FIRST TERM SS3 CHEMISTRY LESSON NOTE

SS3 CHEMISTRY SCHEME OF WORK

WEEK	TOPIC	SPECIFIC TOPIC
1	Metals and their compounds	• Definition and classes of metals • General properties of alkali metals • Principles of extraction of metals • Extraction of sodium, properties and uses of sodium.
2	Metals and their compounds	• General properties of alkaline-earth metals • Extraction of calcium, properties of calcium. • Extraction of aluminum, uses and properties of calcium and aluminum • Compounds of calcium and aluminum
3	Metals and their compounds	• Extraction of Tin • Properties of Tin and uses • Compounds of Tin and Tin metals
4	Metals and their compounds	• Extraction of Copper • Different Methods of extraction of copper and uses • Compounds of copper and Test Analysis of copper.
5	Metals and their compounds	• Extraction of Iron • Properties and uses of Iron • Rusting of Iron and reactivities • Method of preparation.
6	Fats and Oils	• Sources and structure of fats and oils. • Saponification and Sulphonation, properties, reactions and uses of fats and oils.
7	Soaps and detergent	• Properties of soap, action of soap as emulsifying agent • Properties of detergent

		• Differences between soap and detergent • Soap and detergent: properties; physical and chemical properties
8	Giant molecules	• Sources and classification of sugars. • Hydrolysis of sucrose and starch • Test for starch, sugars and glucose • Sources, structure and test for proteins.
9	Revision	
10−12	Examination/ vacation	

METALS AND ITS COMPOUNDS 1

INTRODUCTION

A metal is an element whose atoms ionize by electron loss, while a non-metal is an element whose atoms ionize by electron gain.

DIFFERENCES IN THE PHYSICAL PROPERTIES OF METALS AND NON-METALS

s/n	**Metals**	**Non-metals**
i.	They are good conductors of heat and electricity (e.g. Sodium, silver and aluminum).	They are poor conductors of heat and electricity except graphite.
ii.	Generally solid (except mercury) never gaseous	May be a solid, liquid or gas.
iii.	High melting points(except Na and K)	Low melting points(except C and Si)
iv	Malleable and ductile e.g. iron and copper	Non-malleable and brittle e.g. Sulphur

DIFFERENCES IN THE CHEMICAL PROPERTIES OF METALS AND NON-METALS

s/n	Metals	Non-metals
i.	Their ions migrate to the cathode during electrolysis; hence they form cations.	Their ions migrate to the anode during electrolysis; hence they form anions
ii.	They react with oxygen to form basic oxides.	They react with oxygen to form acidic oxides.
iii.	Water-soluble oxide forms alkaline solution.	Water-soluble oxide forms acidic solutions.
iv.	Form electrovalent chlorides that have high melting points.	Form covalent chlorides that are covalent and volatile.
v.	Do not react with one another but form alloys.	React with one another but do not form alloys.

GENERAL PROPERTIES OF METALS

PHYSICAL PROPERTIES OF METALS

Physical properties of elements are dependent on:

(1). The arrangement of their atoms or molecules in crystal lattices when in the solid state and

(2). The bonds that bind the atoms or molecules in the solid, liquid or gaseous state.

The general characteristic physical properties of metals are as follows:

1. Metals are elements that are solid at room temperature (except mercury, which is a dense liquid).
2. Metals have high melting points (except sodium, melting point 97℃ and potassium, melting point, 63℃).
3. They have characteristic metallic lustre i.e. *reflect light or shine when pure*.
4. They are sonorous i.e. *gives off a note when hit*. (except sodium and potassium).
5. They are good conductor of heat and electricity. *Silver is the best conductor of heat and electricity, followed by copper, then, aluminum*.
6. They have high densities (except magnesium and aluminum, *which are light*; *sodium and potassium are so soft that they can be cut by a knife*).
7. Metals are hard, strong and have high tensile strength.
8. Some metals are malleable, i.e. *can be hammered into thin sheets*.
9. Some metals are ductile, i.e. *can be drawn into thin wires*.

QUESTION 1: Explain why metals are good conductor of electricity. **Answer** = *Atoms of metals have mobile valence electrons, which are able to carry electricity from one of a metal to the other; hence, metals are good conductors.*

CHEMICAL PROPERTIES OF METALS

Chemical properties of elements are dependent on the number of valence electrons present in the atomic structure. An important property that is determined by valence electrons is the tendency of the atoms to an element to ionize.

The general characteristic chemical properties of metals are as follows:

1. When metals react, they ionize by loss of electrons to form cations, such as Na^+, Ca^{2+} and Al^{3+}; hence they are reducing agents.

$$Na \rightarrow Na^+ + e^-$$
$$Ca \rightarrow Ca^{2+} + 2e^-$$
$$Al \rightarrow Al^{3+} + 3e^-$$

2. A metal reacts with oxygen to form one or more oxides. Examples: Na_2O, MgO, FeO, Fe_2O_3, CuO and Cu_2O.
3. Metals that are above hydrogen in the activity series react with dilute mineral acids, such as *hydrochloric acid,* HCl and *tetraoxosulphate(*VI*) acid,* H_2SO_4 to liberate hydrogen gas.

$$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$$
$$Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$$
$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$
$$2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_{2(g)}$$

4. They form electrovalent(ionic) compounds with non-metals, such as chlorine.

$$2Na_{(s)} + Cl_{2(aq)} \rightarrow 2NaCl_{(s)}$$
$$2K_{(s)} + Cl_{2(aq)} \rightarrow 2KCl_{(s)}$$
$$Ca_{(s)} + Cl_{2(aq)} \rightarrow CaCl_{2(s)}$$

ACTIVITY SERIES OF METALS

The reactivity of a metal depends on the ease with which its atoms lose electrons; the ease varies from one metal to another.

The activity series is such that a metal higher up in the series will displace the ions of metals below it in the series.

CONSULT YOUR CHEMISTRY TEXTBOOK AND SEE THE ELECTROCHEMICAL SERIES

REACTIONS OF METALS AND THEIR COMPOUNDS

The relative position of the metals on the reactivity series determines the reactivities of metals and their compounds and methods of extraction.

1. Reaction of Metals with oxygen or air:

K Na Ca	*They burn vigorously in air to form oxides*. $4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$ $2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$
Mg	*Burns in air, when heated, to produce the oxide and nitride;* $2Mg_{(s)} + O_{(2g)} \rightarrow 2MgO_{(s)}$ $3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$
Al Zn Fe Sn	*They burn in air when strongly heated to form oxides*. $4Al_{(s)} + 3O_{(2g)} \rightarrow 2Al_2O_{3(s)}$ $3Fe_{(s)} + 2O_{2(g)} \rightarrow Fe_3O_{4(s)}$ $Sn_{(s)} + O_{2(aq)} \rightarrow SnO_{2(aq)}$
Pb Cu Hg	*They form oxides when strongly heated in air, but do not burn in air*. $2Hg_{(l)} + O_{2(g)} \rightarrow 2HgO_{(s)}$ $2Pb_{(s)} + O_{2(g)} \rightarrow 2PbO_{(s)}$ $3Pb_{(s)} + 2O_{2(g)} \rightarrow Pb_3O_{4(s)}$
Ag Pt Au	*They do not burn or form oxides when heated in air*.

2. Reaction of Metals with water or steam:

K Na Ca	*They react vigorously with cold water with decreasing reactivity to liberate hydrogen gas*. $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$
Mg Al Zn Fe Sn	*They react vigorously with **steam** to liberate hydrogen gas*. $2Zn_{(s)} + H_2O_{(g)} \rightarrow 2ZnO_{(aq)} + H_{2(g)}$ $3Fe_{(s)} + 4H_2O_{(g)} \leftrightarrow Fe_3O_{4(aq)} + 4H_{2(g)}$
Pb Cu Hg	*They do not react with **water** or **steam**; hence, do not displace hydrogen from water or steam.*

Ag Pt Au	

3. Reaction of Metals with acids:

K Na Ca	*They react explosively with dilute acids to liberate hydrogen gas*. $2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_{2(g)}$ $2K_{(s)} + 2HCl_{(aq)} \rightarrow 2KCl_{(aq)} + H_{2(g)}$ $Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$
Mg Al Zn Fe Sn Pb	*They all displace hydrogen from dilute acids with decreasing reactivity to liberate hydrogen gas*. $Zn_{(s)} + H_2SO_{4(l)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$ $Fe_{(s)} + H_2SO_{4(l)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$ $2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$ $Sn_{(s)} + 2HCl_{(aq)} \rightarrow SnCl_{2(aq)} + H_{2(g)}$
Cu Hg Ag Pt Au	*They do not displace hydrogen from dilute acids, since they are below hydrogen in the reactivity series*.

4. Reactions of metals with chlorine gas:

All metals react with chlorine to give the corresponding *anhydrous chloride*. Reactivity of metals with chlorine decreases down the series.

$2Na_{(s)} + Cl_{2(aq)} \rightarrow 2NaCl_{(aq)}$

$Mg_{(s)} + Cl_{2(aq)} \rightarrow MgCl_{2(aq)}$

$Ca_{(s)} + Cl_{2(aq)} \rightarrow CaCl_{2(aq)}$

$2Fe_{(s)} + Cl_{2(aq)} \rightarrow 2FeCl_{3(aq)}$

$2Al_{(s)} + 3Cl_{2(aq)} \rightarrow 2AlCl_{3(aq)}$

$Sn_{(s)} + 2Cl_{2(aq)} \rightarrow SnCl_{4(aq)}$

5. Corrosion of Metals:

Corrosion is the term used to describe the effects produced when metals are exposed to air or water. *It is a slow process*.

Generally, electropositive metals exhibit corrosion: Na, K, and Ca corrode *readily* in air, while Mg, Al, Zn, Pb and Cu corrode slowly. Metals such as gold, silver and platinum are resistant to corrosion. *Corrosion of iron is called rusting*.

USES OF METALS

1. Metals are used for making alloys.
2. They are used in making coins examples: copper, silver and gold.

EXTRACTION OF METALS

DEFINITION

Extraction of metal is the process of obtaining metals from their ores inside the earth's crust. The process involves separating the ores from impurities, converting the ore to its oxide and then reducing it to the metal. *An ore is an impure oxide or sulphide of a metal. The process whereby a metal is obtained from its ore is called smelting.* The unwanted impurities like sand, rocky materials, limestone, mica, etc, ..., present in an ore are called *gangue.*

The process involved in the extraction of metals from their ores and refining are known as metallurgy.

STAGES INVOLVED IN THE EXTRACTION OF METALS

(1). **Concentration or enrichment of ore**:

The concentration of ore is the removal of gangue/impurities from metal ore. The processes used for removing the gangue from the ore are based on the differences between the physical and the chemical properties of the gangue and the ore. Ores are usually concentrated and changed to oxides before extraction.

Ores may be concentrated by one of the following ways:

(i). **Hand-picking**: This is commonly used when the ore particles and the impurities/gangue are different in one of the following properties like colour and size.

(ii). **Hydraulic washing**: Hydraulic washing of the ore is a type of gravity separation between the ore and the gangue particles. *This method is used when there is considerable difference in the density of the ore and the gangue particles*. In this process, the lighter gangue particles are washed away by the stream of water, leaving behind the heavier ore. This method is applied in the concentration of tin ore.

(iii). **Froth-floatation**: This is used mainly for sulphide ores. It takes advantage of the two facts namely:

- Oil can wet the surfaces of the ores
- Oil floats in water.

The Process:

(a). The ore is ground into fine powder; to increase the surface area for upcoming reactions.

(b). It is then mixed with water and a suitable oil detergent. Example pine or eucalyptus;

(c). The mixture is then agitated by blowing compressed air through it.

(d). Small air bubbles attach to the oiled ore particles which are then buoyed up and carried to the surface when they float.

(e). Froth rich in minerals is formed at the top while impurities sink at the bottom. The froth is skimmed off and dried. Froth floatation process is used for copper, lead and zinc metals.

(iv). **Magnetic separation**: This method is used when either the ore particles or the earthly materials are magnetic. A strong magnet is used to attract magnetic components and leaving the non-magnetic materials behind. Examples: in ores like magnetite(Fe_3O_4) and chromite($FeCr_2O_4$) which are magnetic.

(v). **Chemical separation**: Involves the use of chemical reactions to concentrate the ore. Examples: bauxite, the main aluminum ore is chemically concentrated by a process known as Bayer's process. This takes advantage of the amphoteric nature of aluminum oxide which, thus, can react with both acids and bases.

BAYER'S METHOD [used in concentrating bauxite, $Al_2O_3.2H_2O$]

Here, the bauxite ore is crushed and treated with hot concentrated NaOH solution in a tank called digester under high pressure for 2 to 8 hours at 140℃ to 150℃.

The process dissolves sodium tetrahydroxoaluminate(III), $NaAl(OH)_4$, while the earthly impurities such as oxide of iron (Fe_2O_3), silicon(IV) oxide(sand) and calcium trioxosilicate(IV) remain undissolved in aqueous NaOH solution.

$$\mathbf{Al_2O_{3(s)} + 2NaOH_{(l)} + 3H_2O_{(l)} \rightarrow 2NaAl(OH)_{4(aq)}}$$

$$\mathbf{NaAl(OH)_{4(aq)} \rightarrow NaOH_{(aq)} + Al(OH)_{3(aq)}}$$

$$\mathbf{2Al(OH)_{3(aq)} \rightarrow Al_2O_{3(s)} + 3H_2O_{(l)}}$$

Note: $NaAl(OH)_4$ = sodium tetrahydroxoaluminate(III)

(2). **Extraction of crude metal**: After concentration or enrichment of the ore that contained earthen material, we get a concentrated or enriched ore. It is easier to obtain metals from their oxides than from carbonates or sulphides. The sulphide ores are converted into oxides by heating strongly in the presence of excess air. This process is known as *roasting*. Also, the carbonates ores can be changed into oxides by heating strongly in limited air. The process is known as *calcination*.

Sulphide ores + Oxygen[excess] → oxide ore [roasting]

Carbonate ores → oxide ore + $CO_{2(g)}$ [calcination]

$$\mathbf{2ZnS_{(s)} + 3O_{2(g)} \rightarrow 2ZnO_{(s)} + 2SO_{2(g)}}$$ [roasting]

$$\mathbf{2PbS_{(s)} + 3O_{2(g)} \rightarrow 2PbO_{(s)} + 2SO_{2(g)}}$$ [roasting]

$$\mathbf{4FeS_{(s)} + 7O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 4SO_{2(g)}}$$ [roasting]

$$\mathbf{ZnCO_{3(s)} \rightarrow ZnO_{(s)} + CO_{2(g)}}$$ [calcination]

(3). **Refining or purification of the metal**:

PRINCIPLES OF METAL EXTRACTION

The method of extraction of a metal from its ore depends on its position in the electrochemical series. The conversion of metal oxide into metal (by removal of oxygen) is called *reduction*.

The three reduction methods by which metals are extracted from their ores are: electrolysis, use of carbon or carbon(II) oxide and thermal reduction.

ELECTROLYTIC REDUCTION

Metals high up the series (K, Na, Ca, Mg and Al) are strongly electropositive and cannot be obtained from their compounds by heating with carbon or carbon(II) oxide, CO. This is because these metals have more affinity for oxygen than carbon or CO. Hence, they are extracted by electrolysis.

REDUCTION BY COKE OR CO

Metals in the middle of the electrochemical series (Zn, Fe, Sn and Pb) are found as oxides. They are less reactive, hence, they are reduced by coke or carbon(II) oxide at high temperatures to the corresponding metals. Example:

Iron ore: $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$

Tin ore: $SnO_{2(s)} + 2C_{(s)} \rightarrow Sn_{(s)} + 2CO_{(g)}$

THERMAL REDUCTION [Roasting in Air]

Metals lower down in the activity series are unreactive and weakly electropositive e.g. (Cu, Hg, and Pt). They are called *noble metals*. They are found as oxides and sulphides that are readily decomposed by heating in air. For example; copper metal is obtained by heating copper(II) sulphide in air:

$Cu_2S_{(s)} + O_{2(g)} \rightarrow 2Cu_{(s)} + SO_{2(g)}$

$2HgS_{(s)} + O_{2(g)} \rightarrow 2HgO_{(s)} + 2SO_{2(g)}$

Gold, Au is a unique metal; it is unreactive. It occurs naturally as a free metal.

THE ALKALI METALS

General Features:

1. The alkali metals are elements of group 1 of the periodic table. Each element has one valence electron. The early members include lithium, Li, sodium, Na, potassium, K. Others are Rubidium, Rb, Caesium, Cs and Francium, Fr. They are known as the *alkali metals* since they react with water to form alkalis.
2. They have relatively low melting and boiling points compared with metals; their melting and boiling points decrease down the group.
3. The *alkali metals* occur in nature as +1 ions. However, the last element in group 1, francium, Fr, does not exist naturally; it exits as an *unstable radioactive* element.
4. The relative reactivities of the *alkali metals* are reflected by their reactions with water. Thus, potassium reacts *explosively with water*, while sodium *vigorously* and lithium *readily*.

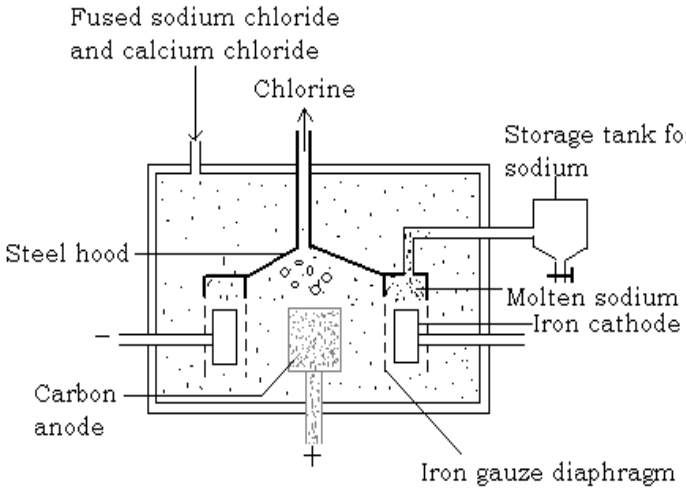
$2K_{(s)} + 2H_2O_{(aq)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$ **[Explosively]**
$2Na_{(s)} + 2H_2O_{(aq)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$ **[Vigorously]**
$2Li_{(s)} + 2H_2O_{(aq)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$ **[Readily]**

5. They exhibit an oxidation number of +1 in all their compounds.

EXTRACTION OF METALS AT THE TOP [HIGHLY REACTIVE] OF THE ACTIVITY SERIES

Metals such as K, Na, Mg, Ca, and Al high up in the activity series are very reactive and cannot be obtained from their compounds by heating with C or CO. *This is because these metals have more affinity for oxygen than* C or CO. The only viable method is to extract these metals by electrolysis of their *fused chlorides* except Al which is extracted from its oxide called *bauxite*.

EXTRACTION OF SODIUM IN THE DOWNS CELL



Cathode: Steel cylinder

Anode: Graphite rod

Electrolyte: Fused sodium chloride, NaCl mixed with calcium chloride, $CaCl_2$, at about 600℃.

Cell used: Downs cell

Procedure:

Sodium is extracted commercially from fused sodium chloride (melting point 801℃) by electrolysis using *downs cell*. During the electrolysis, the molten sodium is collected at the cathode compartment where it rises to the top and is tapped off via a pipe. *Gaseous chlorine, the by-product which forms at the anode*, is guided by a hood and collected from another pipe.

To make the process economically more feasible, calcium chloride, $CaCl_2$, is usually added to lower the melting point of sodium chloride to about 600℃.

Chemistry of the reaction: Fused sodium chloride contains sodium and chloride ions.

At the cathode(-ve): The Na^+ ions receive an electron each and become reduced to metallic sodium.

$Na^+ + e^- \rightarrow Na$ [Reduction reaction]

At the anode(+ve): The Cl^- ions give up an electron each and becomes oxidized to atomic chlorine which then pair up to form gaseous chlorine molecules.

$Cl^- \rightarrow Cl + e^-$ [Oxidation reduction]

$Cl + Cl \rightarrow Cl_{2(g)}$

Overall electrolytic reaction:

$2Na^+_{(l)} + 2Cl^-_{(l)} \rightarrow 2Na_{(s)} + Cl_{2(g)}$

Large amount of electricity is required to keep the ore in molten state. Suitable impurity is added to decrease the melting point of the ore. Fused calcium chloride is added to NaCl to decrease or reduce its melting point. The calcium is removed from the molten sodium by *filtration*.

PHYSICAL PROPERTIES OF SODIUM

Appearance	Slivery solid with metallic lustre
Relative density	$0.98gcm^{-3}$ (floats in water)
Malleability	Very malleable (very soft)
Melting point	97℃ (low for a metal)
Conductivity	Good conductor of electricity and heat

CHEMICAL PROPERTIES OF SODIUM

1. Reaction with air:

When sodium metal is exposed to the atmosphere, it undergoes the following changes:

(a). It tarnishes in air to a white solid, sodium oxide, Na_2O.

$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$ [Limited supply of oxygen]

Sodium oxide is only formed in the presence of limited air. When there is an adequate supply of air, it burns with a golden yellow flame to form peroxide, Na_2O_2.

$2Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{2(s)}$ [Adequate supply of oxygen]

(b). The moisture in air converts the oxide into the hydroxide:

$$Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$$

(c). The sodium hydroxide absorbs carbon(IV) oxide in air to form an aqueous solution of sodium trioxocarbonate(IV).

$$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(s)} + H_2O_{(l)}$$

2. <u>Combination reaction:</u>

On heating, sodium combines directly with hydrogen, halogens, sulphur, phosphorus and most other non-metals (except boron, carbon and nitrogen).

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

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

$$2Na_{(s)} + S_{(s)} \rightarrow Na_2S_{(s)}$$

$$3Na_{(s)} + P_{(s)} \rightarrow Na_3P_{(s)}$$

Sodium dissolves in mercury to give *sodium amalgam*, which reacts quietly with water to form hydrogen and sodium hydroxide.

$$Na_{(s)} + Hg_{(l)} \rightarrow Na.Hg_{(l)}$$

3. <u>Reaction with dilute acids:</u>

Sodium reacts <u>explosively</u> with dilute acids to form hydrogen and a salt. The reaction is extremely dangerous and should not be carried out in the laboratory.

$$2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(s)} + H_{2(g)}$$

$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$

Ionically,

$$2Na_{(s)} + 2H^+_{(aq)} \rightarrow 2Na^+_{(aq)} + H_{2(g)}$$

Caution: (*Never attempt the reaction of sodium and potassium with an acid*).

The reactions of sodium and potassium are similar. Because they react readily with air and water, they are kept under inert solvents, such as liquid paraffin, kerosene and petroleum ether.

<u>QUESTION 2</u>: Give a reason why sodium metal is stored in liquid paraffin. **<u>Answer:</u>** *To prevent sodium from reacting with air and water*.

<u>Reaction with ammonia:</u>

When ammonia gas is bubbled into molten sodium metal, a white solid called *sodamide*, is formed with liberation of hydrogen gas.

$$2Na_{(s)} + 2NH_{3(g)} \rightarrow 2NaNH_{2(s)} + H_{2(g)}$$

Reaction with water:

Sodium reacts *violently* with cold water releasing a lot of heat. If a small piece of sodium is placed in a trough of cold water, coloured with red litmus paper, the metal will dart about on the surface of the water, and melt into silvery ball because of the heat liberated. Hydrogen and sodium hydroxide is produced.

USES OF SODIUM

1. Sodium is used in manufacturing compounds as sodium peroxide, sodamide, and sodium cyanide.
2. It is also used for the production of lead(IV) tetraethyl which serve as an anti-knock agent in petrol.
3. Sodium vapour lamps (which give a bright orange-yellow light) are used for lighting highways and airports.
4. Liquid sodium is used as a coolant in nuclear reactors.

STUDY QUESTIONS

(1). Mention one natural source of sodium.

(2). Sodium metal is always stored under oil. Explain

(3). Describe the changes that occur when sodium is exposed to moist air

(4). List three changes that will be observed when sodium is dropped in a beaker of cold water.

(5). Sodium is extracted in the Downs cell

(a). Name the electrolyte;

(b). Which substance is added to the electrolyte?

(c). Why is graphite used as the anode?

(d). Name the products at the respective electrodes.

COMPOUNDS OF SODIUM

SODIUM HYDROXIDES, NaOH

LABORATORY PREPARATION:

(1). **By the action of slaked lime on washing soda:**

Calcium hydroxide is slowly added into a solution of sodium trioxocarbonate(IV), Na_2CO_3 (washing soda) with stirring and warming, until double decomposition is complete.

$$\mathbf{Na_2CO_{3(aq)} + Ca(OH)_{2(aq)} \rightarrow CaCO_{3(s)} + 2NaOH_{(aq)}}$$

The precipitate white solid, which is $CaCO_3$ is filtered and the filtrate is concentrated to dryness in an iron dish, to give fused NaOH (caustic soda).

(2). **By the action of sodium on water:**

The vigorous reaction between sodium and water is described above. The gas produced is hydrogen, while sodium hydroxide is the resulting *alkaline* solution.

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

INDUSTRIAL PREPARATION OF SODIUM HYDROXIDE

Sodium hydroxide is produced industrially by the electrolysis of brine (*concentrated sodium chloride solution*) using mercury cathode.

Salt ionizes: $\mathbf{NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}}$

Water ionizes: $\mathbf{H_2O_{(l)} \rightarrow H^+_{(aq)} + OH^-_{(aq)}}$

At the anode(+ve): (Cl^- and OH^-)

Cl^- is preferentially discharged, since it is of higher concentration than OH^- from water.

$$\mathbf{2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^- \text{ [Oxidation reaction]}}$$

Chlorine is liberated as a *by-product*.

At the cathode(-ve): (Na^+ and H^+)

Na^+ is discharge preferentially, because hydrogen creates a high over-voltage at the mercury cathode, which prevents its ion being discharged.

$$\mathbf{2Na_{(s)} \rightarrow 2Na^+_{(aq)} + 2e^-_{(aq)} \text{ [Reduction reaction]}}$$

The sodium metal produced at the cathode combines with mercury to form sodium amalgam, which flows into a tank of water, where it decomposes to produce sodium hydroxide and hydrogen gas.

$$\mathbf{2NaHg_{(l)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + 2Hg_{(l)} + H_{2(g)}}$$

PHYSICAL PROPERTIES OF NaOH

(1). Sodium hydroxide is a white *deliquescent* solid. Hence, NaOH (or KOH) should be stored in an air tight container.

(2). It dissolves in water with the liberation of heat to give a colourless solution, which is slippery, soapy to touch and corrodes the skin. (*Never taste a solution of* NaOH or KOH).

(3). An aqueous solution of sodium hydroxide is alkaline. It is also a strong electrolyte.

(4). Sodium hydroxide melts at about 320℃ without decomposing.

CHEMICAL PROPERTIES OF NaOH

Sodium hydroxide, NaOH (or potassium, KOH) undergoes reactions that are typical of a base.

(1). **Exposure to the moist air:**

When sodium hydroxide pellets are exposed to the atmosphere, the following changes occur.

(a). The pellets absorb moisture from air, being *deliquescent*, to form a colourless solution.

$$NaOH_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)}$$

(b). The solution absorbs carbon(IV) oxide in air to produce hydrated sodium trioxocarbonate(IV):

$$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_3.H_2O_{(s)}$$

Or *anhydrous* $NaHCO_3$ as a white solid:

$$NaOH_{(aq)} + CO_{2(g)} \rightarrow NaHCO_{3(s)}$$

(2). Reaction with a dilute acid and acidic oxides/acid anhydrides:

Sodium hydroxide solution is a strong electrolyte which ionizes completely in solution.

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

It is a typical base, neutralizing all acids.

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$$

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Ionically,

$OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)}$

Sodium hydroxide reacts with acidic oxides to form sodium salts. For example, it reacts with sulphur(IV) oxide to form the trioxosulphate(IV) or the hydrogentrioxosulphate(IV) of sodium.

$2NaOH_{(aq)} + SO_{2(g)} \rightarrow Na_2SO_{3(aq)} + H_2O_{(l)}$

$NaOH_{(aq)} + SO_{2(g)} \rightarrow NaHSO_{3(aq)}$

It readily absorbs CO_2, a weak acidic oxide, to form sodium trioxocarbonate(IV).

$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$

(3). <u>Reaction with ammonium salts</u>:

Ammonia gas is liberated when sodium hydroxide is heated with an ammonium salt.

$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$

$2NaOH_{(aq)} + (NH_4)_2CO_{3(aq)} \rightarrow Na_2CO_{3(aq)} + 2H_2O_{(l)} + 2NH_{3(g)}$

$2NaOH_{(aq)} + (NH_4)_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)} + 2NH_{3(g)}$

$NaOH_{(aq)} + NH_4NO_{3(aq)} \rightarrow NaNO_{3(aq)} + H_2O_{(l)} + NH_{3(g)}$

(4). <u>Reaction with metals:</u>

Aluminum and zinc readily dissolve on sodium hydroxide solution to form sodium tetrahydroxoaluminate(III), $NaAl(OH)_4$ and sodium tetrahydroxozincate(II), $Na_2Zn(OH)_4$ respectively, with the evolution of hydrogen gas.

$2Al_{(s)} + 2NaOH_{(aq)} + 6H_2O_{(aq)} \rightarrow 2NaAl(OH)_{4(aq)} + 3H_{2(g)}$

$Zn_{(s)} + 2NaOH_{(aq)} + 2H_2O_{(aq)} \rightarrow Na_2Zn(OH)_{4(aq)} + H_{2(g)}$

(5). <u>To precipitate hydroxides:</u>

Sodium hydroxide solution is often used to precipitate insoluble hydroxides.

SEE QUALITATIVE ANALYSIS TABLE BELOW.

$ZnSO_{4(aq)} + 2NaOH_{(aq)} \rightarrow 2Zn(OH)_{2(s)} + 3Na_2SO_{4(aq)}$

$Pb(NO_3)_{2(aq)} + 2NaOH_{(aq)} \rightarrow Pb(OH)_{2(s)} + 2NaNO_{3(aq)}$

Ionically,

$\mathbf{Zn^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Zn(OH)_{2(s)}}$

$\mathbf{Pb^{2+} + 2OH^-_{(aq)} \rightarrow Pb(OH)_{2(s)}}$

In the case of zinc, aluminum, tin(II) and lead(II) salts, the resulting hydroxide are amphoteric and will redissolve in an excess sodium hydroxide to form complex salts.

$\mathbf{Zn(OH)_{2(s)} + 2NaOH_{(aq)} \rightarrow Na_2Zn(OH)_{4(aq)}}$ [Sodium tetrahydroxozincate(II)]

Ionically,

$\mathbf{Zn(OH)_{2(s)} + 2OH^-_{(aq)} \rightarrow Zn(OH)_4^{2-}{}_{(aq)}}$

(6). Reaction with glass:

Molten and concentrated sodium hydroxide solution attack glass, forming sodium trioxosilicate(IV). This is why glass stoppers become sealed into reagent bottles containing the concentrated alkali and burette taps tend to get stuck after the alkali has been used in the burette. This is called *etching*.

$\mathbf{SiO_{2(s)} + 2NaOH_{(aq)} \rightarrow Na_2SiO_{3(aq)} + H_2O_{(l)}}$

SODIUM TRIOXOCARBONATE(IV), Na_2CO_3

Sodium trioxocarbonate(IV), Na_2CO_3, can exist:

(1). In the anhydrous state known as *soda ash*, as monohydrate, $Na_2CO_3.H_2O$ or

(2). More often as a decahydrate, $Na_2CO_3.10H_2O$, commonly called *washing soda*.

LABORATORY PREPARATION OF Na_2CO_3

Anhydrous sodium trioxocarbonate(IV), Na_2CO_3, is prepared in the laboratory by passing excess of carbon(IV) oxide, CO_2 into sodium hydroxide solution to produce, first, sodium hydrogentrioxocarbonate(IV), $NaHCO_3$(*baking soda*), a white solid, which is only slightly soluble in the cold solution.

$\mathbf{2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}}$ [first stage]

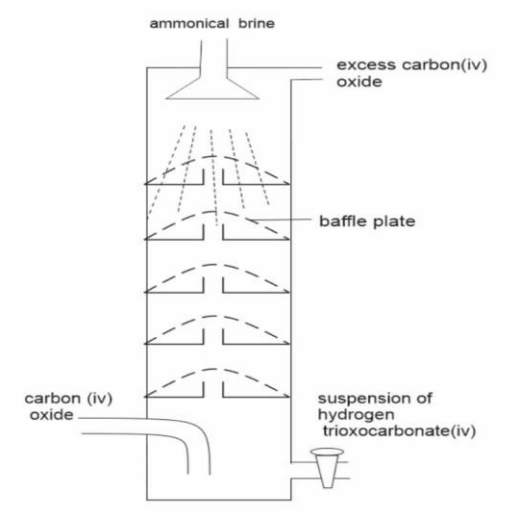
$\mathbf{Na_2CO_{3(aq)} + H_2O_{(l)} + CO_{2(g)} \rightarrow 2NaHCO_{3(s)}}$ [second stage]

The mixture is then *filtered* and the residue is dried; then, strongly heated, to give anhydrous Na_2CO_3, as white powder.

$\mathbf{2NaHCO_{3(s)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}}$

INDUSTRIAL PREPARATION OF Na_2CO_3 [Solvay Process]

Anhydrous sodium trioxocarbonate(IV), Na_2CO_3, is produced on a large scale by the Solvay process. The required raw materials are: **common salt/NaCl**, **limestone/$CaCO_3$**, **ammonia** and **water**.



PROCEDURE:

(1). The limestone, $CaCO_3$ is heated strongly by coke fire in the *kiln* to produce calcium oxide, CaO, a white residue, and CO_2 gas, which is forced up the tower from the bottom.

$$\mathbf{CaCO_{3(s)} \leftrightarrow CaO_{(s)} + CO_{2(g)}}$$

(2). A cold concentrated solution of sodium chloride, brine, which has been saturated with ammonia gas (*ammoniacal brine*) is poured down the tower.

(3). The baffles in the tower ensure that the *ammoniacal brine* absorbs the CO_2 gas, to enable the following reactions to take place.

(a). Ammonium hydrogentrioxocarbonate(IV), NH_4HCO_3, is first formed.

$$\mathbf{NH_{3(g)} + CO_{2(g)} + H_2O_{(l)} \rightarrow NH_4HCO_{3(aq)}}$$

(b). Double decomposition reaction, takes place between ammonium hydrogentrioxocarbonate(IV), NH_4HCO_3 and sodium chloride, NaCl solution; it leads to the *precipitation* of sodium hydrogentrioxocarbonate(IV), $NaHCO_3$, which is slightly soluble in the solution, and settles at the bottom of the tower.

$$\mathbf{NH_4HCO_{3(aq)} + NaCl_{(aq)} \rightarrow NH_4Cl_{(aq)} + NaHCO_{3(s)}} \text{ or}$$

$$\mathbf{NaCl_{(aq)} + NH_{3(g)} + H_2O_{(l)} + CO_{2(g)} \rightarrow NaHCO_{3(s)} + NH_4Cl_{(aq)}} \text{ or}$$

$$\mathbf{NaCl_{(aq)} + NH_3OH_{(aq)} + CO_{2(g)} \rightarrow NaHCO_{3(s)} + NH_4Cl_{(aq)}}$$

(4). The sodium hydrogentrioxocarbonate(IV), $NaHCO_3$, is filtered, washed and *heated* to yield anhydrous sodium trioxocarbonate(IV), Na_2CO_3, steam, $\mathbf{H_2O_{(g)}}$ and CO_2 and used again the tower.

$$\mathbf{2NaHCO_{3(aq)} \rightarrow Na_2CO_{3(s)} + H_2O_{(g)} + CO_{2(g)}}$$

The soda ash can then be dissolved in hot water and *recrystallized* as washing soda.

$$\mathbf{Na_2CO_{3(s)} + 10H_2O_{(l)} \rightarrow Na_2CO_3.10H_2O_{(s)}}$$

ECONOMICS OF THE REACTION [Solvay Process]

Solvay process is continuous, efficient and economical for the following reasons:

(1). The principal raw materials for the process are cheap and plentiful. They are sodium chloride, NaCl and Limestone, $CaCO_3$. The common salt is extracted from deposits as brine; the limestone yields quicklime/calcium oxide and carbon(IV) oxide.

$$\mathbf{CaCO_{3(s)} \leftrightarrow CaO_{(s)} + CO_{2(g)}}$$

(2). Ammonia is recovered and use again in the tower by the action of heat on a mixture of ammonium chloride and calcium oxide.

$$\mathbf{2NH_4Cl_{(aq)} + CaO_{(s)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + 2NH_{3(g)}}$$

(3). The only waste product in the Solvay process is calcium chloride. However, it is used in the industrial production of calcium metal by electrolysis and as a drying agent−being *hygroscopic*.

PHYSICAL PROPERTIES OF Na_2CO_3

(1). Anhydrous sodium trioxocarbonate(IV), Na_2CO_3, exists as a white powder and white crystals when hydrated.

(2). The white hydrated salt efflorescences in air to form a monohydrate salt:

$$\mathbf{Na_2CO_3.10H_2O_{(s)} \rightarrow Na_2CO_3.H_2O_{(s)} + 9H_2O_{(g)}}$$

(3). When heated strongly, the crystalline salt becomes completely powdery and anhydrous:

$$\mathbf{Na_2CO_3.10H_2O_{(s)} \rightarrow Na_2CO_3 + 10H_2O_{(l)}}$$

(4). Both the hydrate and anhydrous salts dissolve in water to give colourless solutions.

CHEMICAL PROPERTIES OF Na_2CO_3

(1). Na_2CO_3 dissolves in water to give an alkaline solution; due to *hydrolysis*:

$$\mathbf{Na_2CO_3 + 2H_2O_{(s)} \rightarrow 2NaOH_{(aq)} + H_2CO_{3(aq)}}$$

(2). When a dilute mineral acid is added to anhydrous or hydrated Na_2CO_3 salt, there is a *brisk effervescence*. The gas evolved is colourless, odourless, changes blue litmus paper to red and turns lime water milky. The gas is CO_2.

$$\mathbf{Na_2CO_3 + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}}$$

$$\mathbf{Na_2CO_3 + 2HNO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}}$$

$$\mathbf{Na_2CO_3.xH_2O + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + (x + 1)H_2O_{(l)} + CO_{2(g)}}$$

$$\mathbf{Na_2CO_3.xH_2O + 2HNO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + (x + 1)H_2O_{(l)} + CO_{2(g)}}$$

USES OF Na_2CO_3

Three of the important uses of sodium trioxocarbonate(IV), Na_2CO_3, are the following:

(1). **Manufacture of glass**: *Glass is a super-cooled liquid*. It is a mixture of trioxosilicate(IV) salts of sodium and calcium.

Ordinary bottle glass is made up by fusing together sodium trioxocarbonate(IV), Na_2CO_3 (or sodium tetraoxosulphate(VI), Na_2SO_4), calcium trioxocarbonate(IV), $CaCO_3$, Silica/Sand/silicon(IV) oxide, SiO_2 and a little carbon (*reducing agent*). Broken glass(cullet) is added to assist fusion.

$$\mathbf{Na_2CO_{3(s)} + SiO_{2(s)} \rightarrow Na_2SiO_{3(s)} + CO_{2(g)}}$$

$$\mathbf{CaCO_{3(s)} + SiO_{2(s)} \rightarrow CaSiO_{3(s)} + CO_{2(g)}}$$

The mixture of silicates (with some unchanged silica) constitutes the glass.

(2). **Manufacture of Water-glass**: Water-glass is sodium trioxosilicate(IV), Na_2SiO_3. Sodium trioxocarbonate(IV), Na_2CO_3, is fused with silica/silicon(IV) oxide to produce sodium trioxosilicate(IV), Na_2SiO_3.

$$\mathbf{Na_2CO_{3(s)} + SiO_{2(s)} \rightarrow Na_2SiO_{3(s)} + CO_{2(g)}}$$

On cooling, a glassy solid is left, which is broken up, boiled with water and evaporated to a colourless, treacly material, known as *water-glass*.

USES OF WATER-GLASS:

Water-glass is used in preserving eggs, in fire-proofing and in producing cements.

(3). **In domestic water-softening**: Calcium ion, Ca^{2+}, which is the principal cause of hardness in water, is precipitated from the water as chalk, $Ca^{2+}CO_3^{2-}$, by the addition of sodium trioxocarbonate(IV), Na_2CO_3.

Note: In the laboratory, Na_2CO_3, is used to standardize acids and as an analytical reagent.