CHAPTER NO: 2. I year

'S' BLOCK ELEMENTS.

Elements having their outermost electrons in 'S' orbital are called 'S' BLOCK ELEMENTS.
'S' block elements include elements of I-A group (ALKALI METALS) and II-A group (alkaline earth metals).

ALKALI METALS:

Elements of group I-A are called ALKALI METALS. The name alkali means "THE ASHES". Because the ashes of Organic compounds mainly contain Na and K. So elements of I-A are called ALKALI METALS.
These are also called alkali metals because their oxides are alkaline.

OCCURRENCE. Alkali metals donot exist ~~in~~ free in nature. these occur in the combined state. Most of earth's crust is composed of alkali metal comp.

SOURCE OF Na & K:

These are most abundant alkali metals each constitute about 2.4% of earth crust. these metals ~~form~~ exist as insoluble alumino silicates.

Table 2.3 Common Minerals of the Most Important Alkali Metals: Table

Element	Name of Mineral	Chemical Formula
Lithium	Spodumene	$LiAl(SiO_3)_2$
Sodium	Rock Salt (Halite)	$NaCl$
	Chile saltpetre	$NaNO_3$ ✓
	Natron	$Na_2CO_3.H_2O$
	Trona	$Na_2CO_3.2NaHCO_3.2H_2O$
	Borax	$Na_2B_4O_7.10H_2O$
Potassium	Carnallite	$KCl.MgCl_2.6H_2O$
	Sylvite	KCl
	Alumite(Alum Stone)	$K_2SO_4.Al_2(SO_4)_3.4Al(OH)_3$

OCCURRENCE OF LITHIUM. Lithium exist in the form of minerals. An important commercial source of lithium is the mineral spodumene ($LiAl(SiO_3)_2$)

OCCURRENCE OF Rb & Cs.

Small amount of Cs and Rb are found in potassium salts deposits.

OCCURRENCE OF FRANCIUM. It has not been found in nature. It is artificially prepared in laboratory.

ALKALINE EARTH METALS.

Elements of II-A group are called "ALKALINE EARTH METALS."

These are called alkaline earth metals because these produce alkalies in water and are widely distributed in earth's crust.

OCCURRENCE

Alkaline earth metals donot occur in free state. Compounds of these metals occur widely in nature

SOURCE OF Mg.

Mg is very abundant in earth crust. Magnesium salicles are found in sea water. Magnesium is also found in chlorophyll. It is an essential constituent of chlorophyll.

Table 2.4 Common Minerals of the Alkaline Earth Metals:

Element	Name of Mineral	Chemical Formula
Beryllium	Beryl	$Be_3Al_2(SiO_3)$
	Chrysoberyl	Al_2BeO_4
Magnesium	Magnesite	$MgCO_3$
	Dolomite	$MgCO_3.CaCO_3$
	Carnallite	$KCl.MgCl_2.6H_2O$
	Epsom Salt	$MgSO_4.7H_2O$
	Soap Stone (talc)	$H_2Mg_3.7H_2O$
	Asbestos	$CaMg_3(SiO_3)_4$
Calcium	Calcite (Lime Stone)	$CaCO_3$
	Gypsum	$CaSO_4.2H_2O$
	Fluorite	CaF_2
	Phosphorite	$Ca_3(PO_4)_2$
Strontium	Strontionite	$SrCO_3$
Barium	Barite	$BaSO_4$

SOURCE OF CALCIUM:- Calcium exist in the form of Phosphate and fluoride minerals ~~and~~ ($Ca_3(PO_4)_2$, CaF_2) It also exist as silicates and aluminosilicates. It is an essential constituent of many living organism. It occurs as skeletal material in bone, teeth. Sea-shells and egg shells also contain calcium.

SOURCE OF RADIUM:- It is rare element. It has radioactive nature. ~~exists~~ - -

WRITE NAMES, SYMBOLS & ELECTRONIC CONFIGURATION OF "S" - BLOCK ELEMENTS. (I-A, II-A).

ALKALI METALS:

Alkali metals have only one electron in "s" orbital of valence shell. Alkali metals lose their one electron of valence shell to form unipositive ion. M^+ They form ionic compounds and show +1 oxidation state. They have very low ionization energy value.

I-A GROUP.

NAME	ATOMIC NO.	ELECTRONIC CONFIGURATION.
Lithium.	$_3Li$	$1s^2, 2s^1$.
Sodium.	$_{11}Na$	$1s^2, 2s^2, 2p^6, 3s^1$.
Potassium.	$_{19}K$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$.
Rubidium.	$_{37}Rb$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^1$.
Cesium.	$_{55}Cs$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 4p^6, 3d^{10}, 5s^2, 5p^6, 6s^1$.
Francium.	$_{87}Fr$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}$
		$5d^{10}, 6p^6, 7s^1$

Table 2.1 Electronic Configuration and Physical Constants of Alkali Metals.

Properties	Li	Na	K	Rb	Cs ✓
Atomic Number	3	11	19	37	55
Electronic configurations	$1s^2,2s^1$	[Na]$3s^1$	[Ar]$4s^1$	[Kr]$5s^1$	[Xe]$6s^1$
Melting Points °C	186.1	97.5	62.7	39.0	26.0
Ionic Radius pm	60	95	133	148	169
Ionization Energy (kJ/ mole)	520	495	420	400	380
Density gm/cm³ at (20°C)	0.53	0.97	0.86	1.53	1.9
Heat of Hydration (kJ/mole)	505	475	384	345	310

ALKALINE EARTH METALS.

These have two electrons in "s" orbital of their valence shell. so these are placed in II-A group. These elements lose two electrons to form dipositive ion M^{2+}. They form ionic compound and show 2+ ox. states.

II-A GROUP.

NAME	ATOMIC NO.	ELECTRONIC CONFIGURATION.
Beryllium	$_4Be$	$1s^2, \underline{2s^2}$.
Magnesium	$_{12}Mg$	$1s^2, 2s^2, 2p^6, \underline{3s^2}$.
Calcium	$_{20}Ca$	$1s^2, 2s^2, 2p^6, 3s^2, 4p^6, \underline{4s^2}$.
Strontium	$_{38}Sr$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, \underline{5s^2}$
Barium	$_{56}Ba$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 4p^6, 3d^{10}, 5s^2, 5p^6, \underline{6s^2}$
Radium	$_{88}Ra$	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6$
		$6s^2, 4f^{14}, 5d^{10}, 6p^6, \underline{7s^2}$

Table 2.2 Electronic Configurations of Alkaline Earth Metals

Properties	Be	Mg	Ca	Sr	Ba
Atomic Number	4	12	20	38	56
Electronic Configurations	$1s^2,2s^2$	[Ne]$3s^2$	[Ar]$4s^2$	[Kr]$5s^2$	[Xe]$6s^2$
Melting Points° C	1289	651	851	771	849
Ionic Radius pm	31	65	99	113	135
2nd Ionization Energy (kJ/mole)	1800	1450	1150	1060	970
Density g/cm^3 (20°C)	1.85	1.74	1.55	2.6	3.5
Heat of Hydration (kJ/mole)	2337	1897	1619	1455	1250

HOW DOES LITHIUM DIFFERS FROM OTHER MEMBERS OF IA FAMILY

PECULIAR BEHAVIOUR OF LITHIUM.

Lithium differs from other alkali metals due to following reasons.

1. Very small size.
2. High charge density.

Most important differences are listed below

1. HARD AND LIGHT METAL.

lithium is much harder and lighter than the other alkali metals.

2. SOLUBILITY IN WATER.

Lithium salts of anions with high charge density are less soluble in water than salts of other alkali metals. e.g; $LiOH$, LiF, Li_2CO_3 etc.

3. FORMATION OF STABLE COMPLEXES.

Lithium forms stable complex compounds. although alkali metals donot form complexes. Most stable complex formed by lithium is $Li(NH_3)_4$.

4. REACTIVITY.

Lithium reacts very slowly with water. while other react with water violently.
Lithium is least reactive metal of all alkali metals

5. STABILITY OF SALTS.

Lithium salts of large anions are less stable than those of other alkali metals. Lithium donot form bicarbonate, tri-iodide or hydrogensulphide at room temperature. Where other alkali metals do so

6. OXIDES.

Lithium forms only normal oxide whereas other alkalimetals form peroxides or superoxides.

7. LITHIUM HYDRIDE.

Lithium hydrides is more stable than the other alkali metal hydride.

8. REACTION WITH ACETYLENE.

Lithium does not produce lithium acetylide on reaction with acetylene. but other alkali metals produce metallic acetylides.

$$2Na + C_2H_2 \longrightarrow \overset{+}{Na}\,\overset{-}{C}\equiv\overset{-}{C}\,\overset{+}{Na} + H_2$$

sod. acetylide.

9. COVALENT NATURE. Lithium compounds are more covalent that's ~~alkali~~ is why its halides are more soluble in organic solvents. aryls and alkyls of lithium are also more stable than those of other alkali metals

10. LITHIUM CARBONATES AND NITRATES.

Lithium has low electropositivity. Thus its carbonates and Nitrates are not stable. and decompose to give lithium oxides. Lithium nitrates on decomposition give different products than other alkali metals

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2.$$

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2.$$

$$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2.$$

11. HYDROXIDE.

Lithium hydroxide on strong heating produce lithium oxide but other alkali metal hydroxides do not do so. $2LiOH \xrightarrow{\text{red hot}} Li_2O + H_2O$

12. REACTION WITH NITROGEN.

Lithium reacts with nitrogen to give nitride but others do not give this reaction.

$$6Li + N_2 \longrightarrow 2Li_3N$$

13. HEAT OF SOLUTION. Lithium chloride has exothermic heat of solution but others have endothermic.

14. CARBIDES:-

~~Lithium carbide is the only alkali~~

Alkali metal carbides cannot be prep. by direct method except lithium carbide.

PECULIAR BEHAVIOUR OF BERYLLIUM.

Beryllium differs from other II-A group elements due to following reasons.

1- its small atomic size.
2. High electronegativity.

Most important differences are as follows.

1. HARDNESS

Beryllium metal is as hard as iron. but other alkaline earth metals are much softer than beryllium.

2. MELTING AND BOILING POINT:

melting and boiling points of beryllium are higher than other alkaline earth metals.

	Be	Mg	Ca	Sr	Ba.
MELTING POINT °C.	1277	650	838	763	714.
BOILING POINT °C.	2770	1107.	1440	1380	1640.

3. REDUCING AGENT

All alkaline earth metals reduce water. these act as powerful reducing agent ~~when~~ on reaction with water But lithium does not reduce water. and form insoluble oxides.

4. OXIDATION.

Beryllium does not under go complete oxidation even by acids. Because it form insoluble oxide (BeO). coating.

5. REACTION WITH ALKALIES:

Beryllium reacts with alkalies. While other members donot react with alkalies

$$Be + NaOH \longrightarrow \underset{\text{(Sod. Beryllate)}}{Na_2BeO_2} + H_2$$

GENERAL BEHAVIOUR OF ALKALI METALS.

Reducing agents can lose electrons. Alkali metals have low ionization energies. These can loose electrons easily So alkali metals are strong reducing agents.
Alkali metals are highly electropositive. They react with halogens to form halides.

TRENDS IN CHEMICAL PROPERTIES OF ALKALI METALS.

1. IONIZATION ENERGIES.-
Alkali metals have low ionization energies So these are the most reactive family in metals.

2. SECOND IONIZATION POTENTIAL
Alkali metals have very high second ionization energies. This show that alkali metals cannot have oxidation No. higher than 1

3. OXIDES OF ALKALI METALS
Alkali metals react with oxygen to form oxides.
Only lithium forms normal oxide, Li_2O.

$$4Li + O_2 \longrightarrow 2Li_2O. \quad \text{(WHITE SOLID)}$$

Oxides formed react with CO_2 in air to form carbonate.

$$Li_2O + CO_2 \longrightarrow Li_2CO_3.$$

Sodium also react in the same mannar if supply of oxygen is limited. In excess of oxygen it forms pale yellow per oxide.

$$2Na + O_2 \longrightarrow Na_2O_2 \quad \text{(PALE YELLOW)}$$

(Sod. per oxide).

K, Rb, and Cs form super oxide.

$$K + O_2 \longrightarrow KO_2. \quad \text{(ORANGE YELLOW)}$$

(Pot. super oxide)

4. REACTION WITH WATER.
Alkali metals react rapidly with water and liberate hydrogen and form alkali metal hydroxide. Reaction is highly exothermic.
Reaction rate increases from lithium to caesium
K, Rb, and Cs can even react with ice at $-100°C$.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2.$$

HYDRIDES OF I-A

Alkali metals form ionic hydride with hydrogen. Rb and

$$2M + H_2 \longrightarrow 2MH$$

Cs react violently at room temp. while Na, K, Li react with hydrogen at relatively high temp.

On treatment with H_2O these hydrides liberate H_2.

$$LiH + H_2O \longrightarrow LiOH + H_2$$

ionic hydrides are used as powerful reducing agents

6. SOLUBILITY IN WATER.

Cations of alkali metals have low charge and large radii. So these have low lattice energy. Most of the salts of these metals are water soluble. Salts are completely dissociated in aqueous solution.

7. NITRIDES & CARBIDES

Only lithium can react with C and nitrogens among alkali metals, to form nitride and carbide.

$$6Li + N_2 \longrightarrow 2Li_3N$$

lithium nitride.

$$4Li + C \longrightarrow Li_4C$$

lithium carbide.

8. REACTION WITH HALOGENS.

Alkali metals react easily with halogens. Li and Na react slowly at room temp. Molten Na burns with yellow flame in a chlorine atmosphere to form NaCl.

$$2Na + Cl_2 \longrightarrow 2NaCl.$$

K, Cs and Rb react vigorously with Halogens.

9. REACTION WITH SULPHER.

All alkali form sulphide on reaction with sulpher.

$$2M + S \longrightarrow M_2S$$

metal sulphide

TRENDS IN CHEMICAL PROPERTIES OF ALKALINE EARTH METALS.

1. OXIDES.

Alkaline earth metals burn with oxygen to produce normal oxides, except ~~beryllium~~ Barium. It forms peroxides.
Beryllium does not undergo complete oxidation at Room temp. At 800°c it oxidizes rapidly.

$$2Be + O_2 \xrightarrow{800°C} 2BeO.$$

Magnesium readily form oxide. It become coated with layer of MgO. So no more reaction occur.

$$2Mg + O_2 \longrightarrow 2MgO.$$

Small amount of nitride is also formed along with MgO.
Barium forms peroxides at 500 – 600°C

$$2Ba + O_2 \longrightarrow Ba_2O_2$$

Bariumperoxide.

2. HYDRIDES

Alkaline earth metals form hydride in molten state and at high pressure.
Mg. reacts with hydrogen at high pressure in presence of MgI_2 catalyst

$$Mg + H_2 \xrightarrow{MgI_2} MgH_2.$$

similarly

$$Ca + H_2 \longrightarrow CaH_2.$$

3. NITRIDES.

These metals react with N_2 to form nitride

$$3Mg + N_2 \longrightarrow Mg_3N_2.$$

nitrides hydrolyse readily with water to give NH_3.

$$Mg_3N_2 + 6H_2O \longrightarrow 2NH_3 + 3Mg(OH)_2.$$

4. SULPHIDES & HALIDES

Alkaline earth metals form sulphide with sulphur

$$Mg + S \longrightarrow MgS.$$

5. They readily react with halogens

$$Ca + Cl_2 \longrightarrow CaCl_2.$$

REACTIVITY:

6. Magnesium is more reactive than beryllium. Beryllium does not react with water. Mg react with boiling water and readily with steam and liberates H_2.

$$Mg + H_2O \xrightarrow{100°C} MgO + H_2$$

but other alkaline earth metals form hydroxide

$$M + 2H_2O \xrightarrow{100°C} M(OH)_2 + H_2.$$

GENERAL TRENDS IN PROPERTIES OF COMPOUNDS OF ALKALI AND ALKALINE EARTH METALS.

1. OXIDES:-

(a)- ALKALI METAL OXIDES.

i - alkali metal oxides are soluble in water they produce strongly alkaline solution in water. e.g.

$$Li_2O + H_2O \longrightarrow 2LiOH.$$

$$2Na_2O_2 + 2H_2O \longrightarrow 4NaOH + O_2.$$

ii - alkali metal oxides react with water. Reaction is an acid base reaction. In this reaction water is decomposed by an oxide ion.

$$O^{2-} + H_2O \longrightarrow 2OH^{\ominus}$$

iii - Basic character of alkali metal oxides increases down the group.

iv - Oxides of alkali metals are used in breathing equipment and in air crafts. KO_2 has ability to absorb CO_2 and giving out oxygen.

$$4KO_2 + 2CO_2 \longrightarrow 2K_2CO_3 + 3O_2.$$

(b)- OXIDES OF ALKALINE EARTH METALS.

i Solubility of alkaline earth metal oxides in water increases down the group. BeO and MgO are insoluble but CaO, SrO, and BaO are soluble.

ii –

These oxides react with water to form hydroxides

iii –

Basic character of alkaline earth metal oxides increases down the group.

Solution of alkaline earth metal oxides are less alkaline than that of alkali metals.

GROUP I-A	GROUP II-A
Lithium oxide (BASIC)	Beryllium oxide (AMPHOTERIC
Sodium oxide (BASIC)	Magensium oxide (BASIC)
Potassium oxide (BASIC)	Calcium oxide (BASIC)

increasing basicity of oxides

← Increasing basicity of oxides.

2. HYDROXIDES.

(a) – ALKALI METAL HYDROXIDES.

i – Alkali metal hydroxides are all crystalline solid
ii. These are soluble in water except LiOH.
iii – These are hygroscopic except LiOH.
iv – Alkali metal hydroxides are very strong bases except lithium hydroxide.
v – Solubility of alkali ~~ear~~ metal hydroxides increases down the group.
vi – These hydroxides are stable to heat except LiOH

$$2LiOH \longrightarrow Li_2O + H_2O.$$

(b) – ALKALINE EARTH METAL HYDROXIDES.

i – Solubility of alkaline earth metal hydroxides in water increases down the group.

$Be(OH)_2$ is insoluble, $Mg(OH)_2$ is sparingly soluble, while $Ba(OH)_2$ is more soluble.

ii - Increase in solubility is due to low lattice energy which is in turn due to higher ionic size.

iii. Alkaline earth metal hydroxides decompose on heating.

$$Mg(OH)_2 \longrightarrow MgO + H_2O$$

iv - Saturated solution of $Ca(OH)_2$ in water is called "LIME WATER". It is used for test of CO_2.

v - A suspension of $Mg(OH)_2$ in water is called "MILK OF MAGNESIA". It is used for treatment of acidity of stomach.

3. NITRATES :-

Nitrates of both alkali and alkaline earth metals are soluble in water. Nitrates of Li, Mg and Ba decompose on heating to give O_2, NO_2 and the oxides. Nitrates of Na, K and Ca decompose to give diff. products.

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$

$$Mg(NO_3)_2 \longrightarrow 2MgO + 4NO_2 + O_2$$

$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

$$Ca(NO_3)_2 \longrightarrow Ca(NO_2)_2 + O_2$$

4. SULPHATES :-

All the alkali metals give sulphates and they are soluble in water.

Solubilities of alkaline earth metal sulphates decreases down the group. $BeSO_4$ and $MgSO_4$ are soluble in water. $CaSO_4$ is slightly soluble. $SrSO_4$ and $BaSO_4$ are almost insoluble.

Calcium sulphate occur in nature as gypsum $CaSO_4 \cdot 2H_2O$. On heating above 100°C, it loses three quarters of its water of crystallization. It will give a white powder called PLASTER OF PARIS.

$$\underset{\text{GYPSUM}}{2CaSO_4 \cdot 2H_2O} \longrightarrow \underset{\text{PLASTER OF PARIS}}{(CaSO_4)_2 \cdot H_2O} + 3H_2O$$

4. CARBONATES.

(a) - CARBONATES OF ALKALI METALS

i - SOLUBILITY

Carbonates of alkali metals are soluble in water except lithium carbonate (Li_2CO_3 insoluble)

ii - HEATING

They donot decompose on heating except Li_2CO_3.

iii - DECOMPOSITION OF Li_2CO_3

lithium carbonate is easily decomposed Because electrostatic attraction increasing in converting Li_2CO_3 to Li_2O.

iv - DECOMPOSITION OF OTHER ALKALI METAL CARBONATES.

In case of carbonates with larger cations as K_2CO_3 gain in electrostatic attraction is quite less so These carbonates are not easily decomposed.

v - USES:-

Na_2CO_3 is very important industrial chemical. Below 35 C, Na_2CO_3 is crystallizes out as $Na_2CO_3 \cdot 10H_2O$. This is called "WASHING SODA". In air it slowly loses water and is changed to a white powder $Na_2CO_3 \cdot H_2O$ (mono hydrated).

vi HYDROLYSIS:-

hydrolysis of alkali metal carbonates produce basic solution

$$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

(b) - ALKALINE EARTH METAL CARBONATES.

i - SOLUBILITY:-

These are slightly soluble in water.

ii - [illegible]

Solubility decreases down the group.

iii - DECOMPOSITION:-

They decompose on heating and ease of decomp decreases down the group.

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2.$$

iv - Ease of decomposition is related to size of metal ion. Smaller the ion more is the lattice energy of resulting oxide and hence greater is the stability of product.

ROLE OF GYPSUM.

(a)- ROLE OF GYPSUM IN AGRICULTURE.

Gypsum ($CaSO_4 \cdot 2H_2O$) is a mineral that occurs in large deposit through out the world.

Gypsum is used as an important source of calcium and sulphur. It is used in fertilizers for calcium and sulphur.

SOURCE OF CALCIUM Gypsum is used as source of calcium. It is added to the soil for better crop product where the soil is subjected to leaching process.

SOURCE OF SULPHER Sulpher is useful for plant for two main purposes. It is used in protein synthesis and for chlorophyll development. Plants turn pale green due to deficiency of sulpher in plants. Gypsum provides sulpher to plants and is helpful in plant production.

DEVELOPMENT OF ROOT SYSTEM:- The application of sulpher greatly enlarges root system of plants. Good crops are produced by using sulpher containing fertilizers such as gypsum.

COMPARE REACTIVITIES OF ALKALI AND ALKALINE EARTH METALS.

Alkali metals are much more reactive than alkaline earth metals. This can be easily explained on the basis of I.P. of these two groups.

The alkali metals have only one electron in their outermost shell, which is loosely held by the nucleus. So alkali metals have lower I.P. The alkaline earth metals have two electrons in outermost shell. As number shells remain same in a period, but nuclear charge increases left to right thus electrons of II-A are relatively tightly bound by nucleus. Further atomic size decreases left to right in a period. Hence II-A group elements have ~~lower~~ higher I.P than I-A group. Thus alkaline earth metals (II A) are less reactive than alkali metals. It is indicated by following examples.

(I-A) ALKALI METALS	II-A. ALKALINE EARTH METALS.
Alkali metals react with oxygen at room temperature. $Na + O_2 \longrightarrow NaO_2$	Alkaline earth metals react with oxygen on heating or on ignition. $Mg + O_2 \xrightarrow{HEAT} 2MgO$
Alkali metals decompose water more vigorously at room temperature. $2K + 2H_2O \longrightarrow 2KOH + H_2$ The reaction is so exothermic that hydrogen liberated catches fire.	The alkaline earth metals decompose water less vigorously and on heating $Mg + 2H_2O \xrightarrow{\Delta} Mg(OH)_2 + H_2$

COMPARE & CONTRAST PHYSICAL & CHEMICAL PROPERTIES OF ALKALI & ALKALINE EARTH METALS.

SIMILARITIES

(1) Alkali and alkaline earth metals have outermost electron is "s" orbital of valence shell. Both are "s" block elements.

(2) The alkali and alkaline earth metals are highly electropositive elements.

(3) The oxide and hydroxides of alkali and alkaline earth elements are alkaline.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

(4) The elements of both groups impart characteristic colour to the bunsen flam.

(5) The elements of both group react with water to produce H_2 gas and hydroxides

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

$$Mg + 2H_2O \xrightarrow{\Delta} 2Mg(OH)_2 + H_2$$

(6) The alkalis and alkaline earth metal elements are obtained by electrolysis of fused chlorides.

(7) The alkali and alkaline earth metals donot occur in free state.

(8) All elements are soft and silvery white metals

(9) All elements of two families are good conductors of heat and electricity.

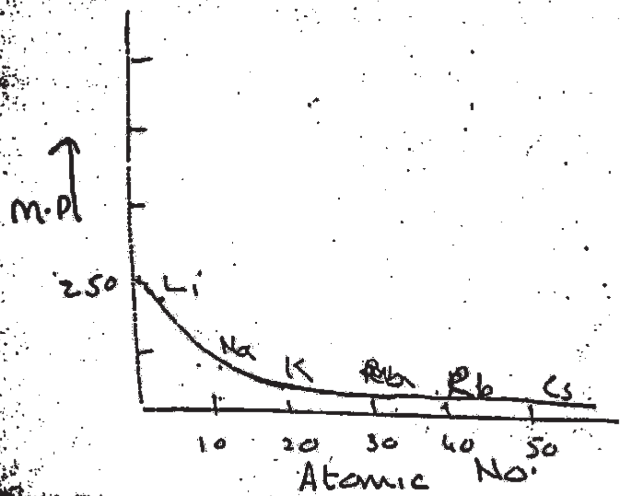
(10) They have similar group trends. i.e m.p. b.p. heat of hydration. heat of sublimation I.P. etc decreases down the group. Atomic size increases down the group.

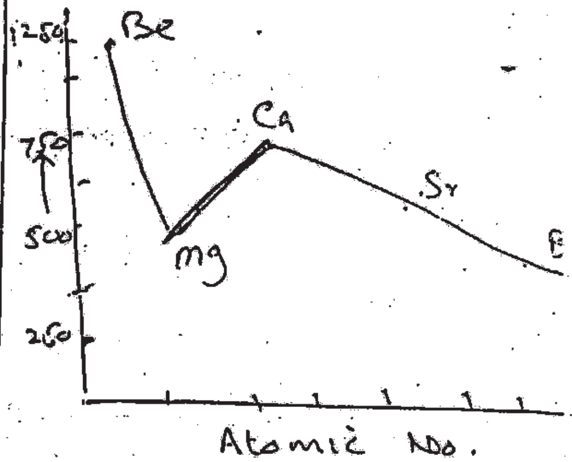
11 The bicarbonates of both families are water soluble and form alkaline solution.

DISSIMILARITIES Alkali and alkaline earth metals differ from each other in following respects

ALKALI METALS	ALKALINE EARTH METALS
(1) These form monovalent cations.	(i) These form divalent cation.
(2) These have one electron in "S" orbital (ns^1 configuration).	These have two electrons in "S" orbital (ns^2 configuration).
(3) The carbonates of alkali metals are water soluble except Li_2CO_3	The carbonates are insoluble in water.
(4) These form peroxides, superoxides and normal oxides, Li_2O, Na_2O_2, KO_2	These form only normal oxides CaO, MgO.
5 The oxides and hydroxides are strongly basic	The oxides and hydroxides are relatively weakly basic.
6 These are lighter than H_2O	These are heavier than H_2O
7 These decompose water vigorously at room temperature. $2K + 2H_2O \longrightarrow 2KOH + H_2$	These decompose water less vigorously (on heating) $Mg + H_2O \longrightarrow MgO + H_2$
8 Their hydroxides are completely soluble in water.	Their hydroxides are sparingly soluble in water.
9 Alkali metals donot form nitrides with Nitrogen.	These form nitrides on heating with nitrogen. $3Mg + N_2 \xrightarrow{\Delta} Mg_3N_2$ $3Ca + N_2 \xrightarrow{\Delta} Ca_3N_2$
10 Alkaline earth metals donot form carbides with Carbon	They form carbides on heating with carbon. $Ca + 2C \xrightarrow{\Delta} CaC_2$
11 Except $LiNO_3$, other alkali metal nitrates decompose to give nitrites and oxygen $2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$ $4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$	The nitrates decompose to give metal oxide, NO_2 & O_2 $2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$
(12) Except Li_2CO_3, other carbonates donot decompose on heating	carbonates decompose on heating $CaCO_3 \rightarrow CaO + CO_2$

(12) The size of alkali metal atom is larges than the corresponding alkaline earth metal atom	They have relatively smaller size.
13 Their I.P. is lower than II-A.	Their I.P. is comparatively higher than I-A.
14 Except Li_3PO_4 other phosphates are water soluble	Phosphates are water insoluble.





CONSTRUCTION OF DOWN'S CELL:-

Down's cell is a circular furnace with outer brick ~~lawg~~ lining. A large block of graphite in the centre of furnace acts as ANODE. A dome above the anode ~~box~~ is used for collection of chlorine gas liberated at anode during electrolysis. CATHODE is a circular bar of iron or copper which surrounds anode. There ~~is~~ is an iron screen in the form of a wire gauze which separates two electrodes. This arrangement avoids contact between sodium and chloride being produced at two electrodes.

There is a sepecial compartment for collection of ~~chlo~~ sodium metal which rises to the surface due to its lesses density.

the following figure.

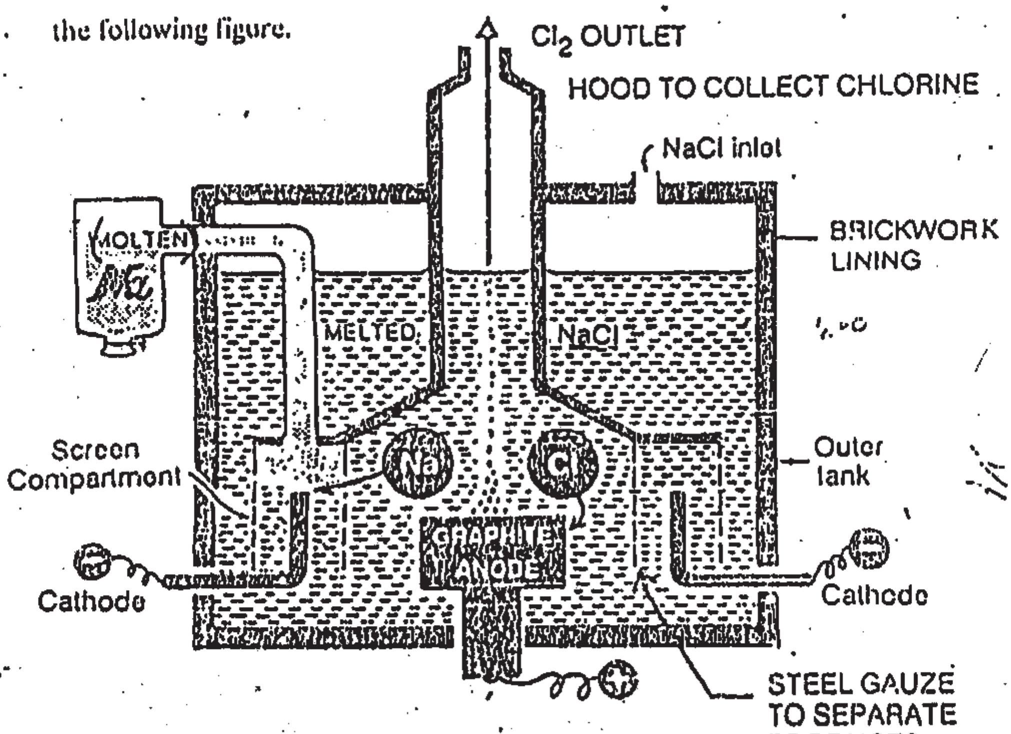


Fig. 3.4 Downs Cell

WORKING :- Molten sodium chloride is added to the furnance. Some calcium chloride is added to it to lower its M.P. from 801°C to 600°C. During electrolysis following reactions take place.

$$2NaCl \xrightleftharpoons{\text{ELECTROLYSIS}} 2Na^{+} + 2Cl^{\ominus}$$

CATHODE REACTION:-

$$2Na^{+} + 2e \longrightarrow 2Na_{(s)}$$

ANODE REACTION:-

$$2Cl^{\ominus} - 2e \longrightarrow Cl_2 \uparrow$$

CELL REACTION:-

$$2NaCl \longrightarrow \cancel{2Na^{+}} + \cancel{2Cl^{\ominus}}$$
$$\cancel{2Na^{\oplus}} + \cancel{2e} \longrightarrow 2Na$$
$$\cancel{2Cl^{-}} - \cancel{2e} \longrightarrow Cl_2$$
$$\overline{2NaCl \longrightarrow 2Na + Cl_2}$$

ADVANTAGES OF DOWN'S PROCESS:-

99.9% pure sodium is obtained.

Liquid sodium can easily be collected at 600°C.

The metallic fog is not formed at 600°C.

Materials of cell are not easily attacked by products of electrolysis.

The side products i.e Chlorine gas is quite useful product.

The raw material i.e NaCl is easily available.

SHOW REACTIONS OF SODIUM WITH FOLLOWING.

(1) **OXYGEN /** ~~ow~~ Sodium reacts with oxygen to form oxides and peroxides.

At room temperature sodium reacts with oxygen to form sodium oxide

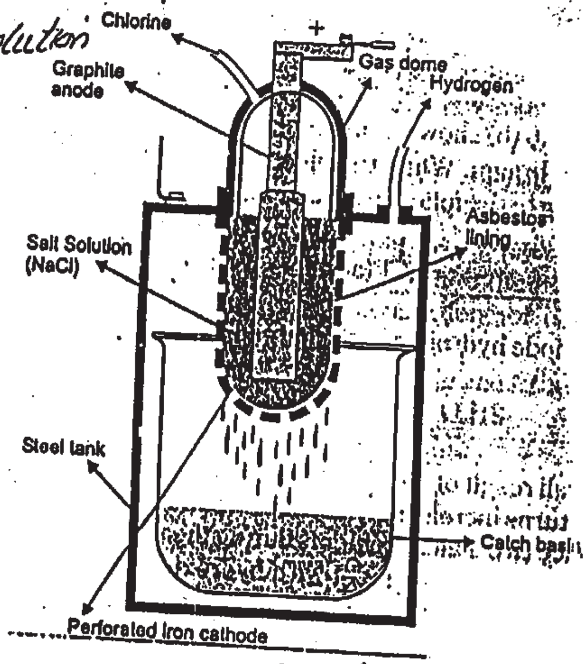
COMMERCIAL METHOD:- There are two commercial methods of preparation of ~~chlorine gas~~ SODIUM HYDROXIDE
(i) ELECTROLYTIC METHOD (ii) DEACON'S PROCESS.

ELECTROLYTIC METHOD (NELSON CELL).
In this method chlorine gas is obtained by electrolysis of aqueous solution of sodium chloride called "BRINE"

CONSTRUCTION OF NELSON CELL

Nelson cell consists of a "U" shaped perforated steel container. It is lined internally by diaphragm. It contains aqueous solution of ~~solution~~ sodium chloride. A graphite anode projects into salt solution. The perforated steel tank acts as cathode (connected to -ve terminal of battery). It is supplied with a constant level device to keep the inner level of solution constant.

The solution seeps out through porous diaphragm. It keeps the steel tank moistened. The "U" shaped tube is fitted in an outer steel tank which is filled with steam. There is a catch basin placed below U-shaped tube to collect NaOH solution.



WORKING OF CELL
Following reactions take place in Nelson Cell.

$$2NaCl \xrightarrow{electrolysis} 2Na^{\oplus} + 2Cl^{\ominus}$$

~~ANODE~~ AT CATHODE

$$2Na^{\oplus} + 2e^{\ominus} \longrightarrow 2Na_{(s)}$$

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

AT ANODE

$$2Cl^{\ominus} - 2e \longrightarrow Cl_2 \uparrow$$

The sodium metal produced at cathode reacts with steam and produce NaOH and H_2 gas

$$2Na + 2H_2O_{(g)} \longrightarrow 2NaOH + H_2\uparrow$$

10-12% NaOH solution is collected at the catch basin. The solution also contains 12-14% NaCl ~~solution~~

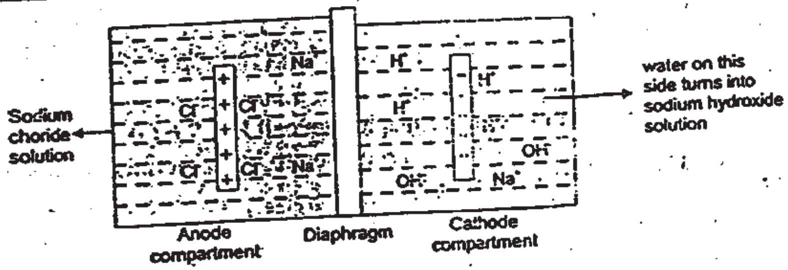


Fig.2.6 (b)

The figure 2.6 (b) shows a simplified version of the cell in order to understand the purpose of diaphragm. When the electrolysis takes place, chlorine is given of at the anode according to the following reaction.

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

At the cathode hydrogen is discharged by the reduction of water.

$$2H_2O + 2e^- \longrightarrow 2OH^-(aq) + H_2(g)$$

The overall result of the above reaction is that the brine loses its chloride ions and the solution turns increasingly alkaline in cathode compartment.

We can face two major problems during the working of the cell.

1. Chlorine produced can react with hydroxide ions in cold giving hypochlorite ions.

$$Cl_2(g) + 2OH^-(aq) \longrightarrow OCl^-(aq) + Cl^-(aq) + H_2O$$

2. Hydroxide ions may be attracted towards anode, where they can be discharged releasing oxygen gas. This oxygen gas may contaminate the chlorine and renders it impure.

The first problem is solved by using asbestos diaphragm. This keeps the two solutions separate while allowing sodium ions to move towards the cathode. This movement of ions keep the current following through the external current.

The second problem is solved keeping the level of brine in anode compartment slightly higher, this keeps the direction of flow of liquid toward the cathode and thus preventing the possibility of hydroxide ions to reach the anode.

The solution that flows out of the cathode compatment contains 11% NaOH

and 16% NaCl. Evaporation of this [illegible]izes the less soluble NaCl which is filtered off, the liquid left conta[illegible] [illegible]2OH and only 1% NaCl as an impurity. For commercial purposes th[illegible] impurity is not important.

2.5 ROLE OF GYPSUM I[illegible]TURE AND INDUSTRY

(a) Role of Gypsum in Agricu[illegible]

Gypsum, a hydrated calciu[illegible] sulphate, is a mineral that occurs in large deposits throughout the world.

Gypsum is applied to the soi[illegible] source of calcium and sulphur. The calcium supplied by gypsum in fertilizers is [illegible] importance in crop production in area where soils are subject to extensive leaching.

Sulphur has been recognised as an essential constituent of plants. For centuries, sulphur compounds had been applied to soils because of their observed beneficial effect on plant growth. Aside from serving as a constituent of protein and various other compounds in plants, sulphur has an influence on chlorophyll development in plant leaves. Although not a constituent of chlorophyll, plants deficient in sulphur exhibits a pale gre[illegible]n colour.

The root system of several plants have been observed to be greatly enlarged by the application of sulphur. It has been reported that good crops are produced by the application of sulphur containing materials such as gypsum.

(b) ROLE OF GYPSUM IN INDUSTRIES

When gypsum is heated und[illegible] carefully controlled conditions, it loses three quarters of water of crystallization. The resulting prod[illegible]ct is called Plaster of Paris. Gypsum must not be heated too stro[illegible]gly as the anhydrous salt is then formed which absorbs water slowly. Such plaster is called 'Dead burnt'.

Plaster of Paris when mixed with half of its weight of water, it forms a plastic type viscous mass and then sets to a hard porous mass. This process is completed within 10 to [illegible]5 minutes. During the process expansion about 1% in volume also occurs, which fills the moulds completely and thus a sharp impression is achieved

Plaster of Paris is used for mak[illegible]ng plaster walls, casts of statuary, coins, etc.

It is used in surgery, Plaster of Paris bandages are used for holding in place fractured bones after they have been set.

Special plasters contain plaster of Paris and other ingredients which vary with the demands of the use to which they are to be put.

Two varieties of plasters are made.

(i) Cement Plaster.

It is plaster of Paris to which usually glue or other oils have been added as retarders to prolong the time of setting.

(2) Hard Finish Plasters

These are made by the calcination of the anhydrous sulphate with alum or borax. These plasters are set very slowly but give a hard finish.

..hen mixed with wood pulp and allowed to set in the form of boards, it forms a material, much used in the construction of buildings as wall boards and partitions.

Gypsum is also used as a filler in paper industries.

Portland cement is made by strongly heating a finely powdered mixture of clay and limestone. The final product, known as clinker, is cooled and then ground into a very fine powder. During the grinding there is added about 2% of gypsum which prevents the cement from hardening too rapidly. The addition of gypsum decreases the setting time of cement.

2.6 ROLE OF LIME IN AGRICULTURE AND INDUSTRY

Lime, (CaO) is a soft, white compound which is obtained by the thermal decomposition of $CaCO_3$

(a) Role of Lime in Agriculture

Large quantities of calcium oxide are used in agriculture for neutralizing acid soils and for preparing sprays.

Calcium oxide is used in large amounts for making lime-sulphur sprays which have a strong fungicidal action. The hydroxide of calcium is obtained when the oxide of the calcium is allowed to react with water. The process is called slaking of lime and it is an exothermic reaction.

$$CaO + H_2O \longrightarrow \underset{\text{Slaked lime}}{Ca(OH)_2}$$

It has been found that application of lime to acidic soils increases the amount of readily soluble phosphorus.

Functions of Calcium in Plant-Growth

The presence of calcium is essential for the normal development of plants. The quantity of calcium required by different plants varies considerably. An adequate supply of calcium appears to stimulate the development of root hairs and, in fact, the entire root system.

Calcium is also necessary for normal leave development and tends to accumulate in leaves as well as in bark. An adequate supply of calcium is also essential for the

optimum activity ... that produce nitrates.

The effect of calci... of available phosphorus in the soil is of special significance. Soils co... calcium are slightly alkaline in nature.

When a deficien... exists various substances such as aluminium and manganese may accumulate in plants in harmful concentrations.

(b) Role of Lime in Industries

1. Large quantities of lime are used in the extraction and refining of metals.
2. Lime is also used in paper industries
3. The ability of lime to react with sand at high temperature forming calcium silicate $CaSiO_3$ serves as an important basis for glass manufacture.
4. Lime is used in ceramic industry for producing different types of sanitary materials.
5. Ordinary mortar, also called lime mortar, is prepared by mixing freshly slaked lime with sand and water to form a thick paste. Mortar is made by mixing slaked lime (one volume) with sand (three or four volumes) and water to make a thick paste. This material when placed between the stones and bricks hardens or sets, thus binding the blocks firmly together. The equations for the chemical reactions that take place when mortar hardens are:

$$CaO + H_2O \longrightarrow Ca(OH)_2$$
$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$
$$Ca(OH)_2 + SiO_2 \longrightarrow CaSiO_3 + H_2O$$

6. Lime is also used in refining of sugar and other food products.
7. Lime is used in the manufacturing of bleaching powder, which is used for the bleaching of the fabric and paper pulp.
8. Lime is also used in leather industry
9. A suspension of the calcium hydroxide is called milk of lime and is used as a white-wash
10. When lime is heated with coke at about 2800°C in an electric furnace, calcium carbide is produced.

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$$CaO + 3C \longrightarrow \underset{\text{Calcium carbide}}{CaC_2} + CO$$

Calcium carbide on hydrolysis produces acetylene.

$$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + \underset{\text{Acetylene}}{C_2H_2}$$

11. Lime is often employed as a dehydrating agent, for example, in the preparation of absolute alcohol and the drying of ammonia gas. A mixture of sodium hydroxide and calcium hydroxide (soda lime) is often employed to remove both water and carbon dioxide from certain gases.

KEY POINTS

1. The elements of group IA except hydrogen are called 'Alkali metals' while those of group IIA are named as alkaline earth metals.
2. Alkali metals have only one electron in s-orbital of their valence shell. They lose one electron of the valence shell forming monovalent positive ions.
3. Alkaline earth metals have two electrons in s-orbital of their valence shell. They lose two electrons forming dipositive ions M^{++}.
4. Spodumene, petalite, halite, natron, alinite are the common minerals of alkali metals.
5. Beryl, magnesite, dolomite, epsom salt, asbestos, calcite, gypsum and barite are the important minerals of alkaline earth metals.
6. Lithium behaves different form the other alkali metals.
7. Lithium forms only normal oxide, whereas the others form higher oxides like peroxides and superoxides.
8. Beryllium is the only member of group II, which reacts with alkalies to give hydrogen. The other member do not react with alkalies.
9. Nitrides hydrolyse vigirously when treated with water, giving NH_3 and respective hydroxides.
10. Nitrates of lithium, magnesium and barium on heating give oxygen, nitrogen peroxide and the corresponding metallic oxides.

11. When gypsum is heated above 100°C, it loses three quarters of its water of crystallization, giving white powder of $CaSO_4.1/2H_2O$ which is called Plaster of Paris.
12. Sodium is prepared by the electrolysis of molten sodium chloride in Down's cell.
13. Calcium is necessary for development of leaves and it tends to accumulate in leaves and bark. An adequate quantity of calcium is essential for the optimum activity of microorganisms that produce nitrates.
14. Lime is used in paper and glass industries. It is also used for refining sugar products.

EXERCISE

Fill in the blanks:

(i) Alkali metals are MORE reactive than alkaline-earth metals.
(ii) Alkali metals decompose water vigorously producing HYDROXIDE and hydrogen.
(iii) When heated in a current of dry hydrogen, alkaline earth metals form white crystalline ionic HYDRIDE of the type MH_2.
(iv) The beryllium hydroxide, like the hydroxide of aluminium is amphoteric, while the hydroxides of the other members of the group IIA are ALKALINE.
(v) The elements of the group IA are termed as alkali metals, because their OXIDES are alkaline.
(vi) Spodumene is an ore of Li metal.
(vii) Alkali metal nitrates on heating give the corresponding NITRITE and oxygen.
(viii) $Na_2CO_3.H_2O$ is the chemical formula of an ore of sodium which is known as NATRON
(ix) Metallic bicarbonates are decomposed on heating into their carbonates, alongwith CO_2 and H_2O.
(x) Metal nitrates other than the alkali metals on heating decompose into the corresponding metal OXIDE alongwith the evolution of nitrogen peroxide and oxygen.

Q2. Indicate True or False.

(i) Group IA elements are called alkali metals because their chlorides are