S block

In 1869 Russian chemist Dimitri Mendeleev started the development of the periodic table, arranging chemical elements by atomic mass. He predicted the discovery of other elements. The elements in which the last electron enters the outermost s-orbital are called s-block elements. The s-block elements have two groups 1 and 2.

The Group 1 elements are called alkali metals. The Group 2 elements are called alkaline earth metals

Alkali Metals - Group 1

Lithium, Sodium, Potassium, Rubidium, Caesium, and Francium.

Alkaline Metals- Group 2

Beryllium, Magnesium, Calcium, Strontium, Barium, and Radium

1 H	→ Alkali Metals
3 Li	4 Be
11 Na	12 Mg
19 K	20 Ca → Alkaline Earth Metals
37 Rb	38 Sr
55 Cs	56 Ba
87 Fr	88 Ra

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They readily lose electron to give monovalent M^+ ions. Hence they are never found in free state in nature.

Atomic And Ionic Radii

Li < Na < K < Rb < Cs

Increases down the group, because value of n (principal quantum number) increases.

Ionization Enthalpy

Li > Na > K > Rb > Cs

As size increases ionization enthalpy decreases.

Hydration Enthalpy

The hydration enthalpies of alkali metal ions decrease with increase in ionic sizes.

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

As Φ increases Hydration Energy increases.

Metal	Li	Na	K	Rb	Cs
Colour	Crimson Red	Yellow	Violet	Red Violet	Blue

(i) Reactivity Towards Water

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

$2M + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2$

(M = an alkali metal)

(ii) Reactivity Towards Halogens

The alkali metals react vigorously with halogens to form ionic halides, M^+X^-. However, lithium halides are somewhat covalent. It is because of the high polarisation capability of lithium ion.

(iii) Reducing Nature

The alkali metals are strong reducing agents, lithium being the most and sodium the least powerful.

(vi) Solutions In Liquid Ammonia

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

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

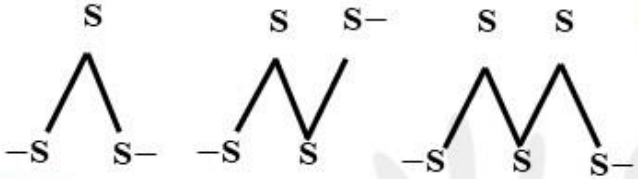
The blue colour of the solution is due to the ammoniated electron which absorbs energy in the visible region of light and thus imparts blue colour to the solution. The solutions are paramagnetic and on standing slowly liberate hydrogen resulting in the formation of NH_2^-.

$M^{+}_{(am)} + e^{-} + NH_3 \longrightarrow MNH_{2(am)} + \frac{1}{2}H_2(g)$

where 'am' denotes solution in ammonia.

(iv) Sulphides

All alkali metals react with S forming sulphides such as Na_2S and Na_2S_n (n = 2, 3, 4, 5 or 6). The polysulphide ions are made from zig-zag chains of sulphur atoms.



Uses

1. Lithium metal is used to make useful alloys, for example with lead to make 'white metal' bearings for motor engines.
2. With aluminium to make aircraft parts, and with magnesium to make armour plates.

Sodium Peroxides (Na_2O_2)

Preparation

It is formed by heating the metal in excess of air or oxygen at 300°, which is free from moisture and CO_2.

$$2Na + O_2 \longrightarrow Na_2O_2$$

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Properties

$$Na_2O_2 + 2H_2O \xrightarrow{\sim 0^\circ C} 2NaOH + H_2O_2$$

$$2Na_2O_2 + 2H_2O \xrightarrow{25^\circ C} 4NaOH + O_2$$

$$Na_2O_2 + H_2SO_4 \xrightarrow{\sim 0^\circ C} Na_2SO_4 + H_2O_2$$

Uses

(i) For preparing H_2O_2, O_2.

(ii) Oxygenating the air in submarines.

(iii) Oxidising agent in the laboratory.

Properties

(i) $2KNO_3 + 10K \xrightarrow{\text{heating}} 6K_2O + N_2$

$K_2O \xrightarrow{\text{heating}} K_2O$

(White) (Yellow)

$K_2O + H_2O \longrightarrow 2KOH$

Halides

The alkali metal halides, MX, (X=F,Cl,Br,I) are all high melting, colourless crystalline solids. They can be prepared by the reaction of the appropriate oxide, hydroxide or carbonate with aqueous hydrohalic acid (HX).

Salts of 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 e.g., carbonic acid, H_2CO_3 ($OC(OH)_2$), sulphuric acid, H_2SO_4 ($O_2S(OH)_2$).

Anomalous Properties of Lithium

The anomalous behaviour of lithium is due to the -

(i) exceptionally small size of its atom and ion.

(ii) high polarising power (i.e., charge/ radius ratio).

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

The Equations For The Complete Process May Be Written as

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

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

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

Properties

Sodium carbonate is a white crystalline solid which exists as a decahydrate, $Na_2CO_3 \cdot 10H_2O$. This is also called washing soda. It is readily soluble in water.

$$Na_2CO_2.10H_2O \xrightarrow{\text{heat}} Na_2CO_3.H_2O + 9H_2O$$

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

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Sodium Hydroxide (Caustic Soda), NaOH

Sodium hydroxide is generally prepared commercially by the electrolysis of sodium chloride in Castner-Kellner cell. A brine solution is electrolysed using a mercury cathode and a carbon anode. Sodium metal discharged at the cathode combines with mercury to form sodium amalgam.

Chlorine gas is evolved at the anode.

Cathode $Na^+ + e^- \xrightarrow{Hg} Na - amalgam$

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

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

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

Uses

(i) The manufacture of soap, paper, artificial silk and a number of chemicals.
(ii) In petroleum refining.
(iii) In the purification of bauxite.
(iv) In the textile industries for mercerising cotton fabrics.
(v) for the preparation of pure fats and oils.
(vi) for the preparation of pure fats and oils.
(vii) As a laboratory reagent.

Sodium Hydrogen Carbonate (Baking Soda), $NaHCO_3$

Sodium hydrogen carbonate is known as baking soda because it decomposes on heating to generate bubbles of carbon dioxide (leaving holes in cakes or pastries and making them light and fluffy).

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

$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$

$KCl + H_2SO_4 \longrightarrow KHSO_4 + HCl$

$KHSO_4 + KCl \longrightarrow K_2SO_4 + HCl$

$K_2SO_4 + CaCO_3 + 2C \longrightarrow K_2CO_3 + CaS + 2CO_2$

It is a white powder , deliquescent in nature. It is highly soluble in water.

Uses It is used in the manufacture of hard glass. The mixture of K_2CO_3 and Na_2CO_3 is used a fusion mixture in laboratory.

Biological Importance of Sodium And Potassium

A typical 70 kg man contains about 90 g of Na and 170 g of K 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.

Group 2 Elements : Alkaline Earth Metals

The group 2 elements comprise beryllium, magnesium, calcium, strontium, barium and radium.

They follow alkali metals in the periodic table. These (except beryllium) are known as alkaline earth metals.

Element	Symbol	Electronic configuration
Beryllium	Be	$1s^2 2s^2$
Magnesium	Mg	$1s^2 2s^2 2p^6 3s^2$
Calcium	Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Strontium	Sr	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$
Barium	Ba	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$ or $[xe]6s^2$
Radium	Ra	$[Rn]7s^2$

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Atomic And Ionic Radii

The atomic and ionic radii of the alkaline earth metals are smaller than those of the Corresponding alkali metals in the same periods.

Atomic and Physical Properties of the Alkaline Earth Metals

Property	Beryllium Be	Magnesium Mg	Calcium Ca	Strontium Sr	Barium Ba	Radium Ra
Atomic number	4	12	20	38	56	88
Atomic mass(g mol^{-1})	9.01	24.31	40.08	87.62	137.33	226.03
Electronic configuration	$[He]2s^2$	$[Ne]3s^2$	$[Ar]4s^2$	$[Kr]5s^2$	$[Xe]6s^2$	$[Rn]7s^2$
Ionization enthalpy(I)/kJ mol^{-1}	899	737	590	549	503	509

Property	Beryllium Be	Magnesium Mg	Calcium Ca	Strontium Sr	Barium Ba	Radium Ra
m.p/K	1560	924	1124	1062	1002	973
b.p/K	2745	1363	1767	1655	2078	1973
Density/g cm^{-3}	1.84	1.74	1.55	2.63	3.59	5.5
Standard potentials E^{-}/V for (M^{+}/M)	-1.97	-2.36	-2.84	-2.89	-2.92	-2.92
Occurrence in lithosphere	2*	2.76**	4.6**	384*	390*	10^{-6}*

Hydration Enthalpies

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

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

Physical Properties

(i) The alkaline earth metals, in general, are silvery white, lustrous and relatively soft but harder than the alkali metals. Beryllium and magnesium appear to be somewhat greyish.

(ii) The melting and boiling points of these metals are higher than the corresponding alkali metals due to smaller sizes. The trend is, however, not systematic.

(iii) Because of the low ionisation enthalpies, they are strongly electropositive in nature. The electropositive character increases down the group from Be to Ba.

(iv) Calcium, strontium and barium impart characteristic brick red, crimson and apple green colours respectively to the flame. In flame the electrons are excited to higher energy levels and when they drop back to the ground state, energy is emitted in the form of visible light.

(v) The alkaline earth metals like those of alkali metals have high electrical and thermal conductivities which are typical characteristics of metals.

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Chemical Properties

The alkaline earth metals are less reactive than the alkali metals. The reactivity of these elements increases on going down the group. Magnesium is more electropositive and burns with dazzling brilliance in air to give MgO and Mg3N2. Calcium, strontium and barium are readily attacked by air to form the oxide and nitride.

Reactivity Towards Water

AEM have lesser tendency to react with water as compared to AM. They form hydroxides and liberate H_2 on reaction with H_2O

$$M + 2H_2O \xrightarrow{\Delta} M(OH)_2 + H_2$$

MgO forms protective layer, that is why it does not react readily unless layer is removed amalgamating with Hg. Other metals react quite readily (Ca, Sr, Ba).

Note

$Be(OH)_2$ is amphoteric but other hydroxides are basic in nature.

Reactivity Towards Halogens

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

$$M + X_2 \longrightarrow MX_2 \ (X = F, Cl, Br, I)$$

Reducing Nature

Like alkali metals, the alkaline earth metals are strong reducing agents. This is indicated by large negative values of their reduction potentials. However their reducing power is less than those of their corresponding alkali metals.

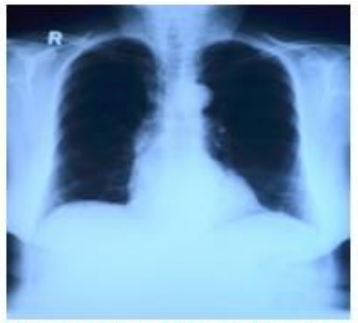
Solution In Liquid Ammonia

Like alkali metals, the alkaline earth metals dissolve in liquid ammonia to give deep blue black solutions forming ammoniated ions.

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

From these solutions, the ammoniates, $[M(NH_3)_6]^{2+}$ can be recovered.

Uses Beryllium is used in the manufacture of alloys. Copper-beryllium alloys are used in the preparation of high strength springs. Metallic beryllium is used for making windows of X-ray tubes.



General Characteristics of 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 but less ionic than the corresponding compounds of alkali metals.

Oxides And Hydroxides

The alkaline earth metals burn in oxygen to form the monoxide, MO which, except for BeO, have rock-salt structure. The BeO is essentially covalent in nature. The enthalpies of formation of these oxides are quite high and consequently they are very stable to heat.

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Oxides And Hydroxides

BeO is amphoteric while oxides of other elements are ionic in nature. All these oxides except BeO are basic in nature and react with water to form sparingly soluble hydroxides.

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

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

Beryllate ion

$Be(OH)_2 + 2HCl \longrightarrow 2H_2O + BeCl_2$

Salts of Oxoacids

The alkaline earth metals also form salts of Oxoacids.

Carbonates

Carbonates of alkaline earth metals are insoluble in water and can be precipitated by addition of a sodium or ammonium carbonate solution to a solution of a soluble salt of these metals.

The thermal stability increases with increasing cationic size.

$\phi \uparrow TS \downarrow$ $\quad MCO_3 \xrightarrow{\Delta} MO + CO_2$

The solubility of carbonates in water decreases as the atomic number of the metal ion increases.

Generally solubility increases down the group, but in case of polyatomic anion it decreases down the group when number of anion is $\geq$ number of cations.

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 and decreasing hydration enthalpy. All of them decompose on heating to give

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

(M: Be, Mg, Ca, Sr, Ba)

Anomalous Behaviour of Beryllium

Beryllium, the first member of the Group 2 metals, shows anomalous behaviour as compared to magnesium and rest of the members. Further, it shows diagonal relationship to Aluminium.

Diagonal Relationship Between Beryllium And Aluminium

The ionic radius of Be^{2+} is estimated to be 31 pm. the charge/radius ratio is nearly the same as that of the Al^{3+} ion. Hence beryllium resembles aluminium in some ways. Some of the similarities are.

Some Important Compounds Of Calcium

Important compounds of calcium are calcium oxide, calcium hydroxide, calcium sulphate, calcium carbonate and cement. These are industrially important compounds.

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Calcium Oxide or Quick Lime, CaO

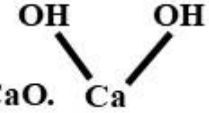
It is prepared on a commercial scale by heating limestone ($CaCO_3$) at 1070-1270 K.

$$CaCO_3 \overset{heat}{\rightleftharpoons} CaO + CO_2 \quad CaO + SiO_2 \rightarrow CaSiO_3$$

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

Calcium Hydroxide (Slaked lime), $Ca(OH)_2$

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



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

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

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

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

Calcium Carbonate, $CaCO_3$

Calcium carbonate occurs in nature in several forms like limestone, chalk, marble etc. It can be prepared by passing carbon dioxide through slaked lime or by the addition of sodium carbonate to calcium chloride.

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

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

When heated to 1200 K, it decomposes to evolve carbon dioxide.

$$CaCO_3 \xrightarrow{1200K} CaO + CO_2$$

Uses

It is used as a building material in the form of marble and in the manufacture of quick lime.

Specially precipitated $CaCO_3$ is extensively used in the manufacture of high quality paper. It is also used as an antacid, mild abrasive in tooth paste, a constituent of chewing gum, and a filler in cosmetics.

Calcium Sulphate (Plaster of Paris), $Caso_4 \cdot \frac{1}{2} H_2O$

It is a hemihydrate of calcium sulphate. It is obtained when gypsum, $CaSO_4 \cdot 2H_2O$, is heated to 393 K.

$$2(CaSO_4.2H_2O) \rightarrow 2(CaSO_4).H_2O + 3H_2O$$

Above 393 K, no water of crystallisation is left and anhydrous calcium sulphate, CaSO4 is formed. This is known as 'dead burnt plaster'.

Uses

The largest use of Plaster of Paris is in the building industry as well as plasters. It is used for immoblising the affected part of organ where there is a bone fracture or sprain. It is also employed in dentistry, in ornamental work and for making casts of statues and busts.

Few Important Points

(i) Magnesium Peroxide (MgO_2) and Calcium Peroxide (CaO_2) are obtained by passing H_2O_2 in a suspension of $Mg(OH)_2$ and $Ca(OH)_2$.

(ii) MgO_2 is used as an antiseptic in tooth paste and as a bleaching agent.

(iii) Preparation of NaOH : Caustication of Na_2CO_3 (Gossage's method)

$$Na_2CO_3 + Ca(OH)_2 \rightarrow 2NaOH + CaCO_3\downarrow$$

(suspension)

Since the K_{sp} $(CaCO_3) < K_{sp}$ $(Ca(OH)_2)$, the reaction shifts towards right.

(vii) On heating $MgCl_2 \cdot 6H_2O$ undergoes hydrolysis as follows