10. S-BLOCK ELEMENTS

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ERNAKULAM

- In s-block elements the last electron enters the outermost s-orbital.
- It includes elements of group 1 and 2 and He in group 18.
- It consists of six elements Li, Na, K, Rb, Cs and Fr.
- These are called alkali metals because they react with water to form soluble hydroxides which are strong bases or alkalies.

- The group 2 includes Be, Mg, Ca, Sr, Ba and Ra.
- These elements with the exception of Be are commonly known as alkaline earth metals.
- They are so called because their oxides are found in the earth's crust.
- The general outer electronic configuration of s-block elements is ns^1 or ns^2 .

DIAGONAL RELATIONSHIP

- The similarity in properties shown by diagonally placed elements of second and third periods in modern periodic table is called diagonal relationship.
- Li shows similarities to Mg.
- Be shows similarities Al.

CAUSE OF DIAGONAL RELATIONSHIP

- As we move from left to right across a period, the electronegativity increases.
- But it decreases when we move from top to bottom.
- As a result of these two opposite changes, as we move diagonally, these two effects tend to cancel each other and there is no marked change in electronegativity.
- Because of the similar values of electronegativities, the diagonal elements have similar chemical properties.

ANOMALOUS BEHAVIOUR OF THE FIRST MEMBER OF A GROUP

- The first member in each group shows certain properties which are different from that of the other elements in their respective groups.
- This anomalous behaviour of the first member in each group is due to
- The very small size of the atoms
- High Ionisation energies
- High electronegativities
- Absence of vacant d orbitals.

GROUP I ELEMENTS
ALKALI METALS

ATOMIC AND PHYSICAL PROPERTIES

ELECTRONIC CONFIGURATION

The general electronic configuration of alkali metals may be represented as [Noble gas] ns^1 .

ATOMIC AND IONIC RADII

- The alkali metals have the largest sizes in a particular period of the periodic table.
- With the increase in atomic number, the atom becomes larger.
- The monovalent ions (M^+) are smaller than the parent atom.
- The atomic and ionic radii of alkali metals increase on moving down the group.

IONISATION ENERGY

- The alkali metals have low ionisation energies.
- Their atomic sizes are quite large and the valence electrons are quite loosely held by the nucleus.
- On moving down the group, the ionisation energy decreases.
- It is due to the increase in size of the atoms.
- Increase in the magnitude of screening effect caused by the increase in the number of intervening electrons.

HYDRATION ENTHALPY

- The hydration enthalpies of alkali metal ions decrease with increase in ionic sizes.
- The smaller the ion, the more is the extent of hydration.
- Due to this, in aqueous solution, the hydration enthalpy decreases in the order $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

PHYSICAL PROPERTIES

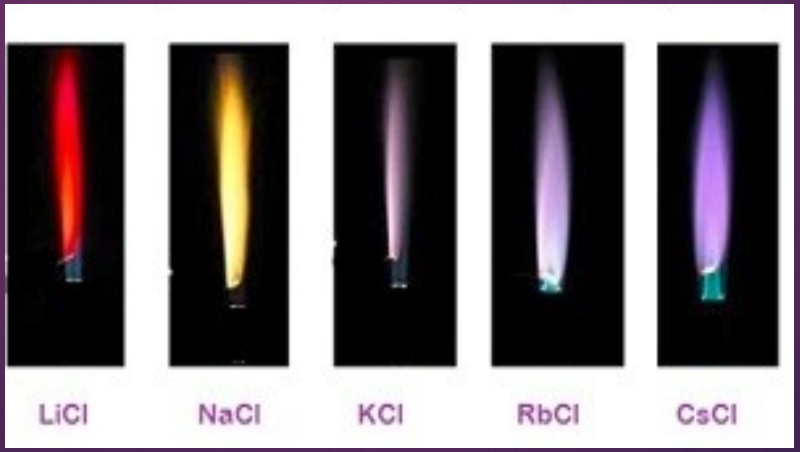
- **1. METALLIC CHARACTER**
- **All the alkali metals are silvery white, soft and light metals.**
- **The metallic character increases down the group.**
- **2. DENSITY**
- **Because of the large size, alkali metals have low density.**
- **The Density generally increases down the group.**

MELTING AND BOILING POINTS

- The melting and boiling points of alkali metals are low.
- It indicates weak metallic bonding due to the presence of only a single valence electron in them.
- The melting and boiling points decreases down the group.

- **When alkali metals or their salts are heated in the flame of a Bunsen burner, they impart characteristic colours to the flame.**

Li	Na	K	Rb	Cs
Crimson red	Golden Yellow	Pale Violet	Reddish Violet	Sky Blue



- Alkali metals have low Ionisation energy.
- Therefore, the electrons present in their atoms get readily excited to higher energy states by absorbing energy from the visible region of the light.
- When these electrons jump back to their ground state, they emit energy in the form of radiations.
- These radiations fall in the visible region, imparting characteristic colouration to the flame.

CHEMICAL PROPERTIES

REACTIVITY TOWARDS AIR

- Alkali metals tarnish in dry air.
- It is due to the formation of their oxides.
- These oxides in turn react with moisture to form hydroxides.
- They burn vigorously in oxygen forming oxides.

- When heated in excess of air, Li forms normal oxide.
- Sodium forms peroxides.
- K, Rb and Cs forms superoxide.

$4Li + O_2 \longrightarrow 2Li_2O$	(Monoxide)
$2Na + O_2 \longrightarrow Na_2O_2$	(Peroxide)
$K + O_2 \longrightarrow KO_2$	(Superoxide)

REACTIVITY TOWARDS WATER

The alkalie metals react with water to form hydroxide and dihydrogen.

$$2M + 2H_2O \longrightarrow 2MOH + H_2$$

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

REACTIVITY TOWARDS DIHYDROGEN

- On heating, alkali metals react with hydrogen to form hydrides of the formula, MH.
- All the alkali metal hydrides are ionic solids with high melting points.

$$2Na + H_2 \longrightarrow 2NaH$$

- The hydrides are strong reducing agents.
- It reduce water with the liberation of hydrogen.

$$LiH + H_2O \longrightarrow LiOH + H_2$$

$$NaH + H_2O \longrightarrow NaOH + H_2$$

REACTIVITY TOWARDS HALOGENS

Alkali metals react with halogens to form halides of the general formula, MX.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

REDUCING NATURE

- The alkali metals are strong reducing agents.
- Lithium, being the most and sodium, the least powerful.
- The reducing character increases down the group from Na to Cs because of the decrease in ionisation energy.

SOLUBILITY IN LIQUID AMMONIA

- The alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature.

$$M + (x + y)\ NH_3 \longrightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-$$

- The blue colour of the solution is due to the ammoniated electrons (solvated electrons).
- Ammoniated electrons absorbs energy in the visible region of light.
- This imparts blue colour to the solution.
- The solutions are paramagnetic.

SALTS OF OXO ACIDS

- Since alkali metals are strongly electropositive, they form salts with oxo acids.
- Oxo acids are those in which the acidic proton is on a hydroxyl group with an oxo group attached to the same atom.

Why is KO_2 paramagnetic?

The superoxide ion, O_2^- is paramagnetic because of one unpaired electron in π *2p molecular orbital.

$$(\sigma 1s)^2\ (\sigma^* 1s)^2\ (\sigma\ 2s)^2\ (\sigma^* 2s)^2\ (\sigma 2p_z)^2$$

$$(\pi 2p_x^2 = \pi 2p_y^2\ )\ (\pi^* 2p_x^2\ = \pi^* 2p_y^1)$$

ANOMALOUS PROPERTIES OF LITHIUM

- Lithium is much harder.
- The melting point and boiling point are higher than the other alkali metals.
- Lithium is the least reactive but the strongest reducing agent among all the alkali metals.

- Lithium reacts with nitrogen to form a nitride, Li_3N.
- LiCl is deliquescent and crystallizes as a hydrate, $LiCl.\ 2H_2O$.
- LiOH is a weak base.
- Lithium does not form acetylide on reaction with acetylene.

DIAGONAL SIMILARITIES BETWEEN Li and Mg

- Both Li and Mg are harder than other elements in the respective groups.
- Both Li and Mg form a nitride by direct combination with nitrogen
- Both Li and Mg react with oxygen to form normal oxides Li_2O and MgO respectively.

- The hydroxides of Li and Mg are weak bases and they decompose on heating.
- Carbonates of Li and Mg decompose easily on heating to form the oxide and CO_2 .
- Both LiCl and $MgCl_2$ are soluble in ethanol.
- Both LiCl and $MgCl_2$ are deliquescent and crystallizes as hydrates $LiCl.2H_2O$ and $MgCl_2 .8H_2O$.

SOME IMPORTANT COMPOUNDS OF SODIUM

- Sodium forms many industrially important compounds like
- sodium carbonate (Na_2CO_3)
- sodium chloride (NaCl)
- sodium hydroxide (NaOH)
- sodium hydrogen carbonate ($NaHCO_3$)

SODIUM CARBONATE (Na_2CO_3)

Sodium carbonate is also known as washing soda or soda ash.

PREPARATION

- Sodium carbonate is generally prepared by Solvay process.
- Prepared by passing CO_2 gas under pressure into a strong sodium chloride solution saturated with ammonia.
- Sodium Hydrogen Carbonate is precipitated.

$$2NH_3 + H_2O + CO_2 \longrightarrow (NH_4)_2CO_3$$

$$(NH_4)_2CO_3 + H_2O + CO_2 \longrightarrow 2NH_4HCO_3$$

$$NH_4HCO_3 + NaCl \longrightarrow NH_4Cl + NaHCO_3$$

- **Sodium hydrogen carbonate on heating gives sodium carbonate.**

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

- **In this process, NH_3 is recovered when the solution containing NH_4Cl is treated with $Ca(OH)_2$.**
- **Calcium chloride is obtained as a by-product.**

$$2NH_4Cl + Ca(OH)_2 \longrightarrow 2NH_3 + CaCl_2 + H_2O$$

K_2CO_3 CANNOT BE PREPARED BY SOLVAY PROCESS

- Potassium carbonate cannot be prepared by Solvay process.
- This is because the solubility of $KHCO_3$ is fairly large.
- Hence it does not precipitate easily in the carbonation tower.

PROPERTIES

- Sodium carbonate is a white crystalline solid.
- It exists as a decahydrate, $Na_2CO_3 .10H_2O$.
- This is also called washing soda.
- It is readily soluble in water.
- On heating, the Decahydrate loses its water of crystallization to form monohydrate.

- Above 373K, the monohydrate becomes completely anhydrous and changes to a white powder called soda ash.

$$Na_2CO_3.10H_2O \longrightarrow Na_2CO_3.H_2O + 9H_2O$$

$$Na_2CO_3.H_2O \longrightarrow Na_2CO_3 + H_2O$$

- The solution of Na_2CO_3 is alkaline due to hydrolysis.

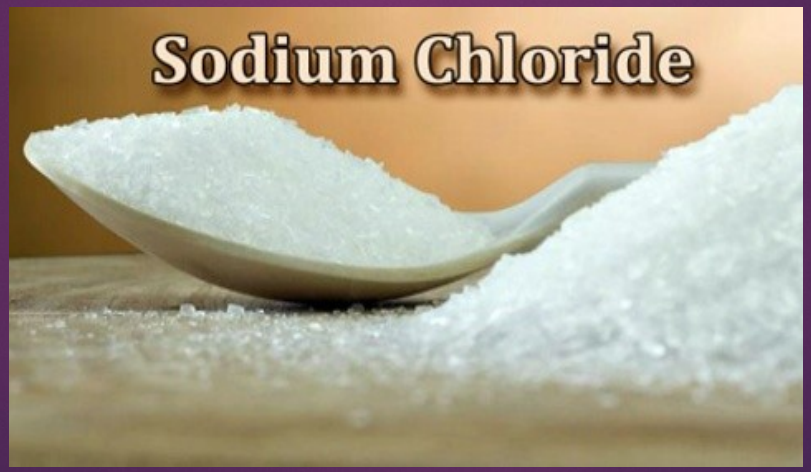
$$Na_2CO_3 + H_2O \longrightarrow NaHCO_3 + NaOH$$

USES

- Sodium Carbonate is used
- in water softening, laundering and cleaning.
- in the manufacture of glass, soap, borax and caustic soda.
- in paper, paints and textile industries.
- It is an important laboratory reagent both in qualitative and quantitative analysis.

SODIUM CHLORIDE (NaCl)

- The most abundant source of sodium chloride is sea water.
- Sea water contains 2.7 to 2.9 % by mass of the salt.



PREPARATION

- In India, common salt is generally obtained by the evaporation of sea water.
- Crude NaCl is generally obtained by crystallization of brine solution.
- It contains sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride as impurities.
- To obtain pure sodium chloride, the crude salt is dissolved in minimum amount of water.

- The insoluble impurities are filtered.
- The solution is then saturated with hydrogen chloride gas.
- Crystals of pure sodium chloride separate out.
- Calcium and magnesium chloride being more soluble than sodium chloride, remain in solution.

PROPERTIES

- Sodium chloride melts at 1082K.
- It has a solubility of 36g in 100g of water at 273K.
- The solubility does not increase appreciably with increase in temperature.

USES

- It is used as common salt or table salt for domestic purpose.
- It is used for the preparation of Na_2O_2 , NaOH and Na_2CO_3 .

SODIUM HYDROXIDE (NaOH)

Sodium hydroxide is also known as caustic soda.

PREPARATION

- Sodium hydroxide is prepared commercially by the electrolysis of sodium chloride in Castner-Kellner cell.
- Sodium hydroxide is obtained by the electrolysis of brine solution.
- Cathode is made up of mercury and anode is made up of carbon.
- Sodium metal is discharged at the cathode.
- It combines with mercury to form sodium amalgam.
- Chlorine gas is evolved at the anode.

Cathode: $Na^+ + e^- \longrightarrow Na - amalgam$

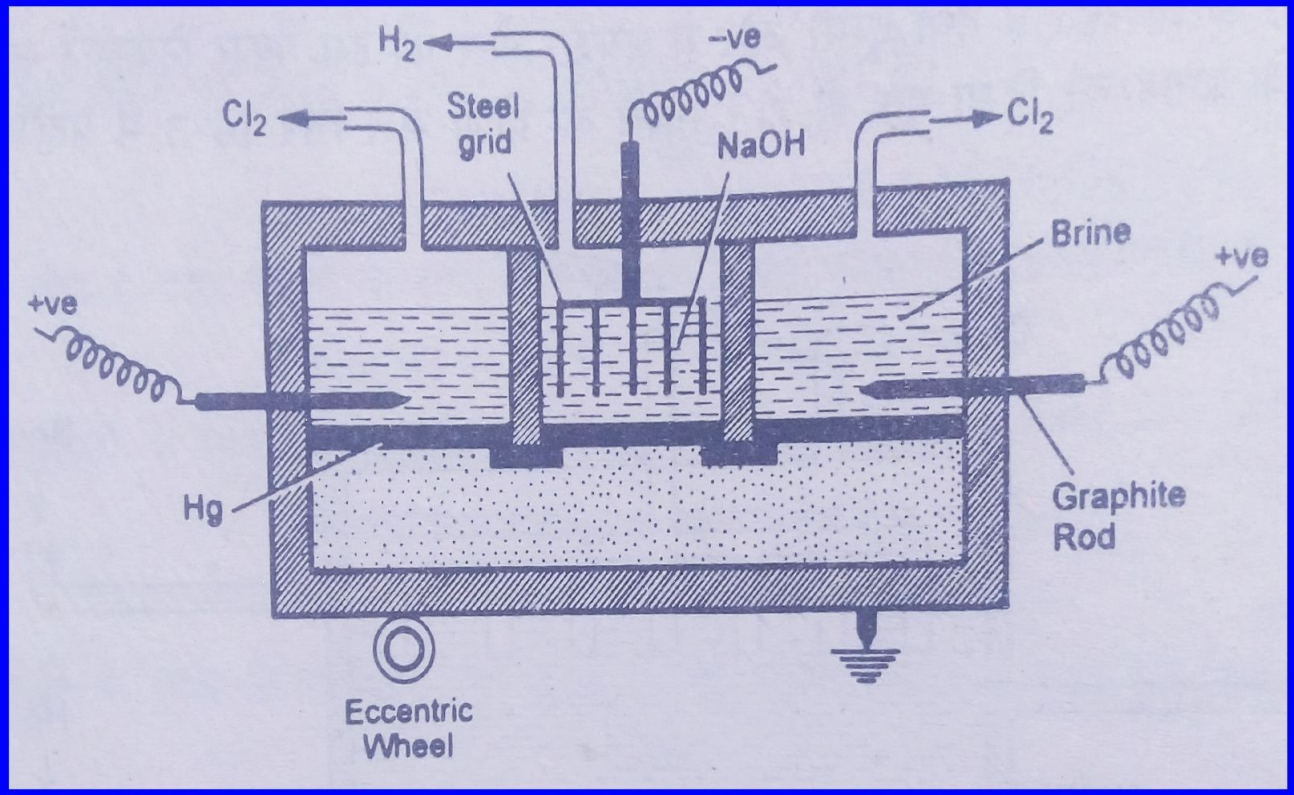
Anode : $Cl^- \longrightarrow Cl + e^-$

Overall : $Cl + Cl \longrightarrow Cl_2$

The amalgam is treated with water to give sodium hydroxide and hydrogen gas.

$$2Na - amalgam + 2H_2O \longrightarrow 2NaOH + 2Hg + H_2$$

Sodium Hydroxide (Castner-Kellener Cell)



PROPERTIES

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

USES

- Sodium hydroxide is used:-
- in the manufacture of paper, soaps, artificial silk etc.
- for the purification of bauxite.
- as an important laboratory reagent.
- in petroleum refining.
- for mercerizing cotton fibres.
- for refining vegetable oils.
- for preparing various sodium salts like sodium chlorate, hypochlorite etc.

SODIUM HYDROGEN CARBONATE

- Sodium Hydrogen Carbonate is known as baking soda.
- It is called so because it decomposes on heating to generate bubbles of CO_2 .

PREPARATION

- It is prepared by saturating a solution of sodium carbonate with carbon dioxide.
- The white crystalline powder of sodium hydrogen carbonate, being less soluble, gets separated out.

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

USES

- $NaHCO_3$ is used
- as a laboratory reagent.
- in baking powder.
- in fire extinguishers.
- as a mild antiseptic for skin infections.

BIOLOGICAL IMPORTANCE OF SODIUM

- Sodium ions are found outside the cells.
- It is located in blood plasma and in the interstitial fluid which surrounds the cells.
- These ions participate in the transmission of nerve signals.
- Regulates the flow of water across cell membranes
- Helps in the transport of sugars and amino acids into cells.

BIOLOGICAL IMPORTANCE OF POTASSIUM

- Potassium ions are the most abundant cations within cell fluids.
- They activate many enzymes.
- Participate in the oxidation of glucose to produce ATP.
- With sodium it is responsible for the transmission of nerve signals.

GROUP II ELEMENTS
ALKALINE METALS

- Group 2 of the periodic table consists of the elements Be, Mg, Ca, Sr, Ba and Ra.
- These elements except beryllium are known as alkaline earth metals.
- Radium, the last member of the group is radioactive in nature.
- Beryllium shows diagonal relationship to Aluminium.

ATOMIC AND PHYSICAL PROPERTIES

ELECTRONIC CONFIGURATION

- The general electronic configuration may be represented as [Noble gas] ns^2
- where n = 2 to 7.

ATOMIC RADII

- The atomic and ionic radii of the alkaline earth metals are smaller than those of the corresponding alkali metals in the same periods.
- This is due to the increased nuclear charge in these elements.

IONISATION ENERGIES

- The alkaline earth metals have low ionisation energies due to fairly large size of the atoms.
- The first ionisation energies of the alkaline earth metals are higher than those of the corresponding group I metals.
- This is due to their small size as compared to the corresponding alkali metals.

HYDRATION ENTHALPIES

- The hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group.
  - $Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$
- The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions.
- Compounds of alkaline earth metals are more extensively hydrated than those of alkali metals.

PHYSICAL PROPERTIES

METALLIC CHARACTER

- The alkaline earth meals are silvery white, lustrous and relatively soft, but harder than the alkali metals.
- Be and Mg appears to be somewhat greyish.

DENSITY

- The alkaline earth metals are denser and harder than the alkali metals of the corresponding periods.
- It is because of the smaller size and stronger metallic bonds resulting in more closely packed crystal lattices.

MELTING AND BOILING POINTS

- Alkaline earth metals have higher melting and boiling points.
- This is because of their smaller size and more closely packed structures.

ELECTROPOSITIVE CHARACTER

- Alkaline earth metals are electropositive in nature.
- Due to higher ionisation energies, these are less electropositive than alkali metals.
- Within the group, the electropositive character increases from Be to Ba.

FLAME COLOURATION

- Ca, Sr and Ba impart characteristic brick red, crimson red and apple green colours to the flame respectively.
- Be and Mg does not impart any colour to the flame due to small size and high ionisation energies.
- The excitation of electrons is not possible by the energy available from the flame.

CHEMICAL PROPERTIES

REACTIVITY TOWARDS AIR

- Be and Mg are kinetically inert to oxygen and water because of the formation of an oxide film on their surface.
- Powdered Be burns brilliantly on ignition in air to give BeO and Be_3N_2.
- Mg is more electropositive and burns with dazzling brilliance in air to give MgO and Mg_3N_2.
- Ca, Sr and Ba are readily attacked by air to form the oxide and nitride.
- They also react with water with increasing vigour even in cold to form hydroxides.

REACTIVITY TOWARDS HYDROGEN

- All the elements except Be combine with hydrogen upon heating to form their hydrides, MH_2.

$$M + H_2 \longrightarrow MH_2$$

$$Ca + H_2 \longrightarrow CaH_2$$

- BeH_2 can be prepared by the reaction of $BeCl_2$ with $LiAlH_4$.

REACTIVITY TOWARDS ACIDS

The alkaline earth metals readily react with acids liberating dihydrogen.

$$M + 2HCl \longrightarrow MCl_2 + H_2$$

REACTIVITY TOWARDS HALOGENS

- All the alkaline earth metals combine with halogen at elevated temperatures forming their halides.

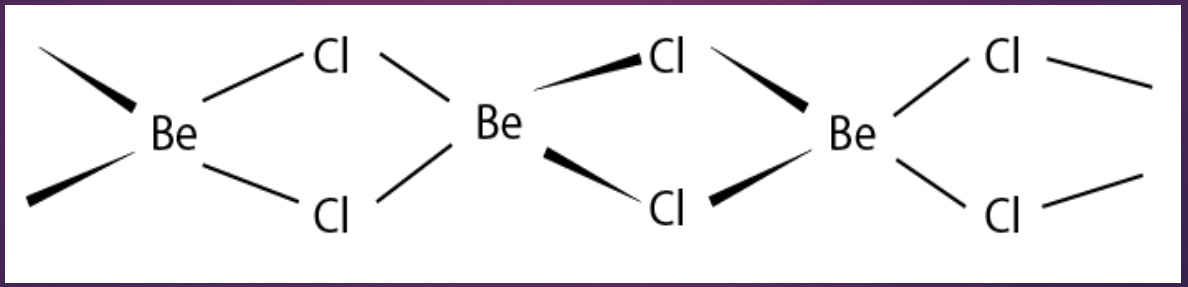
$$M + X_2 \longrightarrow MX_2$$

- $BeCl_2$ is formed by passing chlorine over a heated mixture of Beryllium oxide and carbon.

$$BeO + C + Cl_2 \longrightarrow BeCl_2 + CO$$

STRUCTURE OF $BeCl_2$

- In the solid state, $BeCl_2$ exists as a polymer.
- Each Be atom is surrounded by four chlorine atoms.
- Two of the chlorine atoms are bonded by covalent bonds and the remaining two by coordinate bonds.



REDUCING NATURE

- Alkaline earth metals are strong reducing agents.
- They have large negative values of their reduction potentials.
- Their reducing power is less than those of their corresponding alkali metals.

SOLUTIONS IN LIQUID AMMONIA

Alkaline earth metals dissolve in liquid ammonia to give deep blue black coloured solutions forming ammoniated ions.

$$M + (x + y)NH_3 \longrightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^-$$

FORMATION OF OXIDES

- Alkaline earth metals react with air or oxygen slowly upon heating to form oxides.
- Be, Mg and Ca forms monoxides.
- Sr and Ba forms peroxides.

$2M + O_2 \longrightarrow 2MO$	where M = Be, Mg or Ca.
$M + O_2 \longrightarrow MO_2$	where M = Sr or Ba

FORMATION OF HYDROXIDES

- Elements of group 2 have lesser tendency to react with water.
- On heating, they react with water to form hydroxide and liberate hydrogen.
- The hydroxides are basic in nature, except $Be(OH)_2$ which is amphoteric.
- The basic strength increases down the group.
- It is due to the decrease in the ionisation energies.

ANOMALOUS BEHAVIOUR OF BERYLLIUM

- The properties of Be differs from those of the other elements of the group.
- It is because of
- Its small size
- Relatively high ionisation energy
- Relatively high electronegativity
- Absence of vacant d orbitals in the valence shell.

DIFFERENCES

- Be compounds are more covalent in nature.
- Be does not react with water even at high temperatures.
- BeO is amphoteric.
- Be does not react with H_2 to give hydride.
- Be_2C reacts with water to give methane.

DIAGONAL RELATIONSHIP BETWEEN Be and Al

- Both Be and Al have same electronegativity.
- Both Be and Al form covalent compounds.
- Both Be and Al are resistant to the action of acids.
- The hydroxides of Be and Al are amphoteric.
- Chlorides of both Be and Al have bridged structures in vapour phase.
- Carbides of both Be and Al undergo hydrolysis to give methane.

SOME IMPORTANT COMPOUNDS OF CALCIUM

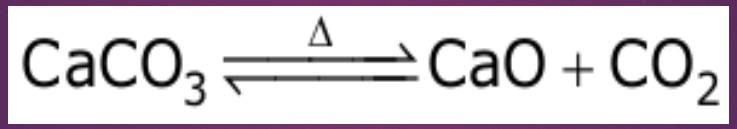
COMPOUNDS OF CALCIUM

- Important compounds of calcium are
- Calcium oxide
- Calcium hydroxide
- Calcium sulphate
- Calcium carbonate
- Cement

CALCIUM OXIDE OR QUICK LIME (CaO)

PREPARATION

- It is prepared on a commercial scale by heating limestone in a rotary kiln at 1070-1270K.

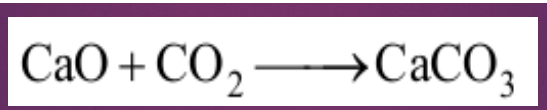


- The CO_2 is removed as soon as it is produced to enable the reaction to proceed to completion.

PROPERTIES

- Calcium oxide is a white amorphous solid.
- It has a melting point of 2870K.
- On exposure to atmosphere, it absorbs moisture and carbon dioxide.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$



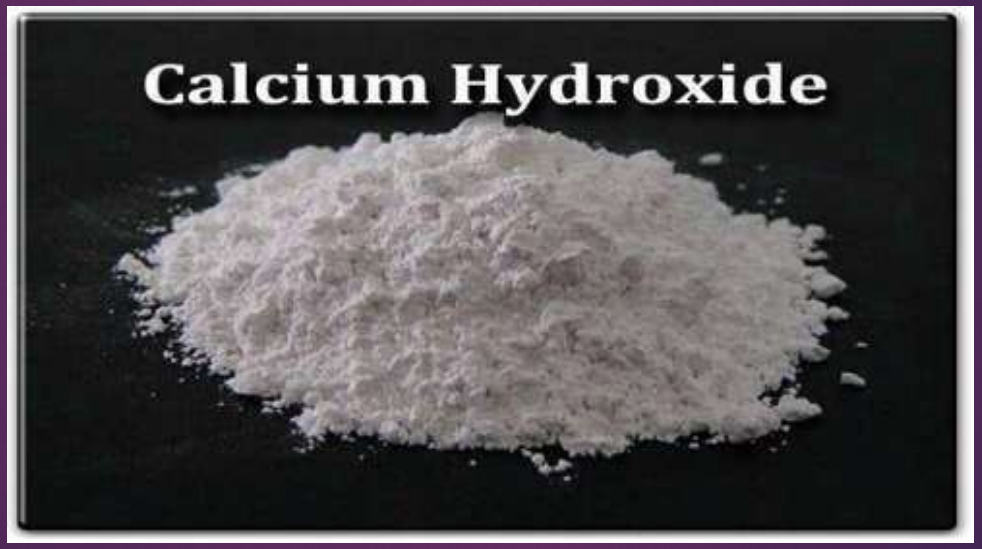
$$CaO + CO_2 \longrightarrow CaCO_3$$

- The addition of limited amount of water breaks the lump of lime.
- This process is called slaking of lime.

USES

- Calcium Oxide is used
- as a building material
- in softening of hard water.
- in the manufacture of calcium carbide.
- as a drying agent.
- in the purification of sugar.

CALCIUM HYDROXIDE OR SLAKED LIME $Ca(OH)_2$



PREPARATION

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

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

PROPERTIES

- Calcium hydroxide is a white amorphous powder.
- It is sparingly soluble in water.
- A suspension of slaked lime in water is called milk of lime.
- The filtered solution is called lime water.

- When CO_2 is passed through lime water, it turns milky due to the formation of $CaCO_3$.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

- On passing excess of CO_2 , the precipitate dissolves to form calcium hydrogen Carbonate.

$$CaCO_3 + CO_2 + H_2O \longrightarrow CaH(CO_3)_2$$

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

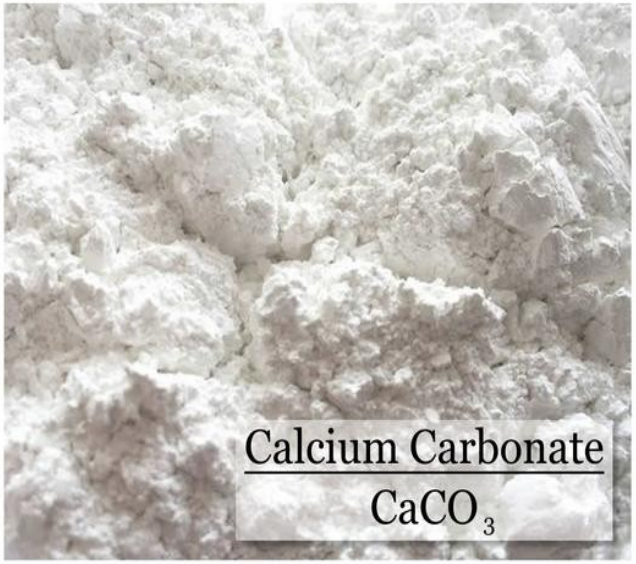
$$2Ca(OH)_2 + 2Cl_2 \longrightarrow CaCl_2 + Ca(OCl)_2 + 2H_2O$$

USES

- Calcium Hydroxide is used
- in the production of mortar, a building material.
- as lime water in laboratories.
- for the manufacture of bleaching powder.
- Milk of lime is used for white washing.

CALCIUM CARBONATE ($CaCO_3$)

It occurs in nature in several forms like limestone, chalk, marble etc.



PREPARATION

Prepared by passing carbon dioxide through slaked lime or by the addition of Na_2CO_3 to $CaCl_2$.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCl_2 + 2NaCl$$

PROPERTIES

- Calcium carbonate is a white fluffy powder.
- It is almost insoluble in water.
- When heated to 1200K, it decomposes to evolve carbon dioxide.

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

- It reacts with dilute acids to liberate carbon dioxide.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

$$CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$$

USES

- Calcium Carbonate is used
- as a building material in the form of marble.
- for the manufacture of quick lime.
- Calcium carbonate along with $MgCO_3$ is used as a flux in metallurgy.
- Specially precipitated $CaCO_3$ is extensively used in the manufacture of high quality paper.
- as an antacid, mild abrasive in tooth paste, a constituent of chewing gum and filler in cosmetics.

CALCIUM SULPHATE ($CaSO_4$)

PREPARATION

- Plaster of Paris is obtained by heating gypsum to 393 K.

$$2CaSO_4.2H_2O \longrightarrow 2CaSO_4.H_2O + 3H_2O$$

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

- When plaster of Paris is mixed with water, it forms a paste which sets to a hard mass in 5 to 15 min.
- This is called setting of Plaster of Paris.
- During setting, it takes up water and forms $CasO_4 .2H_2O$.

USES

- Calcium Sulphate is used
- for setting broken and fractured bones.
- for making statues, models, decorative materials etc.
- for producing moulds for Industries like pottery and ceramics.
- in dentistry and in ornamental work.

CEMENT

- Cement is an important building material.
- It was first introduced in England in 1824 by Joseph Aspidin.
- It is also called Portland cement.
- It resembles the natural lime stone quarried in the Isle of Portland, England.

WHAT IS CEMENT ?

- Cement is essentially a finely ground mixture of calcium and aluminium silicates along with small quantities of gypsum which sets to a hard mass when reacted with water.
- Cement is a product obtained by combining a material, rich in lime, CaO with other material such as clay which contains silica, SiO_2 along with the oxides of aluminium, Iron and Magnesium.

MANUFACTURE OF CEMENT

- The raw materials for the manufacture of cement are limestone and clay.
- When clay and lime are strongly heated together, they fuse and react to form cement clinker.
- This clinker is mixed with 2-3% by weight of gypsum to form cement.
- Thus important ingredients present in Portland cement are dicalcium silicate (Ca_2SiO_4) 26%, tricalcium silicate (Ca_3SiO_5) 51% and tricalcium aluminate ($Ca_3Al_2O_6$) 11%.

SETTING OF CEMENT

- When cement is mixed with water it reacts to form a gelatinous mass which sets to a hard mass.
- The transition of cement from the gelatinous mass to a hard mass is called setting of cement.
- During setting, three dimensional cross links are formed between
- ----Si—O—Si---- and ----Si—O—Al---- chains making the material quite hard.
- The purpose of adding gypsum is to slow down the process of setting of the cement so that it gets sufficiently hardened.

USES OF CEMENT

- Cement is used
- in concrete and reinforced concrete.
- in plastering and in construction of bridges, dams and buildings.

BIOLOGICAL IMPORTANCE OF Mg & Ca

- An adult body contains about 25g of Mg, 1200g of Ca, 5g of Iron and 0.06g of copper.
- The daily requirement in the human body has been estimated to be 200-300mg.
- All enzymes that utilize ATP in phosphate transfer require magnesium as the cofactor.

- The main pigment for the absorption of light in plants is chlorophyll which contains Mg.
- About 99% of Ca is present in bones and teeth.
- It also plays important roles in neuromuscular function, Interneuronal transmission, blood coagulation etc.


THANK
YOU!