

## s-Block Elements

### Section (A) : General facts about elements

The elements in which the last electron enters the outermost s-orbital are called s-block elements.

As the s-orbital can accommodate only two electrons, two groups 1 and 2 belong to the s-block.

#### Flame Test

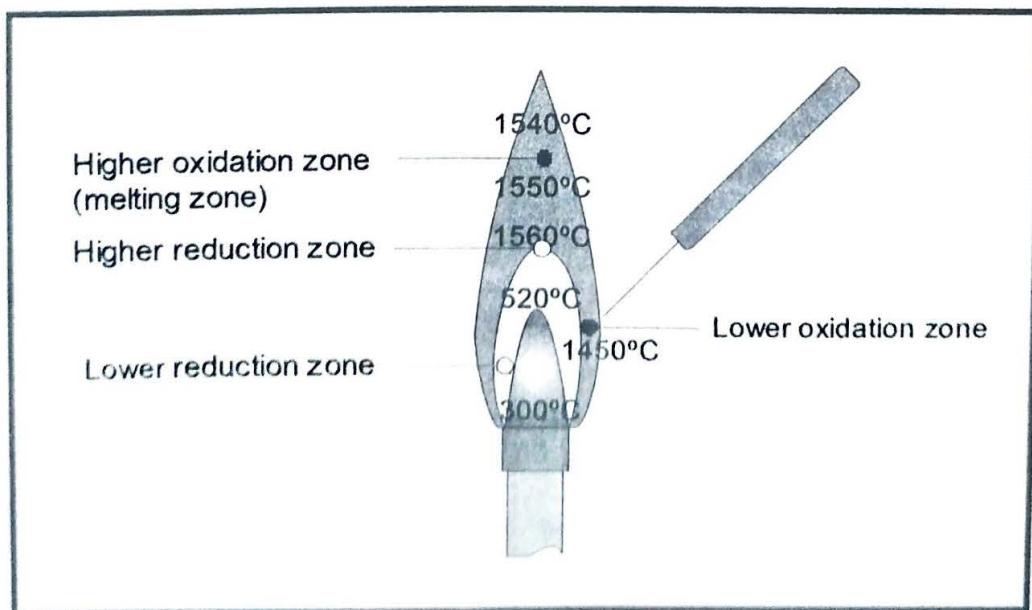
The alkali metals and alkaline earth metals and their salts impart characteristic colour to an oxidizing flame.

**Reason :** This is because the heat from the flame excites the outermost orbital electron to a higher energy level. When they drop back to the ground state, there is emission of radiation in the visible region.

The electrons in beryllium and magnesium are too strongly bound (due to small size) to get excited by flame. Hence, these elements do not impart any colour to the flame.

Metal	Li	Na	K	Rb	Cs
Colour	Crimson red	Yellow	Violet / Lilac	Red violet	Blue

Metal	Be	Mg	Ca	Sr	Ba
Colour	No colour	No colour	Brick red	Crimson red	Apple green



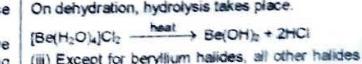
## Section (C) & (D) : Based on Chemical Bonding, Properties of elements

### Properties of Alkali and Alkaline earth metals

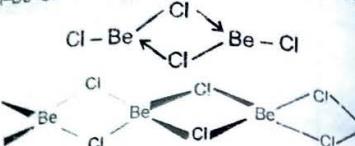
S.No.	Atomic Properties	Alkali metal	Alkaline earth metals
1.	Outer Electronic configuration	$ns^1$	$ns^2$
2.	Oxidation number and valency	(i) These elements easily form univalent +ve ion by losing loosely solitary $ns^1$ electron due to low IP value.	The IP <sub>1</sub> of these metals are much lower than IP and thus it appears that these metals should form univalent ion rather than divalent ions but in actual practice, all these give bivalent ion.
3.	Atomic and Ionic radii	Increase down the group, because value of n (principal quantum number) increases. Order = Li < Na < K < Rb < Cs	The atomic and ionic radii of the alkaline earth metal are smaller than corresponding alkali metals. Reason higher nuclear charge (Zeff) On moving down the group size increase as value of n increases. Be < Mg < Ca < Sr < Ba
4.	Ionisation Energy	As size increases, I.E. decreases down the group (so Cs have lowest I.P.) Order = Li > Na > K > Rb > Cs	Down the group IE decreases due to increase in size. Be > Mg > Ca > Sr > Ba IE <sub>1</sub> of Alkali metal < IE <sub>1</sub> of Alkaline earth metal IE <sub>2</sub> of Alkali metal > IE <sub>2</sub> of Alkaline earth metal Reason IE <sub>1</sub> of Alkaline earth metal is large due to increased nuclear charge in Alkaline earth metal as compared to Alkali metal but IE <sub>2</sub> of Alkali metal is large because second electron in Alkali metal is to be removed from cation which has already acquired noble gas configuration
5.	Electropositive character or metallic character	Alkali metals are strongly electropositive and metallic. Down the group electropositive nature increase so metallic nature also increases. $i.e. M \rightarrow M^+ + e^-$ Metallic Nature : Electropositive character $\propto 1/I.P.$ Order = Li < Na < K < Rb < Cs	Due to low IE they are strong electropositive but not as strong as Alkali metal because of comparatively high IE. The electropositive character increase down the group Order = Be < Mg < Ca < Sr < Ba
6.	Hydration of ions	(i) Hydration represents for the dissolution of a substance in water to get absorb water molecule by weak valency forces Hydration of ions in the process when ions on dissolution in water get hydrated. (ii) Hydration energy $\propto$ charge density on ion Degree of hydration $\propto 1/\text{Cation size} \propto \text{charge} \propto 1/\text{Ionic mobility} \propto 1/\text{conductivity}$ Hydration energy = $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$ (iii) $Li^+$ being smallest in size has maximum degree of hydration and that is why lithium salts are mostly hydrated and moves very slowly under the influence of electric field, e.g. $LiCl \cdot 2H_2O$	Hydration energy = $Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$
7.	Photoelectric effect	The phenomenon of emission of electrons when electromagnetic rays strikes against them is called photoelectric effect. Alkali metal have low IP so show photoelectric effect. Cs and K are used in Photoelectric cells	
8.	Electronegativity	(i) These metals are highly electropositive and thereby possess low values of electronegativities. (ii) Electronegativity of alkali metals decreases down the group. Order = Li > Na > K > Rb > Cs	(i) Their electronegativities are also small but are higher than that of alkali metals (ii) Electronegativity decreases from Be to Ba

S.No.	Physical Property	Alkali metal		Alkaline earth metals
1.	Density	(i) All are light metals. (ii) Density increase down the group but K is lighter than Na. Order = Li < K < Na < Rb < Cs	(i) Heavier than alkali metals. (ii) Density decrease slightly up to Ca after which it increases. (iii) Density of Mg is greater than Ca.	
2.	Hardness	(i) All are silvery white metals. (ii) Light soft, malleable and ductile metals with metallic luster. (iii) Diamagnetic and colour less in form of ions. (iv) These metals are very soft and can be cut with a knife. Lithium is harder than any other alkali metal. The hardness depends upon cohesive energy. Cohesive energy $\propto$ Force of attraction between atoms.	Relatively soft but harder than Alkali metals.	
3.	Melting points/ Boiling points	(i) Lattice energy decreases from Li to Cs and thus Melting points and Boiling points also decrease from Li to Cs. M.P. = Li > Na > K > Rb > Cs B.P. = Li > Na > K > Rb > Cs	They have low Melting points and Boiling points but are higher than corresponding value of group I. Reason They have two valence electrons which may participate in metallic bonding compared with only one electron in Alkali metal. Consequently group II elements are harder and have higher cohesive energy and so, have much higher Melting points / Boiling points than Alkali metal. M.P. = Be > Ca > Sr > Ba > Mg , B.P. = Be > Ba > Ca > Sr > Mg	
4.	Specific heat	It decreases from Li to Cs. Li > Na > K > Rb > Cs (*need not to memorise)	values are lesser than that of alkali metals, decreases down the group. *need not memorise.	

S.No.	Chemical Property	Alkali metals		Alkaline earth metals																								
1.	Action with O <sub>2</sub> and N <sub>2</sub>	(i) They generally form oxides and peroxides. $M + O_2 \rightarrow M_2O$ (Oxide) $\xrightarrow{O_2} M_2O_2$ (Peroxide) The alkali metals tarnish in dry air due to the formation of their oxides on their surface. $4M + O_2 \rightarrow 2M_2O$ They react vigorously in oxygen forming following oxides. $4Li + O_2 \rightarrow 2Li_2O$ (Monoxide) $2Na + O_2 \rightarrow Na_2O_2$ (Peroxide) $M + O_2 \rightarrow MO_2$ (Superoxide) where M = K, Rb, Cs <table border="1"><caption>Principal Combustion Product (Minor Product)</caption><thead><tr><th>Metal</th><th>Oxide</th><th>Peroxide</th><th>Superoxide</th></tr></thead><tbody><tr><td>Li</td><td>Li<sub>2</sub>O</td><td>(Li<sub>2</sub>O<sub>2</sub>)</td><td></td></tr><tr><td>Na</td><td>(Na<sub>2</sub>O)</td><td>Na<sub>2</sub>O<sub>2</sub></td><td></td></tr><tr><td>K</td><td></td><td>KO<sub>2</sub> (Orange/Yellow Crystalline)</td><td></td></tr><tr><td>Rb</td><td></td><td>RbO<sub>2</sub> (Orange/Yellow Crystalline)</td><td></td></tr><tr><td>Cs</td><td></td><td>Cs<sub>2</sub>O<sub>2</sub> (Orange/Yellow Crystalline)</td><td></td></tr></tbody></table> The oxides and peroxides are colourless when pure. (ii) All super oxide are paramagnetic and peroxides are diamagnetic in nature. (iii) The increasing stability of the peroxide or superoxide as the size of the metal ion increases is due to the stabilisation of large anions by larger cations through lattice energy effect. (iv) Since all the alkali metals are highly reactive towards air ; they are kept in kerosene oil. Reactivity increases from Li to Cs. (v) Only Lithium reacts with N <sub>2</sub> (at room temperature) to form ionic lithium nitride Li <sub>3</sub> N because Li being strongest reducing agent converts N <sub>2</sub> into N <sup>3-</sup> . $3Li + 1/2N_2 \rightarrow Li_3N$	Metal	Oxide	Peroxide	Superoxide	Li	Li <sub>2</sub> O	(Li <sub>2</sub> O <sub>2</sub> )		Na	(Na <sub>2</sub> O)	Na <sub>2</sub> O <sub>2</sub>		K		KO <sub>2</sub> (Orange/Yellow Crystalline)		Rb		RbO <sub>2</sub> (Orange/Yellow Crystalline)		Cs		Cs <sub>2</sub> O <sub>2</sub> (Orange/Yellow Crystalline)		(i) Be and Mg are kinetically inert towards oxygen because of formation of a film of oxide on their surface. However powdered Be burn brilliantly. $2Be + O_2 \text{ (air)} \xrightarrow{\Delta} 2BeO$ (amphoteric); $3Be + N_2 \text{ (air)} \xrightarrow{\Delta} Be_3N_2$ (ii) Mg is more electropositive and burns with dazzling brilliance in air give MgO and Mg <sub>3</sub> N <sub>2</sub> . $Mg + O_2 \text{ (air)} \xrightarrow{\Delta} MgO$ ; $Mg + N_2 \text{ (air)} \xrightarrow{\Delta} Mg_3N_2$ Peroxides are coloured due to lattice defect. (Similar property with Li because both shows diagonal relation.) (iii) Ba gives BaO <sub>2</sub> not BaO. (iv) Calcium, strontium and barium are readily attacked by air to form the oxide and nitride. They also react water with increasing vigour even in cold to form hydroxides. (v) BeO, MgO are used as refractory, because they have high M.P. (vi) Other metals (Ba or Sr form peroxide) $M + O_2 \xrightarrow{\Delta} MO_2$	
Metal	Oxide	Peroxide	Superoxide																									
Li	Li <sub>2</sub> O	(Li <sub>2</sub> O <sub>2</sub> )																										
Na	(Na <sub>2</sub> O)	Na <sub>2</sub> O <sub>2</sub>																										
K		KO <sub>2</sub> (Orange/Yellow Crystalline)																										
Rb		RbO <sub>2</sub> (Orange/Yellow Crystalline)																										
Cs		Cs <sub>2</sub> O <sub>2</sub> (Orange/Yellow Crystalline)																										

2.	Action with water	(i) Alkali metals decompose water to form the hydroxides having the formula MOH and dihydrogen. $2M + 2H_2O \rightarrow 2MOH \text{ (aq.)} + H_2(g)$ (M = An alkali metal). (ii) Li decompose water slowly, sodium reacts with water quickly K, Rb and Cs react with water vigorously. (iii) It may be noted that although lithium has most negative E° value (In below table), its reaction with water is less vigorous than that of sodium which has the least negative E° value among the alkali metals. This behaviour of lithium is attributed to its small size and very high hydration energy. Its explanation lies in Kinetics, released energy in case of K, Rb, Cs is sufficient to melt or even vapourise and so more surface area is exposed to the water and kinetically reaction is faster than lithium. Other metals of the group react explosively with water. <table border="1"><caption>Property</caption><thead><tr><th></th><th colspan="6">Alkali metals</th></tr><tr><th></th><th>Li</th><th>Na</th><th>K</th><th>Rb</th><th>Cs</th><th>Fr</th></tr></thead><tbody><tr><td>Standard potentials E°/V for (M<sup>+</sup>/M)</td><td>3.0</td><td>-2.714</td><td>-2.925</td><td>-2.930</td><td>-2.927</td><td>-1</td></tr></tbody></table> (iv) They also react with proton donors such as alcohol, gaseous ammonia and terminal alkynes evolution of hydrogen. $2M + 2C_2H_5OH \rightarrow 2C_2H_5OM + H_2$ Ethyl alcohol Metal ethoxide		Alkali metals							Li	Na	K	Rb	Cs	Fr	Standard potentials E°/V for (M <sup>+</sup> /M)	3.0	-2.714	-2.925	-2.930	-2.927	-1	(i) Ca , Sr , Ba and Ra decompose cold water readily with evolution of hydrogen. $M + 2H_2O \rightarrow M(OH)_2 + H_2$
	Alkali metals																							
	Li	Na	K	Rb	Cs	Fr																		
Standard potentials E°/V for (M <sup>+</sup> /M)	3.0	-2.714	-2.925	-2.930	-2.927	-1																		
3.	Hydrides	(i) They react with H <sub>2</sub> forming metal hydride with formula MH which are of ionic nature. Stability of hydride decreases down the group, since the electropositive character decreases from Cs to Li. $2M + H_2 \rightarrow 2M^0H^0$ (iii) The metal hydrides react with water to give MOH and H <sub>2</sub> (act as reducing agent) $MH + H_2O \rightarrow MOH + H_2$	(i) Except Be, all alkaline earth metals form hydrides (MH <sub>2</sub> ) on heating directly with H <sub>2</sub> . (ii) The stability of hydrides decreases from Be to Ra. (iii) BeH <sub>2</sub> is prepared by the action of LiAlH <sub>4</sub> on BeCl <sub>2</sub> . $BeCl_2 + LiAlH_4 \rightarrow 2BeH_2 + LiCl + AlCl_3$ BeH <sub>2</sub> & MgH <sub>2</sub> is covalent and polymeric but others are ionic. 																					
4.	Halides	(i) The alkali metals react vigorously with halogens to form ionic halides MX <sub>n</sub> . $2M + X_2 \rightarrow 2M^0X^n$ (ii) Alkali metals halides (Cl <sub>2</sub> , Br <sub>2</sub> , I <sub>2</sub> ) formation is increases from Li to Cs due to increase in electropositive character. Order of reactivity towards F <sub>2</sub> Li > Na > K > Rb > Cs (iii) LiX have more covalent character (It is because of the high polarisation capability of Lithium ion (fajans rules)). (iv) Halides having ionic nature have high melting point and are good conductor of current in fused state. These are readily soluble in water. (v) Halides of potassium, rubidium and caesium have property of combining with extra halogen atoms forming polyhalides. $KI + I_2 \rightarrow KI_3$	(i) The alkaline earth metals directly combine with halogens on heating to give metal halides MX <sub>n</sub> (X=F, Cl, Br, I). (ii) Thermal decomposition of (NH <sub>4</sub> ) <sub>2</sub> BeF <sub>6</sub> is the best route for the preparation of BeF <sub>3</sub> and BeCl <sub>2</sub> is conveniently made from the oxide. $BeO + C + Cl \xrightarrow{500-600 K} BeCl_2 + CO$ Anhydrous beryllium halide can not be obtained from materials made in aqueous solution because the hydrated ions [Be(H <sub>2</sub> O) <sub>4</sub> ] <sup>2+</sup> is formed, i.e. [Be(H <sub>2</sub> O) <sub>4</sub> ]Cl <sub>2</sub> On dehydration, hydrolysis takes place. $[Be(H_2O)_4]Cl_2 \xrightarrow{\text{heat}} Be(OH)_2 + 2HCl$ (iii) Except for beryllium halides, all other halides of alkaline earth metals are ionic in nature. Beryllium halides are essentially covalent and soluble in organic solvents. Beryllium chloride has a chain structure in the solid state as shown below : 																					

Cl-Be-Cl



In the vapour phase  $\text{BeCl}_2$  tends to form a chloro-bridged dimer which dissociates into the linear monomer at high temperatures of the order of 1200 K.

(iv) The ionic character of halides increases from Be to Ra.

(v) Beryllium halides have covalent character due to small size and high effective nuclear charge and thus do not conduct electricity in molten state.

(vi) The fluorides are relatively less soluble than the chlorides owing to their high lattice energies.

(vii) The decreases in solubility of halides down the group is due to decrease in hydration energy because of increasing size of metal cation.

(viii) The tendency to form halide hydrates gradually decreases (for example,  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ ,  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ ,  $\text{SrCl}_2 \cdot 6\text{H}_2\text{O}$  and  $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ ) down the group. The dehydration of hydrated chlorides, bromides and iodides of Ca, Sr and Ba can be achieved on heating; however, the corresponding hydrated halides of Be and Mg on heating suffer hydrolysis.

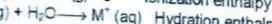
(ix)  $\text{CaCl}_2$  has strong affinity with water and is used as dehydrating agent.

Property	Alkali metals					
	Li	Na	K	Rb	Cs	Fr
Standard potentials $E^\circ/V$ for $(M'/M)$	-3.04	-2.714	-2.925	-2.930	-2.927	—

(i) Reducing agent is electron donor. The alkali metals are strong reducing agents, lithium being the most and sodium the least powerful (above table). The standard electrode potential ( $E^\circ$ ) which measures the reducing power represents the overall change.



Sublimation enthalpy



Ionization enthalpy



Hydration enthalpy

(ii) Lithium is expected to be least reducing agent due to its very high I.E. However, lithium has the highest  $E^\circ$  value and its high reducing power.

Reducing Nature in gas phase

$$= \text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$$

Reducing Nature in aqueous condition

$$= \text{Li} > \text{Cs} > \text{Rb} > \text{K} > \text{Na}$$

The alkaline earth metals are strong reducing agents. This is indicated by large negative values of their reduction potentials (below table). However their reducing power is less than those of their corresponding alkali metals. Beryllium has less negative value compared to other alkaline earth metals.

However, its reducing nature is due to large hydration energy associated with the small size of  $\text{Be}^{2+}$  ion and relatively large value of the atomization enthalpy of the metal.

Property	Alkaline earth metals					
	Be	Mg	Ca	Sr	Ba	Ra
Standard potentials $E^\circ/V$ for $(M'/M)$	-1.97	-2.36	-2.84	-2.89	-2.92	-2.92

(i) These oxides are easily hydrolysed by water to form the hydroxides



(ii) The Hydroxide which are obtained by the reaction of alkali metal hydroxides are the strongest of all bases and dissolve freely in water with evolution of much heat. Basic nature/Solubility in water/Thermal stability

$$= \text{LiOH} < \text{NaOH} < \text{KOH} < \text{RbOH} < \text{CsOH}$$

Basic/thermal stability

$$= \text{Be}(\text{OH})_2 < \text{Mg}(\text{OH})_2 < \text{Ca}(\text{OH})_2 < \text{Sr}(\text{OH})_2 < \text{Ba}(\text{OH})_2$$

7.

Carbonates and bicarbonates

- (i) The carbonates ( $\text{M}_2\text{CO}_3$ ) and bicarbonates ( $\text{MHCO}_3$ ) are highly stable to heat, where M is alkali metals.
- (ii) Group 1 metals are so strongly basic, they (except lithium) also form solid bicarbonates. No other metals form solid bicarbonates. Lithium carbonate is not so stable to heat. Its hydrogen carbonate does not exist as a solid. Although  $\text{NH}_4\text{HCO}_3$  also exists as a solid.
- (iii) The stability of these salts increases with the increasing electropositive character from Li to Cs. It is therefore  $\text{Li}_2\text{CO}_3$  decompose on heating.

Thermal stability/Solubility in water.

$$\text{Li}_2\text{CO}_3 < \text{Na}_2\text{CO}_3 < \text{K}_2\text{CO}_3 < \text{Rb}_2\text{CO}_3 < \text{Cs}_2\text{CO}_3$$

$\text{LiHCO}_3$  does not exist in solid form due to high polarizing power of  $\text{Li}^+$  and uncomparable size of  $\text{Li}^+$  cation and  $\text{HCO}_3^-$  anion.



- (iv) Bicarbonates are decomposed at relatively low temperature.



(v) Hydrolysis of carbonate

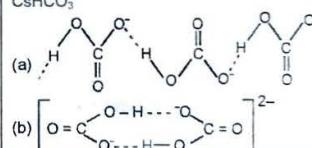


- (vi) The crystal structures of  $\text{NaHCO}_3$  and  $\text{KHCO}_3$  both show hydrogen bonding, but are different.

- (a) In  $\text{NaHCO}_3$ , the  $\text{HCO}_3^-$  ions are linked into an infinite chain.

- (b) In  $\text{KHCO}_3$ ,  $\text{RbHCO}_3$ ,  $\text{CsHCO}_3$ ,  $\text{HCO}_3^-$  forms a dimeric anion.

Solubility in water  $\text{NaHCO}_3 < \text{KHCO}_3 < \text{RbHCO}_3 < \text{CsHCO}_3$



8.

Complex ion formation

- (i) A metal shows complex formation only when it has following characteristics.

- (a) Small size, (b) High nuclear charge, (c) Presence of empty orbitals in order to accept electron pair from ligand (electron pair donor species).

- (ii) Due to small size only Lithium in alkali metals, forms a few complex ions. Rest all alkali metals do not possess the tendency to form complex ion.

$\text{Be}^{2+}$  on account of smaller size forms many complexes such as  $[\text{BeF}_4]^-$ ,  $[\text{BeF}_6]^{4-}$ . Chlorophyll contains  $\text{Mg}^{2+}$  [Photosynthetic pigment in plants] (C.No. = 4)



9.

Reaction with acids

Reacts vigorously with acids



The alkaline earth metals readily react with acids liberating dihydrogen.



10.

Formation of amalgams

- (i) Alkali metals get dissolved in mercury to form amalgams with evolution of heat and the amalgamation is highly exothermic.
- (ii) Alkali metals form alloys themselves as well as with other metals.

Alkaline earth metals get dissolved in mercury to form amalgams with evolution of heat and the amalgamation is highly exothermic.

11.

Sulphates

- (i) All these form sulphates of type  $\text{M}_2\text{SO}_4$ .

- (ii) Except  $\text{Li}_2\text{SO}_4$  rest all are soluble in water.

Thermal stability /solubility in water

$$\text{Li}_2\text{SO}_4 < \text{Na}_2\text{SO}_4 < \text{K}_2\text{SO}_4 < \text{Rb}_2\text{SO}_4 < \text{Cs}_2\text{SO}_4$$

- (iii) These sulphates on fusing with carbon form sulphides.



(i)  $\text{MSc}_4$  type sulphates are formed.

(ii) The solubility of sulphates decreases on moving down the group. The sulphates of the alkaline earth metals are all white solids and stable to heat.  $\text{BeSO}_4$  and  $\text{MgSO}_4$  are readily soluble in water; their solubility decreases from  $\text{CaSO}_4$  to  $\text{BaSO}_4$ . The greater hydration enthalpies of  $\text{Be}^{2+}$  and  $\text{Mg}^{2+}$  ions overcome the lattice enthalpy factor and therefore their sulphates are soluble in water.

Thermal stability



Solubility in water

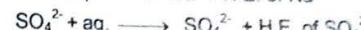
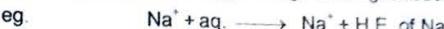


Sulphides	All metals react with S forming sulphides such as $\text{Na}_2\text{S}$ and $\text{Na}_2\text{Sn}$ ( $n = 2, 3, 4, 5$ or $6$ ). The polysulphide ions are made from zig-zag chains of sulphur atoms.
Nitrate	(i) Nitrates of both are soluble in water and decompose on heating. (ii) $\text{LiNO}_3$ decomposes to give $\text{NO}_2$ and $\text{O}_2$ and rest all give nitrites and oxygen. $2\text{MNO}_3 \rightarrow 2\text{MNO}_2 + \text{O}_2$ (except Li) $4\text{LiNO}_3 \rightarrow 2\text{Li}_2\text{O} + 4\text{NO}_2 + \text{O}_2$ $2\text{NaNO}_3 \xrightarrow[500^\circ\text{C}]{\text{M}} 2\text{NaNO}_2 + \text{O}_2$ $2\text{NaNO}_3 \xrightarrow[800^\circ\text{C}]{\text{Na}} \text{Na}_2\text{O} + \text{N}_2 + \text{O}_2$ $2\text{NaNO}_3 \xrightarrow{\text{Na}} \text{Na}_2\text{O} + \text{N}_2 + \text{O}_2$
Nitride	$\text{Li}_3\text{N} + 3\text{H}_2\text{O} \rightarrow 3\text{LiOH} + \text{NH}_3 \uparrow$
Carbide	When Li is heated with carbon, an ionic carbide $\text{Li}_2\text{C}_2$ is formed. $2\text{Li} + 2\text{C} \rightarrow \text{Li}_2\text{C}_2$ Other metals do not react with carbon directly but form carbides when heated with ethyne, or when ethyne is passed through a solution of metal in liquid ammonia. $\text{Na} + \text{C}_2\text{H}_2 \rightarrow \text{NaH} + \text{C}_2 \quad [\text{C}\equiv\text{C}]^- \quad [\text{C}\equiv\text{C}]^+$ $\text{Na}_2\text{C}_2 + 2\text{H}_2\text{O} \rightarrow \text{NaOH} + \text{C}_2\text{H}_2$  The binary compounds of carbon with other elements (less electronegative or of similar electronegativity) are called carbides. They are classified into following 3 categories : (i) Ionic (ii) Covalent (iii) Interstitial (or metallic) (i) <b>Ionic carbides</b> (or salt like carbides) : Generally formed by the most electropositive elements such as alkali and alkaline earth metals and aluminium (Boron is exception). Based on the product obtained on hydrolysis, they are further sub-classified into three types. (a) <b>Methanides</b> : These give $\text{CH}_4$ on reaction with $\text{H}_2\text{O}$ . $\text{Al}_4\text{C}_3 + 12\text{H}_2\text{O} \rightarrow 4\text{Al}(\text{OH})_3 + 3\text{CH}_4$ ; $\text{Be}_2\text{C} + 4\text{H}_2\text{O} \rightarrow 2\text{Be}(\text{OH})_2 + \text{CH}_4$ These carbides contain $\text{C}^{4-}$ ions in their constitution. (b) <b>Acetylides</b> : These give $\text{C}_2\text{H}_2$ on reaction with $\text{H}_2\text{O}$ . $\text{CaC}_2 + 2\text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2 + \text{C}_2\text{H}_2$ $\text{Al}_2(\text{C}_2)_3 + 6\text{H}_2\text{O} \rightarrow 2\text{Al}(\text{OH})_3 + 3\text{C}_2\text{H}_2$ $\text{SrC}_2 + 2\text{H}_2\text{O} \rightarrow \text{Sr}(\text{OH})_2 + \text{C}_2\text{H}_2$ Such compounds contain $\text{C}_2^{2-}$ ions. (c) <b>Allylides</b> : These give 1-propyne on reaction with $\text{H}_2\text{O}$ . $\text{Mg}_2\text{C}_3 + 4\text{H}_2\text{O} \rightarrow 2\text{Mg}(\text{OH})_2 + \text{CH}_3-\text{C}\equiv\text{CH}$ Such compounds contain $\text{C}_3^{4-}$ [ $\text{C}-\text{C}\equiv\text{C}$ ] <sup>4-</sup> ions. <b>Covalent carbides</b> Molecules like $\text{SiC}$ and $\text{B}_4\text{C}$ are also examples of covalent carbides.  <b>Interstitial or metallic carbides</b> Such carbides are formed by transition metals and some of the lanthanides and actinides. Interstitial carbides retain many of the properties of metals. They conduct electricity by metallic conduction and have properties of metals (a lustre like a metal). In these compounds carbon atoms occupy octahedral holes in the closed packed metal lattice. These are generally very hard and have very high melting point (e.g. $\text{WC}$ ). Carbides of Cr, Mn, Fe, Co and Ni are hydrolysed by water or dilute acids.

**Lattice Energy:** Energy change when one mole of crystalline lattice is formed from gaseous ions



**Hydration Energy:** It is the energy change when gaseous ions form aqueous ions.



### Solutions in liquid $\text{NH}_3$

Alkali metals dissolve in liquid ammonia (high conc. 3 M) and give blue solution which is conducting, reducing and paramagnetic in nature.

#### Reason

On dissolving Metal in  $\text{NH}_3$

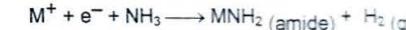


The blue colour is due to  $\text{Ammoniated electron}$

The paramagnetic nature is due to  $\text{Ammoniated electron}$

The conducting nature is due to  $\text{Ammoniated M}^+ + \text{Ammoniated electron}$

On standing the colour fades due to formation of amide After liberating hydrogen.



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

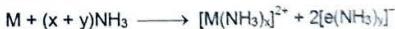


Solutions are of much lower density than the pure solvent, i.e., they occupy greater volume than that expected from the sum of the volumes of metal and solvent

Peroxide and superoxides of Na & K are widely used as Oxidising agent and air purifiers in space capsules, submarines and breathing mask.

### Alkaline metal in liq. $\text{NH}_3$

Like alkali metals, the alkaline earth metals dissolve in liquid ammonia to give deep blue black solutions forming ammoniated ions



From these solutions, the hexa-ammoniates  $[\text{M}(\text{NH}_3)_6]^{2+}$  can be recovered.

### Uses of alkali metal

(1) Lithium metal is used to make useful alloys,

- with lead to make 'white metal' bearings for motor engines.
- with aluminium to make aircraft parts.
- with magnesium to make armour plates.

(2) It is used in thermonuclear reactions.

(3) Lithium is also used to make electrochemical cells.

(4) Sodium is used to make a  $\text{Na}/\text{Pb}$  alloy needed to make  $\text{PbEt}_4$  and  $\text{PbMe}_4$ . These organolead compounds were earlier used as anti-knock additives to petrol, but nowadays vehicles use lead-free petrol.

(5) Liquid sodium metal is used as a coolant in fast breeder nuclear reactors.

(6) Potassium chloride is used as a fertilizer.

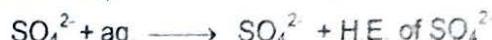
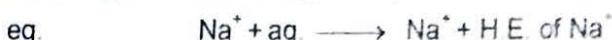
(7) Potassium hydroxide is used in the manufacture of soft soap. It is also used as an excellent absorbent of carbon dioxide

(8) Caesium is used in devising photoelectric cells.

**Lattice Energy:** Energy change when one mole of crystalline lattice is formed from gaseous ions



**Hydration Energy:** It is the energy change when gaseous ions form aqueous ions

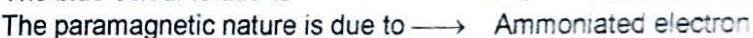
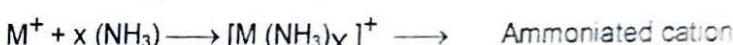
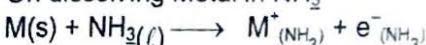


### Solutions in liquid NH<sub>3</sub>

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



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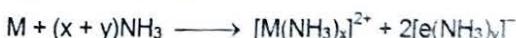


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### Uses of alkali metal

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(8) Caesium is used in devising photoelectric cells.



**Uses of alkaline metal**

- (1) Beryllium is used in the manufacture of alloys.
- (2) Copper-beryllium alloys are used in the preparation of high strength springs.
- (3) Metallic beryllium is used for making windows of X-ray tubes.
- (4) Metallic beryllium is used for making windows of X-ray tubes.
- (5) Magnesium forms alloys with aluminium, zinc, manganese and tin. Magnesium-aluminium alloys being light in mass are used in air-craft construction.
- (6) A suspension of magnesium hydroxide in water (called *milk of magnesia*) is used as antacid in medicine.
- (7) Magnesium carbonate is an ingredient of toothpaste.
- (8) Calcium is used in the extraction of metals from oxides which are difficult to reduce with carbon.
- (9) Calcium and barium metals, owing to their reactivity with oxygen and nitrogen at elevated temperatures, have often been used to remove air from vacuum tubes.
- (10) Radium salts are used in radiotherapy, for example, in the treatment of cancer.

**Biological Importance Of Sodium And Potassium:**

- Sodium ions are found primarily on the outside of cells, being located in blood plasma and in the interstitial fluid which surrounds the cells. These ions participate in the transmission of nerve signals, in regulating the flow of water across cell membranes and in the transport of sugars and amino acids into cells. Sodium and potassium, although so similar chemically, differ quantitatively in their ability to penetrate cell membranes, in their transport mechanisms and in their efficiency to activate enzymes. Thus, potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in the oxidation of glucose to produce ATP and, with sodium, are responsible for the transmission of nerve signals.

A typical 70 kg man contains about 90 g of Na and 170 g of K compared with only 5 g of iron and 0.06 g of copper.

**Biological Importance of Magnesium and Calcium :**

- Monovalent sodium and potassium ions and divalent magnesium and calcium ions are found in large proportions in **biological fluids**. These ions perform important **biological functions** such as maintenance of ion balance and nerve impulse conduction.
- All enzymes that utilise ATP in phosphate transfer require magnesium as the cofactor. The main pigment for the absorption of light in plants is chlorophyll which contains magnesium. About 99 % of body calcium is present in bones and teeth. It also plays important roles in neuromuscular function, interneuronal transmission, cell membrane integrity and blood coagulation.
- The calcium concentration in plasma is regulated at about  $100 \text{ mg L}^{-1}$ . It is maintained by two hormones : calcitonin and parathyroid hormone. Do you know that bone is not an inert and unchanging substance but is continuously being solubilised and redeposited to the extent of 400 mg per day in man? All this calcium passes through the plasma.

An adult body contains about 25 g of Mg and 1200 g of Ca compared with only 5 g of iron and 0.06 g of copper. The daily requirement in the human body has been estimated to be 200–300 mg.

**ANOMALOUS PROPERTIES OF LITHIUM**

The anomalous behavior of lithium is due to the :

- (i) Exceptionally small size of its atom and ion,
- (ii) High polarising power (i.e., charge/ radius ratio).

As a result, there is increased covalent character of lithium compound which is responsible for their solubility in organic solvent. Further, lithium shows diagonal relationship to magnesium.

S.No.	Property	Li
1.	Hardness	Li is much harder.
2.	M P and B.P	Higher M P and B.P
3.	Reactivity	Less reactive
4.	Reducing agent	Strong
5.	Combustion in air	Li form monoxide ( $\text{Li}_2\text{O}$ ) and nitride ( $\text{Li}_3\text{N}$ ), not for other.
6.	Hydration of ion	Favored for $\text{Li}^{+}$ ; not for other. $\text{Li}^{+}$ has maximum degree of hydration for this reason. Lithium salts are mostly hydrated. E.g. $\text{LiCl} \cdot 2\text{H}_2\text{O}$

7.	Hydrogen Carbonate	Li is not obtained in the solid form while all other elements form solid hydrogen carbonates.
8.	Ethyne	Favored for $\text{Li}^{+}$ , not for other
9.	Lithium nitrate	$4\text{LiNO} \rightarrow 2\text{Li}_2\text{O} + 4\text{NO}_2 + \text{O}_2$ Lithium Oxide Where as other alkali metal nitrates decompose to give the corresponding nitrite. $2\text{NaNO}_3 \rightarrow 2\text{NaNO}_2 + \text{O}_2$ Sodium nitrite
10.	$\text{LiF}$ and $\text{Li}_2\text{O}$	These are much less soluble in water. Solubility in water is less than the corresponding compounds of other alkali metal
11.	Carbide	Li reacts directly with carbon to form anionic carbide
12.	Hydroxide	Lithium hydroxide is less basic. $\text{Li}_2\text{CO}_3$ , $\text{LiNO}_3$ and $\text{LiOH}$ all form the oxides on gentle heating.
13.	Carbonate	Less stable.
14.	Nitrite	Less stable.
15.	Bicarbonate	Lithium forms a bicarbonates in solution it does not form a solid bicarbonate. Where as the other all forms stable solid bicarbonates
16.	Complex ion formation	Lithium has a great tendency to form complexes not for other. Due to small size of Lithium.
17.	Reaction with $\text{NH}_3$	Li when heated in $\text{NH}_3$ imide ( $\text{Li}_2\text{NH}$ ) while other alkali metals form amides ( $\text{MNH}_2$ )

**Points of Similarities between Lithium and Magnesium**

The similarity between lithium and magnesium is particularly striking and arises because of the similar size: atomic radii, Li = 152 pm, Mg = 160 pm; ionic radii,  $\text{Li}^{+} = 76 \text{ pm}$ ,  $\text{Mg}^{2+} = 72 \text{ pm}$ . The main points of similarity are :

S.No.	Properties	Li and Mg
1.	Hardness	Li and Mg are much harder
2.	Density	These are lighter than other elements in the respective group
3.	Reaction with water	Both react slowly with water
4.	Solubility of hydroxide and oxide	Less soluble and their hydroxides decompose in acid on heating.
5.	Reaction with $\text{N}_2$	By direct combination with nitrogen both form a nitride $\text{Li}_3\text{N}$ and $\text{Mg}_3\text{N}_2$ .
6.	Oxides	The oxides $\text{Li}_2\text{O}$ and $\text{MgO}$ do not combine with excess oxygen to give any superoxide
7.	Carbonates	Carbonates of both decompose easily on heating to form the oxides and $\text{CO}_2$ . Solid hydrogen carbonates are not formed by Li and Mg.
8.	Solubility of halides in ethanol	Both $\text{LiCl}$ and $\text{MgCl}_2$ are soluble.
9.	Hydration of ion	Both $\text{LiCl}$ and $\text{MgCl}_2$ are deliquescent and crystallise from aqueous solution as hydrates, $\text{LiCl} \cdot 2\text{H}_2\text{O}$ and $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$

**Anamolous Behaviour of Beryllium**

The properties of beryllium the first member of the alkaline earth metal, differ from the rest of the member. Its is mainly because of

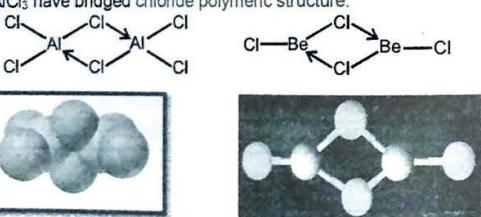
- (i) Its small size and high polarizing power.
- (ii) Relatively high electro negativity and ionization energy as compared to other members.
- (iii) Absence of vacant d-orbitals in its valence shell.

Some important points of difference between beryllium and other members (especially magnesium) given below.

S.No.	Properties	
1.	Hardness	Be is harder than other members of its group
2.	Density	Be is lighter than Mg
3.	M.P. and B.P.	Higher than other members of its group
4.	Reaction with water	Be does not react with water while Mg reacts with boiling water.
5.	Nature of oxides	BeO is amphoteric while MgO is weakly basic.
6.	Nature of compounds	Be forms covalent compounds whereas other members form ionic compounds.
7.	Carbide	Beryllium carbide reacts with water to give methane whereas carbides of other alkaline earth metals gives acetylene gas. $\text{Be}_2\text{C} + 4\text{H}_2\text{O} \rightarrow 2\text{Be}(\text{OH})_2 + \text{CH}_4$ $\text{Mg}_2\text{C} + 2\text{H}_2\text{O} \rightarrow \text{Mg}(\text{OH})_2 + \text{C}_2\text{H}_2$ $\text{Ca}_2\text{C} + 2\text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2 + \text{C}_2\text{H}_2$
8.	Hydride	The beryllium hydride is electron deficient and polymeric, with multi center bonding like aluminium hydride.
9.	Co-ordination number	Beryllium does not exhibit coordination number more than four as it has four orbitals in the valence shell. The other members of this group has coordination number 6.
10.	Reaction with Alkali	Be dissolves in alkalies with evolution of hydrogen $\text{Be} + 2\text{NaOH} + 2\text{H}_2\text{O} \rightarrow \text{Na}_2\text{BeO}_2 \cdot 2\text{H}_2\text{O} + \text{H}_2$ (sodium beryllate)  Other alkaline earth metals don't react with alkalies.

**Resemblance of Beryllium with Aluminium (Diagonal relationship)**

The following points illustrate the anomalous behaviour of Be and its resemblance with Al.

S.No.	Properties	Be and Al
1.	Nature of compounds	Unlike groups-2 elements but like aluminium, beryllium forms covalent compounds.
2.	Nature of hydroxide	The hydroxides of Be, $[\text{Be}(\text{OH})_2]$ and aluminium $[\text{Al}(\text{OH})_3]$ are amphoteric in nature, whereas those of other elements of group - 2 are basic in nature.
3.	Nature of oxide	The oxides of both Be and Al i.e. BeO and $\text{Al}_2\text{O}_3$ are high melting insoluble solids.
4.	Polymeric structure	$\text{BeCl}_2$ and $\text{AlCl}_3$ have bridged chloride polymeric structure. 
5.	Salts	The salts of beryllium as well as aluminium are extensively hydrolysed.
6.	Carbides	Carbides of both the metal reacts with water liberating methane gas. $\text{Be}_2\text{C} + 4\text{H}_2\text{O} \rightarrow 2\text{Be}(\text{OH})_2 + \text{CH}_4$ $\text{Al}_4\text{C}_3 + 12\text{H}_2\text{O} \rightarrow 4\text{Al}(\text{OH})_3 + 3\text{CH}_4$
7.	Oxides and hydroxides	The oxides and hydroxides of both Be and Al are amphoteric and dissolve in sodium hydroxide as well as in hydrochloric acid. $\text{BeO} + 2\text{HCl} \rightarrow \text{BeCl}_2 + \text{H}_2\text{O}$ $\text{BeO} + 2\text{NaOH} \rightarrow \text{Na}_2\text{BeO}_2 + \text{H}_2\text{O}$ $\text{Al}_2\text{O}_3 + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + \text{H}_2\text{O}$ $\text{Al}_2\text{O}_3 + 2\text{NaOH} \rightarrow 2\text{NaAlO}_2 + \text{H}_2\text{O}$
8.	Reaction with acids	Like Al, Be is not readily attacked by acids because of the presence of an oxide film.

**Section (E) : Oxides, Peroxides, Super Oxides, Hydroxides****GROUP -I & II OXIDES****1. Sodium Oxide ( $\text{Na}_2\text{O}$ )**  
**Preparation:**

Name of compound	Name and Brief about the process	Related chemical reaction
Sodium Oxide ( $\text{Na}_2\text{O}$ )	(1) By burning sodium at $180^\circ\text{C}$ in a limited supply of air or oxygen and distilling off the excess of sodium in vacuum.	$2\text{Na} + \frac{1}{2}\text{O}_2 \xrightarrow{-180^\circ\text{C}} \text{Na}_2\text{O}$
	(2) By heating sodium peroxide, nitrate or nitrite with sodium.	$\text{Na}_2\text{O}_2 + 2\text{Na} \longrightarrow 2\text{Na}_2\text{O}$ $2\text{NaNO}_3 + 10\text{Na} \longrightarrow 6\text{Na}_2\text{O} + \text{N}_2$ $2\text{NaNO}_2 + 8\text{Na} \longrightarrow 4\text{Na}_2\text{O} + \text{N}_2$ $3\text{Na}_2\text{N}_2 + \text{NaNO}_2 \longrightarrow 2\text{Na}_2\text{O} + 5\text{N}_2$
	(3) Sodium oxide is formed when the mixture of sodium azide and sodium nitrite is heated	

**Chemical Properties:**

- (1) It is white amorphous substance.
- (2) It dissolve violently in water, yielding caustic soda ( $\text{NaOH}$ ) and evolving a large amount of heat.  
 $\text{Na}_2\text{O} + \text{H}_2\text{O} \longrightarrow 2\text{NaOH}$

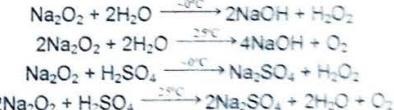
**Uses :** It is used as dehydrating and polymerising agent in organic chemistry

**2. Sodium Peroxide ( $\text{Na}_2\text{O}_2$ )****Preparation**

Name of compound	Name and Brief about the process	Related chemical reaction
Sodium Peroxides ( $\text{Na}_2\text{O}_2$ )	(1) By heating the metal in excess of air or oxygen at $300^\circ\text{C}$ , which is free from moisture and $\text{CO}_2$ .	$2\text{Na} + \text{O}_2 \xrightarrow{300^\circ\text{C}} \text{Na}_2\text{O}_2$
	(2) Industrial method : It is a two stage reaction in the presence of excess air.	$2\text{Na} + \text{O}_2 \longrightarrow \text{Na}_2\text{O}$ $\text{Na}_2\text{O} + \text{O}_2 \longrightarrow \text{Na}_2\text{O}_2$

**Properties:**

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



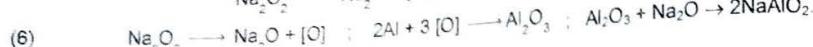
- (3) It reacts with  $\text{CO}_2$ , giving sodium carbonate and oxygen and hence its use for purifying air in a confined space e.g. submarine, ill-ventilated room.



- (4) It is an oxidising agent and oxidises charcoal,  $\text{CO}$ ,  $\text{NH}_3$ ,  $\text{SO}_2$ .



(5) Sulphides are oxidised to corresponding sulphates

Uses :

(1) For preparing  $\text{H}_2\text{O}_2$ ,  $\text{O}_2$

(2) Oxygenating the air in submarines

(3) Oxidising agent in the laboratory

Oxides of Potassium	$\text{K}_2\text{O}$	$\text{K}_2\text{O}_2$	$\text{K}_2\text{O}_3^*$	$\text{KO}_2$	$\text{KO}_3$
Colours	White	White	Red	Bright Yellow	Orange Red Solid

## 3. Potassium sesquioxide (need not memorize).

## Preparation:

Name of compound	Name and Brief about the process	Related chemical reaction
$\text{K}_2\text{O}$ (Potassium oxide)	By heating potassium nitrate with potassium.	$2\text{KNO}_3 + 10\text{K} \xrightarrow{\text{heating}} 6\text{K}_2\text{O} + \text{N}_2$ $\text{K}_2\text{O} \xrightarrow{\text{heating}} \text{K}_2\text{O}$ (White) (Yellow) $\text{K}_2\text{O} + \text{H}_2\text{O} \longrightarrow 2\text{KOH}$

Name of compound	Name and Brief about the process	Related chemical reaction
$\text{K}_2\text{O}_2$ (Potassium peroxide)	By burning potassium at $300^\circ\text{C}$ in a limited supply of air or oxygen.	$2\text{K} + \text{O}_2 \xrightarrow[\text{air at } 300^\circ\text{C}]{\text{Controlled}} \text{K}_2\text{O}_2$

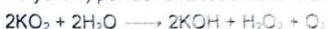
Name of compound	Name and Brief about the process	Related chemical reaction
$\text{KO}_2$ (Potassium superoxide)	Pas (i) Passage of $\text{O}_2$ through a blue solution of K in liquid $\text{NH}_3$ yields oxides $\text{K}_2\text{O}_2$ (white), $\text{K}_2\text{O}_3$ (red) and $\text{KO}_2$ (deep yellow) 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. (ii) It is prepared by burning potassium in excess of oxygen free from moisture	$\text{K} \text{ in liq. } \text{NH}_3 \xrightarrow{\text{O}_2} \text{K}_2\text{O}_2 \longrightarrow \text{K}_2\text{O}_3 \longrightarrow \text{KO}_2$ white red yellow $2\text{KO}_2 + 2\text{H}_2\text{O} \xrightarrow{-\text{O}_2} 2\text{KOH} + \text{H}_2\text{O}_2 + \text{O}_2$ $\text{K} + \text{O}_2 \longrightarrow \text{KO}_2$

Name of compound	Name and Brief about the process	Related chemical reaction
$\text{K}_2\text{O}_3$ (Potassium sesquioxide)	It is obtained when oxygen is passed through liquid ammonia containing potassium	$4\text{K} \text{ (dissolved in liquid } \text{NH}_3) \xrightarrow{\text{O}_2} 2\text{K}_2\text{O}_3$

Name of compound	Name and Brief About the process	Related chemical reaction
$\text{KO}_3$ (Potassium ozonide)	From KOH	$\text{KOH} + \text{O}_3 \text{ (ozonised oxygen)} \xrightarrow[(-70^\circ\text{C})]{(+40^\circ\text{C})} \text{KO}_3$ (Dry powdered) (orange solid)

Properties of Potassium superoxide ( $\text{KO}_2$ )

It is an orange coloured (chrome yellow) powder and reacts with water according to following reaction



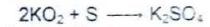
It reacts directly with CO and  $\text{CO}_2$



If more  $\text{CO}_2$  in presence of moisture is present, then



On heating with sulphur, it forms potassium sulphate



Uses : It is used as an oxidising agent and air purifier in space capsules, submarine and breathing mask as it produces  $\text{O}_2$  and removes  $\text{CO}_2$ .

4. Magnesium Oxide ( $\text{MgO}$ ):

Name of compound	Name and Brief about the process	Related chemical reaction
Magnesium Oxide ( $\text{MgO}$ )	It is also called magnesia and obtained by heating natural magnesite	$\text{MgCO}_3 \longrightarrow \text{MgO} + \text{CO}_2$

## Properties :

- It is white powder.
- It's m.p. is  $2850^\circ\text{C}$ . Hence used in manufacture of refractory bricks for furnaces. And it acts as basic flux and facilitates the removal of acidic impurities of Si, P and S from steel through slag formation.
- It is very slightly soluble in water imparting alkaline reaction.

5. Calcium Oxide ( $\text{CaO}$ ):

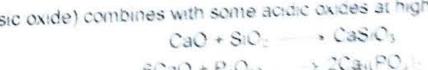
## Preparation

Name of compound	Name and Brief about the process	Related chemical 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 $1000^\circ\text{C}$ .	$\text{CaCO}_3 \xrightarrow{\text{heat}} \text{CaO} + \text{CO}_2$

\* The Carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion.

## Chemical Properties :

- It is white amorphous powder of m.p.  $2570^\circ\text{C}$ . On exposure to atmosphere, it absorbs moisture and carbon dioxide.
- It emits intense light (lime light), when heated in oxygen-hydrogen flame.
- It combines with limited amount of water to produce slaked lime. This process is called slaking of lime. Quick lime slaked with soda gives solid sodalime ( $\text{CaO}_2$ ). Being a basic oxide
- Soda lime (basic oxide) combines with some acidic oxides at high temperature



It combines with limited amount of water to produce slaked lime. This process is called slaking of lime.

Quick lime slaked with soda gives solid sodalime ( $\text{CaO}_2$ ). Being a basic oxide



Quick lime slaked with soda gives solid sodalime ( $\text{CaO}_2$ ). Being a basic oxide



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**Uses :**

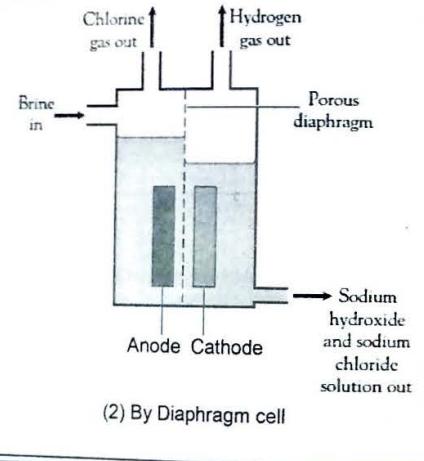
- It is an important primary material for manufacturing cement and is the cheapest form of alkali.
- It is used in the manufacture of sodium carbonate from caustic soda.
- It is employed in the purification of sugar and in the manufacture of dye stuffs.

**Magnesium Peroxide ( $MgO_2$ ) and Calcium Peroxide ( $CaO_2$ )**

These are obtained by passing  $H_2O_2$  in a suspension of  $Mg(OH)_2$  and  $Ca(OH)_2$ .

**Uses :**  $MgO_2$  is used as an antiseptic in tooth paste and as a bleaching agent.

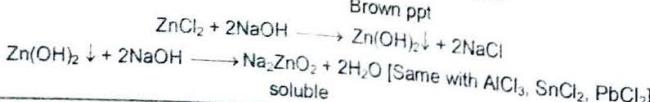
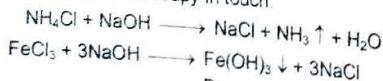
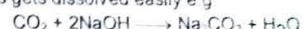
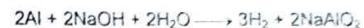
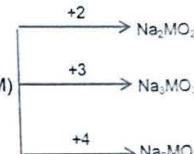
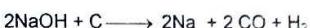
**HYDROXIDES****1. Sodium Hydroxides(Caustic Soda)  $NaOH$  (White) :****Preparation :**

Name of compound	Name and Brief about the process	Related chemical reaction
Sodium Hydroxides ( $NaOH$ )	(1) Electrolysis of Brine : Sodium hydroxide is prepared by the electrolysis of sodium chloride in Caster-Kellner cell. A brine solution is electrolysed using a mercury cathode and a carbon anode. Sodium metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode.  The amalgam is treated with water to give sodium hydroxide and hydrogen gas.	Cathode: $Na^+ + e^- \xrightarrow{Hg} Na\text{-amalgam}$  Anode: $Cl^- \longrightarrow \frac{1}{2} Cl_2 + e^-$  $2\text{Na-amalgam} + 2H_2O \longrightarrow 2NaOH + 2Hg + H_2$    (2) By Diaphragm cell
		$Na_2CO_3 + Ca(OH)_2 \longrightarrow 2NaOH + CaCO_3 \downarrow$ (suspension)  Since the $K_{sp}(CaCO_3) < K_{sp}(Ca(OH)_2)$ , the reaction shifts towards right.
	(3) Caustication of $Na_2CO_3$ (Gossage's method)	

**Properties:**

- Sodium hydroxide is a white, translucent solid. It melts at 591 K. It is readily soluble in water to give a strong alkaline solution. Crystals of sodium hydroxide are deliquescent. The sodium hydroxide solution at the surface reacts with the  $CO_2$  in the atmosphere to form  $Na_2CO_3$ .
- It is white crystalline, deliquescent, highly corrosive solid.
- It is stable towards heat.

It's aqueous solution alkaline in nature and soapy in touch.

**(6) Acidic and amphoteric oxides gets dissolved easily e.g.****(7) Aluminium and Zn metal gives  $H_2$  from  $NaOH$ .****(8) Several non metals such as P, S, Cl etc. yield a hydride instead of hydrogen e.g.****(9)  $NaOH$  is stable towards heat but reduced to metal when heated with carbon**

Above are general reactions of  $NaOH$  with metal oxides having metals Oxidation number +2, +3 & +4 respectively.

**Uses :** It is used in

- The manufacture of soap, paper, artificial silk and a number of chemicals.
- In petroleum refining.
- In the purification of bauxite.
- In the textile industries for mercerising cotton fabrics.
- For the preparation of pure fats and oils.
- As a laboratory reagent.

**2. Potassium Hydroxide ( $KOH$ ):****Preparation:**

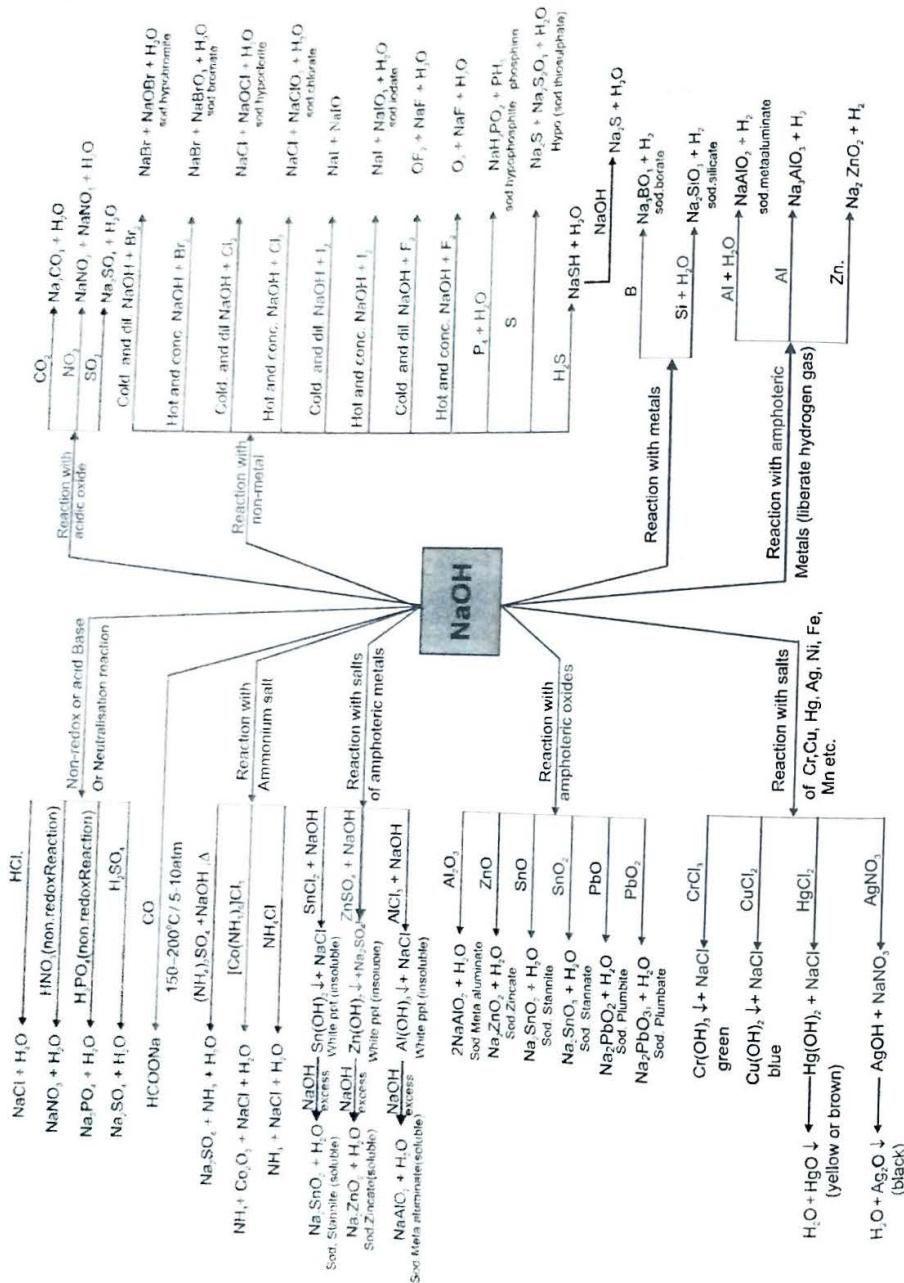
- It is prepared by electrolysis of  $KCl$  solution.
- $KOH$  resembles  $NaOH$  in all its reactions. However  $KOH$  is much more soluble in alcohol. This accounts for the use of alcoholic  $KOH$  in organic chemistry.
- $KOH$  is called caustic potash, because of their corrosive properties (for example on glass or on skin) and its aqueous solution is known as potash lye.



- It is used for the absorption of gases like  $CO_2$ ,  $SO_2$ , etc. It is used for making soft soaps.

**Properties:** Same as  $NaOH$ 

- It is stronger base compared to  $NaOH$ .
- Solubility in water is more compared to  $NaOH$ .
- In alcohol,  $NaOH$  is sparingly soluble but  $KOH$  is highly soluble.
- As a reagent  $KOH$  is less frequently used but in absorption of  $CO_2$ ,  $KOH$  is preferably used compared to  $NaOH$ . Because  $KHCO_3$  formed is soluble whereas  $NaHCO_3$  is insoluble and may therefore choke the tubes of apparatus used.



### 3. Magnesium Hydroxide ( $Mg(OH)_2$ ):

**Preparation:**

Name of compound	Name and Brief about the process	Related chemical reaction
Magnesium Hydroxide ( $\text{Mg(OH)}_2$ )	It can be prepared by adding caustic soda solution to a solution of Magnesium sulphate or chloride solution.	$\text{MgSO}_4 + 2\text{NaOH} \longrightarrow \text{Mg(OH)}_2 + \text{Na}_2\text{SO}_4$ $\text{MgCl}_2 + 2\text{NaOH} \longrightarrow \text{Mg(OH)}_2 + 2\text{NaCl}$ $\text{MgCl}_2 + \text{Ca(OH)}_2 \longrightarrow \text{Mg(OH)}_2 + \text{CaCl}_2$ $\text{MgO} + \text{H}_2\text{O} \longrightarrow \text{Mg(OH)}_2$

#### **Chemical Properties:**

- (1) It can be dried at temperature upto  $100^{\circ}\text{C}$  only otherwise it breaks into its oxide at higher temperature  
 $\text{Mg}(\text{OH})_2 \longrightarrow \text{MgO} + \text{H}_2\text{O}$

(2) It is slightly soluble in water imparting alkalinity.  
(3) 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  $\text{NH}_3 \cdot \text{H}_2\text{O}$  in presence of excess of  $\text{NH}_4\text{Cl}$ .

**Uses :** A suspension of  $Mg(OH)_2$  in water is used in medicine as an antacid (An antacid is substance which neutralizes stomach acidity) under the name milk of magnesia.

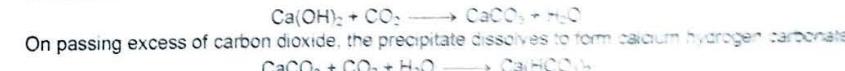
#### **4. Calcium Hydroxide ( $\text{Ca(OH)}_2$ ) (White Powder):**

## **Preparation :**

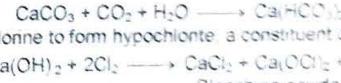
Name of compound	Name and Brief about the process	Related chemical reaction
Calcium Hydroxide ( $\text{Ca}(\text{OH})_2$ )	By spraying water on quicklime.	$\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca}(\text{OH})_2$

## Properties:

- (1) It is a white amorphous powder.  
(2) It is sparingly soluble in water.  
(3) Its solubility in hot water is less than that of cold water. Hence solubility decreases with increase in temperature.  
(4) The aqueous solution is known as lime water and a suspension of slaked lime in water is known as milk of lime.  
(5) When carbon dioxide is passed through lime water it turns milky due to the formation of calcium carbonate.



Milk of lime reacts with chlorine to form hypochlorite, a constituent of bleaching powder.



### Uses:

- It is used in the preparation of mortar, a building material.
  - It is used in white wash due to its disinfectant nature.
  - It is used in glass making, in tanning industry, for the preparation of bleaching powder and for purification of sugar.

## Section (F) : Carbonates, Bicarbonates

CARBONATES1. Sodium Carbonate (Washing soda)  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  (White Solid) :

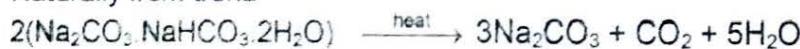
Preparation:

1.

Name of compound	Name and Brief about the process	Related chemical reaction
Sodium Carbonate (Washing soda) $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	(1) Leblanc Process	$\text{NaCl} + \text{H}_2\text{SO}_4(\text{conc.}) \xrightarrow{\text{mild heating}} \text{NaHSO}_4 + \text{HCl}$ $\text{NaCl} + \text{NaHSO}_4 \xrightarrow{\substack{\text{Strongly} \\ \text{heated}}} \text{Na}_2\text{SO}_4 + \text{HCl}$ (Salt Cake) $\text{Na}_2\text{SO}_4 + 4\text{C} \longrightarrow \text{Na}_2\text{S} + 4\text{CO} \uparrow$ $\text{Na}_2\text{S} + \text{CaCO}_3 \longrightarrow \text{Na}_2\text{CO}_3 + \text{CaS}$
	(2) Solvay Process Step-1 (In ammonia absorber) (i) Saturation of brine with ammonia and $\text{CO}_2$ (ii) Ammoniated brine is filtered to remove calcium and magnesium impurities as their insoluble carbonates.	$2\text{NH}_3 + \text{CO}_2 + \text{H}_2\text{O} \longrightarrow (\text{NH}_4)_2\text{CO}_3$ $\text{CaCl}_2 + (\text{NH}_4)_2\text{CO}_3 \longrightarrow \text{CaCO}_3 \downarrow + 2\text{NH}_4\text{Cl}$ $\text{MgCl}_2 + (\text{NH}_4)_2\text{CO}_3 \longrightarrow \text{MgCO}_3 \downarrow + 2\text{NH}_4\text{Cl}$
	Step-2 (In carbonation tower) : (i) Formation of insoluble $\text{NaHCO}_3$ (ii) Reaction is exothermic and hence there is a cooling arrangement. (iii) $\text{NaHCO}_3$ is insoluble in cold brine solution because of the common ion effect. It is separated by filtration and the filtered is used for recovering $\text{NH}_3$ & $\text{CO}_2$ .	$\text{NH}_3 + \text{CO}_2 + \text{H}_2\text{O} \longrightarrow \text{NH}_4\text{HCO}_3$ ; $\text{NH}_4\text{HCO}_3 + \text{NaCl} \xrightarrow{30^\circ\text{C}} \text{NaHCO}_3 + \text{NH}_4\text{Cl}$
	Step-3 (Calcination to get sodium carbonate) :	$2\text{NaHCO}_3 \xrightarrow{150^\circ\text{C}} \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$
	Step - 4 (In recovery tower) : Recovery of ammonia and carbon dioxide. $\text{CaCl}_2$ is obtained as by product	$\text{NH}_4\text{HCO}_3 \xrightarrow{\Delta/\text{Steam}} \text{NH}_3 + \text{CO}_2 + \text{H}_2\text{O}$ $2\text{NH}_4\text{Cl} + \text{Ca}(\text{OH})_2 \xrightarrow{\Delta/\text{Steam}} 2\text{NH}_3 + 2\text{H}_2\text{O} + \text{CaCl}_2$

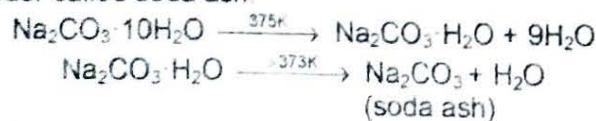
\* advantage is taken of low solubility of  $\text{NaHCO}_3$ , it gets precipitated in the reaction of  $\text{NaCl} + \text{NH}_4\text{HCO}_3$ .

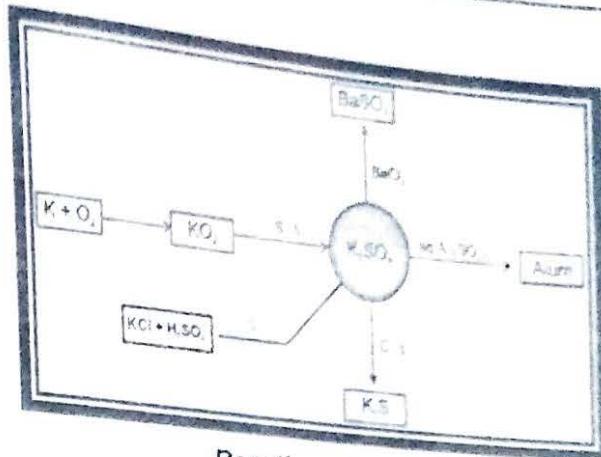
## 2. Naturally from trona



## Properties

- (1) Anhydrous  $\text{Na}_2\text{CO}_3$  is called as soda ash, which does not decompose on heating but melts at  $852^\circ\text{C}$ .  
(2) Sodium carbonate is a white crystalline solid which exists as a decahydrate,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . This is also called washing soda. It is readily soluble in water. On heating, the decahydrate loses its water of crystallisation to form monohydrate. Above  $373\text{K}$ , the monohydrate becomes completely anhydrous and changes to a white powder called soda ash.





## 2. Magnesium Sulphate ( $MgSO_4$ ): Reactions Charts

It occurs in nature as minerals kieserite ( $MgSO_4 \cdot H_2O$ ), epsom salt ( $MgSO_4 \cdot 7H_2O$ ) and kainite ( $KCl \cdot MgSO_4 \cdot 3H_2O$ ).

### Preparation:

- (1) It is obtained by dissolving kieserite,  $MgSO_4 \cdot H_2O$  in boiling water and then crystallising the solution as a hepta hydrate, i.e.  $MgSO_4 \cdot 7H_2O$ . It is called as Epsom salt.
  - (2) It is also obtained by dissolving magnesite in hot dil.  $H_2SO_4$ .
  - (3) By dissolving dolomite ( $CaCO_3 \cdot MgCO_3$ ) in hot dil.  $H_2SO_4$  and removing the insoluble  $CaSO_4$ , by filtration.
  - (4) It is isomorphous with  $FeSO_4 \cdot 7H_2O$ ,  $ZnSO_4 \cdot 7H_2O$
- $$MgCO_3 + H_2SO_4 \longrightarrow MgSO_4 + H_2O + CO_2$$
- $$CaCO_3 \cdot MgCO_3 \text{ (dolomite)} + 2H_2SO_4 \longrightarrow MgSO_4 + CaSO_4 + 2CO_2 + 2H_2O$$

### Chemical Properties:

#### Heating effect:

- (1) When heated to  $150^\circ C$ , it changes to monohydrate. On further heating, it becomes anhydrous at  $200^\circ C$ . On strong heating, it decomposes into  $MgO$ .



- (2) Magnesium sulphate when heated with lamp black at  $800^\circ C$  produces  $SO_2$  and  $CO_2$  gases



- (3) It forms double salts with alkali metal sulphates, e.g.,  $K_2SO_4 \cdot MgSO_4 \cdot 6H_2O$

## 3. Calcium Sulphate (Plaster of paris) $CaSO_4 \cdot \frac{1}{2} H_2O$

It occurs as anhydrite  $CaSO_4$ , hemihydrate  $CaSO_4 \cdot \frac{1}{2} H_2O$  and as the dihydrate  $CaSO_4 \cdot 2H_2O$  gypsum alabaster or satin-spar.

### Preparation:

- (1) It is a hemihydrate of calcium sulphate. It is obtained when gypsum,  $CaSO_4 \cdot 2H_2O$  is heated to 393 K.  

$$2(CaSO_4 \cdot 2H_2O) \xrightarrow{393K} 2(CaSO_4) \cdot H_2O + 3H_2O$$

Above 393 K, no water of crystallisation is left and anhydrous calcium sulphate,  $CaSO_4$ , is formed. This is known as 'dead burnt plaster'.

It has a remarkable property of setting with water. On mixing with an adequate quantity of water, it forms a plastic mass that gets into a hard solid in 5 to 15 minutes.
- (2) It can be prepared by reacting any calcium salt with either sulphuric acid or a soluble sulphate.

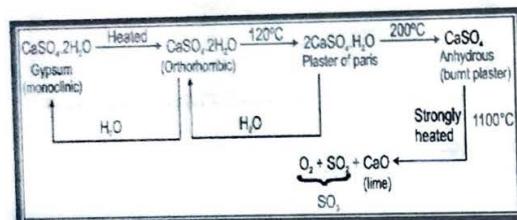


### Properties:

It is a white crystalline solid. It is sparingly soluble in water and solubility decreases as temperature increases.

It dissolves in dilute acids. It also dissolves in ammonium sulphate due to the formation of double sulphate,  $(NH_4)_2SO_4 \cdot CaSO_4 \cdot H_2O$

The setting process is **exothermic**. The process of setting takes place in stages. In the first stage, there is conversion of Plaster of Paris into orthorhombic form of gypsum (setting step) and in the second stage orthorhombic form changes into monoclinic form (hardening step). The setting of Plaster of Paris may be **catalysed by sodium chloride** while it is retarded by borax or alum. Addition of alum to Plaster of Paris makes the setting very hard. The mixture is known as **Keene's cement**.



**Dead plaster** has no setting property as it takes up water only very slowly.

A suspension of gypsum when saturated with ammonia and carbon dioxide forms ammonium sulphate, a nitrogenous fertilizer.



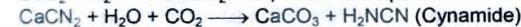
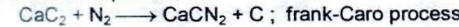
When strongly heated with carbon, it forms calcium sulphide.



**Uses:** For preparing blackboard chalk.  
In anhydrous form as drying agent.

#### Fertilizer

1. **Cyanamide:** It is an organic compound with the formula  $\text{CN}_2\text{H}_2$ . This white solid is widely used in agriculture and the production of pharmaceuticals and other organic compounds. Cyanamide is produced by hydrolysis of calcium cyanamide, which in turn is prepared from calcium carbide via the frank-Caro process.



The main reaction exhibited by cyanamide involves additions of compounds containing an acidic proton. Water, hydrogen sulfide, and hydrogen selenide react with cyanamide to give urea, thiourea, and selenourea, respectively :



2. **Fluorapatite:** It is a phosphate mineral with the formula  $\text{Ca}_5(\text{PO}_4)_3$ .

#### Cement

Cement is a product obtained by combining a material rich in lime,  $\text{CaO}$  with other material such as clay which contains silica,  $\text{SiO}_2$  along with the oxides of aluminium, iron and magnesium.

The raw materials for the manufacture of cement are limestone and clay. When clay and lime are strongly heated together they fuse and react to form cement clinker. This clinker is mixed with 2-3% by weight of gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ) to form cement. Thus important ingredients present in Portland cement are dicalcium silicate ( $\text{Ca}_2\text{SiO}_4$ ) 26%, tricalcium silicate ( $\text{Ca}_3\text{SiO}_5$ ) 51% and tricalcium aluminate ( $\text{Ca}_3\text{Al}_2\text{O}_6$ ) 11%.

**Setting of cement :** When mixed with water, the setting of cement takes place to give a hard mass. This is due to the hydration of the molecules of the constituents and their rearrangement. The purpose of adding gypsum is only to slow down the process of setting of the cement so that it gets sufficiently hardened.

**Uses :** Cement has become a commodity of national necessity for any country next to iron and steel. It is used in concrete and reinforced concrete, in plastering and in the construction of bridges, dams and buildings.

#### Common Names

The names marked with asterisk (\*) should be memorized with formulae. Others are given only for reference. You need not memorize them.

Metal	Ore name	Formula
Lithium (Li)	Spodumene	$\text{LiAl}(\text{SiO}_3)_2$
	Lepidolite	$\text{KLi}_2\text{Al}(\text{Al},\text{Si})_3\text{O}_{10}(\text{F},\text{OH})_2$
	Petalite	$\text{LiAl}(\text{Si}_2\text{O}_5)_2$
Sodium (Na)	*Washing soda	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
	*Baking soda	$\text{NaHCO}_3$
	*Sodium carbonate (soda ash / washing soda)	$\text{Na}_2\text{CO}_3$
	*Sodium chloride (rock salt or halite)	$\text{NaCl}$
	*Sodium nitrate (Chile saltpeter)	$\text{NaNO}_3$
	Salt cake	$\text{Na}_2\text{SO}_4$
	Fusion mixture	$\text{Na}_2\text{CO}_3 + \text{K}_2\text{CO}_3$ (eq. molar mix.)
	Sodium sesquicarbonate (trona)	$\text{Na}_2\text{CO}_3 \cdot \text{NaHCO}_3 \cdot 2\text{H}_2\text{O}$ (it is a double salt)
	*Microcosmic salt	$\text{Na}(\text{NH}_4)\text{HPO}_4 \cdot 4\text{H}_2\text{O}$ (it is obtained by mixing solutions of sodium phosphate and ammonium phosphate or chloride)
	Soda feldspar or sodium feldspar (albite)	$\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$
Potassium (K)	Potash feldspars or orthoclase or microcline or Potassium feldspars	$\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$
	*Hypo	$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
	*Sodium aluminium fluoride (cryolite)	$\text{Na}_3\text{AlF}_6$
	*Borax (Tincal)	$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$
	*Sodium sulphate (glauber's salt)	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ <b>(Sodium sulfate is the sodium salt of sulfuric acid. When anhydrous, it is a white crystalline solid of formula <math>\text{Na}_2\text{SO}_4</math> known as the mineral thenardite; the decahydrate <math>\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}</math> is known as Glauber's salt)</b>
	Sodium aluminium silicate (Soda Feldspar)	$\text{NaAlSi}_3\text{O}_8$
	Sylvite	$\text{KCl}$
	Schonite	$\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$
	Kainite	$\text{MgSO}_4 \cdot \text{KCl} \cdot 3\text{H}_2\text{O}$
	*Carnallite	$\text{MgCl}_2 \cdot \text{KCl} \cdot 6\text{H}_2\text{O}$
Adv SBC - 24	*Indian saltpetre (Nitre)	$\text{KNO}_3$ (used especially as a fertilizer and explosive)
	Pearl ash	$\text{K}_2\text{CO}_3$
	Schonite	$\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$ (it is a double salt)
	Langbeinite	$\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 2\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
	Polyhalite	$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$
	*Potassium Alum	$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 4\text{Al}(\text{OH})_3$
	Alunite or Alumstone	$\text{K}_2\text{O} \cdot 3\text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2 \cdot 2\text{H}_2\text{O}$
	Mica	$\text{KAIS}_3\text{O}_8(\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2)$
	Feldspar	

Beryllium (Be)	Beryl	$3\text{BeO} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$
	Chrysoberyl	$\text{BeO} \cdot \text{Al}_2\text{O}_3$
	Phenacite	$\text{BeSiO}_4$
	Bromalite	$\text{BeO}$
	*Baryta	$\text{Ba}(\text{OH})_2$
Magnesium (Mg)	*Magnesite	$\text{MgCO}_3$
	*Dolomite	$\text{MgCO}_3 \cdot \text{CaCO}_3$
	*Epsom salt	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
	Kiesente	$\text{MgSO}_4 \cdot \text{H}_2\text{O}$
	Asbestos	$\text{CaMg}_3(\text{SiO}_3)_4$
	Talc	$\text{Mg}(\text{Si}_2\text{O}_5)_2 \cdot \text{Mg}(\text{OH})_2$
	Brucite	$\text{Mg}(\text{OH})_2$
	*Magnesia	$\text{MgO}$
	Artinite	$\text{MgCO}_3 \cdot \text{Mg}(\text{OH})_2 \cdot 3\text{H}_2\text{O}$
	*Sorel cement (magnesia cement)	$\text{Mg}_4\text{Cl}_2(\text{OH})_6(\text{H}_2\text{O})_8$
Calcium (Ca)	*Quick lime	$\text{CaO}$
	*Slaked lime	$\text{Ca}(\text{OH})_2$
	*Hydrolith	$\text{CaH}_2$
	*Calcium cyanamide	$\text{CaCN}_2$ OR $\text{CaNCN}$
	*Limestone (Marble / Whiting)	$\text{CaCO}_3$
	Anhydrite	$\text{CaSO}_4$
	*Gypsum	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
	*Fluorspar or Fluorite	$\text{CaF}_2$
	Phosphorite	$\text{Ca}_3(\text{PO}_4)_2$
	*Fluorapatite	$3\text{Ca}_3(\text{PO}_4)_2 \cdot \text{CaF}_2$ OR $\text{Ca}_5(\text{PO}_4)_3\text{F}$
	*Plaster of paris	$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$
	*Bleaching powder	$\text{CaOCl}_2$
	*Rock phosphate	$\text{Ca}_3(\text{PO}_4)_2$
	Wollastonite	$\text{CaSiO}_3$
	Colmanite	$2\text{CaO} \cdot 3\text{Ba}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
Strontium(Sr)	Strontianite	$\text{SrCO}_3$
	Celestite	$\text{SrSO}_4$
	Barytes or Heavy spar	$\text{BaSO}_4$

### Periodic Properties of s-Block

Properties	Order
Thermal stability	$\text{LiH} > \text{NaH} > \text{KH} > \text{RbH} > \text{CsH}$
Basic strength	$\text{BeO} < \text{MgO} < \text{CaO} < \text{SrO}$
Basic Strength or Solubility in water or thermal stability	$\text{LiOH} < \text{NaOH} < \text{KOH} < \text{RbOH} < \text{CsOH}$
Basic Strength or Solubility in water	$\text{Be}(\text{OH})_2 < \text{Mg}(\text{OH})_2 < \text{Ca}(\text{OH})_2 < \text{Ba}(\text{OH})_2$
Thermal stability	$\text{Be}(\text{OH})_2 < \text{Mg}(\text{OH})_2 < \text{Ca}(\text{OH})_2 < \text{Sr}(\text{OH})_2 < \text{Ba}(\text{OH})_2$
Solubility in water or thermal stability	$\text{Li}_2\text{CO}_3 < \text{Na}_2\text{CO}_3 < \text{K}_2\text{CO}_3 < \text{Rb}_2\text{CO}_3 < \text{Cs}_2\text{CO}_3$
Solubility in water	$\text{BaCO}_3 < \text{CaCO}_3 < \text{MgCO}_3 < \text{BeCO}_3$
Thermal stability	$\text{BeCO}_3 < \text{MgCO}_3 < \text{CaCO}_3 < \text{BaCO}_3$
Solubility in water	$\text{BaSO}_4 < \text{SrSO}_4 < \text{CaSO}_4 < \text{MgSO}_4 < \text{BeSO}_4$

### Exercise-1

Marked Questions may have for Revision Questions.

#### PART - I : SUBJECTIVE QUESTIONS

##### Section (A) : General facts about elements

A-1. Why do alkali metals form unipositive ions and impart characteristic colours to flame?

##### Section (B) : Based on Periodic trends

B-1. (a) Explain why is sodium less reactive than potassium?  
(b) IE<sub>1</sub> value of Mg is more than that of Na while its IE<sub>2</sub> value is less. Explain?

B-2. Comment on the order of mobilities of the alkali metal ions in aqueous solution:  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$ .

##### Section (C) : Based on Chemical Bonding

C-1. Why is  $\text{KO}_2$  paramagnetic?

C-2. Draw the structure of  $\text{BeCl}_2$  in solid and vapour state.

C-3. Explain why in anion of  $\text{Na}_2\text{CO}_3$  all bond lengths are equal?

C-4. Order of the ionic character of following:  
 $\text{MgCl}_2, \text{MgBr}_2, \text{Mgl}_2$

C-5. Why  $\text{LiNO}_3$  on heating shows exceptional behaviour than other elements of this group?

C-6. Write the order of thermal stability of following:  
 $\text{BeSO}_4, \text{MgSO}_4, \text{CaSO}_4, \text{SrSO}_4$

C-7. Write the increasing order of basic strength of following:  
 $\text{NaOH}, \text{KOH}, \text{RbOH}, \text{CsOH}$

C-8. Although ionisation potential of Li is very high, then why is it a good reducing agent?

##### Section (D) : Properties of elements

D-1. Alkali metals are soft and can be cut with the help of a knife. Explain.

D-2. We know air mostly contains ( $\text{O}_2, \text{N}_2$ ). What happens when group-I and group-II elements of s-block react with excess of air?

Group-I Elements	+ $\text{O}_2$ (Major product)	+ $\text{N}_2$ (product)	Group-II Elements	+ $\text{O}_2$ (Major product)	+ $\text{N}_2$ (product)
Li	.....	.....	Be	.....	.....
Na	.....	.....	Mg	.....	.....
K	.....	.....	Ca	.....	.....
Rb	.....	.....	Sr	.....	.....
Cs	.....	.....	Ba	.....	.....

D-3. What happens when sodium and calcium metal are dropped in water?



(where M = group-II elements)



(where M = group-II elements)

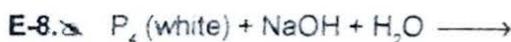
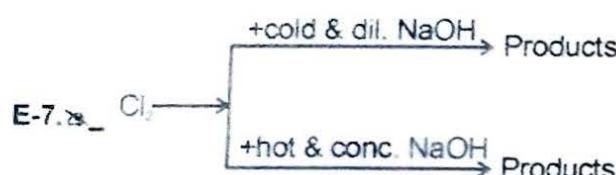
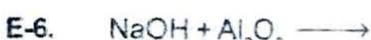
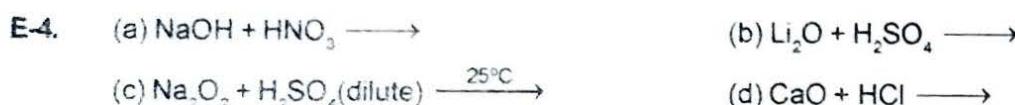
D-5. What happens when sodium metal is dissolved in liquid ammonia?

## Section (E) : Oxides, Peroxides, Super Oxides, Hydroxides

E-1. Lithium forms monoxide, sodium gives peroxide while the rest of the alkali metals form superoxide mainly when treated with excess of air. Explain.

E-2. How NaOH is commercially prepared ? Which cell is used ?

E-3. Write the method of preparation of  $\text{Na}_2\text{O}_2$  &  $\text{KO}_2$  and also give their hydrolysis product ?



## Section (F) : Carbonates, Bicarbonates

F-1. The thermal stability order of following carbonates  
 $\text{BeCO}_3, \text{MgCO}_3, \text{CaCO}_3, \text{SrCO}_3, \text{BaCO}_3$

F-2. Write chemical changes of solvay process.



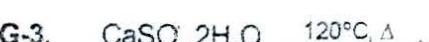
F-4. Write the products of the following reactions :



## Section (G) : Chlorides, Sulphates

G-1. Decreasing order of solubility in water of following sulphates :  
 $\text{BeSO}_4, \text{MgSO}_4, \text{CaSO}_4, \text{SrSO}_4$

G-2. How is  $\text{CaCl}_2$  prepared ?



G-4. How would you explain ?

- (i)  $\text{BeO}$  is insoluble but  $\text{BeSO}_4$  is soluble in water.
- (ii)  $\text{BaO}$  is soluble but  $\text{BaSO}_4$  is insoluble in water.



C-5. Which is amphoteric ?

- (A)  $\text{Li}_2\text{O}$       (B)  $\text{BeO}$       (C)  $\text{BaO}$       (D)  $\text{Cs}_2\text{O}$

C-6. Alkali metals are

- (A) good reductant      (B) good oxidant      (C) Both of these      (D) None of these

**Section (D) : Properties of elements**

D-1. The metallic lustre exhibited by sodium is explained by :

- (A) diffusion of sodium ions      (B) oscillation of mobile valence electrons  
(C) existence of free protons      (D) existence of body centered cubic lattice

D-2. Which of the following will appear silvery white ?

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

D-3. Which of the following s-block metal does not react with water ?

- (A) K      (B) Na      (C) Ca      (D) Be

D-4. Which of the following option is correct for given reaction ?



- (A) It reacts vigorously with acid if M is alkali metal.  
(B) It reacts readily with acid if M is alkaline earth metal.  
(C) metal sulphate and hydrogen gas will form after reaction.  
(D) All are correct.

D-5. Be reacts with excess of caustic soda to form :

- (A)  $\text{Be}(\text{OH})_2$       (B)  $\text{BeO}$       (C)  $\text{Na}_2[\text{Be}(\text{OH})_4]$       (D)  $\text{Be}(\text{OH})_2\text{BeCO}_3$

**Section (E) : Oxides, Peroxides, Super Oxides, Hydroxides**

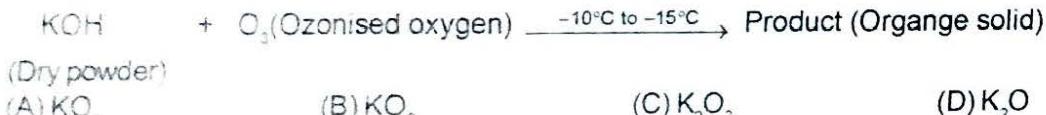
E-1. What is [X] in the following reaction ?



E-2. When magnesium burns in air, compounds of magnesium formed are magnesium oxide and :

- (A)  $\text{Mg}_3\text{N}_2$       (B)  $\text{MgCO}_3$       (C)  $\text{Mg}(\text{NO}_3)_2$       (D)  $\text{Mg}(\text{NO}_2)_2$

E-3. Which product will be formed after the reaction



E-4. Peroxide ion is present in .

- (A)  $\text{KO}_2$       (B)  $\text{CaO}$       (C)  $\text{Li}_2\text{O}$       (D)  $\text{BaO}_2$

E-5. The compound that gives hydrogen peroxide on treatment with a dilute cold acid is :

- (A)  $\text{PbO}_2$       (B)  $\text{Na}_2\text{O}_2$       (C)  $\text{MnO}_2$       (D)  $\text{SnO}_2$

E-6. Products of following reaction .



E-7. The principal products obtained on heating iodine with concentrated caustic soda solution is :

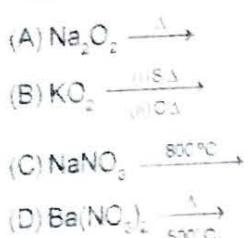
- (A)  $\text{NaIO} + \text{NaI}$       (B)  $\text{NaIO} + \text{NaIO}_3$       (C)  $\text{NaIO}_3 + \text{NaI}$       (D)  $\text{NaIO}_4 + \text{NaI}$

E-8. Products of the following reaction are :



**PART - III : MATCH THE COLUMN**

1. Match the reactions listed in column-I with the characteristic(s) of the products listed in column-II.

**Column - I****Column - II**

- (p) One of the products is diamagnetic.  
 (q) One of the products acts as reducing agent.  
 (r) One of the products acts as oxidising agent.  
 (s) One of the products is a basic oxide.

2. Match the compounds listed in column-I with the characteristic(s) listed in column-II.

**Column-I**

- (A)  $\text{BeO}(s)$   
 (B)  $\text{NaHCO}_3$  (crystalline)  
 (C)  $\text{BeCl}_2(s)$   
 (D)  $\text{CsO}_2(s)$

**Column-II**

- (p) Amphoteric in nature  
 (q) Imparts characteristic colour to Bunsen flame.  
 (r) Produce  $\text{H}_2\text{O}_2$  and  $\text{O}_2$  on reaction with water.  
 (s) Show hydrogen bonding  
 (t) Has a chain structure

**Exercise-2**

Marked Questions may have for Revision Questions.

**PART - I : ONLY ONE OPTION CORRECT TYPE**

1. The element having electronic configuration  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$  will form :  
 (A) Acidic oxide      (B) Basic oxide      (C) Amphoteric oxide      (D) Neutral oxide
2. **Marked** Beryllium has less negative value of reduction potentials compared to other alkaline earth metals due to:  
 (A) the smaller hydration energy of the  $\text{Be}^{2+}$ .  
 (B) the large value of the atomization enthalpy of the Be metal.  
 (C) the large value of ionisation energy of the Be metal.  
 (D) (B) and (C) both.
3. **Marked** The incorrect statement is :  
 (A)  $\text{Be}^{2+}$  cation has largest hydration enthalpy among the alkaline earth metals.  
 (B) The second ionisation enthalpies of alkaline earth metals are smaller than those of the corresponding alkali metals.  
 (C) Li is the strongest reducing agent among all the elements.  
 (D) Both  $\text{LiCl}$  and  $\text{MgCl}_2$  are most covalent in their groups.
4. Select the correct statement with respect to alkali metals.  
 (A) Melting point decrease with increasing atomic number.  
 (B) Potassium is lighter than sodium.  
 (C) Salts of Li to Cs impart characteristic colour to an oxidising flame (of Bunsen burner).  
 (D) All of these
5. On dissolving moderate amount of sodium metal in liquid  $\text{NH}_3$  at low temperature, which one of the following does not occur?  
 (A) Blue coloured solution is obtained  
 (B)  $\text{Na}^+$  ions are formed in the solution  
 (C) Liquid  $\text{NH}_3$  becomes good conductor of electricity  
 (D) Liquid  $\text{NH}_3$  remains diamagnetic



## s-Block Elements



Which of the following option describes the products, reactants and the reaction conditions.

Option	(p)	(q)	(r)	(s)
(A)	Crystallisation	Heat at 393 K	$2 \text{CaSO}_4 \cdot \text{H}_2\text{O}$	$\text{CaSO}_4$
(B)	Crystallisation	Heat at 393 K at high pressure	$2 \text{CaSO}_4 \cdot \text{H}_2\text{O}$	$\text{CaSO}_4$
(C)	Higher temperature	Cool	$\text{CaSO}_4 \cdot \text{H}_2\text{O}$	$\text{CaSO}_4$
(D)	Higher pressure	Heat at 393 K	$\text{CaSO}_4$	$\text{CaSO}_3$

21. Setting of plaster of paris involves:

- (A) the oxidation with atmospheric oxygen.
- (B) the removal of water to form anhydrous calcium sulphate.
- (C) the hydration to form the orthorhombic form of gypsum.
- (D) the reaction with atmospheric carbon dioxide gas.

22. Calcium cyanamide on reaction with steam under pressure gives ammonia and -----.

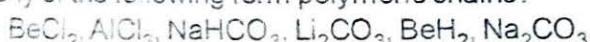
- (A) calcium carbonate
- (B) calcium hydroxide
- (C) calcium oxide
- (D) calcium bicarbonate

## PART - II : SINGLE AND DOUBLE VALUE INTEGER TYPE

1. How many of the following are correctly matched?

Element	Colour in flame test
K	Violet/Lilac
Na	Yellow
Bc	Crimson red
Ca	Brick red
Sr	Apple green
Mg	No colour
Rb	Red violet
Cs	Blue
Li	Crimson red

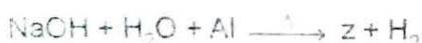
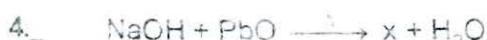
2. How many of the following form polymeric chains?



3. For alkali metal M,



Sum of the number of atoms present in one molecule each of x, y, z.



Sum of the number of atoms present in one molecule each of x, y, z is ..... (Assume no complex formation)

5. How many of the following will turn moist red litmus blue and finally white?



6. The by product of solvay process reacts with  $\text{Na}_2\text{CO}_3$  to form a compound x, which on heating decomposes to give y. y is absorbed by KO<sub>2</sub>. The number of atoms per molecule of y is .....

## **PART - IV : COMPREHENSION**

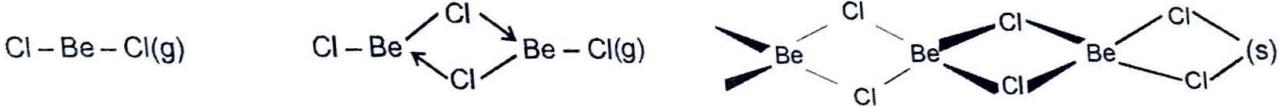
**Read the following passage carefully and answer the questions.**

### Comprehension # 1

All alkali metals dissolve in anhydrous liquid ammonia to give blue colour solution. It is the ammoniated electron which is responsible for the blue colour of the solution, and the electrical conductivity is mainly due to ammoniated electron,  $[e(NH_3)_7]$ . Dilute solutions are paramagnetic due to free ammoniated electrons ; this paramagnetism decreases at higher concentration. Above 3M concentration, the solutions are diamagnetic and no longer blue but are bronze/copper-bronze coloured with a metallic luster.

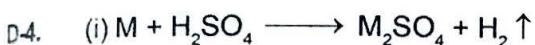
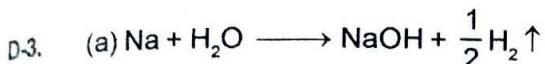
# EXERCISE - 1

## PART - I

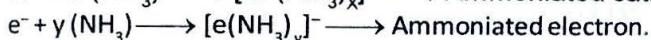
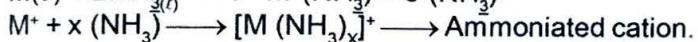
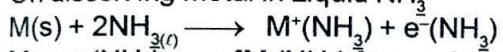
- A-1. After removal of  $1^{\text{st}}$  electron alkali metal occupies inert gas configuration. Now removal of  $2^{\text{nd}}$  electron from inert gas configuration requires very high energy, therefore, they form unipositive ions. As IE<sub>1</sub> of these metals are low, the excitation of electrons can be done by providing less energy. This much of energy can be given by Bunsen flame. When they drop back to the ground state, there is emission of radiation in the visible region.
- B-1. (a) The ionization enthalpy ( $\Delta H$ ) of potassium (419 kJ mol<sup>-1</sup>) is less than that of sodium (496 kJ mol<sup>-1</sup>) or more precisely the standard electrode potential ( $E^\circ$ ) of potassium (-2.925 V) is more negative than that of sodium (-2.714 V) and hence potassium is more reactive than sodium.  
(b) IE<sub>1</sub> of Mg (3s<sup>2</sup>) > Na (3s<sup>1</sup>)  
as Mg has fully filled electronic configuration while Na has one unpaired electron.  
IE<sub>2</sub> of Mg (3s<sup>1</sup>) > Na (2p<sup>6</sup>)  
as Mg<sup>+</sup> (3s<sup>1</sup>) has one unpaired electron and Na<sup>+</sup> has inert gas configuration.
- B-2. Smaller the size of the ion, more highly it is hydrated and hence greater is the mass of the hydrated ion and hence lower is its ionic mobility. Since the extent of hydration decreases in the order :  
 $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$   
therefore, ionic mobility increases in the same order :  
 $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$
- C-1. The superoxide O<sub>2</sub><sup>-</sup> is paramagnetic because of one unpaired electron in  $\pi^* 2p$  molecular orbital.
- C-2. In vapour state it exists as linear or dimeric molecules where as in solid it has polymeric structure,
- 
- C-3. In [CO<sub>3</sub><sup>2-</sup>] all C-O bonds are equal due to resonance with B.O =  $\frac{3}{2}$ , so their bond length will also be equal.
- C-4. There will be more polarisation of big anion due to Fajan's factors, so covalent character will be more in I<sup>-</sup> due to large size and ionic character will be less.  
 $\text{MgCl}_2 > \text{MgBr}_2 > \text{Mgl}_2$
- C-5. Due to small size Li<sup>+</sup>, it has high polarising power while from Na<sup>+</sup> to Cs<sup>+</sup> have bigger size. So they have low polarising power. Li<sup>+</sup> is more similar to Mg<sup>2+</sup> in its properties, which destabilizes a polyatomic anion due to its high polarising power.
- C-6.  $\text{BeSO}_4 < \text{MgSO}_4 < \text{CaSO}_4 < \text{SrSO}_4$
- C-7.  $\text{NaOH} < \text{KOH} < \text{RbOH} < \text{CsOH}$
- C-8. Lithium is expected to be least reducing agent due to its very high I.E. However, lithium has the highest hydration enthalpy due to small size which accounts for its high negative E<sup>⊖</sup> and its high reducing power.
- D-1. Due to large atomic size & only one valence electron per atom, alkali metals have weak metallic bonds as interparticle forces.

D-2.

Group-I Elements	+ O <sub>2</sub> (Major product)	+ N <sub>2</sub> (product )	Group-II Elements	+ O <sub>2</sub> (Major product)	+ N <sub>2</sub> (product) (only on strong heating)
Li	Li <sub>2</sub> O (Oxide)	Li <sub>3</sub> N	Be	BeO	Be <sub>3</sub> N <sub>2</sub>
Na	Na <sub>2</sub> O <sub>2</sub> (Peroxide)	It does not react	Mg	MgO	Mg <sub>3</sub> N <sub>2</sub>
K	KO <sub>2</sub> (superoxide)	It does not react	Ca	CaO	Ca <sub>3</sub> N <sub>2</sub>
Rb	RbO <sub>2</sub> (superoxide)	It does not react	Sr	SrO <sub>2</sub>	Sr <sub>3</sub> N <sub>2</sub>
Cs	CsO <sub>2</sub> (superoxide)	It does not react	Ba	BaO <sub>2</sub>	Ba <sub>3</sub> N <sub>2</sub>



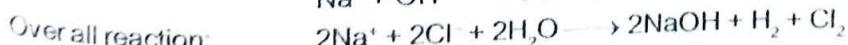
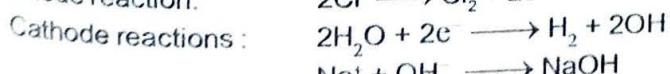
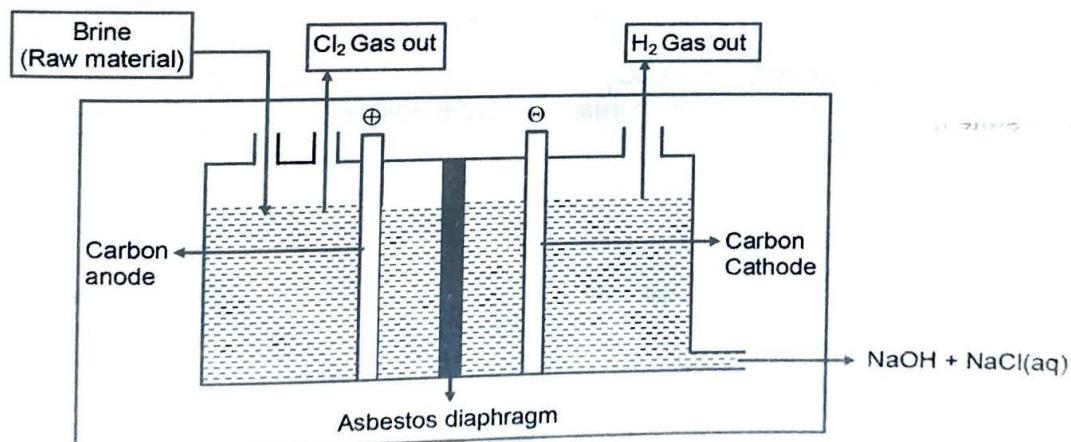
D-5. On dissolving Metal in Liquid NH<sub>3</sub>



E-1. Small cation have high polarizing power therefore it stabilizes monoatomic anion. e.g. Li<sub>2</sub>O

Large cation have less polarizing power therefore it can stabilize polyatomic anion. e.g. Na<sub>2</sub>O<sub>2</sub>; KO<sub>2</sub>

E-2. NaOH is commercially prepared by electrolysis of brine solution in diaphragm cell.



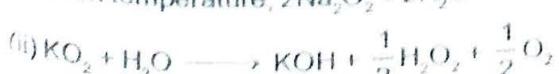
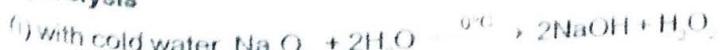
E-3. (i) Industrial method: It is a two stage reaction in presence of excess of air.

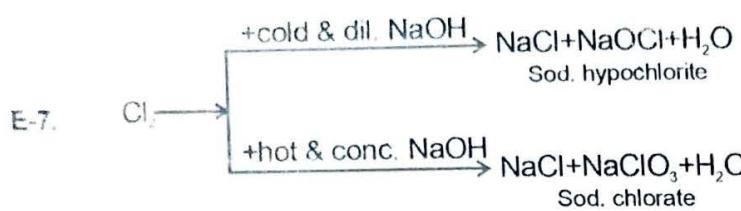
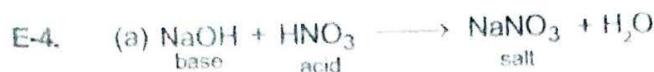


(ii) It is prepared by burning potassium in excess of oxygen free from moisture.

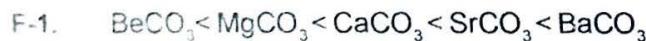


Hydrolysis





E-8. It goes under disproportionation reaction



Stability of carbonates increases with increase in electropositive character and decrease in polarisation power of metal.

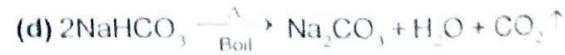
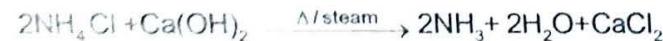
F-2. (i) In ammonia absorber



(ii) Calcination



(iii) In recovery tower:



G-1.  $\text{BeSO}_4 > \text{MgSO}_4 > \text{CaSO}_4 > \text{SrSO}_4$  Bigger cation is stable with bigger anion where as smaller cation is less stable with bigger anion that why  $\text{BeSO}_4$  is more water soluble

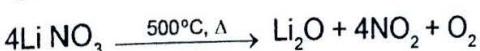
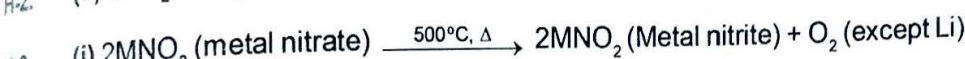
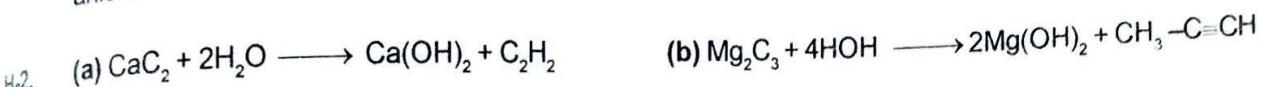
G-2. It is produced in large amount as a by product in solvay process.



## s-block Elements

- G.4. (i)  $\text{Be}^{2+}$  &  $\text{O}^{2-}$  smaller in size & thus higher lattice energy and lattice energy is greater than hydration energy in  $\text{BeO}$  where as in  $\text{BeSO}_4$  lattice energy is less due to bigger sulphate ion and is soluble.  
 Order of solubility :  $\text{BeO} < \text{MgO} < \text{CaO} < \text{SrO} < \text{BaO}$   
 (ii) In  $\text{BaSO}_4$  lattice energy is greater than hydration energy while in  $\text{BaO}$  lattice energy is smaller than hydration energy.  
 Order of solubility :  $\text{BeSO}_4 > \text{MgSO}_4 > \text{CaSO}_4 > \text{SrSO}_4 > \text{BaSO}_4$

- G.5.  $\text{NaOCl} + \text{HOH} \longrightarrow \text{NaOH} + \text{HOCl}$  G-6.  $\text{CaCl}_2 + \text{H}_2\text{SO}_4$  (conc.)  $\xrightarrow{\Delta} \text{CaSO}_4 + 2\text{HCl}$   
 H.1. Order is  $\text{LiH} > \text{NaH} > \text{KH} > \text{RbH} > \text{CsH}$  because small  $\text{Li}^+$  due to high polarisation power will stabilise smaller anion.



## PART - II

A.1. (A)	A-2. (B)	B-1. (B)	B-2. (A)	C-1. (B)
C-2. (D)	C-3. (D)	C-4. (C)	C-5. (B)	C-6. (A)
D-1. (B)	D-2. (D)	D-3. (D)	D-4. (D)	D-5. (C)
E-1. (A)	E-2. (A)	E-3. (B)	E-4. (D)	E-5. (B)
E-6. (B)	E-7. (C)	E-8. (A)	F-1. (B)	F-2. (D)
F-3. (C)	F-4. (A)	F-5. (A)	F-6. (C)	G-1. (D)
G-2. (A)	G-3. (C)	G-4. (C)	G-5. (A)	H-1. (D)
H-2. (A)	H-3. (D)	H-4. (B)	H-5. (D)	H-6. (C)

## PART - III

1. (A - p,r,s); (B - p,q); (C - p,q,r,s); (D - p,q,r). 2. (A - p); (B - p, q, s, t); (C - t); (D - q, r)

## EXERCISE - 2

		PART - I		
		3. (D)	4. (D)	5. (D)
	2. (D)	8. (D)	9. (B)	10. (A)
	7. (B)	13. (A)	14. (D)	15. (C)
	12. (C)	18. (C)	19. (B)	20. (A)
	17. (C)			
	22. (A)			

## PART - II

2.  $3(\text{BeCl}_2, \text{NaHCO}_3, \text{BeI}_2)$   
 4.  $15(x = 5, y = 6, z = 4)$   
 6.  $3(\text{CO}_2)$  7. 4 (1<sup>st</sup> four)  
 9. 10. 3

7 (All except Be & Mg)

9 ( $x = 3, y = 4, z = 2$ )

4 ( $\text{KO}_2, \text{PbO}_2, \text{Cs}_2\text{O}_2, \text{BaO}_2$ )

18 (4, 3, 6, 5)

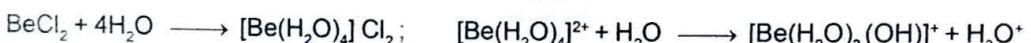
- PART - III**
- |             |               |               |              |             |
|-------------|---------------|---------------|--------------|-------------|
| 1. (A,B)    | 2. (A,B)      | 3. (B,C)      | 4. (A,B,D)   | 5. (A,B,C)  |
| 6. (C,D)    | 7. (A,B,C)    | 8. (A,B,C)    | 9. (A,B,C,D) | 10. (A,B,C) |
| 11. (A,B,D) | 12. (B,C,D)   | 13. (A,B,C,D) | 14. (A,C)    | 15. (A,C)   |
| 16. (B,C,D) | 17. (A,B,C,D) | 18. (A,D)     | 19. (A,C)    | 20. (A,B)   |

**PART - IV**

- |        |        |        |
|--------|--------|--------|
| 1. (D) | 2. (D) | 3. (B) |
|--------|--------|--------|

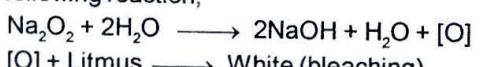
**EXERCISE - 3****PART - I**

- |        |          |            |
|--------|----------|------------|
| 1. (B) | 2. (A,B) | 3. (A,B,D) |
|--------|----------|------------|
4. (i) Beryllium chloride is acidic, when dissolved in water because the hydrated ion hydrolysed producing  $\text{H}_3\text{O}^+$ . This happens because the Be–O bond is very strong, and so in the hydrated ion this weakens the O–H bonds, and hence there is tendency to lose protons.

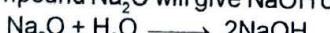


5. A =  $\text{Ca}(\text{OH})_2$ , B =  $\text{NH}_4\text{HCO}_3$ , C =  $\text{Na}_2\text{CO}_3$ , D =  $\text{NH}_4\text{Cl}$ , E =  $\text{CaCl}_2$

6. (i)  $\text{Na}_2\text{O}_2$  is powerful oxidant and bleaching agent and bleaches red litmus paper to white in aqueous solution according to the following reaction,



- (ii) The other compound  $\text{Na}_2\text{O}$  will give NaOH on dissolution in water according to the following reaction.



The red litmus will turn to blue due to stronger alkaline nature of NaOH.

7. (B)
8.  $\text{Na}_2\text{CO}_3 + \text{SO}_2 \xrightarrow{\text{H}_2\text{O}} 2\text{NaHSO}_3$  (A) +  $\text{CO}_2$
- $$2\text{NaHSO}_3 + \text{Na}_2\text{CO}_3 \longrightarrow 2\text{Na}_2\text{SO}_3$$
- (B) +
- $\text{H}_2\text{O} + \text{CO}_2$
- $$\text{Na}_2\text{SO}_3 + \text{S} \xrightarrow{\Delta} \text{Na}_2\text{S}_2\text{O}_3$$
- (C)
- $$2\text{Na}_2\text{S}_2\text{O}_3 + \text{I}_2 \longrightarrow \text{Na}_2\text{S}_4\text{O}_6$$
- (D) +
- $2\text{NaI}$

Oxidation states of S + 4 in  $\text{NaHSO}_3$  [ $1 + 1 + x + 3(-2) = 0$ ] and +4 in  $\text{Na}_2\text{SO}_3$  [ $2 + x + 3(-2) = 0$ ] ; +6 and –2 (or an average +2) in  $\text{Na}_2\text{S}_2\text{O}_3$  and +5 and 0 (or an average +5/2) in  $\text{Na}_2\text{S}_4\text{O}_6$ .

9. (B)      10.\* (A,B)

**PART - II**

- |         |         |         |         |         |
|---------|---------|---------|---------|---------|
| 1. (1)  | 2. (1)  | 3. (1)  | 4. (1)  | 5. (1)  |
| 6. (2)  | 7. (3)  | 8. (1)  | 9. (4)  | 10. (2) |
| 11. (1) | 12. (1) | 13. (3) | 14. (2) | 15. (1) |
| 16. (3) |         |         |         |         |