

**Alkali Metals (Group I elements of modern periodic table):**

Property	Elements						
	Li	Na	K	Rb	Cs	Fr (Radioactive)	
Atomic Number	3	11	19	37	55	87	
Electronic Configuration	2s'	3s'	4s'	5s'	6s'	7s'	
Atomic Mass	6.94	22.99	39.10	85.47	132.91	223	
Metallic radius (pm)	152	186	227	248	265	375	
Ionic radius (M <sup>+</sup> /pm)	76	102	138	152	167	180	
Ionization enthalpy (kJ mol <sup>-1</sup> )	I	520	496	419	403	376	–
	II	7298	3562	3051	2633	2230	–
Electro negativity (Pauling Scale)	0.98	0.93	0.82	0.82	0.79		
Density/g cm <sup>-3</sup> (at 298K)	0.53	0.97	0.86	1.53	1.90		
Melting point/K	454	371	336	312	302		
Boiling point/K	1615	1156	1032	961	944		
E°(V) at 298K for M <sup>+</sup> (aq) + e <sup>-</sup> → M(s)	-3.04	-2.714	-2.925	-2.930	-2.927		
Occurrence in Atmosphere	18*	227**	1.84**	78.12*	2.6*		

\*ppm (parts per million)

\*\* percentage by weight

**Physical Properties of Alkali Metals:**

These are soft in nature and can be cut with the help of knife except the lithium.

The atoms of alkali metals have the largest size in their respective periods.

The first ionization energy of the alkali metals are the lowest as compared to the elements in the other group.

The alkali metals show +1 oxidation state.

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The alkali metals have low values of reduction potential (as shown in table-I) and therefore have a strong tendency to lose electrons and act as good reducing agents.

The melting and boiling points of alkali metals are very low because the intermetallic bonds in them are quite weak.

All the alkali metals form ionic (electrovalent) compounds.

The alkali metals are good conductors of heat and electricity.

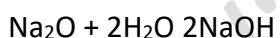
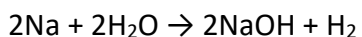
Alkali metals (except Li) exhibit photoelectric effect

The alkali metals and their salts impart a characteristic colour to flame

	Li	Na	K	Rb	Cs
Color	Crimson Red	Golden Yellow	Pale Violet	Violet	Sky Blue
$\lambda/nm$	670.8	589.2	766.5	780.0	455.5

### Hydroxides of Alkali Metals:

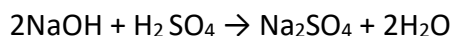
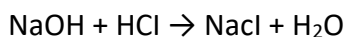
a) All the alkali metals, their oxides, peroxides and superoxides readily dissolve in water to produce corresponding hydroxides which are strong alkalies.



Z) The basic strength of these hydroxides increases as we move down the group Li to Cs.

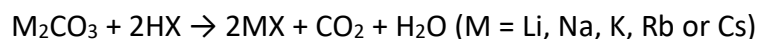
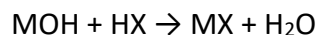
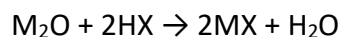
c) All these hydroxides are highly soluble in water and thermally stable except lithium hydroxide.

d) Alkali metals hydroxides being strongly basic react with all acids forming salts.



### Halides of Alkali metals:

For more info visit <http://jeemains2018.in>



(X = F, Cl, Br or I)

**a) Standard enthalpies of formation in (kJ/mol<sup>-1</sup>)**

Element	MF	MCl	MBr	MI
Li	-612	-398	-350	-271
Na	-569	-400	-360	-288
K	-563	-428	-392	-328
Rb	-549	-423	-389	-329
Cs	-531	-424	-395	-

**b) Covalent Character:**

Small cation and large anion favors covalency.

Order: LiCl > NaCl > KCl > RbCl > CsCl & . LiI > LiBr > LiCl > LiF

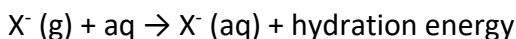
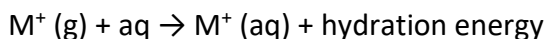
Greater the charge on the cation greater is its polarizing power and hence larger is the covalent character: Na<sup>+</sup>Cl<sup>-</sup> < Mg<sup>+2</sup>Cl<sub>2</sub> < Al<sup>+3</sup>Cl<sub>3</sub>

Greater the charge on the anion, more easily it gets polarized thereby imparting more covalent character to the compound formed eg covalent character increase in the order.  
NaCl < Na<sub>2</sub>SO<sub>4</sub> < Na<sub>3</sub>PO<sub>4</sub>

**c) Lattice Energies:** Amount of energy required to separate one mole of solid ionic compound into its gaseous ions.

Greater the lattice energy, higher is the melting point of the alkali metals halide and lower is its solubility in water

**d) Hydration Energy:** Amount of energy released when one mole of gaseous ions combine with water to form hydrated ions.



Higher the hydration energy of the ions greater is the solubility of the compound in water.

The solubility of the most of alkali metal halides except those of fluorides decreases on descending the group since the decrease in hydration energy is more than the corresponding decrease in the lattice energy.

Due to high hydration energy of  $\text{Li}^+$  ion, Lithium halides are soluble in water except  $\text{LiF}$  which is sparingly soluble due to its high lattice energy.

For the same alkali metal the melting point decreases in the order

fluoride > chloride > bromide > iodide

For the same halide ion, the melting point of lithium halides are lower than those of the corresponding sodium halides and thereafter they decrease as we move down the group from Na to Cs.

The low melting point of  $\text{LiCl}$  (887 K) as compared to  $\text{NaCl}$  is probably because  $\text{LiCl}$  is covalent in nature and  $\text{NaCl}$  is ionic.

#### **Anomalous Behavior of Lithium and diagonal relationship with Magnesium:**

Li has anomalous properties due to

Very small size

High polarizing Power

Lithium show diagonal relationship with magnesium because both elements have almost same polarizing power.

The melting point and boiling point of lithium are comparatively high.

Lithium is much harder than the other alkali metals. Magnesium is also hard metal.

Lithium reacts with oxygen least readily to form normal oxide whereas other alkali metals form peroxides and superoxides.

$\text{LiOH}$  like  $\text{Mg}(\text{OH})_2$  is weak base. Hydroxides of other alkali metals are strong bases.

Due to their appreciable covalent nature, the halides and alkyls of lithium and magnesium are soluble in organic solvents.

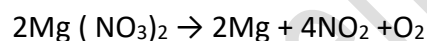
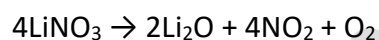
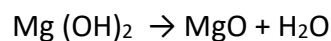
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Unlike elements of group 1 but like magnesium. Lithium forms nitride with nitrogen.  $6\text{Li} + \text{N}_2 \rightarrow 2\text{Li}_3\text{N}$

$\text{LiCl}$  is deliquescent and crystallizes as a hydrate,  $\text{LiCl} \cdot 2\text{H}_2\text{O}$ . Other alkali metals do not form hydrates. also forms hydrate,  $\text{MgCl}_2 \cdot 8\text{H}_2\text{O}$ .

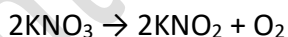
Unlike other alkali metals lithium reacts directly with carbon to form an ionic carbide. Magnesium also forms a similar carbide.

The carbonates, hydroxides and nitrates of lithium as well as magnesium decompose on heating.



The corresponding salts of other alkali metals are stable towards heat.

Lithium nitrate, on heating, decomposes to give lithium oxide,  $\text{Li}_2\text{O}$  whereas other alkali metals nitrate decomposes to give the corresponding nitrite.



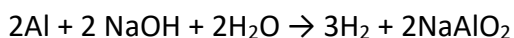
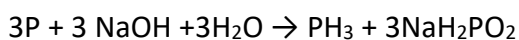
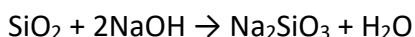
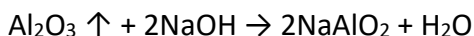
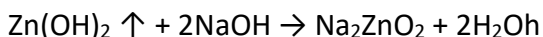
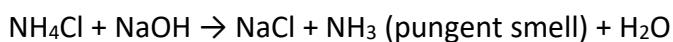
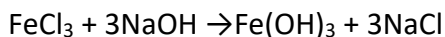
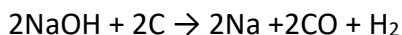
$\text{Li}_2\text{CO}_3$ ,  $\text{LiOH}$ ,  $\text{LiF}$  and  $\text{Li}_3\text{PO}_4$  are the only alkali metal salts which are insoluble in water. The corresponding magnesium compounds are also insoluble in water.

Hydrogen carbonates of both lithium and magnesium can not be isolated in solid state. Hydrogen carbonates of other alkali metals can be isolated in solid state.

**Sodium Hydroxide (NaOH):**

### a. Properties

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



### b) Uses

It is used in the manufacture of paper, soap and artificial silk.

It is used in petroleum refining.

It is used for mercerizing cotton.

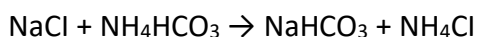
It is used for the preparation of sodium metal and many salts of sodium.

### Sodium Carbonate (Washing soda) ( $\text{Na}_2\text{CO}_3$ ):

#### a) Preparation:

Solvay process:

Carbon dioxide gas is bubbled through a brine solution saturated with ammonia and it results in the formation of sodium hydrogen carbonate.

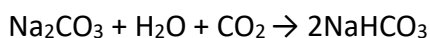


Sodium hydrogen carbonate so formed precipitates out because of the common ion effect caused due to the presence of excess of NaCl. The precipitated NaHCO<sub>3</sub> is filtered off and then ignited to get Na<sub>2</sub>CO<sub>3</sub>.

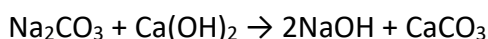
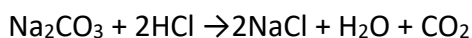


### b) Properties

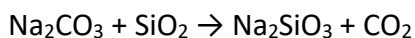
1. The aqueous solution absorbs CO<sub>2</sub> yielding sparingly soluble sodium bicarbonate.



2. dissolves in acids with an effervescence of carbondioxide and is causticised by lime to give caustic soda.



3. Fusion with silica, sodium carbonate yields sodium silicate.



4. Hydrolysis – being a salt of a strong base (NaOH) and weak acid (H<sub>2</sub>CO<sub>3</sub>), when dissolved in water sodium carbonate. Undergoes hydrolysis to form an alkaline solution



### c) Uses

It is used for softening hard water.

A mixture of sodium carbonate & potassium carbonate is used as fusion mixture.

As an important laboratory reagent both in qualitative and quantitative analysis.

It is used in paper, paints and textile industries.

It is used for washing purposes in laundry.

It is used in the manufacture of glass, borax, soap and caustic soda.

### Alkali Earth Metals (Group II elements of modern periodic table):

Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra).

### Physical Properties of Alkali Earth Metals:

#### 1. Alkali earth metals are almost similar in properties to the corresponding alkali metals.

Higher melting points than alkali metals

Higher boiling points alkali metals

Higher densities alkali metals

Harder than the corresponding alkali metals

#### 2. Atomic and ionic radii

The atomic radii as well as ionic radii of the members of the family are smaller than the corresponding members of alkali metals.

**3. Ionization energy:** The alkaline earth metals owing to their large size of atoms have fairly low values of ionization energies as compared to the p – block elements. However within the group, the ionization energy decreases as the atomic number increases. It is because of increase in atomic size due to addition of new shells and increase in the magnitude of screening effect of the electrons in inner shells. Because their  $(IE)_1$  is larger than that of their alkali metal neighbours, the group IIA metals trend to be somewhat less reactive than alkali metals. The general reactivity trend is  $Ba > Sr > Ca > Mg > Be$ .

**4. Oxidation state:** The alkaline earth metal have two electrons in their valence shell and by losing these electrons, these atoms acquire the stable noble gas configuration. Thus, unlike alkali metals, the alkaline earth metals exhibit +2 oxidation state in their compounds.



[noble gas]

#### 5. Characteristic flame colouration:

Ion	Colour
$Ca^{2+}$	Brick-red
$Sr^{2+}$	Crimson
$Ba^{2+}$	Apple green
$Ra^{2+}$	Carmine – red

#### Difference between alkali metals and alkali earth metals:

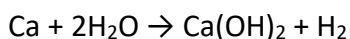
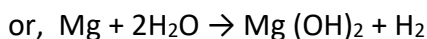
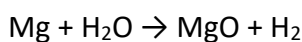
Properties	Alkaline earth metals	Alkali metals
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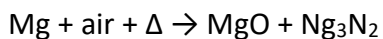
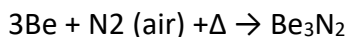
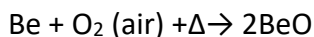
1.	Electronic configuration	Two electrons are present in the valency shell. The configuration is $ns^2$ (bivalent)	One electron is present in the valency shell. The configuration is $ns^1$ (monovalent) more electropositive
2.	Valency	Bivalent	Monovalent
3.	Electropositive nature	Less electropositive	More electropositive
4.	Hydroxides	Weak bases, less soluble and decompose on heating.	Strong bases, highly soluble and stable towards heat.
5.	Bicarbonates	These are not known in free state. Exist only in solution.	These are known in solid state.
6.	Carbonates	Insoluble in water. Decompose on heating.	Soluble in water. Do not decompose on heating ( $LiCO_3$ is an exception)
7.	Action of nitrogen	Directly combine with nitrogen and form nitrides	Do not directly combine with nitrogen except lithium
8.	Action of carbon	Directly combine with carbon and form carbides	Do not directly combine with carbon
9.	Nitrates	Decompose on heating evolving a mixture of $NO_2$ and oxygen	Decompose on heating evolving only oxygen
10.	Solubility of salts	Sulphates, phosphates fluorides, chromates, oxalates etc are insoluble in water	Sulphates, phosphates, fluorides, chromates, oxides etc are soluble in water.
11.	Physical properties	Comparatively harder. High melting points. Diamagnetic.	Soft, low melting points paramagnetic.
12.	Hydration of compounds	The compounds are extensively hydrated. $MgCl_2 \cdot 6H_2O$ , $CaCl_2 \cdot 6H_2O$ , $BaCl_2 \cdot 2H_2O$ are hydrated chlorides.	The compounds are less hydrated. $NaCl$ , $KCl$ , $RbCl$ form non – hydrated chlorides
13.	Reducing power	Weaker as ionization potential values are high and oxidation potential values are low.	Stronger as ionization potential values are low and oxidation potential values are high.

### Chemical Properties of Alkali Earth Metals:

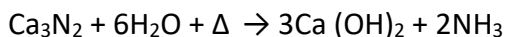
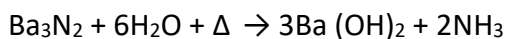
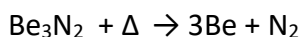
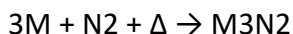
#### 1. Reaction with water :



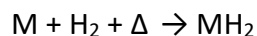
## 2. Formation of oxides and nitrides



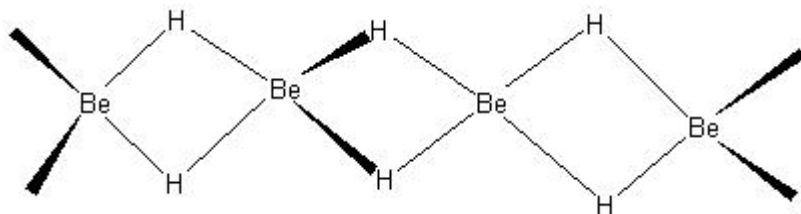
## 3. Formation of Nitrides



## 4. Reaction with hydrogen:

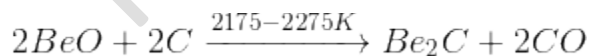


Both  $\text{BeH}_2$  and  $\text{MgH}_2$  are covalent compounds having polymeric structures in which H – atoms between beryllium atoms are held together by three centre – two electron (3C - 2e) bonds as shown below:

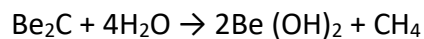


## 5. Reaction with carbon – (Formation of carbides)

When  $\text{BeO}$  is heated with carbon at 2175 – 2275 K a brick red coloured carbide of the formula  $\text{Be}_2\text{C}$  is formed



It is a covalent compound and react water forming methane.



## 6. Reaction with Ammonia:

Like alkali metal, the alkaline earth metals dissolve in liquid ammonia to give deep blue black solution from which ammoniates  $[M(NH_3)_6]^{2+}$  can be recovered.

### Anamolus Behaviour of Beryllium:

Be is harder than other members of its group.

Be is lighter than Mg.

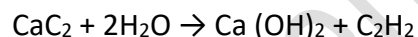
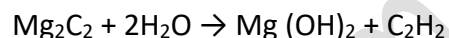
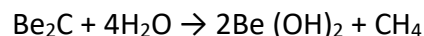
Its melting and boiling points are higher than those of Mg & other members.

Be does not react with water while Mg reacts with boiling water.

BeO is amphoteric while MgO is weakly basic.

Be forms covalent compounds whereas other members form ionic compounds.

Beryllium carbide reacts with water to give methane whereas carbides of other alkaline earth metals gives acetylene gas.



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

### Diagonal relationship of Be with Al:

Unlike groups – 2 elements but like aluminium, beryllium forms covalent compounds.

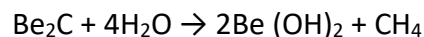
The hydroxides of Be,  $[Be(OH)_2]$  and aluminium  $[Al(OH)_3]$  are amphoteric in nature, whereas those of other elements of group – 2 are basic in nature.

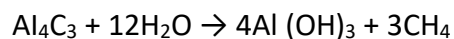
The oxides of both Be and Al i.e. BeO and  $Al_2O_3$  are high melting insoluble solids.

$BeCl_2$  and  $AlCl_3$  have bridged chloride polymeric structure.

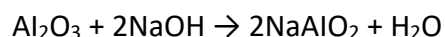
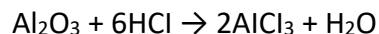
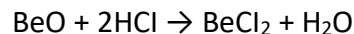
The salts of beryllium as well as aluminium are extensively hydrolysed.

Carbides of both the metal reacts with water liberating methane gas.





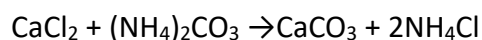
The oxides and hydroxides of both Be and Al are amphoteric and dissolve in sodium hydroxide as well as in hydrochloric acid.



Like Al, Be is not readily attacked by acids because of the presence of an oxide film.

### Calcium Carbonate ( $\text{CaCO}_3$ ):

It occurs in nature as marble, limestone, chalk, coral, calcite, etc. It is prepared as a white powder, known as precipitated chalk, by dissolving marble or limestone in hydrochloric acid and removing iron and aluminium present by precipitating with  $\text{NH}_3$ , and then adding ammonium carbonate to the solution; the precipitate is filtered, washed and dried.

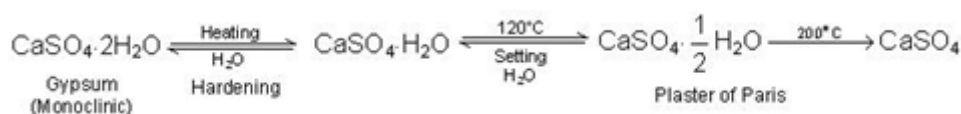
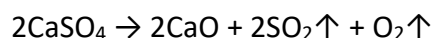


It dissolves in water containing  $\text{CO}_2$ , forming  $\text{Ca}(\text{HCO}_3)_2$  but is precipitated from solution by boiling.

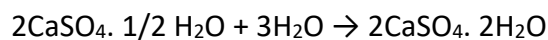


### Plaster of Paris, $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$ or $(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$ :

It occurs in nature as gypsum and the anhydrous salt as anhydrite. It is prepared by precipitating a solution of calcium chloride or nitrate with dilute sulphuric acid. The effect of heat on gypsum or the dihydrate presents a review of interesting changes. On heating the monoclinic gypsum is first converted into orthorhombic form without loss of water. When the temperature reaches  $120^\circ\text{C}$ , the hemihydrate or plaster of paris is the product. The latter loses water, becomes anhydrous above  $200^\circ\text{C}$  and finally above  $400^\circ\text{C}$ , it decomposes into calcium oxide.



The addition of common salt accelerates the rate of setting, while a little borax or alum reduces it. The setting of plaster of paris is believed to be due to rehydration and its reconversion into gypsum.



Plaster of Paris

Gypsum

### Uses

Plaster of paris is used for producing moulds for pottery and ceramics & casts of statues & busts.

It is used in surgical bandages used for plastering broken or fractured bones.

It is also used in dentistry

### Industrial uses of lime and Limestone

#### Uses of lime

Calcium oxide is called lime or quick lime. Its main industrial uses are

It is used in steel industry to remove phosphates and silicates as slag.

It is used to make cement by mixing it with silica, alumina or clay. (iii) It is used in making glass.

It is used in lime soda process for the conversion of  $\text{Na}_2\text{CO}_3$  to  $\text{NaOH}$  & vice versa.

It is used for softening water, for making slaked lime  $\text{Ca(OH)}_2$  by treatment with water and calcium carbide  $\text{CaC}_2$ .

#### Uses of Slaked lime [ $\text{Ca(OH)}_2$ ]

Slaked lime is used as a building material in form of mortar. It is prepared by mixing 3 – 4 times its weight of sand and by gradual addition of water. It sets into a hard mass by loss of  $\text{H}_2\text{O}$  and gradual absorption of  $\text{CO}_2$  from air.

In manufacture of bleaching powder by passing  $\text{Cl}_2$  gas.

In making glass and in the purification of sugar and coal gas.

It is used in softening of hard water.

#### Uses of lime stone ( $\text{CaCO}_3$ )

For more info visit <http://jeemains2018.in>

It is used as building material in form of marble.

In manufacture of quick lime.

It is used as a raw material for the manufacture of  $\text{Na}_2\text{CO}_3$  in solvay – ammonia process.

Commercial limestone contains iron oxide, alumina, magnesia, silica & sulphur with a CaO content of 22 – 56% MgO content upto 21%. It is used as such as a fertilizer.

<http://jeemains2018.in>