

# Chapter 18 s and p-Block Elements

## Alkali Metals and Their Compounds

The group 1 of the periodic table contains six elements, namely lithium (*Li*), sodium (*Na*), potassium (*K*), rubidium (*Rb*), caesium (*Cs*) and francium (*Fr*). All these elements are **typical metals**. Francium is radioactive with longest lived isotope  $^{223}\text{Fr}$  with half life period of only 21 minute. These are usually referred to as alkali metals since their hydroxides form strong bases or alkalies.

### (1) Electronic configuration

Elements	Discovery	Electronic configuration ( $ns^1$ )
${}_3\text{Li}$	Arfwedson (1817)	$[\text{He}]^2 2s^1$
${}_{11}\text{Na}$	Davy (1807)	$[\text{Ne}]^{10} 3s^1$
${}_{19}\text{K}$	Davy (1807)	$[\text{Ar}]^{18} 4s^1$
${}_{37}\text{Rb}$	Bunsen (1861)	$[\text{Kr}]^{36} 5s^1$
${}_{55}\text{Cs}$	Bunsen (1860)	$[\text{Xe}]^{54} 6s^1$
${}_{87}\text{Fr}$	Percy (1939)	$[\text{Rn}]^{86} 7s^1$

(2) **Occurrence** : Alkali metals are very reactive and thus found in combined state some important ores of alkali metals are given ahead.

(i) **Lithium** : Triphylite, Petalite, lepidolite, Spodumene [ $\text{LiAl}(\text{SiO}_3)_2$ ], Amblygonite [ $\text{Li}(\text{AlF})\text{PO}_4$ ]

(ii) **Sodium** : Chile salt petre ( $\text{NaNO}_3$ ), Sodium chloride ( $\text{NaCl}$ ), Sodium sulphate ( $\text{Na}_2\text{SO}_4$ ), Borax ( $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ ), Glauber salt ( $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ )

(iii) **Potassium** : Sylime ( $\text{KCl}$ ), carnallite ( $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ ) and Felspar ( $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$ )

(iv) **Rubidium** : Lithium ores Lepidolite, triphylite contains 0.7 to 3% *RbO*

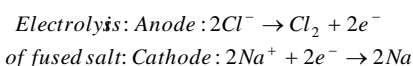
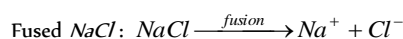
(v) **Caesium** : Lepidolite, Pollucite contains 0.2 to 7% *CsO*

(3) **Extraction of alkali metals** : Alkali metals cannot be extracted by the usual methods for the extraction of metals due to following reasons.

(i) Alkali metals are strong reducing agents, hence cannot be extracted by reduction of their oxides or other compounds.

(ii) Being highly electropositive in nature, it is not possible to apply the method of displacing them from their salt solutions by any other element.

(iii) The aqueous solutions of their salts cannot be used for extraction by electrolytic method because hydrogen ion is discharged at cathode instead of an alkali metal ions as the discharge potentials of alkali metals are high. However, by using *Hg* as cathode, alkali metal can be deposited. The alkali metal readily combines with *Hg* to form an amalgam from which its recovery difficult. The only successful method, therefore, is the electrolysis of their fused salts, usually chlorides. Generally, another metal chloride is added to lower their fusion temperature.



### (4) Alloys Formation

(i) The alkali metals form alloys among themselves as well as with other metals.

(ii) Alkali metals also get dissolved in mercury to form amalgam with evolution of heat and the amalgamation is highly exothermic.

### Physical properties

#### (1) Physical state

(i) All are silvery white, soft and light solids. These can be cut with the help of knife. When freshly cut, they have bright lustre which quickly tarnishes due to surface oxidation.

(ii) These form diamagnetic colourless ions since these ions do not have unpaired electrons, (*i.e.* *M* has *ns* configuration). That is why alkali metal salts are colourless and diamagnetic.

#### (2) Atomic and ionic radii

(i) The alkali metals have largest atomic and ionic radii than their successive elements of other groups belonging to same period.

(ii) The atomic and ionic radii of alkali metals, however, increases down the group due to progressive addition of new energy shells.

No doubt the nuclear charge also increases on moving down the group but the influence of addition of energy shell predominates

	<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>	<i>Fr</i>
Atomic radius (pm)	152	186	227	248	265	375
Ionic radius of <i>M</i>	60	95	133	148	169	-

ions (pm)

### (3) Density

(i) All are light metals, *Li*, *Na* and *K* have density less than water. Low values of density are because these metals have high atomic volume due to larger atomic size. On moving down the group the atomic size as well as atomic mass both increase but increase in atomic mass predominates over increase in atomic size or atomic volume and therefore the ratio mass/volume i.e. density gradually increases down the groups

(ii) The density increases gradually from *Li* to *Cs*, *Li* is lightest known metal among all.

$Li = 0.534$ ,  $Na = 0.972$ ,  $K = 0.86$ ,  $Rb = 1.53$  and  $Cs = 1.87$  g/ml at 20°C.

(iii) *K* is lighter than *Na* because of its unusually large atomic size.

(iv) In solid state, they have body centred cubic lattice.

### (4) Melting point and Boiling point

(i) All these elements possess low melting point and boiling point in comparison to other group members.

	<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>	<i>Fr</i>
melting point (K)	453.5	370.8	336.2	312.0	301.5	–
boiling point (K)	1620	1154.4	1038.5	961.0	978.0	–

(ii) The lattice energy of these atoms in metallic crystal lattice relatively low due to larger atomic size and thus possess low melting point and boiling point on moving down the group, the atomic size increases and binding energy of their atoms in crystal lattice decreases which results lowering of melting point.

(iii) Lattice energy decreases from *Li* to *Cs* and thus melting point and boiling also decreases from *Li* to *Cs*.

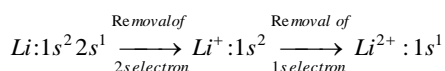
### (5) Ionisation energy and electropositive or metallic character

(i) Due to unpaired lone electron in *ns* sub-shell as well as due to their larger size, the outermost electron is far from the nucleus, the removal of electron is easier and these low values of ionisation energy. (*IE*)

(ii) Ionisation energy of these metal decreases from *Li* to *Cs*.

Ionisation energy	<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>	<i>Fr</i>
$IE_1$	520	495	418	403	376	–
$IE_2$	7296	4563	3069	2650	2420	–

A jump in 2nd ionisation energy (huge difference) can be explained as,



Removal of 1s electrons from *Li* and that too from completely filled configuration requires much more energy and a jump in 2nd ionisation is noticed.

(iii) Lower are ionisation energy values, greater is the tendency to lose *ns* electron to change in *M* ion (i.e.  $M \rightarrow M^+e$ ) and therefore stronger is electropositive character.

(iv) Electropositive character increases from *Li* to *Cs*.

Due to their strong electropositive character, they emit electrons even when exposed to light showing **photoelectric effect**. This property is responsible for the use of *Cs* and *K* in photoelectric cell.

### (6) Oxidation number and valency

(i) Alkali metals are univalent in nature due to low ionisation energy values and form ionic compounds. Lithium salts are, however, covalent.

(ii) Further, the *M* ion acquires the stable noble gas configuration. It requires very high values of energy to pull out another electron from next to outer shell of *M* ion and that is why their second ionisation energy is very high. Consequently, under ordinary conditions, it is not possible for these metals to form *M* ion and thus they show +1 oxidation state.

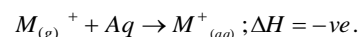
(iii) Since the electronic configuration of *M* ions do not have unpaired electron and thus alkali metal salts are diamagnetic and colourless. Only those alkali metal salts are coloured which have coloured anions e.g.

$KCr_2O_7$  is orange because of orange coloured  $Cr_2O_7^{2-}$  ion,  $KMnO_4$  is violet because of violet coloured  $MnO_4^-$  ion.

### (7) Hydration of ions

(i) Hydration represents for the dissolution of a substance in water to get adsorb water molecule by weak valency force. Hydration of ions is the exothermic process (i.e. energy is released during hydration) when ions on dissolution water get hydration.

(ii) The energy released when 1 mole of an ion in the gaseous state is dissolved in water to get it hydrated is called hydration energy



(iii) Smaller the cation, greater is the degree of hydration. Hydration energy is in the order of,  $Li > Na > K > Rb > Cs$

(iv) *Li* being smallest in size has maximum degree of hydration and that is why lithium salts are mostly hydrated,  $LiCl \cdot 2H_2O$  also lithium ion being heavily hydrated, moves very slowly under the influence of electric field and, therefore, is the poorest conductor current among alkali metals ions. It may, therefore, be concluded that it is the degree of hydration as well as the size of ion is responsible for the current carried by an ion.

Relative ionic radii  $Cs^+ > Rb^+ > K^+ > Na^+ > Li^+$

Relative hydrated ionic radii  $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

Relative conducting power  $Cs^+ > Rb^+ > K^+ > Na^+ > Li^+$

### (8) Electronegativity, Electro positivity and metallic character.

(i) These metals are highly electropositive and thereby possess low values of electronegativities. Metallic character and electro positivity increase from *Li* to *Cs* ( $Li < Na < K < Rb < Cs$ )

(ii) Electronegativity of alkali metals decreases down the group as the trend of numerical values of electronegativity given below on Pauling scale suggests.

	<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>	<i>Fr</i>
Electronegativity	0.98	0.93	0.82	0.82	0.79	–

*Fr* being radioactive elements and thus studies on physical properties of this element are limited.

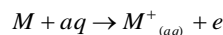
(9) **Specific heat** : It decreases from *Li* to *Cs*.

	<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>	<i>Fr</i>
Specific heat (Cal/g)	0.941	0.293	0.17	0.08	0.049	–

(10) **Conduction power** : All are good conductors of heat and electricity, because of loosely held valence electrons.

### (11) Standard oxidation potential and reduction properties

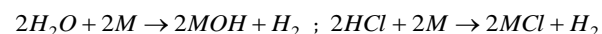
(i) Since alkali metals easily lose *ns* electron and thus they have high values of oxidation potential i.e.,



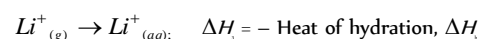
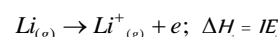
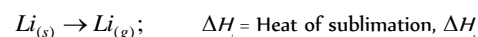
(ii) The standard oxidation potentials of a alkali metals (in volts) are listed below,

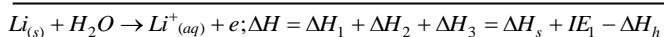
<i>Li</i>	<i>Na</i>	<i>K</i>	<i>Rb</i>	<i>Cs</i>
+3.05	+2.71	+2.93	+2.99	+2.99

(iii) More is oxidation potential, more is the tendency to get oxidized and thus more powerful is reducing nature in aqueous medium that is why alkali metals liberate  $H_2$  from  $H_2O$  and  $HCl$ .

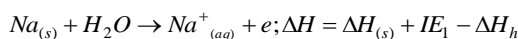


(iv) However, an examination of ionisation energy for alkali metals reveals that *Li* should have the minimum tendency to lose electron and thus its reducing nature should be minimum. The greatest reducing nature of *Li* in aq. medium is accounted due to the maximum hydration energy of *Li* ion. For Lithium





Similarly, for sodium,



$\Delta H_i$  for  $Li > \Delta H_i$  for  $Na$ . Therefore, large negative  $\Delta H$  values are observed in case of  $Li$  and this explains for more possibility of  $Li$  to get itself oxidized or have reducing nature.

**(12) Characteristic flame colours :** The alkali metals and their salts give characteristic colour to Bunsen flame. The flame energy causes and excitation of the outermost electron which on reverting back to its initial position gives out the absorbed energy as visible light. These colour differ from each other  $Li$  –crimson,  $Na$ –Golden yellow,  $K$  – Pale violet ,  $Rb$ –Red violet and  $Cs$  –Blue violet. These different colours are due to different ionisation energy of alkali metals. The energy released is minimum in the case of  $Li$  and increases in the order.

Energy released :  $Li < Na < K < Rb < Cs$

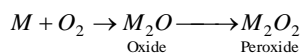
$\lambda$  released :  $Li > Na > K > Rb > Cs$

Frequency released :  $Li < Na < K < Rb < Cs$

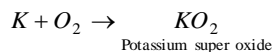
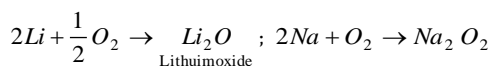
### Chemical properties

#### (1) Formation of oxides and hydroxides

(i) These are most reactive metals and have strong affinity for  $O$ , quickly tranish in air due to the formation of a film of their oxides on the surface. These are, therefore, kept under kerosene or paraffin oil to protect them from air,

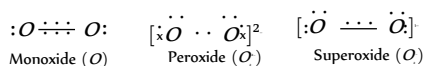


(ii) When burnt air ( $O$ ), lithium forms lithium oxide ( $LiO$ ) sodium forms sodium peroxide ( $Na_2O$ ) and other alkali metals form super oxide ( $MO$ , i.e.  $KO, RbO$ , or  $CsO$ )



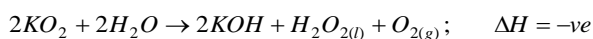
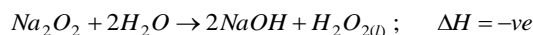
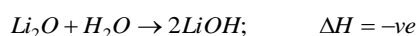
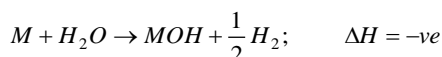
The reactivity of alkali metals towards oxygen to form different oxides is due to strong positive field around each alkali metal cation.  $Li$  being smallest, possesses strong positive field and thus combines with small anion  $O$  to form stable  $LiO$  compound. The  $Na$  and  $K$  being relatively larger thus exert less strong positive field around them and thus reacts with larger oxygen anion i.e.  $O_2^{2-}$  and  $O_2^{1-}$  to form stable oxides.

The monoxide, peroxides and superoxides have  $O$  and  $O_2^{2-}$ ,  $O_2^{1-}$  ions respectively. The structures of each are,



The  $O_2^{1-}$  ion has a three electron covalent bond and has one electron unpaired. It is therefore superoxides are paramagnetic and coloured  $KO_2$  is light yellow and paramagnetic substance.

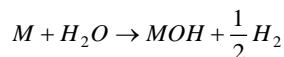
(iii) The oxides of alkali metals and metal itself give strongly alkaline solution in water with evolution of heat



The peroxides and superoxides act as strong oxidising agents due to formation of  $H_2O$ .

(iv) The reactivity of alkali metals towards air and water increases from  $Li$  to  $Cs$  that is why lithium decomposes  $H_2O$  very slowly at  $25^\circ C$

whereas  $Na$  does so vigorously,  $K$  reacts producing a flame and  $Rb, Cs$  do so explosively.



(v) The basic character of oxides and hydroxides of alkali metals increases from  $Li$  to  $Cs$ . This is due to the increase in ionic character of alkali metal hydroxides down the group which leads to complete dissociation and leads to increase in concentration of  $OH$  ions.

#### (2) Hydrides

(i) These metals combine with  $H$  to give white crystalline ionic hydrides of the general of the formula  $MH$ .

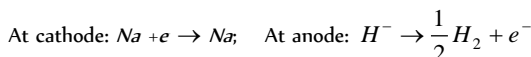
(ii) The tendency to form their hydrides, basic character and stability decreases from  $Li$  to  $Cs$  since the electropositive character decreases from  $Cs$  to  $Li$ .

$2M + H_2 \rightarrow 2MH$ ; Reactivity towards  $H_2$  is  $Cs < Rb < K < Na < Li$ .

(iii) The metal hydrides react with water to give  $MOH$  and  $H_2$ ;  $MH + H_2O \rightarrow MOH + H_2$

(iv) The ionic nature of hydrides increases from  $Li$  to  $Cs$  because of the fact that hydrogen is present in the these hydrides as  $H$  and the smaller cation will produce more polarisation of anion (according to Fajans rule) and will develop more covalent character.

(v) The electrolysis of fused hydrides give  $H_2$  at anode.  $NaH_{fused}$  Contains  $Na^+$  and  $H^-$  i.e.,



(vi) Alkali metals also form hydrides like  $NaBH_4$ ,  $LiAlH_4$  which are good reducing agent.

#### (3) Carbonates and Bicarbonates

(i) The carbonates ( $MCO_3$ ) & bicarbonates ( $MHCO_3$ ) are highly stable to heat, where  $M$  stands for alkali metals.

(ii) The stability of these salts increases with the increasing electropositive character from  $Li$  to  $Cs$ . It is therefore  $Li_2CO_3$  decompose on heating,  $Li_2CO_3 \rightarrow Li_2O + CO_2$

(iii) Bicarbonates are decomposed at relatively low temperature,  $2MHCO_3 \xrightarrow{300^\circ C} M_2CO_3 + H_2O + CO_2$

(iv) Both carbonates and bicarbonates are soluble in water to give alkaline solution due to hydrolysis of carbonate ions or bicarbonate ions.

#### (4) Halides

(i) Alkali metals combine directly with halogens to form ionic halide  $M^+ X^-$ .

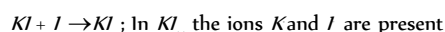
(ii) The ease with which the alkali metals form halides increases from  $Li$  to  $Cs$  due to increasing electropositive character from  $Li$  to  $Cs$ .

(iii) Lithium halides however have more covalent nature. Smaller is the cation, more is deformation of anion and thus more is covalent nature in compound. Also among lithium halides, lithium iodide has maximum covalent nature because of larger anion which is easily deformed by a cation. Thus covalent character in lithium halides is,  $LiI > LiBr > LiCl > LiF$

(iv) These are readily soluble in water. However, lithium fluoride is sparingly soluble. The low solubility of  $LiF$  is due to higher forces of attractions among smaller  $Li$  and smaller  $F$  ions (high lattice energy).

(v) Halides having ionic nature have high m.pt. and good conductor of current. The melting points of halides shows the order,  $NaF > NaCl > NaBr > NaI$

(vi) Halides of potassium, rubidium and caesium have a property of combining with extra halogen atoms forming polyhalides.

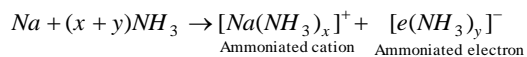


### (5) Solubility in liquid $NH_3$

(i) These metals dissolve in liquid  $NH_3$  to produce blue coloured solution, which conducts electricity to an appreciable degree.

(ii) With increasing concentration of ammonia, blue colour starts changing to that of metallic copper after which dissolution of alkali metals in  $NH_3$  ceases.

(iii) The metal atom is converted into ammoniated metal in i.e.  $M(NH_3)$  and the electron set free combines with  $NH_3$  molecule to produce ammonia solvated electron.



(iv) It is the ammoniated electron which is responsible for blue colour, paramagnetic nature and reducing power of alkali metals in ammonia solution. However, the increased conductance nature of these metals in ammonia is due to presence of ammoniated cation and ammonia solvated electron.

(v) The stability of metal-ammonia solution decreases from  $Li$  to  $Cs$ .

(vi) The blue solution on standing or on heating slowly liberates hydrogen,  $2M + 2NH_3 \rightarrow 2MNH_2 + H_2$ . Sodamide ( $NaNH_2$ ) is a waxy solid, used in preparation of number of sodium compounds.

(6) **Nitrates** : Nitrates of alkali metals ( $MNO_3$ ) are soluble in water and decompose on heating.  $LiNO_3$  decomposes to give  $NO_2$  and  $O_2$  and rest all give nitrites and oxygen.



### (7) Sulphates

(i) Alkali metals' sulphate have the formula  $MSO_4$ .

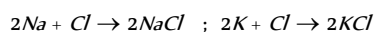
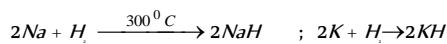
(ii) Except  $Li_2SO_4$ , rest all are soluble in water.

(iii) These sulphates on fusing with carbon form sulphides,  $MSO_4 + 4C \rightarrow MS + 4CO$

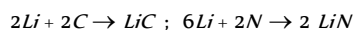
(iv) The sulphates of alkali metals (except  $Li$ ) form double salts with the sulphate of the trivalent metals like  $Fe$ ,  $Al$ ,  $Cr$  etc. The double sulphates crystallize with large number of water molecules as alum. e.g.  $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ .

### (8) Reaction with non-metals

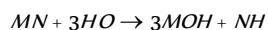
(i) These have high affinity for non-metals. Except carbon and nitrogen, they directly react with hydrogen, halogens, sulphur, phosphorus etc. to form corresponding compounds on heating.



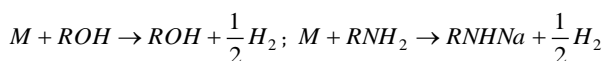
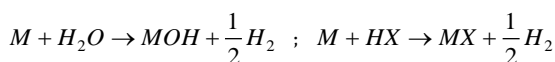
(ii)  $Li$  reacts, however directly with carbon and nitrogen to form carbides and nitrides.



(iii) The nitrides of these metals on reaction with water give  $NH_3$ .



(9) **Reaction with acidic hydrogen** : Alkali metals react with acids and other compounds containing acidic hydrogen (i.e.  $H$  atom attached on  $F$ ,  $O$ ,  $N$  and triply bonded carbon atom, for example,  $HF$ ,  $H_2O$ ,  $ROH$ ,  $RNH_2$ ,  $CH \equiv CH$ ) to liberate  $H_2$ .



(10) **Complex ion formation** : A metal shows complex formation only when it possesses the following characteristics, (i) Small size (ii) High nuclear charge (iii) Presence of empty orbitals in order to accept electron pair ligand. Only Lithium in alkali metals due to small size forms a few complex ions. Rest all alkali metals do not possess the tendency to form complex ion.

### Anomalous behaviour of Lithium

Anomalous behaviour of lithium is due to extremely small size of lithium its cation on account of small size and high nuclear charge, lithium exerts the greatest polarizing effect out of all alkali metals on negative ion. Consequently lithium ion possesses remarkable tendency towards solvation and develops covalent character in its compounds.  $Li$  differs from other alkali metals in the following respects,

(1) It is comparatively harder than other alkali metals.  $Li$  can't be stored in kerosene as it floats to the surface, due to its very low density.  $Li$  is generally kept wrapped in paraffin wax.

(2) It can be melted in dry air without losing its brilliance.

(3) Unlike other alkali metals, lithium is least reactive among all. It can be noticed by the following properties,

(i) It is not affected by air. (ii) It decomposes water very slowly to liberate  $H_2$ . (iii) It hardly reacts with bromine while other alkali metals react violently.

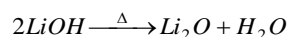
(4) Lithium is the only alkali metal which directly reacts with  $N_2$  to form Lithium nitride ( $Li_3N$ )

(5) Lithium when heated in  $NH_3$  forms amide,  $LiNH_2$  while other metals form amides,  $MNH_2$ .

(6) When burnt in air, lithium form  $Li_2O$  sodium form  $Na_2O$  and  $Na_2O_2$ , other alkali metals form monoxide, peroxide and superoxide.

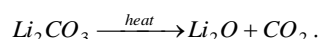
(7)  $Li_2O$  is less basic and less soluble in water than other alkali metals.

(8)  $LiOH$  is weaker base than  $NaOH$  or  $KOH$  and decomposes on heating.



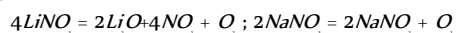
(9)  $LiHCO_3$  is liquid while other metal bicarbonates are solid.

(10) Only  $Li_2CO_3$  decomposes on heating



$Na_2CO_3$ ,  $K_2CO_3$  etc. do not decompose on heating.

(11)  $LiNO_3$  and other alkali metal nitrates give different products on heating



(12)  $LiCl$  and  $LiNO_3$  are soluble in alcohol and other organic solvents. These salts of other alkali metals are, however, insoluble in organic solvents.

(13)  $LiCl$  is deliquescent while  $NaCl$ ,  $KBr$  etc. are not. Lithium chloride crystals contain two molecules of water of crystallisation ( $LiCl \cdot 2H_2O$ ). Crystals of  $NaCl$ ,  $KBr$ ,  $KI$  etc do not contain water of crystallisation.

(14)  $Li_2SO_4$  does not form alums like other alkali metals.

(15)  $Li$  reacts with water slowly at room temperature  $Na$  reacts vigorously. Reaction with  $K$ ,  $Rb$  and  $Cs$  is violent.

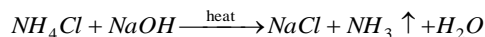
(16)  $Li$  reacts with  $Br_2$  slowly. Reaction of other alkali metals with  $Br_2$  is fast.

(17)  $Li_2CO_3$ ,  $Li_2C_2O_4$ ,  $LiF$ ,  $Li_3PO_4$  are the only alkali metal salts which are insoluble or sparingly soluble in water.

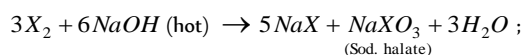
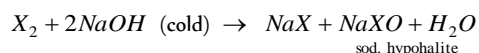
### Diagonal Relationship of $Li$ with $Mg$

Due to its small size lithium differs from other alkali metals but resembles with  $Mg$  as its size is closer to  $Mg$ . Its resemblance with  $Mg$  is known as diagonal relationship. Generally the periodic properties show either increasing or decreasing trend along the group and vice versa along



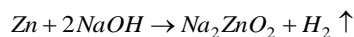


(b) **Reaction with halogens :**

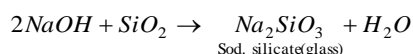


( $X = Cl, Br, I$ )

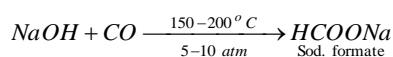
(c) **Reaction with metals :** Weakly electropositive metals like *Zn, Al* and *Sn* etc.



(d) **Reaction with sand, *SiO* :**



(e) **Reaction with *CO* :**



*NaOH* breaks down the proteins of the skin flesh to a pasty mass, therefore it is commonly known as caustic soda.

**Caustic property :** sodium hydroxide breaks down the proteins of the skin flesh to a pasty mass, therefore, it is commonly known as caustic soda.

**Uses :** Sodium hydroxide is used :

(a) in the manufacture of sodium metal, soap (from oils and fats), rayon, paper, dyes and drugs,

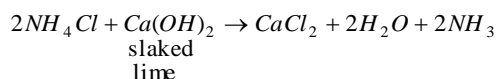
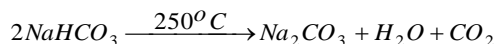
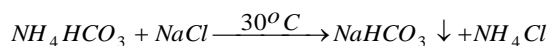
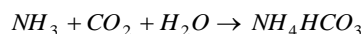
(b) for mercuring cotton to make cloth unshrinkable and

(c) as a reagent in the laboratory.

(iii) **Sodium carbonate or washing soda,  $Na_2CO_3$**

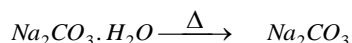
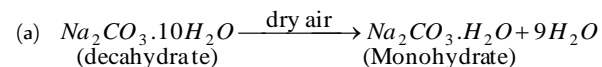
It exists in various forms, namely anhydrous sodium carbonate *NaCO* (soda-ash); monohydrate  $Na_2CO_3 \cdot H_2O$  (crystal carbonate); heptahydrate  $Na_2CO_3 \cdot 7H_2O$  and decahydrate  $Na_2CO_3 \cdot 10H_2O$  (washing soda or sal soda).

**Preparation :** (a) **Solvay process :** In this process, brine (*NaCl*),  $NH_3$  and  $CO_2$  are the raw materials.



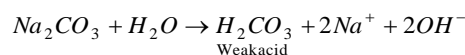
$CaCl_2$  so formed in the above reaction is a by product of solvay process.

**Properties**

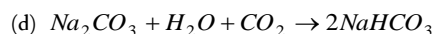


It does not decompose on further heating even to redness (m.pt.  $853^\circ C$ )

(b) It is soluble in water with considerable evolution of heat.



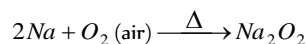
(c) It is readily decomposed by acids with the evolution of  $CO_2$  gas.



**Uses :** In textile and petroleum refining, Manufacturing of glass, *NaOH* soap powders etc.

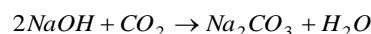
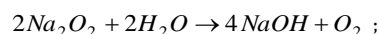
(iv) **Sodium peroxide ( $Na_2O$ )**

**Preparation :** It is manufactured by heating sodium metal on aluminium trays in air (free from  $CO_2$ )



**Properties :** (a) When pure it is colourless. The faint yellow colour of commercial product is due to presence of small amount of superoxide ( $NaO_2$ ).

(b) On coming with moist air it become white due to formation of *NaOH* and  $Na_2CO_3$ .

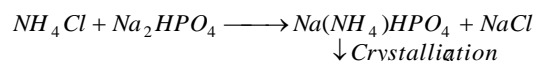


(c) It is powerful oxidising agent. It oxidises *Cr* (III) hydroxide to sodium chromate, *Mn* (II) to sodium manganate and sulphides to sulphates.

**Uses :** As a bleaching agent and it is used for the purification of air in confined spaces such as submarines because it can combine with  $CO_2$  to give  $Na_2CO_3$  and oxygen,  $2CO_2 + 2Na_2O_2 \rightarrow 2Na_2CO_3 + O_2$ .

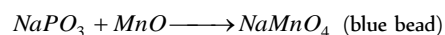
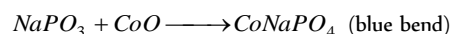
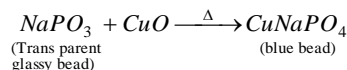
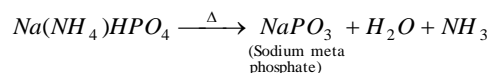
(v) **Micro cosmic salt [ $Na(NH)HPO_4 \cdot 4H_2O$ ]**

Prepared by dissolving equimolar amounts of  $Na_2HPO_4$  and  $NH_4Cl$  in water in 1 : 1 ratio followed by crystallization



**Chemical properties :**

On heating M.C.S,  $NaPO_3$  is formed.  $NaPO_3$  forms coloured beads with oxides of transition metal cloudy  $SiO_2$



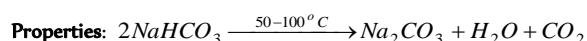
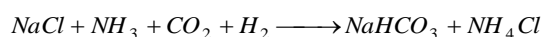
**Uses :** (a) For the formation of sodium meta phosphate and copper sodium phosphate

(b) It is used for the detection of colored ion

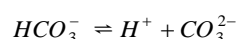
(c) It is especially used for testing silica with which a cloudy bead containing floating properties of silica is obtained.

(vi) **Sodium bi Carbonate ( $NaHCO_3$ , Baking soda)**

**Preparation :** It is an intermediate compound in manufacture of sodium carbonate by the solvay's process



It is amphiprotic  $HCO_3^- + H^+ \rightleftharpoons H_2CO_3$



**Uses :** (a) Baking powder contains  $NaHCO_3$ ,  $Ca(H_2PO_4)_2$  and starch.

Improved Baking powder contains 40% starch 30%  $NaHCO_3$ , 20%  $NaAl(SO_4)_2$  and 10%  $CaH_2(PO_4)$

(b) In pharmaceutical industry (Antacids etc.)

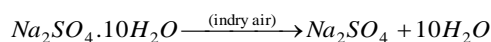
(c) Fire extinguishers.

(vii) **Sodium Sulphate  $Na_2SO_4$  or salt cake**

**Preparation :** It is the by-product of  $HCl$  industry



**Properties :** When aqueous solution of  $Na_2SO_4$  is cooled below  $32^\circ C$  Glauber's salt ( $Na_2SO_4 \cdot 10H_2O$ ) gets crystallised and if cooled to  $12^\circ C$ ,  $Na_2SO_4 \cdot 7H_2O$  crystals are formed.



**Uses :**  $Na_2SO_4$  finds use in paper industry detergent and glass manufacturing.

## Alkaline Earth Metals and Their Compounds

The group 2 of the periodic table consists of six metallic elements. These are beryllium ( $Be$ ), magnesium ( $Mg$ ), calcium ( $Ca$ ), strontium ( $Sr$ ), barium ( $Ba$ ) and radium ( $Ra$ ). These (except  $Be$ ) are known as alkaline earth metals as their oxides are alkaline and occur in earth crust.

### (i) Electronic configuration

Element	Electronic configurations ( $ns^2$ )
${}_4 Be$	$[He]2s^2$
${}_{12} Mg$	$[Ne]3s^2$
${}_{20} Ca$	$[Ar]4s^2$
${}_{38} Sr$	$[Kr]5s^2$
${}_{56} Ba$	$[Xe]6s^2$
${}_{88} Ra$	$[Rn]7s^2$

Radium was discovered in the ore pitch blende by madam Curie. It is radioactive in nature.

(2) **Occurrence :** These are found mainly in combined state such as oxides, carbonates and sulphates  $Mg$  and  $Ca$  are found in abundance in nature.  $Be$  is not very abundant,  $Sr$  and  $Ba$  are less abundant.  $Ra$  is rare element. Some important ores of alkaline earth metals are given below,

(i) **Beryllium :** Beryl ( $3BeO \cdot Al_2O_3 \cdot 6SiO_2$ ); Phenacite ( $Be_2SiO_5$ )

(ii) **Magnesium :** Magnesite ( $MgCO_3$ ); Dolomite ( $CaCO_3 \cdot MgCO_3$ ); Epsomite ( $MgSO_4 \cdot 7H_2O$ ); Carnallite ( $MgCl_2 \cdot KCl \cdot 6H_2O$ ); Asbestos [ $CaMg(SiO_3)_2$ ]

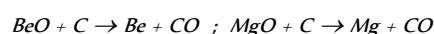
(iii) **Calcium :** Limestone ( $CaCO_3$ ); Gypsum : ( $CaSO_4 \cdot 2H_2O$ ), Anhydrite ( $CaSO_4$ ); Fluorapatite [ $(3Ca_3(PO_4)_2 \cdot CaF_2)$ ]; Phosphorite rock [ $Ca_3(PO_4)_2$ ]

(iv) **Barium :** Barytes ( $BaSO_4$ ) ; witherite ( $BaCO_3$ )

(v) **Radium :** Pitch blende ( $UO_2$ ); ( $Ra$  in traces); other radium rich minerals are carnotite [ $K_2UO_7$ ] ( $VO_3$ ),  $8H_2O$  and antamite [ $Ca(UO_2)_2$ ]

### (3) Extraction of alkaline earth metals

(i)  $Be$  and  $Mg$  are obtained by reducing their oxides carbon,



(ii) The extraction of alkaline earth metals can also be made by the reduction of their oxides by alkali metals or by electrolysis of their fused salts.

(4) **Alloy formation :** These dissolve in mercury and form amalgams.

### Physical properties

(1) **Physical state :** All are greyish-white, light, malleable and ductile metals with metallic lustre. Their hardness progressively decrease with increase in atomic number. Although these are fairly soft but relatively harder than alkali metals.

### (2) Atomic and ionic radii

(i) The atomic and ionic radii of alkaline earth metals also increase down the group due to progressive addition of new energy shells like alkali metals.

	$Be$	$Mg$	$Ca$	$Sr$	$Ba$	$Ra$
Atomic radius (pm)	112	160	197	215	222	—
Ionic radius of $M^{2+}$ ion (pm)	31	65	99	113	135	140

(ii) The atomic radii of alkaline earth metals are however smaller than their corresponding alkali metal of the same period. This is due to the fact that alkaline earth metals possess a higher nuclear charge than alkali metals which more effectively pulls the orbit electrons towards the nucleus causing a decrease in size.

### (3) Density

(i) Density decreases slightly upto  $Ca$  after which it increases. The decrease in density from  $Be$  to  $Ca$  might be due to less packing of atoms in solid lattice of  $Mg$  and  $Ca$ .

$Be$	$Mg$	$Ca$	$Sr$	$Ba$	$Ra$
1.84	1.74	1.55	2.54	3.75	6.00

(ii) The alkaline earth metals are more denser, heavier and harder than alkali metal. The higher density of alkaline earth metals is due to their smaller atomic size and strong intermetallic bonds which provide a more close packing in crystal lattice as compared to alkali metals.

### (4) Melting point and Boiling point

(i) Melting points and boiling points of alkaline earth metals do not show any regular trend.

	$Be$	$Mg$	$Ca$	$Sr$	$Ba$	$Ra$
melting points (K)	1560	920	1112	1041	1000	973
boiling point (K)	2770	1378	1767	1654	1413	—

(ii) The values are, however, more than alkali metals. This might be due to close packing of atoms in crystal lattice in alkaline earth metals.

### (5) Ionisation energy and electropositive or metallic character

(i) Since the atomic size decreases along the period and the nuclear charge increases and thus the electrons are more tightly held towards nucleus. It is therefore alkaline earth metals have higher ionisation energy in comparison to alkali metals but lower ionisation energies in comparison to p-block elements.

(ii) The ionisation energy of alkaline earth metals decreases from  $Be$  to  $Ba$ .

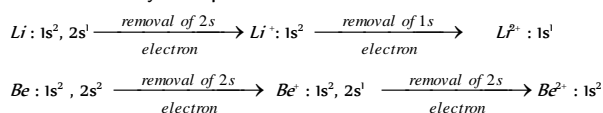
	$Be$	$Mg$	$Ca$	$Sr$	$Ba$	$Ra$
First ionisation energy ( $kJ mol^{-1}$ )	899	737	590	549	503	509
Second ionisation energy ( $kJ mol^{-1}$ )	1757	1450	1146	1064	965	979

(iii) The higher values of second ionisation energy is due to the fact that removal of one electron from the valence shell, the remaining electrons are more tightly held in which nucleus of cation and thus more energy is required to pull one more electron from monovalent cation.

(iv) No doubt first ionisation energy of alkaline earth metals are higher than alkali metals but a closer look on 2nd ionisation energy of alkali metals and alkaline earth metals reveals that 2nd ionisation energy of alkali metals are more

	Li	Be
1st ionisation energy ( $kJ\ mol^{-1}$ )	520	899
2nd ionisation energy ( $kJ\ mol^{-1}$ )	7296	1757

This may be explained as,



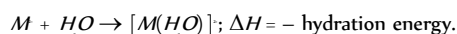
The removal of  $2s$  electron from alkali metals takes place from  $1s$  sub shell which are more closer to nucleus and exert more nuclear charge to hold up  $1s$  electron core, whereas removal of  $2nd$  electron from alkaline earth metals takes from  $2s$  sub shell. More closer are shells to the nucleus, more tightly are held electrons with nucleus and thus more energy is required to remove the electron.

(v) All these possess strong electropositive character which increases from  $Be$  to  $Ba$ .

(vi) These have less electropositive character than alkali metals as the later have low values of ionisation energy.

### (6) Oxidation number and valency

(i) The IE of the these metals are much lower than IE and thus it appears that these metals should form univalent ion rather than divalent ions but in actual practice, all these give divalent ions. This is due to the fact that  $M$  ion possesses a higher degree of hydration or  $M$  ions are extensively hydrated to form  $[M(H_2O)_n]^+$ , a hydrated ion. This involves a large amount of energy evolution which counter balances the higher value of second ionisation energy.



(ii) The tendency of these metals to exist as divalent cation can thus be accounted as,

(a) Divalent cation of these metals possess noble gas or stable configuration.

(b) The formation of divalent cation lattice leads to evolution of energy due to strong lattice structure of divalent cation which easily compensates for the higher values of second ionisation energy of these metals.

(c) The higher heats of hydration of divalent cation which accounts for the existence of the divalent ions of these metals in solution state.

### (7) Hydration of ions

(i) The hydration energies of alkaline earth metals divalent cation are much more than the hydration energy of monovalent cation.

	$Mg^{2+}$	$Mg^{2+}$
Hydration energy or Heat of hydration ( $kJ\ mol^{-1}$ )	353	1906

The abnormally higher values of heat of hydration for divalent cations of alkaline earth metals are responsible for their divalent nature.  $MgCl_2$  formation occurs with more amount of heat evolution and thus  $MgCl_2$  is more stable.

(ii) The hydration energies of  $M$  ion decreases with increase in ionic radii.

	$Be^{2+}$	$Mg^{2+}$	$Ca^{2+}$	$Sr^{2+}$	$Ba^{2+}$
Heat of hydration $kJ\ mol^{-1}$	2382	1906	1651	1484	1275

(iii) Heat of hydration are larger than alkali metals ions and thus alkaline earth metals compounds are more extensively hydrated than those

of alkali metals e.g  $MgCl_2$  and  $CaCl_2$  exists as  $MgCl_2 \cdot 6H_2O$  and  $CaCl_2 \cdot 6H_2O$  which  $NaCl$  and  $KCl$  do not form such hydrates.

(iv) The ionic mobility, therefore, increases from  $Be$  to  $Ba$ , as the size of hydrated ion decreases.

### (8) Electronegativities

(i) The electronegativities of alkaline earth metals are also small but are higher than alkali metals.

(ii) Electronegativity decreases from  $Be$  to  $Ba$  as shown below,

	Be	Mg	Ca	Sr	Ba
Electronegativity	1.57	1.31	1.00	0.95	0.89

(9) **Conduction power** : Good conductor of heat and electricity.

### (10) Standard oxidation potential and reducing properties

(i) The standard oxidation potential (in volts) are,

	Be	Mg	Ca	Sr	Ba
	1.69	2.35	2.87	2.89	2.90

(ii) All these metals possess tendency to lose two electrons to give  $M$  ion and are used as reducing agent.

(iii) The reducing character increases from  $Be$  to  $Ba$ , however, these are less powerful reducing agent than alkali metals.

(iv) Beryllium having relatively lower oxidation potential and thus does not liberate  $H_2$  from acids.

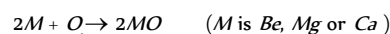
### (11) Characteristic flame colours

The characteristic flame colour shown are :  $Ca$  - brick red;  $Sr$  - crimson ;  $Ba$ -apple green and  $Ra$ - crimson.

### Chemical Properties

#### (1) Formation of oxides and hydroxides

(i) The elements (except  $Ba$  and  $Ra$ ) when burnt in air give oxides of ionic nature  $MO$  which are crystalline in nature.  $Ba$  and  $Ra$  however give peroxide. The tendency to form higher oxides increases from  $Be$  to  $Ra$ .



(ii) Their less reactivity than the alkali metals is evident by the fact that they are slowly oxidized on exposure to air, However the reactivity of these metals towards oxygen increases on moving down the group.

(iii) The oxides of these metals are very stable due to high lattice energy.

(iv) The oxides of the metal (except  $BeO$  and  $MgO$ ) dissolve in water to form basic hydroxides and evolve a large amount of heat.  $BeO$  and  $MgO$  possess high lattice energy and thus insoluble in water.

(v)  $BeO$  dissolves both in acid and alkalies to give salts i.e.  $BeO$  possesses amphoteric nature.



(vi) The basic nature of oxides of alkaline earth metals increases from  $Be$  to  $Ra$  as the electropositive Character increases from  $Be$  to  $Ra$ .

(vii) The tendency of these metal to react with water increases with increase in electropositive character i.e.  $Be$  to  $Ra$ .

(viii) Reaction of  $Be$  with water is not certain, magnesium reacts only with hot water, while other metals react with cold water but slowly and less energetically than alkali metals.

(ix) The inertness of  $Be$  and  $Mg$  towards water is due to the formation of protective, thin layer of hydroxide on the surface of the metals.

(x) The basic nature of hydroxides increase from  $Be$  to  $Ra$ . It is because of increase in ionic radius down the group which results in a decrease in strength of  $M-O$  bond in  $M-(OH)_2$ , to show more dissociation of hydroxides and greater basic character.

(xi) The solubility of hydroxides of alkaline earth metals is relatively less than their corresponding alkali metal hydroxides Furthermore, the



solubility of hydroxides of alkaline earth metals increases from *Be* to *Ba*. *Be(OH)* and *Mg(OH)* are almost insoluble, *Ca(OH)* (often called lime water) is sparingly soluble whereas *Sr(OH)* and *Ba(OH)* (often called baryta water) are more soluble.

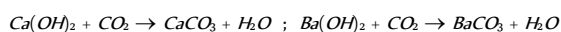
The trend of the solubility of these hydroxides depends on the values of lattice energy and hydration energy of these hydroxides. The magnitude of hydration energy remains almost same whereas lattice energy decreases appreciably down the group leading to more  $-Ve$  values for  $\Delta H_{hyd}$  down the group.

$$\Delta H_{hyd} = \Delta H_{lattice} + \Delta H_{hydration}$$

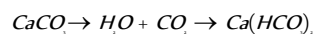
More negative is  $\Delta H_{hyd}$  more is solubility of compounds.

(xii) The basic character of oxides and hydroxides of alkaline earth metals is lesser than their corresponding alkali metal oxides and hydroxides.

(xiii) Aqueous solution of lime water [*Ca(OH)*]<sub>2</sub> or baryta water [*Ba(OH)*]<sub>2</sub> are used to qualitative identification and quantitative estimation of carbon dioxide, as both of them gives white precipitate with *CO*<sub>2</sub> due to formation of insoluble *CaCO*<sub>3</sub> or *BaCO*<sub>3</sub>.



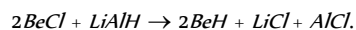
*SO*<sub>2</sub> also give white ppt of *CaSO* and *BaSO* on passing through lime water or baryta water. However on passing *CO*<sub>2</sub> in excess, the white turbidity of insoluble carbonates dissolve to give a clear solution again due to the formation of soluble bicarbonates,



## (2) Hydrides

(i) Except *Be*, all alkaline earth metals form hydrides (*MH*) on heating directly with *H*<sub>2</sub>.  $M + H_2 \rightarrow MH_2$ .

(ii) *BeH*<sub>2</sub> is prepared by the action of *LiAlH*<sub>4</sub> on *BeCl*<sub>2</sub>



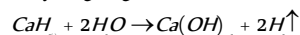
(iii) *BeH*<sub>2</sub> and *MgH*<sub>2</sub> are covalent while other hydrides are ionic.

(iv) The ionic hydrides of *Ca*, *Sr*, *Ba* liberate *H*<sub>2</sub> at anode and metal at cathode.



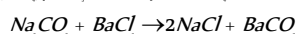
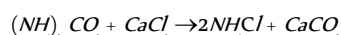
(v) The stability of hydrides decreases from *Be* to *Ba*.

(vi) The hydrides having higher reactivity for water, dissolves readily and produce hydrogen gas.

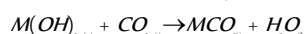


## (3) Carbonates and Bicarbonates

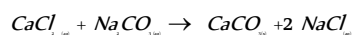
(i) All these metal carbonates (*MCO*) are insoluble in neutral medium but soluble in acid medium. These are precipitated by the addition of alkali metal or ammonium carbonate solution to the solution of these metals.



(ii) Alkaline earth metal carbonates are obtained as white precipitates when calculated amount of carbon dioxide is passed through the solution of the alkaline metal hydroxides.



and sodium or ammonium carbonate is added to the solution of the alkaline earth metal salt such as *CaCl*<sub>2</sub>.



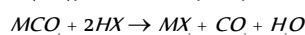
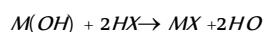
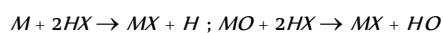
(iii) Solubility of carbonates of these metals also decreases downward in the group due to the decrease of hydration energy as the lattice energy remains almost unchanged as in case of sulphates.

(vi) The carbonates of these metals decompose on heating to give the oxides, the temperature of decomposition increasing from *Be* to *Ba*. Beryllium carbonate is unstable.

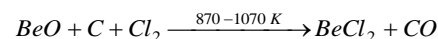


## (4) Halides

(i) The alkaline earth metals combine directly with halogens at appropriate temperatures forming halides, *MX*. These halides can also be prepared by the action of halogen acids (*HX*) on metals, metal oxides, hydroxides and carbonates.



Beryllium chloride is however, conveniently obtained from oxide



(ii) *BeCl*<sub>2</sub> is essentially covalent, the chlorides *MgCl*<sub>2</sub>, *CaCl*<sub>2</sub>, *SrCl*<sub>2</sub> and *BaCl*<sub>2</sub> are ionic; the ionic character increases as the size of the metal ion increases. The evidence is provided by the following facts,

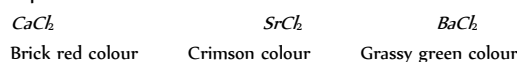
(a) Beryllium chloride is relatively low melting and volatile whereas *BaCl*<sub>2</sub> has high melting and stable.

(b) Beryllium chloride is soluble in organic solvents.

(iii) The halides of the members of this group are soluble in water and produce neutral solutions from which the hydrates such as *MgCl*<sub>2</sub>·6*H*<sub>2</sub>O, *CaCl*<sub>2</sub>·6*H*<sub>2</sub>O, *BaCl*<sub>2</sub>·2*H*<sub>2</sub>O can be crystallised. The tendency to form hydrated halides decreases with increasing size of the metal ions.

(iv) *BeCl*<sub>2</sub> is readily hydrolysed with water to form acid solution,  $BeCl_2 + 2H_2O \rightarrow Be(OH)_2 + 2HCl$ .

(v) The fluorides are relatively less soluble than the chlorides due to high lattice energies. Except *BeCl*<sub>2</sub> and *MgCl*<sub>2</sub> the chlorides of alkaline earth metals impart characteristic colours to flame.



**Structure of *BeCl*<sub>2</sub>** : In the solid phase polymeric chain structure with three centre two electron bonding with *Be-Cl-Be* bridged structure is shown below,



In the vapour phase it tends to form a chloro-bridged dimer which dissociates into the linear triatomic monomer at high temperature at nearly 1200 K.

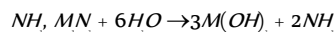
(5) **Solubility in liquid ammonia** : Like alkali metals, alkaline earth metals also dissolve in liquid ammonia to form coloured solutions. When such a solution is evaporated, hexammoniate, *M(NH)*<sub>6</sub>, is formed.

## (6) Nitrides

(i) All the alkaline earth metals direct combine with *N*<sub>2</sub> give nitrides, *MN*<sub>2</sub>.

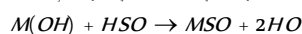
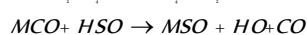
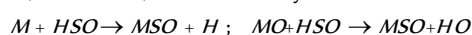
(ii) The ease of formation of nitrides however decreases from *Be* to *Ba*.

(iii) These nitrides are hydrolysed water to liberate



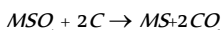
## (7) Sulphates

(i) All these form sulphate of the type *MSO*<sub>4</sub> by the action of *H*<sub>2</sub>*SO*<sub>4</sub> on metals, their oxides, carbonates or hydroxides.

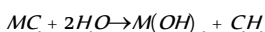


(ii) The solubility of sulphates in water decreases on moving down the group. *BeSO*<sub>4</sub> and *MgSO*<sub>4</sub> are fairly soluble in water while *BaSO*<sub>4</sub> is completely insoluble. This is due to increases in lattice energy of sulphates down the group which predominates over hydration energy.

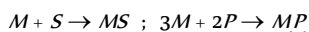
(iii) Sulphate are quite stable to heat however reduced to sulphide on heating with carbon.



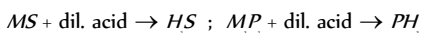
(8) **Action with carbon** : Alkaline metals (except *Be*, *Mg*) when heated with carbon form carbides of the type  $MC_2$ . These carbides are also called acetylides as on hydrolysis they evolve acetylene.



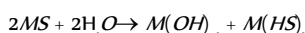
(9) **Action with sulphur and phosphorus** : Alkaline earth metals directly combine with sulphur and phosphorus when heated to form sulphides of the type  $MS$  and phosphides of the type  $MP_3$  respectively.



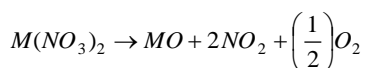
Sulphides on hydrolysis liberate  $H_2S$  while phosphides on hydrolysis evolve phosphine.



Sulphides are phosphorescent and are decomposed by water



(10) **Nitrates** : Nitrates of these metals are soluble in water. On heating they decompose into their corresponding oxides with evolution of a mixture of nitrogen dioxide and oxygen.



#### (11) Formation of complexes

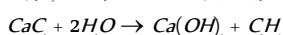
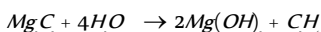
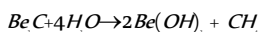
(i) Tendency to show complex ion formation depends upon smaller size, high nuclear charge and vacant orbitals to accept electron. Since alkaline metals too do not possess these characteristics and thus are unable to form complex ion.

(ii) However, *Be* on account of smaller size forms many complex such as  $(BeF)_2$ ,  $(BeF)_4^{2-}$ .

#### Anomalous behaviour of Beryllium

Beryllium differs from rest of the alkaline earth metals on account of its small atomic size, high electronegativity *Be* exerts high polarizing effect on anions and thus produces covalent nature in its compounds. Following are some noteworthy difference of *Be* from other alkaline earth metals,

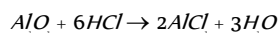
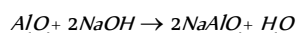
- (1) *Be* is lightest alkaline earth metal.
- (2) *Be* possesses higher m.pt. and b.pt than other group members.
- (3) *BeO* is amphoteric in nature whereas oxides of other group members are strong base.
- (4) It is not easily effected by dry air and does not decompose water at ordinary temperature.
- (5) *BeSO\_4* is soluble in water.
- (6) *Be* and *Mg* carbonates are not precipitated by  $(NH_4)_2CO_3$  in presence of  $NH_4Cl$ .
- (7) *Be* and *Mg* salts do not impart colour to flame.
- (8) *Be* does not form peroxide like other alkaline earth metals.
- (9) It does not evolve hydrogen so readily from acids as other alkaline earth metals do so.
- (10) It has strong tendency to form complex compounds.
- (11)  $BeN_2$  is volatile whereas nitrides of other alkaline earth metals are non-volatile.
- (12) It's salts can never have more than four molecules of water of crystallization as it has only four available orbitals in its valence shell.
- (13) Beryllium carbide reacts water to give methane whereas magnesium carbide and calcium carbide give propyne and acetylene respectively.



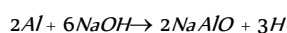
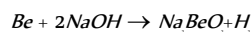
#### Diagonal relationship of *Be* with *Al*

Due to its small size *Be* differs from other earth alkaline earth metals but resembles in many of its properties with *Al* on account of diagonal relationship.

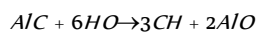
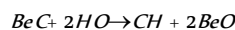
- (1) *Be* and *Al* have almost same and smaller size and thus favour for covalent bonding.
- (2) Both these form covalent compounds having low m. pt and soluble in organic solvent.
- (3) Both have same value of electronegativity (i.e. 1.5).
- (4) The standard O.P of these elements are quite close to each other ; *Be*=1.69 volts and *Al*= 1.70 volts.
- (5) Both become passive on treating with conc.  $HNO_3$  in cold.
- (6) Both form many stable complexes e.g.  $(BeF)_4^{2-}$ ,  $(AlH)_4^-$ .
- (7) Like *BeO*,  $Al_2O_3$  is amphoteric in nature. Also both are high melting point solids.



(8) *Be* and *Al* both react with  $NaOH$  to liberate  $H_2$  forming beryllates and alluminates.



(9) *Be*, *C* and  $AlCl_3$  both give  $CH_4$  on treating with water.



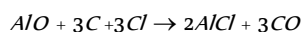
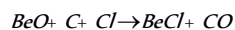
(10) Both occur together in nature in beryl ore,  $3BeO \cdot Al_2O_3 \cdot 6SiO_2$ .

(11) Unlike other alkaline earths but like aluminium, beryllium is not easily attacked by air (Also *Mg* is not attacked by air)

(12) Both *Be* and *Al* react very slowly with dil.  $HCl$  to liberate  $H_2$ .

(13) Both *Be* and *Al* form polymeric covalent hydrides while hydrides of other alkaline earth are ionic.

(14) Both  $BeCl_2$  and  $AlCl_3$  are prepared in similar way.



(15) Both  $BeCl_2$  and  $AlCl_3$  are soluble in organic solvents and act as catalyst in Friedel-Crafts reaction.

(16) Both  $Be(OH)_2$  and  $Al(OH)_3$  are amphoteric whereas hydroxides of other alkaline earths are strong alkali.

(17) The salts of *Be* and *Al* are extensively hydrated.

(18)  $BeCl_2$  and  $AlCl_3$  both have a bridged polymeric structure.

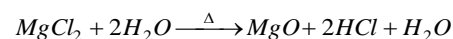
(19) *Be* and *Al* both form fluoro complex ions  $[BeF_4]^{2-}$  and  $[AlF_6]^{3-}$  in solution state whereas other members of 2nd group do not form such complexes.

#### Magnesium and its compounds

(1) **Ores of magnesium** : Magnesite ( $MgCO_3$ ), Dolomite ( $MgCO_3 \cdot CaCO_3$ ), Epsomite (epsom salt) ( $MgSO_4 \cdot 7H_2O$ ) Carnallite ( $MgCl_2 \cdot KCl \cdot 6H_2O$ ) Asbestos ( $CaMg_3(SiO_3)_4$ ), Talc ( $Mg_3(Si_2O_5)_2 \cdot Mg(OH)_2$ ).

(2) **Extraction of magnesium** : It is prepared by the electrolysis of fused magnesium chloride which in turn is obtained from carnallite and magnesite.

Carnallite ( $MgCl_2 \cdot KCl \cdot 6H_2O$ ) can't be directly converted into anhydrous  $MgCl_2$  by heating because all the water of crystallisation cannot be removed by heating. Moreover, strong heating may change it to  $MgO$ .



In Dow's process, magnesium chloride is obtained from sea water as  $MgCl_2 \cdot 6H_2O$ . It is rendered anhydrous by heating it in a current of dry

$HCl$  gas. The anhydrous magnesium chloride is fused with  $NaCl$  (to provide conductivity to the electrolyte and to lower the fusing temperature of anhydrous  $MgCl_2$ ) and then electrolysed at  $700^\circ C$ .

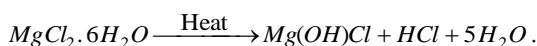
### (3) Compounds of magnesium

(i) **Magnesia ( $MgO$ )** : It is used as magnesia cement. It is a mixture of  $MgO$  and  $MgCl_2$ . It is also called Sorel's cement.

(ii) **Magnesium hydroxide** : Its aqueous suspension is used in Medicine as an antacid. Its medicinal name is milk of magnesia.

(iii) **Magnesium sulphate or Epsom salt ( $MgSO_4 \cdot 7H_2O$ )** : It is isomorphous with  $ZnSO_4 \cdot 7H_2O$ . It is used as a purgative in medicine, as a mordant in dyeing and as a stimulant to increase the secretion of bile.

(iv) **Magnesium chloride ( $MgCl_2 \cdot 6H_2O$ )** : It is a deliquescent solid. Hydrated salt on heating in air undergoes partial hydrolysis.



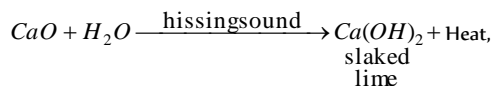
### Calcium and its compounds

(1) **Ores of calcium** : Lime stone or marble or chalk ( $CaCO_3$ ), Gypsum ( $CaSO_4 \cdot 2H_2O$ ), Dolomite ( $CaCO_3 \cdot MgCO_3$ ), Fluorspar ( $CaF_2$ ), phosphorite  $Ca_3(PO_4)_2$ . Calcium phosphate is a constituent of bones and teeth.

(2) **Manufacture** : It is manufactured by the electrolysis of a molten mixture of calcium chloride containing some calcium fluoride. Calcium chloride is obtained as a by product of the solvay process.

### (3) Compounds of calcium

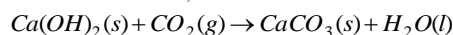
(i) **Calcium oxide or Quick lime or Burnt lime ( $CaO$ )** : Its aqueous suspension is known as slaked lime.



When exposed to oxy-hydrogen flame, it starts emitting light called lime light.

$CaO$  is used as basic flux, for removing hardness of water, as a drying agent (for  $NH_3$  gas) for preparing mortar ( $CaO$  + sand + water).

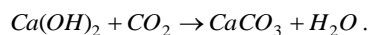
**Mortar** : Mortar used in making buildings is a mixture of lime ( $CaO$ ) and sand in the ratio 1 : 3 with enough water to make a thick paste. When the mortar is placed between bricks, it slowly absorbs  $CO_2$  from the air and the slaked lime reverts to  $CaCO_3$ .



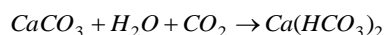
Although the sand in the mortar is chemically inert, the grains are bound together by the particles of calcium carbonate and a hard material results.

(ii) **Calcium chloride ( $CaCl_2 \cdot 6H_2O$ )** : Fused  $CaCl_2$  is a good desiccant (drying agent). It can't be used to dry alcohol or ammonia as it forms additional products with them.

(iii) **Calcium carbonate ( $CaCO_3$ )** :

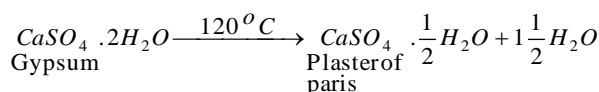


It is insoluble in water but dissolves in the presence of  $CO_2$  due to the formation of calcium bicarbonate.

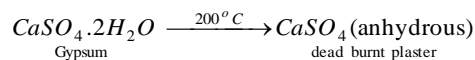
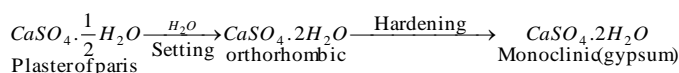


It is a constituent of protective shells of marine animals.

(iv) **Gypsum ( $CaSO_4 \cdot 2H_2O$ )** : On partially dehydrates to produce plaster of paris.

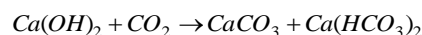
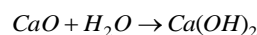


### Plaster of paris :

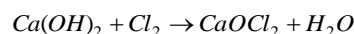


Gypsum when heated to about  $200^\circ C$  is converted into anhydrous calcium sulphate. The anhydrous form (anhydrite) is known as dead burnt plaster because it does not set like plaster of paris when moistened with water.

(v) **Calcium Hydroxide  $Ca(OH)_2$**  (slaked lime)



Suspension of  $Ca(OH)_2$  in water is called milk of lime.



(vi) **Cement** : (a) It is essentially a mixture of lime stone and clay. It is also called Portland cement because in presence of water it sets to a hard stone-like mass resembling with the famous Portland rock, a famous building stone of England. The approximate composition of cement is

Calcium oxide ( $CaO$ )	50 – 60 %
Silica ( $SiO_2$ )	20 – 25%
Alumina ( $Al_2O_3$ )	5 – 10%
Magnesia ( $MgO$ )	1 – 3%
Ferric oxide ( $Fe_2O_3$ )	1 – 3%

The above compounds are provided by the two raw materials, namely lime stone (which provides  $CaO$ ) and clay which provides  $SiO_2$ ,  $Al_2O_3$  and  $Fe_2O_3$ . In cement, almost entire amount of lime is present in the combined state as calcium silicates ( $2CaO \cdot SiO_2$  and  $3CaO \cdot SiO_2$ ) and calcium aluminates ( $3CaO \cdot Al_2O_3$  and  $4CaO \cdot Al_2O_3$ ).

(b) Cement containing excess amount of lime cracks during setting; while cement containing less amount of lime is weak in strength.

(c) Cement with excess of silica is slow-setting and that having an excess of alumina is quick-setting.

(d) Cement containing no iron oxide is white but hard to burn.

Cement is manufactured by two processes, viz, wet and dry. A small amount (2–3%) of gypsum is added to slow down the setting of the cement so that it gets sufficiently hardened. Setting of cement is an exothermic process and involves hydration of calcium aluminates and calcium silicates.

## Boron Family

Group 13 of long form of periodic table (previously reported as group III A according to Mendeleev's periodic table) includes boron ( $B$ ); aluminium ( $Al$ ), gallium ( $Ga$ ), indium ( $In$ ) and thallium ( $Tl$ ). Boron is the first member of group 13 of the periodic table and is the only non-metal of this group. The all other members are metals. The non-metallic nature of boron is due its small size and high ionisation energy. The members of this family are collectively known as boron family and sometimes as aluminium family.

### (1) Electronic configuration

Element	Electronic configuration ( $ns^2 np^1$ )
${}_5B$	$[He]2s^2 2p^1$
${}_{13}Al$	$[Ne]3s^2 3p^1$
${}_{31}Ga$	$[Ar]3d^{10} 4s^2 4p^1$
${}_{49}In$	$[Kr]4d^{10} 5s^2 5p^1$
${}_{81}Tl$	$[Xe]4f^{14} 5d^{10} 6s^2 6p^1$

(2) **Occurrence** : The important of this group elements are given below,

**Boron** : Borax (Tincal) ( $Na_2B_4O_7 \cdot 10H_2O$ ), Colemanite ( $Ca_2B_6O_{11} \cdot 5H_2O$ )

Boracite ( $2Mg_3B_2O_7 \cdot MgCl_2$ ), Boronatro calcite ( $CaB_2O_7 \cdot NaBO_2 \cdot 8H_2O$ ),

Kernite ( $Na_2B_4O_7 \cdot 4H_2O$ ), Boric acid ( $H_3BO_3$ )

**Aluminium** : Corundum ( $Al_2O_3$ ), Diaspore ( $Al_2O_3 \cdot H_2O$ ), Bauxite ( $Al_2O_3 \cdot 2H_2O$ ), and Cryolite ( $Na_3AlF_6$ ).

#### Physical properties

(1) A regular increasing trend in density down the group is due to increase in size.

(2) Melting points do not vary regularly and decrease from  $B$  to  $Ga$  and then increase.

(3) Boron has very high melting point because it exist as giant covalent polymer in both solid and liquid state.

(4) Low melting point of  $Ga$  ( $29.8^\circ C$ ) is due to the fact that consists of only  $Ga$  molecule; it exist as liquid upto  $2000^\circ C$  and hence used in high temperature thermometry.

(5) Boiling point of these elements however show a regular decrease down the group.

(6) The abrupt increase in the atomic radius of  $Al$  is due to greater screening effect in  $Al$  (it has 8 electrons in its penultimate shell) than in  $B$  (it has 2 electrons in its penultimate shell)

(7) The atomic radii of group 13 elements are smaller than the corresponding  $s$ -block elements. This is due to the fact that when we move along the period, the new incoming electron occupy the same shell whereas the nuclear charge increases regularly showing more effective pull of nucleus towards shell electrons. This ultimately reduces the atomic size.

(8) The atomic radius of  $Ga$  is slightly lesser than of  $Al$  because in going from  $Al$  to  $Ga$ , the electrons have already occupied  $3d$  sub shell in  $Ga$ . The screening effect of these intervening electrons being poor and has less influence to decrease the effective nuclear charge, therefore the electrons in  $Ga$  experience more forces of attractions towards nucleus to result in lower size of  $Ga$  than  $Al$

#### (9) Oxidation state

(i) All exhibit +3 oxidation state and thus complete their octet either by covalent or ionic union.

(ii) Boron being smaller in size cannot lose its valence electrons to form  $B^3+$  ion and it usually show +3 covalence. The tendency to show +3 covalence however decreases down the group even  $Al$  shows +3 covalence in most of its compounds.

(iii) Lower elements also show +1 ionic state e.g.  $Tl$ ,  $Ga$ . This is due to inert pair effect. The phenomenon in which outer shell ' $s$ ' electrons ( $ns$ ) penetrate to ( $n-1$ )  $d$ -electrons and thus become closer to nucleus and are

more effectively pulled the nucleus. This results in less availability of  $ns$  electrons pair for bonding or  $ns$  electron pair becomes inert. The inert pair effect begins after  $n \geq 4$  and increases with increasing value of  $n$ .

(iv) The tendency to form  $M^+$  ion increases down the gp.  $Ga < Tl$

(10) **Hydrated ions** : All metal ions exist in hydrated state.

#### (11) Ionisation energy

(i) In spite of the more charge in nucleus and small size, the first ionisation energies of this group elements are lesser than the corresponding elements of  $s$  block. This is due to the fact that removal of electron from a  $p$ -orbitals (being far away from nucleus and thus less effectively held than  $s$ -orbitals) is relatively easier than  $s$ -orbitals.

(ii) The ionisation energy of this group element decrease down the group due to increases in size like other group elements.

(iii) However, ionisation energy of  $Ga$  are higher than that of  $Al$  because of smaller atomic size of  $Ga$  due to less effective shielding of  $3d$  electrons in  $Ga$ . Thus valence shell exert more effective nuclear charge in  $Ga$  to show higher ionisation energies.

#### (12) Electropositive character

(i) Electropositive character increases from  $B$  to  $Tl$ .

(ii) Boron is semi metal, more closer to non-metallic nature whereas rest all members are pure metals.

(iii) Furthermore, these elements are less electropositive than  $s$ -block elements because of smaller size and higher ionisation energies.

#### (13) Oxidation potential

(i) The standard oxidation potentials of these element are quite high and are given below,

	$B$	$Al$	$Ga$	$In$	$Tl$
$E^{\circ}_{op}$ for $M \rightarrow M^{3+} + 3e$	-	+1.66	+0.56	+0.34	+1.26
$E^{\circ}_{op}$ for $M \rightarrow M^+ + e$	+0.55	-	+0.18	+0.34	

(ii) However Boron does not form positive ions in aqueous solution and has very low oxidation potential.

(iii) The higher values of standard oxidation potentials are due to higher heats of hydration on account of smaller size of trivalent cations.

(iv) Aluminium is a strong reducing agent and can reduce oxides which are not reduced even by carbon. This is due to lower ionisation energy of aluminium than carbon. The reducing character of these elements is  $Al > Ga > In > Tl$ .

(14) **Complex formation** : On account of their smaller size and more effective nuclear charge as well as vacant orbitals to accept elements, these elements have more tendency to form complexes than  $s$  block elements.

#### Chemical properties

##### (1) Hydrides

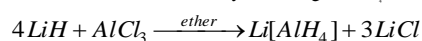
(i) Elements of group 13 do not react directly with hydrogen but a number of polymeric hydrides are known to exist.

(ii) Boron forms a large no. of volatile covalent hydrides, known as boranes e.g.  $BH_3$ ,  $B_2H_6$ ,  $B_4H_{10}$ ,  $B_{10}H_{12}$ . Two series of boranes with general formula  $B_nH_n$  and  $B_nH_{n+2}$  are more important.

(iii) Boranes are electron deficient compounds. It is important to note that although  $BX_3$  are well known,  $BH_3$  is not known. This is due of the fact that hydrogen atoms in  $BH_3$  have no free electrons to form  $p\pi-p\pi$  back bonding and thus boron has incomplete octet and hence  $BH_3$  molecules dimerise to form  $B_2H_6$  having covalent and three centre bonds.

(iv)  $Al$  forms only one polymeric hydride ( $AlH_3$ ) commonly known as alane. It contains  $Al \cdots H \cdots Al$  bridges.

(v)  $Al$  and  $Ga$  forms anionic hydrides e.g.  $LiAlH_4$  and  $LiGaH_4$ ,



##### (2) Reactivity towards air

(i) Pure boron is almost unreactive at ordinary temperature. It reacts with air to form  $B_2O_3$  when heated. It does not react with water. Al burns in air with evolution of heat to give  $Al_2O_3$ .

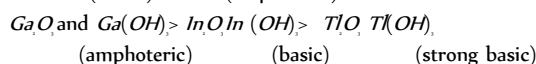
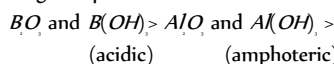
(ii)  $Ga$  and  $In$  are not affected by air even when heated whereas  $Tl$  is little more reactive and also forms an oxide film at surface. In moist air, a layer of  $Tl(OH)_3$  is formed.

(iii)  $Al$  decomposes  $H_2O$  and reacts readily in air at ordinary temperature to form a protective film of its oxides which protects it from further action.

### (3) Oxides and hydroxides

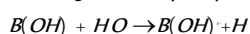
(i) The members of boron family form oxide and hydroxides of the general formula  $MO_2$  and  $M(OH)_3$ , respectively.

(ii) The acidic nature of oxides and hydroxides changes from acidic to basic through amphoteric from  $B$  to  $Tl$ .

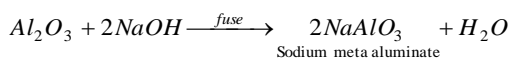
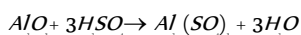


$B(OH)_3$  or  $H_3BO_3$  is weak monobasic Lewis acid.

(iii) Boric acid,  $B(OH)_3$ , is soluble in water as it accepts lone pair of electron to act as Lewis acid. Rest all hydroxides of group 13 are insoluble in water and form a gelatinous precipitate.



(iv)  $Al_2O_3$  being amphoteric dissolves in acid and alkalis both.



(v) One of the crystalline forms of alumina ( $Al_2O_3$ ) is called corundum. It is very hard and used as an abrasive. It is prepared by heating the amorphous form of  $Al_2O_3$  to 2000 K.

### (4) Action of Acids

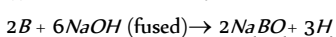
(i) Boron does not react with non-oxidizing acids, however, it dissolves in nitric acid to form boric acids.

(ii)  $Al$ ,  $Ga$  and  $In$  dissolve in acids forming their trivalent cations; however,  $Al$  and  $Ga$  become passive due to the formation of a protective film of oxides.

(iii) Thallium dissolves in acids forming univalent cation and becomes passive in  $HCl$  due to the formation of water-insoluble  $TlCl$ .

### (5) Action of Alkalies

(i) Boron dissolves only in fused alkalis,

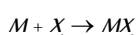


(ii)  $Al$  and  $Ga$  dissolve in fused as well as in aqueous alkalis,  $2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2$

(iii) Indium remains unaffected in alkalis even on heating.

### (6) Halides

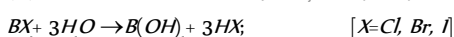
(i) All the group 13 elements from the trihalides,  $MX_3$ , on directly combining with halogens.



(ii) All the trihalides of group 13 elements are known except  $Tl(III)$  iodide.

(iii) Due to small size and high electronegativity of boron, all boron halides are covalent and Lewis acids. These exist as monomeric molecules having plane triangular geometry ( $sp^2$  hybridization).

(iv) All boron trihalides except  $BF_3$  are hydrolysed to boric acid.



However,  $BF_3$  forms an addition product with water,



$BF_3$  having less tendency for hydrolysis as well as Lewis acid nature, is extensively used as a catalyst in organic reactions e.g. Friedel-Crafts reaction.

(v) Boron atom, in  $BX_3$ , has six electrons in the outermost orbit and thus it can accept a pair of electrons from a donor molecule like  $NH_3$  to complete its octet. Hence boron halides act as very efficient Lewis acids. The relative Lewis acid character of boron trihalides is found to obey the order;  $BI_3 > BBr_3 > BCl_3 > BF_3$ .

However, the above order is just the reverse of normally expected order on the basis of relative electronegativities of the halogens. Fluorine, being the most electronegative, should create the greatest electron deficiency on boron and thus  $B$  in  $BF_3$  should accept electron pair from a donor very rapidly than in other boron trihalides. But this is not true.

This anomalous behaviour has been explained on the basis of the relative tendency of the halogen atom to back-donate its unutilised electrons to the vacant  $p$  orbitals of boron atom. In boron trifluoride, each fluorine has completely filled unutilised  $2p$  orbitals while boron has a vacant  $2p$  orbital. Now since both of these orbitals belong to the same energy level ( $2p$ ) they can overlap effectively as a result of which fluorine electrons are transferred into the vacant  $2p$  orbital of boron resulting in the formation of an additional  $p\pi-p\pi$  bond. This type of bond formation is known as **back bonding** or **back donation**. Thus the  $B-F$  bond has some double bond character. Back bonding may take place between boron and one of the three fluorine atoms and thus boron trifluoride is regarded as a resonance hybrid of some structures.

Resonance in boron trifluoride is also evidenced by the fact that the three boron-fluorine bonds are *identical* and are shorter than the usual single boron-fluorine bond. As a result of back bonding, the electron deficiency of boron is reduced and hence Lewis acid nature is decreased. The tendency for the formation of back bonding ( $p\pi-p\pi$  bond) is maximum in  $BF_3$  and decreases very rapidly from  $BF_3$  to  $BI_3$ . This is probably due to the fact that overlapping of the vacant  $2p$  orbitals of boron cannot take place easily with the  $p$ -orbitals of high energy levels ( $3p$  in  $Cl$ ,  $4p$  in  $Br$  and  $5p$  in iodine). Thus  $BI_3$  and  $BCl_3$  are stronger Lewis acids than the  $BF_3$ .

(vi) Lewis acid character of halides of the group 13 elements decreases in the order,  $B > Al > Ga > In$ .

(vii) Boron halides form complex halides of the type,  $[BF_4]^-$ , in which boron atom extends its coordination number to four by utilising empty  $p$ -orbital. It cannot extend its coordination number beyond four due to non-availability of  $d$ -orbitals. However, the other trihalides of this group form complex halides of the type  $(AlF_6)^-$ ,  $(GaCl_6)^-$  and  $(InCl_6)^-$ , etc where the central atom extends its coordination number to 6 by the use of  $d$ -orbitals.

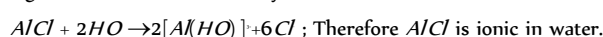
(viii) The fluorides of  $Al$ ,  $Ga$ ,  $In$  and  $Tl$  are ionic and have high melting points. The high melting points of metal fluorides can be explained on the basis that their cations are sufficiently large and have vacant  $d$ -orbitals for attaining a coordination number of six towards the relatively small fluorine atom.

(ix) Other halides of  $Al$ ,  $Ga$ ,  $In$  and  $Tl$  are largely covalent in anhydrous state and possess low melting points. These halides do not show backbonding because of increases in the size of the element. However, they make use of vacant  $p$ -orbitals by co-ordinate bond *i.e.* metal atoms complete their octet by forming dimers. Thus aluminium chloride, aluminium bromide and indium iodide exist as dimers, both in the vapour state and in non-polar solvents.

The dimer structure for  $Al_2Cl_6$  is evidenced by the following facts,

(a) Vapour density of aluminium chloride measured at 400°C corresponds to the formula  $Al_2Cl_6$ .

(b) Bond distance between aluminium-chlorine bond forming bridge is greater (2.21 Å) than the distance between aluminium-chlorine bond present in the end (2.06 Å). The dimeric structure disappears when the halides are dissolved in water. This is due to high heat of hydration which splits the dimeric structure into  $[M(H_2O)_6]^{3+}$  and  $3X^-$  ions and the solution becomes a good conductor of electricity.



The dimeric structure may also split by reaction with donor molecules e.g.  $R_3N$ . This is due to the formation of complexes of the type  $R_3NAlCl_2$ . The dimeric structure of  $AlCl_3$  exist in vapour state below  $473K$  and at higher temperature it dissociates to trigonal planar  $AlCl_3$  molecule.

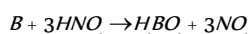
Boron halides do not exist as dimer due to small size of boron atom which makes it unable to co-ordinate four large-sized halide ions.

(x)  $BF_3$  and  $AlCl_3$  acts as catalyst and Lewis acid in many of the industrial process.

#### Anomalous Behaviour of Boron

Like  $Li$  and  $Be$ , Boron – the first member of group 13 also shows anomalous behaviour due to extremely low size and high nuclear charge/size ratio, high electronegativity and non-availability of d electrons. The main point of differences are,

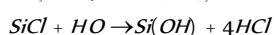
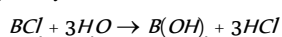
- (1) Boron is a typical non-metal whereas other members are metals.
- (2) Boron is a bad conductor of electricity whereas other metals are good conductors.
- (3) Boron shows allotropy and exists in two forms – crystalline and amorphous. Aluminium is a soft metal and does not exist in different forms.
- (4) Like other non-metals, the melting point and boiling point of boron are much higher than those of other elements of group 13.
- (5) Boron forms only covalent compounds whereas aluminium and other elements of group 13 form even some ionic compounds.
- (6) The hydroxides and oxides of boron are acidic in nature whereas those of others are amphoteric and basic.
- (7) The trihalides of boron ( $BX_3$ ) exist as monomers. On the other hand, aluminium halides exist as dimers ( $Al_2X_6$ ).
- (8) The hydrides of boron i.e. boranes are quite stable while those of aluminium are unstable.
- (9) Dilute acids have no action on boron. Others liberate  $H_2$  from them.
- (10) Borates are more stable than aluminates.
- (11) Boron exhibit maximum covalency of four e.g.,  $BH_4^-$  ion while other members exhibit a maximum covalency of six e.g.,  $[Al(OH)_6]^-$ .
- (12) Boron does not decompose steam while other members do so.
- (13) Boron combines with metals to give borides e.g.  $Mg_2B$ . Other members form simply alloys.
- (14) Concentrated nitric acid oxidises boron to boric acid but no such action is noticed other group members.



#### Diagonal relationship between Boron and Silicon

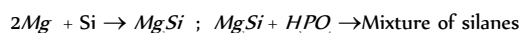
Due to its small size and similar charge/mass ratio, boron differs from other group 13 members, but it resembles closely with silicon, the second element of group 14 to exhibit diagonal relationship. Some important similarities between boron and silicon are given below,

- (1) Both boron and silicon are typical non-metals, having high m.pt. b.pt nearly same densities ( $B=2.35g/ml$   $Si=2.34 g/ml$ ). low atomic volumes and bad conductor of current. However both are used as semiconductors.
- (2) Both of them do not form cation and form only covalent compounds.
- (3) Both exists in amorphous and crystalline state and exhibit allotropy.
- (4) Both possess closer electronegativity values ( $B=2.0$ ;  $Si=1.8$ ).
- (5) Both form numerous volatile hydrides which spontaneously catch fire on exposure to air and are easily hydrolysed.
- (6) The chlorides of both are liquid, fume in moist air and readily hydrolysed by water.



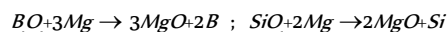
(7) Both form weak acids like  $H_3BO_3$  and  $H_4SiO_4$ .

(8) Both form binary compounds with several metals to give borides and silicides. These borides and silicide react with  $H_2PO_4$  to give mixture of boranes and silanes.

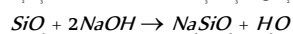
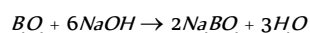
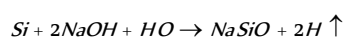
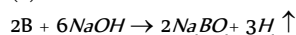


(9) The carbides of both Boron and silicon ( $BC$  and  $SiC$ ) are very hard and used as abrasive.

(10) Oxides of both are acidic and can be reduced by limited amount of  $Mg$ . In excess of  $Mg$  boride and silicide are formed.



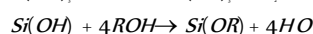
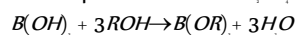
(11) Both the metals and their oxides are readily soluble in alkalis.



Both borates and silicates have tetrahedral structural units

$BO_4^{n-}$  and  $SiO_4^{n-}$  respectively. Boro silicates are known in which boron replaces silicon in the three dimensional lattice. Boron can however form planar  $BO$  units.

(12) Acids of both these elements form volatile esters on heating with alcohol in presence of conc.  $H_2SO_4$ .



#### Boron and its compounds

Boron is the first member of group -13 (IIIA) of the periodic table. Boron is a non-metal. It has a small size and high ionization energy due to which it can not lose its valence electrons to form  $B^{+3}$  ion. Its compounds especially the hydrides and halides are electron deficient and behave as Lewis acid.

##### (1) Ores of boron

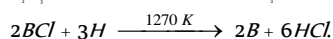
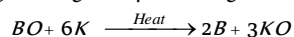
(i) *Borax or tincal*:  $Na_2B_4O_7 \cdot 10H_2O$

(ii) *Kernite or Rasorite*:  $Na_2B_4O_7 \cdot 4H_2O$

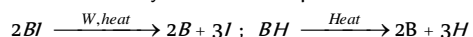
(iii) *Colemanite*:  $Ca_2B_6O_{11} \cdot 5H_2O$

(iv) *Orthoboric acid*:  $H_3BO_3$  (It occurs in the jets of steam called soffioni escaping from ground in the volcanic region of the *Tuscany*). Boron is present to a very small extent (0.001%) in earth's crust.

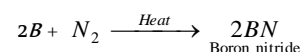
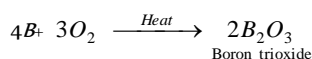
(2) **Isolation**: Elemental boron in the form of dark brown powder is obtained either by reduction of boric oxide with highly electropositive metals like  $K$ ,  $Mg$ ,  $Al$ ,  $Na$ , etc. in the absence of air and boron halides with hydrogen at high temperature eg.

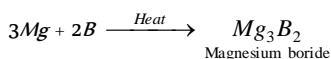
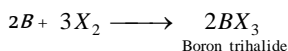


By thermal decomposition of boron triiodide over red hot tungsten filament and boron hydrides for example,

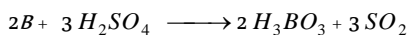
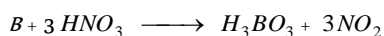


(3) **Properties**: It exists in mainly two allotropic forms i.e. amorphous dark brown powder and crystalline black very hard solid. It occurs in two isotopic forms, i.e.,  $_{5}B^{10}$  (20% abundance) and  $_{5}B^{11}$  (80% abundance). With air, boron forms  $B_2O_3$  and  $BN$  at  $973K$ , with halogens, trihalides ( $BX_3$ ) are formed, with metals borides are formed. eg.

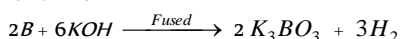




Water, steam and  $HCl$  have no action on  $B$ . oxidising acids ( $HNO_3$ ,  $H_2SO_4$ ) convert boron to  $H_3BO_3$ .



Fused alkalis ( $NaOH$ ,  $KOH$ ) dissolve boron forming borates, liberating hydrogen.



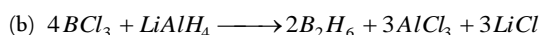
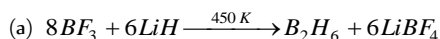
(4) **Uses of Boron** : Boron is used in atomic reactors as protective shields and control rods, as a semiconductors for making electronic devices in steel industry for increasing the hardness of steel and in making light composite materials for air crafts.

#### (5) Compounds of Boron

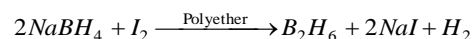
##### (i) Boron Hydrides

Boron forms hydrides of the types  $B_nH_{n+4}$  and  $B_nH_{n+6}$  called boranes. Diborane is the simplest boron hydride which is a dimer of  $BH_3$ .

##### Preparation

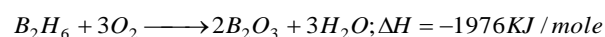


(c) In the laboratory, it is prepared by the oxidation of sod. Borohydride with  $I_2$ .

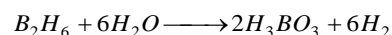


**Properties** : (a) Since Boron in boranes never complete its octet of electrons hence all boranes are called as electron-deficient compounds or Lewis acids.

(b) All boranes catch fire in the presence of oxygen to liberated a lot of heat energy. Thus, they can also be used as high energy fuels.



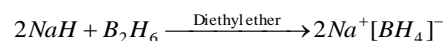
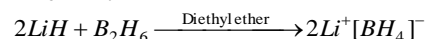
(c) Boranes are readily hydrolysed by water.



(d) With carbon monoxide



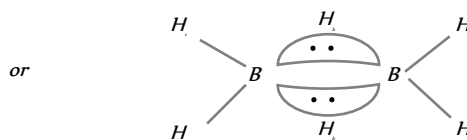
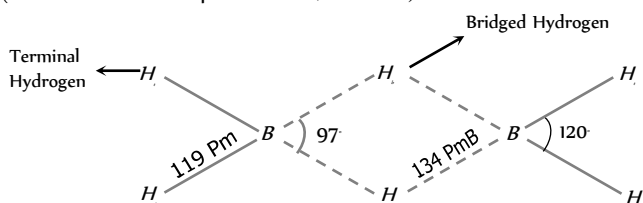
(e) Boranes are used for formation of hydroborates or borohydrides such as  $LiBH_4$  or  $NaBH_4$ , which are extensively used as reducing agents in organic synthesis.



**Structure of diborane** :  $B_2H_6$  has a three centre electron pair bond also called a banana shape bond.

(a)  $B-H_t$  : It is a normal covalent bond (two centre electron pair bond i.e.,  $2c-2e$ ).

(b)  $B-H_b$  : This is a bond between three atoms,  $B-H_b-B$ , (three centre electron pair bond i.e.,  $3c-2e$ ).

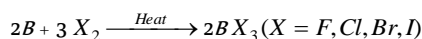


#### Structure of diborane ( $B_2H_6$ )

The other boron hydrides are  $B_5H_9$ ,  $B_4H_{10}$ ,  $B_5H_{11}$  etc.

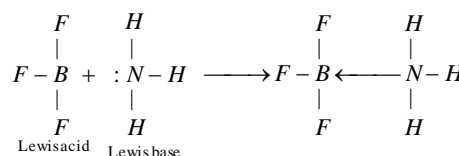
##### (ii) Boron Halides

Boron reacts with halogens on strong heating to form boron halides



$BF_3$  and  $BCl_3$  are gases,  $BBR_3$  is a volatile liquid while  $BI_3$  is a solid.

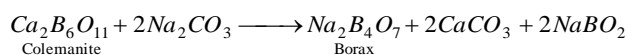
In these halides, the central boron atom has three shared pairs of electrons with the halogen atoms. Therefore, these have two electrons less than the octet and are electron deficient compounds. They acts as Lewis acids.



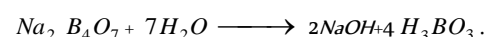
The relative acidic strength of boron trihalides decreases as :  $BI_3 > BBr_3 > BCl_3 > BF_3$ .

##### (iii) Borax ( $Na_2B_4O_7 \cdot 10H_2O$ )

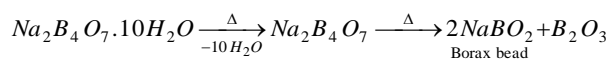
It occurs naturally as **tincal** (Suhaga) which contains about 50% borax in certain land, lakes. It is also obtained from the mineral colemanite by boiling it with a solution of  $Na_2CO_3$ .



**Properties** : (a) Its aqueous solution is alkaline due to hydrolysis,



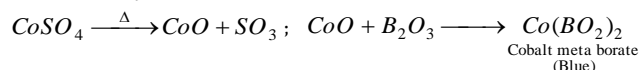
(b) On heating borax loses its water of crystallization and swells up to form a fluffy mass. On further heating, it melts to give a clear liquid which solidifies to a transparent glassy bead consisting of sodium metaborate ( $NaBO_2$ ) and boric anhydride ( $B_2O_3$ ),



**Borax bead** is used for the detection of coloured basic radicals under the name borax bead test.

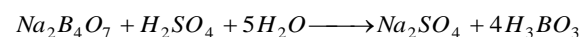
**Borax bead test** : Borax bead is a mixture of  $NaBO_2$  and  $B_2O_3$ .

$B_2O_3$  on heating combines readily with a number of coloured transition metal oxides such as  $Co$ ,  $Ni$ ,  $Cr$ ,  $Cu$ ,  $Mn$ , etc. to form the corresponding metaborates which possess characteristic colours,



**Colours of some important metaborates** are : Cupric metaborate,  $Cu(BO_2)_2$  is dark blue, chromium metaborate,  $Cr(BO_2)_2$  is green, nickel metaborate,  $Ni(BO_2)_2$  is brown and manganese metaborate,  $Mn(BO_2)_2$  is pink violet.

(c) When heated with  $C_2H_5OH$  and conc.  $H_2SO_4$  it gives volatile vapours of triethyl borate which burns with a green edged flame.







In alum crystals, 6 water molecules are held by monovalent ion, 6 water molecules are held by trivalent ion, 12 water molecules are held in the crystal structure.

(b) All alums are isomorphous. Aqueous solutions of alums are acidic due to cationic hydrolysis of trivalent cation.

(c) Double sulphates of divalent ions and trivalent ions with 24 water molecules in their crystals are known as Pseudoalums. General formula is  $MSO_4 \cdot M'(SO_4)_3 \cdot 24H_2O$

$M$  = Bivalent metal;  $M'$  = Trivalent metal

(d) Pseudoalums are not isomorphous with alums.

(e) Feather alum or 'Hair-salt'  $Al_2SO_4 \cdot 18H_2O$  is a native form of aluminium sulphate.

(f) Potash alum is used for tanning of leather, as mordant in dyeing and calico printing, for sizing paper, as a syptic to stop bleeding and purification of water.

#### Some important alums are

Potash alum  $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$

Sodium alum  $Na_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$

Ammonium alum  $(NH_4)_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$

Chrome alum  $K_2SO_4 \cdot Cr_2(SO_4)_3 \cdot 24H_2O$

## Carbon Family

Carbon is the first member of group 14 or IVA of the periodic table. It consists of five elements carbon (*C*), silicon (*Si*), germanium (*Ge*), tin (*Sn*) and lead (*Pb*). Carbon and silicon are nonmetals, germanium is metalloid and tin and lead are metals.

#### (i) Electronic configuration

Elements	Electronic configuration ( $ns^2 np^2$ )
${}_6C$	$[He]2s^2 2p^2$
${}_{14}Si$	$[Ne]3s^2 3p^2$
${}_{32}Ge$	$[Ar]3d^{10} 4s^2 4p^2$
${}_{50}Sn$	$[Kr]4d^{10} 5s^2 5p^2$
${}_{82}Pb$	$[Xe]4f^{14} 5d^{10} 6s^2 6p^2$

### Physical properties

(1) **Non-metallic nature** : The non-metallic nature decreases along the group.

<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
Non-metals		metalloid	metal	metal

(2) **Abundance** : Carbon and silicon are most abundant elements in earth's crust whereas germanium occurs only as traces. Tin and lead also occur in small amounts. Only carbon occurs in free state as coal, diamond and graphite and in combined state as carbonates,  $CO$ , petroleum and natural gas. Silicon is the second most abundant element after oxygen in earth's crust in form of silicates and silica. Germanium found in traces in coal and in certain deposits. It important constituent for making conductors and transistors. The important ore of tin is tin stone ( $SnO$ ) or cassiterite. Lead is found in form of galena ( $PbS$ ), anglesite ( $PbSO_4$ ) and cerussite ( $PbCO_3$ ). The abundance ratio in earth's crust is given below,

(3) **Density** : The density of these elements increases down the group as reported below

Element	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
Density (g/ml)	3.51 (for diamond) 2.22 (for graphite)	2.34	5.32	7.26	11.34

### (4) Melting point and boiling points

(i) The melting point and boiling point of this group members decrease down the group.

Element	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
m.pt.(K)	4373	1693	1218	505	600
b.pt.(K)	–	3550	3123	2896	2024

(ii) The melting point and boiling point of group 14 elements are however, higher than their corresponding group 13 elements. This is due to the formation of four covalent bonds on account of four electrons in their valence shells which results in strong binding forces in between their atoms in solid as well as in liquid state.

### (5) Atomic radii and atomic volume

(i) Both atomic radii and atomic volume increases gradually on moving down the group due to the effect of extra shell being added from member to member.

	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
Atomic radius (pm)	0.77	111	122	141	144
Atomic volume ( $m^3$ )	3.4	11.4	13.6	16.3	18.27

(ii) The atomic radii of group 14 elements are than their corresponding group 13 elements due to increase in nuclear charge in the same period.

(iii) Some of the ionic radii involving six co-ordination of these group elements are given below,

	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
Ionic radius ( $M^4+$ ) in pm	–	–	73	118	119
Ionic radius ( $M^2+$ ) in pm	–	40	53	69	78

(6) **Electronegativity** : The electronegativity decreases from *C* to *Si* and then becomes constant.

	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
Electronegativity on Pauling scale	2.5	1.8	1.8	1.7	1.6

The electronegativity from silicon onwards is almost constant or shows a comparatively smaller decrease due to screening effects of *d* electrons in elements from *Ge* onwards.

### (7) Ionisation energy

(i) The ionisation energy decreases regularly down the group; *Pb* however shows a higher value than *Sn* due to poor shielding of inner *f* orbitals as a result of which effective nuclear charge experienced by outer shell electrons becomes more in *Pb*.

Ionisation energy ( $kJ\ mol^{-1}$ )	<i>C</i>	<i>Si</i>	<i>Ge</i>	<i>Sn</i>	<i>Pb</i>
IE <sub>1</sub>	1086	786	761	708	715
IE <sub>2</sub>	2352	1577	1537	1411	1450

(ii) The first ionisation energies of group 14 elements are higher than their corresponding group 13 elements because of smaller size.

(iii) The electropositive character of these elements increases down the group because of decrease in ionisation energy.

### (8) Oxidation state

(i) Presence of four electrons in outermost shell of these elements reveals that the members of this family can gain four electrons forming  $M^4+$  or  $M^4-$  ions to show ionic nature or exhibit tetravalent covalent nature by sharing of four electron pairs in order to attain stable configuration.

(ii) The formation of  $M^4+$  or  $M^4-$  ions require huge amount of energy which is normally not available during normal course of reactions, therefore, these elements usually do not form  $M^4+$  or  $M^4-$  ions, but they usually form compounds with covalence of four.

(iii) *Ge*, *Sn* and *Pb* also exhibit +2 oxidation state due to inert pair effect.

(iv) *Sn* and *Pb* show ionic nature.

(v) The tendency to form +2 ionic state increases on moving down the group due to inert pair effect.

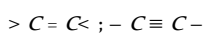
### (9) Catenation

(i) The tendency of formation of long open or closed atom chains by the combination of same atoms in themselves is known as catenation.

(ii) The catenation is maximum in carbon and decreases down the group.

(iii) This is due to high bond energy of catenation.

(iv) Only carbon atoms also form double or triple bonds involving  $p\pi-p\pi$  multiple bond within itself.



(v) Carbon also possesses the tendency to form closed chain compounds with *O*, *S* and *N* atoms as well as forming  $p\pi-p\pi$  multiple bonds with other elements particularly nitrogen and oxygen e.g.  $C=O$ ;  $C=N$ ;  $C \equiv N$ ;  $C=S$  are the functional groups present in numerous molecules due to this reason.

(vi) Carbon can form chain containing any number of carbon atoms. *Si* and *Ge* cannot extend the chain beyond 6 atoms, while *Sn* and *Pb* do not form chains containing more than one or two atoms.

(vii) The reason for greater tendency of carbon for catenation than other elements in the group may further be explained by the fact that the  $C-C$  bond energy is approximately of the same magnitude as the energies of the bond between *C* and other elements. On the other hand, the  $Si-Si$  bond is weaker than the bond between silicon and other elements.

Bond	Bond energy (kJ/mol)	Bond	Bond energy (kJ/mol)
C-C	348	Si-Si	180
C-O	315	Si-O	372
C-H	414	Si-H	339
C-Cl	326	Si-Cl	360
C-F	439	Si-F	536

#### (10) Allotropy

The phenomenon of existence of a chemical element in two or more forms differing in physical properties but having almost same chemical nature is known as allotropy. If an element or compound exists in two or more forms, it is also known as polymorphism e.g. zinc blende and wurtzite are polymorphs of  $ZnS$ .

Kinds of allotropy. Allotropy is of three types :

(i) **Enantiotropy** : When two forms of a solid substance exist together in equilibrium with each other at a particular temperature under normal pressure it is called enantiotropy.

For example, at normal pressure and temperature between  $368.6\text{ K}$  and  $285\text{ K}$ , sulphur (solid) exist in two forms (rhombic sulphur),  $S_R$  and (monoclinic sulphur),  $S_M$  in equilibrium with each other.  $S_R \rightleftharpoons S_M$

(ii) **Monotropy** : It is the type of allotropy in which only one allotrope is stable, under normal conditions the other being unstable e.g., diamond and graphite, oxygen and ozone etc.

(iii) **Dynamictropy** : It is the type of allotropy in which there is a true equilibrium between the two allotropes, one changing into the other at exactly the same rate as the reverse occurs. Both allotropes are stable over a wide range of temperature.

For example, liquid sulphur exist in two forms, the pale yellow mobile form called  $S_\lambda$  and dark viscous form called  $S_\mu$  in equilibrium with each other.  $S_\lambda \rightleftharpoons S_\mu$

With increase in temperature, the later form is formed at the expense of the former but when the liquid is cooled, the reverse change occur. Thus sulphur shows both enantiotropy and dynamictropy.

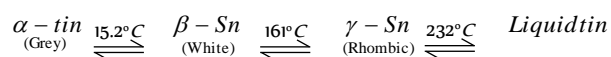
**Cause of allotropy** : (i) In general the allotropy among solid substances is due to the difference in crystalline structure. (ii) It may also be due to the presence of different number of atoms e.g.,  $O_2$  and  $O_3$ ,  $S_8$ ,  $S_2$  etc. (iii) It may be due to the difference in nuclear spins e.g., ortho and para hydrogen.

**Different allotropic forms** Except lead, all elements show allotropy.

(i) Carbon has two crystalline allotropic forms i.e., diamond and graphite. In diamond  $C$  atom is  $sp^3$  hybridised and it has a three dimensional network structure. Since no valence electron is available, hence diamond is a bad conductor of electricity. However in graphite  $C$  atom is  $sp^2$  hybridised and has a delocalised  $\pi$ -electron cloud responsible for its high electrical conductivity. It may be noted that diamond is thermodynamically less stable than graphite at ordinary temperatures.

(ii) Silicon has both crystalline and amorphous forms.

(iii) Tin has three crystalline modifications with the following equilibrium temperature



The conversion of white tin to grey tin is accompanied by an increase in volume and the latter, being very brittle, easily crumbles down to powder. This phenomenon is called tin disease tin pest or tin plague.

#### Chemical properties

(1) **Hydrides** : All the elements of group 14 combine with hydrogen directly or indirectly to form the covalent hydrides,  $MH_4$  ( $M = C, Si, Ge, Sn$  or  $Pb$ ). The number of hydrides and the ease of preparation decrease on going from carbon to lead.

The hydrides of silicon are called *silanes* having the general formula  $Si_nH_{2n+2}$ . The hydrides of germanium are called *germanes* while those of tin are called the *stannanes*. Only lead forms an unstable hydride of the formula,  $PbH_4$  called the *plumbane*.

Three hydrides of germanium, i.e.,  $GeH_4$ ,  $Ge_2H_6$  and  $Ge_3H_8$  and only two hydrides of tin i.e.,  $SnH_4$  and  $Sn_2H_6$  are well known.

(2) **Oxides** : Carbon forms five oxides  $CO, CO_2, C_3O_2$  (carbon suboxide),  $C_5O_2$  and  $C_{12}O_9$ ,  $C_3O_2$  is the anhydride of malonic acid and  $CO_2$  is the anhydride of  $H_2CO_3$  (carbonic acid)  $CO_2$  is a non-polar linear molecule due to maximum tendency of C to form  $p\pi-p\pi$  multiple bond with oxygen. Si forms  $SiO_2$ . Pb forms a number of oxides.  $PbO$  can be obtained by heating  $Pb(NO_3)_2$ ,  $2Pb(NO_3)_2 \xrightarrow{\text{Heat}} 2PbO + 4NO_2 + O_2$ .

The red form of  $PbO$  is called *litharge* and the yellow form is *massicot*.  $Pb_3O_4$  (Red lead, or Sindur) is prepared by heating litharge in air at  $470^\circ\text{C}$ ,  $6PbO + O_2 \xrightarrow{470^\circ\text{C}} 2Pb_3O_4$ ,  $Pb_3O_4$  is a mixed oxide of  $PbO_2 \cdot 2PbO$ .  $Pb_2O_3$  is called lead sesquioxide.  $GeO_2, SnO_2$  etc. are also network solids.

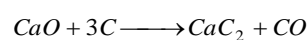
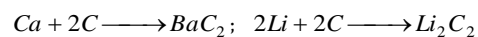
$CO_2$  and  $SiO_2$  is acidic,  $GeO_2$  is weakly acidic while  $SnO_2$  and  $PbO_2$  are amphoteric in nature.

All the elements of group 14 except silicon form monoxides e.g.,  $CO, GeO, SnO$  and  $PbO$ . Out of these monoxides only  $CO$  is neutral, while all other monoxides are basic.

(3) **Halides** : Elements of group 14 react with halogens directly to form tetrahedral and covalent halides except C where its halide is produced by the action of halogens on hydrocarbons.  $PbBr_4$  and  $PbI_4$  do not exist because  $Pb^{4+}$  is a strong oxidant and  $Br^-$  and  $I^-$  are strong reductants. Hence  $Pb^{4+}$  ion is difficult to survive in presence of strong reductants  $Br^-$  and  $I^-$  and is immediately reduced to  $Pb^{2+}$ .

(4) **Carbides** : Carbides are binary compounds of carbon with elements of lower or about equal electronegativity.

**Preparation** : Carbides are generally prepared by heating the elements or its oxide with carbon or hydrocarbon at very high temperatures.



Carbides are classified into three types on the basis of chemical bonding.

(i) **Salt like carbides** : These carbides are formed by the metals of groups IA, IIA, IIIA (except boron), coinage metals, Zinc, cadmium & some lanthanides.

(ii) **Acetylides** : These are ionic carbides which yield acetylene on hydrolysis. The alkali metals and copper, silver and gold form  $M_2C$  type compounds. These contain  $C_2^{2-}$  ions.

(iii) **Methanides** : These carbides evolve methane on hydrolysis.  $Al_4C_3, Be_2C, Mn_3C$  etc are some are of methanides. These contains  $C^{4-}$  groups.

(iii) **Allylides** : These carbides evolve allylene (methyl acetylene) on hydrolysis. This type of the carbides is only  $Mg_2C_3$  it contains  $C_3^{4-}$  discrete groups.

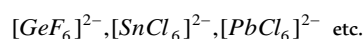
(2) **Mixed carbides** : These carbides yield a mixture of hydrocarbons on hydrolysis, carbides of iron group,  $UC_2$  and  $ThC_2$  belong to this group.

(3) **Covalent carbides** : The only true covalent carbides are those of  $SiC$  (carborundum) and  $B_4C, B_{13}C_2$  etc. These are chemically inert so become hard.

On account of hardness, these are used as abrasives.

(4) **Metallic or interstitial carbides** : If these carbides possess metallic lustre high electrical conductivity and chemically inert. These are extremely hard like diamond and possess very high melting points.

**Ability to form complexes** : The ability of group 14 elements to form complexes is highly favoured by a high charge, small size and availability of empty orbitals of the right energy. The compounds in which carbon shows a covalency of four possess a closed shell electronic configuration of a noble gas and therefore carbon does not form complexes. Silicon and other heavier elements, however, can form complexes due to the availability of energetically suitable empty  $d$ -orbitals and a coordination number of six is found in these complexes. For example, in the formation of  $[SiF_6]^{2-}$ , four covalent and two co-ordinate bonds are formed as a result of  $sp^3d^2$  hybridisation. As such the resulting ion has an octahedral geometry. Thus elements like  $Si, Ge, Sn$  and  $Pb$  have an ability to increase their co-ordination number from four to six. Other examples of hexa co-ordinated species are :



#### Anomalous behaviour of Carbon

Carbon is found differ in many properties from the rest of the members of group 14. This is because of the following : (i) Its smallest size (ii) Its high electronegativity (iii) Its property to catenate (iv) Absence of  $d$ -orbitals in it.

Some of the properties in which it differs from other members are,

(1) The melting and boiling points of carbon are very high as compared to the rest to the members of the family.

(2) Carbon in its diamond form is one of the hardest substance known.

(3) It has maximum tendency to show catenation.

(4) Carbon has high tendency to form  $P\pi - P\pi$  multiple bonds with other elements like nitrogen, oxygen, sulphur etc. Other members of the family form  $P\pi - d\pi$  bonds and that also to a lesser extent.

(5)  $CO_2$  is a gas while the dioxides of all other members are solids.

(6) Carbon is not affected by alkalis whereas other members react on fusion. For example, silicon form silicates,  $Si + 2NaOH + 1/2 O_2 \rightarrow Na_2SiO_3 + H_2$ .  
Sodium silicate

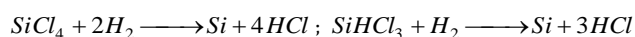
#### Silicon and its compounds

Silicon, being a second member of group - 14, has a much larger size and lower electronegativity than that of carbon. As a result silicon does not form double bond with itself or with oxygen. Thus  $SiO$  bonds are much stronger than  $Si - Si$  and  $Si - H$  bonds. Silicon has vacant  $3d$ -orbitals in its valence shell due to which it can extend its covalency from four to five and six.

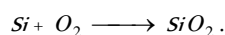
(1) **Occurrence** : Silicon is the second most abundant element ( 27.7%) in earth's crust next to oxygen .It does not occur in free state. It occurs mainly in the form of Silica and silicates. Silicates are formed in rocks and clay as silicates of  $Mg, Al, K$  or  $Fe$ . e.g. Feldspar ;  $K_2Al_2O_3 \cdot 6SiO_2$ , Kaolinite;  $Al_2O_3 \cdot 2SiO_2 \cdot 2H_2O$ .

(2) **Preparation** : Elemental silicon is obtained by reduction of silica with high purity coke in an electric furnace using excess of silica e.g.  $SiO_2 + 2C \longrightarrow Si + 2CO$

Very high purity silicon required for making semiconductors is obtained by reduction of highly purified  $SiCl_4$  form ( $SiHCl_3$ ) with hydrogen followed by purification by zone refining eg.



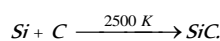
(3) **Properties** : (i) Silicon exists in three isotopes  $_{14}Si^{29}$  (most common),  $_{14}Si^{30}$  with air at high temperature  $SiO_2$  form,



(ii) With steam,  $Si$  reacts when heated to redness to liberate hydrogen,  $Si + 2H_2O \xrightarrow{\text{Redness}} SiO_2 + 2H_2$ .

(iii) With halogens,  $Si$  reacts at elevated temperature forming  $SiX_4$  except fluorine which reacts at room temperature.

(iv) Silicon combines with  $C$  at  $2500K$  forming Silicon Carbide ( $SiC$ ) known as carborundum (an extremely hard substance),



(v) It reacts with metals like  $Ca, Mg$  etc in an electric arc furnace to form Silicides ( $Ca_2Si, Mg_2Si$  etc.)

(vi) Silicon dissolves in hot aqueous alkalis liberating hydrogen,  $Si + 4NaOH \xrightarrow{\text{Heat}} Na_4SiO_4 + 2H_2 \uparrow$

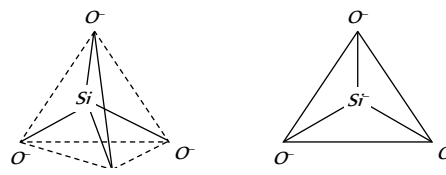
(vii) It also dissolves in fused  $Na_2CO_3$  displacing carbon  $Na_2SiO_3 + C$ .

(4) **Uses of silicon** : (i) It is added to steel as ferrosilicon ( an alloy of  $Fe$  and  $Si$ ) to make it acid resistant.

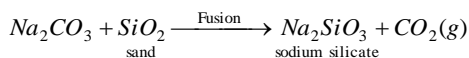
(ii) It is used in the pure form as a starting material for production of silicon polymers (Silicones).

#### (5) Compounds of silicon








(i) **Silicates** : Silicates are the metal derivatives of silicic acid ( $H_4SiO_4$ ). The basic of all silicates is the  $SiO_4^{4-}$  anion. In  $SiO_4^{4-}$  anion,  $Si$  is  $sp^3$  hybridised, and it forms four covalent bonds with four negatively charged oxygen atoms.  $SiO_4^{4-}$  anion has a tetrahedral shape.



Alkali metal silicates are commonly prepared by fusing metal oxides or metal carbonates with sand ( $SiO$ ) at high temperatures. For examples, sodium silicate can be prepared by fusing sand with sodium carbonate.



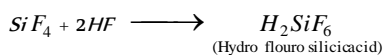
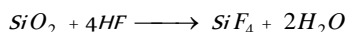
**Table 18.1 Classification of silicates**

No. of corners of $SiO_4$ tetrahedra shared	No. of common oxygen atoms	Structure	Net charge and Anion in silicate structure	Example
		Diagram and Description		
Zero	Zero	 Discrete $SiO_4^{4-}$ anion	$Si = +4$ $O = -8$ $Net = -4$ $SiO_4^{4-}$	Ortho-silicates $Mg_2SiO_4$
1	1	 Island structure	$Si = +8$ $O = -14$ $Net = -6$ $(Si_2O_7)^{6-}$	Pyro-silicates
2	2	 Ring anion	$Si = +12$ $O = -18$ $Net = -6$ $(Si_3O_9)^{6-}$	Wollastonite $Ca_3Si_3O_9$
2	2	 Ring anion	$Si = +24$ $O = -36$ $Net = -12$ $(Si_6O_{18})^{12-}$	Beryl, $Be_3Al_2Si_6O_{18}$
2	2	 Chain anion	$Si = +4$ $O = -6$ $Net = -2$ $(SiO_3^{2-})_n$	Pyroxenes, e.g., $MgCaSi_2O_6$ Asbestos
3	3	 Two dimensional sheet structure	$Si = +8$ $O = -10$ $Net = -2$ $(Si_2O_5^{2-})_n$	Clays, talc kaolinite
4	4	 Three dimensional network	$Si = +4$ $O = -4$ $(SiO_2)_n$	Quartz, Tridymite and Cristobalite

(ii) **Silica or silicon dioxide ( $SiO_2$ )**

It occurs in nature in various forms such as sand, quartz and flint. It is also a constituent of various rocks. It is solid at room temperature. It is insoluble in water.

Silica has a three dimensional network structure in which each  $Si$  is bonded to four oxygen atoms which are tetrahedrally disposed around silicon atom. Each  $O$  atom is shared by two  $Si$  atoms. It may be noted that  $CO_2$  is a gas, while  $SiO_2$  is hard solid with very high melting point.



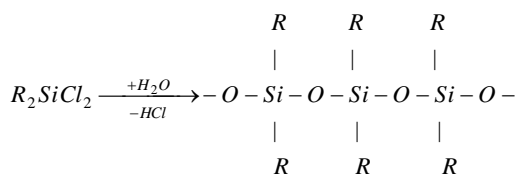
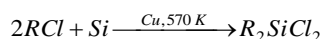
$HF$  readily dissolves Silica, therefore  $HF$  can not be store in glass bottles which contain Silica.

It is used in large amount to form mortar which is a building material. It is also used in the manufacture of glass and lenses.

(iii) **Silicones**

Polymeric organo-silicon compounds containing  $Si-O-Si$  bonds are called silicones. These have the general formula  $(R_2SiO)_n$ . Where  $R$  is  $CH_3$ -group (majority cases) or  $C_6H_5$ -group.

**Preparation :** The preparation of silicones is generally carried out by the hydrolysis of dialkyldichlorosilanes ( $R_2SiCl_2$ ) or diaryldichlorosilanes ( $Ar_2SiCl_2$ ), which are prepared by passing vapours of  $RCl$  or  $ArCl$  over silicon at  $570K$  with copper as a catalyst.

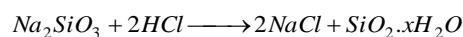


Silicones may be obtained in the form of oils, rubber on resins depending upon the extent of polymerisation which depends upon reaction conditions and nature of alkyl groups.

**Properties and Uses :** Silicones are water repellent and quite inert chemically. These resist oxidation, thermal decomposition and attack by organic reagents. These are also good electrical insulators and antifoaming agents. These have found the following uses :

- Silicones have been used for making water-proof papers, wools, textiles, wood etc., after coating these articles with silicones.
- The viscosities of silicones do not change with changes in temperature, therefore, these are used as all weather lubricants.
- As antifoaming agent in industrial processes.
- As a mould releasing agent in rubber industry and foundry. It avoids the sticking of the castings to the mould.
- For making body implants in cosmetic surgery due to its inert nature.
- Silicones are now incorporated in paints for resisting dampness and for water proofing.
- Due to their water repellent nature and high dielectric constant, silicones are used in electrical condensers.

(iv) **Silica gel :** When a mineral acid (Such as  $HCl$ ) is added to a concentrated solution of a silicate, gelatinous white ppt. of hydrated silica (silicic acid) separate out.



The white ppt. thus obtained is heated to lose water. When the water content is very low, the solid product is called silica gel. It possesses excellent absorptive properties due to its porous nature and is used for absorbing moisture and an adsorbent in chromatography.

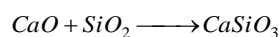
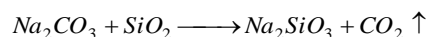
(v) **Silanes** : The hydrides of silicon are called silanes. For example;  $SiH_4$  Silane,  $Si_2H_6$  disilane,  $Si_3H_8$  Trisilane  $Si_4H_{10}$  Tetrasilane.

Silanes are poisonous. These are much less stable than the corresponding alkanes and are decomposed into elements on heating above  $450^\circ C$ . Their thermal stability decreases with increase in molecular mass. Unlike alkanes, silanes are reducing agents.

(vi) **Glass**

Glass is an amorphous and transparent solid which is obtained by solidification of various silicates and borates of potassium and calcium.

**Preparation** : Ordinary glass is a mixture of sodium and calcium silicates and is produced by fusing together a mixture of sodium carbonate, calcium oxide and silicon dioxide (Silica) in a furnace at about 1700K



On continuously heating the entire amount of  $CO_2$  is driven out and clear viscous fused mass is obtained. It is poured into moulds to get different types of articles, which are allowed to cool gradually.

This typed of glass is called soda glass or soft glass which has the approximate composition,  $Na_2SiO_3, CaSiO_3, 4SiO_2$ .

**Various varieties of glass** : The different varieties of glasses and their special constituents are given below,

Table 18.2

Type of glass	Constituents	Special use
Soft glass	$Na_2CO_3, CaCO_3, SiO_2$	Ordinary glass for window panes, test tubes, bottles, etc.
Hard glass	$K_2CO_3, CaCO_3, SiO_2$	For combustion tubes and chemical glassware
High refractive index glass (Flint glass)	Lead oxide, $K_2CO_3$	For making lenses cut glasses
Pyrex glass	$Na_2CO_3, Al_2O_3, B_2O_3$ or borax, sand	For high quality glass apparatus cooking utensils
Crook's glass	$K_2CO_3, PbCO_3, CeO_2$ , sand	Absorbs ultra violet rays, for making lenses
Jena glass	Zinc and Barium Boro silicates	It is resistant to heat shock and common reagent. It is used for making good quality of glass wares.

**Coloured glass** : Addition of transition metal compounds to glass give coloured glasses. Small amounts of  $Cr(III)$ ,  $Mn(IV)$ ,  $Co(II)$  and  $Fe(III)$  compounds impart green, violet blue or brown colour respectively

Table 18.3

Compound added – Colour imparted	Compound added – Colour imparted
Cobalt oxide ( $CoO$ ) – Blue	Chromium oxide ( $Cr_2O_3$ ) – Green
Cuprous oxide ( $Cu_2O$ ) – Red	Auric chloride ( $AuCl_3$ ) – Ruby
Cadmium sulphide ( $CdS$ ) – Lemon yellow	Manganese dioxide ( $MnO_2$ ) – Purple

**Etching of glass** : Glass is attacked by hydrofluoric acid. This property is used in the etching of glass. The glass to be etched is coated with a thin layer of wax and the design to be produced is scratched with a needle. An aqueous solution of  $HF$  is applied to the exposed part. After some time it is placed in hot water and wax is removed from the surface. The marks are engraved on the exposed parts.

#### Tin and its Compounds

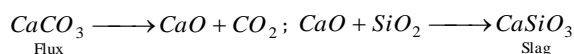
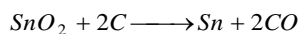
(1) **Important ore** : Cassiterite (tin stone)  $SnO_2$

(2) **Extraction of tin from tin stone.**

(i) **Concentration** : The powdered tin stone is concentrated by gravity separation and the magnetic impurities like wolframite etc., are separated from tin stone by magnetic separators.

(ii) **Roasting** : The concentrated ore is heated in a current of air when impurities like  $S$  and  $As$  are oxidised to volatile  $SO_2$  and  $As_2O_3$ . Iron pyrites change to their oxides and sulphates.

(iii) **Leaching and washing** : The roasted ore is treated with water when  $CuSO_4$  and  $FeSO_4$  are washed away from the main ore. Further lighter ferric oxide is washed away leaving behind heavier ore particles known as black tin containing 60 to 70%  $SnO_2$ .



Molten tin is drawn into blocks. It contains 99.5 percent of tin metal and is called block tin.

**Refining of tin** : It is purified by liquation, poling and electrolytic refining.

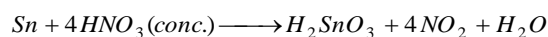
For very high purity, it is purified by electrolytic method. The electrolyte consists of tin sulphate containing a small amount of hydrofluorosilicic acid ( $H_2SiF_6$ ) and sulphuric acid. Impure tin makes anode while pure tin sheet serves as cathode.

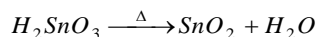
(3) **Compounds of Tin**

(i) **Stannic oxide,  $SnO_2$**  : It is prepared by heating tin strongly in air.



It can also be prepared by heating metastannic acid obtained by the action of conc.  $HNO_3$  on tin.

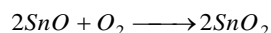




It occurs in nature as tin stone. It is a white solid insoluble in water and is amphoteric in nature. With  $NaOH$  it forms  $Na_2SnO_3$ . It is used for making enamels and glazes for tiles, pottery etc. it is also used as a polishing powder.

(ii) **Stannous oxide,  $SnO$** : It is prepared by heating stannous oxalate  $SnC_2O_4 \xrightarrow{\Delta} SnO + CO + CO_2$

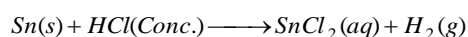
Oxidation of  $SnO$  to  $SnO_2$  is checked by  $CO$ . It is a grey solid which oxidises readily to  $SnO_2$  when heated in air.



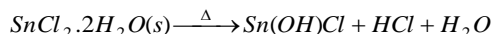
It is amphoteric in nature and reacts both with acids and alkalis. With  $NaOH$  it forms  $Na_2SnO_2$ .

(iii) **Stannous sulphide,  $SnS$** : It is insoluble in water but soluble in hot conc.  $HCl$ . In yellow ammonium polysulphide it gets converted to ammonium thiostannate  $(NH_4)_2SnS_3$ .

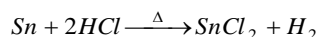
(iv) **Stannous chloride,  $SnCl_2$** : When  $Sn$  is heated with  $HCl$  (conc.)  $SnCl_2$  is formed.



On concentrating the resulting solution, crystals of  $SnCl_2 \cdot 2H_2O$  are obtained. When it is heated, basic tin chloride is obtained.

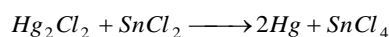
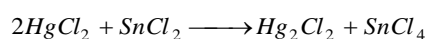


To obtain anhydrous  $SnCl_2$ , heat  $Sn$  in dry  $HCl$  gas.



(a) It exists as a anhydrous (white powder, m.p. = 520K, rhombic solid) as well as dihydrate  $SnCl_2 \cdot 2H_2O$  (white, m.p. = 480K, monoclinic) and is used as a strong reducing agent in conc.  $HCl$  in laboratory.

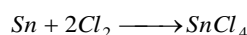
$SnCl_2$  also reduces  $HgCl_2$



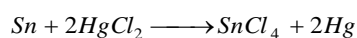
(b) It is precipitated as hydroxide by an alkali.

(c) It forms addition compounds with  $NH_3$  such as  $SnCl_2 \cdot NH_3$  and  $SnCl_2 \cdot 2NH_3$ .

(v) **Stannic chloride,  $SnCl_4$** : It is obtained by the action of  $Cl_2$  on molten  $Sn$



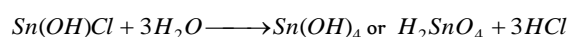
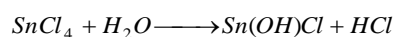
It can also be obtained by distilling tin with mercuric chloride.



(a) It is a colourless fuming liquid (b.p 388K) soluble in water.

It is used as a strong reducing agent in laboratory. It is also used as a mordant in dyeing.

(b) It can exist as  $SnCl_4 \cdot 5H_2O$  and with excess water it is hydrolysed to form basic chloride and ultimately stannic acid ( $H_2SnO_4$ ).



Its hydrolysis is prevented by  $HCl$  which forms complex anion  $[SnCl_6]^{2-}$

(c) It forms double salts with  $NH_3, N_2O, PCl_5$  e.g.,  $SnCl_4 \cdot 4NH_3$ .

It is used as a mordant and tinning agent.

(vi) **Stannous fluoride,  $SnF_2$** : It is obtained by dissolving  $SnO$  in  $HF$



It is a white crystalline solid insoluble in water. It is used in tooth pastes to help in controlling dental decay.

**Tinning**: During cooking, organic acids present in food stuff attack the household utensils made of copper, brass etc. in the presence of air. since tin is not attacked by organic acids, the utensils are protected by tinning.

### Lead

#### (1) Some important ores

Galena;  $-PbS$  (Main); Cerussite  $-PbCO_3$

Anglesite  $-PbSO_4$ , lararkite  $PbO \cdot PbSO_4$

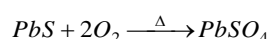
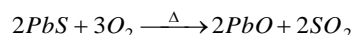
#### (2) Extraction from galena

(i) **Concentration**: The finely powdered ore is concentrated by froth floatation process.

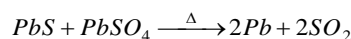
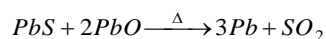
#### (ii) Reduction process

##### (a) Self reduction process

• **Roasting**: The concentrated ore is heated in air. lead sulphide is partially converted into lead oxide and lead sulphate.

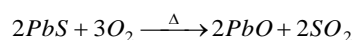
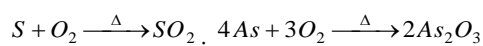


• **Self reduction**: The supply of air is cut off and the temperature is raised to melt the charge. The galena reduces both  $PbO$  and  $PbSO_4$  to metallic lead.

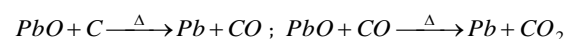


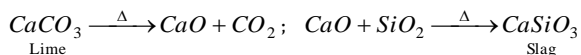
##### (b) Carbon reduction process

• **Roasting**: The powdered ore is mixed with lime and roasted in excess of air. the impurities like  $S$  and  $As$  are oxidised to their volatile oxides while  $PbS$  is converted to  $PbO$ .



• **Smelting**: Roasted ore is mixed with coke and lime (flux) and smelted in a blast furnace.  $PbO$  is reduced to  $Pb$ .





(iii) **Purification** : It is purified electrolytically. The electrolyte consists of lead silicofluoride ( $\text{PbSiF}_6$ ) and hydrofluosilicic acid. Impure lead is made anode and sheet of pure lead serves as cathode.

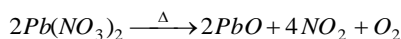
#### Properties of Lead

With oxygen lead form oxides, with chlorine it forms chloride  $\text{PbCl}_2$ , with sulphur it gives sulphide  $\text{PbS}$  and with  $\text{H}_2\text{SO}_4$  the corresponding sulphate  $\text{PbSO}_4$ . With  $\text{NaOH}$  it forms plumbate.



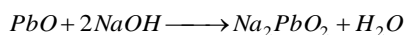
#### (3) Compounds of Lead

(i) **Lead oxide (Litharge),  $\text{PbO}$**  : It is prepared by heating the nitrate.



It exist in two varieties yellow form (messicol) and red form (litharge). Yellow form is prepared by gently heating lead in air while fusion yield red form. It is insoluble in water and amphoteric in nature.

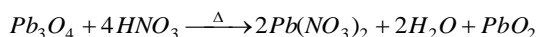
It dissolves in  $\text{NaOH}$  to form sod. Plumbite.



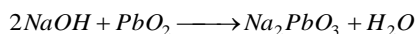
It can be reduced with various reducing agents ( $\text{C}, \text{H}_2, \text{CO}$  etc.) to lead.

It is used in paints and varnishes, for making flint glass, for making lead (II) salts and for glazing pottery.

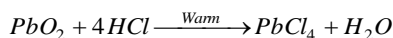
(ii) **Lead dioxide,  $\text{PbO}_2$**  : It is prepared by heating  $\text{Pb}_3\text{O}_4$  ( $2\text{PbO} + \text{PbO}_2$ ) with dilute  $\text{HNO}_3$



It is amphoteric in nature and dissolve in  $\text{NaOH}$  to form sodium plumbate.

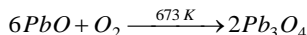


It is a powerful oxidising agent. It reacts with conc.  $\text{HCl}$  on warming to give  $\text{PbCl}_4$



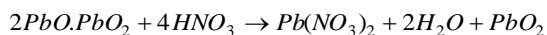
It is a chocolate brown solid insoluble in water and nitric acid. It is a powerful oxidizing agent. It is amphoteric in nature and is used in lead storage batteries and in safety matches.

(iii) **Minium or sindhur or Red lead,  $\text{Pb}_3\text{O}_4$**  : It is prepared by heating  $\text{PbO}$  in air to above  $673\text{K}$ .

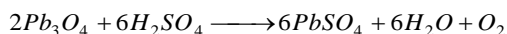
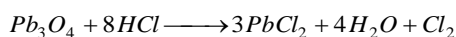


It is a red crystalline solid insoluble in water.

It is a mixed oxide  $\text{PbO}_2 + 2\text{PbO}$  and reacts with  $\text{HNO}_3$  to form  $\text{Pb}(\text{NO}_3)_2$  and  $\text{PbO}_2$



It is a strong oxidising agent. It liberates chlorine with conc.  $\text{HCl}$  and  $\text{O}_2$  with conc.  $\text{H}_2\text{SO}_4$

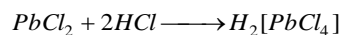


It is used as a protective paint in iron, steel and silver mirrors and in glass industry.

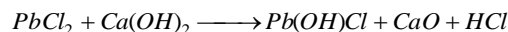
(iv) **Lead chloride,  $\text{PbCl}_2$**  : It can be prepared by treating a salt of lead with dil.  $\text{HCl}$



It can also be obtained by dissolving lead (II) oxide to lead (II) carbonate in  $\text{HCl}$ . It is soluble in hot water but precipitate out in cold water. it is soluble in conc.  $\text{HCl}$  due to the formation of a complex, tetrachloroplumbate (II) ion.

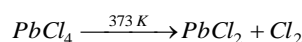


It also reacts with hot lime water to give  $\text{Pb}(\text{OH})\text{Cl}$  which is used as white pigment.

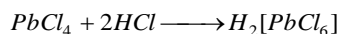


(v) **Lead tetrachloride,  $\text{PbCl}_4$**  : It is obtained by heating of  $\text{PbO}_2$  with conc.  $\text{HCl}$ .

It is a yellow oily fuming liquid which decomposes into  $\text{PbCl}_2$  at  $373\text{K}$ .



It also combines with  $\text{HCl}$  to form complex hexachloroplumbate (IV) ion.



## Nitrogen Family

Nitrogen is the first member of group 15 or VA of the periodic table. It consists of five elements nitrogen ( $N$ ), phosphorus ( $P$ ), arsenic ( $As$ ), antimony ( $Sb$ ) and bismuth ( $Bi$ ). The elements of this group are collectively called **pnictogens** and their compounds as pniconides. The name is derived from Greek word "Pniomigs" meaning suffocation. Pniconide contain  $M^{3-}$  species.

#### (1) Electronic configuration

Elements	Electronic configuration ( $ns^2 np^3$ )
${}_7N$	$[\text{He}]2s^2 2p^3$
${}_{15}P$	$[\text{Ne}]3s^2 3p^3$
${}_{33}As$	$[\text{Ar}]3d^{10} 4s^2 4p^3$
${}_{51}Sb$	$[\text{Kr}]4d^{10} 5s^2 5p^3$
${}_{83}Bi$	$[\text{Xe}]4f^{14} 5d^{10} 6s^2 6p^3$

#### Physical properties

(1) **Physical state** : Nitrogen – (gas), phosphorus – (solid) (vaporises easily),  $As, Sb, Bi$  – solids.

Nitrogen is the most abundant gas in the atmosphere. It constitutes about 78% by volume of the atmosphere. Phosphorus is the most reactive element in this group and its yellow form is always kept under water.

(2) **Atomic radii** : Atomic radii increases with atomic number down the group i.e., from  $N$  to  $Bi$  due to addition of extra principal shell in each succeeding elements.

(3) **Ionisation energy** : The ionisation values of the elements of this group decreases down the group due to gradual increases in atomic size.



(4) **Electronegativity** : Generally the elements of nitrogen family have high value of electronegativity. This value shows a decreasing trend in moving down the group from nitrogen to bismuth.

(5) **Non-metallic and metallic character** : Nitrogen and phosphorus are non-metals, arsenic and antimony are metalloids (semi-metal) and bismuth a typical metal.

(6) **Molecular state** : Nitrogen readily forms triple bond (two  $p\pi - p\pi$  bonds) and exists as discrete diatomic gaseous molecule ( $N \equiv N$ ) at room temperature. Phosphorus, arsenic and antimony exist in the form of discrete tetra atomic molecules such as  $P_4, As_4, Sb_4$  in which the atoms are linked to each other by single bonds.

(7) **Melting and boiling points** : The melting points and boiling points of group 15 elements do not show a regular trend.

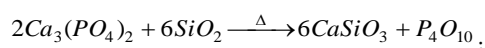
M.pt. first increases from  $N$  to  $As$  and then decreases from  $As$  to  $Bi$ . Boiling point increases from  $N$  to  $Sb$ . Boiling point of  $Bi$  is less than  $Sb$ .

(8) **Allotropy** : All the members of group 15 except  $Bi$  exhibit the phenomenon of allotropy.

(i) Nitrogen exists in two solid and one gaseous allotropic forms.

(ii) Phosphorus exists in several allotropic forms such as white, red, scarlet, violet and black form.

(a) **White or yellow phosphorus** : White phosphorus is prepared from rock phosphate  $Ca_3(PO_4)_2, SiO_2$  and coke which are electrically heated in a furnace.



When exposed to light, it acquires a yellow colour.

(b) **Red phosphorus** : It is obtained by heating yellow phosphorus, between  $240 - 250^\circ C$  in the presence of an inert gas. Yellow phosphorus can be separated from red phosphorus by reaction with  $NaOH(aq)$  or  $KOH(aq)$  when the former reacts and the latter remains unreacted.

(iii) Arsenic exists in three allotropic forms namely grey, yellow and black. Antimony also exists in three forms, viz., metallic, yellow and explosive.

(9) **Oxidation state** : The members of the group 15 exhibit a number of positive and negative oxidation states.

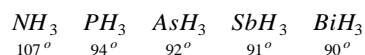
(i) **Positive oxidation states** : The electronic configuration ( $ns^2np^3$ ) for the valence shell of these elements shows that these elements can have +3 and +5 oxidation states. In moving down this group, the stability of +3 oxidation state increases. It may be pointed out here that nitrogen does not exhibit an oxidation state of +5, because it fails to expand its octet due to nonavailability of vacant  $d$ -orbitals.

(ii) **Negative oxidation states** : For example oxidation state of nitrogen is -3. The tendency of the elements to show -3 oxidation state decreases on moving down the group from  $N$  to  $Bi$ .

(10) **Catenation** (self linkage) : Elements of group 15 also show some tendency to exhibit catenation. This tendency goes on decreasing in moving down the group due to gradual decrease in their bond ( $M-M$ ) energies.

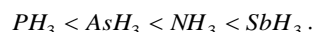
#### Chemical properties

(i) **Hydrides** : All the members form volatile hydrides of the type  $AH_3$ . All hydrides are pyramidal in shape. The bond angle decreases on moving down the group due to decrease in bond pair-bond pair repulsion.



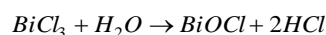
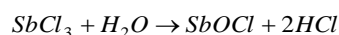
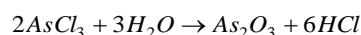
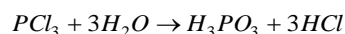
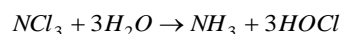
The decreasing order of basic strength of hydrides is as follows :  $NH_3 > PH_3 > AsH_3 > SbH_3 > BiH_3$ .

The increasing order of boiling points is as follows :



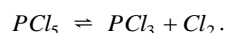
$NH_3$  is thermally most stable and  $BiH_3$  is least stable. This is because in  $NH_3$ ,  $N-H$  covalent bond is the strongest due to small size of  $N$  atom. Hence, the decomposition temperature of  $NH_3$  will be the highest. The increasing order of reducing character is as follows,  $NH_3 < PH_3 < AsH_3 < SbH_3 < BiH_3$ .

(2) **Halides** : The members of the family form trihalides ( $MX_3$ ) and pentahalides ( $MX_5$ ). The trihalides are  $sp^3$ -hybridized with distorted tetrahedral geometry and pyramidal shape while pentahalides are  $sp^3d$ -hybridized and are trigonal bipyramidal in shape. The trihalides are hydrolysed by water and ease of hydrolysis decreases when we move down the group. Hence,  $NCl_3$  is easily hydrolysed but  $SbCl_3$  and  $BiCl_3$  are partly and reversibly hydrolysed.  $NF_3$  is not hydrolysed due to lack of vacant  $d$ -orbital with nitrogen.  $PF_3$  and  $PF_5$  are also not hydrolysed because the  $P-F$  bond is stronger than  $P-O$  covalent bond. The hydrolysis products of the halides are as follows :



Their basic character follows this decreasing order as  $Ni_3 > NBr_3 > NCl_3 > NF_3$ . Except  $NF_3$ , the trihalides of nitrogen are unstable and decompose with explosive violence.  $NF_3$  is stable and inert.  $NCl_3$  is highly explosive. Trifluorides and trichlorides of phosphorus and antimony act as Lewis acid. The acid strength decreases down the group. For example, acid strength of tri-chlorides is in the order ;  $PCl_3 > AsCl_3 > SbCl_3$ .

Nitrogen does not form pentahalides due to non-availability of vacant  $d$ -orbitals. The pentachloride of phosphorus is not very stable because axial bonds are longer (and hence weaker) than equatorial bond. Hence,  $PCl_5$  decomposes to give  $PCl_3$  and  $Cl_2$ ;



The instability of  $PCl_5$  makes it a very good chlorinating agent. All pentahalides act as Lewis acids since they can accept a lone pair of electron from halide ion.

Solid  $PCl_5$  is an ionic compound consisting of  $[PCl_4]^+$   $[PCl_6]^-$ ,  $[PCl_4]^+$  has a tetrahedral structure, while  $[PCl_6]^-$  has an octahedral structure.

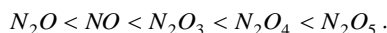
Since,  $PCl_5$  reacts readily with moisture it is kept in well stoppered bottles.

$PI_5$  does not exist due to large size of  $I$  atoms and lesser electronegativity difference between phosphorus and iodine.

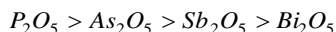
Down the group, the tendency to form pentahalides decreases due to inert pair effect. e.g.,  $BiF_5$  does not exist.

(3) **Oxides** : These elements form oxides of the type  $X_2O_3$ ,  $X_2O_4$  and  $X_2O_5$ .

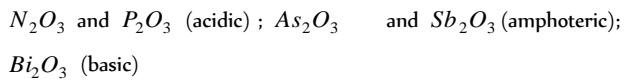
The acidic strength of oxides :



The decreasing order of stability of oxides of group 15 follows as,



The nature of oxides of group 15 elements is as follows,



(4) **Oxyacids** :  $N_2$  and  $P_4$  of this group forms oxyacids which are discussed further. In this chapter.

#### Anomalous behaviour of Nitrogen

Nitrogen is known to differ from other members of the family because of the following facts,

(1) Its small size (2) Its high electronegativity (3) Its high ionisation energy (4) non-availability of  $d$ -orbital in the valence shell. (5) Its capacity to form  $p\pi-p\pi$  multiple bonds.

The main points of difference are,

(i) Nitrogen is a gas while other members are solids.  
(ii) Nitrogen is diatomic while other elements like phosphorus and arsenic form tetra-atomic molecules ( $P_4, As_4$ ).

(iii) Nitrogen form five oxides ( $N_2O, NO, N_2O_3, N_2O_4$  and  $N_2O_5$ ) while other members of the family form two oxides (tri and pentaoxides).

(iv) Hydrides of nitrogen show H-bonding while those of other elements do not.

(v) Nitrogen does not show pentacovalency because of absence of  $d$ -orbitals while all other elements show pentacovalency.

(vi) Nitrogen does not form complexes because of absence of  $d$ -orbitals while other elements show complex formation e.g.,  $[PCl_6]^-$ ,  $[AsCl_6]^-$  etc.

(vii) The hydride of nitrogen ( $NH_3$ ) is highly basic in nature while the hydrides of other elements are slightly basic.

(viii) Except for  $NF_3$ , other halides of nitrogen e.g.,  $NCl_3, NBr_3$  and  $NI_3$  are unstable.

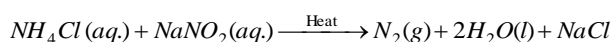
#### Nitrogen and its compounds

$N_2$  was discovered by **Daniel Rutherford**. It is the first member of group 15 in the periodic table.

**Occurrence** :  $N_2$ , occurs both in the free state as well as in the combined state.  $N_2$  occurs in atmosphere to the extent of 78% by volume in free state.  $N_2$  is present in many compounds such as potassium nitrate (nitre). Sodium nitrate (Chile salt peter) and many ammonium compounds.  $N_2$  is an important constituent of proteins in plants and animals in combined state.

**Preparation** : It is prepared by the following methods,

(1) **Laboratory method** : In the laboratory  $N_2$  is prepared by heating an aqueous solution containing an equivalent amounts of  $NH_4Cl$  and  $NaNO_2$ .



(2) **Commercial preparation** : Commercially  $N_2$  is prepared by the fractional distillation of liquid air.

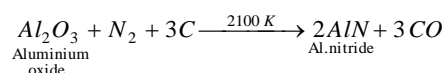
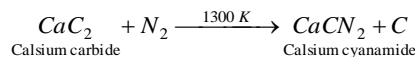
**Physical properties** :  $N_2$  is a colourless, odourless and tasteless gas. It is a non-toxic gas. Its vapour density is 14. It has very low solubility in water.

#### Chemical properties

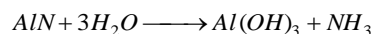
(1)  $N_2$  is neutral towards litmus. It is chemically unreactive at ordinary temp. It is neither combustible nor it supports combustion.

(2) The  $N-N$  bond in  $N_2$  molecule is a triple bond ( $N \equiv N$ ) with a bond distance of 109.8 pm and bond dissociation energy of 946 kJ mol<sup>-1</sup>

(3) **Combination with compounds** :  $N_2$  combines with certain compounds on strong heating . eg



Both these compounds are hydrolysed on boiling with water to give ammonia.



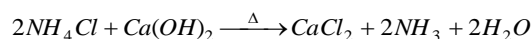
Therefore, calcium cyanamide is used as a fertilizer under the name nitrolim ( $CaCN_2 + C$ )

**Uses of nitrogen** :  $N_2$  is mainly used in the manufacture of compounds like  $NH_3, HNO_3, CaCN_2$  etc.

#### Compounds of nitrogen

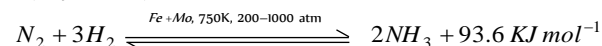
##### (1) Hydrides of nitrogen – Ammonia

**Preparation of ammonia** : Ammonia is prepared in the laboratory by heating a mixture of  $NH_4Cl$  and slaked lime,  $Ca(OH)_2$



Moist  $NH_3$  gas is dried over quick lime,  $CaO$ . However, it cannot be dried over conc.  $H_2SO_4, P_2O_5$  because being basic it forms salts with them. Anhydrous  $CaCl_2$  also cannot be used because it forms a complex  $CaCl_2 \cdot 8NH_3$  with it.

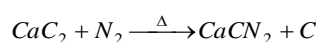
**Manufacture** : (i) Ammonia is manufactured by Haber's process. A mixture of pure  $N_2$  and  $H_2$  (in the ratio 1 : 3 by volume) is compressed to 200 – 1000 atmospheres and passed over finely divided  $Fe$  (as catalyst) and  $Mo$  (as promoter) at 750K

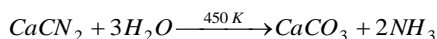


Favourable conditions for maximum yield of  $NH_3$  are :

(a) excess of reactants ( $N_2$  and  $H_2$ ) (b) high pressure (c) low temperature and (d) use of catalyst and a promoter.

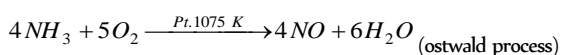
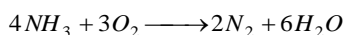
(ii) By the hydrolysis of calcium cyanamide ( $CaCN_2$ ) with superheated steam at 450K.  $CaCN_2$  itself is obtained by heating  $CaC_2$  and  $N_2$  at 1270K.



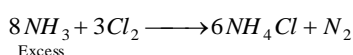


**Properties of  $NH_3$  :** It is a colourless gas with pungent smell, highly soluble in  $H_2O$  and basic in nature. It liquefies on cooling under pressure to give liquid ammonia (bp. 240K). On heating, it causes intense cooling and hence is used as a refrigerant in ice, factories and cold storages.

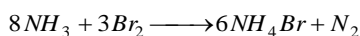
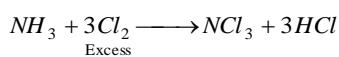
It burns in excess of air to give  $N_2$  and  $H_2O$  and is oxidised to  $NO$  when passed over heated  $Pt$  at 1075K.



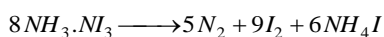
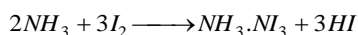
It reduces heated  $CuO$  to  $Cu$  and  $Cl_2$  to  $HCl$  (which combines with  $NH_3$  to give  $NH_4Cl$ ).



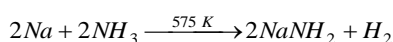
With excess of  $Cl_2$ , it gives  $NCl_3$ . With  $Br_2$  it gives  $NH_4Br$  and  $N_2$  is set free.



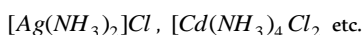
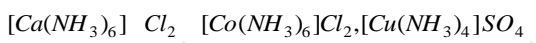
With  $I_2$ , it gives nitrogen triiodide ammonia (brown ppt) which is explosive in dry state and decomposes when struck



It forms amides with active metals like  $Na, K$  etc.



It forms complexes with many substances, e.g.,

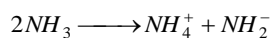


Its aqueous solution is weakly basic due to the formation of  $OH^-$  ions,  $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$

With sodium hypochlorite in presence of glue or gelatine, excess of ammonia gives hydrazine

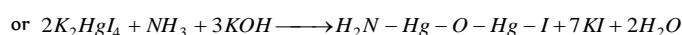
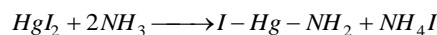
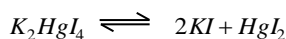


It undergoes self ionization in liquid state and acts as a solvent.



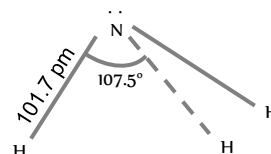
Many polar compounds are soluble in liquid ammonia.

With Nessler's reagent (an alkaline solution of  $K_2HgI_4$ ), ammonia and ammonium salts give a brown precipitate due to the formation of Millon's base.



It is used as a refrigerant and in the manufacture of fertilizers.

**Structure of  $NH_3$  :** The  $N$  atom in  $NH_3$  is  $sp^3$ -hybridized containing a lone pair of electrons due to which the  $H-N-H$  bond angle is  $107.5^\circ$ . As a result  $NH_3$  molecule is pyramidal.



(2) **Oxides of nitrogen :** Nitrogen combines with  $O_2$  under

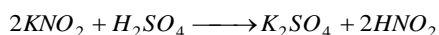
different conditions to form a number of binary oxides which differ with respect to the oxidation state of the nitrogen atom. The important oxides are  $N_2O, NO, N_2O_3, NO_2, N_2O_4$  and  $N_2O_5$ .  $N_2O$  and  $NO$  both are neutral. Nitrous oxide ( $N_2O$ ) has a sweet taste and its main use is as anaesthetic. When inhaled in mild quantities it causes hysterical laughter so it is also called Laughing gas. Nitric oxide ( $NO$ ) can be obtained by treating a mixture of sodium nitrite and ferrous sulphate with dil.  $H_2SO_4 \cdot N_2O_5$  is the strongest oxidising agent.

Table : 18.4 Oxides of Nitrogen

Oxide	Structure	Physical appearance	Preparation
Nitrous oxide ( $N_2O$ ) +1	$N \equiv N \rightarrow O$	Colourless gas	By heating ammonium nitrate upto $240^\circ C$ $NH_4NO_3 \xrightarrow{\Delta} N_2O + 2H_2O$ , It is Collected over hot water
Nitric oxide ( $NO$ ) +2	$N = O$	Colourless	(a) By the action of cold dil. $HNO_3$ on copper turnings (Laboratory method) $3Cu + 8 \text{ dil. } HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$ (b) By the action of $H_2SO_4$ on a mixture of $FeSO_4$ and $KNO_3$ (4:1) $2KNO_3 + 5H_2SO_4 + 6FeSO_4 \rightarrow 2KHSO_4 + 3Fe_2(SO_4)_3 + 4H_2O + 2NO$ (c) By catalytic oxidation of ammonia. $4NH_3 + 5O_2 \xrightarrow[850^\circ C]{Pt} 4NO + 6H_2O$
Dinitrogen trioxide ( $N_2O_3$ ) +3	$O = N - N \begin{matrix} \nearrow O \\ \searrow O \end{matrix}$	Blue solid	(a) By the action of 50% $HNO_3$ on arsenious oxide. $2HNO_3 + As_2O_3 + 2H_2O \rightarrow NO + NO_2 + 2HAsO_2$ $\downarrow 250 K$ $NO_2$
Dinitrogen tetroxide ( $N_2O_4$ ) +4	$O = N - N \begin{matrix} \nearrow O \\ \searrow O \\ \swarrow O \\ \nwarrow O \end{matrix}$	Colourless liquid	(a) By heating nitrates of heavy metals, e.g., lead nitrate. $2Pb(NO_3)_2 \xrightarrow{673 K} 4NO_2 + 2PbO + 2O_2$
Nitrogen dioxide ( $NO_2$ ) +4	$O = N \begin{matrix} \nearrow O \\ \searrow O \end{matrix}$	Brown gas	(b) By heating copper turnings with conc. $HNO_3$ . $Cu + 4 HNO_3 (\text{conc.}) \rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$
Dinitrogen pentoxide ( $N_2O_5$ ) +5	$O = N \begin{matrix} \nearrow O \\ \searrow O \end{matrix} - O - N \begin{matrix} \nearrow O \\ \searrow O \end{matrix}$	Colourless gas	(a) By dehydrating $HNO_3$ with phosphorus pentoxide $4HNO_3 + P_2O_5 \rightarrow 2N_2O_5 + 4HPO_3$

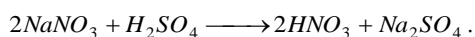
(3) **Oxyacids of nitrogen** : Oxyacids of nitrogen are  $HNO_2$ ,  $HNO_3$ ,  $H_4N_2O_4$  and  $HNO_4$ , which are explosive.  
(Nitroxyl acid) (Pernitric acid)

(i) **Nitrous acid ( $HNO_2$ )** : It is prepared by adding ice cold dil.  $HCl$  or dil.  $H_2SO_4$  to a well cooled solution of any nitrite ( $NaNO_2$ ,  $Ba(NO_2)_2$  etc.).

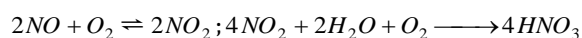
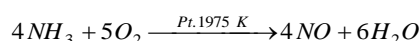


It oxidises  $H_2S$  to  $S$ ,  $KI$  to  $I_2$  and acts as a reducing agent in presence of strong oxidising agent, i.e., it reduces acidified  $KMnO_4$ ,  $K_2Cr_2O_7$ ,  $H_2O_2$  etc. to  $Mn^{2+}$ ,  $Cr^{3+}$  and  $H_2O$  respectively.

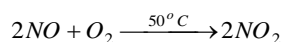
(ii) **Nitric acid ( $HNO_3$ )** :  $HNO_3$  is called aqua fortis. It is prepared in the laboratory by distillation of nitre with conc.  $H_2SO_4$ .



Commercially, it is obtained by Ostwald's process. In this process,  $NH_3$  is first catalytically oxidised to  $NO$  which is cooled to about  $300K$  and then oxidised by air to  $NO_2$ . Absorption of  $NO_2$  in water in presence of oxygen gives  $HNO_3$



From air (Birkeland Eyde electric arc process)



**Properties** : It is a very strong acid and decomposes on boiling or in presence of sunlight. It acts as a strong oxidising agent. It oxidises nonmetals and metalloids to their respective oxy-acids, i.e.,  $C$  to  $H_2CO_3$ ,  $S$  to  $H_2SO_4$ ,  $P$  to  $H_3PO_4$ ,  $I_2$  to  $HIO_3$ ,  $As$  to  $H_3AsO_4$  (arsenic acid) and  $Sb$  to  $H_3SbO_4$  (antimonic acid), while nitric acid itself is reduced to  $NO_2$ .



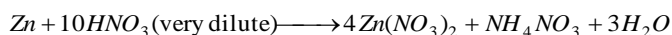
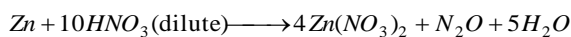
Nitric acid reacts with metals to form nitrates and is itself reduced to  $NO$ ,  $N_2O$ ,  $NO_2$  or  $NH_3$  (which further reacts with  $HNO_3$  to give



$NH_4NO_3$ ) depending upon the concentration of the acid, activity of the metal and the temperature of the reaction.

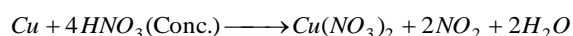
(i) Very active metals such as  $Mn, Mg, Ca$ , etc. give  $H_2$  on treatment with very dilute  $HNO_3$  (2%).

(ii) Less active metals like  $Cu, Hg, Ag, Pb$  etc. give  $NO$  with dil.  $HNO_3$ . Zinc, however, gives  $N_2O$  with dil  $HNO_3$  and  $NH_4NO_3$  with very dilute  $HNO_3$ .



Similarly,  $Fe$  and  $Sn$  react with dilute nitric acid to give  $NH_4NO_3$ .

(iii) Conc.  $HNO_3$  gives  $NO_2$  both with active metals ( $Zn, Pb$  etc.) and less active metals ( $Cu, Hg, Ag$  etc.)

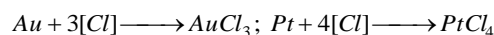
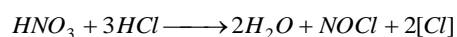


Tin is, however, oxidized by conc.  $HNO_3$  to metastannic acid ( $H_2SnO_3$ ).

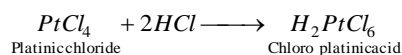
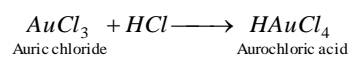


**Passivity** :  $Fe, Cr, Ni$  and  $Al$  become passive in conc.  $HNO_3$  (i.e., lose their normal reactivity) due to the formation of a thin protective layer of the oxide on the surface of the metal which prevents further action.

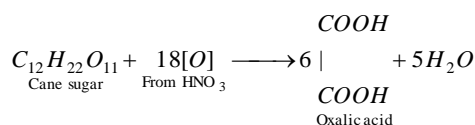
Nitric acid has no action on noble metals ( $Au, Pt$ ) but these metals dissolve in aqua regia (3 vol.  $HCl$  + 1 vol.  $HNO_3$ ) forming their respective chlorides.



These chlorides subsequently dissolve in excess of  $HCl$  forming their corresponding soluble complexes. Thus,



Sugar on oxidation with nitric acid gives oxalic acid. Nitric acid reacts with glycerine to give glycerol trinitrate or nitro glycerine, with toluene it gives 2, 4, 6-trinitrotoluene (T.N.T.) and with cellulose (cotton) it gives cellulose trinitrate (gun cotton). All these are used as explosives.



**Table : 18.5 Oxyacids of nitrogen**

Name of oxoacid	F.	Structure	Oxidation State of N	Basicity	pK <sub>a</sub>	Nature
Hyponitrous acid	$HNO$	$\begin{array}{c} \cdot\cdot \\   \\ N - OH \\    \\ HO - N \\ \cdot\cdot \end{array}$	+1	2 (dibasic)	Very weak	Highly explosive
Nitrous acid	$HNO$	$\begin{array}{c} H - N = O \\   \\ O \end{array}$	+3	1 (monobasic)	3.3	Unstable, Weak acid
Nitric acid	$HNO$	$\begin{array}{c} H - O - N = O \\   \\ O \end{array}$	+5	1 (monobasic)	-3.0	Stable, Strong acid
Pernitric acid	$HNO$	$\begin{array}{c} O = N - O - O - H \\   \\ O \end{array}$	+5	1 (monobasic)		Unstable and explosive

### Phosphorus and its compounds

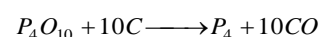
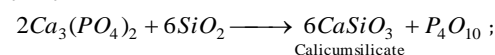
It is the second member of group 15 (VA) of the Periodic table. Due to larger size of  $P$ , it can not form stable  $P\pi - P\pi$  bonds with other phosphorous atoms where as nitrogen can form  $P\pi - P\pi$  bonds.

(i) **Occurrence** : Phosphorous occurs mainly in the form of phosphate minerals in the crust of earth. Some of these are :

- (i) Phosphorite  $Ca_3(PO_4)_2$ , (ii) Fluorapatite  $Ca_5(PO_4)_3F$ ,  
 (iii) Chlorapatite  $3Ca_3(PO_4)_2 \cdot CaCl_2$ , (iv) Hydroxyapatite;

$Ca_5(PO_4)_3OH$ . Phosphates are essential constituents of plants and animals. It is mainly present in bones, which contains about 58% calcium phosphate.

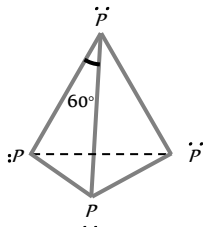
(2) **Isolation** : Elemental phosphorus is isolated by heating the phosphorite rock with coke and sand in an electric furnace at about 1770K,



(3) **Allotropic forms of phosphorus** : Phosphorus exists in three main allotropic forms,

(i) White phosphorus, (ii) Red phosphorus, (iii) Black phosphorus

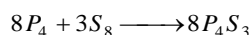
(i) **White or yellow phosphorus** : It is obtained from phosphate rock or phosphorite as explained above. It exists as  $P_4$  units where four  $P$  atoms lie at the corners of a regular tetrahedron with  $\angle PPP = 60^\circ$ . Each  $P$  atom is linked to three other  $P$  atoms by covalent bonds. there are total six bonds and four lone pairs of electrons present in a  $P_4$  molecule of white phosphorus.



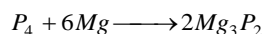
**Properties** : White phosphorus is extremely reactive due to strain in the  $P_4$  molecule, poisonous, soft, low melting ( $317K$ ) solid, soluble in  $CS_2$ , alcohols and ether. It has a garlic odour. Persons working with white  $P$  develop a disease known as Phossy jaw in which jaw bones decay. It turns yellow on exposure to light. Hence, it is also called yellow phosphorus.

It spontaneously catches fire in air with a greenish glow which is visible in the dark ( $P_4 + 3O_2 \rightarrow P_4O_6$ ). This phenomenon is called phosphorescence. Because of its very low ignition temperature ( $303K$ ), it is always kept under water.

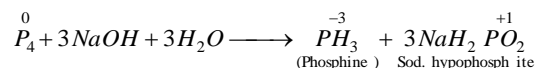
With sulphur it gives tetraphosphorus trisulphide with explosive violence which is used in "strike anywhere matches".



With metals phosphorus forms phosphides. For example,

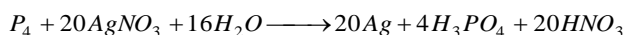
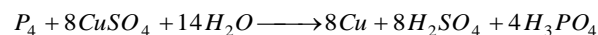


With aqueous alkalis, on heating, white phosphorus gives phosphine

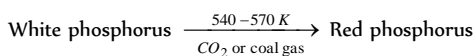


It is an example of a disproportionation reaction where the oxidation state of  $P$  decreases from 0 to  $-3$  (in  $PH_3$ ) and increases to  $+1$  (in  $NaH_2PO_2$ )

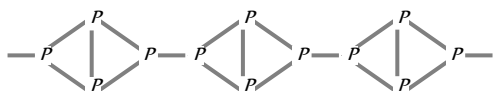
White phosphorus acts as a strong reducing agent. It reduces  $HNO_3$  to  $NO_2$  and  $H_2SO_4$  to  $SO_2$ . It also reduces solutions of  $Cu$ ,  $Ag$  and  $Au$  salts to their corresponding metals. For examples,



(ii) **Red phosphorus** : It is obtained by heating white phosphorus at  $540-570K$  out of contact with air in an inert atmosphere ( $CO_2$  or coal gas) for several hours.



Red phosphorus exists as chains of  $P_4$  tetrahedra linked together through covalent bonds to give a polymeric structure as shown.



Due to its polymeric structure, red phosphorus is much less reactive and has m.p. much higher than that of white phosphorus.

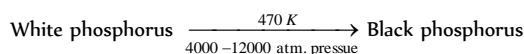
**Properties** : Red phosphorus is a hard, odourless, non poisonous solid, insoluble in organic solvents such as  $CS_2$ , alcohol and ether. Its ignition temperature is much higher than that of white phosphorus and thus does not catch fire easily. It does not show phosphorescence.

It sublimes on heating giving vapours which condense to give white phosphorus. It is denser than white phosphorus and is a bad conductor of electricity.

It burns in oxygen at  $565K$  to give phosphorus pentoxide, reacts with halogens, sulphur and alkali metals only when heated forming their corresponding salts.

It does not react with caustic alkalis and this property is made use in separating red phosphorus from white phosphorus.

(iii) **Black phosphorus** : It is obtained by heating white phosphorus at  $470K$  under high pressure ( $4000-12000\text{atm}$ ) in an inert atmosphere.

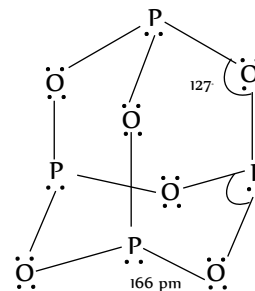


It has a double layered structure. Each layer is made up of zig-zag chains with  $P-P-P$  bond angle of  $99^\circ$ . Since it is highly polymeric, it has high density. It is the most stable (inactive) form of phosphorus and has a black metallic luster. It is a good conductor of heat and electricity.

#### (4) Compounds of phosphorus

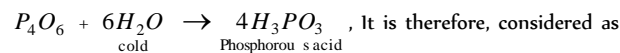
(i) **Oxides and oxyacids of phosphorus** : Phosphorus is quite reactive and forms number of compounds in oxidation states of  $-3$ ,  $+3$  and  $+5$ . Phosphorus forms two common oxides namely, (a) phosphorus trioxide ( $P_4O_6$ ) and (b) phosphorus penta oxide ( $P_4O_{10}$ ).

(a) **Phosphorus (III) oxide ( $P_4O_6$ )** :



It is formed when phosphorus trioxide is burnt in a limited supply of air,  $P_4 + 3O_2 \rightarrow P_4O_6$ . (limited)

It is a crystalline solid with garlic odour. It dissolves in cold water to give phosphorous acid,



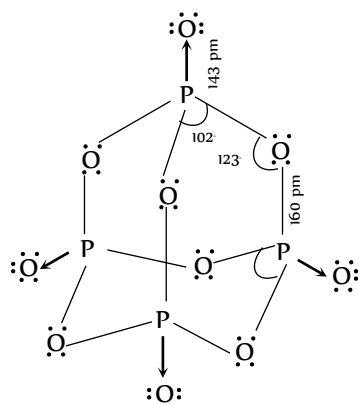
anhydride of phosphorus acid.

With hot water, it gives phosphoric acid and inflammable phosphine,  $P_4O_6 + 6H_2O \text{ (hot)} \rightarrow 3H_3PO_4 + PH_3$   
Phosphoric acid

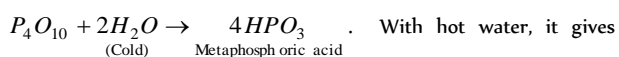
It reacts vigorously with  $Cl_2$  to form a mixture of phosphoryl chloride and meta phosphoryl chloride.



(b) **Phosphorus (V) oxide ( $P_4O_{10}$ )** :

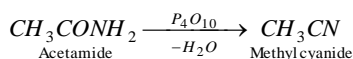
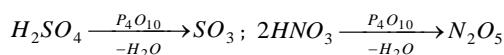


It is prepared by heating white phosphorus in excess of air,  $P_4 + 5O_2(\text{excess}) \rightarrow P_4O_{10}$ . It is snowy white solid. It readily dissolves in cold water forming metaphosphoric acid.



With hot water, it gives phosphoric acid,  $P_4O_{10} + 6H_2O \xrightarrow{\text{Hot}} 4H_3PO_4$ .

$P_4O_{10}$  is a very strong dehydrating agent. It extracts water from many compounds including  $H_2SO_4$  and  $HNO_3$ ,



(ii) **Oxyacids of phosphorus**: Phosphorus forms a number of oxyacids which differs in their structure and oxidation state of phosphorus. These are  $H_3PO_2, H_3PO_3, H_4P_2O_5, H_3PO_4, (HPO_3)_n, H_4P_2O_5, H_4P_2O_7$ . From these  $H_3PO_2, H_3PO_3$  are reducing agents.  $H_4P_2O_5$  (pyrophosphoric acid) is dibasic acid.

$(HPO_3)_n$  is formed by dehydration of  $H_3PO_4$  at  $316^\circ C$ .

Table : 18.6 Oxyacids of phosphorus

Name	Oxidation state of P and Basicity	Structure
Hypophosphorous acid $H_3PO_2$	+1 Monobasic	
Phosphorous acid $H_3PO_3$	+3 Dibasic	
Hypophosphoric acid $H_4P_2O_6$	+4 Tetrabasic	
Orthophosphoric acid $H_3PO_4$	+5 Tribasic	

Metaphosphoric acid $(HPO_3)_n$	+5 Monobasic	
Pyrophosphoric acid (Diphosphoric acid) $H_4P_2O_7$	+5 Tetrabasic	

(5) **Chemical Fertilizers**: The chemical substances which are added to the soil to keep up the fertility of soil are called fertilizers.

**Types of fertilizers**: Chemical fertilizers are mainly of four types,

(i) **Nitrogenous fertilizers**: e.g. Ammonium sulphate  $(NH_4)_2SO_4$ , Calcium cyanamide  $CaCN_2$ , Urea  $NH_2CONH_2$  etc.

(ii) **Phosphatic fertilizers**: e.g.  $Ca(H_2PO_4)_2 \cdot H_2O$  (Triple super phosphate), Phosphatic slag etc.

(iii) **Potash fertilizers**: e.g. Potassium nitrate  $(KNO_3)$ , Potassium sulphate  $(K_2SO_4)$  etc.

(iv) **Mixed fertilizers**: These are made by mixing two or more fertilizers in suitable proportion. e.g. *NPK* (contains nitrogen, phosphorus and potassium).

*NPK* is formed by mixing ammonium phosphate, super phosphate and some potassium salts.

## Oxygen Family

Oxygen is the first member of group 16 or VIA of the periodic table. It consists of five elements Oxygen (*O*), sulphur (*S*), selenium (*Se*), tellurium (*Te*) and polonium (*Po*). These (except polonium) are the ore forming elements and thus called **chalcogens**.

### (1) Electronic configuration

Elements	Electronic configuration ( $ns^2 np^4$ )
$8 O$	$[He] 2s^2 2p^4$
$16 S$	$[Ne] 3s^2 3p^4$
$34 Se$	$[Ar] 3d^{10} 4s^2 4p^4$
$52 Te$	$[Kr] 4d^{10} 5s^2 5p^4$
$84 Po$	$[Xe] 4f^{14} 5d^{10} 6s^2 6p^4$

### Physical properties

(1) **Physical state**: Oxygen is gas while all other are solids.

(2) **Atomic radii**: Down the group atomic radii increases because the increases in the number of inner shells overweighs the increase in nuclear charge.

(3) **Ionisation energy**: Down the group the ionisation energy decrease due to increase in their atomic radii and shielding effect.

(4) **Electronegativity**: Down the group electronegativity decreases due to increase in atomic size.

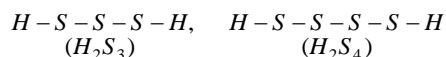
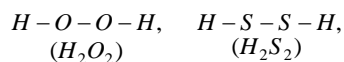
(5) **Electron affinity**: Element of this group have high electron affinity, electron affinity decreases down the group.

(6) **Non – metallic and metallic character**: These have very little metallic character because of their higher ionisation energies.

(7) **Nature of bonding** : Compound of oxygen with non metals are predominantly covalent. *S*, *Se*, and *Te* because of low electronegativities show more covalent character.

(8) **Melting and boiling points** : The melting point and boiling points increases on moving down the group.

(9) **Catenation** : Oxygen has some but sulphur has greater tendency for catenation.



(10) **Allotropy**

Oxygen	-	$O_2$ and $O_3$
Sulphur	-	Rhombic, monoclinic, plastic sulphur
Selenium	-	Red (non-metallic) grey (metallic)
Tellurium	-	Non-metallic and metallic (more stable)
Polonium	-	$\alpha$ and $\beta$ (both metallic)

(ii) **Oxidation states** : Oxygen shows - 2, + 2 and -1 oxidation states. Other elements show +2, +4 and +6 oxidation states.

**Chemical properties**

(i) **Hydrides** : The elements of this group form hydrides such as  $H_2O, H_2S, H_2Se, H_2Te$  and  $H_2Po$ . Following are their characteristics.

(i) **Physical states** : Water is colourless and odourless while hydrides of the rest of the elements of this group are colourless, unpleasant smelling poisonous gases.

(ii) **Volatile nature** : Volatility increases from  $H_2O$  to  $H_2S$  and then decreases. The low volatility and abnormally high boiling point of water is due to the association of water molecules on account of hydrogen bonding because of strongly electronegative oxygen atom linked to hydrogen atom. thus, water is liquid while  $H_2S$  and other hydrides are gases under normal condition of temperature and pressure.

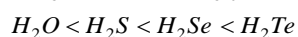
(iii) **Acidic character** : The hydrides of this group behave as weak diprotic acids in aqueous solution, the acidic character increasing from  $H_2S$  to  $H_2Te$  when  $H_2O$  is neutral.

(iv) **Thermal stability** : The thermal stability decreases from  $H_2O$  to  $H_2Po$  because the size of the central atom (from *O* to *Po*) increases resulting in longer and weaker *M-H* bond consequently the bond strength decreases. This results in the decrease of the thermal stability.

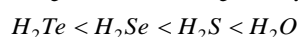
(v) **Reducing character** : The reducing power of the hydrides increases from  $H_2O$  to  $H_2Po$  due to the decreasing bond strength from  $H_2O$  to  $H_2Po$ .

(vi) **Bond angle** : All these hydrides are angular molecules and the bond angle  $H-X-H$  (*X* is *O, S, Se, Te*) decreases from  $H_2O$  to  $H_2Te$ .

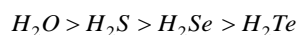
Increasing order of reducing power of hydrides :



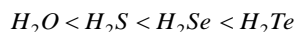
Increasing order of bond angles in hydrides :



The order of stability of hydrides :



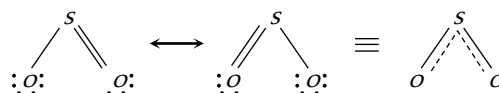
The order of increasing acidic nature of hydrides :



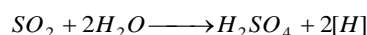
(2) **Oxides** : These elements form monoxides (*MO*), dioxides ( $MO_2$ ) and trioxides ( $MO_3$ ).

(i) **Dioxides** : Sulphur, selenium and tellurium burn in air to form  $SO_2, SeO_2$  and  $TeO_2$ . The dioxide molecules contain  $p\pi-p\pi$  bonds which become weaker with increase in atomic number because of the increase in the bond length.

(a) Sulphur dioxide,  $SO_2$  is a gas at room temperature and exists as individual molecules even in the solid state. Its molecule has bent structure and is a resonance hybrid of the following canonical structures.



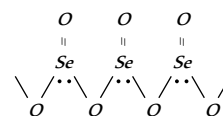
$SO_2$  is acidic in nature and also called the anhydride of sulphurous acid. It can act as reducing and oxidising agent.  $SO_2$  also acts as a bleaching agent in the presence of moisture, but in contrast to  $Cl_2$ , its bleaching action is temporary.



Colouring matter +2[H] = Colourless compound

Hence,  $SO_2$  bleaches due to reduction and the bleaching action is temporary.

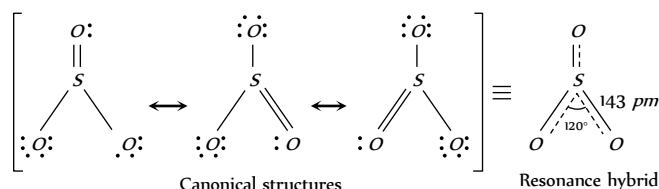
(b) Selenium dioxide,  $SeO_2$  is a solid with polymeric zig-zag structure at room temperature however it exist as discrete molecules in the gaseous phase.



(c) Tellurium dioxide,  $TeO_2$  is also a solid with polymeric zig-zag structure at room temperature very similar to that of selenium dioxide.

(ii) **Trioxides** : Sulphur, selenium and tellurium can form trioxides also.

(a) **Sulphur trioxide,  $SO_3$**  : In the gaseous state monomeric  $SO_3$  has a planar structure with *S-O* bond distance of 143 pm and *O-S-O* bond angle of  $120^\circ$ .  $SO_3$  molecule is a resonance hybrid of following structures.



In the solid phase sulphur trioxide polymerises to cyclic trimer or to a stable linear chain structure.  $SO_3$  is the anhydride of  $H_2SO_4$ . It is acidic in nature and acts as oxidising agent.

(b) **Selenium trioxide,  $SeO_3$**  : it is a solid substance which exists as a cyclic tetramer, however in the vapour phase it exists as a monomer

(c) **Tellurium trioxide,  $TeO_3$**  : It is a solid at room temperature existing as a polymer.

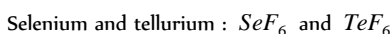
The increasing order of acidic nature of oxides is  $TeO_3 < SeO_3 < SO_3$ .



(3) **Oxyacids :**



(4) **Halides :** Oxygen :  $OF_2, Cl_2O, Br_2O$



**Anomalous Behaviour of Oxygen**

Oxygen is the first member of the group 16 family and differs from the other members of the family because of

- (1) Its small size
- (2) Its high electronegativity
- (3) Its high ionisation energy
- (4) Absence of  $d$ -orbitals in the valence shell

It differs from the other members of the family as follows

(1) **Elemental state :** Oxygen is a diatomic gas while others are octatomic solids with eight membered puckered ring structure.

(2) **Oxidation states :** Oxygen shows O.S. of  $-2$  in most of its compounds. It also shows an O. S. of  $+2$  in  $F_2O$  and  $-1$  in  $H_2O_2$  or other peroxides. It cannot show O.S. beyond 2. Other elements show oxidation states of  $+2, +4$  and  $+6$  because these elements have vacant  $d$ -orbitals so that their valence shell can expand.

(3) **Hydrogen-bonding :** Oxygen atom is very small and has quite high nuclear charge. therefore, it has high value of electronegativity and is able to form  $H$ -bonds. the other elements, because of their large size, cannot form  $H$ -bonds. As a result,  $H_2O$  is liquid while  $H_2S$  is a gas and  $H_2Se$  etc., are solids.

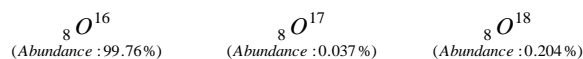
(4) **Maximum covalency :** Oxygen has a maximum covalency of two while other elements can show a maximum covalency of six. This is because these elements have vacant  $d$ -orbitals while oxygen has not.

(5) **Types of compounds :** The compounds of oxygen are mainly ionic and polar covalent due to high electronegativity of oxygen while those of others are not.

(6) **Magnetic character :** Oxygen is paramagnetic while others are not.

**Oxygen and its compounds**

Oxygen is the most abundant element in the earth crust (46.5%). It was discovered by **Karl Scheele** and **Joseph Priestley**. It occurs in three isotopic forms :

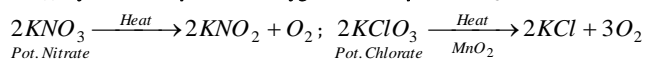


Out of the three isotopes,  ${}_8O^{18}$  is radioactive.

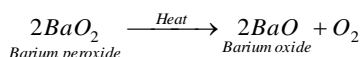
**Occurrence :** In free state, it occurs in air and constitutes 21% by volume of air.

**Preparation of Dioxygen :** Oxygen is prepared by the following methods.

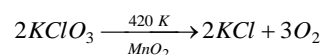
(1) **By the decomposition of oxygen rich compounds :** e.g.



(2) **By heating dioxides, Peroxides and higher oxides :** e.g.



(3) **Laboratory Method :** In the laboratory,  $O_2$  is prepared by thermal decomposition of potassium chlorate.



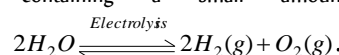
In the absence of  $MnO_2$  catalyst, the decomposition takes place at 670-720 K. Therefore,  $MnO_2$  acts as a catalyst and also lowers the temperature for the decomposition of  $KClO_3$ .

(4)  $O_2$  can also be prepared by the action of water on sodium peroxide as,  $2Na_2O_2 + 2H_2O \rightarrow 4NaOH + O_2$ .

(5) **Industrial preparation :** The main sources for the industrial preparation of dioxygen are air and water.

(i) **From air :**  $O_2$  is prepared by fractional distillation of air. During this process,  $N_2$  with less boiling point (78 K) distills as vapour while  $O_2$  with higher boiling point (90 K) remains in the liquid state and can be separated.

(ii) **From water :**  $O_2$  can also be obtained by the electrolysis of water containing a small amount of acid or alkali,



**Physical properties of  $O_2$  :** It is a colourless, tasteless and odourless gas. It is slightly soluble in water and its solubility is about  $30\text{ cm}^3$  per litre of water at 298 K.

**Table : 18.7 Physical properties of atomic and molecular oxygen**

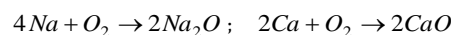
Atomic properties	Molecular properties
Atomic radius (pm) – 73	Bond length (pm) – 120.7
Ionic radius O (pm) – 140	Bond energy (kJ mol <sup>-1</sup> ) – 493
Electronegativity – 3.5	Density at S.T.P. (gcm <sup>-3</sup> ) – 1.429
Ionisation energy (kJ mol <sup>-1</sup> ) – 1310	Melting point (K) – 54.4
Electron affinity (kJ mol <sup>-1</sup> ) – 140	Boiling point (K) – 90.2

**Chemical properties of  $O_2$  :** It does not burn itself but helps in burning. It is quite stable in nature and its bond dissociation energy is very high. Therefore, it is not very reactive as such,  $O_2 \rightarrow O + O$ .

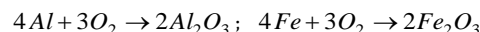
Therefore, dioxygen reacts at higher temperatures. However, once the reaction starts, it proceeds of its own. This is because the chemical reactions of dioxygen are exothermic and the heat produced during the reaction is sufficient to sustain the reactions.

(1) **Action with litmus :** Like dihydrogen, it is also neutral and has no action on blue or red litmus.

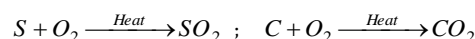
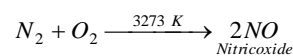
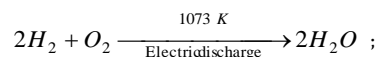
(2) **Reaction with metals :** Active metals like  $Na, Ca$  react at room temp. to form their respective oxides.



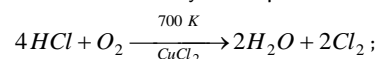
It reacts with  $Fe, Al, Cu$  etc. metals at high temperature

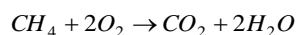
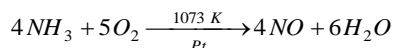


(3) **Action with Non-metals :** It form oxides.



(4) **Reaction with compounds :** Dioxygen is an oxidising agent and it oxidises many compounds under specific conditions. e.g.





### Uses of dioxygen

- (1) It is used in the oxy-hydrogen or oxy-acetylene torches which are used for welding and cutting of metals.
- (2) It is used as an oxidising and bleaching agent,
- (3) Liquid  $O_2$  is used as rocket fuel.
- (4) It is used in metallurgical processes to remove the impurities of metals by oxidation.

### Compounds of Oxygen

(i) **Oxides** : A binary compound of oxygen with another element is called oxide. On the basis of acid-base characteristics, the oxides may be classified into the following four types,

(i) **Basic oxides** : Alkali, alkaline earth and transition metals form basic oxides -  $Na_2O, MgO, Fe_2O_3$  etc. their relative basic character decreases in the order : alkali metal oxides > alkaline earth metal oxides > transition metal oxides.

(ii) **Acidic oxides** : Non-metal oxides are generally acidic -  $CO_2, SO_2, SO_3, NO_2, N_2O_5, P_4O_{10}, Cl_2O_7$  etc.

(iii) **Amphoteric oxides** :  $Al_2O_3, SnO_2$  etc.

(iv) **Neutral oxides** :  $H_2O, CO, N_2O, NO$  etc.

**Trends of oxides in the periodic Table** : On moving from left to the right in periodic table, the nature of the oxides change from basic to amphoteric and then to acidic. For example, the oxides of third period has the following behaviour,

$Na_2O$ strongly basic	$MgO$ basic	$Al_2O_3$ amphoteric	$SiO_2$ weakly acidic	$P_4O_{10}$ acidic	$SO_2$ strongly acidic	$Cl_2O_7$ very strongly acidic
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Basic to acidic character increases

However, on moving down a group, acidic character of the oxides decreases. For example in the third group, the acidic character of oxides decreases as:

$B_2O_3$ acidic	$Al_2O_3$ amphoteric	$Ga_2O_3$ (weakly basic)	$In_2O_3, Tl_2O_3$ basic
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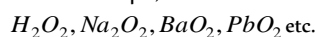
Acidic to basic character increases

On the basis of oxygen content the oxides may be classified into the following types,

**Normal oxides** : These contain oxygen atoms according to the normal oxidation number *i.e.* - 2. For example,  $MgO, H_2O, CaO, Li_2O, Al_2O_3$  etc.

**Polyoxides** : These contain oxygen atoms more than permitted by the normal valency. Therefore, these contain oxygen atoms in oxidation state different than -2.

**Peroxides** : These contain  $O_2^{2-}$  ion having oxidation number of oxygen as -1. For example,



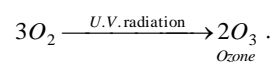
**Superoxides** : These contain  $O_2^-$  ion having oxidation number of oxygen as -1/2. For example,  $KO_2, PbO_2$ , etc.

**Suboxides** : These oxides contain less oxygen than expected from the normal valency. For example,  $N_2O$ .

**Mixed oxides** : These oxides are made up of two simple oxides. For example, red lead  $Pb_3O_4 (2PbO_2 + PbO_2)$ , magnetic oxide of iron,  $Fe_3O_4 (FeO + Fe_2O_3)$  and mixed oxide of manganese,  $Mn_3O_4 (MnO_2 + 2MnO)$ .

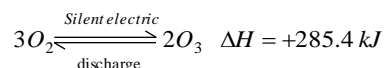
### Ozone or trioxygen

Ozone is an allotrope of oxygen. It is present in the upper atmosphere, where it is formed by the action of U. V. radiations on  $O_2$ ,



$O_3$  protects us from the harmful U. V. radiations which causes skin cancer. Now a days, ozone layer in the atmosphere is depleting due to  $NO$  released by supersonic aircrafts and chlorofluoro carbons (CFC'S) *i.e.* freon which is increasingly being used in aerosols and as a refrigerant.

**Preparation** : Ozone is prepared by passing silent electric discharge through pure, cold and dry oxygen in a specially designed apparatus called ozoniser. The formation of ozone from oxygen is an endothermic reaction.



Ozone is prepared in the laboratory by the following two types of ozonisers,

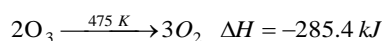
- (a) Siemen's ozoniser, (b) Brodie's ozoniser

**For the better yield of ozone** : (a) Only pure and dry oxygen should be used. (b) The ozoniser must be perfectly dry. (c) A fairly low temperature ( $\approx 273 K$ ) must be maintained. (d) The electric discharge must be sparkless.

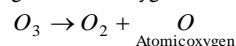
**Physical properties** : Ozone is a light blue coloured gas, having pungent odour. It is heavier than air. Its vapour density is 24. It is slightly soluble in water.

**Chemical properties** : The important chemical properties of ozone are discussed below,

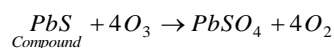
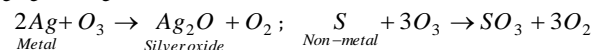
(1) **Decomposition** : Pure ozone decomposes on heating above 475 K to form  $O_2$  gas.



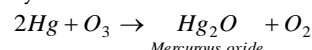
(2) **Oxidising agent** : Ozone is one of the most powerful oxidising agent with the liberation of dioxygen. In fact, ozone is a stronger oxidising agent than molecular oxygen because ozone has higher energy content and decomposes to give atomic oxygen as:



Therefore, ozone oxidises a number of non-metals and other reducing agents. *e.g.*

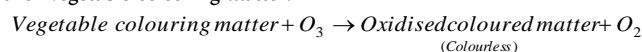


Mercury is oxidised to mercurous oxide,



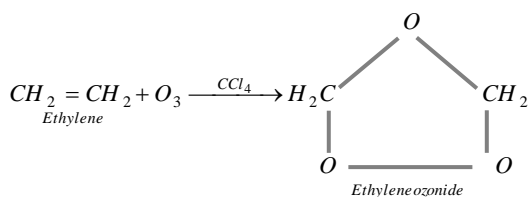
During this reaction mercury loses its meniscus and starts sticking to the sides of the glass. This is known as tailing of mercury. Mercurous oxide formed in this reaction dissolves in mercury and starts sticking to the glass surface.

(3) **Bleaching agent** : Due to the oxidising action of ozone, it acts as a mild bleaching agent as well as a sterilizing agent. It acts as a bleaching agent for vegetable colouring matter.

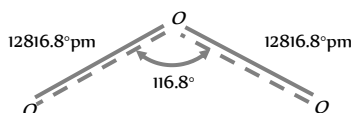


For example, ozone bleaches indigo, ivory, litmus, delicate fabrics etc.

(4) **Formation of ozonides** : Ozone reacts with alkenes in the presence of  $CCl_4$  to form an ozonide. e.g.



**Structure of  $O_3$**  : The structure of  $O_3$  molecule is angular as shown in fig. The  $O-O-O$  bond angle is  $116.8^\circ$  and  $O-O$  bond length is 128 pm.



#### Uses of ozone

- (1)  $O_3$  is used for disinfecting water for drinking purposes because ozone has germicidal properties.
- (2) It is used for purifying air of crowded places such as cinemas, under ground railway, auditoriums, tunnels, mines etc.
- (3) It is used in industry for the manufacture of  $KMnO_4$ , artificial silk, synthetic camphor etc.

#### Sulphur and its compounds

Sulphur is the second member of oxygen family and belongs to group-16 (VI A) of the periodic table.

**Occurrence** : Sulphur occurs in the earth's crust to the extent of 0.05%. It occurs in the free state as well as in combined state. Sulphur occurs mainly as sulphides and sulphates. eg.

Table 18.8

Sulphide Ores	Sulphate Ores
Iron pyrites (fool's gold) - $FeS_2$	Gypsum - $CaSO_4 \cdot 2H_2O$
Galena - $PbS$	Epsom salt - $MgSO_4 \cdot 7H_2O$
Copper pyrites - $CuFeS_2$	Barytes - $BaSO_4$
Cinnabar - $HgS$	Zinc blende - $ZnS$

**Extraction of sulphur** (Frasch process) : Sulphur is generally extracted from underground deposits by drilling three concentric pipes upto the beds of sulphur (700 - 1200 feet deep).

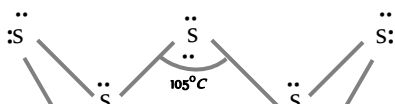
**Allotropy in sulphur** : Sulphur exists in four allotropic forms,

(1) **Rhombic or octahedral or  $\alpha$ -sulphur** : It is a bright yellow solid, soluble in  $CS_2$  and stable at room temp. All other varieties of sulphur gradually change into this form on standing.

(2) **Monoclinic sulphur or prismatic or  $\beta$ -sulphur** : It is prepared by melting the sulphur and then cooling it till a  $\beta$  is formed. On removing the crust, needle shaped crystals of monoclinic sulphur separate out. It is dull yellow in colour, soluble in  $CS_2$  and stable only above 369K. Below this temperature it changes into rhombic form.

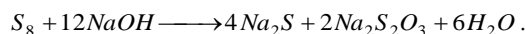
Thus, at 369K both these varieties co-exist. This temperature is called transition temperature and the two sulphurs are called enantiotropic substances. It also exist as molecules similar to that of rhombic sulphur but the symmetry of the crystals is different.

(3) **Plastic or amorphous or  $\gamma$ -sulphur** : It is a super cooled liquid insoluble in  $CS_2$ , soft and amorphous. It consists of long zig-zag chains of S-atoms.



(4) **Colloidal or  $\delta$ -sulphur** : It is prepared by passing  $H_2S$  through a solution of an oxidizing agent or water or by treating sodium thiosulphate with dil.  $HCl$ .

**Properties of sulphur** : It burns in air with, a blue flame forming  $SO_2$ , gives sulphur hexafluoride with  $F_2$  and sulphur mono chloride with  $Cl_2$ , sulphides with metals like  $Na$ ,  $Ca$ ,  $Zn$ ,  $Hg$ ,  $Fe$ ,  $Cu$  etc., reduces  $HNO_3$  to  $NO_2$  and  $H_2SO_4$  to  $SO_2$ . With  $NaOH$  solution on heating,

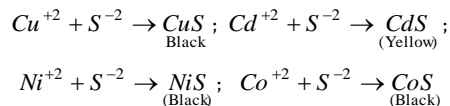


It gives sodium sulphide and sodium thiosulphate, with excess of sulphur,  $2Na_2S + S_8 \longrightarrow 2Na_2S_5$ .

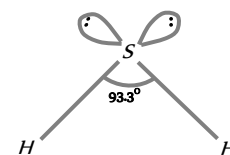
**Uses of sulphur** : It is used in the manufacture of matches, gun powder (mixture of charcoal, sulphur and potassium nitrate), explosives and fire works  $SO_2$ ,  $H_2SO_4$ ,  $CS_2$  and dyes, sulphur drugs and ointment for curing skin diseases and in the vulcanization of rubber.

#### Compounds of Sulphur

(1) **Hydrogen Sulphide** : It is prepared in the laboratory by the action of dil.  $H_2SO_4$  on ferrous sulphide in kipp's apparatus,  $FeS + H_2SO_4 \rightarrow FeSO_4 + H_2S$ . It is colourless gas having foul smell resembling that of rotten eggs. It reacts with many cations (of group II and IV) to give coloured sulphides,

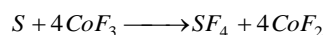


The solubility of sulphides can be controlled by the  $H^+$  ions concentration and therefore,  $H_2S$  finds extensive use in qualitative analysis of cation radicals.



(2) **Halides of sulphur** : Two important halides of sulphur are  $SF_4$  and  $SF_6$ .

(i) **Sulphur tetrafluoride** :  $SF_4$  is formed by the reaction of sulphur with  $CoF_3$ .

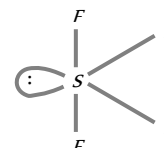


It is a colour gas which is quite reactive. It is hydrolysed with water.



It is used for fluorinating inorganic and organic compounds.

**Structure** : It has see-saw structure with  $sp^3d$  hybridization and is derived from trigonal bipyramid geometry in which an equatorial position is occupied by a lone pair of electrons.

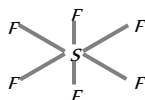


(ii) **Sulphur hexafluoride** :  $SF_6$  is prepared by burning sulphur in a stream of fluorine.  $OF_6$  is not known though sulphur forms  $SF_6$ . This is because oxygen has no  $d$ -orbitals in its valence shell.

$SF_6$  is a colourless gas. It is extremely inert substance even at red heat. It does not react with water. on account of its chemical inertness and dielectric strength, it is used as an insulator in high voltage generators and switch-gears.

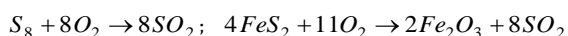
**Structure** : It has an octahedral structure with  $sp^3d^2$ -hybridisation around the central sulphur atom.

Therefore, all  $S-F$  bond distances are equal in its structure.



(3) **Oxides of sulphur** : Sulphur forms several oxides of which sulphur dioxide ( $SO_2$ ) and sulphur trioxide ( $SO_3$ ) are most important.

(i) **Sulphur dioxide ( $SO_2$ )** : It is prepared by burning sulphur or iron pyrites in air.

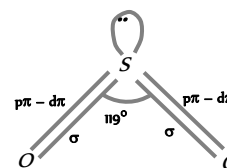


In laboratory, it is prepared by heating copper turnings with conc.  $H_2SO_4$

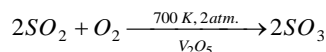


It is a colourless gas with irritating and suffocating smell.

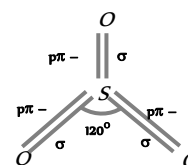
$SO_2$  molecule has a bent structure with a  $O-S-O$  bond angle of  $119^\circ$ . Sulphur is  $sp^2$  hybridized.



(ii) **Sulphur trioxide ( $SO_3$ )**: It is formed by the oxidation of  $SO_2$ .



In the gaseous phase, it exists as planar triangular molecular species involving hybridization of the  $S$ -atom. It has three  $S-O$   $\sigma$  bonds and three  $S-O$   $\pi$  bonds. The  $O-S-O$  bond angle is of  $120^\circ$ .



(4) **Oxyacids of sulphur** : Sulphur forms many oxyacids. Some of these are,

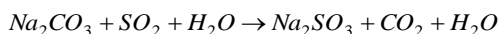
**Table : 18.9 Oxyacids of sulphur**

Formula	Name	Important properties	Structural formula
$H_2SO_3$ (+4)	Sulphurous acid	Free acid does not exist diprotic, strong reducing agent	$O = \overset{\cdot\cdot}{S} - OH$   $OH$
$H_2SO_4$ (+6) (Oil of vitriol)	Sulphuric acid	Stable diprotic, dehydrating agent	$O$    $O = S - OH$   $OH$
$H_2S_2O_3$ (-2 and +6)	Thiosulphuric acid	Free acid does not exist but its salts e.g. $Na_2S_2O_3$ All quite stable reducing agent	$S$    $O = S - OH$   $OH$
$H_2S_2O_4$ (+3)	Dithionous acid		$O \quad O$       $HO - S - S - OH$
$H_2S_2O_6$ (+5)	Dithionic acid	Free acid is moderately stable but its salts are quite stable.	$O \quad O$       $O = S - S = O$     $OH \quad OH$
$H_2S_2O_7$ (+6) (Oleum)	Disulphuric acid (Pyrosulphuric acid)	Strong oxidising agent	$O \quad O$       $O = S - O - S = O$     $OH \quad OH$
$H_2SO_5$ (+6) (Caro's acid)	Peroxomonosulphuric acid (Its salts known as persulphates)	Stable crystalline solid, powerful oxidising agent	$O$    $HO - S - OOH$   $O$
$H_2S_2O_8$ (+6) (Marshals acid)	Peroxodisulphuric acid (its salts are known as disulphates)	Strong oxidising agent.	$O \quad O$       $O = S - O - O - S = O$     $OH \quad OH$

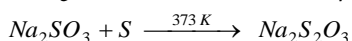


**Uses :**  $H_2SO_4$  is used (i) in the preparation of fertilizers like  $(NH_4)_2SO_4$  and super phosphate of lime, (ii) in lead storage batteries (iii) in preparation of dyes, paints and explosives (iv) in textile and paper industry (v) for tanning (vi) as a dehydrating agent.

(5) **Sodium thiosulphate**  $Na_2S_2O_3 \cdot 5H_2O$  : It is manufactured by saturating a solution of sodium carbonate with  $SO_2$  which gives a solution of sodium sulphite,



The resulting solution is boiled with powdered sulphur as,



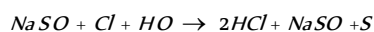
The solution is then cooled to get crystals of sodium thiosulphate.

**Physical properties :** (i) Sodium thiosulphate is a colourless crystalline solid. In the hydrated form, it is called hypo. (ii) It melts at 320 K and loses its water molecules of crystallization on heating to 490K.

#### Chemical properties

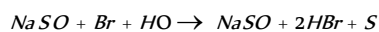
(i) **Action with halogens :** It reacts with halogens as,

(a) Chlorine water oxidizes sodium thiosulphate to sodium sulphate and sulphur is precipitated,

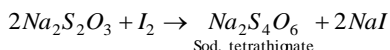


This property enables it to act as an antichlor in bleaching i.e. it destroys the unreacted chlorine in the process of bleaching.

(b) Bromine water also oxidizes sodium thiosulphate to sodium sulphate and sulphur,

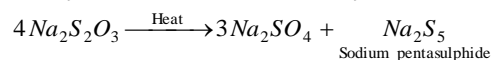


(c) With iodine it forms a soluble compound called sodium tetrathionate,

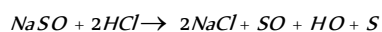


Therefore, hypo is commonly used to remove iodine stains from the clothes.

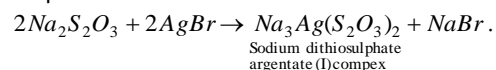
(ii) **Action of heat :** Upon heating, sodium thiosulphate decomposes to form sodium sulphate and sodium pentasulphide,



(iii) **Action with acids :** Sodium thiosulphate reacts with dilute hydrochloric acid or Sulphuric acid forming sulphur dioxide and sulphur. The solution turns milky yellow due to sulphur.



(iv) **Action with silver halides :** Sodium thiosulphate forms soluble complex when treated with silver chloride or silver bromide,



This property of hypo is made use in photography.

#### Uses of sodium thiosulphate

(i) It is largely used in photography as a fixing agent.

(ii) It is used as a preservative for fruit products such as jams and squashes.

(iii) It is used as an antichlor in bleaching.

(iv) It is used as a volumetric agent for the estimation of iodine.

(v) It is used in medicine.

## Halogen Family

Fluorine is the first member of group 17 or VIIA of the periodic table. It consists of five elements Fluorine ( $F$ ), Chlorine ( $Cl$ ), bromine ( $Br$ ), iodine ( $I$ ) and astatine ( $At$ ). These are known as halogen because their salts are found in sea water. Halogen is a greek word meaning a sea salt.

#### (1) Electronic configuration

Elements	Electronic configuration ( $ns^2 np^5$ )
${}_9F$	$[He]2s^2 2p^5$
${}_{17}Cl$	$[Ne]3s^2 3p^5$
${}_{35}Br$	$[Ar]3d^{10} 4s^2 4p^5$
${}_{53}I$	$[Kr]4d^{10} 5s^2 5p^5$
${}_{85}At$	$[Xe]4f^{14} 5d^{10} 6s^2 6p^5$

#### Physical properties

(1) **Atomic and ionic radii :** A halogen atom has the smallest radius as compared to any other element in its period. This is due to the increased effective nuclear charge which results in greater attraction of the electrons by the nucleus. The atomic radii increase from fluorine to iodine down the group due to increase in number of shells.

Element	$F$	$Cl$	$Br$	$I$
Covalent radius (pm)	72	99	114	133
Ionic radius (pm)	133	184	196	220

(2) **Ionization energy :** Ionization energy of these elements are higher than those of the corresponding elements of group 16 due to increased nuclear charge. These values decrease systematically as we move down the group from  $F$  to  $I$ .

Element	$F$	$Cl$	$Br$	$I$
I.E. ( $kJ mol^{-1}$ )	1680	1256	1142	1008

Thus iodine which has a comparatively low value of I.E., has a tendency to lose an electron to form positive iodinium ion,  $I^+$  and thus shown electropositive or metallic character.

(3) **Electronegativity :** Fluorine is the most electronegative element in the periodic table. With increase in atomic number down the group, the electronegativity decreases.

Element	$F$	$Cl$	$Br$	$I$	$At$
Electronegativity	4.0	3.2	3.0	2.7	2.2

The decreasing order of electronegativity is  $F > Cl > Br > I$

(4) **Electron affinity :** Electron affinity of chlorine, bromine and iodine decrease as the size of the atom increases. The electron affinity of fluorine is, however, lower than that of  $Cl$  and  $Br$ , because of its small size as a result of which inter-electronic repulsions present in its  $2p$  subshell are comparatively large. Thus chlorine has the highest electron affinity.

Element	$F$	$Cl$	$Br$	$I$
Electron affinity ( $kJ mol^{-1}$ )	333	348	325	296

The decreasing order of electron affinity is  $Cl > F > Br > I$

(5) **Oxidation states :** All the halogens show an oxidation state of  $-1$ . Fluorine being the most electronegative element always shows an oxidation state of  $-1$  while other halogens also show positive oxidation states up to a maximum of  $+7$  (i.e.  $+1, +3, +5$  and  $+7$ ) due to the availability of vacant  $d$  -

orbitals in the valence shell of these atoms. Some halogens also show +4 and +6 oxidation states in oxides and oxy acids.

(6) **Nature of bonds** : All the halogens have seven electrons in the valence shell and hence require one more electron to acquire the nearest inert gas configuration either by gaining an electron from the metallic atom to form halide,  $X^-$  ion, or by sharing an electron with an electronegative element. Thus, halogens form both ionic and covalent compounds. The halides of highly electropositive metals are ionic while those of weakly electropositive metals and non-metals are covalent. The tendency to form ionic compounds decreases from  $F$  to  $I$ . Thus,  $F$  because of its high electronegativity forms ionic compounds even with less electropositive metals like  $Hg, Bi, Sn$  etc. while other halogens form only covalent compounds.

(7) **Non-metallic character** : All the halogens are non-metallic in nature due to their high ionization energies. The non-metallic character gradually decreases down the group. However, iodine is, solid and has metallic lustre.

(8) **Atomicity and physical state** : All the halogens exist as diatomic covalent molecules ( $F_2, Cl_2, Br_2$  and  $I_2$ ).  $F_2$  and  $Cl_2$  are gases at room temperature,  $Br_2$  is corrosive liquid and  $I_2$  is volatile solid.

(9) **Colour** : All the halogens have characteristic colours.  $F_2$  is light yellow,  $Cl_2$  is greenish yellow,  $Br_2$  is reddish brown and  $I_2$  is deep violet. The colour of halogens is due to the reason that their molecules absorb light in the visible region as a result of which electron are excited to higher energy levels. The amount of energy needed for excitation decreases progressively from  $F_2$  to  $I_2$  and consequently there is a progressive deepening of colour of the halogens from  $F_2$  to  $I_2$ . Since, fluorine atom requires large amount of energy for excitation of electrons and therefore absorbs violet light and appears yellow. On the other hand, iodine requires low energy for excitation of electrons (absorbs yellow light) and appears deep violet.

(10) **Bond dissociation energy** : Bond dissociation energies of chlorine, bromine and iodine decrease down the group as the size of the atom increases. The bond dissociation energy of fluorine, is however, lower than those of chlorine and bromine because of inter electronic repulsions present in the small atom of fluorine

$X$	$F$	$Cl$	$Br$	$I$
Bond dissociation energy ( $kJ\ mol^{-1}$ )	158	243	192	151

Hence bond energy decreases in the order  $Cl_2 > Br_2 > F_2 > I_2$

(11) **Bond length in  $X_2$  molecule** : As the size of the halogen atom increase, the bond length of  $X-X$  bond in  $X_2$  molecule increases from  $F_2$  to  $I_2$

$X-X$ bond	$F-F$	$Cl-Cl$	$Br-Br$	$I-I$
Bond length (pm)	143	199	228	266

Thus, the bond length increases in the order  $F_2 < Cl_2 < Br_2 < I_2$ .

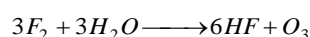
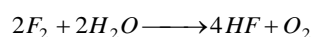
(12) **Melting points and boiling points** : Melting points and boiling points of these elements increase as we move down the group from  $F$  to

$I$  due to an increase in the vander Waals forces of attraction which increase down the group as the size of the atom increases.

Element	$F$	$Cl$	$Br$	$I$
Melting point ( $K$ )	54	172	266	386
Boiling point ( $K$ )	85	239	332	458

Hence, the melting points and boiling points show the order as  $F < Cl < Br < I$ .

(13) **Solubility** : Halogens, being non polar in nature do not readily dissolve in a polar solvent like water. however, fluorine reacts with water vigorously even at low temperature (exothermally) forming a mixture of ozone and oxygen



Chlorine and bromine are fairly soluble but iodine is very little soluble in water. chlorine, bromine and iodine are more soluble in organic solvents like  $CCl_4, CS_2$  or  $CHCl_3$  and produce coloured solutions. Thus  $Cl_2, Br_2$  and  $I_2$  give yellow, brown and violet colour respectively. It is believed that in non-polar solvents, halogens exist as free molecules just as in the gas phase.

In nucleophilic (electron donating) polar solvents like alcohols, ketones or liquid  $SO_2$ , halogens produce brown solution. This colour is due to the complex formation (solvent  $\rightarrow$  halogen) which are charge transfer compounds.

The solubility of iodine ( $I_2$ ) in water increases with addition of  $KI$  or  $NaI$  due to the formation of polyhalide (triiodide,  $I_3^-$ ) ion,  $KI + I_2 \rightleftharpoons KI_3$

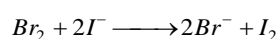
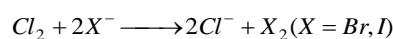
However, this solution behaves as a simple mixture of  $KI$  and free  $I_2$  and contains  $K^+$  and  $I^-$  ions and free  $I_2$  molecules. It has a brown colour. The solution of iodine in water (due to its very little solubility) is also brown.

(14) **Oxidising power** : All the halogens acts as strong oxidising agents since they have a strong tendency to attract electrons and have positive values of electrode potentials ( $E^\circ$ ). The oxidising power, however, decreases as we move down the group from  $F$  to  $I$ . i.e.,  $F_2 > Cl_2 > Br_2 > I_2$

Since  $F_2$  is the strongest oxidising agent, it will oxidise all other halide ions to halogens.



Similarly,  $Cl_2$  will displace  $Br^-$  and  $I^-$  ions from their solutions while  $Br_2$  will displace  $I^-$  ions only.



Hence  $F_2$  is the strongest and  $I_2$  is the weakest oxidising agent.

This is also indicated by the decrease in the electrode potential ( $E^\circ$ ) for the reaction  $X_2(aq) + 2e^- \longrightarrow 2X^-(aq)$  on moving down the group.

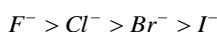
$X_2$	$F_2$	$Cl_2$	$Br_2$	$I_2$	$At_2$
$E^\circ$ (volts)	2.87	1.36	1.09	0.53	0.3

The electron affinity of fluorine is less than that of chlorine but still it is the strongest oxidising agent. This is because of its low bond dissociation energy ( $158 \text{ kJ mol}^{-1}$ ) and high heat hydration ( $510 \text{ kJ mol}^{-1}$ ) as compared to chlorine (for which the values are 243 and  $372 \text{ kJ mol}^{-1}$ , respectively).

(15) **Heat of hydration** : The heat of hydration of the halide ion ( $X^-$ ) decreases as the size of the halogen decreases down the group from  $F$  to  $I$ .

$X^-$ ion	$F^-$	$Cl^-$	$Br^-$	$I^-$
Heat of hydration ( $\text{kJ mol}^{-1}$ )	510	372	339	301

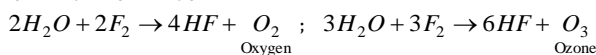
Thus, the decreasing order of heat of hydration of halides is as follows :



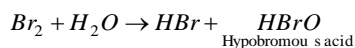
#### Chemical properties

(1) **Reactivity** : The halogen are most reactive elements due to their low bond dissociation energy, high electron affinity and high enthalpy of hydration of halide ion.  $F > Cl > Br > I$

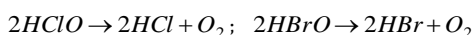
(2) **Reaction with  $H_2O$**  : Halogens readily decomposes water. This tendency decreases on moving down the group. Fluorine decomposes water very energetically to give oxygen and ozone,



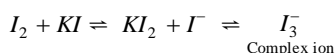
Fluorine gives fumes in moist air. This is due to the formation of  $HF$ , which is a liquid and can absorb moisture to form liquid droplets and therefore, gives fumes with moist air. Chlorine and bromine react less vigorously,



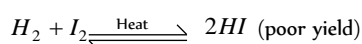
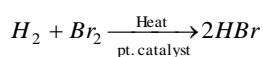
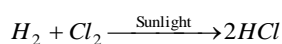
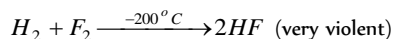
In the presence of sunlight,  $HClO$  (hypochlorous acid)  $HBrO$  (hypobromous acid) liberate oxygen.



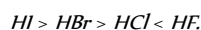
Iodine is only slightly soluble in water. However, it dissolves in 10% aqueous solution of  $KI$  due to the formation of  $I_3^-$  ions.



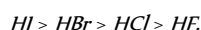
(3) **Reaction with hydrogen** : Form covalent halides.



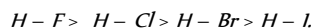
• Acidic strength in aqueous solution is in the order,



• Reducing character of hydrides follow the order,



• Boiling point  $HF > HI > HBr > HCl$ . Thermal stability,



$HCl$  is also called Muriatic acid.

(4) **Hydrides** : All the halogens combine directly with hydrogen to form halogen acids but their reactivity progressively decreases from fluorine to iodine,  $H_2 + X_2 \rightarrow 2HX$  ( $X = F, Cl, Br$  or  $I$ ).

(i) **Boiling points or volatility** : In other words volatility decreases in the order :  $HCl > HBr > HI > HF$  as the boiling points increase in the order :  $HCl$  (189K) <  $HBr$  (206K) <  $HI$  (238K) <  $HF$  (292.5K).

(ii) **Thermal stability** : Thermal stability of the hydrides decrease from  $HF$  to  $HI$  i.e.,  $HF > HCl > HBr > HI$ .

(iii) **Acidic strength** : The acidic strength of halogen acids decreases from  $HI$  to  $HF$  i.e.,  $HI > HBr > HCl > HF$ .

(iv) **Reducing properties** : Since the stability of hydrides decreases from  $HF$  to  $HI$ , their reducing properties increase in the order  $HF < HCl < HBr < HI$ .

(v) **Dipole moments** : The dipole moments of hydrogen halides decrease in the order :  $HF > HCl > HBr > HI$  as the electro negativity of the halogen atom decreases from  $F$  to  $I$ .

$HX$	$HF$	$HCl$	$HBr$	$HI$
Dipole moment (D)	1.74	1.07	0.78	0.38

(5) **Oxides** : Halogens (except  $F$ ) do not combine readily with oxygen. However, a number of compounds of halogens with oxygen have been prepared by indirect methods. Only two compounds of fluorine with oxygen, i.e. oxygen difluorine ( $OF_2$ ) and oxygen fluoride ( $O_2F_2$ ) are known. Chlorine forms largest number of oxides i.e.  $ClO$ ,  $Cl_2O$ ,  $ClO_2$  and  $Cl_2O_7$ , while iodine forms the least, i.e.  $IO_2$ . Bromine, however, forms three oxides ( $Br_2O$ ,  $BrOCl$   $BrO_2$ ). In all these compounds, bonds are largely covalent. All the oxides of halogens are powerful oxidizing agents. These compounds are very reactive and are unstable towards heat. The stability of oxides is greatest for iodine while bromine oxides are the least stable. For a particular halogen, higher oxides are more stable than the lower ones.

Iodine-oxygen bond is stable due to greater polarity of the bond (due to larger electro negativity difference between  $I$  and  $O$ ) while in chlorine-oxygen bond, the stability is gained through multiple bond formation involving the  $d$ -orbital of chlorine atom. Bromine lacks both these characteristics and hence forms least stable oxides.

Oxides of chlorine, bromine and iodine are acidic and the acidic character increases as the percentage of oxygen increases in them.

Iodine also forms  $IO_2$  and  $IO_3$  compounds which are believed not to be true oxides but are basic iodyliodate,  $IO_2(O)$  and normal iodine triodate,  $I(O)_3$ , having tripositive iodine as the cation.

$OF_2$  is V-shaped having bond angle  $103^\circ$ ,  $ClO_2$  is also V-shaped with bond angle  $117^\circ$  while  $ClO$  is angular with bond angle  $118^\circ$ . It is paramagnetic due to odd number of electrons having three-electron bond. It is regarded as a mixed anhydride of chloric and chlorous acids.  $2ClO_2 + H_2O \rightarrow HClO_3 + HClO$ .

(6) **Oxoacids of halogens** : Fluorine does not form any oxoacid since it is the strongest oxidizing agent. Chlorine, bromine and iodine mainly form four series of oxoacids namely hypohalous acid ( $HXO$ ), halous acid ( $HXO_2$ ), halic acid ( $HXO_3$ ) and perhalic acid ( $HXO_4$ ) as given below :

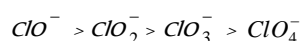


Table 18.10

Oxidation state	Chlorine	Bromine	Iodine	Thermal stability and acid strength	Oxidising power
+1	$HClO$	$HBrO$	$HIO$	Increases ↓	Decreases ↓
+3	$HClO_2$	—	—		
+5	$HClO_3$	$HBrO_3$	$HIO_3$		
+7	$HClO_4$	$HBrO_4$	$HIO_4$		
	Acidity decreases →				

(i) **Hybridized ion** : In all these oxoacids, the halogen atom is  $sp^3$  - hybridized.

(ii) **Acidic character** : All these acids are monobasic containing an  $-OH$  group. The acidic character of the oxoacids increases with increase in oxidation number, i.e.,  $HClO < HClO_2 < HClO_3 < HClO_4$  and the strength of the conjugate bases of these acids follows the order,



(iii) **Oxidising power and thermal stability** : The oxidizing power of these acids decreases as the oxidation number increases, i.e.,  $HClO < HClO_2 < HClO_3 < HClO_4$ . Stability of oxoacids of chlorine in the increasing order is,  $HClO < HClO_2 < HClO_3 < HClO_4$  and the increasing stability order of anions of oxoacids of chlorine is,  $ClO^- < ClO_2^- < ClO_3^- < ClO_4^-$ .

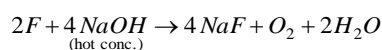
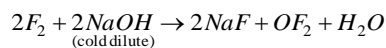
As the number of oxygen atoms in an ion increases there will be a greater dispersal of negative charge and thus greater will be the stability of ion formed. For different halogen having the same oxidation number, the thermal stability decreases with increase in atomic number i.e., it is in the order  $HClO > HBrO > HIO$  and  $ClO > BrO > IO$ . However, in  $HXO_3$  is most stable. The stability order being  $HClO_3 < HBrO_3 < HIO_3$ .

(iv) Perhalates are strong oxidizing agents, the oxidizing power is in the order,  $BrO_4^- > IO_4^- > ClO_4^-$ .

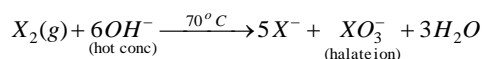
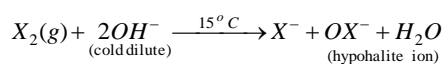
Thus  $BrO_4^-$  is the strongest oxidizing agent (though its reaction is quite slow) and  $ClO_4^-$  is the weakest.

(v) The acidity of oxoacids of different halogens having the same oxidation number decreases with increase in the atomic size of the halogen i.e.  $HClO_4 > HBrO_4 > HIO_4$ .

#### (7) Reaction with alkalis :

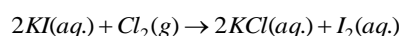
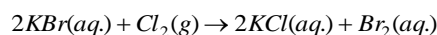


Halogen other than fluorine ( $Cl_2, Br_2, I_2$ ) react with  $NaOH$  as follows,



(8) **Bleaching action of halogen** :  $Cl_2$  acts as bleaching agent, its bleaching action is permanent.  $Cl_2$  water can also act as ink remover.

#### (9) Reaction with other halides



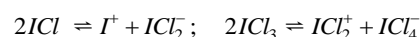
(10) **Inter halogen compounds** : The compounds of one halogen with the other are called inter halogens or inter halogen compounds. The main reason for their formation is the large electronegativity and the size differences between the different halogens. Taking  $A$  as the less electronegative and  $B$  as the more electronegative halogen, they are divided into the following four types the less electronegative halogen ( $A$ ) is always written first.

$AB$	$AB_2$	$AB_3$	$AB_4$
$ClF$ $BrF, BrCl, ICl$ $IBr, IF$	$ClF_3, BrF_3$ $IF_3, ICl_3$	$BrF_5, IF_5$	$IF_7$

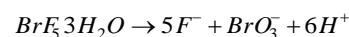
These interhalogen compounds are unstable and more reactive

#### (i) General properties

- Largest halogen always serves the central atom.
- The highest interhalogen compound i.e.  $IF_7$  is obtained with iodine, the largest halogen attached to the smallest one
- The bonds in interhalogen compounds are essentially covalent.
- Thermal stability decreases as the size difference decreases and increases as the polarity of the bond increases. Thus  $ClF$  is thermally more stable as compared to  $IBr$ .
- They ionize in solution or in the liquid state,

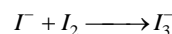


(f) Hydrolysis of interhalogen compounds always produces a halide ion derived from smaller halogen and oxyhalide derived from larger halogen,  $ICl + H_2O \rightarrow Cl^- + IO^- + 2H^+$ ;



- They are strong oxidizing agents.
- Largest number of interhalogens are formed by fluorine due to its smaller size and higher electronegativity or oxidizing power.
- Structure** : Interhalogen compounds are,
  - $AB$  type i.e.  $ICl, IBr, IF$  etc, are linear
  - $AB_2$  type i.e.  $IF_2, ClF_2, BrF_2$  have distorted trigonal bipyramidal ( $dsp$ -hybridization) structures of T-shape due to two lone pairs in equatorial positions  $ICl_2$  is dimeric,  $ICl_2$  and has a planar structure.
  - $AB_3$  types i.e.  $BrF_3, IF_3$  have distorted octahedral ( $dsp$ -hybridization) shapes or square pyramidal due to a lone pair one of the axial positions.
  - $AB_4$  type i.e.  $IF_4$  have pentagonal bipyramidal ( $dsp$ -hybridization) structures.

(ii) **Polyhalide ions** : Halogens or interhalogens combine with halide ions to form polyhalide ions. The most common example of polyhalide ion formation is furnished by the increase in solubility of iodine in water in the presence of  $KI$  which is due to the formation of tri iodide ion,  $I_3^-$



Many other examples of polyhalides ions are

(i)  $Cl_3^-, Br_3^-, ICl_2^-, IBr_2^-$  including  $I_3^-$ . In these ions, one of the halogen atoms (in case of similar atoms) or halogen atom larger in size undergoes  $sp^3d$ -hybridization giving a linear shape with three lone pairs at equatorial positions.

(ii)  $Cl_3^+, Br_3^+, I_3^+, ICl_2^+, IBr_2^+$ . Here we find central atom  $sp^3$  hybridized giving a bent shape with two lone pairs of electrons on the central atom.

(iii)  $ICl_4^-, BrF_4^-, I_5^-$ . Here central atom involves  $sp^3d^2$  hybridization giving square planar shape with two lone pairs of electrons on axial positions.

(iv)  $ICl_4^+, BrF_4^+, I_5^+$ . In these ions central atom involves  $sp^3d$  hybridization giving a distorted tetrahedral structure with one lone pair of electrons on equatorial position.

(v)  $I_7^-, IF_6^-$ . The central atom  $I$  undergoes  $sp^3d^3$  hybridization giving a distorted octahedral structure with one lone pair of electrons.

(vi)  $I_7^+$ . Here central  $I$  atom involves  $sp^3d^2$  hybridization giving an octahedral structure.

Fluorine due to its highest electronegativity (and only  $-1$  oxidation state) does not form polyhalide ions where it acts as a central atom.

#### (12) Pseudohalogen and pseudohalides

Pseudohalogen	Pseudohalide
Cyanogen – $(CN)_2$	Cyanide – $CN^-$
Oxocyanogen – $(OCN)_2$	Cyanate – $OCN^-$
Thiocyanogen – $(SCN)_2$	Thiocyanate – $SCN^-$
Selenocyanogen – $(SeCN)_2$	Selenocyanate – $SeCN^-$

(13) **Anomalous behaviour of fluorine** : Fluorine differs from rest of the elements of its family due to (i) its small size (ii) highest electronegativity, (iii) low bond dissociation energy and (iv) absence of  $d$ -orbitals in the valence shell. The main points of difference are :

(1) Fluorine is most reactive of all the halogens due to lower value of  $F-F$  bond dissociation energy ( $F_2 = 158, Cl_2 = 243$ , bromine = 192 and iodine = 151  $kJ mol^{-1}$ ).

(2) Being the most electronegative element, it shows only an oxidation state of  $-1$  and does not show positive oxidation states due to absence of  $d$ -orbitals in its valence shell. Other halogens show positive oxidation states of  $+1, +3, +5$  and  $+7$ .

(3) Due to small atomic size and high electronegativity of  $F, HF$  undergoes strong  $H$ -bonding while other halogen acids do not. As a result,

(i)  $HF$  is a liquid (boiling point  $292.5K$ ), while other halogen acids are gases at room temperature (boiling point of  $HCl = 189K, HBr = 206K, HI = 238K$ ).

(ii)  $HF$  is weakest of all the halogen acids due to high strength of  $H-F$  bond.

(iii) Due to  $H$ -bonding,  $HF$  can form acid salts of the type  $KHF_2$ , i.e.,  $K^+[H-F \cdots F^-]$  while  $HCl, HBr$  and  $HI$  do not form such salts (i.e., no  $KHCl_2, KHBr_2$  and  $KHI_2$  are known).

(4) Fluorides have the maximum ionic character. For example  $AlF_3$  is ionic while other halides of  $Al$  are covalent.

(5) Of all the halogens, fluorine has the highest positive electrode potential ( $F_2 = 2.87, Cl_2 = 1.36, Br_2 = 1.09$  and  $I_2 = 0.53$  volt) i.e., it is most easily reduced and hence acts as the strongest oxidising agent. It brings about the highest oxidation of other elements with which it combines. For example with  $S$ , it gives  $SF_6$ , with  $I_2$  it gives  $IF_7$ . Other halogens do not always bring about the highest oxidation state. For example, with sulphur  $Cl_2$  gives  $SCl_4, Br_2$  gives  $SBr_2$  while  $I_2$  does not react at all.  $F_2$  is so powerful oxidising agent that it can even oxidise inert-gases.

(6)  $HF$  cannot be stored in glass bottles since it reacts with silicates to form fluorosilicates.



While other halogen acids ( $HCl, HBr$  and  $HI$ ) do not react with silicates and hence can be stored in glass bottles.

(7)  $AgF$  is soluble in  $H_2O$  while all other silver halides i.e.,  $AgCl, AgBr$  and  $AgI$  are insoluble in water. In contrast,  $CaF_2$  is insoluble while other calcium halides i.e.,  $CaCl_2, CaBr_2, CaI_2$  are soluble in  $H_2O$ .

(8) Due to absence of  $d$ -orbitals, fluorine, does not form polyhalide ions while other halogens form polyhalides of the type  $I_3^-, Br_3^-, I_5^-$  etc.

#### Preparation of halogens and its uses

##### (1) Fluorine

(i) **Occurrence of fluorine** : Fluorine does not occur free in nature but occurs mostly as fluor spar  $CaF_2$ , cryolite,  $Na_3AlF_6$  and fluorapatite,  $CaF_2 \cdot 3Ca_3(PO_4)_2$ . Traces of fluoride occur in sea water, bones, teeth, blood, milk etc.

(ii) **Difficulties encountered during its isolation** : (a)  $F_2$  attacks all the materials of the apparatus such as glass, platinum, carbon and other metals, (b)  $F_2$  is the strongest oxidising agent and hence no oxidising agent can oxidise  $F^-$  ions to  $F_2$ . (c)  $F_2$  cannot be prepared even by electrolysis of an aqueous solution of  $HF$  because  $F_2$  formed reacts violently with water. It also cannot be prepared by electrolysis of anhydrous  $HF$  because it is not only poisonous, corrosive and volatile but also is a bad conductor of electricity.

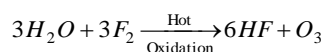
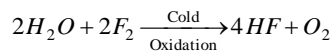
(iii) **Preparation** :  $F_2$  is now prepared by electrolysis of a solution of  $KHF_2$  (1 part) in anhydrous  $HF$  (5 parts) in a vessel (modern method) made of  $Ni-Cu$  alloy or  $Ni-Cu-Fe$  alloy called the monel metal using carbon electrodes. During the electrolysis following reactions occur.



At cathode :  $K^+ + e^- \longrightarrow K ; 2K + 2HF \longrightarrow 2KF + H_2 \uparrow$

At anode :  $F^- \longrightarrow F + e^- ; F + F \longrightarrow F_2$

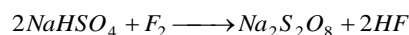
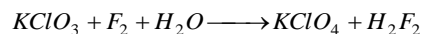
(iv) **Properties** : It is the most reactive of all the halogens. It combines with metals as well as non-metals to form fluorides. It decomposes water forming  $O_2$  and  $O_3$  and reacts vigorously with hydrogen of hydrocarbons leaving behind fluorinated hydrocarbons.



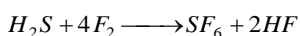
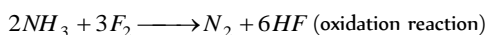
( $HF$  being a volatile liquid fumes in air)



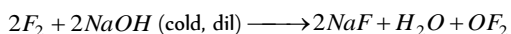
It is a strong oxidising agent and oxidises  $KClO_3$  to  $KClO_4, KIO_3$  to  $KIO_4$  and bisulphates to peroxy sulphates.



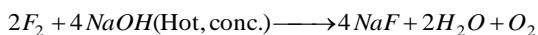
It reacts with  $NH_3$  to form nitrogen and with  $H_2S$  forming  $SF_6$ .



Fluorine reacts with cold and dilute sodium hydroxide solution to give oxygen difluoride ( $OF_2$ )



However, with hot and concentrated sodium hydroxide solution it gives oxygen



Since  $F_2$  is the strongest oxidising agent, it is always reduced and hence does not show disproportionation reactions while others halogens do.

$F_2$  oxidises all other halide ions to the corresponding halogens ( $F_2 + 2X^- \longrightarrow 2F^- + X_2$ ); ( $X = Cl, Br$  or  $I$ )

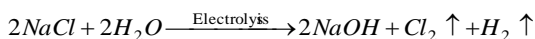
(v) **Uses of fluorine** : Fluorine is used in the manufacture of  $UF_6$  (which is used for nuclear power generation),  $SF_6$  (which is used as an electrical insulator), chlorofluorocarbons, teflon, cryolite and  $HF$ .

(vi) **Fluorocarbons** are the derivatives of hydrocarbons in which  $H$ -atoms are replaced by  $F$ -atoms. these are obtained by fluorination of hydrocarbons with  $F_2$  diluted with an inert gas such as  $N_2$  in presence of  $CuF_2$  as catalyst. Fluorocarbons are widely used in industry because of their extreme inertness (non-in-flammability and extreme stability). Freon ( $CF_2Cl_2$ ) is used as a refrigerant, tetrafluoroethylene ( $F_2C = CF_2$ ) is used for the manufacture of teflon which is highly non-inflammable, has high thermal stability and is chemically inert *i.e.*, is not attacked by acids and corrosive chemicals. It is used for making pipes, surgical tubes, non-stick utensils and as an electrical insulator.

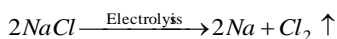
## (2) Chlorine

(i) **Occurrence** : Chlorine mainly occurs as rock salt ( $NaCl$ ) Carnallite, ( $KCl, MgCl_2 \cdot 6H_2O$ ) and Calcium chloride. ( $CaCl_2$ ).

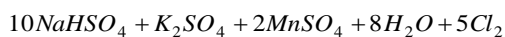
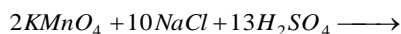
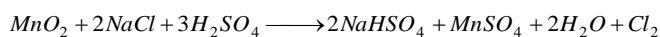
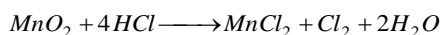
(ii) **Preparation** : On a commercial scale chlorine is prepared by electrolysis of an aqueous solution of sodium chloride (brine solution) (Nelson cell, Castner and Kellner's cell for the manufacture of  $NaOH$ ) when  $Cl_2$  is evolved at the anode and  $H_2$  is evolved at the cathode.



It can also be prepared by electrolysis of molten  $NaCl$  (Down's cell for the manufacture of metallic sodium). When  $Cl_2$  is evolved at the anode and sodium metal at the cathode.

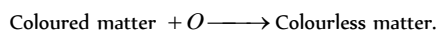


In the laboratory,  $Cl_2$  is prepared by the action of  $MnO_2$  or  $KMnO_4$  or  $K_2Cr_2O_7$  on conc.  $HCl$  or a mixture of  $NaCl$  and conc.  $H_2SO_4$



Other oxidising agents such as  $PbO_2, Pb_3O_4, CaOCl_2, O_3$  etc. also react with  $HCl$  to liberate  $Cl_2$ .

(iii) **Properties** : It combines with metals and non metals to form chlorides. it decomposes water forming  $HCl$  and  $HClO$  (hypochlorous acid) which is unstable and decomposes giving nascent oxygen which is responsible for oxidising and bleaching action of chlorine.

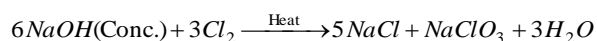
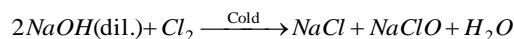


The bleaching action is permanent and colour is not restored on standing. However, it cannot be used for bleaching delicate articles such as straw, silk, wool etc. which are damaged by it.

$Cl_2$  oxidises  $Br^-$  and  $I^-$  ions to  $Br_2$  and  $I_2$  respectively.

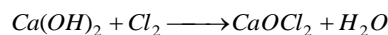


It combines with alkalis forming hypochlorite and chlorate salts in cold and hot conditions respectively.

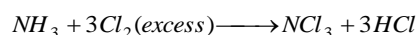
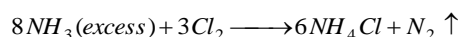


During these reactions, halogen is simultaneously reduced to  $X^-$  ion and is oxidised to either hypohalite ( $XO^-$ ) or halate ( $XO_3^-$ ) ion. Such reactions are called disproportionation reactions.

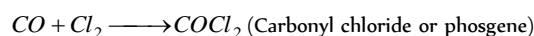
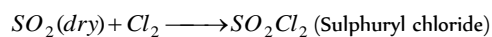
With slaked lime,  $Cl_2$  gives bleaching powder ( $CaOCl_2$ )



With ammonia,  $Cl_2$  reacts as follows :



With  $SO_2$  and  $CO$ , addition compounds are formed



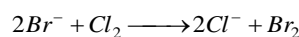
$Cl_2$  is strong oxidising agent. It oxidises  $FeCl_2$  to  $FeCl_3$ , moist  $SO_2$  to  $H_2SO_4, SO_3^{2-}$  to  $SO_4^{2-}$ , thiosulphate to sulphate and sulphur.

(iv) **Uses of chlorine** : It is used in the manufacture of  $HCl, NaOCl$ , bleaching powder, chlorates, vinyl chloride, insecticides such as DDT, chlorinated organic solvents like  $CHCl_3, CCl_4$ . It is also used in sterilisation of drinking water, in the extraction of  $Au$  and  $Pt$  and as a bleaching agent for paper, pulp and textiles.

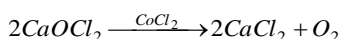
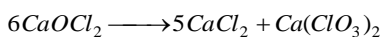
## (3) Bromine

(i) **Occurrence** : It mainly occurs in sea water and salt lakes as  $NaBr, KBr$  and  $MgBr_2$ .

(ii) **Preparation** : On a commercial scale, bromine is prepared either from sea water (containing  $NaBr, KBr$  and  $MgBr_2$ ) or the mother liquor (containing  $MgBr_2$ ) left after crystallisation of chlorides from carnallite. On passing  $Cl_2$  gas through these solutions, bromides get oxidised to bromine which is cooled and condensed to  $Br_2$  liquid.







It is used for bleaching cotton, wood pulp etc., as a disinfectant, as a germicide for sterilization of drinking water, in the manufacture of chloroform and for making wood unshrinkable.

## Noble Gases

Helium is the first member of group 18 or zero of the periodic table. It consists of six elements helium (*He*), Neon (*Ne*), argon (*Ar*), krypton (*Kr*), xenon (*Xe*) and radon (*Rn*). Zero group occupies the intermediate position between the elements of VIIA (17th) and IA (1st) groups. These are collectively called as inactive gases or inert gases. However, these are now called noble gases as some compounds of these gases have been obtained under certain specific conditions.

### (1) Electronic configuration

Elements	Discovery	Electronic configuration ( $ns^2 np^6$ )
${}_2 He$	Lockyer and Janssen (1868)	$1s^2$
${}_{10} Ne$	Ramsay	$1s^2, 2s^2 2p^6$
${}_{18} Ar$	Rayleigh and Ramsay (1894)	$1s^2, 2s^2 2p^6, 3s^2 3p^6$
${}_{36} Kr$	Ramsay and Travers (1898)	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6$
${}_{54} Xe$	Ramsay and Travers (1898)	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6 4d^{10} 4f^{14}, 5s^2 5p^6$
${}_{86} Rn$	Dorn (1900)	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6 4d^{10} 4f^{14}, 5s^2 5p^6 5d^{10}, 6s^2 6p^6$

(2) **Occurrence** : Due to the inert nature of noble gases, they always occur in the free state. Except radon, all these gases are present in atmosphere in the atomic state.

Element	<i>He</i>	<i>Ne</i>	<i>Ar</i>	<i>Kr</i>	<i>Xe</i>
Abundance (Volume %)	$5.2 \times 10^{-4}$	$1.8 \times 10^{-3}$	$9.3 \times 10^{-1}$	$1.4 \times 10^{-3}$	$8.7 \times 10^{-6}$

*He* is also present in natural gas to the extent of 2 to 7%.

### (3) Isolation

(i) **Helium** : It is commercially obtained from natural gas. The natural gas contains hydrocarbons (methane etc.), *CO*, *H<sub>2</sub>S* and *He* as the main constituents.

The natural gas is compressed to about 100 atm and cooled to 73K. *He* remains unliquefied while other gases get liquefied. About 99% pure *He* is prepared by this method.

(ii) **Argon, Neon, Krypton and Xenon** : These gases are prepared by the fractionation distillation of liquid air. Fractional distillation of air gives *O<sub>2</sub>*, *N<sub>2</sub>* and mixture of noble gases. The individual gases may be obtained by adsorption of air on coconut charcoal. The charcoal adsorbs different gases at different temperatures and can be collected.

(iii) **Radon** : It can be obtained by radio active disintegration of radium ( ${}_{88}Ra$ )  $\rightarrow$   ${}_{86}Rn + {}_2\alpha^4$ .

### Properties :

(1) **Atomic radii** : The atomic radii of noble gases increases on moving down the group and their atomic radii correspond to the vander Waal's radii.

(2) **Boiling points** : The m.pt. and b.pt. increases from *He* to *Rn*, because of increase in magnitude of vander Waal's forces.

(3) **Polarizability** : The polarizability increases down the group,  $He < Ne < Ar < Kr < Xe$

(4) **Ionisation energy and electron affinity** : Noble gases have stable  $ns^2 np^6$  fully filled electronic configuration, so these have no tendency to add or lose electron. Therefore, ionisation energy of noble gases is very high. On the other hand their electron affinity is zero.

(5) **Heat of vaporisation** : They possess very low values of heat of vapourisation, because of presence of very weak vander Waal's forces of attraction between their monoatomic molecules. However the value of heat of vaporisation increases with atomic number down the group and this shows that there is an increasing polarizability of the larger electronic clouds of the elements with higher atomic number.

(6) **Solubility in water** : They are slightly soluble in water. Their solubility generally increases with the increase in atomic number down the group.

(7) **Adsorption by charcoal** : All of them except helium are adsorbed by coconut charcoal at low temperature. The extent of adsorption increases down the group.

(8) **Characteristic spectra** : All of them give characteristic spectra, by which they can be identified.

(9) **Liquification of gases** : It is difficult to liquify noble gases as their atoms are held by weak vander Waal's forces. Ease of liquification increases down the group from *He* to *Rn*. Helium has the lowest boiling point (4.18 K) of any known substance. The ease of liquification increases down the group due to increase in intermolecular forces.

The elements helium (*He*), neon (*Ne*), argon (*Ar*), krypton (*Kr*), xenon (*Xe*) and radon (*Rn*), constitute zero group of the periodic table. These are gases at ordinary temperature and do not have chemical reactivity and therefore, these are called inert gases.

### Compounds of Xenon

In 1962, N. Bartlett noticed that *PtF<sub>6</sub>* is a powerful oxidizing agent which combines with molecular oxygen to form ionic compound, dioxygenyl hexafluoro platinate (v)  $O_2^+[PtF_6]^-$ ,  $O_{2(g)} + PtF_{6(g)} \rightarrow O_2^+[PtF_6]^-$ . This indicates that *PtF<sub>6</sub>* has oxidized *O*, to  $O_2^+$ . Now, oxygen and xenon have some similarities,

(i) The first ionization energy of *Xe* gas (1170 kJ mol<sup>-1</sup>) is fairly close to that of oxygen (1166 kJ mol<sup>-1</sup>).

(ii) The molecular diameter of oxygen and atomic radius of *Xe* are similar (4Å)

On this assumption, Bartlett reacted Xenon and *PtF<sub>6</sub>* in gas phase and an orange yellow solid of the composition *XePtF<sub>6</sub>* was obtained,  $Xe_{(g)} + PtF_{6(g)} \rightarrow Xe^+[PtF_6]^-$ .  
Orange yellow

Some important stable compounds of *Xe* are,

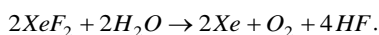
+2	+4	+6
<i>XeF<sub>2</sub></i>	<i>XeF<sub>4</sub></i> , <i>XeOF<sub>2</sub></i>	<i>XeF<sub>6</sub></i> , <i>XeOF<sub>4</sub></i> , <i>XeO<sub>3</sub></i>

**Fluorides** : Xenon forms three compounds with fluorine. These are : Xenon difluoride ( $XeF_2$ ), Xenon tetrafluoride ( $XeF_4$ ) and Xenon hexafluoride ( $XeF_6$ ).

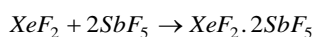
(1) **Xenon difluoride ( $XeF_2$ )** is formed when a mixture of Xenon and fluorine in the ratio 1 : 3 by volume is passed through a nickel tube at 673 K,  $Xe + Fe \xrightarrow{Ni, 673K} XeF_2$

**Structure** :  $XeF_2$  has trigonal bipyramid geometry due to  $sp^3d$ -hybridization of  $Xe$ . Three equatorial positions are occupied by lone pairs of electrons giving a linear shape to the molecule.

**Properties** :  $XeF_2$  is a colourless crystalline solid, reacts with  $H_2$  to give  $Xe$  and  $HF$ . It is hydrolysed completely by water,



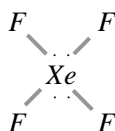
It also forms addition compounds with reactive pentafluorides like  $SbF_5$ ,  $TaF_5$  etc.



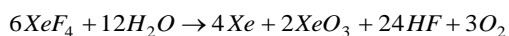
It is a mild fluorinating agent and hence reacts with benzene to give fluorobenzene.

(2) **Xenon tetrafluoride ( $XeF_4$ )** is prepared by heating a mixture of xenon and fluorine in the ratio 1 : 5 in a nickel vessel at 673 K and then suddenly cooling it in acetone.  $XeF_4$  is also formed when an electric discharge is passed through a mixture of xenon and excess of fluorine,  $Xe + 2F_2 \xrightarrow{Ni, 673K} XeF_4$

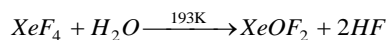
**Structure** :  $XeF_4$  has square planar shape due to  $sp^3d^2$  hybridization of  $Xe$  giving octahedral geometry with two trans positions occupied by lone pairs of electrons.



**Properties** :  $XeF_4$  is a colourless, crystalline solid, soluble in anhydrous  $HF$ , reacts with  $H_2$  to form  $Xe$  and  $HF$  and reacts with water to give highly explosive solid,  $XeO_3$ . (complete hydrolysis),



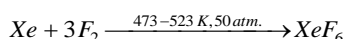
Partial hydrolysis yields  $XeOF_4$ ,



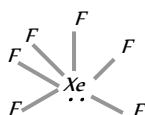
It also forms addition compounds with  $SbF_5$ ,  $XeF_4 + SbF_5 \rightarrow [XeF_3]^+ [SbF_6]^-$ .

It also acts as a strong fluorinating agent.

(3) **Xenon hexafluoride ( $XeF_6$ )** is prepared by heating a mixture of xenon and fluorine in the ratio 1 : 20 at 473—523K under a pressure of 50 atmospheres.

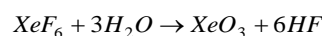
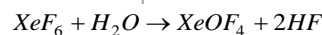


**Structure** :  $XeF_6$  has pentagonal bipyramid geometry due to  $sp^3d^2$  hybridization. One trans position is occupied by a lone pair giving a distorted octahedral shape.



**Properties** : It is colourless, crystalline solid, highly soluble in anhydrous  $HF$  giving solution which is a good conductor of electricity,  $HF + XeF_6 \rightarrow XeF_5^+ + HF_2^-$ .

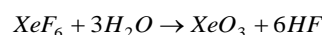
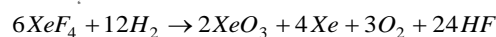
It is the most powerful fluorinating agent and reacts with  $H_2$  to give  $Xe$  and  $HF$ . Partial hydrolysis of  $XeF_6$  yields  $XeOF_4$ , a complete hydrolysis yields xenon trioxide,  $XeO_3$ .



It forms addition compounds with alkali metal fluorides (except  $LiF$ ) of the formula  $XeF \cdot MF$  where  $M$  represents the alkali metal.

**Oxides** : Xenon forms two oxides such as xenon trioxide ( $XeO_3$ ) and xenon tetraoxide ( $XeO_4$ ).

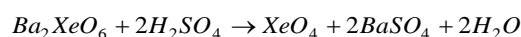
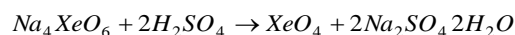
(1) **Xenon trioxide ( $XeO_3$ )** is prepared by complete hydrolysis of  $XeF_6$  and  $XeF_4$ .



**Structure** :  $XeO_3$  has tetrahedral geometry due to  $sp^3$  hybridization of  $Xe$ . One of the hybrid orbitals contains a lone pair of electrons giving a trigonal pyramidal shape. The molecule has three  $Xe = O$  double bonds containing  $p\pi - d\pi$  overlapping.

**Properties** : It is a colourless solid, highly explosive and powerful oxidizing agent.

(2) **Xenon tetraoxide ( $XeO_4$ )** is prepared by the action of conc.  $H_2SO_4$  on sodium or barium xenate ( $Na_4XeO_6$ ;  $Ba_2XeO_6$ ) at room temperature,



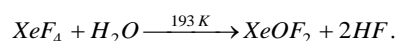
$XeO_4$  is purified by vacuum sublimation at 195 K.

**Structure** :  $XeO_4$  has tetrahedral structure due to  $sp^3$  hybridization of  $Xe$ . There are four  $Xe-O$  double bonds containing  $p\pi - d\pi$  overlapping.

**Properties** : It is quite unstable gas and decomposes to xenon and oxygen,  $XeO_4 \rightarrow Xe + 2O_2$ .

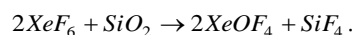
**Oxyfluorides** : Xenon forms three types of oxy fluorides such as xenon oxydifluoride ( $XeOF_2$ ), xenon oxytetrafluoride  $XeOF_4$  and xenon dioxydifluoride ( $XeO_2F_2$ ).

(1) **Xenon oxydifluoride ( $XeOF_2$ )** is formed by partial hydrolysis of  $XeF_4$  at 193 K,



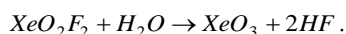
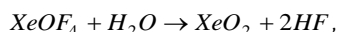
**Structure** :  $XeOF_2$  has trigonal bipyramid geometry due to  $sp^3d$  hybridization of  $Xe$ . Two equatorial positions are occupied by lone pairs of electrons giving a T-shape to the molecule. There is one  $Xe-O$  double bond containing  $p\pi - d\pi$  overlapping.

(2) **Xenon oxytetrafluoride ( $XeOF_4$ )** is prepared by partial hydrolysis of  $XeF_6$ ;  $XeF_6 + H_2O \rightarrow XeOF_4 + 2HF$ . It can also be prepared by the reaction of  $SiO_2$  with  $XeF_6$ ,

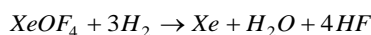


**Structure** :  $XeOF_4$  has octahedral geometry due to  $sp^3d^2$  hybridization of  $Xe$ . One trans position is occupied by a lone pair giving pyramid shape to the molecule. There is one  $Xe-O$  double bond containing  $p\pi - d\pi$  overlapping.

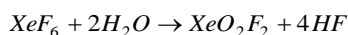
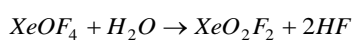
**Properties :** It is a colourless volatile liquid which melts at 227 K. It reacts with water to give  $XeOF_4$  and  $XeO_3$ ,



It is reduced by  $H_2$  to  $Xe$ ,



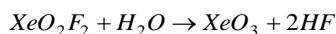
(3) **Xenon dioxydifluoride ( $XeOF_2$ )** is formed by partial hydrolysis of  $XeOF_4$  or  $XeF_6$ ,



It can also be prepared by mixing  $XeO$  and  $XeOF_4$  at low temperature (195K). The product is purified by fractional distillation,  
$$XeO_3 + XeOF_4 \xrightarrow{195K} 2XeO_2F_2$$

**Structure :**  $XeOF_4$  has trigonal bipyramid geometry due to  $sp^3d$ -hybridization of  $Xe$ . One equatorial position is occupied by a lone pair of electrons giving a see-saw structure (shape) to the molecule. There are two  $Xe-O$  double bonds containing  $p\pi-d\pi$  overlapping.

**Properties :** It is a colourless solid which melts at 303K. It is easily hydrolysed to give  $XeO$ ,



#### Uses of noble gases

(1) He is used for filling of balloons and air ships because of its non-inflammability and high power (which is 92.6% to that of hydrogen).

(2) Oxygen-helium (1 : 4) mixture is used for treatment of asthma and for artificial respiration in deep sea diving because unlike nitrogen, helium is not soluble in blood even under high pressure.

(3) Helium is also used for creating inert atmosphere in chemical reactions.

(4) Liquid helium is used as a cryogenic fluid to produce and maintain extremely low temperatures for carrying out researches and as a coolant in atomic reactors and super conducting magnets.

(5) It is also used in low temperature gas thermometry and as a shield gas for arc welding.

(6) Argon is used for creating inert atmosphere in chemical reactions, welding and metallurgical operations and for filling in incandescent and fluorescent lamps. It is also used in filling Geiger-Counter tubes and thermionic tubes.

(7) Krypton and Xenon are also used in gas filled lamps. A mixture of Krypton and Xenon is also used in some flash tubes for high speed photography.

(8) Radon is used in radioactive research and therapeutics and in the non-surgical treatment of cancer and other malignant growths.

by sodium chloride while it is retard by borax or alum. Addition of alum to plaster of paris makes the setting very hard. The mixture is known as keen's cement.

✍  $B_4C$  (boron carbide) is one of the hardest known artificial substance and is called norbia.

✍ Bitter almonds contain  $HCN$  in free state. It is produced by the action of water on amygdalin (present in bitter almonds) in the presence of enzyme emulsion (also present in bitter almonds).

✍ Industrial lubricant oiltag is a suspension of graphite in oil and colloidal solution of graphite is called aquadag.

✍ Broken glass pieces added during glass making is known as cullet.

## Tips & Tricks

✍ Among all the alkalimetals  $Li$  is the strongest and  $Na$  is the weakest reducing agent.

✍ Lindlar's catalyst is  $Pd$  poisoned with  $BaSO_4$  in quinoline.

✍ **Keen's cement :** The setting of plaster of paris may be catalysed