV shows and theatres.
7. As carbon dioxide neither burns nor supports combustion, it is used for extinguishing fires.

The Fire Extinguisher

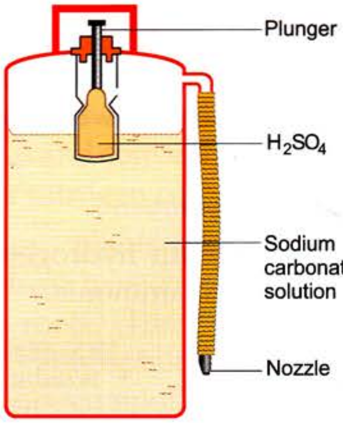

Fig. 4.12 A fire extinguisher

A fire extinguisher contains an aqueous solution of sodium carbonate or sodium hydrogencarbonate. A sealed glass bottle containing sulphuric acid is also placed in it. When a fire breaks out, the bottle of acid is broken by striking the plunger. The acid comes in contact with the solution, and a vigorous reaction takes place. As a result, carbon dioxide, along with water, gushes out and is directed towards the fire through the nozzle. The burning object loses contact with air and the fire is extinguished.

$$\underset{\text{sodium carbonate}}{Na_2CO_3} + \underset{\text{sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{CO_2}\uparrow$$

$$\underset{\text{sodium hydrogencarbonate}}{2NaHCO_3} + \underset{\text{sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{water}}{2H_2O} + \underset{\text{carbon dioxide}}{2CO_2}\uparrow$$

The action of vinegar on baking soda

Carbon dioxide is evolved when an acid acts on any carbonate or hydrogencarbonate. Vinegar is a dilute solution of acetic acid (CH_3COOH), and baking soda is sodium hydrogencarbonate ($NaHCO_3$). Vinegar reacts with baking soda, and carbon dioxide is liberated with effervescence.

Activity Prepare a cake mix or a batter by using eggs, milk and flour. Divide it into three parts. Leave one part (A) as it is. Mix some baking soda in another part (B), and baking soda and a teaspoonful of vinegar in the third (C). Leave them for an hour or two.

Observation The cake mix in C rises much better than in A or B.

Explanation Carbon dioxide, which evolves in the reaction, makes the mix rise.

Baking powder

The reaction of $NaHCO_3$ with a weak acid is utilised in raising a cake or a batter. For this, baking powder is commonly used. Baking powder is a mixture of baking soda and tartaric acid.

The Greenhouse Effect

The earth receives heat from the sun during the day. The heated surface of the earth in turn radiates heat. Atmospheric gases, such as carbon dioxide and water vapour, trap the radiated heat. Had heat not been trapped in this way, the earth would have become too cold at night to sustain life. Such trapping of heat is called the greenhouse effect because a similar thing happens in a greenhouse used for keeping plants warm. The walls and roof of a greenhouse are made of glass, which does not allow the heat trapped inside to escape.

In recent times, an increase in the carbon dioxide present in the air has increased the greenhouse effect. As a result, temperatures all over the world have risen. This is called global warming. Scientists fear that continued rise in temperatures will lead to melting of the polar ice caps. This will raise the sea level and submerge coastal areas.

Nitrogen

In 1772, Scheele proved that air contains a gas that does not support combustion and respiration. Lavoisier called it azote (*a* = not, *zote* = life). Later, J A Chaptal named the gas nitrogen.

Preparation of Nitrogen

1. From liquid air

Nitrogen is obtained on a large scale by the fractional distillation of liquid air.

2. From ammonium nitrite

Nitrogen is most easily obtained in the laboratory by heating a solution of ammonium nitrite.

$$\underset{\text{ammonium nitrite}}{NH_4NO_2} \rightarrow \underset{\text{nitrogen}}{N_2\uparrow} + \underset{\text{water}}{2H_2O}$$

It is collected by the downward displacement of water.

Physical Properties of Nitrogen

1. Nitrogen is a colourless and odourless gas.
2. It is sparingly soluble in water.
3. Liquid nitrogen (colourless) boils at –195.8°C.
4. Solid nitrogen (colourless) melts at –210.5°C.

Chemical Properties of Nitrogen

Nitrogen gas is chemically inert, i.e., unreactive under ordinary conditions.

1. Reaction with oxygen

Nitrogen combines with oxygen to form nitric oxide when an electric spark is passed through a mixture of the two gases.

$$\underset{\text{nitrogen}}{N_2} + \underset{\text{oxygen}}{O_2} \xrightarrow[\text{spark}]{\text{electric}} \underset{\text{nitric oxide}}{2NO}$$

This reaction also takes place in the upper atmosphere when there is lightning. The nitric oxide formed reacts with oxygen to produce nitrogen dioxide, which in turn reacts with water to form nitric acid. Nitric acid comes down with rain.

$$\underset{\text{nitric oxide (colourless)}}{2NO} + \underset{\text{oxygen}}{O_2} \rightarrow \underset{\text{nitrogen dioxide (red-brown)}}{2NO_2}$$

$$\underset{\text{nitrogen dioxide}}{3NO_2} + \underset{\text{water}}{H_2O} \rightarrow \underset{\text{nitric acid}}{2HNO_3} + \underset{\text{nitric oxide}}{NO}$$

The nitric acid that comes down with rain forms nitrate salts in the soil, which are taken up by plants.

2. Reaction with hydrogen

Nitrogen reacts with hydrogen under special conditions to form ammonia.

$$\underset{\text{nitrogen}}{N_2} + \underset{\text{hydrogen}}{3H_2} \xrightarrow[\text{Fe catalyst}]{500\,^{\circ}C,\ 200\text{ atm}} \underset{\text{ammonia}}{2NH_3}$$

Tests for Nitrogen

1. A burning splinter is extinguished by nitrogen.
2. It does not turn limewater milky nor does it act on litmus. This distinguishes it from carbon dioxide.

Uses of Nitrogen

1. As liquid nitrogen has a very low boiling point, it is used to preserve blood, corneas and donated organs.
2. Being chemically inert, it is used for filling food packages. In an inert atmosphere, the food does not go bad easily.
3. Atmospheric nitrogen is used to manufacture ammonia, which has a large number of industrial uses.
4. It is required in large amounts in the manufacture of urea—a very important nitrogenous fertiliser.
5. Some of its compounds, like hydrazine (N_2H_4), are used as rocket fuel.
6. Some explosives like trinitrotoluene (TNT) and nitroglycerine (dynamite) contain nitrogen. These are prepared using nitric acid.

Use of Nitrogen by Plants

Cells contain proteins and what are called nucleic acids. These compounds are rich in nitrogen. However they are not found in minerals or the atmosphere. Plants manufacture them from the nitrogen of the atmosphere. Plants get nitrogen in the following ways.

1. The nitric acid coming down with rain forms nitrate salts in the soil. Plants take up nitrates and convert them into proteins and other nitrogen compounds.
2. Leguminous plants (like beans, peas and other pulses) have certain bacteria living in their roots. These bacteria help to convert atmospheric nitrogen into nitrogenous compounds that the plants can use.

We depend heavily on plants for proteins and nucleic acids. Proteins are responsible for the growth of our muscles, blood cells, nails, hair, etc. Nucleic acids play an important role in reproduction and heredity.

If a soil is deficient in nitrogen, we add a nitrogen-rich fertiliser such as urea to the soil. This helps the crop grow better.

Noble Gases

Helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn) are collectively called noble gases because they are unreactive. Noble gases were discovered by Lord Rayleigh, William Ramsay and others in the last decade of the nineteenth century.

Argon is present in dry air to the extent of 0.93% by volume. Helium, neon, krypton and xenon are found in air only in traces. Radon, a radioactive element, is not present in air.

Noble gases have different boiling points and are, therefore, obtained by the fractional distillation of liquid air.

Uses of Noble Gases

Noble gases are used for various purposes.

Helium

1. It is a very light gas. Among gases, only hydrogen is lighter than it. So it is used for filling balloons—both ordinary and those used for weather study. Helium-filled balloons do not run the risk of catching fire, while hydrogen-filled ones do.

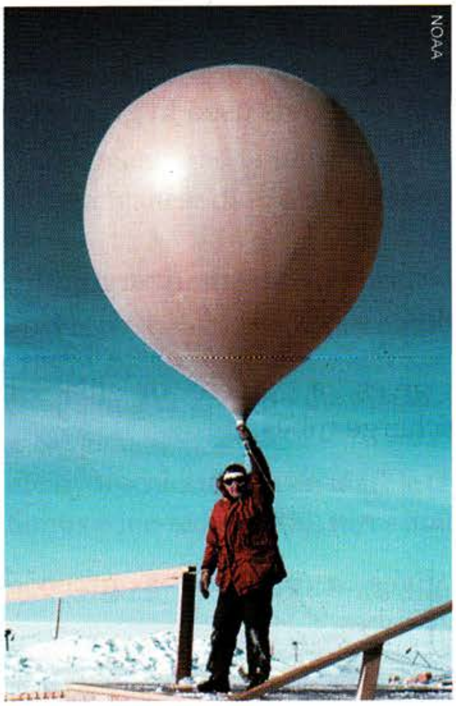

Fig. 4.13 A weather balloon

2. A mixture of helium and oxygen is carried by deep-sea divers for respiration.

Neon

Neon is used mainly for illumination. Neon signs are used for advertisements. The gas gives a beautiful orange-red light when an electric spark is passed through it at a high voltage. The sign is made of a glass tube, called discharge tube, containing the gas at a very low pressure. At the two ends of the tube there are electrical leads for the passage of electricity. Mixed with a little mercury vapour in the tube, neon gives a blue light.

Fig. 4.14 A neon sign

Argon

1. Argon is used to fill electric bulbs. An electric bulb should not contain air as the filament will then get oxidised.
2. Mixed with a little mercury vapour in a discharge tube, argon gives a beautiful green light. It is used in green signs.

Krypton

1. It works better than argon in electric bulbs.
2. It gives a green or lilac light in discharge tubes.

Xenon

Xenon gives a blue or a green light in discharge tubes.

Water Vapour

When there is a great deal of water vapour, i.e., moisture, in air, we say the air is humid. If air holds the maximum moisture that it can at a given temperature, it is said to have 100% humidity. Humidity is expressed relative to this maximum.

Relation of Humidity with Season and Location

Humidity is high in the rainy season, but low in other seasons. When the humidity is low, the air accepts a greater amount of moisture, and hence, evaporation from water surfaces is fast. On the other hand, when the humidity is high, the evaporation is slow. So things dry more slowly during humid weather (e.g., in the rainy season) than in dry weather.

Humidity is also dependent on location. It is higher at places near the sea, e.g., Kolkata, Puri, Chennai, Mumbai or Karachi, than at places far from the sea, e.g., Chandigarh, Delhi or Bhopal. Further, humidity decreases with altitude because water vapour is heavier than air.

Precipitation

The coming down of water in any form—liquid or solid—from the sky to the earth is called precipitation.

In the sky, water vapours condense on dust particles and form droplets of water or ice crystals. Millions of these droplets or crystals float together in the air as clouds. When small droplets of water combine to form bigger and heavier ones, they come down in the form of rain.

Under certain conditions, raindrops freeze to form hailstones. These come down as hail.

The small ice crystals in clouds may come together to form larger and heavier snowflakes. In cold regions snowflakes come down as snow.

The Water Cycle

The amount of water on the surface of the earth remains the same. This is due to the water cycle.

Water evaporates from large water bodies like seas, rivers and lakes. In the sky, they condense on dust particles to form clouds which drift from one area to another. The clouds bring rain or snow, and give the water back to the earth. The water vapours are also directly converted to snow over high mountains. The snow melts and the water flows down through the rivers to the sea. This entire process is known as the water cycle.

Points to Remember

- Air is a mixture of gases.
  By volume, it contains about 78% nitrogen and 21% oxygen. The remaining 1% consists of several gases like argon and other noble gases, carbon dioxide, water vapour, etc. Besides, smoke and dust are also present.
- Air containing harmful substances is called polluted air. The common pollutants of air are carbon monoxide, sulphur dioxide, oxides of nitrogen, soot and dust.
- Acid rain contains mainly sulphuric and nitric acids.
- Acid rain has harmful effects on plants, animals, soil, sculptures, monuments and buildings.
- Oxygen can be prepared by decomposing hydrogen peroxide using manganese dioxide as a catalyst.
- A reaction in which a substance adds on oxygen is called *oxidation*.
- When a substance combines with oxygen giving out heat and light, the phenomenon is called *combustion*.
- The slow conversion of iron into reddish brown hydrated iron(III) oxide by the action of moist air is known as *rusting*. The reddish brown product $Fe_2O_3 \cdot xH_2O$ is known as *rust*.
- Rusting can be prevented by painting, greasing, galvanising, electroplating or alloying the metal.
- A glowing splinter burns with a flame in oxygen. This is a test for oxygen.
- Carbon dioxide is prepared in the laboratory by the action of dilute hydrochloric acid on marble chips.
- Carbon dioxide (i) extinguishes a burning splinter, (ii) turns blue litmus wine-red, and (iii) turns limewater milky.
- Carbon dioxide extinguishes fires. A *fire extinguisher* produces carbon dioxide by the action of a dilute acid on a carbonate or a hydrogencarbonate.
- Nitrogen is obtained on a large scale by the fractional distillation of liquid air.
- Nitrogen extinguishes a burning splinter but does not act on litmus or limewater.
- Helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn) are collectively known as *noble gases*. Except radon, they are present in air in small amounts.
- Helium is the lightest element after hydrogen. It is used for filling balloons and is mixed with oxygen for respiration by deep-sea divers.
- Argon is used in filling electric bulbs.
- Neon, argon, krypton and xenon are used for illumination and in signs meant for advertisements.
- When there is too much water vapour present in the atmosphere, we say the air is *humid*.
- Humidity changes with season and location. It is high in the rainy season and at places on the seashore. It decreases with altitude.
- Rain, hail and snow are forms of precipitation.

Exercise

Short-Answer Questions

1. Arrange oxygen, nitrogen, carbon dioxide and argon in decreasing order of proportion in air.
2. Is air a mixture or a compound?
3. Name the method by which the constituents of liquid air may be separated.
4. Name three gaseous pollutants of air.
5. Name the gas that combines with haemoglobin to form carboxyhaemoglobin.
6. Why is CNG preferred to petrol or diesel as a fuel?
7. What is a catalyst?

8. What does a catalytic converter do when used in an automobile?
9. Give the formula of rust.
10. What are the following called?
    (a) The phenomenon in which a substance combines with oxygen giving out heat and light
    (b) The reaction in which a substance adds on oxygen
11. If a diver inhales air deep under the sea, a gas dissolves in his blood, and this gas bubbles out when he comes up. Name the gas.
12. By which natural process is the oxygen of air renewed?
13. What is the common name of solid carbon dioxide and why is it called so?
14. What happens when nitrogen dioxide reacts with water?
15. Name the gas used for filling food packages.
16. Name three forms of precipitation.

Long-Answer Questions

1. Describe a method of preparing oxygen without using a source of heat. Mention a test for oxygen.
2. What are the conditions essential for rusting? Give two methods used to prevent rusting.
3. Describe the laboratory method of preparing carbon dioxide.
4. By which tests can you distinguish between carbon dioxide and nitrogen?
5. Write a few lines to show how the amount of water on the earth is constant.

Objective Questions

Choose the correct option.

1. Which of the following constituents of air helps plants manufacture carbohydrates?
   (a) Nitrogen (b) Carbon dioxide
   (c) Helium (d) Neon
2. Which of the following gases is essential for respiration?
   (a) Nitrogen (b) Oxygen
   (c) Carbon dioxide (d) Argon
3. Electrostatic precipitators are used to remove
   (a) soot and dust from waste gases in industries
   (b) sulphur dioxide from waste gases in industries
   (c) lead from petrol
   (d) carbon monoxide from air
4. Which of the following gaseous mixtures is carried by deep-sea divers for respiration?
   (a) Liquid air (b) Oxygen and carbon monoxide
   (c) Oxygen and carbon dioxide (d) Oxygen and helium
5. Which of the following gases can be used along with oxygen in a torch for cutting and welding metals?
   (a) Acetylene (b) Carbon dioxide
   (c) Nitrogen (d) Water vapour
6. Which of the following is a noble gas?
   (a) Nitrogen (b) Hydrogen
   (c) Helium (d) Methane

Fill in the blanks.

1. Carbon monoxide is produced by the ...... combustion of fuels like coal, petrol or diesel. (complete/incomplete)

2. Combustion is a/an ...... reaction. (oxidation/decomposition)
3. Rusting does not take place in ...... air. (moist/dry)
4. A burning splinter is extinguished by .... (oxygen/nitrogen)
5. During lightning, nitrogen combines with oxygen to form .... (nitric oxide/nitrogen dioxide)
6. Urea supplies ...... to soil or plants. (nitrogen/oxygen)
7. Humidity ...... with altitude. (increases/decreases)
8. ...... is preferred to hydrogen in weather balloons. (Helium/Argon)
9. A neon sign glows with a beautiful ...... colour. (blue/green/orange-red)
10. Complete the following equations and balance them.

    (a) $\underset{\text{(burning)}}{Mg} + O_2 \longrightarrow ......$

    (b) $CO + O_2 \xrightarrow{\text{ignite}} ......$

    (c) $CaCO_3 + HCl \longrightarrow CaCl_2 + ... + ...$

    (d) $Ca(OH)_2 + CO_2 \longrightarrow ...... + H_2O$

    (e) $CH_4 + O_2 \xrightarrow{\text{ignite}} ...... + H_2O$

Indicate which of the following statements are true and which are false.

1. Sulphur dioxide does not cause acid rain.
2. Acid rain damages sculptures and monuments.
3. Nitrogen is a poisonous gas.
4. Upon dissolving in water, carbon dioxide forms carbonic acid.

You can make your own fire extinguisher

To make a fire extinguisher, you would need a jar with a plastic cap (about the size of a 50-g coffee jar), a 2-mL disposable syringe, a drinking straw that bends at the neck, baking soda and hydrochloric acid used for cleaning bathrooms.

Heat a large nail, pointed screwdriver or a knitting needle (made of metal) over a candle flame. Use it to make two holes in the plastic cap. One hole should be just large enough for the straw and the other, for the plastic base of the needle of the syringe. Fit the needle and the straw, and use adhesive to seal them in place. The straw should extend about 3–4 cm below the cap.

Now dissolve about 2 teaspoons of baking soda in a cup of water. Pour the solution in the jar. Attach the syringe to the needle and draw acid into it. Screw on the cap. We are now all set to test the fire extinguisher. Point the straw 'nozzle' towards a candle flame and press the plunger of the syringe to inject the acid into the jar. The acid will react with the baking soda to form carbon dioxide. The gas will come out of the nozzle and put out the flame.

$$\underset{\text{sodium hydrogencarbonate (baking soda)}}{NaHCO_3} + \underset{\text{hydrochloric acid}}{HCl} \rightarrow \underset{\text{sodium chloride}}{NaCl} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{CO_2}$$

You could use vinegar (acetic acid) instead of hydrochloric acid. Since vinegar is a weak acid, you would need a greater volume of vinegar. This means a much larger syringe. You could also use washing soda in place of baking soda. Experiment and find out what works best for you.

❑