

## S – Block

The elements in which the last electron enters the outer most s - orbital are called s-block elements. The group I and II of periodic table belongs to the s-block.

## ALKALI METAL

## Physical State

- (a) One electron in outermost shell & General formula  $ns^1$ .
- (b) Francium is radioactive element.
- (c) All are silvery white
- (d) Light soft, malleable and ductile metals with metallic lustre.
- (e) Alkali metals are paramagnetic, diamagnetic and colourless in form of ions.

## Atomic Size

- (a) Biggest in their respective period (except noble gas element)
- (b) Size increases from Li to Fr due to addition of an extra shell.  
 $Li < Na < K < Rb < Cs < Fr$

## Softness

- (a) Alkali metals are soft because of
  - (i) Large atomic size
  - (ii) BCC crystal structure (HCP in Li)
  - (iii) Loose packing (68 % packing efficiency)
  - (iv) Weak metallic bond
- (b) Cs is the softest metal in s-block

$$\text{Atomic size} \propto \frac{1}{\text{strength of metallic bond}} \propto \text{softness} \propto \frac{1}{\text{Melting \& Boiling point}}$$

## Melting point and boiling point

- (a) 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.
- (b) Decreasing order of melting point and boiling point is  
 $Li > Na > K > Rb > Cs$
- (c) With the increase in the size of metal atom, the repulsion of the non-bonding electrons increases and therefore melting point and boiling point decreases from Li to Cs.

## Electro positive character or metallic character

Electropositive  $\propto 1/\text{Ionisation Potential}$

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

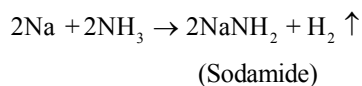
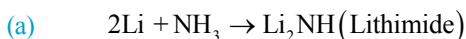
## Flame Test

Alkali metals and their salts give 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 visible light

<b>Ex.</b>	Li-Crimson red	Na-Golden yellow	K-Violet
	Rb-Red violet	Cs-Blue	

## CHEMISTRY FOR JEE MAIN & ADVANCED

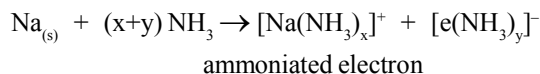
### Reaction with $\text{NH}_3$



(b) Solubility in liquid ammonia

(i) All the alkali metals dissolve in  $\text{NH}_3$  (Liq.) and produce a blue solution.

(ii) This blue solution conducts electricity and possesses strong reducing power, due to the presence of ammoniated electrons.



(iii) This dilute solution is paramagnetic in nature.

### Photo Electric Effect

(a) Atomic size of K, Rb and Cs is quite large, so their ionisation potential is very low

(b) Due to very low ionisation potential their valence shell electrons get excited even by absorbing visible light. That's why Cs is used in photo cells.

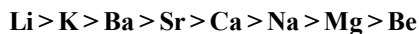
### Standard oxidation potential

(a) All the alkali metals have high +ve values of standard oxidation potential (tendency of releasing electrons in water or self ionic solutions)

(b) So these are good reducing agents, having upper most position in the electrochemical series.

(c) Li has the highest standard oxidation potential (+3.05 eV) due to its high hydration energy. Such that it converts into  $\text{Li}^+$  ion by losing one electron.

#### Order of standard oxidation potential of s-block elements

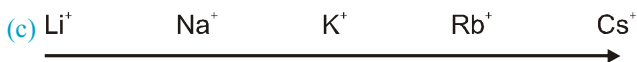


Hydration energy  $\propto$  charge density on ion

### Hydration energy (Heat of hydration)

(a) Alkali metal salts are generally soluble in water due to hydration of cations by water molecules.

(b) Smaller the cation, greater is the degree of its hydration.



\* Degree of hydration decreasing

\* Hydration energy decreasing

\* Hydrated ion size decreasing

\* Ionic conductance increasing

### Reducing property

(a) Since alkali metals have high standard oxidation potential, so these are the strongest reductants.

(b) Reducing property increases down the group in gaseous or molten state



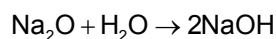
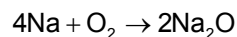
9 (c) But in aqueous solution order is



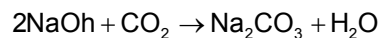
**Reaction with air**

(a) alkali metals get tarnish in air due to the formation of oxide at their surface hence they are kept in kerosene or paraffin oil.

(b) These elements react with moist air to form carbonates



(moist)



(in air)

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

**Reaction with oxygen**

Oxide ion  $[\text{O}^{2-}]$

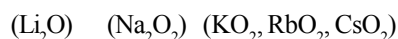
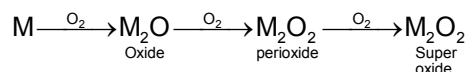
Li forms only  $\text{Li}_2\text{O}$  (Lithium oxide)

Peroxide  $[\text{O}_2]^{-2}$

Na reacts with  $\text{O}_2$  to form peroxide ( $\text{Na}_2\text{O}_2$ )

Super oxide  $[\text{O}_2^-]$

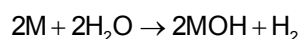
K, Rb and Cs form  $\text{MO}_2$  type oxides (super oxides) in excess of  $\text{O}_2$ . So super oxides are paramagnetic and coloured.

**Their stability order is**

Normaloxide > Peroxide > Superoxide

**Reaction with water**

(a) Alkali metals react vigorously with water forming hydroxides with the liberation of  $\text{H}_2$ .



(b) Reactivity with water increases from Li to Cs.

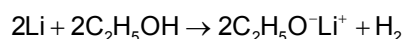
Li → least reactive towards water

Na → reacts vigorously

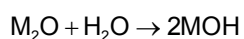
K → reacts producing a flame

Rb, Cs → reacts explosively

(c) These metals also react with alcohol giving alkoxide and  $\text{H}_2$ .



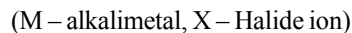
(d) Monoxides give strongly alkaline solution with water



## CHEMISTRY FOR JEE MAIN & ADVANCED

### Halides

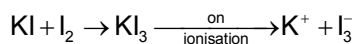
(a) Alkali metals react directly with halogen to form MX



(b) Ionic properties of MX increase from LiCl to CsCl

(c) LiCl is covalent in nature (due to polarisation of  $\text{Cl}^-$  ion by small  $\text{Li}^+$  ion). Hence it hydrolyses with water while others are ionic so do not hydrolyse.

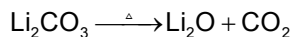
(d) K, Rb and Cs halides react with more halogens to give polyhalides.



### Carbonates

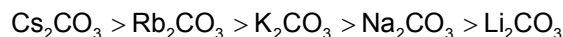
(a) All the alkali metals form  $\text{M}_2\text{CO}_3$  type carbonates.

(b) Except  $\text{Li}_2\text{CO}_3$ , all the carbonates are stable towards heat



(c) Thermal stability of carbonates  $\propto 1/\phi$  (Ionic potential)

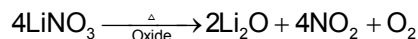
Order of stability is



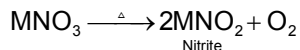
### Nitrates

(a) Alkali metals form  $\text{MNO}_3$  type nitrates (M - alkali metal)

(b) Stability increases from  $\text{LiNO}_3$  to  $\text{CsNO}_3$ .  $\text{LiNO}_3$  decomposes into Lithium oxide &  $\text{NO}_2$  on heating

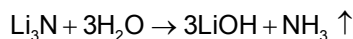
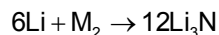


(c) Other nitrates, on heating, give nitrite and oxygen.



### Nitrides

Only Li reacts directly with  $\text{N}_2$  to form nitride which gives  $\text{NH}_3$  on reacting with water.



### Formation of Amalgam

(a) Alkali metals give amalgam with Hg.

(b) These metals react with other metals to give mixed metals (alloys)

### Sulphates

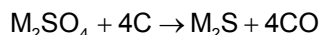
(a) Alkali metals form  $\text{M}_2\text{SO}_4$  type sulphates

(b) All alkali metal sulphates are ionic. Ionic properties increase from Li to Cs.



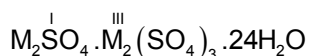
(c)  $\text{Li}_2\text{SO}_4$  is least soluble in water.

(d) These sulphates on burning with C form sulphides



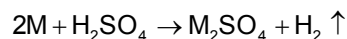
(e) Except lithium, sulphates of IA group react with

sulphates of trivalent metals like  $\text{Fe}^{+3}$ ,  $\text{Cr}^{+3}$ ,  $\text{Al}^{+3}$  etc. give double salts called alum.



**Reaction with acids**

Reacts vigorously with acids



**Ex.** The first three elements of Group 1 have the following atomic structures

- (a) Lithium : 3p, 4n ; 2, 1 electrons
- (b) Sodium : 11p, 12n; 2, 8, 1 electrons
- (c) Potassium : 19p, 20n; 2, 8, 8, 1 electrons

Which of the following features causes them to have similar properties

- (A) The same number of protons
- (B) More protons than electrons
- (C) Two electrons in the first shell
- (D) One electron in the outermost shell

**Ans** (D)

**Sol.** One electron in the outermost shell.

**Ex.** Considering greater polarization in LiCl compared to that in NaCl, which of the following statements you would expect to be wrong ?

- (A) LiCl has lower melting point than that of NaCl
- (B) LiCl dissolves more in organic solvents
- (C) LiCl will ionize in water more than NaCl
- (D) Fused LiCl would be less conducting than fused NaCl.

**Ans** (C)

**Sol.** According to Fajan's rules, higher the polarization, higher is the covalent character. Thus LiCl is covalent while NaCl is ionic. As a result LiCl will not ionize in water more than NaCl.

**Ex.** Sodium ordinarily does not show an oxidation state of +2, because of its

- (A) High first ionization potential
- (B) High second ionization potential
- (C) Large ionic radius
- (D) High electronegativity

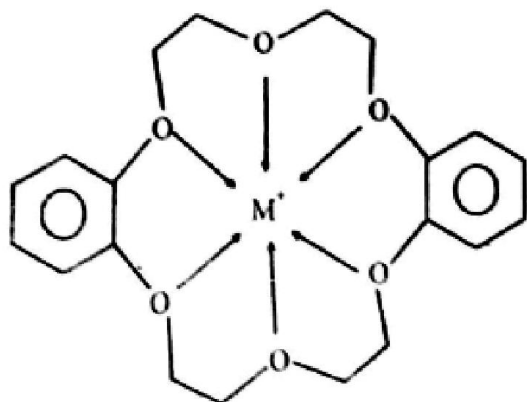
**Ans.** (B)

**Ex.** Mark the false statement

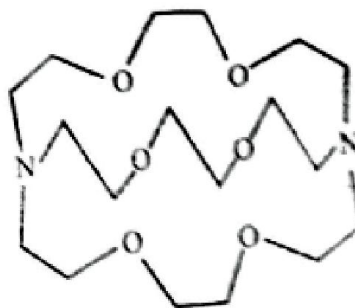
- (A) The electropositive character of alkali metals decreases with increase in atomic number
- (B) Lithium is hard metal and cannot be cut with a knife
- (C) Alkali metals are strong reducing agents
- (D) Electronegativities of all alkali metals lie between 1.0 to 0.7.

**Ans** (A)

Crown Ethers and Cryptands



Dibenzo-18-Crown-6  
 $[\text{Na}(\text{Cryptand } 222)]^+ \text{Na}^-$   
 $\text{Cs}^+(\text{Cryptand} - 222) [(\text{Cryptand} - 222)\text{e}^-]$



Cryptand - 222  
 [Contains  $\text{Na}^-$  (sodide ion)]  
 [electride]

Diagonal Relationship

Lithium shows diagonal relationship with magnesium since they have almost the same polarizing power, i.e., charge/size ratio.  $\rightarrow \infty$   $\downarrow \checkmark$

Lithium resembles magnesium in the following respects :

- (i) Atomic size of Li (1.34 Å) and Mg (1.36 Å) are not much different.
- (ii) ionic size of  $\text{Li}^+$  (0.60 Å) and  $\text{Mg}^{2+}$  (0.65 Å) are almost identical
- (iii) polarizing power (i.e., ionic charge to ionic radius ratio) of  $\text{Li}^+$  and  $\text{Mg}^{2+}$  are about the same
- (iv) electronegativities of Li (1.00) and Mg (1.20) are not much different
- (v) both Li and Mg are hard metals
- (vi) both decompose water only on heating
- (vii) both combine with oxygen to form monoxides,
- (viii) both  $\text{LiOH}$  and  $\text{Mg}(\text{OH})_2$  are weak bases
- (ix) both  $\text{LiCl}$  and  $\text{MgCl}_2$  are predominantly covalent and hence are soluble in organic solvents, such as alcohol and pyridine
- (x) both Li and Mg combine with nitrogen to form their respective nitrides,  $\text{Li}_3\text{N}$  and  $\text{Mg}_3\text{N}_2$ .
- (xi) carbonates, chlorides, oxalates and phosphates of both Li and Mg are sparingly soluble in  $\text{H}_2\text{O}$ .
- (xii) the hydroxides and carbonates of both Li and Mg decompose on heating forming their respective oxides.
- (xiii) both lithium and magnesium nitrate on heating evolve  $\text{NO}_2$  and  $\text{O}_2$  leaving behind their oxides.

Ore of alkali metals

Since alkali metals are highly reactive metals, they do not occur in free state in nature sodium and potassium are quite abundant in the earth's crust with 7th and 8th position in the order of abundance of elements

(A) Ores of Lithium

- (i) Spodumene,  $\text{LiAlSi}_2\text{O}_6$
- (ii) Lepidolite,  $(\text{Li})_2\text{Al}_2(\text{SiO}_3)_3 \cdot (\text{FOH})_2$

(B) Ores of Sodium

- (i) Common salt or rock salt,  $\text{NaCl}$
- (ii) Chile salt patre,  $\text{NaNO}_3$
- (iii) Albite (Soda Feldspar),  $\text{NaAlSi}_3\text{O}_8$
- (iv) Glauber's salt,  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

(C) Ores of Potassium

- (i) Camallite,  $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$
- (ii) Feldspar,  $\text{KAlSi}_3\text{O}_8$
- (iii) Sylvine,  $\text{KCl}$

**Extraction of alkali metals**

(i) **Lithium:** Extraction of lithium involves two steps

1. Preparation of LiCl from the ore 2. Electrolysis of LiCl

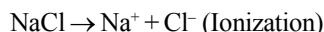
During electrolysis the electrolyte used is 55% LiCl and 45% KCl. The electrolyte is maintained in molten state by heating to about 723 K.

(ii) **Sodium**

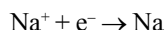
**Down's Process**

In this method, Sodium is obtained by the electrolysis of a mixture of sodium chloride (40%) and calcium chloride (60%) in fused state. The function of calcium chloride is to lower the operating temperature from 1080 K (m.pt. of NaCl) to about 850 K.

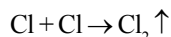
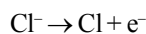
The following reactions take place:



At cathode :

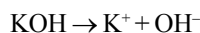


At anode :

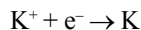


The sodium metal obtained by this method is about 99% pure. Chlorine is the by-product in this process.

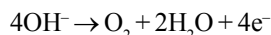
(iii) **Potassium :** potassium is obtained by electrolysis of fused potassium hydroxide (KOH)



At cathode :



At anode :



**Ex.** Alkali metals readily dissolve in liquid ammonia to give blue coloured solution. The blue colour is believed to be due to

(A) Ammoniated cations

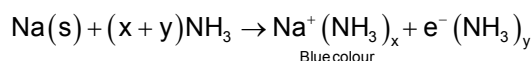
(B) Ammoniated anions

(C) Ammoniated electrons

(D) Ammoniated cations and ammoniated electrons

**Ans.** (C)

**Sol.** Ammoniated electrons, i.e.



**Ex.** Choose the compound which does not possess a peroxide group

(A)  $\text{Na}_2\text{O}_2$

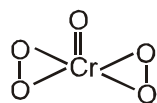
(B)  $\text{CrO}_5$

(C)  $\text{Fe}_2\text{O}_3$

(D)  $\text{BaO}_2$

**Ans.** (C)

**Ex.**  $\text{Na}_2\text{O}_2$  is  $\text{Na}^+\text{O}^- \text{O}^- \text{Na}^+$ ;  $\text{CrO}_5$  is



$\text{BaO}_2$  is  $\text{Ba}^{2+} (\text{O}^- \text{O}^-)$  while  $\text{Fe}_2\text{O}_3$  consists of only  $\text{Fe}^{3+}$  and  $\text{O}^{2-}$  ions. Thus  $\text{Fe}_2\text{O}_3$  does not contain a peroxide ( $\text{O}^- \text{O}^-$ ) linkage

**Ex.** The addition of  $\text{Na}_2\text{CO}_3$  to the aqueous solution of an oxide produces  $\text{CO}_2$ . This reaction indicates that

(A) Oxide is basic

(B) Oxide is amphoteric

(C) Oxide is that of a metal

(D) Oxide is that of a non-metal

**Ans** (D)

- Ex.** Alkali metal when exposed to air tarnish quickly due to the  
**(A)** Formation of their hydroxides **(B)** Formation of their carbonates  
**(C)** Formation of their oxides **(D)** All the above
- Ans** **(D)**

## COMPOUNDS OF ALKALI METALS

### 1. SODIUM (NA), NATRIUM

**(a) Extraction : Down's Process**

By Electrolysis of fused  $\text{NaCl} + \text{CaCl}_2 + \text{NaF}$

At cathode (Iron Vessel) :  $\text{Na}^+ + e^- \longrightarrow \text{Na(s)}$

At Anode (Graphite) :  $2\text{Cl}^- \longrightarrow \text{Cl}_2 + 2e^-$

- (i)** ( $\text{CaCl}_2 + \text{NaF}$ ) is used to lower Melting point ( $800^\circ\text{C}$ ) of  $\text{NaCl}$  to about  $600^\circ\text{C}$ .  
**(ii)** Aqueous sodium chloride cannot be used for preparing sodium by electrolysis. Because instead of metallic sodium, hydrogen gas will be liberated at cathode.

**(b) Properties**

- (i)** It is a crystalline soft metal.  
**(ii)** Highly reactive, so kept in kerosene.  
**(iii)** Na dissolves in liquid  $\text{NH}_3$  to give blue solution.

**(c) Uses**

- (i)** In the preparation of sodium amalgam (used as reducing agent)  
**(ii)** In sodium vapour lamp, which emits monochromatic yellow light.  
**(iii)** As heat transfer medium in nuclear reactors.

### 2. SODIUM CHLORIDE $\text{NaCl}$

**(a) Occurrence :** Sea water is the main source and also found in salt lakes.

**(b) Preparation**

**(i)** Sea water  $\text{NaCl}$ (2.7 – 2.9%)  $\xrightarrow[\text{by solar heat}]{\text{Evaporation}}$  crude  $\text{NaCl}$

**(ii)** It contains impurities –  $\text{Na}_2\text{SO}_4$ ,  $\text{MgCl}_2$ ,  $\text{CaCl}_2$  etc.

**(iii)** Insoluble impurities removed by filtration.

**(iv)** Filtrate  $\xrightarrow{\text{HCl gas passed}}$  Pure  $\text{NaCl}$  precipitation (Common ion effect)



Ionic product of  $[\text{Na}^+][\text{Cl}^-] >$  solubility product of  $\text{NaCl}$  hence it precipitates out.

**(v)**  $\text{MgCl}_2$  and  $\text{CaCl}_2$  are more soluble in water so left in solution.

**(c) Properties**

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

**(ii)** Reaction with  $\text{AgNO}_3$



Reaction with  $\text{K}_2\text{Cr}_2\text{O}_7 + \text{conc. H}_2\text{SO}_4$

**(iii)**  $4\text{NaCl} + \text{K}_2\text{Cr}_2\text{O}_7 + 5\text{H}_2\text{SO}_4 \xrightarrow{\Delta} 4\text{NaHSO}_4 + \text{K}_2\text{SO}_4 + 2\text{CrO}_2\text{Cl}_2 + 3\text{H}_2\text{O}$   
 (orange red)

**(d) Uses**

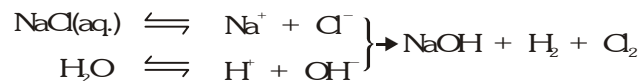
- (i)** As a preservative for pickles, meat and fish.  
**(ii)** For making freezing mixture with Ice.



### 3. SODIUM HYDROXIDE (NaOH), CAUSTIC SODA

(a) **Manufacture** : By electrolysis of NaCl.

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



At cathode (Perforated steel) :  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$  At anode (Carbon) :  $2\text{Cl}^-(\text{aq.}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$

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

Electrolite (Brine)  $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$

**On electrolysis –**

At Cathode (Hg)

$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$  and  $\text{Na} + \text{Hg} \rightarrow \text{Na.Hg}$  (amalgum)

At anode (Graphite)

$2\text{Cl}^- \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$  and  $2\text{Na.Hg} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2 + 2\text{Hg}$

(d) **Properties**

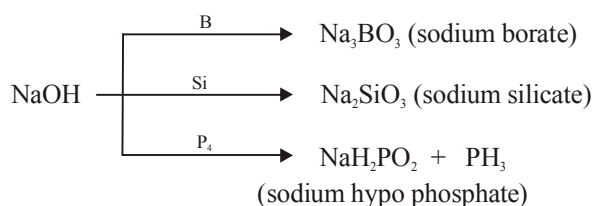
(i) It is deliquescent white crystalline solid.

(ii) It absorbs  $\text{CO}_2$  from air forming  $\text{Na}_2\text{CO}_3$ .

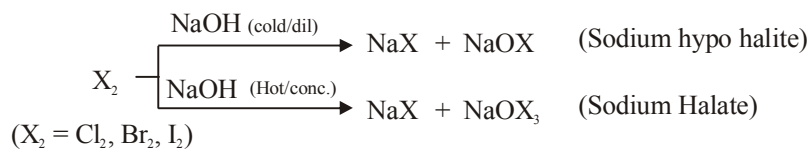
(iii) NaOH is **strong base**



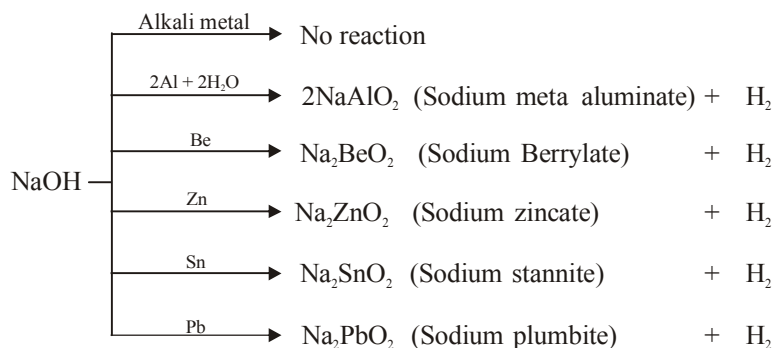
(iv) **Reaction with non metals** : no reaction with  $\text{H}_2$ ,  $\text{N}_2$  and C



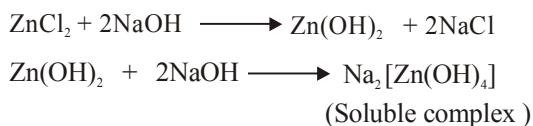
(v) **Reaction with halogens**



(vi) **Reaction with Metal**



(vii) **Reaction with  $ZnCl_2$  or  $ZnSO_4$**



(viii) 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.

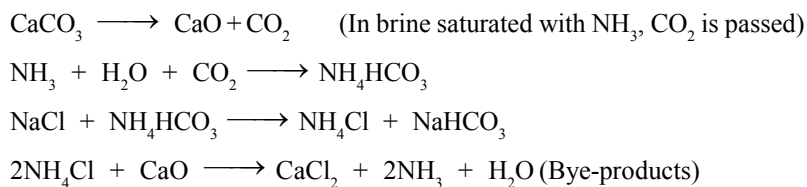


(e) **Uses**

- (i) In the manufacture of soap, rayon, dyes, paper and drugs.
- (ii) In petroleum refining.

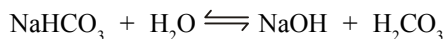
4. **SODIUM BICARBONATE OR BAKING SODA ( $NaHCO_3$ )**

(a) **Preparation : Solvay process** (Commercial Scale)

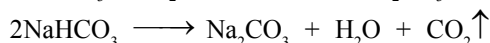


(b) **Properties**

**Hydrolysis**

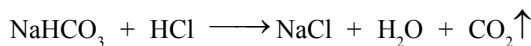


**Effect of heat** (temp. > 100°C)

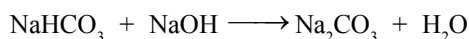


(Process occurs during preparation of cake)

**Reaction with acids** – gives  $CO_2$



**Reaction with base**



(c) **Uses**

- (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)

### 5. SODIUM CARBONATE OR WASHING SODA ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ )

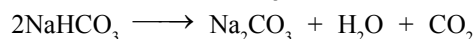
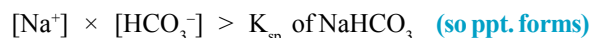
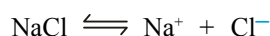
(a) **Occurrence** :  $\text{Na}_2\text{CO}_3$ —Soda ash.

(b) **Manufacture** : By solvay process

(i) Concentrated aqueous solution of NaCl is saturated with  $\text{NH}_3$ .

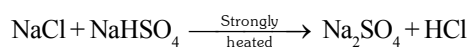
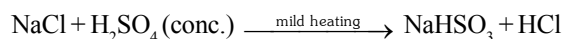
(ii) Current of  $\text{CO}_2$  passed through the solution.

(iii)  $\text{NaHCO}_3$  precipitated—

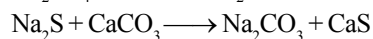
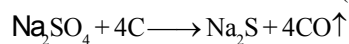


(iv) Potassium bicarbonate ( $\text{KHCO}_3$ ) cannot be prepared by solvay process as it is soluble in water.

(c) **Leblanc Process**



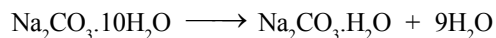
(Salt Cake)



(d) **Properties**

(i) **Efflorescence**

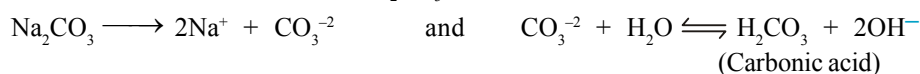
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  when exposed to air it gives out nine out of ten  $\text{H}_2\text{O}$  molecules.



(Monohydrate)

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

(ii) **Hydrolysis** : Aqueous solution of  $\text{Na}_2\text{CO}_3$  is alkaline in nature due to anionic hydrolysis.



(e) **Uses**

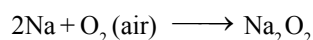
(i) For making fusion mixture ( $\text{Na}_2\text{CO}_3 + \text{K}_2\text{CO}_3$ )

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

(iii) In laundries and softening of water.

### 6. SODIUM PEROXIDE $\text{Na}_2\text{O}_2$

(a) Sodium peroxide is manufactured by heating sodium metal on aluminium trays in air (free from  $\text{CO}_2$ )

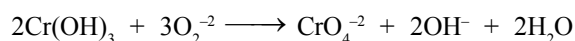


(i) When pure it is colourless, and the faint yellow colour of the usual product arises from the presence of a small amount of  $\text{NaO}_2$ .

(ii) When it is exposed, it comes in contact with moist air and turns white due to formation of  $\text{NaOH}$  and  $\text{Na}_2\text{CO}_3$ . Thus



(iii) Sodium peroxide is a powerful oxidizing agent and oxidizes chromium (III) hydroxide to sodium chromate, manganese (II) to sodium manganate and sulphides to sulphates.



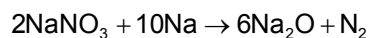
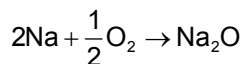
**(b) Uses**

- (i) Sodium peroxide is widely used as an oxidizing agent yielding in inorganic chemistry; its reaction with organic compounds are dangerously violent.
- (ii) Sodium readily combines with carbon dioxide, sodium carbonate and oxygen, it may be used for the purification of air in confined spaces such as submarines.
- (iii) It is also used as a bleaching agent because of its oxidizing property.
- (iv) Sodium peroxide is used in the manufacture of dyes, and many other chemicals such as benzoyl peroxide, sodium perborate etc.

**7. SODIUM MONOXIDE**

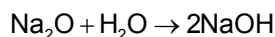
**Sodium Monoxide (Na<sub>2</sub>O)**

**Preparation** : It is obtained by burning sodium at 180°C in a limited supply of air or oxygen and distilling off the excess of Na in vacuum or by heating Na<sub>2</sub>O<sub>2</sub> (sodium peroxide), sodium nitrate (NaNO<sub>3</sub>) with sodium.

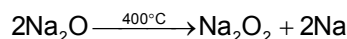


**Properties**

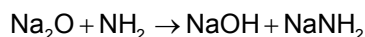
- (i) It is a white amorphous substance.
- (ii) It reacts with water violently forming NaOH.



- (iii) On heating above 400°C, it disproportionate to give peroxide and metallic sodium.



- (iv) It reacts with liquid ammonia forming sodamide and NaOH.



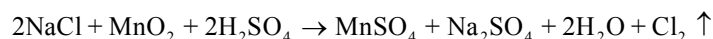
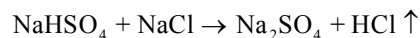
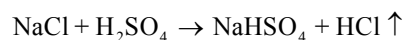
**8. SODIUM CHLORIDE (NaCl)**

**Preparation**

It is also called common salt occurs abundantly in nature as rock salt or halite. The most abundant source is sea-water where sodium chloride occurs to the extent of 2.6 – 2.9 percent. The sea water is exposed to the sun and air in large shallow pits. The gradual evaporation of water lead to the crystallization of the salt. The purification is done by dissolving the salt in minimum volume of water and filtering if necessary, to remove insoluble impurities. The solution in the saturated with a current of dry hydrogen chloride whereby crystals of pure sodium chloride separate out.

**Properties**

- (i) NaCl is a colourless crystalline salt, almost insoluble in alcohol and highly soluble in water.
- (ii) It gives rise to HCl when heated with conc. H<sub>2</sub>SO<sub>4</sub> and Cl<sub>2</sub> with MnO<sub>2</sub> plus H<sub>2</sub>SO<sub>4</sub>



**Ex.** Give reason for decreasing order of conductivity of following



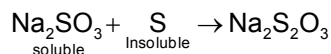
**Sol.** Ions are hydrate in solution. Since Li is very small it is heavily hydrate. This make the radius of the dydrated ions large and hence it moves only slowly (although  $\text{Li}^+$  is very small) and the radius of hydrated  $\text{Cs}^+$  ion is smaller than the radius of hydrated  $\text{Li}^+$ .

### 9. SODIUM THIOSULFATE ( $\text{Na}_2\text{S}_2\text{O}_3$ ) $\cdot$ 5 $\text{H}_2\text{O}$

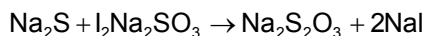
Sodium thiosulfate is a salt of an unstable acid  $\text{H}_2\text{S}_2\text{O}_3$  (Thiosulphuric acid)

#### Preparation

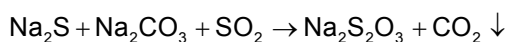
(i) sodium sulfite and flowers of S



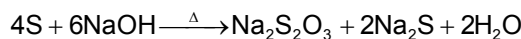
(ii) Spring's Reaction



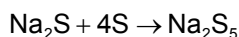
(iii) sodium carbonate solution with  $\text{SO}_2$



(iv) sulfur and caustic soda

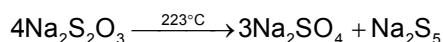


With excess pentasulfide forms

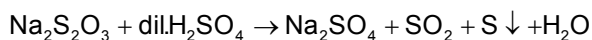
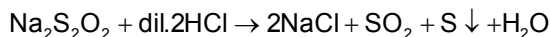


#### Chemical Properties

(i) Heating effect -  $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} \xrightarrow{215^\circ\text{C}} \text{Na}_2\text{S}_2\text{O}_3 + 5\text{H}_2\text{O}$

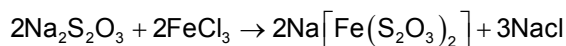


(ii) Acidification – Liberates  $\text{SO}_2$  and S ppt.



(iii) Reduction -  $\text{Na}_2\text{S}_2\text{O}_3 + \text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{HCl} + \text{S}$

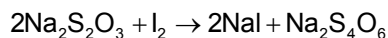
chlorine  
water



↓ (Purple)

$\text{FeCl}_2$  (green)

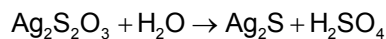
(iv) Oxidation - (oxidises by  $\text{I}_2$  quantitatively)



sodium

tetrathionate

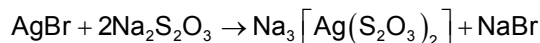
## CHEMISTRY FOR JEE MAIN & ADVANCED



Thus white ppt. of silver thiosulfate changes as

White  $\rightarrow$  yellow  $\rightarrow$  brown  $\rightarrow$  black

(vi) Silver halide action

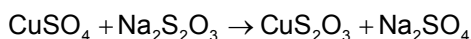


sodium argento

thiosulfate(colourless)

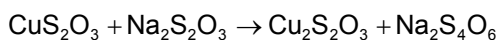
this property is used in photography fixing.

(vii) Action on  $\text{CuSO}_4$



cupric

thiosulfate

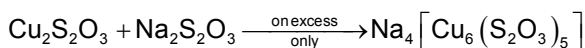


cuprous

sodium tetra

thiosulfate

thionate

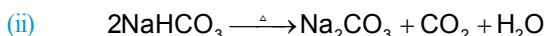


sodium cupro thiosulfate(sol)

### 10. SODIUM BICARBONATE ( $\text{NaHCO}_3$ )

**Preparation** : Sodium bicarbonate is prepared by the solvay process. It is infact formed in between the manufacture of washing sodaer

(i) Aqueous solution gives no colour with phenolphthalein but yellow colour with Metyl orange and hence is weakly basic.



(iii) Salts which give basic carbonates with washing soda give normal salts with the bicarbonate.



(iv) **It is used** as "SODA BICARB" to neutralisc stomach acidity.

(v) Sodium/Potassium salt of tartaric acid and  $\text{NaHCO}_3$  mixture is used as Baking Soda.

**Ex.** Based on lattice energy and other considerations which one of the following alkali metal chlorides is expected to have the highest melting point

(A) LiCl

(B) NaCl

(C) KCl

(D) RbCl

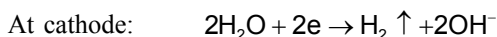
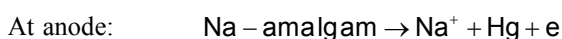
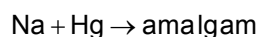
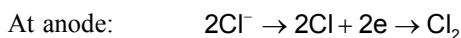
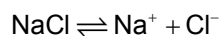
**Ans.** (B)

**Sol.** Although lattice energy of LiCl higher than NaCl but LiCl is covalent in nature and NaCl ionic there, after the melting point decreases as we move NaCl because the lattice energy decreases as a size of alkali metal atom increases (lattice energy  $\propto$  melting point of alkali metal halide)

- Ex.** NaOH is prepared by the method  
 (A) Down's cell (B) Castner cell  
 (C) Solvay process (D) Castner Kellnear cell

**Ans.** (D)

**Sol.** The cell involves the following reaction,



- Ex.** Sodium thiosulphate ( $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ ) is used in photography to  
 (A) Reduce silver bromide to silver  
 (B) Convert metallic silver to silver salt  
 (C) Remove unreacted AgBr as a soluble silver thiosulphate complex  
 (D) Remove unreacted silver

**Ans.** (C)

- Ex.** Which of the following pair can't exist in solution  
 (A)  $\text{NaHCO}_3$  and  $\text{NaOH}$  (B)  $\text{Na}_2\text{CO}_3$  and  $\text{NaOH}$   
 (C)  $\text{Na}_2\text{CO}_3$  and  $\text{NaCl}$  (D)  $\text{NaHCO}_3$  and  $\text{NaCl}$

**Ans.** (A)

## POTASSIUM

### Potassium and its compounds

- Atomic Number : 19
- Electronic Configuration : 2, 8, 8, 1
- Valency and Oxidation state: + 1
- Occurrence** : Potassium is highly reactive metal. It does not occur in nature in free form. Its important minerals are:  
 (a) Pearl ash,  $\text{K}_2\text{CO}_3$  (b) Sylvine,  $\text{KCl}$  (c) Indian Salt petre,  $\text{KNO}_3$  (d) Carnallite,  $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$
- Extraction** : It may be prepared by any of the following methods:  
 (a) By electrolysis of fused potassium cyanide  
 (b) by heating potassium fluoride with calcium.  
 (c) By heating potassium carbonate with carbon (coke)  
 (d) By electrolysis of molten  $\text{KCl}$  containing a little  $\text{KF}$ .

6. **Properties**

- (a) Potassium is a silvery white metal (M.P. 65°).
- (b) It is lighter and softer than sodium. It resembles sodium in its chemical properties, but is more reactive than it.
- (c) Potassium decomposes cold water more vigorously than sodium. The heat of reaction is so high that the hydrogen gas evolved catches fire when excess of metal is burnt in air or oxygen under reduced pressure,  $K_2O$  is formed. When heated in excess of air, potassium forms superoxide ( $KO_2$ )
- (d) Potassium (40) is feebly radioactive. It is a  $\beta$ -emitter. It has 3 isotopes of mass 39, 40 and 41 of which the first one is present in larger proportions. The radioactivity is due to the presence of isotope of mass number 40 which makes up about 0.012% of the metal.

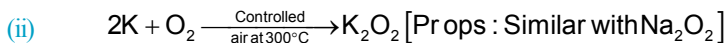
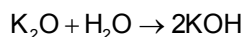
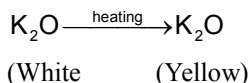
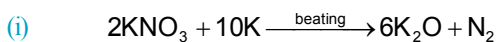
7. **Uses**

- (a) Potassium is used in photoelectric cell.
- (b) An alloy of sodium and potassium which is liquid is used in high temperature thermometers.

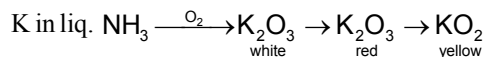
(i) **Oxides of potassium**

	$K_2O$ ,	$K_2O_2$ ,	$K_2O_3$ ,	$KO_2$	and	$KO_3$
<b>Colours</b>	White	White	Red	Bright Yellow		Orange Solid

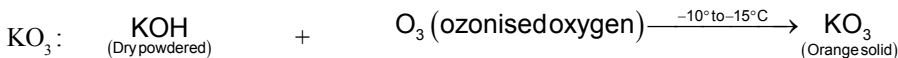
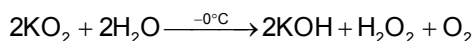
**Preparation**



(iii) Passage of  $O_2$  through a blue solution of K in liquid  $NH_3$  yields oxides  $K_2O_2$  (white),  $K_2O_3$  (red), and  $KO_2$  (deep yellow) i.e.

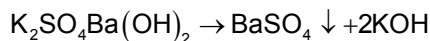


$KO_2$  reacts with  $H_2O$  and produces  $H_2O_2$  and  $O_2$  both



(ii) **Potassium hydroxide (Caustic Potash) KOH**

**Preparation:** It may be prepared like that of NaOH, i.e. the electrolysis of KCl solution and by the action of lime on potassium carbonate. It may also be obtained by the action of baryta  $Ba(OH)_2$  on potassium sulphate

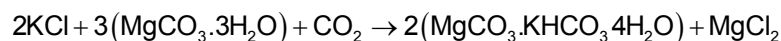


**Properties:** Its properties are similar to those of sodium hydroxide. However, it is a stronger alkali and more soluble in alcohol and is therefore, used in organic reactions instead of caustic soda. Moreover, it is a better absorber of  $CO_2$  than NaOH because potassium carbonate is more soluble and does not separate out. Its aqueous solution is known as potash lye.

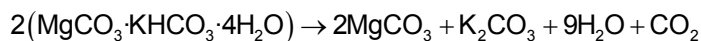


**(iii) Potassium carbonate ( $K_2CO_3$ )** : It is also known as pearl ash.

**Preparation:** It is made by passing  $CO_2$  into a conc. solution of the chloride, containing hydrated magnesium carbonate in suspension at  $20^\circ C$  when an insoluble potassium hydrogen magnesium carbonate is precipitated.



The precipitate is separated by filtration, and then decomposed either by heating with water under pressure at  $140^\circ C$  or by the action of magnesium oxide below  $20^\circ C$ .



### Properties

(i) It is white, deliquescent solid

(ii)  $K_2CO_3$  resembles  $Na_2CO_3$  in properties, but is more alkaline and more soluble than  $Na_2CO_3$

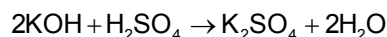
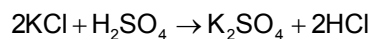
### (iv) Potassium Bicarbonate ( $KHCO_3$ )

**Preparation** : It resembles sodium bicarbonate in all respects except that it is more soluble in water. It is used in medicine and in baking powders.

(v) Potassium Sulphate ( $K_2SO_4$ )

### Preparation

(i) By treating KCl or KOH with  $H_2SO_4$



(ii) By treating naturally occurring mineral, schonite ( $K_2SO_4 \cdot MgSO_4 \cdot 6H_2O$ ) with sylvine (KCl)



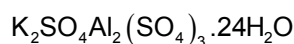
### Properties

(i) It is a white crystalline (m.p.  $1050^\circ C$ ) solid, not very soluble in water.

(ii) Unlike sodium sulphate, its crystals do not contain water of crystallisation.

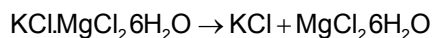
(iii) When heated with carbon, it is reduced to potassium sulphide.

(iv) It forms a series of double salts with the sulphates of trivalent metals, e.g. potash alum,



### (vi) Potassium Chloride (KCl)

**Preparation:** KCl is prepared from fused carnallite. Nearly pure KCl separates from the melt, leaving fused  $MgCl_2$  behind.



**Properties:** It is a colourless cubic crystal soluble in water. Its solubility increases almost linearly with temperature.

**Ex.** Which is an ore of potassium

(A) Carnallite                      (B) Cryolite                      (C) Bauxite                      (D) Dolomite

**Ans.** (A)

**Sol.** Carnallite -  $KCl, MgCl_2, 6H_2O$                       Cryolite -  $Na_3AlF_6$

Bauxite -  $(Al_2O_3 \cdot 2H_2O)$                       Dolomite -  $MgCO_3 \cdot CaCO_3$

## CHEMISTRY FOR JEE MAIN & ADVANCED

- Ex.** KF combination with HF to form  $\text{KHF}_2$ . The compound contains the species  
(A)  $\text{K}^+\text{F}^-$  and  $\text{H}^+$       (B)  $\text{K}^+$ ,  $\text{F}^-$  and HF      (C)  $\text{K}^+$  and  $[\text{HF}_2]^-$       (D)  $[\text{KHF}]^+$  and  $\text{F}^-$
- Ex.** When potassium dichromate crystal are heated with conc. HCl  
(A)  $\text{O}_2$  is evolved      (B) Chromyl chloride vapours are evolved  
(C)  $\text{Cl}_2$  is evolved      (D) No reaction takes place
- Ans.** (C)
- Ex.** When potassium ferrocyanide crystals are heated with concentrated sulphuric acid, the gas evolved is  
(A) Ammonia      (B) Sulphur dioxide      (C) Carbon dioxide      (D) Carbon monoxide
- Ans.** (D)

### ALKALINE EARTH METAL

#### Physical state

- (a) Two electrons in outer most shell & General formula  $ns^2$ .
- (b) Radium is radioactive element
- (c) All are greyish white.
- (d) These metals are harder than alkali metals.
- (e) These are diamagnetic and colourless in form of ions or in metal states.

#### Atomic size

Smaller than IA group elements, since extra charge on nucleus attracts the electron cloud.

- (a) Size increases gradually from Be to Ba  
 $\text{Be} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$
- (b) In s-block elements  
Be is the smallest, Cs is the biggest

#### Softness

- (a) These metals are slightly harder than IA group because of -
  - (i) Smaller atomic size
  - (ii) FCC, HCP crystal structures
  - (iii) Packing capacity 74%
  - (iv) Stronger metallic bond due to presence of two electrons in valence shell.
- (b) Be is the hardest metal in s-block.

#### Melting point and boiling point

- (a) 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.
- (b) Decreasing order of melting point and boiling point is  
 $\text{Be} > \text{Ca} > \text{Sr} > \text{Ba} > \text{Mg}$
- (c) Melting point and Boiling point of Ca, Sr and Ba is higher than Mg because of presence of d-orbitals in the outer most shell, which forms stronger metallic bond.

### Electro positive character or metallic character

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

### Flame Test

- (a) Be and Mg atoms, due to small size, bind their electrons more strongly, so are not excited to higher level, hence no flame test.
- (b) Other elements give characteristic colour to flame

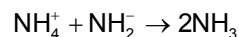
Ca-Brick red

Sr-crimson red

Ba-Apple green

### Reaction With $\text{NH}_3$

- (a) On increasing metal ion concentration solution converts into bronze colour due to cluster formation of metal ions.
- (b) Solubility in liquid ammonia
- (i) Only Ca, Sr and Ba give blue solution of ammoniated electron.
- (ii) Be and Mg are small in size and have high ionisation potential so do not dissolve in liquid  $\text{NH}_3$ .
- (iii) Dark blue colour of solution becomes faded if it is allowed to stand for a long time, it is because of metal amide formation.
- (iv) Blue colour of solution disappears on addition of ammonium salt, due to  $\text{NH}_3$  formation

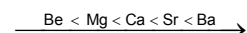


### Photo Electric Effect

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

### Standard oxidation potential

- (a) They have lower values of standard oxidation potential due to their small size
- (b) Increasing order of standard oxidation potential is



- (c) Tendency of losing electron increases

### Hydration energy (heat of hydration)

- (a) Due to smaller ionic size and higher charge density their hydration energy is high
- (b) Its decreasing order is



- (c) hydration energy  $\propto 1/\text{cation size}$

### Reducing property

- (a) Less reductant than alkali metals
- (b) Order of reducing property in aqueous and gaseous medium is



### Reaction with AIR

- (a) Except Be, these metals are easily tarnished in air, as a layer of oxide is formed on the surface.
- (b) Barium in powdered form, bursts into flame on exposure to air.
- (c) In moist air, except Be all the elements convert into carbonates.
- (d) In dry air Be and Mg give nitride and oxide both while others give only oxides.

### Reactivity with oxygen

(a) Alkaline earth metals react with  $O_2$  to form 'MO' Type oxides

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

(b) But Ca, Sr and Ba due to low ionisation potential and more reactivity, forms  $MO_2$  (peroxides) at low temperature.

Ex.  $CaO_2, SrO_2, BaO_2$

(c) Peroxides are coloured due to Lattice defect.

(d) BeO shows amphoteric property.

MgO → weak base

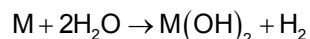
CaO, SrO & BaO → Strong base

(e) Basic properties increase from Be to Ba

(f) Its stability order general oxide > Peroxide > super oxide

### Reaction with water

(a) These metals react slowly with water giving  $H_2$  and metal hydroxides



(b) Be does not react with water

(c) Mg reacts only with hot water

(d) Ca, Sr, Ba react with cold water but not as energetically as alkali metals, order of reactivity

Ba > Sr > Ca > Mg > Be

(e) From  $Be(OH)_2$  to  $Ba(OH)_2$  basic property and stability increase.

### Halides

(a) Alkaline earth metals react with X (Halogen) to form  $MX_2$ .

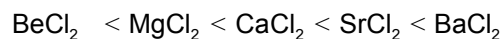
Ex. ( $BeCl_2, MgCl_2, CaCl_2$  etc.)

(b) Ionic nature of  $MX_2$  increases from  $BeCl_2$  to  $BaCl_2$

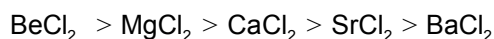
(c) Ba burns in contact with  $Cl_2$

(d) Hydrolytic nature of these halides decreases from  $BeCl_2$  to  $BaCl_2$

(e)  $BeCl_2$  and  $MgCl_2$  are covalent in nature. Order of ionic nature



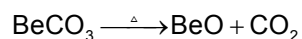
Solubility in water



### Carbonates

(a) All the alkaline earth metals form  $MCO_3$  type carbonates

(b) Except  $BeCO_3$ , all the carbonates are stable towards heat



(c) Order of decreasing stability

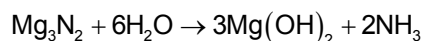
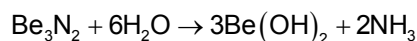


### Nitrates

- (a) Alkaline earth metals form  $M(\text{NO}_3)_2$  type nitrates. (M – Alkaline earth metal).
- (b) Stability increases from  $\text{Be}(\text{NO}_3)_2$  to  $\text{Ba}(\text{NO}_3)_2$  but these are less stable than IA group, due to smaller atomic size.
- (c) All alkaline earth metal nitrates on heating give oxides and  $\text{NO}_2 + \text{O}_2$
- $$M(\text{NO}_3)_2 \xrightarrow{\Delta} \text{Oxide} + \text{NO}_2 + \text{O}_2$$
- (d)  $\text{Be}(\text{NO}_3)_2$  forms a layer of BeO on its surface so reaction stops.

### Nitrides

Only Be and Mg burn in  $\text{N}_2$  to give  $\text{M}_3\text{N}_2$  ( $\text{Be}_3\text{N}_2, \text{Mg}_3\text{N}_2$ )



### Formation of amalgam

Shows same properties as alkali metals.

### Sulphates

- (a) Alkaline earth metals form  $\text{M}_2\text{SO}_4$  type sulphates
- (b) Ionic nature of alkaline earth metal sulphate increases from Be to Ba



(c) Solubility decreases from  $\text{BeSO}_4$  to  $\text{BaSO}_4$  as  $\text{Be}^{+2}$  and  $\text{Mg}^{+2}$  are of small size so their hydration energy is high. Hydration energy > Lattice energy.

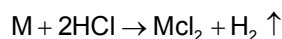
(d) Order of solubility -



Ionic nature increases, Thermal stability increases

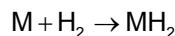
### Reaction with acids

Freely reacts with acids and displaces hydrogen

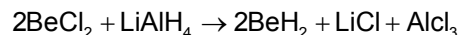


### Hydrides

(a) Except Be, all alkaline earth metals form hydrides ( $\text{MH}_2$ ) on heating directly with  $\text{H}_2$ .



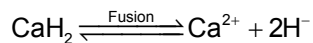
(b)  $\text{BeH}_2$  is prepared by the action of  $\text{LiAlH}_4$  on  $\text{BeCl}_2$



(c)  $\text{BeH}_2$  and  $\text{MgH}_2$  are covalent while other hydrides are ionic

## CHEMISTRY FOR JEE MAIN & ADVANCED

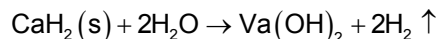
(d) The ionic hydrides of Ca, Sr, Ba liberated  $H_2$  at anode and metal t cathode.



(e) The stability of hydrides decreases from Be to Ba.

(f) The hydrides having higher reactivity for water, dissolves readily and produce hydrogen gas.

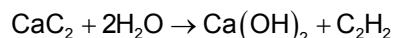
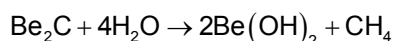
(g) The hydrides having higher reactivity for water, dissolves readily and produce hydrogen gas.



**Carbides :** BeO when heated with C at about  $2000^\circ C$ ,  $Be_2C$  is formed. The metals Mg, Ca, Sr and Ba form carbides for the formula  $MC_2$ . These carbides are formed when the metal or the metal oxide is heated with carbon in an electric furnace.

These carbides are ionic in nature.

$Be_2C$  yields methane on hydrolysis whereas carbides of other metals yield acetylene



**Complex compounds :** Alkaline earth metals have a tendency to form stable complexes. Their ability to form complexes is more as compared to alkali metals. This is because of the smaller size and higher charge of the alkaline earth metal ions as compared to alkali metal ions. Among the elements of group-2, beryllium and magnesium, have greater tendency of complexation.

**Ex.** In alkaline earth metals, the electrons are more firmly held to the nucleus and hence

(A) Ionization energy of alkaline earth metals is greater than that of alkali metals

(B) Alkaline earths are less abundant in nature

(C) Reactivity of alkaline earth metals is greater than that of alkali metals

(D) Atoms of alkaline earth metals are bigger in size than alkali metals.

**Ans.** (A)

**Sol.** Ionization energies of alkaline earth metals are greater than those of alkali metals due to increased nuclear charge.

**Ex.** The ion having highest mobility in aqueous solution is

(A)  $Be^{2+}$

(B)  $Mg^{2+}$

(C)  $Ca^{2+}$

(D)  $Ba^{2+}$

**Ans.** (D)

**Sol.** Though  $Ba^{2+}$  is the heaviest ion but it is least hydrated. As a result, its size becomes lesser than any other hydrated alkaline earth metal cation and hence has the highest ionic mobility.

**Ex.** Alkaline earth metal show typical characteristics of

(A) Inner transition elements

(B) Noble gases

(C) Representative elements

(D) Transition elements

**Ans.** (C)

- Ex.** Which of the following is true for magnesium ?
- (A) It is more electropositive than sodium  
 (B) It is manufactured by electrolysis of aqueous magnesium chloride  
 (C) It is a strong reducing agent  
 (D) It resembles, in chemical properties, with its diagonally placed element Boron in 13 group of the periodic Table.
- Ans.** (C)

## EXTRACTION OF ALKALINE METALS

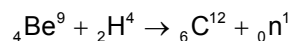
### (1) Beryllium

**Occurrence :** (1) Beryl-3BeO, Al<sub>2</sub>O<sub>3</sub>, 6 SiO<sub>2</sub>      (2) Chrysoberyl – BeO, Al<sub>2</sub>O<sub>3</sub>

Extraction

The powdered beryl is treated with sulphuric acid when sulphates of beryllium and aluminium are formed from which aluminium sulphate is removed double sulphate by the addition of (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub>. Beryllium sulphate remaining solution is separated and calcined at high temperature to give BeO. The oxide is finally reduced to the metal by heating with carbon.

- Uses:** (1) In making alloys /0 a very small quantity of Be added to Mg increases casting power  
 (2) It is highly permeable to X-rays and hence used for windows of x-ray tubes.  
 (3) In nuclear reactions as a source of neutrons, Bombarded with α-particles, Beryllium gives neutrons.



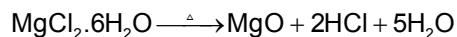
### (2) Magnesium

Importance ores of magnesium are

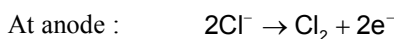
- (i) Magnesite, MgCO<sub>3</sub>      (ii) Dolomite, MgCO<sub>3</sub>·CaCO<sub>3</sub>  
 (iii) Carnallite, KCl·MgCl<sub>2</sub>·6H<sub>2</sub>O      (iv) Kieserite, MgSO<sub>4</sub>·H<sub>2</sub>O

Sea water also contains magnesium salts. Magnesium content in sea - water is about 0.13%. Magnesium is extracted by electrolysis of a molten mixture of anhydrous MgCl<sub>2</sub>, NaCl and CaCl<sub>2</sub>. Anhydrous MgCl<sub>2</sub> may be obtained from sea water or from carnallite.

Anhydrous MgCl<sub>2</sub> is prepared from hydrated MgCl<sub>2</sub>(MgCl<sub>2</sub>·6H<sub>2</sub>O) by passing a current of dry HCl gas over it. Anhydrous MgCl<sub>2</sub> cannot be prepared by direct heating because it results in formation of MgO which is refractory



During electrolysis, the following reactions take place.

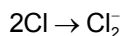


3. Calcium

Ores of Calcium

- (i) Gypsum,  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$       (ii) Lime stone,  $\text{CaCO}_3$       (iii) Marble,  $\text{CaCO}_3$   
 (iv) Fluorspar,  $\text{CaF}_2$       (v) Anhydrite,  $\text{CaSO}_4$       (vi) Dolomite,  $\text{CaCO}_3, \text{MgCO}_3$

Calcium is prepared by the electrolysis of a fused mixture of anhydrous calcium chloride and calcium fluoride. On passing electric current, calcium is liberated at the cathode.



**Anomalous behaviour of Be and its diagonal relationship with Al**

Anomalous behaviour of Be and its diagonal relationship with Al is illustrated by the following points :

- (i) Unlike group -2 elements but like aluminium, beryllium forms covalent compounds.
- (ii) The hydroxides of beryllium,  $[\text{Be}(\text{OH})_2]$  and aluminium,  $[\text{Al}(\text{OH})_3]$  are amphoteric in nature, whereas those of other elements of group - 2 are basic in nature.
- (iii) Beryllium and aluminium forms fluoro complexes in solution.
- (iv) The oxides of both Be and Al, i.e,  $\text{BeO}$  and  $\text{Al}_2\text{O}_3$  are high melting insoluble solids.
- (v)  $\text{BeCl}_2$  like  $\text{Al}_2\text{Cl}_6$  has a bridged polymeric structure.
- (vi) The salts of beryllium as well as aluminium are extensively hydrolyzed.
- (vii) Be like Al, is rendered passive by concentrated  $\text{HNO}_3$
- (viii) Be forms many complex like Al
- (ix) Be and Al react with  $\text{NaOH}$  solution liberating  $\text{H}_2$ .
- (x)  $\text{Be}_2\text{C}$  and  $\text{Al}_4\text{C}_3$  react with water to give methane.
- (xi) Halides of the two elements have similar solubilities.
- (xii) The two metals do not impart colour to the bunsen flame.

**Ex.** Which of the following elements will form a compound with oxygen in which one atom of the element is combined with one atom of oxygen to form an oxide of formula  $\text{MO}$ ?

- (A) Be      (B) Ne      (C) Al      (D) Cl

**Ans** (A)

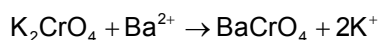
**Sol.** Be forms  $\text{BeO}$ , Al forms  $\text{Al}_2\text{O}_3$ , Cl forms  $\text{Cl}_2\text{O}_7, \text{Cl}_2\text{O}, \text{ClO}_2$  etc. while Ne does not react.

**Ex.** The cation which gives a yellow precipitate with potassium chromate is

- (A)  $\text{SrCO}_3$       (B)  $\text{BaCO}_3$       (C)  $\text{CaCO}_3$       (D)  $\text{MgCO}_3$

**Ans** (B)

**Sol.** Barium gives a yellow ppt. of barium chromate with pot. chromate



**Ex.** The atom with atomic number 12 will most likely combine chemically with the atom whose atomic number is

- (A) 3      (B) 10      (C) 11      (D) 16

**Ans.** (D)



**Ex.** A piece of magnesium ribbon was heated to redness in an atmosphere of nitrogen and then cooled with water. The gas evolved is

- (A) Ammonia                      (B) Hydrogen                      (C) Nitrogen                      (D) Oxygen

**Ans.** (A)

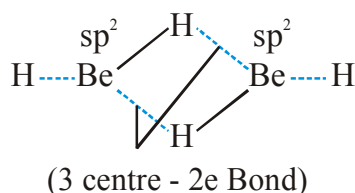
## COMPOUND OF ALKALINE EARTH METALS

### (1) Beryllium

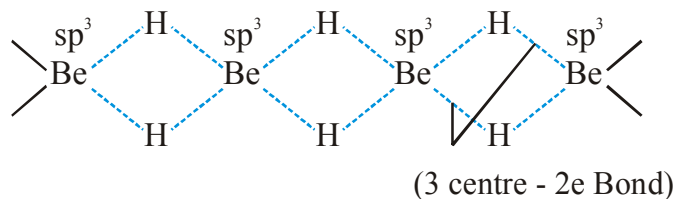
(i) **Oxides and hydroxides** : Be forms only the monoxides: all others from monoxide and peroxide (eg. BaO, BaO<sub>2</sub>). The basic nature of the oxides and hydroxides of the metals increases down the group. BeO is amphoteric, CaO, SrO and BaO are basic. All these oxides dissolve in H<sub>2</sub>O giving the hydroxide (BeO is nearly insoluble). BeO dissolves in acids and alkalis to give salts.

$(\text{BeO}) + 2\text{HCl} \rightarrow \text{BeCl}_2 + \text{H}_2\text{O}$  ;  $\text{BeO} + 2\text{NaOH} \rightarrow \text{Na}_2\text{BeO}_2 + \text{H}_2\text{O}$  The solubility of the hydroxides increase from Be to Ba. Ba(OH)<sub>2</sub> (Baryta) is moderately soluble in water.

(ii) **Hydrides** : BeH<sub>2</sub> is covalent MgH<sub>2</sub> is partially covalent and the remaining hydrides are largely ionic. BeH<sub>2</sub> is polymeric solid, involving hydrogen bridging.

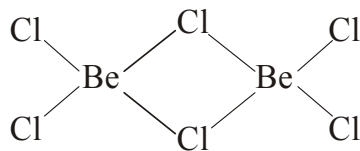


### Dimer form of BeH<sub>2</sub> in vapour phase



### Polymeric form of BeH<sub>2</sub> in solid state

(iii) **Halides** : II group metal when heated with halogens, give the halides. Beryllium halides (eg. BeF<sub>2</sub> and BeCl<sub>2</sub>) are covalent due to the small size of Be<sup>2+</sup> (see Fajan's rules) and hence have comparatively low melting points and boiling points. The chlorides and fluorides of the other metals of the group are ionic solid BeCl<sub>2</sub> in the solid state is polymeric containing chains of the form.



In the vapour state at high temperature BeCl<sub>2</sub> is a monomer.

The halides are deliquescent and they form hydrated salts. The solubility of halides in water decrease down the group from Ca to Ba.

## CHEMISTRY FOR JEE MAIN & ADVANCED

- (iv) **Sulphates** :  $\text{BeSO}_4$  is slightly soluble in water. The solubility of the sulphates of the metals decreases down the group  $\text{BaSO}_4$  and  $\text{RaSO}_4$  are practically insoluble.

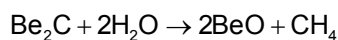
[**Note** : The solubility of an ionic solid in water depends on the lattice energy of the solid and the hydrate ion energy for the ions of the solid. A large value of the former would lower the solubility, hydration energy factor has a dominant role. Since the hydration of ion decreases from,  $\text{Be}^{2+}$ , to hydroxides of this group, the lattice energy has a dominant role and hence the solubility of these salts increases down the group.]

- (v) **Carbonates and bicarbonates** : The carbonates of alkaline earth metals are insoluble in water: the bicarbonates are soluble. The carbonates and bicarbonates decompose on heating to give  $\text{CO}_2$  gas.

- (vi) **Carbides** of II group metal are ionic carbides, which may be in a methanide or acetylide or allylide.

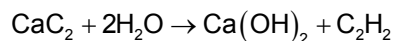
### Beryllium Carbide

$\text{Be}_2\text{C}$  (methanide) reacts with water to give methane.



### Calcium Carbide

$\text{CaC}_2$  (acetylide) is decomposed by  $\text{H}_2\text{O}$  to give acetylene.



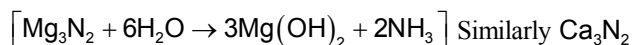
### Magnesium Carbide

$\text{Mg}_2\text{C}_3$  (allylide) reacts with water to give allylene (methyl acetylene)



- (vii) **Nitrides**

Example :  $\text{Mg}_3\text{N}_2$  decomposed by  $\text{H}_2\text{O}$  liberating  $\text{NH}_3$ .



Complexes : Be is the best complexing agent in this group. The complexing tendency decrease down the group.  $\text{Mg}(\text{Oxine})_2$  is the complex formed by  $\text{Mg}^{2+}$  with 8 hydroxy quinoline(Oxine) and this forms the basis of quantitative estimation of Mg.

- Ex.** A metal M readily forms water soluble sulphate  $\text{MSO}_4$  water insoluble hydroxide  $\text{M}(\text{OH})_2$  and oxide  $\text{M}_2\text{O}$  which becomes inert on heating. The hydroxide is soluble in NaOH. The M is

(A) Be (B) Mg (C) Ca (D) Sr

**Ans.** (A)

**Sol.**  $\text{Be}(\text{OH})_2$  is amphoteric in nature so it is soluble in NaOH

- Ex.** Amongst the following hydroxides, the one which has the lowest value of  $K_{sp}$  at ordinary temperature is :

(A)  $\text{Mg}(\text{OH})_2$  (B)  $\text{Ca}(\text{OH})_2$  (C)  $\text{Ba}(\text{OH})_2$  (D)  $\text{Be}(\text{OH})_2$

**Ans.** (D)

## (2) Magnesium

- (i) **Magnesium Sulphate ( $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ ) epsom salt**

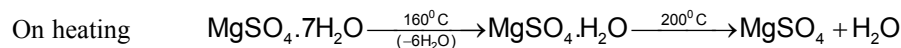
### Preparation

From magnesite or dolomite (Epsom salt)

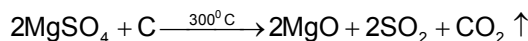


**Physical properties**

Colourless, efflorescent crystalline substance which forms a number of hydrate

**Chemical properties**

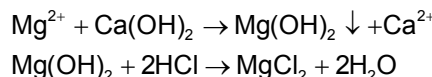
It is reduced by lamp black

**USES**

1.  $\text{MgSO}_4$  is manufactured by dissolving keiserite ( $\text{MgSO}_4 \cdot \text{H}_2\text{O}$ ) in boiling water & then crystallising the solution.
2.  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  is called as Epsom salt & shows deliquescence efflorescence & has purgative action.
3. Platinized  $\text{MgSO}_4$  is used as catalyst in Grillo's process for manufacture of  $\text{H}_2\text{SO}_4$ .

**(ii) Magnesium Chloride  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$** 

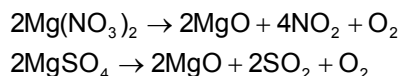
Magnesium chloride can be prepared by fractional crystallization of carnallite,  $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ . It can also be obtained from sea water. Magnesium salts present in sea-water are precipitated as  $\text{Mg}(\text{OH})_2$  by adding lime. The precipitate is filtered and dissolved in HCl solution. The solution on crystallization yields crystals of  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ .



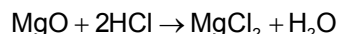
A mixture of magnesium chloride and magnesium oxide is known as Sorel cement or magnesium cement. This mixture reacts with water and sets into a solid mass having composition

**(iii) Magnesium Oxide ( $\text{MgO}$ )**

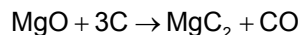
**Preparation** : It is prepared by burning Mg in air, by heating hydroxide, nitrate, sulphate or carbonate. Oxide prepared by heating magnesite ( $\text{MgCO}_3$ ) is called calcined magnesia.

**Properties**

- (i) Magnesium oxide is a light white powder, which is highly infusible (m.p.  $2800^\circ\text{C}$ ), and only slightly soluble in water.
- (ii) It is a basic oxide and hence reacts with acids to form salts.



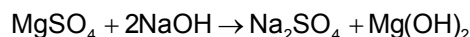
- (iii) It is reduced by carbon at very high temperature to form magnesium carbide.



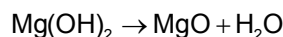
- (iv) It dissolves in aqueous solution of magnesium chloride or bromide forming basic salts such as  $\text{MgCl}_2 \cdot 9\text{Mg}(\text{OH})_2 \cdot 5\text{H}_2\text{O}$ .

**(iv) Magnesium Hydroxide ( $\text{Mg}(\text{OH})_2$ )**

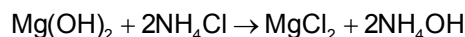
**Preparation** : It is obtained by adding caustic soda solution to a solution of magnesium sulphate or chloride.

**Properties**

- (i) It is converted into its oxide on heating.



- (ii) It dissolves in  $\text{NH}_4\text{Cl}$  solution easily.



## CHEMISTRY FOR JEE MAIN & ADVANCED

- (v) **Magnesium Carbonate, (MgCO<sub>3</sub>):** It occurs in nature as magnesite, MgCO<sub>3</sub> and dolomite, MgCO<sub>3</sub>.CaCO<sub>3</sub>.

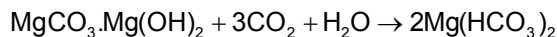
(i) By adding sodium bicarbonate solution to a hot solution of magnesium salt.



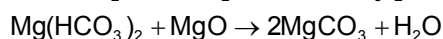
In case Na<sub>2</sub>CO<sub>3</sub> is used in place of NaHCO<sub>3</sub>, a basic carbonate is obtained.



When CO<sub>2</sub> gas is passed through the suspension of basic carbonate, magnesium bicarbonate is formed which on heating forms MgCO<sub>3</sub>.



(ii) Pure magnesium bicarbonate and hence carbonate can be prepared by passing CO<sub>2</sub> through suspension of magnesium oxide in water.



### Properties

(i) Magnesium carbonate is a white powder, insoluble in water.

(ii) It dissolves in acids forming salts.

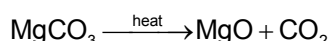


(iii) Its aqueous solution also dissolves on passing a current of carbon dioxide because of the formation of Mg(HCO<sub>3</sub>)<sub>2</sub>, a temporary hardness causing substance.



A solution of magnesium bicarbonate containing 12 gm. of the salt per 100 ml is known as fluid magnesia.

(iv) It decomposes on heating to form magnesium oxide.



### 6. Sorel cement

Magnesium oxide when mixed with a saturated solution of magnesium chloride sets to a hard mass like cement known as magnesia cement or Sorel cement. The composition is



**Mg(ClO<sub>4</sub>)<sub>2</sub> is known as anhydron**

**Ex.** Magnesium burns in CO<sub>2</sub> to form

- (A) MgO+C                      (B) MgO+CO                      (C) MgCO<sub>3</sub>                      (D) MgO

**Ans.** (A)

**Sol.** Mg burns in CO<sub>2</sub> to give MgO and C.

**Ex.** which of the following is incorrect?

- (A) Mg burns in air releasing dazzling light rich in UV rays  
(B) CaCl<sub>2</sub>.6H<sub>2</sub>O when mixed with ice gives freezing mixture  
(C) Mg cannot form complexes  
(D) Be can form complexes due to its very small.

**Ans.** (C)

**Sol.** Mg can form complexes and chlorophyll is one of such examples.

- Ex.** Flame test is not given by  
 (A) Be (B) Mg (C) Ca (D) Sr
- Ans.** (A,B)

- Ex.** Pure anhydrous  $\text{MgCl}_2$  can be prepared from the hydrated salt by  
 (A) Heating the hydrate with coke (B) Heating the hydrate with Mg ribbon  
 (C) Melting the hydrate  
 (D) Heating the hydrate to red heat in an atmosphere of HCl gas
- Ans.** (D)

### (3) Calcium

#### (i) Calcium oxide Quick lime, Burnt lime (CaO)

##### Preparation

It is prepared by heating lime stone ( $\text{CaCO}_3$ ) at  $800^\circ\text{C}$

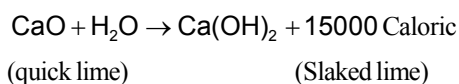


##### Condition for Good Yield

- (i) Since the reaction is reversible, therefore the carbon dioxide formed must be removed as soon as it is formed so that the reaction remain in the forward direction.
- (ii) Temperatone will react with CaO to form  $\text{CaSiO}_3$ .

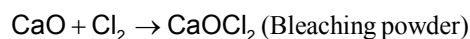
##### Properties

###### (i) Action of water

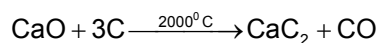


Paste of lime in water is called milk of lime, while its hydrate is known as lime water.

###### (ii) With moist chlorine :



###### (iii) When heated with carbon, it forms calcium carbide.

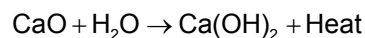


##### Uses

- (i) In purification of sugar  
 (ii) As basic lining in furnances.
- (ii) Calcium hydroxide, Slaked lime  $\text{Ca(OH)}_2$**

##### Preparation

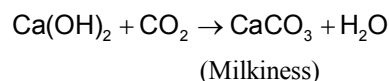
By the action of water on quick lime.



Means it is an exothermic reaction.

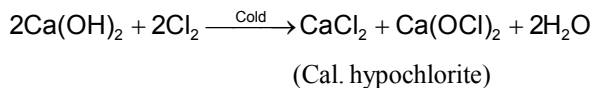
##### Properties

- (i) Sparingly soluble in water and its solubility decreases with increases in temperature.  
 (ii) Action of  $\text{CO}_2$  : Lime water turns milky on passing  $\text{CO}_2$  gas



**(iii) Action of chlorine**

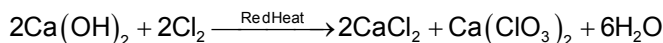
(a) Cold Condition :



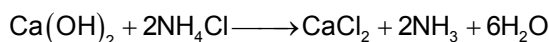
(b) Below 35° :  $3\text{Ca}(\text{OH})_2 + 2\text{Cl}_2 \xrightarrow[35^\circ\text{C}]{\text{below}} \text{CaOCl}_2 + \text{H}_2\text{O}$

(c) On gentle Heating :  $6\text{Ca}(\text{OH})_2 + 6\text{Cl}_2 \xrightarrow{\text{Heat}} 5\text{CaOCl}_2 + \text{H}_2\text{O}$

(d) On Red Hot :



(iv) Action of Ammonia

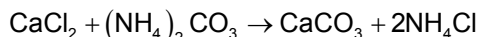


**Uses**

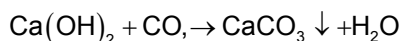
- (i) For softening of hard water.
- (ii) For purification of sugar and coal gas
- (iii) In preparation of bleaching powder, white wash, plaster etc.

**(iii) Calcium Carbonate, (CaCO<sub>3</sub>)** : Calcium carbonate is found in nature as limestone, marble, coral ice land spar, calcite, chalk, dolomite, etc.

**Preparation:** It is prepared as a white powder, known as precipitated chalk by dissolving marble or limestone in hydrochloric acid followed by precipitation with sod. or ammonium carbonate.

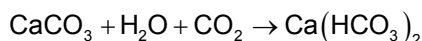


In the laboratory, it is prepared by passing CO<sub>2</sub> through lime-water or by adding sod. or ammo carbonate solution of CaCl<sub>2</sub>.



**Properties**

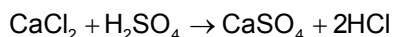
- (i) It is a white powder and exists in two crystalline forms: calcite and aragonite.
- (ii) It is insoluble in water but dissolves in the presence of CO<sub>2</sub> due to the formation of calcium bicarbonate.



**(iv) Calcium Bicarbonate, (Ca(HCO<sub>3</sub>)<sub>2</sub>)** : It does not exist in solid state. however, its solution can be prepared by passing CO<sub>2</sub> gas through a suspension of calcium carbonate in water. On warming, it decomposes to calcium carbonate along with the evolution of carbon dioxide gas.

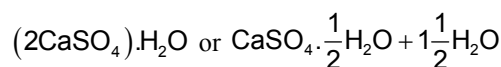
**(v) Calcium Sulphate, (CaSO<sub>4</sub>)** : It occurs in nature as anhydride (CaSO<sub>4</sub>) and gypsum (CaSO<sub>4</sub>·2H<sub>2</sub>O)

**Preparation:** It may be prepared by adding dilute sulphuric acid or soluble sulphate (e.g. Na<sub>2</sub>SO<sub>4</sub>) to the solution of a calcium salt.

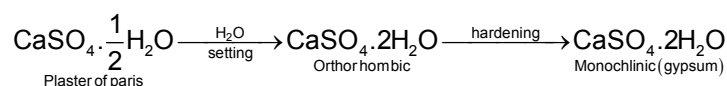


**Properties**

- (i) It is a white crystalline solid. Like slaked lime it is sparingly soluble in cold water and its solubility decreases further with the increase in temperature.
- (ii) It dissolves much more readily in dilute acids, even in acetic acid than in water. It dissolves also in ammonium sulphate solution forming  $\text{CaSO}_4 \cdot (\text{NH}_4)\text{SO}_4 \cdot \text{H}_2\text{O}$
- (iii) Monoclinic crystals of gypsum when heated, first change into orthorhombic form without any loss of water. On further heating to  $120^\circ\text{C}$ , it loses three-fourth of its water of crystallization and forms the hemihydrate,



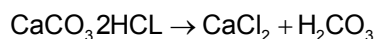
Plaster of Paris is a white powder which when mixed with water takes up its water of crystallization again, thus converted back into the dihydrate and sets to a hard mass with slight expansion. Setting takes about 10–15 minutes and may be catalysed by the addition of common salt or delayed by the use of alum or borax.



Plaster of Paris or gypsum when heated to about  $200^\circ\text{C}$  is converted into anhydrous calcium sulphate. The anhydrous form (anhydrite) is known as dead burnt because it does not set like Plaster of Paris when moistened with water.

**(vi) Calcium Chloride ( $\text{CaCl}_2$ )**

Preparation: It separates out as deliquescent crystals when a solution of lime or calcium carbonate in HCl is evaporated.



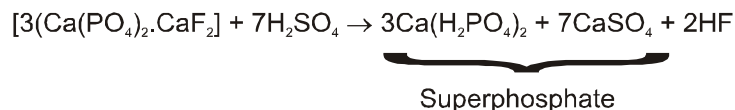
It separates out from the reaction mixture as  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ . The anhydrous salt is obtained on heating above  $200^\circ\text{C}$ .

Properties: It is a colourless, deliquescent salt, highly soluble in water. The anhydrous salt is an excellent drying agent.

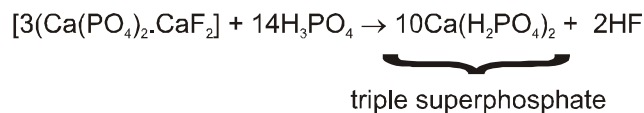
**(vii) Superphosphate**

Phosphate rocks such as fluorapatite  $[\text{3Ca}_3(\text{PO}_4)_2 \cdot \text{CaF}_2]$  are very insoluble, and thus are of no use to plants.

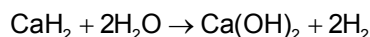
Superphosphate is made by treating phosphate rock with concentrated  $\text{H}_2\text{SO}_4$ . The acid salt  $\text{Ca}(\text{H}_2\text{PO}_4)_2$  is more soluble, and over a period of weeks the superphosphate will dissolve in the soil water.



“Triple superphosphate” is made in similar way, using  $\text{H}_3\text{PO}_4$  to avoid the formation of the waste product  $\text{CaSO}_4$

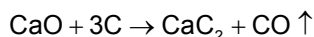
**(viii) Hydrolith**

$\text{CaH}_2$  is technically called hydrolith and is used on a large scale for the production of hydrogen

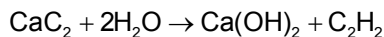


(ix) **Calcium Carbide, CaC<sub>2</sub>**

Calcium carbide is prepared by heating quick lime with coke in an electric furnace.



CaC<sub>2</sub> on reaction with water produces acetylene gas



**Ex.** Which of the following substance is used as dehydrating agent in laboratory

- (A) Calcium chloride (B) Sodium chloride  
(C) Sodium carbonate (D) Potassium nitrate

**Ans.** (A)

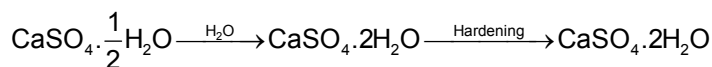
**Sol.** CaCl<sub>2</sub> because it is hygroscopic.

**Ex.** Setting of plaster of paris is

- (A) Oxidation with atmospheric CO<sub>2</sub> (B) Combination with atmospheric CO<sub>2</sub>  
(C) Dehydration (D) Hydration to yield another

**Ans.** (D)

**Sol.** Setting of plaster of paris is



orthorhombic

Mono orthorhombic Gypsum

The setting is due to formation of another hydrate.

**Ex.** Which of the following statements is false

- (A) CaOCl<sub>2</sub> gives OH<sup>-</sup>, Cl<sup>-</sup> and OCl<sup>-</sup> in aqueous solution  
(B) Diamond and graphite are allotropes  
(C) Bleaching action of Cl<sub>2</sub> in moist condition is not permanent  
(D) Calomel is HgCl<sub>2</sub>

**Ans.** (C)

**SIMILARITIES BETWEEN LITHIUM AND MAGNESIUM**

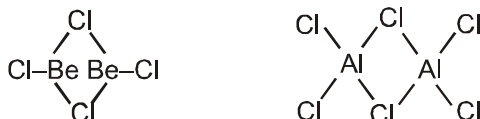
- (a) Both lithium and magnesium are harder and lighter than other elements in the respective groups.  
(b) Lithium and magnesium react slowly with cold water. Their oxides and hydroxides are much less soluble and their hydroxides decompose on heating. Both form a nitride by direct combination with nitrogen, Li<sub>3</sub>N and MgN<sub>2</sub>.  
(c) The oxides, Li<sub>2</sub>O and MgO do not combine with excess oxygen to give a peroxide or a super oxide.  
(d) The carbonates of lithium and magnesium decompose easily on heating to form the oxide and CO<sub>2</sub>. Solid bicarbonates are not formed by lithium and magnesium.  
(e) Both LiCl and MgCl<sub>2</sub> are soluble in ethanol.  
(f) Both LiCl and MgCl<sub>2</sub> are deliquescent and crystallise from aqueous solution as hydrates, LiCl·2H<sub>2</sub>O and MgCl<sub>2</sub>·8H<sub>2</sub>O.



**DIAGONAL SIMILARITY BETWEEN BERYLLIUM AND ALUMINIUM**

In many of its properties, beryllium resembles aluminium. Thus-

- (a) The two elements have same electronegativity and their charge / radius ratios.
- (b) Both metals are fairly resistant to the action of acids due to a protective film of oxide on the surface. Both metals are acted upon by strong alkalis to form soluble complexes, beryllates  $[\text{Be}(\text{OH})_4]^{2-}$  and aluminates,  $[\text{Al}(\text{OH})_4]^-$ .
- (c) The chlorides of both beryllium and aluminium



have bridged chloride structures in vapour phase.

- (d) Salts of these metals form hydrated ions, Ex.  $[\text{Be}(\text{OH}_2)_4]^{2+}$  and  $[\text{Al}(\text{OH}_2)_6]^{3+}$  in aqueous solutions. Due to similar charge/ radius ratios of beryllium and aluminium ions have strong tendency to form complexes. For example beryllium forms tetrahedral complexes such as  $\text{BeF}_4^{2-}$  and  $[\text{Be}(\text{C}_2\text{O}_4)_2]^{2-}$  and aluminium forms octahedral complexes like  $\text{AlF}_6^{3-}$  and  $[\text{Al}(\text{C}_2\text{O}_4)_3]^{3-}$

## Tips

1.	Active nitrogen	:	N(atomic nitrogen)
2.	Alums	:	$M_2'SO_4 \cdot M_2''(SO_4)_3 \cdot 24H_2O$  $M' = K^+, NH_4^+, Na^+ \text{ etc.}$  $M'' = Cr^{+3}, Al^{+3}, Fe^{+3} \text{ etc.}$
3.	Asbestos	:	$CaMg_3(SiO_3)_4$
4.	Arsine	:	$AsH_3$
5.	Aquaregia	:	Conc. $HNO_3$ + Conc. $HCl$ (1 : 3 part)
6.	Anhydrone	:	$Mg(CLO_4)_2$
7.	Argentiferrous galena	:	$PbS + Ag_2S$
8.	Borax	:	$Na_2B_4O_7 \cdot 10H_2O$
9.	Blue vitriol	:	$CuSO_4 \cdot 5H_2O$
10.	Barytes	:	$BaSO_4$
11.	Baryta water	:	$Ba(OH)_2$ solution
12.	Baryta	:	$BaO$
13.	Baking soda	:	$NaHCO_3$
14.	Bleaching powder	:	$CaOCl_2$
15.	Boranes	:	Hydride of borone
16.	Brine	:	$NaCl$ solution
17.	Calgon	:	$Na_2[Na_4(PO_3)_6]$
18.	Coinage metals	:	$Cu, Ag$ and $Au$
19.	Carborundum	:	$SiC$
20.	Cementite	:	$FeC$
21.	Caliche	:	$NaNO_3 + NaClO_3$
22.	Caustic soda	:	$NaOH$
23.	Caustic potash	:	$KOH$
24.	Calomel	:	$Hg_2Cl_2$
25.	Corrosive sublimate	:	$HgCl_2$
26.	Deuterium	:	${}_1H^2$ of D
27.	D.D.T.	:	p-dichloro, diphenyl, trichloroethane
28.	Dry ice	:	Solid $CO_2$
29.	Freon	:	$CF_2Cl_2$
30.	Ferric Alum	:	$K_2SO_4 \cdot Fe_2(SO_4)_3 \cdot 24H_2O$
31.	Fenton's reagent	:	$H_2O_2$ + few drops of $FeSO_4$
32.	Fusion's mixutre	:	$Na_2CO_3 + K_2CO_3$
33.	Fluid magnesia	:	12% solution of $Mg(HCO)_2$

34.	Fehling solution	:	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O} + \text{NaOH} + \text{Na, K tartarate}$
35.	King of metals	:	Gold
36.	Horn Silver	:	AgCl
37.	Green vitriol	:	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
38.	Graphite	:	An allotrope of carbon
39.	Gun powder	:	75% $\text{KNO}_3$ + 12% S + 13% charcoal (explosive)
40.	Glauber salt	:	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
41.	Hydrolith	:	$\text{CaH}_2$
42.	Heavy water	:	$\text{D}_2\text{O}$
43.	Hypo (sodium thiosulphate)	:	$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
44.	Heavy hydrogen	:	$\text{D}_2$
45.	King of chemicals	:	$\text{H}_2\text{SO}_4$
46.	Lime (quick lime or burnt lime)	:	CaO
47.	Lead pencil	:	Graphite
48.	Lime water	:	$\text{Ca}(\text{OH})_2$
49.	Laughing gas	:	$\text{N}_2\text{O}$
50.	Lunar Caustic	:	$\text{AgNO}_3$
51.	Litharge (Masscote)	:	PbO
52.	Lithopone	:	$(\text{ZnS} + \text{BaSO}_4)$ , a pigment
53.	Mortar	:	Slaked lime + sand (1 : 3 in water)
54.	mica	:	$\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2 \cdot 2\text{H}_2\text{O}$
55.	Mohr salt	:	$\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$
56.	Matte	:	$\text{Cu}_2\text{S} + \text{FeS}$
57.	Milk of lime	:	$\text{Ca}(\text{OH})_2$ in water
58.	Minium	:	$\text{Pb}_3\text{O}_4$
59.	Micro cosmic salt	:	$\text{NaNH}_4 \cdot \text{HPO}_4 \cdot 4\text{H}_2\text{O}$ (used in test of silicates)
60.	Milk of magnesia	:	Paste of $\text{Mg}(\text{OH})_2$ in water (Antacid)
61.	Magnesia	:	MgO
62.	Marsh gas	:	$\text{CH}_4$
63.	Nitrolim	:	$\text{CaCN}_2 + \text{C}$ (a fertilizer)
64.	Nascent Hydrogen	:	H at the moment of generation
65.	Nessler's reagent	:	$(\text{K}_2\text{HgI}_4 + \text{KOH})$ aqueous solution
66.	Indian saltpetre, Bengal salt petre	:	$\text{KNO}_3$
67.	Oil of vitriol	:	Conc. $\text{H}_2\text{SO}_4$
68.	Ozone	:	$\text{O}_3$
69.	Oleum	:	$\text{H}_2\text{S}_2\text{O}_7$
70.	Permutit (Zeolite)	:	$\text{Na}_2\text{Al}_2\text{SiO}_8 \cdot \text{XH}_2\text{O}$
71.	Pearl ash (Potash)	:	$\text{K}_2\text{CO}_3$
72.	Plaster of paris	:	$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$ or $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$

## CHEMISTRY FOR JEE MAIN & ADVANCED

73.	Philosopher's wool (chinese white)	:	ZnO (Zinc white)
74.	Phosgene	:	$\text{COCl}_2$
75.	Phosphine	:	$\text{PH}_3$
76.	Pig iron	:	Impure form of iron
77.	Producer gas	:	A mixture of $\text{CO} + \text{N}_2 + \text{H}_2$
78.	Quartz	:	$\text{SiO}_2$
79.	Refrigerant	:	$\text{CO}_2, \text{NH}_3, \text{CF}_2\text{Cl}_2$ etc.
80.	Red lead	:	$\text{Pb}_3\text{O}_4$
81.	Rochelle salt	:	Sodium - potassium tartarate
82.	Rust	:	$\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$
83.	Sorel's cement (Magnesia cement)	:	$\text{MgCl}_2 \cdot 5\text{MgO} \cdot x\text{H}_2\text{O}$
84.	Soda - lime	:	$\text{NaOH} + \text{CaO}$
85.	Soda ash	:	$\text{Na}_2\text{CO}_3$ (anhydrous)
86.	Slaked lime	:	$\text{Ca(OH)}_2$
87.	Stainless steel	:	An alloy of Fe, Cr and C
88.	Salt cake	:	$\text{Na}_2\text{SO}_4$ (anhydrous)
89.	Super phosphate	:	$\text{Ca(H}_2\text{PO}_4)_2 + 2\text{CaSO}_4$
90.	TNT	:	Trinitro toluene (an explosive)
91.	TNB	:	Trinitro benzene (an explosive)
92.	Tincal	:	$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$
93.	Talc	:	$3\text{MgO} \cdot 4\text{SiO}_2 \cdot \text{H}_2\text{O}$ or $\text{Mg}_2(\text{Si}_2\text{O}_3)_2 \cdot \text{Mg(OH)}_2$
94.	Tritium	:	${}_1\text{H}^3$ (an isotope of H)
95.	Water glass	:	$\text{Na}_2\text{SiO}_3$
96.	water gas	:	$\text{CO} + \text{H}_2$
97.	White vitriol	:	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
98.	Wrought iron	:	Pure form of iron
99.	Washing soda	:	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
100.	Willemite \ Zincite	:	ZnO
101.	Zinc white	:	ZnO
102.	Zinc blend	:	ZnS

**INDUSTRIALLY IMPORTANT PROCESS**

Ammonia Soda process (Solvay process)	:	Manufacture of $\text{NaHCO}_3$ , $\text{Na}_2\text{CO}_3$
Birkeland - Eyde process	:	Manufacture of $\text{HNO}_3$
Bosch process	:	Manufacture of $\text{H}_2$
Castner process	:	Manufacture of Na
Caster - Kellner Cell process	:	Manufacture of NaOH
Contact process	:	Manufacture of $\text{H}_2\text{SO}_4$
Down process	:	Manufacture of Na
Dow's process	:	Manufacture of phenol
Deacon's process	:	Manufacture of $\text{Cl}_2$
Haber process	:	Manufacture of $\text{NH}_3$
Hasenclever process	:	Manufacture of Bleaching powder
L.D. process	:	Manufacture of steel
Lead chamber process	:	Manufacture of $\text{H}_2\text{SO}_4$
Nelson cell process	:	Manufacture of NaOH
Ostwald process	:	Manufacture of $\text{HNO}_3$