

# P-BLOCK ELEMENTS

## GENERAL CHARACTERISTICS :

1. General Electronic Configuration :  $ns^2np^{1-6}$ . The inner core of the electronic configuration influence the physical properties (such as atomic & ionic radii, IE,  $-\Delta_{eg}H$ , electronegativity etc) and the chemical properties.
2. **Oxidation States** : The oxidation states of elements are closely related to the number of electrons in the valence shell of their elements. The highest oxidation state = (gp no. -10).

INERT PAIR EFFECT : It is the tendency of having less availability of 'ns' electron in bonding on account of their penetration in (n-1)d sub shell. On descending the group, the inert pair effect increases & a lower oxidation state which is two less than the highest oxidation state becomes more stable in group 13 to 16.

Group	13	14	15	16	17	18
General E.C.	$ns^2np^1$	$ns^2np^2$	$ns^2np^3$	$ns^2np^4$	$ns^2np^5$	$ns^2np^6$
First member of the group	B	C	N	O	F	Ne
Group Ox. state	+3	+4	+5	+6	+7	+8
Other Ox. state	+1	+2, -4	+3, -3	+4, +2, -2	+5, +3, +1 -1	+6, +4, +2

3. **Metallic & non-metallic character** : The p block contains metallic & non-metallic elements. The non metals & metalloids exist only in the p block. The non metallic character decreases down the group. The most metallic elements are located in the lower left portion while the most nonmetallic character is present at the top right portion of the periodic table. The common metalloids are B, As, Ge, Sb, Te, Po, At etc.
4. **IE & Electro negativity** of nonmetals is higher & thus they readily form anions.
5. The nonmetal oxides are acidic or neutral whereas metal oxides are basic in nature, ( $Al_2O_3$  is amphoteric)
6. The first member of each group differ in many respect from the other elements of the same group due to
  - (i) Small size of the atom & its ions
  - (ii) high electro negativity
  - (iii) absence of d-orbital in their valence shell

These factors have significant effect on the chemistry of the first element :

- (a) **Covalence up to four** : First member has one 2s & three 2p orbitals (& no vacant d orbital) available for taking part in chemical reaction. Thus, they may have maximum covalence of four Eg, Boron Forms  $BF_4$  whereas Al form  $AlF_6$  Carbon can form tetrahalides ( $CCl_4$ ) whereas other members form hexahalides ( $SF_6$ ,  $SiCl_6^{2-}$  etc), Nitrogen forms  $NF_3$  while phosphorus forms  $PF_5$ .
- (b) **Reactivity** : Due to availability of d-orbitals of the elements of 3rd period, they can show higher covalence (beyond 4)
- (c) **Tendency to form multiple bonds** : Because of the combined effect of the smaller size & unavailability of d orbitals, the 1st member of each group show greater tendency to form  $p\pi - p\pi$  multiple bonds either with itself ( $C = C, C \equiv C, N \equiv N, O = O$ ) or with other members of the second period of the elements ( $C = O, C \equiv N, N = O$ ). Heavier elements do not form  $p\pi - p\pi$  bonds because their atomic orbitals are too large and diffuse to have effective overlapping.

## GENERAL CHARACTERISTICS OF GROUP 13 & 14

Characteristics	Group 13 : BORON FAMILY	Group 14 : CARBON FAMILY
Occurrence	Orthoboric acid ( $H_3BO_3$ ), Borax ( $Na_2B_4O_7 \cdot 10H_2O$ ) Bauxite ( $Al_2O_3 \cdot 2H_2O$ ) Cryolite ( $Na_3AlF_6$ )	Coal, Graphite & diamond Cassiterite ( $SnO_2$ ) Galena ( $PbS$ )
Electronic Configuration	B : $[He]2s^2 2p^1$ (has noble gas core) Al : $[Ne]3s^2 3p^1$ (has noble gas core) Ga : $[Ar]3d^{10} 4s^2 4p^1$ (has noble gas + 10d) In : $[Kr]4d^{10} 5s^2 5p^1$ (has noble gas + 10d core) Tl : $[Xe] 4f^{14} 5d^{10} 6s^2 6p^1$ (has noble gas + 14f + 10d core)	C : $(He) 2s^2 2p^2$ Si : $[Ne] 3s^2 3p^2$ (has noble gas core) Ge : $[Ar]4s^2 4p^2$ (has noble gas + 10d core) Sn : $[Kr]5s^2 5p^2$ (has noble gas + 10d core) Pb : $[Xe] 6s^2 6p^2$ (has noble gas + 14f + 10d core)
Atomic Radii	On moving down the group both atomic & ionic radii are expected to increase due to addition of new shells. But atomic radii of Al > Ga because of low screening & poor penetration of d orbitals in Ga.	The increase in covalent radii from C to Si is considerable but from Si to Pb is small increase due to presence of completely filled d & f orbitals
IE	It decreases sharply from B to Al but IE of Ga & Al as well as In & Tl is almost the same due to low screening of d & f electrons.	The 1st IE is higher. The decrease is small from Si to Ge to Sn & slightly increase from Sn to Pb (poor shielding & increasing size.)
Electronegativity	It is decreasing from B to Al but then increases due to discrepancies in atomic size.	More electro-ve than gp 13
M.P & B.P	Decreases on moving down the group but the decrease in M.P is not regular (decreases from B to Ga & increases from Ga to Tl) as in B.P. B has unusually high MP due to very strong crystalline lattice & Ga has unusually low MP exists as discrete diatomic molecules.	They form covalent bonds, thus stronger binding forces. Thus higher MP & BP than gp 13. BP & MP decreases down the group.
Density	Increase down the group because increase in mass outweighs the effect of increase in size.	Same as group 13 elements
Electropositivity (metallic character)	Less electropositive than group 2 elements. On moving down the group, metallic character increases because IE decreases (eg B is a non metal.)	Metallic character increases because IE decreases Ge is metalloid.
Oxidation state	Maximum oxidation state +3 ( $2e^-$ in s & $1e^-$ in p subshell). From B to Al, the sum of the three IE decreases, thus Al forms $Al^{3+}$ ions. Except boron, other elements also show +1 oxidation state which is more stable as we move down the group (inert pair effect). The relative stability of +1 : $Al < Ga < In < Tl$ .  +3 state is highly oxidizing in nature & +1 more ionic	Maximum oxidation state +4 ( $2e^-$ in s & $2e^-$ in p subshell) Ge, Sn & Pb also show +2 which is more stable on moving down (inert pair effect) $Ge < Sn < Pb$ .

**CHEMICAL REACTION :**

Group 13	Group 14
<p>1. Boron is non metallic &amp; the crystalline form is very hard, unreactive &amp; non conducting (reacts only at very high temp)</p> <p>2. Al is reactive but gets covered with a protective oxide layer.</p> <p>3. Ga, In &amp; Tl are relatively soft &amp; reactive metals which readily dissolve in acids.</p> <p>4. The sum of the three IE for Boron is much greater than that compensated by the lattice (solid) or hydration (aqueous) energies. Thus ionic compounds of boron are not easily formed. Other atoms of gp 13 form ionic compounds in aqueous solution. eg <math>\text{AlCl}_3</math> is covalent which hydrolyse to ionic <math>[\text{Al}(\text{H}_2\text{O})_6]^{3+}</math>.</p> <p>5. Trichlorides on hydrolysis in water form tetrahedral <math>[\text{M}(\text{OH})_4]</math> species.</p> $\text{sp}^2\text{BCl}_3 \xrightarrow{\text{Hy}} [\text{B}(\text{OH})_4]^- \text{sp}^{3-}$ <p>6. In +3 state, the number of electrons around the central atom in a molecule of these elements are only six (<math>\text{BF}_3</math>) and are <b>ELECTRON DEFICIENT</b> molecules and behave as <b>LEWIS ACIDS</b>. <math>\text{AlCl}_3</math> stabilizes by forming dimer. But <math>\text{BCl}_3</math> exists as molecule. (B atom is so small can accommodate 4 large sized atoms around it)</p> <div style="display: flex; justify-content: space-around; align-items: flex-start;"> <div style="text-align: center;"> <p>Planar</p> </div> <div style="text-align: center;"> <p>Tetrahedral</p> </div> </div>	<p>1. Carbon is a typical non metal and exhibit allotropy.</p> <p>2. Silicon also behaves as non metal but in certain physical properties behave as semi metal.</p> <p>3. Ge is a metalloid while Sn &amp; Pb are typical metal.</p> <p>4. In tetravalent state the number of electrons around the central atom in a molecule (eg <math>\text{CCl}_4</math>, <math>\text{SiCl}_4</math>) is eight, thus they are <b>ELECTRON PRECISE</b> molecules.</p> <p>5. Except carbon other members have d orbitals &amp; can exceed its covalence more than 4. Due to this their halides can undergo hydrolysis &amp; have tendency to form complexes by accepting electron pairs from donor species. Eg : <math>[\text{SiF}_6]^{2-}</math>, <math>[\text{GeCl}_6]^{2-}</math>, <math>[\text{Sn}(\text{OH})_6]^{2-}</math> where the central atom has <math>\text{sp}^3\text{d}^2</math> hybridisation. central atom can accept <math>e^-</math> pair from <math>\text{H}_2\text{O}</math> due to availability of d-orbitals.</p> <p>6. Carbon has a tendency to link with one another through covalent bonds to form chains and rings. This property is called <b>catenation</b>. Down the group catenation decreases as size increases and electronegativity decreases. Pb does not show catenation.</p> <p>7. Carbon shows allotropic forms due to catenation and formation of <math>p\pi - p\pi</math> bonds.</p>

Property	Group 13	Group 14
Reactivity towards $\text{O}_2$ (Nature of oxides)	$4\text{B} + 3\text{O}_2 \xrightarrow{\Delta} 2\text{B}_2\text{O}_3$ & $4\text{Al} + 3\text{O}_2 \xrightarrow{\Delta} 2\text{Al}_2\text{O}_3$ $2\text{B} + \text{N}_2 \xrightarrow{\Delta} 2\text{BN}$ & $2\text{Al} + \text{N}_2 \xrightarrow{\Delta} 2\text{AlN}$ Boron forms acidic. Al & Ga form amphoteric and In & Tl form basic oxides.	All elements when heated with oxygen form oxides ( $\text{MO}$ & $\text{MO}_2$ ) Oxides of higher oxidation state are more acidic than those in lower ox. state $\text{CO}_2$ , $\text{SiO}_2$ , $\text{GeO}_2$ are acidic whereas $\text{SnO}_2$ & $\text{PbO}_2$ are amphoteric CO is neutral, $\text{GeO}$ is acidic while $\text{SnO}$ and $\text{PbO}$ are amphoteric
Reactivity towards $\text{H}_2\text{O}$	B & Al do not react but Al amalgam reacts with cold water. $\therefore$ oxide coating not formed $2\text{Al} + 6\text{H}_2\text{O} \rightarrow 2\text{Al}(\text{OH})_3 + 3\text{H}_2$	C, Si & Ge are not affected. Sn reacts with steam $\text{Sn} + 2\text{H}_2\text{O} \xrightarrow{\Delta} \text{SnO}_2 + 2\text{H}_2$ Lead is unaffected due to formation of protective layer

Reactivity towards Halogen	Form trihalides $2E + 3X_2 \rightarrow 2EX_3$ $BBr_3 > BCl_3 > BF_3$ (strength of Lewis acids) reason:- Due to back bonding $BF_3$ is no longer remain electron deficient but in case of $BCl_3$ and $BBr_3$ , there is no back bonding between B & Cl or B & Br as size of Cl and Br is big enough, so that effective back bonding is not possible	Form halides $MX_2$ & $MX_4$ Tetra halides are mostly covalent. Except carbon, all other members react directly with halogens under suitable conditions. $PbI_4$ does not exist because Pb-I bond initially formed does not release enough energy to unpair $6s^2$ electrons and excite one of them to higher orbital to have 4 unpaired $6s^2$ electrons and excite one of them to higher orbital to have 4 unpaired $e^-$ . Stability of dihalides increases down the group $PbX_2$ more stable than $PbX_4$ , but $GeX_4$ more stable than $GeX_2$ due to inert pair effect.
Reactivity towards Acids & Alkalis	B does not react & Al shows amphoteric nature. $2Al + 6HCl \rightarrow 2Al^{3+} + 6Cl^-$ $2Al + 2NaOH + 6H_2O \rightarrow 2Na[Al(OH)_4] + 3H_2$ Conc $HNO_3$ Makes Al passive by forming a protective layer.	C, Si & Ge are unaffected by acids. Sn dissolve in dil $HNO_3$ to form nitrate. Pb dissolve slowly in dil HCl forming $PbCl_2$ but dissolve readily in dil $HNO_3$ $4Sn + 10HNO_3 \rightarrow 4Sn(NO_3)_2 + NH_4NO_3 + 3H_2O$ $Pb + HNO_3 \rightarrow Pb(NO_3)_2 + NO_2 + H_2O$

### SOME IMPORTANT COMPOUNDS :

1. BOROX :  $Na_2B_4O_7 \cdot 10H_2O$  : a) It contains tetranuclear units  $[B_4O_5(OH)_4]^{2-}$  thus the formula is  $Na_2[B_4O_5(OH)_4] \cdot 8H_2O$ .

(b) Effect of heat :  $Na_2B_4O_7 \cdot 10H_2O \xrightarrow{\Delta} Na_2B_4O_7 + 10H_2O \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$   
(Sodium metaborate) (Boric anhydride)  
*Borax Bead*

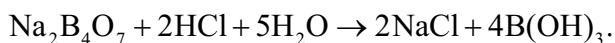
(c)  $NaBO_2 + B_2O_3$  form transparent glassy bead which when touched with some colored cations as  $Ni^{2+}$ ,  $Co^{2+}$ ,  $Cu^{2+}$  etc & heated again gives characteristic colored beads. This is known as BOROX BEAD TEST.

e.g.  $CoO + B_2O_3 \xrightarrow{\Delta} Co(BO_2)_2$   
Blue

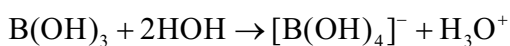
$NiO + B_2O_3 \xrightarrow{\Delta} Ni(BO_2)_2$   
Brown

2. ORTHOBORIC ACID ( $H_3BO_3$ ) :

(a) It is prepared by acidifying an aqueous solution of borax.



(b) It is a weak monobasic acid which is not a protonic acid but acts as lewis acid by accepting electrons from a hydroxyl ion.



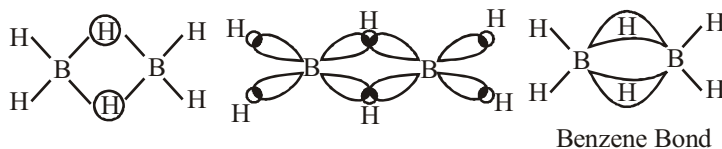
(c) Effect of heat :  $H_3BO_3 \xrightarrow{\Delta} HBO_2 \xrightarrow{\Delta} B_2O_3$   
Meta boric acid Boric oxide

### 3. DIBORANE (B<sub>2</sub>H<sub>6</sub>) :

- (a) Reaction of NH<sub>3</sub> with diborane gives initially B<sub>2</sub>H<sub>6</sub>. 2NH<sub>3</sub> formulated as [BH<sub>2</sub>(NH<sub>3</sub>)<sub>2</sub>]<sup>+</sup>[BH<sub>4</sub>]<sup>-</sup>. This gives borazine B<sub>2</sub>N<sub>2</sub>H<sub>6</sub> (inorganic benzene) on heating.



- (b) Structure of diborane : There are two types of hydrogen : (i) four terminal hydrogen are bonded to B by normal covalent bonds sharing one e<sup>-</sup> each by B & H atom (ii) The bridged hydrogen are bonded by three centre e<sup>-</sup> pair bonds, which involve one e<sup>-</sup> pair only (2e<sup>-</sup>) but binds three atoms B, H & B known as three centre electron pair bond.



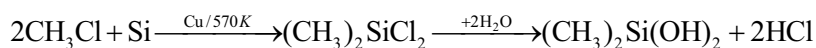
- (c) Boron also forms a series of hydridoborates; the most important one is the tetrahedral [BH<sub>4</sub>]<sup>-</sup> ion. Lithium and sodium tetrahydridoborates, also known as borohydrides, are prepared by the reaction of metal hydrides with B<sub>2</sub>H<sub>6</sub> in diethyl ether. MBH<sub>4</sub> (metal borohydrides) used as reducing agent in organic synthesis

### 4. SILICON DIOXIDE (SiO<sub>2</sub>)

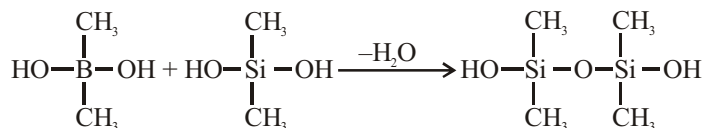
- (a) Also called silica, forms 95% of the earth's crust. The crystalline forms are quartz, cristobalite & tridymite. They are interconvertible at suitable temperature.
- (b) It is a covalent, three dimensional network solid in which each Si atom is covalently bonded in a tetrahedral manner to four Oxygen atoms. Each Oxygen atom in turn is covalently bonded in another Si atom. The entire crystal is a giant molecule in which eight member rings are formed. Silica in normal form is almost non-reactive because of very high Si-O bond enthalpy, however, it is attacked by HF and NaOH.

### 5. SILICONES

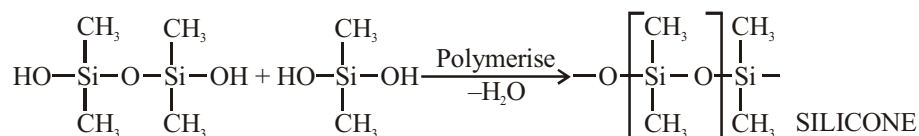
- (a) They are polymeric compounds containing Si-O linkage with general formula R<sub>2</sub>SiO as repeating unit.
- (b) The starting material for the manufacture of silicones is alkyl substituted chlorosilanes.



The polymers are obtained by the hydrolysis of the chloro derivative. When two molecules of dialkyl silinol combine, dimer is obtained with the elimination of the molecule of water.



Since an active -OH group is left on each end of the chain, polymerization reaction continues & the length of the chain increases.



- (c) Silicones are surrounded by non polar alkyl groups which are water repelling. They are used as sealant, greases, electrical insulators & for water proofing of the fabric. They are also used in surgical & cosmetic plants

### 6. SILICATES : (SiO<sub>4</sub><sup>4-</sup>) and ZEOLITE

Do Uses of SiO<sub>4</sub><sup>4-</sup> (silicates), SiO<sub>2</sub> (Silicon dioxide) **from N.C.E.R.T. BOOK.**

Allotropes of Carbon: Allotrope forms or allotropes are the different forms of the same elements having different physical properties but similar chemical properties.

1. Crystalline form: Diamond and graphite are two crystalline forms of carbon having well defined structure
2. Amorphous form: Coal, wood charcoal, animal charcoal, lamp black, coke, etc. are many amorphous forms of carbon such as

Structure of different allotropes of carbon:

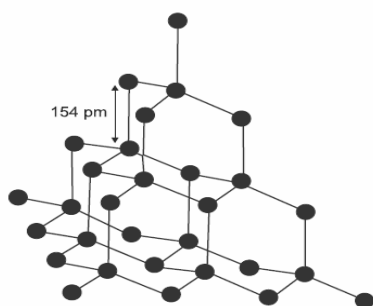
(a) Diamond: Diamond has a network structure consisting of a very large number of carbon atoms bonded to each other. Each carbon atom is  $sp^3$  hybridized and is bonded to four other carbon atoms by single covalent bonds. There is a three – dimensional network of strong covalent bonds in diamond. This makes diamond an extremely hard crystal with very high melting point (about 3843 K).

Since all the valence electrons of carbon are strongly held in carbon – carbon bonds, diamond is a poor conductor of electricity.

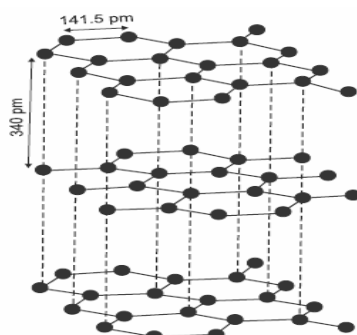
(b) Graphite: In graphite, each carbon atom undergoes  $sp^2$  hybridisation and is covalently bonded to three other carbon atoms by single bonds. The fourth electron on each carbon atom forms  $\pi$  bonds. In this way, graphite consists of hexagonal rings in two dimensions. The C–C covalent distance in rings is 141.5 pm indicating strong bonding. These arrays of ring form layers. The layers are separated by a distance of 340 pm. The large distance between these layers indicates that only weak van der Waals' forces hold these layers together. The weak van der Waals' forces which hold these layers together are responsible for soft nature of graphite.

Because of the weak Van der Waals' forces between layers, one layer can move over the other layer and this account for the slippery nature of graphite. Therefore, graphite is used as lubricant. The electrons forming  $\pi$  bonds in the rings of graphite are delocalized and are relatively free to move under the influence of electric field. Therefore, graphite is a good conductor of electricity.

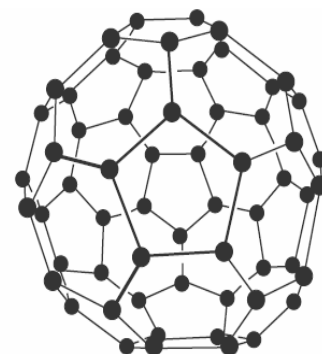
(c) Fullerene: It contains 20 six- membered rings and 12 five membered rings. All the carbon atoms are equal and they undergo  $sp^2$  hybridisation. Each carbon atom forms three sigma bonds with other three carbon atoms. The remaining electron at each carbon is delocalised in molecular orbitals, which in turn give aromatic character to molecule. This ball shaped molecule has 60 vertices and each one is occupied by one carbon atom and it also contains both single and double bonds with C–C distances of 143.5 pm and 138.3 pm respectively. Spherical fullerenes are also called bucky balls in short.



Structure of diamond



Structure of graphite

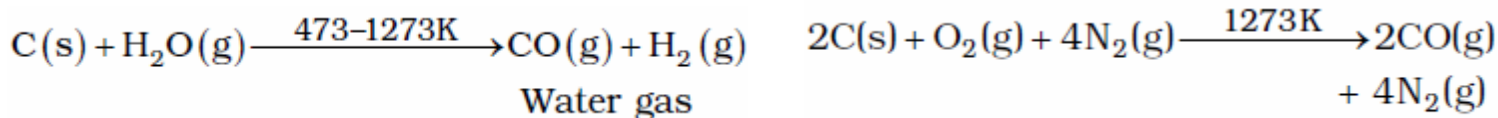
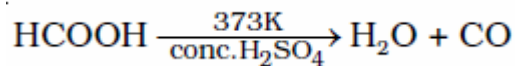
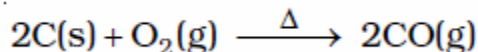


Structure of fullerene

## Some important compounds of Carbon

(i) CARBON MONOXIDE (CO)

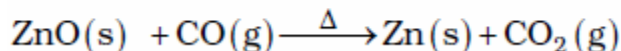
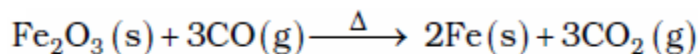
(a) Preparation of carbon monoxide (CO):



(b) Properties of carbon monoxide (CO):

(i) It is colourless, odourless and almost water insoluble gas

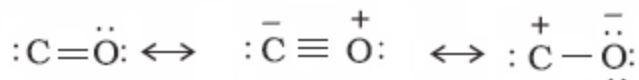
(ii) It is a powerful reducing agent and reduces almost all metal oxides other than those of the alkali and alkaline earth metals, aluminium and a few transition metals



(iii) CO is highly poisonous because it has ability to form a complex with haemoglobin, which is about 300 times more stable than the oxygen-haemoglobin complex. This prevents haemoglobin in the red blood corpuscles from carrying oxygen round the body and ultimately resulting in death.

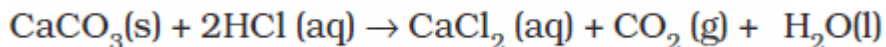
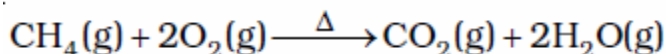
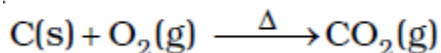
(c) Structure of carbon monoxide (CO): In CO molecule, there are one  $\sigma$  and two  $\pi$  bonds between carbon and oxygen:  $\text{C} \equiv \text{O}:$ . Because of the presence of a lone pair on carbon, CO molecule acts as a donor and reacts with certain metals when heated to form metal carbonyls.

Carbon monoxide is regarded as resonance hybrid of the following structures:



(ii) Carbon dioxide (CO<sub>2</sub>)

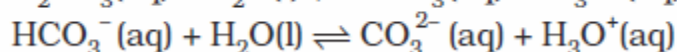
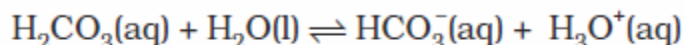
(a) Preparation of carbon dioxide (CO<sub>2</sub>):



(b) Properties of carbon dioxide (CO<sub>2</sub>):

(i) It is a colourless and odourless gas

(ii) With water, it forms carbonic acid, H<sub>2</sub>CO<sub>3</sub> which is a weak dibasic acid and dissociates in two steps:

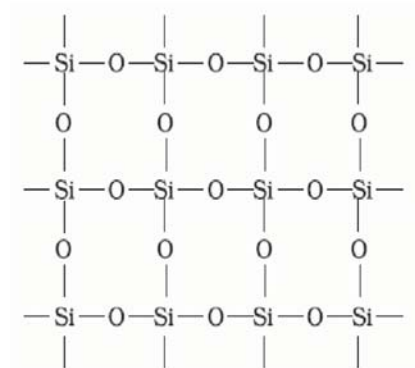
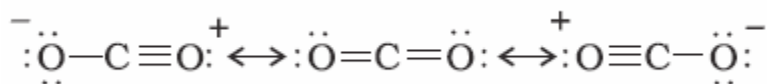


H<sub>2</sub>CO<sub>3</sub>/HCO<sub>3</sub><sup>-</sup> buffer system helps to maintain pH of blood between 7.26 to 7.42

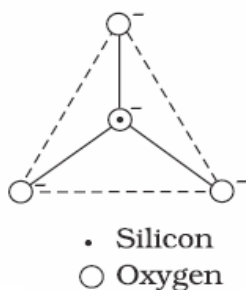
(iii) Being heavy and non-supporter of combustion carbon dioxide is used as fire extinguisher

(c) Structures of carbon dioxide:

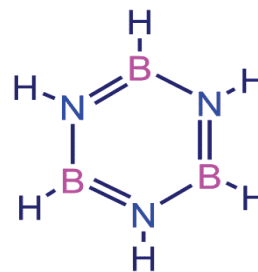
In  $\text{CO}_2$  molecule carbon atom undergoes  $sp$  hybridisation. Two  $sp$  hybridised orbitals of carbon atom overlap with two  $p$  orbitals of oxygen atoms to make two sigma bonds while other two electrons of carbon atom are involved in  $p\pi - p\pi$  bonding with oxygen atom. This results in its linear shape [with both C-O bonds of equal length (115 pm)] with no dipole moment. The resonance structures are shown below:



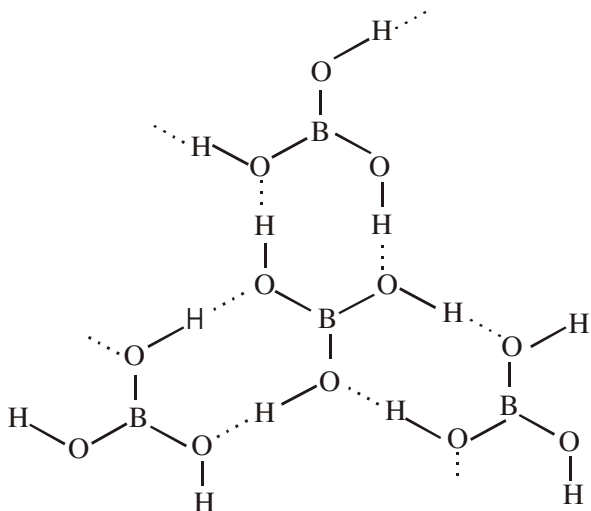
Structure of Silicon dioxide



Representation of  $\text{SiO}_4^{4-}$  unit



borazine



Structure of boric acid; the dotted lines represent hydrogen bonds

**Zeolites:**

Aluminosilicates are formed when few Si atoms are replaced by Al atoms in three-dimensional network of silicon dioxide. Zeolites are widely used as a catalyst in petrochemical industries for cracking of hydrocarbons and isomerisation, e.g., ZSM-5 (A type of zeolite) used to convert alcohols directly into gasoline.



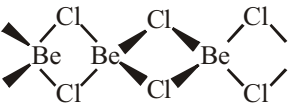
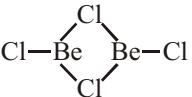
## S BLOCK ELEMENTS

### GENERAL TRENDS IN PROPERTIES :

Properties	Alkali metals (group 1) → Why	Alkali Earth Metals (group 2)
Electronic configuration	General configuration $ns^1$ ( $n = 2 - 7$ ) Li : [He]2s <sup>1</sup> Na : [Ne]2s <sup>1</sup> K. [Ar] 4s <sup>1</sup> Rb : [Kr]5s <sup>1</sup> Cs : [Xe]6s <sup>1</sup> Fr : [Rn]7s <sup>1</sup>	General configuration $ns^2$ Be. [He]2s <sup>2</sup> Mg : Ne]3s <sup>2</sup> Ca : [Ar] 4s <sup>2</sup> Sr : [Kr]5s <sup>2</sup> Ba : [Xe]6s <sup>2</sup> Ra : [Rn]7s <sup>2</sup>
Atomic & ionic radii	They have largest size in their respective period. It increases on moving down the group due to addition of new energy shells. The monovalent ions (M <sup>+</sup> ) are smaller than the parent atom.	They are smaller than alkali metals. It increases on moving down the group due to addition of new energy shells. The divalent ions (M <sup>2+</sup> ) are smaller than the parent atom $M^{2+} < M^+ < M$ (due to decreased effective nuclear charge)
Ionisation enthalpy	The 1 <sup>st</sup> IE is quite low due to large size. The value decreases down the group. After losing one e <sup>-</sup> , it attains stable noble gas configuration.	The 1 <sup>st</sup> IE of alkaline earth metals are higher than that of alkali metals due to smaller size. But the 2 <sup>nd</sup> IE are smaller than corresponding alkali metals I.E. <sub>2</sub> (group 1) is greater than I.E. <sub>2</sub> (group 2).
Hydration Enthalpy	Decreases with increase in ionic size $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+ > Fr^+$ Li <sup>+</sup> has maximum degree of hydration thus its salts exist as hydrated salts e.g. LiCl. 2H <sub>2</sub> O.	Larger than alkali metals due to greater charge density $Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$ Salts are more extensively hydrated eg. MgCl <sub>2</sub> 6H <sub>2</sub> O, CaCl <sub>2</sub> . 6H <sub>2</sub> O.
Density	Low density due to large atomic volume & can be cut with a knife. The density increases down the group from Li to Cs but K is lighter than Na as increase in size is more than mass.	Denser, heavier & harder than alkali metals due to small size & strong metallic bonds. Show regular trend.
State	Silvery white, soft & light metals.	Silvery white, lustrous & relatively harder than alkali metals.
Conductors	Good conductors of heat & electricity.	Good conductors of heat & electricity.
Melting & Boiling points	Low MP & BP due to weak metallic bonds (Presence of only one valence electron) & decreases down the group	Higher than alkali metals due to smaller size. Does not show a regular trend. (2 valence e <sup>-</sup> → metallic bonding strong.)
Flame Coloration (due to excitation & de-excitation of valence electron)	Li : Crimson red, Na : Yellow K : Violet, Rb : Red Violet Cs : Blue due to low I.E. e <sup>-</sup> excited by light of particular $\lambda$ ; while coming back complementary colour is observed by emitting $\lambda$ .	Be & Mg do not impart color as the electrons are strongly bound to get excited by flame (due to small size) Ca : Brick red Sr : Crimson red Ba : Apple green

Reducing Properties	Strong reducing agent as they have low IE (tendency to lose e <sup>-</sup> ) E <sup>o</sup> is less	They are weaker reducing agents than alkali metals because of higher I.E. + 2 (noble gas config. & high lattice enthalpy.)
Oxidation State	+1	+2
Photoelectric effect	Cs shows maximum due to low I.E.	
Basic Strength		

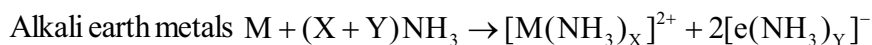
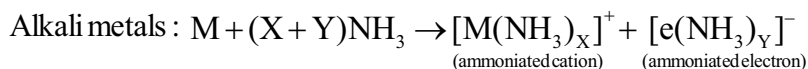
<b>Reaction with Air</b>	<ul style="list-style-type: none"> <li>Tarnish in air due to formation of a layer of oxides which in turn form hydroxides</li> <li>Reactivity with oxygen increases from Li to Cs, (they are stored under kerosene oil)           <math display="block">4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}(\text{Oxide})</math> <math display="block">2\text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}_2(\text{Peroxide})</math> <math display="block">\text{K}(\text{or Rb or Cs}) + \text{O}_2 \rightarrow \text{KO}_2(\text{superoxide})</math>           The stabilities of these oxides are linked with the relative sizes of cations &amp; anions involved &amp; also upon the charges present on them. A smaller cation can stabilize a smaller anion while a larger cation can stabilize larger anion. Thus Li<sup>+</sup> forms oxide (O<sup>2-</sup>) but Na<sup>+</sup>, K<sup>+</sup> being larger form peroxide &amp; superoxide.         </li> <li>Li react directly with nitrogen to form nitrides <math>6\text{Li} + \text{N}_2 \rightarrow 2\text{Li}_3\text{N}</math></li> </ul>	<ul style="list-style-type: none"> <li>Be &amp; Mg are kinetically inert to O<sub>2</sub> due to formation of an oxide layer (BeO, MgO)</li> <li>With the exception Ba &amp; Ra which form peroxides (MO<sub>2</sub>) rest of the elements from normal oxides on heating with excess Oxygen.           <math display="block">2\text{Be} + \text{O}_2 \rightarrow 2\text{BeO}</math> <math display="block">2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}</math> <math display="block">2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}</math> <math display="block">2\text{Ba} + \text{O}_2 \rightarrow \text{BaO}_2</math>   <math display="block">\left[ \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \cdot\cdot \\ \cdot\cdot \end{array} \right]^{2-} \quad \left[ \begin{array}{c} \cdot\cdot \\ \cdot\cdot - \cdot\cdot \\ \cdot\cdot \\ \cdot\cdot \end{array} \right]^{2-} \quad \left[ \begin{array}{c} \cdot\cdot \\ \cdot\cdot \cdots \cdot\cdot \\ \cdot\cdot \\ \cdot\cdot \end{array} \right]^{-}</math>           Oxide (O<sup>2-</sup>) Peroxide (O<sub>2</sub><sup>2-</sup>) Superoxide (O<sub>2</sub><sup>-</sup>)         </li> <li>Superoxide are Paramagnetic due to the presence of unpaired e<sup>-</sup>.</li> </ul>
Reaction towards Acids	<ul style="list-style-type: none"> <li>Readily react with acids to liberate H<sub>2</sub></li> <li><math>2\text{M} + 2\text{HCl} \rightarrow 2\text{MCl} + \text{H}_2 \uparrow</math></li> </ul>	<ul style="list-style-type: none"> <li>Readily react with acids to liberated H<sub>2</sub></li> <li><math>\text{M} + 2\text{HCl} \rightarrow \text{MCl}_2 + \text{H}_2 \uparrow</math></li> </ul>
Nature of Oxides	<ul style="list-style-type: none"> <li>Oxides &amp; peroxides are colorless but superoxide is yellow or orange. The superoxides are paramagnetic and used as oxidizing agents.</li> <li>All oxides are basic in nature &amp; are easily hydrolysed by water to form hydroxides.           <math display="block">\text{M}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{M}^+ + 2\text{OH}^-</math> <math display="block">\text{M}_2\text{O}_2 + \text{H}_2\text{O} \rightarrow 2\text{M}^+ + 2\text{OH}^- + \text{H}_2\text{O}_2</math> <math display="block">2\text{MO}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{M}^+ + 2\text{OH}^- + \text{H}_2\text{O}_2 + \text{O}_2</math> </li> </ul>	<ul style="list-style-type: none"> <li>BeO is a covalent solid, the oxides of the rest of the metals are white crystalline ionic solids</li> <li>Except for BeO which is amphoteric in nature, rest of the oxides are basic &amp; the basic strength increases down the group.           <math display="block">\text{MO} + \text{H}_2\text{O} \rightarrow \text{M}(\text{OH})_2</math> <math display="block">\text{BeO} + 2\text{NaOH} + \text{H}_2\text{O} \rightarrow \text{Na}_2[\text{Be}(\text{OH})_4]</math>           Sodium beryllate         </li> <li><math>\text{BeO} + 2\text{HCl} \rightarrow \text{BeCl}_2 + \text{H}_2\text{O}</math></li> </ul>
Reaction with water	<ul style="list-style-type: none"> <li>React with water to form hydroxides &amp; release H<sub>2</sub>.           <math display="block">2\text{M} + 2\text{H}_2\text{O} \rightarrow 2\text{MOH} + \text{H}_2</math> </li> <li>Reaction of Li with water is less vigorous than Na due to its small size &amp; high hydration energy.</li> <li>Other metals react vigorously.</li> </ul>	<ul style="list-style-type: none"> <li>Be &amp; Mg are kinetically inert due to formation of an oxide layer.</li> <li>Ca, Sr &amp; Ba are readily react with water with increasing vigour even in cold to form hydroxides.</li> </ul>

<b>Nature of Hydroxides</b>	<ul style="list-style-type: none"> <li>• Strongest of all bases.</li> <li>• Dissolve easily in water with evolution of heat due to intense hydration.</li> <li>• They are corrosive to skin &amp; are thus called as Caustic alkalis.</li> </ul>	<ul style="list-style-type: none"> <li>• Less basic and less stable than alkali metals</li> <li>• Due to low IE of metals, M-OH bond ionizes, releasing OH<sup>-</sup>;</li> <li>• Basic strength increases down the group due to decrease in IE.</li> <li>• Be(OH)<sub>2</sub> is amphoteric in nature.</li> </ul> $\text{Be(OH)}_2 + 2\text{OH}^- \rightarrow [\text{Be(OH)}_4]^{2-}$ <p style="text-align: center;">(Beryllate ion)</p> $\text{Be(OH)}_2 + 2\text{HCl} + 2\text{H}_2\text{O} \rightarrow [\text{Be(OH)}_4]\text{Cl}_2$
<b>Reactivity towards halogens</b>	<ul style="list-style-type: none"> <li>• React vigorously with halogens to form ionic halides.</li> <li>• Li forms covalent halides due to high polarisability.</li> <li>• Among halides, the covalent character increase from F to I (Fajan's Rules)</li> <li>• MP &amp; BP follow : F &gt; Cl &gt; Br &gt; I increase in lattice enthalpy)</li> <li>• Low solubility of LiF in water is due to high lattice enthalpy.</li> <li>• Low solubility of CsI is due to small hydration enthalpy of two ions</li> </ul>	<ul style="list-style-type: none"> <li>• Combine with halogens at elevated temperature.</li> <li>• Ionic character increases from Be to Ra.</li> <li>• BeCl<sub>2</sub> in solid state has a chain structure.</li> </ul>  <p>(as these are e<sup>-</sup>-deficient therefore to complete octet, they remain in polymeric form)</p> <ul style="list-style-type: none"> <li>• In vapour state BeCl<sub>2</sub> forms dimer with Cl as bridge which dissociates into linear monomer at high temperature.</li> </ul>  <ul style="list-style-type: none"> <li>• Tendency to form hydrated halides decreases down the group (MgCl<sub>2</sub>.8H<sub>2</sub>O, CaCl<sub>2</sub>.6H<sub>2</sub>O, BaCl<sub>2</sub>.2H<sub>2</sub>O)</li> </ul>
<b>Reactivity towards hydrogen</b>	<ul style="list-style-type: none"> <li>• All metals react to form ionic hydrides which have high melting points</li> <li>• Ionic nature increases from Li to Cs</li> </ul>	<ul style="list-style-type: none"> <li>• All the elements except Be combine with H<sub>2</sub> to form MH<sub>2</sub></li> <li>• BeH<sub>2</sub> &amp; MgH<sub>2</sub> are covalent &amp; polymeric whereas others are ionic.</li> </ul>
<b>Salts of Oxo-acids (H<sup>+</sup> on hydroxyl group) (Carbonates &amp; bicarbonates)</b>	<ul style="list-style-type: none"> <li>• Forms salts with all oxo acids (H<sub>2</sub>CO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>).</li> <li>• Salts are soluble in water &amp; thermally stable.</li> <li>• Thermal stability increases down the group.</li> <li>• Li<sub>2</sub>CO<sub>3</sub> decomposes on heating &amp; its bicarbonates does not exist as a solid,</li> </ul> $\text{Li}_2\text{CO}_3 \rightarrow \text{Li}_2\text{O} + \text{CO}_2$ <ul style="list-style-type: none"> <li>• Bicarbonates of other metals decomposes at low temperature.</li> <li>• Nitrates : Lithium nitrate on heating gives oxide whereas other alkali metals give nitrites.</li> </ul> $4\text{LiNO}_3 \rightarrow 2\text{Li}_2\text{O} + 4\text{NO}_2 + \text{O}_2$ $2\text{NaNO}_3 \rightarrow 2\text{NaNO}_2 + \text{O}_2$	<ul style="list-style-type: none"> <li>• Carbonates are insoluble in water &amp; their solubility decreases with increases in at no of metal ion.</li> <li>• Thermal stability increases with increase in cationic size BeCO<sub>3</sub> is unstable &amp; is stored only in the atmosphere of CO<sub>2</sub>.</li> <li>• <b>Sulphates</b> are thermally stable &amp; white in color.</li> <li>• The solubility of sulphates decreases down the group because SO<sub>4</sub><sup>2-</sup> is large &amp; masks the cation in lattice thus the size of cation does not influence lattice energy but hydration energy decreases from Be<sup>2+</sup> to Ba<sup>2+</sup>.</li> <li>• <b>Nitrates</b> are prepared by dissolution of carbonates in dil HNO<sub>3</sub>.</li> </ul>

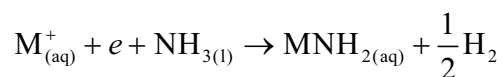
		<ul style="list-style-type: none"> <li>The tendency to form hydrated nitrates decreases with increase in size of cation. (BaO is soluble in H<sub>2</sub>O but BaSO<sub>4</sub> does not)</li> <li>All nitrates decompose on heating to give oxide.  <math display="block">2M(NO_3)_2 \rightarrow 2MO + 4NO_2 + O_2</math> </li> </ul>
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### Solutions in liquid NH<sub>3</sub>

- Alkali metals as well as alkali earth metals dissolve in liquid NH<sub>3</sub> giving deep blue solutions which are conducting in nature.



- When light falls on the ammoniated electron, they absorb energy corresponding to red color and the transmitted light has a blue color.
- The blue solution is paramagnetic which on standing slowly liberate H<sub>2</sub> & formation of metal amide.



- The electrical conductivity of the solution is due to ammoniated cation as well as ammoniated electron.
- In concentrated solution the color changes from blue to bronze due to pairing of electron.
- The concentrated solutions are diamagnetic in nature.

### ANOMALOUS BEHAVIOUR OF HEAD ELEMENTS

The elements belonging to second period sometimes called the head elements of their respective groups. The first element in each group (1, 2, 13, 14, 15, 16, 17) exhibit certain properties which are different from that of other elements in their respective groups. This anomalous behavior of the first element in each group is due to :

- |                                      |                                    |
|--------------------------------------|------------------------------------|
| (i) Smaller size of their atoms      | (ii) Higher ionisation enthalpies  |
| (iii) High polarizing power          | (iv) Low electropositive character |
| (v) Non availability of 'd' orbitals |                                    |

Difference between Lithium & other alkali metals (Anomalous behaviour of Li)	Difference between Beryllium & other alkali earth Metals : (Anomalous behavior of Be)
<ol style="list-style-type: none"> <li>Li is much harder with high MP &amp; BP than other alkali metals</li> <li>Li is less reactive but strongest reducing agent (due to high hydration energy)</li> <li>Li forms Lithium monoxide with O<sub>2</sub> &amp; reacts directly with N<sub>2</sub> to form nitrides.</li> <li>LiCl is deliquescent &amp; forms hydrates like LiCl.2H<sub>2</sub>O</li> <li>Lithium hydrogen carbonate is not formed in solid state.</li> <li>Lithium nitrate on heating gives oxide whereas other alkali metals give nitrites  <math display="block">4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2</math> <math display="block">2NaNO_3 \rightarrow 2NaNO_2 + O_2</math> </li> <li>LiF &amp; Li<sub>2</sub>O are less soluble in water.</li> </ol>	<ol style="list-style-type: none"> <li>Be has higher MP &amp; BP than other alkali earth metals</li> <li>BeO is amphoteric while others are strong bases.</li> <li>Be &amp; Mg do not impart color to the flame.</li> <li>Be does not form peroxides.</li> <li>BeSO<sub>4</sub> is soluble in water.</li> <li>Be does not liberate H<sub>2</sub> from acids whereas others do</li> </ol>

### Important Compounds of alkali metals:

Name of compound	Name of process & Brief about the process	Related chemical equations
Sodium Carbonate (Washing Soda), $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Solvay Process: When $\text{CO}_2$ gas is passed through a brine solution saturated with ammonia, sodium bicarbonate is formed. Sodium bicarbonate on heating forms sodium carbonate	$2\text{NH}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow (\text{NH}_4)_2\text{CO}_3$ $(\text{NH}_4)_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow 2\text{NH}_4\text{HCO}_3$ $\text{NH}_4\text{HCO}_3 + \text{NaCl} \rightarrow \text{NH}_4\text{Cl} + \text{NaHCO}_3$ $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$
Sodium Chloride, $\text{NaCl}$	Evaporation of sea water: Evaporation of sea water gives crude salt which contains impurities of $\text{CaSO}_4$ , $\text{Na}_2\text{SO}_4$ etc. To obtain pure sodium chloride, the crude salt is dissolved in minimum amount of water and filtered to remove insoluble impurities. The solution is then saturated with $\text{HCl}$ gas. Crystals of pure sodium chloride separate out.	-
Sodium Hydroxide (Caustic Soda), $\text{NaOH}$	Electrolysis of $\text{NaCl}$ in Castner- Kellner cell: A brine solution is electrolysed using a mercury cathode and a carbon anode. $\text{Na}$ metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode. Sodium amalgam on treatment with water forms sodium hydroxide	Cathode : $\text{Na}^+ + \text{e}^- \xrightarrow{\text{Hg}} \text{Na} - \text{amalgam}$ Anode : $\text{Cl}^- \rightarrow \frac{1}{2}\text{Cl}_2 + \text{e}^-$ $2\text{Na} - \text{amalgam} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + 2\text{Hg} + \text{H}_2$

Sodium Hydrogencarbonate (Baking Soda), $\text{NaHCO}_3$	$\text{NaHCO}_3$ is made by saturating a solution of sodium carbonate with carbon dioxide	$\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow 2\text{NaHCO}_3$
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**Important Compounds of alkaline earth metals:**

Name of compound	Name of process & Brief about the process	Related chemical equations
Calcium Oxide or Quick Lime, $\text{CaO}$	It is prepared by heating limestone ( $\text{CaCO}_3$ ) in a rotary kiln at 1070-1270 K. The carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion	$\text{CaCO}_3 \xrightleftharpoons{\text{heat}} \text{CaO} + \text{CO}_2$
Calcium Hydroxide (Slaked lime), $\text{Ca(OH)}_2$	It is prepared by adding water to quicklime, $\text{CaO}$	$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
Calcium Carbonate, $\text{CaCO}_3$	It is prepared by passing carbon dioxide through slaked lime. Excess of carbon dioxide should be avoided since this leads to the formation of water soluble calcium hydrogen-carbonate	$\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$
	It is prepared by the addition of sodium carbonate to calcium chloride.	$\text{CaCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow \text{CaCO}_3 + 2\text{NaCl}$
Calcium Sulphate of (Plaster Paris), $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$	It is obtained when gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ , is heated to 393 K	$2(\text{CaSO}_4 \cdot 2\text{H}_2\text{O}) \rightarrow 2(\text{CaSO}_4) \cdot \text{H}_2\text{O} + 3\text{H}_2\text{O}$

### ❖ Diagonal relationship between Li and Mg

Li resembles Mg mainly due to similarity in sizes of their atoms and ions. The main points of similarity are:

- i) Both are quite hard.
- ii) Both LiOH and  $\text{Mg}(\text{OH})_2$  are weak bases.
- iii) Carbonates of both on heating decompose to produce oxides and carbon dioxide.
- iv) Both react with nitrogen to give ionic nitrides.
- v) Nitrates of both decompose on heating to give oxides.
- vi) Both Li and Mg do not form solid bicarbonates.
- vii) Because of covalent character LiCl and  $\text{MgCl}_2$  are soluble in ethanol.
- viii) The hydroxides, bicarbonates and fluorides of both Li and Mg are sparingly soluble in water.

### ❖ Biological importance of Na and K

- i) Sodium ions participate in the transmission of nerve signals.
- ii) Sodium ions also regulate flow of water across the cell membranes and in transport of sugars and amino acids into the cells.
- iii) Potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in oxidation of glucose to produce ATP.
- iv) Potassium ions in combination with sodium ions are responsible for transmission of nerve signals.
- v) The functional features of nerve cells depend upon the sodium potassium ion gradient that is established in the cell.

### ❖ Biological importance of Ca and Mg

- i) Magnesium ions are concentrated in animal cells and Calcium ions are concentrated in body fluids, outside the cell.
- ii) All enzymes that utilize ATP in phosphate transfer require magnesium ion as cofactor.
- iii) In green plants magnesium is present in chlorophyll.
- iv) Calcium and magnesium ions are also essential for the transmission of impulses along nerve fibres.
- v) Calcium ions are important in blood clotting and are required to trigger the contraction of muscles.
- vi) Calcium ions also regulate the beating of the heart.

## S BLOCK AND P BLOCK ELEMENTS REASONING

- 1. Explain why is sodium less reactive than potassium?**  
Ionization energy of  $K < Na$  i.e. the outermost electron in potassium can be lost easily as compared to sodium
- 2. Why is the Li is strongest reducing agent or SRP is most -ve, i.e. = -3.0 V**  
Due to high heat of hydration which compensate its high heat of ionization make total heat to be more exothermic
- 3. Why is Group I elements known as the most electropositive element**  
Ionization energy is least for 1st gp. The loosely held s-electron in the outermost valence shell of these elements makes them the most electropositive metals. They readily lose electron to give monovalent  $M^+$  ions.
- 4. What do you mean by diagonal relationship and its reason in the periodic table,?**  
Similarities in few property of two elements placed diagonally in second and 3<sup>rd</sup> period, The diagonal relationship is due to the similarity in ionic sizes and /or charge / radius ratio of the elements.
- 5. Why alkali metal hydroxides are make the strongest bases?**  
The alkali metal hydroxides are the strongest of all bases because they dissolve freely in water with evolution of much heat on account of intense hydration.
- 6. What makes lithium show properties different from rest of the alkali metals?**  
Lithium is a small atom and it forms smaller  $Li^+$ . As a result, it has very high charge to radius ratio. This is primarily responsible for the anomalous behavior of lithium.
- 7. Explain why alkali and alkaline earth metals cannot be obtained by chemical reduction methods?**  
Alkali metals and alkaline earth metals are among the strongest reducing agents and the reducing agents that are stronger than them are not available
- 8. Why are potassium and cesium, rather than lithium used in photoelectric cells?**  
because as compared to Cs and K, Li is smaller in size and therefore, requires high energy to lose an electron. While on the other hand, K and Cs have low ionization energy
- 9. When an alkali metal dissolves in liquid ammonia the solution can acquire different colours. Explain the reasons for this type of colour change**  
When an alkali metal is dissolved in liquid ammonia, it results in the formation of a deep blue coloured solution.  
$$M + (x+y) NH_3 \longrightarrow M^+(NH_3)_x + e^-(NH_3)_y$$
- 10. Why are peroxides and super oxides stable in comparison to other oxides down the group?**  
The stability of peroxides and super oxides is due to the stabilization of large anions by larger cations through lattice energy effects.
- 11. Why are lithium compounds soluble in organic solvents**  
Due to high polarizing power, there is increased covalent character of lithium compounds which is responsible for their solubility in organic solvents.
- 12. Why does table salt get wet in rainy season?**  
Table salts contains impurities of  $CaCl_2$  and  $MgCl_2$  which being deliquescent compounds absorb moisture from the air in rainy season.
- 13. The atomic radii of alkaline earth metals are smaller than those of the corresponding alkali metals. Explain why?**  
Beryllium, Magnesium, Calcium, Strontium, Barium and Radium.
- 14. The second ionization enthalpy of calcium is more than the first. How is that calcium forms  $CaCl_2$  and not  $CaCl$  give reasons.**  
The higher value of second ionization enthalpy is more than compensated by the higher enthalpy of lattice formation of  $Ca^{2+}$ . Therefore formation of  $CaCl_2$  becomes more favorable than  $CaCl$  energetically.
- 15. Compounds of alkaline earth metals are more extensively hydrated than those of alkali metals. Give reason**  
The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions because of smaller size.
- 16. Why does the solubility of alkaline earth metal hydroxides in water increase down the group?**  
Among alkaline earth metal hydroxides, the anion being common the cationic radius will influence the lattice enthalpy. Since lattice enthalpy decreases much more than the hydration enthalpy with increasing ionic size, the solubility increases as we go down the group.



- 17. Why does the solubility of alkaline earth metal carbonates and sulphates in water decrease down the group?**  
The size of anions being much larger compared to cations, the lattice enthalpy will remain almost constant within a particular group. Since the hydration enthalpies decrease down the group, solubility will decrease as found for alkaline earth metal carbonates and sulphates.
- 18. Beryllium and magnesium do not give colour to flame whereas other alkaline earth metals do so. Why?**  
In Be and Mg, the electrons are strongly bound. The energy required to excite these electrons is very high
- 19. Potassium carbonate cannot be prepared by Solvay process. Why?**  
**This is because unlike sodium bicarbonate, potassium bicarbonate is fairly soluble in water and does not precipitate out.**
- 20. Why is  $\text{KO}_2$  paramagnetic ?**  
The superoxide  $\text{O}_2^-$  is paramagnetic because of one unpaired electron in  $\pi^*2p$  molecular orbital.
- 21. Why is  $\text{Li}_2\text{CO}_3$  decomposed at a lower temperature whereas  $\text{Na}_2\text{CO}_3$  at higher temperature?**  
This is because lithium carbonate is covalent. Lithium ion, being very small in size, polarizes a large carbonate ion,
- 22. The hydroxides and carbonates of sodium and potassium are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water. Explain.**  
the lattice energies of carbonates and hydroxides formed by calcium and magnesium are much more than those of sodium and potassium
- 23. Why are lithium salts commonly hydrated and those of the other alkali metal ions usually anhydrous?**  
Lithium is the smallest in size among the alkali metals. Hence,  $\text{Li}^+$  ion can polarize water molecules more easily than other alkali metals.
- 24. Why is LiF almost insoluble in water whereas LiCl soluble not only in water but also in acetone?**  
This is mainly because of the greater ionic character of LiF as compared to LiCl
- 25. The mobilities of the alkali metal ions in aqueous solution are  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$**   
Smaller the size of an ion, the more highly is it hydrated. Since  $\text{Li}^+$  is the smallest, it gets heavily hydrated in an aqueous solution. On the other hand,  $\text{Cs}^+$  is the largest and so it is the least hydrated, Greater the mass of a hydrated ion, the lower is its ionic mobility
- 26. Lithium is the only alkali metal to form a nitride directly.**  
Li reacts directly with nitrogen to form lithium nitride. This is because  $\text{Li}^+$  is very small in size and so its size is the most compatible with the  $\text{N}^{3-}$  ion. Hence, the lattice energy released is very high. This energy also overcomes the high amount of energy required for the formation of the  $\text{N}^{3-}$  ion.
- 27. a solution of  $\text{Na}_2\text{CO}_3$  is alkaline ?**  
its hydrolyses to give sodium bicarbonate and sodium hydroxide (a strong base). As a result, the solution becomes alkaline.  $\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} \longrightarrow \text{NaHCO}_3 + \text{NaOH}$
- 28. alkali metals are prepared by electrolysis of their fused chlorides ?**  
as they themselves are very strong reducing agents. They cannot be prepared by displacement reactions Neither can electrolysis of aqueous solutions be used to extract these elements. This is because the liberated metals react with water.
- 29. sodium is found to be more useful than potassium ?**  
Sodium ions are involved in the transmission of nerve signals, in regulating the flow of water across the cell membranes,
- 30. Why does beryllium show similarities with Al?**  
Because of their similarity in charge / radius ratios
- 31. Why is beryllium carbonate unusually unstable thermally as compared to the other carbonates of this group?**  
This is due to strong polarizing effect of small  $\text{Be}^{2+}$  on the large  $\text{CO}_3^{2-}$  anion and leading to the formation of more stable BeO
- 32. BeO is almost insoluble but  $\text{BeSO}_4$  is soluble in water,**

The size compatibility of  $\text{Be}^{2+}$  and  $\text{O}^{2-}$  is high. Therefore, the lattice energy released during their formation is also very high and can not be overcome by Hyd. Energy. On the other hand,  $\text{SO}_4^{2-}$  ion is a large anion. Hence,  $\text{Be}^{2+}$  can easily polarise  $\text{SO}_4^{2-}$  ions, making  $\text{BeSO}_4$  unstable. Thus, the lattice energy of  $\text{BeSO}_4$  is not very high and so it is soluble in water.

**33. BaO is soluble but  $\text{BaSO}_4$  is insoluble in water,**

The size compatibility of  $\text{Ba}^{2+}$  and  $\text{O}^{2-}$  is not high. As a result,  $\text{BaO}$  is unstable. The lattice energy released during its formation is also not very large. It can easily be overcome by the hydration energy of the ions.

Therefore,  $\text{BaO}$  is soluble in water. In  $\text{BaSO}_4$ ,  $\text{Ba}^{2+}$  and  $\text{SO}_4^{2-}$  are both large-sized. The lattice energy released is high. Hence, it is not soluble in water.

**34. LiI is more soluble than KI in ethanol.**

As a result of its small size, the lithium ion has a higher polarising power than the potassium ion. It polarises the electron cloud of the iodide ion to a much greater extent than the potassium ion. This causes a greater covalent character in LiI than in KI. Hence, LiI is more soluble in ethanol.

**35. What is 'inert pair effect'?**

Inertness of the electron pair of ns orbital which occurs down the group due to poor overlapping bond energy can not compensate for excitation energy, which fails to promote e- from ns orbital and oxidation states decrease by one unit than the group oxidation states are attributed to the 'inert pair effect'.

**36. How can you explain higher stability of  $\text{BCl}_3$  as compared to  $\text{TlCl}_3$ ?**

In Tl, the +3 state is highly oxidising and it reverts back to the more stable +1 state. due to inert pair effect, +3 In Tl, is unstable.

**37. Why does boron trifluoride behave as a Lewis acid?**

It has three electrons in its valence shell. Thus, it can form only three covalent bonds and one p orbital is vacant. This means that there are only six electrons around boron and its octet remains incomplete.

**38. Consider the compounds,  $\text{BCl}_3$  and  $\text{CCl}_4$ . How will they behave with water? Justify**

Being a Lewis acid,  $\text{BCl}_3$  readily undergoes hydrolysis. Boric acid is formed as a result

$\text{BCl}_3 + 3\text{H}_2\text{O} \longrightarrow 3\text{HCl} + \text{B}(\text{OH})_3$   $\text{CCl}_4$  completely resists hydrolysis. Carbon does not have any vacant orbital

**39. Is boric acid a protic acid? Explain.**

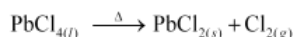
Boric acid is not a protic acid. It is a weak monobasic acid, behaving as a Lewis acid.  $\text{B}(\text{OH})_3 + 2\text{HOH} \longrightarrow [\text{B}(\text{OH})_4]^- + \text{H}_3\text{O}^+$  It behaves as an acid by accepting a pair of electrons from  $\text{OH}^-$  ion.

**40. Lead is known not to form an iodide,  $\text{PbI}_4$ .**

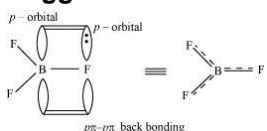
Due to inert pair effect  $\text{Pb}(\text{IV})$  is highly unstable and can reduce to  $\text{Pb}(\text{II})$ .  $\text{PbI}_4 \longrightarrow \text{PbI}_2 + \text{I}_2$

**41. Lead(IV) chloride is highly unstable towards heat.**

On moving down group IV, the higher oxidation state becomes unstable because of the inert pair effect.  $\text{Pb}(\text{IV})$  is highly unstable and when heated, it reduces to  $\text{Pb}(\text{II})$ .



**42. Suggest reasons why the B-F bond lengths in  $\text{BF}_3$  (130 pm) and  $\text{BF}_4^-$  (143 pm) differ.**



With a vacant p-orbital on boron, the fluorine and boron atoms undergo  $\pi\pi$ - $\pi\pi$  back-bonding to remove this deficiency. This imparts a double-bond character to the B-F bond. This double-bond character causes the bond length to shorten in  $\text{BF}_3$  (130 pm). However, when  $\text{BF}_3$  coordinates with the fluoride ion, a change in

hybridisation from  $sp^2$  (in  $\text{BF}_3$ ) to  $sp^3$  (in  $\text{BF}_4^-$ ) occurs. Boron now forms 4  $\sigma$  bonds and the double-bond character is lost.

This accounts for a B-F bond length of 143 pm in  $\text{BF}_4^-$  ion.

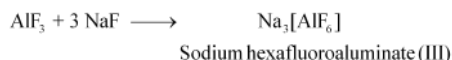
**43. If B-Cl bond has a dipole moment, explain why  $\text{BCl}_3$  molecule has zero dipole moment.**

This is because  $\text{BCl}_3$  is trigonal planar in shape. It is a symmetrical molecule. Hence, the respective dipole-moments of

the B–Cl bond cancel each other, thereby causing a zero-dipole moment.

**44. Aluminium trifluoride is insoluble in anhydrous HF but dissolves on addition of NaF**

(HF) is a covalent compound and has a very strong intermolecular hydrogen-bonding. Thus, it does not provide ions and aluminium fluoride (AlF) does not dissolve in it. Sodium fluoride (NaF) is an ionic compound and when it is added to the mixture, AlF dissolves. This is because of the availability of free F<sup>-</sup>. The reaction involved in the process is:



**45. Aluminium trifluoride precipitates out of the resulting solution when gaseous BF<sub>3</sub> is bubbled through. Give reasons.**

This happens because the tendency of boron to form complexes is much more than that of aluminium. Therefore, when BF<sub>3</sub> is added to the solution, B replaces Al from the complexes according to the following reaction:



**46. Suggest a reason as to why CO is poisonous.**

Carbon monoxide is highly-poisonous because of its ability to form a complex with haemoglobin. The CO–Hb complex is more stable than the O<sub>2</sub>–Hb complex. The former prevents Hb from binding with oxygen. Thus, a person dies because of suffocation on not receiving oxygen. It is found that the

**47. How is excessive content of CO<sub>2</sub> responsible for global warming?**

Carbon dioxide has the property of trapping the heat provided by sunrays. Higher the level of carbon dioxide, higher is the amount of heat trapped. This results in an increase in the atmospheric temperature, thereby causing global warming.

**48. Conc. HNO<sub>3</sub> can be transported in aluminium container.**

Conc. HNO<sub>3</sub> form a thin protective oxide layer on the aluminium surface. This oxide layer renders aluminium passive.

**49. A mixture of dilute NaOH and aluminium pieces is used to open drain.**

Sodium hydroxide and aluminium react to form sodium tetrahydroxoaluminate(III) and hydrogen gas. The pressure of the produced hydrogen gas is used to open blocked drains.



**50. Graphite is used as lubricant.**

different layers of graphite are bonded to each other by weak van der Waals' forces. These layers can slide over each other. Graphite is soft and slippery

**51. Aluminium utensils should not be kept in water overnight.**

However, when water is kept in an aluminium vessel for long periods of time, some amount of aluminium oxide may dissolve in water. As aluminium ions are harmful, water should not be stored in aluminium vessels overnight.

**52. How would you explain the lower atomic radius of Ga as compared to Al?**

Although Ga has one shell more than Al, its size is lesser than Al. This is because of the poor shielding effect of the 3d<sup>-</sup> electrons. The shielding effect of d<sup>-</sup> electrons is very poor and the effective nuclear charge experienced by the valence electrons in gallium is much more than it is in the case of Al.

**53. Classify following oxides as neutral, acidic, basic or amphoteric: CO, B<sub>2</sub>O<sub>3</sub>, SiO<sub>2</sub>, CO<sub>2</sub>, Al<sub>2</sub>O<sub>3</sub>, PbO<sub>2</sub>, Ti<sub>2</sub>O<sub>3</sub>**

(1) CO = Neutral

(2) B<sub>2</sub>O<sub>3</sub> = Acidic

Being acidic, it reacts with bases to form salts. It reacts with NaOH to form sodium metaborate.



(3) SiO<sub>2</sub> = Acidic

Being acidic, it reacts with bases to form salts. It reacts with NaOH to form sodium silicate.

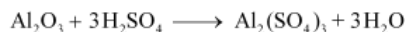
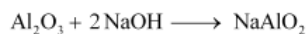


(4) CO<sub>2</sub> = Acidic

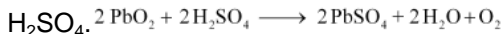
Being acidic, it reacts with bases to form salts. It reacts with NaOH to form sodium carbonate.  $\text{CO}_2 + 2 \text{NaOH} \longrightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$

(5)  $\text{Al}_2\text{O}_3 = \text{Amphoteric}$

Amphoteric substances react with both acids and bases.  $\text{Al}_2\text{O}_3$  reacts with both  $\text{NaOH}$  and  $\text{H}_2\text{SO}_4$ .



(6)  $\text{PbO}_2 = \text{Amphoteric}$  Amphoteric substances react with both acids and bases.  $\text{PbO}_2$  reacts with both  $\text{NaOH}$  and



(7)  $\text{Tl}_2\text{O}_3 = \text{Basic}$  Being basic, it reacts with acids to form salts. It reacts with  $\text{HCl}$  to form thallium chloride.



54. Aluminium forms  $[\text{AlF}_6]^{3-}$  whereas  $[\text{BF}_6]^{3-}$  is not formed why?

Due to presence of vacant d-orbital's, Al can expand its octet to form bonds with six fluoride ions whereas B cannot. Boron does not have d-orbital's

55. **Some halides of group 14 elements form complexes of the type  $[\text{MX}_6]^{2-}$ . Give reason.**

the central atom can increase its coordination number from 4 to 6 due to availability of vacant d-orbital's.

56. **Why is carbon different from other member of the group?**

Carbon differs from rest of the members of its group due to its smaller size, higher electro negativity, higher ionization enthalpy and unavailability of d-orbital's.

57. **Why does the covalence of carbon not expand beyond four?**

In carbon, only s and p orbital's are available for bonding and therefore it can accommodate only four pairs of electrons around it. This limit the maximum covalence to four whereas other members can expand their covalence due to the presence of d-orbital's.

58. **Why does the heavier elements do not form  $\text{p}\pi - \text{p}\pi$  multiple bond as carbon do?**

because their atomic orbital's are too large and diffuse to have effective overlapping.

59. **Why solid  $\text{CO}_2$  is called dry ice?**

Solid  $\text{CO}_2$  on heating sublimates without melting

60. **Why does  $\text{CO}_2$  have a linear shape with no dipole moment**

In  $\text{CO}_2$  molecule carbon atom undergoes sp hybridization. Two sp hybridized orbital of carbon atom overlap with two p-orbital's of oxygen atoms to make two sigma bonds while other two electrons of carbon atom are involved in  $\text{p}\pi - \text{p}\pi$  bonding with oxygen atom. This results in its linear shape [with both c-o bond of equal length (115 pm)] with no dipole moment.

61. **Why does graphite conduct electricity?**

Graphite forms hexagonal ring and undergoes sp<sup>2</sup> hybridization. The electrons are delocalized over the whole sheet. Electrons are mobile and therefore graphite conducts electricity over the sheet.

62. **White fumes appear around the bottle of anhydrous aluminium chloride. Give reason.**

Anhydrous aluminium chloride is partially hydrolysed with atmospheric moisture to liberate  $\text{HCl}$  gas. Moist  $\text{HCl}$  appears white in colour.

63. **Boron is unable to form  $\text{BF}_6^{3-}$  ion. Explain.**

Due to non-availability of d orbitals, boron is unable to expand its octet. Therefore, the maximum covalence of boron cannot exceed 4.

64. **Diamond is covalent, yet it has high melting point. Why ?**

Diamond has a three-dimensional network involving strong C—C bonds, which are very difficult to break and, in turn has high melting point.