8. OXYGEN AND ITS COMPOUNDS

8.1 Introduction

Oxygen occurs naturally in the free state and in combination with other elements. In the free state, it constitutes 21% of air by volume. Oxygen is one of the most abundant elements in the earthâ€™s crust where it exists in a variety of forms such as water, metallic and non-metallic oxides, salts, and complex silicates.

Oxygen is a very reactive diatomic non-metal. It is highly electronegative and forms ionic compounds readily with metals, and covalent compounds with nearly all non-metals. These compounds are called oxides. The binary ionic compounds have strong ionic bonds such as are present in Fe_2O_3, MgO and CuO. Hence, they have high melting points, and are basic oxides. The covalent compounds have covalent bonds and exist as discrete molecules such as CO,SO_2, etc, with low melting points. However, some covalent oxides exist as giant molecules with very high melting points as in SiO_2.

8.2 Laboratory Preparation of Oxygen

Oxygen may be prepared in the laboratory by several methods as described in the following experiments.

Experiment 8.1 : Laboratory preparation of oxygen from hydrogen peroxide.

Put some manganese(IV) oxide in a flat-bottomed flask carrying a cork with a delivery tube and a thistle funnel as in Figure 8.1. Fill the funnel with 20 volume hydrogen peroxide. Open the tap and run in the peroxide solution drop by drop to react with the manganese(IV) oxide.

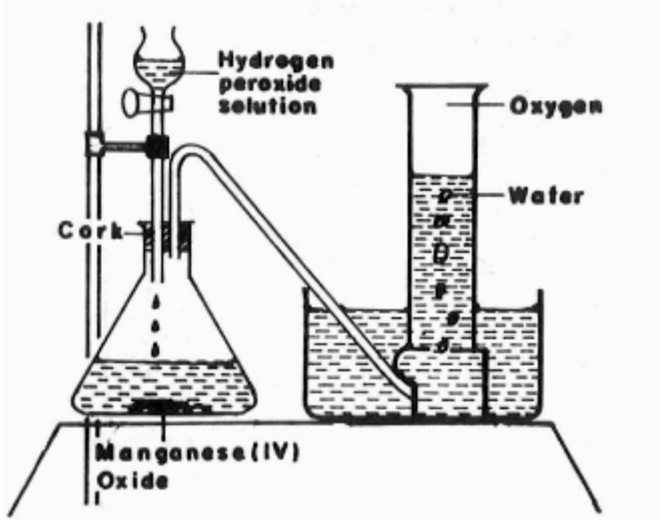

Figure 8.1. Preparation of oxygen from hydrogen peroxide.

The reaction produces oxygen gas.

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O\ (l) + O_2(g)$$

The manganese (IV) oxide acts as a catalyst.

If dry oxygen is required, the wet gas is first passed through a wash bottle of concentrated tetraoxosulphate(VI) acid, and then collected over mercury because it has about the same density as air.

Experiment 8.2: Preparation of oxygen from potassium tetraoxomanganate(VII) and hydrogen peroxide.

The manganese(IV) oxide in Experiment 8.1 may be replaced with potassium tetraoxomanganate(VII) dissolved in excess dilute tetraoxosulphate(VI) acid. When the tap is opened, hydrogen peroxide reacts with it, liberating oxygen. The liberation of oxygen gas continues until the purple colour of the potassium tetraoxomanganate(VII) disappears, indicating that the salt has been used up.

$5H_2O_2(aq) + 2KMnO_4(aq) + 3H_2SO_4$ â†’ $K_2SO_4(aq) + 2MnSO_4(aq) + 8H_2O(l) + 5O_2(g)$

Experiment 8.3: Preparation of oxygen from potassium trioxochlorate(V).

Set up the apparatus as in Figure 8.2. Mix potassium trioxochlorate(V) with manganese(IV) oxide in the ratio of 4 to 1 by mass, and put the mixture in the hard test-tube. Heat the mixture gently at first, and then strongly Oxygen gas is readily liberated. Collect it by displacement of water. The manganese(lV) oxide acts as a catalyst.

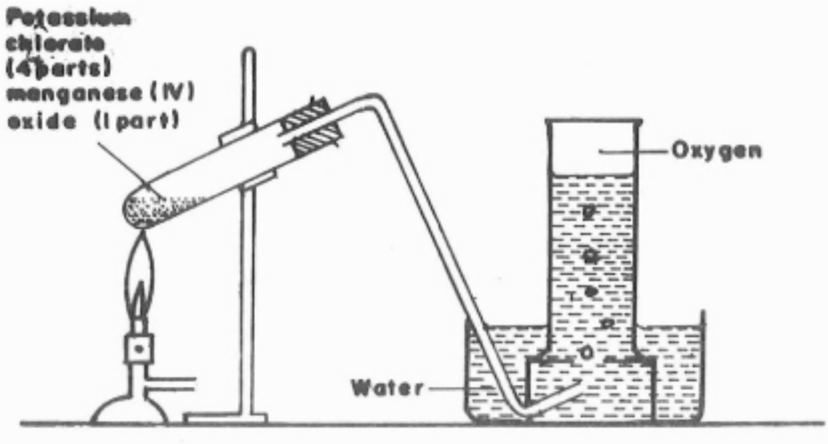

Figure 8.2 Preparation of oxygen from potassium trioxochlorate(V)

$$2KClO_3(s) \xrightarrow{MnO_2} 2KCl(s) + 3O_2(g)$$

At the end of the reaction, disconnect the delivery tube from the water trough before removing the source of heat. Otherwise, the reduction in pressure will cause water to rise into the hot tube and cause an explosion!

8.3 Properties and Reactions of Oxygen

Experiment 8.4: Investigating the properties and reactions of oxygen.

Using the gas jars of oxygen collected from Experiment 8.1., perform the following tests.

1. Observe the colour of the gas and note its smell.
2. Put moist blue and red litmus papers into the gas jar, and note any change of colour.
3. Plunge a glowing splint into one of the gas jars containing oxygen. The glowing splint is rekindled (bursts into flame)!
4. Lower a deflagrating spoon containing burning sulphur into a gas jar of oxygen. The burning continues more vigorously (Figure 8.3).
5. Repeat (4) using phosphorus.

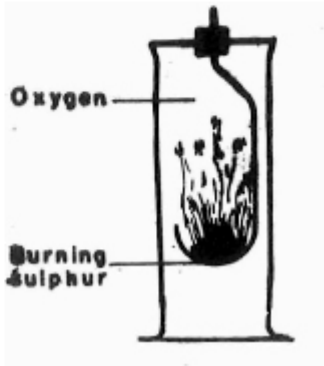

Figure 8.3 Lowering burning sulphur into a gas jar of oxygen

6. Pour water into the gas jar used in tests 4 and 5. Cover these with gas jar covers and shake well. Dip a piece of blue litmus paper into each of them and observe any colour change. The paper turns red in both cases showing that the oxides are acidic.
7. Lower a deflagrating spoon containing a piece of burning sodium into a gas jar of oxygen. It burns more rapidly. Repeat the experiment with magnesium ribbon. It also burns with a very bright flame.
8. Pour some water into the gas jars containing the burnt sodium and magnesium, shake well and test with red litmus paper. The paper turns blue showing that the oxides are alkaline.

From the above tests, it can be concluded that oxygen is a neutral gas and a very active element. It is a strong supporter of combustion, because burning materials burn more vigorously and with brighter flames, in oxygen. For example, the glowing splint bursts into flame when put into the gas jar of oxygen.

The tests also show that aqueous solutions of oxides of non-metals are acidic, while those of metals such as sodium and magnesium are alkaline.

Physical properties of oxygen

1. Oxygen is a colourless and odourless gas.
2. It is slightly less dense than air.
3. It is slightly soluble in water, hence aquatic animals like fish have oxygen for their respiration.

Chemical properties of oxygen

1. It is a strong supporter of combustion.
2. Non-metals burn readily in oxygen to form acidic oxides.

$C(s) + O_2(g) \rightarrow CO_2(g)$

$S(s) + O_2(g) \rightarrow SO_2(g)$

$4P(s) + 5O_2(g) \rightarrow 2P_2O_5(g)$

If these non-metallic oxides are dissolved in water, they form acidic solutions.

CO_2(g) + H_2O(l) â†’ H_2CO_3(aq)
SO_2(aq) + H_2O(l) â†’ H_2SO_3(aq)
$2P_2O_5$(g) + $6H_2O$(l) â†’ $4H_3PO_4$(aq)

3. Metals also burn readily in oxygen, forming basic oxides.

4Na(s) + O_2(g) â†’ $2Na_2O$(s)
2Mg(s) + O_2(g) â†’ 2MgO(s)

If the metallic oxide is soluble in water, it forms an alkaline solution of the metal hydroxide.

Na_2O(s) + H_2O(l) â†’ 2NaOH(aq)

In general, heated metals are readily oxidized by oxygen, the vigour of the reaction decreasing as we move down the activity series from potassium to copper. The noble metals such as mercury, silver and gold are not easily oxidized. Their oxides are not stable to heat, and readily yield oxygen on heating.

Test for oxygen

Oxygen may be identified as a colourless, odourless, slightly soluble gas that causes a glowing splint to burst into a flame. It also turns colourless nitrogen(II) oxide brown, with the formation of nitrogen(IV) oxide.

8.4 Industrial Preparation of Oxygen

In industry, oxygen is obtained from air because of its abundance in air. The cheapest method is by the fractional distillation of liquid air.

Air is first filtered to remove dust and then compressed to about 200 atmosphere pressure. The compression raises the temperature of the air. The hot air is cooled by passing it through water-cooled pipes. Carbon(IV) oxide and water vapour solidify under the high pressure and are removed. The cooled and compressed air is allowed to escape through a fine jet into an expansion chamber which is almost a vacuum. This rapid expansion through the fine jet cools the air greatly as energy is used in separating the gas molecules. The cooled air is used to further cool more incoming compressed, water-cooled air. Repeated compression and expansion of the air lowers its temperature till it is cool enough to liquefy.

The main constituents of liquid air, nitrogen and oxygen, are then separated by fractional distillation. Nitrogen distills off first at â€“196Â°C (77K), while oxygen boils off at -183Â° C (90K). The oxygen is sold as liquid oxygen or as the compressed gas in strong steel cylinders.

8.5 Allotropes of Oxygen

Oxygen exists as oxygen, O_2, and ozone, O_3. They are allotropes, being forms of the same element whose atoms have different molecular arrangements. Ozone is formed by passing oxygen through silent electrical discharges. Ozone is a poisonous gas with a characteristic smell. Both oxygen and ozune are oxidizing agents, but ozone is a more powerful oxidizing agent. Ozone, unlike oxygen, is able to oxidize potassium iodide to iodine in acidic solution.

$2I^-(aq) + 2H^+(aq) + O_3(g)$ â†’ $I_2(aq) + O_2(g) + H_2O(l)$

Ozone is present in the upper atmosphere. Its presence there is important because it filters out the sunâ€™s dangerous radiations. These radiations therefore do not reach the earth.

8.6 Hydrides of Oxygen

Oxygen forms two types of compounds (hydrides) with hydrogen.

These are: (i) water, H_2O, and (ii) hydrogen peroxide, H_2O_2. When hydrogen is burnt in air, water is formed.

$2H_2(g) + O_2(g)$ â†’ $2H_2O(l)$

Oxygen has an oxidation number of â€“2 in water. The chemistry of this very important compound will be treated in detail in Chapter 9. Hydrogen peroxide, unlike water, does not occur naturally. Oxygen exhibits an oxidation state of-1 in hydrogen peroxide, which is a covalent compound.

Experiment 8 5: Laboratory Preparation of Hydrogen Peroxide.

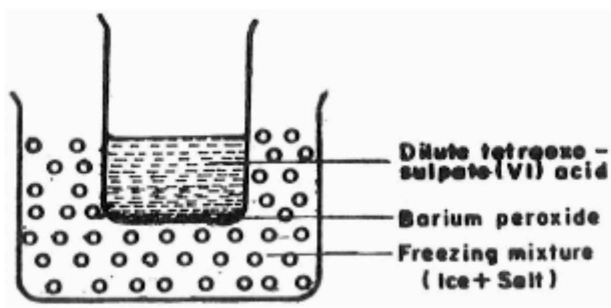

Figure 8.4 Preparation of hydrogen peroxide

Put a freezing mixture of ice and salt in a large beaker, and powdered barium peroxide in another beaker placed inside the freezing mixture (Figure 8.4). Add dilute tetraoxosulphate(VI) acid gradually, with stirring. A precipitate of barium tetraoxosulphate(VI) is formed. Filter off the precipitate. The filtrate is hydrogen peroxide.

$BaO^2(s) + H_2SO_4 (aq)$ â†’ $BaSO_4(s) + H_2O_2 (aq)$

Add a few drops of barium hydroxide solution to the filtrate until the solution is neutral. Put the aqueous solution in a stoppered bottle for the following tests.

1. Put some potassium iodide solution in a test-tube, and add hydrogen peroxide solution to it. What do you observe?
2. Put iron(II) chloride solution in a test-tube, and add hydrogen peroxide solution to it. What do you observe? Explain your observations.
3. Put some sodium trioxosulphate(IV) solution into a test-tube. Add hydrogen peroxide solution to it. Record your observations.
4. Add hydrogen peroxide solution to an acidified solution of potassium tetraoxomanganate(VII). What do you observe?

If you carry out the above reactions, you will observe that hydrogen peroxide behaves both as an oxidizing agent and as a reducing agent.

In test 1, violet vapour of iodine is given off. The iodine vapour can be tested using starch paper (Figure 8.5a).

$21^{-}(aq) + 2H^{+}(aq) + H_2O_2(aq)$ â†’ $2H_2O(l) + I_2(aq)$

The iodide ion is oxidised by electron loss and hydrogen peroxide is reduced.

$$2H^{+}(aq) + H_2O_2(aq) + 2e^{-} \rightarrow 2H_2O(l)$$

$$2I^{-}(aq) \rightarrow I_2(aq) + 2e^{-}$$

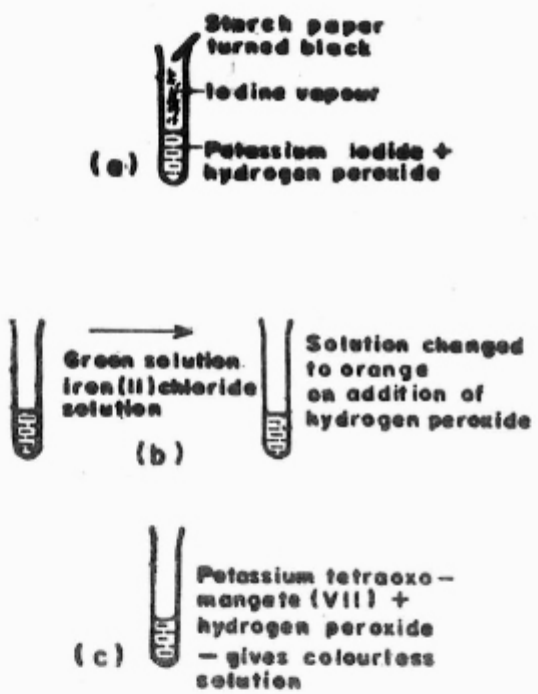

Figure 8.5 Reactions of hydrogen peroxide.

In test 2, iron(II) ion is oxidised to iron(III) ion, hence the colour of the solution changes from green to orange (Figure 8.5b).

$2Fe^{2+}(aq) + H_2O_2(aq) + 2H_+(aq)$ â†’ $2Fe^{3+}(aq) + 2H_2O(l)$

In test 3, sulphur in SO_3^{2-} is oxidised from the +4 to +6 state.

$SO_3^{2-}(aq) + H_2O_2(aq)$ â†’ $SO_4^{2-}(aq) + H_2O(l)$

In test 4, the purple colour of acidified potassium tetraoxomanganate(VII) solution is decolorised by hydrogen peroxide (Figure 8.5c). The hydrogenperoxide reduces manganese from the +7 state in MnO_4 ion to the +2 state in Mn^{2+}.

$2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq)$ â†’ $2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$

Hydrogen peroxide is normally used in the dilute aqueous solution form.

The pure compound is a syrupy liquid. When it is exposed to the atmosphere it decomposes to give water and oxygen.

$2H_2O_2(aq)$ â†’ $2H_2O(l) + O_2(g)$

It is also decomposed catalytically by manganese(IV) oxide, in the preparation of oxygen.

Hydrogen peroxide is sold in shops in â€œ10 volumeâ€ and â€œ20 volumeâ€ solutions. â€œ10 volumeâ€ hydrogen peroxide means that at s.t.p., 1 volume of hydrogen peroxide will produce 10 volumes of oxygen.

Uses of hydrogen peroxide

(i) Hydrogen peroxide is used for water purification.

(ii) It is used as a bleaching agent for the bleaching of pulp, cotton and other fibres.

(iii) Dilute hydrogen peroxide is used as a disinfectant for treating wounds.

(iv) Hydrogen peroxide solution in high concentration (about 90% and above) is used as a propellant in rockets and torpedoes.

8.7 Metallic and Non-metalic Oxides

When metals and non-metals are heated or burnt in oxygen, oxides are formed. Oxides are classified as acidic, basic, neutral, or amphoteric. Peroxides, super oxides, mixed oxides and other higher oxides also exist.

Acidic oxides

In a previous experiment, we burnt some non-metals in oxygen and found that the oxides produced, when dissolved in water, formed acidic solutions. In general, acidic oxides are the oxides of non-metals.

For example,

$SO_2(g) + H_2O(l)$ â†’ $H_2SO_3(aq)$
$H_2SO_3(aq)$ â†’ $2H^+(aq) + SO_3^{2-}$

Other acidic oxides are CO_2, SO_3, NO_2, SiO_2, P_4O_{10} and P_4O_6. Acidic oxides react with bases to form salts and water only.

Basic oxides

We have also observed that when a metal burns in oxygen, the oxide produced, if soluble in water, produces an alkaline solution. In general, oxides of most metals are basic oxides.

$CuO(s) + 2HCl(aq)$ â†’ $CuCl_2(aq) + H_2O(l)$
$Na_2O(s) + 2HNO_3(aq)$ â†’ $2NaNO_3(aq) + H_2O(l)$

Most metallic oxides are insoluble in water, but those of group I and some of group II metals, are soluble.

$Na_2O(s) + H_2O(l)$ â†’ $2NaOH(aq)$
$CaO(s) + H_2O(l)$ â†’ $Ca(OH)_2(aq)$

Neutral oxides

These are oxides of some non-metals which do not show basic or acidic properties. That is, they do not react with acidic or alkaline solutions. Examples are carbon(II) oxide, CO; nitrogen(I) oxide, N_2O; and nitrogen(H) oxide, NO.

Amphoteric oxides

These are oxides of some metals which react with both acids and bases to form salt and water only.

$ZnO(s) + 2HCl(aq)$ â†’ $ZnCl_2(aq) + H_2O(l)$
$ZnO(s) + 2NaOH(aq)$ â†’ $Na_2ZnO_2(aq) + H_2O(l)$

Other amphoteric oxides are Al_2O_3, BeO, and PbO.

Peroxides

Peroxides are metallic oxides of group I and heavier members of group II metals which contain the O_2^{2-} ion. In these compounds, oxygen has a combining power of â€“1. These peroxides react with acid or water to produce hydrogen peroxides.

$Na_2O_2(s) + 2H_2O(l)$ â†’ $2NaOH(aq) + H_2O_2(aq)$

Peroxides are powerful oxidizing agents.

Superoxides

These are formed by only the most electropositive metals in group I with the exception of lithium. They contain the superoxide ion, O_2^{2-}. They are very powerful oxidizing agents. They react with carbon(IV) oxide to form a trioxocarbonate(IV) and liberate oxygen, hence they are used in submarines for producing oxygen.

$4KO_2(s) + 2CO_2(g)$ â†’ $2K_2CO_3(aq) + 3O_2(g)$

8.8 Uses of Oxygen

1. It is used in hospitals, mountaineering, and underwater diving, to aid breathing.
2. It is used in the production of steel in the L-D process.
3. Liquid oxygen is used in rockets as a fuel.
4. It is used in oxy-acetylene flame for welding and cutting metals. The oxy-acetylene flame when properly adjusted, gives a temperature of more than 2000Â°C.

Plate 8.1 A welder using oxy-acetylene flame

Chapter Summary

1. Oxygen makes up 21% of the volume of air.
2. It is prepared in the laboratory by the thermal decomposition of some compounds rich in oxygen, notably hydrogen peroxide or potassium trioxochlorate(V). Manganese(IV) oxide acts as a catalyst for the decomposition.

3. The most important property of oxygen is that it is a strong supporter of combustion.

4. There are two allotropes of oxygen, the O_2 and O_3 molecules. O_3 is known as ozone and is poisonous.

5. Oxides of most metals are basic, while oxides of non-metals are usually acidic. Exceptions include oxides of zinc, aluminium and lead which are amphoteric, and the oxides of hydrogen, (water and hydrogen peroxide), carbon(II) oxide and nitrogen(II) oxide, which are neutral.

Assessment

1. (a) How is oxygen prepared in the laboratory?
   (b) Give two uses of oxygen.
   (c) What are the disadvantages of the presence of oxygen in air?
2. Describe the laboratory preparation of dry oxygen from hydrogen peroxide. How else can oxygen be prepared in the laboratory? Describe the test for oxygen.
3. (a) Ozone is an allotrope of oxygen. What do you understand by allotropy? What is the importance of ozone in the atmosphere?
   (b) Give two examples of each of the following:
      (i) acidic oxides
      (ii) basic oxides
      (iii) amphoteric oxides
      (iv) neutral oxides.
      Write equations to illustrate the typical reactions of each.
4. How is oxygen produced on a commercial scale?
   What are its uses? What would happen if the percentages of oxygen and nitrogen in the atmosphere were reversed?