



Alkali and Alkaline Earth Metals



Learning Objectives

After studying this unit, students will be able to

- ♦ Explain the properties of alkali metals and alkaline earth metals
- ♦ Recognise the anomalous properties of Li and Be
- ♦ List the uses of alkali metals and alkaline earth metals
- ♦ Describe the general characteristics of compounds of alkali metals and alkaline earth metals
- ♦ Appreciate the biological importance of sodium and potassium, Magnesium and Calcium
- ♦ Explain the preparation, properties and uses of calcium oxide, calcium hydroxide, gypsum and plaster of paris.



Rock salt



Sylvite



Spodumene

5.1 s-Block Elements:

The elements belonging to the group 1 and 2 in the modern periodic table are called s-block elements. The elements belonging to these two groups are commonly known as alkali and alkaline earth metals respectively. In this unit, we study their properties, uses, important compounds and biological importance.

5.2 Alkali metals:

The word “alkali” is derived from the word al-qaliy meaning the plant ashes, referring to the original source of alkaline substances. A water-extract of burnt plant ashes, called potash contain mainly potassium carbonate. Alkali metal group consists of the elements: lithium, sodium, potassium, rubidium, caesium and francium. They are all metals, generally soft and highly reactive. They form oxides and hydroxides and these compounds are basic in nature.

5.2.1 General characteristics of alkali metals:

Alkali metals are highly reactive and are found in nature only as compounds. Rubidium and caesium are found associated in minute quantities with minerals of other alkali metals. Francium is radioactive and does not occur appreciably in nature. Francium is highly radioactive; its longest-lived isotope has a half-life of only 21 minutes.

Table 5.1 Abundance of important alkali metals and their sources

Elements	Abundance in earth crust (%)	Relative Abundance	Mineral source
Lithium	0.0018	35	Spodumene [$\text{LiAl}(\text{SiO}_3)$]
Sodium	2.27	7	Rock Salt [NaCl]
Potassium	1.84	8	Sylvite [KCl]
Rubidium	0.0078	23	No convenient Source (obtained as by product of lithium processing)
Cesium	0.00026	46	



Figure 5.1 Alkali metals Li, Na and K stored under oil

Electronic configuration

The general valence shell electronic configuration of alkali metals is ns^1 , where 'n' represents the period number.

Table 5.2 Electronic configuration of alkali metals

Element	Symbol	Atomic No.	Electronic configuration
Lithium	Li	3	[He]2s ¹
Sodium	Na	11	[Ne]3s ¹
Potassium	K	19	[Ar]4s ¹
Rubidium	Rb	37	[Kr]5s ¹
Caesium	Cs	55	[Xe]6s ¹
Francium	Fr	87	[Rn]7s ¹

Common oxidation state

All these elements are highly electropositive in nature. They readily lose their valence electron to give monovalent cations (M^+). Alkali metals have only one oxidation state which is +1.

Atomic and ionic radii

Being the first element of each period, alkali metals have the largest atomic and ionic radii in their respective periods. On moving down the group, there is an increase in the number of shells and, therefore, atomic and ionic radii increase. The monovalent ions (M^+) are smaller than the respective parent atoms as expected.

Table 5.3 Physical properties of alkali metals

Physical property	Li	Na	K	Rb	Cs
Atomic radius / Metallic radius (Å)	1.52	1.86	2.27	2.48	2.65
Ionic radius (Å)	0.76	1.02	1.38	1.52	1.67
Melting point (°C)	181	98	63	39	28.5
Boiling point (°C)	1347	881	766	688	705
First ionization enthalpy (kJ mol ⁻¹)	520.2	495.8	418.8	403.0	375.7
Electronegativity (Paulings scale)	1.0	0.9	0.8	0.8	0.7
Density (g cm ⁻³)	0.54	0.97	0.86	1.53	1.90
Standard potential E^0 for M^+/M (V)	-3.04	-2.71	-2.92	-2.93	-2.93
Hydration enthalpy (kJ mol ⁻¹)	-506	-406	-330	-310	-276

Ionisation enthalpy

Alkali metals have the lowest ionisation enthalpy compared to other elements present in the respective period. As we go down the group, the ionisation enthalpy decreases due to the increase in atomic size. In addition, the number of inner shells also increases, which in turn increases the magnitude of screening effect and consequently, the ionisation enthalpy decreases down the group.

The second ionisation enthalpies of alkali metals are very high. The removal of an electron from the alkali metals gives monovalent cations having stable electronic configurations similar to the noble gas. Therefore, it becomes very difficult to remove the second electron from the stable configurations already attained.

Hydration enthalpy

Lithium salts are more soluble than the salts of other metals of group 1. eg. LiClO_4 is upto 12 times more soluble than NaClO_4 . Other salts KClO_4 , RbClO_4 and CsClO_4 have solubilities only 10^{-3} times of that of LiClO_4 . The high solubility of Li salts is due to strong solvation of small size Li^+ ion.

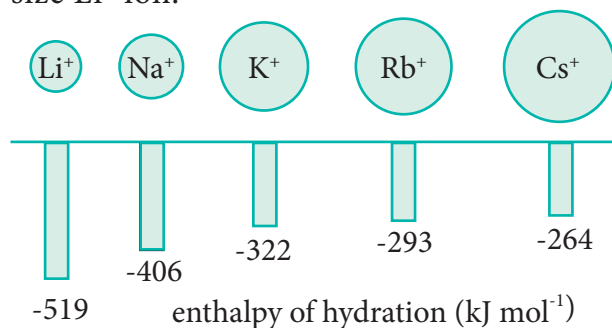


Figure 5.2 Hydration enthalpy of alkali metals

Electronegativity:

Alkali metals have comparatively smaller value of electronegativity than the other elements in the respective period. When they react with other elements, they usually produce ionic compounds. For example, they react with halogens to form ionic halides.

Flame colour and the spectra:

When the alkali metal salts moistened with concentrated hydrochloric acid are heated on a platinum wire in a flame, they show characteristic coloured flame as shown below.

Table 5.4 Flame colour and wavelength

Element	Colour	Wavelength (nm)
Lithium	Crimson red	670.8
Sodium	Yellow	589.2
Potassium	Lilac (violet)	766.5
Rubidium	Reddish violet	780.0
Caesium	Blue	455.5

The heat in the flame excites the valence electron to a higher energy level. When it drops back to its actual energy level, the excess energy is emitted as light, whose wavelength is in the visible region as shown in the above table.

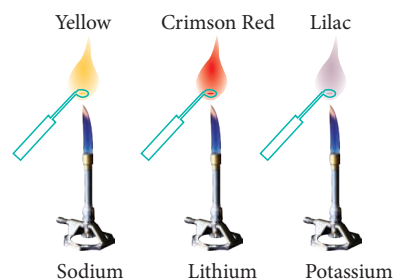


Figure 5.3 Flame colours of alkali metal salts

5.2.2 Distinctive behavior of lithium

The distinctive behaviour of Li^+ ion is due to its exceptionally small size, high polarising power, high hydration energy and non availability of d-orbitals.

Table 5.5 Comparison of properties of lithium with other elements of the group:

Lithium	Other elements of the family
Hard, high melting and boiling point	Soft and Lower melting and boiling point
Least reactive (For example it reacts with oxygen to form normal oxide, forms peroxides with great difficulty and its higher oxides are unstable)	More reactive
Reacts with nitrogen to give Li_3N $6\text{Li} + \text{N}_2 \rightarrow 2\text{Li}_3\text{N}$	No reaction
Reacts with bromine slowly	React violently
Reacts directly with carbon to form ionic carbides. For example $2\text{Li} + 2\text{C} \rightarrow \text{Li}_2\text{C}_2$	Do not react with carbon directly, but can react with carbon compounds. $2\text{Na} + \text{C}_2\text{H}_2 \rightarrow \text{Na}_2\text{C}_2 + \text{H}_2$
Lithium nitrate decomposes to give an oxide	Decompose to give nitrites

Table 5.6 Similarities between lithium and Magnesium

S.No.	Properties
1	Both lithium and magnesium are harder than other elements in the respective groups
2	Lithium and magnesium react slowly with water. Their oxides and hydroxides are much less soluble and their hydroxides decompose on heating.
3	Both form a nitride, Li_3N and Mg_3N_2 , by direct combination with nitrogen
4	They do not give any superoxides and form only oxides, Li_2O and MgO
5	The carbonates of lithium and magnesium decompose upon heating to form their respective oxides and CO_2 .
6	Lithium and magnesium do not form bicarbonates.
7	Both LiCl and MgCl_2 are soluble in ethanol and are deliquescent. They crystallise from aqueous solution as hydrates, $\text{LiCl} \cdot 2\text{H}_2\text{O}$ and $\text{MgCl}_2 \cdot 8\text{H}_2\text{O}$

Diagonal Relationship:

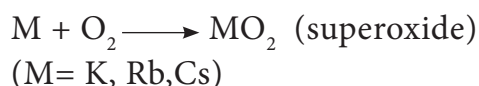
Similarity between the first member of group 1 (Li) and the diagonally placed second element of group 2 (Mg) is called diagonal relationship. It is due to similar size ($r_{\text{Li}^+} = 0.766 \text{ \AA}$ and $r_{\text{Mg}^{2+}} = 0.72 \text{ \AA}$) and comparable electronegativity values (Li = 1.0; Mg = 1.2).

5.2.3 Chemical properties of alkali metals

Alkali metals exhibit high chemical reactivity. The reactivity of alkali metals increases from Li to Cs, since the ionisation energy decreases down the group. All alkali metals are highly reactive towards the more electronegative elements such as oxygen and halogens. Some characteristic chemical properties of alkali metals are described below.

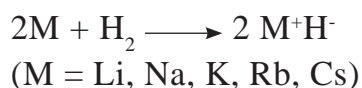
Reaction with oxygen

All the alkali metals on exposure to air or oxygen burn vigorously, forming oxides on their surface. Lithium forms only monoxide, sodium forms the monoxide and peroxide and the other elements form monoxide, peroxide, and superoxides. These oxides are basic in nature.



Reaction with hydrogen

All alkali metals react with hydrogen at about 673 K (lithium at 1073 K) to form the corresponding ionic hydrides. Reactivity of alkali metals with hydrogen decreases from Li to Cs.

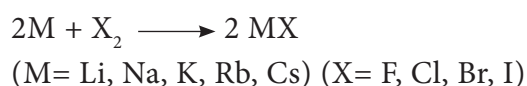


The ionic character of the hydrides increases from Li to Cs and their stability decreases. The hydrides behave as strong

reducing agents and their reducing nature increases down the group.

Reaction with halogen

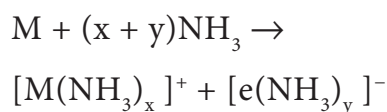
Alkali metals combine readily with halogens to form ionic halides MX. Reactivity of alkali metals with halogens increases down the group because of corresponding decrease in ionisation enthalpy.



All metal halides are ionic crystals. However Lithium iodide shows covalent character, as it is the smallest cation that exerts high polarising power on the iodide anion. Additionally, the iodide ion being the largest can be polarised to a greater extent by Li^+ ion.

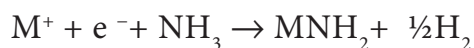
Reaction with liquid ammonia:

Alkali metals dissolve in liquid ammonia to give deep blue solutions that are conducting in nature. The conductivity is similar to that of pure metals (The specific conductivity of Hg is $10^4 \Omega^{-1}$ and for sodium in liquid ammonia is $0.5 \times 10^4 \Omega^{-1}$). This happens because the alkali metal atom readily loses its valence electron in ammonia solution. Both the cation and the electron are ammoniated to give ammoniated cation and ammoniated electron.



The blue colour of the solution is due to the ammoniated electron which

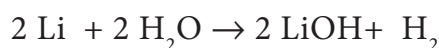
absorbs energy in the visible region of light and thus imparts blue colour to the solution. The solutions are paramagnetic and on standing slowly liberate hydrogen resulting in the formation of an amide.



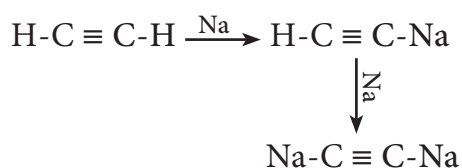
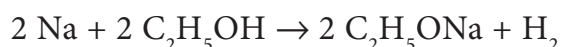
In concentrated solution, the blue colour changes to bronze colour and become diamagnetic.

Reaction with water:

Alkali metals react with water to give corresponding hydroxides with the liberation of hydrogen.



They also react with alcohol, and alkynes which contain active hydrogens.



Reducing activity:

Alkali metals can lose their valence electron readily hence they act as good reducing agents.



Reaction with carbon:

Lithium directly reacts with carbon to form the ionic compound, lithium carbide. Other metals do not react with carbon directly. However, when they are treated with compounds like acetylene they form acetelydes.



5.2.4 Uses of alkali metals:

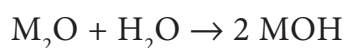
- Lithium metal is used to make useful alloys. For example with lead it is used to make 'white metal' bearings for motor engines, with aluminium to make aircraft parts, and with magnesium to make armour plates. It is used in thermonuclear reactions.
- Lithium is also used to make electrochemical cells.
- Sodium is used to make Na/Pb alloy needed to make $Pb(Et)_4$ and $Pb(Me)_4$. These organolead compounds were earlier used as anti-knock additives to petrol, but nowadays lead-free petrol in use.
- Liquid sodium metal is used as a coolant in fast breeder nuclear reactors.
- Potassium has a vital role in biological systems. Potassium chloride is used as a fertilizer.
- Potassium hydroxide is used in the manufacture of soft soap. It is also used as an excellent absorbent of carbon dioxide.
- Caesium is used in devising photoelectric cells.

5.3 General characteristics of the compounds of alkali metals

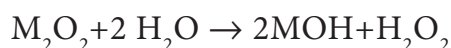
All the common compounds of the alkali metals are generally ionic in nature. General characteristics of some of their compounds are discussed here.

Oxides and Hydroxides

On combustion in excess of air, alkali metals form normal oxides with formula M_2O . They react with water to form corresponding hydroxides which are basic in nature.

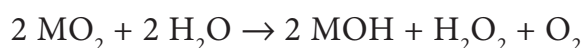


Alkali metals apart from lithium form peroxides in addition to normal oxides upon combustion with excess air. These peroxides produce hydroxides and H_2O_2 upon reacting with water.



(M = Na, K, Rb, Cs)

Except lithium and sodium, all the other alkali metals form superoxides also. These superoxides also give basic hydroxides upon treatment with water.



(M = K, Rb, Cs)

Under appropriate conditions pure compounds M_2O , M_2O_2 or MO_2 may be prepared.

Properties of oxides and hydroxides:

The oxides and the peroxides are colourless when pure, but the superoxides are yellow or orange in colour. The peroxides are diamagnetic while the superoxides are paramagnetic. Sodium peroxide is widely used as an oxidising agent. The hydroxides which are obtained by the reaction of the oxides with water are all white crystalline solids. The alkali metal hydroxides are strong bases. They

dissolve in water with evolution of heat on account of intense hydration.

Halides:

The alkali metal halides, MX , (X=F, Cl, Br, I) are colourless crystalline solids with high melting points. They can be prepared by the reaction of the appropriate oxide, hydroxide or carbonate with aqueous hydrohalic acid (HX). As the electropositive character of alkali metal increases from Li to Cs, the ease with which the metals form halides increases from Li to Cs. All halides are ionic in nature except LiBr and LiI. Except LiF, all other halides are soluble in water. The low solubility of LiF in water is due to its high lattice enthalpy (small size of Li^+ and F^-). Due to the presence of covalent nature both LiBr and LiI are soluble in organic solvents.

Salts of oxo-acids

Alkali metals form salts with all the oxo-acids. Most of these salts are soluble in water and are thermally stable. As the electropositive character increases down the group, the stability of the carbonates and bicarbonates increases. This is due to the decrease in polarising power of alkali metal cations. The carbonates (M_2CO_3) of alkali metals are remarkably stable up to 1273 K, above which they first melt and then eventually decompose to form oxides. However, Li_2CO_3 is considerably less stable and decomposes readily.



This is presumably due to large size difference between Li^+ and CO_3^{2-} which

makes the crystal lattice unstable. Being strongly basic, alkali metals except lithium form solid bicarbonates. No other metal forms solid bicarbonates.



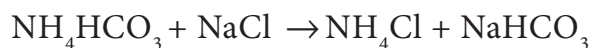
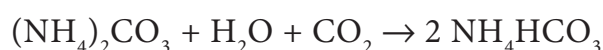
(M = Na, K, Rb, Cs)

All the carbonates and bicarbonates are soluble in water and their solubilities increase rapidly on descending the group. This is due to the reason that lattice energies decrease more rapidly than their hydration energies on moving down the group.

5.3.1 Important compounds of alkali metals:

Sodium Carbonate $Na_2CO_3 \cdot 10H_2O$ (Washing soda):

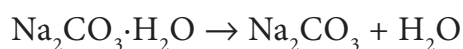
Sodium carbonate is one of the important inorganic compounds used in industries. It is prepared by Solvay process. In this process, ammonia is converted into ammonium carbonate which then converted to ammonium bicarbonate by passing excess carbon dioxide in a sodium chloride solution saturated with ammonia. The ammonium bicarbonate thus formed reacts with the sodium chloride to give sodium bicarbonate and ammonium chloride. As sodium bicarbonate has poor solubility, it gets precipitated. The sodium bicarbonate is isolated and is heated to give sodium carbonate. The equations involved in this process are,



The ammonia used in this process can be recovered by treating the resultant ammonium chloride solution with calcium hydroxide. Calcium chloride is formed as a by-product.

Properties:

Sodium carbonate, commonly known as washing soda, crystallises as decahydrate which is white in colour. It is soluble in water and forms an alkaline solution. Upon heating, it loses the water of crystallisation to form monohydrate. Above 373 K, the monohydrate becomes completely anhydrous and changes to a white powder called soda ash.



Uses:

- Sodium carbonate known as washing soda and is used mainly for laundering.
- It is also used in water treatment to convert the hard water to soft water.
- It is used in the manufacturing of glass, paper, paint etc...

Sodium chloride NaCl (Cooking salt or Table salt):

Sodium chloride is isolated by evaporation from sea water which contains

2.7 to 2.9% by mass. Approximately 50 lakh tons of salt are produced annually in India by solar evaporation. Crude sodium chloride can be obtained by crystallisation of brine solution which contains sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride as impurities. Pure sodium chloride can be obtained from crude salt as follows. Firstly removal of insoluble impurities by filtration from the crude salt solution with minimum amount of water. Sodium chloride can be crystallised by passing HCl gas into this solution. Calcium and magnesium chloride, being more soluble than sodium chloride, remain in solution.

Sodium chloride melts at 1081K. It has a solubility of 36.0 g in 100 g of water at 273 K. The solubility does not increase appreciably with increase in temperature.

Uses :

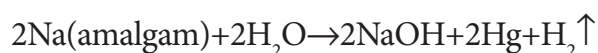
- (i) It is used as a common salt or table salt for domestic purpose.
- (ii) It is used for the preparation of many inorganic compounds such as NaOH and Na_2CO_3

Sodium hydroxide:

Sodium hydroxide is prepared commercially by the electrolysis of brine solution in Castner-Kellner cell using a mercury cathode and a carbon anode. Sodium metal is discharged at the cathode and combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode. The sodium amalgam thus obtained is treated with water to give sodium hydroxide.

At cathode : $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}(\text{amalgam})$

At anode : $\text{Cl}^- \rightarrow \frac{1}{2} \text{Cl}_2 \uparrow + \text{e}^-$



Sodium hydroxide is a white, translucent and deliquescent solid, that dissolves in water to give a strong alkaline solution. It melts at 591 K. The sodium hydroxide solution at the surface reacts with the CO_2 in the atmosphere to form Na_2CO_3

Uses:

- Sodium hydroxide is used in the purification of bauxite (ore of Aluminium) and petroleum refining
- It is used in the textile industries for mercerising cotton fabrics
- It is used in the manufacture of soap, paper and artificial silks.

Sodium bicarbonate NaHCO_3 (Baking soda):

Sodium hydrogen carbonate or sodium bicarbonate is used in baking cakes pastries etc. It is called so because it decomposes on heating to generate bubbles of carbon dioxide, leaving holes in cakes or pastries and making them light and fluffy. This compound is prepared by saturating a solution of sodium carbonate with carbon dioxide. The white crystalline powder of sodium bicarbonate, being less soluble, precipitated out.

Uses:

- Primarily used as an ingredient in baking.
- Sodium hydrogen carbonate is a mild antiseptic for skin infections.
- It is also used in fire extinguishers.

5.4 Biological importance of sodium and potassium

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

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.

Sodium-potassium pump play an important role in transmitting nerve signals.

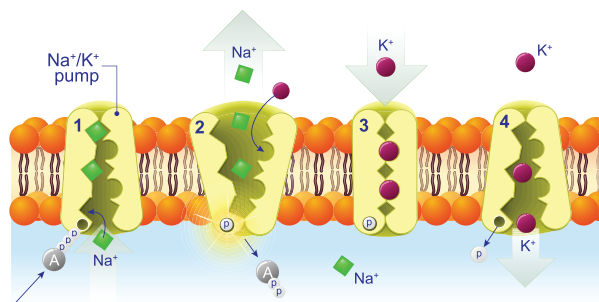


Figure 5.4 Sodium-potassium pump

5.5 Alkaline earth metals

Group 2 in the modern periodic table contains the elements beryllium, magnesium, calcium, strontium, barium and radium. These elements with the exception of beryllium are commonly known as the alkaline earth metals because their oxides and hydroxides are alkaline in nature and these metal oxides are found in the earth's crust.



Figure 5.5 Alkaline earth metals

Table 5.7 Abundance of important alkaline earth metals and their sources

Element	Abundance in Earth crust by weight (ppm)	Mineral source
Be	2.0	beryl $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$
Mg	27640	Carnallite ($\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$) Dolomite $\text{MgCO}_3 \cdot \text{CaCO}_3$
Ca	1.84	Fluorapatite $\text{Ca}_5(\text{PO}_4)_3\text{F}$
Sr	384	Celestite SrSO_4
Ba	390	barytes BaSO_4

5.5.1 General characteristics of alkaline earth metals

Physical state

Beryllium is rare and radium is the rarest of all comprising only 10 % of igneous rocks. Magnesium and calcium are very common in the earth's crust, with calcium the fifth-most-abundant element, and magnesium the eighth. Magnesium and calcium are found in many rocks and minerals: magnesium in carnallite, magnesite, dolomite and calcium in chalk, limestone, gypsum. Most strontium is found in the minerals celestite and strontianite. Barium is slightly less common, much of it in the mineral barite. Radium, being a decay product of uranium, is found in all uranium-bearing ores.

FIREWORK



Many alkaline and alkaline earth metals are used in creating colours, such as strontium and barium, are the colourful stars of a fireworks show. Combined with the element chlorine, barium sends up a green spark, calcium gives orange and lithium gives medium red. Strontium carbonate gives a bright red colour. Nitrates of sodium gives orange, potassium and rubidium gives violet colour and caesium gives indigo colour. The burning "excites" the electrons, pushing them into higher than normal energy level; they release their extra energy as a colourful burst of light



The blue fireworks are the hardest to make, since the compound copper chloride breaks down in a hot flame. In recent years, fireworks experts have used magnalium- a mixture of the alkaline earth metal magnesium and aluminium - to boost all firework colours. Magnalium has made the blues brighter, but pyro technicians are still searching for a blue as brilliant as the red, green and yellow colours.

Electronic configuration

These elements have two electrons in the valence shell of their atoms, preceded by the noble gas configuration. Their general electronic configuration is written

as [Noble gas] ns^2 where 'n' represents the valence shell.

Table 5.8 Electronic configuration of alkaline earth metals

Element	Atomic No.	Electronic configuration
Be	4	[He] $2s^2$
Mg	12	[Ne] $3s^2$
Ca	20	[Ar] $4s^2$
Sr	38	[Kr] $5s^2$
Ba	56	[Xe] $6s^2$
Ra	88	[Rn] $7s^2$

Atomic and ionic radii

The atomic and ionic radii of alkaline earth metals are smaller than the corresponding members of the alkali metals. This is due to the fact the Group 2 elements having a higher nuclear charge that allows electrons to be attracted more strongly towards the nucleus. On moving down the group, the radii increases due to gradual increase in the number of the shells and the screening effect.

Common oxidation state

The group 2 elements have two electrons in their valence shell and by losing these electrons, they acquire the stable noble gas configuration. So these elements exhibit +2 oxidation state in their compounds.

Ionisation enthalpy

Due to a fairly large size of the atoms, alkaline earth metals have low ionisation enthalpies when compared to 'p' block elements. Down the group the ionisation enthalpy decreases as atomic size increases. This is due to the addition of new shells as well as increase in the magnitude of the screening effect of inner shell electrons. Members of group 2 have higher ionization enthalpy values than group 1 because of their smaller size, with electrons being more attracted towards the nucleus of the atoms. Correspondingly they are less electropositive than alkali metals.

Table 5.9 Physical properties of alkaline earth metals

Physical property	Be	Mg	Ca	Sr	Ba
Atomic radius -non bonded (Å)	1.12	1.60	1.97	2.15	2.22
Ionic radius (Å)	0.27(0.31)	0.72	1.00	1.18	1.35
First ionization energy (kJ mol ⁻¹)	899.5	737.8	589.8	549.5	502.9
Second ionization energy (kJ mol ⁻¹)	1757.1	1450.7	1145.5	1064.2	965.2
Hydration enthalpy (kJ mol ⁻¹)	- 2494	- 1921	-1577	- 1443	- 1305
Melting Point (°C)	1287	651	851	789	729
Boiling Point (°C)	2472	1090	1494	1382	1805
Density (g cm ⁻³)	1.84	1.74	1.55	2.63	3.59
Standard Potential E^\ominus for M^+/M (V)	-1.97	-2.36	-2.84	-2.89	-2.92
Electronegativity (Paulings scale)	1.6	1.2	1.0	1.0	0.9

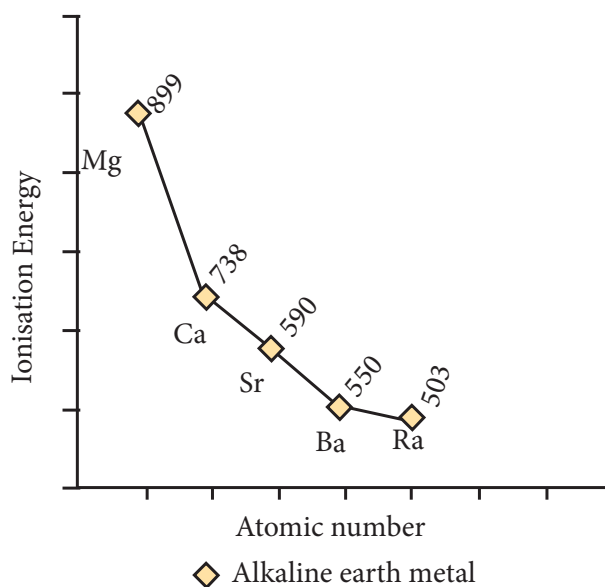


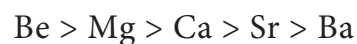
Figure 5.6 Variation of ionisation energy - Alkaline earth metals.

Although IE_1 values of alkaline earth metals are higher than that of alkali metals, the IE_2 values of alkaline earth metals are much smaller than those of alkali metals. This occurs because in alkali metals the second electron is to be removed from a cation, which has already acquired a noble gas configuration. In the case of alkaline earth metals, the second electron is to be removed from a monovalent cation, which still has one electron in the outermost shell. Thus, the second electron can be removed more easily in the case of group 2 elements than in group 1 elements.

Hydration Enthalpies

Compounds of alkaline earth metals are more extensively hydrated than those of alkali metals, because the hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions.

Like alkali metal ions, the hydration enthalpies of alkaline earth metal ions also decrease with increase in ionic size down the group.



e.g., Magnesium chloride and calcium chloride exist as their hydrated crystals $MgCl_2 \cdot 6H_2O$ and $CaCl_2 \cdot 6H_2O$ respectively, whereas NaCl and KCl do not form such hydrates.

Electronegativity

In alkaline earth metals the electronegativity values decrease as we go down the group as seen in the alkali metals.

Flame colour and the spectra:

When the alkaline earth metal salts moistened with concentrated hydrochloric acid are heated on a platinum wire in a flame, they show characteristic coloured flame as shown below.

Table 5.10 Flame Colour and wavelength

Element	Colour	Wavelength (nm)
Calcium	Brick - Red	622
Strontium	Crimson- Red	689
Barium	Apple Green	554

The heat in the flame excites the valence electron to a higher energy level. when it drops back to its actual energy level, the excess energy is emitted as light, whose wavelength is in the visible region as shown in the above table.

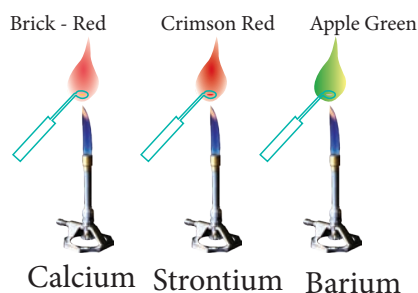


Figure 5.7: Flame colours of alkaline earth metal salts

5.5.2 Distinctive behavior of beryllium

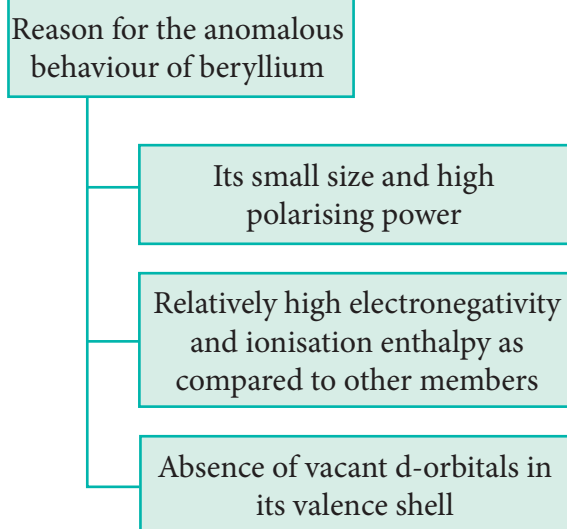


Figure 5.8 Distinctive behaviour of beryllium

The anomalous properties of beryllium is mainly due to its small size, high electronegativity, high ionisation energy and high polarising power compared to the other elements in the block. The anomalous properties of beryllium compared to other elements of the group are mentioned in Table 5.11

Table 5.11 Comparison of Properties of Beryllium with other elements of the group

Beryllium	Other elements of the family
Forms covalent compounds	form ionic compounds
High melting and boiling point	Low melting and boiling point
Does not react with water even at elevated temperature	React with water
Does not combine directly with hydrogen	Combine directly with hydrogen
Does not combine directly with halogens. Halides are covalent.	Combine directly with halogens Halides are electrovalent.

Hydroxide and oxides of beryllium are amphoteric in nature	Basic in nature.
It is not readily attacked by acids because of the presence of an oxide film	Readily attacked by acids
Beryllium carbide evolves methane with water.	evolve acetylene with water.
Salts of Be are extensively hydrolysed	Hydrolysed

Diagonal Relationship:

As observed in alkali metals, beryllium (the first member of group 2) shows a diagonal relationship with aluminium. In this case, the size of these ions ($r_{\text{Be}^{2+}} = 0.45 \text{ \AA}$ and $r_{\text{Al}^{3+}} = 0.54 \text{ \AA}$) is not as close. However, their charge per unit area is closer ($\text{Be}^{2+} = 2.36$ and $\text{Al}^{3+} = 2.50$). They also have same electronegativity values ($\text{Be} = 1.5$; $\text{Al} = 1.5$).

Table 5.12 Similarities between Beryllium and Aluminium

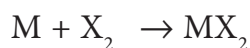
S.No.	Properties
1	Beryllium chloride forms a dimeric structure like aluminium chloride with chloride bridges. Beryllium chloride also forms polymeric chain structure in addition to dimer. Both are soluble in organic solvents and are strong Lewis acids.
2	Beryllium hydroxide dissolves in excess of alkali and gives beryllate ion and $[\text{Be}(\text{OH})_4]^{2-}$ and hydrogen as aluminium hydroxide which gives aluminate ion, $[\text{Al}(\text{OH})_4]^-$.
3	Beryllium and aluminium ions have strong tendency to form complexes, BeF_4^{2-} , AlF_6^{3-} .
4	Both beryllium and aluminium hydroxides are amphoteric in nature.
5	Carbides of beryllium (Be_2C) like aluminium carbide (Al_4C_3) give methane on hydrolysis.
6	Both beryllium and aluminium are rendered passive by nitric acid.

5.5.3 Chemical properties of alkaline earth metals

The alkaline earth metals are less reactive than the alkali metals. The reactivity of these elements increases on going down the group.

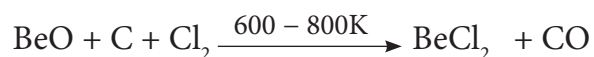
Reactivity towards the halogens:

All the alkaline earth metals combine with halogen at elevated temperatures to form their halides.



(M = Be, Mg, Ca, Sr, Ba, Ra ,
X = F, Cl, Br, I)

Thermal decomposition of $(\text{NH}_4)_2\text{BeF}_4$ is the best route for the preparation of BeF_2 . BeCl_2 is conveniently made from the oxide.



Reactivity towards hydrogen:

All the elements except beryllium, combine with hydrogen on heating to form their hydrides with general formula MH_2 . BeH_2 can be prepared by the reaction of BeCl_2 with LiAlH_4 .



5.5.4 Uses of alkaline earth metals

Uses of beryllium

1. Because of its low atomic number and very low absorption for X-rays, it is used as radiation windows for X-ray tubes and X-ray detectors.
2. The sample holder in X-ray emission studies usually made of beryllium
3. Since beryllium is transparent to energetic particles, it is used to build the 'beam pipe' in accelerators.
4. Because of its low density and diamagnetic nature, it is used in various detectors.

Uses of magnesium

1. Removal of sulphur from iron and steel.

2. Used as photoengrave plates in printing industry.
3. Magnesium alloys are used in aeroplane and missile construction.
4. Mg ribbon is used in synthesis of Grignard reagent in organic synthesis.
5. It alloys with aluminium to improve its mechanical, fabrication and welding property.
6. As a desiccant .
7. As sacrificial anode in controlling galvanic corrosion.

Uses of calcium

1. As a reducing agent in the metallurgy of uranium, zirconium and thorium.
2. As a deoxidiser, desulphuriser or decarboniser for various ferrous and non-ferrous alloys.
3. In making cement and mortar to be used in construction.
4. As a getter in vacuum tubes.
5. In dehydrating oils
6. In fertilisers, concrete and plaster of paris.

Uses of strontium

1. ^{90}Sr is used in cancer therapy.
2. ^{87}Sr / ^{86}Sr ratios are commonly used

in marine investigations as well as in teeth, tracking animal migrations or in criminal forensics.

3. Dating of rocks.
4. As a radioactive tracer in determining the source of ancient archaeological materials such as timbers and coins.

Uses of Barium

1. Used in metallurgy, its compounds are used in pyrotechnics, petroleum mining and radiology.
2. Deoxidiser in copper refining.
3. Its alloys with nickel readily emits electrons hence used in electron tubes and in spark plug electrodes.
4. As a scavenger to remove last traces of oxygen and other gases in television and other electronic tubes.
5. An isotope of barium ^{133}Ba , used as a source in the calibration of gamma ray detectors in nuclear chemistry.

Uses of Radium

Used in self-luminous paints for watches, nuclear panels, aircraft switches, clocks and instrument dials.

5.6. General characteristics of the compounds of the alkaline earth metals

The dipositive oxidation state (M^{2+}) is the predominant valence of group 2 elements. The alkaline earth metals form compounds which are predominantly ionic. However, they are less ionic than

the corresponding compounds of alkali metals. This is due to increased nuclear charge and smaller size. The general characteristics of some of the compounds of alkaline earth metals are described below.

(a) Oxides

Generally alkaline earth metals form monoxides and peroxides.

Monoxides

Monoxides are obtained by heating the metals in oxygen. BeO and MgO are almost insoluble in water. On the other hand, oxides of other elements form hydroxides. BeO is amphoteric; MgO is weakly basic while CaO , SrO and BaO are strongly basic.

BeO oxide is covalent due to the small size of Be^{2+} ion, while other oxides are ionic in nature.

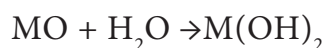
Peroxides

Except beryllium, all the remaining metals form peroxides. It is prepared by heating monoxides with oxygen at high temperature.



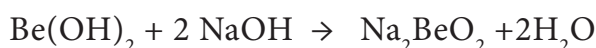
b) Hydroxides:

All the oxides except BeO are basic in nature and react with water to form sparingly soluble hydroxides.



The solubility, thermal stability

and the basic character of the hydroxides increase down the group. The alkaline earth metal hydroxides are, however, less basic and less stable than alkali metal hydroxides. Beryllium hydroxide is amphoteric in nature as it reacts with both acid and alkali.



c) Halides:

Alkaline earth metals form halides with general formula MX_2 . They can be prepared by heating metals with halogens on heating.



Beryllium halides are covalent on account of smaller size of Be^{+2} . Beryllium halides are hygroscopic, fume in moist air and soluble in organic solvents. Beryllium chloride has a chain structure in the solid state as shown in figure 5.9 (structure-a). In the vapour phase BeCl_2 tends to form a chloro-bridged dimer (structure-c) which dissociates into the linear monomer at high temperatures of the order of 1200 K. (structure-b).

Except beryllium halides, all the other halides of alkaline earth metals are ionic in nature. Chloride and fluorides of the other metals are ionic solids. These are good conductors of electricity in fused state and in aqueous solutions. The tendency to form halide hydrates gradually decreases (for example, $\text{MgCl}_2 \cdot 8\text{H}_2\text{O}$, $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$, $\text{SrCl}_2 \cdot 6\text{H}_2\text{O}$ and $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$) down the group.

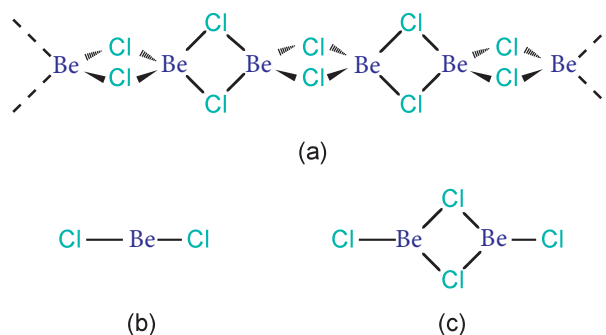


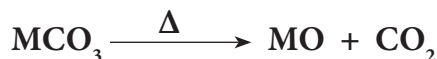
Figure 5.9 Structure of beryllium chloride

Salts of oxo acids

The alkaline earth metals form salts of oxo acids. Some of these are given below:

Carbonates:

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



- The solubility of carbonates in water decreases down the group.
- The thermal stability increases down the group with increasing cationic size.

Table 5.13 Decomposition temperature of alkaline metal carbonates and sulphates

Element	Decomposition temp for carbonates (in °C)	Decomposition temp for sulphates (in °C)
Be	25	500
Mg	540	895
Ca	900	1149
Sr	1290	1374
Ba	1360	-

Sulphates:

The sulphates of the alkaline earth metals are all white solids and stable to heat. BeSO_4 and MgSO_4 are readily soluble in water; the solubility decreases from CaSO_4 to BaSO_4 . The greater hydration enthalpies of Be^{2+} and Mg^{2+} ions overcome the lattice enthalpy factor and therefore their sulphates are soluble in water.

Nitrates:

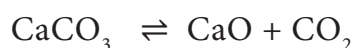
The nitrates are made by dissolution of the carbonates in dilute nitric acid. Magnesium nitrate crystallises with six molecules of water, whereas barium nitrate crystallises as the anhydrous salt. This again shows a decreasing tendency to form hydrates with increasing size. All of them decompose on heating to give the oxide.

5.6.1 Important compounds of calcium

Quick lime, CaO

Preparation

It is produced on a commercial scale by heating limestone in a lime kiln in the temperature range 1070-1270K.



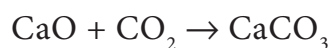
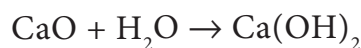
The reaction being reversible, carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion.

Properties

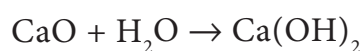
Calcium oxide is a white amorphous solid.

It has a melting point of 2870 K.

(i) It absorbs moisture and carbon dioxide on exposure to atmosphere.



(ii) The addition of limited amount of water breaks the lump of lime. This process is called slaking of lime and the product is slaked lime.



(iii) The mixture of Quick lime (CaO) and sodium hydroxide is called soda lime.

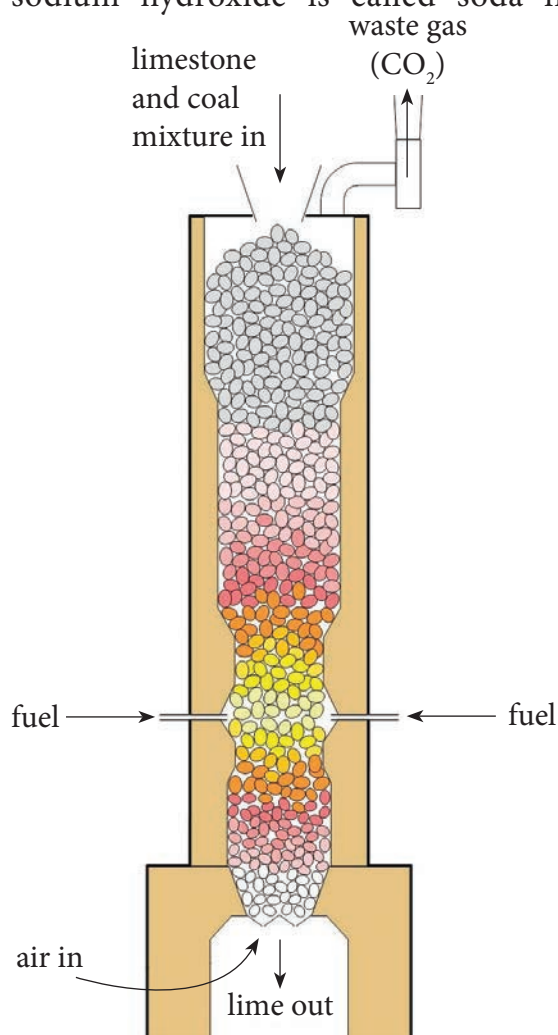
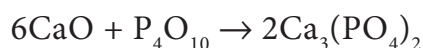
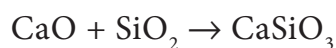


Figure 5.10 Preparation of Quick Lime

(iv) It combines with acidic oxides such as SiO_2 and P_4O_{10} to form CaSiO_3 and $\text{Ca}_3(\text{PO}_4)_2$, respectively.



Uses

Calcium oxide is used

- (i) to manufacture cement, mortar and glass.
- (ii) in the manufacture of sodium carbonate and slaked lime.
- (iii) in the purification of sugar.
- (iv) as a drying agent.

5.6.2 Calcium hydroxide

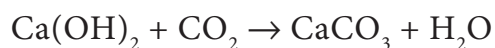
Preparation

Calcium hydroxide is prepared by adding water to quick lime, CaO .

Properties

It is a white powder. It is sparingly soluble in water. The aqueous solution is known as lime water and a suspension of slaked lime in water is known as milk of lime.

When carbon dioxide is passed through lime water, it turns milky due to the formation of calcium carbonate.

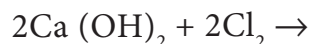


On passing excess of carbon dioxide, the precipitate dissolves to form calcium hydrogen carbonate.



Milk of lime reacts with chlorine

to form hypochlorite, a constituent of bleaching powder.



Uses:

Calcium hydroxide is used

- (i) in the preparation of mortar, a building material.
- ii) in white wash due to its disinfectant nature.
- (iii) in glass making, in tanning industry, in the preparation of bleaching powder and for the purification of sugar.

5.6.3 Gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$)

Gypsum beds were formed due to the evaporation of water from the massive prehistoric sea basins. When water evaporates, the minerals present in it become concentrated, and crystallise.



Figure 5.11 A Gypsum Quarry

Properties of Gypsum

- Gypsum is a soft mineral, which is moderately soluble in water. The solubility of this mineral in water is



affected by temperature. Unlike other salts, gypsum becomes less soluble in water as the temperature increases. This is known as retrograde solubility, which is a distinguishing characteristic of gypsum.

- Gypsum is usually white, colorless, or gray in color. But sometimes, it can also be found in the shades of pink, yellow, brown, and light green, mainly due to the presence of impurities.
- Gypsum crystals are sometimes found to occur in a form that resembles the petals of a flower. This type of formation is referred to as 'desert rose', as they mostly occur in arid areas or desert terrains.
- Gypsum is known to have low thermal conductivity, which is the reason why it is used in making drywalls or wallboards. Gypsum is also known as a natural insulator.



Figure 5.12 -The Alabaster Variety of Gypsum

- Gypsum has hardness between 1.5 to 2 on Moh's Hardness Scale. Its specific gravity is 2.3 to 2.4.
- ### Uses of Gypsum
- The alabaster variety of gypsum was used in ancient Egypt and Mesopotamia by the sculptors. The ancient Egyptians knew how to turn gypsum into plaster of Paris about 5,000 years ago. Today, gypsum has found a wide range of uses and applications in human society, some of which are enlisted below.
 - Gypsum is used in making drywalls or plaster boards. Plaster boards are used as the finish for walls and ceilings, and for partitions.
 - Another important use of gypsum is the production of plaster of Paris. Gypsum is heated to about 300 degree Fahrenheit to produce plaster of Paris, which is also known as gypsum plaster. It is mainly used as a sculpting material.
 - Gypsum is used in making surgical and orthopedic casts, such as surgical splints and casting moulds.
 - Gypsum plays an important role in agriculture as a soil additive, conditioner, and fertilizer. It helps loosen up compact or clay soil, and provides calcium and sulphur, which are essential for the healthy growth of a plant. It can also be used for removing sodium ion from soils having excess salinity.
- Alabaster is a variety of gypsum, that is highly valued as an ornamental stone. It has been used by the sculptors for centuries. Alabaster is granular and opaque.

- Gypsum is used in toothpastes, shampoos, and hair products, mainly due to its binding and thickening properties.
- Gypsum is a component of Portland cement, where it acts as a hardening retarder to control the speed at which concrete sets.
- To sum up, gypsum is one of the most abundant minerals that have endless uses and applications. Mining of gypsum is simple and easy, as the mineral occurs in large thick beds near the Earth's surface. However, large-scale mining of gypsum involves considerable damage to the environment. Gypsum can also be recycled, but not much importance has been given to recycle this mineral due to its abundance.

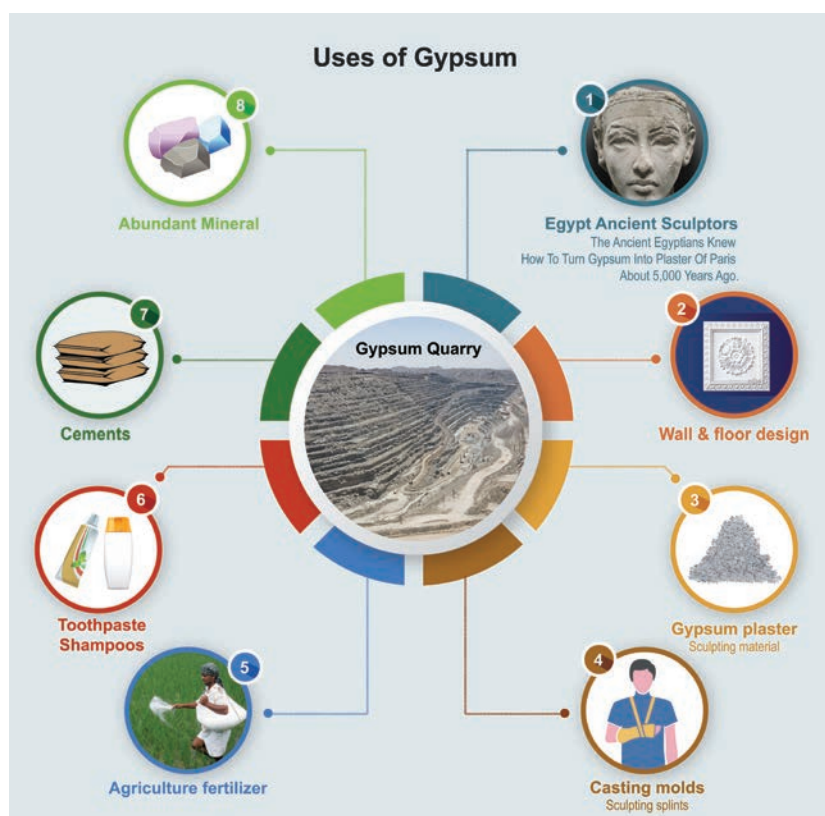


Figure 5.13 -Uses of Gypsum

5.6.4 Plaster of paris

Calcium Sulphate hemihydrate : $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$ (Plaster of Paris)

It is a hemihydrate of calcium sulphate. It is obtained when gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, is heated to 393 K.



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.

Uses:

Plaster of Paris is used as/in,

1. The building industry as well as plasters.
2. For immobilising the affected part of organ where there is a bone fracture or sprain.
3. Employed in dentistry, in ornamental work and for making casts of statues and busts.



5.7 Biological importance of magnesium and calcium

Magnesium and calcium also plays a vital role in biological functions. A typical adult human body contains about 25 g of magnesium and 1200 g of calcium. Magnesium plays an important role in many biochemical reactions catalysed by enzymes. It is the co-factor of all enzymes that utilize ATP in phosphate transfer and energy release. It also essential for DNA synthesis and is responsible for the stability and proper functioning of DNA. It is also used for balancing electrolytes in our body. Deficiency of magnesium results into convulsion and neuromuscular irritation.

Calcium is a major component of bones and teeth. It is also present in in blood and its concentration is maintained by hormones (calcitonin and parathyroid hormone). Deficiency of calcium in blood causes it to take longer time to clot. It is also important for muscle contraction.

The main pigment that is responsible for photosynthesis, chlorophyll, contains magnesium which plays an important role in photosynthesis.



SUMMARY

The elements belonging to groups 1 and 2 of the modern periodic table are called s-block elements. They are called so because the valence electron occupies the s orbitals. The group 1 elements have a general outer electronic configuration ns^1 and are called alkali metals. The group 2 elements have a general outer electronic configuration ns^2 and these are called alkaline earth metals as they are found in earth's crust and their oxides and hydroxides are alkaline in nature. Elements belonging to group 1 and 2 are highly reactive and forms M^+ and M^{2+} cations respectively. Their physical and chemical properties of both groups show a regular trend as we move down the group. The atomic and ionic radii increase as we move down the group while their ionisation enthalpies decrease.

The first element in each of these groups, lithium in Group 1 and beryllium in Group 2 shows some difference in behaviour with the elements in rest of their groups and show similarities in properties to the second member of the next group. This behaviour is known as the 'diagonal relationship' in the periodic table.

The alkali metals are soft and silvery white in colour with low melting points. They are highly reactive. The compounds of alkali metals are predominantly ionic. They form metal hydrides and halides with hydrogen and halogens respectively. Their oxides and hydroxides are soluble in water forming

strong alkalies. Important compounds of sodium include sodium carbonate, sodium chloride, sodium hydroxide and sodium hydrogen carbonate. Sodium hydroxide is manufactured by Castner-Kellner process and sodium carbonate by Solvay process.

The chemistry of alkaline earth metals is similar to alkali metals. However, we observe some differences because of their reduced atomic and ionic sizes and increased cationic charges. Their oxides and hydroxides are less basic than the alkali metal oxides and hydroxides. They also form hydrides and halides with hydrogen and halogens respectively. Industrially important compounds of calcium include calcium oxide (lime), calcium hydroxide (slaked lime), calcium sulphate hemihydrate (Plaster of Paris), calcium carbonate (limestone) and cement. Portland cement is an important constructional material. It is manufactured by heating a pulverised mixture of limestone and clay in a rotary kiln. The clinker thus obtained is mixed with some gypsum (2-3%) to give a fine powder of cement. All these substances find variety of uses in different areas.

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.





EVALUATION



1. For alkali metals, which one of the following trends is incorrect ?
 - a) Hydration energy : $\text{Li} > \text{Na} > \text{K} > \text{Rb}$
 - b) Ionisation energy : $\text{Li} > \text{Na} > \text{K} > \text{Rb}$
 - c) Density : $\text{Li} < \text{Na} < \text{K} < \text{Rb}$
 - d) Atomic size : $\text{Li} < \text{Na} < \text{K} < \text{Rb}$
2. Which of the following statements is incorrect ?
 - a) Li^+ has minimum degree of hydration among alkali metal cations.
 - b) The oxidation state of K in KO_2 is +1
 - c) Sodium is used to make Na / Pb alloy
 - d) MgSO_4 is readily soluble in water
3. Which of the following compounds will not evolve H_2 gas on reaction with alkali metals ?
 - a) ethanoic acid b) ethanol
 - c) phenol d) none of these
4. Which of the following has the highest tendency to give the reaction
$$\text{M}^+(\text{g}) \xrightarrow[\text{Medium}]{\text{Aqueous}} \text{M}^+(\text{aq})$$
 - a) Na b) Li c) Rb d) K
5. sodium is stored in
 - a) alcohol b) water c) kerosene d) none of these
6. RbO_2 is
 - a) superoxide and paramagnetic b) peroxide and diamagnetic
 - c) superoxide and diamagnetic d) peroxide and paramagnetic



7. Find the wrong statement
- a) sodium metal is used in organic qualitative analysis
 - b) sodium carbonate is soluble in water and it is used in inorganic qualitative analysis
 - c) potassium carbonate can be prepared by solvay process
 - d) potassium bicarbonate is acidic salt
8. Lithium shows diagonal relationship with
- a) sodium
 - b) magnesium
 - c) calcium
 - d) aluminium
9. In case of alkali metal halides, the ionic character increases in the order
- a) $MF < MCl < MBr < MI$
 - b) $MI < MBr < MCl < MF$
 - c) $MI < MBr < MF < MCl$
 - d) none of these
10. In which process, fused sodium hydroxide is electrolysed for extraction of sodium ?
- a) Castner's process
 - b) Cyanide process
 - c) Down process
 - d) All of these
11. The product obtained as a result of a reaction of nitrogen with CaC_2 is (NEET - Phase I)
- a) $Ca(CN)_3$
 - b) CaN_2
 - c) $Ca(CN)_2$
 - d) Ca_3N_2
12. Which of the following has highest hydration energy
- a) $MgCl_2$
 - b) $CaCl_2$
 - c) $BaCl_2$
 - d) $SrCl_2$
13. Match the flame colours of the alkali and alkaline earth metal salts in the bunsen burner
- (p) Sodium
 - (1) Brick red



- (q) Calcium (2) Yellow
(r) Barium (3) Lilac (violet)
(s) Strontium (4) Apple green
(t) Cesium (5) Crimson red
(u) Potassium (6) Blue

- a) p - 2, q - 1, r - 4, s - 5, t - 6, u - 3
b) p - 1, q - 2, r - 4, s - 5, t - 6, u - 3
c) p - 4, q - 1, r - 2, s - 3, t - 5, u - 6
d) p - 6, q - 5, r - 4, s - 3, t - 1, u - 2

14. Assertion : Generally alkali and alkaline earth metals form superoxides

Reason : There is a single bond between O and O in superoxides.

- a) both assertion and reason are true and reason is the correct explanation of assertion
b) both assertion and reason are true but reason is not the correct explanation of assertion
c) assertion is true but reason is false
d) both assertion and reason are false

15. Assertion : BeSO_4 is soluble in water while BaSO_4 is not

Reason : Hydration energy decreases down the group from Be to Ba and lattice energy remains almost constant.

- a) both assertion and reason are true and reason is the correct explanation of assertion
b) both assertion and reason are true but reason is not the correct explanation of assertion
c) assertion is true but reason is false
d) both assertion and reason are false





16. Which is the correct sequence of solubility of carbonates of alkaline earth metals ?
- a) $\text{BaCO}_3 > \text{SrCO}_3 > \text{CaCO}_3 > \text{MgCO}_3$
 - b) $\text{MgCO}_3 > \text{CaCO}_3 > \text{SrCO}_3 > \text{BaCO}_3$
 - c) $\text{CaCO}_3 > \text{BaCO}_3 > \text{SrCO}_3 > \text{MgCO}_3$
 - d) $\text{BaCO}_3 > \text{CaCO}_3 > \text{SrCO}_3 > \text{MgCO}_3$
17. In context with beryllium, which one of the following statements is incorrect ?
(NEET Phase - 2)
- a) It is rendered passive by nitric acid
 - b) It forms Be_2C
 - c) Its salts are rarely hydrolysed
 - d) Its hydride is electron deficient and polymeric
18. The suspension of slaked lime in water is known as (NEET Phase - II)
- a) lime water
 - b) quick lime
 - c) milk of lime
 - d) aqueous solution of slaked lime
19. A colourless solid substance (A) on heating evolved CO_2 and also gave a white residue, soluble in water. Residue also gave CO_2 when treated with dilute HCl.
- a) Na_2CO_3
 - b) NaHCO_3
 - c) CaCO_3
 - d) $\text{Ca}(\text{HCO}_3)_2$
20. The compound (X) on heating gives a colourless gas and a residue that is dissolved in water to obtain (B). Excess of CO_2 is bubbled through aqueous solution of B, C is formed. Solid (C) on heating gives back X. (B) is
- a) CaCO_3
 - b) $\text{Ca}(\text{OH})_2$
 - c) Na_2CO_3
 - d) NaHCO_3
21. Which of the following statement is false ? (NEET - Phase - I)
- a) Ca^{2+} ions are not important in maintaining the regular beating of the heart
 - b) Mg^{2+} ions are important in the green parts of the plants
 - c) Mg^{2+} ions form a complex with ATP
 - d) Ca^{2+} ions are important in blood clotting



22. The name 'Blue John' is given to which of the following compounds ?
- a) CaH_2 b) CaF_2 c) $\text{Ca}_3(\text{PO}_4)_2$ d) CaO
23. Formula of Gypsum is
- a) $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ b) $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$
- c) $3 \text{CaSO}_4 \cdot \text{H}_2\text{O}$ d) $2\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
24. When CaC_2 is heated in atmospheric nitrogen in an electric furnace the compound formed is
- a) $\text{Ca}(\text{CN})_2$ b) CaNCN
- c) CaC_2N_2 d) CaNC_2
25. Among the following the least thermally stable is
- (a) K_2CO_3 b) Na_2CO_3
- (c) BaCO_3 d) Li_2CO_3

II. Write brief answer to the following questions:

26. Why sodium hydroxide is much more water soluble than sodium chloride?
27. Write the chemical equations for the reactions involved in solvay process of preparation of sodium carbonate.
28. An alkali metal (x) forms a hydrated sulphate, $\text{X}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$. Is the metal more likely to be sodium (or) potassium.
29. Write balanced chemical equation for each of the following chemical reactions.
- (i) Lithium metal with nitrogen gas
- (ii) heating solid sodium bicarbonate
- (iii) Rubidium with oxygen gas
- (iv) solid potassium hydroxide with CO_2
- (v) heating calcium carbonate
- (vi) heating calcium with oxygen



30. Discuss briefly the similarities between beryllium and aluminium.
31. Give the systematic names for the following
 - (i) milk of magnesia (ii) lye (iii) lime (iv) Caustic potash
 - (v) washing soda (vi) soda ash (v) trona
32. Substantiate Lithium fluoride has the lowest solubility among group one metal fluorides.
33. Mention the uses of plaster of paris
34. Beryllium halides are covalent whereas magnesium halides are ionic why?
35. Alkaline earth metal (A), belongs to 3rd period reacts with oxygen and nitrogen to form compound (B) and (C) respectively. It undergoes metal displacement reaction with AgNO_3 solution to form compound (D). Identify A,B.C and D.
36. Write balanced chemical equation for the following processes
 - (a) heating calcium in oxygen
 - (b) heating calcium carbonate
 - (c) evaporating a solution of calcium hydrogen carbonate
 - (d) heating calcium oxide with carbon
37. Explain the important common features of Group 2 elements.
38. Discuss the similarities between beryllium and aluminium.
39. Why alkaline earth metals are harder than alkali metals.
40. How is plaster of paris prepared?
41. Give the uses of gypsum.
42. Describe briefly the biological importance of Calcium and magnesium.
43. Which would you expect to have a higher melting point, magnesium oxide or magnesium fluoride? Explain your reasoning.

CONCEPT MAP

