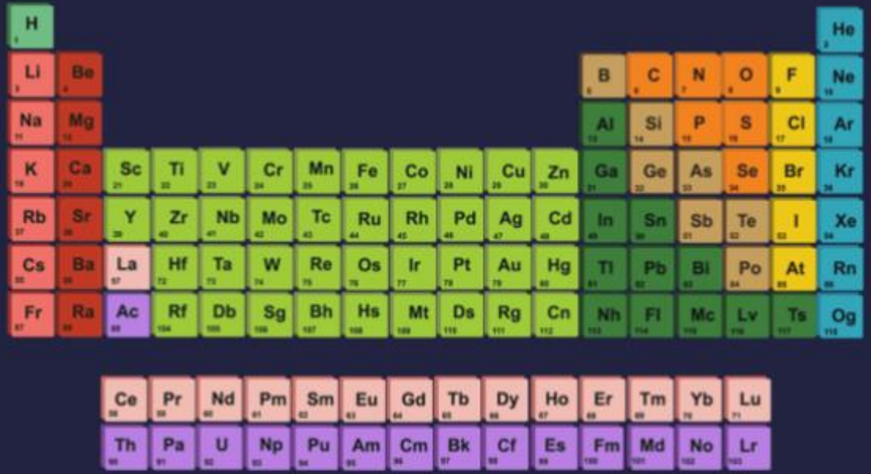
Welcome to

Aakash + BYJU'S LIVE

The s-Block Elements

s-block



The first two columns constitutes s-Block

Group 1 (s-block)

Collectively known as the **alkali metals**

They form strongly-alkaline hydroxides with water.

Lithium (Li)
Sodium (Na)
Potassium (K)
Rubidium (Rb)
Caesium (Cs)
Francium (Fr)

Group 2 (s-block)

These elements (except beryllium) are known as the Alkaline earth metals.

Their **oxides & hydroxides** are **alkaline** in nature & these metal oxides are found in the **earth's crust.**

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

Properties of Group I Elements

1. Electronic Configuration:
[Noble gas] **ns^1**

- Most **electropositive** metals
  - Readily lose electron to give unipositive **M^+ ion**
  - Never found in free state in nature

2. Atomic & Ionic radius

Down the group → Atomic & ionic radius increases

Properties of Group I Elements

3 Ionisation Enthalpy

Down the group → Ionisation enthalpy ↓

Li > Na > K > Rb > Cs

Properties of Group I Elements

4 Hydration Enthalpy

Down the group → Hydration enthalpy ↓

Li^+ has the maximum degree of hydration. Thus, lithium salts are mostly hydrated.
E.g.: $LiCl.2H_2O$.

Smaller the ion, **higher** is the **charge density,** higher is the **hydration enthalpy.**

$Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

Properties of Group I Elements

5 Melting point & Boiling point

Low M.P. & B.P. due to **weak metallic bonding** (single valence electron).

All the alkali metals are silvery white, soft and light metals.

Li	>	Na	>	K	>	Rb	>	Cs

Flame Test

Alkali metals and their salts impart characteristic color to **oxidising flame.**

Li	Na	K	Rb	Cs
Crimson red	Yellow	Violet/ Lilac	Red-Violet	Blue

Chemical Properties

Highly reactive due to their low ionisation enthalpy.

The reactivity of these metals **increases** down the group.

Because of their high reactivity towards air and water, they are normally **kept in kerosene oil.**

Cs and K are used in photoelectric cell

Reaction with Air

a Reaction with air

Alkali metals **tarnish in dry air** due to the formation of their **oxides.**

Li	Oxide	$4Li + O_2 \longrightarrow 2Li_2O$
Na	Peroxide	$2Na + O_2 \longrightarrow Na_2O_2$
Others	Superoxide	$M + O_2 \longrightarrow MO_2$

M = K, Rb, Cs

Oxides

Size of the metal ion ↑

Stability of the peroxide or superoxide ↑

Due to the stabilization of large anions by larger cations through **lattice energy** effects.

Hydroxides

Oxides are easily hydrolysed by water to form the **hydroxides.**

$$M_2O + H_2O \longrightarrow 2MOH$$

$$M_2O_2 + 2H_2O \longrightarrow 2MOH + H_2O_2$$

$$2MO_2 + 2H_2O \longrightarrow 2MOH + H_2O_2 + O_2$$

Reaction with Air

Li shows **exceptional behaviour** in reacting directly with nitrogen (at room temperature).

$$6Li + N_2 \longrightarrow 2Li_3N$$

Other members of this group do not react with nitrogen directly.

Chemical Properties

b Reaction with water

$$2M\,(s) + 2H_2O\,(l) \longrightarrow 2MOH\,(aq) + H_2\,(g)$$

M = Alkali Metal

Lithium has the **most negative E°** value ($E^\circ_{(M^+/M)}$ = -3.05 V)

But its reaction with water is **less vigorous** than that of **sodium,** which has the **least negative E°** ($E^\circ_{(M^+/M)}$ = -2.71 V) value among the alkali metals. The reaction of K is even more vigorous than sodium.

Reaction with Water

Since the **melting point** decreases down the group, the reaction with water becomes more and more **vigorous**.

Surface area exposed to water ↑ → Reaction is kinetically faster

Chemical Properties

c Reaction with dihydrogen

$$2M + H_2 \longrightarrow 2M^+H^-$$

Alkali metals react with H_2 to form **ionic hydrides**

Stability of hydrides **decreases** down the group

d Reaction with halogens

Alkali metals **react vigorously** with halogens to form ionic halides, **M^+X^-.**

$$2M + X_2 \longrightarrow 2M^+X^-$$

Alkali Metal Halides

Lithium halides have more covalent character because of the **high polarisation** capability of Li^+ ion.

Halides having **ionic nature** have **high M.P.** and are **good conductors of electricity** in fused state.

Since anions with **large size** can easily be distorted, **lithium iodide** is the most covalent in nature.

These are readily **soluble in water.**

Alkali Metal Halides

Low solubility in water

LiF → High lattice enthalpy

CsI → Smaller hydration enthalpy

For a given metal (M), $\Delta_f H°$ (MX) always becomes **less negative** on going from **MF to MI.**

Electropositive character (Li to Cs) ↑

$|\Delta_f H°|$ of Alkali metals halides (Cl_2, Br_2, I_2) ↑

Alkali metal **fluorides** follow the **reverse** order.

Chemical Properties

e Reducing nature

Alkali metals are **strong reducing agents.**

Li is the **most powerful** & Na is the **least powerful.**

$$M(s) \longrightarrow M(g)$$

Sublimation enthalpy

$$M(g) \longrightarrow M^{+}(g) + e^{-}$$

Ionisation enthalpy

$$M^{+}(g) + H_2O\,(l) \longrightarrow M^{+}(aq)$$

Hydration enthalpy

Reducing nature depends on **reduction potential** which is resultant of **sublimation, ionisation** and **hydration enthalpy** of elements.

Reducing Nature

Due to the small size of **Li^+** ion, it has the highest hydration enthalpy.

Accounts for its **high negative $E^\circ_{(M^+/M)}$ value** & therefore, has **high reducing power.**

Reducing nature in **gaseous** phase	Li < Na < K < Rb < Cs
Reducing nature in aqueous phase	Li > Cs > Rb > K > Na

Chemical Properties

f Solutions in liquid ammonia

Alkali metals dissolve in liquid ammonia giving **deep blue** solution which is **conducting, reducing, & paramagnetic** in nature.

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

Ammoniated cation

$$M^+ + x\,(NH_3) \longrightarrow [\,M(NH_3)_x\,]^+$$

Ammoniated electron

$$e^- + y\,(NH_3) \longrightarrow [e(NH_3)_y]^-$$

Solutions in Liquid Ammonia

Properties	Reason
Blue colour	Ammoniated electron
Paramagnetic	Ammoniated electron
Conducting	Ammoniated M^+ ion & Ammoniated electron

The **ammoniated electrons** in the solution absorb energy in the **visible region**, imparting blue colour.

Solutions in Liquid Ammonia

On standing, the colour fades due to **formation of amide** after liberating hydrogen.

$$M^+ + e^- + NH_3 \longrightarrow MNH_2\,(\text{amide}) + \frac{1}{2}H_2\,(g)$$

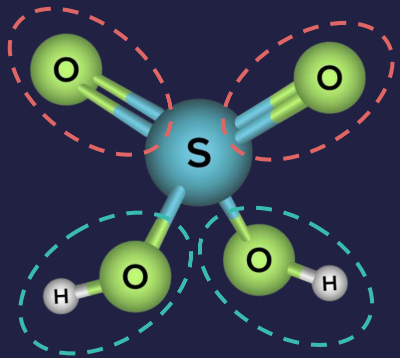
Concentrated metal-ammonia solutions have a **metallic bronze colour** & are **diamagnetic.**

$$2e^- + 2(NH_3)_y \longrightarrow [e^-(NH_3)_y]_2$$

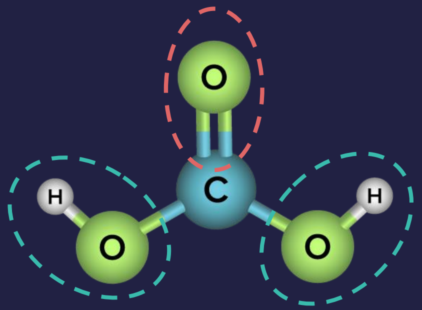
Salts of Oxo-Acids

Oxo-acids

Compounds which contain **oxygen**, at least one **hydrogen bound to oxygen**, and which produce a conjugate base by proton loss.



Sulphuric acid
H_2SO_4



Carbonic acid
H_2CO_3

Examples

$HClO_4$, H_2SO_4, H_2CO_3 etc.

Carbonates

Hydrolysis of carbonate

$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$

Carbonates (M_2CO_3) are **highly stable to heat.**

M = Alkali Metal

Electropositive character (from Li to Cs) ↑

Stability of salt ↑

Thermal stability

$Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3$

Bicarbonates

Bicarbonates are decomposed at **relatively low temperatures.**

$$2MHCO_3 \xrightarrow{300^\circ C} M_2CO_3 + H_2O + CO_2$$

$LiHCO_3$ does not exist in solid form due to

High polarising power of Li^+

Uncomparable size of Li^+ cation and HCO_3^- anion

Carbonates & Bicarbonates

Solubility in water

Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3

$NaHCO_3$ < $KHCO_3$ < $RbHCO_3$ < $CsHCO_3$

Nitrates and Sulphates

Lithium nitrate when heated gives lithium oxide, Li_2O, whereas other alkali metal nitrates decompose to give the corresponding nitrite.

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$$

$$2MNO_3 \xrightarrow{\Delta} 2MNO_2 + O_2$$

Lithium sulphate when heated gives lithium oxide, Li_2O, whereas other alkali metal sulphates decompose to give the corresponding sulphite.

Anomalous Behaviour of Lithium

Anomalous behaviour

Lithium belongs to **group I** but some of its properties are different as compared to other elements in the group.

Anomalous behaviour of lithium is due to:

1. Exceptionally small size of its atom and ion
2. High polarising power (i.e. **charge/radius ratio**)

Results in the **increased covalent character** of lithium compounds.

Anomalous Behaviour of Lithium

1 Li is much **harder**. M.P. and B.P. are **higher** than other alkali metals.

2 LiCl is deliquescent and crystallises as a hydrate, **$LiCl.2H_2O$.**

Whereas, other alkali metal chlorides **do not** form hydrates.

3 Li is the **least reactive** but is the strongest **reducing agent.**

4 **Combustion** in air

Li forms monoxide (**Li_2O**) and nitride (**Li_3N**).

Anomalous Behaviour of Lithium

5

Lithium hydrogen carbonate is not obtained in the solid form.

All other elements form solid hydrogen carbonates.

6

Lithium forms no **ethynide** on reaction with ethyne.

Other alkali metals form ethynide on reaction with ethyne.

Anomalous Behaviour of Lithium

7 Lithium nitrate when heated gives lithium oxide, **Li_2O.**

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$$

Other alkali metal nitrates decompose to give the corresponding **nitrites**.

$$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$$

8 **LiF** and **Li_2O** are comparatively much less soluble in water than the corresponding compounds of other alkali metals.

Similarities Between Li and Mg

Lithium shows some similarities with magnesium (Group 2 element , diagonal to Li in periodic table). It is called **diagonal relationship.**

Similarity arises because of their **similar ionic sizes**

Atomic radii		Ionic radii	
Li	152 pm	Li^{+}	76 pm
Mg	160 pm	Mg^{2+}	72 pm

Similarities Between Li and Mg

1. Both Li and Mg are **harder and lighter**.

2. Li and Mg **react slowly** with water.

3. **Li_2O** and **MgO** do not combine with excess oxygen to give any super oxides.

4. Their carbonates decompose easily on heating to form **oxides** and **CO_2**.

Solid hydrogen carbonates are **not** formed by Li and Mg.

Similarities Between Li and Mg

5 Both **LiCl** and **$MgCl_2$** are deliquescent and crystallise as hydrates, **$LiCl.2H_2O$** & **$MgCl_2.6H_2O$**.

6 Both **LiCl** and **$MgCl_2$** are soluble in ethanol.

7 Both form nitrides, **Li_3N** and **Mg_3N_2**, by direct combination with nitrogen.

Compounds of Sodium & Potassium

Sodium Hydroxide (NaOH): Preparation

NaOH is prepared by the electrolysis of brine solution in **Castner-Kellner cell.**

Cathode — **Mercury**

Anode — **Carbon**

Sodium Hydroxide (NaOH): Preparation

Anode (+)

$$Cl^- \longrightarrow \frac{1}{2} Cl_2 + e^-$$

Cathode (-)

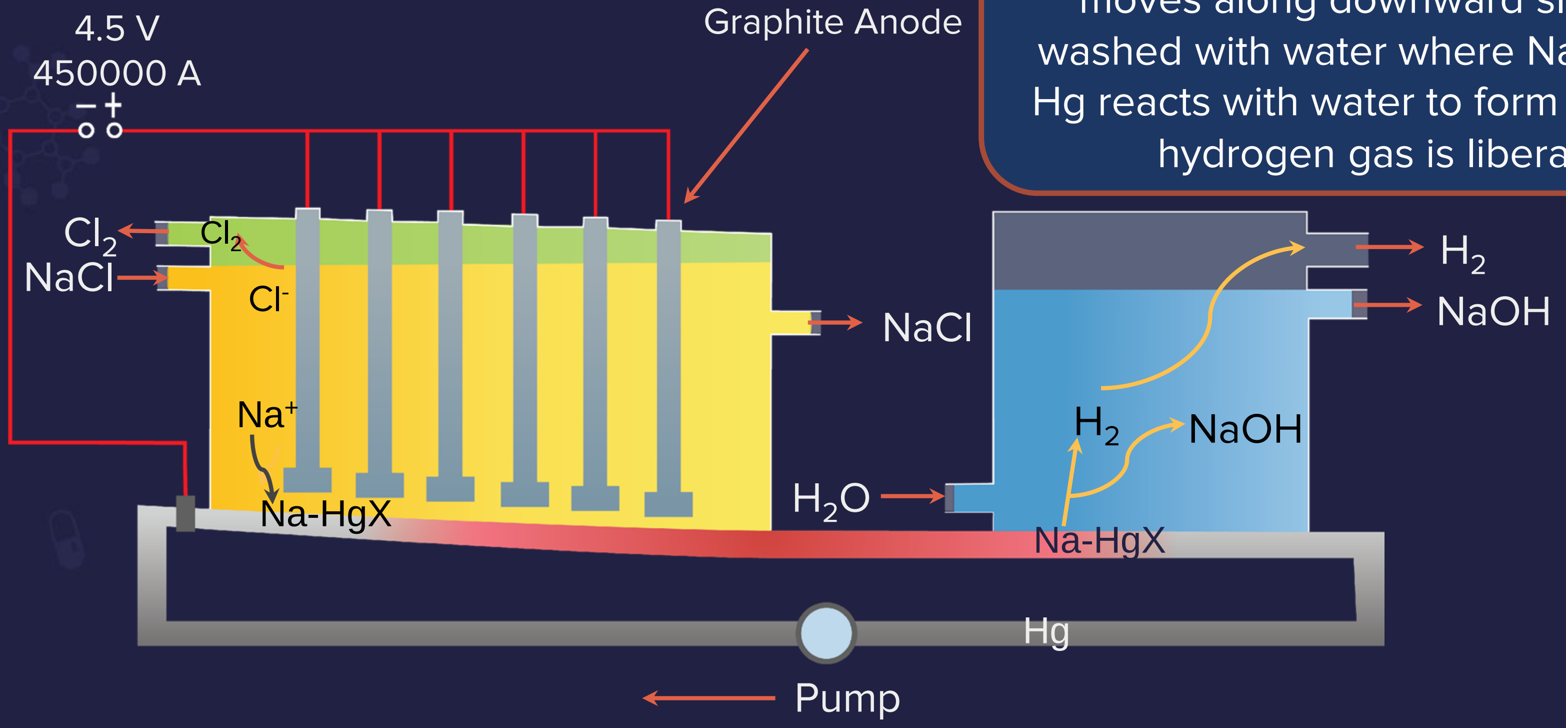
$$Na^+ + e^- \longrightarrow Na$$

$$Na + Hg \longrightarrow \text{Na-amalgam}$$

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

Castner-Kellner Cell

Na$^+$ reduced to Na at cathode and mix with Hg forming sodium amalgam. It moves along downward slope and washed with water where Na free from Hg reacts with water to form NaOH and hydrogen gas is liberated.



Properties of Sodium Hydroxide (NaOH)

1. **White** translucent solid
2. Melting point is **591 K**
3. **Highly soluble** in water to give strong alkaline solution
4. Crystals of NaOH are **deliquescent**

Chemical Properties of NaOH

a) $CO_2 + 2NaOH \longrightarrow Na_2CO_3 + H_2O$

b) $Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$

c) $4P + 3NaOH + 3H_2O \longrightarrow PH_3 + 3NaH_2PO_2$

d) $6NaOH + 4S \xrightarrow{\Delta} Na_2S_2O_3 + 2Na_2S + H_2O$

Chemical Properties of NaOH

e) $2Al + 2NaOH + 2H_2O \longrightarrow 3H_2 + 2NaAlO_2$

f) $FeCl_3 + 3NaOH \longrightarrow \underset{\text{Brown ppt.}}{Fe(OH)_3\downarrow} + 3NaCl$

g) $NaOH + CO \xrightarrow{450\ K} HCOONa$

Uses of sodium hydroxide

- Petroleum refining
- Preparation of pure fats and oils
- In textile industries for mercerising cotton fabrics
- Manufacturing of soap, paper
- Manufacturing of artificial silk
- Purification of bauxite

Sodium Carbonate
(Washing Soda, $Na_2CO_3.10H_2O$)

Preparation by Solvay Process

a **In ammonia absorber**

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

$$CaCl_2 + (NH_4)_2CO_3 \longrightarrow CaCO_3\downarrow + 2NH_4Cl$$

$$MgCl_2 + (NH_4)_2CO_3 \longrightarrow MgCO_3\downarrow + 2NH_4Cl$$

NH_3 is highly soluble in water. With CO_2, it forms ammonium bicarbonate. Calcium and magnesium salts are precipitated as carbonates and removed from the reaction.

Preparation by Solvay Process

b In carbonation tower

$$NH_3 + CO_2 + H_2O \longrightarrow NH_4HCO_3$$

$$NH_4HCO_3 + NaCl \xrightarrow{30^oC} NaHCO_3 + NH_4Cl$$

Ammonium bicarbonate on reacting with NaCl, gives $NaHCO_3$ (very less soluble in water).

Preparation by Solvay Process

c **Calcination** to get sodium carbonate

$$2NaHCO_3 \xrightarrow{150^{o}C} Na_2CO_3 + CO_2 + H_2O$$

d **In recovery tower**

$$NH_4HCO_3 \xrightarrow[\text{steam}]{\Delta} NH_3 + CO_2 + H_2O$$

$$2NH_4Cl + Ca(OH)_2 \xrightarrow[\text{steam}]{\Delta} 2NH_3 + 2H_2O + CaCl_2$$

NH_3 and CO_2 produced during reactions is utilised to produce more $NaHCO_3$

Remember

Solvay process cannot be extended to the manufacture of K_2CO_3 because $KHCO_3$ is soluble in water.

Properties

1 Sodium carbonate is a white crystalline solid which exists as a decahydrate, **$Na_2CO_3 \cdot 10H_2O$**.

On heating, the decahydrate loses its water of crystallisation to form monohydrate, **$Na_2CO_3.H_2O$.**

$$Na_2CO_3.10H_2O \xrightarrow{373\ K} Na_2CO_3.H_2O + 9H_2O$$

Properties

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

$$Na_2CO_3.H_2O \xrightarrow{> 373\ K} Na_2CO_3 + H_2O$$

Soda ash

2 It is readily **soluble** in water

Carbonate part of sodium carbonate gets **hydrolysed** by water to form an alkaline solution.

$$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

Uses of sodium carbonate

In water softening, laundering, and cleaning

In the manufacture of glass, soap, borax, and caustic soda

Used in qualitative and quantitative analysis

In paper, paints and textile industries

Preparation and Properties of $NaHCO_3$

Preparation

$NaHCO_3$ can be prepared by **bubbling carbon dioxide** through a **saturated solution** of the **carbonate**.

$$Na_2CO_3\,(aq) + CO_2\,(g) + H_2O\,(l) \longrightarrow 2NaHCO_3\,(s)$$

Properties

1. **White crystalline** solid
2. **Less soluble** than sodium carbonate in water
3. On heating, it loses CO_2 and H_2O forming **Na_2CO_3**

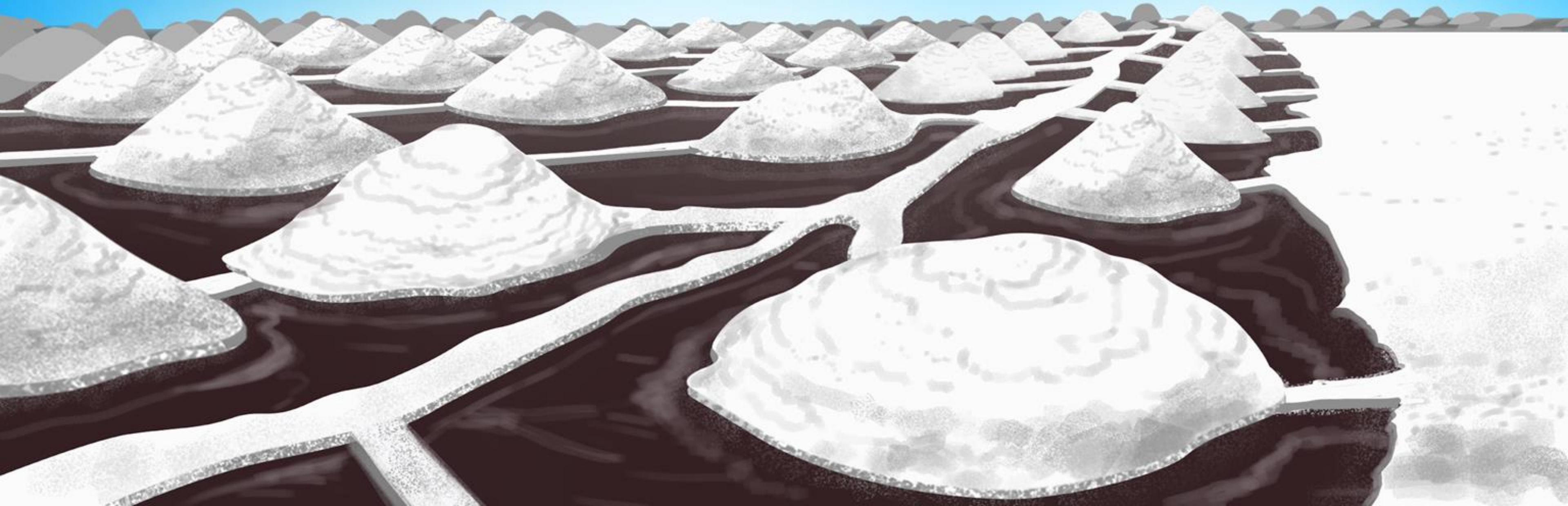
$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

Uses

1. Used for making **baking powder.**
2. Mild antiseptic for skin infections.
3. Used in fire extinguishers.

Sodium Chloride (NaCl)

Found in nature as **rock salt** or in **sea water**.



Preparation of NaCl

1 **Evaporation of Sea Water**

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

2 **Crystallization of Brine solution**

Crude NaCl contains sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride as impurities.

Properties

1. It is **non-hygroscopic,** but the presence of $MgCl_2$ in common salt renders it hygroscopic.

2. **NaCl** melts at **1073 K.**

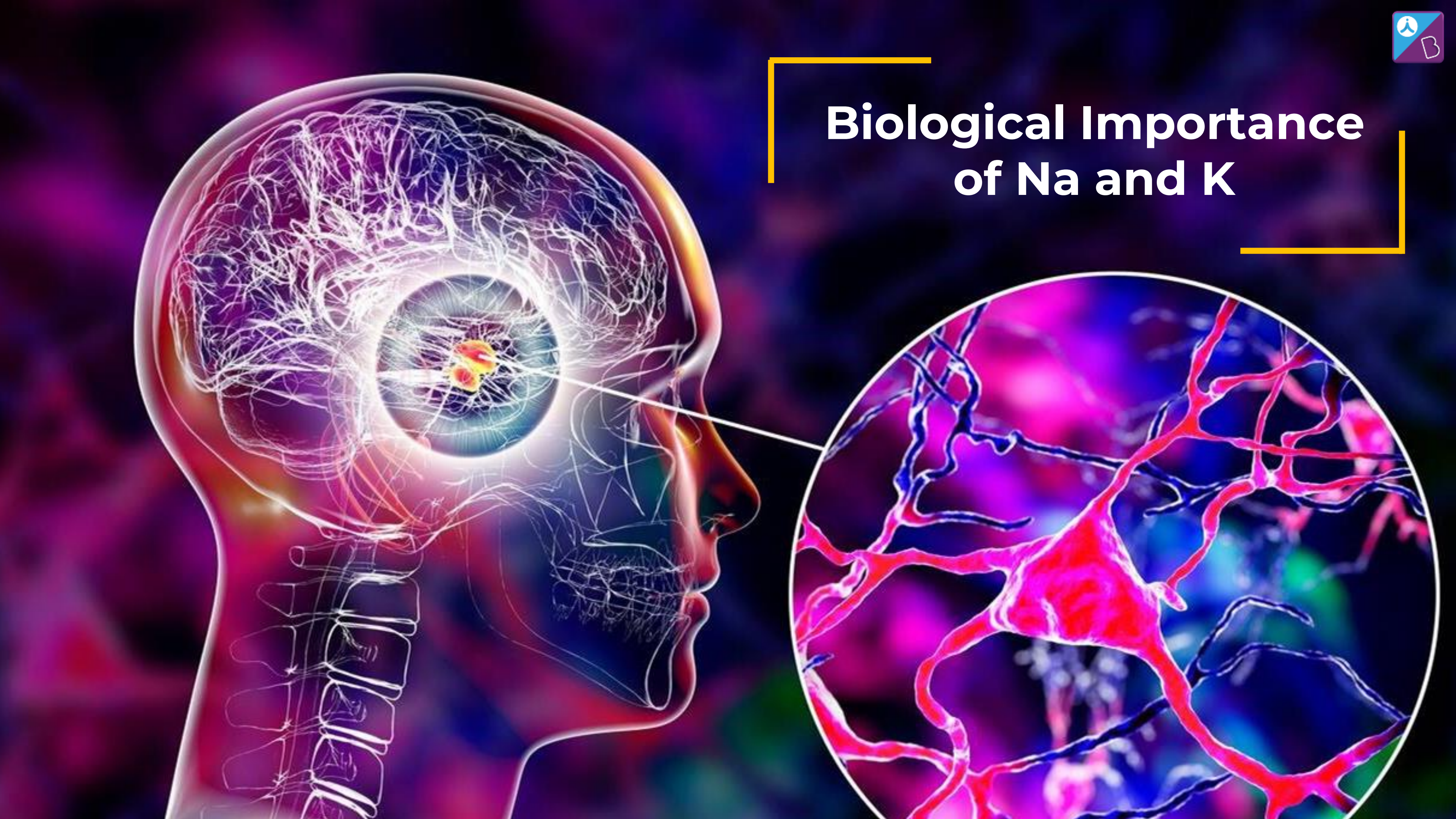
3. It has a solubility of 36.0 g in 100 g of water at 298 K.

Uses

1 Used as common salt for **domestic purpose**.

2 Used for the **preparation of Na_2O_2, $NaOH$ and Na_2CO_3**

Biological Importance of Na and K



Biological Importance of Na and K

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

Na^+ participates in the **transmission of nerve signals, in regulating the flow of water across cell membranes** and in **transporting sugars and amino acids into cells**.

Biological Importance of Na and K

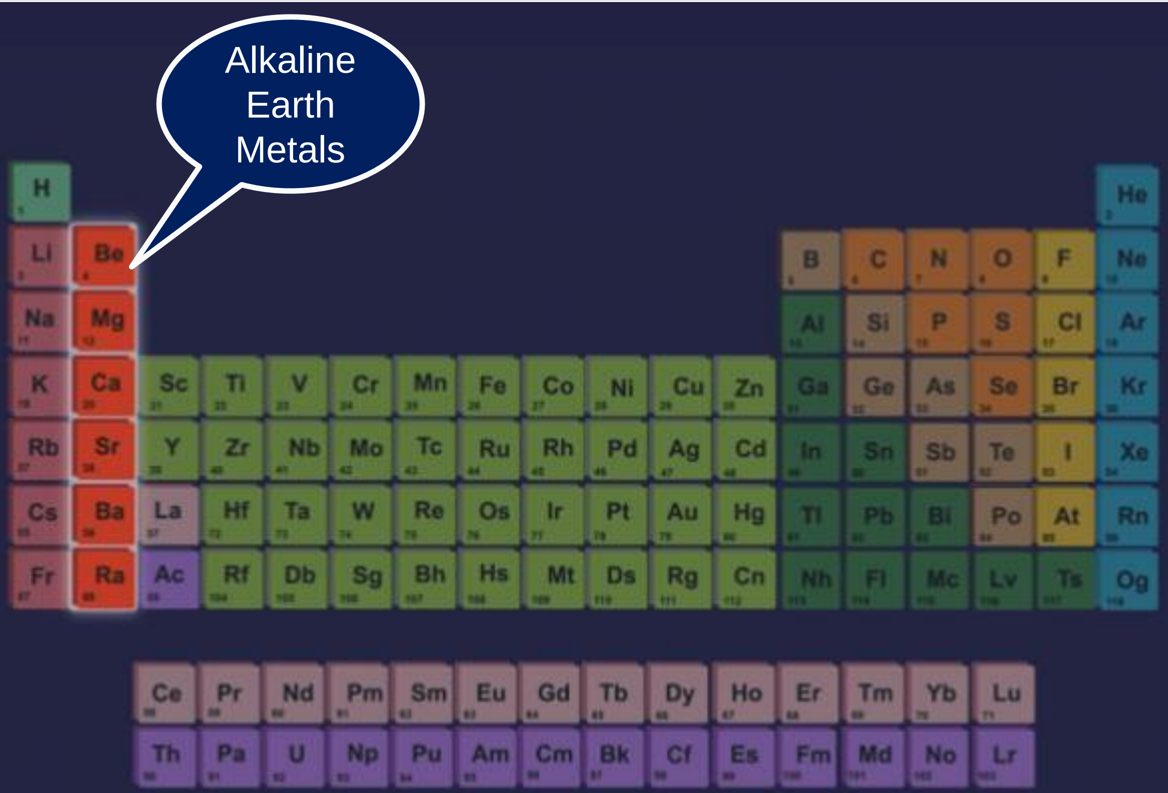
Most abundant **cations** within cell fluids where they activate many enzymes

K^+ ions

With **sodium**, responsible for the transmission of **nerve signals**

Participate in the **oxidation** of **glucose** to produce **ATP**

Alkaline Earth Metals


Alkaline Earth Metals
H He
Li Be B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba La Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Ac Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og
Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

Electronic Configuration

Electronic configuration:
[Noble gas] ns^2

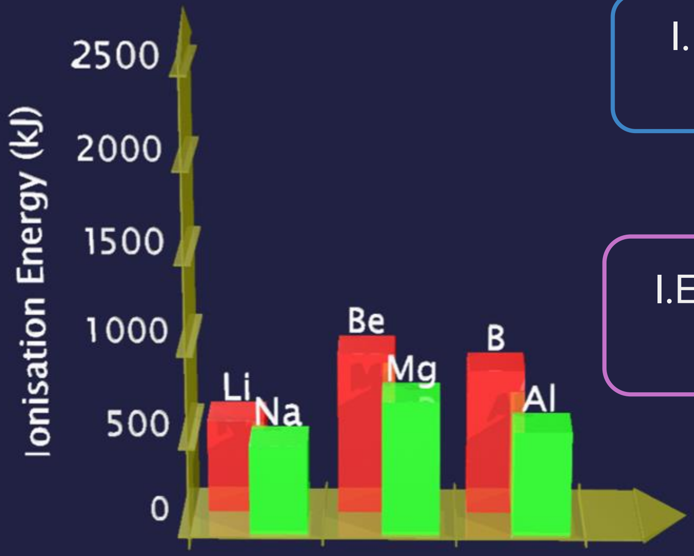
The compounds of these elements are predominantly **ionic** (except Be).

Atomic and Ionic Radii

Alkaline earth metals < Corresponding **alkali metals**

Down the group → Atomic/Ionic radii ↑

Ionisation Enthalpy



I.E.$_1$ of alkali metal < I.E.$_1$ of alkaline earth metal

I.E.$_2$ of alkali metal > I.E.$_2$ of alkaline earth metal

Hydration Enthalpy

Hydration enthalpies of alkaline earth metal ions **decrease** with the **increase in ionic size** down the group.

Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}

Compounds of alkaline earth metals are **more extensively hydrated** than those of alkali metals.

$MgCl_2$ and $CaCl_2$ exist as **$MgCl_2.6H_2O$** and **$CaCl_2.6H_2O$**

While **NaCl** & **KCl** do **not** form such **hydrates**

Physical Properties

Physical Properties

Silver-coloured (Be & Mg are greyish)

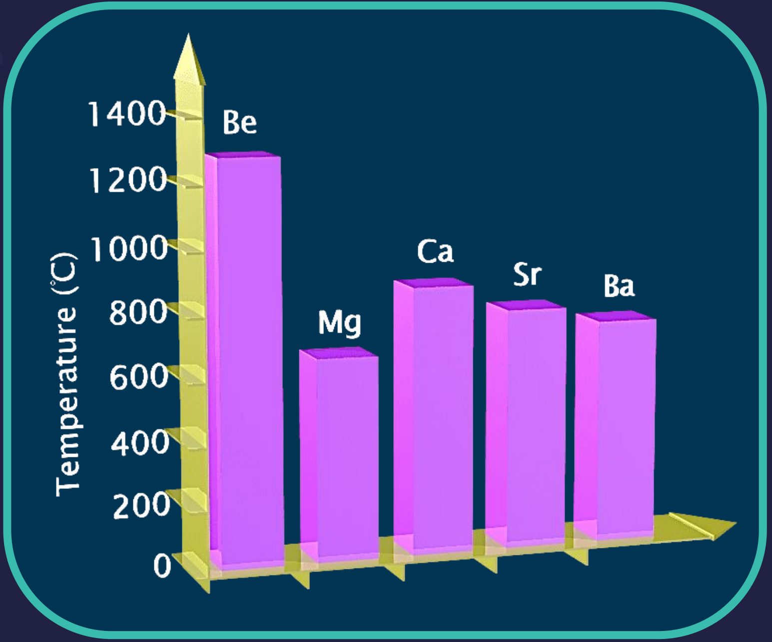
Relatively soft (Harder than alkali metals)

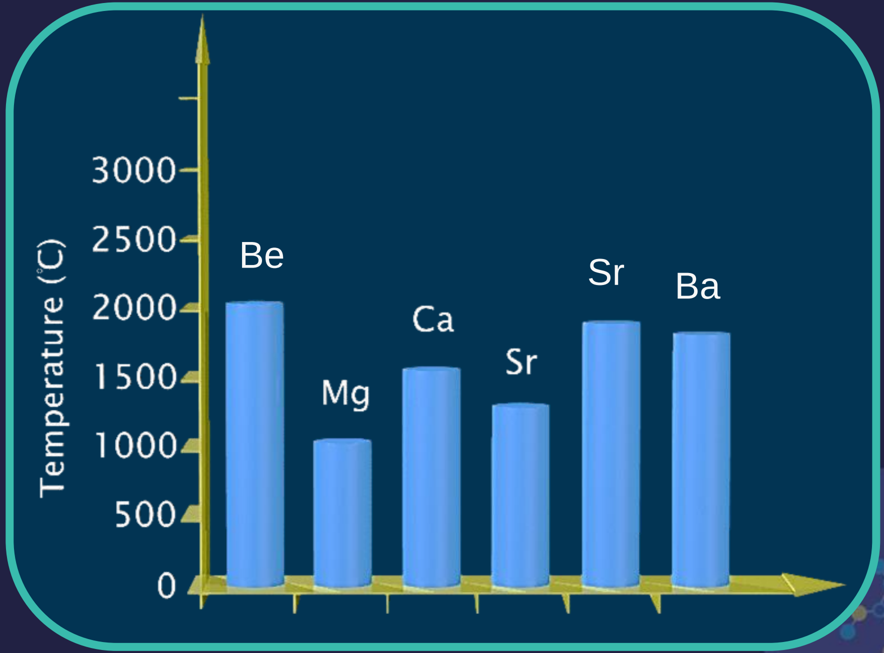
Strongly **electropositive** in nature

High **electrical** and **thermal conductivities**

Physical Properties

They have **low M.P.** and **B.P.** but are higher than the corresponding alkali metals.





Flame Test

The alkaline earth metals and their salts impart characteristic color to an **oxidising flame**.

Brick red	Crimson red	Apple-green

The electrons in **Be** & **Mg** are **strongly bound** to get excited by flame.

Do not impart any colour to the flame.

Chemical Properties

a Reaction with Air

Alkaline earth metals are **less reactive** than the alkali metals.

The reactivity of these elements **increases** down the group.

Be is inert in air as its surface is passivated by the formation of a thin layer of BeO.

Mg & Ca also tarnish in air with the formation of an **oxide layer**, but will burn completely to their oxides and nitrides when heated.

Sr and **Ba** are readily attacked by air.

Chemical Properties

Powdered Be burns brilliantly

$$2Be + O_2 \xrightarrow{\Delta} 2BeO \text{ (amphoteric)}$$

$$3Be + N_2 \xrightarrow{\Delta} Be_3N_2$$

All the Group 2 elements form normal oxides with oxygen **except Ba**, which forms the peroxide.

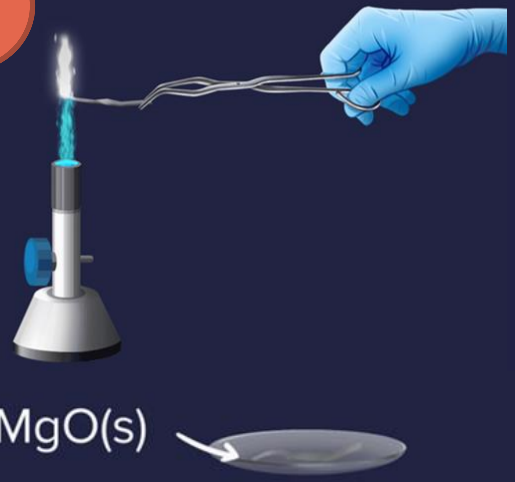
Reaction with Air

1

2

Mg burns with dazzling brilliance in air to give **MgO** and **Mg_3N_2.**

3



$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

Oxides

Oxides of alkaline earth metals are **basic** in nature (**Except BeO).**

They react with water to form **hydroxides.**

$$MO + H_2O \longrightarrow M(OH)_2$$

(M = Alkaline earth metal)

Nitrides

They react with water to form **hydroxides.**

$$Be_3N_2 + 6H_2O \longrightarrow 3Be(OH)_2 + 2NH_3 \uparrow$$

$$Mg_3N_2 + 6H_2O \longrightarrow 3Mg(OH)_2 + 2NH_3 \uparrow$$

Chemical Properties

b Reaction with Water

Mg decomposes **boiling water** but **Be** is not attacked by **water** even at high temperatures.

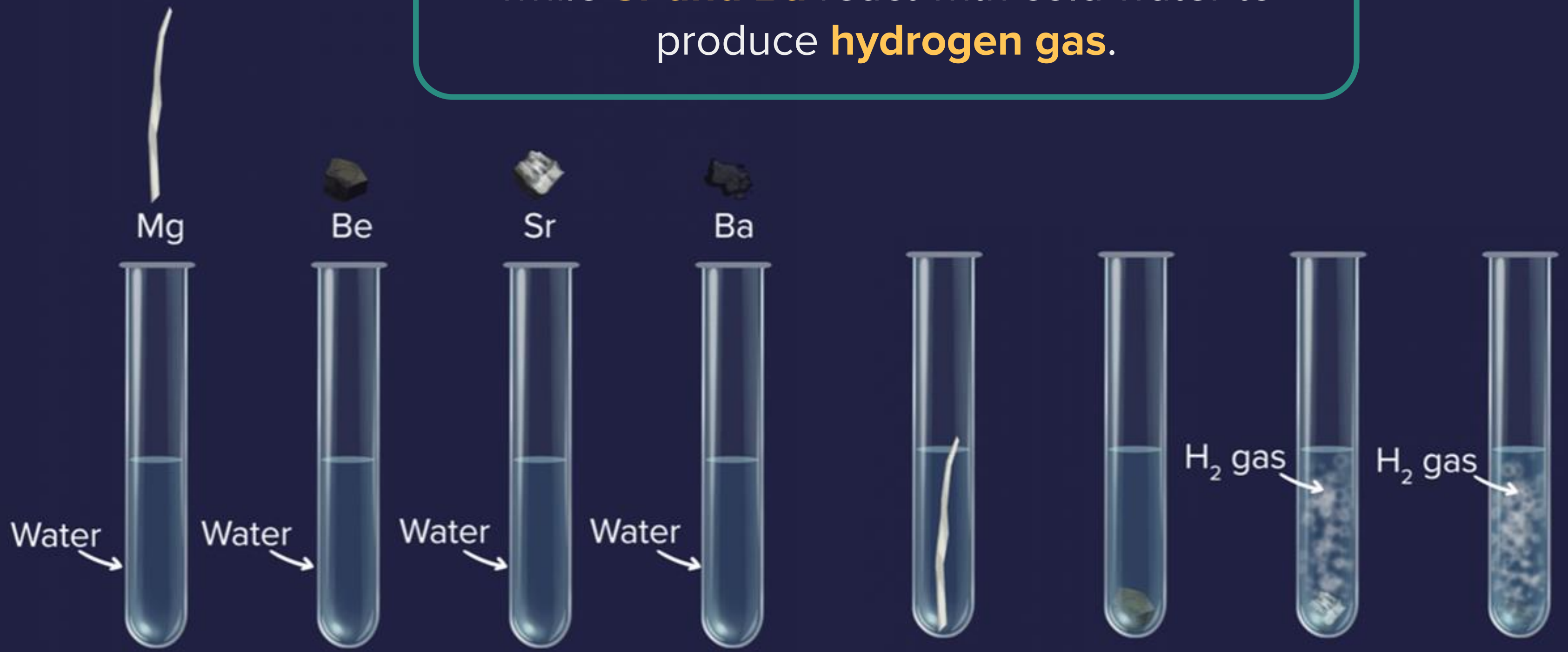
$$Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2$$

Ca, Sr, Ba, and **Ra** decompose cold water readily with the evolution of **hydrogen.**

$$M + 2H_2O \longrightarrow M(OH)_2 + H_2$$

Reaction with Water

Mg and Be do not react with cold water while **Sr and Ba** react with cold water to produce **hydrogen gas**.



Hydroxides

$Mg(OH)_2$ < $Ca(OH)_2$ < $Sr(OH)_2$ < $Ba(OH)_2$

Order of solubilities, thermal stabilities, and basic character.

Alkaline earth metal hydroxides are **less basic and less stable** than alkali metal hydroxides.

Hydroxides

Beryllium hydroxide is **amphoteric** in nature.

$$Be(OH)_2 + 2OH^- \longrightarrow [Be(OH)_4]^{2-}$$

$$Be(OH)_2 + H_2SO_4 \longrightarrow BeSO_4 + 2H_2O$$

Chemical Properties

C Reaction with Halogens

Group 2 elements directly combine with halogens on heating to give **metal halides.**

$$M + X_2 \longrightarrow MX_2$$

(X = F, Cl, Br, I)
(M = Alkaline earth metal)

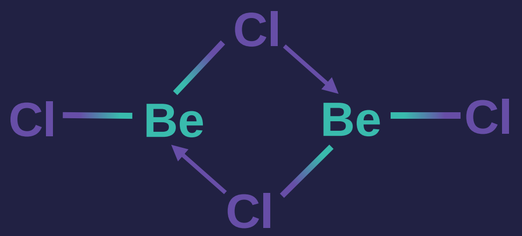
BeF_2 is best formed by the thermal decomposition of **$(NH_4)_2BeF_4$.**

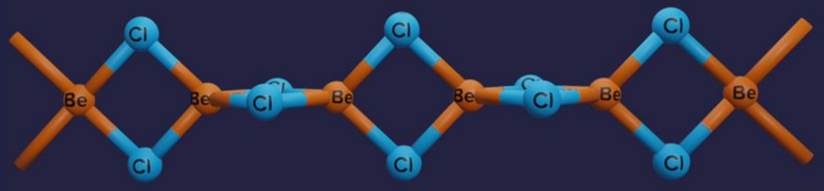
$BeCl_2$ is conveniently made from the oxide.

$$BeO + C + Cl_2 \xrightleftharpoons{600\text{-}800\ K} BeCl_2 + CO$$

Beryllium Chloride

$BeCl_2$ forms **chloro-bridged dimer** in the **vapour** phase.





Beryllium chloride has a **chain structure** in the solid state.

Halides

Halides of alkaline earth metals are **ionic** in nature (except for **BeX_2**).

Ionic character of halides **increases** from Be to Ra.

The tendency to form halide hydrates gradually **decreases down the group.**

Mg > Ca > Sr > Ba

Halides

The dehydration of hydrated chlorides, bromides, and iodides of **Ca**, **Sr**, and **Ba** can be achieved on **heating.**

The corresponding hydrated halides of **Be** and **Mg** on heating undergo hydrolysis.

$$[Be(H_2O)_4]Cl_2 \xrightarrow{\Delta} Be(OH)_2 + 2H_2O + 2HCl$$

Halides

Down the group → Size of the metal ion ↑ → Hydration energy ↓ → **Solubility of halides** ↓

The **fluorides** are relatively **less soluble** than the chlorides due to their high lattice energies.

Chemical Properties

d Reaction with dihydrogen

Except Be, all the alkaline earth metals form hydrides (**MH_2**) on heating directly with H_2.

BeH_2 can be prepared by the action of **$LiAlH_4$** on **$BeCl_2$**

$$2BeCl_2 + LiAlH_4 \longrightarrow 2BeH_2 + LiCl + AlCl_3$$

Chemical Properties

e

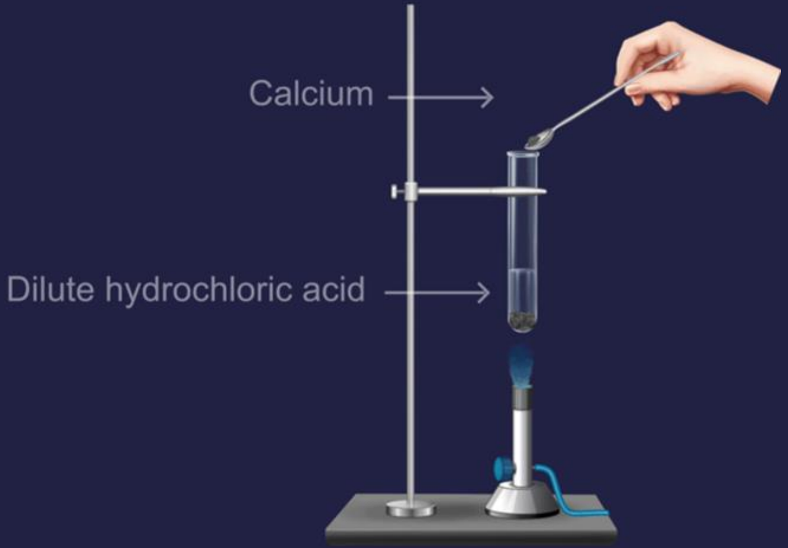
Reaction with acids

Alkaline earth metals readily react with **acids** liberating **dihydrogen**.

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

(M = Alkaline earth metal)

Reaction with Acids



$$Ca\,(s) + 2HCl\,(aq) \xrightarrow{\text{Heat}} CaCl_2\,(aq) + H_2\,(g)$$

Calcium liberates hydrogen gas when reacted with acids like HCl.

Chemical Properties

f

Reducing nature

Alkaline earth metals are **strong reducing agents**

Indicated by the **large negative** values of their **reduction potentials**

Reducing Nature

Reducing power of alkaline earth metals is **less** than that of their **corresponding alkali metals.**

Be has a less negative value compared to other alkaline earth metals.

Its reducing nature is due to the **large hydration energy** associated with the small size of Be^{2+} ion and relatively **large value of the atomisation enthalpy** of the metal.

Chemical Properties

g Solutions in liquid ammonia

Dissolve in liquid ammonia to give **deep blue black** solutions forming **ammoniated ions.**

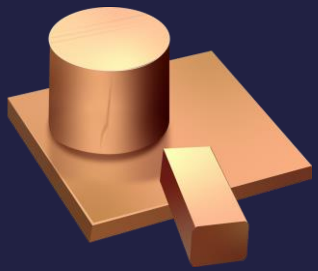
$$M + (x + 2y)NH_3 \longrightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^-$$

Uses of Alkaline Earth Metals

Uses of Beryllium

Used in the manufacture of **alloys**

For making windows of **X-ray tubes**



Uses of Magnesium

Milk of magnesia is used as an **antacid.**

$MgCO_3$ is an ingredient of **toothpaste.**

Uses of Alkaline Earth Metals

Uses of Calcium

In the **extraction of metals from oxides** which are difficult to reduce with carbon.

Used to remove air from **vacuum tubes.**

Radium salts are used in **radiotherapy.**

General Characteristics of Compounds of the Alkaline Earth Metals

General Characteristics

Group 2 metals form compounds which are **predominantly ionic,** but **less ionic** than the corresponding compounds of alkali metals.

Carbonates

Carbonates of alkaline earth metals are **insoluble** in water.

Their solubility **decreases down the group.**

Carbonates

Solubility

$BeCO_3$ > $MgCO_3$ > $CaCO_3$ > $SrCO_3$ > $BaCO_3$

Order of Thermal Stability

$BeCO_3$ < $MgCO_3$ < $CaCO_3$ < $SrCO_3$ < $BaCO_3$

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

Bicarbonates

Bicarbonates of alkaline earth metals **do not exist in solid state** but are known to exist in solution.

$$M(HCO_3)_2 \xrightarrow{\Delta} MCO_3 + CO_2 + H_2O$$

Sulphates

Sulphates

The sulphates of the alkaline earth metals are all **white solids** and stable to **heat**.

Thermal Stability

$BeSO_4 < MgSO_4 < CaSO_4 < SrSO_4 < BaSO_4$

Solubility in water

$BeSO_4 > MgSO_4 > CaSO_4 > SrSO_4 > BaSO_4$

General Characteristics

Nitrates

Hydrated nitrates, such as **$Ca(NO_3)_2.4H_2O$**, can be obtained by treating the oxides, hydroxides, & carbonates with **nitric acid** & **crystallising the salt** from the resulting aqueous solution.

Nitrates decompose on **heating** to give the corresponding **oxides** with evolution of a mixture of nitrogen dioxide and oxygen.

$$2M(NO_3)_2 \longrightarrow 2MO + 4NO_2 + O_2$$

(M = Group 2 metals)

Anomalous Behaviour of Be

Properties of Be differ from the rest of the group 2 elements because of:

- Its **small size** and **high polarising power**.
- Relatively **high E.N.** and **I.E.** as compared to other members.
- Absence of **vacant d-orbitals** in its valence shell.

Anomalous Behaviour of Be

1 Reaction with water

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

2 Nature of oxides

BeO is amphoteric while **MgO is weakly basic**.

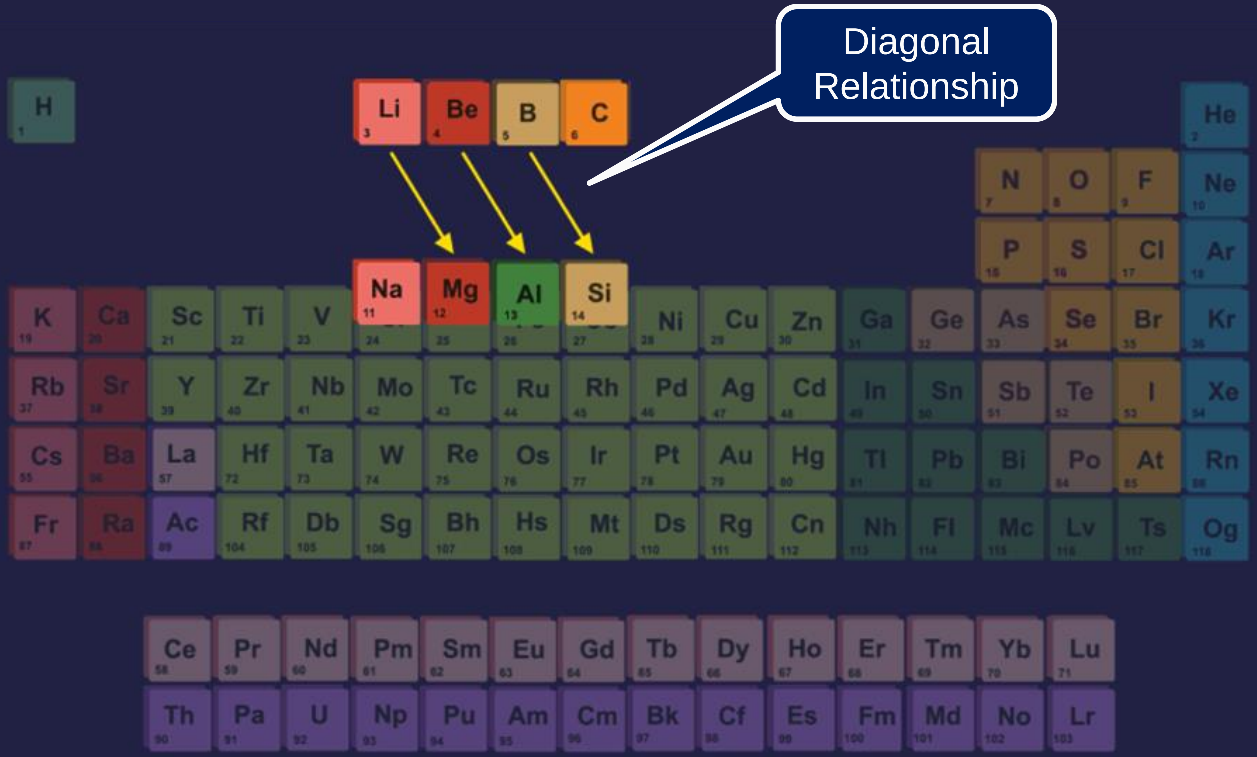
3 Nature of compounds

Be forms covalent compounds, whereas **other members form ionic** compounds.

4 Coordination number (C.N.)

Be does not exhibit C.N. more than 4 as it has only four orbitals in its valence shell while **other members** of this group can have **C.N. = 6**

Diagonal Relationship between Be and Al



Diagonal Relationship

1 Reaction with acids

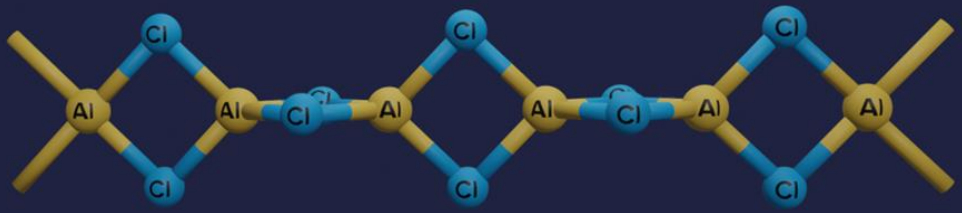
Like Al, Be is **not readily attacked by acids** because of the presence of an oxide film.

2 Nature of hydroxide

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

Diagonal Relationship

3 Polymeric structure



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

Some Important Compounds of Calcium

Calcium Carbonate ($CaCO_3$) : Preparation

a It can be prepared by passing **carbon dioxide** through slaked lime.

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

Excess of CO_2 should be avoided as it leads to the formation of water soluble **$Ca(HCO_3)_2$**

b By the addition of **sodium carbonate** solution to **$CaCl_2$**.

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaCl$$

Properties of $CaCO_3$

a It is a **white fluffy powder,** almost **insoluble in water.**

b It decomposes to give CO_2 when heated at a high temperature.

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

c 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 of calcium carbonate

- Building material (marble)
- Manufacture of quick lime
- Manufacture of high quality paper
- Antacid
- Mild abrasive in toothpaste
- Filler in cosmetics

Calcium oxide (CaO) : Preparation

Also known as **quick lime**

It can be obtained by **decomposing limestone** at a high temperature.

$$CaCO_3 \overset{\Delta}{\rightleftharpoons} CaO + CO_2$$

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

Properties of CaO

a. It is a **white amorphous powder** of melting point **2843 K.**

b. On exposure to atmosphere, it absorbs moisture and CO_2.

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

d. It combines with some **acidic oxides** at **high temperatures.**

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

$$CaO + CO_2 \longrightarrow CaCO_3$$

$$CaO + SiO_2 \longrightarrow CaSiO_3$$

$$6CaO + P_4O_{10} \longrightarrow 2Ca_3(PO_4)_2$$

Properties of CaO

$$CaO + 3C \xrightarrow{2300\ K} CaC_2 + CO$$

$$CaC_2 \xrightarrow{2H_2O} Ca(OH)_2 + C_2H_2$$

$$CaC_2 \xrightarrow[\Delta]{N_2} Ca(NCN) + C$$

$$Ca(NCN) \xrightarrow{3H_2O} 2NH_3 + CaCO_3$$

Uses of calcium oxide

Manufacturing cement

Purification of sugar

Manufacture of Na_2CO_3 from caustic soda

Manufacture of dye stuffs

Calcium Hydroxide($Ca(OH)_2$) :Preparation

By spraying water on quick lime

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

Also called as slaked lime.

Properties of $Ca(OH)_2$

a It is a **white amorphous powder.**

b It is **sparingly soluble** in water.

c Its solubility in **hot water** is **less** than that in **cold water**.

d The **aqueous solution** is known as **lime water** and a **suspension** of **slaked lime** in water is known as **milk of lime.**

e When CO_2 is passed through lime water, it turns **milky** due to the formation of calcium carbonate.

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

Properties of $Ca(OH)_2$

On passing excess of CO_2, **calcium hydrogen carbonate** is formed.

$$CaCO_3 + CO_2 + H_2O \longrightarrow Ca(HCO_3)_2$$

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

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

Bleaching powder

Uses of calcium hydroxide

Preparation of mortar (Building material)

Preparation of sugar

White wash (Due to its disinfectant nature)

Glass making (Tanning industry)

Calcium Sulphate (Plaster of Paris)

$$[CaSO_4.\frac{1}{2}H_2O]$$

Preparation of Plaster of Paris

Obtained when gypsum, is heated at **120°C (393 K).**

$$\underset{\textbf{Gypsum}}{CaSO_4.2H_2O} \xrightarrow[\Delta]{393\ K} \underset{\textbf{Plaster of Paris}}{CaSO_4.\frac{1}{2}H_2O} + \frac{3}{2}H_2O$$

Properties of Plaster of Paris

a

It has the **property of setting** with water.

b

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

Uses

1. For immobilising the affected part of organ where there is a **bone fracture.**

2. For making **casts** of statues, etc.

3. In making **blackboard chalks**.

Cement

Cement is a product obtained by combining a material rich in **lime**, **CaO** with other materials such as **clay** which contain silica, **SiO_2** along with the **oxides of Al, Fe,** & **Mg.**

Raw materials

- **Limestone**
- **Clay**

Preparation

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 ($CaSO_4.2H_2O$)** to form cement.

Cement

Important ingredients present in **Portland cement**

- Dicalcium silicate (Ca_2SiO_4) 26%
- Tricalcium silicate (Ca_3SiO_5) 51%
- Tricalcium aluminate ($Ca_3Al_2O_6$) 11%

Setting of Cement

When mixed with **water**, the setting of cement takes place to give a **hard mass.**

This is due to the **hydration of the molecules** of the constituents & their rearrangements.

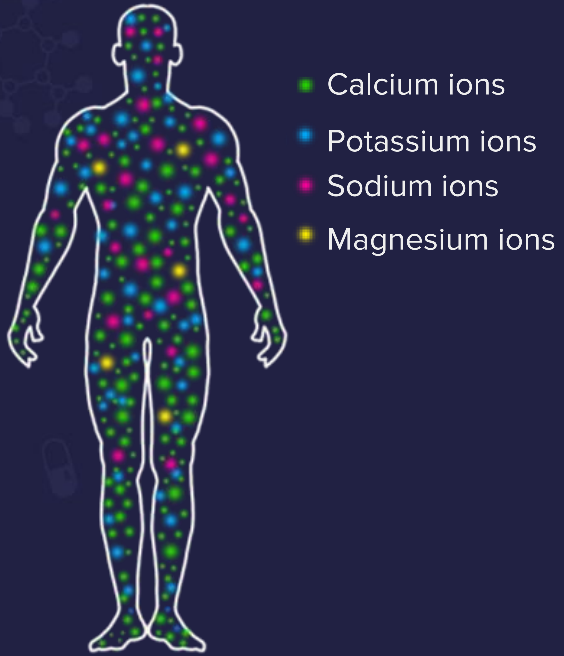
The purpose of adding gypsum is only to **slow down** the process of setting of the cement, so that it gets sufficiently hardened.

Uses of Cement

- Construction of bridges, dams, and buildings
- Concrete & reinforced concrete
- Plastering

Biological Importance of Mg and Ca

An adult body contains about **25 g of Mg** and **1200 g of Ca** compared to only **5 g of Fe** and **0.06 g of Cu.**



All enzymes that utilise **ATP in phosphate** transfer, require **magnesium** as the **cofactor**.

The main pigment for the absorption of light in plants is **chlorophyll**, which contains **magnesium**.

Biological Importance of Mg and Ca

The calcium concentration in **plasma** is regulated at about **100 mg/L**. It is maintained by **calcitonin** and **parathyroid hormone**.

About 99% of calcium in a human body is present in **bones and teeth**.