Chapter # 02
s-Block Elements

2.1 INTRODUCTION

S-BLOCK ELEMENTS

The elements in which s-orbitals are being filled in their outermost shell are called s-block elements.

Generally, elements of group IA and IIA are called **s**-block elements.

ALKALI METALS

Elements of group IA are called alkali metals

Origin of name

Word alkali is an Arabic word means 'ashes'. This word was used by Arabs because they found that ashes of plant mainly consist of Na and K.

Alkali metals include Li, Na, K, Rb, Cs and Fr. They produce strong alkaline solution with water.

ALKALINE EARTH METALS

Elements of group IIA are called Alkaline earth metals

Origin of name

Alkaline earth metals are given this name because they form alkalies in water and are widely distributed in earth crust. These include Be, Mg, Ca, Sr, Ba & Ra.

Both alkali and alkaline earth metals are highly reactive

2.1.1 Electronic Configuration of S-Block Elements

ALKALI METALS

- *These have* one electron in 's' orbital of their valence shell.
- Their ionization energies are very low. Thus they form monopositive ions (M^+) by losing one electron from their valence shell.
- They form ionic compounds.
- They show +1 oxidation state.

Properties	Li	Na	K	Rb	Cs
Atomic number	3	11	19	37	55
Electronic configurations	$1s^2 2s^1$	$[Ne]3s^1$	$[Ar]4s^1$	$[Kr]5s^1$	$[Xe]6s^1$
Ionization energy (kJ/mol)	520	496	419	4.3	376
Electron affinity (kJ/mol)	–60	–53	–48	–47	–46
Electronegativity	1.0	0.9	0.8	0.8	0.7
Atomic radius	123	158	203	216	235
Ionic radius of 1+ ion (pm)	60	95	133	148	169
Melting points (^{0}C)	187.0	97.5	63.6	39.0	28.5
Boiling points (^{0}C)	1325	889	774	688	690
Density gm/cm^3 at (20^{0}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.
- Their ionization energies are low. Thus they form dipositive ions (M^{++}) by loosing two electrons from their valence shell.
- They form ionic compounds.
- They show +2 oxidation state.

Properties	Be	Mg	Ca	Sr	Ba
Atomic number	4	12	20	38	56
Electronic configurations	$1s^2 2s^2$	$[Ne]3s^2$	$[Ar]4s^2$	$[Kr]5s^2$	$[Xe]6s^2$
Ionization energy (kJ/mol)	899	738	590	549	503
Electron affinity (kJ/mol)	240	230	156	168	52
Electronegativity	1.5	1.2	1.0	1.0	0.9
Atomic radius	89	136	174	191	198
Ionic radius of 2+ ion (pm)	31	65	99	113	135
Melting points (^{0}C)	1289	649	839	769	725
Boiling points (^{0}C)	2970	1107	1484	1384	1640
Density gm/cm^3 (20^{0}C)	1.85	1.74	1.55	2.6	3.5
Heat of hydration (kJ/mole)	2337	1897	1619	1455	1250

Down the group number of shells increases. The number of shells is equal to the period number of element.

2.1.2 Occurrence of Alkali Metals

Alkali metals are highly reactive. Therefore they do not occur free in nature. They are found in combined state.

Most of earth crust is made up of insoluble alumino silicates of alkali metals.

- Na and K are more abundant and constitute about 2.4% of earth crust.
- Li is found as complex minerals which are widely distributed. An important commercial source of Li is the mineral spodumene i.e $LiAl(SiO_3)_2$
- Small amounts of Rb and Cs are found in K salts deposits.
- Fr does not occur in nature. It is radioactive element. It is produced in laboratory by nuclear reaction. It is very unstable, therefore, its chemistry is not well known.

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
Alunite (Alum Stone)	$K_2SO_4.Al_2(SO_4)_3.4Al(OH)_3$

2.1.3 Occurrence Alkaline Earth Metals

These are very reactive. Therefore, they do not occur free in nature. They are found in combined state. Their compounds occur widely in nature.

- Mg and Ca are very abundant in earth crust. Outer portion of the earth was originally in the form of silicates and aluminosilicates of alkaline earth metals.
- Ca and Mg along with Na and K are present in rocks as cations.
- Magnesium halides are found in sea water. Mg is essential part of chlorophyll.
- Calcium phosphate $Ca_3(PO_4)_2$ and calcium fluoride CaF_2 are also found as mineral.
- Ca is an important constituent of living organisms. It is found in bones, teeth, sea shells and egg shells.
- Ra is a rare element. It is highly radioactive.

Name of Mineral	Chemical Formula
Beryllium	
Beryl	$Be_3Al_2(SiO_3)_6$
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	$CaCO_4.2H_2O$
Fluorite	CaF_2
Phosphorite	$Ca_3(PO_4)_2$
Strontium	
Strotionite	$SrCO_3$
Barium	
Barite	$BaSO_4$

2.1.4 PECULIAR BEHAVIOUR OF LITHIUM

(Anomalous behaviour of Li) OR (Differences of Li from its family members)

First member of each main group of periodic table does not follow the regular trends of the Group.

Similarly Li also shows difference from its family members. This behaviour is due to following reasons

- Both Li and Li^+ ions have very small size.
- Li^+ has high charge density.
- Li has less electropositivity than sodium.

DIFFERENCES

1. Li is much harder and lighter than the other alkali metals
2. Many Li salts are insoluble in water especially which have anions with high charge density. e.g. LiOH, LiF, Li_3PO_4, Li_2CO_3. The corresponding salts of other alkali metals are soluble in water.
3. Li can form stable complexes while other alkali metals do not have this property. e.g Li form $Li(NH_3)_4$.
4. Li reacts with water very slowly while other metals react violently.
5. Li salts with large palarizable ions are less stable than other alkali metals. Thus Li does not form bicarbonate, tri-iodide or hydrogen sulphide at room temperature, while other alkali metals form these compounds.
6. Li form normal oxide on burning in air while other alkali metals form superoxides and peroxides.

$$4Li + 2O_2 \longrightarrow 2Li_2O$$
$$2Na + O_2 \longrightarrow Na_2O_2$$
$$K + O_2 \longrightarrow KO_2$$

7. Lithium hydride is more stable than hydrides of other alkali metals
8. Li compounds are more covalent in nature. Thus its halides are more soluble in organic solvents. Also its alkyls and aryls are more stable than those of other alkali metals.
9. Li is the least reactive of all the alkali metals
10. Li does not form acetylides when acetylene is passed over heated Li, while other alkali metals form metallic acetylides.

$$2Na + C_2H_2 \longrightarrow \underset{\text{(Sodium acetylide)}}{Na^+C^-{\equiv}C^-Na^+} + H_2$$

11. *Li is* less electropositive than other alkali metals. Thus its carbonate decomposes to give lithium oxide while carbonates of other alkali metals are stable

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

12. Nitrate of Li also decomposes but it gives different product than the decomposition of nitrates of other alkali metals.

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$
$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

13. Lithium hydroxide decomposes on strong heating to give lithium oxide, while other alkali metals do not show this behaviour.

$$2LiOH \xrightarrow{\text{red hot}} Li_2O + H_2O$$

14. Lithium reacts with nitrogen directly to give lithium nitride while other alkali metals do not give this reaction.

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

15. LiCl has exothermic heat of solution while chlorides of Na and K have endothermic heats of solution

16. Li can form lithium carbide readily by direct reaction while other alkali metals do not give this reaction.

$$4Li + C \longrightarrow Li_4C$$

2.1.5 Peculiar Behaviour Of Beryllium

(Anomalous behaviour of Be) OR (Differences off Be from its family members)

Be shows different behaviour than its family members

It is due to following reasons

- Size of Be is very small, and
- Electronegativity of Be is quite high than its family members.

Differences

1. Be is as hard as iron. It is quite hard to scratch glass, while other alkaline earth metals are softer than Be but still harder than alkali metals.

2. Melting points and boiling points of Be are higher than other alkaline earth metals

3. In principal all group IIA metals can reduce water. But Be forms insoluble oxide coating over the surface that protects it from further attack.

4. Be is more resistant to complete oxidation as compare to its family members, because of its BeO coating.

5. Only Be reacts with alkalies to give H_2 gas while other alkaline earth metals do not give this reaction.

$$Be + 2NaOH \longrightarrow \underset{\text{(Sodium beryllate)}}{Na_2BeO_2} + H_2$$

6. Be mostly forms covalent compounds while other alklaline earth metals form ionic compounds

7. Oxides and hydroxides of Be are amphoteric while other alkaline earth metals form basic oxides and hydroxides

2.2 GENERAL BEHAVIOUR OF ALKALI METALS

2.2.1 TRENDS IN CHEMICAL PROPERTIES OF ALKALI METALS:

1. REACTIVITY

Alkali metals have very low ionization energies. Thus they are very reactive.

2. REDUCING PROPERTY

Alkali metals can easily loose electrons i.e. they are easily oxidised. Thus they can reduce substances and act as excellent reducing agents.

3. OXIDATION STATE

They have high second I.E. Therefore, second electron can not be removed easily. Hence they do not show oxidation state greater than +1.

4. SOLUBILITIES AND LATTICE ENERGIES

Cations of alkali metals have low charge and large radii than the radius of any cation from the same period. Thus lattice energies of their salts are low. Hence, most of the simple salts of alkali metals are water-soluble. Most of the salts are ionized completely in aqueous solution.

5. HYDROXIDES

Hydroxides of alkali metals are among the strongest bases. These are generally hygroscopic.

6. REACTION WITH OXYGEN

- They react with oxygen to form oxides and their surface is tarnished.
- Only lithium burns in air to form the normal oxide, Li_2O (white solid).

$$4Li + O_2 \longrightarrow 2Li_2O$$ *White solid (Lithium oxide)*

The exposed metal surfaces are oxidized immediately by oxygen in air, and in the presence of moisture. The oxides formed react with CO_2 in the atmosphere to form carbonates.

$$\underset{\text{Lithium oxide}}{Li_2O + CO_2} \longrightarrow \underset{\text{Lithium carbonate}}{Li_2CO_3}$$

- Sodium gives similar reaction only in limited supply of oxygen.

$$4Na + O_2 \longrightarrow 2Na_2O$$

In excess oxygen, sodium forms pale yellow sodium peroxide.

$$2Na + O_2 \longrightarrow Na_2O_2$$ *Sodium peroxide (Pale yellow)*

- Potassium, rubidium and caesium react with oxygen to form super oxides.

$$K + O_2 \longrightarrow KO_2$$ *Potassium superoxide (Orange yellow)*

$$Rb + O_2 \longrightarrow RbO_2$$ *Rubidium superoxide*

- Caesium explodes spontaneously when it is in contact with air or oxygen.

$$Cs + O_2 \longrightarrow CsO_2$$ *(Ceasium superoxide)*

7. REACTION WITH WATER

- Alkali metals react with water rapidly.
- A small piece of sodium or potassium or lithium floats on water and it reacts vigorously with water to liberate hydrogen and form metal hydroxide. The reaction is highly exothermic. The heat released may ignite the hydrogen.

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

- The reaction becomes increasingly vigorous form lithium to caesium. Potassium, rubidium and caesium are so highly reactive that they react with ice even at -100^0C.

8. Reaction with H_2

- Alkali metals react with hydrogen to form ionic hydrides.

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

- Rubidium and caesium react violently with hydrogen at room temperature.
- Other three metals react with hydrogen at high temperature to form hydrides.

$$2Li + H_2 \longrightarrow 2LiH$$
$$2Na + H_2 \longrightarrow 2NaH$$
$$2K + H_2 \longrightarrow 2KH$$

- Lithium hydride produces hydrogen when treated with water.

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

Thus ionic hydrides are used as powerful reducing agents due to presence of hydride ion (H^-).

9. REACTION WITH NITROGEN

Only Li of alkali metals combines with nitrogen to form nitride.

$$6Li + N_2 \longrightarrow 2Li_3N \text{ (Lithium nitride)}$$

10. REACTION WITH CARBON

Only Li of alkali metals combines with carbon to form carbide.

$$4Li + C \longrightarrow Li_4C \text{ (Lithium carbide)}$$

11. REACTION WITH HALOGENS

- Alkali metals are highly electropositive. They react easily with halogens to give halides.
- Lithium and sodium react slowly with chlorine at room temperature.
- Molten sodium burns with brilliant yellow flame in a chlorine atmosphere to form sodium chloride.

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

- Potassium, rubidium and caesium react vigorously with all halogens to give metal halides.

12. REACTION WITH S

All alkali metals give sulphides when treated with molten sulphur.

$$2M + S \longrightarrow M_2S$$
$$2Li + S \longrightarrow Li_2S$$
$$2Na + S \longrightarrow Na_2S$$

13. REACTION WITH ACIDS

Reaction of alkali metals with acids is violent. The liberated hydrogen may ignite.

2.2.2 TRENDS IN CHEMICAL PROPERTIES OF ALKALINE EARTH METALS:

1. REACTIVITY

- Alkaline earth metals are very reactive elements due to low ionization energies.
- The reactivity increases from top to bottom.
- However alkaline earth metals are less reactive than alkali metals.

2. REACTION WITH OXYGEN

- Alkaline earth metals burn in oxygen to form oxides. Barium forms peroxide.
- Be is the least reactive alkaline earth metals. It is resistant to complete oxidation due to formation of oxide layer over the surface. However it is completely oxidized at 800°C.

$$2Be + O_2 \xrightarrow{800^\circ C} 2BeO$$

Due to less reactivity Be is not tarnished by atmospheric attack. However metal soon loses the silvery appearance.

- Magnesium quickly forms coating of MgO when exposed to atmosphere. This layer protects the surface from further corrosion at ordinary temperature.

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

When magnesium is burnt in air a small amount of nitride is also formed alongwith magnesium oxide

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

- Barium on heating in air or oxygen at 500 – 600°C forms peroxide.

$$Ba + O_2 \xrightarrow{500-600\,C} BaO_2$$

Barium peroxide

3. REACTION WITH HYDROGEN

- When molten alkaline earth metals are reacted with hydrogen, hydrides are produced. The reaction usually occurs under high pressure.
- Magnesium reacts with hydrogen at high pressure and in the presence of a catalyst (MgI_2) forming magnesium hydride.

$$Mg + H_2 \xrightarrow[MgI_2]{Pressure} MgH_2$$

$$Ca + H_2 \longrightarrow CaH_2$$

4. **REACTION WITH NITROGEN**

- All alkaline earth metals react with nitrogen on heating to give nitrides.

$$3Mg + N_2 \longrightarrow Mg_3N_2 \text{ (Magnesium nitride)}$$
$$3Ca + N_2 \longrightarrow Ca_3N_2$$

- The nitrides hydrolyse vigorously with water to give ammonia and respective hydroxide.

$$Mg_3N_2 + 6H_2O \longrightarrow 2NH_3 + 3Mg(OH)_2$$
$$Ca_3N_2 + 6H_2O \longrightarrow 2NH_3 + 3Ca(OH)_2$$

5. **REACTION WITH S**

Alkaline earth metals react with S to give sulphides. Magnesium gives magnesium sulphide with S.

$$Mg + S \longrightarrow MgS \quad \text{(Magnesium sulphide)}$$

6. **REACTION WITH HALOGENS**

All alkaline earth metals react directly with halogens to give halides of the type MX_2,

$$Ca + Cl_2 \longrightarrow CaCl_2$$

7. **REACTION WITH WATER**

- Magnesium is more reactive than beryllium although it is not attacked by cold water. Magnesium reacts slowly with boiling water. It reacts rapidly with steam to liberate hydrogen.

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

- *Beryllium does not react with water* even at red hot temperature.
- Remaining alkaline earth metals produce hydroxide with water.

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

8. **CARBONATES**

Carbonates of alkaline earth metals are insoluble in water and decompose on heating.

$$CaCO_3 \longrightarrow CaO + CO_2$$
$$MgCO_3 \longrightarrow MgO + CO_2$$
$$BaCO_3 \longrightarrow BaO + CO_2$$

2.2.3 GENERAL TRENDS IN PROPERTIES OF COMPOUNDS OF ALKALI AND ALKALINE EARTH METALS:

(I) OXIDES

ALKALI METALS

- Alkali metals oxides dissolve in water to give strongly alkaline solution.

$$Li_2O + H_2O \longrightarrow 2LiOH$$

- Super oxides and peroxides also give O_2 in addition to alkaline solution.

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

Reaction of an alkali metal oxide with water is an acid-base reaction and not an oxidation reduction reaction because oxidation number of elements is not changed. The reaction simply involves the decomposition of water molecule by an oxide ion.

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

- Basic character of alkali metal oxides increases down the group.
- Potassium superoxide (KO_2) is used in breathing equipments for mountaineers and in space craft. It can absorb carbon dioxide while giving out oxygen at the same time.

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

ALKALINE EARTH METALS

- Solubility of alkaline earth metal oxides in water increases down the group.
- BeO and MgO are insoluble.
- CaO, SrO and BaO are soluble and react with water to form hydroxides.

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

- Basic character of the oxides of alkaline earth metals increases down the group.
- Group IIA oxides are less basic than group IA oxides.
- BeO is amphoteric in nature since it reacts with both acids and bases.

$$BeO + H_2SO_4 \longrightarrow BeSO_4 + H_2O$$
$$BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O$$

Sodium beryllate

Group I-A	Group II-A
Lithium oxide (Basic)	Beyllium oxide (Basic)
Sodium oxide (Basic)	Magensium oxide (Basic)
Potassium oxide (Basic)	Calcium oxide (Basic)

↓ Increasing basicity of oxides

← Increasing basicity of oxides

(II) HYDROXIDES

ALKALI METALS

- Alkali metal hydroxides are all crystalline solids.
- These are very soluble in water except LiOH, which is slightly soluble.
- They are generally hygroscopic.
- These are very strong bases except LiOH.
-

ALKALINE EARTH METALS

- Solubility of alkaline earth metal hydroxides in water increases down the group.
  - ✓ $Be(OH)_2$ is quite insoluble.
  - ✓ $Mg(OH)_2$ is sparingly soluble.
  - ✓ $Ba(OH)_2$ is more soluble.

  This increase in solubility is due to low lattice energy of hydroxides which, in turn, is due to higher ionic size.
- Alkali metal hydroxides are stable to heat except LiOH, while alkaline earth metals hydroxide like $Mg(OH)_2$ and $Ca(OH)_2$ decompose on heating.

$$2LiOH \longrightarrow Li_2O + H_2O$$
$$Mg(OH)_2 \longrightarrow MgO + H_2O$$

- A saturated solution of $Ca(OH)_2$ in water is called lime water and is used as a test for CO_2.

$$Ca(OH)_2 + CO_2 \longrightarrow \underset{\textit{(white ppt.)}}{CaCO_3} + H_2O$$

- A suspension of $Mg(OH)_2$in water is called milk of magnesia and it is used for treatment of acidity in stomach.

(III) CARBONATES

ALKALI METALS

- Carbonates of alkali metals are all soluble in water except Li_2CO_3.
- These are stable on heating except Li_2CO_3.
- Li_2CO_3 is insoluble in water and also decompose on heating to give lithium oxide. The decomposition is made easy because the electrostatic attraction in converting from carbonate to oxide is considerable. In case of large cation like K^+ in K_2CO_3, the gain in electrostatic attraction is relatively much less and the decomposition is difficult.
- Sodium carbonate is very important industrial chemical.
  - ✓ Below 35.2°C, Na_2CO_3 crystallizes out from water as $Na_2CO_3.10H_2O$ which is called washing soda.
  - ✓ Above this temperature it crystallizes as $Na_2CO_3.H_2O$.
  - ✓ On standing in air, $Na_2CO_3 .10H_2O$ slowly loses water and converted to a white powder $Na_2CO_3. H_2O$.
- The solution of Na_2CO_3 in water is basic due to hydrolysis of carbonate ion.

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

Alkaline Earth Metals

- Alkaline earth metal carbonates are only slightly soluble in water.
- The solubility decreasing down the group.
- They decompose on heating.

$$CaCO_3 \longrightarrow CaO + CO_2$$
$$MgCO_3 \longrightarrow MgO + CO_2$$
$$BaCO_3 \longrightarrow BaO + CO_2$$

- The ease of decomposition decreases down the group. The ease of decomposition is due to the size of the metal ion. Smaller the ion, more is the lattice energy of the resulting oxide and hence higher the stability of the products.

(IV) Nitrates

- Nitrates of both alkali and alkaline earth metals are soluble in water.
- Nitrates of Li, Mg, Ca and Ba completely decompose on heating to give O_2, NO_2 and the metallic oxide.

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$
$$2Mg(NO_3)_2 \longrightarrow 2MgO + 4NO_2 + O_2$$
$$2Ca(NO_3)_2 \longrightarrow 2CaO + 4NO_2 + O_2$$

- Nitrates of Na and K partially decompose to give different products i.e. nitrites and oxygen.

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

(V) Sulphates

- All the alkali metals sulphates are soluble in water.
- Solubility of sulphates of alkaline earth metals gradually decrease down the group.
  - ✓ $BeSO_4$ and $MgSO_4$ are fairly soluble in water.
  - ✓ $CaSO_4$ is slightly soluble.
  - ✓ $SrSO_4$ and $BaSO_4$ are almost insoluble.
- Calcium sulphate occurs in nature as gypsum $CaSO_4 . 2H_2O$.

  On heating above 100ºC, it loses ¾ of its water of crystallization, to give white powder called Plaster of Paris.

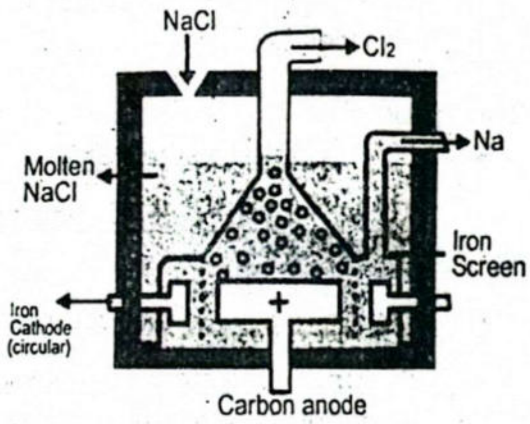
$$\underset{\textit{Gypsum}}{2CaSO_4 . 2H_2O} \longrightarrow \underset{\textit{Plaster of Paris}}{(CaSO_4)_2 . H_2O} + 3H_2O$$

2.3 COMMERCIAL PREPARATION OF SODIUM BY DOWN'S CELL

Most of the Na is prepared by the electrolysis of fused NaCl in Down's cell.

CONSTRUCTION

- Down cell was designed by J.C. Down. It is shown in the fig.
- It consists of a circular furnance. There is a large block of graphite in the centre which acts as anode. Above it is a dome for the collection of chlorine.
- A circular bar of copper surrounds the anode. This bar acts as cathode. It is separated from anode by an iron screen which ends in a gauze.
- The cell is so designed, that electric current easily pass through it. But Na and Cl_2 produced make no contact with each other.



Fig(2.1) Down's cell.

WORKING

- NaCl is fed into the cell. Some $CaCl_2$ is also added to lower the m.p. of NaCl from 801°C to 600°C. The cell is operated at this temperature.
- When electricity is passed, Na rises in a special compartment from where it is withdrawn from time to time.
- Dry Cl_2 gas is collected above anode.
- The process is carried out at 600°C which has following adavantages
  - ✓ The metallic fog is not produced
  - ✓ Liquid sodium can be easily collected at 600°C.
  - ✓ Products formed do not react with material of the cell.

REACTIONS

$$NaCl \longrightarrow Na^{+} + Cl^{-}$$

At Anode (oxidation)

$$Cl^{-} \longrightarrow \frac{1}{2} Cl_2 + 1e^{-}$$

At Cathode (reduction)

$$Na^{+} + e^{-} \longrightarrow Na$$

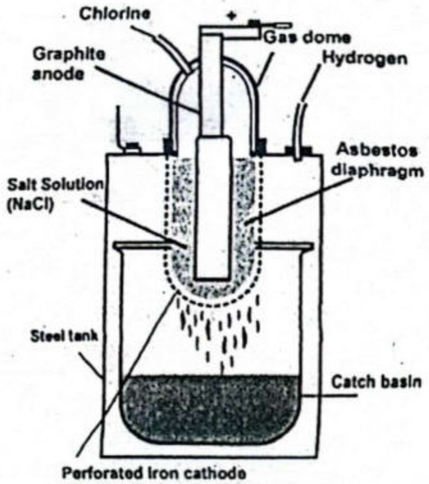
Net Reaction $$NaCl \longrightarrow Na + \frac{1}{2} Cl_2$$

A conc. solution of NaCl is called brine.

In industry NaOH is prepared by electrolysis of brine in diaphragm cell called Nelson's cell.

CONSTRUCTION

- It is an oblong steel tank.
- There is a U-shaped diaphragm in it which holds the salt solution.
- Graphite anode project into the salt solution.
- A current of steam is passed through the cell to keep the salt solution warm and keep perforation open.
- Cathode is made up of oblong perforated steel internally lined with asbestos diaphragm.
- There is a special device, which keeps constant level of salt solution in the cell.



Fig(2.2) Nelson cell for the production of NaOH.

WORKING

- Brine is introduced into the cell and its constant level is maintained during the process.
- When electricity is passed Cl_2 rises into the dome at the top of the anode while H_2 is liberated at cathode and escapes through a pipe, where they are collected.
- NaOH solution slowly moves downward through the diaphragm and is collected in a catch basin below diaphragm.

REACTIONS

$$2NaCl_{(s)} \longrightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)}$$

At Anode (oxidation)

$$2Cl^{-}_{(aq)} \longrightarrow Cl_{2(g)} + 2e^{-}$$

At Cathode (reduction)

$$2H_2O_{(l)} + 2e^{-} \longrightarrow 2OH^{-}_{(aq)} + H_{2(g)}$$

Net Reaction $$2NaCl + 2H_2O \longrightarrow 2Na^{+}_{(aq)} + 2OH^{-}_{(aq)} + Cl_2 + H_{2(g)}$$

PROBLEMS

Two problems may occur during the process.

- **First Problem**

Cl⁻ produced can react with OH⁻ ions to give OCl⁻ (hypochlorite) ions.

$$Cl_{2(g)} + 2OH^-_{(aq)} \longrightarrow OCl^-_{(aq)} + Cl^-_{(aq)} + H_2O$$

This problem can be removed by using asbestos diaphragm which keeps the two solutions separate and allow sodium ions to move to cathode. The movement of ions keeps the current flowing through the external circuit.

- **Second Problem**

OH^- may be attracted to anode where it can discharge to give O_2 gas which makes the liberated Cl_2 impure.

$$4OH^- \longrightarrow 2H_2O + O_2 + 4e^-$$

It is removed by keeping the level of brine in anode compartment slightly higher. This keeps the direction of flow of liquid toward the cathode. Thus it prevents the movement of OH^- towards anode.

COLLECTION OF NaOH

Solution obtained from cathode contains 11% NaOH and 16% NaCl. This solution is evaporated and less soluble NaCl crystallizes out which is filtered off. The remaining liquid contains 50% NaOH and only 1% NaCl which is not harmful for commercial use.

2.5 ROLE OF GYPSUM IN AGRICULTURE AND INDUSTRY

Chemical formula of Gypsum is $CaSO_4.2H_2O$. This mineral is widely deposits throughout the world.

ROLE OF GYPSUM IN AGRICULTURE

1. Gypsum is an important source of Ca and S.
2. In fertilizers, gypsum supplied Ca to the soil for crop production. It is particularly important where soils are subject to extensive leaching.
3. S is also an important element for plants growth.
4. S is an important constituent of protein.
5. S has some influence on chlorophyll development in plant leaves. It is not part of chlorophyll. But plants with deficiency of S shows pale green colour.
6. Root system of plants is greatly enlarged by the application of S.

In general good crops are produced by the application of sulpher containing materials e.g. gypsum.

ROLE OF GYPSUM IN INDUSTRY

1. PLASTER OF PARIS

When gypsum is heated carefully, it looses ¾th (3 quarters) of its water of crystallization. The resulting product is called plaster of paris.

$$\underset{\text{Gypsum}}{2CaSO_4 . 2H_2O} \longrightarrow \underset{\text{Plaster of Paris}}{(CaSO_4)_2 . H_2O} + 3H_2O$$

However, if gypsum is heated strongly, then anhydrous product is obtained called dead burnt gypsum.

$$CaSO_4 . 2H_2O \longrightarrow CaSO_4 + 2H_2O$$

It absorbs water very slowly.

Setting of Plaster Of Paris

When plaster of paris is mixed with half of its weight of water, it forms a plastic type viscous mass which then sets to a hard porous mass. This process completes in 10 to 15 minutes. During this process, solid mass expand 1% by volume and fills the mould completely. Thus it gives a sharp impression.

Uses of Plaster of Paris

It is used

- for making plaster walls.
- for casts of statuary, coins etc
- in surgery. Its bandages are used to keep the fractured bone in place after setting.

2. SPECIAL PLASTERS

Special plasters are made by adding certain ingredients

Cement Plasters

It is plaster of paris in which glue or other oils have been added. These acts as retarders and increases the setting time.

Hard Finish Plasters

- It is produced by calcination of anhydrous sulphate with alum or borax.
- These plasters set very slowly but give hard finish.
- When plaster of paris is mixed with wood pulp and allowed to set in the form of boards, then a material is obtained which is used in the construction of buildings as wall boards and partitions.

3. Gypsum is also used as filler in paper industry.

4. Gypsum is also used in cement industry.

Portland cement is obtained by heating a mixture of clay and lime stone. The resulting mass called clinker is cooled and crushed to fine powder. During crushing about 2% gypsum is also added. It prevents the cement from hardening quickly and increases the setting time of cement so that it can be used easily.

2.6 ROLE OF LIME IN AGRICULTURE AND INDUSTRY

Formula of lime is CaO. It is soft, white compound obtained by decomposition of $CaCO_3$

Slaked lime

When lime is mixed with water it form calcium hydroxide called slaked lime. This process is called slaking of lime. It is an exothermic process.

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

ROLE OF LIME IN AGRICULTURE

1. CaO is used for neutralizing acidic soils.
2. Application of lime to acidic soils increases the amount of readily soluble phosphorous.
3. CaO is used for making lime sulpher sprays which have a strong fungicidal action.

Function of Ca in plant growth

1. Ca is essential for normal growth of plant. Different plants require different amounts of Ca.
2. Sufficient supply of Ca stimulates the development of root hairs and entire root system.
3. Ca is also necessary for normal development of leaves. Sufficient Ca accumulates in leaves and bark.
4. Sufficient supply of Ca is needed for optimum activity of micro-organisms that produce nitrates.
5. Presence of Ca in soil controls the availability of phosphorous in the soil.
6. Soils containing Ca are alkaline in nature.
7. With deficiency of Ca many substances like Al, Mn etc. accumulates in plants. Their higher concentrations are harmful to plants.

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. It is used in glass making. Lime reacts with sand at high temperature forming calcium silicate $CaSiO_3$ which is the 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), sand (three or four volumes) and water to make thick paste. This material is hardened or sets when placed between stones and bricks. Thus it binds the blocks firmly together.

   The reactions for this process 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.

$$CaO + 3C \longrightarrow CaC_2 + CO$$

   Calcium carbide on hydrolysis produces acetylene.

$$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$$

11. Lime is often used as a dehydrating agent. e.g. rectified spirit contains 5% water which is removed by distilling with lime to produce absolute alcohol. It is also used in drying of ammonia gas.
12. A mixture of sodium hydroxide and calcium hydroxide (soda lime) is often used to remove both water and carbon dioxide from certain gases.

OBJECTIVE AND SHORT ANSWER, QUESTIONS (EXERCISE)

Q.1 Fill in the blanks.

(i) Alkali metals are ______ reactive than alkaline-earth metals.

(ii) Alkali metals decompose in water vigorously producing __________ and hydrogen.

(iii) When heated in a current of dry hydrogen, alkaline earth metals form white crystalline_______of the type MH_2.

(iv) The beryllium hydroxide, like the hydroxide of aluminum is amphoteric, while the hydroxides of the other members of the group IIA are ______.

(v) The elements of the group IA are termed as alkali metals, because their _____________ are alkaline.

(vi) Spodumene is an ore of _______ metal.

(vii) Alkali metal nitrates on heating give the corresponding ______ and oxygen.

(viii) $Na_2CO_3.H_2O$ is the chemical formula of an ore of sodium, which is known as __________.

(ix) Metallic bicarbonates are decomposed on heating into their carbonates, alongwith ______ and __________.

(x) Metal nitrates other than the alkali metals on heating decompose in to the corresponding metal _______ along with the evolution of nitrogen peroxide and oxygen.

Answers:

(i) More	(ii) hydroxides	(iii) solid	(iv) basic	(v) hydroxides
(vi) lithium	(vii) nitrite	(viii) natron	(ix) CO_2 , H_2O	(x) oxides

Q.2 Indicate True or False.

(i) Group IA elements are called alkali metals because their chlorides are alkaline in nature.

(ii) Alkali metals are very good conductor of electricity.

(iii) The hydroxides of alkali metals and alkaline earth metals are soluble in water.

(iv) Plaster of Paris is a hemihydrate.

(v) Alkali metals have low melting and boiling points as compared to those of alkaline earth metals.

(vi) Lithium carbonate is decomposed to its oxide, but the carbonates of the other are stable towards heat.

(vii) All alkali metal sulphates are insoluble in water.

(viii) Lithium combines with nitrogen to form lithium nitrides but other alkali metals do not react with nitrogen.

(ix) Trona is an ore of lithium.

(x) Alkaline earth metals are stronger reducing agents than alkali metals.

Answers:				
(i) False	(ii) True	(iii) False	(iv) True	(v) True
(vi) True	(vii) False	(viii) True	(ix) False	(x) False

Q.3 Multiple choice questions. Encircle the correct answer.

(i) Which one of the following does not belong to alkaline earth metals?
(a) Be (b) Ra (c) Ba d) Rn

(ii) The oxides of beryllium are:
(a) acidic (b) basic (c) amphoteric (d)non of these

(iii) Which ion will have the maximum value of heat of hydration?
(a) Na^+ (b) Cs^+ (c) Ba^{2+} (d) Mg^{2+}

(iv) Which one of the following is not an alkali metal?
(a) Francium (b) Caesium (c) Rubidium (d) Radium

(v) Which of the following sulphates is not soluble in water?
(a) Sodium sulphate (b) Potassium sulphate
(c) Zinc sulphate (d) Barium sulphate

(vi) The element caesium bears resemblance with
(a) Ca (b) Cr (c) Both of the above (d) None of these

(vii) Chile saltpetre had the chemical formula:
(a) $NaNO_3$ (b) KNO_3 (c) $Na_2B_4O_7$ (d) $Na_2CO_3.H_2O$

(viii) The mineral $CaSO_4.2H_2O$ has the general name:
(a) gypsum (b) dolomite (c) calcite (d) epsom salt

(ix) Down's cell is used to prepare:
(a) Sodium carbonate (b) Sodium bicarbonate
(c) Sodium metal (d) Sodium hydroxide

(x) Which element is deposited at the cathode during the electrolysis of brine in Nelson's cell?
(a) H_2 (b) Na (c) Cl_2 (d) O_2

ANSWERS TO MULTPLE CHOICE QUESTIONS

(i) Ans: (d)	(ii) Ans: (c)
Rn does not belong to the alkaline earth metals. It is a noble gas. Choice (d) is correct.	The oxides of beryllium are amphoteric and most of the properties of beryllium resembles group III-A elements. Choice (c) is correct.

(iii) Ans: (d)	**(iv) Ans: (d)**
The Mg^{2+} ion has maximum charge density and it has smallest possible size with +2 charge, so that it has maximum heat of hydration. Choice (d) is correct.	Radium is alkaline earth metal not an alkali metal. Choice (d) is correct.
(v) Ans: (d)	**(vi) Ans: (d)**
The solubilities of group II-A suphates decreases down the group. So that $BaSO_4$ is not soluble in water. Choice (d) is correct.	Caesium is an alkali metal. The chromium belongs to VI-B group and calcium belongs to II-A group. So caesium does not resemble with Ca and Cr. Choice (d) is correct.
(vii) Ans: (a)	**(viii) Ans: (a)**
The formula of Chile saltpetre is $NaNO_3$. Choice (a) is correct.	The general name of $CaSO_4.2H_2O$ is gypsum. Choice (a) is correct.
(ix) Ans: (c)	**(x) Ans: (b)**
For the preparation of sodium metal at commercial scale the Down's cell is used. Choice (c) is correct.	In nelson cell the electrolysis of NaCl produce Na metal which is deposited at cathode and Cl_2 gas produced at anode. Choice (b) is correct.

Q.4 **(a) Give the names, electronic configurations and occurrence of s-block elements.**

See Section 2.1.1

(b) Discuss the peculiar behaviour of lithium with respect to the other members of alkali metals.

See Section 2.1.4

Q.5 **Discuss the trends in chemical properties of compounds like oxides, hydroxides, carbonates, nitrates and sulphates of IA and IIA group elements.**

See Section 2.2.3

Q.6 **Compare the chemical behaviour of lithium with magnesium.**

(Diagonal relationship between lithium and magnesium)

Diagonal relationship:

Some elements of the second period show similar behaviour to the elements of next group belonging to the third period. This is called Diagonal relationship.

Examples

(i) Li shows behaviour to Mg

(ii) Be shows similar behaviour to Al etc.

2nd period	**Li** ↘	Be	B	C
3rd period	Na	**Mg**	Al	Si

Diagonal relationship between Li & Mg

1. Both Li and Mg are more electropositive than Be and less electropositive than Na.
2. Li^+ and Mg^{2+} ions have almost same size.

3. On heating in air, both Li and Mg forms normal oxides, while other alkali metals form higher oxides.

$$4Li + O_2 \rightarrow 2Li_2O$$
$$2Mg + O_2 \rightarrow 2MgO$$

4. Carbonates and phosphates of both Li and Mg are insoluble in water. While other alkali metals have water-soluble carbonates and phosphates.
5. On heating, carbonates of both Li and Mg decompose to form oxides, while carbonates of alkali metal are stable.

$$MgCO_3 \rightarrow MgO + CO_2$$
$$Li_2CO_3 \rightarrow Li_2O + CO_2$$

6. Both metals show strong hydration of their ions.
7. Some of the compounds of both Li and Mg are soluble in organic solvents e.g. chlorides.
8. Nitrates of both decompose completely.

$$2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$$
$$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$$

9. Both form nitrides directly.

$$6Li + N_2 \rightarrow 2Li_3N$$
$$3Mg + N_2 \rightarrow Mg_3N_2$$

10. Both form carbides directly

$$4Li + C \rightarrow Li_4C$$
$$2Mg + C \rightarrow Mg_2C$$

Q.7 (a) Mention the properties of beryllium in which it does not resemble with its own family.

See Section 2.1.5

(b) Why the aqueous solution of Na_2CO_3 is alkaline in nature?

Na_2CO_3 is hydrolyzed by water to give alkaline solution.
In water, Na_2CO_3 ionizes to give carbonate anion.

$$Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{2-}$$

Hydrolysis of carbonate anion gives strong alkali NaOH, which is greatly ionized. While carbonic acid produced, is a weak acid and is not greatly ionized. Thus, solution of Na_2CO_3 shows alkaline nature.

$$CO_3^{2-} + 2H_2O \longrightarrow H_2CO_3 + 2OH^-$$

Overall reaction is

$$Na_2CO_3 + 2H_2O \longrightarrow \underset{\textit{Strong alkali}}{2NaOH} + \underset{\textit{Weaker acid}}{H_2CO_3}$$

Q.8 (a) Describe with diagram the manufacture of sodium by Down's cell.

See Section 2.3

(b) Point out the three advantages of this process.

See Section 2.3

Q.9 (a) Compare the physical and chemical properties of alkali metals with those of alkaline-earth metals.

Similarities

1. Both alkali and alkaline earth metals are s-block elements. Both have their outermost electron in s-orbital.
2. Elements of both groups are highly electropositive.
3. Elements of both groups do not occur free in nature.
4. Elements of both groups are prepared by the electrolysis of their fused salts.
5. Hydroxides of both alkali and alkaline earth metals are strongly basic. However, hydroxides of alkaline earth metals sparingly soluble in water.
6. On heating in Bunsen flame, elements of both groups impart characteristic colour to flame.
7. All the elements of both groups are soft and silvery white metal.
8. All the elements of two groups can conduct heat and electricity.
9. Elements of both groups react with water to give hydroxides and hydrogen.
10. Elements of both groups form water soluble bicarbonates.
11. In both groups, properties very similarly.
    (i) m.p., b.p. hardness, ionization energy decreases down the group.
    (ii) Electropositivity, atomic radii, atomic volume increases down the group.

Differences

ALKALI METALS	ALKALINE EARTH METALS
1. They have one electron in their outermost s-orbital.	1. They have two electrons in their outermost s-orbital.
2. They are lighter than water	2. They are heavier than water
3. They have low m.p and b.p. than alkaline earth metals.	3. They have relatively high m.p. and b.p.
4. They have relatively larger atomic size.	4. They have relatively smaller atomic size
5. The alkali metals are relatively softer than alkaline earth metals	5. They are relatively harder.
6. They have relatively low values of heats of hydration and ionization energies.	6. They have relatively high values of heats of hydration and ionization energies
7. Their oxides and hydroxides are more ionic in nature.	7. Their oxides and hydroxides are less ionic in nature.
8. They decompose water vigorously at room temperature. $2Na + 2H_2O \rightarrow 2NaOH + H_2$	8. They decompose water less vigorously. $Mg + H_2O \rightarrow MgO + H_2$
9. Their carbonates, sulphates and phosphates are soluble in H_2O except Li	9. Their carbonates, sulphates and phosphates are mostly insoluble in H_2O
10. They are highly electropositive.	10. They are relatively less electropositive.
11. Their hydroxides are strongly basic except that of Li.	11. Their hydroxides are relatively less basic.

12. On heating, their nitrates give nitrites and oxygen $2NaNO_3 \rightarrow 2NaNO_2 + O_2$	12. On heating, their nitrates give oxides, nitrogen peroxide and oxygen $2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$
13. They don not form nitrides directly.	13. They form nitrides directly on heating with N_2. $3Mg + N_2 \rightarrow Mg_3N_2$
14. They do not form carbides directly.	14. They form carbides directly on heating with C. $Ca + 2C \rightarrow CaC_2$

(b) What happens when:

(i) Lithium carbonate is heated.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

(ii) Lithium hydroxide is heated to red hot.

$$2LiOH \xrightarrow{\text{red hot}} Li_2O + H_2O$$

(iii) Beryllium is treated with sodium hydroxide.

$$Be + 2NaOH \longrightarrow Na_2BeO_2 + H_2$$

(iv) Lithium hydride is treated with water.

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

Q.10 Give formulas of the following ores.

(a) Dolomite (b) Asbestos

(c) Halite (d) Natron

(e) Beryl (f) Sylvite

(g). Poosphorite (h) Chile salt paper

See Table 2.3 & 2.4

Q.11 Answer the following questions briefly.

(a) Why alkali and alkaline earth meals are among the reactive elements of the periodic table?

The strength of metal depends upon the ease with which it loses electron and change into positive ion. Alkali and alkaline earth metals have bigger size due to which their valence electrons are not strongly attracted. A small amount of energy is required to remove their valence electrons.

Therefore, they have low ionization energy and are highly electropositive. They lose electrons easily change to positive ions and act as reducing agents.

(b) *Why lime water turns milky with CO_2 but become clear with excess CO_2?*

Lime water is solution of lime CaO in water. When CaO is dissolved in water it reacts with water to produce $Ca(OH)_2$. When CO_2 is passed through lime water it reacts with $Ca(OH)_2$ and produce $CaCO_3$.

Therefore, lime water turns milky due to the formation of $CaCO_3$

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

This milkyness disappears when excess CO_2 is passed through it due to the formation of $Ca(HCO_3)_2$, which is soluble in water.

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

$$CaCO_3 + CO_2 + H_2O \rightarrow Ca(HCO_3)_2$$

(c) *How gypsum is converted into plaster of paris?*

When gypsum is heated at 100°C it loses 3/4th of its water and changes to plaster of Rairs.

$$2CaSO_4.2H_2O \rightarrow (CaSO_4)_2.\ H_2O + 3H_2O$$

or

$$CaSO_4.\ H_2O \rightarrow CaSO_4.\ \tfrac{1}{2}H_2O + H_2O$$

(d) *Why 2% gypsum is added in the cement?*

Gypsum prevents the cement from hardening too rapidly. The addition of gypsum increases the setting time of cement increases interlocking property.

(e) *Why lime is added to an acidic soil?*

Lime is added to the acidic soil. It control the pH of soil and increase the readily soluble phosphorous.

(f) *How lime and sand are used to make glass.*

Lime and sand react at high temperature to form calcium silicate ($CaSiO_3$). Calcium silicate is further used in the formation of glass.

(g) *How lime mortar is prepared?*

When freshly prepared slaked lime (one volume) is mixed with sand (three or four volume) and water to form a thick paste. This material when placed between stones and bricks hardens or sets firmly. It binds the stones or bricks firmly. The equations when mortar hardens are

$$CaO + HO_2 \rightarrow Ca(OH)_2$$

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

$$Ca(OH)_2 + SiO_2 \rightarrow CaSiO_3 + H_2$$

TEST YOUR SKILLS — Marks: 85

OBJECTIVE

Time: 20 Minutes — **Marks: 17**

Note: Over writing, cutting, erasing, using lead pencil will result in loss of marks.

Q1. Each question has four possible answers. Choose the correct answer and encircle it.

(i) Sodium hydroxide is prepared by electrolysis of
(a) Molten NaCl (b) Aqueous NaCl (c) Aqueous Na_2CO_3 (d) Molten Na_2CO_3

(ii) Sodium metal can not be stored under
(a) Benzene (b) Kerosene (c) Alcohol (d) Toluene

(iii) Magnesium burns in air to form
(a) MgO (b) Mg_3H_2 (c) $MgCO_3$ (d) MgO and Mg_3N_2

(iv) Setting of plaster of Paris is
(a) oxidation of atmospheric oxygen. (b) combination with atmospheric CO_2.
(c) dehydration (d) Hydration to yield another hydrate.

(v) The hydroxide which is most soluble in water is
(a) $Ba(OH)_2$ (b) $Mg(OH)_2$ (c.) $Sr(OH)_2$ (d) $Ca(OH)_2$

(vi) Which mineral was used in isolation of radium?
(a) Lime stone (b) Pitchblende (c) Rutile (d) Haematite

(vii) Which one of the following is most basic?
(a) Al_2O_3 (b) MgO (c) SiO_2 (d) P_2O_5

(viii) In manufacture of caustic soda, what is formed as a by-product?
(a) O_2 (b) Cl_2 (c) NaCl (d) N_2

(ix) Lime is used in
(a) Refining of sugar (b) Manufacture of bleaching powder
(c) Ceramic industry (d) All of above

(x) Natron is the mineral of
(a)Na (b) B (c)Ca (d) K

(xi) Down's cell is used to prepare
(a) Sodium carbonate (b) Sodium metal (c) Sodium bicarbonate (d) Sodium hydroxide

(xii) All of alkali metals burn in jar of
(a) Air (b) Cl_2 (c) Both a and b (d) Neither a nor b

(xiii) The alkaline earth metals Ba, Sr, Ca and Mg can be arranged in order of their decreasing first ionization potential as
(a) Mg, Ca, Sr, Ba (b) Ca, Sr, Ba, Mg (c) Sr, Ba, Mg, Ca (d) Ba, Mg, CA, Sr

(xiv) Epsom salt is
(a) $MgSO_4.7H_2O$ (b) $MgSO_4.2H_2O$ (c) $2CaSO_4.H_2O$ (d) $BaSO_4.2H_2O$

(xv) Elements of IIA group give colour in Bunsen burner due to
(a) Low ionization potential (b) Low melting points
(c) Softness (d) One electron in outermost orbit

(xvi) Which of the following metal is resistant to complete oxidation
(a) Na (b) K (c) Be (d)None

(xvii) Dead burnt is
(a) hydrated Calcium sulphate (b) anhydrous Calcium sulphate
(c) hydrated Copper sulphate (d) anhydrous Copper sulphate

SUBJECTIVE

Time: 2:10 Hours — **Marks: 68**

Note: Attempt any TWENTY TWO(22) questions from section-I and any THREE questions from section-II.

Section – I

Q2. Attempt any TWENTY TWO (22) questions. **(22x 2)=44**

(i) Why the group IA elements are called alkali metals?
(ii) Why the group IIA Elements are called alkaline Earth metals?
(iii) Explain why lithium differs from the other elements of group IA?
(iv) Explain why nitrates and carbonates of Li are not stable?
(v) Why m.p's and b.p's of group IA and IIA decreases down the group?
(vi) Why alkali metals are strong reducing agents?
(vii) Differentiate the behaviour of Li and Na with Atmospheric oxygen.
(viii) The reaction of sodium with water is an exothermic reaction. Explain.
(ix) Most of the simple salts of alkali metals are water soluble explain.
(x) Explain the reaction of alkali metals oxides with water is an acid bases reaction but not a redox reaction .
(xi) Why beryllium differs from other elements of the same group?
(xii) Why Beryllium is not oxidised completely in water as compared to other elements?
(xiii) Alkali metal carbonates are more soluble than alkaline earth metal carbonates. Why?
(xiv) What is the significance of KO_2 for mountaineers in space air crafts?
(xv) Justify that BeO is an amphoteric oxide.
(xvi) What is lime water give its one use?
(xvii) Explain why stability and solubility of alkaline earth metal carbonates decreases down the group
(xviii) What are the two major problems faced during working with nelson's cell?
(xix) How the problem working with Nelson cell can be solved?
(xx) What is plaster of paris?
(xxi) What is dead burnt gypsum?
(xxii) Why alkali and alkaline earth metals are among the reactive elements of periodic table?
(xxiii) Why lime water turns milky with CO_2, but becomes clear with excess of CO_2?
(xxiv) How is gypsum converted into plaster of Paris?
(xxv) Why 2% gypsum is added to cement?
(xxvi) Why lime is added to the acidic soil?
(xxvii) How lime and sand are used to make glass?
(xxviii) How lime mortar is prepared?
(xxix) How lime is employed as a dehydrating agent?
(xxx) What is meant by slaking of lime?
(xxxi) What happens when lithium carbonate and lithium hydroxide is heated to red hot?
(xxxii) What happens when beryllium is treated with sodium hydroxide?
(xxxiii) What happens when lithium hydride is treated with water?

Section – II (Attempt any three questions) (8x 3)=24

Q. 3. (a) What are alkali and alkaline earth metals? Give their occurrence and electronic configuration. **(04)**

(b) Give the names and formulas of minerals of Magnesium, Calcium, Sodium and Potassium? **(04)**

Q. 4. (a) Discuss the peculiar behaviour of beryllium with respect to other alkaline earth metals? **(04)**

(b) Discuss the nitrates and sulphates of alkali and alkaline earth metals. **(04)**

Q. 5. (a) How sodium hydroxide is prepared by diaphragm cell commercially? **(04)**

(b) Give the reactions involved in the preparation of sodium by Down's cell. **(02)**

(c) Why calcium is important for plant growth? **(02)**

Q. 6. (a) What is role of lime in agriculture and industry? **(06)**

(b) How lithium reacts with **(02)**

(i) Oxygen *(ii) Carbon*

Q. 7. (a) How the lithium shows different behaviour from its family members? **(04)**

(b) What are hard finish plasters? **(02)**

(c) What are the advantages of operating Down's cell at 600^0C? **(02)**