

## S Blocks Elements

### Group – I (Alkali Metals)

The group 1 of the periodic table contains six elements, namely lithium (*Li*), sodium (*Na*), potassium (*K*), rubidium (*Rb*), caesium (*Cs*) and francium (*Fr*). All these elements are typical metals. These are usually referred to as alkali metals since their hydroxides forms strong bases or alkalies.

#### General characteristics.

##### (1) Electronic configuration

Elements	Discovery	Electronic configuration ( $ns^1$ )
${}_3\text{Li}$	Arfwedson (1817)	$1s^2 2s^1$ or $[\text{He}]^2 2s^1$
${}_{11}\text{Na}$	Davy (1807)	$1s^2 2s^2 2p^6 3s^1$ or $[\text{Ne}]^{10} 3s^1$
${}_{19}\text{K}$	Davy (1807)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or $[\text{Ar}]^{18} 4s^1$
${}_{37}\text{Rb}$	Bunsen (1861)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ or $[\text{Kr}]^{36} 5s^1$
${}_{55}\text{Cs}$	Bunsen (1860)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ or $[\text{Xe}]^{54} 6s^1$
${}_{87}\text{Fr}$	Percy (1939)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^6 7s^1$ or $[\text{Rn}]^{86} 7s^1$

**Note** :  $\square$  Francium is radioactive with longest lived isotope  ${}^{223}\text{Fr}$  with half life period of only 21 minute.

##### (2) Physical properties

(i) **Physical state** : All elements are silvery white, soft and light metals. These are highly malleable and ductile. When freshly cut, they have a bright lustre which quickly tarnishes on exposure to air.

(ii) **Atomic and ionic radii** : Atomic as well as ionic size increases from *Li* to *Fr* due to the presence of an extra shell of electrons. Atomic volume (At. wt./Density) also increases in moving down from *Li* to *Cs*.

**Note** : □ Hydrogen having one electron in s-orbital can also be regarded as a s-block element but due to its peculiar behaviour it is always discussed separately.

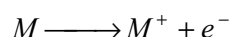
(iii) **Density** : Their densities are quite low and increase from lithium to caesium. *Li* is the lightest known metal. *Na*, *Li* and *K* are lighter than water.

**Note** : □ All elements of group 1st elements are stored in kerosene except lithium because it is a lightest so it float to the surface of kerosene. It is kept wrapped in paraffin wax.

(iv) **Melting and boiling points** : The m. pt. and b. pt. are very low because of the weak bonding in the crystal lattice of the metals. The m. pt. and b. pt. increases moving down the group.

(v) **Ionisation energy** : The ionisation energy decreases on moving down the group.

(vi) **Electropositive character** (the tendency of the element to lose electron) : On account of their low ionisation energies, these metals have a great tendency to lose the  $ns^1$  electron and form positive ions,



Ionisation energy decreases from *Li* to *Cs*, the electropositive character increases in going down from *Li* to *Cs*.

**Note** : □ Due to their strong electropositive nature, they emit electrons even when exposed to light (photo electric effect) (*Cs* and *K*).

(vii) **Oxidation state** : They show only +1 oxidation state. Absence of unpaired electron they are diamagnetic and colourless.

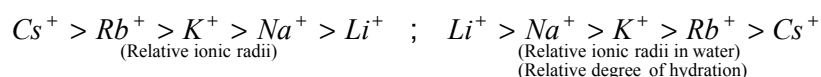
(viii) **Electronegativity** : Their electronegativity values are very low. Electronegativity decreases on moving down the group . (*Li* → *Cs*)

(ix) **Conductivity** : The alkali metals are good conductors of heat and electricity.

(x) **Flame colouration** : The alkali metals and their salts when introduced into the flame, give characteristic colour to flame.

<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>
Crimson Red	Golden Yellow	Pale Violet	Violet	Violet

(xi) **Hydration of ions** : The degree of hydration depends upon the size of the cation. Smaller the size of a cation, greater is its charge density and hence greater is its tendency to withdraw electron from molecules which are thus polarised, Lithium ion, being smallest in size among alkali metal ions, is the most extensively hydrated while  $Cs^+$  ion, the largest alkali metal ion, is the least hydrated.



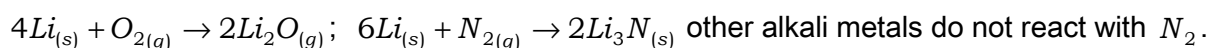
**Note** :  $\square$  Lithium ion, being highly hydrated moves very slowly under the effect of electric current and is thus the poorest conductor of electricity.  $Cs^+ > Rb^+ > K^+ > Na^+ > Li^+$   
Relative electrical conductivity

(xii) **Hydration energy** : Hydration of ions is an exothermic process. The energy released in the hydration of ions is known as hydration energy. Since the degree of hydration of  $M^+$  ions decreases as we go down the group, the hydration energy of alkali metal ions decreases from  $Li^+$  to  $Cs^+$ .

### (3) Chemical properties

(i) **Action of air** : Alkali metals are so reactive that they tarnish rapidly when exposed to air because of the formation of oxides, hydroxides and carbonates at the surface. Hence they are kept under inert liquid like kerosene oil.  $4M_{(s)} + O_{2(g)} \rightarrow 2M_2O_{(s)}$ ;  $M_2O_{(s)} + H_2O_{(l)} \rightarrow 2MOH_{(s)}$ ;  $2MOH_{(s)} \rightarrow M_2CO_{3(s)} + H_2O_{(l)}$

**Note** :  $\square$  When burnt in air  $Li$  reacts with  $O_2$  as well as  $N_2$ :



(ii) **Action of oxygen** :  $4Li + O_2 \rightarrow 2Li_2O$  ;  $2Na + O_2 \rightarrow Na_2O_2$  ;  $K + O_2 \rightarrow KO_2$   
Lithium monoxide Sodium Peroxide Potassium superoxide

The increasing stability of peroxides and superoxides of alkali metals from  $Li$  to  $Cs$  is due to stabilisation of larger anions by larger cations through lattice energy.

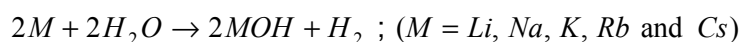
**Note** :  $\square$  The peroxides and superoxides are strong oxidising agents.

$\square$  Peroxides react with water to give  $O_2$ ;  $2Na_2O_2 + 2H_2O \rightarrow 4NaOH + O_2$

$\square$  The peroxides are colourless and diamagnetic while the superoxides are paramagnetic and coloured. Superoxides are orange or yellow solids at room temperature.

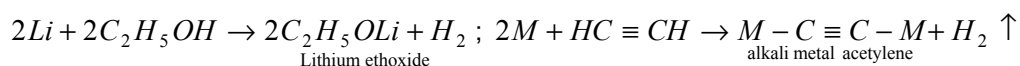
□ The superoxides when heated with sulphur form sulphates;  $2KO_2 + S \rightarrow K_2SO_4$ .

(iii) **Action with water and other compounds containing acidic hydrogen:** Alkali metals react readily and vigorously with water forming hydroxide with the liberation of hydrogen.



Order of reactivity with water :  $Cs > Rb > K > Na > Li$

Alkali metal also react with alcohols and acetylene with the liberation of  $H_2$  e.g.



(iv) **Action with hydrogen :**  $2M + H_2 \xrightarrow{673 K} 2MH$  ( $M = Li, Na, K$  etc.)

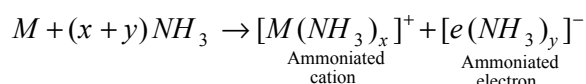
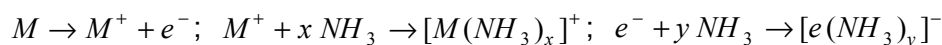
These ionic hydrides have high melting points. Since electropositive character increases from  $Li$  to  $Cs$ , the ease of formation of hydrides decreases from  $Li$  to  $Cs$ .  $Li$  reacts with  $H_2$  at about  $1073 K$ .

**Note** : □ The alkali metal hydrides are attacked by water (or any other proton donor like alcohols,  $NH_{3(g)}$  etc.) to give back hydrogen;  $MH + H_2O \rightarrow MOH + H_2 \uparrow$ .

(v) **Action with halogen :**  $2M + X_2 \rightarrow 2M^+X^-$  ( $X = F, Cl, Br, I$ )

Reactivity of alkali metals towards a particular halogen,  $Cs > Rb > K > Na > Li$ . Reactivity of halogens towards a particular alkali metals,  $F_2 > Cl_2 > Br_2 > I_2$ . All halides of alkali metals (except  $LiF$ ) are highly soluble in water. The poor solubility of  $LiF$  in water is due to its high lattice energy. On the basis of this, the covalent character of various halides is in the order;  $Na^+Cl^- < Mg^{2+}Cl_2^- < Al^{3+}Cl_3^-$ .

(vi) **Solubility in liquid ammonia :** The blue solution of alkali metal in ammonia has some characteristic properties due to formation of ammoniated metal cations and ammoniated electrons in the solution.



The blue colour of the solution is due to the excitation of free ammoniated electron. The free ammoniated electron make the solution a very powerful reducing agent. The blue solution is highly conducting because of the presence of ammoniated electrons and ammoniated cations. The conductivity decreases with increases in

temperature. The blue solution is paramagnetic due to the presence of unpaired electrons in the cavities in ammoniacal solution.

(vii) **Reducing agent** : All alkali metals are strong reducing agent i.e., they have a strong tendency to lose electrons. *Li* to *Cs*, the reducing property should increase which is found to be so except for lithium which is found to be the strongest reducing agent.

(viii) **Alloy formation** : The alkali metals dissolve readily in mercury forming amalgam, the process is highly exothermic.

(ix) **Complex formation** : Lithium has a small size, it forms certain complexes.

(x) **Nature of hydroxides** : Alkali metal hydroxides (except lithium hydroxide) are strongest of all bases, highly soluble in water and are not decomposed on heating;  $2LiOH \xrightarrow{\Delta} Li_2O + H_2O$

**Note** : □ The basic character of alkali metal hydroxides increases on going down the group :



(xi) **Nature of carbonates and bicarbonates** :  $Li_2CO_3$  is unstable towards heat and decomposes to give  $CO_2$ ;  $Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$ .  $LiHCO_3$  does not exist in solid state. The thermal stability of carbonates and bicarbonates increases on moving down the group. The solubility of carbonates and bicarbonates increases when we move down the group. The increasing order of solubility is,  $Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3$

(xii) **Nature of nitrates** :  $4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2 \uparrow$ ;  $2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2 \uparrow$

(xiii) **Nature of sulphates** :  $Li_2SO_4$  is insoluble in water whereas the other sulphates i.e.  $K_2SO_4$ ,  $Na_2SO_4$  are soluble in water.

### Anomalous behaviour of lithium

Lithium, the first member of the alkali metal family shows an anomalous behaviour because of the following reasons :

(a) It has the smallest size (152 pm) in the group. Also size of  $Li^+$  ion is very small.

(b) It has very high ionisation energy and low electropositivity.

- (c) High polarizing power of  $Li^+$  is due to high charge density giving rise to covalent character in bonds.
- (d) Strong metallic bond in lithium due to small size.
- (e) High heat of hydration of  $Li^+$ .
- (f) It has the highest electronegativity (0.98) in the group.
- (g) It has no d-orbitals in the valence shell.

As a result, it differs from the other members of the alkali metal family in following respects :

- (i) Lithium is **harder** than other alkali metals, due to strong metallic bond.
- (ii) Lithium combines with oxygen to form lithium oxide,  $Li_2O$  whereas the other alkali metals form preferably the peroxides,  $M_2O_2$  and superoxides,  $MO_2$ .

- (iii) Lithium unlike the other alkali metals, reacts with nitrogen to form the nitride;  $6Li + N_2 \rightarrow 2Li_3N$   
Lithium nitride

This property is used for removing nitrogen from a mixture of gases.

- (iv)  $LiOH$  is a weak base and decomposes to give the corresponding oxide while the hydroxides of alkali metals are stable to heat and sublime as such;  $2LiOH \xrightarrow{900^\circ C} Li_2O + H_2O$

- (v)  $Li_2O, LiOH, Li_2CO_3$  and  $LiF$  are less soluble in water as compared to the corresponding salts of  $Na$  and  $K$ .

- (vi)  $Li^+$  being very small in size is heavily hydrated as compared to other alkali metal ions. The effective size of  $Li^+$  in aqueous solution is the largest.

## Group – II (Alkaline earth Metals)

The group 2 of the periodic table consists of six metallic elements. These are beryllium ( $Be$ ), magnesium ( $Mg$ ), calcium ( $Ca$ ), strontium ( $St$ ), barium ( $Ba$ ) and radium ( $Ra$ ). These (except  $Be$ ) are known as alkaline earth metals as their oxides are alkaline and occur in earth crust.

**General characteristics.****(1) Electronic configuration**

Element	Electronic configurations ( $ns^2$ )
${}_4\text{Be}$	$1s^2 2s^2$ or $[\text{He}]2s^2$
${}_{12}\text{Mg}$	$1s^2 2s^2 2p^6 3s^2$ or $[\text{Ne}]3s^2$
${}_{20}\text{Ca}$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ or $[\text{Ar}]4s^2$
${}_{38}\text{Sr}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$ or $[\text{Kr}]5s^2$
${}_{56}\text{Ba}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$ or $[\text{Xe}]6s^2$
${}_{88}\text{Ra}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^6 7s^2$ or $[\text{Rn}]7s^2$

**Note** :  $\square$  Radium was discovered in the ore pitch blende by madam Curie. It is radioactive in nature.

**(2) Physical properties**

(i) **Physical state** : They are silvery white and harder than alkali metals.

(ii) **Atomic and ionic radii** – The atomic as well as ionic radii go on increasing down the group due to the gradual addition of extra energy level and also because of the screening effect.

(iii) **Density** : The density decreases slightly on moving down the group up to calcium then it increases there after upto radium.

**Note** :  $\square$  Density of *Ca* is less than that of *Mg* due to the presence of vacant 3*d*-subshell leading to much increase in atomic volume.

(iv) **Melting and boiling points** : The melting point and boiling points decrease down the group with the exception of magnesium. This is because due to increase in atomic size the strength of the metallic bond decreases.

(v) **Ionisation energy** : On moving down the group due to increase in atomic size the magnitude of ionisation energy decreases.

(vi) **Electropositive character** : On moving down the group, the electropositive character increases due to increase in atomic radii and decrease in ionisation energy .

**Note** : □ Unlike alkali metals, these elements do not emit electrons on exposure to light.

(vii) **Oxidation state** : Alkaline earth metals, exhibit only +2 oxidation state. The divalent ions of alkaline earth metals have stable noble gas configuration. In the solid state, the divalent ions  $M^{2+}$  form strong lattice due to their small size and high charge (i.e., high lattice energy).

**Note** : □ Since the divalent ions have no unpaired electron, these are diamagnetic and colourless. Their compounds are colourless provided their anions are also colourless.

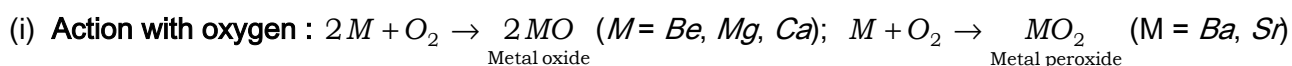
(viii) **Conductivity** : Due to the presence of two loosely held valence electrons, which are free to move throughout the crystal structure, the alkaline earth metals, are good conductors of heat and electricity.

(ix) **Flame colouration** : Alkaline earth metals, except *Be* and *Mg* impart characteristic colours to the flame.

<i>Be</i>	<i>Mg</i>	<i>Ca</i>	<i>Sr</i>	<i>Ba</i>	<i>Ra</i>
No colour	No colour	Brick red	Crimson red	Grassy green	Crimson

(x) **Nature of the compounds** : Beryllium and magnesium generally form covalent compounds due to their smaller atomic size and high ionisation energy (i.e. lesser electron releasing tendency) while the compounds of other members of the family are ionic in character and this character increases down the group.

### (3) Chemical properties



The reactivity with oxygen increases as we move down the group 2 elements because their electropositive character increases.

(ii) **Action with water** : Group 2 elements are less reactive with water as compared to alkali metals. The reaction with water becomes increasingly vigorous on moving down the group. This is supported by the fact that beryllium does not react with water, magnesium reacts only with hot water while other metals *Ca*, *Sr* and



*Ba* react with cold water but not as vigorously as the alkali metals;  $Mg + H_2O \xrightarrow{\text{boil}} \underset{\text{Magnesium oxide}}{MgO} + H_2O$ .

Order of reactivity with water :  $Ba > Sr > Ca > Mg$ .

(iii) **Reaction with hydrogen** : Except beryllium, all other elements combine with hydrogen on heating to form hydride ( $MH_2$ );  $M + H_2 \xrightarrow{\text{Heat}} MH_2$ . Beryllium hydride and magnesium hydride are covalent in nature while *Ca*, *Sr* and *Ba* hydrides are ionic in nature.

Note : □ The hydride of beryllium can be prepared indirectly by reducing beryllium chloride with lithium aluminium hydride  $2 BeCl_2 + LiAlH_4 \rightarrow 2 BeH_2 + LiCl + AlCl_3$ .

(iv) **Reaction with halogen** : All elements of group II combine with halogens at high temperature, forming their halides ( $MX_2$ );  $M + X_2 \xrightarrow{\text{Heat}} MX_2$ . The nature of beryllium halide is covalent while the rest of the elements of this group form ionic halides. The halides are soluble in water. The solubility of halides in water decreases down the group fluorides of alkaline earth metals are almost insoluble in water.

(v) **Formation of amalgams and alloys** : Alkaline earth metals form amalgams with mercury and alloys with other metals.

(vi) **Solubility in liquid ammonia** : Like alkali metals, alkaline earth metals dissolve in liquid ammonia giving coloured solutions. When the metal ammonia solutions are evaporated, hexammoniates,  $M(NH_3)_6$  are formed. The ammoniates are good conductors of electricity and decompose at high temperatures. The tendency to form ammoniated decreases with increase in size of the metal atom *i.e.*, on moving down the group.

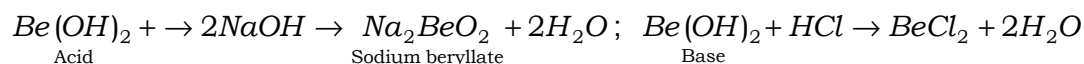
(vii) **Reducing agents** : All the alkaline earth metals are strong reducing agents as these can release the two valence electrons and change into dipositive ions with noble gas configuration;  $M \rightarrow M^{2+} + 2e^-$ .

Since the ionisation energy decreases (*i.e.*, electropositive character increases) from *Be* to *Ba* the reducing character of alkaline earth metals increases from *Be* to *Ba*.

(viii) **Complex formation** : This tendency is greatest, however, in case of beryllium since the size of  $Be^{+2}$  ion is the smallest in this group.

(ix) **Basic nature of hydroxides** : Except  $Be(OH)_2$ , all other hydroxides are basic in nature and their basic character goes on increasing on moving down the group. The basic character of hydroxides of group 2

elements is lesser than those of group 1 hydroxides. Beryllium hydroxide is *amphoteric*. It reacts with caustic alkalis and mineral acids and behaves both as an acid and as base.

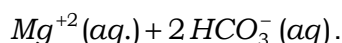


(x) **Solubility of the salts** : (Hydroxides, sulphates, carbonates and halides). The solubility of a salt in water depends upon *Lattice energy and Hydration energy*.

The solubility of hydroxides of group II metals increases from *Be* to *Ba*.

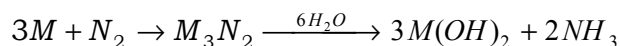
The solubility of the sulphates of group 2 metals in water decreases from *Be* to *Ba*. The solubility decreases with increase in the ionic size.

The solubility of group II metal carbonates in water decreases down the group. It is due to the decrease in the magnitude of hydration energy. However these insoluble metal carbonates dissolve in water in the presence of carbon dioxide,  $\underset{\text{insoluble}}{\text{MgCO}_3(\text{s})} + \text{CO}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \underset{\substack{\text{Magnesium} \\ \text{bicarbonate} \\ \text{(soluble)}}}{\text{Mg}(\text{HCO}_3)_2(\text{aq})}; \quad \text{Mg}(\text{HCO}_3)_2 \rightleftharpoons$

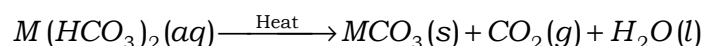


The halides of beryllium ( $\text{BeX}_2$ ) are covalent in nature on account of small size of  $\text{Be}^{2+}$  ion. These halides are soluble in organic solvents whereas the halides of other alkaline earth metals are ionic solids so these are water soluble. The solubility in water decreases from *Mg* to *Ba*.

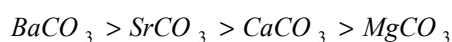
(xi) **Action with nitrogen** : These metals burn in nitrogen to form nitrides of the type  $\text{M}_3\text{N}_2$  which are hydrolysed with water to evolve  $\text{NH}_3$ .  $\text{Be}_3\text{N}_2$  is volatile in nature.



(xii) **Nature of bicarbonates and carbonates** : Bicarbonates of these metals do not exist in solid state but are known in solution only. When these solutions are heated, bicarbonates get decomposed to evolve  $\text{CO}_2$ .



Thermal stability of carbonates and bicarbonates decreases down the group :



### Anomalous behaviour of beryllium

Beryllium, the first member of group 2 differs from the rest of the members of its group due to the following reason :

- (a) Beryllium has a small atomic size as well as small ionic size. (b) Its hydration energy is high.
- (c) It has high first and second ionisation energies in comparison to other members of the family.
- (d) It has no vacant  $d$ -orbitals. (e) It has high electronegativity value. (f) It has a high charge density.

### Some important points.

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#### Group – I (Alkali Metals)

- (1) Lithium does not form alum due to its small size.
- (2)  $Li$  metal is used as scavenger in metallurgy to remove  $O_2$  and  $N_2$  gases.
- (3) All alkali metals form body centered cubic lattice with coordination number 8.
- (4)  $K^+$  is detected in the laboratory by using picric acid or tartaric acid.
- (5) With  $NH_3$  lithium forms imide  $Li_2NH$  while other alkali metals form amides  $MNH_2$ .
- (6) Francium has 19 isotopes.

#### Group – II (Alkaline earth Metals)

- (1) From  $Be$  to  $Ba$  : density decreases, Density decreases, Metallic nature increases, Reactivity increases, Reducing nature increases.
- (2)  $Ba(NO_3)_2$  is used in the preparation of green fire.
- (3) The ionisation potential of radium is higher than that of barium.
- (4) The polarising power decreases from  $Be^{2+}$  to  $Ba^{2+}$  ion. Thus, beryllium has a tendency to form covalent compounds.
- (5) Magnesium exists as a natural complex, chlorophyll where it is complexed with pyrrole rings of porphyrin.
- (6) Anhydrous  $BeCl_2$  acts as a lewis acid.
- (7) Aqueous suspension of  $Mg(OH)_2$  is called **Milk of magnesia** and it is used as antacid.
- (8) Aqueous  $Ba(OH)_2$  is known as **baryta water**.

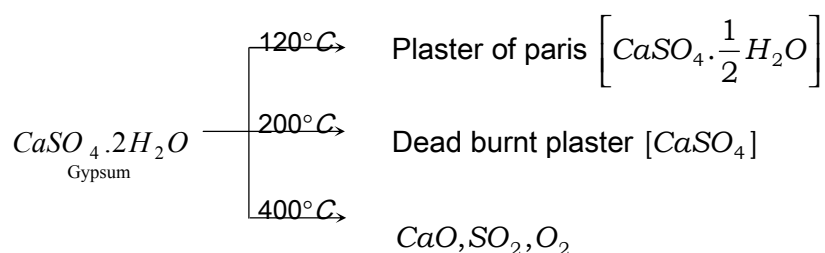
(9) Halides of group 2 elements are hygroscopic and form hydrates like  $MgCl_2 \cdot 6H_2O$ ,  $CaCl_2 \cdot 6H_2O$ ,  $SrCl_2 \cdot 6H_2O$ ,  $BaCl_2 \cdot 2H_2O$  etc.

(10) Compounds of  $Mg$  are both covalent and ionic.  $MgBr_2$  and  $MgI_2$  due to their covalent nature are soluble in acetone.

(11) Burning  $Mg$  wire continues to burn in atmosphere of  $N_2$  and  $CO_2$  due to its high electropositive nature and release sufficient amount of heat to decompose the carbon dioxide.

(12) Chlorides and sulphates of  $Ca$  and  $Mg$  are responsible for permanent hardness of natural water whereas their bicarbonates cause temporary hardness.

(13) Gypsum when heated to different temperature gives different products.

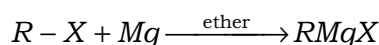


(14) Anhydrous  $CaCl_2$  is a good desiccant (drying agent) due to its hygroscopic nature ( $CaCl_2 \cdot 2H_2O$ ). It cannot be used to dry alcohol or ammonia as it forms addition products with them.

(15) On hydrolysis  $Mg_2C_3$  (magnesium allylide) gives  $C_3H_4$  (allylene).

(16)  $Mg$  burns in air releasing dazzling light rich in U. V. rays.

(17)  $Mg$  dissolves in dry ether in presence of alkyl halide ( $R - X$ ) to form **Grignard's reagent**.



(18) Bicarbonates of alkaline earth metals do not exist in solid state but are known in solution only.

(19)  $CaCl_2 \cdot 6H_2O$  when mixed with ice gives a freezing mixture ( $-54^\circ C$ ).

(20)  $CaC_2O_4 \cdot 2H_2O$  is Insoluble in bases but soluble in dilute strong acids. This oxalate is a component of most kidney stones.

(21)  $Be$  does not form a peroxide.

(22) Barium sulphate is insoluble in water. The insolubility of barium sulphate is used for detecting obstruction in the digestive system by the technique commonly known as Barium meal. The presence of  $BaSO_4$  in the stomach helps in getting X-rays pictures because of the greater scattering power of it.

(23)  $CaH_2$  is known as hydrolith.

(24)  $Be$  reacts slowly with acids,  $Mg$  reacts at a faster rate while  $Ca$ ,  $Sr$  and  $Ba$  react explosively with acids.

(25) Only  $Mg$  displaces hydrogen from a very dilute  $HNO_3$ .

(26) Beryllium due to small size forms complexes of type  $[BeF_4]^-$ ,  $[BeF_4]^{2-}$ ,  $[Be(H_2O)_4]^{2+}$ .

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