

The s-Block Elements

Syllabus

Introduction, General Characteristic, Group-1 (Alkali Metals), Physical Properties, Chemical Reactivity, Group-2 (Alkaline Earth Metals), Reducing Nature

INTRODUCTION

- (a) Elements of IA and IIA group of the periodic table are called s-block elements.
- (b) For these elements outer s-orbital is in the process of filling.
- (c) IA [ns^1] group elements are called **alkali metals** and IIA [ns^2] group elements are called **alkaline earth metals**.

GENERAL CHARACTERISTIC

- (a) They are good conductors of heat and electricity.
- (b) They are malleable and ductile.
- (c) Exhibit group valency of 1 and 2 for IA and IIA groups respectively.
- (d) They are prepared by the electrolysis of their fused salts.
- (e) They are very reactive as their last shell contains 1 or 2 electrons which can be given off easily (low ionization potential).
- (f) They form colourless compounds except chromates, dichromates etc.
- (g) Their cations are diamagnetic.
- (h) They form ionic compounds (except Li and Be).
- (i) Their solutions in liquid ammonia are good conductor of electricity and are good reductant.
- (j) Oxides are basic in nature.

GROUP-1 (ALKALI METALS)

Elements	Symbol	Atomic number
Lithium	Li	3
Sodium	Na	11
Potassium	K	19
Rubidium	Rb	37
Caesium	Cs	55
Francium	Fr	87

PHYSICAL PROPERTIES

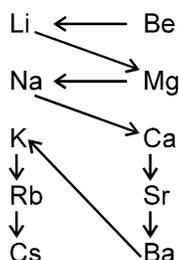
Important physical properties of Alkali metals are given below :

1. Physical State

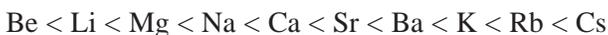
- Soft, silvery white metal having a high and bright lusture when freshly cut.
- They all form body centred lattices.
- Softness increases with increase of atomic number because there is a continuous decrease of metallic bond strength on account of an increase in atomic size.

2. Atomic Size :

These elements are largest in size in the period and the atomic size increase in going downwards in the group.



Order of size :



3. Oxidation State :

These metals exhibit + 1 oxidation state, difference of their second and third ionisation potentials is more than 16 eV. Therefore, their + 1 oxidation state is more stable.

4. Density :

$$\text{Density} = \frac{\text{Atomic weight}}{\text{Volume}} = \frac{M}{V}$$

Atomic weight increase from Li to Cs in the group and volume also increase, but increase in atomic weight is more as compared to volume. Therefore, density

increases from Li to Cs.

Exception :

Density of Na is more than that of K.

Density :



5. Tendency of forming ionic Bond :

One electron is present in the outermost shell of these metals. They form cation by the loss of this electron, i.e., they form ionic bond in their compounds.

6. Standard Electrode Potential or Standard Oxidation Potential :

The measure of the tendency of donating electrons of a metal in water is called its electrode potential. If concentration of metal ions is unity, then it is called standard electrode potential.

Standard electrode potential \propto \propto Atomic size

7. Colourless and Diamagnetic Ions :

The property of an ion as being colourless or coloured, depends on the number of unpaired electrons present in the ion. If unpaired electrons are more in an ion, then these electrons get excited by the atmospheric energy and show colour on coming back to the ground state.

Intensity of the colour \propto Number of unpaired electrons

The ions which have unpaired electrons, show magnetic properties. Whereas, the ions having paired electrons nullify the magnetic fields of each other. Such ions are called diamagnetic ions.

8. Flame Test :

Alkali metals have large size. When they are heated in the flame of Bunsen burner, the electrons present in the valence shell move from lower energy level to higher energy level by absorption of heat from the flame (ns^1 or $ns^2 n^0p$). When they come back to the ground state, they emit the extra energy in the form of visible light to provide colour to the flame. Elements and their respective colours imparted to the flame are given below.

Element	Li	Na	K	Rb	Cs
Colour	Red	Golden	Violet	Red-violet	Blue

Size of Cs is large and one electron is present in its outermost shell. Due to this, electron of outermost shell gets excited by absorption of visible light. Therefore, Cs shows photoelectric effect. This is the reason that it is used in the cells.

10. Solubility in Liquefied Ammonia :

Ionisation potential is low due to large size of these metals, i.e., they readily dissolve in liquefied ammonia to form blue coloured solution, which is a good conductor of electricity and a strong reducing agent.



Ammoniated metal ion

11. Hydration Energy :

Hydration energy decreases on going downwards in the group, due to increase in the size of metal ion.



Lithium gets more hydrated due to high hydration energy of Li^+ and the charge present on it gets protected.

Thus,

$$\text{Hydration energy} \propto \frac{1}{\text{Ionic size}} \propto$$

12. Reactivity :

Due to large size of these metals, the electron of the outermost shell is weakly attracted towards the nucleus.

- (1) Na is very reactive, and is kept in kerosene, so that air does not come directly in contact with sodium.
- (2) Li is stable in air due to small size, Na and K become neutral and Rb and Cs burn spontaneously in air.
- (3) Li hardly reacts with steam, whereas, Cs reacts even with cold water.
- (4) Li forms only one of oxide (Li_2O), because ionisation potential of Li is high.

Superoxide are paramagnetic and coloured due to the presence of unpaired electron. Order of their stability is as follows :

Normal oxide > Peroxide > Superoxide

13. Lustrous Surface :

Lustre is due to mobile electrons in the metallic lattice. Valence electrons generated vibration in the electrical field of the light waves. The vibrating electrons emit electromagnetic energy in the form of light, and thus the surface of these metals starts shining.

14. Tendency of Forming Complex compounds :

A complex compound is a compounds which gives a complex ion on ionisation. For example – $\text{K}_4\text{Fe}(\text{CN})_6$ gives K^+ and a complex ion. $[\text{Fe}(\text{CN})_6]^{-4}$, on ionisation. Complex compounds are formed by the metal which has :

- (1) Very small size of the cation.
- (2) Maximum charge on the cation
- (3) Vacant d orbitals in the cation.

15. Strength of metallic Bonds (Softness)

Metallic bond is weak due to presence of one electron in the valence shell and the BCC structure. The packing efficiency is 68%. Thus, packing of atoms is loose and these elements are soft.

Strength of metallic bond \propto

These metals are soft because one electron is present in their valence shell, which participates in bond formation. Thus, metallic bond is weak.

Atomic size increases in the group from Li to Cs, due to which strength of metallic bond decreases. This is the reason why Li is hard, but Na and K are soft, whereas, Rb and Cs are liquid due to weak metallic bonds. Sheets and wires can be prepared from Li because of its hardness.

16. Melting point and Boiling Point :

Their melting and boiling points are low due to weak metallic bonds. Strength of metallic bond decreases in the group from Li to Cs, due to which hardness from Li to Cs.

$\text{Li} > \text{Na} > \text{K} > \text{Rb} > \text{Cs}$

Thus, melting and boiling points \propto Strength of metallic bond.

17. Reducing Power

The reducing power of a metal is related to its oxidation potential which represents the tendency of element to lose electron and get oxidised.

All alkali metals have low I.E. leading to a high oxidation potential.

(a) Reducing nature (in solution is) :-



(b) In gaseous state



Elements	Electronic	IP (eV)	At. Vol.	At. Radii 'Å'	Ionic radii 'Å'	m.pt. °C	b.pt. °C	Standard electrode potential (volt)	Electronegativity
Li (3)	[He]2s ¹	5.4	13.1	1.23	0.60	181	1336	-3.04	1.0
Na (11)	[Ne] 3s ¹	5.1	23.7	1.54	0.95	98	883	-2.71	0.9
K (19)	[Ar] 4s ¹	4.3	45.3	2.03	1.33	64	760	-2.92	0.8
Rb (37)	[Kr] 5s ¹	4.2	55.9	2.16	1.48	39	688	-2.92	0.8
Cs (55)	[Xe] 6s ¹	3.9	70.0	2.35	1.69	29	690	-2.92	0.7

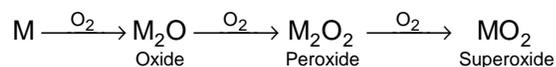
CHEMICAL PROPERTIES

Metal	Li	Na	K	Rb	Cs
Melting Point, °C	180	98	63	38	29
Boiling Point, °C	1320	880	760	700	670

Alkali metals are highly reactive due to low ionisation energy.

1. Reaction with Oxygen

(a) Alkali metal ignites in oxygen and form oxides.



(b) Li forms stable oxide (Li₂O), Na forms peroxide (Na₂O₂) and rest of the metal forms superoxides.

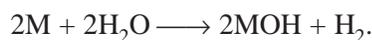
(c) Oxides of alkali metals are basic in nature and basic character increases from Li to Cs as ionic character increases.

(d) Peroxides and superoxides behave as strong oxidising agents. Superoxides on treatment with dil. acids form H₂O₂, O₂ and hydroxide.



2. Reaction with Water

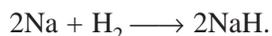
(a) Hydrogen is liberated.



(b) Basicity of hydroxides increases down the group, due to increase in electropositive character.

3. Reaction with Hydrogen

- (a) Alkali metal hydrides are formed when metals are heated with
- H_2
- .



- (b) Metal hydrides are ionic. They are good reducing agents. Reducing power increases down the group.

4. Reactivity with Halogen

- (a) Halides are ionic compounds having negative enthalpies of formation.

- (b) The most negative enthalpy of formation occur with fluorides. The negative value decreases as

Fluorides > Chlorides > Bromides > Iodides.

Thus fluorides are most stable.

- (c) LiF is insoluble in water due to very high lattice energy.

CsI is insoluble in water due to very low hydration energy.

Rest of halides are soluble in H_2O .

5. Metal Hydroxides

- (a) Basic strength of hydroxide increases with the increasing electropositivity of metal.

$CsOH > RbOH > KOH > NaOH > LiOH$.

- (b) Solubility of hydroxides increases with increasing ionic character.

$CsOH > RbOH > KOH > NaOH > LiOH$.

6. Reaction with dilute acids :

Due to alkaline nature, these metals react rapidly with dilute acids and the rate of reaction increases from Li to Cs, because of increase in basic character.

METAL CARBONATES

- (1) All these metals from
- M_2CO_3
- type carbonates.

(Li_2CO_3 , Na_2CO_3 , K_2CO_3 , Rb_2CO_3 , Cs_2CO_3)

- (2) Basic character, ionic character, melting point, boiling point these carbonates increase from carbonates of Li to Cs.

- (3)
- Li_2CO_3
- is least stable out of all these carbonates, because it is covalent and decomposes to
- Li_2O
- and
- CO_2
- at low temperature. Order of their stability is as follows :

$Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3$

- (4) Stability of carbonates of IA group metals > stability of carbonates of IIA group metals.

NITRIDES

Among all alkali metals, only lithium directly combines with nitrogen to form nitride. Other alkali metals combine indirectly with nitrogen, because Li_3N is covalent and as the metallic character increases, the tendency of donating electron and forming ionic bond increases. Due to which strength of metal nitrogen bond decreases.

SULPHATES

Basic character, ionic character, melting point, boiling point, solubility, thermal stability and reactivity increases from Li to Cs.



NITRATES

Their basic character, ionic character, solubility, melting point boiling point and thermal stability increase from Li to Cs.

LiNO_3 decomposes to Li_2O at low temperature, whereas NaNO_3 gets decomposed to NaNO_2 .

HYDRIDES

- (1) Lithium reacts with hydrogen due to its low electropositive character.
- (2) Li is less electropositive and therefore, thermal stability on LiH is high.



They are ionic hydrides and their stability depends of lattice energy.

BICARBONATES

These metals form MHCO_3 type bicarbonates. Basic character, ionic character, melting point, boiling point reactivity and thermal stability of these bicarbonates increase from Li to Cs.

FORMATION OF AMALGAMS

Alkali metals form amalgams with mercury and alloys with other metals.

ANOMALOUS BEHAVIOUR OF LITHIUM

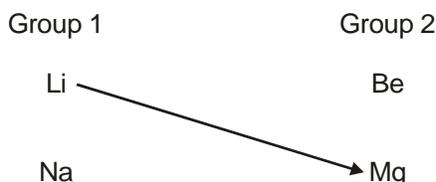
Due to small size of Li, it has high tendency of polarization and due to high density of electrical charge. It shows difference with other alkali metals.

- (1) Li is hard, due to which its melting and boiling points are higher as compared to other metals.
- (2) LiOH is weak base compared to other hydroxides.
- (3) Li forms single type of oxide (Li_2O), whereas, Na (M_2O and M_2O_2), K, Rb and Cs (M_2O , M_2O_2 and MO_2) form more types of oxides.
- (4) LiCl is insoluble in water, whereas, other chlorides are soluble, LiCl gets dissolved in benzene, petrol and ether.
- (5) Due to small size of Li^{+1} , its hydration energy is high.
- (6) Li does not get affected easily by moist air. Therefore, it can be kept open in the air, whereas, other metals form oxides.
- (7) Due to high hydration energy of Li, its conductivity is low.
- (8) Li directly combines with N_2 to form Li_3N whereas, other metal do not form nitrides.
- (9) Phosphate, oxalate, chloride, fluoride, sulphate and carbonate of Li are insoluble in water, whereas the above compounds of other alkali metals are soluble.

- (10) Due to small size of Li, its ionisation potential electronegativity and electron affinity are higher than those of the other alkali metals.
- (11) Due to covalent nature of LiCl, their melting and boiling points are lower than those of the other alkali metal halides.

DIAGONAL RELATIONSHIP : SIMILARITIES WITH MAGNESIUM

Lithium shows resemblance with magnesium, an element of group 2. This resemblance is termed as diagonal relationship



Reasons for the diagonal relationship are the following :

- Electronegativities of Li and Mg are quite comparable (Li = 1.00, Mg = 1.20).
- Atomic radii and ionic radii of Li and Mg are not very much different.

Atomic radii (Å) Li 1.52 Mg 1.60

Ionic radii (Å) Li⁺ 0.76 Mg²⁺ 0.72

- (i) Atomic volumes of Li and Mg are quite similar.

Li 12.97 mL/mole Mg 13.97 mL/mole

- Both have high polarising power (ionic potential)

$$\text{Polarising power } (\Phi) = \frac{\text{Ionic charge}}{(\text{ionic radius})^2}$$

Cations with large ionic potentials have a tendency to polarise the anions and to give partial covalent character to compounds.

- Nitrates of lithium like magnesium decompose to give oxide, whereas all other alkali metal nitrates give nitrite.

GROUP-2. ALKALINE EARTH METALS

Elements	Symbol	Atomic number
Beryllium	Be	4
Magnesium	Mg	12
Calcium	Ca	20
Strontium	Sr	38
Barium	Ba	56
Radium	Ra	88

PHYSICAL PROPERTIES

1. Atomic Size :

Size of these elements are small as compared to alkali metals and it increases on going downwards in the group.

Order of size :



2. Oxidation State :

These metals exhibit + 2 oxidation state, difference of their second and first ionisation potentials is 11 eV. Therefore, their + 2 oxidation state of these metals is more stable.

3. Density :

Atomic weight increase from Be to Ba in a group and volume also increases, but increase in atomic weight is more as compared to volume. Therefore, density increases from Be to Ba.

Exception :

Density of Mg is more as compared to Ca.

Density :



4. Tendency of forming ionic Bond :

There are two electrons in the outermost shell of these metal, which are donated to form ionic compounds. For example, BaCl_2 , CaCl_2 , MgCl_2 , etc. Due to small size of cations of Be and Mg, their compounds have covalent character.

5. Standard Electrode Potential or Standard Oxidation Potential :

Size of these metals is smaller than that of alkali metals. Therefore, their ionisation potential will be higher than those of alkali metals, i.e. they have low tendency of donating electron as compared to alkali metals. Their standard electrode potentials (oxidation potentials) are lower than those of alkali metals and increases in the group with the increases in atomic size from Be to Ba.



6. Colourless and Diamagnetic ions :

These metals form diapositive ions (M^{+2}). These diapositive ions have noble gas configuration and their compounds are colourless, because all the electrons are paired. Their ions are diamagnetic due to the presence of paired electrons. For example, BeCl_2 , CaCl_2 , CaCO_3 , BaSO_4 , etc., are colourless compounds.

7. Flame Test :

Size of Be and Mg is very small and their electrons are strongly bonded to the nucleus. These electrons cannot be excited to higher energy level by the flame of the burner. Thus, Be and Mg do not impart any colour to the flame. Elements and their respective colour imparted to the flame are given below :

Element	Be	Mg	Ca	Sr	Ba
Flame	—	—	Brick red	Blood red	Apple green

8. Photoelectric effect :

Due to small size of these metals as compared to alkali metals, their ionisation potential is high. Thus, electrons can be released only by high energy radiations.

9. Solubility in Liquefied Ammonia :

Due to small size of Be and Mg. Their ionisation potential is high. Therefore, they do not dissolve in liquefied ammonia. Ca, Sr and Ba give ammoniated electron by getting dissolved in liquefied ammonia due to large size, due to which the solution turns blue.



The solution is a good conductor of electricity and a strong reducing agent.

10. Hydration Energy :

Hydration energy of the metal ions (M^{+2}) is higher than that of the elements of IA group, because the size of these cations is small and charge is high. Hydration energy decreases on going downwards in the group, due to increase in the size of cations.

**11. Reactivity :**

Due to small size of these metals as compared to alkali metals, the alkaline earth metals are less reactive than the alkali metals. Their reactivity increases from Be to Ba with increase in the size of the metal.

- (i) Beryllium does not react with hot water, Mg reacts with hot water, whereas Ca, Sr and Ba react even with cold water.
- (ii) All these metals react with oxygen to form MO type oxides ($M = \text{Be, Mg, Ca, Sr and Ba}$), but due to low ionisation potential and high reactivity, Ca, Sr and Ba form peroxides also at low temperature.
- (iii) Be and Mg are less reactive due to their high ionisation potential and they form normal oxides because of breaking of O = O bond.

12. Lustrous Surface :

Lustre is due to mobile electron in the metallic lattice. Valence electrons generated vibration in the electrical field of the light waves. The vibrating electrons emit electromagnetic energy in the form of light, and thus the surface of these metals starts shining.

13. Tendency of Forming Complex compounds :

These metals have weak tendency of forming complex compounds due to large size, low charge density and absence of vacant orbitals. But these metals have higher tendency of forming complex compounds as compared to alkali metals, due to their relatively smaller size. This tendency decreases from Be to Ba.

14. Strength of metallic Bonds (Softness)

There are two electrons in the outermost shell of these metals, which participate in bond formation. Therefore, metallic bond is weak, but a little bit stronger than the elements of IA group. Their atomic size is smaller as compared to elements of IA group. Therefore, these metals form strong metallic bonds as compared to metals of IA group. Thus, these metals are harder than the metals of IA group.

Order of their hardness is $\text{Be} > \text{Mg} > \text{Ca} > \text{Sr} > \text{Ba}$

They have BCC, HCP and FCC structures, i.e., packing efficiency is more than that of the elements of IA group.

15. Melting point and Boiling Point :

Melting and boiling points of these metals are low, but these metals are harder as compared to metals of IA group. Thus, their melting and boiling points are higher as compared to metals of IA group.

Hardness decreases from Be to Ba, due to which melting and boiling points decrease.



Higher melting and boiling points of Ca than Mg are due to the presence of d-orbitals in its outermost shell, which starts from Ca. The d-orbital forms strong metallic point.

Metal	Be	Mg	Ca	Sr	Ba
Melting Point, °C	1280	647	850	770	710
Boiling Point, °C	1500	1135	1490	1360	1535

16. Reducing Nature

- (a) The alkaline earth metals have the tendency to lose electrons and change into bivalent cation.



Hence they act as strong reducing agents.

- (b) The reducing nature increases as the atomic number increases.

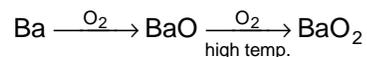
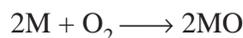
Electronic Configuration	IP (eV)	At. Vol.	At. Radii 'Å'	Ionic radii 'Å'	m.pt. °C	b.pt. °C	Standard electrode potential (volt)	Electronegativity
[He] 2s ²	9.3	5.0	0.9	0.31	1277	2970	-1.80	1.5
[Ne] 3s ²	7.6	14.0	1.36	0.65	650	1100	-2.37	1.2
[Ar] 4s ²	6.1	29.9	1.74	0.99	838	1440	-2.87	1.0
[Kr] 5s ²	5.7	33.7	1.91	1.13	768	1380	-2.89	1.0
[Xe] 6s ²	5.2	39.0	1.98	1.35	714	1640	-2.90	0.9

CHEMICAL PROPERTIES

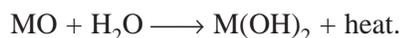
The important chemical properties of alkaline earth metals are discussed below

1. Reaction with Oxygen

- (a) Except Ba and Ra, all other metals when burnt in oxygen form oxide of type MO.



- (b) BeO is amphoteric in nature while other oxides are basic in nature.

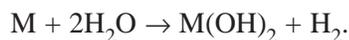


- (c) Basic nature of oxide increases from Be to Ba.

2. Action of Water

- (a) Mg decomposes boiling water.
 (b) Be is not attacked by water at any temperature.

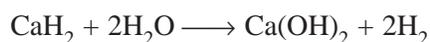
- (c) Ca, Sr, Ba, Ra decomposes cold water with evolution of H₂ gas.



- (d) Be(OH)₂ is amphoteric, but other hydroxides are basic. The basic strength increase from Be to Ba.
- (e) Solubility of hydroxides also increases from Be(OH)₂ to Ba(OH)₂.

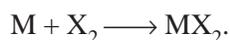
3. Hydride formation

- (a) Except Be, all other metals directly combines with hydrogen to form hydrides of the type MH₂.
- (b) BeH₂ and MgH₂ are covalent hydride while other hydrides are ionic in nature.
- (c) Ionic hydrides are violently decomposed by water evolving hydrogen.



4. Halides

- (a) Metals directly combines with X₂ to form halides



- (b) BeX₂ is covalent in nature, whereas other metal halides are ionic solids.
- (c) These halides are hygroscopic in nature and readily form hydrates *e.g.*, MgCl₂.6H₂O, BaCl₂.2H₂O etc.

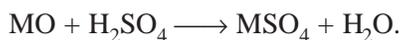
5. Carbonates and Bicarbonates

- (a) Carbonates of metals are insoluble in water, they dissolve in presence of CO₂ only.



- (b) Solubility of carbonates decreases on moving down the group, while stability increases.

6. Sulphates



- (a) Solubility of sulphates decreases on moving down the group.



- (b) Solubility depends upon hydration energy of cation which is maximum in Be²⁺.

7. Complex Formation

Be²⁺ and Mg²⁺ due to their small size show some tendency of complex formation.

Ex. [BeF₃]⁻, [BeF₄]²⁻ and [Be(H₂O)₄]²⁺, chlorophyll (Complex of Mg²⁺)

8. Organometallic Compounds

Be and Mg forms a number of compounds with M – C bonds.

e.g., Grignard reagent.

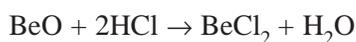
BEHAVIOUR OF BERYLLIUM DIFFERENT FROM OTHER ALKALINE EARTH METALS

Beryllium exhibits different behaviour due to small size as compared to other elements of its group.

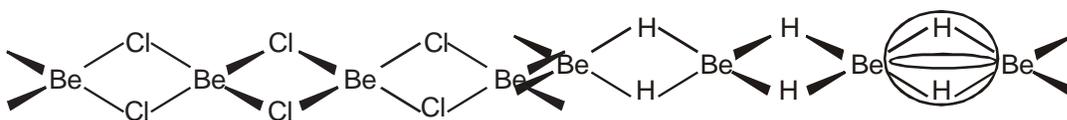
- (1) Ionisation potential and electronegativity of Be are higher than those of other metals.
- (2) BeCl_2 is insoluble in water, due to its covalent nature, but soluble in organic solvents. Other chlorides (CaCl_2 , SrCl_2 and BaCl_2) get dissolved in water.
- (3) BeCl_2 gets hydrolysed due to its covalent nature. Chlorides of Ca, Sr and Ba get dissociated.
- (4) BeO and $\text{Be}(\text{OH})_2$ are amphoteric in nature. Therefore they react with acids as well as bases. Other oxides react only with acids due to their alkaline nature



Sodium berylate



- (5) Beryllium forms single type of oxide (MO), Ca Sr and Ba form peroxides also.
- (6) Hydration energy of Be is higher than that of other metals.
- (7) Beryllium does not give flame test, Ca, Sr and Ba impart characteristic colours to the flame.
- (8) Due to small size, Be forms complex compounds, Ca, Sr and Ba are insoluble in water.
- (9) Hydrides and halides of Be get polymerizes.



Polymerized BeH_2 in which tricentric forces are present.

Polymerized BeCl_2

- (10) Beryllium does not react with water even at high temperatures. Others (Ca, Sr and Ba) react even with cold water.

DIAGONAL RELATIONSHIP SIMILARITIES BETWEEN BERYLLIUM AND ALUMINIUM

Beryllium shows some similarities in properties with aluminium, the second typical element of group 13 of the next higher period.

	Group 2	Group 13
Second period	Be	B
Third period	Mg	Al

↘

The following are the causes of diagonal relationship :

- The polarising power of Be^{2+} and Al^{3+} ions is high. Consequently the compounds develop covalent nature.
- The standard oxidation potential values of both the elements are very close to each other.
- The electronegativity values of both the elements are the same.

SODIUM (Na)

Extraction :

Castner's process : In this process fused sodium hydroxide is electrolysed at 320°C



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

At anode (Nickel) : $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 \uparrow + 4\text{e}^-$

Down's process

Fused NaCl mixed with $\text{CaCl}_2 + \text{NaF}$ act as electrolyte at 600°C .

[$\text{CaCl}_2 + \text{NaF}$ are mixed to decrease the melting point of sodium chloride (815°C)]



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

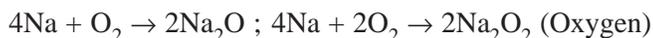
At anode : $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

PHYSICAL PROPERTIES

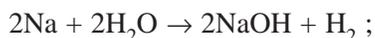
- It is a soft silvery white metal and is kept in kerosene.
- It is lighter than water. It melts at 97.8°C and boils at 883°C .
- It is malleable, ductile and good conductor of heat and electricity.
- Gives blue coloured solution with ammonia (good reductant paramagnetic and good conductor).
- Forms amalgam with mercury.

CHEMICAL PROPERTIES

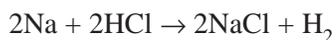
(i) **Heating in air or oxygen :**



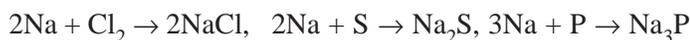
(ii) **Action of water**

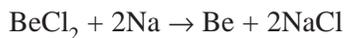
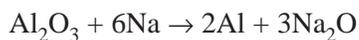
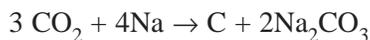


(iii) **Action of acid**



(iv) **Action of non metals**



(v) **Action of NH₃**(vi) **Reducing agent :****USES**(i) Sodium is used in the manufacture of Na₂O₂, NaNH₂, NaCN etc.

(ii) For making sodium amalgam (used as reducing agent).

(iii) For making TEL (used as antiknock in petrol).



(iv) For making sodium vapour lamps which emits yellow light.

(v) For making high temperature thermometer, use is made in the form of Na – K alloy.

(vi) For the extraction of Be, Mg, Si, C.

SODIUM CHLORIDE (NaCl) :**Occurrence:** Sodium chloride or common salt is found in sea water, in salt wells, lakes and in deposits of rock salt.

high heat
(evaporation) →

MANUFACTURE

Hot countries : sea water or lake water

Crude salt

Crude salt contains : [Na₂SO₄, MgCl₂, CaCl₂ as impurities]

Filtered solution Precipitation of NaCl

[Common ion effect

**PHYSICAL PROPERTIES**

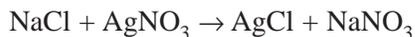
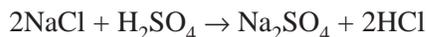
(i) NaCl is white crystalline solid.

(ii) Hygroscopic (due to the presence of MgCl₂ and CaCl₂)

(iii) M.P. 800°C, density 2.17 g/ml.

(iv) Dissolves with the absorption of heat. Solubility does not change appreciably with rise of temperature.

(v) FCC structure.

CHEMICAL PROPERTIES**USES**

Essential constituent of food.

- (i) Used in the manufacture of Na_2CO_3 , HCl , Cl_2 , NaHCO_3 , Na_2SO_4 etc.
- (ii) For the preservation of fish, meat etc.
- (iii) For salting out of soap.
- (iv) For making freezing mixtures.

SODIUM HYDROXIDE (NaOH)**PREPARATION**

Sodium hydroxide is prepared mainly by following processes.

1. Soda lime process, Causticisation process or Gossage process**2. Castner Kellner process - Electrolytic process**

Electrolyte \rightarrow Brine (NaCl Solution in water)

Anode Graphite

Cathode Iron rods and Mercury (acts as intermediate cathode by induction).

At anode :



At cathode



(sodium amalgam)

PROPERTIES

Sodium hydroxide is white hygroscopic solid used in many processes as a solvent. Its important properties are discussed below:

SODIUM CARBONATE OR WASHING SODA [Na₂CO₃·10H₂O]

DIFFERENT FORMS

Na₂CO₃ · H₂O – Crystal carbonate

Na₂CO₃ · 7H₂O – Heptyhydrate

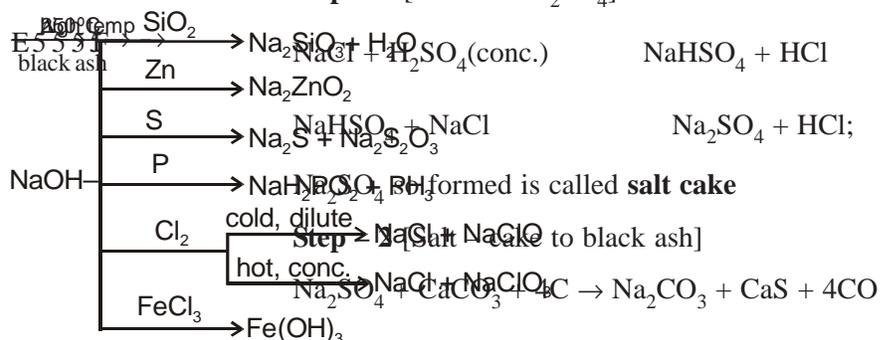
Na₂CO₃ – soda ash (anhydrous form)

MANUFACTURE

(i) Le - Blanc process :

Le-Blanc process involves the following steps :

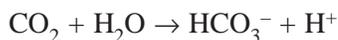
Step – 1 [NaCl to Na₂SO₄]



The solid residue is called black ash. It contains about 45% sodium carbonate. Sodium carbonate is recovered by treating the black ash with water removing CaS and evaporating the solution.

(ii) Solvay ammonia soda process :

Brine (NaCl) solution is saturated with ammonia and carbon dioxide to get the precipitate of least soluble NaHCO₃.



[Na⁺] x [HCO₃⁻] (ionic product) > solubility product of NaHCO₃

NaHCO₃ is heated to obtain sodium carbonate.



NH_4Cl produced (shown above) is heated with lime to obtain ammonia from saturating the brine solution.



CO_2 produced in the process [$\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3$] and also obtained by heating lime stone is used for saturating the brine solution.



CaO is used to produce ammonia by the reaction with NH_4Cl

PHYSICAL PROPERTIES

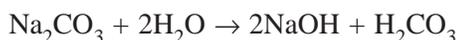
- (i) $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ is white crystalline solid.
- (ii) Exhibit efflorescence of exposure to dry air [i.e. releases water molecules to atmosphere]



- (iii) Anhydrous Na_2CO_3 is stable to heat even to redness.

CHEMICAL PROPERTIES

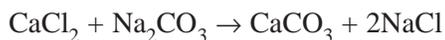
(i) Hydrolysis



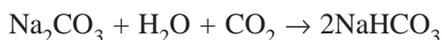
(ii) With acid



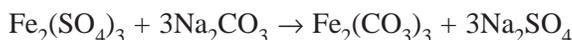
(iii) Formation of normal and basic salts



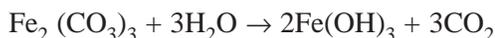
(iv) Formation of NaHCO_3



(v) Formation of carbonates



Carbonate so formed in the process immediately hydrolyses to hydroxide

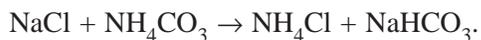


USES

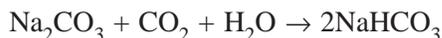
- (i) Used in laundries and softening water.
- (ii) Used in the manufacture of glass, sodium silicate, borax, caustic soda etc.
- (iii) Used as a laboratory reagent and fusion mixture ($\text{Na}_2\text{CO}_3 + \text{K}_2\text{CO}_3$)
- (iv) Used in textile and petroleum refining.
- (v) For preparing metal carbonates.

SODIUM BICARBONATE (BAKING SODA) NaHCO_3

Manufacture : It is manufactured by solvay's process as described in Na_2CO_3



Laboratory method



PHYSICAL PROPERTIES

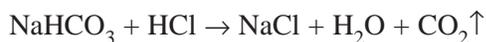
- (i) It is white crystalline solid sparingly soluble in water.
- (ii) Aqueous solution is alkaline. The solution does not give red colour with phenolphthalein, yellow colour is obtained with methyl orange.

CHEMICAL PROPERTIES

Action of heat



With acid



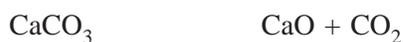
USES

- (i) Used as medicine (Sodabcarb) to neutralize acidity.
- (ii) Used in making effervescent drinks.
- (iii) Used in fire extinguishers.
- (iv) Used in making baking soda (mixture of NaHCO_3 and potassium acid tartarate)

CALCIUM OXIDE, QUICK LIME (CaO)

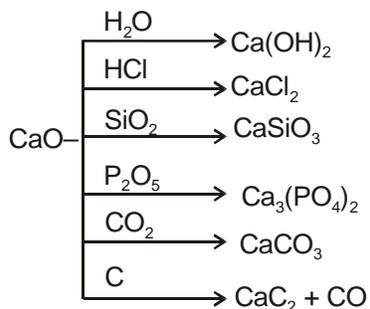
PREPARATION

Calcium oxide is mostly prepared by decomposition of calcium carbonate.



PROPERTIES

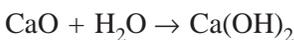
Calcium oxide is white amorphous powder. The important reactions of CaO are given below:



CALCIUM HYDROXIDE, SLAKED LIME, (Ca(OH)₂)

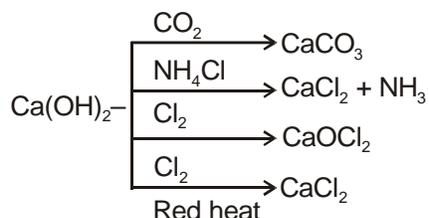
PREPARATION

Calcium hydroxide is mostly prepared from quick lime



PROPERTIES

Calcium hydroxide is a white amorphous solid. The important properties of calcium hydroxide are given below



CALCIUM CARBONATE, LIME STONE OR MARBLE (CaCO₃)

PREPARATION

Calcium carbonate is prepared by



PROPERTIES

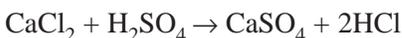
Calcium carbonate is a white powder. It is insoluble in water



Calcium Sulphate (CaSO₄) :

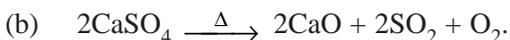
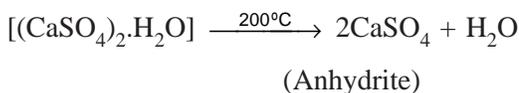
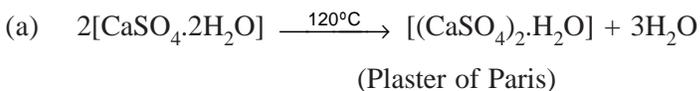
PREPARATION

Calcium sulphate is obtained by adding calcium chloride in sulphuric acid or sodium sulphate



PROPERTIES

Calcium sulphate is a white crystalline solid. It is sparingly soluble in water. It is found in nature as anhydrite (CaSO₄) and gypsum (CaSO₄·2H₂O). Gypsum when heated converts into CaO.



Cement : Composition-CaO-50-60%; MgO-2-3%, SiO₂-20-25%; Fe₂O₃-1-2%; Al₂O₃-5-10%; SO₃ 1-2%.

In general silica and alumina should be present in the ratio of 2.5 to 4.0. Similarly the ratio of lime to the total mixture of SiO₂, Al₂O₃ and Fe₂O₃ should be 2 : 1.

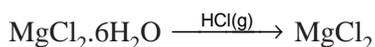
MAGNESIUM CHLORIDE, MgCl₂·6H₂O

PREPARATION

1. By heating and cooling the carnallite when whole of KCl is deposited while MgCl₂·6H₂O remains in the fused state.
2. By dissolving Mg, MgO, Mg(OH)₂ or MgCO₃ in dil. HCl acid.

PROPERTIES

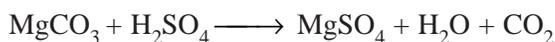
1. Highly deliquescent solid
2. The molecule of hexahydrate can be removed by heating



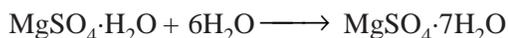
MAGNESIUM SULPHATE, EPSOM SALT, MgSO₄·7H₂O

PREPARATION

1. By heating dolomite with dil. H₂SO₄

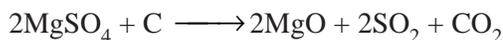


2. By boiling Kieserite in water and cooling the resulting solution



PROPERTIES

1. Upon heating to 150°C, it changes to monohydrate which on further heating changes to anhydrous state at 200°C.
2. It is reduced by lamp black.



3. Forms double salts with alkalimetal sulphates

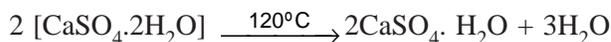
BIOLOGICAL ROLE OF SODIUM, POTASSIUM, MAGNESIUM AND CALCIUM

1. Sodium and potassium are the most common cations in biological fluids.
2. Na⁺ is an extracellular ion and K⁺ an intracellular ion.
3. Na⁺ takes parts in maintenance of electric potential across the cell membrane, also help in retaining water in blood.
4. The main pigment for absorption of light in plants is chlorophyll which contains magnesium.
5. Calcium ions are also involved in enzymatic system and play roles in regulating muscle contraction, transmitting nervous pulses and acts as an agent of blood coagulation.

PLASTER OF PARIS [$\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$]

PREPARATION

It is obtained when gypsum is heated to a temperature of 120°



PROPERTIES

- (i) It is a white crystalline solid.
- (ii) It has a property of setting to hard mass when paste with water is allowed to stand.
- (iii) It gives dead burnt plaster (calcium sulphate) when heated to a temperature of 200°C



USES

- (i) Used in making black board chalks.
- (ii) Used in making statues, toys etc.
- (iii) Used in surgery for setting broken bones.

COMPOSITE FORM OF IA AND IIA GROUP

Properties	Order
Atomic size	$\text{Be} < \text{Li} < \text{Mg} < \text{Na} < \text{Ca} < \text{Sr} < \text{Ba} < \text{K} < \text{Rb} < \text{Cs}$
Ionisation potential	$\text{Cs} < \text{Ba} < \text{Rb} < \text{Sr} < \text{K} < \text{Ca} < \text{Na} < \text{Mg} < \text{Li} < \text{Be}$
Electronegativity	$\text{Cs} < \text{Ba} < \text{Rb} < \text{Sr} < \text{K} < \text{Ca} < \text{Na} < \text{Mg} < \text{Li} < \text{Be}$
Density	$\text{Li} < \text{Be} < \text{K} < \text{Ca} < \text{Na} < \text{Mg} < \text{Rb} < \text{Sr} < \text{Cs} < \text{Ba}$
Order of hydration of ions	$\text{Cs}^+ < \text{Ba}^{+2} < \text{Rb}^+ < \text{Sr}^{+2} < \text{K}^+ < \text{Ca}^{+2} < \text{Na}^+ < \text{Mg}^{+2} < \text{Li}^+ < \text{Be}^{+2}$
Thermal stability of hydrides	Group – IA $\text{LiH} > \text{NaH} > \text{KH} > \text{RbH} > \text{CsH}$
Order of basic character of hydroxides	Group – IIA $\text{BeH}_2 > \text{MgH}_2 > \text{CaH}_2 > \text{SrH}_2 > \text{BaH}_2$ Group – IA $\text{CsOH} > \text{RbOH} > \text{KOH} > \text{NaOH} > \text{LiOH}$ Group – IIA $\text{Ba}(\text{OH})_2 > \text{Sr}(\text{OH})_2 > \text{Ca}(\text{OH})_2 > \text{Mg}(\text{OH})_2 > \text{Be}(\text{OH})_2$
Thermal stability of metal carbonates	Group – IA $\text{Rb}_2\text{CO}_3 > \text{K}_2\text{CO}_3 > \text{Na}_2\text{CO}_3 > \text{Li}_2\text{CO}_3$ Group – IIA $\text{BaCO}_3 > \text{SrCO}_3 > \text{CaCO}_3 > \text{MgCO}_3 > \text{BeCO}_3$
Solubility of hydroxides in water	Group – IA $\text{CsOH} > \text{RbOH} > \text{KOH} > \text{NaOH} > \text{LiOH}$ Group – IIA $\text{Ba}(\text{OH})_2 > \text{Sr}(\text{OH})_2 > \text{Ca}(\text{OH})_2 > \text{Mg}(\text{OH})_2$

SOME IMPORTANT COMPOUNDS

Na_2CO_3 – Soda ash

$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ – Washing soda

$\text{Na}_2\text{CO}_3 \cdot \text{NaHCO}_3 \cdot \text{Na}_2\text{SO}_4$ – Sazzi clay

$[\text{Na}_2\text{O}_2 + \text{dil HCl}]$ – Soda bleach

Na_2SO_4 – Salt cake.

SOME IMPORTANT POINTS (s - BLOCK ELEMENTS)

1. KNO_3 is used in gun powder.
2. KO_2 is used to obtain oxygen on high mountains.
3. BaCO_3 is a medicine for killing rats.
4. CaOCl_2 is a germicide and a bleaching agent.
5. BaSO_4 is used in X-ray examination of intestines.
6. $\text{MgCO}_3 \cdot \text{Mg}(\text{OH})_2 \cdot 3\text{H}_2\text{O}$ is called **magnesium alba**, which is used for the filling up to cavities in the teeth.
7. CaCN_2 is a fertilizer.
8. $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ is called borax, which is a mineral of Na.
9. Complex compounds of Mg and Fe are chlorophyll and haemoglobin, respectively.
10. Na and K is kept in kerosene because they are reactive.
11. Li is hard and cannot be cut by knife. Na, K and Rb are soft solids and Cs is a liquid.
12. Order of stability of alkali metals is $\text{M}_2\text{O} > \text{M}_2\text{O}_2 > \text{MO}_2$.
13. Hydration energy of Li^{+1} is maximum. Therefore, its conductivity is low.
14. Strength of metallic bond \propto \propto Melting and boiling points \propto Hardness
15. Ionic character \propto Size of cation \propto Boiling and melting points of compounds.
16. Superoxides are paramagnetic and coloured due to the presence of unpaired electrons.
17. Formulae and constituents of portland cement are : $\text{Ca}_3\text{SiO}_5 + \text{CaSiO}_4 + \text{Ca}_3\text{Al}_2\text{O}_6$
18. Formula of superphosphate of lime is : $\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} + \text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
19. Stability of hydrides is : $\text{LiH} > \text{NaH} > \text{KH} > \text{RbH} > \text{CsH}$
20. H_2SO_4 , Na_2CO_3 and NaHCO_3 are used as fire extinguisher.
21. LiCO_3 is used in mental disorders.
22. LiNO_3 and NaNO_3 melt on absorbing moisture of air, KNO_3 does not have this character. Therefore, it is used in gun powder.
23. CaH_2 is called Hydrolith. It is used in the transportation of H_2 , because it is light and reacts with ice to release hydrogen.

1
Atomic size

24. Mixture of water + sand + slaked lime is called **mortar**.
25. $\text{Mg}(\text{ClO}_4)_2$ (Magnesium perchlorate) is called **anhydrone**.
26. A mixture of MgO and MgCl_2 , (Mg_2OCl_2) is called **sorrel cement**.
27. Calcium nitrate is called **Norwegian salt petre**.
28. 15% K_2CO_3 is found in wood ash.
29. KClO_3 (Potassium chlorate) is used in an explosive substance, in ammunitions and fireworks.
- ◆ Cs is the most electropositive element in the periodic table.
 - ◆ Rb and Cs because of low ionisation energies show photoelectric effect.
 - ◆ The ease of formation of ionic hydrides increases from lithium to caesium.
 - ◆ The alkali metals are soluble in ammonia.
The solution is (i) paramagnetic (ii) highly conducting (iii) good reductant (iv) blue in colour
 - ◆ All alkali metals possess **body centrad cubic** lattice in the solid state.
 - ◆ The most abundant alkali metal in earth's crust is sodium.
 - ◆ The alkali metal hydrides are strong reducing agents and reducing property increases with decrease in stability.
[Li –H bond is only 25% ionic]. The stability decreases from LiH to CsH.
 - ◆ Under the trade name **oxone**, Na_2O_2 is used as a bleaching agent.
Calcium is the most abundant element of the IIA group.
 - ◆ The solubility of some halides in water follows the order
(i) $\text{BeF}_2 > \text{MgF}_2 > \text{CaF}_2 > \text{SrF}_2 > \text{BaF}_2$
(ii) $\text{BeCl}_2 < \text{MgCl}_2 < \text{CaCl}_2 < \text{SrCl}_2 < \text{BaCl}_2$
(iii) $\text{CaF}_2 < \text{CaCl}_2 < \text{CaBr}_2 < \text{CaI}_2$
 - ◆ Hydroxides of alkaline earth metals are only slightly soluble and the solubility increases from $\text{Mg}(\text{OH})_2$ to $\text{Ba}(\text{OH})_2$.
 - ◆ Both BeO and $\text{Be}(\text{OH})_2$ are amphoteric. $\text{Be}(\text{OH})_2$ is a feeble base and its aqueous suspension is known as milk of magnesia.
 - ◆ Setting of plaster of Paris is due to hydration as well as transition.
 - ◆ Quick lime reacts vigorously with water of form strong base $\text{Ca}(\text{OH})_2$ which is much less soluble than $\text{Ba}(\text{OH})_2$.
 $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$, $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$
 $\text{Ca}(\text{OH})_2$ is known as slaked lime.

IMPORTANT COMPOUNDS :	
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Washing soda
NaHCO_3	Baking soda
$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$	Plaster of paris
CaOCl_2	Bleaching powder
$\text{Ca}_3(\text{PO}_4)_2$	Rock phosphate
$\text{MgCO}_3 \cdot \text{Mg}(\text{OH})_2 \cdot 3\text{H}_2\text{O}$	Compound used for Dental work
$\text{Na}_2\text{CO}_3 + \text{K}_2\text{CO}_3$	Fusion mixture
$\text{Na}(\text{NH}_4)\text{HPO}_4$	Microcosmic salt
$\text{MgSO}_4 \cdot \text{H}_2\text{O}$	Magnesite
$3\text{BeO} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$	Beryl
SrCO_3	Strontianite
SrSO_4	Celestine
$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$	Alum
$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 4\text{Al}(\text{OH})_3$	Alunite or Alumstone
$\text{KCl} \cdot \text{MgSO}_4 \cdot \text{MgCl}_2 \cdot 3\text{H}_2\text{O}$	Canite
$\text{LiAl}(\text{SiO}_3)_2$	Spodumene
$\text{K}_2\text{O} \cdot 3\text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2 \cdot 2\text{H}_2\text{O}$	Mica
BaSO_4	Baryte or Heavy spar
$\text{Ba}(\text{OH})_2$	

POTASSIUM (K)	
Name	Formula
KCl	Sylvin
$\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$	Carnalite
$\text{KAISi}_3\text{O}_8 (\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2)$	Felspar
KNO_3	Salt petre (Nitre)
K_2CO_3	Pearl ash
$\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$	Schonite
$\text{K}_2\text{SO}_4 \cdot 2\text{MgSO}_4$	Longebenite
K_2SO_4	Salt cake
$\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 2\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	Polyhalite

MAGNESIUM (Mg)	
Formula	Name
MgCO_3	Magnesite
$\text{MgCO}_3 \cdot \text{CaCO}_3$	Dolomite
$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	Epsom salt
$\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$	Carnalite
$\text{MgSO}_4 \cdot \text{H}_2\text{O}$	Kieserite
$\text{CaMg}_3(\text{SiO}_3)_4$	Asbestos
$\text{Mg}(\text{Si}_2\text{O}_5) \cdot \text{Mg}(\text{OH})_2$	Talc

ANAND ACADEMY BHILAI 9827464311, 8109132711

SODIUM (Na) :

CALCIUM (Ca)

CBSE SECTION

SECTION - 1

A. VERY - SHORT - ANSWER QUESTION

1. Name the elements belonging to the Group 1 of the periodic table.
2. Write the general electronic configuration of alkali metals.
3. Which of the group 1 elements has the highest and the lowest value of ionisation enthalpy?
4. Why do the group 1 elements show +1 oxidation state in their compounds?
6. Why does lithium behave differently than the other elements in this group?
7. What happens when sodium is burnt in excess of oxygen.
8. Which alkali metal gives golden yellow colour to the flame?
9. How can you obtain sodium from fused sodium chloride.
10. Why are alkali metals normally kept under kerosene?
11. Write two physical properties of sodium.
12. Write one major difference in the chemical behaviour of sodium and potassium.
13. What is Nafion membrane? Where is it used?
14. Why is the solution of an alkali metal in ammonia blue?
15. Which of the sodium carbonate and sodium bicarbonate is more soluble in water?
16. Mention one typical use of each: (a) sodium carbonate, (b) sodium bicarbonate.
17. How is crude sodium chloride purified?
18. Why is sodium carbonate solution basic in nature?
19. How is potassium chloride obtained from carnallite?
20. To which group of the periodic table the metals calcium and magnesium belong?
21. Write the electronic configurations of Mg and Ca.
22. Why do Ca and Mg act as good reducing agents?
23. Salts of which metal give brick-red colour to the flame?
24. Why are the oxides of alkali metals more basic than those of alkaline earth metals?
25. Name the minerals of magnesium which are commonly used for the manufacture of magnesium.
26. Write the chemical equations to describe the extraction of magnesium from sea water.
27. Why is magnesium not isolated by chemical reduction method?
28. Why is anhydrous calcium chloride used as a dehydrating agent?
29. Which of the bases, NaOH, and $\text{Mg}(\text{OH})_2$ is more basic?
30. Ionic hydrides of only alkali and alkaline earth metals are known? Give reason.
31. How is anhydrous magnesium chloride obtained from $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$?
32. What is Grignard's reagent?
33. What is the composition of magnesia cement?
34. Write down the formula of epsom salt.

35. How does limewater differ from milk of lime?
36. What are the various forms in which CaCO_3 occur in nature?
37. Write the chemical equation describing the effect of heat on gypsum.
38. What is the formula of plaster of paris. Mention one of its use.
39. Name the raw material required for the manufacture of cement.
40. Describe the average composition of cement.
41. What changes occur during the setting of cement?
42. Why is small amount of gypsum added before cement clinker is finely ground?

B. SHORT - ANSWER - QUESTIONS

1. Which of the alkali metals has
 - (a) highest reducing power
 - (b) smallest ionic size
 - (c) highest melting point
 - (d) least electropositive character?
2. Why does lithium resemble magnesium?
3. Sodium amalgam is used as a reducing agent. Give reason
4. Sodium metal cannot be obtained by the electrolysis of aqueous solution of sodium chloride. Give reason.
5. Can we store sodium metal under water? Give reason.
6. The solutions of alkali metals in ammonia are blue coloured and conducting. Explain
7. State what happens when
 - (i) Sodium burns in air
 - (ii) Sodium is exposed to air
 - (iii) Sodium reacts with water
8. What happens when sodium peroxide is reacted with
 - (a) Water at 0°C
 - (b) water at 40°C
 - (c) chromium (III) hydroxide
9. Name the raw materials used in the manufacture of sodium carbonate by Solvay process. Name the by-product of this process. Why is potassium carbonate not prepared by this method?
10. What are the products formed when a concentrated aqueous solution of sodium chloride is electrolysed using, (a) platinum cathode, (b) mercury cathode?
11. Why are the hydration enthalpies of alkaline earth metal ions much larger than those for the alkali metal ions?
12. BeO is amphoteric, MgO is weakly basic, CaO is more basic, while SrO and BaO are extremely basic. What do you conclude from this?
13. How do the following properties vary in the group?
 - (a) Solubility of hydroxides of alkaline earth metals.
 - (b) Solubility of sulphates of alkaline earth metals.
14. Draw the structures of
 - (a) Solid beryllium chloride
 - (b) Solid aluminium chloride
15. The complex-forming tendency of magnesium is reflected in the formation of chlorophylls. Give reason.
16. Arrange the following in the increasing order of the ionic character : KCl , MgCl_2 , CaCl_2 , BeCl_2 .
17. What happens when
 - (i) Magnesium is burnt in air.

- (ii) Magnesium is heated in steam,
(iii) A piece of burning magnesium continues to burn in sulphur dioxide?
18. How will you detect the presence of Mg^{2+} ions in any solution?
19. Explain the following phenomenon by means of balanced equations.
- (i) When exhaling is made through a tube passing into a solution of limewater, solution becomes turbid.
(ii) The turbidity of solution in (i) eventually disappears when continued exhaling is made through it.
(iii) When the solution obtained in (ii) is heated, turbidity reappears.
20. Write the balanced chemical equations showing the preparation of the following compounds starting from calcium carbonate.
- (i) CaO (ii) $Ca(OH)_2$ (iii) $CaSO_4$ (iv) $Ca(HCO_3)_2$

C. LONG - ANSWER - QUESTIONS

- Alkali metals have low ionisation enthalpies. Why is it so?
- How can sodium be extracted from (i) fused sodium hydroxide, (ii) fused sodium chloride. Give the theory of the process employed.
- Why cannot sodium and potassium be prepared by the electrolysis of aqueous solutions of its salts?
- Describe the chemistry of Solvay process for the manufacture of sodium carbonate.
- Draw a flow-sheet diagram of Solvay process for the manufacture of sodium carbonate.
- How is sodium hydroxide manufactured by Castner-Kellner process?
- Mention some important uses of the following compounds:
(a) Sodium bicarbonate (b) Sodium hydroxide (c) Sodium peroxide.
- Why are calcium and magnesium metals not isolated by chemical reduction methods?
- Name the important ores of magnesium. How is magnesium extracted from fused magnesium chloride?
- How is anhydrous magnesium chloride obtained from the sea water?
- Name two important ores of calcium. How is calcium obtained from calcium chloride?
- How is lime (quicklime) obtained from limestone? What precautions should one take while converting limestone into lime (quicklime)?
- How does slaked lime react with,
(a) chloride (b) carbon dioxide (c) hydrochloric acid?
- How is epsom salt obtained? Mention two of its uses.
- Explain why during the preparation of plaster of paris, the temperature is maintained around $120^\circ C$?
- Write the sequence of changes that take place during setting of the plaster of paris.
- What is the effect of heat on the following compounds? Write equations for the reactions.
(1) Calcium carbonate (2) Magnesium chloride hexahydrate,
(3) Gypsum, (4) Magnesium sulphate heptahydrate.
- Name the raw materials used for manufacturing cement. How is cement manufactured?