

## s-BLOCK ELEMENTS

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### 1. ALKALI METALS (GROUP 1)

The group 1 elements have  $ns^1$  electronic configuration and are highly reactive metals

Elements	Atomic Number	Electronic Configuration
Lithium (Li)	3	[He] 2s <sup>1</sup>
Sodium (Na)	11	[Ne] 3s <sup>1</sup>
Potassium (K)	19	[Ar] 4s <sup>1</sup>
Rubidium (Rb)	37	[Kr] 5s <sup>1</sup>
Cesium (Cs)	55	[Xe] 6s <sup>1</sup>
Francium (Fr)	87	[Rn] 7s <sup>1</sup>

#### 1.1 Physical Properties

##### 1.1.1 Atomic Size

The atoms are largest in their corresponding periods. Atomic size increases in going down the group.

##### 1.1.2 Oxidation State

The group 1 elements exhibit +1 oxidation state

##### 1.1.3 Density

Alkali metals have large size which accounts for their low density

$$\text{Density} = \frac{\text{Atomic mass}}{\text{Atomic volume}}$$

**Atomic weight** increases from Li to Cs in the group and volume also increases but increase in atomic weight is more than increase in volume. Therefore density increases from Li to Cs.

**Exception :** Density of Na is more than that of K

order : Li < K < Na < Rb < Cs

##### 1.1.4 Nature of Bonds

The electronegativity values being low, they combine with other elements to form Ionic bond.

#### 1.1.5 Ionization Energy

The first ionisation energies for the atoms in this group are lower than those for any other group in the periodic table. The atoms are very large so the outer electron are held weakly by the nucleus hence the ionisation energy is not large. Ionization energy decreases on moving down the group.

#### 1.1.6 Flame Test

Alkali metals have large size, when heated on the flame the electrons present in the valence shell move from lower energy level to higher energy level by absorption of heat from the flame. When they come back to the ground state, they emit the extra energy in the form of visible light to provide colour to the flame

Element	Colour
Li	Red
Na	Golden yellow
K	Violet
Rb	Red Violet
Cs	Blue

#### 1.1.7 Standard Oxidation Potential

The measure of the tendency of donating electrons of a metal in water is called its electrode potential. If the concentration of metal ions is unity, then it is called as standard electrode potential. Lithium has the highest ionization potential but has the highest electrode potential due to highest hydration energy.

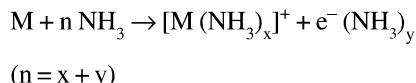
#### 1.1.8 Hydration of Ions

The ions are heavily hydrated. The smaller the size of the ion, the greater is the degree of hydration. Thus the degree of hydration decreases down the group from  $\text{Li}^+$  to  $\text{Cs}^+$ . Thus with the increase in hydration electrical conductivity decreases.

#### 1.1.9 Lattice Energy

Salts of alkali metals are ionic solids. Lattice energy of salts of alkali metals having common anion decreases on descending down the group.

### 1.1.10 Solubility in Liquid Ammonia



Dilute solutions of alkali metals in liquid ammonia are dark blue in colour and the main species present are solvated metal ions and solvated electrons. If the blue solution is allowed to stand, the colour fades until it disappears owing to the formation of metal amide. The solutions of metal in liquid conduct electricity because of the presence of solvated electrons. The dilute solutions are paramagnetic because they contain free electrons.

### 1.1.11 Electronegativity Values

The electronegativity values are small which decrease from Li to Cs.

### 1.1.12 Reactivity

The reactivity of alkali metals goes on increasing in the following order.



### 1.1.13 Colourless and Diamagnetic ions

The property of an ion being colourless or coloured depends on the number of unpaired electrons present in the ion. If unpaired electrons are present in anion then these electrons can be excited by energy from light and show colour on coming back to the ground state. The ion which have unpaired electrons show magnetic properties whereas the ions having paired electrons nullify the magnetic fields of each other. Such ions are called diamagnetic ions.

Super oxides are para magnetic and coloured due to the presence of unpaired electrons.

### 1.1.14 Melting and Boiling Point

The cohesive energy is the force holding the atoms or ions together in the solid. The cohesive energy depends on the number of electrons that can participate in bonding. The cohesive force decreases down the group in alkali metals group as they have only one valence electron which

participates in bonding and of the large size and diffusing nature of the outer bonding electron. The atoms become larger on descending down the group, so the bonds are weaker, the cohesive energy decreases and the softness of metal increases. Hence the melting point decreases down the group. Boiling point also decreases down the group.

### 1.2 Chemical Properties

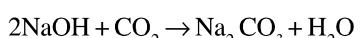
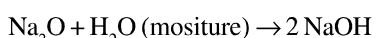
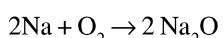
#### Some common reactions of Group 1 metals

Reaction	Comment
$M + H_2O \rightarrow MOH + H_2$	Hydroxides are strongest base known
$Li + O_2 \rightarrow Li_2O$	Monoxide formed by Li and to a small extent by Na
$Na + O_2 \rightarrow Na_2O_2$	Peroxide formed by Na and to a small extent by Li
$K + O_2 \rightarrow KO_2$	Superoxide formed by K, Rb, Cs.
$M + H_2 \rightarrow MH$	Ionic salt like hydrides
$Li + N_2 \rightarrow Li_3N$	Nitride formed only by Li
$M + S \rightarrow M_2S$	All metals form sulphides
$M + X_2 \rightarrow MX$	All the metals form halide
$M + NH_3 \rightarrow MNH_2$	All the metals form amides.

#### 1.2.1 Reaction with Air

Group 1 elements are very reactive and tarnish rapidly in air. Cs burns spontaneously in air.

These metals form alkaline carbonates in moist air.



#### 1.2.2 Reaction with O<sub>2</sub>

Li forms  $\text{Li}_2\text{O}$ , Na forms two type of oxide

$(\text{M}_2\text{O}, \text{M}_2\text{O}_2)$  and K, Rb, Cs forms superoxides ( $\text{MO}_2$ ).

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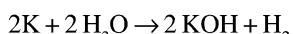
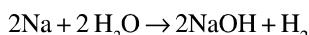
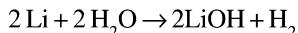
### 1.2.2.1 Basic Nature, Ionic Nature of the Oxides

Basic nature of oxides increases from Li to Cs due to increase in the size of cation

- Size of cation increases from Li to Cs. According to Fajan's Rule, ionic character of these oxides increases from Li to Cs.
- Solubility in water increases from Li to Cs oxides, due to increase in ionic character of these metal oxides.

### 1.2.3 Reaction with water

Group 1 metals react with water liberating H<sub>2</sub> and forming hydroxides.



### 1.2.4 Reaction with Hydrogen

Group 1 metals reacts with H<sub>2</sub> to form ionic hydrides. Thermal stability of LiH is high.

Stability of hydrides :



### 1.2.5 Reaction with dilute acids

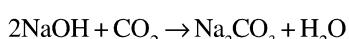
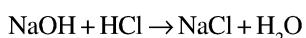
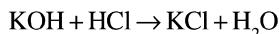
Due to alkaline nature, these metals react rapidly with dilute acids and the rate of reaction increases from Li to Cs because of increase in basic character.

## 1.3 Compounds of Alkali Metal

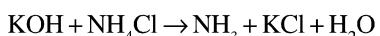
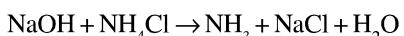
### 1.3.1 Hydroxides

NaOH is often called as caustic soda. KOH is called caustic potash because of their corrosive properties. The caustic alkali are the strongest base in aqueous solution. The solubility of hydroxides increases down the group.

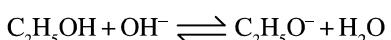
The bases react with acids to form salt and H<sub>2</sub>O.



The bases liberate ammonia from ammonium salts



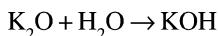
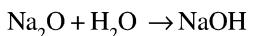
KOH resembles NaOH in all its reaction. but as KOH is much more expensive it is seldom used. However KOH is more soluble in alcohol, thus producing C<sub>2</sub>H<sub>5</sub>O<sup>-</sup> ions by the equilibrium



This accounts for the use of alcoholic KOH in organic chemistry Group 1 hydroxides are thermally stable.

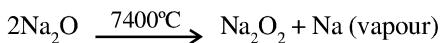
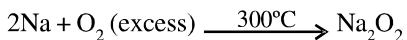
### 1.3.2 Oxides, Peroxides, Superoxides

**Normal oxides - monoxide :** The monoxides are ionic. They are strongly basic oxides and they react with water form strong bases.

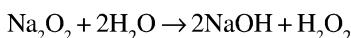
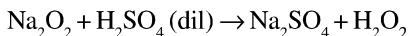


### Peroxides

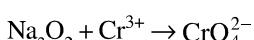
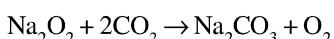
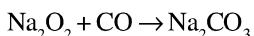
#### Preparation :



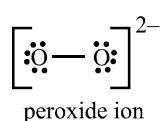
### Properties



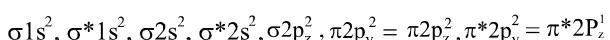
Na<sub>2</sub>O<sub>2</sub> is a powerful oxidant Because it reacts with CO<sub>2</sub> in the air it has been used to purify the air in submarines



### Structure



O is sp<sup>3</sup> hybridised. The peroxide ion has 18 electrons which occupy the molecular orbitals as shown.



Bond order is 1 and it is diamagnetic

### Superoxides ( $O_2^-$ )

Superoxides are ionic oxides  $M^+O_2^-$

#### Preparation :



(M = K, Rb, Cs)

Superoxides are stronger oxidizing agents than peroxides.

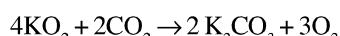
The stability of these superoxides is in the order



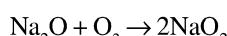
#### Reactions



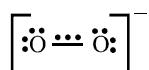
$KO_2$  is used in space capsules, submarines and breathing masks because it produces  $O_2$  and removes  $CO_2$



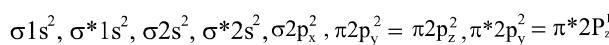
sodium superoxide cannot be prepared by burning metal in oxygen but it can be prepared by reacting sodium peroxide with  $O_2$  at high temperature and pressure



#### Structure



The presence of one unpaired electron in 3 electron bond explains paramagnetic character. The superoxide has 17 electrons which give a bond order of 1.5 which occupy the molecular orbitals as shown



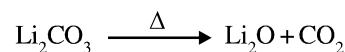
**Stability of oxides :** Normal oxide > peroxide > superoxide.

### 1.3.3 Carbonates and Bicarbonates

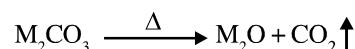
Group 1 metals form solid bicarbonates ( $MHCO_3$ ). All alkali metals form carbonates of type  $M_2CO_3$ . Due to highly electro positive nature of the alkali metals their carbonates and bicarbonates are highly stable to heat ( $Li_2CO_3$  decomposes easily by heat).

The exceptional behaviour of  $Li_2CO_3$  can be explained by

- (a) small size and strong polarisation of Li distorts the  $e^-$  cloud of the near by oxygen atom of the large  $CO_3^{2-}$  to such an extent that the C—O bond gets weakened.

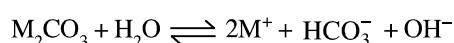
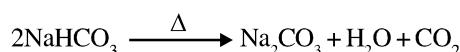
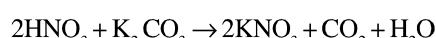


- (b) Replacement of the larger carbonate ion by a smaller ion leads to increased lattice energy and thus favours the decomposition.



$Na_2CO_3$  is used as **washing soda**.  $NaHCO_3$  is used as **baking soda**. The crystal structure of  $NaHCO_3$  and  $KHCO_3$  both show hydrogen bonding. In  $NaHCO_3$ , the  $HCO_3^-$  linked into an infinite chain whilst in  $KHCO_3$  a dimeric anion is formed.

#### Reactions



They hydrolyze to give basic solution.

### 1.3.4 Halides

All the metals in this group form halides of type  $MX$ .

$Li^+$  is the smallest ion in the group, it would be expected to form hydrated salts more readily than other metals.

#### Properties :

As evident from their following properties, alkali metal halides are ideal ionic compounds.

- (i) All alkali halides except lithium fluoride are freely soluble in water (LiF is soluble in non-polar solvents).
- (ii) They have high melting and boiling points.
- (a) For the same alkali metal, the melting and boiling points decrease regularly in the order

Fluoride > chloride > bromide > iodide

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This is explained on the basis of **lattice energy\*** of these metal halides. For the same metal, lattice energy decreases with the decrease in electronegativity of the halogen. For example,

Metal Halide :	NaF	NaCl	NaBr	NaI
Lattice energy (kJ/mole) :	910	769	732	682
Melting point (K) :	1298	1081	1028	934

- (b) For the same halide ion, the melting point of lithium halides are lower than those of the sodium halides. However, after sodium the melting points of halides decrease as we move down the group from Na to Cs. This abnormal behaviour shown by lithium halides is probably due to its covalent nature whereas sodium and other halides are ionic in nature. Amongst ionic halides, melting point decreases as lattice energy decreases as we move down the group i.e.,



1081 K	1045 K	990 K	918 K
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- (iii) **Solubility of halides of alkali metals :-** The solubilities of alkali metal halides show a gradation. For example, solubility of alkali metal fluorides in water increases regularly from lithium to caesium.

Metal fluoride :	LiF	NaF	KF	RbF	CsF
Solubility in water at 298 K (gm/litre) :	2-7	42	1020	1310	3700

In case of chlorides, LiCl has much higher solubility in water than NaCl. This is due to small size of  $\text{Li}^+$  ion and much higher hydration energy. However, from NaCl to CsCl, solubility in water increases regularly due to decrease in their lattice energy.

- (iv) They are good conductors of electricity in the fused state.
- (v) They have ionic crystal structure. However, lithium halides have partly covalent character due to polarising power of  $\text{Li}^+$  ions.

The structure and stability (solubility) of alkali metal halides are explained by the lattice energy, and polarising power.

- (a) **Lattice energy :-** Lattice energy is the energy released during the formation of a crystal lattice from the respective gaseous cations and anions; or it is the energy required to separate one mole of the solid ionic compound into its gaseous ions. Thus lattice energy (the force of attraction among the ions) is a direct measure of the stability of ionic crystals; higher the lattice energy of a compound lower will be its solubility in water.

When a crystal of an ionic compound comes in contact with a polar solvent such as water, the hydrogen end (positive pole) of the water molecule is attracted to a negative ion while the oxygen end (negative pole) is attracted to a positive ion. This attachment of polar solvent molecules to the ions is known as **solvation (or hydration)**, if the solvent is water) of the ions. With the stabilization of the ions by solvation, a large amount of **solvation energy (or hydration energy)** is released which if exceeds the lattice energy of the crystal causes the dissolution of the ionic compound in the solvent. On the other hand, if the solvation energy is not enough to counteract the lattice energy, the substance remains insoluble as in case of lithium fluoride. The high lattice energy of lithium fluoride is due to the combination of small lithium ion with small fluoride ion. In general, for a given ion, the lattice energy increases as the size of the oppositely charged ion decreases.

- (b) **Polarising power and polarisability (Fajan's rule)** :- Although an ionic bond in a compound like  $\text{M}^+\text{X}^-$  is considered to be 100% ionic, in some cases (e.g., lithium halides) it is found to have significant covalent character. According to Fajan, when the two oppositely charged ions approach each other, the nature of the bond between them depends upon the effect of one ion on the other.

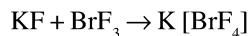
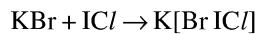
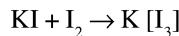
When two oppositely charged ions approach each other, the positive ion attracts electrons present on the outermost shell of the anion and repels its positively charged nucleus. This results in the distortion, deformation or polarisation of the anion. Thus the power of a cation to distort the anion is known as its polarisation power and the tendency of the anion to get polarised by the cation is known as its **polarisability**. If the polarisation is quite small,

an ionic bond is formed, while if the degree of polarisation is large, electrons are drawn from the anion to the cation by electrostatic attraction with the result the electron density between the two ions is increased and the resulting bond becomes covalent in character. In general, greater the polarisation power or polarisability of an ion greater will be its tendency to form covalent bond. Since polarisation power increases with the decrease in the size of the cation while polarisability increases with the increase in the size of the anion, in a compound consisting of large negative ions and small positive ions the polarization may be so marked that the bond becomes covalent. Thus lithium iodide, consisting of  $\text{Li}^+$  ions (the smallest alkali metal ion) and  $\text{I}^-$  ions (the greatest halide ion), is found to be markedly covalent in nature.

Other examples of such ionic-covalent compounds are  $\text{AlCl}_3$ ,  $\text{FeCl}_3$ ,  $\text{SnCl}_4$ , etc.

### Reactions

The alkali metal halides react with the halogen and interhalogen compounds forming ionic polyhalides.



### 1.3.5 Sulphates

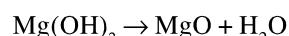
They form sulphates of type  $\text{M}_2\text{SO}_4$

### 1.4 Anomalous Behaviour of Lithium

Although lithium exhibits most of the characteristic properties of the group I elements, it differs, at the same time, in many respects from them. This anomalous behaviour of lithium is due to extremely small size of lithium atom and its ion. The small size of the  $\text{Li}^+$  ion leads to its high charge density. Lithium ion, therefore, possesses the greatest polarising power out of all the alkali metal ions. Hence it exerts a great distorting effect on a negative ion. Consequently, the  $\text{Li}^+$  ion has a remarkable tendency towards solvation and covalent bond formation. Further it is important to note that the

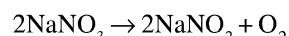
polarising power of the  $\text{Li}^+$  ion is similar to that of  $\text{Mg}^{2+}$  ion, hence the two elements (placed diagonally in the periodic table) resemble very much in their properties.

- (i) Lithium is much harder than the other elements of group I (similarity with Mg which is also a hard metal).
- (ii) The m.p. and b.p. of lithium are comparatively high.
- (iii) Unlike other elements of this group, lithium is the least reactive as represented by the following points.
  - (a) Unlike others, it is not affected by air.
  - (b) Unlike others, it decomposes water very slowly (resemblance with Mg).
  - (c) Unlike others, it hardly reacts with bromine.
  - (d) On burning in oxygen, it forms only the monoxide  $\text{Li}_2\text{O}$ , while the others form peroxides ( $\text{M}_2\text{O}_2$ ) too. Further, K, Rb and Cs form even the superoxide,  $\text{MO}_2$ .
- (iv) Unlike other elements, it directly combines with nitrogen to form nitride,  $\text{Li}_3\text{N}$  (similarity with Mg):
- (v) Lithium is much less electropositive and, therefore, several of its compounds ( $\text{Li}_2\text{CO}_3$  and  $\text{LiOH}$ ) are less stable (similarity with Mg). For example,



Hydroxides of other alkali metals sublime unchanged.

- (vi) Lithium nitrate, on heating, gives nitrogen dioxide and oxygen leaving behind lithium oxide (similarity with  $\text{MgNO}_3$ ), while sodium and potassium nitrates evolve only oxygen, thus leaving nitrites.



- (vii) Most of the lithium salts (e.g., hydroxide, carbonate, oxalate, phosphate and fluoride) are sparingly soluble in water (similarity with Mg.) The corresponding salts of sodium and potassium are freely soluble in water.
- (viii) Lithium halides and lithium alkyls are soluble in organic solvents, while those of Na and K are insoluble;  $\text{MgCl}_2$  is also soluble in alcohol.

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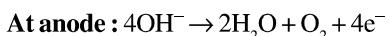
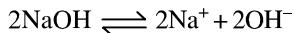
- (ix) Lithium chloride (like  $MgCl_2$ ) undergoes hydrolysis in hot water though to a small extent; while  $NaCl$  and  $KCl$  do not hydrolyse at all.
- (x) Unlike sulphates of other alkali metals, lithium sulphate does not form alums.
- (xi) Lithium compounds, particularly, lithium halides are partially covalent in nature. This is due to the tendency of  $Li^+$  to draw electrons towards itself (polarising power). This explains the lower Value (e.g., 6.25 D in  $LiI$ ) of the dipole moment of lithium compounds than the expected (e.g., 11.5 D in  $LiI$ ).
- (xii) The ions and its compounds are more heavily hydrated than those of the other alkali metals (similarity with  $Mg$ .)

### 1.5 Extraction of Sodium

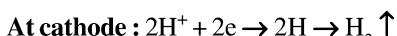
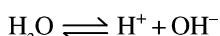
Sodium is obtained on large scale by two processes :

#### 1.5.1 Castner's process

In this process, electrolysis of fused sodium hydroxide is carried out at  $330^\circ C$  using iron as cathode and nickel as anode.



During electrolysis, oxygen and water are produced. Water formed at the anode gets partly evaporated and is partly broken down and hydrogen is discharged at cathode.

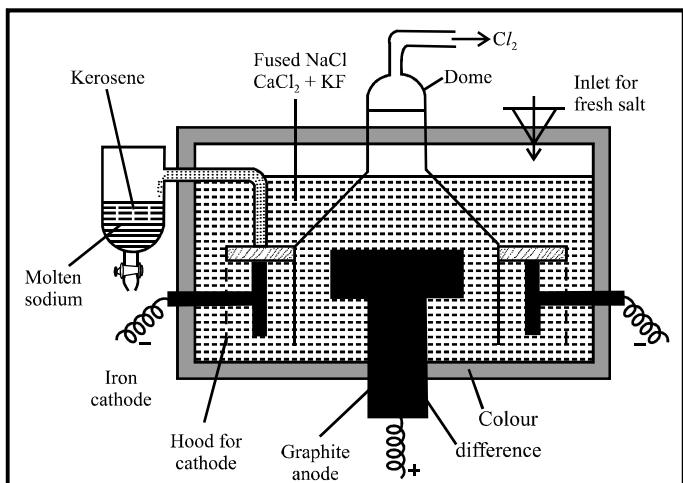


#### 1.5.2 Down's process

Now-a-days the metal is manufactured by Down's process. It involves the electrolysis of fused sodium chloride containing calcium chloride and potassium fluoride, using iron as cathode and graphite as anode, at about  $600^\circ C$ .

The cell consists of a steel tank lined with heat-resistant bricks. A circular graphite anode is placed in the centre of the cell which is surrounded by a cylindrical iron cathode. The anode and cathode are separated by a steel gauze cylinder through which fused charge can

pass. The anode is covered by a dome-shaped steel hood which provides the outlet for the escape of chlorine gas. The molten metal liberated at the cathode moves up and flows into the receiver containing kerosene.



Sodium obtained by this method is 99.5% pure.

The electrolysis of pure sodium chloride presents the following difficulties :

- (i) The fusion temperature of  $NaCl$  is high, i.e.,  $803^\circ C$  (1076 K) which is difficult to maintain.
- (ii) Sodium is volatile at this temperature and therefore, a part of it vapourises and forms a metallic fog.
- (iii) At this temperature, the products of electrolysis, sodium and chlorine are corrosive and may attack the material of the cell.

To remove the above difficulties, pure sodium chloride is mixed with calcium chloride and potassium fluoride. Calcium chloride and potassium fluoride do not decompose at the voltage employed however, they lower the fusion temperature. The fusion temperature of a mixture containing 40%  $NaCl$  and 60% calcium chloride and a very small amount of potassium fluoride becomes about  $600^\circ C$ . The electrolysis of this mixture at  $600^\circ C$  is done in the electrolytic cell.

**EXAMPLE-1** Alkali metals are paramagnetic but their salts are diamagnetic. Explain.

**Sol.** In metals, the outermost energy shell is singly occupied, but in cations, all the orbitals are doubly occupied (inert gas configuration).

e.g.,  $\text{Na}$ ,  $1s^2, 2s^2 2p^6, 3s^2 3p^6, (4s^1)$  paramagnetic

$\text{Na}^+$   $1s^2, 2s^2 2p^6, (3s^2 3p^6)$  Diamagnetic

**EXAMPLE-2** Alkali metals are good reducing agents. Explain.

**Sol.** Alkali metals act as strong reducing agents because they can lose valence electrons readily on account of low ionisation enthalpy values and high values of oxidation potential.

**EXAMPLE-3** Which alkali metal ion has the maximum polarising power and why ?

**Sol.**  $\text{Li}^+$  ion has the maximum polarising power among the alkali metal ions. This is due to small size of  $\text{Li}^+$  ion.

**EXAMPLE-4**  $\text{Li}^+$  ion is far smaller than other alkali metal ions but it moves through a solution less rapidly than the others. Explain.

Or

The conductance of lithium salts is less in comparison to the salts of other alkali metals. Explain.

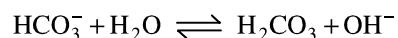
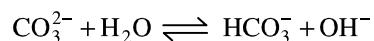
**Sol.** The dense charge of  $\text{Li}^+$  attracts several water molecules around it, i.e.,  $\text{Li}^+$  ion has maximum degree of hydration. Thus, the size of the hydrated lithium ion is largest in comparison to the size of the other alkali metal ions which affects its movement in solution and the conductance is less.

Size :  $[\text{Li(aq)}]^+ > [\text{Na(aq)}]^+ > [\text{K(aq)}]^+$

**EXAMPLE-5** Sodium salts in aqueous solutions are either neutral or alkaline in nature. Explain.

**Sol.** The anions in sodium salts are either from strong acids or weak acids. When anions are from strong acids, there is no hydrolysis and aqueous solutions are neutral ( $\text{NaCl}$ ,  $\text{NaNO}_3$ ,  $\text{Na}_2\text{SO}_4$  solutions are neutral).

On the other hand, when anions are from weak acids, there is hydrolysis and the solutions are alkaline in nature. For example, in the case of sodium carbonate or bicarbonate, solns. are alkaline.

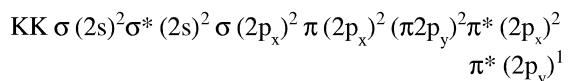


**EXAMPLE-6** Why do K, Rb and Cs form superoxides in preference to oxides and peroxides on being heated in excess supply of air ?

**Sol.**  $\text{K}^+, \text{Rb}^+$  and  $\text{Cs}^+$  are large cations in size and superoxide ion ( $\text{O}_2^-$ ) is larger in size in comparison to oxide ( $\text{O}^{2-}$ ) and peroxide ( $\text{O}_2^{2-}$ ) ion. A larger cation can stabilise a large anion and therefore, these metals form superoxides rather than oxides and peroxides.

**EXAMPLE-7** Why is  $\text{KO}_2$  paramagnetic ?

**Sol.** The superoxide  $\text{O}_2^-$  is paramagnetic because of one unpaired electron in  $\pi^* 2p$  molecular orbital.



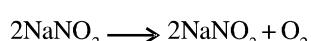
**EXAMPLE-8** Among the alkali metals which element has :  
 (i) highest melting point  
 (ii) highest size of hydrated ion in solution  
 (iii) strongest reducing agent in solution  
 (iv) least electronegative

**Sol.** (i) Li, (ii)  $[\text{Li(aq)}]^+$  (iii) Li, (iv) Cs

**EXAMPLE-9** What happens when following compounds are heated ?

- |                              |   |
|------------------------------|---|
| (a) $\text{Li}_2\text{CO}_3$ | (b) $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ |
| (c) $\text{LiNO}_3$          | (d) $\text{NaNO}_3$                                     |

**Sol.**  $\text{Li}_2\text{CO}_3 \longrightarrow \text{Li}_2\text{O} + \text{CO}_2$



## s-BLOCK ELEMENTS

**EXAMPLE – 10** (a) Arrange LiF, NaF, KF, RbF and CsF in order of increasing lattice energy.

(b) Arrange the following in order of the increasing covalent character.

MCl, MBr, MF, MI (where M = alkali metal)

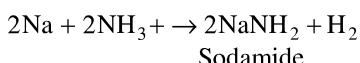
**Sol.** (a) CsF < RbF < KF < NaF < LiF  
 (b) MF < MCl < MBr < MI  
 with increasing size of the anion, covalent character increases.

**EXAMPLE – 11** Why a standard solution of sodium hydroxide cannot be prepared by weighing ?

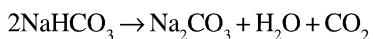
**Sol.** NaOH is a deliquescent substance. It absorbs moisture and reacts with CO<sub>2</sub> of the atmosphere and both increase its mass. Thus, accurate weighing is difficult.

**EXAMPLE – 12** What happens when :  
 (a) fused sodium reacts with dry ammonia.  
 (b) sodium hydrogen carbonate is heated.  
 (c) sodium hydroxide is heated with sulphur ?

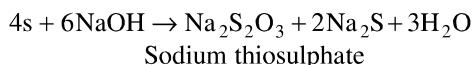
**Sol.** (a) Sodamide is formed with evolution of hydrogen



(b) Sodium carbonate is formed.



(c) Sodium thiosulphate is formed.



**EXAMPLE – 13** Give reasons for the following :

- (i) LiCl is more covalent than NaCl
- (ii) LiI has lower melting point than LiCl
- (iii) MgCl<sub>2</sub> is more covalent than NaCl
- (iv) CuCl is more covalent than NaCl

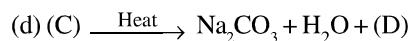
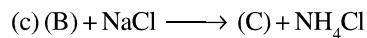
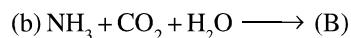
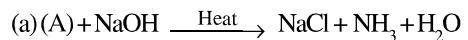
**Sol.** (i) Due to smaller size, Li<sup>+</sup> ion is more polarising than Na<sup>+</sup> and hence LiCl is more covalent than NaCl

(ii) Due to bigger size, I<sup>-</sup> is more polarisable than Cl<sup>-</sup> and hence LiI is more covalent than LiCl. Therefore, LiI has lower melting point than LiCl.

(iii) Due to higher charge, Mg<sup>2+</sup> is more polarising than Na<sup>+</sup> and hence MgCl<sub>2</sub> is more covalent than NaCl.

(iv) Due to pseudo inert gas configuration, Cu<sup>+</sup> is more polarising than Na<sup>+</sup>, and hence CuCl is more covalent than NaCl.

**EXAMPLE – 14** Identify (A), (B), (C) and (D) and give their chemical formulae.

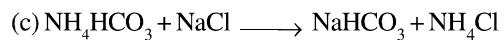


**Sol.** (a) NH<sub>4</sub>Cl + NaOH  $\xrightarrow{\text{Heat}}$  NH<sub>3</sub> + NaCl + H<sub>2</sub>O

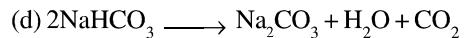
(A) is ammonium chloride (NH<sub>4</sub>Cl).



(B) is ammonium bicarbonate (NH<sub>4</sub>HCO<sub>3</sub>).



(C) is sodium bicarbonate (NaHCO<sub>3</sub>).



(D) is carbon dioxide (CO<sub>2</sub>).

**EXAMPLE – 15** Arrange the following as specified :

- (i) MgO, SrO, K<sub>2</sub>O and Cs<sub>2</sub>O (increasing order of basic character)
- (ii) LiCl, LiBr, LiI (decreasing order of covalent character)
- (iii) NaHCO<sub>3</sub>, KHCO<sub>3</sub>, Mg(HCO<sub>3</sub>)<sub>2</sub>, Ca(HCO<sub>3</sub>)<sub>2</sub> (decreasing solubility in water)
- (iv) LiF, NaF, RbF, KF and CsF (in order of increasing lattice energy)
- (v) Li, Na, K (in order of decreasing reducing nature in solution)

**Sol.** (i) MgO < SrO < K<sub>2</sub>O < Cs<sub>2</sub>O

(ii) LiI > LiBr > LiCl

(iii) NaHCO<sub>3</sub> < KHCO<sub>3</sub> < Mg(HCO<sub>3</sub>)<sub>2</sub> < Ca(HCO<sub>3</sub>)<sub>2</sub>

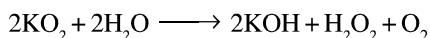
(iv) CsF < RbF < KF < NaF < LiF

(v) Li > K > Na

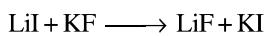
**EXAMPLE-16** (a) What happens when  $KO_2$  reacts with water ? Give the balanced chemical equation.  
 (b) Predict giving reason the outcome of the reaction :

$$LiI + KF \longrightarrow$$

**Sol.** (a) When  $KO_2$  reacts with water, oxygen is evolved and an alkaline solution containing potassium hydroxide and  $H_2O_2$  is formed ?



(b) LiI reacts with KF i.e., anions are exchanged



The exchange occurs as stable compounds are formed i.e., larger cation stabilizes larger anion and smaller cation stabilizes smaller anion.

## 2. ALKALINE EARTH METALS

### 2.1 Introduction

The elements, **beryllium, magnesium, calcium, strontium, barium** and **radium**, constitute Group-II of the periodic table.

All these elements are metals. The oxides of calcium, strontium and barium were known much earlier than the metals themselves and they were called alkaline earths since they were alkaline and existed in the earth. Later, when the elements were discovered, they were named as alkaline earth metals. Radium corresponds to the alkaline earth metals in its chemical properties but, being a radioactive element, it is studied separately along with the other radioactive elements.

### 2.2 Physical Properties

#### 2.2.1 Atomic size

Atomic size increases down the group.

#### 2.2.2 Oxidation State

The group 2 elements exhibit +2 oxidation state.

#### 2.2.3 Density

The size of group 2 elements are smaller than those of group 1 thus these elements have higher density than group 1 elements. Density increases from Be to Ra

**Exception :** Density of Ca is less than Mg and density of Mg is less than Be.

### 2.2.4 Nature of bonds

Be mainly forms covalent compound. The rest of the elements in group 2 forms ionic bond.

### 2.2.5 Hydration energy

The hydration energies of the group 2 ions are four or five times greater than for group 1 ions due to their smaller size and increased charge.  $\Delta H_{hydration}$  decreases down the group as the size of the ions increases.

### 2.2.6 Lattice Energy

Lattice energy of salts of alkali metals having common anion decreases on descending down the group.

### 2.2.7 Ionization Energy

Since the atoms are smaller than those in group 1, the electrons are more tightly held so that the energy required to remove the first electron (first ionization energy) is greater than for group 1. Energy required to remove the second electron is nearly double than the first. Therefore energy required to produce divalent ions for group 2 elements is four times greater than the energy required to produce  $M^+$  from group 1 metals.

### 2.2.8 Flame Test

When energy is supplied to these elements in a flame, their electrons are excited to higher energy states, as is the case with alkali metals under similar conditions. As the electrons drop back to the original energy level, the extra energy is emitted in the form of visible light with characteristic colours as given below :

Element	Colour
Ca	brick red
Sr	crimson red
Ba	grassy green
Ra	crimson

Beryllium and magnesium atoms are smaller. The electrons in these atoms are, therefore, more strongly bound. Hence these are not excited by the energy of the flame to higher

## s-BLOCK ELEMENTS

energy states. These elements, therefore, do not give any colour in bunsen flame.

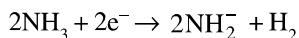
### 2.2.9 Standard oxidation potential

#### Standard Oxidation Potential of Alkaline Earth Metals

Element	Oxidation Reaction	Standard Oxidation Potential (volt)
Be	$\text{Be} \rightarrow \text{Be}^{2+} + 2\text{e}^-$	1.85
Mg	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$	2.37
Ca	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$	2.87
Sr	$\text{Sr} \rightarrow \text{Sr}^{2+} + 2\text{e}^-$	2.89
Ba	$\text{Ba} \rightarrow \text{Ba}^{2+} + 2\text{e}^-$	2.90

### 2.2.10 Solubility in liquid $\text{NH}_3$

The metals dissolve in liquid ammonia as do group 1 metals. Dilute solutions are blue in colour due to the formation of solvated  $\text{e}^-$ . The solution decomposes forming amides and evolving  $\text{H}_2$



### 2.2.11 Electronegative values

The electronegativity values of group 2 are low but are higher than the values of group 1. The value of decreases down the group.

### 2.2.12 Colourless and diamagnetism

The elements of alkaline earth metal group form  $\text{M}^{2+}$  ions, since it does not have any unpaired electron, it is diamagnetic and colourless

### 2.2.13 Melting and Boiling point

Since the cohesive force decreases down the group the melting point of elements of group 2 decreases down the group.

**Exception :** Mg has the lowest melting point.

Boiling points do not show regular trends. They are harder than alkali metals .

### 2.2.14 Metallic properties

Elements of Group-II have typical metallic properties. They show good metallic luster and high electrical as well as thermal conductivity.

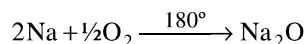
## GROUP – I & II

### OXIDES

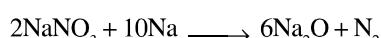
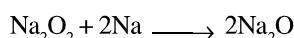
#### Sodium Oxide ( $\text{Na}_2\text{O}$ ) :

##### Preparation :

- (i) 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 sodium.

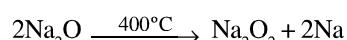


- (ii) By heating sodium peroxide, nitrate or nitrate with sodium.

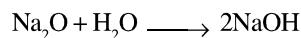


##### Properties :

- (i) It is white amorphous mass.  
(ii) It decomposes at  $400^\circ\text{C}$  into sodium peroxide and sodium



- (iii) It dissolve violently in water, yielding caustic soda.



#### Sodium Peroxides ( $\text{Na}_2\text{O}_2$ ) :

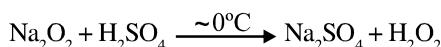
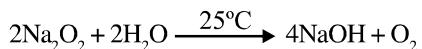
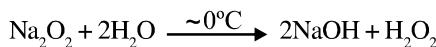
**Preparation :** It is formed by heating the metal in excess of air or oxygen at  $300^\circ\text{C}$ , which is free from moisture and  $\text{CO}_2$ .



##### Properties :

- (i) It is a pale yellow solid, becoming white in air from the formation of a film of  $\text{NaOH}$  and  $\text{Na}_2\text{CO}_3$ .

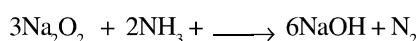
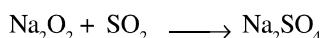
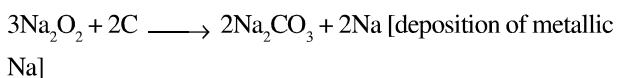
- (ii) In cold water ( $\sim 0^\circ\text{C}$ ) produces  $\text{H}_2\text{O}_2$  but at room temperature produces  $\text{O}_2$ . In ice-cold mineral acids also produces  $\text{H}_2\text{O}_2$ ,



- (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,  $\text{NH}_3$ ,  $\text{SO}_2$ .



- (v) It contains peroxide ion  $[\text{---O---O---}]^{2-}$

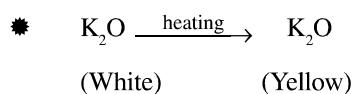
#### Uses :

- (i) For preparing  $\text{H}_2\text{O}_2$ ,  $\text{O}_2$
- (ii) Oxygenating the air in submarines
- (iii) Oxidising agent in the laboratory.

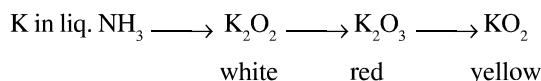
#### Oxides of Potassium :

	Colours
$\text{K}_2\text{O}$	White
$\text{K}_2\text{O}_2$	White
$\text{K}_2\text{O}_3$	Red
$\text{KO}_2$	Bright Yellow
$\text{KO}_3$	Reddish brown needles

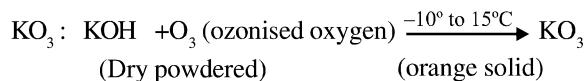
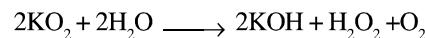
#### Preparations :



- (iii) Passage of  $\text{O}_2$  through a blue solution of K in liquid  $\text{NH}_3$  yields oxides  $\text{K}_2\text{O}_2$  (white),  $\text{K}_2\text{O}_3$  (red) and  $\text{KO}_2$  (deep yellow) i.e.

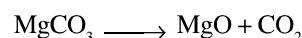


- \*  $\text{KO}_2$  reacts with  $\text{H}_2\text{O}$  and produces  $\text{H}_2\text{O}_2$  and  $\text{O}_2$  both



#### Magnesium Oxide ( $\text{MgO}$ ) :-

It is also called magnesia and obtained by heating natural magnesite.

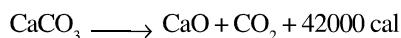


#### Properties :

- (i) It is white powder.
- (ii) Its m.p. is  $2850^\circ\text{C}$ . Hence used in manufacture of refractory bricks for furnaces.
- (iii) It is very slightly soluble in water imparting alkaline reaction.

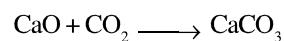
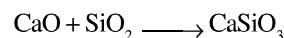
#### Calcium Oxide ( $\text{CaO}$ ) :

It is commonly called as quick lime or lime and made by decomposing lime stone at a high temperature about  $1000^\circ\text{C}$ .



#### Properties :

- (i) It is white amorphous powder of m.p.  $2570^\circ\text{C}$ .
- (ii) It emits intense light (lime light), when heated in oxygen-hydrogen flame.
- (iii) It is a basic oxide and combines with some acidic oxide e.g.



## s-BLOCK ELEMENTS

- (iv) It combines with water to produce slaked lime.



### Magnesium Peroxide ( $\text{MgO}_2$ ) and Calcium Peroxide ( $\text{CaO}_2$ ) :-

These are obtained by passing  $\text{H}_2\text{O}_2$  in a suspension of  $\text{Mg}(\text{OH})_2$  and  $\text{Ca}(\text{OH})_2$ .

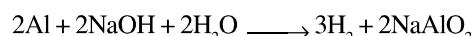
#### Uses :

$\text{MgO}_2$  is used as an antiseptic in tooth paste and a bleaching agent.

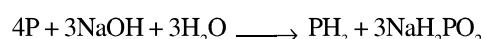
- (v) Acidic and amphoteric oxides gets dissolved easily e.g.



- (vi) Aluminium and Zn metal gives  $\text{H}_2$  from  $\text{NaOH}$



- (vii) Several non metals such as P, S, Cl etc. yield a hydride instead of hydrogen e.g.



### Potassium Hydroxide :

**Preparation :** Electrolysis of aqueous solution of KCl.

**Properties :** Same as NaOH

- \*\* (a) It is stronger base compared to NaOH.
- (b) Solubility in water is more compared to NaOH.
- (c) In alcohol, NaOH is sparingly soluble but KOH is highly soluble.
- (d) As a reagent KOH is less frequently used but in absorption of  $\text{CO}_2$ , KOH is preferably used compared to NaOH. Because  $\text{KHCO}_3$  formed is soluble whereas  $\text{NaHCO}_3$  is insoluble.

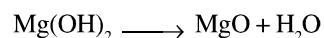
**Magnesium Hydroxide :** It occurs in nature as the mineral brucite.

**Preparation :** It can be prepared by adding caustic soda solution to a solution of Mg-sulphate or chloride solution.



### Properties :

- (i) It can be dried at temperature upto 100°C only otherwise it breaks into its oxide at higher temperature.



- (ii) It is slightly soluble in water imparting alkalinity.  
(iii) It dissolves in  $\text{NH}_4\text{Cl}$  solution



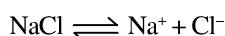
- \* Thus,  $\text{Mg}(\text{OH})_2$  is not therefore precipitated from a solution of  $\text{Mg}^{+2}$  ions by  $\text{NH}_4\text{OH}$  in presence of excess of  $\text{NH}_4\text{Cl}$ .

## HYDROXIDES

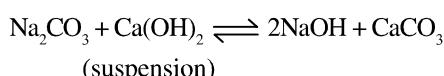
### Sodium Hydroxides :

#### Preparation :

- (i) Electrolysis of Brine :



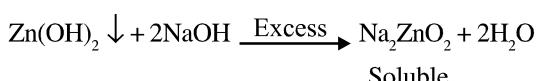
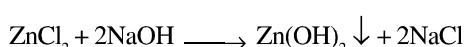
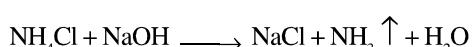
- (ii) Caustication of  $\text{Na}_2\text{CO}_3$  (Gossage's method) :



Since the  $K_{sp}$  ( $\text{CaCO}_3$ ) <  $K_{sp}$  ( $\text{Ca}(\text{OH})_2$ ), the reaction shifts towards right.

#### Properties :

- (i) It is white crystalline, deliquescent, highly corrosive solid.
- (ii) It is stable towards heat.
- (iii) Its aqueous solution is alkaline in nature and soapy in touch.
- (iv)  $\text{FeCl}_3 + 3\text{NaOH} \longrightarrow \text{Fe}(\text{OH})_3 \downarrow + 3\text{NaCl}$



[Same with  $\text{AlCl}_3$ ,  $\text{SnCl}_2$ ,  $\text{PbCl}_2$ ]

### **Calcium Hydroxide :**

**Preparation :** By spraying water on quicklime.



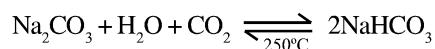
### **Properties :**

- (i) It is sparingly soluble in water.
- (ii) Its solubility in hot water is less than that of cold water. Hence solubility decreases with increase in temperature.
- (iii) It readily absorbs  $\text{CO}_2$  as used as a test for the gas.
- (iv) It is used as a mortar.

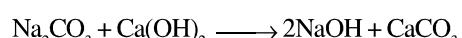
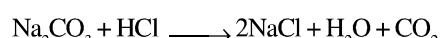
[Mortar is a mixture of slaked lime (1 Part) and sand (3 Parts) made into paste with water.]

- (iii) Hydrated  $\text{Na}_2\text{CO}_3$  is called washing soda ( $\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$ ) and is prepared by Le Blanc process, Solvay process and electrolytic process.

- (iii)  $\text{Na}_2\text{CO}_3$  absorbs  $\text{CO}_2$  yielding sparingly soluble sodium bicarbonate which can be calcined at  $250^\circ\text{C}$  to get pure sodium carbonate.



- (iv) It dissolves in acid with effervescence of  $\text{CO}_2$  and causticised by lime to give caustic soda.



**Uses :** It is widely used in glass making as smelter.

### **Potassium Carbonate :**

By Leblanc process, it can be prepared but by Solvay process it cannot be prepared because  $\text{KHCO}_3$  is soluble in water.

**Properties :** It resembles with  $\text{Na}_2\text{CO}_3$ , m.p. is  $900^\circ\text{C}$  but a mixture of  $\text{Na}_2\text{CO}_3$  and  $\text{K}_2\text{CO}_3$  melts at  $712^\circ\text{C}$ .

**Uses :** It is used in glass manufacturing.

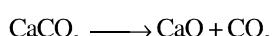
### **Calcium Carbonate :**

It occurs in nature as marble, limestone, chalk, calcite etc. It is prepared by dissolving marble or limestone in HCl and removing iron and aluminium present, by precipitating with  $\text{NH}_3$  and then adding  $(\text{NH}_4)_2\text{CO}_3$  to the solution.



### **Properties :**

- (i) It dissociates above  $1000^\circ\text{C}$  as follows :



- (ii) It dissolves in water containing  $\text{CO}_2$  forming  $\text{Ca}(\text{HCO}_3)_2$  but is precipitated from the solution by boiling.



### **Magnesium Carbonate :**

It occurs in nature as magnesite, isomorphous with calcite. It is obtained as a white precipitate by adding sodium bicarbonate to a solution of a magnesium salt but only basic carbonate, called magnesia alba, having the approximate composition  $\text{MgCO}_3 \cdot \text{Mg(OH)}_2 \cdot 3\text{H}_2\text{O}$  is precipitated.

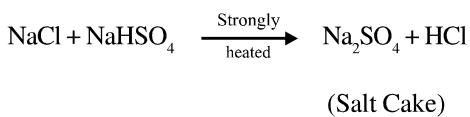
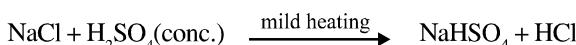
**Properties :** Same as  $\text{CaCO}_3$ .

## **CARBONATES**

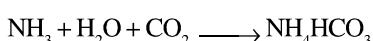
### **Sodium Carbonate :**

#### **Preparation :**

- (i) Leblanc Process :



- (ii) Solvay Process :



#### **Properties :**

- (i) Anhydrous  $\text{Na}_2\text{CO}_3$  is called as soda ash, which does not decompose on heating but melts at  $852^\circ\text{C}$ .
- (ii) It forms number of hydrates.

## BICARBONATES

### Sodium bicarbonates :

**Preparation :** By absorption of  $\text{CO}_2$  in  $\text{Na}_2\text{CO}_3$  solution.



**Uses :** It is used in medicine and as baking powder.

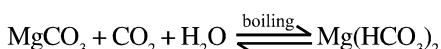
### Potassium bicarbonates :

**Preparation :** Same as  $\text{NaHCO}_3$

**Properties :** Same as  $\text{NaHCO}_3$

But it is more alkaline and more soluble in water compared to  $\text{NaHCO}_3$ .

### Magnesium bicarbonate :



### Calcium bicarbonate :



## CHLORIDES

**Sodium Chloride :** Prepared from brine containing 25% NaCl.

### Properties :

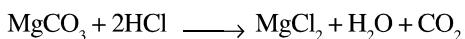
- (i) It is nonhygroscopic but the presence of  $\text{MgCl}_2$  in common salt renders it hygroscopic.
- (ii) It is used to prepare freezing mixture in laboratory [Ice-common salt mixture is called freezing mixture and temperature goes down to  $-23^\circ\text{C}$ .]
- (iii) For melting ice and snow on road.

**Potassium Chloride :** It also occurs in nature as sylvite ( $\text{KCl}$ ) or carnallite  $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ .

**Uses :** It is used as fertiliser.

### Magnesium Chloride :

**Preparation :** By dissolving  $\text{MgCO}_3$  in dil. HCl.

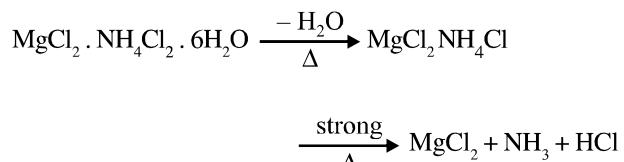


### Properties :

- (i) It crystallises as hexahydrate.  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ .
- (ii) It is deliquescent solid.
- (iii) This hydrate undergoes hydrolysis as follows :



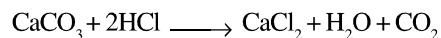
- Hence, anhy.  $\text{MgCl}_2$  cannot be prepared by heating this hydrate.
- Because of the formation of HCl sea water cannot be used in marine boilers which corrodes the iron body.
- (iv) Anhydrous  $\text{MgCl}_2$  can be prepared by heating a double salt like.  $\text{MgClO}_2 \cdot \text{NH}_4\text{Cl} \cdot 6\text{H}_2\text{O}$  as follows :



**Sorel Cement :** It is a mixture of  $\text{MgO}$  and  $\text{MgCl}_2$  (paste like) which sets to hard mass on standing. This is used in dental filling, flooring etc.

### Calcium Chloride :

- (i) It is the by-product in solvay process.
- (ii) It may also be prepared by dissolving the carbonate in HCl



### Properties :

- (i) It is deliquescent crystals.
- (ii) It gets hydrolysed like  $\text{MgCl}_2$  hence anhydrous  $\text{CaCl}_2$  cannot be prepared.



Hence, anhy.  $\text{CaCl}_2$  is prepared by heating  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$  in a current of HCl (dry).

- (iii) Anhy.  $\text{CaCl}_2$  is used in drying gases and organic compounds but not  $\text{NH}_3$  or alcohol due to the formation of  $\text{CaCl}_2 \cdot 8\text{NH}_3$  and  $\text{CaCl}_2 \cdot 4\text{C}_2\text{H}_5\text{OH}$ .

## SULPHATES

### Sodium Sulphate :

#### Preparation :

It is formed in the I<sup>st</sup> step of leblanc process by heating common salt with sulphuric acid.



Thus the salt cake formed is crystallised out from its aqueous solution as  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ . This is called as Glauber's salt.

- One interesting feature of the solubility of glauber's salt is ; when crystallised at below 32.4°C, then  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  is obtained but above 32.4°C,  $\text{Na}_2\text{SO}_4$  (anhy.) comes out.

**Properties :** it is reduced to  $\text{Na}_2\text{S}$  when fused with carbon.



**Uses :** It is used in medicine.

#### Potassium Sulphate :

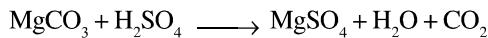
It occurs in stassfurt potash beds as schonite  $\text{K}_2\text{SO}_4$ ,  $\text{MgSO}_4 \cdot 6\text{H}_2\text{O}$  and Kainite,  $\text{KCl} \cdot \text{MgSO}_4 \cdot 3\text{H}_2\text{O}$  from which it is obtained by solution in water and crystallisation. It separates from the solution as crystals whereas  $\text{Na}_2\text{SO}_4$  comes as decahydrate.

**Uses :** It is used to prepare alum.

#### Magnesium Sulphate :

##### Preparation :

- It is obtained by dissolving kieserite  $\text{MgSO}_4 \cdot \text{H}_2\text{O}$  in boiling water and then crystallising the solution as a hepta hydrate. i.e.  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ . It is called as **Epsom salt**.
- It is also obtained by dissolving magnesite in hot dil.  $\text{H}_2\text{SO}_4$ .



or by dissolving dolomite ( $\text{CaCO}_3 \cdot \text{MgCO}_3$ ) in hot dil.  $\text{H}_2\text{SO}_4$  and removing the insoluble  $\text{CaSO}_4$  by filtration.

- It is isomorphous with  $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ ,  $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$

**Calcium Sulphate :** It occurs as anhydrite  $\text{CaSO}_4$  and as the dihydrate  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ , gypsum, alabaster or satin-spar.

#### Properties :

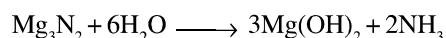
- Gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ )  $\xrightarrow[+H_2O]{120^\circ C}$   $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$  (Plaster of paris)
 

$\downarrow$   
 200°C  
 $\text{CaSO}_4$  (anhydrous)  
 Dead burnt plaster
- Solubility of  $\text{CaSO}_4$  at first increases upto a certain point and then decreases with rise of temperature.
- Plaster of paris is used in wood making due to its porous body.

- EXAMPLE-17** (a)  $\text{Mg}_3\text{N}_2$  when reacted with water, gives off  $\text{NH}_3$  but  $\text{HCl}$  is not obtained from  $\text{MgCl}_2$  on reaction with water at room temperature. Why ?

- (b) The crystalline salts of alkaline earth metals contain more water of crystallization than corresponding alkali metal salts. Why ?

- Sol.** (a)  $\text{Mg}_3\text{N}_2$  is a salt of a strong base and weak acid ( $\text{NH}_3$ ), hence its hydrolysis is possible.



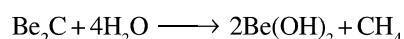
$\text{MgCl}_2$  is a salt of a strong base,  $[\text{Mg}(\text{OH})_2]$ , and a strong acid, ( $\text{HCl}$ ), and hence, its hydrolysis is not possible.

- (b) Due to small size and high nuclear charge, alkaline earth metal ions have higher tendency of hydration in comparison to alkali metal ions. Thus, the salts of alkaline earth metals contain more water of crystallisation than the salts of alkali metals.

- EXAMPLE-18** What happens when :

- Beryllium carbide reacts with water.
- Magnesium nitrate is heated.
- Quick lime is heated in electric furnace with powdered coke.
- $\text{NaOH}$  solution is added to  $\text{ZnCl}_2$  solution.

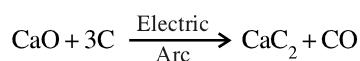
- Sol.** (i) Methane gas is evolved.



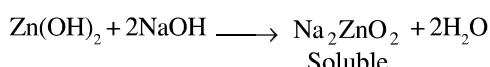
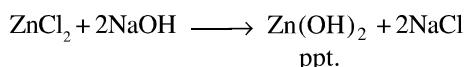
- (ii) Brown coloured gas,  $\text{NO}_2$ , is evolved.



- (iii) Calcium carbide is formed with evolution of  $\text{CCl}_4$ .



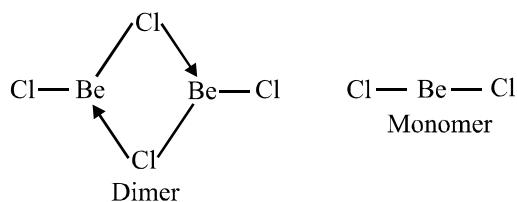
- (iv) While precipitate of  $\text{Zn}(\text{OH})_2$  is formed which dissolves in excess of  $\text{NaOH}$  forming sodium zincate.



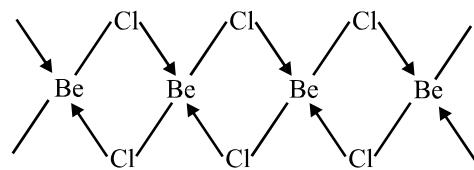
## s-BLOCK ELEMENTS

**EXAMPLE – 19** Draw the structure of (i)  $\text{BeCl}_2$  (vapour state)  
(ii)  $\text{BeCl}_2$  (solid).

**Sol.** In vapour state, it has chlorobridged dimer structure which dissociates into linear monomer at  $1000^\circ\text{C}$ .



In the solid phase, it has a polymeric structure with chlorobridges in which a halogen atom bonded to one beryllium atom uses a lone pair of electrons to form a coordinate bond and to another atom by a covalent bond.



### GENERAL TENDENCY OF DIFFERENT PROPERTIES IN THE PERIOD AND GROUPS

