

## s-BLOCK

The elements in which the last electron enters the outermost s-orbital are called s - block elements. The group 1 and 2 of periodic table belong to the s-block.

### ALKALI METALS

#### **ALKALI METAL AND THEIR COMPOUNDS**

- (i) **The Elements :** are Li , Na, K, Rb, Cs, Fr (Radioactive :  $t_{1/2}$  of Fr<sup>233</sup> = 21 minutes)  
group - I elements are called alkali metals because they form hydroxides on reaction with water, which are alkaline in nature.

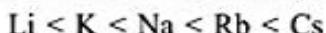
*Outer Electronic configuration : ns<sup>1</sup>*

- (ii) **Atomic Radii.** The atomic radii of alkali metals are largest in their respective periods. The atomic radii increase on moving down the group from top to bottom.



- (iii) **Ionic Radii.** Alkali metals change into positively charged ions by losing their outermost electron. These ions are considerably smaller than the parent atoms. Ionic radii increase on moving down the group.

- (iv) **Density.** Alkali metals have low densities due to their large atomic size. Densities increase on going down the group from top to bottom. Potassium is, however lighter than sodium.



**Exception :** Density of K < Na due to abnormal increment in size of K.

- (v) **Ionization Energy.** The first ionization energies of alkali metals are very low as compared with the other elements of the same period.

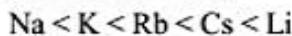


- (vi) **Electropositive Character.** On account of their low ionization energies, these metals have a strong tendency to lose their valence electrons and thus change into positive ions. Consequently, **alkali metals are strongly electropositive or metallic in character.** As this tendency for losing electrons increases down the group, the electropositive character increases.



- (vii) **Oxidation State.** Alkali metals exhibit an oxidation state of +1.

- (viii) **Reducing Character :** All the alkali metals are good reducing agents due to their low ionization energies. Their reducing character, however, follows the order :



The reducing character of any metal is best measured in terms of its electrode potential which among other things depends upon its (i) heat of vaporization (ii) ionization energy and (iii) heat of hydration. Since  $\text{Li}^+$  ion has the smaller size, its heat of hydration has the highest value. Therefore, among the alkali metals Li has the highest negative electrode potential ( $E_{\text{Li}^+/\text{Li}}^0 = -3.05$  volts) and hence is the strongest reducing agent.

Element	Li	Na	K	Rb	Cs	Fr
$E_{\text{M}^+/\text{M}}^0$ (V) at 298 K for	-3.05	-2.71	-2.93	-2.93	-2.92	-
$\text{M}^+(\text{aq}) + \text{e}^- \rightarrow \text{M(s)}$						

- (ix) **Melting and Boiling point.** The generally low values for cohesive energy are reflected in the very low values of melting and boiling points in the group. The cohesive energy decreases down the group, and the melting points decreases correspondingly.

	Melting point (°C)	Boiling point (°C)
Li	181	1347
Na	98	881
K	63	766
Rb	39	688
Cs	28.5	705

	Melting point (°C)	Boiling point (°C)
Li	181	1347
Na	98	881
K	63	766
Rb	39	688
Cs	28.5	705

- (x) **Softness:** These metals are soft and Malleable and have metallic lusture when they are freshly cut due to oscillation of loosely binded electrons.

- (xi) **Ionic Mobility:** Ionic mobility  $\propto \frac{1}{\text{Size of hydrated ion}}$

Size of the hydrated ion is =  $\text{Li}^+(\text{aq}) > \text{Na}^+(\text{aq}) > \text{K}^+(\text{aq}) > \text{Rb}^+(\text{aq}) > \text{Cs}^+(\text{aq})$

**Order of ionic mobility**

$$\text{Li}^+(\text{aq}) < \text{Na}^+(\text{aq}) < \text{K}^+(\text{aq}) < \text{Rb}^+(\text{aq}) < \text{Cs}^+(\text{aq})$$

- (xii) **Crystal Structure :** All the Alkali metals have their body centre cubic (BCC) structure with coordination Number 8.

- (xiii) **Flame Colouration.** Alkali metals impart characteristic colours to the flame when they are heated in a Bunsen burner flame.

Li-Cannine Red, Na-Golden Yellow, K- violet (Lilac), Rb-Reddish Violet; Cs-Blue violet

- (xiv) **Colour** - The compounds of alkali metals are typically white

- (xv) **Magnetic behaviour** - The compounds of alkali metals are diamagnetic. Superoxides of alkali metals are, however, paramagnetic.

- (xvi) **Hydration.** Most of alkali metal salts dissolve in water. In solution alkali metal ions are hydrated. Since  $\text{Li}^+$  ion is smallest in size it is most heavily hydrated. Salts of lithium such as  $\text{LiF}$ ,  $\text{Li}_2\text{CO}_3$ ,  $\text{Li}_3\text{PO}_4$  are insoluble in water.

---

***Illustration***

---

1. 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.

2. 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.

---



---

***Exercise***

---

1. 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)**

2. Mark the false statement ?

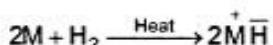
- (A) The electropositive character of alkali metals decreases with increase in atomic number
- (B) Lithium is a hard metal and cannot be cut with a knife
- (C) Alkali metals are strong reducing agents
- (D) Electronegativities of all alkali metals lie between 1.0 to 0.7.

**Ans. (A)**

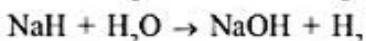
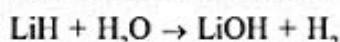
Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231139.jpg

## CHEMICAL PROPERTIES

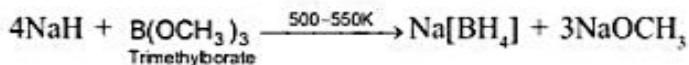
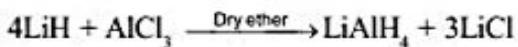
- The alkali metals are highly reactive elements. The reactivity of alkali metals is due to low value of ionization energy; and low heat of atomisation.
- The reactivity of alkali metals increases from Li to Cs
- (i) Reaction with hydrogen. Alkali metals react with dry hydrogen to form hydrides.



- These hydrides are ionic in nature and exist as crystalline solids.
- The hydrides of alkali metals react with water to form corresponding hydroxides and hydrogen gas



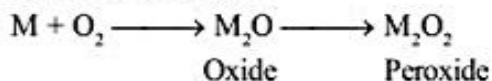
- These hydrides are **strong reducing agents** and their reducing nature increases down the group.
- Alkali metals also form complex hydrides such as  $LiAlH_4$  and  $NaBH_4$ , which are also good reducing agents. Alkali metal hydrides do not exist in water and this reaction with any other reagent is carried out in protic solvent.



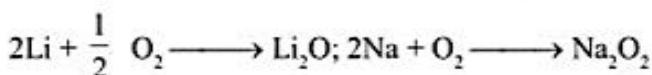
- Fused alkali metal hydrides on electrolysis produces  $H_2$  gas at anode

### (ii) Formation of oxides and hydroxides :

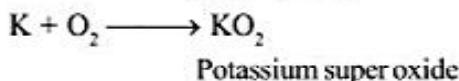
- (a) These are most reactive metals and have strong affinity for  $O_2$  quickly tarnish in air due to the formation of a film of their oxides on the surface. These are, therefore, kept under kerosene or paraffin oil to protect them from air.



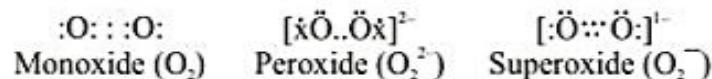
- (b) When burnt in air ( $O_2$ ), lithium forms lithium oxide ( $Li_2O$ ) sodium forms sodium peroxide ( $Na_2O_2$ ) and other alkali metals form super oxide ( $MO_2$  i.e.  $KO_2$ ,  $RbO_2$  or  $CsO_2$ )



Lithium oxide

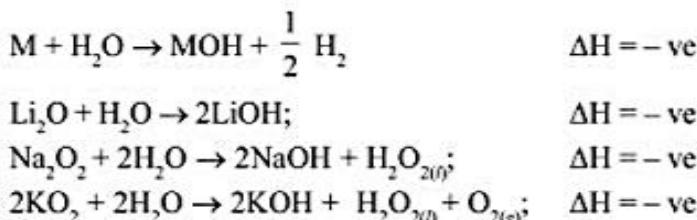


The reactivity of alkali metals towards oxygen to form different oxides is due to strong positive field and thus combines with small anion  $O^{2-}$  to form stable  $Li_2O$  compound. The  $Na^+$  and  $K^+$  being relatively larger thus exert less strong positive field around them and thus reacts with larger oxygen anion i.e.,  $O_2^{2-}$  and  $O_2^{1-}$  to form stable oxides. The monoxide, peroxides and superoxides have  $O_2$  and  $O_2^{2-}$  and  $O_2^{1-}$  ions respectively. The structures of each are,

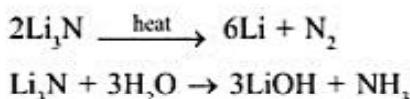


The  $O_2^-$  ion has one three electron covalent bond and has one electron unpaired. It is therefore superoxides are paramagnetic and coloured.  $KO_2$  is light yellow and paramagnetic substance.

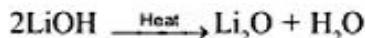
The oxides of alkali metals and metal itself give strongly alkaline solution in water with evolution of heat



- (iii) **Reaction with nitrogen.** Lithium is the only element in the group that reacts with dinitrogen to form a nitride. Lithium nitride,  $Li_3N$ , is ionic ( $3Li^+$  and  $N^{3-}$ ), and is ruby red. Two reactions of the nitride are of interest. First, on heating to a high temperature it decomposes to the elements, and second, it reacts with water, giving ammonia.



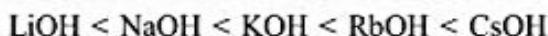
- (iv) **Reaction with water.** The hydroxides of alkali metals are strong bases, all of which are highly soluble in water as well as in alcohol and are stable towards heat. Lithium hydroxide, however, is an exception. It is slightly soluble in water and loses a water molecule on heating.



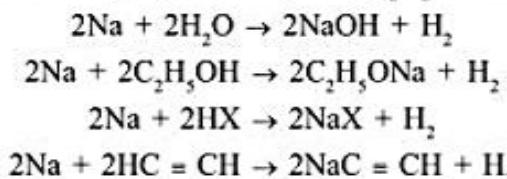
The basic character of alkali metal hydroxides increases in going down the group. This can be explained in terms of decreasing polarising power of alkali metal ions.

$LiOH$  is relatively covalent. As moving down the group ionic nature of hydroxides increases. More the ionic nature of hydroxide; basic nature of hydroxide will increase.

Thus, the basic character of alkali metal hydroxides is in the order:



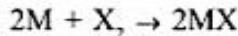
- (v) **Reaction with Water.** Alkali metals react with water and other compounds containing acidic hydrogen atoms such as hydrogen halides ( $HX$ ) and acetylene ( $C_2H_2$ ) and liberate hydrogen gas.



On descending the group from lithium to caesium, the vigour of the reaction increases.

All the alkali metals when exposed to atmosphere react with oxygen and moisture to form oxides and hydroxides and therefore their surface gets tarnished. In order to protect from atmospheric oxygen and water, these metals are stored under kerosene oil.

- (vi) **Reaction with Halogens** - Alkali metals react with halogens to form metal halides, which are ionic crystalline solids having general formula  $M^+X^-$ .



All the halides of alkali metals except lithium fluoride are freely soluble in water. The low solubility of lithium fluoride is attributed to greater force of attraction between lithium ions and fluoride ions in the crystal lattice.

- (vii) **Reaction with Non Metals** - Alkali metals, on heating react with non – metals such as sulphur and phosphorus to form sulphides and phosphides respectively.

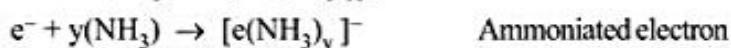
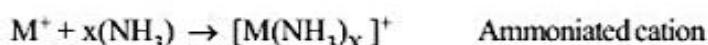


- (viii) **Solubility in Liquid Ammonia.**

Alkali metals dissolve in liquid ammonia (high conc. 5 M) and give blue solution which is conducting, reducing and paramagnetic in nature.

**Reason**

On dissolving Metal in NH<sub>3</sub>,



The blue colour is due to → Ammoniated electron

The paramagnetic nature is due to → Ammoniated electron

The conducting nature is due to → Ammoniated M<sup>+</sup> + Ammoniated electron

\* On standing the colour fades due to formation of amide



In the absence of impurities like. Fe, Pt, Zn etc, the solutions are stable.

\* In concentrated solution, the blue colour changes to bronze colour and diamagnetic due to the formation of metal clusters and ammoniated electrons also associate to form electron pairs



- (ix) **Complex Compounds.** The alkali metal ions form fewer complex compounds than any other group of metal ions. This is due to their large atomic size and weak effective nuclear charge. The complex forming ability decreases in the order :



- (x) **Nature of carbonate and Bicarbonate salts:** Except Li<sub>2</sub>CO<sub>3</sub>, other metal carbonates do not decompose even at high temperature due to their high ionic character.

Order of Solubility in water : Li<sub>2</sub>CO<sub>3</sub> < Na<sub>2</sub>CO<sub>3</sub> < K<sub>2</sub>CO<sub>3</sub> < Rb<sub>2</sub>CO<sub>3</sub> < Cs<sub>2</sub>CO<sub>3</sub>

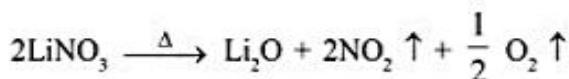
Except LiHCO<sub>3</sub>, all the alkali metal bicarbonate exist.

NaHCO<sub>3</sub> exist in polymeric form and KHCO<sub>3</sub> exist in dimeric form due to the presence of intermolecular hydrogen bonding.

Order of solubility

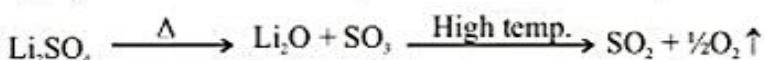


- (xi) **Nature of Nitrate Salt:**

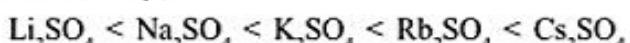


(xii) **Nature of sulphate Salts:**

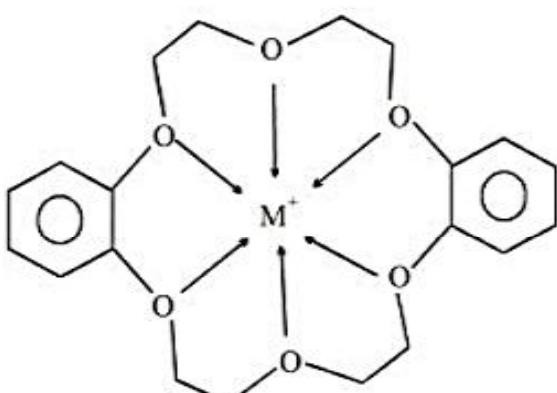
Except  $\text{Li}_2\text{SO}_4$  other alkali metal sulphate salts do not decompose even at high temperature



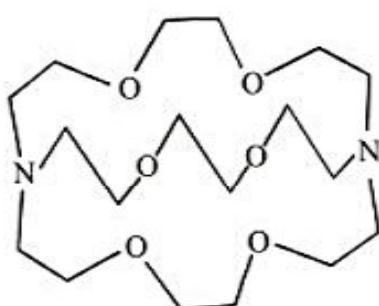
**Order of solubility:**



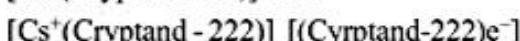
(xiii) **Crown Ethers and Cryptands:**



Dibenzo-18-Grown-6



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.

Lithium resembles magnesium in the following respects :

- (i) atomic size of Li ( $1.34 \text{ \AA}$ ) and Mg ( $1.36 \text{ \AA}$ ) are not much different.
- (ii) ionic size of  $\text{Li}^+$  ( $0.60 \text{ \AA}$ ) and  $\text{Mg}^{2+}$  ( $0.65 \text{ \AA}$ ) 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  $\text{Mg(OH)}_2$  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 respectively nitrides,  $\text{Li}_3\text{N}$  and  $\text{Mg}_3\text{N}_2$

- (xi) carbonates, fluorides, oxalates and phosphates of both Li and Mg are sparingly soluble in H<sub>2</sub>O
- (xii) the hydroxides and carbonates of both Li and Mg decompose on heating forming their respective oxides
- (xiii) both lithium and magnesium nitrates on heating evolve NO<sub>2</sub> and O<sub>2</sub> leaving behind their oxides.

## ORES 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 7<sup>th</sup> and 8<sup>th</sup> position in the order of abundance of elements.**

### (A) Ores of Lithium

- (i) Spodumene, LiAlSi<sub>2</sub>O<sub>6</sub>
- (ii) Lepidolite, (Li)<sub>2</sub>Al<sub>2</sub>(SiO<sub>3</sub>)<sub>3</sub>·(FOH)<sub>2</sub>

### (B) Ores of Sodium

- (i) Common salt or rock salt, NaCl
- (ii) Chile salt patre, NaNO<sub>3</sub>
- (iii) Albite (Soda Feldspar), NaAlSi<sub>3</sub>O<sub>8</sub>
- (iv) Glauber's salt, Na<sub>2</sub>SO<sub>4</sub>·10H<sub>2</sub>O
- (v) Borax, Na<sub>2</sub>B<sub>4</sub>O<sub>7</sub>·10H<sub>2</sub>O

### (C) Ores of Potassium

- (i) Carnallite, KCl·MgCl<sub>2</sub>·6H<sub>2</sub>O
- (ii) Feldspar, KAlSi<sub>3</sub>O<sub>8</sub>
- (iii) Sylvine, KCl

## EXTRACTION OF ALKALI METALS

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

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

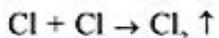
During electrolysis the electrolyte used is 55% LiCl and 45% KCl. The electrolyte is maintained in molten state by heating to about 723 K.

### (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.

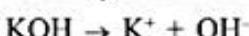
The following reactions take place:



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 :

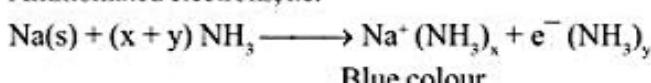


### Illustration

1. Alkali metals readily dissolve in liquid ammonia to give blue coloured solutions. 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.

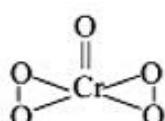


2. Choose the compound which does not possess a peroxide group

(A)  $\text{Na}_2\text{O}_2$     (B)  $\text{CrO}_5$     (C)  $\text{Fe}_2\text{O}_3$     (D)  $\text{BaO}_2$

**Ans.** (C)

**Sol.**  $\text{Na}_2\text{O}_2$  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.

### Exercise

1. 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)

2. Alkali metals 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)

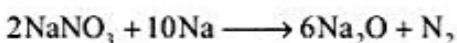
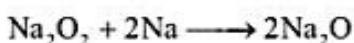
Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231156.jpg

## COMPOUND OF ALKALI METALS

### SODIUM

#### (1) Sodium Monoxide ( $\text{Na}_2\text{O}$ )

**Preparation:** It is obtained by burning sodium at  $180^\circ\text{C}$  in a limited supply of air or oxygen and distilling off the excess of Na in vacuum or by heating  $\text{Na}_2\text{O}_2$  (sodium peroxide), sodium nitrate ( $\text{NaNO}_3$ ) with sodium.

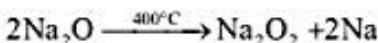


**Properties:**

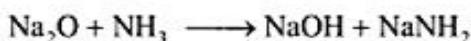
- (i) It is a white amorphous substance.
- (ii) It reacts with water violently forming  $\text{NaOH}$ .



- (iii) On heating above  $400^\circ\text{C}$ , it disproportionate to give peroxide and metallic sodium.



- (iv) It reacts with liquid ammonia forming sodamide and  $\text{NaOH}$ .



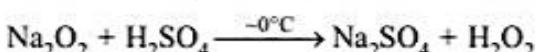
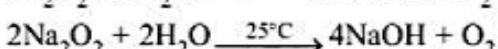
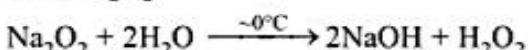
#### (2) Sodium Peroxide ( $\text{Na}_2\text{O}_2$ )

**Preparation:** It is formed by heating the metal in excess of air or oxygen at  $300^\circ$ , which is free from moisture and  $\text{CO}_2$ .

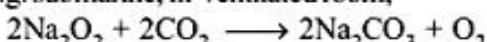


**Properties:**

- (i) It is a pale yellow solid, becoming white in air from the formation of a film of  $\text{NaOH}$  and  $\text{Na}_2\text{CO}_3$ .
- (ii) In cold water ( $\sim 0^\circ\text{C}$ ) produces  $\text{H}_2\text{O}_2$  but at room temperature produces  $\text{O}_2$ . In ice-cold mineral acids also produces  $\text{H}_2\text{O}_2$ .



- (iii) It reacts with  $\text{CO}_2$ , giving sodium carbonate and oxygen and hence its use for purifying air in a confined space e.g. submarine, ill-ventilated room,



- (iv) It is an oxidising agent and oxidises charcoal, CO, NH<sub>3</sub>, SO<sub>2</sub>.
- $$3\text{Na}_2\text{O}_2 + 2\text{C} \longrightarrow 2\text{Na}_2\text{CO}_3 + 2\text{Na}$$
- [deposition of metallic Na]
- $$\text{CO} + \text{Na}_2\text{O}_2 \longrightarrow \text{Na}_2\text{CO}_3$$
- $$\text{SO}_2 + \text{Na}_2\text{O}_2 \longrightarrow \text{Na}_2\text{SO}_4$$
- $$2\text{NH}_3 + 3\text{Na}_2\text{O}_2 \longrightarrow 6\text{NaOH} + \text{N}_2$$
- (v) It contains peroxide ion [-O-O-]<sup>-2</sup>

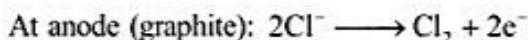
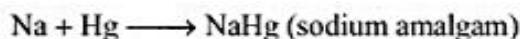
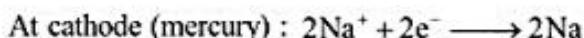
**Uses:**

- (i) For preparing H<sub>2</sub>O<sub>2</sub>, O<sub>2</sub>
- (ii) Oxygenating the air in submarines
- (iii) Oxidising agent in the laboratory.

**(3) Sodium Hydroxide (Caustic Soda), NaOH**

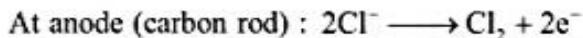
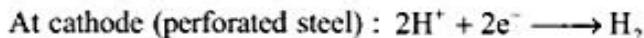
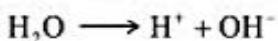
**Preparation:** Sodium hydroxide is commonly called caustic soda therefore used for air purify and it is prepared by electrolytic process. There are two electrolytic methods commonly used for preparing sodium hydroxide.

**(a) Castner Kellner cell:**



Sodium amalgam is removed from the cell. It is then decomposed in a separate cell by water giving NaOH, hydrogen and mercury. Mercury is recirculated to the cell.

**(b) Nelson cell (or Diaphragm cell)**



**Note:** This cell is used for the industrial production of Cl<sub>2</sub>. As well as this cell can be used for production of Na<sub>2</sub>CO<sub>3</sub>, for which CO<sub>2</sub> is mixed with steam.

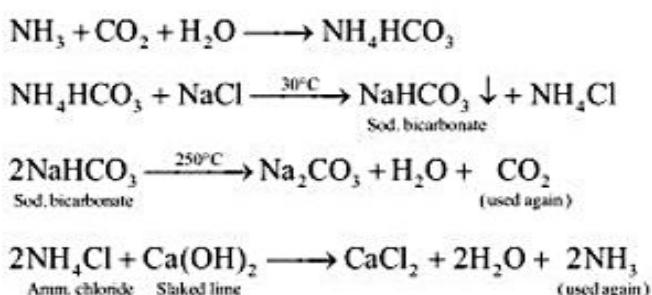
**Properties:**

NaOH is stable towards heat but is reduced to metal when heated with carbon



- (4) Sodium Carbonate (Na<sub>2</sub>CO<sub>3</sub>) :** It exists in various forms, namely anhydrous sodium carbonate Na<sub>2</sub>CO<sub>3</sub> (soda-ash); monohydrate, Na<sub>2</sub>CO<sub>3</sub>.H<sub>2</sub>O (crystal carbonate); heptahydrate, Na<sub>2</sub>CO<sub>3</sub>.7H<sub>2</sub>O and decahydrate, Na<sub>2</sub>CO<sub>3</sub>.10H<sub>2</sub>O (washing soda). It occurs in the form of reh and sajji matti in U.P. and Bihar.

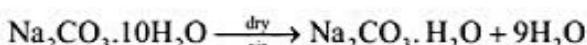
**Manufacture: (Solvay or ammonia soda process).** In this process, brine ( $\text{NaCl}$ ), ammonia and carbon dioxide are the raw materials. The chemical reactions involved are as below.



**Note:** Solvay process can not be used for the production of  $\text{K}_2\text{CO}_3$ , since  $\text{KHCO}_3$  has much greater solubility than  $\text{NaHCO}_3$ . So,  $\text{KHCO}_3$  cannot be precipitated.

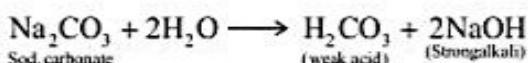
#### Properties:

- (i) Sodium carbonate crystallizes from water as decahydrate. It is a white solid which efflorescence on exposure to dry air forming the monohydrate.

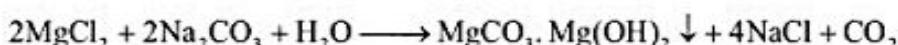


On heating, the monohydrate changes to anhydrous salt (m.p.  $853^\circ\text{C}$ ) which does not decompose on further heating even to redness.

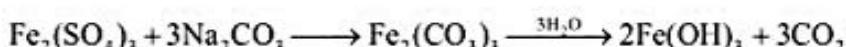
- (ii) It is soluble in water with considerable evolution of heat. The resulting solution is alkaline due to hydrolysis.



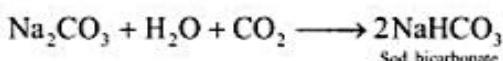
- (iii) It is readily decomposed by acids with the evolution of  $\text{CO}_2$  gas.
- (iv) It reacts with metal salts to form insoluble normal or basic carbonate.



Carbonates of metals like Al, Fe, Sn, etc. When formed are immediately hydrolysed to hydroxides.



- (v) When  $\text{CO}_2$  gas is passed through aqueous solution of sodium carbonate, sodium bicarbonate is formed.



#### (5) Sodium Sulphate ( $\text{Na}_2\text{SO}_4 \cdot 10 \text{ H}_2\text{O}$ ) (Glauber's salt)

**Preparation -** By heating a mixture of common salt & conc.  $\text{H}_2\text{SO}_4$  (Le blanc process )

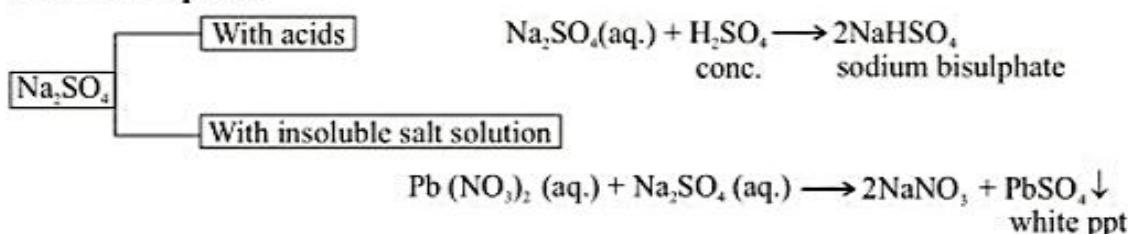


#### Properties

##### Physical properties :

- (i) Crystalline solid, which is soluble in water.
- (ii) Crystallizes as decahydrate & on heating becomes anhydrous.

### Chemical Properties



### (6) 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 is then saturated with a current of dry hydrogen chloride whereby crystals of pure sodium chloride separate out.

#### Properties:

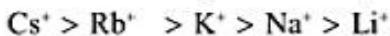
- (i) NaCl is a colourless crystalline salt, almost insoluble in alcohol and highly soluble in water.
- (ii) It gives rise to HCl when heated with conc.  $\text{H}_2\text{SO}_4$  and  $\text{Cl}_2$ , with  $\text{MnO}_2$  plus  $\text{H}_2\text{SO}_4$ .
 
$$\text{NaCl} + \text{H}_2\text{SO}_4 \longrightarrow \text{NaHSO}_4 + \text{HCl} \uparrow$$

$$\text{NaHSO}_4 + \text{NaCl} \longrightarrow \text{Na}_2\text{SO}_4 + \text{HCl} \uparrow$$

$$2\text{NaCl} + \text{MnO}_2 + 2\text{H}_2\text{SO}_4 \longrightarrow \text{MnSO}_4 + \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} + \text{Cl}_2 \uparrow$$

### Illustration

Give reason for decreasing order of conductivity of following



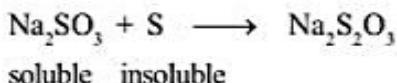
- Sol. Ions are hydrated in solution. Since Li is very small it is heavily hydrated. This make the radius of the hydrated ions large and hence it moves only slowly (although  $\text{Li}^+$  is very small) and the radius of hydrated  $\text{Cs}^+$  ion is smaller than the radius of hydrated  $\text{Li}^+$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

### (7) 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



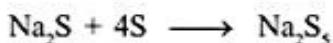
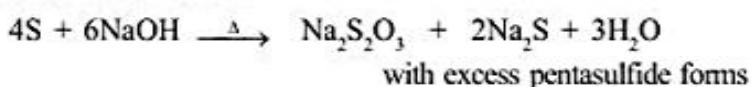
- (ii) Spring's Reaction



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



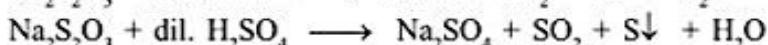
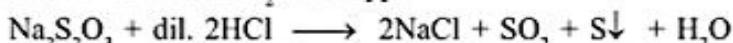
(iv) sulfur and caustic soda



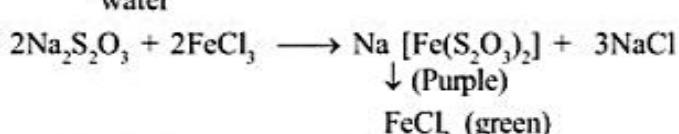
### Chemical Properties

(i) **Heating effect** -  $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} \xrightarrow{215^\circ\text{C}} \text{Na}_2\text{S}_2\text{O}_3 + 5\text{H}_2\text{O}$   
 $\text{4Na}_2\text{S}_2\text{O}_3 \xrightarrow{223^\circ\text{C}} 3\text{Na}_2\text{SO}_4 + \text{Na}_2\text{S}_5$

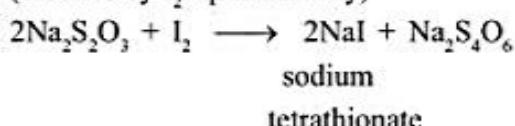
(ii) **Acidification** - liberates  $\text{SO}_2$  and S ppt.



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



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

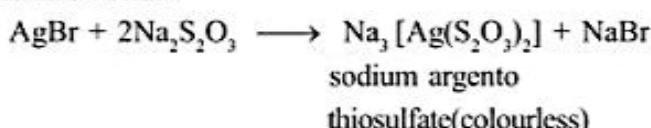


(v) **AgNO<sub>3</sub> action** -  $2\text{AgNO}_3 + \text{Na}_2\text{S}_2\text{O}_3 \longrightarrow \text{Ag}_2\text{S}_2\text{O}_3 + 2\text{NaNO}_3$   
 $\text{Ag}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} \longrightarrow \text{Ag}_2\text{S} + \text{H}_2\text{SO}_4$

Thus white ppt. of silver thiosulfate changes as

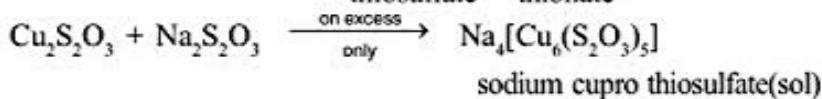
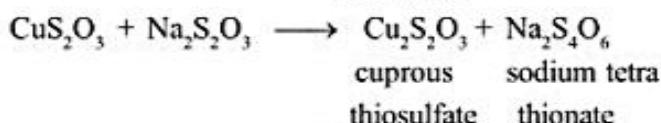
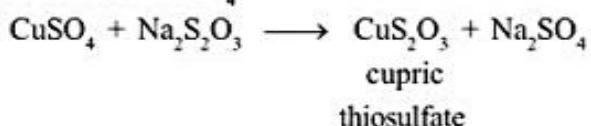
White  $\longrightarrow$  yellow  $\longrightarrow$  brown  $\longrightarrow$  black

(vi) **Silver halide action**



this property is used in photography fixing.

(vii) **Action on CuSO<sub>4</sub>**



### (8) Sodium Bicarbonate ( $\text{NaHCO}_3$ )

**Preparation :** Sodium bicarbonate is prepared by the Solvay Process. It is infact formed in between the manufacture of washing soda.

- \* Aqueous solution gives no colour with phenolphthalein but yellow colour with Methyl orange and hence is weakly basic.
- \*  $2\text{NaHCO}_3 \xrightarrow{\Delta} \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$
- \* Salts which give basic carbonates with washing soda give normal salts with the bicarbonate.  
 $\text{ZnSO}_4 + 2\text{NaHCO}_3 \longrightarrow \text{ZnCO}_3 + \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + 2\text{CO}_2$
- \* It is used as "SODA BICARB" to neutralise stomach acidity.
- \* Sodium/Potassium salt of tartaric acid and  $\text{NaHCO}_3$  mixture is used as Baking Soda.

### **Illustration**

---

1. 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)**

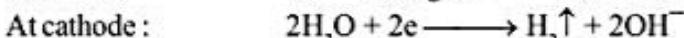
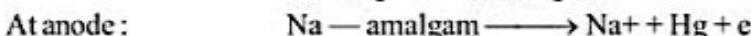
**Sol.** Although lattice energy of LiCl higher than NaCl but LiCl is covalent in nature and NaCl ionic there, after, the melting point decreases as we move NaCl because the lattice energy decreases as a size of alkali metal atom increases (lattice energy  $\propto$  melting point of alkali metal halide)

2. NaOH is prepared by the method

(A) Down's cell            (B) Castner cell            (C) Solvay process            (D) Castner Kellner cell

**Ans. (D)**

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



### **Exercise**

---

1. 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 metallic silver  
 (B) Convert metallic silver to silver salt  
 (C) Remove undecomposed AgBr as a soluble silver thiosulphate complex  
 (D) Remove unreduced silver

**Ans. (C)**

2. Which of the following pair can't exist in solution

(A)  $\text{NaHCO}_3$  and NaOH                    (B)  $\text{Na}_2\text{CO}_3$  and NaOH  
 (C)  $\text{Na}_2\text{CO}_3$  and NaCl                    (D)  $\text{NaHCO}_3$  and NaCl

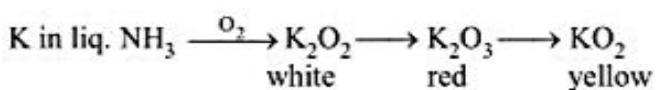
**Ans. (A)**

Saved /storage/emulated/0/Pictures/TouchShot/  
 20170808\_231212.jpg

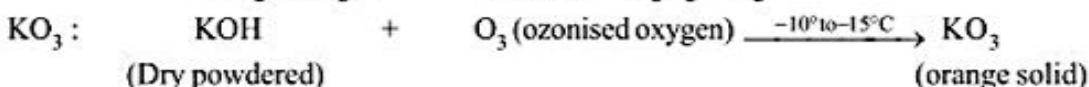
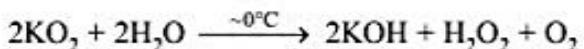
## POTASSIUM

### POTASSIUM AND ITS COMPOUNDS

1. Atomic Number : 19
  2. Electronic Configuration : 2, 8, 8, 1
  3. Valency and Oxidation state : + 1
  4. **Occurrence :** Potassium is highly reactive metal. It does not occur in nature in free form. Its important minerals are :
    - (a) Pearl ash,  $K_2CO_3$
    - (b) Sylvine,  $KCl$ .
    - (c) Indian Salt petre,  $KNO_3$
    - (d) Carnallite,  $KClMgCl_2 \cdot 6H_2O$
  5. **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  $KCl$  containing a little  $KF$ .
  6. **Properties :**
    - (a) Potassium is a silvery white metal (M.P.  $65^\circ$ ).
    - (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 cells.
    - (b) An alloy of sodium and potassium which is liquid is used in high temperature thermometers.
- (i) Oxides of Potassium:**
- |                            |                     |                   |                         |     |                        |
|----------------------------|---------------------|-------------------|-------------------------|-----|------------------------|
| $K_2O$ ,<br>Colours: White | $K_2O_2$ ,<br>White | $K_2O_3$ ,<br>Red | $KO_2$<br>Bright Yellow | and | $KO_3$<br>Orange Solid |
|----------------------------|---------------------|-------------------|-------------------------|-----|------------------------|
- Preparation:**
- (i)  $2KNO_3 + 10K \xrightarrow{\text{heating}} 6K_2O + N_2$
  - \*\*  $K_2O \xrightarrow{\text{heating}} K_2O$   
(White) (Yellow)
  - \*\*  $K_2O + H_2O \longrightarrow 2KOH$
  - (ii)  $2K + O_2 \xrightarrow[\text{air at } 300^\circ C]{\text{Controlled}} K_2O_2$  [Props: Similar with  $Na_2O_2$ ]
  - (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<sub>2</sub> reacts with H<sub>2</sub>O and produces H<sub>2</sub>O<sub>2</sub> and O<sub>2</sub> both



### (ii) Potassium hydroxide (Caustic Potash), KOH

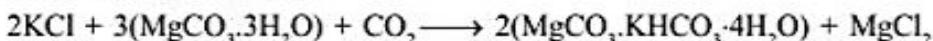
**Preparation:** It may be prepared like that of NaOH, i.e., by the electrolysis of KCl solution and by the action of lime on potassium carbonate. It may also be obtained by the action of baryta, Ba(OH)<sub>2</sub>, on potassium sulphate.



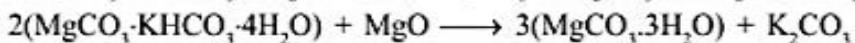
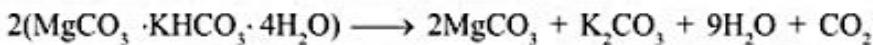
**Properties:** Its properties are similar to those of sodium hydroxide. However, it is a stronger alkali and more soluble in alcohol and is therefore, used in organic reactions instead of caustic soda. Moreover, it is a better absorber of CO<sub>2</sub> than NaOH because potassium carbonate is more soluble and does not separate out. Its aqueous solution is known as potash lye.

### (iii) Potassium carbonate (K<sub>2</sub>CO<sub>3</sub>): It is also known as pearl ash.

**Preparation:** It is made by passing CO<sub>2</sub> into a conc. solution of the chloride, containing hydrated magnesium carbonate in suspension at 20°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°C or by the action of magnesium oxide below 20°C.



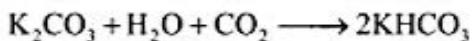
**Properties:**

(i) It is white, deliquescent solid

(ii) K<sub>2</sub>CO<sub>3</sub> resembles Na<sub>2</sub>CO<sub>3</sub> in properties, but is more alkaline and more soluble than Na<sub>2</sub>CO<sub>3</sub>.

### (iv) Potassium Bicarbonate (KHCO<sub>3</sub>)

**Preparation:** It is prepared by passing CO<sub>2</sub> through a cold saturated solution of potassium carbonate.



**Properties:** 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<sub>2</sub>SO<sub>4</sub>)

**Preparation:**

(i) By treating KCl or KOH with H<sub>2</sub>SO<sub>4</sub>



- (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,  $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ .

**(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.

**Illustration**

1. Which is an ore of potassium

- (A) Camallite      (B) Cryolite      (C) Bauxite      (D) Dolomite

**Ans. (A)**

**Sol.** Carnallite -  $KCl \cdot MgCl_2 \cdot 6H_2O$

Cryolite -  $Na_3AlF_6$

Bauxite -  $Al_2O_3 \cdot 2H_2O$

Dolomite -  $MgCO_3 \cdot CaCO_3$

2. KF combination with HF to form  $KHF_2$ . The compound contains the species

- (A)  $K^+$ ,  $F^-$  and  $H^+$       (B)  $K^+$ ,  $F^-$  and HF      (C)  $K^+$  and  $[HF_2]^-$       (D)  $[KHF]^+$  and  $F^-$

**Ans. (C)**



**Exercise**

1. When potassium dichromate crystal are heated with conc. HCl

- (A)  $O_2$  is evolved      (B) Chromyl chloride vapours are evolved  
 (C)  $Cl_2$  is evolved      (D) No reaction takes place

**Ans. (C)**

2. 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)**

Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231220.jpg

## ALKALINE EARTH METALS

### ALKALI EARTH METAL & THEIR COMPOUNDS

- (a) The group IIA consists of six elements - Be, Mg, Ca, Sr, Ba, Ra.  
Collectively called alkali earth metals.
- (b) Ra was discovered in the ore pitch blende by madam curie, it is radioactive in nature.
- (c) The first member Beryllium is less active than other members and shows abnormal properties like lithium in IA group.

### PHYSICAL PROPERTIES

- (i) **Electronic configuration :** *The Elements* are Be, Mg, Ca, Sr, Ba, Ra,  
**Outermost Electronic configuration :** - ns<sup>2</sup>
- (ii) **Atomic Radii.** The atomic radii of alkaline earth metals are quite larger but are smaller than those of alkali metals.  

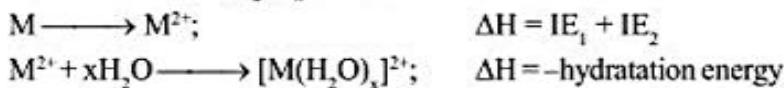
$$\text{Be} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba} < \text{Ra}$$
- (iii) **Density.** The densities of alkaline earth metals are larger than those of alkali metals. This is due to stronger metallic bond. Density of alkaline earth metals first decreases from Be to Ca and then increases from Ca to Ra.  

$$\text{Be} > \text{Mg} > \text{Ca} < \text{Sr} < \text{Ba}$$
- (iv) **Cohesive Forces.** Cohesive forces in alkaline earth metals are stronger than alkali metals due to smaller atomic radii and greater nuclear charge. Alkaline earth metals are harder than alkali metals due to stronger metallic bond.
- (v) **Melting and Boiling points**  
They have low m.p. and b.p. but are higher than corresponding value of group I.

**Reason :** They have two valency electrons which may participate in metallic bonding compared with only one electron in AM. Consequently group II elements are harder and have higher cohesive energy and ∴ have much higher m.p./b.p. than A.M.

$$\begin{array}{ll} \text{M.P.} & \text{Be} > \text{Ca} > \text{Sr} > \text{Ba} > \underline{\text{Mg}} \\ \text{B.P.} & \text{Be} > \text{Mg} > \text{Ca} > \text{Ba} > \text{Sr} \end{array}$$

- (vi) **Ionization energy.** The alkaline earth metals owing to their large size of atoms have fairly low values of **ionization energies** as compared to the p block elements. However, within the group, the ionization energy, decreases as the atomic number increases.
- (vii) **Electropositive character.** These are less electropositive than alkali metals.
- (viii) **Oxidation Number :** The tendency of these metals to exist as divalent cation as IE<sub>1</sub> is much lower than IE<sub>2</sub> so they should be form univalent in rather than divalent ion but actually these give bivalent ions.
- (ix) This is due the fact that M<sup>2+</sup> ion possesses a higher degree of hydration or M<sup>2+</sup> ions are extensively hydrated form [M(H<sub>2</sub>O)<sub>x</sub>]<sup>2+</sup> a hydrated ion.



- (x) **Reducing properties.** The members of this group are weaker reducing agents than the alkali metals.  
 $\text{Be} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$
- (xi) **Flame colouration.** Among the members of this group except Be and Mg all other members give characteristic colouration to the flame. Beryllium and magnesium do not impart any colour to the flame because of their small size and higher ionization energies. Due to this, the energy of flame is insufficient to cause the excitation of their valence electrons to higher shells.  
 Ca – Brick Red, Sr – Crimson Red, Ba – Apple green
- (xii) **Solubility of Liquid Ammonia.** All these metals dissolve in liquid  $\text{NH}_3$ .

### *Illustration*

1. 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.

2. The ion having highest mobility in aqueous solution is

(A)  $\text{Be}^{2+}$                     (B)  $\text{Mg}^{2+}$                     (C)  $\text{Ca}^{2+}$                     (D)  $\text{Ba}^{2+}$

**Ans.** (D)

**Sol.** Though  $\text{Ba}^{2+}$  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.

### *Exercise*

1. Alkaline earth metal show typical characteristics of  
 (A) Inner transition elements                    (B) Noble gases  
 (C) Representative elements                    (D) Transition elements
- Ans.** (C)
2. 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)

Saved /storage/emulated/0/Pictures/TouchShot/  
 20170808\_231225.jpg

## CHEMICAL PROPERTIES

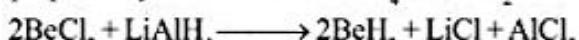
- Alkaline earth metals are quite reactive elements due to their low ionization energies and high electropositive character. The reactivity of these elements increases with increase in atomic number.
- Alkaline earth metals are less reactive than alkali metals.

(i) **Hydrides :**

- (a) Except Be, all alkaline earth metals form hydrides ( $MH_2$ ) on heating directly with  $H_2$ .

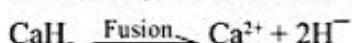


- (b)  $BeH_2$  is prepared by the action of  $LiAlH_4$  on  $BeCl_2$



- (c)  $BeH_2$  and  $MgH_2$  are covalent while other hydrides are ionic.

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



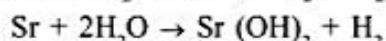
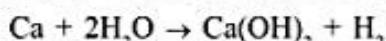
- (e) The stability of hydrides decreases from Be to Ba.

- (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.



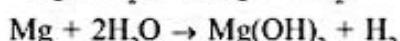
- (ii) **Reaction with water.** Ca, Sr and Ba react with cold water, liberating hydrogen gas



Magnesium decomposes hot water



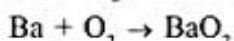
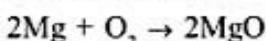
Or



Beryllium does not react with water.

Ordinary of Reactivity :  $Ba > Sr > Ca > Mg$

- (iii) **Reaction with Oxygen.** All the alkaline earth metals burn in oxygen to form oxides Be, Mg and Ca form oxides whereas Ba and Sr form peroxides.



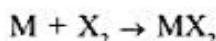
- (iv) **Hydroxides.** The hydroxides of these elements can be formed either by dissolving metal oxides in water or by reaction of these elements with water.

$Be(OH)_2$  is amphoteric. The hydroxides of magnesium, calcium, strontium and barium are bases and their strength increases from magnesium to barium.

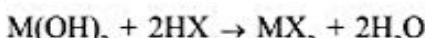
These hydroxides are less soluble in water as compared to the alkali metal hydroxides.

The solubility of the hydroxides in water increases with the increase in atomic number.  $Be(OH)_2$  and  $Mg(OH)_2$  are almost insoluble,  $Ca(OH)_2$  is sparingly soluble while  $Sr(OH)_2$  and  $Ba(OH)_2$  are increasingly more soluble.

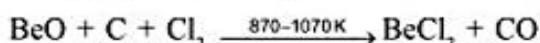
- (v) **Halides.** Alkaline earth metals react with halogens directly to form halides having general formula,  $MX_2$ .



These halides can also be obtained by the action of halogen acids on metals, their oxides, carbonates and hydroxides.

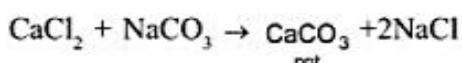


Beryllium chloride is prepared indirectly, from its oxide as follows:



Beryllium chloride exists as monomer or dimer in vapour state but exists as polymer in solid state.

- (vii) **Carbonates.** Carbonates of alkaline earth metals are insoluble in water. These can be precipitated by addition of sodium or ammonium carbonate solution to the solution of salts of these metals. For example,

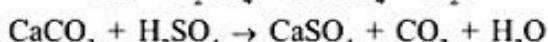
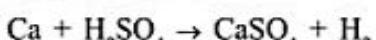


All the carbonates decompose on heating to give carbon dioxide and meals oxide.



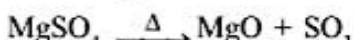
However, as the atomic number increases, the stability of the carbonates towards heat increases. Beryllium carbonate is unstable and can be kept only in an atmosphere of  $\text{CO}_2$ .

- (viii) **Sulphates.** Sulphates of alkaline earth metals can be prepared by the reaction of  $\text{H}_2\text{SO}_4$  with the metals, their oxides, hydroxides or carbonates.



The sulphates of alkaline earth metals are less soluble than the corresponding salts of alkali metals. Their solubilities decrease on going down the group.

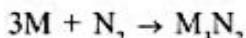
All these sulphates decompose on heating



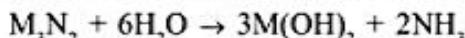
The thermal stability of sulphates increases on moving down the group from top to bottom.

**Lithopone:** Mixture of  $\text{BaSO}_4$  and  $\text{ZnS}$  is known as lithopone.

- (ix) **Nitrides.** Alkaline earth meals burn in nitrogen to form nitrides,  $\text{M}_3\text{N}_2$  which are ionic in nature.



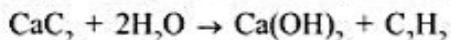
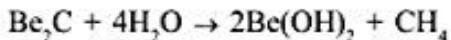
Nitrides on reactions with water are hydrolyzed and ammonia is released.



- (x) **Carbides.**  $\text{BeO}$  when heated with  $\text{C}$  at about  $2000^\circ\text{C}$ ,  $\text{Be}_2\text{C}$  is formed. The metals Mg, Ca, Sr and Ba form carbides of the formula  $\text{MC}_2$ . 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.

$\text{Be}_2\text{C}$  yields methane on hydrolysis whereas carbides of other metals yield acetylene



- (xi) **Complex Compounds.** Alkaline earth metals have a tendency to form state 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.

## 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 – 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  $\alpha$ -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.



At cathode :  $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$

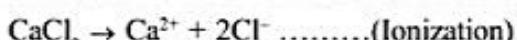
At anode :  $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

### (3) Calcium

#### Ores of Calcium

- |   |                                    |  |
|---|------------------------------------|--|
| (i) Gypsum, CaSO <sub>4</sub> , 2H <sub>2</sub> O | (ii) Lime stone, CaCO <sub>3</sub> | (iii) Marble, CaCO <sub>3</sub>                      |
| (iv) Fluorspar, CaF <sub>2</sub>                  | (v) Anhydrite, CaSO <sub>4</sub>   | (vi) Dolomite, CaCO <sub>3</sub> , MgCO <sub>3</sub> |

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.



At cathode :  $\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$

At anode :  $\text{Cl}^- \rightarrow \text{Cl} + \text{e}^-$



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

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

- Unlike group -2 elements but like aluminium, beryllium forms covalent compounds.
- The hydroxides of beryllium,  $[Be(OH)_2]$  and aluminium,  $[Al(OH)_3]$  are amphoteric in nature, whereas those of other elements of group -2 are basic in nature.
- Beryllium and aluminium forms fluoro complexes in solution.
- The oxides of both Be and Al, i.e., BeO and  $Al_2O_3$  are high melting insoluble solids.
- $BeCl_2$  like  $Al_2Cl_6$  has a bridged polymeric structure.
- The salts of beryllium as well as aluminium are extensively hydrolyzed.
- Be, like Al, is rendered passive by concentrated  $HNO_3$ .
- Be forms many complex like Al
- Be and Al react with NaOH solution liberating  $H_2$ .
- $Be_2C$  and  $Al_4C_3$  react with water to give methane.
- Halides of the two elements have similar solubilities.
- The two metals do not impart colour to the Bunsen flame.

### Illustration

1. 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 MO ?

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

**Ans.** (A)

**Sol.** Be forms BeO, Al forms  $Al_2O_3$ , Cl forms  $Cl_2O_7$ ,  $Cl_2O$ ,  $ClO_2$  etc. while Ne does not react.

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

(A)  $SrCO_3$                     (B)  $BaCO_3$                     (C)  $CaCO_3$                     (D)  $MgCO_3$

**Ans.** (B)

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



### Exercise

1. 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)

2. 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)

Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231236.jpg

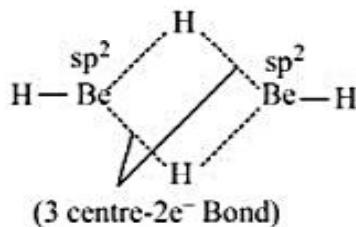
## COMPOUND OF ALKALINE EARTH METALS

### (1) Beryllium

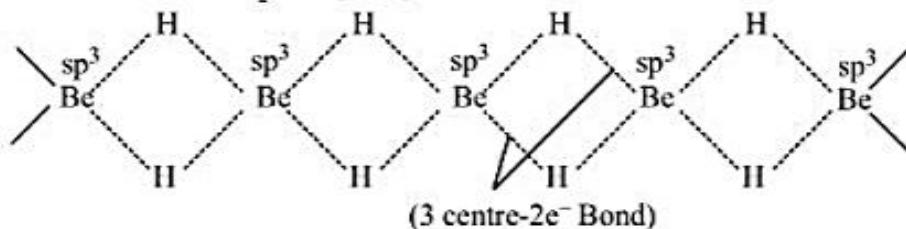
(i) **Oxides and hydroxides :** Be forms only the monoxides: all others form monoxide and peroxide (eg. BaO, BaO<sub>2</sub>). The basic nature of the oxides and hydroxides of the metals increases down the group. BeO is amphoteric, CaO, SrO and BaO are basic. All these oxides dissolve in H<sub>2</sub>O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

(BeO) + 2HCl → BeCl<sub>2</sub> + H<sub>2</sub>O; BeO + 2NaOH → Na<sub>2</sub>BeO<sub>2</sub> + H<sub>2</sub>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 a polymeric solid, involving hydrogen bridging.

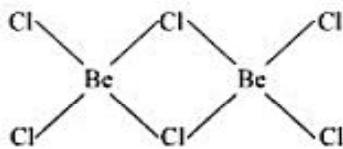


Dimer form of BeH<sub>2</sub> in vapour phase



Polymeric Form of BeH<sub>2</sub>, in solid state

(iii) **Halides :** II group metals 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 solids. 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.

(iv) **Sulphates :** BeSO<sub>4</sub> is slightly soluble in water. The solubility of the sulphates of the metals decreases down the group. BaSO<sub>4</sub> and RaSO<sub>4</sub> are practically insoluble.

[Note : The solubility of an ionic solid in water depends on the lattice energy of the solid and the hydration energy for the ions of the solid. A large value of the former would lower the solubility, while a large value of the latter would increase the solubility. In the case of II group sulphates, the hydration energy factor has a dominant role. Since the hydration of ion decreases from,  $\text{Be}^{2+}$ , to  $\text{Ba}^{2+}$ , the solubility of the II group sulphates decreases down the group. For the fluorides and 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 acetylidy or allylidy.

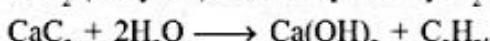
**Beryllium Carbide :**

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



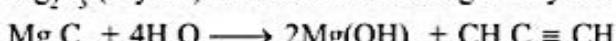
**Calcium Carbide :**

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



**Magnesium Carbide :**

$\text{Mg}_2\text{C}_3$  (allylidy) 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.

### Illustration

1. A metal M readily forms water soluble sulphate  $\text{MSO}_4$ , water insoluble hydroxide  $\text{M}(\text{OH})_2$  and oxide  $\text{MO}$  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.

### Exercise

1. Amongst the following hydroxides, the one which has the lowest value of  $K_{\text{sp}}$  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)

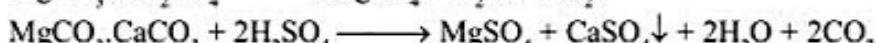
Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231241.jpg

(2) **Magnesium**

- (i). Magnesium Sulphate ( $MgSO_4 \cdot 7H_2O$ ) epsum salt

**Preparation :**

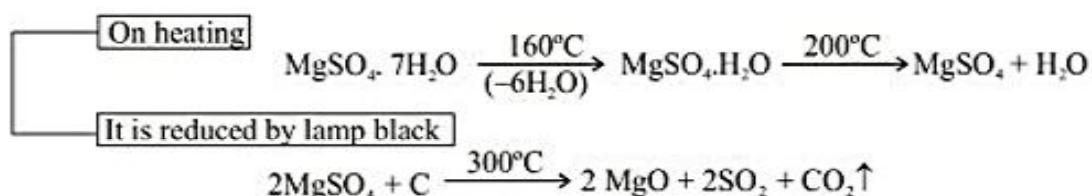
From magnesite or dolomite (Epsom salt)



**Physical properties :**

Colourless, efflorescent crystalline substance which forms a number of hydrate

**Chemical properties :**



**USES :**

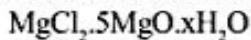
1.  $MgSO_4$  is manufactured by dissolving keiserite ( $MgSO_4 \cdot H_2O$ ) in boiling water & then crystallising the solution.
2.  $MgSO_4 \cdot 7H_2O$  is called as Epsom salt & shows deliquescecence efflorescence & has purgative action.
3. Platinized  $MgSO_4$  is used as catalyst in Grillo's process for manufacture of  $H_2SO_4$ .

(ii) **Magnesium Chloride  $MgCl_2 \cdot 6H_2O$**

Magnesium chloride can be prepared by fractional crystallization of carnallite, KCl.  $MgCl_2 \cdot 6H_2O$ . It can also be obtained from sea water. Magnesium salts present in sea-water are precipitated as  $Mg(OH)_2$  by adding lime. The precipitate is filtered and dissolved in HCl solution. The solution on crystallization yields crystals of  $MgCl_2 \cdot 6H_2O$ .

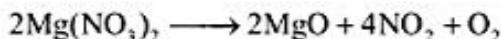


A mixture of magnesium chloride and magnesium oxide is known as **Sorel cement** or **magnesium cement**. This mixture reacts with water and sets into a solid mass having composition



(iii) **Magnesium Oxide ( $MgO$ )**

**Preparation:** It is prepared by burning Mg in air, by heating hydroxide, nitrate, sulphate or carbonate. Oxide prepared by heating magnesite ( $MgCO_3$ ) is called **calcined magnesia**.



**Properties:**

- (i) Magnesium oxide is a light white powder, which is highly infusible (m.p. 2800°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.



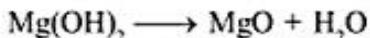
- (iv) It dissolves in aqueous solution of magnesium chloride or bromide forming basic salts such as  $\text{MgCl}_2 \cdot 9\text{Mg(OH)}_2 \cdot 5\text{H}_2\text{O}$ .

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

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

**Properties:**

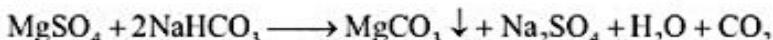
- (i) It is converted into its oxide on heating.



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

**(v) Magnesium Carbonate, ( $\text{MgCO}_3$ ):** It occurs in nature as magnesite,  $\text{MgCO}_3$ , and dolomite,  $\text{MgCO}_3 \cdot \text{CaCO}_3$ .**Preparation:**

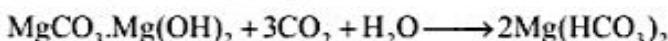
- (i) By adding sodium bicarbonate solution to a hot solution of magnesium salt.



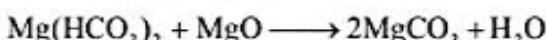
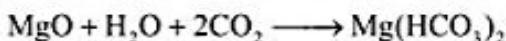
In case  $\text{Na}_2\text{CO}_3$  is used in place of  $\text{NaHCO}_3$ , a basic carbonate is obtained.



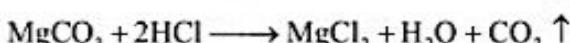
When  $\text{CO}_2$  gas is passed through the suspension of basic carbonate, magnesium bicarbonate is formed which on heating forms  $\text{MgCO}_3$ .



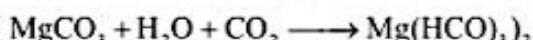
- (ii) Pure magnesium bicarbonate and hence carbonate can be prepared by passing  $\text{CO}_2$  through suspension of magnesium oxide in water.

**Properties:**

- (i) Magnesium carbonate is a white powder, insoluble in water.
- (ii) It dissolves in acids forming salts.

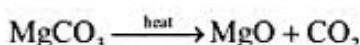


- (iii) Its aqueous solution also dissolves on passing a current of carbon dioxide because of the formation of  $Mg(HCO_3)_2$ , a temporary hardness causing substance.



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 Soral cement. The composition is



\*  $Mg(ClO_4)_2$  is known as anhydronite

---

#### ***Illustration***

1. Magnesium burns in  $CO_2$  to form  
 (A)  $MgO + C$       (B)  $MgO + CO$       (C)  $MgCO_3$       (D)  $MgO$

**Ans. (A)**

**Sol.** Mg burns in  $CO_2$  to give  $MgO$  and C.

2. Which of the following is incorrect ?

- (A) Mg burns in air releasing dazzling light rich in UV rays  
 (B)  $CaCl_2 \cdot 6H_2O$  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.

---

#### ***Exercise***

1. Flame test is not given by  
 (A) Be      (B) Mg      (C) Ca      (D) Sr

**Ans. (A, B)**

2. Pure anhydrous  $MgCl_2$  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)**

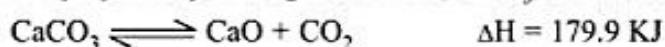
**Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231249.jpg**

### (3) Calcium

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

##### Preparation

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

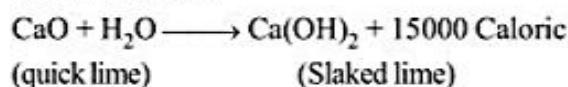


##### Condition for Good Yield :

- (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) Temperature should not be allowed to rise above  $1270\text{K}$  otherwise silica ( $\text{SiO}_2$ ) present as impurity in limestone will react with CaO to form  $\text{CaSiO}_3$ .

##### Properties

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

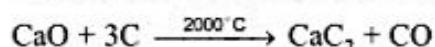


Paste of lime in water is called milk of lime, while its hydrate is known as lime water.

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



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



##### Uses

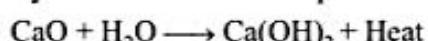
###### (i) In purification of sugar

###### (ii) As basic lining in furnaces.

#### (ii) Calcium hydroxide, Slaked lime $\text{Ca}(\text{OH})_2$ ,

##### Preparation

By the action of water on quick lime.

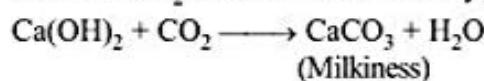


means it is an exothermic reaction.

##### Properties

###### (i) Sparingly soluble in water and its solubility decreases with increases in temperature.

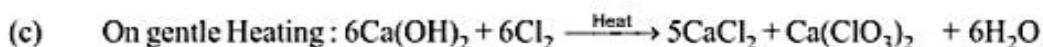
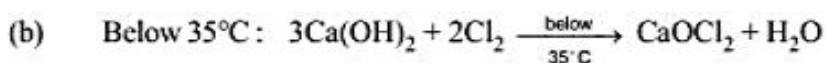
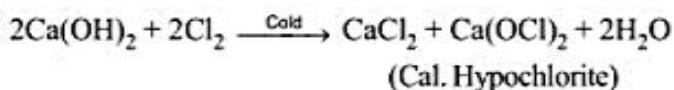
###### (ii) Action of $\text{CO}_2$ : Lime water turns milky on passing $\text{CO}_2$ gas



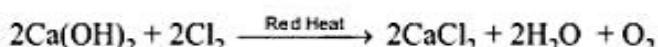
Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231251.jpg

(iii) Action of chlorine :

(a) Cold Condition :



(d) On Red Hot :



(iv) Action of Ammonia :



### Uses

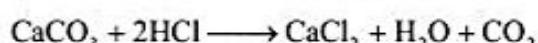
(i) For softening of hard water.

(ii) For purification of sugar and coal gas.

(iii) In preparation of bleaching powder, white wash, plaster etc.

(iii) **Calcium Carbonate, ( $\text{CaCO}_3$ )**: 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.



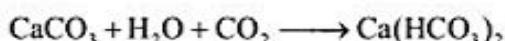
In the laboratory, it is prepared by passing  $\text{CO}_2$  through lime-water or by adding sod. or ammo. carbonate solution to  $\text{CaCl}_2$ .



### 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  $\text{CO}_2$  due to the formation of calcium bicarbonate.



(iv) **Calcium Bicarbonate, ( $\text{Ca}(\text{HCO}_3)_2$ )**: It does not exist in solid state. However, its solution can be prepared by passing  $\text{CO}_2$  gas through a suspension of calcium carbonate in water. On warming, it decomposes to calcium carbonate along with the evolution of carbon dioxide gas.

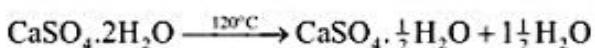
(v) **Calcium Sulphate, ( $\text{CaSO}_4$ )**: It occurs in nature as anhydride ( $\text{CaSO}_4$ ) and gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ).

**Preparation:** It may be prepared by adding dilute sulphuric acid or soluble sulphate (e.g.  $\text{Na}_2\text{SO}_4$ ) to the solution of a calcium salt.

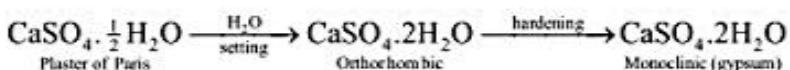


**Properties:**

- (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 changes 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,  $(2\text{CaSO}_4) \cdot \text{H}_2\text{O}$  or  $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$  which is commonly known as Plaster of Paris.



Plaster of Paris is a white powder which when mixed with water takes up the water of crystallization again, thus converted back into the dehydrate 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  $\text{HCl}$  is evaporated.

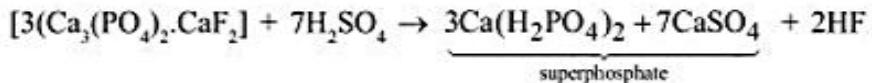


But it separates out from the reaction mixture as  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ . The anhydrous salt is obtained on heating above  $200^\circ\text{C}$ .

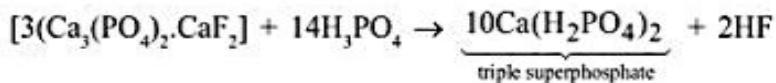
**Properties:** It is a colourless, deliquescent salt, highly soluble in water. The anhydrous salt is an excellent drying agent.

**(vii) Superphosphate**

Phosphate rocks such as fluoroapatite  $[3\text{Ca}_3(\text{PO}_4)_2 \cdot \text{CaF}_2]$  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 a similar way, using  $\text{H}_3\text{PO}_4$  to avoid the formation of the waste product  $\text{CaSO}_4$ .

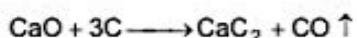


**(viii) Hydrolith**

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

**(ix) Calcium Carbide,  $\text{CaC}_2$** 

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



$\text{CaC}_2$  on reaction with water produces acetylene gas

***Illustration***

1. 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.**  $\text{CaCl}_2$  because it is hygroscopic.

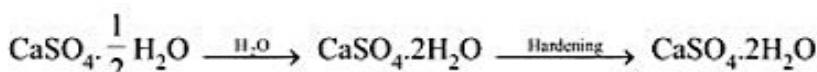
2. Setting of plaster of paris is

(A) Oxidation with atmospheric  $\text{CO}_2$  (B) Combination with atmospheric  $\text{CO}_2$

(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.

***Exercise***

1. Which of the following statements is false

(A)  $\text{CaOCl}_2$  gives  $\text{OH}^-$ ,  $\text{Cl}^-$  and  $\text{OCl}^-$  in aqueous solution

(B) Diamond and graphite are allotropes

(C) Bleaching action of  $\text{Cl}_2$  in moist condition is not permanent

(D) Calomel is  $\text{Hg}_2\text{Cl}_2$

**Ans. (C)**

Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231259.jpg

## SOLVED EXAMPLES

**Q.1** NaOH gives disproportionation reaction with

- (A) S                      (B) CO<sub>2</sub>                      (C) SO<sub>2</sub>                      (D) SO<sub>3</sub>

**Ans.** (A)



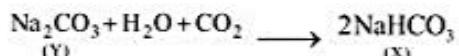
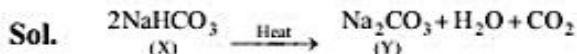
with excess pentasulfide forms



**Q.2** CO<sub>2</sub> gas along with solid (Y) is obtained when sodium salt (X) is heated. (X) is again obtained when CO<sub>2</sub> gas is passed into (Y). X & Y are -

- (A) Na<sub>2</sub>CO<sub>3</sub>, Na<sub>2</sub>O    (B) Na<sub>2</sub>CO<sub>3</sub>, NaOH    (C) NaHCO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>    (D) Na<sub>2</sub>CO<sub>3</sub>, NaHCO<sub>3</sub>

**Ans.** (C)



**Q.3** Commercial common salt becomes slightly damp on keeping. This is due to the fact that -

- (A) Common salt is hygroscopic  
 (B) Common salt contains some impurity which is hygroscopic  
 (C) Salt is efflorescent  
 (D) Salt is crystalline

**Ans.** (B)

**Sol.** Commercial common salt commonly becomes slightly damp on keeping because common salt contains some impurity MgCl<sub>2</sub> and CaCl<sub>2</sub> which is hygroscopic in nature and absorbs moisture from the atmosphere.

**Q.4** The highest melting point among alkali metal of -

- (A) Li                      (B) Na                      (C) K                      (D) Rb

**Ans.** (A)

**Sol.** Li has highest melting point among alkali metals. All alkali metals have low M.P. The M.P. decrease down the group. The low M.P. are attributed to their larger atomic size due to which the binding energies of their atoms in the crystal lattice are low.

**Q.5** Li has the maximum value of ionisation potential among alkali metals i.e. lithium has the minimum tendency to ionise to give Li<sup>+</sup> ion. lithium is -

- |                               |                             |
|-------------------------------|-----------------------------|
| (A) Strongest reducing        | (B) Poorest reducing agent  |
| (C) Strongest oxidising agent | (D) Poorest oxidising agent |

**Ans.** (A)

**Saved /storage/emulated/0/Pictures/TouchShot/  
20170808\_231301.jpg**

**Sol.** The ionisation potential value of Lithium is maximum among alkali metals i.e., its tendency to ionise to give  $\text{Li}^+$  ions should be the minimum i.e. Li should be the poorest reducing agent. But, lithium is the strongest reducing agent. This is due to the largest value of hydration energy of  $\text{Li}^+$  ions.

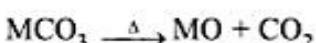
**Q. 6** Which of the following compound decomposes at highest temperature -

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

**Ans.** (B)

**Sol.**  $\text{BaCO}_3$  decomposes at highest temp.

All the carbonates decompose on heating to give  $\text{CO}_2$  and metal oxide.

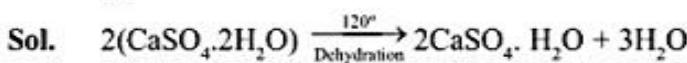


The stability of carbonate towards heat depends upon the stability of the resulting metal oxide. More is the stability of the resulting metal oxide lesser is the stability of the carbonate towards heat and vice versa.

**Q.7** Gypsum  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$  on heating to about  $120^\circ\text{C}$  forms a compound which has the chemical composition represented by

- (A)  $\text{CaSO}_4$       (B)  $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$       (C)  $\text{CaSO}_4 \cdot \text{H}_2\text{O}$       (D)  $2\text{CaSO}_4 \cdot 3\text{H}_2\text{O}$

**Ans.** (B)



Gypsum                          Plaster of paris

**Q.8** Portland cement is manufactured by using -

- |                                    |                                 |
|------------------------------------|---------------------------------|
| (A) Lime stone, clay and sand      | (B) Lime stone, gypsum and sand |
| (C) Lime stone, gypsum and alumina | (D) Lime stone, clay and gypsum |

**Ans.** (D)

**Sol.** Lime stone -  $\text{CaCO}_3$ ,

Clay - silica and alumina

Gypsum -  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

**Q.9** The first ionization potential of Mg is

- (A) Less than Al      (B) More than Al      (C) Equal to Al      (D) Zero

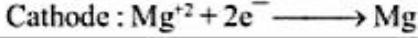
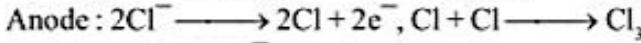
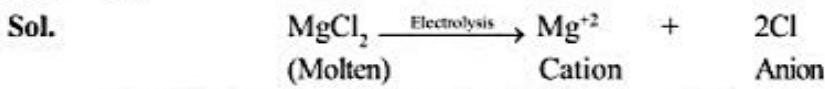
**Ans.** (B)

**Sol.** The first ionization potential of Mg is more than Al since the electron has to be removed from completely filled svalence shell of Mg.

**Q.10** Metallic magnesium is prepared by

- (A) Reduction of  $\text{MgO}$  by coke  
 (B) Electrolysis of aqueous solution of  $\text{Mg}(\text{NO}_3)_2$   
 (C) Displacement of Mg by iron from  $\text{MgSO}_4$  solution  
 (D) Electrolysis of molten  $\text{MgCl}_2$

**Ans.** (D)



**Q.11** Which liberates ammonia when treated with

- (A)  $\text{Li}_3\text{N}$       (B)  $\text{Mg}_3\text{N}_2$       (C)  $\text{CaCN}_2$       (D) All

**Ans.** (D)

**Sol.** All nitrides react with  $\text{H}_2\text{O}$  to yield  $\text{NH}_3$ . Calcium cyanamide ( $\text{CaCN}_2$ ) on hydrolysis also gives  $\text{NH}_3$ .

$$\text{CaCN}_2 + 3\text{H}_2\text{O} \longrightarrow \text{CaCO}_3 + \text{NH}_3$$

**Q.12** 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.



**Q.13** If  $\text{Na}^+$  ion is larger than  $\text{Mg}^{2+}$  ion and  $\text{S}^{2-}$  ion is larger than  $\text{Cl}^-$  ion, which of the following will be least soluble in water?

- (A) Sodium chloride    (B) Sodium sulphide    (C) Magnesium chloride    (D) Magnesium sulphide

**Ans.** (D)

**Sol.** Magnesium sulphide. Higher the lattice energy lower the solubility. Out of the four combinations possible, the lattice energy of  $\text{MgS}$  (Bi-bivalent ionic solid) is higher than those of  $\text{Na}_2\text{S}$ ,  $\text{MgCl}_2$  (uni-bivalent or biunivalent ionic solids) and  $\text{NaCl}$  (uniunivalent ionic solids) and hence  $\text{MgS}$  is the least soluble.

**Q.14** Which of the following is used as barium meal for getting the X-ray spectrum of the human digestive system?

- (A)  $\text{BaSO}_4$       (B)  $\text{BaCl}_2$       (C)  $\text{BaF}_2$       (D)  $\text{BaCO}_3$

**Ans.** (A)

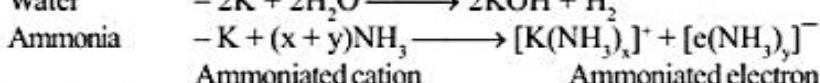
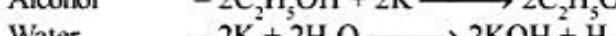
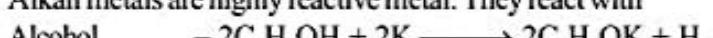
**Sol.**  $\text{BaSO}_4$  is both insoluble in water and opaque to X-rays and hence is used to get the X-ray spectrum of the digestive system?

**Q.15** Potassium is kept in

- (A) Alcohol      (B) Water      (C) Kerosene      (D) Liquid ammonia

**Ans.** (C)

**Sol.** Alkali metals are highly reactive metals. They react with



But they do not react with Kerosene.