S. Block elements.

- Elements in which the last electron enters the S-orbital are called S-block elements. Since S-subshell has only one orbital which can accommodate only two electrons therefore there are only two group of S-block elements Group I or IA (Alkali metals) and group 2 or IIA (Alkali earth metals).
Electronic Configuration of group 1 may be represented as [Nobel gas] ns^1 and group 2 has [Nobel gas] ns^2
Both alkali and alkaline are highly reactive and hence do not occur in the free state but are widely distributed in nature in the combined state.

- Anomalous behaviour of first element of a group —
The first element of a group differs considerably from the rest of the same group due to.
1. Small atomic and ionic radius.
2. High electronegativity and ionization enthalpy.
3. High Polarising Power of its Cation
4. Absence of d-electrons in its valance shell.

So Li differs from Na, K, Rb and Cs.
and Be differs from Mg, Ca, Sr, and Ba.

- Polarization and Polarizing Power — when a cation and a anion approach each other, they attract the electron cloud of each other, but the effect of a small cation on a large anion is more pronounced than that of the large anion on the small cation. The distortion of the electron cloud of an anion by a cation is called Polarization and the capacity to Polarize an anion is known its Polarizing Power. Polarizing Power increases as the charge on the cation increases and the size decreases.

$$\text{Polarizing Power} = \frac{\text{ionic charg.}}{\frac{4}{3}\pi(\text{ionic radius})^2}$$

- Diagonal relationship – Some elements of certain groups in the second period resemble with the certain elements of the next higher group in the third period. This is called diagonal relationship.

ex. ($Li \rightarrow Mg$) ($Be \rightarrow Al$) ($B \rightarrow Si$)

The main reason for the diagonal relationship are.

1. Similarity in electropositive character –
2. Similarity in atomic or ionic radii
3. Similarity in Polarizing Power.

- Trends in Atomic and Physical Properties of Alkali metals :-

1. Large atomic Size – The atomic radius of alkali metals are the largest in their respective Periods. These increases as we go down the group from Li to Cs. This is due to addition of a new energy shell of inner filled shell on the valence s-electron.
2. Large ionic radius – The ionic radius of the cation formed by them are smaller in size then the corresponding atoms. However these are largest in their respective Periods. The ionic radius increases as we move down the group.
3. Low ionization enthalpy –

(i) The first ionization enthalpies of the alkali metals are the lowest as compared to the elements in the other groups. The ionization enthalpies decreases as we move down the group.

(ii) The second ionization enthalpies of alkali metals are very high. Like the first ionization enthalpies, second ionization enthalpies decreases as we move down the group due to increasing size and increasing shielding effect.

4. Hydration enthalpy – The hydration enthalpy of alkali metals decreases with the increase in ionic radius.

Because of the smallest size, Li^+ has the maximum degree of hydration and it is because of this lithium salts are mostly hydrated.

5. +1 oxidation state — The alkali metals exhibit oxidation state of (+1) in their compounds and are strongly electro-positive in character. This character increases as we move down the group.

6. Metallic character — The elements of this group are typical metals and are soft. These can be easily cut with knife. When freshly cut, they are silvery white and on exposure to air they are become dull. The metallic character increase as we move down the group.

7. Melting and Boiling Points — The melting and boiling points of alkali metals are very low and decreases as we move down the group.

8. Nature of Bond formed — All the alkali metals formed ionic compounds. The ionic character increases from Li to Cs as we move down.

9. Density — The density of alkali metals are quite low as compared to other metals. These densities increases as we move down from Li to Cs. But Potassium is lighter than Sodium due to an abnormal increase in atomic size of Potassium.

10. Flame Colouration — All the alkali metals and their salts impart a characteristic colour to the flame.

Li	Na	K	Rb	Cs
crimson	Yellow	Pale violet	Red violet	Blue

The colour actually arises from electronic transitions in short lived species which are formed momently in the flame. The flame is a rich source of electrons. The yellow flame of sodium commonly called Sodium D-line.

11. Photoelectric effect — Alkali metals exhibit Photoelectric effect. It is the phenomenon of ejection or emission of

electrons from the metal surfaces when electromagnetic radiations are made to strike against them.
Due to low ionization enthalpies all alkali metals except Lithium show Photoelectric effect.

chemical Properties.

1 Reactivity and electrode Potential – All the alkali metals are highly reactive metals, this is due to their strong tendency to lose the single valance electron. Alkali metals also acts as strong reducing agent. Their reducing character or reactivity in the gaseous state increases from Li to Cs.

$$Li < Na < K < Rb < Cs.$$

However in aqueous solution, the reducing character of alkali metals follow the sequence: $Na < K < Rb < Cs < Li$.
This may be explained in term of electrode Potential (E°).
Electrode Potential is a measure of the tendency of an element to lose electrons in the aqueous solution. Thus more negative is the electrode Potential, higher is the tendency of the element to lose electrons and hence stronger is the reducing agent.

2. Reactivity towards water – Alkali metals reacts with water liberating H_2 and forming their hydroxides. The reaction becomes more and more violent as we move down the group. Sodium melts on the surface of water and the molten metal moves around vigorously and may sometimes catch fire. Thus, Lithium is the least reactive while the reactivity of other alkali metals towards water and other acidic hydrogen containing compounds increases on moving down the group from Na to Cs.

3. Reactivity towards oxygen – Alkali metals form different types of oxides depending upon the nature of the metal when heated in excess of air.
The reactivity of metals with oxygen increases down the

group. Further the increasing stability of Peroxide or Super oxide, as the size of the metal ion increases, is due to the stabilization of larger anions by larger Cations through higher lattic energies.
It may be noted that Super oxide ion has a three electron bond. It has one unpaired electron which makes it Coloured and Paramagnetic. The normal oxides of alkali metals are however Colourless and diamagnetic.

4. Reactivity towards air and moisture :- All the alkali metals on exposure to atmosphere get Converted into oxides, hydroxides and finally to Carbonates. So alkali metals are Stored in inert hydrocarbon solvents like Petroleum ether and kerosene oil.

5. Reactivity towards hydrogen. — All the alkali metals react with hydrogen to form Colourless Crystalline ionic hydrides.

a. The order of reactivity of the alkali metals towards hydrogen decreases as we move down the group. This is due to the reason that the lattice energies of these hydrides decreases as the size of the metal Cation increases and thus stability of these hydrides decreases from LiH to CsH.

b. All the alkali metals hydrides are ionic Solids with high melting Point.

c. All the hydrides behave as Strong reducing agents and their reducing nature increases down the group.

d. Since these hydrides Contain the hydride ion (H^-). So they liberate hydrogen at anode on electrolysis.

e. All these hydrides react with Proton donors such as water, alcohols, gaseous ammonia and alkynes liberating H_2 gas.

6. Reactivity towards halogens :- Alkali metals react vigorously with halogens to form ionic metal halides. The reactivity of

alkali metals towards a particuler halogen increases as we move down the group from Li to Cs. due to a decrease in the ionization enthalpy or electropositive character of the metal. On the other hand, reactivity of halogens towards a particuler alkali metal decreases from F_2 to I_2

$$F_2 > Cl_2 > Br_2 > I_2$$

7. Solubility in liquid ammonia.— All alkali metals dissolve in liquid ammonia (solubility may be as high as 5 M) giving highly conducting deep blue solution.

Some important properties of blue solution.

a. Dilute solution of alkali metals in liquid ammonia are dark blue but as the concentration increases above 3M, the colour changes to copper-bronze and the solution acquire metalic lustre due to formation of metal ion clusters.

b. The blue colour solution are paramagnetic due to the presence of large number of unpaired of electrons but bronze solution are diamagnetic due to formation of electron clusters.

c. The solutions of alkali metals in liquid ammonia are good conductors of electricity due to the presence of ammoniated cations and ammoniated electrons. However the conductivity decreases as the concentration increases.

d. These solutions are stronger reducing agent then hydrogen and hence will react with water to liberate hydrogen.

e. In presence of impurities or catalyst, the blue coloured solution decomeseto form metal amides with the liberation of H_2.

— General characteristics of the compounds of alkali metals.

1 Oxides and Hydroxides :- All the alkali metals, their oxides peroxides and superoxides readily dissolve in water to produce corresponding hydroxides which are strong alkalies. Peroxides and superoxides also act as oxidising agent since they react with H_2O forming H_2O_2 and O_2 respectively.

basic strength — The hydroxides of alkali metals behave as strong base due to their low ionization enthalpies. As we move down the group, the ionization energy decreases and basic strength increases.

(b) Solubility and Stability — All these hydroxides are highly soluble in water in water and thermally stable except LiOH.

(c) Formation of salts with acids — Alkali metal hydroxides being strongly basic react with all acids forming salts. The salts are colourless ionic solids which are soluble in water.

2. Halides — The alkali metals combine directly with halogens under appropriate conditions forming halides. All these halides are colourless, high melting crystalline solids having high negative enthalpies of formation.

Fluorides > Chloride > Bromide > Iodide.

a. Polarization effects — Comparison of ionic and covalent character of alkali metal halides — The covalent character of any compound, in general, depends upon.

(I) Size of cation — Smaller the cation, greater is its polarizing power and hence larger is the covalent character. The covalent character decreases as the size of the cation increases.

$LiCl > NaCl > KCl > RbCl > CsCl$

(II) Size of anion — Larger the anion, greater is its polarizability

$LiI > LiBr > LiCl > LiF$.

(III) Charge on the ion — Greater the charge on the cation, greater is its polarizing power and hence larger is the covlant character.

(IV) Electronic configuration of the cation — If two cations have the same charge and size, the one with a nobel gas configuration having 18 electrons in the outermost shell has greater polarizing power than a cation with nobel gas configuration having 8 electrons in the outermost shell.

These four factors are commonly reffered to as Fajan's rules.

b. Lattice enthalpies – Lattice enthalpies as the amount of energy required to separate one mole of solid ionic compound into its gaseous ions. Evidently greater the lattice enthalpy higher is the melting point of the alkali metal halide and lower is its solubility in water.

c. Hydration enthalpy – It is the amount of enthalpy released when one mole of gaseous ions combine water to form hydrated ions. Higher the hydration enthalpy of the ions greater is the solubility of the enthalpy in water.

The solubility of most of the alkali metal halides except those of fluorides decreases on decending the group.

Because of the small size and higher electronegativity, lithium hadies except LiF are predominantly covalent and hence are soluble in organic solvents.

Due to high hydration enthalpy of Li^+ ion Lithium halides are soluble in water except LiF which is spairingly soluble due to its high lattic enthalpy.

— Anomalous behaviour of Lithium –

a. Very small size of Lithium atom and ion.

b. Higher polarizing power of Li^+ resulting in increased covalent character of its compounds which is responsible for their solubility in organic solvents.

c. Comparatively high ionization enthalpy and low electropositive character of lithium as compared to other alkali metals.

d. non-availability of d-orbitals in its valance shell.

e. Strong intermetallic bonding.

— Some of the properties in which Lithium differs from the other members of its group.

1. Lithium is harder while other alkali metals are soft.
2. The melting and boiling points of Lithium are high.
3. Lithium forms monoxides with oxygen while other alkali metals form peroxides as well as superoxides.
4. Lithium combines with nitrogen to form nitride while other alkali metal do not.
5. LiOH is a weak base while the hydroxides of other alkali metals are strong base.
6. The hydride of Lithium is more stable than hydrides of other alkali metals.
7. Lithium hydroxide on heating decompose to form Lithim oxide while other do not.
8. Lithium carbonate on heating decompose to give CO_2 while other alkali metal carbonates do not.
9. Lithium bicarbonate exists in solution while other alkali metals form solid bicarbonates.
10. Lithium does not react with ethyne (Acetylene) to form lithium acetylide while other alkali metals react to form the corresponding acetylides.
11. Lithium nitrate on heating decomposes to give Li_2O, NO_2, O_2 while other alkali metal nitrates decomposes to give the corresponding nitrite and oxygen.
12. The oxide, Hydroxide, Carbonate, Phosphate, oxalate, and Flouride of lithium are sparing soluble in water where as the corresponding salts of other alkali metals are soluble in water.
13. Lithium is the lightest alkali metal, so it can not stored in kerosene oil because it floats on the surface so it is kept wrapped in paraffin wax.

— Diagonal relationship of Lithium with magnesium.

Lithium resembles magnesium mainly due to the similarity in sizes of their atoms and ions and have almost equal properties

Some other Properties of alkali metals.

1 Cs is the most electropositive element due to its lowest ionization energy.

2. Lithium Cannot be used in making Photoelectric cell because it has highest ionization energy and Cannot emit electrons.

3. The Compounds of alkali metals are Colourless and diamagnetic. This is because they have noble gas Configuration with no unpaired electron.

4. All alkali metals exists as body centered cubic lattice with a coordination number of 8.

5. Due to small size, Lithium does not form alums.

6. Alkali metals Combine with murcury to form compounds known as Amalgams. This reaction is highly exothermic.

— Difficulties encountered during extraction of alkali metals.

1 Alkali metals are strong reducing agents and hence Can not be extracted by reduction of their oxides or chlorides.

2 Alkali metals being highly electropositive Cannot be displaced from the aqueous solutions of their salts by other metals.

3. Alkali metals Can not be isolated by electrolysis of the aqueous Solution of their salts since hydrogen is liberated at the Cathode instead of the alkali metals because the discharge Potential of alkali metals are much higher than that of the hydrogen. Therefore the only successful method is the electrolysis of their molten salts usually chlorides.

But the melting Points of these chlorides are very high then this difficulty Can be over come to some extent by lowering the melting Points of these chlorides by the addition of suitable salts such as $CaCl_2$, KF etc.

— Lithium :—

It is Prepared by electrolysis of a fused mixture of dry lithium.

Chloride (55%) and Potassium chloride (45%). Potassium Chloride is added to increase the Conductivity of lithium chloride and to lower the fusion temperature.

Uses.

1 It is used in the manufacture of alloys

a. Lithium lead alloy (·05% Li) or white metal which is used for making toughened bearing

b. Lithium Aluminium alloy has the great tensile strength and elasticity. It is used for aircraft construction.

c. Lithium magenisum alloy (14% Li) is extremely tough and Corrosion-resistant which is used for Aerospace Components.

2. It is used for Producing thermonuclear energy required for Propelling rockets and guided missiles.

3. It is used to make both Primary and secondary batteries.

4. It is used as a getter or scavenger since it combines with oxygen and nitrogen.

5. Lithium Carbonate is used in making a special veriety of glass which is very strong and is weather Proff.

— Sodium

It is extracted by the electrolysis of fused sodium chloride by a Process Called Downs Process.

- Difficulties

1 Sodium chloride melts at 1074 K and it is difficult to attain and maintain this high temperature.

2. Sodium boils at about 1156 K and hence at the temperature of electrolysis, the metal liberates will vaporise.

3 Molten sodium forms a metallic fog with fused sodium chloride.

4. The Products of electrolysis, sodium and chlorine, corrode the material of the cell at this high temperature.

- The addition of Calcium chloride and Potassium fluoride.

lower the melting point of sodium chloride to 850-825K.

— Biological importance of Sodium and Potassium — Although Na^+ and K^+ ions have similar chemical properties but their biological functions are quite different. Na^+ ions are found outside the cell in blood plasma. K^+ ions are present inside the cell. these ions help in transmission of nerve signals, in regulating the flow across cell membrane and in the transport of suger and amino acids into the cell.

Since K^+ ions are the most abundant cations within the cell fluid. they activate many enzymes and participate in the oxidation of glucose to produce ATP (Adenosine Triphosphate)

Alkaline earth metals.

Like alkali metals, alkaline earth metals are also highly reactive. and hence donot occur in the free state but are widely distributed in nature in the combined state.

The atoms of all the alkaline earth metals have two s-electron in their outermost shell (ns^2). As these elements have similer valance shell electronic configuration, they show similar physical and chemical properties.

— Trends in atomic and Physical Properties

1 Atomic radii — The atomic and ionic radii of alkaline earth metals are fairly though smaller than the corresponding alkali metals and these increase down the group. this is due to addition of an extra shell of electrons in each succeeding element and the increasing screening effect.

2. Ionization enthalpy — The alkaline earth metals have fairly low ionization enthalpies though greater than those of the corresponding elements of group I and they decrease down the group. because of the smaller size and higher nuclear charge than those of alkali metals.

The second ionization enthalpies of the elements of group I are higher than those of the elements of group 2.

3. Hydration enthalpy — Like alkali metal ions, the hydration enthalpies of alkaline earth metal ions decreases as the size of the metal ions increases down the group.

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

Due to smaller size of alkaline earth metal ions as compared to alkali metal ions, the hydration enthalpies of alkaline earth metal ions are larger than alkali metal ions.

4. Dipositive oxidation state (M^{+2}) — The chemistry of alkaline earth metals is dominated by the dipositive oxidation state just as unipositive oxidation state is the predominant oxidation state of group 1 elements. Alkaline earth metals always form divalent cations.

5. Electropositive or Metallic Character — The alkaline earth metals are highly electropositive and hence metallic and their electropositive or metallic character increases down the group. However they are less electropositive or metallic than the alkali metals. These metals are less soft than alkali metals.

6. Melting and Boiling Points — The alkaline earth metals have higher melting and boiling points as compared to those of alkali metals, However down the group there is no regular trend in their melting and boiling points.

7. Nature of Bond formed — Like alkali metals, alkaline earth metals predominantly form ionic compounds which are, less ionic than corresponding alkali metals compounds. The tendency to form ionic compounds increases down the group.
The first member, Be forms covalent compounds. Mg also show some tendency for covalency. All other elements forms ionic compounds.

8. Density — The alkaline earth metals are denser than the alkali metals, However the densities of alkaline earth metals show

any reguler trend with increasing atomic number. The density of these metals fir st decrease from Be to Ca and than increase from Ca to Ba. The decrease in density from Be to Ca may be due to decrease in the packing of atoms in their solid lattice.

9. Flame Colouration – Alkaline earth metals impart a characteristics Colour to the flame.

Ca.	Sr	Barium	Ra.
Brick Red	Crimson	Apple green.	Crimson.

— Reactivity and Electrode Potential – All the alkaline earth metals are highly reactive elements since they have a strong tendency to lose the two valance electron to form the corresponding dipositive ions. Beryllium (Be) is the least reactive while Ba is the most reactive element. Alkaline earth metals are less reactive than corresponding alkali metals.

1 Reducing character – The alkaline earth metals are weaker reducing than the alkali metals. the reducing character also increases down the group.

2 Reactivity towards water – formation of hydrogen :– The alkaline earth metals rectivity increases as we move down the group. The reaction of alkaline earth metals is less vigorous as compared to alkali metals.

The electrode Potential of Be is least negative amongest all the alkaline earth metals so Be is less electropositive and hence does not react with water or steam even at red hot.

3. Reactivity towards air.

a Formation of oxides – The alkaline earth metals being less electropositive than alkali metals reacts with air or oxygen slowly upon heating, to form oxides. The reactivity towards oxygen increases as we go down the group.

b. Formation of nitrides – All the alkaline earth metals burns in dinitrogen to form ionic nitrides.

c. Formation of Peroxides — Since Cations stablize larger anions. therefore tendency to form Peroxide increases as the size of the metal ion become larger. All Peroxides are white crystalline ionic solids Containing the Peroxide ions.

4. Reactivity towards halides — All the alkaline earth metals Combine with halogens at elevated temperature forming their halides.

5. Reactivity towards hydrogen — All the alkaline earth metals except beryllium Combine with hydrogen directly on heating to form metal hydrides. CaH_2 is Called hydrolith and is used for Production of H_2 by action of water on it.

6. Reactivity towards Carbon — when beryllium oxide is heated with Carbon at 2175–2225°K, a brick Red Coloured Carbide is formed. The rest of the alkaline earth metals form Carbides of the general formula MC_2 when the metal is heated with Carbon In an electric furnace or when their oxides are heated with Carbon.

when CaC_2 is heated in an electric furnace with atmospheric dinitrogen, it forms Calcium cyanamide. The mixture of Calcium cyanamide and Carbon is Called nitrolim and is used as a slow acting nitrogeneous fertilizer as it hydrolysis Slowly over a Period of months evolving NH_3 gas.

7 Reactivity towards Acids — All alkaline earth metals react with acids liberating H_2.

— Solution of liquid Ammonia — All alkaline earth metals dissolve in liquid ammonia. The dilute Solutions are bright blue in Colour due to Solvated electrons but Concentrated solutions are bronze Coloured due to the formation of metal clusters.

Tendency to form Complexes – Be and Mg have the maximum tendency to form Complexes. This is due to their small size and higher charge density. Chlorophyll is an important Complex of magnesium.

– General characteristics of Compounds of the alkaline earth metals

1 oxides and hydroxides – The oxides of alkaline earth metals are obtained either by heating the metals in dioxygen or by thermal decomposition of their carbonates.
These have high melting Points, have very low vapour Pressures, are very good conductors of heat, are chemically inert and act as electrical insulators. These oxides are used as refractory materials for lining furnaces.

– Hydroxides – The hydroxides of Ca, Sr, Ba are obtained by either treating the metal with cold water or by reacting the corresponding oxides with water. The reaction of these oxides with H_2O is also sometimes called as slaking.

Properties of hydroxides –

a Basic character – Due to small size and high ionization enthalpy, $Be(OH)_2$ is amphoteric. The hydroxides of Mg, Ca, Sr and Ba are basic. Their basic strength increases as we move down the group. this is due to their low ionization energy.
These hydroxides are less basic than the corresponding alkali metal hydroxides.

b. Solubility in water – Alkaline earth metals hydroxides are less soluble in water as compared to the alkali metal hydroxides. The solubility in water increases with increases in atomic number down the group. This is due the reason that both lattice energy and hydration energy decrease down the group as the size of the cation increases but lattice energy decreases more rapidly than the hydration energy.

and hence their solubility increases down the group.

2. Halides — The alkaline earth metals combine directly with halogens at appropriate temperatures forming halides.
These halides can also be prepared by the action of halogen acids on metals, metal oxides, hydroxides and carbonates.
Properties of halides.

a. All beryllium halides are essentially covalent and are soluble in organic solvents. they are hydroscopic and fume in air due to hydrolysis. On hydrolysis, they produce acidic solution.

b. The halides of all other alkaline earth metals are ionic. Their ionic character increases as the size of the metal ion increases. As the ionic character increases or the covalent character decreases, their tendency towards hydrolysis decreases.

c. Except $BeCl_2$, all other chlorides of group 2 form hydrates but their tendency to form hydrates decreases.

d. The hydrated chloride, bromide and iodides of Ca, Sr, Ba, can be dehydrated on heating but those of Be and Mg undergo hydrolysis.

e. BeF_2 is highly soluble in water due to the high hydration energy of the small Be^{+2} ion. The other flourides are soluble in water. on decending the group lattice energy decreases more rapidly than the hydration energy, therefore, their solubility increases slightly with the increases in cation size.
The chlorides, bromides and iodides of all other elements Mg Ca, Sr, Ba, are ionic have lower melting point than flouride and are readily soluble in water. The solubility decreases some what with increasing atomic number.

f. Except $BeCl_2$ and $MgCl_2$, the other chlorides of alkaline earth metals impart characteristic colour to flame.

$CaCl_2$.	$SrCl_2$.	$BaCl_2$.
Brick Red	Crimson.	Grassy Green.

Solubility and thermal stability of oxo salts —

The salts Containing one or more atoms of oxygen such as oxides, hydroxides, Carbonates, bicarbonates, nitrites and etc. are Called oxo salts.

a. Sulphates — The sulphates of alkaline earth metals are prepared by the action of sulphuric acid on metals, metal oxides hydroxides and Carbonates.

Properties —

1 The sulphates of alkaline earth metals are all white solids. Be, Mg and Ca sulphates crystallise in the hydrated form but Sulphates of Sr and Ba crystallise without water of crystallization.

(ii) Solubility — The solubility of the sulphates in water decreases down the group.

(iii) Stability — The sulphates of alkaline earth metals decompose on heating giving the oxides and SO_3.

The temperature of decompose increases as the electropositive Character of the metal or the basicity of the metal hydroxide, increases down the group.

uses.

- $BaSO_4$ is both insoluble in H_2O and opaque to x-rays, therefore barium meal is used to obtain a shadow of the stomach on an x-ray film.

2. Carbonates and Bicarbonates. —

(i) Calculated amount of Carbon dioxide is passed through the solution of the alkaline metal hydroxides a white precipitates is obtained it is alkaline earth metal Carbonates.

(ii) Sodium or ammonium Carbonate is added to the solution of the alkaline earth metal salt again a white precipitate is obtain it is alkaline earth metal carbonate.

- Property.

(i) Solubility — The Carbonates of alkaline earth metals are sparingly soluble in water and their solubility decreases down the group from Be to Ba.

(ii) Stability — The Carbonates of all alkaline earth metals decompose on heating to form the corresponding metal oxide and CO_2. thermal stability of these Carbonates increases down the group as the basicity of metal hydroxides increases.

- Bicarbonates. — The bicarbonates of alkaline earth metals are Prepared by Passing CO_2 through a suspension of metal Carbonates in water.

All the bicarbonates of alkaline earth metals are stable only in solution and have not been isolated in the pure state.

3. Nitrates — Alkaline earth metal nitrates are Prepared in solution and Can be crystallized as hydrated salts by the action of HNO_3 on oxides, hydroxides and Carbonates.

The tendency to form hydrates decrease with increasing size and decreasing hydration enthalpy down the group. All nitrates on heating give the Corresponding oxides.

4. Oxalates — The oxalates of Ca, Sr, and sparingly soluble in water but their solubility increases from Ca to Ba. Be oxalate is however, soluble in water.

— Anomalous behaviour of Beryllium. —

Beryllium, the first member of alkaline earth metals, Shows an anomalous behaviour differ from the rest of the members of its family. Due to.

a. Exceptionally small atomic and ionic size.

b. High ionization enthalpy.

Absence of d-orbitals in its valance shell.
Importent Properties.

a. Be is harder than other members of its group
b. Be doesnot react with water even at high temperatures
c. It has higher meltig and boiling Points
d. Be forms Covalent Compounds due to high Charge density and hence greater Polarizing Power.
e. BeO and $Be(OH)_2$ are amphoteric
f. Be_2C reacts with water to Produce methane gas.
g. Be doesnot exhibit Coordination number of more than four since its valance shell has only four orbitals.
h. Be forms fluro complex anion whereas other members of the group do not form fluoro Complex anions.

— Resemblance between Beryllium and Aluminium —
Beryllium and Aluminium have similar Porperties.