UNIT 2 NON-METALS

Structure

2.1 INTRODUCTION

In the previous unit, you have seen that non-metals are located in the upper right hand corner of the periodic table. They have diverse appearance and exhibit various colours, e.g. sulphur is a yellow solid, nitrogen and oxygen are colourless gases while bromine is a brown liquid.

Non-metals are non-lustrous. They are brittle solids and are poor conductors of heat and electricity. The oxides of non-metals are acidic in nature.

Although metals are very important in our daily life about which you will study in the next unit, non-metals are also very important in their own right as life would not be possible at all without them. In this unit, you will study about some non-metallic elements. Atmosphere is a rich source for two such elements – *nitrogen* and *oxygen*. However, in the form of water – *hydrogen* and *oxygen* are abundant in oceans, seas and other water bodies.

In addition to these non-metallic gases, we would also discuss another important non-metallic element namely chlorine. The discussion would include occurrence, isolation, preparation, properties and uses of these elements and some of their important compounds such as ammonia, nitric acid and sulphuric acid.

Objectives

After studying this unit, you should be able to

- list some important characteristics of non-metals,
- state various zones of atmosphere,
- give various components of atmosphere,
- describe the occurrence of hydrogen and its isotopes,
- explain various methods of preparation of hydrogen,
- describe the properties and uses of hydrogen,

- discuss the occurrence and preparations of oxygen,
- explain important properties and uses of oxygen,
- discuss the important oxides and hydrides of nitrogen,
- explain the Haber process of synthesis of ammonia,
- describe the preparation, important reactions and uses of chlorine,
- discuss the preparation and properties of nitric acid and sulphuric acid, and
- give important uses of nitrogen, chlorine, ammonia, nitric acid and sulphuric acid.

2.2 ATMOSPHERE

The atmosphere which is a mixture of gases surrounding earth, has a very important role in the sustenance of life on earth. The air we breathe in mainly contains nitrogen and oxygen. However, small amounts of argon, carbon dioxide and water vapours are also present in it. The typical composition of air that we inhale and exhale is given in Table 2.1.

Table 2.1 : Composition of Inhaled and Exhaled Air

Sl. No.	Compound	Inhaled Air (%)	Exhaled Air (%)
1	Nitrogen	78	75
2	Oxygen	21	16
3	Carbon dioxide	0.03	4
4	Water	0	4
5	Argon	0.9	0.9

There are two interseting points to be noted here :

(i) the nitrogen and oxygen gases together constitute ~ 99% of the inhaled air, and

(ii) the exhaled air still contains 16% oxygen.

You can also see that there is significant amount of water vapours present in exhaled air. In fact, amount of water vapours present in atmosphere varies from place to place. Because of this variation, generally the composition of dry air is reported in the literature data. An average percentage of various gases present in dry air is given in Table 2.2.

As we move upwards, we see that the composition and physical characteristics of the atmosphere change with altitude. This results in various regions of atmosphere as shown in Figure 2.1.

You can see in Figure 2.1 that the region of atmosphere closest to the earth is called **troposphere**. All atmospheric phenomena such as clouds, rain, snow, wind etc. occur in this regions. As we move up, the temperatures drops to – 40^{o}C. The region next to troposphere is called **stratosphere** which is clear and cold region having temperature of about – 55^{o}C. The stratosphere contains the ozone layer.

Table 2.2 : Composition of Dry Air

Component	Average %
N_2	78.084
O_2	20.9476
Ar	0.934
CO_2	0.0314
Ne	0.001818
He	0.000524
CH_4	0.0002
Kr	0.000114
H_2	0.00005
Xe	0.0000087
N_2O	0.00005

Figure 2.1 : Regions of Atmosphere

Above the stratosphere, the temperature increases with altitude and then drops as we reach **mesosphere**. This is followed by a region called **ionosphere** where the gases present in the atmosphere are ionized by solar radiation.

SAQ 1

(a) List three important characteristics of non-metals.

(b) What is atmosphere?

(c) Name two important zones of atmosphere.

(d) Name the most abundant component of atmosphere.

2.3 HYDROGEN

Pure hydrogen was first isolated by Henry Cavendish. It was named **hydrogen** (from Greek *hydor* meaning water and *gennao* meaning to produce) by Lavoisier since it produced water on burning in air.

Hydrogen is the first element in the periodic table. It has a unique position in the periodic table. According to its electronic configuration, $1s^1$, it is generally placed above alkali metals but it is a non-metal.

Occurrence

Hydrogen is the most abundant element in the universe. It is present on the sun and other stars. On earth, hydrogen is present as a constituent of water, petroleum, cellulose, starch, fats and other materials. It constitutes 0.87% mass of the earth. Hydrogen is rare in the atmosphere as it escapes the gravitational attraction of the earth.

Hydrogen occurs as three different isotopes which are given in Table 2.3.

Table 2.3 : Isotopes of Hydrogen

Name	Symbol	Atomic Mass/amu	Natural Abundance
Hydrogen (protium)	${}^{1}_{1}H$ (H)	1.0078	99.985
Deuterium	${}^{2}_{1}H$ (D)	2.0141	0.0148
Tritium	${}^{3}_{1}H$ (T)	3.0160	Trace amount

Tritium is radioactive and emits β particles.

$${}^{3}_{1}H \rightarrow {}^{3}_{2}He + e \quad t_{1/2} = 12.4 \text{ yr}$$

It is produced in the upper atmosphere by the reaction of nitrogen with cosmic ray neutrons.

$${}^{14}_{7}N + {}^{1}_{0}n \rightarrow {}^{3}_{1}H + {}^{12}_{6}C$$

Tritium can be synthesized in nuclear reactors by neutron bombardment of ${}^{6}_{3}Li$.

$${}^{6}_{3}Li + {}^{1}_{0}n \rightarrow {}^{3}_{1}H + {}^{4}_{2}He$$

Preparation

(i) In the laboratory, small quantities of hydrogen can be prepared by the reaction of diluted acids such as hydrochloric acid or sulphuric acid on active metals.

$$Zn(s) + 2H^{+}(aq) \rightarrow Zn^{2+} + H_2(g)$$

It is collected by downward displacement of water as is shown in Figure 2.2.

Figure 2.2 : Apparatus for Preparation of Hydrogen

(ii) Hydrogen for commercial and laboratory use can also be prepared by electrolysis of aqueous solutions of sulphuric acid using platinum or nickel electrodes.

$$2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$$

(iii) The industrial method of preparation of hydrogen involves the reaction of hydrocarbons with steam at high temperature in the presence of nickel catalyst. The hydrocarbons can be obtained from natural gas or oil refinery.

$$CH_4(g) + H_2O \xrightarrow[1000^{o}C]{Ni} CO(g) + 3H_2(g)$$

The mixture of gases obtained above as products is also known as **synthesis gas**.

Hydrogen can also be prepared industrially by **water gas** reaction as follows :

$$C_{coke} + H_2O(g) \xrightarrow[400^{o}C]{Fe_2O_3} \underbrace{CO(g) + H_2(g)}_{\text{water gas}}$$

Water gas so obtained is used as an industrial fuel. Carbon monoxide obtained in both of the above reactions can be reacted with steam to yield additional amount of hydrogen.

$$CO(g) + H_2O(g) \xrightarrow[400^{o}C]{Fe_2O_3} CO_2(g) + H_2(g)$$

(iv) Hydrogen is also obtained as a by-product of **cracking** of hydrocarbons. The hydrocarbons of high molecular weight occurring in petroleum can be broken down to smaller ones and this is accompanied by formation of hydrogen as a side product.

Properties

Hydrogen is a colourless, odourless diatomic gas at ordinary conditions of temperature and pressure. It is non-polar in nature and hence, its melting point ($-260^{o}C$) and boiling point ($-253^{o}C$) are both low. It has very low density (0.070 g cm^{-3}).

Hydrogen can form both the cation (H^+) or the anion (H^-). The ionisation energy of hydrogen is 1310 kJ mol^{-1}. The H – H single bond is strong and has a bond dissociation energy of 436 kJ mol^{-1}.

Hydrogen forms strong covalent bonds with many elements. The strong covalent bond between hydrogen and oxygen (with a bond dissociation energy 464 kJ mol^{-1}) is responsible for the reducing property of hydrogen. It can reduce metal oxides to metals by abstracting oxygen from them.

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Hydrogen reacts with metals of Groups 1 and 2 to form hydrides.

$$2Na(s) + H_2(g) \rightarrow 2NaH(s)$$

$$Ca(s) + H_2(g) \rightarrow CaH_2(s)$$

On ignition in air, hydrogen forms water by a vigorous reaction.

$$2H_2O(g) + O_2(g) \rightarrow 2H_2O(l)$$

Air containing very small amounts of hydrogen is potentially explosive.

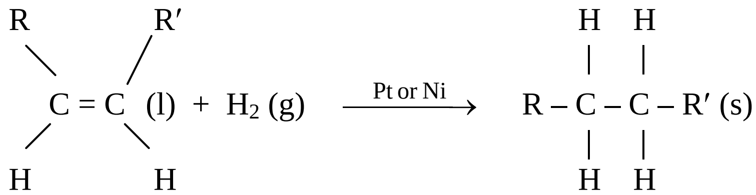
Uses

(i) The major industrial use of hydrogen is in the synthesis of ammonia by **Haber process**.

$$N_2(g) + 3H_2(g) \quad 2NH_3(g)$$

You will study Haber processes in detail later under the synthesis of ammonia.

(ii) Large quantities of hydrogen are used for catalytic hydrogenation of unsaturated liquid vegetable oils to produce saturated solid fats such as *ghee*.



where R and R′ are long hydrocarbon chains.

(iii) Hydrogen is used in the manufacture of many organic compounds particularly methanol.

$$2H_2(g) + CO(g) \xrightarrow{Co} CH_3OH(l)$$

(iv) Hydrogen is used in the oxy-hydrogen torch which has a flame temperature around 2500^{o}C.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

The oxy-hydrogen torch is used for cutting and welding metals.

(v) Hydrogen is used as reducing agent. In metallurgical operations, it is used for the reduction of oxides to pure metals.

(vi) Because of its low mass and very exothermic reaction with oxygen, liquid hydrogen is used as a fuel in large booster rockets.

It may become the fuel of the future as it has the highest **specific enthalpy** (enthalpy of combustion per gram) of any known fuel. It is a very clean fuel as it produces only water on combustion. The natural abundance of hydrogen is also very large. Hence, the challenge today is to obtain it cheaply from its natural sources.

(vii) Deuterium, the isotope of hydrogen, is used for isotopic labelling of many compounds in the study of the mechanism and kinetics of reactions.

(viii) Heavy water, D_2O, is used as moderator and coolant.

SAQ 2

(a) Write the isotopes of hydrogen.

(b) Which isotopes of hydrogen is radioactive?

(c) Explain one industrial method of preparation of hydrogen gas.

(d) What is synthesis gas?

(e) Complete the following reactions :

(i) $Ca + H_2 \longrightarrow ?$

(ii) $? + 3H_2 \longrightarrow NH_3$

(iii) $2H_2 + CO \xrightarrow{Co} ?$

(f) Give two uses of hydrogen.

2.4 OXYGEN

Oxygen was discovered by Joseph Priestley in 1774. Lavoisier named it **oxygen** meaning *acid former*.

Joseph Priestley

Occurrence

Oxygen is the most abundant element on earth and the third most abundant element in the universe. It constitutes 89% of water by mass which is present in oceans. Two-thirds of the mass of human body is oxygen. Most rocks and limestone contain oxygen in the combined form. Silicon dioxide (SiO_2) present in sand consists of more than 50% oxygen by mass.

Most of the oxygen present in the atmosphere is generated by **photosynthesis**. The overall reaction of photosynthesis can be represented as follows :

$$CO_2(g) + H_2O(g) \xrightarrow[\text{chlorophyll}]{\text{sunlight}} \text{carbohydrates} + O_2(g)$$

Preparation

(i) Oxygen can be obtained by fractional distillation of liquid air. The normal boiling point of oxygen is – 183°C whereas that of nitrogen is – 196°C. Hence, when liquefied air is warmed, nitrogen boils off first leasing oxygen containing a small amount of nitrogen and argon.

(ii) In the laboratory, oxygen can be obtained by thermal decomposition of potassium chlorate, $KClO_3$ with manganese dioxide, MnO_2 which acts as a catalyst.

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

(iii) Alternatively, oxygen can also be prepared by adding sodium peroxide, Na_2O_2 to water

$$2Na_2O_2(s) + 2H_2O(l) \longrightarrow 4NaOH(aq) + O_2(g)$$

(iv) Oxygen is also obtained by electrolysis of water.

$$2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$$

Properties

(i) Oxygen has two allotropic forms – O_2 and O_3 which are respectively known as **dioxygen** (or molecular oxygen) and **ozone**.

(ii) Dioxygen is a colourless, odourless and tasteless gas which condenses to a pale blue liquid at – 183°C and freezes to a solid at – 218°C.

(iii) It is slightly soluble in water and hence is very important for the survival of marine life.

(iv) It is paramagnetic in nature as it has two unpaired electrons.

(v) The electronic configuration of oxygen is $1s^2\ 2s^2\ 2p^4$. It forms oxides with many elements. The oxides contain O^{2-} as the anion. The electronic configuration of oxide ion is $1s^2\ 2s^2\ 2p^6$. The oxides formed by non-metals are **acidic** in nature e.g. sulphur dioxide, SO_2, is acidic. It dissolves in water to give sulphurous acid.

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

On the other hand, the oxides formed by metals are **basic** in nature. For example, barium oxide, BaO_2, is basic. It dissolves in water to give barium hydroxide.

$$BaO(s) + H_2O(l) \longrightarrow Ba(OH)_2(aq)$$

(vi) Hydrocarbons burn in oxygen to give carbon dioxide and water.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2(g)$$

A large amount of heat is liberated in such processes. Hence, hydrocarbons are used as fuels.

(vii) In oxy-acetylene torch, a mixture of acetylene and oxygen is burnt according to the following reaction :

$$\underset{\text{acetylene}}{2C_2H_2(g)} + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$$

The temperature of this torch is about 2400°C and it is used for welding purposes.

(viii) Some more reactive metals such as Na and K, form peroxide and superoxide, respectively with oxygen.

$$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$$

$$K(s) + O_2(g) \longrightarrow KO_2(s)$$

Peroxides contain O_2^{2-} as the anion. Hydrogen peroxide is an important compound. It is a colourless, syrupy liquid. It is a very strong oxidizing agent and is used as an antiseptic, hair bleach and an oxidizing agent.

Uses

(i) Oxygen is used in the blast furnace in the manufacture of steel.

(ii) It is used in oxy-hydrogen and oxy-acetylene torches.

(iii) It is used in hospitals and at high altitudes for facilitating breathing.

SAQ 3

(a) How can oxygen be prepared in the laboratory?

(b) Give the products of reaction of oxygen with

(i) Sodium

(ii) Potassium.

2.5 NITROGEN AND AMMONIA

(1) **Nitrogen**

Nitrogen was discovered by Scottish botanist Daniel Rutherford in 1772. Lavoisier named it **azote** meaning *lifeless*.

Occurrence

Nitrogen occurs in air as a major component (78%) but it is not abundant in the earth's crust. It occurs in the form of its compounds as KNO_3 (Saltpeter) in India and $NaNO_3$ (Chile Saltpeter) in Chile. Conversion of nitrogen to nitrogen compounds is called **nitrogen fixation**. One natural nitrogen fixation processes is **lightening** in which the nitrogen oxides are formed which later get converted to nitrates. Nitrogen fixation is also done by nitrogen fixing bacteria such as rhizobium which convert it into ammonia and other nitrogen containing compounds useful for plants. These bacteria are present in the root nodules of leguminous plants such as peas, beans, alfa alfa, etc.

Haber process which is used for the synthesis of ammonia is also a nitrogen fixation process.

Preparation

(i) Pure nitrogen is obtained by fractional distillation of air.

(ii) In the laboratory, nitrogen can be obtained by heating an aqueous solution of ammonium nitrite. Ammonium nitrite is prepared by the reaction of ammonium chloride and sodium nitrite. Ammonium nitrate has a potential explosive nature and hence must be handled with care.

(iii) Another method of laboratory preparation of nitrogen is by passing ammonia gas over hot copper (II) oxide.

$$2NH_3(g) + 3CuO(s) \xrightarrow{400^{o}C} N_2(g) + 3H_2O(g) + 3Cu(s)$$

Properties

Nitrogen is a colourless, odourless and tasteless gas. It is very unreactive because of the presence of a triple bond between two nitrogen atoms. Its bond dissociation energy is 941 kJ mol $^{-1}$.

Nitrogen reacts with magnesium and lithium to form respective nitrides.

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

$$6Li(s) + N_2(g) \longrightarrow 2Li_3N(s)$$

The electronic configuration of nitrogen is $1s^2\ 2s^2\ 2p^3$ and, therefore, it exhibits all oxidation states ranging from + 5 to – 3 (Table 2.4).

Table 2.4 : Compounds Showing various Oxidation States of Nitrogen

Oxidation State	Example
+ 5	HNO_3, NO_3^-, N_2O_5
+ 4	NO_2, N_2O_4
+ 3	HNO_2, NO_2^-
+ 2	NO (Nitrogen monoxide or nitric oxide)
+ 1	N_2O (nitrous oxide or dinitrogen monoxide)
0	N_2
– 1	NH_2OH (hydroxyl amine)
– 2	N_2H_4 (hydrazine)
– 3	NH_4^+, NH_3, NH_2^-

Table 2.4 shows that nitrogen forms a series of *oxides*. All these oxides are acidic in nature.

(i) **Dinitrogen monoxide, N_2O**

It is commonly called nitrous oxide, is prepared by heating ammonium nitrate.

$$NH_4NO_3(s) \xrightarrow{\Delta} N_2O(g) + 2H_2O(g)$$

Ammonium nitrate can explode and hence is to be handled carefully.

(ii) **Nitric oxide, NO**

It is prepared by catalytic oxidation of ammonia.

$$4NH_3(g) + 5O_2(g) \xrightarrow[Pt]{1000^oC} 4NO(g) + 6H_2O(g)$$

It contributes to the destruction of ozone layer by the following reactions.

$$NO + O_3 \longrightarrow NO_2 + O_2$$

$$NO_2 + O \longrightarrow NO + O_2$$

NO is getting regenerated in the above reactions and hence acts as a catalyst for the following change :

$$O_3 + O \longrightarrow 2O_2$$

NO gets readily oxidized to NO_2 on exposure to air.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

(iii) **Nitrogen dioxide**, $\mathbf{NO_2}$

It is a poisonous brown gas contributing the colour and odour to smog.

It exists in equilibrium with its dimer, N_2O_4.

$$2NO_2(g) \quad N_2O_4(g)$$

While NO and NO_2 are paramagnetic, N_2O_4 is diamagnetic in nature.

N_2O_4 is a colourless solid.

NO_2 disproportionates into nitric acid and nitrogen oxide.

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$$

(iv) **Dinitrogen pentaoxide**, $\mathbf{N_2O_5}$

It can be viewed as a mixture of NO_2^+ and NO_3^- ions. It exists in the gas phase but readily dissociates to nitrogen dioxide and oxygen.

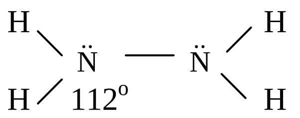
$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

Dinitrogen pentaoxide is the anhydride of nitric acid.

$$N_2O_5(s) + H_2O(l) \longrightarrow 2HNO_3(aq)$$

With hydrogen, nitrogen forms **hydrides** such as **ammonia** and **hydrazine**. Ammonia will be discussed a little later.

Hydrazine, N_2H_4 can be represented as follows :



Hydrazine is a colourless liquid. It is a powerful reducing agent.

$$N_2H_4(l) + O_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

It is used as a rocket propellant. Many hydrazine derivatives are also used as rocket fuels.

Two common **oxyacids** formed by nitrogen are **nitrous acid** and **nitric acid**.

Nitrous acid, HNO_2 is a weak acid. It is often prepared by the action of strong acids such as H_2SO_4 on nitrite salts such as $NaNO_2$.

It dissociates as follows :

$$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$$

Nitrous acid disproportionates into NO and HNO_3.

It forms yellow *nitrite* salts. These salts are prepared by bubbling equal moles of nitric oxide and nitrogen dioxide into appropriate aqueous metal hydroxide solution.

$$NO(g) + NO_2(g) + 2NaOH(aq) \longrightarrow 2NaNO_2(aq) + H_2O(l)$$

Sodium nitrite is used as a preservative for meat. The nitrite ion combines with the hemoglobin present in the meat to produce deep red colour. Nitrites also react with amines present in the body to produce nitrosamines which are carcinogenic in nature. Hence, their use in preservation is raising many questions.

Uses of Nitrogen

Low reactivity of nitrogen provides inert atmosphere during processing and packaging of foods. Liquid nitrogen is also used as a coolant for freezing foods.

Nitrogen is used for providing a non-oxidizing atmosphere in the electronics industry.

Nitrogen containing compounds are used in the synthesis of medicines, fertilizers, plastics and explosives.

(2) **Ammonia**

Small amounts of ammonia result from bacterial decomposition of organic matter in the absence of air. Ammonia is synthesized by **Haber process**. It was developed by the German chemist Fritz Haber in 1912. The reaction involved in this process can be represented as follows :

$$N_2(g) + 3H_2(g) \underset{\substack{500^{o}C,\\ \text{catalyst}}}{\overset{300\,\text{atm}}{\rightleftharpoons}} 2NH_3(g)$$

Fritz Haber **Carl Bosch**

Note that the process of synthesis of ammonia is carried out at high temperature and high pressure. The equipment required for carrying out the process under these conditions was developed by Karl Bosch and hence, Haber process is also referred to as **Haber-Bosch process**.

In the laboratory, ammonia is obtained by the reaction of ammonium salts and NaOH.

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow NH_3(g) + H_2O(l) + NaCl(aq)$$

Ammonia is a colourless gas with a pungent small. It is extremely soluble in water. It is a weak base and reacts with acids to produce salts.

$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(g)$$

Ammonia itself also acts as a solvent for many substances.

Ammonium nitrate (a compound of ammonia) is used as a fertilizer because of its high nitrogen content.

SAQ 4

(a) What do you understand by the term nitrogen fixation?

(b) Name the following oxides :

(i) N_2O

(ii) NO

(iii) NO_2

(iv) N_2O_5

(c) Give the structure of hydrazine.

(d) Write two important uses of hydrazine.

(e) Explain Haber process.

2.6 CHLORINE

Chlorine is one of group 7 (Group 17 in the new nomenclature) elements of the periodic table which are called *halogens*. It is a diatomic green gas. The name chlorine was derived from *chloros* meaning green.

Chlorine was discovered in 1774 by Carl Wilhelm Scheele by combining pyrolusite (MnO_2) with hydrochloric acid. He mistakenly thought that the gas produced in the experiment contained oxygen. But Sir Humphry Davy proved in 1810 that it was a distinct element and he named it as chlorine.

In nature, chlorine occurs in the combined form mainly as common salt (NaCl) in the form of *halite*. It also occurs as *sylvite* (KCl). In the form of anion, it also occurs in seawater.

Chlorine is obtained by the electrolysis of aqueous solution of sodium chloride or molten rock salt.

$$2NaCl(l) \xrightarrow{\text{electrolysis}} 2Na(l) + Cl_2(g)$$

A diaphragm cell is used for electrolysis of aqueous sodium chloride. You can see in Figure 2.3 that the diaphragm prevents the mixing of chlorine with hydrogen and sodium hydroxide. The cell contains compartments containing steel and titanium electrodes separated by porous diaphragms.

$$2Na^+(aq) + 2Cl^-(aq) + 2H_2O(l) \xrightarrow{\text{electrolysis}} 2Na^+(aq) + 2OH^-(aq) + Cl_2(g) + H_2(g)$$

Figure 2.3 : Diaphragm Cell

In the laboratory, chlorine can be prepared by heating a mixture of manganese dioxide and hydrochloric acid.

$$MnO_2(s) + 4H^+(aq) + 2Cl^-(aq) \longrightarrow Mn^{2+}(aq) + Cl_2(g) + 2H_2O(l)$$

The two main stable isotopes of chlorine having mass 35 and 37 are found in the relative proportions of 3 : 1, respectively. Hence, its atomic mass is 35.5 (more correctly 35.453).

Properties

Chlorine reacts with rarely all elements except, carbon, nitrogen, oxygen and noble gases.

It reacts with metals to form ionic chlorides :

$$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(g)$$

With non-metals, chlorine forms covalent chlorides.

$$2Sb(s) + 3Cl_2(g) \longrightarrow 2SbCl_3(s)$$

$$2Sb(s) + 5Cl_2(g) \longrightarrow 2SbCl_5(s)$$

Chlorine is a strong oxidizing agent and oxidizes metals to their higher oxidation states.

$$Cl_2\,(aq) + 2Fe^{2+}\,(aq) \longrightarrow 2Cl^-\,(aq) + 2Fe^{3+}\,(aq)$$

It also oxidizes Br^- ions to produce bromine.

$$2Br^-\,(aq) + Cl_2\,(g) \longrightarrow Br_2\,(l) + 2Cl^-\,(aq)$$

The oxidation of I^- ions to iodine also takes place by chlorine.

$$2I^-\,(aq) + Cl_2\,(g) \longrightarrow I_2\,(aq) + 2Cl^-\,(aq)$$

Chlorine reacts slowly with water to form relatively stable aqueous solutions of hydrochloric acid (HCl) and hypochlorous acid (HOCl).

Chlorine reacts with hydrogen to form hydrogen chloride which on forming aqueous solutions gives, hydrochloric acid.

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

The mixture explodes when exposed to light.

Uses

Chlorine is used as a germicide and as a disinfectant. It is added in small amounts to drinking water to keep it clean.

Chlorine also finds use in the manufacture of dyes, insecticides, paints, plastics, medicines, textiles, solvents and other products. PVC, Teflon, neoprene etc. are important polymers containing chlorine.

Chlorine forms other important compounds such as chloroform ($CHCl_3$), carbon tetrachloride (CCl_4), lithium chloride (LiCl) and chlorine dioxide (ClO_2). ClO_2 is used for bleaching paper pulp because it can oxidize pigments without affecting the wood fibres.

Some more compounds of chlorine and their uses are listed in Table 2.5.

Table 2.5 : Important Chlorine containing Compounds and their Uses

Compound	Their Uses
Hydrochloric acid (HCl)	Metallurgy, boiler scale removed, food processing, as a general acid
Calcium chloride ($CaCl_2$)	Drying agent, paper and pulp industry, fire proof fabrics, sizing and finishing of cotton fabrics
Potassium chlorate ($KClO_3$)	Oxidizing agent, matches, textile printing, bleaching agent
Calcium hypochlorite ($Ca(ClO)_2$)	Bactericide, deoderant, swimming pool disinfectant, fungicide, bleaching agent for paper and textiles
Sodium hypochlorite (NaClO)	Water purification, fungicide, laundry agent, bleaching agent, disinfectant

Many chlorinated hydrocarbons pose a serious health hazard to humans, fishes and birds. Their examples are DDT and polychlorinated biphenyls (PCBs). They are stable compounds and are not biodegraded.

SAQ 5

(a) How can chlorine be prepared in the laboratory?

(b) Name two important compounds of chlorine.

(c) Give important uses of calcium hypochlorite.

2.7 ACIDS

In this section, we would discuss two important acids, namely, nitric acid and sulphuric acid. Both these acids are very important industrially.

2.7.1 Nitric Acid

Nitric acid is produced by the **Ostwald process**. The first step of this process involves the oxidation of ammonia to nitric oxide.

$$4NH_3(g) + 5O_2(g) \xrightarrow[825^oC]{Pt} 4NO(g) + 6H_2O(g)$$

In the second step, the nitric oxide is reacted with oxygen to give nitrogen dioxide.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

In the final step, nitrogen dioxide is dissolved in water to yield nitric acid.

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$$

The gaseous NO is recycled for oxidation to NO_2. The aqueous nitric acid so obtained is 50% HNO_3 by mass which can be concentrated to 68% HNO_3 by distilling out water. It can be further concentrated to 95% HNO_3 by treatment with H_2SO_4 which further absorbs water from it.

Laboratory grade nitric acid has a density 1.42 g mL^{-1} and contains 70% HNO_3 by weight. Its concentration is 16 M.

Properties

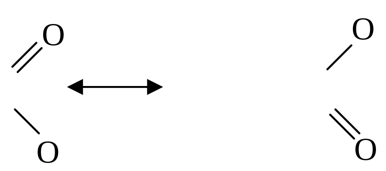
Nitric acid is a colourless fuming liquid having a boiling point of 83^{o}C. It has pungent order.

It turns yellow because it decomposes in sunlight by the following reaction;

$$4HNO_3(l) \xrightarrow{h\upsilon} 4NO_2(g) + 2H_2O(l) + O_2(g)$$

and the NO_2 produced gets dissolved in it.

The structure of nitric acid can be represented follows :



H – O – N H – O – N

It produces nitrate salts an reaction with oxides, hydroxides and carbonates of metals.

$$Ca(OH)_2(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$$

Nitric acid is a very strong oxidizing agent.

Uses

(i) It is used as an oxidizing agent.

(ii) It is used for the production of explosives such as trinitrotoluene (TNT), nitroglycerine and nitrocellulose (gun cotton).

(iii) It is used in the manufacture of ammonium nitrate which is used as a fertilizer.

(iv) It is also used in the production of plastics, dyes and drugs.

2.7.2 Sulphuric Acid

Sulphuric acid is commercially produced by **Contact process**. The process is so named because sulphur dioxide and oxygen react in *contact* with the surface of solid vanadium oxide (V_2O_5).

In the first step of this process, sulphur is burned in oxygen to produce sulphur dioxide.

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

The sulphur dioxide is then converted to sulphur trioxide in presence of vanadium pentaoxide catalyst.

$$2SO_2(g) + O_2(g) \xrightarrow{V_2O_5(s)} 2SO_3(g)$$

The sulphur trioxide so produced is absorbed into pure liquid sulphuric acid to form fuming sulphuric acid also known as **oleum** which can then be diluted to the desired strength.

$$SO_3(g) + H_2SO_4(l) \longrightarrow \underset{\text{oleum}}{H_2S_2O_7(l)}$$

$$H_2S_2O_7(l) + H_2O(l) \longrightarrow 2H_2SO_4(l)$$

Properties

It is a colourless, corrosive, syrupy liquid having melting point as 10°C and boiling point as 290°C. The conc. sulphuric acid in the laboratory is 18 M and is 98% H_2SO_4.

It is a powerful dehydrating agent. It is also a strong oxidizing agent and a strong acid.

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) \quad H^+(aq) + SO_4^{2-}(aq)$$

Considerable amount of heat is evolved when sulphuric acid is diluted with water. Hence, the dilution is always done by adding acid to water and not otherwise. This dissipates the evolved heat and also avoids the spattering of the acid.

Uses

(i) Sulphuric acid is used in the production of fertilizers and other industrial chemicals.

(ii) It is also used in the production of synthetic fibres, dyes and detergents.

(iii) It also finds use in metallurgical processes and storage batteries.

(iv) It is used in the manufacture of plastics and insecticides.

SAQ 6

(a) Describe the Ostwald process of synthesis of nitric acid.

(b) Which catalyst is used in contact process?

(c) Give important uses of sulphuric acid.

2.8 SUMMARY

In this unit, you have learnt that

- Non-metals have characteristics properties.
- The gases present in the atmosphere are non-metallic in nature and they can be isolated from the atmosphere.
- Hydrogen has three isotopes – *protium*, *deuterium* and *tritium*.
- Oxygen and nitrogen gases are very important for the existence of life on earth. You also studied about their occurrence, preparation, properties and important uses.
- Nitrogen forms a variety of oxides which have their characteristics properties and uses.
- Chlorine is an important non-metal. It reacts with many other elements and forms useful compounds.
- Ammonia which is a compound of nitrogen, can be synthesis by Haber process. Haber process is an important method of nitrogen fixation.
- Nitric acid and sulphuric acid are very important industrial chemicals. They have diverse uses.

2.9 ANSWERS TO SAQs

SAQ 1

(a) non-lustrous, brittle solids, poor conditions of heat and electricity.

(b) Mixture of gases surrounding the earth.

(c) Any two from troposphere, stratosphere, mesosphere, ionosphere.

(d) Nitrogen.

SAQ 2

(a) Protium, deuterium, tritium

(b) Tritium.

(c) (i) Reaction of hydrocarbons with steam at high temperature in presence of nickel catalyst, or

(ii) By water gas reaction.

(d) Mixture of CO(g) and $3H_2$(g)

(e) (i) CaH_2

(ii) N_2

(iii) CH_3OH

(f) Synthesis of ammonia, manufacture of ghee, oxy-hydrogen torch etc.

SAQ 3

(a) By thermal decomposition of potassium chlorate with manganese dioxide.

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

(b) (i) Na_2O_2 ; (ii) KO_2

SAQ 4

(a) Conversion of nitrogen to nitrogen compounds is called nitrogen fixation.

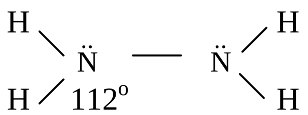
(b) (i) N_2O - nitrous oxide

(ii) NO - nitric oxide

(iii) NO_2 - nitrogen dioxide

(iv) N_2O_5 - dinitrogen pentaoxide

(c)



(d) Reducing agent and rocket propellant.

(e) $N_2 + 3H_2 \xrightarrow[500^{o}C,\ catalyst]{300\ atm} 2NH_3$

SAQ 5

(a) By heating a mixture of manganese dioxide and hydrochloric acid.

(b) Any two from HCl, $CaCl_2$, $KClO_3$, $Ca(ClO)_2$, NaClO.

(c) Bactericide, deoderant, disinfectant, fungicide etc.

SAQ 6

(a) (i) $4NH_3(g) + 5O_2(g) \xrightarrow[825^o]{Pt} 4NO(g) + 6H_2O(g)$

(ii) $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

(iii) $3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$

(b) V_2O_5

(c) Used in the production of fertilizers, dyes, storage batteries etc.