

p-BLOCK ELEMENTS

GROUP 15 ELEMENTS : THE NITROGEN FAMILY

Group 15 includes nitrogen phosphorus, arsenic, antimony and bismuth. As we go down the group, there is a shift from non-metallic to metallic through metalloidal character. Nitrogen and phosphorus are non-metal, arsenic and antimony metalloid and bismuth is a typical metal.

Electronic Configuration :

The valence shell electronic configuration of these element is $ns^2 np^3$ the s orbital in these element is completely filled and p orbitals are half- filled, making their electronic configuration extra stable.

Atomic and Ionic Radii :

Covalent and ionic (in a particular state) radii increase in size down the group. There is a considerable increase in covalent radius from N to P. However, from As to Bi only a small increase in covalent radius is observed. This is due to the presence of completely filled d and / or f orbitals in heavier members.

Ionisation Enthalpy:

Ionisation enthalpy decreases down the group due to gradual increase in atomic size. Because of the extra stable half- filled p-orbital electronic configuration and smaller size, the ionisation enthalpy of the group 15 element is much greater than of group 14 elements in the corresponding periods. The order of successive ionisation enthalpies, as expected is

$$\Delta_i H_1 < \Delta_i H_2 < \Delta_i H_3$$

Electronegativity :

The electronegativity value, in general, decreases down the group with increasing atomic size. However, amongst the heavier elements, the difference is not that much pronounced.

Physical Properties:

All the elements of this group are polyatomic. Dinitrogen is a diatomic gas while all others are solids. Metallic character increases down the group. Nitrogen and phosphours are non – metals , arsenic and antimony metalloids and bismuth is a metal. This is due to decrease in ionisation enthalpy and increase in atomic size. The boiling points , in general , increase from top to bottom in the group but the melting point increases upto arsenic and then decreases upto bismuth. Except nitrogen , all the elements show allotropy.

ATOMIC & PHYSICAL PROPERTIES

Element		N	P	As	Sb	Bi
Atomic Number		7	15	33	51	83
Atomic Mass		14.01	30.97	74.92	121.76	208.98
Electronic configuration		[He] $2s^2 2p^3$	[Ne] $3s^2 3p^3$	[Ar] $3d^{10} 4s^2 4p^3$	[Kr] $4d^{10} 5s^2 5p^3$	[Xe] $4f^{14} 5d^{10} 6s^2 6p^3$
Covalent Radius / pm		70	110	120	140	150
Ionic Radius / pm $a = M^{3-}$, $b = M^{+3}$		171 ^a	212 ^a	222 ^a	76 ^b	103 ^b
Ionization enthalpy / (kJ mol ⁻¹)	I	1402	1012	947	834	703
	II	2856	1903	1798	1595	1610
	III	4577	2910	2736	2443	2466
Electronegativity		3.0	2.1	2.0	1.9	1.9

Chemical Properties :

Oxidation States and trends in a chemical reactivity :

The common oxidation states of these elements are -3 , $+3$ and $+5$. The tendency to exhibit -3 oxidation state decreases down the group , bismuth hardly forms any compound in -3 oxidation state. The stability of $+5$ oxidation state decreases down the group. The only well characterised Bi (V) compound is BiF_5 . The stability of $+5$ oxidation state decreases and that of $+3$ state increases (due to inert pair effect) down the group. Nitrogen exhibits $+1$, $+2$, $+4$ oxidation states also when it reacts with oxygen. Phosphours also shows $+1$ and $+4$ oxidation states in some oxoacids.

In the case of nitrogen , all oxidation states from $+1$ to $+4$ tend to disproportionate in acid solution. For example ,



Similarly , in case of phosphorus nearly all intermediate oxidation states disproportionate into $+5$ and -3 both in alkali and acid. However $+3$ oxidation state in case of arsenic , antimony and bismuth become increasingly stable with respect to disproportionation.

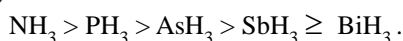
Nitrogen is restricted to a maximum covalency of 4 since only four (one s and three p) orbitals are available for bonding. The heavier elements have vacant d orbitals in the outermost shell which can be used for bonding (covalency) and hence, expand their covalence as in PF_6^- .

Anomalous properties of nitrogen :

Nitrogen differs from the rest of the members of this group due to its smaller size, high electronegativity, high ionisation enthalpy and non-availability of d orbitals. Nitrogen has unique ability to form $p\pi - p\pi$ multiple bonds with itself and with other elements having small size and high electronegativity (e.g., C, O). Heavier elements of this group do not form $p\pi - p\pi$ bonds as their atomic orbitals are so large and diffuse that they cannot have effective overlapping. Thus, nitrogen exists as a diatomic molecule with a triple bond (one s and two p) between the two atoms. Consequently, its bond enthalpy ($941.1 \text{ kJ mol}^{-1}$) is very high. On the contrary, phosphorus, arsenic and antimony form metallic bonds in elemental state. However, the single N - N bond is weaker than the single P - P bond because of high interelectronic repulsion of the non-bonding electrons, owing to the small bond length. As a result the catenation tendency is weaker in nitrogen. Another factor which affects the chemistry of nitrogen is the absence of d orbitals in its valence shell. Besides restricting its covalency to four, nitrogen cannot form $d\pi - p\pi$ bonds as the heavier elements can e.g., $\text{R}_3\text{P} = \text{O}$ or $\text{R}_3\text{P} = \text{CH}_2$ (R = alkyl group). Phosphorus and arsenic can form $d\pi - p\pi$ bond also with transition metals when their compounds like $\text{P}(\text{C}_2\text{H}_5)_3$ and $\text{As}(\text{C}_6\text{H}_5)_3$ act as ligands.

(i) Reactivity towards hydrogen :

All the elements of Group 15 form hydrides of the type EH_3 where E = N, P, As, Sb or Bi. Some of the properties of these hydrides are shown in Table. The hydrides show regular gradation in their properties. The stability of hydrides decreases from NH_3 to BiH_3 which can be observed from their bond dissociation enthalpy. Consequently, the reducing character of the hydrides increases. Ammonia is only a mild reducing agent while BiH_3 is the strongest reducing agent amongst all the hydrides. Basicity also decreases in the order



Properties of Hydrides of Group 15 Elements

Property	NH_3	PH_3	AsH_3	SbH_3	BiH_3
Melting point / K	195.2	139.5	156.7	185	–
Boiling point / K	238.5	185.5	210.6	254.6	290
(E - H) Distance / pm	101.7	141.9	151.9	170.7	–
HEH angle ($^\circ$)	107.8	93.6	91.8	91.3	–
$\Delta_f H^- / \text{kJ mol}^{-1}$	-46.1	13.4	66.4	145.1	278
$\Delta_{\text{diss}} H^-(\text{E} - \text{H}) / \text{kJ mol}^{-1}$	389	322	297	255	–

(ii) Reactivity towards oxygen :

All these elements form two types of oxides : E_2O_3 and E_2O_5 . The oxide in the higher oxidation state of the element is more acidic than that of lower oxidation state. Their acidic character decreases down the group. The oxides of the type E_2O_3 of nitrogen and phosphorus are purely acidic, that of arsenic and antimony amphoteric and those of bismuth is predominantly basic.

(iii) Reactivity towards halogens :

These elements react to form two series of halides : EX_3 and EX_5 . Nitrogen does not form pentahalide due to non-availability of the d-orbitals in its valence shell. Pentahalides are more covalent than trihalides. All the trihalides of these elements except those of nitrogen are stable. In case of nitrogen, only NF_3 is known to be stable. Trihalides except BiF_3 are predominantly covalent in nature.

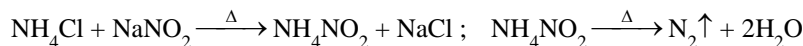
(iv) Reactivity towards metals :

These elements react with metals to form their binary compounds exhibiting -3 oxidation state, such as, Ca_3N_2 (calcium nitride) Ca_3P_2 (calcium phosphide), Na_3As_2 (sodium arsenide), Zn_3Sb_2 (zinc antimonide) and Mg_3Bi_2 (magnesium bismuthide).

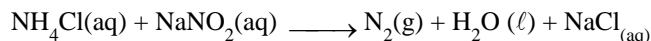
● **NITROGEN (N):**

○ **PREPARATION:**

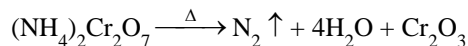
- (i) By heating a mixture of NH_4Cl and NaNO_2 , N_2 is collected by the downward displacement of water.



- * By treating an aqueous solution of ammonium chloride with sodium nitrate. It is laboratory method of preparation.

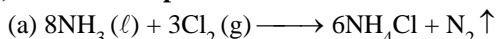


- (ii) By heating ammonium dichromate:

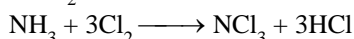


- (iii) By oxidation of ammonia

(A) At lower temperature

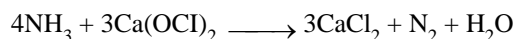


If excess of Cl_2 is used in this reaction, nitrogen trichloride is formed as per the following reaction,



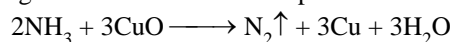
Nitrogen trichloride is an explosive substance.

- (b) By reaction of ammonia with calcium hypochlorite or Br_2

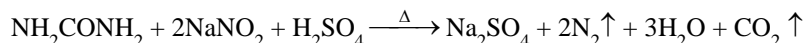


(B) At higher temperature

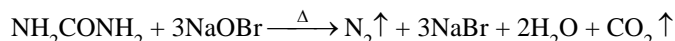
By passing ammonia over heated cupric oxide or PbO :



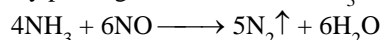
- (iv) By heating urea with a nitrite in presence of dilute H_2SO_4 :



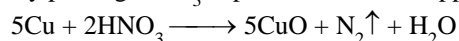
- (v) By heating urea solution with sodium hypobromite:



- (vi) By passing nitric oxide and NH_3 over red hot copper gauze:



- (vii) By passing HNO_3 vapours on red hot copper:



- (viii) Very pure nitrogen ; $\text{Ba}(\text{N}_3)_2 \xrightarrow{\Delta} \text{Ba} + 3\text{N}_2$

Sodium azide also gives N_2 on heating.

INDUSTRIAL METHODS OF PREPARATION:

- (i) From liquefied air by fractional distillation : The bp of N_2 is -196°C and that of oxygen is -183°C and hence they can be separated.

- (ii) From producer gas from furnaces : Producer gas is a mixture of CO and N_2 . When the mixture of CO and N_2 is passed over heated CuO , the CO gas is oxidized to CO_2 which is absorbed in alkalis & N_2 remains which is collected in gas cylinders.

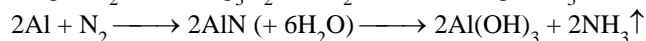
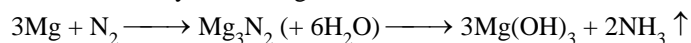
○ **PROPERTIES:**

- (i) N_2 is a colourless, odourless gas insoluble in water.

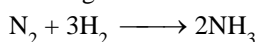
- (ii) It is non-polar covalent and neutral molecule.

- (iii) It is neither combustible nor a supporter of combustion.

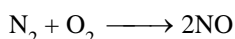
- (iv) It is absorbed by heated Mg and Al . The nitrides formed thus react with water to form NH_3 .



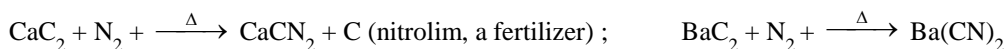
- (v) Reaction with H_2 : At 200 atm and 500°C , and in the presence of iron catalyst and molybdenum promoter, N_2 combines with H_2 reversibly to form ammonia. The process is called Haber's Process and is the industrial method of manufacturing ammonia. The reaction is exothermic.



- (vi) Reaction with oxygen: When air free from CO_2 and moisture is passed over an electric arc at about 2000 K, nitric oxide is formed. This reaction is endothermic.



- (vii) Reaction with CaC_2 and BaC_2 : At 1100°C, these carbides react with N_2 forming CaCN_2 and Ba(CN)_2 respectively.



CaCN_2 reacts with H_2O in the soil to produce NH_3 gas. NH_3 gas is converted by the nitrating bacteria present in soil into nitrates. (The nitrates are readily absorbed by the plants and meet their requirement of the element nitrogen.)

○ **USES :**

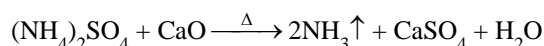
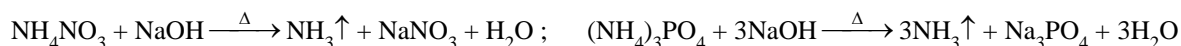
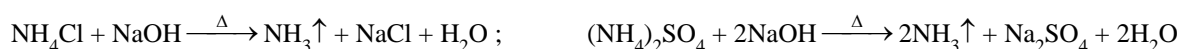
- for providing an inert atmosphere during many industrial processes where presence of air or O_2 is to be avoided.
- for manufacture of NH_3 by the Haber's process.
- for manufacture of HNO_3 by the Birkeland-Eyde process.
- for manufacture of nitrolim.

● **COMPOUNDS OF NITROGEN :**

○ **AMMONIA :**

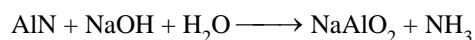
○ **PREPARATION :**

- (i) By the action of any base or alkali on any ammonium salt :

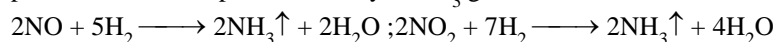


This is a general method and is used as a test for ammonium salts.

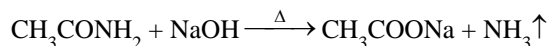
- (ii) By the hydrolysis of metal nitrides like AlN or Mg_3N_2 .



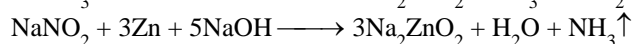
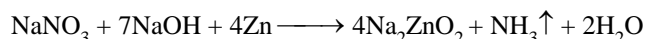
- (iii) From oxides of nitrogen: When oxides of nitrogen are mixed with H_2 and the mixture is passed over heated platinum catalyst, NH_3 gas is evolved.



- (iv) From organic amides: When an organic amide is heated with NaOH solution ammonia is evolved.

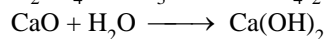
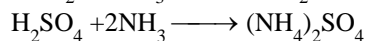
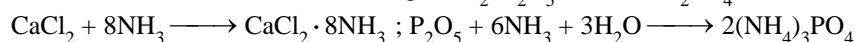


- (v) From nitrates and nitrites: When a metal nitrate or nitrite is heated with zinc powder and concentrated NaOH solution ammonia is obtained. The reactions are



Thus a nitrite or a nitrate can be identified by this reaction but this test cannot make distinction between them.

☞ The ammonia evolved is passed through quick lime to dry it and collected by the downward displacement of air. Ammonia cannot be dried using CaCl_2 , P_2O_5 , or conc. H_2SO_4 because NH_3 reacts with all of these.



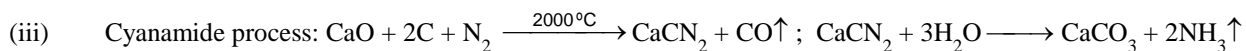
INDUSTRIAL METHODS OF PREPARATION:

- (i) Haber's process : $\text{N}_2 + 3\text{H}_2 \xrightarrow[\text{Iron oxide} + \text{K}_2\text{O} \& \text{Al}_2\text{O}_3]{500^\circ\text{C}, 200\text{atm.}} 2\text{NH}_3$

- (ii) From destructive distillation of coal: When coal is heated at a high temperature in an iron retort and the distillate is bubbled in water, three substances are obtained:

(a) Tarry black pitch, (b) Liquor ammonia & (c) Coal gas

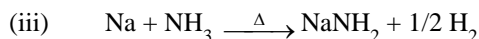
The liquor ammonia is a concentrated solution of ammonia and ammonium salts. When heated it gives out ammonia. When all the free NH_3 is obtained, the residual liquid is heated with Ca(OH)_2 when ammonium salts get decomposed to liberate further quantity of ammonia.



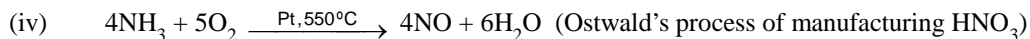
PROPERTIES:

(i) Colourless gas lighter than air. Being basic in nature turns red litmus blue. It acts as a Lewis base

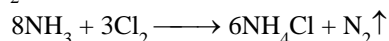
(ii) Highly soluble in water. The solution is called ammonium hydroxide solution.



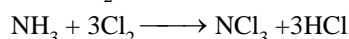
Amides decompose back with water to form NH_3 and NaOH .



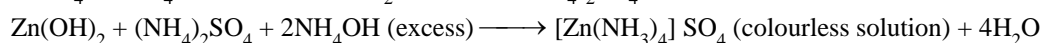
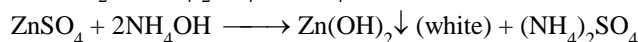
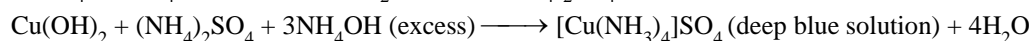
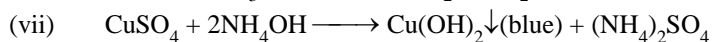
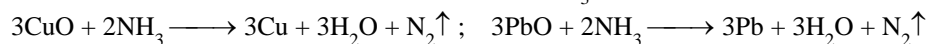
(v) When Cl_2 is bubbled in liquor ammonia, nitrogen gas is formed.



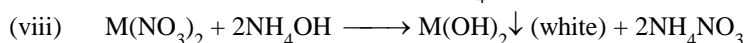
With excess of Cl_2 it is converted to an explosive substance, nitrogen trichloride.



(vi) Cupric oxide and PbO are reduced to metal when NH_3 is passed over heated CuO and PbO .

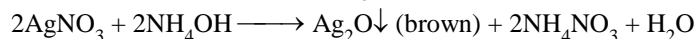


Similar reactions occur with CdSO_4 solution also.

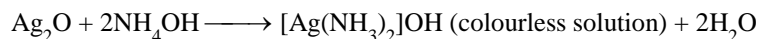


($\text{M} = \text{Mg}, \text{Ca}, \text{Sr}, \text{Ba}, \text{Ra}, \text{Sn}, \text{Pb}$)

(ix) When NH_4OH solution is added to AgNO_3 solution, a brown ppt. is obtained



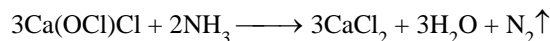
☞ Brown ppt of silver oxide formed dissolves in excess ammonium hydroxide forming a soluble complex



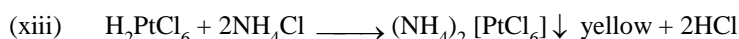
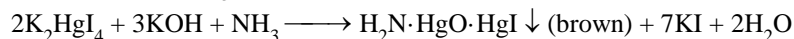
Similarly with mercuric salts, NH_4OH forms a white precipitate



(x) When liquor ammonia is dropped on heated bleaching powder, nitrogen gas is formed



(xii) When NH_3 gas is passed into the colourless solution of Nessler's reagent a brown precipitate or coloration is formed. This is a test for NH_3 gas.



USES: (i) Used as a refrigeration fluid. (ii) For the production of ammonium fertilizers such as ammonium sulphate, ammonium phosphate, ammonium nitrate, urea etc. (iii) For removing grease because NH_4OH dissolves grease. (iv) For manufacture of HNO_3 by the Ostwald process. (v) As a laboratory reagent. (vi) In the production of artificial rayon, silk, nylon etc.

(2) OXIDES OF NITROGEN :

Nitrogen forms a number of oxides, N_2O , NO , N_2O_3 , NO_2 or N_2O_4 and N_2O_5 , and also very unstable NO_3 and N_2O_6 . All these oxides of nitrogen exhibit *pf-pf* multiple bonding between nitrogen and oxygen.

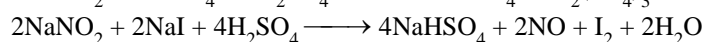
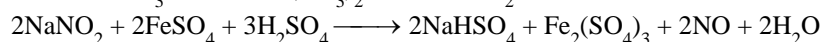
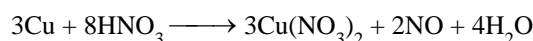
Name	Formula	Colour	Remarks
Nitrous oxide	N ₂ O	Colourless (g)	Neutral, +1 oxidation state
Nitric oxide	NO	Colourless (g)	Neutral, +2 oxidation state
Dinitrogen trioxide	N ₂ O ₃	Dark blue (s)	Acidic, +3 oxidation state
Nitrogen dioxide	NO ₂	Brown (g)	Acidic, +4 oxidation state
Dinitrogen tetroxide	N ₂ O ₄	Colourless (s)	Extensively dissociated to NO ₂ as gas and partly dissociated as liquid, acidic, +4 oxidation state
Dinitrogen pentoxide	N ₂ O ₅	Colourless (s)	Unstable as gas; ionic solid NO ₂ ⁺ , NO ₃ ⁻ , acidic, +5 oxidation state

PREPARATION :

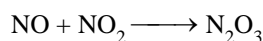
(i) N₂O is obtained generally by heating NH₄NO₃ with caution.



(ii) NO is best prepared by the reduction of 8 M HNO₃ with reducing agents like Cu or by reduction of nitrous acid or nitrites by Fe²⁺ or I⁻ ions.

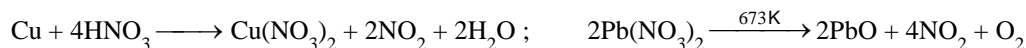


(iii) N₂O₃ is obtained as an intense blue liquid or a pale blue solid on cooling an equimolar mixture of NO and NO₂ to 250 K.



On warming, its colour fades due to its dissociation into these two oxides.

(iv) NO₂ can be prepared by reduction of concentrated HNO₃ with Cu or by heating heavy metal nitrates.



(v) N₂O₅ is an anhydride of HNO₃. It is best prepared by dehydrating HNO₃ by P₄O₁₀ at low temperatures.

**PROPERTIES:**

Oxides of nitrogen are all oxidizing agents, N₂O even supporting the combustion of S and P. NO which is thermally more stable, supports the combustion of Mg and P but not of S. Sulphur flame is not hot enough to decompose it.

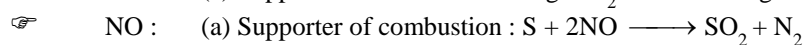
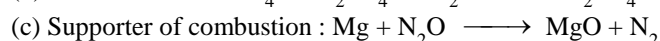
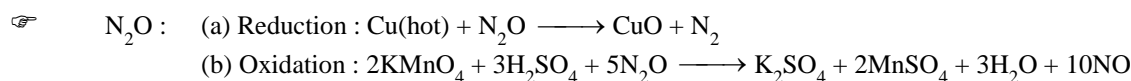
N₂O is isoelectronic with CO₂ and also has a linear structure. However, unlike CO₂, N₂O has a small dipole moment.

NO has a total of 15 electrons. It is impossible for all of them to be paired and hence this is an odd electron molecule. In the gaseous state, it is paramagnetic. However, the liquid and the solid states are diamagnetic because loose dimers are formed canceling out the magnetic effects of unpaired electrons.

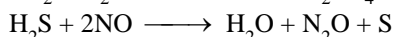
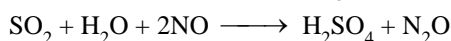
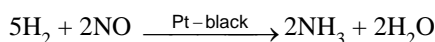
The brown ring formed in the test for nitrates is due to the formation of a complex of iron, [Fe(H₂O)₅NO]²⁺.

NO₂ with 23 electrons is again an odd electron molecule. In the gaseous state it is paramagnetic. On cooling, the gas condenses to a brown liquid and eventually to a colourless solid both of which are diamagnetic due to dimerisation.

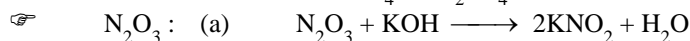
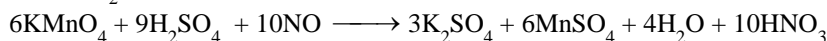
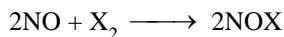
Liquid N₂O₄ undergoes self-ionization to form NO⁺ and NO₃⁻ ions and therefore, it has been extensively studied as a non-aqueous solvent.



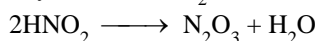
(b) Oxidising properties (Reduction of NO) :



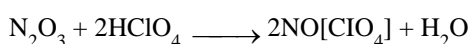
(c) Reducing properties (oxidation of NO) :



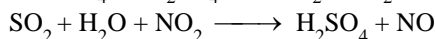
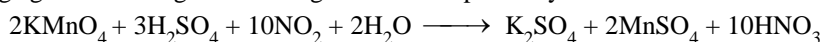
(b) It is anhydride of HNO_2 :



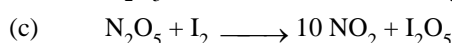
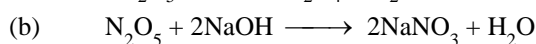
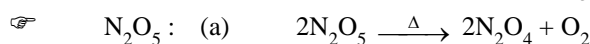
(c) with concentrated acids, form nitrosyl salts



☞ NO_2 : It behaves both like HNO_2 and HNO_3 . It behaves like HNO_2 as a reducing agent and like HNO_3 as an oxidising agent according to following reactions respectively.



* N_2O_4 is mixed anhydride of HNO_3 and HNO_2



(3) Oxyacids of nitrogen :

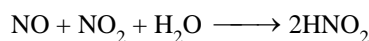
(A) NITROUS ACID (HNO_2) :

PREPARATION:

(i) By acidifying an aqueous solution of a nitrite



(ii) By passing an equimolar mixture of NO and NO_2 into water:



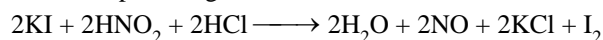
PROPERTIES :

(i) It is an unstable, weak acid which is known only in aqueous solution.

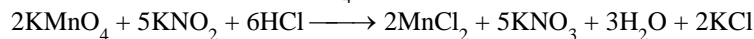
(ii) On trying to concentrate, the acid decomposes as given below.



(iii) Nitrous acid and nitrites are good oxidizing agents and convert iodides to iodine, ferrous salts to ferric, stannous to stannic and sulphites to sulphates eg.



(iv) With strong oxidizing agents like KMnO_4 nitrous acid and nitrites function as reducing agents and get oxidized to NO_3^- ions:

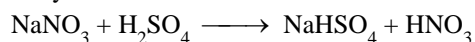


(iii) Nitrite ion is a good coordinating agent. Both nitrogen and oxygen have lone pairs capable of forming coordinate bond with metal ions. Nitrite ion can coordinate either through N or through O. (It is an ambidentate ligand) This generates linkage isomerism. Analogous organic derivatives are also known, the nitrites, RONO and the nitro compounds RNO_2 where R is any alkyl or aryl group.

(B) NITRIC ACID (HNO_3) :

PREPARATION :

(i) Laboratory method :



(ii) HNO_3 is now almost exclusively manufactured by the Ostwald process. In this process NH_3 is catalytically oxidized to NO over a Pt-Rh catalyst at 1200K.



About 96 to 98 % of NH_3 is converted into NO . The mixture is then diluted with air. NO combines with O_2 to give NO_2 which is absorbed in water to give HNO_3 and NO , which is then recycled.

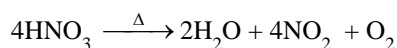


Nitric acid can be concentrated to 68 % by distillation, when a **constant boiling mixture** is formed. More concentrated acid can be made by distilling the mixture with concentrated sulphuric acid.

PROPERTIES :

(i) Pure nitric acid is colourless liquid (bp 359°C). It decomposes readily in light giving a yellow colour due to the formation of nitrogen dioxide. It is a strong acid and is almost completely dissociated into ions in solution.

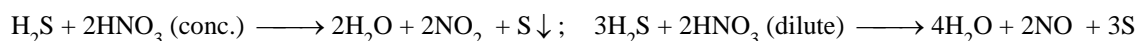
(ii) Thermal stability



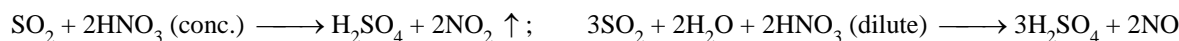
(iii) Oxidising properties



(a) Oxidises H_2S to sulphur

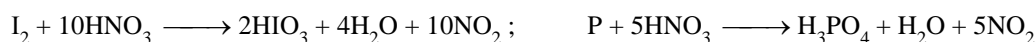
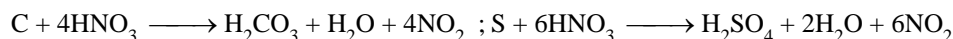


(b) Oxidises SO_2 to H_2SO_4



☞ Similarly oxidises ferrous salts to ferric salts, halogen acids to respective halogens.

(iv) Reaction with non-metals.

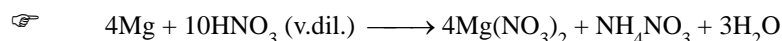
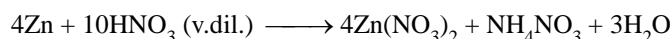
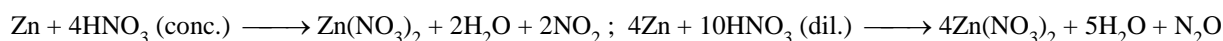


☞ These reactions are given with conc. HNO_3 .

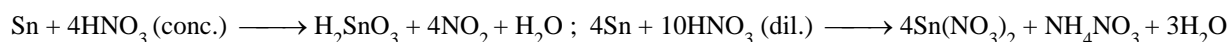
(v) Reaction with metals :

(A) Metals that are more electropositive than hydrogen

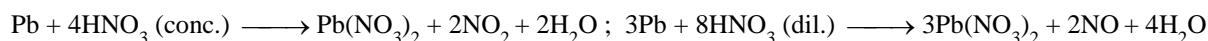
(a) Action on zinc or Fe :



(b) Action on tin :

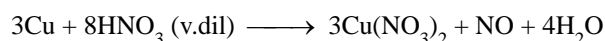
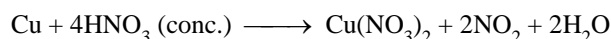


(c) Action on lead :



(B) Metals which are less electropositive than hydrogen.

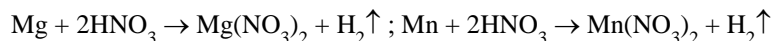
(i) Action on copper



(C) Metalloids : Sb and As.



☞ Mg & Mn only metals which produce hydrogen gas with cold (1 – 2%) HNO_3 .

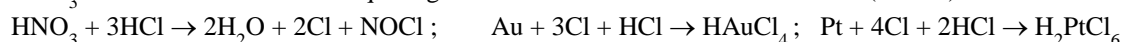


p-BLOCK ELEMENTS

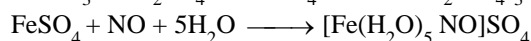
☞ As concentrated nitric acid (80%) behaves as an oxidizing agent and metals such as Al, Fe, Cr etc are rendered passive due to the formation of a tenacious layer of insoluble oxide on the metal surface.

☞ Noble metals such as Au, Pt, Rh, and Ir are not attacked by nitric acid. However a 1:3 mixture of conc.

HNO_3 and conc. HCl known as aqua regia dissolves Au and Pt as it contains free(atomic) chlorine:



Brown ring test :



2. PHOSPHORUS :

It is a very reactive non-metal. It catches fire in air. It occurs in nature in the form of stable phosphates. (Animal bones also contain calcium phosphate (58 %)). The important minerals are:

(i) Phosphorite, $\text{Ca}_3(\text{PO}_4)_2$

(ii) Chlorapatite, $\text{Ca}_3(\text{PO}_4)_2\text{CaCl}_2$

(iii) Fluorapatite, $\text{Ca}_3(\text{PO}_4)_2\text{CaF}_2$

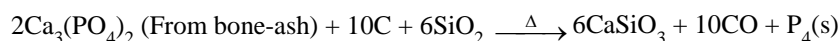
(iv) Vivianite, $\text{Fe}_3(\text{PO}_4)_2 \cdot 8\text{H}_2\text{O}$

(v) Redonda phosphate, AlPO_4

ALLOTROPIC FORMS OF PHOSPHORUS:

(i) **White or yellow phosphorus (P_4):**

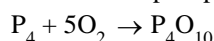
PREPARATION:



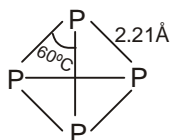
PROPERTIES:

It is white-to-transparent and soft waxy solid. Its density is 1.8 g/cc at 20°C. Its mp and bp are 44°C and 287°C respectively.

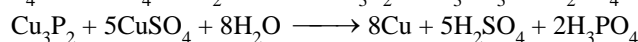
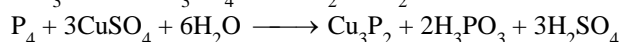
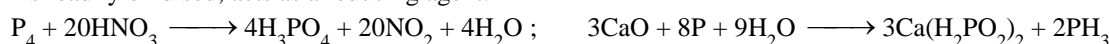
It is soluble in CS_2 but insoluble in water. It glows in dark due to slow oxidation producing yellowish-green light. This phenomenon is called phosphorescence



White phosphorus is poisonous. It turns yellow after some time; it is called yellow phosphorus. It undergoes oxidation in the presence of air which slowly raises its temperature and due to its low ignition temperature (~ 30°C) after a few moments it catches fire spontaneously. Due to this reason, it is stored under water.



* As readily oxidised, acts as a reducing agent.

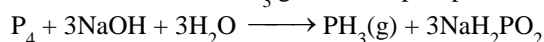


Colloidal solution of gold may be prepared by reducing a solution of gold chloride with phosphours dissolved in ether.

(ii) **Red phosphorus:**

PREPARATION :

When white phosphorus is heated in the atmosphere of CO_2 or coal gas at 573 K red phosphorus is produced. This red phosphorus may still contain some white phosphours which is removed by boiling the mixture with NaOH where white phosphorus is converted into PH_3 gas but red phosphorus remains inert.

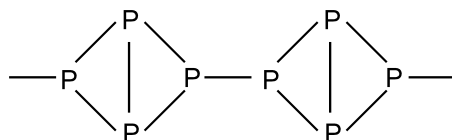


☞ It is also prepared by heating white phosphorus with a few crystals of iodine catalyst at 250°C under high pressure in absence of air.

PROPERTIES :

It is a red crystalline solid having a density of 2.2 g/cc. It is less reactive than white phosphorus and does not dissolve in liquid CS_2 . It does not catch fire at room temperature because its ignition temperature is 260°C.

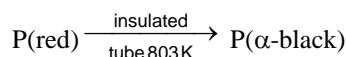
It is a polymeric substance forming linear chains like this.



(iii) Black phosphorus:

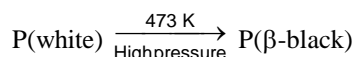
It has two forms α -black phosphorus and β -black phosphorus

(a) α -black phosphorus

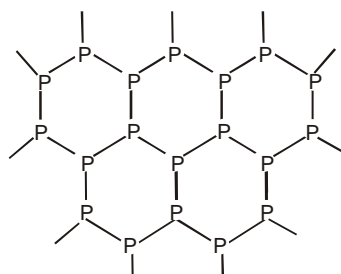


α -black phosphorus structure is not definite and is non conductor of electricity.

(b) β -black phosphorus



β -black phosphorus is an electrical conductor resembling graphite in this respect and also in its flakiness and luster. It is insoluble in CS_2 . It has a layered structure like graphite.



(iv) Brown phosphorus:

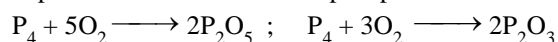
Above 1600°C , P_4 molecules begin to dissociate into P_2 molecules. Rapid cooling of this vapour gives brown phosphorus which probably contains P_2 molecules.

CHEMICAL PROPERTIES OF PHOSPHORUS :

☞ Reactivity of the various allotropic forms of phosphorus towards other substances decreases in the order: Brown > white > red > black, the last one being almost inert. Apart from their reactivity difference, all the forms are chemically similar.

(i) Action of air :

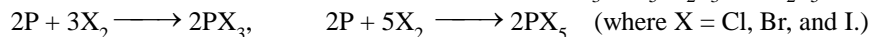
White phosphorus burns in air to form phosphorus trioxide and pentoxide.



Red and other forms of phosphorus also burn in air or oxygen but on heating.

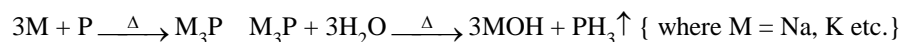
(ii) Action of non-metals:

When heated with non-metals phosphorus forms compounds PX_3 , PX_5 , P_2S_3 and P_2S_5 .



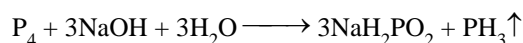
(iii) Action with metals:

Alkali metals when heated with white phosphorus in vacuum produce alkali metal phosphide, which react with water to form phosphine gas.



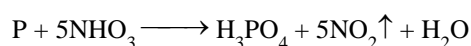
(iv) Action of NaOH:

When white phosphorus is heated with NaOH solution, phosphine gas is evolved.



(v) Action of conc. HNO_3 :

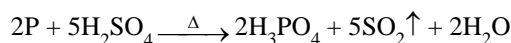
When heated with conc. HNO_3 , phosphorus is oxidized to H_3PO_4 .



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(vi) Action of conc. H_2SO_4 :

When heated with conc. H_2SO_4 phosphorus is oxidized to H_3PO_4 .



☞ Conversion of white phosphorus to red phosphorus and red phosphorus to white phosphorus.

Heat in presence of I_2 catalyst at 250°C and high pressure in absence of air

White phosphorus \rightleftharpoons Red phosphorus

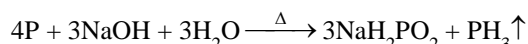
Heat above 250°C in presence of inert gas and condense in water

COMPOUNDS OF PHOSPHORUS :

(1) PHOSPHINE :

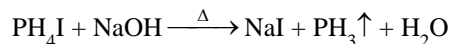
PREPARATION:

(i) By heating white phosphorus with NaOH solution in presence of coal gas. Phosphine gas is collected by the downward displacement of water.

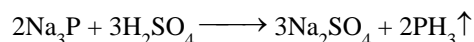


☞ Coal gas is used to prevent oxidation of phosphine. The phosphine gas is contaminated with a combustible gas P_2H_4 . It is separated from PH_3 by passing the gaseous mixture through a freezing mixture whereupon P_2H_4 condenses to a liquid and PH_3 is collected by the downward displacement of air. Pure PH_3 does not burn in air.

(ii) By the action of alkalies on phosphonium salts:



(iii) By the action of dil. HCl or dil. H_2SO_4 on metal phosphides.



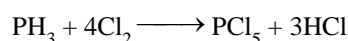
(iv) $\text{H}_3\text{PO}_2 + 4\text{H} \xrightarrow{\text{Zn}/\text{HCl}} \text{PH}_3 + 2\text{H}_2\text{O}$

PROPERTIES :

(i) It is a colourless gas with a smell of rotten fish and is neutral to litmus paper, It is heavier than air and only slightly soluble in water, It is a poisonous gas and acts as a Lewis base.

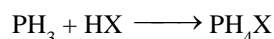
(ii) Action of chlorine:

It reacts with Cl_2 to form PCl_5 .



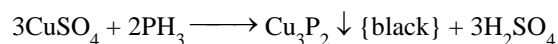
☞ Similar reaction occurs between heated phosphine and Br_2 and I_2 vapours.

(iii) Reaction with hydrogen halides: When PH_3 is bubbled in solution of HX , phosphonium halides are formed.



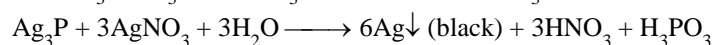
(iv) Action on CuSO_4 solution:

When PH_3 is bubbled in acidic solution of copper sulphate, a black precipitate of copper phosphide is formed.



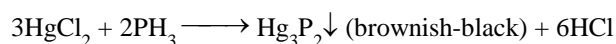
(v) Reaction with AgNO_3 solution:

When PH_3 gas is bubbled in a solution of AgNO_3 , a yellow precipitate of silver phosphide, Ag_3P , is first formed which later decomposes to black Ag .



(vi) Reaction with mercuric chloride solution:

When mercuric chloride solution is treated with PH_3 gas, a brownish black precipitate of mercuric phosphide is formed.



* Samples of PH_3 can be dried using quick lime, or NaOH sticks. It cannot be dried by conc. H_2SO_4 because of its reaction with it.

USES :

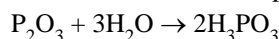
It is used in making smoke signals and producing smoke screens in battle-fields.

2. OXIDES OF PHOSPHORUS :**(A) PHOSPHORUS TRIOXIDE (P_2O_3) :****PREPARATION :**

It is prepared by burning phosphorus in a limited supply of oxygen when gaseous P_4O_{10} and P_4O_6 are formed. On lowering the temperature using a condenser, P_4O_6 remains in gaseous form whereas P_4O_{10} condenses as a solid which is stopped by glasswool. On passing the remaining gaseous mixture through freezing mixture, it converts into colourless crystals of P_4O_6 .

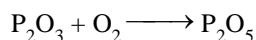
PROPERTIES :

- (i) It is colourless crystalline solid having mp 23.8°C and bp 178°C .
- (ii) It dissolves in cold water to form phosphorus acid. It is thus the anhydride of phosphorus acid.

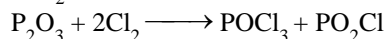


- (iii) It dissolves in hot water liberating PH_3
- $$2\text{P}_2\text{O}_3 + 6\text{H}_2\text{O} \longrightarrow 3\text{H}_3\text{PO}_4 + \text{PH}_3$$

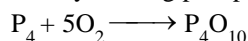
- (iv) It slowly gets oxidized in air to form P_2O_5



- (v) It burns in Cl_2 gas forming phosphorus oxytrichloride (POCl_3) and phosphoryl chloride (PO_2Cl)

**(B) PHOSPHORUS PENTOXIDE (P_2O_5) :****PREPARATION :**

It is obtained by burning phosphorus in air.

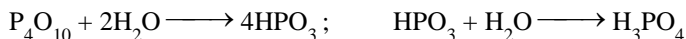
**PROPERTIES:**

- (i) It is a white powder acidic in nature and is the anhydride of orthophosphoric acid. Its empirical formula is P_2O_5 and its molecular formula is P_4O_{10} .

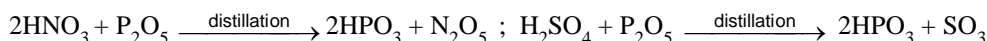
- (ii) It sublimes on heating at 250°C .

(iii) Action of water:

It dissolves in water with hissing sound forming metaphosphoric acid and finally orthophosphoric acid.

**(iv) Dehydrating power:**

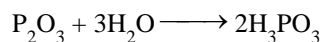
It dehydrates conc. H_2SO_4 and conc. HNO_3 to SO_3 and N_2O_5 respectively.

**USES :**

- (i) For drying acidic gases
- (ii) As a dehydrating agent
- (iii) For the preparation of SO_3 and N_2O_5
- (iv) For the preparation of phosphoric acid

3. OXY-ACIDS OF PHOSPHORUS :**(A) PHOSPHORUS ACID (H₃PO₃):****PREPARATION:**

- (i) By dissolving P
- ₂
- O
- ₃
- in water:

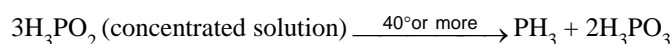


- (ii) By hydrolysis of PCl
- ₃
- with water:

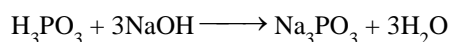


The solution containing H₃PO₃ and HCl is heated to 180°C and HCl gas is driven out. The resulting solution on crystallization gives white crystals of H₃PO₃.

- (iii) By heating hypophosphorous acid :

**PROPERTIES:**

- (i) It is a white crystalline solid, soluble in water and having melting point of 74°C.
 (ii) It is a weak acid and a reducing agent
 (iii) When neutralized with bases or alkalis, it forms neutral salts called phosphites which are unstable.

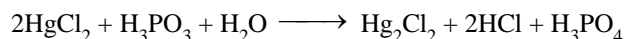
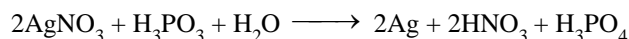


Its basicity is 2 because it has only two replaceable H atoms. The third H atom is not replaceable because it is not connected to oxygen.

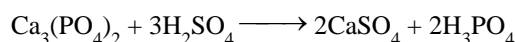
- (iv)
- $4H_3PO_3 \xrightarrow{\Delta} 3H_3PO_4 + PH_3$
- (Disproportionation)

- (v)
- $H_3PO_3 + 3PCl_5 \longrightarrow PCl_3 + 3POCl_3 + 3HCl$

- (vi) It is a strong reducing agent :

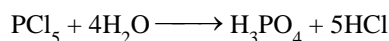
**(B) ORTHOPHOSPHORIC ACID (H₃PO₄):****PREPARATION:**

- (i) By heating calcium phosphate with conc. H
- ₂
- SO
- ₄

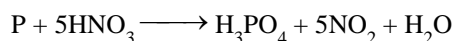


CaSO₄ is insoluble. Solution of H₃PO₄ is separated from CaSO₄. It is then concentrated by evaporating it at 180°C and dehydrated by conc. H₂SO₄ placed in a vacuum desiccator cooled by freezing mixture. White crystals of H₃PO₄ are thus formed.

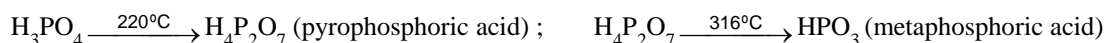
- (ii) By hydrolysis of PCl
- ₅
- :

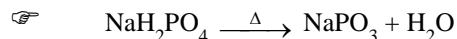


- (iii) By heating white phosphorus with conc. HNO
- ₃
- :

**PROPERTIES :**

- (i) Pure orthophosphoric acid is a white crystalline solid highly soluble in water having melting point of 42°C. It is a weak acid. It forms two acid salts and one normal salt. NaH₂PO₄ is sodium dihydrogen phosphate, Na₂HPO₄ is disodium hydrogen phosphate & Na₃PO₄ is sodium orthophosphate.
 (ii) **Action of heat :**



(iii) Neutralization with alkalis or bases:**USES:**

It is used as a laboratory reagent and in manufacture of medicines.

GROUP SIXTEEN ELEMENTS : THE OXYGEN FAMILY

Oxygen, sulphur, selenium, tellurium and polonium constitute group 16 of the periodic table. This is sometimes known as group of chalcogens. The name is derived from the greek word for brass and points to the association of sulphur and its congeners with copper. Most copper minerals contain either oxygen or sulphur and frequently the other members of the group.

Occurrence :

Oxygen is the most abundant of all the elements on the earth. Oxygen forms about 46.6% by mass of earth's crust . Dry air contains 20.946% oxygen by volume.

However, the abundance of sulphur in the earth's crust is only 0.03-0.1%. Combined sulphur exists primarily as sulphates such as gypsum $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, epsom salt $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, baryta BaSO_4 and sulphides such as galena PbS , zinc blende ZnS , copper pyrites CuFeS_2 . Traces of sulphur occur as hydrogen sulphide in volcanoes.

Selenium and tellurium are also found as metal selenides and tellurides in sulphide ores. Polonium occurs in nature as a decay product of thorium and uranium minerals.

Electronic Configuration :

The elements of group 16 have six electrons in the outermost shell and have $ns^2 np^4$ general electronic configuration.

Atomic and Ionic Radii :

Due to increase in the number of shells , atomic and ionic radii increase from top to bottom in the group. The size of oxygen atoms is however, exceptionally small .

Ionisation Enthalpy :

Ionisation enthalpy decrease down the group. It is due to increase in size. However, the element of this group have lower ionisation enthalpy values compared to those of group 15 in the corresponding periods. This is due to the fact that group 15 element have extra stable half-filled p orbitals electronic configurations.

Electron Gain Enthalpy :

Because of the compact nature of oxygen atom, it has less negative electron gain enthalpy than sulphur. However from sulphur onwards the value again becomes less negative upto polonium.

Electronegativity :

Next to fluorine, oxygen has the highest electronegativity value amongst the elements. Within the group, electronegativity decrease with an increase in atomic number. This implies that the metallic character increase from oxygen to polonium.

Physical Properties :

Oxygen and sulphur are non-metal, selenium and tellurium metalloids, whereas polonium is a metal. Polonium is radioactive and is short lived (Half-life 13.8 days). All these element exhibit allotropy. The melting and boiling points increase with an increase in atomic number down the group. The larger difference between the melting and boiling points of oxygen and sulphur may be explained on the basis of their atomicity; oxygen exist as diatomic molecules (O_2) whereas sulphur exists as polyatomic molecule (S_8).

p-BLOCK ELEMENTS

ATOMIC & PHYSICAL PROPERTIES :

Element		O	S	Se	Te
Atomic Number		8	16	34	52
Atomic Mass		16	32.06	78.96	127.6
Electronic configuration		[He] 2s ² 2p ⁴	[Ne] 3s ² 3p ⁴	[Ar] 3d ¹⁰ 4s ² 4p ⁴	[Kr] 4d ¹⁰ 5s ² 5p ⁴
Covalent Radius / pm		74	103	119	142
Ionic Radius X ⁻² / pm		140	184	198	221
Ionization enthalpy / (kJ mol ⁻¹)	I	1314	1000	941	869
	II	3388	2251	2045	1790
Electronegativity		3.5	2.44	2.48	2.01
Density/[g cm ⁻³ (293 K)]		1.32	2.06	4.19	6.25
Melting point / K		54	393	490	725
Boiling point / K		90	718	958	1260

Chemical Properties :

Oxidation states and trends in chemical reactivity :

The elements of group 16 exhibit a number of oxidation states. The stability of -2 oxidation state decreases down the group. Polonium hardly shows -2 oxidation states. Since electronegativity of oxygen is very high, it shows only negative oxidation states as -2 except in the case of OF₂ where its oxidation states is + 2. Other elements of the group exhibit + 2 + 4 + 6 oxidation states but + 4 and + 6 are more common. Sulphur, selenium and tellurium usually show + 4 oxidation in their compounds with oxygen and +6 oxidations state with fluorine. The stability of +6 oxidation state decreases down the group and stability of + 4 oxidation state increases (inert pair effect). Bonding in + 4 and + 6 oxidation states are primarily covalent.

Anomalous behaviour of oxygen :

The anomalous behaviour of oxygen, like other member of p-block present in second period is due to its small size and high electronegativity. One typical example of effects of small size and high electronegativity is the presence of strong hydrogen bonding in H₂O which is not found in H₂S.

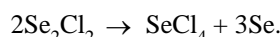
The absence of d orbitals in oxygen limits its covalency to four and in practice, rarely exceeds two. On the other hand, in case of other elements of the group, the valence shell can be expanded and covalence exceeds four.

- (i) **Reactivity with hydrogen :** All the elements of group 16 form hydrides of the type H₂E (E = S, Se., Te, Po). Some properties of hydrides are given in Table. Their acidic character increases from H₂O to H₂Te. The increase in acidic character can be explained in terms of decrease in bond (H-E) dissociation enthalpy down the group. Owing to the decrease in bond (H-E) dissociation enthalpy down the group, the thermal stability of hydrides also decreases from H₂O to H₂Po. All the hydrides except water possess reducing property and this character increases from H₂S to H₂Te.

Table : Properties of Hydrides of Group 16 Elements

Property	H ₂ O	H ₂ S	H ₂ Se	H ₂ Te
m.p./K	273	188	208	222
b.p./K	373	213	232	269
H-E distance/pm	96	134	146	169
HEH angle (°)	104	92	91	90
Δ _f H/kJ mol ⁻¹	-286	-20	73	100
Δ _{diss} H (H-E)/kJ mol ⁻¹	463	347	276	238
Dissociation constant ^a	1.8 × 10 ⁻¹⁶	1.3 × 10 ⁻⁷	1.3 × 10 ⁻⁴	2.3 × 10 ⁻³

- (ii) **Reactivity with oxygen :** All these elements form oxides of the EO_2 and EO_3 types where $E = S, Se, Te$ or Po . Ozone (O_3) and sulphur dioxide (SO_2) are gases while selenium dioxide (SeO_2) is solid. Reducing property of dioxide decreases from SO_2 to TeO_2 ; SO_2 is reducing while TeO_2 is an oxidising agent. Besides EO_2 type sulphur, selenium and tellurium also form EO_3 type oxide (SO_3, SeO_3, TeO_3). Both types of oxides are acidic in nature.
- (iii) **Reactivity toward the halogens :** Elements of group 16 form a larger number of halides of the type EX_6, EX_4 and EX_2 where E is an element of the group and X is a halogen. The stabilities of the halides decrease in the order $F > Cl > Br > I$. Amongst hexahalides, hexafluorides are the only stable halides. All hexafluorides are gaseous in nature. They have octahedral structure. Sulphur hexafluoride SF_6 is exceptionally stable for steric reasons. Amongst tetrafluorides, SF_4 is a gas, SeF_4 liquid and TeF_4 a solid. These fluorides have sp^3d hybridisation and thus, have trigonal bipyramidal structure in which one of the equatorial positions is occupied by a lone pair of electrons. This geometry is also regarded as see-saw geometry. All elements except selenium form dichlorides and dibromides. These dihalides are formed by sp^3 hybridisation and thus have tetrahedral structure. The well known monohalides are dimeric in nature, Examples are $S_2F_2, S_2Cl_2, S_2Br_2, Se_2Cl_2$ and Se_2Br_2 . These dimeric halides undergo disproportionation as given below :



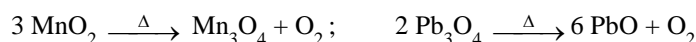
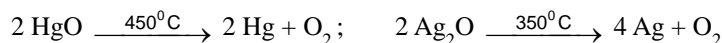
1. DIOXYGEN (O_2) :

It differs from the remaining elements of the VIth group because of the following properties.

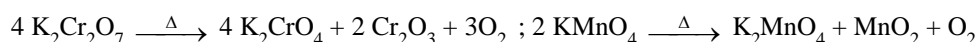
(A) small size (B) high electronegativity and (C) non-availability of d-orbitals.

PREPARATION :

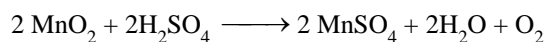
- (i) By thermal decomposition of oxides of metals.



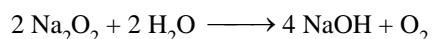
- (ii) By thermal decomposition of oxygen rich compounds.



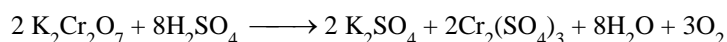
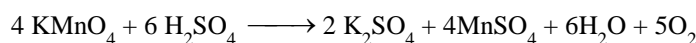
- (iii) By the action of conc. H_2SO_4 on MnO_2 .



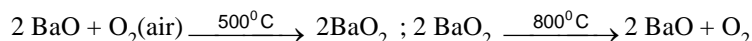
- (iv) By the action of water on Na_2O_2 .



- (v) By the action of conc. H_2SO_4 on $KMnO_4$ or $K_2Cr_2O_7$.



- (vi) **By Brins process (mfg.) :**



- (vii) **From air (mfg.) :**

Oxygen is obtained by liquefaction of air and then its fractional distillation.

PROPERTIES:

Colourless, odourless and tasteless gas. It is paramagnetic and exhibits allotropy. Three isotopes of oxygen are $^{16}_8O$, $^{17}_8O$ and $^{18}_8O$. Oxygen does not burn but is a strong supporter of combustion.

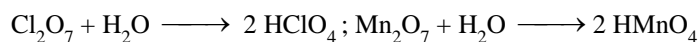
USE :

- (i) Oxygen mixed with helium or CO_2 is used for artificial respiration. (ii) Liquid oxygen is used as oxidising agent in rocket fuels. (iii) Oxygen is used for production of oxy-hydrogen or oxy-acetylene flames employed for cutting and welding.

2. OXIDES :

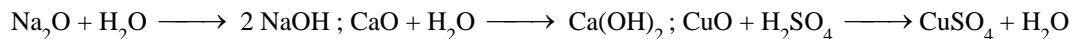
(i) Acidic oxides

They dissolve in water forming oxyacids, e.g., CO_2 , SO_2 , SO_3 , N_2O_5 , N_2O_3 , P_4O_6 , P_4O_{10} , Cl_2O_7 , CrO_3 , Mn_2O_7 , V_2O_5 .



(ii) Basic oxides

They either dissolve in water to form alkalis or combine with acids to form salts and water or combine with acidic oxides to form salts; e.g., Na_2O , CaO , CuO , FeO , BaO etc.

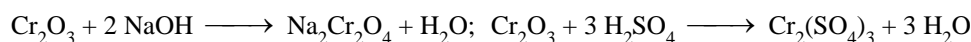
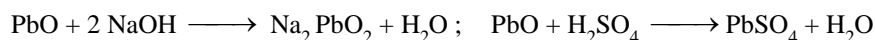


(iii) Neutral Oxides

They neither combine with acids nor with the bases to form salts e.g., CO , N_2O , NO etc.

(iv) Amphoteric Oxides

These can combine with acids as well as bases e.g., ZnO , Al_2O_3 , BeO , Sb_2O_3 , Cr_2O_3 , PbO etc.



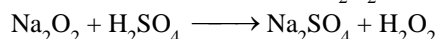
(v) Mixed Oxides

They behave as mixture of two simple oxides,

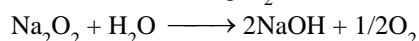
e.g., Pb_3O_4 ($2\text{PbO} + \text{PbO}_2$), Fe_3O_4 ($\text{FeO} + \text{Fe}_2\text{O}_3$), Mn_3O_4 ($2 \text{MnO} + \text{MnO}_2$)

(vi) Peroxides

They react with dilute acids and form H_2O_2 , e.g., Na_2O_2 , K_2O_2 , BaO_2 etc.



They react with water forming O_2 .



(vii) Dioxides

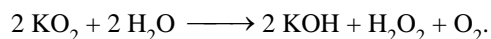
Like peroxide, they contain excess of oxygen but do not yield H_2O_2 with dil. acids e.g. PbO_2 , MnO_2 etc.

They evolve Cl_2 with conc. HCl and O_2 with conc. H_2SO_4 .



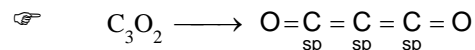
(viii) Super Oxides

They contain O_2^- ion, e.g., KO_2 , RbO_2 and CsO_2 . These oxides react with water forming H_2O_2 and O_2



(ix) Sub Oxides

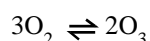
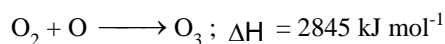
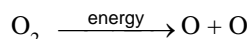
They contain less oxygen than expected from the normal valency of the elements e.g., C_3O_2 , N_2O , Pb_2O , Hg_2O etc.



2. OZONE (O_3) :

PREPARATION:

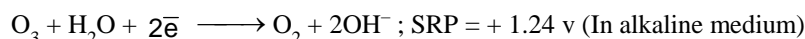
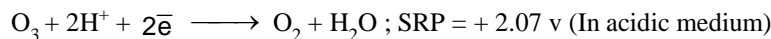
It is prepared by passing silent electric discharge through pure and dry oxygen



Mixture obtained contains 5-10% ozone by volume and this mixture is called ozonised oxygen. The apparatus used for this is called as ozoniser ((i) Simen's and (ii) Brodie's ozonisers).

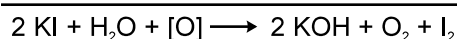
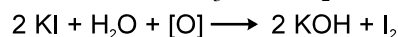
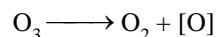
PROPERTIES :

- (i) Pale blue gas which forms a blue liquid on cooling and on solidification forms violet black crystals. It has a strong fish-like smell and is slightly soluble in water but more in turpentine oil, glacial acetic acid or CCl_4 . O_3 molecule is diamagnetic but O_3^- is paramagnetic.

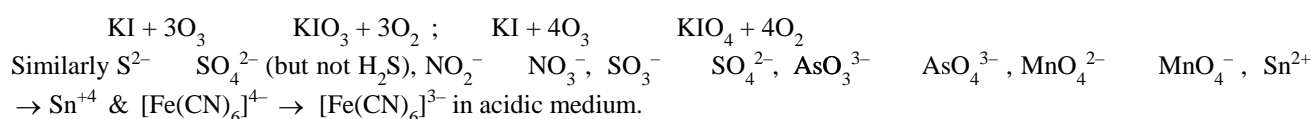
(i) **Oxidising agent :**

Therefore, ozone is a strong oxidising agent in acidic medium.

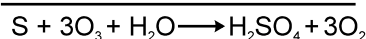
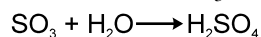
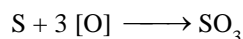
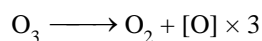
- (a) It oxidises I^- to I_2 (from neutral solution of KI)



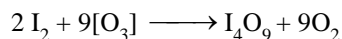
☞ Alkaline KI is oxidised to potassium iodate & periodate.



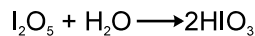
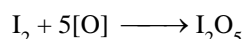
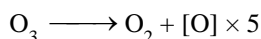
- (b) It oxidises moist S, P, As into their oxy acids.



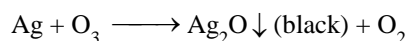
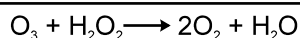
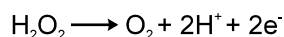
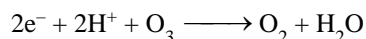
- (c) It oxidises H_2S to S

(ii) **Reaction with dry I_2 :**

☞ I_4O_9 yellow solid has the composition $\text{I}^{+3}(\text{IO}_3^-)_3$. Formation of this compound is a direct evidence in favour of basic nature of I_2 (i.e. its tendency to form cations).

(iii) **Reaction with moist iodine :**(iv) **Reaction with Silver :**

Silver articles become black in contact with ozone.

(v) **Reaction with H_2O_2 :**

Oxidising agent Reducing agent

p-BLOCK ELEMENTS

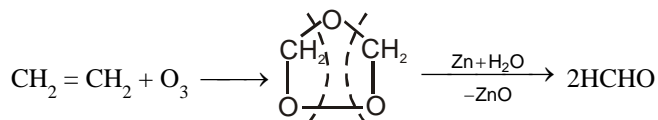
It is supported by the fact that SRP of ozone is higher (+2.07) than SRP of hydrogen peroxide (+1.77). Therefore, ozone is stronger oxidising agent than hydrogen peroxide.

(vi) **Bleaching Action :**

O₃ also bleaches coloured substances through oxidation

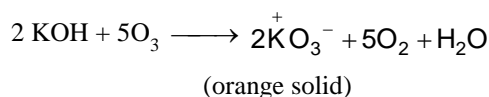
(vii) **Ozonolysis :**

Alkenes, alkynes react with ozone forming ozonides.



(viii) **Reaction with KOH :**

Forms orange coloured compound, potassium ozonide.

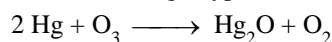


TESTS FOR OZONE

(i) A filter paper soaked in an alcoholic benzidine $\left[\text{H}_2\text{N}-\text{C}_6\text{H}_4-\text{C}_6\text{H}_4-\text{NH}_2 \right]$ becomes brown when brought in contact with O₃ (this is not shown by H₂O₂)

(ii) **Tailing of mercury**

Pure mercury is a mobile liquid but when brought in contact with O₃ its mobility decreases and it starts sticking to glass surface forming a type of tail due to the dissolution of Hg₂O (mercury sub-oxide) in Hg.



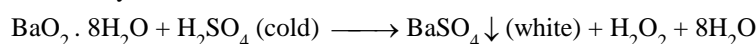
USES :

- (i) As a germicide and disinfectant for sterilising water and improving the atmosphere of crowded places.
- (ii) For detecting the position of double bond in the unsaturated organic compounds.
- (iii) In mfg. of artificial silk, synthetic camphor, KMnO₄ etc.

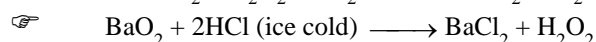
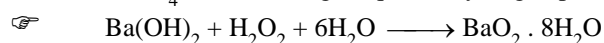
3. HYDROGEN PEROXIDE (H₂O₂) :

PREPARATION :

(i) Laboratory method



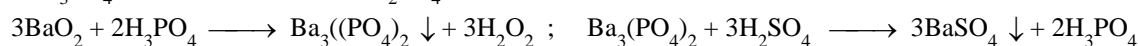
BaSO₄ is filtered to get aqueous hydrogen peroxide.



Not possible to isolate H₂O₂ from this solution as BaCl₂ is soluble in water

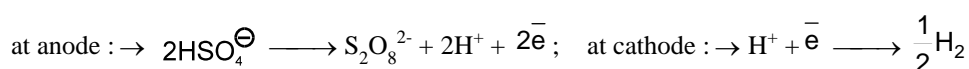
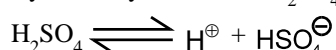
The reaction between anhydrous BaO₂ and H₂SO₄ is slow and practically ceases after sometimes due to the formation of a protective layer of BaSO₄ on BaO₂.

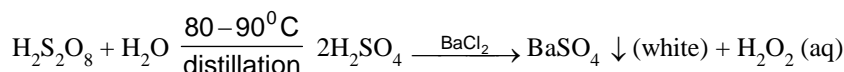
Since H₂SO₄ can decompose H₂O₂ at a higher temperature, therefore, the reaction should be carried out at low temperature or H₃PO₄ can be used in place of H₂SO₄



H₃PO₄ can be used again.

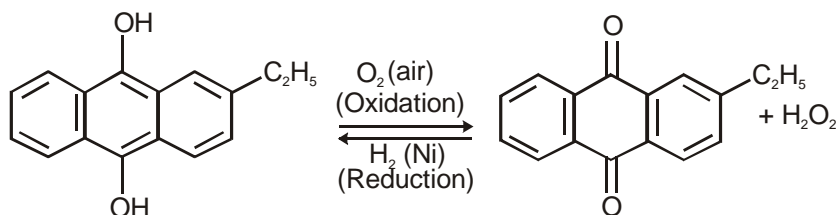
(ii) By electrolysis of conc. H₂SO₄ at 0°C using inert electrodes (platinum).





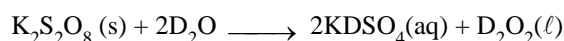
BaSO_4 is removed by filtration to get aqueous H_2O_2 .

(iii) **Industrial method (Auto oxidation):**



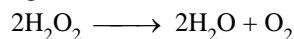
2-Ethyl anthraquinol

2-Ethylanthraquinone



PROPERTIES :

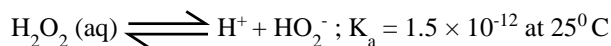
- (i) Colourless viscous liquid which appears blue in the larger quantity and is soluble in water (due to H-bonding) in all proportions and form a hydrate $\text{H}_2\text{O}_2 \cdot \text{H}_2\text{O}$ (mp 221 K)
- (ii) Its boiling point 144°C more than water but freezing point (-4) less than water. Density and dielectric constant are also higher than H_2O
- (iii) Its aqueous solution is more stable than the anhydrous liquid where it decomposes into water and O_2 slowly on exposure to light.



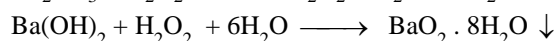
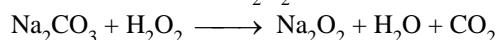
H_2O_2 is not kept in glass containers because traces of alkali metal ions from the glass can catalyse the explosive decomposition of H_2O_2 . Therefore, aqueous solution is stored in plastic or wax-lined glass containers and some urea or phosphoric acid or glycerol is added to that solution because these compounds have been found to behave as negative catalyst for the decomposition of H_2O_2 .

(iv) **Acidic nature :**

Behaves as a weak acid according to the following equation

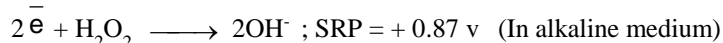


Aqueous solution of H_2O_2 turns blue litmus red which is then bleached by the oxidising property of H_2O_2



☞ A 30% H_2O_2 solution has pH = 4.0

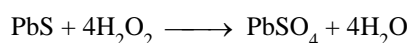
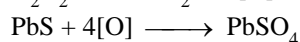
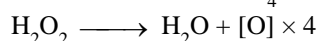
(v) **Oxidising Agent :**



☞ On the basis of the above potentials, we can say that H_2O_2 is strong oxidising agent in acidic medium but kinetically it is found that reactions are faster in basic medium.

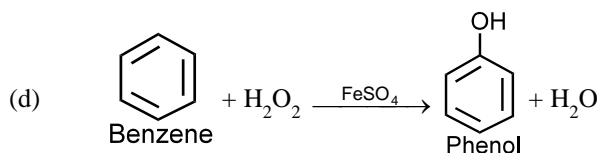
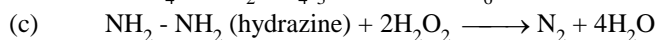
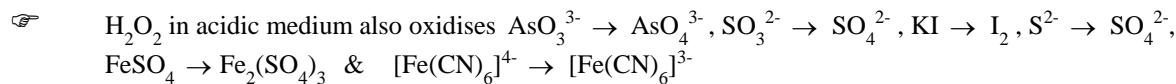
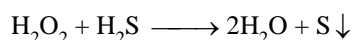
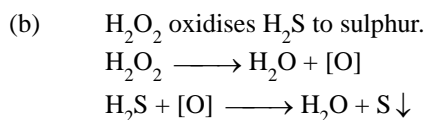
(A) **In acidic medium :**

(a) It oxidises PbS to PbSO_4 .

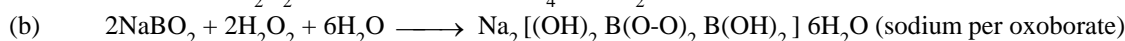
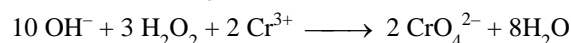
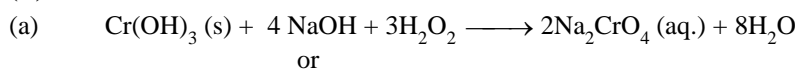


This property is utilised in restoring the white colours in old paintings which turns black due to the formation of PbS by the action of atmospheric H_2S .

p-BLOCK ELEMENTS



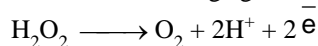
(B) In alkaline medium :



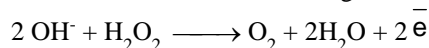
☞ Used as a brightner in washing powder.

(vi) Reducing Agent :

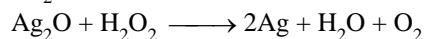
It acts as a reducing agent towards powerful oxidising agent.



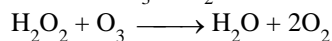
☞ In alkaline solution, its reducing character is more than in acidic medium.



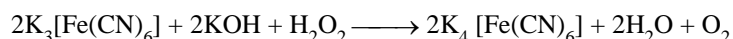
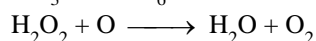
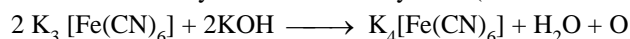
(a) Ag_2O is reduced to Ag.



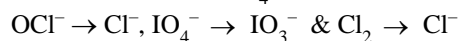
(b) It reduces O_3 to O_2 .



(c) It reduces ferric cyanide to ferrous cyanide (basic medium).

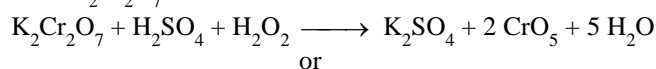


☞ It also reduces $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$ (acidic medium), $\text{MnO}_4^- \rightarrow \text{MnO}_2$ (basic medium),

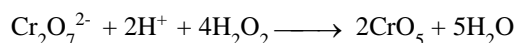


TESTS FOR H_2O_2 :

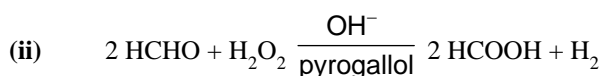
(i) With $\text{K}_2\text{Cr}_2\text{O}_7$



or



☞ CrO_5 bright blue coloured compound soluble in ether.



When this reaction is carried out in dark, it is accompanied by emission of light (yellow coloured). It is an example of chemiluminescence.

- (iii) An acidified solution of titanium salt gives yellow or orange colour with H_2O_2 .
 $\text{Ti}^{+4} + \text{H}_2\text{O}_2 + 2\text{H}_2\text{O} \longrightarrow \text{H}_2\text{TiO}_4$ (yellow/orange) + 4H^+

USES :

- In bleaching of delicate materials such as silk, wool, cotton, ivory etc.
- As a valuable antiseptic and germicide for washing wounds, teeth and ears under the name perhydrol.
- As 'antichlor' to remove traces of chlorine and hypochlorite.
- As oxidising agent in rocket fuels

4. SULPHUR (S) :

Sulphur Allotropic Forms :

Sulphur forms numerous allotropes of which the **yellow rhombic** (α - sulphur) and **monoclinic** (β - sulphur) forms are the most important. The stable form at room temperature is rhombic sulphur, which transforms to monoclinic sulphur when heated above 369 K.

Rhombic sulphur (α - sulphur) :

This allotrope is yellow in colour, m.p. 385.8 K and specific gravity 2.06. Rhombic sulphur crystals are formed on evaporating the solution of roll sulphur in CS_2 . It is insoluble in water but dissolved to some extent in benzene, alcohol and ether. It is readily soluble in CS_2 .

Monoclinic sulphur (β - sulphur) :

Its m.p. is 393 K and specific gravity 1.98. It is soluble in CS_2 . This form of sulphur is prepared by melting rhombic sulphur in a dish and cooling till crust is formed. Two holes are made in the crust and the remaining liquid poured out. On removing the crust, colourless needle shaped crystals of β - sulphur are formed. It is stable above 369 K and transforms into α - sulphur below it. Conversely, α - sulphur is stable below 369 K and transforms into β - sulphur above this. At 369 K both the forms are stable. This temperature is called transition temperature.

Both rhombic and monoclinic sulphur have S_8 molecules these S_8 molecules are packed to give different crystal structures. The S_8 ring in both the forms is puckered and has a crown shape. The molecular dimensions are given in figure.

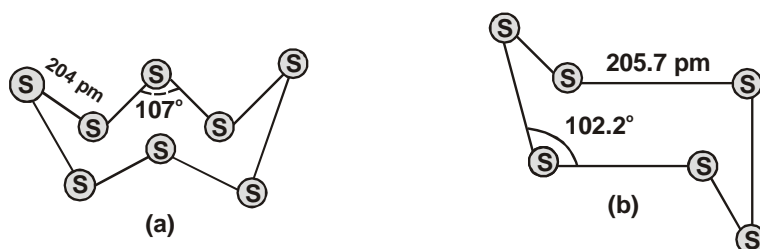


Fig. : The structures of (a) S_8 ring in rhombic sulphur and (b) S_6 form

Several other modifications of sulphur containing 6-20 sulphur atoms per ring have been synthesised in the last two decades. In cyclo- S_6 , the ring adopts the chair form and the molecular dimensions are as shown in fig. (b) At elevated temperatures (~ 1000 K), S_2 is the dominant species and is paramagnetic like O_2 .

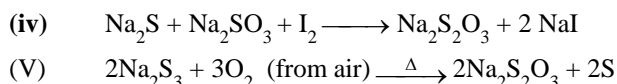
3. COMPOUNDS OF SULPHUR :

(A) SODIUM THIOSULPHATE ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$) :

PREPARATION :

- $\text{Na}_2\text{SO}_3 + \text{S} \xrightarrow[\text{in absence of air}]{\text{boiled}} \text{Na}_2\text{S}_2\text{O}_3$
- $\text{Na}_2\text{CO}_3 + 2\text{SO}_2$ (excess) + $\text{H}_2\text{O} \longrightarrow 2\text{NaHSO}_3 + \text{CO}_2$; $2\text{NaHSO}_3 + \text{Na}_2\text{CO}_3 \longrightarrow 2\text{Na}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} + \text{CO}_2$
- $2 \text{NaHS} + 4\text{NaHSO}_3 \longrightarrow 3\text{Na}_2\text{S}_2\text{O}_3 + 3\text{H}_2\text{O}$

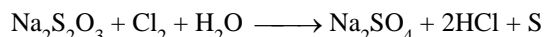
p-BLOCK ELEMENTS



PROPERTIES :

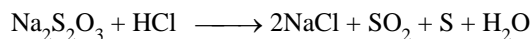
- (i) It is a colourless crystalline substance soluble in water which loses water of crystallisation on strong heating
(ii) **As antichlor :**

It removes the chlorine from the surface of fibres (while dyeing) according to following reaction.



Therefore, it is known as antichlor

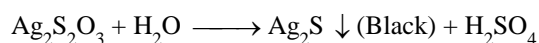
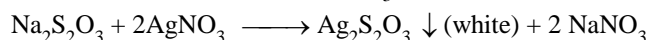
- (iii) **Reaction with HCl :**



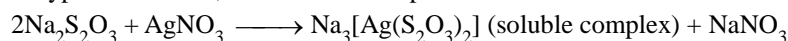
This test is used for distinction between $\text{S}_2\text{O}_3^{2-}$ and SO_3^{2-} ions as SO_3^{2-} ions give only SO_2 with HCl.

- (iv) **Complex formation reactions :**

- (a) **Reaction with silver salts (AgNO_3 , AgCl , AgBr or AgI) :**



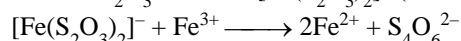
If hypo is in excess, then soluble complex is formed.



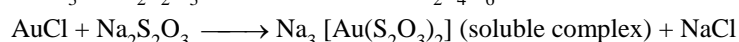
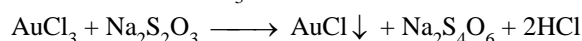
This reaction is utilized in photography where hypo is used as fixer.

- (b) **Reaction with FeCl_3 :**

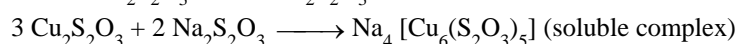
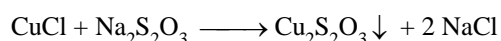
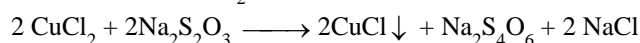
It develops a pink or violet colour which soon vanishes according to following reaction.



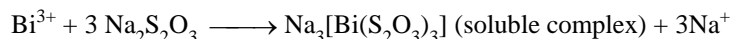
- (c) **Reaction with AuCl_3 (Soluble in water) :**



- (d) **Reaction with CuCl_2 :**

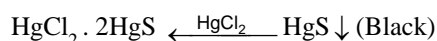
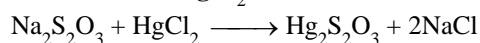


- (e) **Reaction with bismuth :**

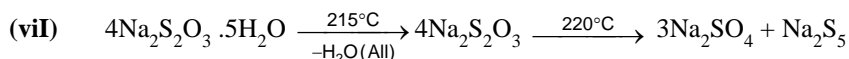
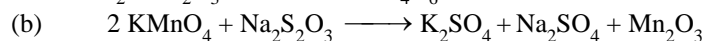
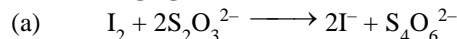


But it soon decomposes to give black ppt. of Bi_2S_3

- (v) **Reaction with HgCl_2 :**



- (vi) **As reducing agent In iodometric titration :**



USES :

- (i) As an 'antichlor' to remove excess of chlorine from bleached fabrics.
(ii) In photography as fixer.
(iii) As a reagent in iodometric and idiometric titrations.

(B) HYDROGEN SULPHIDE (H₂S) :**PREPARATION:**

- (i) $\text{FeS} + \text{H}_2\text{SO}_4 \longrightarrow \text{FeSO}_4 + \text{H}_2\text{S}$
It is prepared in kipp's apparatus
- (ii) Preparation of pure H₂S gas
 $\text{Sb}_2\text{S}_3 \text{ (pure)} + 6 \text{HCl (pure)} \longrightarrow 2 \text{SbCl}_3 + 3 \text{H}_2\text{S}$

PROPERTIES :

- (i) Colourless gas with rotten egg smell
- (ii) Moderately soluble in water but solubility decreases with increasing temperature.
- (iii) **Reducing Agent :**

Acts as a strong reducing agent as it decomposes evolving hydrogen.

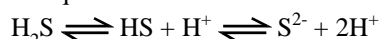
- (a) $\text{H}_2\text{S} + \text{X}_2 \longrightarrow 2 \text{HX} + \text{S}$;
- (b) $\text{H}_2\text{S} + \text{SO}_2 \xrightarrow{\text{moisture}} \text{H}_2\text{O} + \text{S}$;
- (c) $\text{H}_2\text{O}_2 + \text{H}_2\text{S} \longrightarrow \text{H}_2\text{O} + \text{S} + \text{O}_2$
- (d) $2\text{HNO}_3 \longrightarrow \text{H}_2\text{O} + 2\text{NO}_2 + [\text{O}]$
 $\text{H}_2\text{S} + [\text{O}] \longrightarrow \text{H}_2\text{O} + \text{S}$

 $2\text{HNO}_3 + \text{H}_2\text{S} \longrightarrow 2\text{H}_2\text{O} + 2\text{NO}_2 + \text{S}$

☞ It also reduces $\text{KMnO}_4 \rightarrow \text{Mn}^{2+}$, $\text{H}_2\text{SO}_4 \rightarrow \text{SO}_2$ & $\text{K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{Cr}^{3+}$

(iv) Acidic Nature :

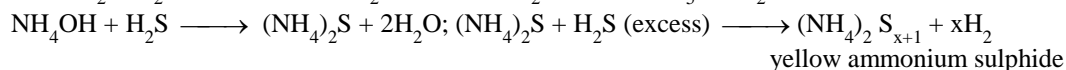
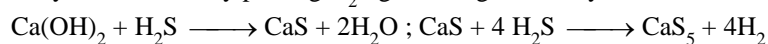
Its aqueous solution acts as a weak dibasic acid according to following reaction



Therefore, It forms two series of salts as given below

**(v) Formation of Polysulphides :**

They are obtained by passing H₂S gas through metal hydroxides.

**TESTS FOR H₂S :**

- (i) Turns acidified lead acetate paper black
- (ii) Gives violet or purple colouration with sodium nitropruside solution.

USES :

- (i) As a laboratory reagent for the detection of basic radicals in qualitative analysis.
- (ii) As reducing agent.

(C) SULPHUR DIOXIDE :**PREPARATION :**

- (i) $\text{S} + \text{O}_2 \text{ or air} \xrightarrow{\text{Burn}} \text{SO}_2$
- (ii) $\text{S} + 2\text{H}_2\text{SO}_4 \text{ (conc.)} \xrightarrow{\Delta} 3\text{SO}_2 + 2\text{H}_2\text{O}$
- (iii) By heating Cu or Ag with conc. H₂SO₄
 $\text{Cu} + \text{H}_2\text{SO}_4 \longrightarrow \text{CuSO}_4 + 2\text{H}_2\text{O} + \text{SO}_2$
- (iv) By reaction of metal sulphites with dil.HCl
 $\text{Na}_2\text{SO}_3 + 2\text{HCl} \longrightarrow 2\text{NaCl} + \text{SO}_2 + \text{H}_2\text{O}$
Similarly bisulphites also give SO₂ with dil. HCl
 $\text{NaHSO}_3 + \text{HCl} \longrightarrow \text{NaCl} + \text{SO}_2 + \text{H}_2\text{O}$

p-BLOCK ELEMENTS

(v) By heating sulphides in excess of air



(vi) CaSO_4 (gypsum) + C $\xrightarrow[1000^\circ\text{C}]{\Delta}$ 2 CaO + SO₂ + CO₂

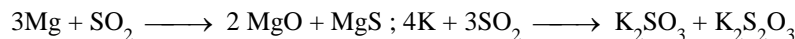
☞ By this method SO₂ is obtained in large scale

PROPERTIES :

(i) Colourless gas with burning sulphur smell.

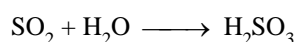
(ii) It is heavier than air and is highly soluble in water\

(iii) Neither burns nor helps in burning but burning magnesium and potassium continue to burn in its atmosphere.

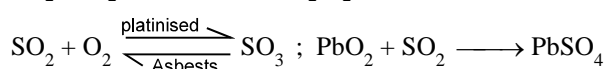
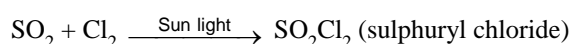


(iv) **Acidic Nature :**

Acidic oxide and thus dissolve in water forming sulphurous acid.



(v) **Addition Reaction :**

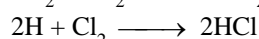
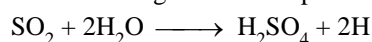


(vi) **Reducing Nature :**

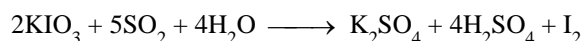
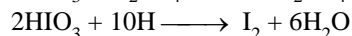
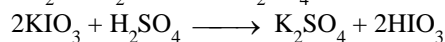
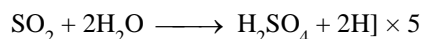


☞ Reducing character is due to the liberation of nascent hydrogen

(a) Reduces halogens to corresponding halides



(b) Reduces acidified iodates to iodine



☞ It also reduces acidified KMnO₄ \longrightarrow Mn²⁺ (decolourises),

Acidified K₂Cr₂O₇ \longrightarrow Cr³⁺ (green coloured solution) & Ferric Sulphate \longrightarrow Ferrous sulphate

(vii) **Oxidising nature :**

Acts as oxidising agent with strong reducing agent

(a) $2\text{H}_2\text{S} + \text{SO}_2 \xrightarrow{\text{moisture}} 2\text{H}_2\text{O} + 3\text{S}$

(b) $2\text{SnCl}_2 + \text{SO}_2 + 4\text{HCl} \longrightarrow 2\text{SnCl}_4 + 2\text{H}_2\text{O} + \text{S}$

(c) $2\text{Hg}_2\text{Cl}_2 + \text{SO}_2 + 4\text{HCl} \longrightarrow 2\text{HgCl}_2 + 2\text{H}_2\text{O} + \text{S}$

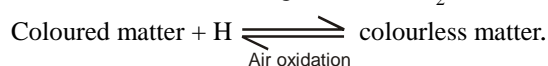
(d) $2\text{CO} + \text{SO}_2 \longrightarrow 2\text{CO}_2 + \text{S}$

(e) $2 \text{Fe} + \text{SO}_2 \longrightarrow 2\text{FeO} + \text{FeS}$

(viii) **Bleaching Action :**



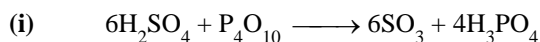
This is due to the reducing nature of SO₂



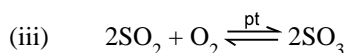
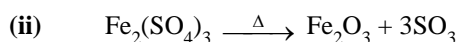
Therefore, bleaching is temporary

USES :

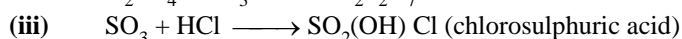
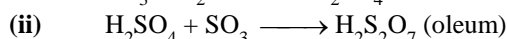
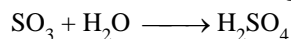
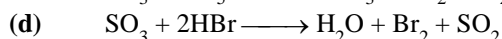
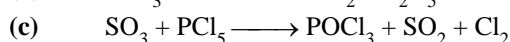
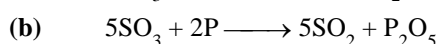
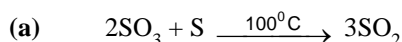
- (i) Used in manufacture of H_2SO_4 & paper from wood pulp.
- (ii) As a bleaching agent for delicate articles like wool, silk and straw.
- (iii) Used in refining of petroleum and sugar.

(D) SULPHUR TRIOXIDE (SO_3) :**PREPARATION:**

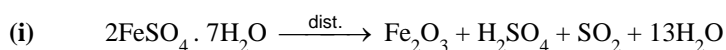
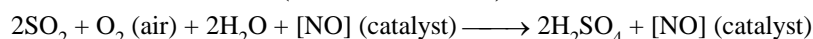
☞ P_4O_{10} is dehydrating agent

**PROPERTIES:****(i) Acidic Nature :**

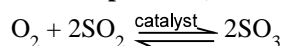
Dissolves in water forming sulphuric acid

**(iv) Oxidising Nature :****USES:**

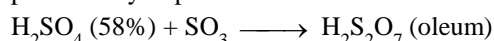
- (i) Used in manufacture of H_2SO_4 and oleum.
- (ii) Used as a drying agent for gases.

(E) SULPHURIC ACID (H_2SO_4) :**PREPARATION:****(ii) Lead Chamber Process (Industrial method) :**

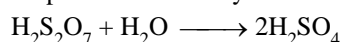
Acid obtained is 80% pure and is known as brown oil of vitriol

(iii) Contact process (Industrial method) :

The commonly used catalysts are platinum, ferric oxide or vanadium pentoxide. V_2O_5 is preferred as it is cheaper and is not poisoned by impurities



Sulphuric acid of any desired concentration can be obtained from oleum by dilution with water.



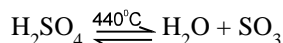
Oxidation of sulphur dioxide is reversible and exothermic. According to Le-chatelier principle, the favourable conditions for the greater yield of sulphur trioxide are.

- (a) An excess of air- SO_2 and oxygen are taken in 2:3 molecular proportion
- (b) Low tempt-optimum tempt. 450°C
- (c) Higher pressure-one atmosphere

PROPERTIES :

- (i) A colourless syrupy liquid (it is H-bonded)
 (ii) It fumes strongly in moist air and is highly corrosive in nature.

(iii) **Thermal decomposition :**



(iv) **Acidic Nature :**

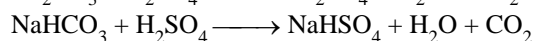
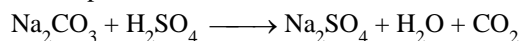
It is a strong dibasic acid and ionises as



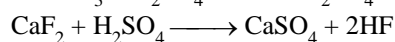
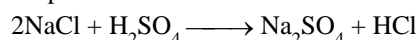
- (a) Forms two series of salts



- (b) Decomposes carbonates and bicarbonates into carbon dioxide

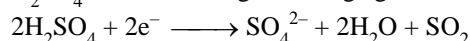


- (c) Displaces more volatile acids from their metal salts.

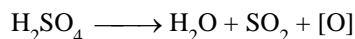


(v) **Oxidising Nature :**

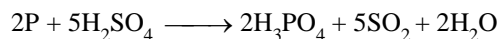
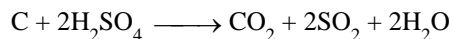
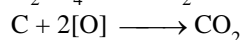
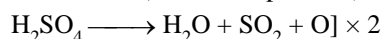
H_2SO_4 acts as a strong oxidising agent



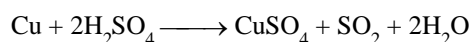
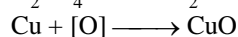
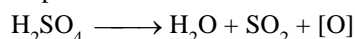
or



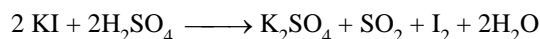
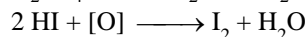
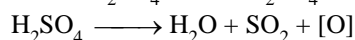
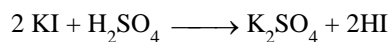
- (a) Non-metals (carbon, sulphur etc) are oxidised to their oxides.



- (b) Metals (copper, silver, mercury etc.) are oxidised to their oxides which then combine with acid to form corresponding sulphates.



- (c) Iodine is liberated from KI

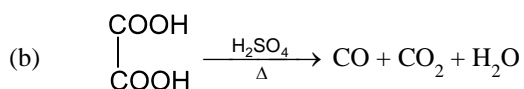
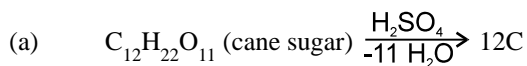
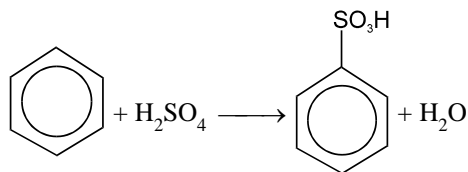


Here HI is oxidised to I_2 . Similarly bromine is liberated from KBr

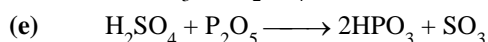
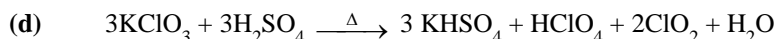
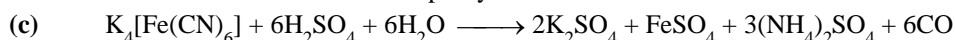
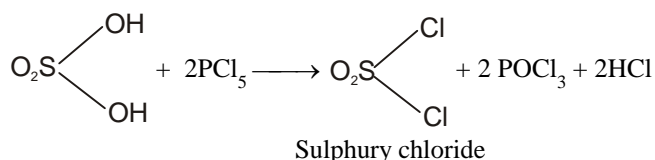
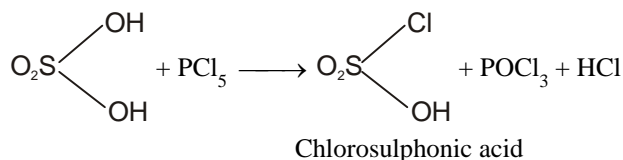
- (d) C_{10}H_8 (naphthalene) + $9\text{H}_2\text{SO}_4 \xrightarrow[\text{catalyst}]{\text{Hg as}}$ $\text{C}_8\text{H}_6\text{O}_4$ (phthalic acid) + $10\text{H}_2\text{O}$ + 9SO_2 + 2CO_2

(vi) Dehydrating agent :

Sulphuric acid acts as a powerful dehydrating agent because it has a great affinity for water

**(vii) Miscellaneous reactions :****(a) Sulphonation of aromatic compounds**

Benzene Benzene sulphonic acid

(b) Reaction with PCl₅ :**USES :**

- (i) For the manufacture of fertilizer such as ammonium sulphate and super phosphate of lime.
- (ii) As an important laboratory reagent.
- (iii) In storage batteries.
- (iv) In leather, textile, paper and dyeing industries.

GROUP 17 ELEMENTS : THE HALOGEN FAMILY

Fluorine, chlorine, bromine, iodine and astatine are members of Group 17. These are collectively known as the halogens (Greek halo means salt and genes born i.e., salt producers). The halogens are highly reactive non-metallic elements.

Electronic Configuration :

All these elements have seven electrons in their outermost shell ($ns^2 np^5$) which is one electron short of the next noble gas.

Atomic and Ionic Radii :

The halogens have the smallest atomic radii in their respective periods due to maximum effective nuclear charge. Atomic and ionic radii increase from fluorine to iodine due to increasing number of quantum shells.

Ionisation Enthalpy :

They have little tendency to lose electron. Thus they have very high ionisation enthalpy. Due to increase in atomic size, ionisation enthalpy decreases down the group.

Electron Gain Enthalpy :

Halogen have maximum negative electrons gain enthalpy in the corresponding period. This is due to the fact that the atoms of these elements have only one electron less than stable noble gas configurations. Electron gain enthalpy of the elements of the group becomes less negative down the group. However, the negative electron gain enthalpy of fluorine is less than that of chlorine. It is due to small size of fluorine atom. As a result, there are strong interelectronic repulsions in the relatively small 2p orbitals of fluorine and thus, the incoming electron does not experience much attraction.

p-BLOCK ELEMENTS

Electronegativity :

They have very high electronegativity. The electronegativity decreases down the group. Fluorine is the most electronegative element in the periodic table

Physical Properties :

Fluorine and chlorine are gases, bromine is a liquid and iodine is a solid. Their melting and boiling points steadily increase with atomic number. All halogens are coloured. This is due to absorption of radiations in visible region which results in the excitation of outer electrons to higher energy level. By absorbing different quanta of radiation, they display different colours. For example, F_2 , has yellow, Cl_2 , greenish yellow, Br_2 , red and I_2 , violet colour. Fluorine and chlorine react with water. Bromine and iodine are only sparingly soluble in water. But are soluble in organic solvents such as chloroform, carbon tetrachloride, carbon disulphide and hydrocarbons to give coloured solutions. Except the smaller enthalpy of dissociation of F_2 compared to that of Cl_2 whereas X-X bond dissociation enthalpies from chlorine onwards show the expected trend: $Cl-Cl > Br-Br > I-I$. A reason for this anomaly is the relatively larger electron-electron repulsion among the lone pairs in F_2 molecule where they are much closer to each other than in case of Cl_2 .

ATOMIC & PHYSICAL PROPERTIES :

Element	F	Cl	Br	I
Atomic Number	9	17	35	53
Atomic Mass	19	35.45	79.90	126.90
Electronic configuration	[He] $2s^2 2p^5$	[Ne] $3s^2 3p^5$	[Ar] $3d^{10} 4s^2 4p^5$	[Kr] $4d^{10} 5s^2 5p^5$
Covalent Radius / pm	64	99	114	133
Ionic Radius X^- / pm	133	184	196	220
Ionization enthalpy / ($kJ mol^{-1}$)	1680	1256	1142	1008
Electron gain enthalpy / ($kJ mol^{-1}$)	-333	-349	-325	-296
Distance X-X/pm	143	199	229	266
Enthalpy of dissociation (X_2)/ $kJ mol^{-1}$	158.8	242.6	192.8	151.1
Electronegativity	4	3.2	3.0	2.7
Melting point / K	54.4	172	265.8	386.6
Boiling point / K	84.9	239.0	332.5	458.2

Chemical Properties

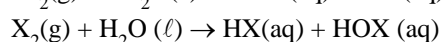
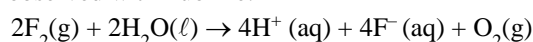
Oxidation states and trends in chemical reactivity

All the halogens exhibit -1 oxidation state. However, chlorine, bromine and iodine exhibit +1, +3, +5 and +7 oxidation states also. The higher oxidation states of chlorine, bromine and iodine are realised mainly when the halogens are in combination with the small and highly electronegative fluorine and oxygen atoms e.g., in interhalogens, oxides and oxoacids.

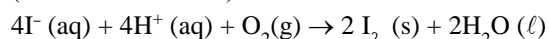
The fluorine atom has no d orbitals in its valence shell and therefore cannot expand its octet. Being the most electronegative, it exhibits only -1 oxidation state.

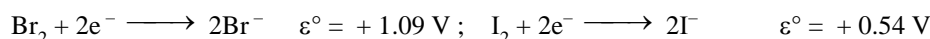
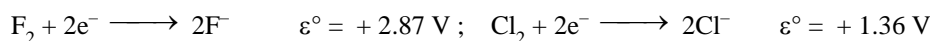
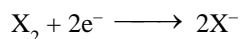
All the halogens are highly reactive. They react with metals and non-metals to form halides. The reactivity of the halogens decreases down the group.

The ready acceptance of an electron is the reason for the strong oxidising nature of halogens. F_2 is the strongest oxidising halogen and it oxidises other halide ions in solution or even in the solid phase. The decreasing oxidising ability of the halogen in aqueous solution down the group is evident from their standard electrode potentials. Fluorine oxidises water to oxygen whereas chlorine and bromine react with water to form corresponding hydrohalic and hypohalous acids. The reactions of iodine with water is non-spontaneous. I^- can be oxidised by oxygen in acidic medium; just the reverse of the reaction observed with fluorine.



(where X = Cl or Br)



Standard Reduction Potential (SRP)

☞ More the value of the SRP, more powerful is the (algebraically) oxidising agent. Hence the order of oxidising power is $F_2 > Cl_2 > Br_2 > I_2$

Since SRP is the highest for F_2 (among all elements of P.T.), it is a strongest oxidising agent.

☞ F_2 is more powerful oxidising agent than O_3 [In spite of 3 'O's in O_3]

Note : E.A. and I.E. values pertain to atoms in gas phase where as redox phenomena occurs in gaseous medium.

Hence properties in the gas phase cannot reflect parallelly in solution phase.

Electrode potential values would be the monitoring parameter in solution phase because they are experimental (based on the correct situation).

Hydration energy of X^-

Smaller the ion, higher is the hydration energy

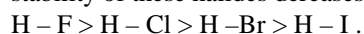
F^-	Cl^-	Br^-	I^-
515 kJ/mol	381	347	305

Anomalous behaviour of fluorine :

The anomalous behaviour of fluorine is due to its small size, highest electronegativity, low F-F bond dissociation enthalpy, and non availability of d orbitals in valence shell. Most of the reactions of fluorine are exothermic (due to the small and strong bond formed by it with other elements). It forms only one oxoacid while other halogen form a number of oxoacids. Hydrogen fluoride is liquid (b.p. 293 K) due to strong hydrogen bonding. Other hydrogen halides are gases.

(i) Reactivity towards hydrogen:

They all react with hydrogen to give hydrogen halides but affinity for hydrogen decreases from fluorine to iodine. They dissolve in water to form hydrohalic acids. The acidic strength of these acids varies in the order : $HF < HCl < HBr < HI$. The stability of these halides decreases down the group due to decrease in bond (H-X) dissociation enthalpy in the order :

**(ii) Reactivity towards oxygen :**

Halogens form many oxides with oxygen but most of them are unstable. Fluorine forms two oxides OF_2 and O_2F_2 . However, only OF_2 is thermally stable at 298 K. These oxides are essentially oxygen fluorides because of the higher electronegativity of fluorine than oxygen. Both are strong fluorinating agents. O_2F_2 oxidises plutonium to PuF_6 and the reaction is used in removing plutonium as PuF_6 from spent nuclear fuel.

Chlorine, bromine and iodine form oxides in which the oxidation states of these halogens range from +1 to +7. A combination of kinetic and thermodynamic factors lead to the generally decreasing order of stability of oxides formed by halogens, $I > Cl > Br$. The higher oxides of halogens tend to be more stable than the lower ones.

Chlorine oxides, Cl_2O , ClO_2 , Cl_2O_6 and Cl_2O_7 are highly reactive oxidising agents and tend to explode. ClO_2 is used as a bleaching agent for paper pulp and textiles and in water treatment.

The bromine oxides, Br_2O , BrO_2 , BrO_3 are the least stable halogen oxides and exist only at low temperature. They are very powerful oxidising agents.

The iodine oxides, I_2O_4 , I_2O_5 , I_2O_7 are insoluble solids and decompose on heating. I_2O_5 is a very good oxidising agent and is used in the estimation of carbon monoxide.

(iii) Reactivity towards metals:

Halogens react with metals to form metal halides. For e.g., bromine reacts with magnesium to give magnesium bromide.

(iv) Reactivity of halogen towards other halogens :

Halogens combine amongst themselves to form a number of compounds known as interhalogens of the types AB , AB_3 , AB_5 and AB_7 where A is a larger size halogen and B is a smaller size halogen.

1. FLUORINE (F₂) :

PREPARATION:

(i) Electrolytic method :

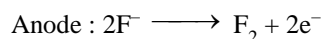
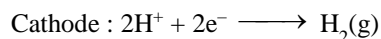
Electrolyte : Molten KHF₂ (1 part) + HF (5 part)

Anode : Carbon

Cathode : Steel

Vessel : Monel metal

On Electrolysis

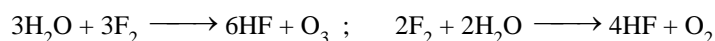


The F₂ gas thus evolved must be free from HF which is more corrosive than fluorine.

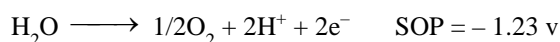
In order to make fluorine free from HF, the gas is passed through NaF which absorbs HF.

☞ Anode of carbon should be free from graphide because F₂ reacts with graphite easily to form a polymeric substance known as graphite fluoride.

There should be no moisture present in the vessel otherwise fluorine will react with water.



☞ It is not possible to prepare fluorine by electrolysis of aq. solution of NaF or KF. It is because when aq. solution of KF is subjected to electrolysis, there will be following two oxidation in competition at anode,

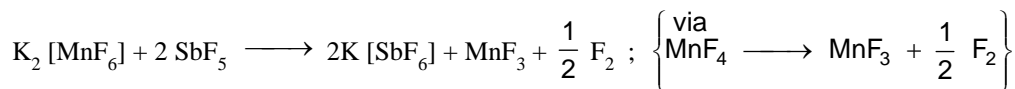


and



As a matter of rule that substance will be oxidise whose SOP is higher therefore water gets oxidise at anode and not F⁻.

(ii) Chemical method : From K₂ [MnF₆] - potassium hexafluoromanganate (IV)

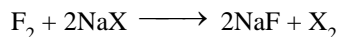


☞ In this reaction, the stronger Lewis acid SbF₅ displaces the weaker one, MnF₄ from its salt. MnF₄ is unstable and readily decomposes to give MnF₃ and fluorine.

PROPERTIES :

(i) Diatomic, Pale green-yellow gas which appears to be almost colourless. It is heavier than air. It condenses to yellow liquid at -188°C and yellow solute at -223°C. It has pungent odour and is highly poisonous.

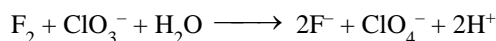
(ii) **Oxidising character :** It is the most powerful oxidising agent.



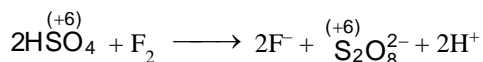
(X = Cl, Br, I)

(a) It can oxidise all other halide ions into halogen molecules

(b) It can oxidise ⁺⁵ClO₃⁻ into ⁺⁷ClO₄⁻ and ⁺⁵IO₃⁻ to ⁺⁷IO₄⁻

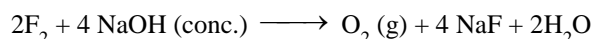
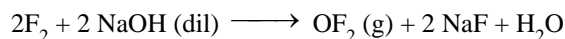


(c) It can oxidise HSO₄⁻ into S₂O₈²⁻

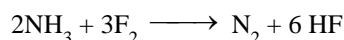


☞ Some of the O²⁻ changes to O⁻ in the persulphate ions (having the O⁻). Hence oxygen is getting oxidised.

(iii) **Reaction with NaOH solution :** With dilute alkali forms oxygen difluoride and with concentrated alkali O_2 .

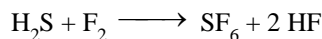


(iv) **Reaction with NH_3 :** (Distinction from other halogens)



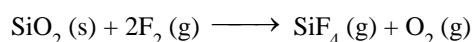
☞ Other halogens form explosive NX_3 , with conc. NH_3 (liquor ammonia)

(v) **Reaction with H_2S :**



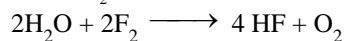
☞ H_2S burns

(vi) **Reaction with SiO_2 :** It attacks glass at about $100^{\circ}C$.

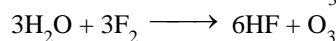


The reaction is slow with dry F_2 .

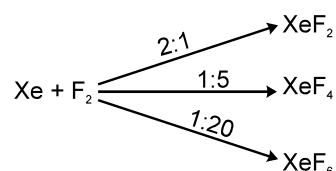
(vii) **Reaction with H_2O :**



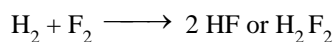
sometime a little amount of O_3 also forms



(viii) **Reaction with Xe :**

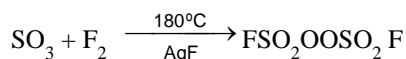


(ix) **Reaction with H_2 :**

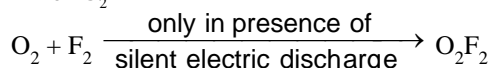


☞ This reactions occurs even in dark

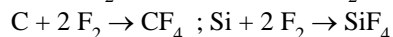
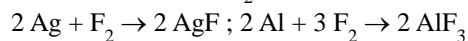
(x) **Reaction with SO_3 :**



(xi) **Reaction with O_2 :**



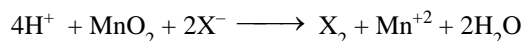
(xii) **Reaction with metals and non-metals :** It combines with most of the metals. Almost all non-metals except O_2 & N_2 ignite spontaneously in presence of F_2 .

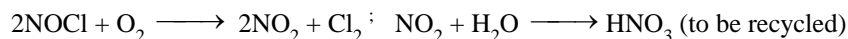
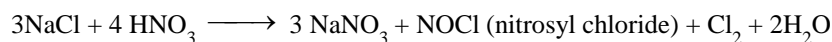
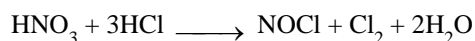
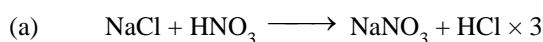


2. CHLORINE (Cl_2) :

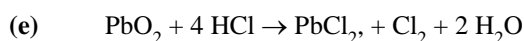
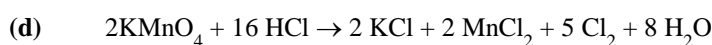
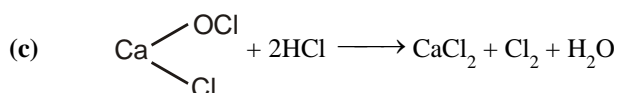
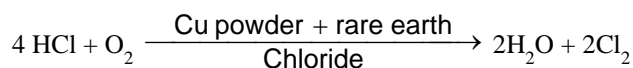
PREPARATION:

(i) **Common method (Cl_2 , Br_2 , I_2) :**

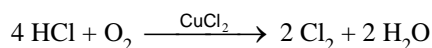


p-BLOCK ELEMENTS**(ii) Only for Cl₂ :**

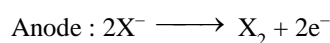
(b) When Cl₂ is used for the chlorination of hydrocarbon the byproduct is HCl. The HCl is catalytically oxidised into H₂O & Cl₂ using copper powder mixed with rare earth chlorides.

**(iii) Manufacture of chlorine :**

(a) **Deacon's process** : By oxidation of hydrogen chloride gas by atmospheric oxygen in the presence of CuCl₂ (catalyst) at 723 K.



(b) **Electrolytic process** : Chlorine is obtained by the electrolysis of brine (concentrated NaCl solution). Chlorine is liberated at anode. It is also obtained as a by-product in many chemical industries.

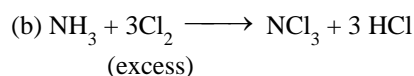
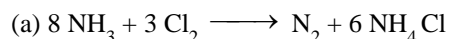
**PROPERTIES :**

(i) It is a greenish-yellow gas with pungent and suffocating odour. It is about 2–5 times heavier than air. It can be liquefied into greenish-yellow liquid which boils at 239 K. It is soluble in water.

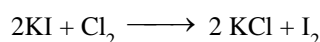
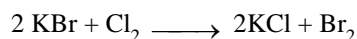
(ii) At low temperature it forms a hydrate with water having formula Cl₂ · 8H₂O which is in fact a clathrate compound.

(iii) $\text{H}_2 + \text{Cl}_2 \xrightarrow{\hspace{2cm}} 2\text{HCl (g)}$, [$\text{H}_2 + \text{Br}_2 \rightarrow 2\text{HBr}$ is not a zero order reaction]
(a zero order reaction)

(iv) **Reaction with NH₃ (common for Cl₂ & Br₂) :**



(v) **Reaction with alkali metal halides (KX) :**

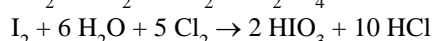
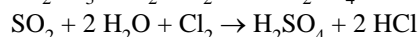
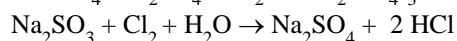
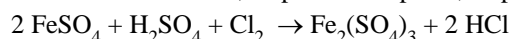


☞ Cl₂ can oxidise both Br⁻ & I⁻ but Br₂ can oxidise I⁻ only.

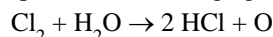
☞ F₂ is not used in aqueous reaction since it itself reacts with water.

(vi) **Oxidising & bleaching properties** : Chlorine water on standing loses its yellow colour due to the formation of HCl and HOCl. Hypochlorous acid (HOCl) so formed, gives nascent oxygen which is responsible for oxidising and bleaching properties of chlorine.

(i) It oxidises ferrous to ferric, sulphite to sulphate, sulphur dioxide to sulphuric acid and iodine to iodic acid.



(ii) It is a powerful bleaching agent ; Bleaching action is due to oxidation.

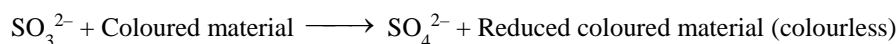
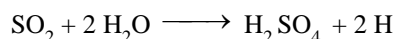


Coloured substance + O → Colourless substance

It bleaches vegetable or organic matter in the presence of moisture. Bleaching effect of chloride is permanent.



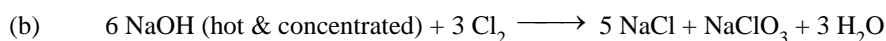
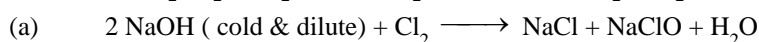
But the bleaching action of SO_2 is temporary because it takes place through reduction.



Reduced Coloured material (colourless) $\xrightarrow{\text{O}_2 \text{ of air}}$ Coloured material

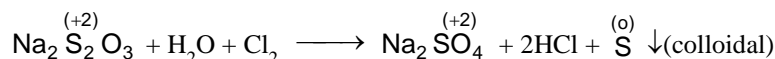
(vii) Reaction with NaOH :

Common to Br_2 , I_2 & Cl_2 (but with F_2 it is different, OF_2 or O_2 is obtained)



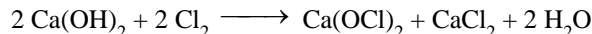
(viii) Reaction with Hypo solution :

This reaction is common with Cl_2 & Br_2 but with I_2 it is different.

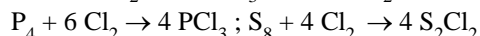
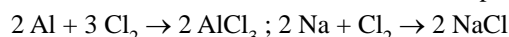


In this reaction thiosulphate ions undergo disproportionation into SO_4^{2-} and S. Cl_2 is reduced to Cl^- .

(ix) Reaction with dry slaked lime, $\text{Ca}(\text{OH})_2$: It gives bleaching powder.



(x) Reaction with metals & non-metals : Forms corresponding chlorides.



It has great affinity for hydrogen. It reacts with compounds containing hydrogen to form HCl.



Uses : Cl_2 is used

(i) for bleaching woodpulp (required for the manufacture of paper and rayon). bleaching cotton and textiles,

(ii) in the manufacture of dyes, drugs and organic compounds such as CCl_4 , CHCl_3 , DDT, refrigerants, etc.

(iii) in the extraction of gold and platinum.

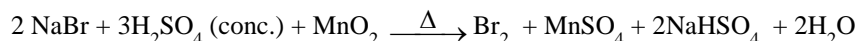
(iv) in sterilising drinking water and

(v) preparation of poisonous gases such as phosgene (COCl_2), tear gas (CCl_3NO_2), mustard gas ($\text{ClCH}_2\text{CH}_2\text{SCH}_2\text{CH}_2\text{Cl}$).

3. BROMINE (Br_2) :

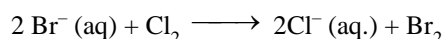
PREPARATION:

(i) Common method :



(ii) From Sea-water :

NaCl is main component but NaBr is also present in some quantity in sea water. Cl_2 gas is passed through sea water when vapours of bromine are evolved.



PROPERTIES :

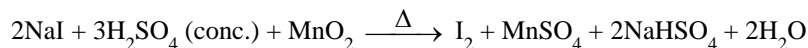
(i) Reddish brown liquid, fairly soluble in water. It also forms hydrate like Cl_2
 $(\text{Br}_2 \cdot 8\text{H}_2\text{O}) \longleftarrow$ Clathrate compound

(ii) Rest reactions are same as with Cl_2

4. IODINE (I_2) :

PREPARATION:

(i) **Common method :**



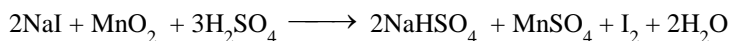
(ii) **From Caliche or Crude chile salt petre :**

The main source of iodine is NaIO_3 (sodium iodate) which is found in nature with NaNO_3 (chile saltpetre). NaIO_3 is present in small amount. After crystallisation of NaNO_3 , the mother liquor left contains NaIO_3 (soluble). To this solution NaHSO_3 is added where upon I_2 is precipitated.



(iii) **From sea-weeds :**

Sea-weeds are dried and burnt in shallow pits, ash left is called kelp. Ash on extraction with hot water dissolves out chlorides, carbonates, sulphates and iodides of sodium and potassium. The solution on concentration separates out all leaving behind iodide in the solution. Solution is mixed with MnO_2 and concentrated H_2SO_4 in iron retorts. Liberated iodine is condensed in series of earthen-ware known as aludels.



(iv) $2\text{KI} + \text{Cl}_2 \longrightarrow 2\text{KCl} + \text{I}_2$

(v) $2\text{KI} + \text{H}_2\text{O}_2 \longrightarrow 2\text{KOH} + \text{I}_2$

(vi) $\text{CuSO}_4 + 2\text{KI} \longrightarrow \text{K}_2\text{SO}_4 + \text{CuI}_2$; $2\text{CuI}_2 \longrightarrow \text{Cu}_2\text{I}_2 + \text{I}_2$

☞ This I_2 gets dissolved into KI forming KI_3 , since I_3^- ions are yellow, therefore solution develops yellow colour.

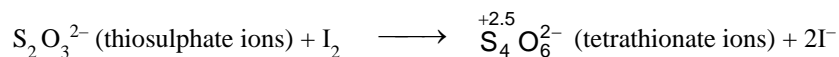
PROPERTIES :

(i) It is a dark violet solid, undergoes sublimation, least soluble (among halogens) in water but much more soluble in $\text{KI}(\text{aq.})$ due to formation of KI_3

☞ KF_3 cannot be formed similarly since F does not have 'd' orbitals. So sp^3d hybridisation is not possible with F.

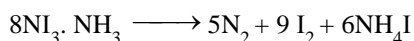
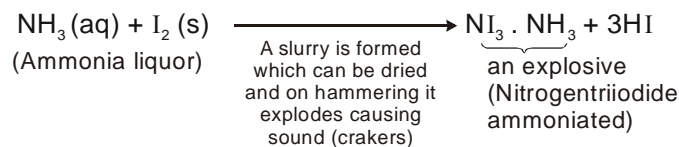
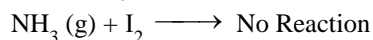
(ii) It is soluble in organic solvents like CHCl_3 , CCl_4 etc. to get violet solutions.

(iii) **Reaction with hypo :**
iodometric titrations

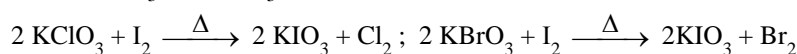


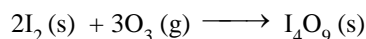
This reaction is the basis of iodometric (direct I_2 titration) titration, which is carried out for the estimation of iodine using starch indicator.

(iv) **Reaction with NH_3 :**



(v) **Reaction with KClO_3 or KBrO_3 :**

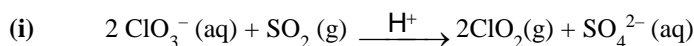


(vi) Reaction with ozone (dry) :

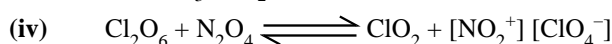
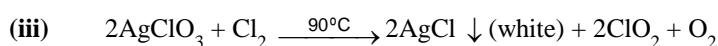
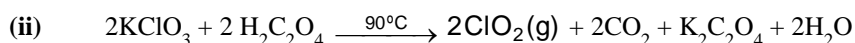
I_4O_9 is an ionic compound consisting of I^{3+} & $(\text{IO}_3^-)_3$ indicative of metallic character of I (low I.E, low E.N.)



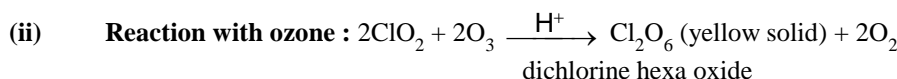
With NaOH there is common reaction as with Cl_2 and Br_2 . With H_2 there is reversible reaction.

OXIDES OF CHLORINE :**Chlorine dioxide (ClO_2) :****PREPARATION:**

Chlorates of sodium and potassium can be used

**PROPERTIES :**

(i) Yellow gas at room temp, soluble in water and explosive substance. It also behaves as an oxidising agent. (It kill bacteria better than Cl_2)



In the reaction O_3 is behaving as an oxidising agent.



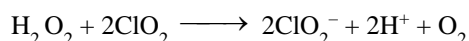
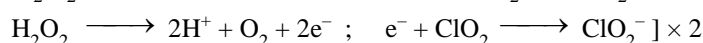
$\text{Cl}_2\text{O}_6(\text{s})$ is a mixed anhydride of HClO_3 & HClO_4 because on dissolving in water it gives a mixture of these two acids.



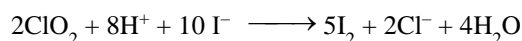
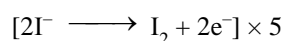
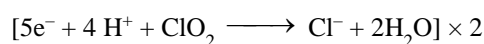
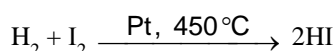
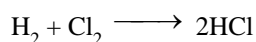
In solid state, Cl_2O_6 exists as ClO_2^+ & ClO_4^- .

(iii) Reaction with alkaline H_2O_2 :

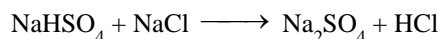
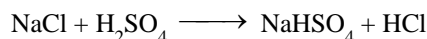
In this reaction H_2O_2 acts as a reducing agent. It reduces ClO_2 into ClO_2^-

**(iv) Reaction with HI :**

In this reaction HI behaves as a reducing agent where it reduces ClO_2 into Cl^- and itself is oxidised to I_2 .

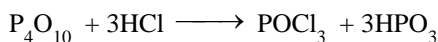
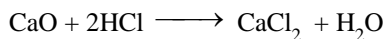
**HYDRA ACIDS (HALOGEN ACIDS) :****HCl, HBr & HI :****PREPARATION:****(i) By direct combination of elements :**

(ii) By heating a halide with acid :



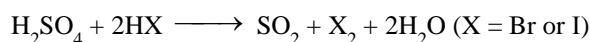
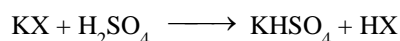
For HCl we use H_2SO_4 as an acid, while for HBr and HI we use H_3PO_4 .

- (a) HCl cannot be dried over P_2O_5 (P_4O_{10}) or quick lime since they react with gas chemically.

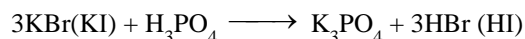


HCl is, hence dried by passing through conc. H_2SO_4 .

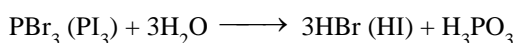
- (b) HBr (or HI) cannot be prepared by heating bromide (iodide) with conc. H_2SO_4 because HBr and HI are strong reducing agents and reduce H_2SO_4 to SO_2 and get themselves oxidised to bromine and iodine respectively.



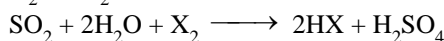
Hence, HBr and HI are prepared by heating bromides and iodides respectively with conc. H_3PO_4 .



(iii) By reaction of P_4 (Laboratory Method) :



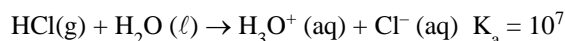
(iv) By passing $\text{H}_2\text{S}/\text{SO}_2$ into solutions of halogens :



PROPERTIES :

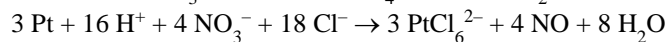
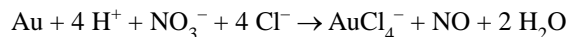
- (i) These are colourless, pungent smelling gases with acidic tastes.
- (ii) It is heavier than air, can be liquified to colourless liquids.
- (iii) These are neither combustible nor supporter of combustion.
- (iv) When perfectly dry, they have no action on litmus, but in presence of moisture, they turn blue litmus red, showing acidic nature. Among HX, HI is the strongest and HF is the weakest acid.
- (v) These are quite soluble in water.

HCl ionises as below.



Its aqueous solution is called hydrochloric acid. High value of dissociation constant (K_a) indicates that it is a strong acid in water.

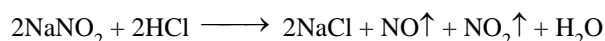
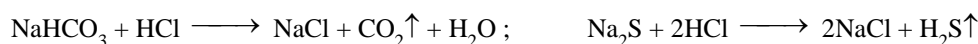
When three parts of concentrated HCl and one part of concentrated HNO_3 are mixed, aqua regia is formed which is used for dissolving noble metals, e.g., gold, platinum.



- (v) Reaction with metals oxides, hydroxides and bicarbonates

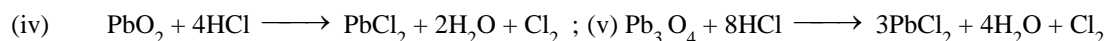
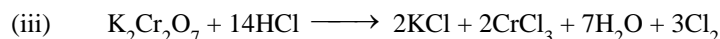
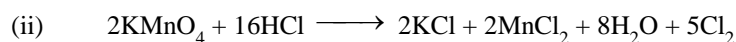
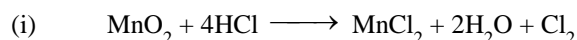


- (vi) Reaction with salts, HCl decomposes salts of weaker acids.



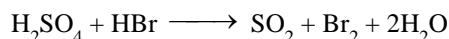
(vii) Reducing property and stability of hydrazids :

HCl : It is quite stable and hence is oxidised by strong oxidising agents like MnO_2 , KMnO_4 , $\text{K}_2\text{Cr}_2\text{O}_7$, PbO_2 , Pb_3O_4 .

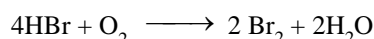


Therefore, HCl is a weak reducing agent.

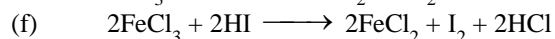
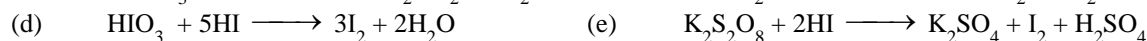
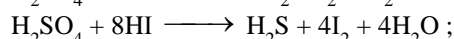
HBr : It is not very stable and hence more easily oxidised or acts as a strong reducing agents. In addition to above reducing properties of HCl, it also reduces H_2SO_4 to SO_2 which is not done by HCl.



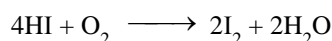
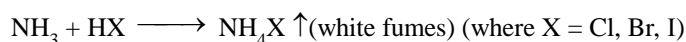
Aqueous HBr on exposure to atmospheric oxygen is oxidised to bromine (yellow)



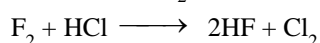
HI : It is least stable hydrogen halide. It is readily oxidised and thus acts as a powerful reducing agent. In addition to reaction shown by HCl, it shows following reactions also.



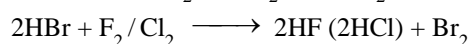
☞ Aqueous solution of acid, if exposed to O_2 is oxidised to iodine.

**(viii) Reaction with ammonia :****(ix) Action of halogens :**

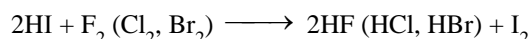
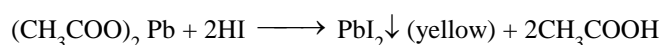
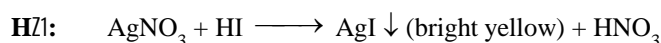
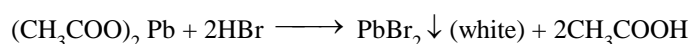
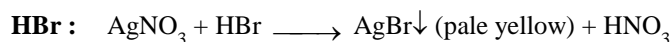
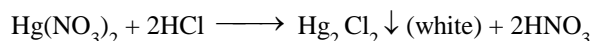
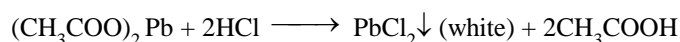
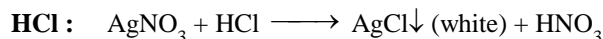
(a) Cl_2 is liberated from HCl by F_2 alone.



(b) Br_2 is liberated from HBr by F_2 and Cl_2 (not by I_2)

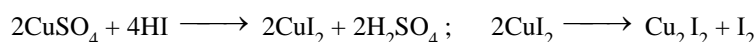


(c) I_2 is liberated from HI by F_2 , Cl_2 and Br_2

**(x) Detection of cation:**

p-BLOCK ELEMENTS

HI reacts with CuSO_4 liberating iodine via the formation of cupric iodide (not by HCl or HBr).



(xi) **Formation of aqua-regia :**

3 parts of conc. HCl and 1 part of conc. HNO_3 is known as aqua-regia. This is used for dissolving noble metals like Au (Gold) and Pt (Platinum).



USES :

- (i) HCl is used in preparation of Cl_2 , chlorides, aqua regia, glucose, (from corn starch), medicines, laboratory reagents, cleaning metal surfaces before soldering or electroplating. It is also used for extracting glue from bones and purifying bone black.
- (ii) HBr is used as laboratory reagent for preparing bromo derivatives like sodium bromides and potassium bromide. These are used in medicines as sedatives.
- (iii) HI is used as reducing agent in organic chemistry.

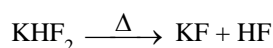
HYDROFLUORIC ACID [H_2F_2 , HF] :

PREPARATION :

H_2 and F_2 combine with each other very violently (even in dark) to form HF. So simple reaction cannot be used for its preparation, special methods are employed for its preparation.

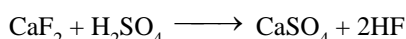
(i) **Laboratory Method :**

Anhydrous HF is obtained by heating dry potassium hydrogen fluoride in a copper retort connected with copper condenser.



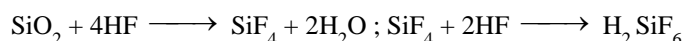
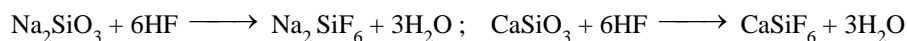
(ii) **Industrial Method :**

HF is prepared by heating fluorspar (CaF_2) with conc H_2SO_4 .



☞ Aqueous HF being corrosive to glass, is stored in wax lined bottles or vessel made of copper or monel.

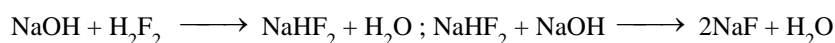
In glass or silica bottles, it attacks them as follows:



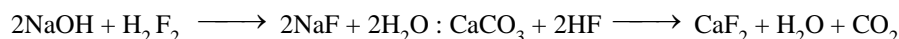
This action of HF on silica (silicates) is used for etching glass. The glass surface to be etched is coated with wax, the design, is scratched on glass through wax coating this is then treated with 40% solution.

PROPERTIES:

- (i) It is colourless, corrosive liquid with pungent smell with high boiling point due to hydrogen bonding.
- (ii) Dry HF does not attack metals under ordinary conditions (except K), but in presence of water, it dissolves metals with liberation of hydrogen gas.
- (iii) It is a weak dibasic acid (due to strong HF bond) and forms two series of salt.



- (iv) Concentrated acid reacts with oxides, hydroxides and carbonates.



OXY-ACIDS OF CHLORINE :

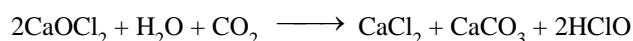
HYPO-CHLOROUS ACID [HClO] :

PREPARATION:

- (i) The acid is known only in solution, It is obtained by shaking precipitate of HgO with chlorine water.

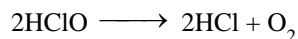


- (ii) Commercially, it is obtained by passing CO_2 through suspension of bleaching powder and then distilling.

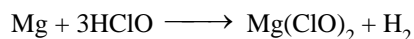


PROPERTIES:

- (i) It is a weak acid. Its concentrated solution is yellow in colour while dilute solution is colourless. It is unstable and decomposes.

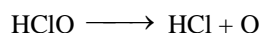


- (ii) It dissolves magnesium with evolution of hydrogen.



- (iii) With alkalies, it forms salts called hypochlorites.

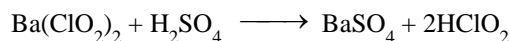
- (iv) It acts as a powerful oxidising and bleaching agent. This is due to release of nascent oxygen easily.



CHLOROUS ACID [HClO_2] :

PREPARATION:

It is obtained in aqueous solution when barium chlorite suspension in water is treated with H_2SO_4 . The insoluble barium sulphate is filtered off.

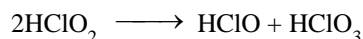


PROPERTIES:

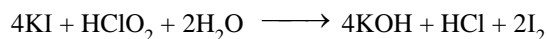
- (i) The freshly prepared solution is colourless but it soon decomposes to ClO_2 which makes the solution yellow.



- (ii) The acid undergoes auto-oxidation.



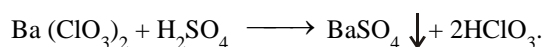
- (iii) The acid liberates iodine from KI.



CHLORIC ACID [HClO_3] :

PREPARATION:

This acid is only known in solution. The acid is prepared by the action of the dilute H_2SO_4 on barium chlorate.



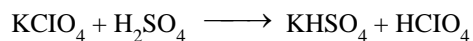
PROPERTIES:

Concentrated acid is colourless and pungent smelling liquid. It decomposes in light. However, it is stable in dark. It acts as a strong oxidising and bleaching agent in light. Organic substances like paper, cotton, wool, etc., catch fire in contact with the acid.

Per-Chloric Acid [HClO_4] :

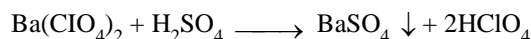
PREPARATION:

It is the most stable oxy-acid of chlorine. Anhydrous HClO_4 is obtained by doing distillation of KClO_4 (potassium perchlorate), with 96-97.5% H_2SO_4 under low pressure at 90-160°C.



p-BLOCK ELEMENTS

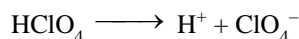
An aqueous solution of the acid is obtained by reacting barium perchlorate with calculated quantity of dilute H_2SO_4 . The insoluble barium sulphate is removed by filtration.



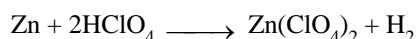
PROPERTIES:

(i) Anhydrous HClO_4 is a colourless liquid which turns dark on keeping. It fumes in moist air.

(ii) It is one of the strongest acid and ionises as follows :



(iii) It dissolves most of the metals.



(iv) It is an oxidising agent and explodes in contact with organic matter.

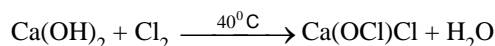
(v) $2\text{HClO}_4 + \text{P}_2\text{O}_5 \longrightarrow 2\text{HPO}_3 + \text{Cl}_2\text{O}_7$

BLEACHING POWDER ($\text{CaOCl}_2 \cdot \text{H}_2\text{O}$)

Bleaching powder is also called calcium chlorohypochlorite because it is considered as a mixed salt of hydrochloric acid and hypochlorous acid.



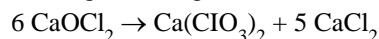
PREPARATION:



Properties

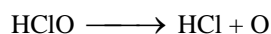
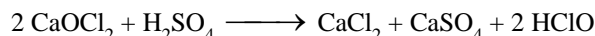
(i) It is a pale yellow powder. It has a strong smell of chlorine. It is soluble in water but a clear solution is never formed due to the presence of impurities.

(ii) On long standing, it undergoes auto-oxidation into calcium chlorate and calcium chloride.



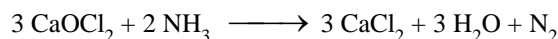
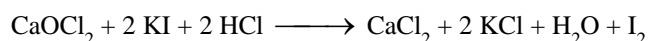
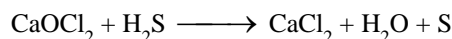
(iii) $2 \text{CaOCl}_2 \xrightarrow{\text{CoCl}_2} 2 \text{CaCl}_2 + \text{O}_2$

(iv) In presence of a little amount of a dilute acid, it loses oxygen.



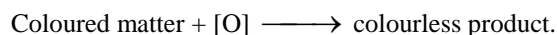
On account of the liberation of nascent oxygen, it shows oxidising and bleaching properties.

(a) Oxidising properties

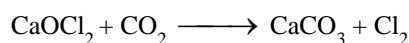
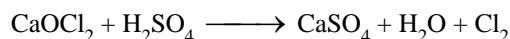
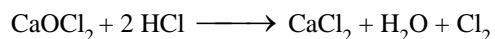


☞ It oxidises NO_2^- to NO_3^- , AsO_3^{3-} to AsO_4^{3-} and Fe^{2+} to Fe^{3+} (in acidic medium)

(b) Bleaching action



(v) When bleaching powder reacts with dilute acids or CO_2 it liberates chlorine which is known as available chlorine.

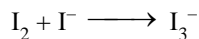


☞ HNO_3 is a strong oxidising acid to be avoided, here.

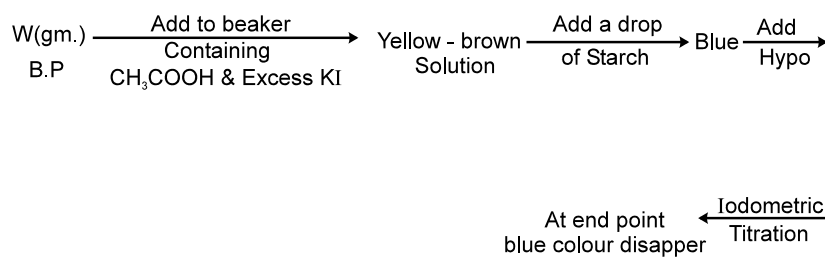
ESTIMATION OF AVAILABLE CHLORINE :

Let the wt. of sample of B.P. = W gm.

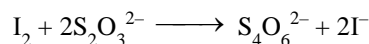
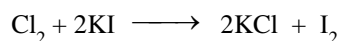
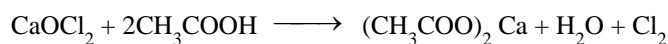
Add into a beaker containing acetic acid solution and excess KI. A yellow brown solution is formed (I_3^-)



Now few drops of starch solution is added into it. An intensive blue color is observed. Now hypo is used as the titrant. Note the volume where the blue colour disappear.



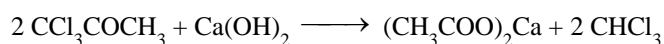
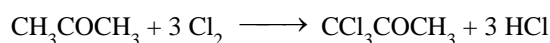
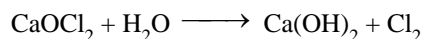
∅ Reaction involved



Calculation :

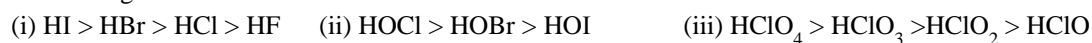
$$\% \text{Cl} = \frac{[\text{M}_{\text{hypo}} \times \text{V}_{\text{hypo}}] \times \frac{1}{2} \times 71}{\text{W}} \times 100$$

(vi) Bleaching powder converts acetone or ethyl alcohol into CHCl_3

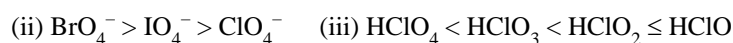
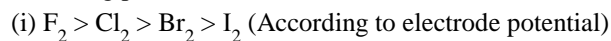


∅ Some important order

(a) Acid strength :



(b) Oxidising power :



(c) Order of disproportionation :



PSEUDO HALOGENS & PSEUDO HALIDES:

Some inorganic compounds like $(\text{CN})_2$ cyanogen, $(\text{SCN})_2$ thiocyanogen, $(\text{SeCN})_2$ selenocyanogen, $(\text{SCSN}_3)_2$ azidocarbonyl disulphide have been found to behave like halogens & they are known as pseudo halogen solids. Similarly few ions are known, consisting of two or more atoms of which at least one is N, that have properties similar to those of the halide ions. They are therefore called pseudohalide ions, e.g. (CN^-) cyanide ion, (SCN^-) thiocyanate ion, (SeCN^-) selenocyanate ion, (OCN^-) cyanate ion, (NCN^{2-}) cyanamide ion, (N_3^-) azide ion etc.

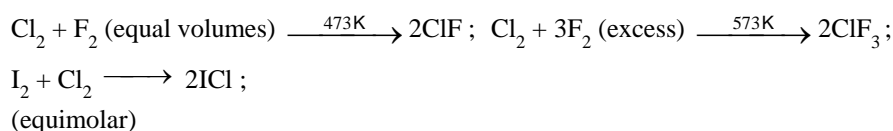
INTERHALOGEN COMPOUNDS :

We know that halogen atoms have different electronegativity. Due to this difference in electronegativity the halogen atoms combine with each other and give rise to the formation of binary covalent compounds, which are called interhalogen compounds. These are of four types.

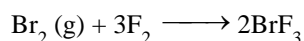
AB	AB ₃	AB ₅	AB ₇
ClF	ClF ₃	ClF ₅	IF ₇
BrF	BrF ₃	BrF ₅	
ICl	ICl ₃	IF ₅	
IF	IF ₃		

PREPARATION:

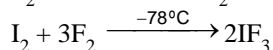
(i) By the direct combination of halogens :



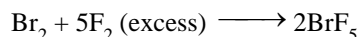
☞ Diluted with water.



☞ F₂ is diluted with N₂

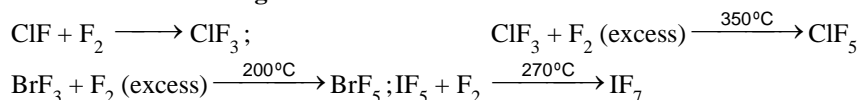


☞ F₂ is taken in freon.



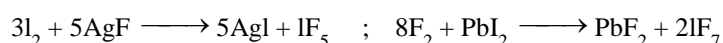
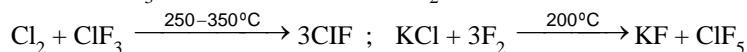
☞ IF₇ can not be prepared by direct combination of I₂ & F₂.

(ii) **From lower interhalogens :**



☞ This method is generally applied for the preparation of halogen fluorides.

(iii) **Other methods :**



PROPERTIES:

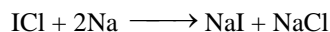
(i) These compounds may be gases, liquids or solids.

Gases : ClF, BrF, ClF₃, IF₇; Liquids : BrF₃, BrF₅; Solids : ICl, IBr, IF₃, ICl₃.

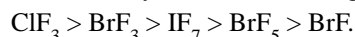
(ii) Interhalogens containing fluorine are colourless but inter halogens consisting of heavier halogens are coloured. The intensity of colour increases with increase in the molecular weight of the compounds.

- (iii) All interhalogens are covalent molecules and are diamagnetic in nature since all the valence electrons present as bonding or non-bonding electrons are paired.
- (iv) The boiling points increases with the increase in the electronegativity difference between A and B atoms.
- (v) Thermal stability of AB type interhalogen compounds decreases with the decrease in electronegativity difference between A and B atoms. $IF > BrF > ClF > ICl > IBr > BrCl$.
- ☞ More polar is the A – B bond more is the stability of interhalogen.

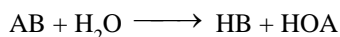
- (vi) Interhalogen compounds are more reactive than the parent halogens but less reactive than F_2 .



The order of reactivity of some interhalogens is as follows :



- (vii) Hydrolysis : All these undergo hydrolysis giving halide ion derived from the smaller halogen and a hypohalite (when AB), halite (when AB_3), halate (when AB_5), and perhalate (when AB_7) anion derived from the larger halogen.



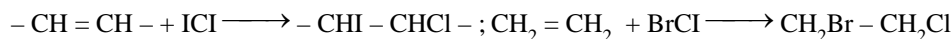
☞ Hydrolysis products are halogen acid and oxy-halogen acid. Oxy-halogen acid is of larger central halogen atom.

☞ Oxidation state of A atom does not change during hydrolysis.

- (viii) Reaction with non-metallic and metallic oxides :



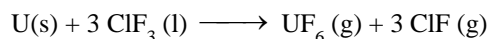
- (ix) Addition reaction (of AB type compounds) :



- (x) Reaction with alkali metal halides :



Uses : These compounds can be used as non aqueous solvents. Interhalogen compounds are very useful fluorinating agents. ClF_3 and BrF_3 are used for the production of UF_6 in the enrichment of ^{235}U .

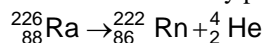


(F)GROUP 18 ELEMENTS : (THE ZERO GROUP FAMILY)

Group 18 Consists of six elements: helium, neon, argon, krypton, xenon and radon. All these are gases and chemically unreactive. They form very few compounds. Because of this they are termed noble gases.

Occurrence :

All the noble gases except radon occur in the atmosphere. Their atmospheric abundance in dry air is ~ 1% by volume of which argon is the major constituent. Helium and sometimes neon are found in minerals of radioactive origin e.g., pitchblende, monazite, cleveite. The main commercial source of helium is natural gas. Xenon and radon are the rarest elements of the group. Radon is obtained as a decay product of ^{226}Ra .



☞ Most abundant element in air is Ar. Order of abundance in the air is $Ar > Ne > Kr > He > Xe$.

Electronic Configuration :

All noble gases have general electronic configuration ns^2np^6 except helium which has $1s^2$. Many of the properties of noble gases including their inactive nature are ascribed to their closed shell structures.

p-BLOCK ELEMENTS

Ionisation Enthalpy :

Due to stable electronic configuration these gases exhibit very high ionisation enthalpy. However, it decreases down the group with increases in atomic size.

Atomic Radii :

Atomic radii increase down the group with increase in atomic number.

Electron Gain Enthalpy :

Since noble gases have stable electronic configurations, they have no tendency to accept the electron and therefore, have larger positive values of electron gain enthalpy.

Physical properties :

All the noble gases are monoatomic. They are colourless, and tasteless. They are sparingly soluble in water. They have very low melting and boiling points because the only type of interatomic interaction in these elements is weak dispersion forces. Helium has the lowest boiling point (4.2K) of any known substance. It has a unusual property of diffusing through most commonly used laboratory materials such as rubber, glass or plastics.

ATOMIC & PHYSICAL PROPERTIES :

Element	He	Ne	Ar	Kr	Xe
Atomic Number	2	10	18	36	54
Atomic Mass	4	20.18	39.10	83.80	131.30
Electronic configuration	$1s^2$	$[\text{He}] 2s^2 2p^6$	$[\text{Ne}] 3s^2 3p^6$	$[\text{Ar}] 3d^{10} 4s^2 4p^6$	$[\text{Kr}] 4d^{10} 5s^2 5p^6$
Atomic Radius (pm)	120	160	190	200	220
Ionization enthalpy / (kJ mol^{-1})	2372	2080	1520	1351	1170
Density (at STP)/ g cm^{-3}	1.8×10^{-4}	9.0×10^{-4}	1.8×10^{-3}	3.7×10^{-3}	5.9×10^{-3}
Melting point / K	–	24.6	83.8	115.9	161.3
Boiling point / K	4.2	27.1	87.2	119.7	165.0

Chemical Properties

In general, noble gases are least reactive. Their inertness to chemical reactivity is attributed to the following reasons:

- The noble gases except helium ($1s^2$) have completely filled $ns^2 np^6$ electronic configuration in their valence shell.
- They have high ionisation enthalpy and more positive electron gain enthalpy.

The reactivity of noble gases has been investigated occasionally ever since their discovery, but all attempt to force them to react to form the compounds were unsuccessful for quite a few years. In March 1962, Neil Bartlett, then at the University of British Columbia, observed the reaction of a noble gas. First, he prepared a red compound which is formulated as $\text{O}_2^+ \text{PtF}_6^-$. He, then realised that the first ionisation enthalpy of molecular oxygen (1175 kJ mol^{-1}) was almost identical with that xenon (1170 kJ mol^{-1}). He made efforts to prepare same type of compound with $\text{Xe}^+ \text{PtF}_6^-$ by mixing PtF_6 and Xenon. After this discovery, a number of xenon compounds mainly with most electronegative elements like fluorine and oxygen, have been synthesised.

The compounds of krypton are fewer. Only the difluoride (KrF_2) has been studied in detail. Compounds of radon have not been isolated but only identified (e.g., RnF_2) by radiotracer technique. No true compounds of Ar, Ne or He are yet known.

☞ If Helium is compressed and liquified it forms He(I) liquid at 4.2 K. This liquid is a normal liquid like any other liquid. But if it is further cooled then He(II) is obtained at 2.2 K, which is known as super fluid, because it is a liquid with properties of gases. It climbs through the walls of the container & comes out. It has very high thermal conductivity & very low viscosity. Compounds of inert gases are of following two types.

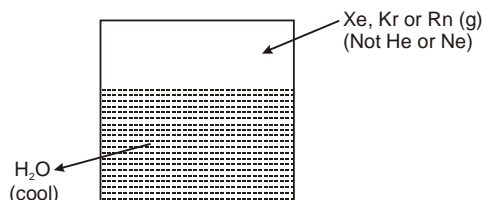
- Physical compounds (possess no proper bonding)
Physical compounds may be (A) compounds whose existence is on the basis of spectroscopic studies (temporary phase not isolated) and (B) clathrate compounds.

Clathrate compounds :

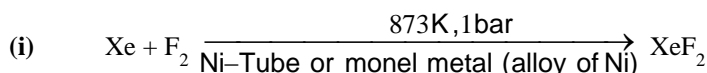
Inert gas molecules get trapped in the cages formed by the crystal structure of water.

During the formation of ice Xe atoms will be trapped in the cavities (or cages) formed by the water molecules in the crystal structure of ice. Compounds thus obtained are called clathrate compounds.

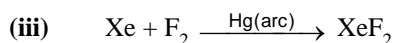
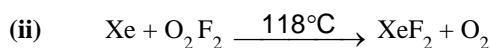
There are no chemical bonds. They do not possess an exact chemical formula but approx it is 6 water molecules : 1 inert gas molecule. The cavity size is just smaller than the atom of the noble gas. Such compounds are also formed by the other organic liquids like dihydroxybenzene (for example quinol).



- (ii) True chemical compounds (posses proper bonding).

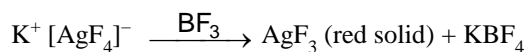
COMPOUNDS OF XENON :**(a) XENON DIFLUORIDE (XeF₂):****PREPARATION:**

☞ Volume ratio should be 2:1 otherwise other higher fluorides tend to form.

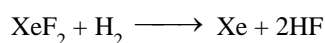


- (iv) Recently discovered method :

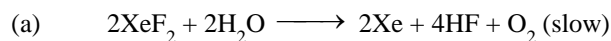
$\text{K}^+ [\text{AgF}_4]^-$ [potassium tetrafluoroargentate (III)] is first prepared and this is reacted with BF_3 .

**PROPERTIES:**

- (i) Colorless crystalline solid and sublimates at 298 K.
 (ii) Dissolves in water to give a solution with a pungent odour. Much soluble in HF liquid.
 (iii) This is stored in a vessel made up of monel metal which is a alloy of nickel.
 (iv) **Reaction with H₂ :** It reacts with hydrogen gas at 400°C

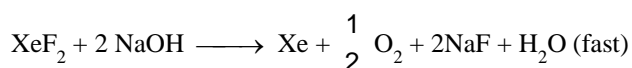


- (v) **Hydrolysis :**



The above is neither a cationic hydrolysis nor an anionic hydrolysis as seen in ionic equilibrium. It is a covalent compound and hydrolysis is like that of PCl_5 .

- (b) Hydrolysis is more rapid with alkali.

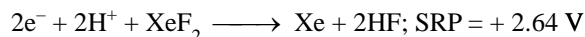


The reaction (a) is slower probably due to dissolution of XeF_2 in HF.

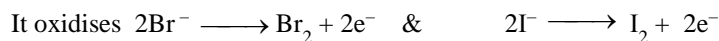
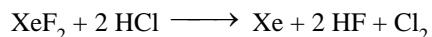
(vi) Oxidising properties :

Higher the value of SRP better is the oxidising property of the species.

The standard reduction potential for XeF_2 is measured to be + 2.64 V. Therefore it acts as a strong oxidising agent.



This oxidises halides to their respective halogens.

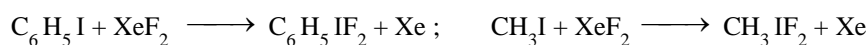


In this manner XeF_2 will oxidise halide ions (except F^-) into free halogens.

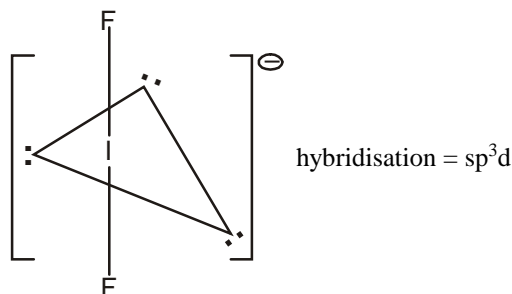
Similarly it can oxidise BrO_3^- (bromate) which are themselves good oxidising agents to BrO_4^- (perbromate ions) and Ce^{+3} to Ce^{+4} ion.

(vii) Oxidising as well as fluorinating properties :

It can act as strong oxidising agent as well as fluorinating agent.

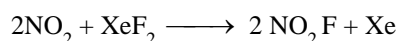
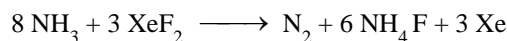
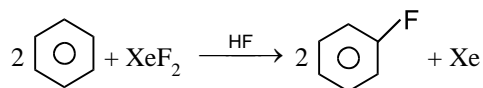
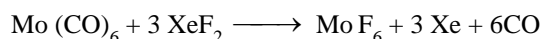


CH_3IF_2 exists as $\text{CH}_3^+ \text{IF}_2^-$, IF_2^- is analogous to I_3^-



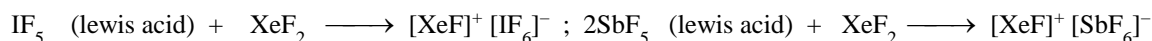
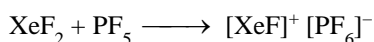
F_3^- can not be formed as it has no d-orbitals to attain sp^3d hybridisation.

(viii) Reactions of XeF_2 + HF (anhydrous) :

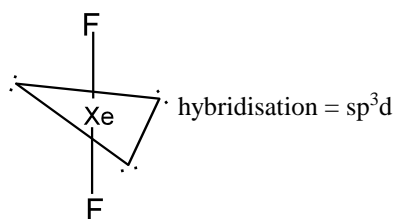


(nitronium fluoride)

(ix) Formation of addition compounds : XeF_2 reacts with fluoride ion acceptors to form cationic species and fluoride ion donors to form fluoroanions.

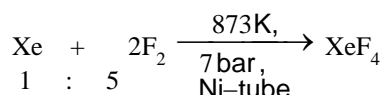


- ☞ Similar behaviour is shown by PF_5 and AsF_5
Structure : Shape linear and geometry trigonal bipyramidal.



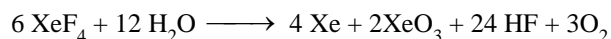
(b) XENON TETRAFLUORIDE (XeF_4) :

PREPARATION:



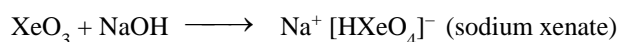
PROPERTIES:

- (i) It is a colorless crystalline solid and sublimes at 298 K.
 (ii) It undergoes sublimation, soluble in CF_3COOH . It undergoes hydrolysis violently hence no moisture must be present during its preparation.
 (iii) **Reaction with H_2O :**

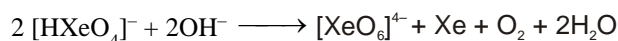


- ☞ XeO_3 is white solid and explosive compound (dry), soluble in water (well behaved in water)

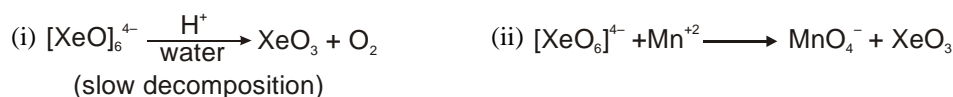
- ☞ XeO_3 reacts with NaOH forming sodium xenate



- ☞ It disproportionates into perxenate ion in basic medium.

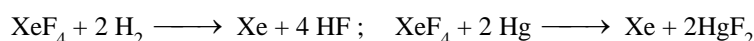


Xenic acid (H_2XeO_4) is a very weak acid.

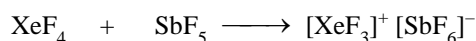


$[\text{XeO}_6]^{4-}$ is obtainable as $\text{Na}_4\text{XeO}_6 \cdot 8\text{H}_2\text{O}$ (sodium perxenate)

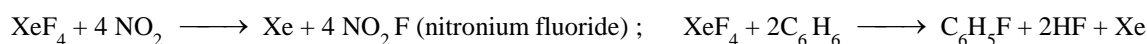
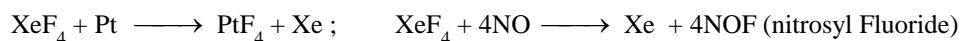
- (iv) **Oxidising properties of XeF_4 :**



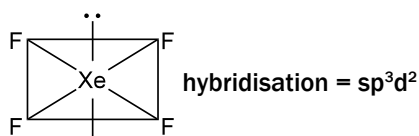
- (v) **Addition reactions :** XeF_4 reacts with fluoride ion acceptors to form cationic species and fluoride ion donors to form fluoroanions.



- (vi) **Fluorinating agent :**

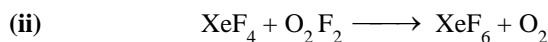
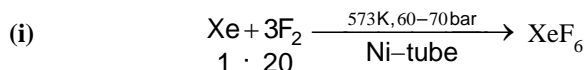


Structure : Shape square planar & geometry octahedral



(c) XENON HEXAFLUORIDE (XeF₆) :

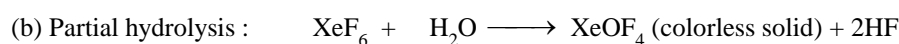
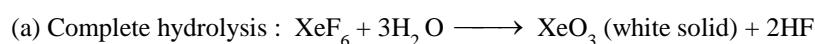
PREPARATION:



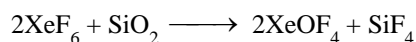
PROPERTIES :

- (i) Colourless crystalline solid and sublimes at 298 K.
- (ii) It gives yellow liquid on melting where as other form white liquids on melting (a point of difference)
- (iii) HF is a good solvent for all three fluorides.

(iv) Hydrolysis :



(v) Reaction with silica (SiO₂) :

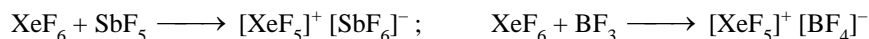


(vi) Thermal decomposition (effect of heat) :

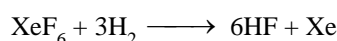


☞ XeF₂ & XeF₄ do not undergo decomposition

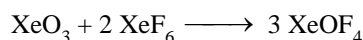
(vii) Formation of addition compounds :



(viii) Reaction With H₂ :

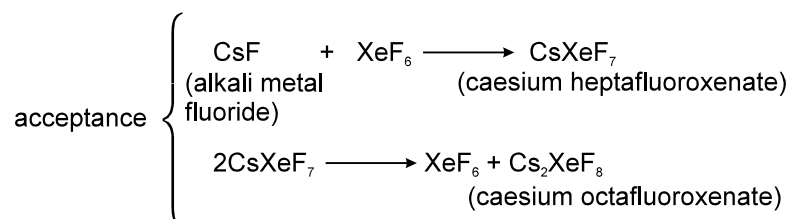
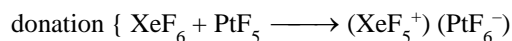


(ix) Reaction of XeF₆ with XeO₃ :



(x) F⁻ donating/ F⁻ accepting properties :

XeF₆ reacts with fluoride ion acceptors to form cationic species and fluoride ion donors to form fluoroanions.



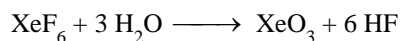
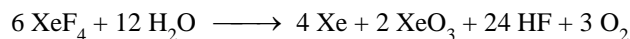
☞ **Order of oxidising power :**



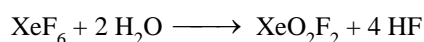
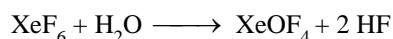
STRUCTURE : Shape capped octahedral (distorted octahedral)

(d) XENON-OXYGEN COMPOUNDS :

Hydrolysis of XeF_4 and XeF_6 with water gives XeO_3 .



Partial hydrolysis of XeF_6 gives oxyfluorides, XeOF_4 and XeO_2F_2 .



XeO_3 is a colourless explosive solid and has a pyramidal molecular structure. XeOF_4 is a colourless volatile liquid and has a square pyramidal molecular structure.

Uses :

Helium is a non-inflammable and light gas. Hence, it is used in filling balloons for meteorological observations. It is also used in gas-cooled nuclear reactors. Liquid helium (b.p. 4.2 K) finds use as cryogenic agent for carrying out various experiments at low temperatures. It is used to produce and sustain powerful superconducting magnets which form an essential part of modern NMR spectrometers and Magnetic Resonance Imaging (MRI) systems for clinical diagnosis. It is used as a diluent for oxygen in modern diving apparatus because of its very low solubility in blood.

Neon is used in discharge tubes and fluorescent bulbs for advertisement display purposes. Neon bulbs are used in botanical gardens and in green houses.

Argon is used mainly to provide an inert atmosphere in high temperature metallurgical process (arc welding of metals or alloys) and for filling electric bulbs. It is also used in the laboratory for handling substances that are air-sensitive.

Xenon and Krypton are used in light bulbs designed for special purposes.