Chapter 9

THE CHEMISTRY OF SELECTED METALS AND NON-METALS

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LEARNING OBJECTIVES: By the end of this Chapter readers should be able to:

i. define Metals and Non-metals
ii. state the differences in physical and chemical properties between metals and non-metals
iii. show the distribution of metals, non-metals and metalloids in the periodic table
iv. describe how the following metals are extracted from their ores;
(a) Sodium (b) Aluminium
v. state five chemical properties each of (a) Sodium (b) Aluminium
vi. describe the chemical test for (a) sodium (b) aluminium
vii. explain how nitrogen can be obtained industrially
viii. show how chlorine can be prepared in the laboratory
ix. state the chemical properties of chlorine
x. outline the uses of chlorine.
xi. describe, using appropriate diagram the laboratory preparation of hydrogen
xii. state five chemical properties of hydrogen
xiii. outline the uses of hydrogen.

INTRODUCTION: The majority of the elements in the periodic table are metals, with about twenty (20) being non-metals. The metals are found to the left, and non-metals to the right. However, between these two extremes in the periodic table there are elements that display properties that have characteristics of both metals and non-metals. These are called **METALLOIDS** or **SEMIMETALS**. Elements of this type occur mainly in groups 3A and 5A and they tend to be **SEMI-CONDUCTORS -** that is they conduct electricity although not as well as metals. The semiconductors are used in making transistors and diodes which have revolutionised the electronic industry. Each of these main categories of elements-metals, non-metals and metalloids have their characteristic properties which shall be analytically x-rayed in this chapter.

WHAT ARE METALS AND NON-METALS?

Metals are elements whose atoms ionize by electron loss. Such elements can easily lose the valence electrons and have electron configuration resembling those of the nearest noble gas atom.

Na: $[Ne]\ 3s^1 \longrightarrow Na^+: [Ne] + e^-$
Mg: $[Ne]\ 3s^2 \longrightarrow Mg^{2+}: [Ne] + 2e^-$

This is particularly true of elements whose atoms have a small number of s and p electrons in the outermost shells and are commonly referred to as **METALS**.

On the other hand, non-metals are elements whose atoms ionize by electron gain. For instance, the halogens and the oxygen family can acquire a few electrons to have the electron configuration of the nearest noble gas atom.

CL: $[Ne]\ 3s^2 3p^5 + e^- \longrightarrow Cl^-: [Ar]$
S: $[Ne]\ 3s^2 3p^4 + 2e^- \longrightarrow S^{2-}: [Ar]$

This group of elements is referred to as **NON-METALS**. Between these two extremes-metals and non-metal, there are elements that exhibit the characteristics of both metals and non- metals. These elements are called **METALLOIDS**.

DIFFERENCES IN PHYSICAL AND CHEMICAL PROPERTIES BETWEEN METALS AND NON-METALS

Table 1: Table showing differences between metals and non-metals

	METALS	NON-METALS
i.	All metals except mercury are generally solids	Non-metals may be gases eg. Oxygen or liquid eg. Bromine, or solid eg. Carbon.
ii.	Metals have high melting-point and building point	Non-metals except carbon have low melting and boiling points
iii.	Metals are good conductors of heat and electricity eg aluminum	Non-metals except graphite are poor conductors of heat and electricity
iv.	Metals are malleable and ductile	Non-metals are brittle, ie cannot be hammered
v.	They possess characteristic lustre	They do not possess characteristic lustre except carbon
vi.	They exist as crystal lattices held by strong metallic bonds	Non-metals except diamond exist as covalent molecules held together by weak vander waals forces
vii.	Metals have positive ions	Non-metals have negative ions
viii.	Metals are good reducing agents	Non-metals are oxidizing agents
ix.	They have low electronegativities	They have high electronegativities

Activity 1.1: State eight differences between metals and non-metals

THE DISTRIBUTION OF METALS, NON-METALS AND METALLOIDS IN THE PERIODIC TABLE

An interesting feature at the boundary between the metals and the non-metals is the almost diagonal demarcation that separates metals from non-metals. The diagonal dividing line runs through boron, silicon and astatine (and includes germanium, antimony and polonium). On the basic of the definition we have adopted the elements to the left of this demarcation line are predominantly metallic. Among the elements in the demarcation line, boron and silicon are regarded as non-metals while the rest are metalloids.

The metals as we have already noted form basic oxides and display positive oxidation states in their compounds. They hardly combine chemically among themselves but they can form alloys. Non-metals combine with themselves, with other non-metals and with metals. In their binary compounds with metals, they display negative oxidation states but when they combine with other non-metals the oxidation state may be positive or negative depending on the relative negativities of the elements.

The distribution of metals, non-metals and metalloids is illustrated in the periodic table below. Metals found in groups I and II are called s-block metals. Especially, the metals to the bottom of Group I are the most powerful reducing agents. This is as a result of the outermost s electron being extremely well shielded from the nuclear charge. The electron is only weakly held to the atom, so it is easily lost. The Group I metals react with water to give strong alkalis; hence they are known as ***ALKALI METALS.*** The metals in Group II tend to make weaker alkalis, and are known as the ***ALKALINE EARTH METALS.***

The ***HALOGENS***, in Group VII, show the most complete set of properties of non- metals. Fluorine, at the top the Group is the most powerful oxidizing agent. It has a nuclear charge that is not very well shielded by the seven electrons it possesses. Fluorine will readily take an electron from a metal and make a fluoride ion, *F*⁻.

Going across a period there is a point at which the properties of the elements change from being primarily metallic to primarily non-metallic. Hence, we can safely say that metallic nature of the elements increases going down any group.

Table 2: The Periodic Table of the Elements:
The **DOUBLE LINE** marks the division between those elements normally regarded as metals (on the left) and those thought of as non-metals (on the right).

PERIOD \ GROUP	I	II											III	IV	V	VI	VII	O
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V.	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Os	Os	Ir	Pt	Au	Hg	Ti	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ad	Unq	Unp	Unh	Uno	Uno	Une	Uun	Uuu	Uub	Uut	Uuq	Unp	Uup	Uns	Uuo

metals and show close physical and chemical similarities among the members of each group.

PREPARATION OF METALS:
Because of their high reactivity the metals occur in nature as their compounds. Sodium and magnesium occur as their chlorides in sea water while calcium is mainly found as the carbonate and some other occur as silicates. Because of their large negative reduction potentials, the metals cannot be extracted from their ores by conventional reducing agents or by the electrolysis of their aqueous solutions. They are usually produced commercially by the electrolysis by the molten salt.

PHYSICAL PROPERTIES OF METALS
In both Groups the melting points, building points and hardness decrease with increasing atomic number. They are good conductors of heat and electricity but generally have low densities when compared to other metals like the transition metals. In comparison to the 1A metal ions, the IIA metal ions have considerably higher ratio of ionic charge to ionic radius. For this reason, the hydration energy of a given alkaline earth ion is about five times that of the alkali ion in the same period.

Table 3: Physical properties of the elements in group 1

	Lithium Li	Sodium Na	Potassium K	Rubidum Rb	Caesium Cs
Electron structure	[He]2s	[Ne] 3s	[Ar]4s	[Kr]5s	[Xe]6s
Electronegativity	1.0	0.9	0.8	0.8	0.7
I.E KJ/mol	520	513	419	400	380
Melting point/^{0}C	181	98	63	9	29
Boiling point/°C	1331	890	766	701	685
Atomic radius/pm	123	157	203	216	235
Ionic radius/pm	68	98	133	148	167

Principal oxid.no:	+1	+1	+1	+1	+1
$\in^{\theta}\ m^{+}$/m/V	-3.03	-2.71	-2.92	-2.93	-3.08

Source: (Matthews, 2006)

THE EXTRACTION OF REACTIVE METALS

We shall look at two processes. The first is for the extraction of sodium and the second is for the extraction of Aluminium.

[a] **THE DOWNS PROCESS FOR THE EXTRACTION OF SODIUM FROM SODIUM CHLORIDE**

The electrolysis is carried out in a DOWNS cell. The sodium is mixed with calcium chloride in a ratio of about 2:3. The melting point of the mixture is about 600°c (around 200°c lower than that of pure sodium chloride).

Sodium is discharged at the steel cathode, and chloride released at the graphite anode.

at the anode $2\ Cl^{-} - 2e^{-} \longrightarrow Cl_2$

at the cathode $Na^{+} + e^{-} \longrightarrow Na$

a large current is passed through the cell, but at a low voltage. This has the effect of both discharging the sodium effectively and heating the mixture so that it does not crystallize.

The sodium collects in inverted troughs above the cathode ring, and can be drawn off when necessary.

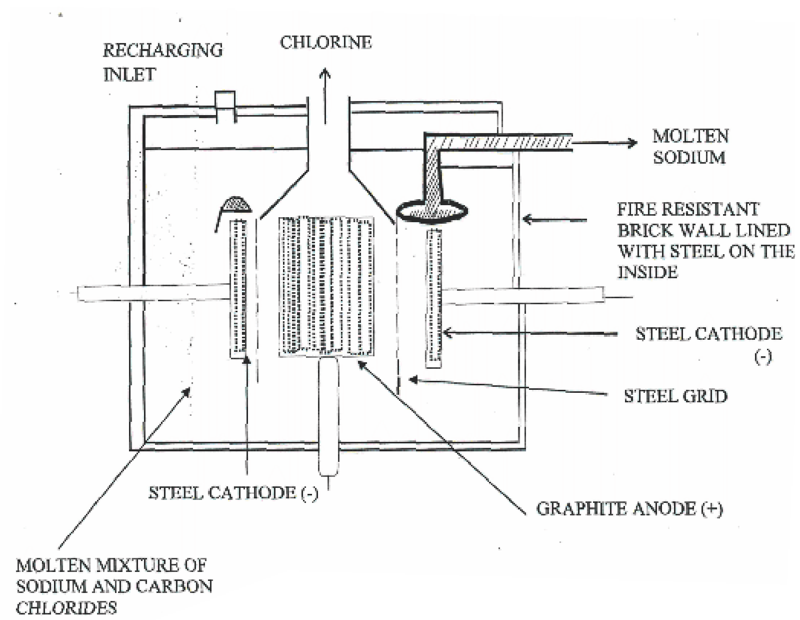

Fig. 2.1: Downs cell for extracting sodium from sodium chloride

Source: (Matthews, 2006)

PHYSICAL PROPERTIES OF SODIUM

i. Sodium is silvery soft solid with a metallic lustre.
ii. It has a melting point of 98°c and boiling point of 883°c.
iii. It is a good conductor of heat and electricity.
iv. It has a density of 0.97gcm^{-3}.

CHEMICAL PROPERTIES OF SODIUM

i. REACTION WITH AIR: Sodium metal tarnishes rapidly when exposed to air because it is readily oxidized by atmospheric oxygen to form sodium oxide.

$4Na_{(s)} + 0_{2(g)} \longrightarrow 2Na_20_{(s)}$

The oxide reacts with water vapour in the air to form sodium hydroxide which slowly absorbs atmospheric carbon (iv) oxide to form crystals of hydrated sodium trioxocarbonate IV.

$Na_2O_{(s)} + H_20_{(g)} \longrightarrow 2NaOH_{(aq)}$

$2NaOH_{(aq)} + CO_{2(g)} \longrightarrow Na_2CO_{3(s)} + H_20_{(l)}$

When heated in a plentiful supply of air sodium burns with a golden yellow flame to form sodium peroxide.

$2Na_{(s)} + 02(g) \longrightarrow Na_2O_{2(s)}$
excess

Thus, sodium is always stored under paraffin oil, toluene or naphthalene to prevent its oxidation by the atmospheric air.

ii. **REACTIONS WITH COLD WATER**: Sodium reacts vigorously with cold water, releasing a tremendous amount of heat: hydrogen and sodium hydroxide are produced.

$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$

iii. **REACTION WITH ACIDS:** Sodium reacts explosively with dilute acids to form hydrogen and salt. (The reaction is dangerous and should not be carried out in the laboratory)

$2Na_{(s)} + 2HCl_{(aq)} \longrightarrow 2NaCl(aq) + H2(g)$

v. **REACTION WITH AMMONIA:** Sodium reacts with ammonia to form sodamide and hydrogen

$2Na(s) + 2NH_{3(g)} \longrightarrow 2NaNH_{2(s)} + H_{2(g)}$

vi. **REACTION WITH NON-METALS:** On heating, sodium combines directly with most non-metals except boron, carbon and nitrogen.

$2Na_{(s)} + S_{(s)} \longrightarrow Na_2S_{(s)} + H_{2(g)}$
Sodium Sulphide

$3Na_{(s)} + P_{(s)} \longrightarrow Na_3P_{(s)}$
Trisodium Phosphide III

TEST FOR SODIUM IONS

FLAME TEST: Sodium compounds give a golden-yellow colour to a non-luminous flame. If the golden-yellow colour cannot be seen through a blue glass, the presence of sodium ions in the unknown compound is confirmed.

USES OF SODIUM

i. Sodium is used in manufacturing compounds such as sodium peroxide, sodamide and sodium cyanide.
ii. It is used in the production of tetra ethyl lead IV which serves as an anti-knock agent in petrol.
iii. It is used in sodium vapour lamps which give a bright-yellow light used for lighting highways, airports and streets.
iv. Sodium is used as a reducing agent in the extraction of titanium.
v. Liquid sodium is used as a coolant in nuclear reactors.
vi. Mixtures of sodium and ethanol or sodium amalgam and water are used as reducing agents in organic chemistry.

COMPOUNDS OF SODIUM

Important compounds of sodium include sodium oxide (Na_2O), sodium peroxide ($Na_2\ O_2$), sodium hydroxide (caustic soda) (NaOH), sodium chloride (common salt). Nacl, sodium tetraoxosulphate VI, sodium hydrogen trioxocarbonate IV NaHCO3.

ALUMINIUM

The major ore of Aluminium is BAUXITE, consisting of up to 60% Aluminium oxide, Al_2O_3, commonly known as ALUMINA. Aluminium is the most abundant metal and the third most plentiful element in the earth's crust. Other important sources of aluminium include ***CRYOLITE, KAOLIN, CORUNDUM and MICA.***

EXTRACTION OF ALUMINUM

Aluminum is extracted by the electrolysis of solution of aluminum oxide in molten cryolite. The pure aluminum oxide is obtained from the bauxite. The extraction proceeds in two stages.

STATE 1 – PURIFICATION OF BAUXITE: The bauxite is first heated with concentrated sodium hydroxide solution under pressure to form soluble aluminate III, NaAl $(OH)_4$

$Al_2 0_{3(s)} + 2NaOH_{aq} + 3H_2O \longrightarrow 2NaAL(OH)_{4(aq)}$

The solution is filtered. The filtrate is the sodium aluminate III while the residue contains the insoluble iron III oxide, trioxosilicates IV and other impurities which do not dissolve. The filtrate is then seeded with pure aluminium hydroxide crystals to induce the precipitation of aluminium hydroxide.

$NaAl(OH)_{4(s)} \longrightarrow Al(OH)_{3(aq)} + NaOH_{(aq)}$

The aluminium hydroxide precipitate is then filtered off, washed dried and heated strongly to yield pure aluminium oxide or alumina, Al_2O_3

$2Al(OH)_{3(s)} \longrightarrow Al_2O_{3(s)} + 3H_2O_{(1)}$

The sodium hydroxide is recovered and used again.

STAGE II-ELECTROLYSIS OF ALUMINA (REDUCTION): The electrolytic cell is a rectangular ion container lined with a graphite which serves as the cathode of the cell. The anodes consist of graphite rods dipping into the electrolyte which is solution of pure alumina in molten cryolite, Na_3AlF_6. The cryolite lowers the temperature to about 1000°C and this is maintained by the heating effect of a large current. During the process, oxygen is given off at the anodes. As a result, the anodes are slowly burnt away as carbon IV oxide has to be replaced. Molten (pure) aluminum is collected at the bottom of the cell and tapped off at intervals. The process is however expensive. CHEMISTRY OF THE REACTION: Alumina consists of aluminium and oxygen ions. AT THE CATHODE: The aluminium ions gain three electrons each to become deposited as metallic aluminium.

$2Al^{3+} + 6e^- \longrightarrow 2A1_{(s)}$ (REDUCTION)

AT THE ANODE: The oxygen ions donate two electrons each and become oxidized to atomic oxygen which then pair up to form gaseous oxygen molecules.

$O^{2-} \longrightarrow O + 2e^-$

$O + O \longrightarrow O_{2(g)}$

Or $2O_2^- \longrightarrow O_{2(g)} + 4e^-$

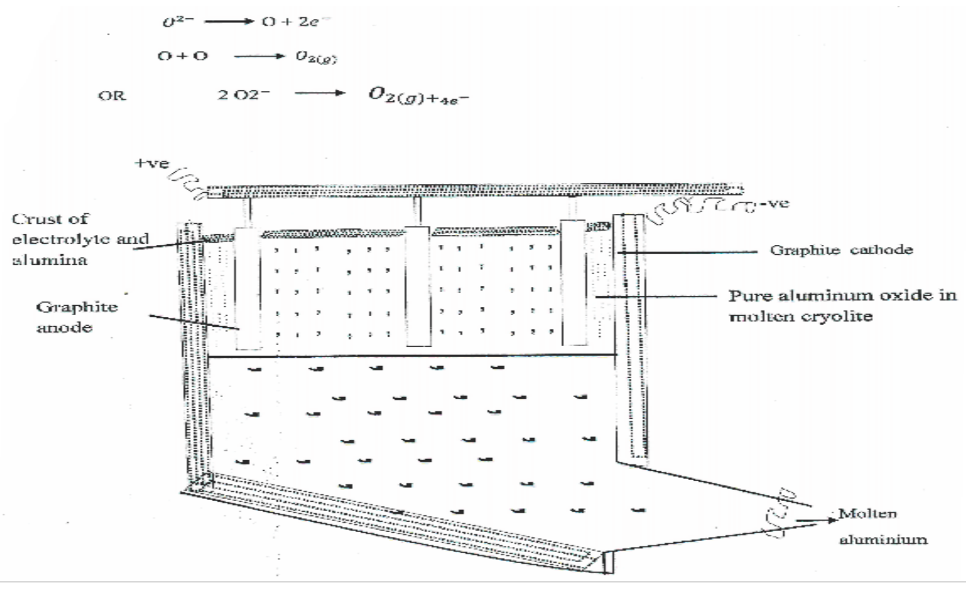

Fig. 4.1: Design of an electrolytic cell for the extraction of aluminium Source: (Odesina, 2012)

PHYSICAL PROPERTIES OF ALUMINIUM

1. Aluminium is a silvery-white metal
2. It is ductile and malleable, can be made into foil, wire and sheets
3. It is a good conductor of heat and electricity
4. It has a moderate tensile strength
5. It has a melting point of 660°C and boiling point of 2450°C.

CHEMICAL PROPERTIES OF ALUMINIUM

1. REACTION WITH AIR: When a piece of aluminium sheet is exposed to moist air, it acquires a thin, continuous coating of aluminium oxide, which prevents further attack of the metal by atmospheric oxygen, water or steam under normal conditions. Aluminium is thus, said to be corrosion free. Aluminium metal burns in air at a high temperature of about 800°C to form the oxide and nitride.

$4\ Al_{(s)} + 30_{2(g)} \longrightarrow 2AL_2O_3$

$2\ Al + N_{2(g)} \longrightarrow 2ALN()$

2. REACTION WITH NON-METALS: On heating, aluminium combines directly with the halogens, sulphur, nitrogen, phosphorus and carbon with the evolution of heat

$2Al_{(s)} + 3Cl_{2(g)} \longrightarrow 2AlCl_{3(s)}$

$2Al_{(s)} + N_{2(g)} \longrightarrow 2AlN_{(s)}$

3. REACTION WITH ACIDS: Aluminium reacts slowly with dilute hydrochloric acid, but more rapidly with the concentrated acid to displace hydrogen.

$2Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$

Aluminium does not react with dilute tetraoxosulphate VI acid, but hot concentrated H_2SO_4 attacks aluminium to liberate sulphur IV oxide.

$2Al_{(s)} + 6H_2SO_4 \longrightarrow Al_2(SO_4)_3(aq)^{+}3SO_{2(g)}$

Here, aluminium behaves as a reducing agent, reducing H_2SO_4 acid to sulphur IV oxide.. Aluminium does not react with HNO_3 acid at any concentration due to the formation of a protective layer of aluminium oxide

4. REACTION WITH ALKALI: Aluminium dissolves in both sodium and potassium hydroxide to form a soluble aluminate III, with the evolution of hydrogen gas.

$2Al(s) + 2NaOH(aq) + 6H_2O_{(L)} \longrightarrow 2NaAl(OH)_{4(aq)} + 3H_{2(g)}$

5. REACTION WITH IRON III OXIDE: Aluminium reduces iron III oxide to molten iron. This is the reaction that is employed during the termite process.

$2Al_{(s)} + Fe_2O_{3(s)} \longrightarrow Al_2O_{3(s)} + 2Fe_{(s)}$

TEST FOR ALUMINIUM IONS (Al^{3+})

Add few drops of sodium hydroxide solution to the unknown salt solution. The formation of a white gelatinous precipitate which dissolves in excess sodium hydroxide solution indicates the presence of aluminium ions.

$Al^{3+}_{(aq)} + NaOH_{(aq)} \longrightarrow Al(OH)_{3(s)} + 3Na^{+}_{(aq)}$

$Al(OH)_3 + \underset{excess}{NaOH(aq)} \longrightarrow NaAl(OH)_{4(aq)}$

As a confirmatory test, add few drops of aqueous ammonia to the unknown salt solution. The formation of a while gelatinous precipitate which is insoluble in excess aqueous ammonia indicates the presence of aluminium ions.

USES OF ALUMINIUM

1. Aluminium foils are used as packaging materials
2. Aluminium ions, Al" are used as coagulating agents in water treatment.

3. It is used in making overhead electric cables because it is light and has good electrical conductivity.
4. Aluminium is used in making cooking utensils.
5. Aluminium powder suspended in oil is used in paints and mirrors because of its high reflectivity.
6. Aluminium powder is used in the thermite process of machines and equipment.
7. It is used in making alloys such as ***DURALUMIN*** (Al, Cu, Mg, Mn); aluminium bronze (cu, Al) and magnallium (Al, mg) which are stronger and harder than pure aluminium but retain its properties of lightness and resistance to corrosion.

COMPOUNDS OF ALUMINIUM
Important compounds of aluminium inclued Aluminium oxide (Al_2O_3), Alumini hydroxide ($Al(OH)_3$, Aluminium chloride ($AlCl_3$) and Aluminium testraoxosulphate IV.

NON-METAL (NITROGEN)
A BRIEF CHEMISTRY OF NITROGEN
Free nitrogen exists in air as a diatonic molecule with triple covalent bonds between its atoms (N=N). The high bond energy of the triple bond makes the bond very stable and accounts for the unreactive nature of nitrogen under ordinary conditions. However, at very high temperature and pressure, nitrogen combines directly with certain non-metals and metals.

NITROGEN
Nitrogen makes up almost 78% of the atmosphere, and is therefore the most common gas. Nitrogen is chemically unreactive in nature. Free nitrogen in the atmosphere is important because it dilutes the oxygen to the point where combustion and oxidation of metals are reasonably slow

INDUSTRIAL PREPARATION OF NITROGEN
Nitrogen is obtained industrially by the fractional distillation of liquid air. Upon distillation, nitrogen, nitrogen gas is evolved first at-196°C at standard pressure and is separated from oxygen which boils at -183°C at standard pressure. The separated nitrogen is reliquefied and stored in steel cylinders and marked as liquid nitrogen.

PHYSICAL PROPERTIES OF NITROGEN
1. It is a colourless, odourless and tasteless gas.
2. Pure nitrogen is slightly less dense than air
3. Nitrogen has a melting point o -210C and a boiling point of -196°C.
4. It is only slightly soluble in water.
5. Nitrogen exhibits oxidation states varying from +3 to +5 in its various compounds.

CHEMICAL PROPERTIES
i. WITH NON-METALS: Nitrogen combines with hydrogen produce ammonia.
$N_{2(g)} + H_{2(g)} \rightleftharpoons 2NH_{3(g)}$
ii. WITH METALS: Nitrogen combines directly with some metals at high temperature to form nitrides. E.g;
$3Mg_{(s)} + N_{2(g)} \longrightarrow Mg_3N_{2(s)}$
Magnesium Nitride
The nitride hydrolyzes readily when warmed with water to give ammonia gas.
$Mg_3N_{2(s)} + 6H_2O_{(l)} \xrightarrow{\text{Warm}} 3Mg(OH)_{2(s)} + 2NH_{3(g)}$

USES OF NITROGEN
1. It is used in the manufacture of ammonia, cyanide and fertilizer.
2. Liquid nitrogen is used as a cooling agent

3. Nitrogen is used as a carrier gas in gas chromatography due to its inert nature

CHLORINE
Chlorine was first isolated by Scheele in 1774, but its nature as an element was established by Davy. Chlorine does not occur freely in nature, because it is very reactive but is usually found in the combined state as chlorides.

LABORATORY PREPARATION
Chlorine is prepared in the lab by the oxidation of concentrated HCI acid with a strong oxidizing agent, such as MnO_2 or $KMnO_4$

INDUSTRIAL PREPARATION
Chlorine is prepared industrially by the electrolysis of.

1. Brine
2. The chlorides of molten sodium, magnesium or calcium. A specially designed cell, developed by Castrer, Kellener and Solvay is used in the process. The chlorine gas is then liquefied and stored under pressure in steel cylinders.

PHYSICAL PROPERTIES
1. Chlorine is a greenish-yellow gas with choking, unpleasant, irritating smell.
2. It is moderately soluble in water
3. It is denser than air
4. It is poisonous gas
5. It can easily be liquefied under pressure of about

CHEMICAL PROPERTIES
1. **DISPLACEMENT OF OTHER HALOGENS:** With the exception of fluorine. chlorine displaces other halogens from solutions of their acids and salts.
   $Cl_{2(g)} + 2HI_{(ag)} \longrightarrow 2HCl_{(ag)} + I_{2(g)}$
2. **COMBINATION WITH OTHER ELEMENTS**
   Chlorine combines directly with most other elements to form chlorides. Eg.
   $2Na + Cl_2 \longrightarrow 2NaCl$
3. **AS AN OXIDIZING AGENT:** Chlorine is a powerful oxidizing agent because of its ability to remove hydrogen and readiness to accept electrons from reducing agents to form chlorides. Eg. With iron II salts:
   $2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$
4. **REACTION WITH HYDROGEN:** Chlorine has a very strong affinity for hydrogen. It readily removes hydrogen from its compounds to form hydrogen chloride.
   $C_{10}H_{16(g)} + 8C1_{2(g)} \longrightarrow 10C_{(s)} + 16HCl_{(g)}$
5. **AS A BLEACHING AGENT:** Chlorine is a common bleaching agent. It bleaches most dyes and inks in the presence of water, except those containing carbon.

USES OF CHLORINE
1. It is used as a bleaching agent for cotton, linen and wood pulp.
2. It is used as a germicide- for sterilization.
3. It is used for the manufacture of important organic compounds like CHCI), CCI. PVC, KCIO), HCl etc.
4. Chlorine is used as in aerosol propellants.

HYDROGEN AND HYDRIDES

Although air contains almost no tree hydrogen, hydrogen atoms are very abundant in nature. One reason for this is that they are found in water, which is widely distributed over the earth; another is that, combined with carbon, they are found in all living matter. Hydrogen is also found in space. It plays a crucial role in stars (such as our sun) in nuclear fusion reactions.

Hydrogen has three isotopes: $^{1}_{1}\mathrm{H}$ (Protium, H) $^{2}_{1}\mathrm{H}$ (deueterium, D) and $^{3}_{1}\mathrm{H}$ (tritium, T). Deuterium occurs naturally but only to the extent of 0.015%. Tritium can be made by causing the nuclear reaction:

$$^{1}_{1}\mathrm{H}+^{2}_{1}\mathrm{H} \longrightarrow ^{3}_{1}\mathrm{H}+^{1}_{1}\mathrm{H}$$

Tritium is radioactive, decaying by beta emission.

LARGE SCALE EXTRACTION OF HYDROGEN

One of the vital uses of hydrogen is in the manufacture of ammonia by the Haber process. A very important source of Hydrogen for the process is natural gas, methane. The key idea is to react methane with steam at a high pressure (35 atm) and temperature (800°C) in the presence of a catalyst (nickel). A mixture of carbon II oxide (Co) and carbon (iv) oxide (CO_2) and hydrogen results:

$$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{2(g)} + 3H_{2(g)}$$

Synthesis gas

$$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{2(g)} + H_{2(g)}$$

The carbon II oxide and carbon (iv) oxide are removed, leaving the hydrogen for further reaction.

Hydrogen can also be obtained from the oil refining industry. It is made in many reactions that involve cracking long-chain hydrocarbons into smaller fractions, example

$$C_6H_{12(g)} \longrightarrow C_6H_{12(g)} + 3H_{2(g)}$$

It is possible to make hydrogen by the BOSCH REACTION. The process takes place in three stages:

(i) Stage 1: Steam is passed over white hot coke:
$H_2O_{(g)} + C_{(s)} \longrightarrow CO_{(g)} + H_{2(g)}$
Water gas

(ii) Stage 2: The mixture of carbon II oxide and hydrogen, know as WATER GAS, is mixed with more steam and passed over an iron catalyst. Only the carbon II oxide reacts:
$H_2O_{(g)} + CO_{(g)}\ CO_{2(g)} + H_{2(g)}$

(iii) Stage 3: Under pressure the carbon IV oxide is dissolved in water, thus leaving the hydrogen available for further use.

Another method of obtaining hydrogen commercially is ELECTROLYSIS. Hydrogen is obtained as a by-product in the electrolysis of brine. When a cheap source of electricity is available, hydrogen is also obtained by the electrolysis of acidified water.

LABORATORY PREPARATION OF HYDROGEN

The three method available for the preparation of hydrogen in the laboratory are as follows:

1. **Action of dilute acid on zinc:** Dilute hydrochloric or tetraoxosulphate (vi) acid reacts with granulated metallic zinc to liberate hydrogen gas

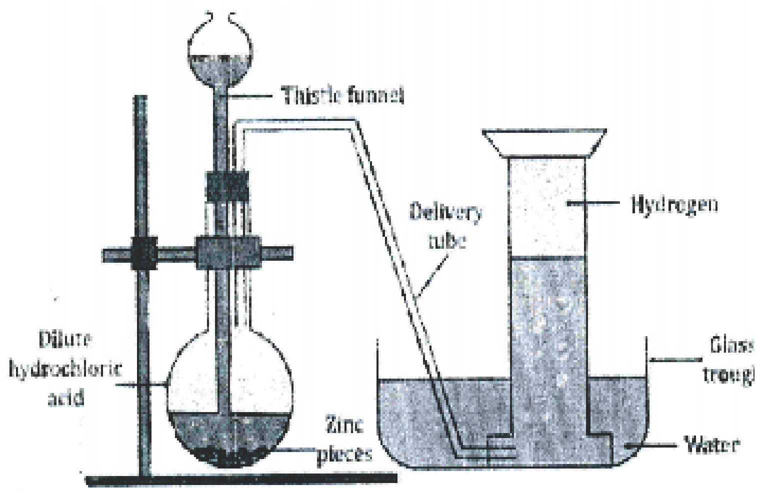

Fig. 7.1: Preparing hydrogen by the action of a dilute acid on zinc
Source: (Odesina, 2012)

The hydrogen gas is given off in a gas-jar over water. If the gas is required dry, it is passed through fused calcium chloride or concentrated tetraoxosulphate (vi) acid and collected by downward displacement of air, since it is lighter than air.

N/B: Dilute trioxonitrate (v) acid is not used for the preparation of hydrogen because it is a strong oxidizing acid and will produce hydrogen only with magnesium

2. **ACTION OF COLD WATER ON ACTIVE METALS:** Sodium and potassium react rapidly with cold water, liberating hydrogen. This reaction is very vigorous and should be carried out with extreme care using only a small piece of the metal.

   $2Na_{(s)} + 2H_2O_{(1)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$

   $2K_{(s)} + 2H_2O_{(1)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$

3. **ACTION OF STEAM ON TRON**: Iron at red beat liberate hydrogen from stream.
   Tri Iron-tetraoxide, Fe_3O_4, is formed at the same time

   $3Fe_{(s)} + 4H_2O_{(1)} \rightleftharpoons 3FeO_{4(s)} + 4H_{2(g)}$

PHYSICAL PROPERTIES OF HYDROGEN

(i)Pure hydrogen is a colourless, odourless and tasteless gas
(ii) It is neutral to litmus paper.
(iii) It is the lightest known substance, about 14.4 times less dense than air
(iv) It has a very low boiling point of-253°C
(v) It is relatively insoluble in water.

CHEMICAL PROPERTIES OF HYDROGEN

- **COMBINATION REACTIONS**
  - **With metals:** Hydrogen combines directly with the more reactive metals to form ionic hydrides.

    $2Na_{(s)} + H_{2(g)} \longrightarrow 2NaH_{(s)}$

    $Ca_{(s)} + H_{2(g)} \longrightarrow CaH_{2(s)}$
  - **With oxygen**: Pure hydrogen burns with a pale blue flame as it combines with oxygen to produce steam.

    $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$

- **With Halogens**: Hydrogen combines directly with the halogens to produce halides.
  $H_{2(g)} + Cl_2 \longrightarrow 2HCl_{(g)}$
  $H_{2(g)} + Br_{2(g)} \longrightarrow 2HBr_{(g)}$

- **With chlorine**, the reaction is spontaneous in bright sunlight but slower in diffused light. However, the reaction cannot take place in darkness.

- **With nitrogen**: Hydrogen combines directly with nitrogen to produce ammonia

  $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

- **REDUCING ACTION**
  Hydrogen is a strong reducing agent. When passed through heated oxides of copper. lead, iron, and zinc, it reduces the oxides to their metals.
  CuO(s) + H2(g) ⟶ Cu(s) + H2O(g)
  PbO(s) + H2(g) ⟶ Pb(s) + H2O(g)
  Fe2O3(s) + 3H2(g) ⟶ 2Fe(s) + 3H2(g)
  ZnO(s) + H2(g) ⟶ Zn(s) + H2O(g)

USES OF HYDROGEN

i. Hydrogen is used in the synthesis of ammonia
ii. It is used in hardening vegetable and animal oil for the manufacture of margarine, soap and candles.
iii. Liquid hydrogen is used as a rocket fuel
iv. It is used in oxy-hydrogen flames to produce high temperatures (over 2000°C) that can melt metals
v. Hydrogen is used in atomic hydrogen flames
vi. It is used for filling balloons
vii. Hydrogen can be used to inflate airships, but it is highly inflammable

TEST FOR HYDROGEN

Hydrogen gas burns with a *pop* sound when a lighted splinter is inserted into a test - tube containing the gas.

HYDRIDES

Active metals like sodium, potassium and calcium form ionic hydrides with hydrogen. These hydrides share some properties with other ionic compounds That is, they are crystalline solids with high melting and boiling points. They conduct electricity when molten and, react readily with water to form hydroxides and liberate hydrogen gas.

$$CaH_{(s)} + 2H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + 2H_{2(g)}$$

Aluminum and boron form complex covalent hydrides which are important reducing agent especially is organic chemistry, e.g. Lithium tetrahydridoaluminate III, $LiAIH_4$ and Sodium tetrabydrido borate III, $NaBH_4$ are the common complex hydrides.

Non-metallic elements like chlorine and nitrogen form simple covalent hydrides. These are volatile compounds that are gaseous at room temperature. The hydrides of fluorine, HF and oxygen, H_2O are exceptions because they are liquids due to the presence of hydrogen bonding. The hydrides of the more electronegative elements like chlorine and sulphur form acidic solutions when dissolved in water.

Students' Activity

(a) What are ALLOYS

(b) Mention eight (8) examples of alloys, stating their constituents and their uses.

SUMMARY

Metals are elements whose atoms ionize by electron loss while non-metals are elements whose atoms ionize by electron gain. The principles of metal extraction include ***ELECTROLYTIC PROCESS, CHEMICAL REDUCTION and THERMAL REDUCTION***. Most metals and non-metals occur freely in nature, while others occur in combined states with other elements.

REVISION QUESTIONS

1. Define metals and non-metals
2. Outline, in a tabular form, the differences in physical and chemical properties between metals and non-metals
3. Show the distribution of metals, non-metals and metalloids in the periodic table
4. Describe how the following metals are extracted from their ores.
   (a) Sodium (b) Aluminum
5. State, five chemical properties each of (a) sodium (b) Aluminum
6. Describe the chemical test for (a) sodium (b) aluminum
7. Explain how nitrogen can be obtained industrially
8. With the help of diagram, describe how hydrogen gas can be prepared in the laboratory.
9. State five chemical properties of hydrogen gas.
10. Outline five uses of hydrogen gas
11. Show how chlorine can be prepared in the laboratory
12. State the chemical properties of chlorine
13. Outline the uses of chlorine

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