s-BLOCK ELEMENTS

INTRODUCTION

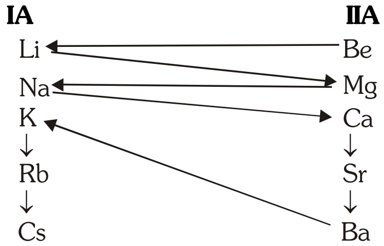
The s-block elements of the Periodic Table are those in which the last electron enters in the outermost s-orbital. As the s-orbital can accommodate only two electrons, two groups (1 & 2) belong to the s-block of the Periodic Table. Group 1 of the Periodic table consists of the elements: lithium, sodium, potassium, rubidium, cesium and francium. They are collectively known as the alkali metals. They are so called because they form hydroxides on reaction with water which are strongly alkaline in nature. The elements of Group 2 include Beryllium, Magnesium, Calcium, Strontium, Barium and Radium. These elements with the exception of **beryllium** are commonly known as the alkaline earth metals. They are so called because their oxides and hydroxides are alkaline in nature and these metal oxides are found in the earth's crust.

PHYSICAL PROPERTIES OF S-BLOCK ELEMENTS

The atomic, physical and chemical properties of alkali metals are discussed below.

Atomic and Physical Properties of the Alkaline Metals

physical properties of s-Block Elements

ALKALI METALS	ALKALINE EARTH METALS
Physical state	
● One electron in outermost shell & **General formula ns^1**. ● Francium is radioactive element. ● All are silvery white ● Light soft, malleable and ductile metals with metallic lustre.	● Two electrons in outer most shell & **General formula ns^2.** ● Radium is radioactive element. ● All are silvery white. ● These metals are harder than alkali metals.
Atomic size	
● Largest in their respective period (except noble gas element) ● Size increases from Li to Cs due to addition of an extra shell. $\mathrm{Li} < \mathrm{Na} < \mathrm{K} < \mathrm{Rb} < \mathrm{Cs}$ 	● Smaller than IA group elements, since extra charge on nucleus attracts the electron cloud. ● Size increases gradually from Be to Ba $\mathrm{Be} < \mathrm{Mg} < \mathrm{Ca} < \mathrm{Sr} < \mathrm{Ba}$ **In s-block elements** Be has smallest size, while Cs has largest size.
Melting point and Boiling point	
● Weak interatomic bonds are due to their large atomic radii and presence of only one valence electron hence melting point and boiling point are low. ● Decreasing order of melting point and boiling point is $\mathrm{Li} > \mathrm{Na} > \mathrm{K} > \mathrm{Rb} > \mathrm{Cs}$	● Metallic bond is stronger than IA group due to smaller atomic size and two electrons in valence shell hence melting point and boiling point are higher. ● Decreasing order of melting point $\mathrm{Be} > \mathrm{Ca} > \mathrm{Sr} > \mathrm{Ba} > \mathrm{Mg}$ Boiling point $\mathrm{Be} > \mathrm{Ba} > \mathrm{Ca} > \mathrm{Sr} > \mathrm{Mg}$
Melting point & Boiling point $\propto$ Strength of metallic bond $\propto$ Number of valence shell e^-	

Ionisation energy (I.E.)

- First ionisation energy (I.E.) is very less because of larger atomic size and only one electron in outer most shell.
- Decreasing order of ionisation energy-

  Li > Na > K > Rb > Cs
- Second ionisation energy of alkali metals is very high because by loosing one electron they achieve inert gas configuration.

- First ionisation energy is higher than IA group because of smaller atomic size and completely filled s-orbital (stable electronic configuration)
- Decreasing order of ionisation energy–

  Be > Mg > Ca > Sr > Ba
- Second ionisation energy is lesser than IA group.

Oxidation state

- The alkali metals shows only + 1 oxidation state.

  (difference between IE_1 and $IE_2 > 16eV$)

- Alkaline earth metal shows +2. Oxidation state

  (difference between IE_1 and $IE_2 < 11eV$)

Electro positive character or metallic character

- **Electropositivity $\propto$ 1/Ionisation energy**

  Due to their larger size electron can easily be removed to form M^+ ion. Electro positive property increases from Li to Cs.

- Their atomic size is smaller than IA group so these are lesser electro positive than IA group. Electropositivity increases from Be to Ba

Density (D = M /V)

- In a group atomic volume also increase along with atomic weight but atomic weight increases more than atomic volume, so density increases from Li to Cs

  Increasing order of density Li < K < Na < Rb < Cs

Exception : Density of K is less than Na. Why ?

Ans. This is due to presence of vacant d-orbital in the inner shells of K (volume increases, density decreases)

- Density increases from Be to Ba

  Increasing order of density Ca < Mg < Be < Sr < Ba

Conductivity

- Due to the presence of loosely held valence electrons which are free to move in a metal structure, these elements are good conductor of heat and electricity.

- These are also good conductor of heat and electricity due to presence of two free electrons.

  Conductivity of IA < Conductivity of IIA

Flame test

- Alkali metals and their salts gives characteristic colour to bunsen flame. The flame energy causes an excitation of the outer most electron which on dropping back to ground state emits absorbed energy as a visible light

  Li-Crimson red Na-Golden yellow K-Violet
  Rb-Red violet Cs-Blue

- Due to small size of Be & Mg outer most electrons are tightly bounded. So not excited to higher level, hence they do not give flame test.
- Other elements gives characteristic colour to flame **Ca-Brick red Sr-Crimson red Ba-apple green**

Photo electric effect

- Atomic size of K, Rb and Cs is quite large, so their ionisation energy is very low
- Due to very low ionisation energy their valence shell electrons gets excited even by absorbing visible light. That's why Cs is used in photoelectric cells.

- These elements do not show this property as their atomic size is small hence ionisation energy is higher than IA group.

Standard oxidation potential

- All the alkali metals have high +ve values of standard oxidation potential (tendency of releasing electrons in water or self ionic solutions)
  So they are good reducing agent, having upper most positions in the electro chemical series.
- Li has highest standard oxidation potential (+3.05 eV) due to its high hydration energy. Such that it converts into. Li^+ ion by loosing one electron.
- Order of standard oxidation potential is -
  $Li > Cs \approx Rb \approx K > Na$

- They have lower values of standard oxidation potential due to their high IE.
- Increasing order of standard oxidation potential is -
  $Be < Mg < Ca < Sr < Ba$

Complex formation tendency

- Only those elements can form complex compounds which have
  (a) Small cation size
  (b) High charge density
  (c) Vacant orbitals to accept electrons.
- Only Li^+ can form complex compound, due to its small size rest alkali metals have very less tendency to form complex compounds.

- Less tendency to form complex compound, but due to small size of cations Be and Mg forms complex compounds like

$Be–(BeF_4)^{-2}$; $Be_4O(CH_3COO)_6$; Mg – Chlorophyll

Chemical properties of s-block elements

Reactivity

- These elements are very reactive, so do not found in free state in nature.

Reactivity $\propto$ 1/Ionisation potential

order of reactivity – $Li < Na < K < Rb < Cs$

- Less reactive than alkali metals.

Order of reactivitity :-

$Be < Mg < Ca < Sr < Ba$

Reaction with air

- Alkali metals gets tarnish in air due to the formation of oxide at their surface hence they are kept in kerosene or paraffin wax.
- These elements reacts with moist air to form carbonates

$4Na + O_2 \rightarrow 2Na_2O$

$Na_2O + \underset{\text{(moist)}}{H_2O} \rightarrow 2NaOH$

$2NaOH + \underset{\text{(in air)}}{CO_2} \rightarrow Na_2CO_3 + H_2O$

In dry air only Li gives nitride and oxide both while other elements gives only oxides.

- Except Be,these metals easily tarnished in air.
- Beryllium in powdered form, burns brilliantly on ignition in air.
- In moist air, except Be all the elements converts into carbonates.
- In dry air all elements of II-A give nitride and oxide both.

Reaction with oxygen

Oxide ion [O^{2-}] :

- Li forms mainly Li_2O (Lithium oxide).

Peroxide [O_2^{-2}] :

- Na reacts with O_2 to form mainly peroxide (Na_2O_2).

Super oxide [O_2^{-}] :

- K, Rb and Cs forms MO_2 type oxides (super oxides) in excess of oxygen. Super oxides are paramagnetic and coloured.

$$M \xrightarrow{O_2} \underset{\text{Oxide}}{M_2O} \xrightarrow{O_2} \underset{\text{peroxide}}{M_2O_2} \xrightarrow{O_2} \underset{\text{super oxide}}{MO_2}$$

(Li_2O) (Na_2O_2) (KO_2, RbO_2, CsO_2)

- Stability order of different oxide of a metal is due to Lattice Energy

  Normal oxide > Peroxide > Superoxide

- Alkaline earth metals reacts with O_2 to form 'MO' type oxides

  (M = Be, Mg, Ca, Sr, Ba)

- In IIA only Ca, Sr, Ba form peroxide.

  Ex. CaO_2, SrO_2, BaO_2

- BeO shows amphoteric property.

  MgO → weak basic

  CaO, SrO & BaO → Strong basic

- Basic properties increases from BeO to BaO.

Reaction with hydrogen

- Alkali metals combine with H_2 forming ionic hydrides

  $2M + H_2 \rightarrow 2MH$

- Hydrides of alkali metals are attacked by water to give back hydrogen

  $MH + H_2O \rightarrow MOH + H_2$

$$\xrightarrow[\text{Thermal stability decrease, Basic property increases}]{\text{LiH, NaH, KH, RbH, CsH}}$$

- Except Be all the alkaline metals forms MH_2 type hydrides, (MgH_2, CaH_2, SrH_2, BaH_2) on heating directly with H_2

- BeH_2 is prepared by action of $BeCl_2$ with $LiAlH_4$

  $$2BeCl_2 + \underset{\text{reducing agent}}{LiAlH_4} \rightarrow 2BeH_2 + LiCl + AlCl_3$$

- BeH_2 and MgH_2 are covalent, other are ionic.

Reaction with water

- Alkali metals react vigorously with water forming hydroxides with the liberation of H_2.

  $2M + 2H_2O \rightarrow 2MOH + H_2$

- Reactivity with water increases from Li to Cs.

  Li → least reactive towards water

  Na → reacts vigorously

  K → reacts producing a flame

  Rb, Cs → reacts explosively.

- Monoxides gives strongly alkaline solution with water

  $M_2O + H_2O \rightarrow 2MOH$

- These metals reacts slowly with water gives H_2 and metals hydroxides.

  $M + 2H_2O \rightarrow M(OH)_2 + H_2$

- Be does not reacts with water

- Mg reacts only with hot water

- Ca, Sr, Ba reacts with cold water but not as energetically as alkali metals.

- from $Be(OH)_2$ to $Ba(OH)_2$ basic nature increases.

Halides

- Alkali metals reacts directly with halogen to form MX

  (M – alkalimetal, X – Halide ion)
- Ionic nature of MX increases from LiCl to CsCl
- LiCl is covalent in nature (due to polarisation of Cl^- ion by small Li^+ ion). hence its tendency of hydrolysis is more.
- K, Rb and Cs halides reacts with more halogens to gives polyhalides.

  $KI + I_2 \rightarrow KI_3$

  $CsBr + Br_2 \rightarrow CsBr_3$

- Alkaline metals reacts with X (Halogen) to form MX_2.
- Order of Ionic nature

  $BeCl_2 < MgCl_2 < CaCl_2 < SrCl_2 < BaCl_2$
- Hydrolysis tendency of these halides decreases from $BeCl_2$ to $BaCl_2$ due to decrease in covalent nature.
- $BeCl_2$ and $MgCl_2$ are covalent in nature.

  $BeO + C + Cl_2 \xrightarrow{\Delta} BeCl_2 + CO$

  $(NH_4)_2 BeF_4 \xrightarrow{\Delta} BeF_2 + NH_3 + HF$

Nitrides

- Only Li reacts directly with N_2 to form nitride which gives NH_3 on reacting with water.

  $6Li + N_2 \rightarrow 2Li_3N$

  $Li_3N + 3H_2O \rightarrow 3LiOH + NH_3\uparrow$

- All elements of II-A burns in N_2 to give M_3N_2 (nitrides)

For example :

$Be_3N_2 + 6H_2O \rightarrow 3Be(OH)_2 + 2NH_3$

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

Formation of amalgam

- Alkali metals gives amalgam with Hg.
- These metals reacts with other metals to give mixed metals (alloys)

- Shows same properties.

Solubility in liquid ammonia

- All the alkali metals dissolves in NH_3 (liq.) and produces blue coloured solution.
- The blue colour and reducing nature of solution is due to presence of ammoniated electron.
- Solution is a good conductor due to presence of both ammoniated ion and ammoniated electron.

$$Na_{(s)} + (x+y)NH_3 \rightarrow \underset{\text{ammoniated ion}}{\left[Na(NH_3)_x\right]^+} + \underset{\text{ammoniated electron}}{\left[e(NH_3)_y\right]^-}$$

- This dilute solution is paramagnetic in nature.

- Only Ca, Sr and Ba gives blue solution of ammoniated electron.
- Be and Mg are small in size and have high ionisation energy so do not dissolves in liquid NH_3.
- Dark blue colour of solution becomes fade if it is allowed to stand for a long time, it is because of metal amide formation.

  $2Na + 2NH_3 \rightarrow \underset{\text{(Sodamide)}}{2NaNH_2} + H_2\uparrow$

Exception ⇒ Li forms imide

$2Li + NH_3 \rightarrow Li_2NH + H_2$

- On increasing the concentration of metal in solution, it converts into bronze colour due to cluster formation of metal ions.

COMPOUNDS OF s-BLOCK ELEMENTS

SODIUM CHLORIDE NaCl

Occurrence : Sea water is the main source and also found in salt lakes.

Preparation :

(i) Sea water NaCl(2.7 – 2.9%) $\xrightarrow[\text{by solar heat}]{\text{Evaporation}}$ crude NaCl
(ii) It contains impurities – Na_2SO_4, $MgCl_2$, $CaCl_2$ etc.
(iii) Insoluble impurities removed by filtration.
(iv) Filtrate $\xrightarrow{\text{HCl gas passed}}$ Pure NaCl precipitation (Common ion effect)
(v) $MgCl_2$ and $CaCl_2$ are more soluble in water so left in solution.

Properties :

i. Table salt is slightly hygroscopic due to the presence of magnesium and calcium chlorides in small amounts.

ii. Reaction with $AgNO_3$

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl\text{(white ppt.)}$$

Uses i. As a preservative for pickles, meat and fish.

ii. For making freezing mixture with Ice.

SODIUM HYDROXIDE [Caustic Soda(NaOH)]

Manufacture : By electrolysis of NaCl.

(a) **Nelson Cell or Diaphragm Cell :** The following reactions takes place –

$$\left.\begin{aligned} NaCl(aq.) &\rightleftharpoons Na^+ + Cl^- \\ H_2O &\rightleftharpoons H^+ + OH^- \end{aligned}\right\} \rightarrow NaOH + H_2 + Cl_2$$

At cathode (Perforated steel) : $2H^+ + 2e^- \rightarrow H_2(g)$ At anode (Carbon) : $2Cl^-(aq.) \rightarrow Cl_2(g) + 2e^-$

(b) **Castner – Kellner Cell :** (Hg – Cathode Process)

Electrolyte (Brine) $NaCl \rightleftharpoons Na^+ + Cl^-$

On electrolysis –

At Cathode (Hg)

$Na^+ + e^- \rightarrow Na.$ and $Na + Hg \rightarrow Na.Hg$ (amalgam)

At anode (Graphite)

$2Cl^- \rightarrow Cl_2(g) + 2e^-$ and $2Na.Hg + 2H_2O \rightarrow 2NaOH + H_2 + 2Hg$

Properties :

i. It is deliquescent white crystalline solid.

ii. It absorbs CO_2 from air forming Na_2CO_3.

iii. NaOH is **strong base**

$$NaOH \xrightarrow{SiO_2} Na_2SiO_3 + H_2O$$
$$NaOH \xrightarrow{Al_2O_3} 2NaAlO_2 + H_2O$$

iv. **Reaction with non metals :** No reaction with H_2, N_2 and C

$$NaOH \xrightarrow{B} Na_3BO_3 \text{ (sodium borate)}$$
$$NaOH \xrightarrow{Si} Na_2SiO_3 \text{ (sodium silicate)}$$

v. **Reaction with Metal :**

$NaOH \xrightarrow{\text{Alkali metal}}$ No reaction

$NaOH \xrightarrow{2Al + 2H_2O} 2NaAlO_2$ (Sodium meta aluminate) + H_2

$NaOH \xrightarrow{Be} Na_2BeO_2$ (Sodium Berrylate) + H_2

$NaOH \xrightarrow{Zn} Na_2ZnO_2$ (Sodium zincate) + H_2

$NaOH \xrightarrow{Sn} Na_2SnO_2$ (Sodium stannite) + H_2

$NaOH \xrightarrow{Pb} Na_2PbO_2$ (Sodium plumbite) + H_2

vi. The hydroxides of aluminium, zinc, lead and tin, however, dissolve in excess of sodium hydroxide giving clear solution which can also be obtained when these metals are acted upon by the concentrated solution of sodium hydroxide.

$Zn(OH)_2 + 2OH^- \longrightarrow [Zn(OH)_4]^{2-}$ Zincate ion

$Al(OH)_3 + 3OH^- \longrightarrow [Al(OH)_6]^{3-}$ Aluminate ion

vii. **Reaction with $ZnCl_2$ or $ZnSO_4$**

(a) $ZnCl_2 + 2NaOH \longrightarrow Zn(OH)_2 \downarrow + 2NaCl$

(b) $Zn(OH)_2 + 2NaOH \longrightarrow Na_2[Zn(OH)_4]$ (Soluble complex)

Uses (i) In the manufacture of soap, rayon, dyes, paper and drugs.

(ii) In petroleum refining.

SODIUM BICARBONATE OR BAKING SODA ($NaHCO_3$) and SODIUM CARBONATE OR WASHING SODA [$Na_2CO_3.10H_2O$]

Preparation : Solvay Process (Commercial Scale)/Ammonia Soda Process

(i) $CaCO_3 \longrightarrow CaO + CO_2$

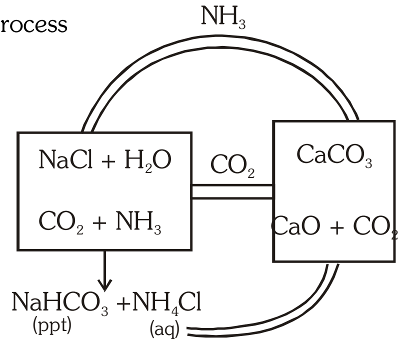
(ii) $NH_3 + H_2O + CO_2 \longrightarrow NH_4HCO_3$

(iii) $NaCl + NH_4HCO_3 \longrightarrow NH_4Cl + NaHCO_3$ (ppt)

(iv) Sodium carbonate is prepared by heating of $NaHCO_3$

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

(v) $2NH_4Cl + CaO \longrightarrow \underset{\text{(Bye-product)}}{CaCl_2} + 2NH_3 + H_2O$



Note : Potassium bicarbonate ($KHCO_3$) cannot be prepared by Solvay process as it is more soluble in water as compared to $NaHCO_3$.

Properties of $NaHCO_3$:

(i) **Hydrolysis** $NaHCO_3 + H_2O \rightleftharpoons NaOH + H_2CO_3$

(ii) **Effect of heat** (temp. > 100°C) $2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2\uparrow$

(iii) **Reaction with acids** – gives CO_2 $NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2\uparrow$

(iv) **Reaction with bases** $NaHCO_3 + NaOH \longrightarrow Na_2CO_3 + H_2O$

Note : Reaction (iii) and (iv) explain **amphoteric** behaviour of $NaHCO_3$.

Properties of Na_2CO_3

(i) Efflorescence :

$Na_2CO_3.10H_2O$ when exposed to air it gives out nine out of ten H_2O molecules.

$$Na_2CO_3.10H_2O \longrightarrow \underset{\text{(Monohydrate)}}{Na_2CO_3.H_2O} + 9H_2O$$

This process is called efflorescence. Hence washing soda loses weight on exposure to air.

(ii) Hydrolysis : Aqueous solution of Na_2CO_3 is alkaline in nature due to anionic hydrolysis.

$$Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{-2} \quad \text{and} \quad CO_3^{-2} + H_2O \rightleftharpoons \underset{\text{(Carbonic acid)}}{H_2CO_3} + 2OH^-$$

Uses of $NaHCO_3$

i. In the preparation of baking powder.

ii. In the preparation of effervescent drinks.

iii. In the fire extinguishers.

iv. As antacid medicine (removing acidity)

Uses of Na_2CO_3

i. For making fusion mixture (Na_2CO_3 + K_2CO_3)

ii. In the manufacture of glass, caustic soda, soap powders etc.

iii. In laundries and softening of water.

CALCIUM OXIDE [Quick lime (CaO)]

Preparation : By heating limestone at 800⁰C.

$$CaCO_3 \overset{800^\circ C}{\rightleftharpoons} CaO + CO_2$$

Properties :

(i) **Action of water :** $\underset{\text{(quick lime)}}{CaO} + H_2O \longrightarrow \underset{\text{(Slaked lime)}}{Ca(OH)_2}$

(ii) **Basic Nature :**

$$CaO + SiO_2 \xrightarrow{\Delta} \underset{\text{(Calcium silicate)}}{CaSiO_3}$$

$$CaO + P_4O_{10} \xrightarrow{\Delta} 2Ca_3(PO_4)_2 \text{ (Calcium phosphate)}$$

(iii) Reaction with carbon :

$$CaO + 3C \xrightarrow{2000^\circ C} \underset{\text{(Calcium carbide)}}{CaC_2} + CO\uparrow$$

$$CaC_2 + NH_2 \longrightarrow \underset{\text{Nitrolime}}{CaCN_2 + C}$$

Uses of CaO :

(i) In the manufacture of bleaching powder, cement, glass, calcium carbide etc.

(ii) In the purification of sugar

(iii) As a drying agent for NH_3 and C_2H_5OH

(iv) As basic lining in furnaces

(v) For making Soda lime [NaOH + CaO]

CALCIUM HYDROXIDE [Slaked limeCa(OH)$_2$]

Preparation : By the action of water on quick lime

$$CaO + H_2O \longrightarrow Ca(OH)_2 + \text{heat (slaking of lime)}$$

Properties of $Ca(OH)_2$

(i) **Action of CO_2 :** Lime water turns milky on passing CO_2 gas.

$$Ca(OH)_2 + CO_2 \longrightarrow \underset{\text{Milkiness}}{CaCO_3} + H_2O$$

$$CaCO_3 \xrightarrow[CO_2 + H_2O]{\text{Excess of}} \underset{\text{(soluble)}}{Ca(HCO_3)_2} \xrightarrow{\Delta} CaCO_3$$

(ii) **Action of Chlorine :**

$$\underset{\text{dry}}{Ca(OH)_2} + Cl_2 \xrightarrow[\text{below } 35^\circ C]{\text{(Room temp.)}} CaOCl_2 + H_2O$$

$$2Ca(OH)_2 + 2Cl_2 \xrightarrow{\text{red hot}} 2CaCl_2 + 2H_2O + O_2\uparrow$$

Uses of $Ca(OH)_2$

(i) For softening of hard water.

(ii) For purification of sugar and Coal gas.

(iii) In the manufacture of bleaching powder, Caustic soda and soda lime

(iv) In preparation of mortar, plaster and white washing.

CALCIUM SULPHATE [Gypsum$CaSO_4.2H_2O$]

Preparation : $CaSO_4.2H_2O$ is naturally occuring calcium sulphate. It can be obtained by the action of dil.H_2SO_4 on a soluble calcium salt below 60^0C.

$$CaCl_2 + \underset{\text{dilute}}{H_2SO_4} \longrightarrow 2HCl + \underset{\text{white ppt.}}{CaSO_4\downarrow}$$

Properties of Gypsum

(i) **Action of heat :**

$$\underset{\text{(Gypsum)}}{2(CaSO_4.2H_2O)} \xrightarrow[-3H_2O]{120^\circ C} \underset{\text{(Plaster of paris)}}{2(CaSO_4).H_2O} \xrightarrow{200^\circ C} \underset{\text{(Anhydride)}}{2CaSO_4} + H_2O$$

(ii) It forms an important fertilizer $(NH_4)_2SO_4$

$$CaSO_4 + 2NH_3 + CO_2 + H_2O \longrightarrow CaCO_3\downarrow + (NH_4)_2SO_4$$

Uses of Gypsum

(i) In the preparation of plaster of paris

(ii) Anhydrous $CaSO_4$ used as drying agent.

(iii) Anhydrite ($CaSO_4$) is used for manufacture of sulphuric acid, ammonium sulphate.

PLASTER OF PARIS [POP] $2(CaSO_4).H_2O$ or $CaSO_4 \cdot \frac{1}{2} H_2O$

Preparation : It is obtained when gypsum is heated at 120^0C

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

Properties of POP

(i) It is a white powder.
(ii) It has the property of setting to a hard mass when a paste with water is allowed to stand aside for sometime.
(iii) When it is heated at 200°C, anhydrous $CaSO_4$ is formed.

Uses of POP

(i) In surgery for setting broken bones
(ii) In making casts for toys, statues etc.
(iii) In making blackboard chalks.

SOME IMPORTANT POINTS TO REMEMBER

ANOMALOUS BEHAVIOR OF LITHIUM

- On account of its small size it exerts the greatest polarising effect out of all the alkali metals and ions, consequently covalent character is developed.
- Li has the highest ionisation energy and electronegativity as compared to other alkali metals.
- It is not affected by air easily and does not lose its lusture even on melting.
- It is more harder and lighter than other alkali metal.
- It reacts slowly with water to liberate hydrogen.
- When burnt in air or oxygen, it forms only monoxide, Li_2O. However, the rest of the alkali metals give peroxide or superoxides.
- Li_2O is less basic oxides than of other alkali metals.
- Lithium hydroxide decomposes when red heated to form Li_2O. Hydroxides of other alkali metals do not decompose.

  $2LiOH \xrightarrow{\Delta} Li_2O + H_2O$
- Li_2CO_3 is less stable, as it decomposes on heating. $Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$
- Li_2SO_4 is the only alkali metal sulphate, which does not form double salts Ex. Alum.
- Li when heated in NH_3 forms imide Li_2NH while other alkali metals form amides. MNH_2
- Lithium shows resemblance with magnesium, an element of group IIA.
  This resemblance is termed as diagonal relationship.

Similariteis between lithium and Magnesium

(i) Both lithium and magnesium are harder and lighter than the other elements in their respective groups.
(ii) Lithium and magnesium reacts slowly with cold water. Their oxides and hydroxides are much less soluble and their hydroxides decomposes on heating. They both form nitride by direct combination with nitrogen, Li_3N and Mg_3N_2.
(iii) The oxides, Li_2O and MgO do not combine with excess oxygen to give a peroxide or a superoxide.
(iv) The carbonates of lithium and magnesium decomposes easily on heating to form the oxide and CO_2. Solid bicarbonates are not formed by lithium and magnesium.
(v) Both LiCl and $MgCl_2$ are soluble in ethanol.
(vi) Both LiCl and $MgCl_2$ are deliquescent and crystallise from their aqueous solution as hydrates, $LiCl.2H_2O$ and $MgCl_2.8H_2O$.

ANOMALOUS BEHAVIOR OF BERYLLIUM

- It is the hardest of all alkaline earth metal as maximum metallic bonding is there due to it's smallest size.
- The melting and boiling points of the beryllium are the highest.
- It is least reactive due to highest ionisation potential.
- Due to high charge density its polarising effect is highest and it forms covalent bond.
- It dissolves in alkalies with the evolution of hydrogen

  $$Be + 2NaOH + 2H_2O \longrightarrow \underset{\text{Sodium beryllate}}{Na_2BeO_2.2H_2O} + H_2$$

  other alkaline earth metals do not react with alkalies.
- Oxides and hydroxides of beryllium are amphoteric in nature.

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

  $Be(OH)_2 + 2HCl \rightarrow BeCl_2 + 2H_2O$ $Be(OH)_2 + 2NaOH \rightarrow Na_2BeO_2 + 2H_2O$

  The hydroxide is Insoluble in water and is covalent in nature.
- Like Al_4C_3, its carbide (Be_2C) on hydrolysis evolves methane.

- Due to its small size it has strong tendency to form complex.
- It shows diagonal relationship with Al.

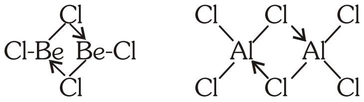
DIAGONAL SIMILARITY BETWEEN BERYLLIUM AND ALUMINIUM :

In many of its properties, beryllium resembles with aluminium. Thus –

(i) The two elements have same electronegativity and their charge/ radius ratios.

(ii) Both metals are fairly resistant to the action of acids due to the formation of a protective film of oxide on their surface. Both metals are acted upon by strong alkalies to form soluble complexes, beryllates $[Be(OH)_4]^{2-}$ and aluminates, $[Al(OH)_4]^-$.

(iii) The chlorides of both beryllium and aluminium have bridged chloride structures in vapour phase.



(iv) Salts of these metals form hydrated ions, Ex. $[Be(OH_2)_4]^{2+}$ and $[Al(OH_2)_6]^{3+}$ in aqueous solutions. Due to similar charge/ radius ratios of beryllium and aluminium ions they have a strong tendency to form complexes. For example beryllium forms tetrahedral complexes such as BeF_4^{2-} and $[Be(C_2O_4)_2]^{2-}$ and aluminium forms octahedral complexes like AlF_6^{3-} and $[Al(C_2O_4)_3]^{3-}$.

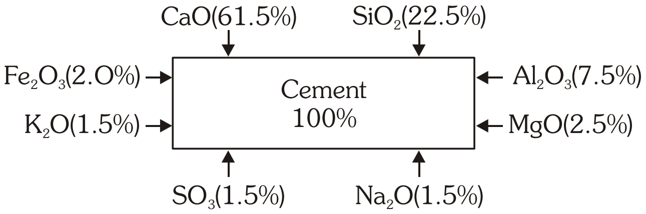
PORTLAND CEMENT

It is a light grey, heavy fine powder, It is a homogenous mixture of silicates and aluminates of calcium, which form more than 90% of the cement are –

(i) Tricalcium silicate - $3CaO.SiO_2$
(ii) Dicalcium silicate (slowest setting component) - $2CaO.SiO_2$
(iii) Tricalcium aluminate (fastest setting component) - $3CaO.Al_2O_3$
(iv) Tetracalcium alummino ferrite - $4CaO.Al_2O_3.Fe_2O_3$

Important Points :-

(i) Composition of Cement



For high quality

⇒ SiO_2 to Al_2O_3 Ratio = 2.5–4

⇒ CaO to Rest all oxide = 2

(ii) Raw Materials

- Lime Stone – It provides CaO
- Clay – It provides Al_2O_3 and silica (SiO_2)
- Gypsum – $CaSO_4 \cdot 2H_2O$

(iii) Setting of cement : When water is mixed to cement and the mixture is left it becomes very hard. This property of cement is called setting.

(iv) Mortar : It is a mixture of cement, sand and water to give a proper consistency.

(v) Concrete : A mixture of cement, Sand gravel and water is known as concrete.

(vi) Reinforced concrete cement (RCC) : When concrete is filled in beams made of iron bars, it is called RCC. Iron imparts extra strength to the structure.

BIOLOGICAL ROLE OF SODIUM, POTASSIUM, MAGNESIUM AND CALCIUM (FROM BIOLOGY PORTION)

Normally % of abundance in human body – K > Na > Fe > Cu