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p-block elements

Introduction:

• In the long form of periodic table G-13 to G-18 are p-block elements.

B	C	N	O	F	He
Al	Si	P	S	Cl	Ne
Ga	Ge	As	Se	Br	Ar
In	Sn	Sb	Te	I	Kr
Tl	Pb	Bi	Po	At	Xe
					Rn

- Those elements in which last electron enters in p-orbital.
- Only "He" is the element which does not have any electron in p-orbital.
- The properties of p-block elements depends upon atomic size, ionization energy, e^- gain enthalpies and electronegativities.

- The absence of d-orbital in second period and presence of d and f orbital in heavier elements has significant effect on the properties of elements.

- p-block elements have all type of elements: metals
non-metals
metalloids.

Group-15 Elements

[Nitrogen family]

- Elements have general electronic configuration is $ns^2 np^3$

Element	Symbol	At. No.	Configuration
Nitrogen	N	7	[He] $2s^2 2p^3$
Phosphorous	P	15	[Ne] $3s^2 3p^3$
Arsenic	As	33	[Ar] $3d^{10} 4s^2 4p^3$
Antimony	Sb	51	[Kr] $4d^{10} 5s^2 5p^3$
Bismuth	Bi	83	[Xe] $4f^{14} 5d^{10} 6s^2 6p^3$

- General electronic configuration is $ns^2 np^3$
- Also known as Pnictogens
- Number of valence electrons = 5
- These elements have a half filled stable configuration.

Occurrence ⇒

- Nitrogen is present in atmosphere and it is 78% by volume. 33rd most abundant element in earth crust
- In living beings, it is present as proteins and amino acids.
- Phosphorous is present as phosphates rocks. It is present in milk and eggs as phosphoproteins.
- In living beings, it is present in nucleic acids.
- Arsenic, Antimony and Bismuth are found as sulphide minerals.

⇒ Physical Properties of Nitrogen Family:

(i) Atomic and ionic radii ⇒

small atomic radii than carbon family because of higher nuclear charge.

- They have
- As we move down the group, the atomic radii goes on increasing because screening effect increases and effective nuclear charge decreases.

• From "Sb" and "Bi", the increase in atomic radius is small because they have completely filled d and f-orbital which cause less screening effect

N	P	As	Sb	Bi
70	110	120	140	150
				(pm)

Ionic Radius (pm)

N^{3-}	P^{3-}	As^{3-}	Sb^{3+}	Bi^{3+}
171	212	222	76	103

- The ionic radii of -ve ions increases down the group because the effective nuclear charge decreases.

But in case of ~~anion~~ Sb^{3+} to Bi^{3+} , there is little increase in radius because antimony and bismuth have completely filled 'd' and 'f' orbitals which cause screening effect.

Ionization enthalpy \Rightarrow

Nitrogen has quite higher 1st ionization enthalpy than carbon because of half filled stable configuration.

As we move down the group, the ionization energy goes on decreasing because atomic size increases and effective nuclear charge decreases

- The successive ionization energies of all ~~the~~ elements increase i.e

$$\Delta I.E_1 < \Delta I.E_2 < \Delta I.E_3$$

It is so because it is more difficult to remove an electron from a +ve ion as it has more effective nuclear charge.

Electronegativity \Rightarrow

• Nitrogen and phosphorus have high electronegativity because of small size and high nuclear charge

◦ As we move down the group, the electronegativity goes on decreasing because atomic size increases and effective nuclear charge decreases. However, the decrease down the group is not regular.

Metallic and non-metallic character →

N P
└───┘
Non
Metals

As Sb
└───┘
Metalloids

Bi
└───┘
Metals

The metallic character increases down the group because ionization enthalpy decreases and so electro-positive character increases.

Allotropy ⇒

Except nitrogen (in gaseous and liquid state) and bismuth, other three elements have many allotropic form.

Phosphorus ⇒ white P, Red P, Black P

Arsenic ⇒ Grey, yellow and black

Antimony ⇒ Yellow and Black.

[Bi, is a metal so it does not have any allotropic form]

Physical State:

Nitrogen is diatomic, while other elements are poly-atomic.

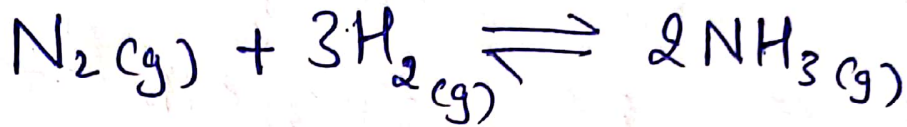
Reason ⇒

N atom is small in size. It completes its octet by forming 1- σ & 2 π bond by $p\pi-p\pi$ overlapping.

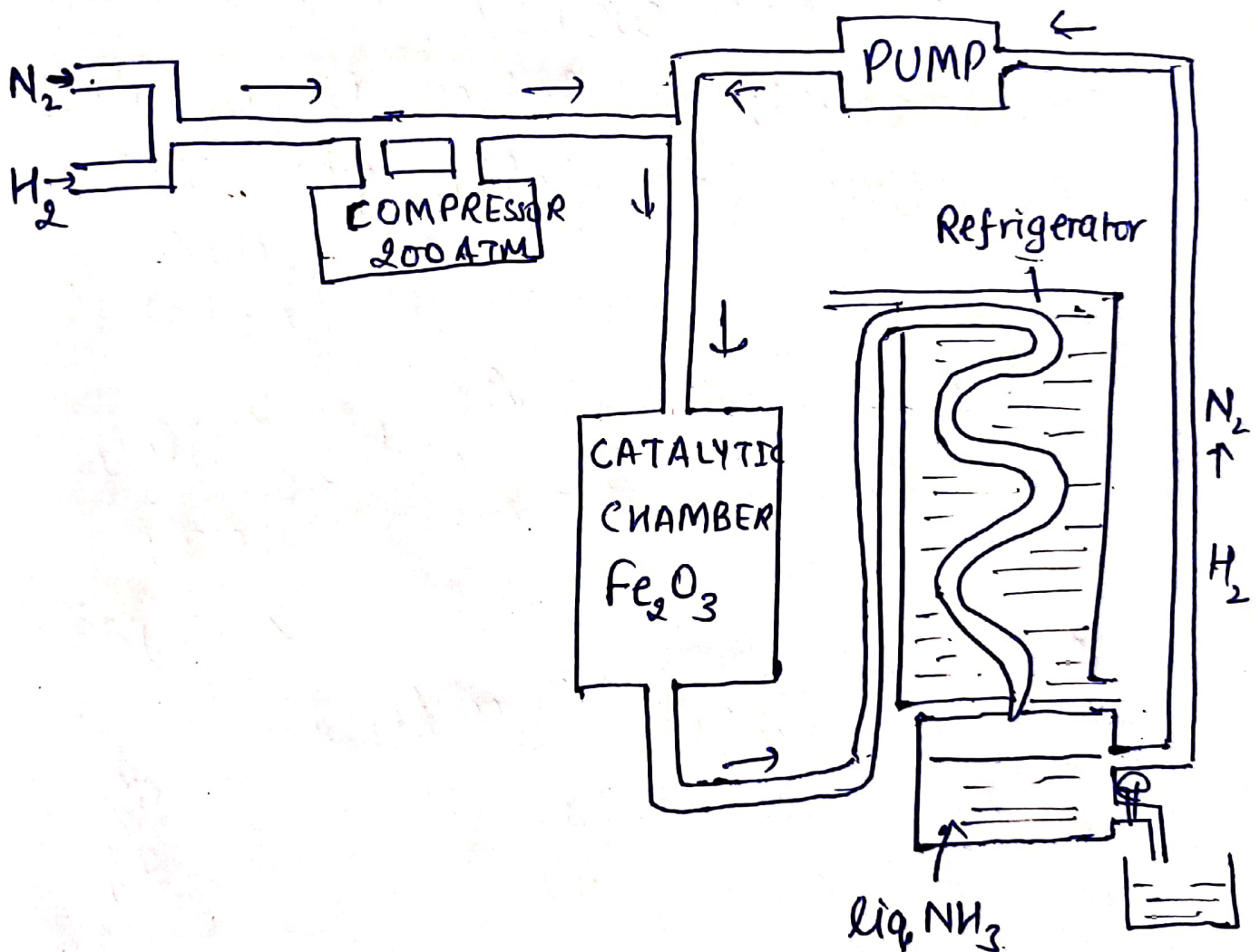
Other atoms have large size, they are unable to form $p\pi-p\pi$ multiple bond. They complete their octet by forming three different bond with diff. element

Ammonia (NH₃)

On commercial scale,
ammonia is prepared by "HABER
PROCESS"



$$\Delta H = -92.4 \text{ KJ/mol}$$



Since the reaction is reversible
So Le-Chatelier Principle is
followed for the maximum
yield.

- Presence of catalyst
- Low Temperature
- High Pressure

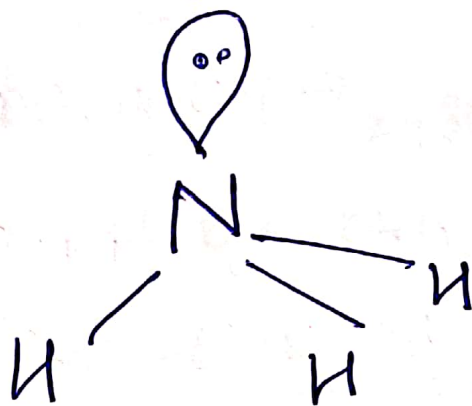
Low Temperature \Rightarrow The production
of ammonia, is an exothermic
process. So temperature should
be low. So optimum temp.
is 700K.

High Pressure \Rightarrow Number of moles
of product is lower than no.
of moles of reactant. So high
pressure is required i.e. 200 atm
pressure is required (applied)

Presence of catalyst \Rightarrow

Rate of R_{xH} is increased by catalyst i.e. iron oxide (Fe_2O_3) with the traces of K_2O and Al_2O_3 .

Structure of ammonia (NH_3) \Rightarrow



Nitrogen is sp^3 hybridised and has three bond pairs and one lone pair. Its geometry is tetrahedral. Bond angle 107.8°

Physical Properties

(i) It is a colourless gas with ammoniacal smell.

(ii) It is highly soluble in water because they can form H-Bond with water.

Uses ⇒

- ⇒ It is used in the preparation of nitrogenous fertilizers
- ⇒ It is used in the manufacture of nitric acid & sodium benzoate
- ⇒ It is used as refrigerant
- ⇒ It is used in the manufacture of rayon

Physical Properties

(i) It is lighter than air

(ii) Ammonia can be easily liquefied under pressure

- It is colourless gas with ammoniacal smell
- It is highly soluble in water because they can form H-Bond

Oxides and oxoacids of nitrogen ⇒

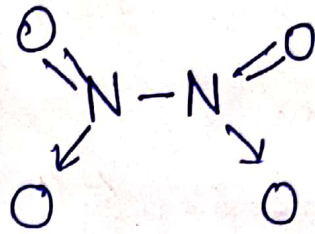
[A] Oxides of nitrogen ⇒

Nitrogen forms six different oxides by combination with oxygen. The oxidation states range from +1 to +5.

Formula	Oxi. State	Lewis formula	Colour
N_2O Nitrous oxide	+1	$:N \equiv N \rightarrow \ddot{O}:$	No
NO Nitric oxide	+2	$:\dot{N} = \ddot{O}:$	No
N_2O_3 Dinitrogen trioxide	+3	$\begin{array}{c} O \\ // \\ N - N \rightarrow O \\ // \\ O \end{array}$	Blue (Acid)
NO_2 Nitrogen dioxide	+4	$\begin{array}{c} \cdot \\ \cdot \\ \cdot \\ \cdot \\ O = \dot{N} \rightarrow O \end{array}$	Red Brown (Acid)

N_2O_4
Dinitrogen
tetroxide

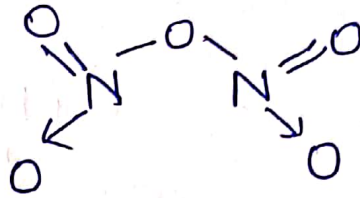
+4



No
(Acid)

N_2O_5
Dinitrogen
pentoxide

+5



No
(Acid)

Q Is there any coloured oxide
of Nitrogen?

Ans. NO_2 is brown gas

N_2O_3 is coloured liquid

Oxoacids of nitrogen

Nitric Acid

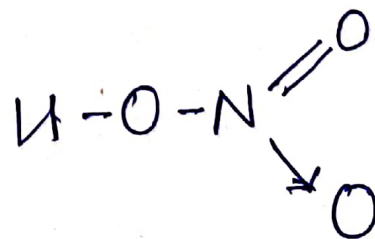
Anhydride

formula $\rightarrow HNO_3$

N_2O_5

Oxidation
State $\rightarrow +5$

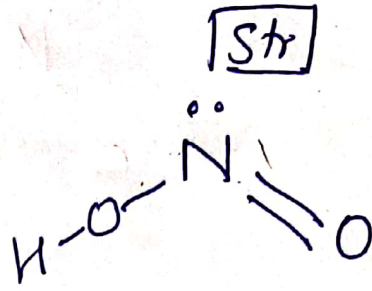
Str.



Nitrous Acid (HNO_2)

Oxi. State $\rightarrow +3$

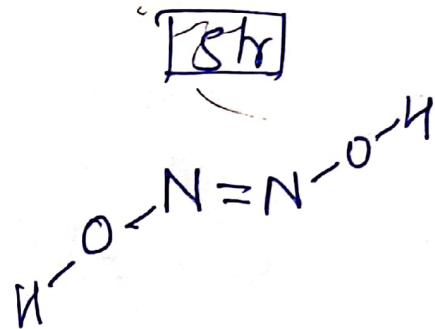
Anhydride $\rightarrow \text{N}_2\text{O}_3$



Hyponitrous Acid ($\text{H}_2\text{N}_2\text{O}_2$)

Oxi. State $\rightarrow +1$

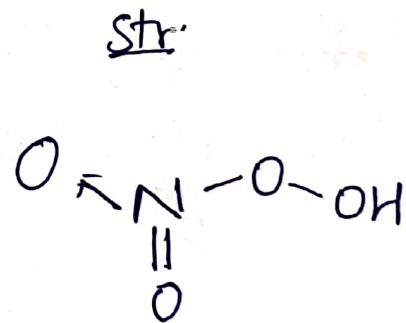
Anhydride $\rightarrow \text{N}_2\text{O}$



Pernitric Acid (HNO_5)

Oxi. State $\rightarrow +5$

Anhydride $\rightarrow \text{N}_2\text{O}_9$



7.1.3 Anomalous Behaviour of Nitrogen

- Nitrogen differs from the rest of the members of group-15. This is called anomalous behaviour of nitrogen

Reason for anomalous behaviour \Rightarrow

- Smaller size
- High ionisation enthalpy.
- High electronegativity
- Absence of d-orbital in the valence shell.

Example of anomalous behaviour

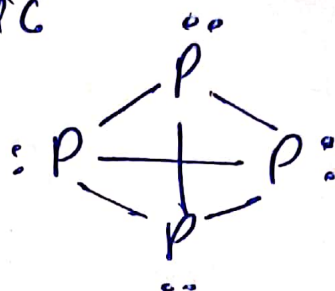
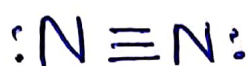
- Nitrogen can form $p\pi-p\pi$ multiple bonds with itself and also with other elements of small size and high electronegativity

eg $N \equiv N$, $N = O$, $C \equiv N$

Other elements of this group can not form $p\pi-p\pi$ bond because their atomic orbitals are diffused and large

- Nitrogen exist as diatomic molecule while other elements of this group are tetra-atomic

e.g



- Nitrogen can not expand its octet and hence can not show valency beyond four while other elements of this group show higher valency using vacant d-orbitals
 - Hydrides of nitrogen form H-bonds. These are more soluble in water and have higher boiling points. Other elements can not form H-bond.
- e.g NH_3 is highly soluble in water but PH_3 is insoluble.

- Bond dissociation energy of nitrogen is very high i.e. 941.4 kJ/mol . Hence it is inert while other elements of this group are very reactive because of
 - large size
 - large single bond
 - low bond dissociation energy

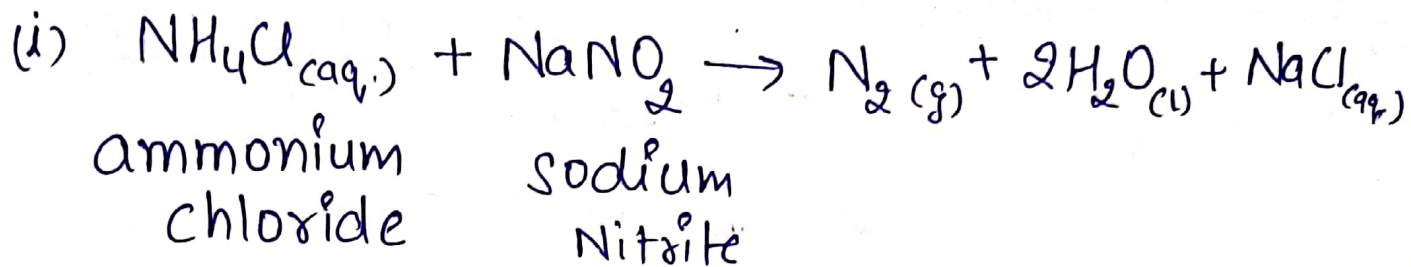
- Nitrogen atom has less catenation property because of inter electronic repulsion of non bonding electrons in nitrogen is high.

As a result N-N single bond is weaker than P-P bond. Other elements of this group have some catenation property.

- Nitrogen does not have vacant d-orbital so it can not form $d\pi-p\pi$ bond while other elements can form $d\pi-p\pi$ as well as $d\pi-d\pi$
 - P form $p\pi-p\pi$ bond in $R_3P=O$
 - P & As form $d\pi-d\pi$ bond with $(T-E)$

Dinitrogen (N_2) \Rightarrow

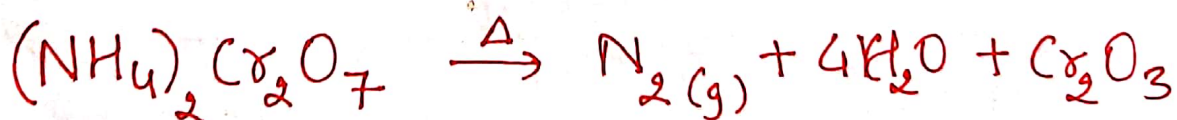
\hookrightarrow Laboratory preparation:



Note

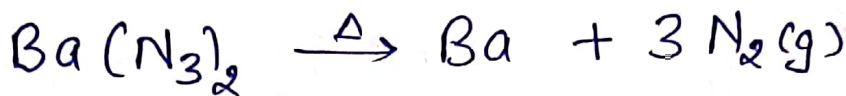
Small amounts of NO and HNO_3 are also formed in this reaction, which are removed by the gas through potassium dichromate and H_2SO_4 .

(ii) Dinitrogen can also be obtained by thermal decomposition of ammonium dichromate $[(NH_4)_2Cr_2O_7]$



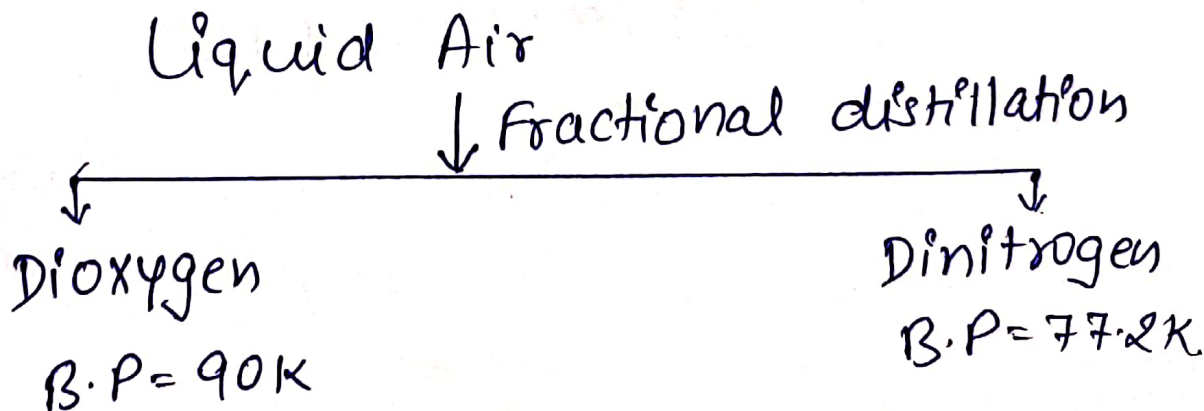
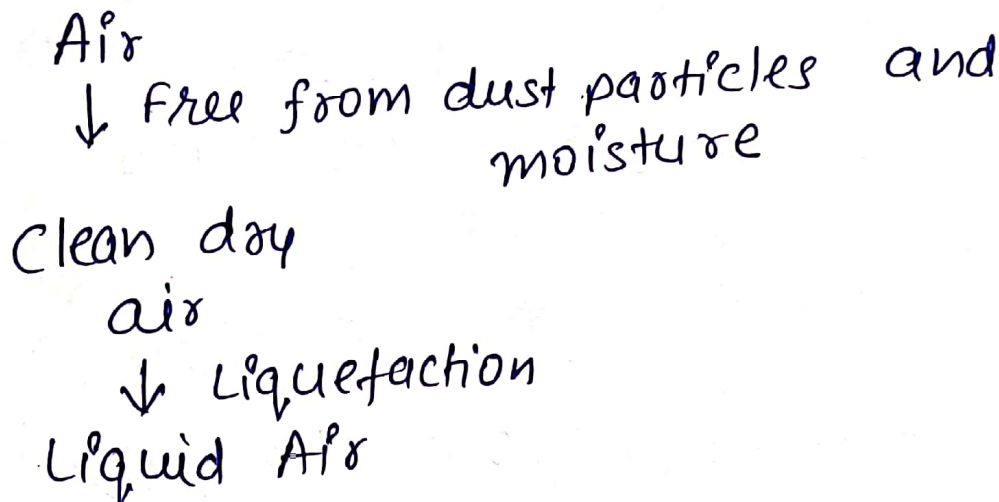
(iii) Very pure dinitrogen ⇒

By thermal decomposition of sodium or barium azide



(iv) Commercial production ⇒

By fractional distillation of liquid air at 77.2K.



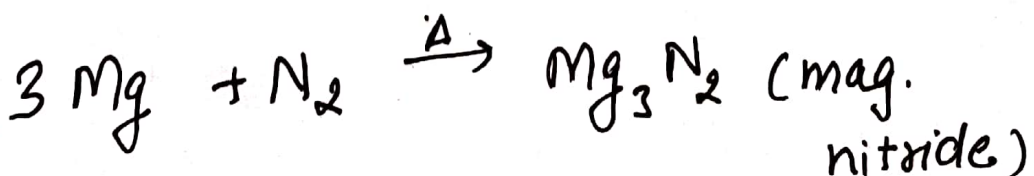
Properties of dinitrogen ⇒

- Dinitrogen is colourless, odourless, tasteless and non toxic gas.
- It is very low soluble in water.
- It has very low M.P and B.P i.e 63 K and 77.2 K respectively.
- Natural dinitrogen has 2 stable isotopes



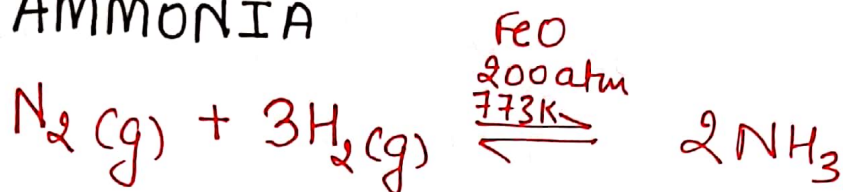
- Dinitrogen is inert at room temp. because bond dissociation energy of $\text{N}\equiv\text{N}$ is very high i.e 941.4 kJ/mol
- Action of metals ⇒

combine with metals directly and form ionic nitride (at high temp).



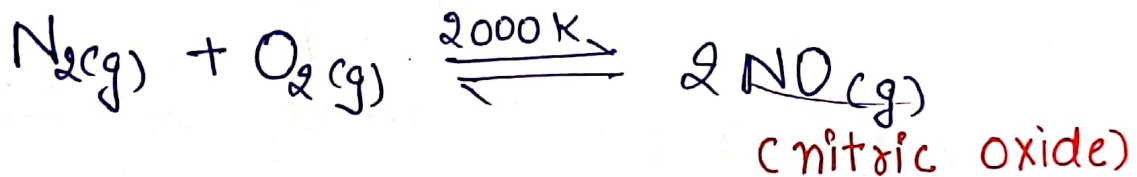
• Action of hydrogen ⇒

Dinitrogen combines with dihydrogen at 773 K and 200 atm. pressure in the presence of iron oxide catalyst to give AMMONIA



(HABER PROCESS)

Action of oxygen ⇒



This reaction is used in the manufacture of nitric acid by

BIRKELAND EYDE PROCESS

▲ Uses of dinitrogen ⇒

- It is used for producing inert atmosphere in lab. and industry e.g in iron and steel industry.

• In the manufacture of ammonia and many other nitrogen containing compounds. such as.

→ Nitric acid (HNO_3)

→ Nitrolim (CaCN_2)

→ Nitrates (NO_3^-)

⇒ Liquid dinitrogen is used as a refrigerant for preserving food items, biological materials.

⇒ It is used ~~for producing~~ as inert diluent for reactive chemicals.

Density ⇒

Density increase regularly from N to Bi.

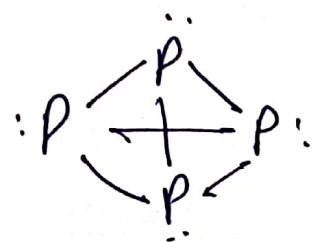
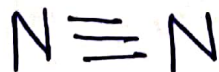
Boiling Point ⇒

Boiling point increases from N to Sb regularly, and then decrease slightly in Bi.

Nitrogen is gas at room temperature because molecular mass of N_2 is small. The molecules are held by weak van der Waals forces.

Other elements of group-15 are solids because they have large molecular masses. They are held by strong van der Waals forces.

Hence these are solids at room temperature



Phosphorus

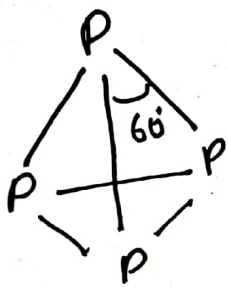
Allotropic forms of Phosphorous \Rightarrow

- (i) white or yellow phosphorous (P_4)
- (ii) Red Phosphorous - $(P_4)_x$
- (iii) Violet phosphorous - $(P_4)_x$
- (iv) Scarlet phosphorous - $(P_4)_x$
- (v) Black phosphorous. - $(P_4)_x$

Reactivity decreases and colour depends on increasing value of "x"

White Phosphorous \Rightarrow

- most common variety.



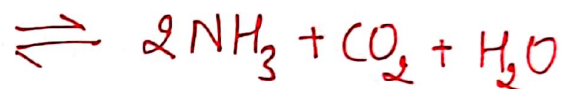
Properties \Rightarrow

- Soft, translucent white, waxy solid with garlic smell.
- It can be cut with knife.
- It melts at 317 K and boils at 553 K.
- It is insoluble in water but soluble in benzene, CS_2 , NH_3 (liq.)

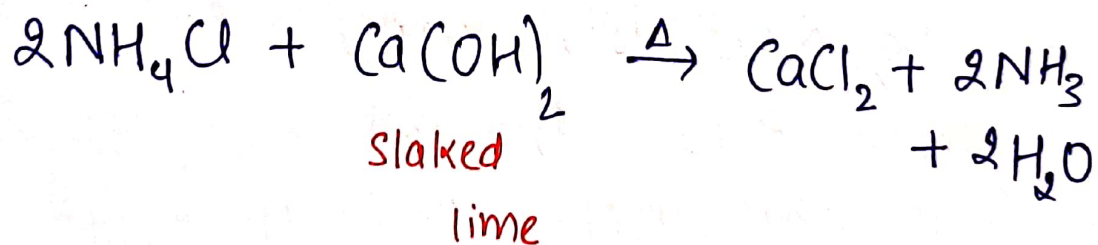
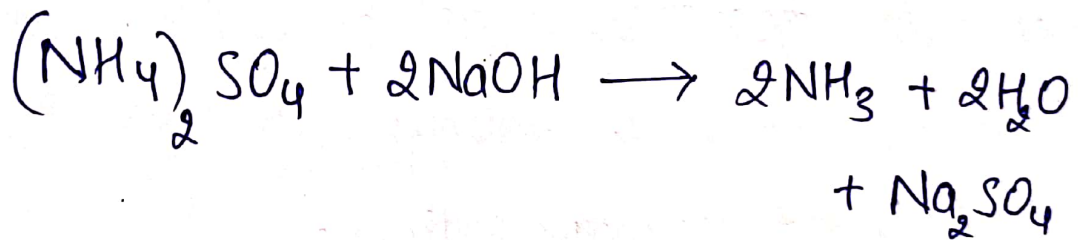
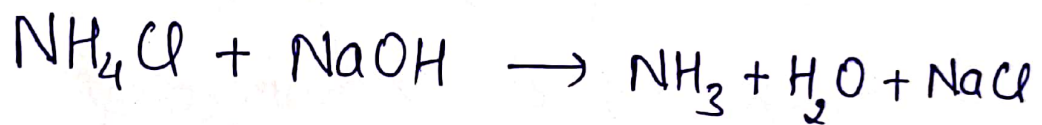
Ammonia (NH₃)

ammonia is the most important compound of nitrogen

In soil, it is formed by the decay of nitrogenous organic matter e.g urea



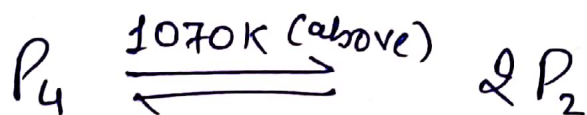
Laboratory Preparation \Rightarrow



- It is very poisonous.
- It glows in dark (a property known as chemiluminescence)

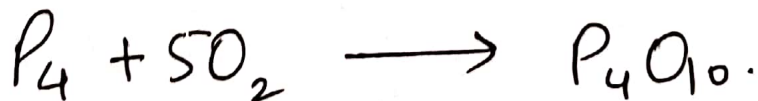
- White phosphorous, P_4 molecules are stable upto 1070 K.

When heated above 1070 K, P_4 molecules begins to dissociate into P_2 molecules.



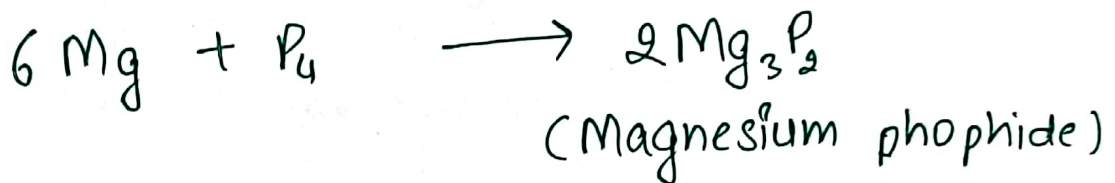
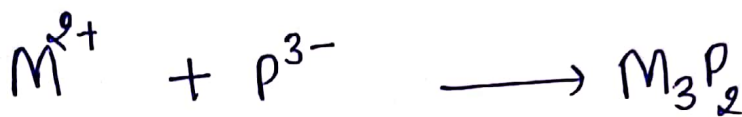
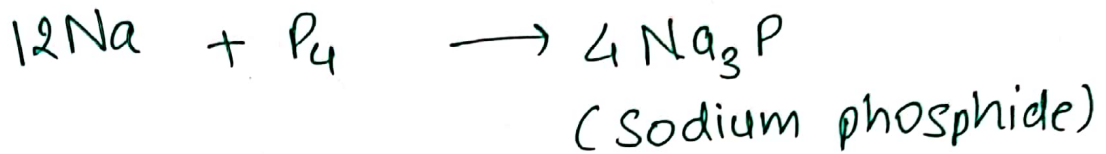
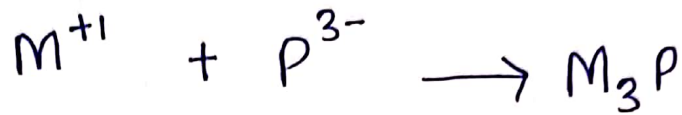
- White phosphorous is less stable and therefore, more reactive than the other solid phases under normal conditions.

Its ignition temp. is very low i.e. 303 K. therefore it catches fire in air to form dense white fumes of P_4O_{10} . It is generally stored under water



Reaction with metals ⇒

Phosphorous combine with metal and form their phosphide

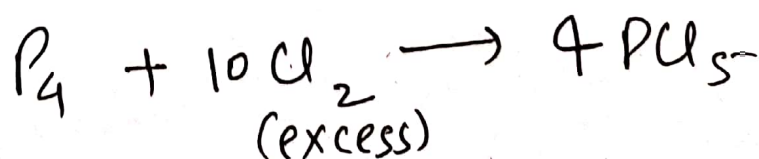


- Reaction with sodium or potassium is very vigorous.

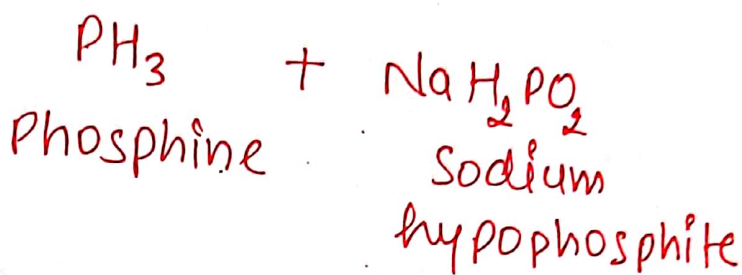
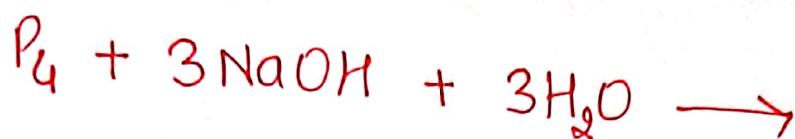
Reaction with Halogens

It readily combines with halogens to form trihalides (PX_3)

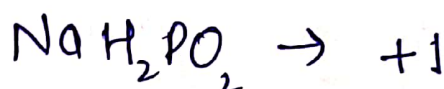
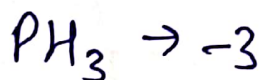
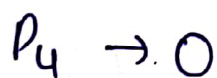
On prolonged treatment forms pentahalides (PX_5)



- On heating with caustic soda solution, it forms phosphine

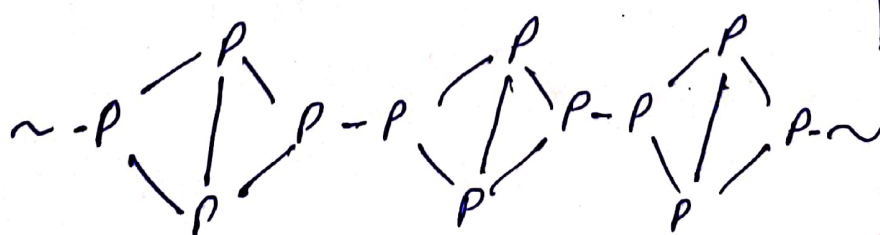
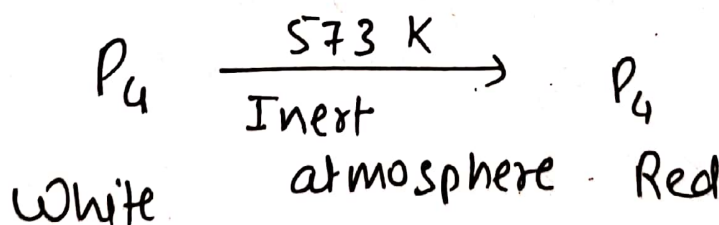


This is an example of "disproportionation reaction"



⇒ Red phosphorous ⇒

white phosphorous obtained by heating in an inert atmosphere for several days.



Str.

Important characteristics ⇒

- It is a hard crystalline solid without any smell. It possesses iron grey lusture.
- It is non poisonous in nature.
- It is insoluble in water as well as carbon di sulphide.
- It does not glow in the dark.
- It is a bad conductor of electricity.
- Red P, is quite stable and its ignition temp is 543 K.
- Being less reactive than white P, it reacts with halogens, sulphur and alkali metals only when heated.
- Red P, does not react with caustic soda.
- Red P, can be converted into white P. by boiling (in inert atmosphere) and then condensing the vapours of white P, formed under water.

Black Phosphorous \Rightarrow

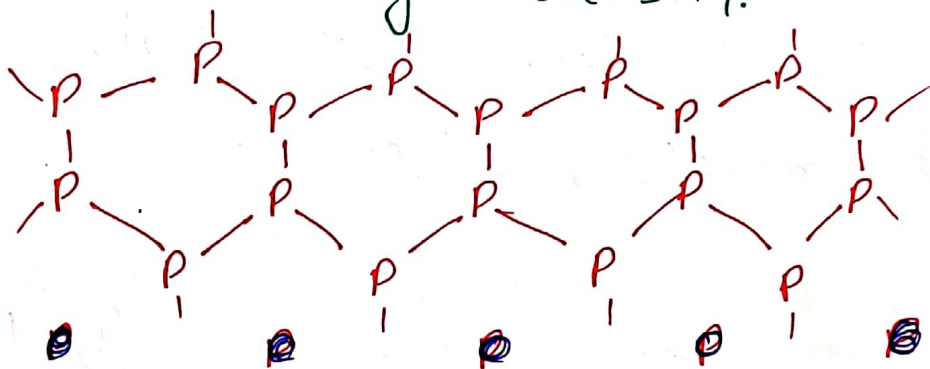
• α -black P

• β -black P

Red P $\xrightarrow[\text{sealed tube}]{803\text{ K}}$ α -black P

white P $\xrightarrow[\text{-12000 atm. press.}]{473\text{ K}}$ β -black P

It has highly polymeric str. As a result high density.



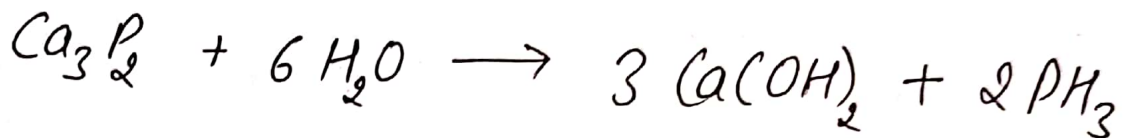
- It has a black metallic lusture
- It has sharp melting point of 860 K.
- It is a moderate conductor of heat and electricity.
- It is thermodynamically the most stable and inactive form of P.
- Its specific gravity is 2.69.

Phosphine (PH_3)

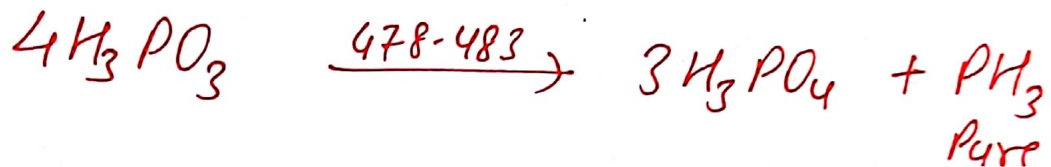
It is hydride of phosphorous

Preparation

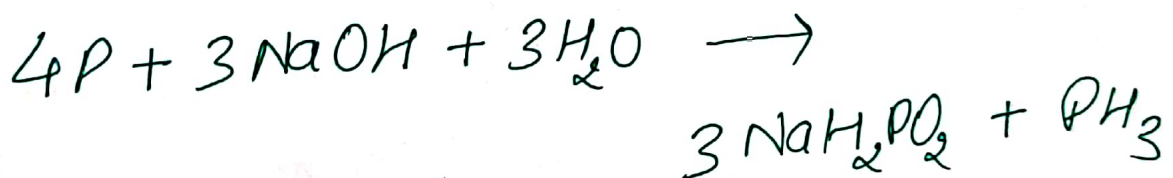
(i) From phosphides:-



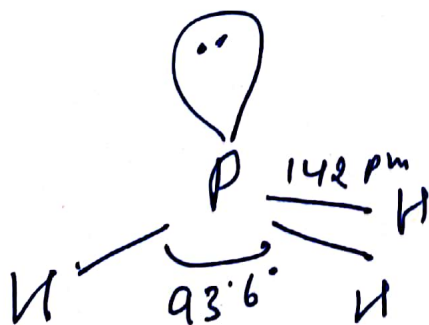
(ii) From phosphorous acid:-



(iii) Laboratory preparation \Rightarrow



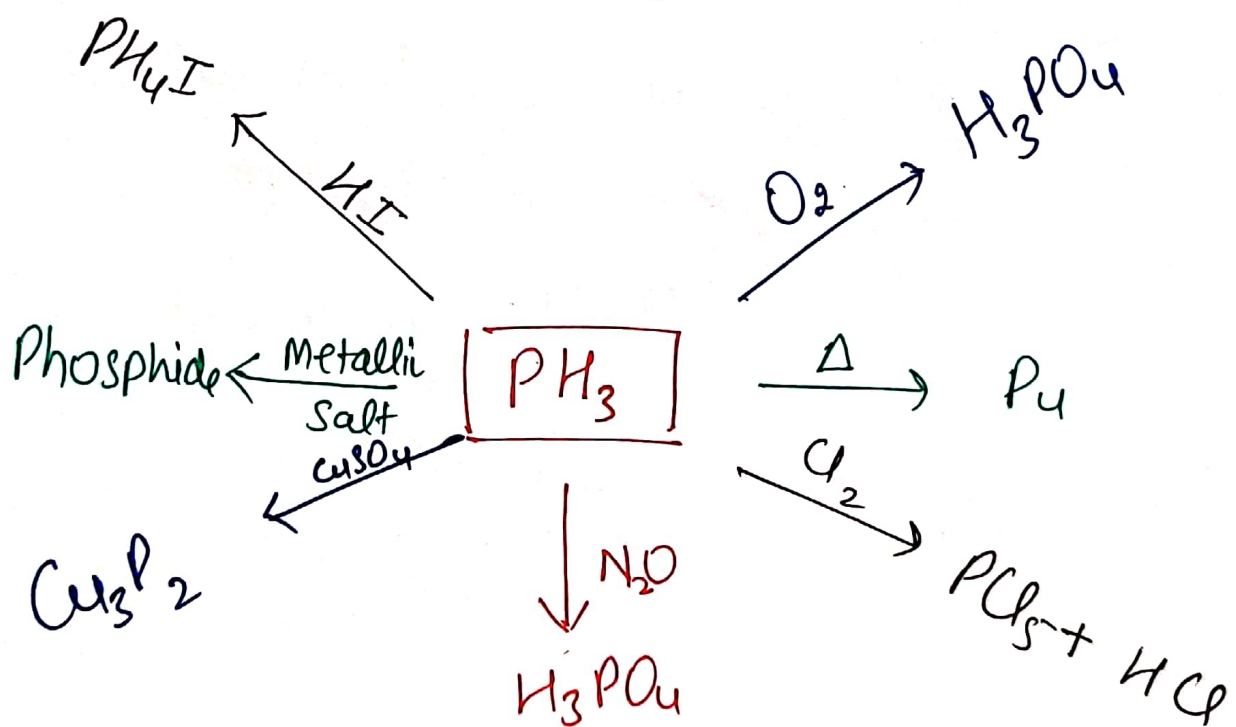
Structure of phosphine



Physical Properties

- (i) colourless gas with unpleasant smell
- (ii) highly poisonous in nature
- (iii) heavier than air
- (iv) slightly soluble in water

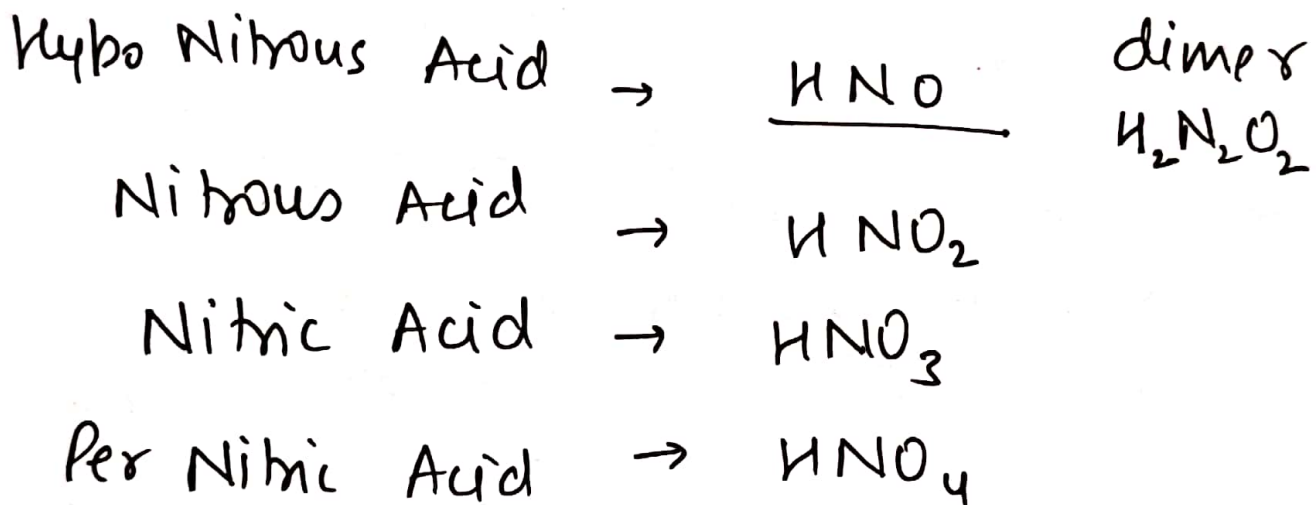
Chemical Properties



Uses

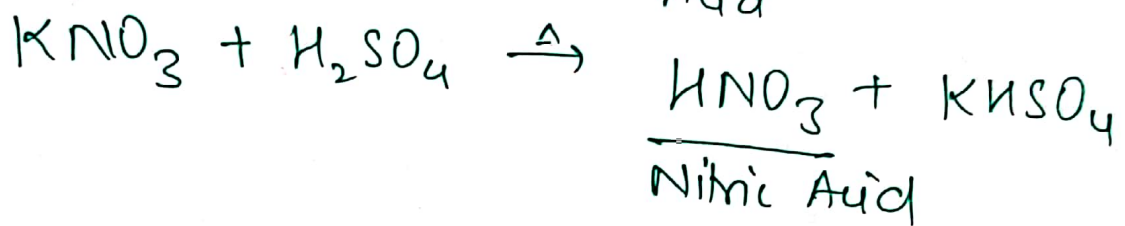
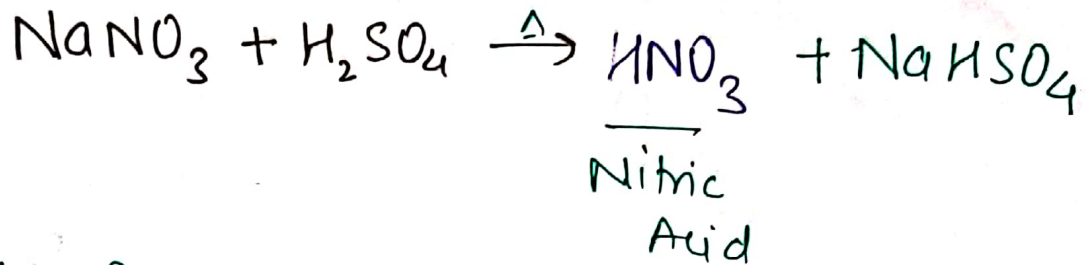
- (i) It is used in preparing Holme's signals.
- (ii) It is used to prepare smoke screens.

Oxoacid of nitrogen.



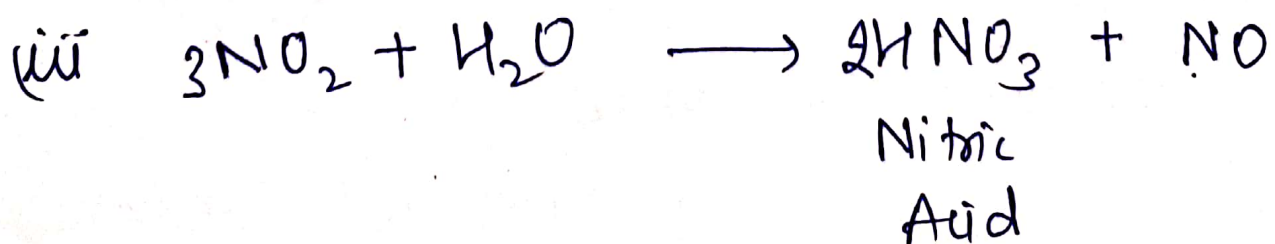
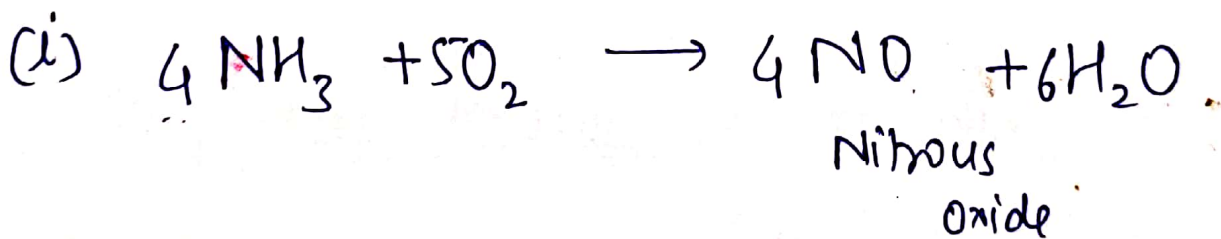
Nitric Acid (HNO₃)

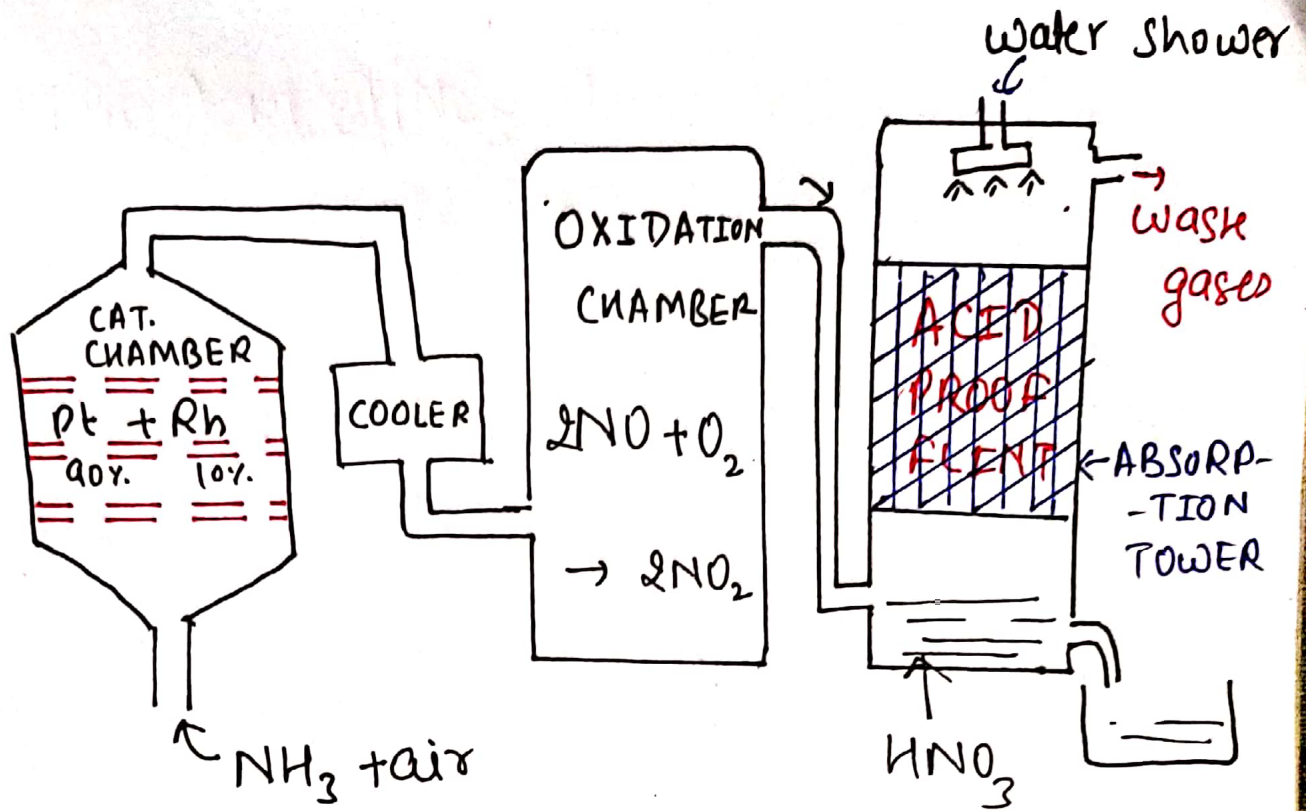
Laboratory Preparation:-



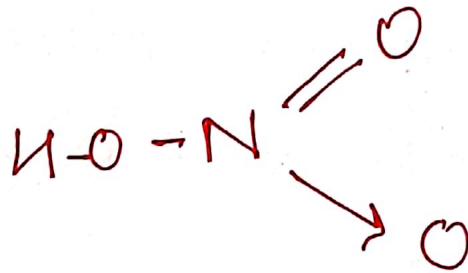
Commercial scale:-

On commercial scale
HNO₃ is prepared by OSTWALD
PROCESS.





Structure of HNO₃



Physical Properties of Nitric Acid

- (i) It is a colourless fuming liquid with pungent odour. However impure nitric acid is yellow due to the presence of dissolved oxides of nitrogen.

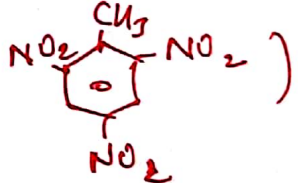
(*) Pure nitric acid freezes at 231.4 K and boils at 355.6 K .

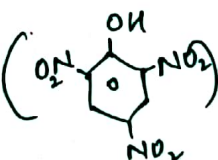
(*) It has a specific gravity of 1.504

(*) Laboratory grade nitric acid is azeotrope (constant boiling mixture) i.e. 394 K

Uses

(i) These are used in the preparation of fertilizers eg ammonium nitrate, basic calcium nitrate.

(ii) In the manufacture of explosives eg T.N.T (Tri Nitro Toluene )

Picric Acid 

(iii) In the manufacture of perfumes, medicines, dyes

(iv) In the manufacture of artificial silk.

- (*) for purification of gold and silver
- (*) In rocket fuels (as an oxidiser)
- (*) As a reagent in laboratory
- (*) For pickling (cleaning) of stainless steel and etching of metals.

Oxides of Phosphorous →

forms two common ^{phosphorous} oxides

- (i) phosphorous trioxide (P_4O_6)
- (ii) phosphorous pentoxide (P_4O_{10})

1) Phosphorous (III) oxide (P_4O_6)

Preparation

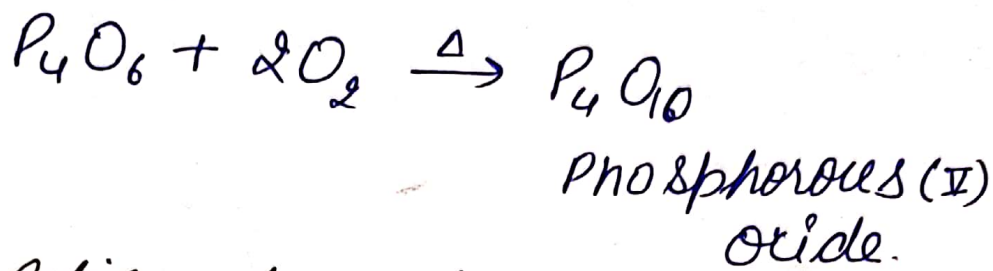


Properties

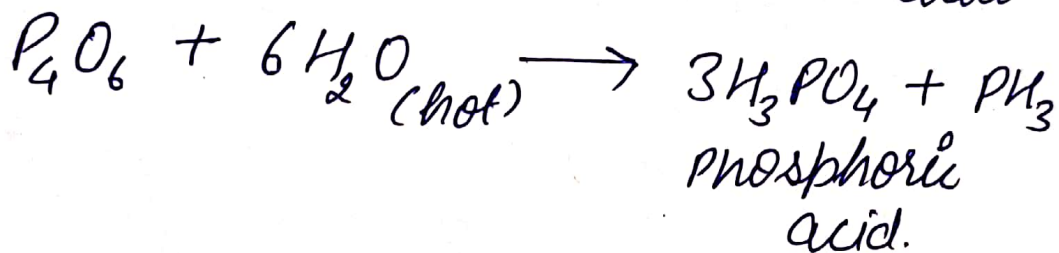
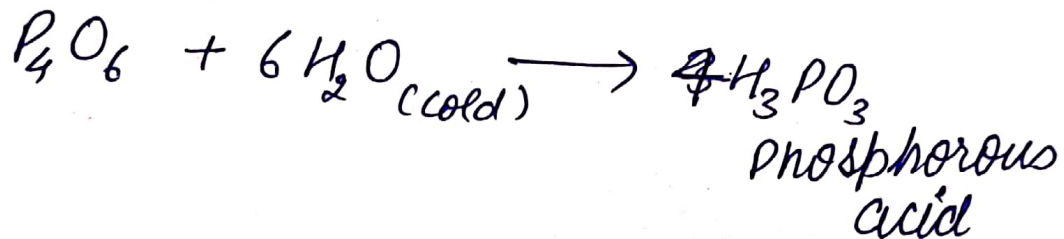
- It is a crystalline solid with garlic odour

It is soluble in CS_2 , $R-O-R$, $CHCl_3$.

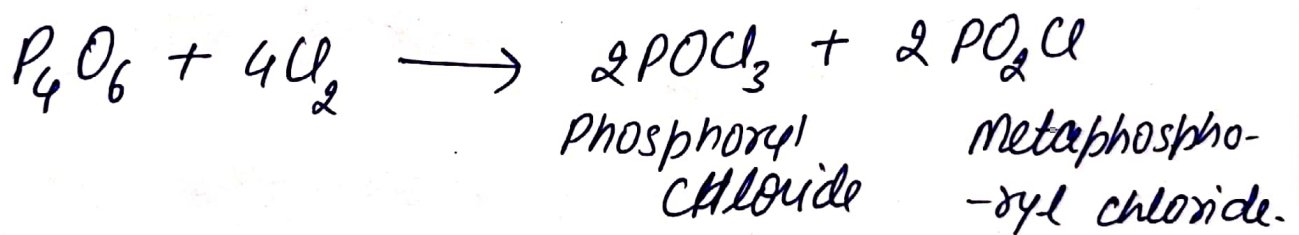
Heating in air \Rightarrow



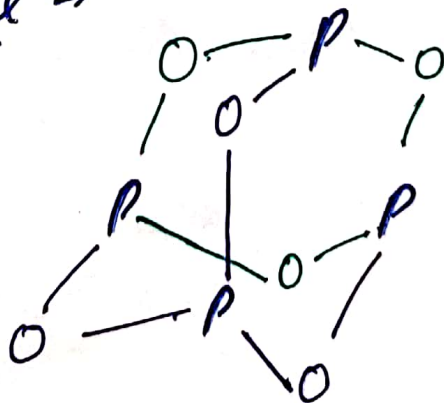
Action of water \Rightarrow



Action with chlorine \Rightarrow

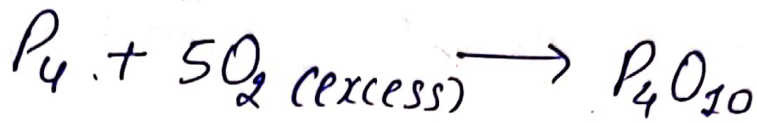


Structure \Rightarrow



2.) phosphorous (V) oxide [P_4O_{10}]

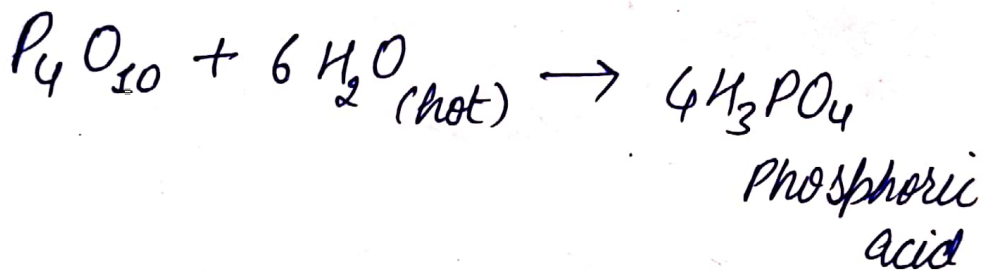
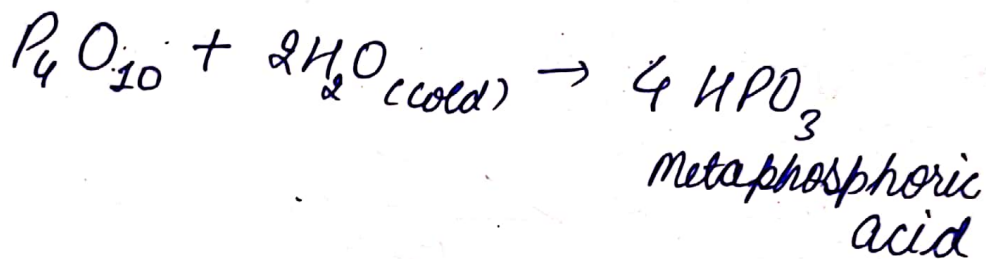
◦ Preparation \Rightarrow



◦ Properties \Rightarrow

(*) It is snowy white solid

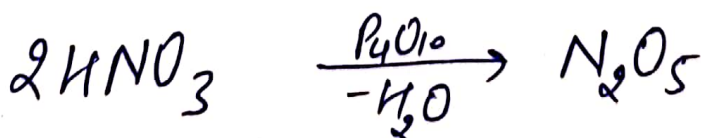
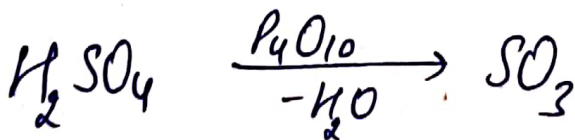
(*) Action with water \Rightarrow

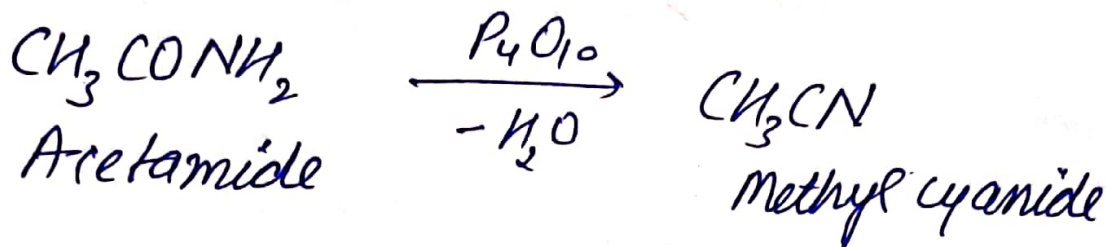
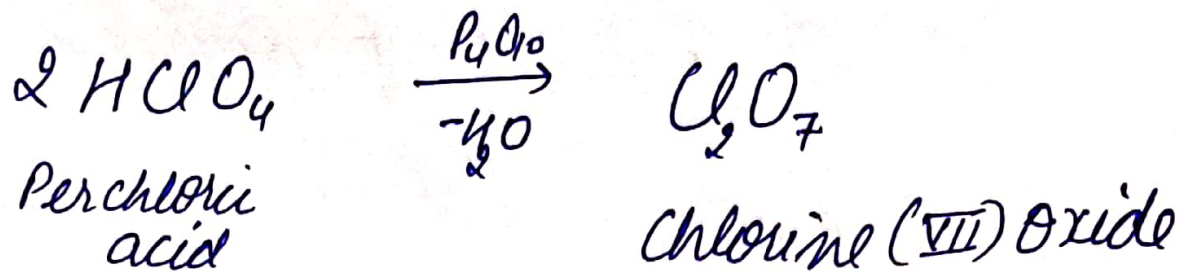


Dehydrating nature:

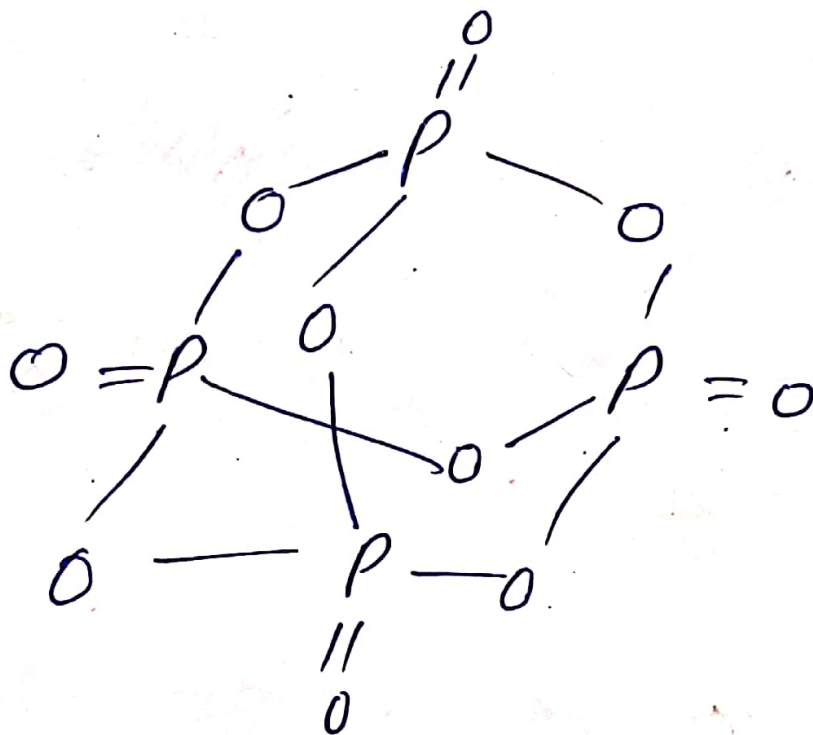
phosphorous pentoxide

has strong affinity for water





Structure of Phosphorus



Group-16 Elements

Group-16

		At.No.	Electronic Configuration
O	- Oxygen	8	[He] $2s^2 2p^4$
S	- Sulphur	16	[Ne] $3s^2 3p^4$
Se	- Selenium	34	[Ar] $3d^{10} 4s^2 4p^4$
Te	- Tellurium	52	[Kr] $4d^{10} 5s^2 5p^4$
Po	- Polonium	84	[Xe] $4f^{14} 5d^{10} 6s^2 6p^4$

Group 16 elements are named as oxygen family. Sometimes first four members are also known as Chalcogens. (means ore forming). First four members are non metallic while Polonium is metallic in nature.

Occurrence

- Oxygen is most abundant of all elements
- Sulphur is 16th most abundant element
- O_2 makes 20.946 % by volume of the atmosphere and most of it has been produced by photosynthesis
- Oxygen makes up 46.6 % by mass of earth crust.

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- Oxygen makes up 46.6% by mass of earth crust.
- Sulphur occurs in the combined form as sulphide ores or sulphate ores e.g Galena (Pbs), zinc blende (Zns), Copper pyrites ($CuFeS_2$), Gypsum ($CaSO_4 \cdot 2H_2O$) epsom salt ($MgSO_4 \cdot 7H_2O$), Traces of sulphur occur as hydrogen sulphide in volcanoes.
- Organic materials such as eggs, proteins, garlic, onion, mustard, hair and wool contain sulphur.
- 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

Group 16 elements have six e⁻ in the valence shell and have $ns^2 np^4$ general electronic configuration.

Atomic and ionic radii

The atomic and ionic radii of the elements of this group are smaller than corresponding elements of group 15. The atomic and ionic radii of elements of G-16, as expected, increase on moving down the group.

Explanation

- (i) due to increased effective nuclear charge of Group-16 elements
- (ii) due to the increase in the number of electron shells

Ionisation enthalpy

The ionisation enthalpies of the elements of oxygen family are less than those of nitrogen family

As we move down the group from oxygen to polonium, the ionisation enthalpy decreases

Explanation

First I.E of the elements of G-16 are lower than the elements of G-15 due to higher stability of exactly half filled p-orbital in G-15.

Second I.E of G-16 elements is higher than G-15 elements due to higher stability of p^3 in G-16.

	N	O	(KJ/mole)
I.E ₁	1402	1314	
I.E ₂	2856	3388	

Melting and Boiling Point

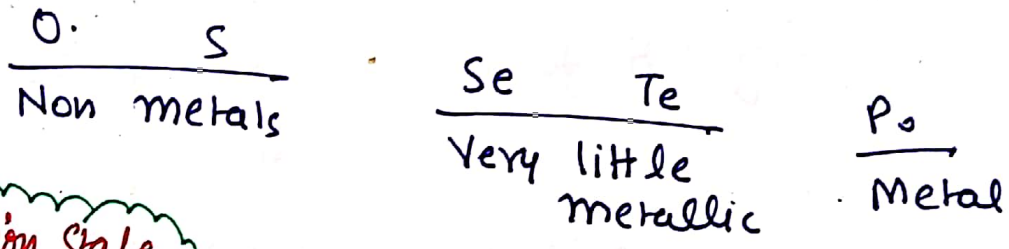
Melting and boiling point increase regularly from O to Te due to increase in size and hence greater van der Waal forces. The last element Po, has lower m.pt and b.pt than Te because of the maximum inert pair effect

Electronegativity

These are more electronegative than G-15 elements because of decrease in size. It decreases down the group because of increase in atomic size. Oxygen is second most electronegative element after fluorine

Metallic and Non metallic character

These elements have little metallic character because of High I.E. Down the group, the metallic character increases due to decrease in I.E

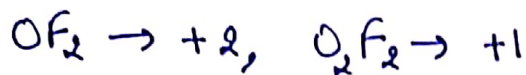


Oxidation State

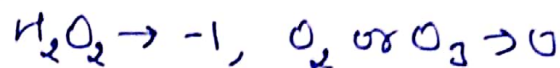
All the elements of this group ns²np⁴ oxidation state

Generally show -2 Oxidation state, as well they can show +2, +4 and +6 Oxidation state

Note:- Oxygen does not show positive oxidation state except in OF_2 and O_2F_2



exception

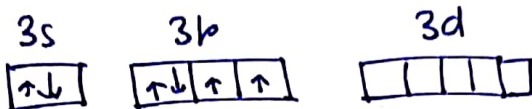


Oxygen atom in ground state



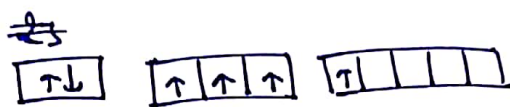
There are no d-orbitals, excitation not possible

Sulphur atom in ground state



Two unpaired e⁻ accounts for +2 O.S

Sulphur atom in first excitation state



Four unpaired e⁻ accounts for +4 O.S

Sulphur atom in 2nd excitation state



Six unpaired e⁻ accounts for +6 O.S

with fluorine they can show +6 O.S.

S, Se, Te with O generally +4 O.S. They show both oxidising and reducing properties.

Allotropy

All the elements of the group exhibit allotropy

Oxygen :- O_2 and O_3

Sulphur :- Yellow ortho rhombic α and β -monoclinic form

Electron Gain enthalpy

The elements of this family have negative electron gain enthalpies. The value decrease from Sulphur to polonium

Oxygen, unexpectedly has low negative electron gain enthalpy

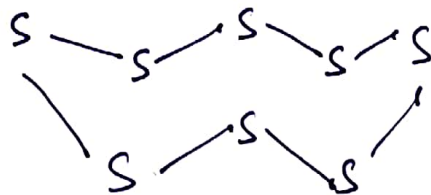
Reason due to small size of oxygen atom, its e^- cloud is distributed over a small region of space and therefore it repels incoming electron

Elemental state or physical state

Oxygen is a gas while others are solids. Oxygen molecule is diatomic while the molecule of other elements are more complex

Sulphur, selenium and tellurium exist as staggered 8-atoms rings.

Str. of S_8 molecule



Reason

The oxygen atom has a tendency to form multiple bond ($\pi-\pi$) with the other oxygen atom on account of small size and high electronegativities ($O=O$). The molecules are held together by weak Van der Waal forces of attraction, hence oxygen exist as a gas at room temperature.

Rest elements do not form $\pi-\pi$ multiple bonds due to their large size and does not exist as diatomic molecule. They prefer to form single bonds and possess complex structure crown type

Catenation

Oxygen and sulphur show the property of catenation. The property is more pronounced in sulphur, S-S bond is much stronger (213 KJ/mol) and O-O bond (138 KJ/mol) because of small size of oxygen atom, the lone pairs of e^- on the oxygen atom repel the bond pair of e^- on the oxygen atom to a greater extent than the lone pairs of e^- on the sulphur atoms in S-S bond.

Chemical Properties

Anomalous Behaviour of Oxygen

Like other elements of 2nd period oxygen shows anomalous behaviour from rest of elements of group-16 due to.

- Small size
- High electronegativity
- absence of vacant d-orbital

(absence of vacant d-orbital limits its covalency to four)

Reaction with Hydrogen/Hydrides

All the elements of group-16 form hydrides of the type H_2E ($E = O, S, Se, Te, Po$)

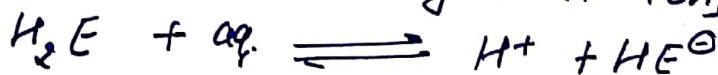
Properties

(a) Thermal stability

It decreases down the group because as the size of atom increases the bond dissociation enthalpy decreases.

(b) Acidic character

It increases down the group because they can dissociate more easily to give H^+ ions.



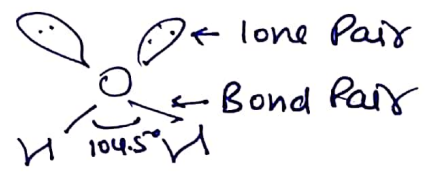
Reducing Nature

Except H_2O , all are reducing agents. The reducing character increases down the group.

Boiling Point/volatility

H_2O has the highest B.Pt because of H-bonding. After a sudden drop from H_2O to H_2S . boiling point gradually increases from H_2S to H_2Te because of increase of molecular mass and hence Van der Waals forces.

shape



Geo - Tetrahedral
Shape - Bent/v-shape

Anomalous Behaviour of Oxygen

Oxygen differs from its family members. The anomalous behaviour of oxygen is due to its

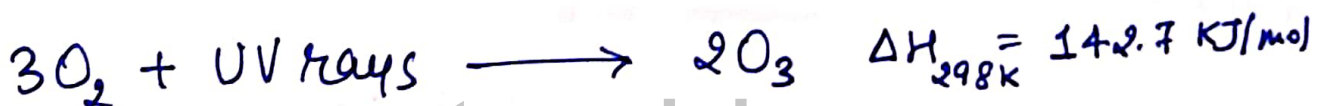
- (i) small size
- (ii) high value of ionization enthalpy.
- (iii) high electronegativity
- (iv) absence of vacant d-orbital in its valence shell

Different Properties

- Oxygen is a gas at ordinary temp while other elements exist as solid at room temperature
- Oxygen molecule is diatomic while the molecules of other elements are more complex
- Molecular oxygen is paramagnetic whereas the molecules of other members are diamagnetic in nature
- Oxygen can show only -2 oxidation state while other elements can show +2, +4, +6 oxidation state in addition to -2.
- The hydride of oxygen (H_2O) is a liquid at room temperature while the hydrides of other members are gases
- Oxygen can form pπ-pπ multiple bond while other elements can't

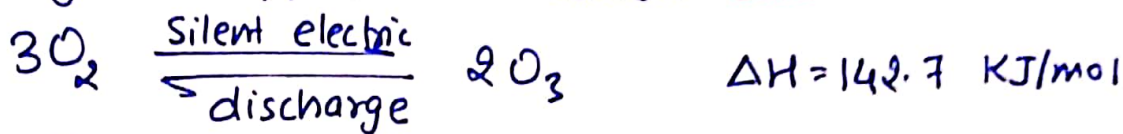
Ozone

Ozone is an allotropic form of oxygen. It is present in upper atmosphere about 20 km above the surface of the earth. It is formed by action of UV rays on oxygen as



Preparation of Ozone

Ozone is prepared by passing silent electric discharge through pure, cold and dry oxygen in a specially designed apparatus called **ozoniser**



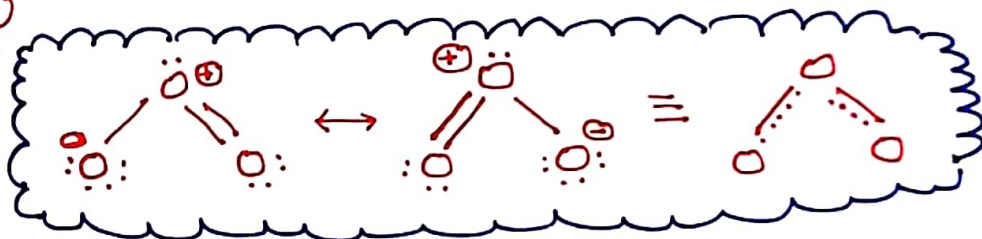
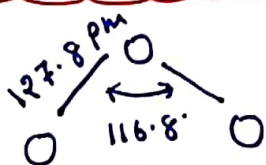
in lab it is prepared by two type of ozonisers
- siemen's ozoniser - Brodie's ozoniser

Physical Properties

Sparkless electric discharge is called silent electric discharge

- It is pale blue gas having pungent odour
- It is heavier than air. It is about 1.5 times heavier than air.
- Its b.pt is 385 K and m.pt is 24 K

Structure of ozone

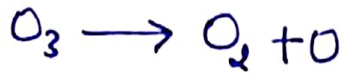


Resonating str. of ozone

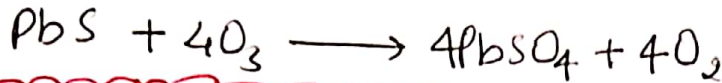
Uses Of Ozone

- used as germicide, disinfectant and for sterilising water.
- It is used for bleaching, flour delicate fabrics, oils, flour, starch, ivory etc.
- used for purifying air of crowded places such as cinemas, underground railways, auditoriums, tunnels, mines, etc. and for destroying objectionable odours in slaughter house
- It is used for manufacture of KMnO_4
- In lab. for the ozonolysis of organic compounds.

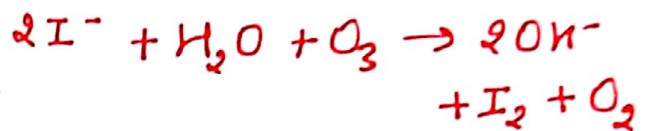
Due to the ease with which it liberates atoms of nascent oxygen. It acts as a powerful oxidising agent



-lead sulphide to lead sulphate



iodide ions to iodine

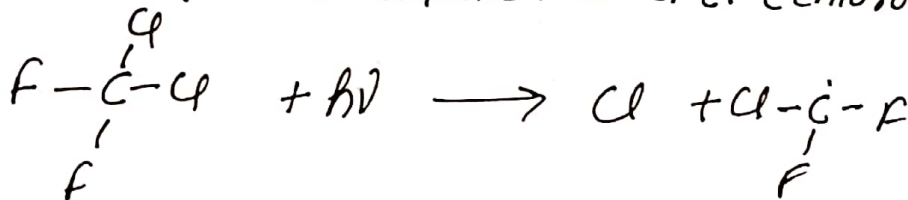


Depletion of Ozone layer

thick layer^{of ozone} in upper atmosphere is called ozone blanket.

Without this protective ozone layer, more U.V rays reach the surface of earth can cause damage to plant and animal life. UV radiations may cause skin cancer, damage to immune system

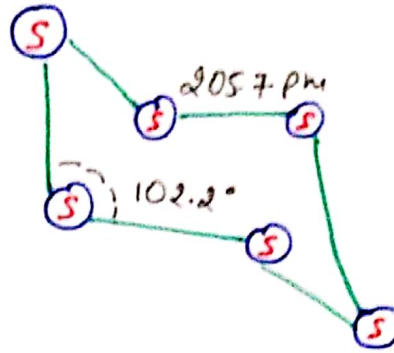
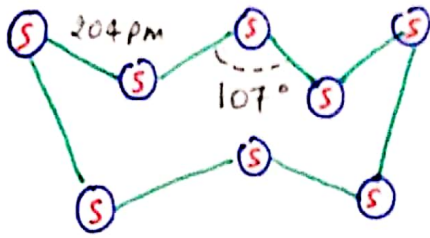
main cause of ozone depletion is CFC. (Chloro fluoro Carbon)



another cause of ozone depletion is because of release of NO_x (Nitrogen Oxide) into the stratosphere by the exhaust system of supersonic jet aeroplanes.



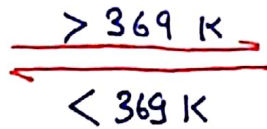
Allotropes of Sulphur



Three allotropes of sulphur

- α -sulphur (Rhombic)
 - β -sulphur (Monoclinic)
 - γ -sulphur (Plastic)
- most important

Rhombic
Sulphur
(α)



Monoclinic
Sulphur
(β)

Rhombic Sulphur

Common form of sulphur and is formed by slowly evaporating the roll sulphur in CS_2

Properties

- It is bright yellow in colour
- Its specific gravity is 2.06 g/cm^3
- Its m.pt is 385.8 K
- It is insoluble in water but dissolve to some extent in benzene, alcohol and ether.
- It is readily soluble in CS_2
- It has low thermal and electrical conductivity

Monoclinic Sulphur

It is prepared by melting rhombic sulphur in a dish. The molten sulphur is allowed to cool till a crust is formed. Two holes are made in the crust and remaining liquid is poured out. On removing the crust, colourless needle shaped crystal of β -sulphur are formed.

Properties

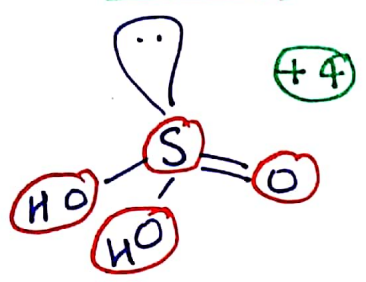
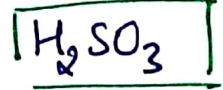
- It is dull yellow in colour solid
- It is insoluble in water but soluble in CS_2
- Its m.p.t is 393 K and specific gravity is 1.98
- It is stable above 369 K above this temp. transform slowly in α -sulphur

At 369 K, both the forms are stable. This temp. is called transition temperature

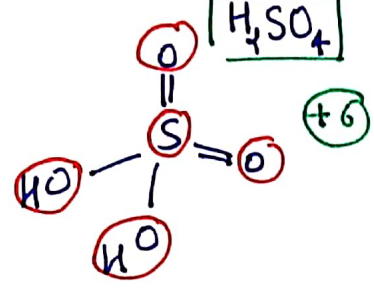
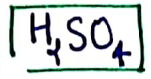
Oxoacids of Sulphur

Sulphur form a number of oxoacids

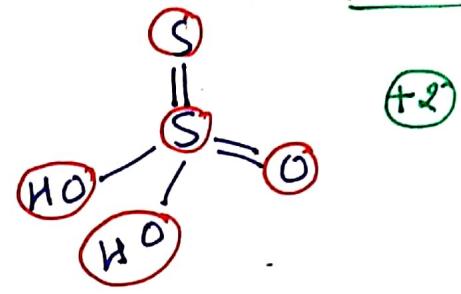
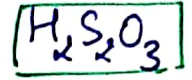
Sulphurous Acid



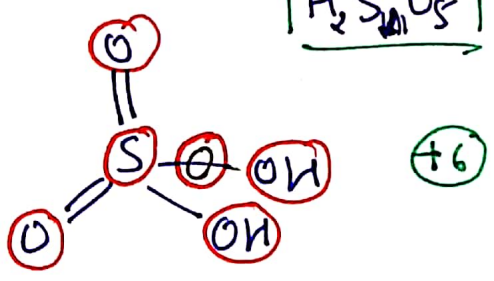
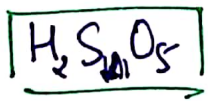
Sulphuric Acid



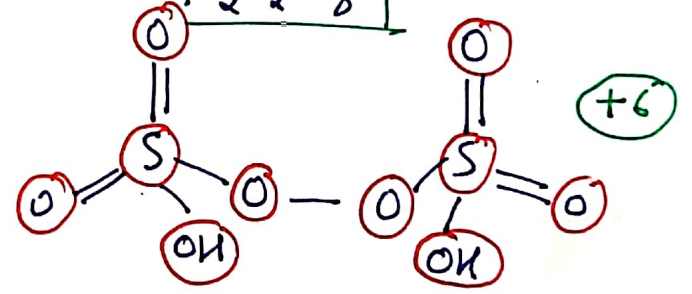
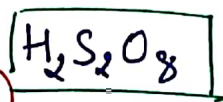
Thio sulphuric Acid



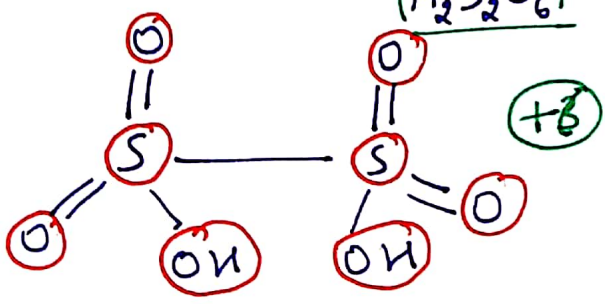
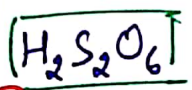
Peroxomonosulphuric Acid (Caro's Acid)



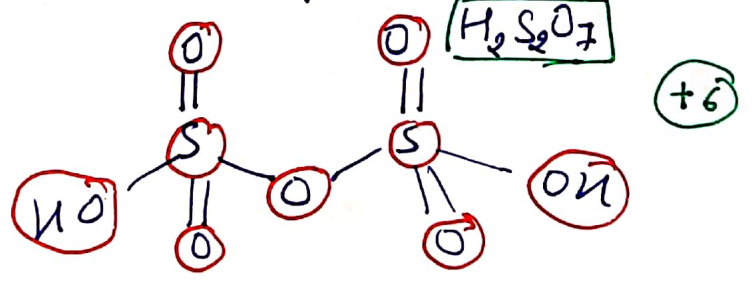
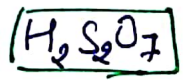
Peroxodisulphuric Acid (Marshall acid)



Dithionic Acid



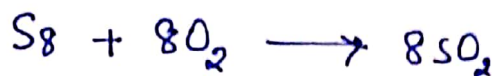
Pyrosulphuric Acid



Sulphur Dioxide, SO_2

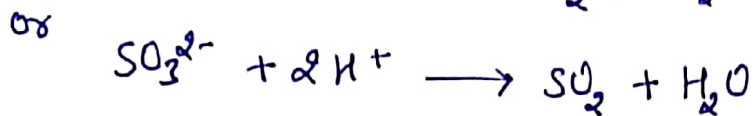
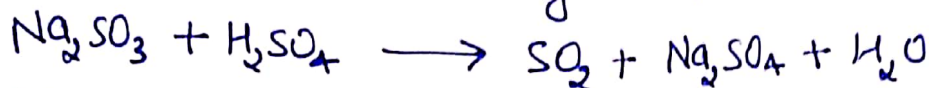
Preparation

Sulphur dioxide is formed together with a little (6-8%) sulphur trioxide when sulphur is burnt in air or oxygen

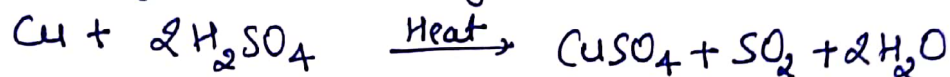


In laboratory

Prepared by treating a sulphite with dil. sulphuric acid



- Also by heating copper turnings with conc. H_2SO_4



In Industry

By product of roasting of sulphide ores such as iron pyrites.



Properties

- colourless, toxic gas with pungent & suffocating odour
- heavier than air
- Highly soluble in water
- It can be easily liquefied at room temperature

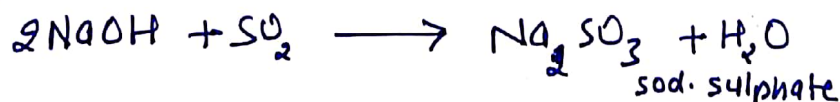
Acidic Nature

Sulphur dioxide dissolve in water giving sulphurous acid.



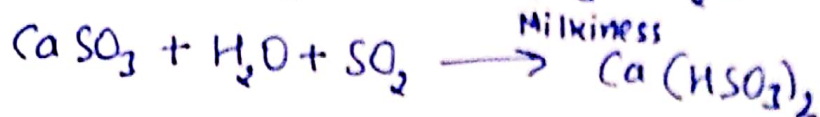
aq. solution is acidic turning blue litmus into red.

Because of its acidic nature, it readily reacts with sodium hydroxide solution forming sodium sulphite which then reacts with more SO_2 to form sodium hydrogen sulphite



sod. bisulphate

when gas is bubbled through alkalis



Cal. bisulphite

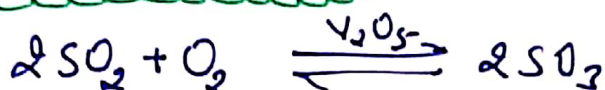
(Milky disappears)

Combustibility

The gas is non-combustible. However certain substances such as carbon, magnesium etc extract oxygen from the gas and burnt in it when ignited

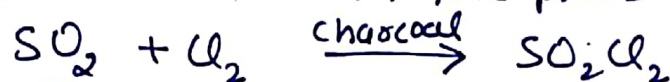


Combination with oxygen

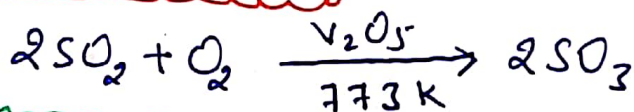


Combination with Halogens

reacts with halogens (F, Cl, Br) forming sulphonyl halide in the presence of charcoal.



Combination with oxygen



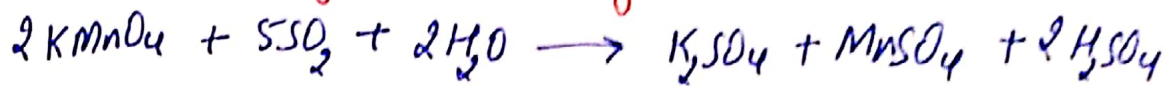
Reducing Properties

SO_2 act as good reducing agent

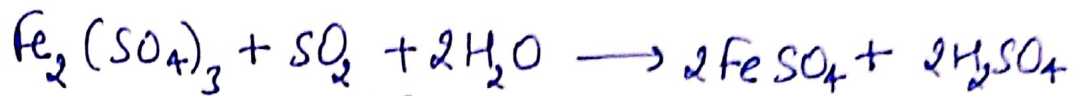
→ Potassium dichromate to chromium sulphate



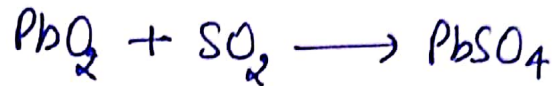
(b) potassium permanganate to manganese sulphate



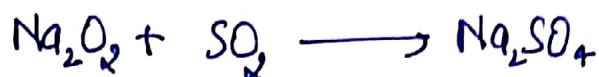
⇒ Ferric sulphate to ferrous sulphate ($\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$)



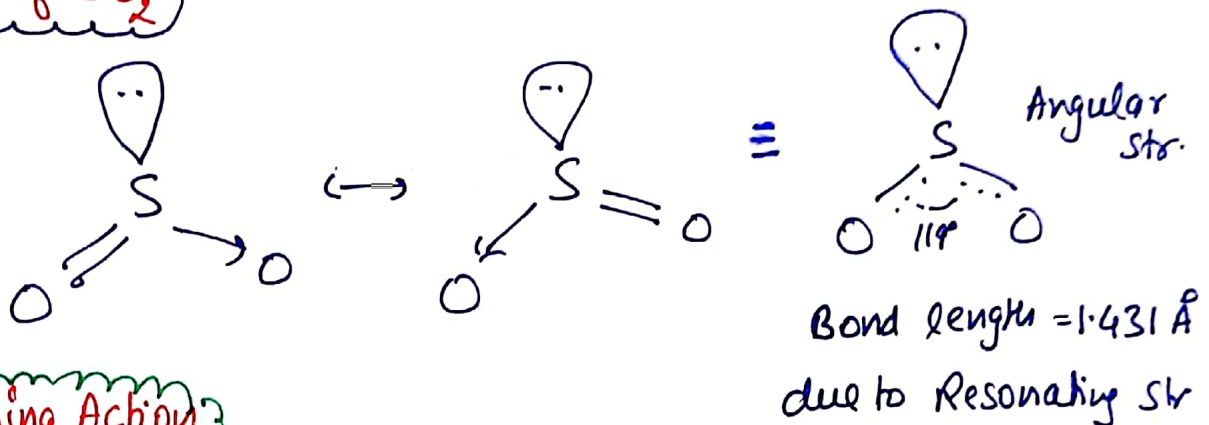
⇒ lead Oxide to lead sulphate
or $2\text{Fe}^{3+} + \text{SO}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{Fe}^{2+} + \text{SO}_4^{2-} + 4\text{H}^+$



⇒ sodium peroxide to sodium sulphate

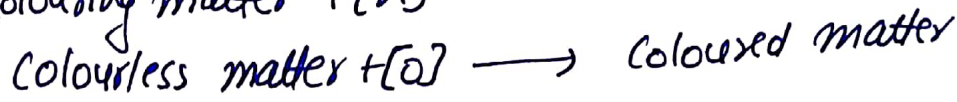
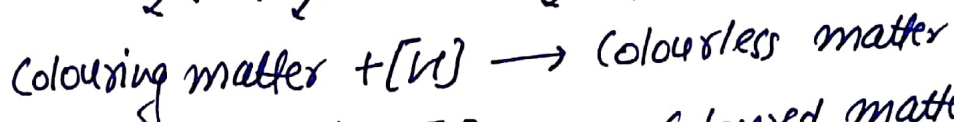
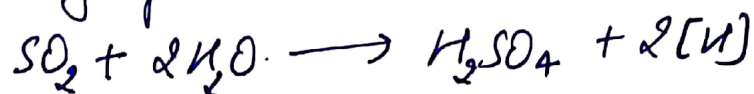


Str. of SO_2



Bleaching Action

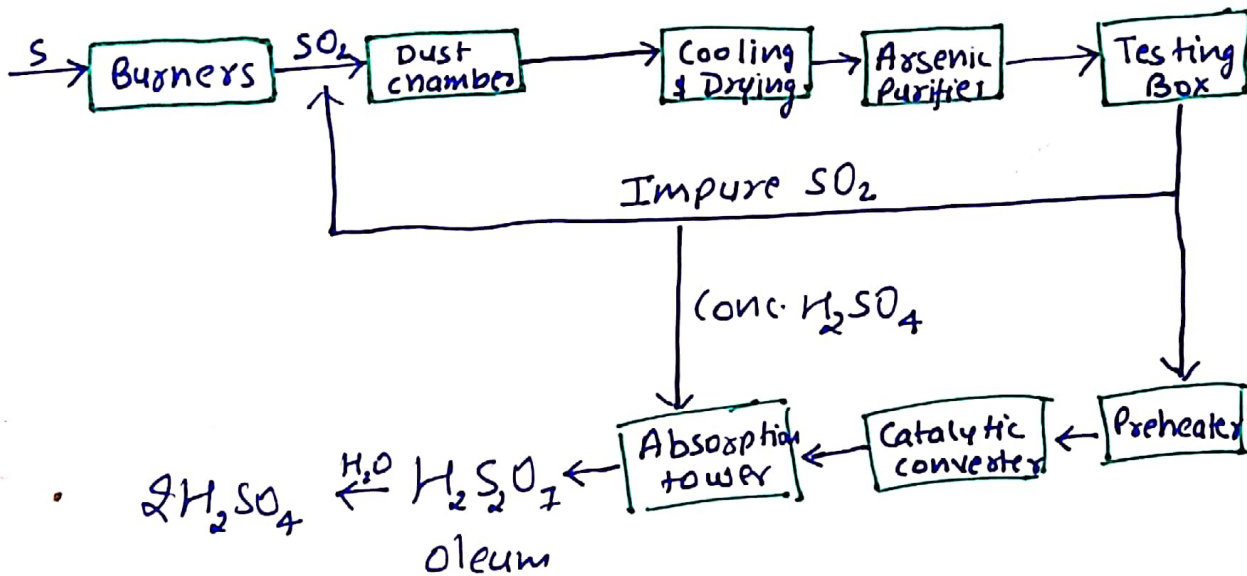
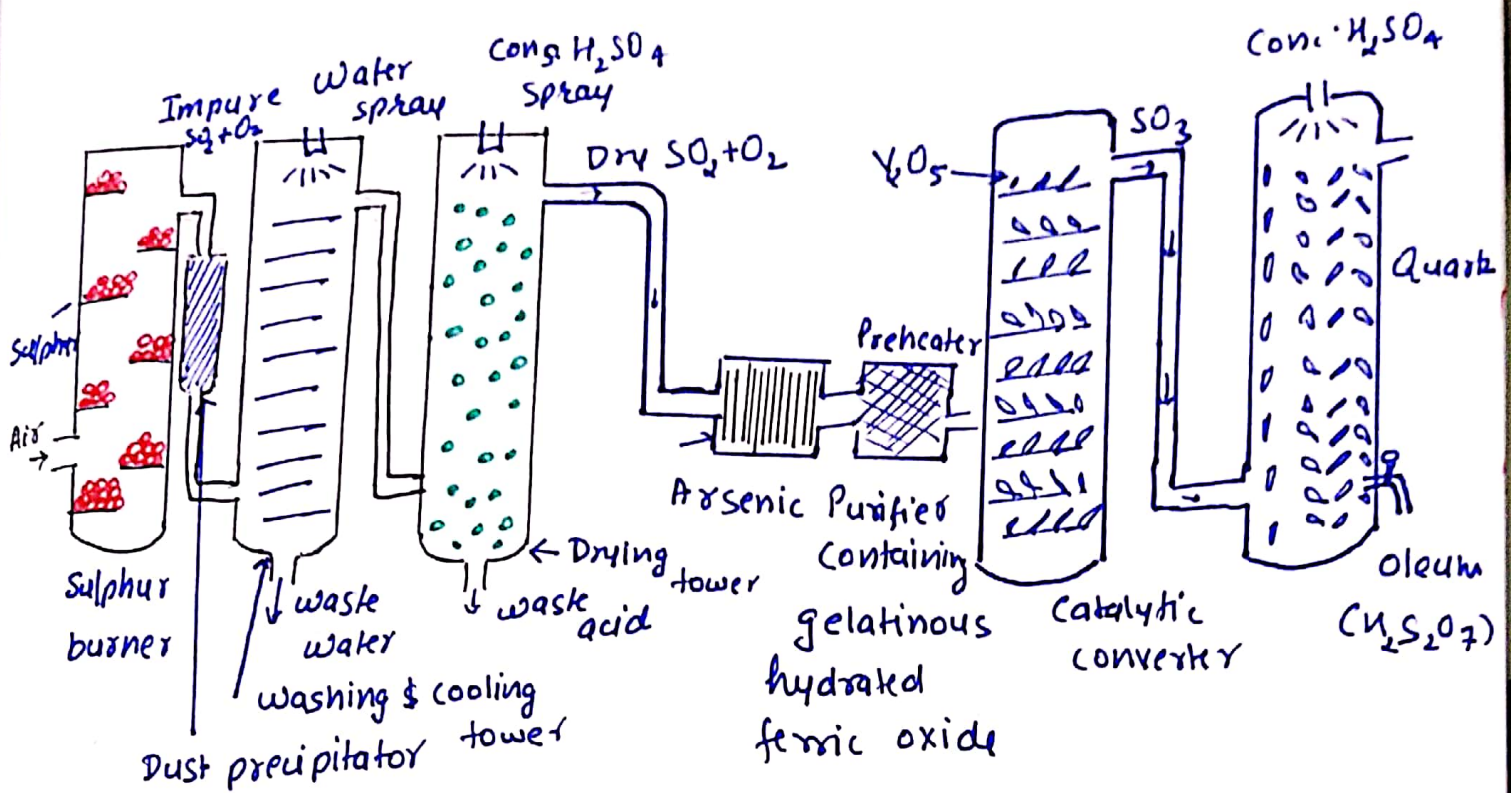
in presence of moisture SO_2 can be used as bleaching agent



Uses

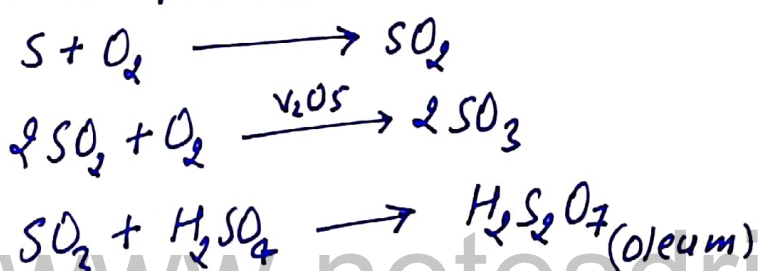
- used for petroleum and sugar refining
- used for bleaching delicate articles such as wool, silk, straw
- used as a germicide and disinfectant
- used for manufacture of sulphuric acid, sodium hydrogen sulphite, calcium hydrogen sulphite
- used as preservative for jams, pickles, jellies and squashes

Sulphuric Acid (H₂SO₄)



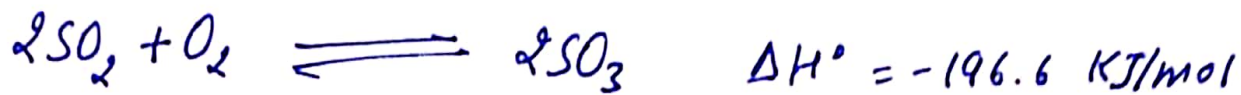
Contact Process

Sulphuric Acid is prepared by contact process



oleum is diluted with water to get H₂SO₄

$$H_2S_2O_7 + H_2O \longrightarrow 2H_2SO_4$$



This step is key step in the manufacture of H_2SO_4
According to le-Chatelier principle

High Pressure

Low pressure will favour the reaction.
A pressure of about 2 bar is used.

Low temperature

Since the forward reaction is exothermic, therefore, low temperature will favour the reaction
Therefore the reaction is carried out at an optimum temperature of 720K.

Use Of a Catalyst

To increase the rate of reaction at low temperature, a catalyst is to be used. V_2O_5 is used because it is cheaper and not easily poisoned.

Purity of gas

To prevent poisoning of catalyst, the gases must be free from impurities of As_2O_3 , dust, moisture

Excess of gases

To have maximum yield of SO_3 , O_2 is used in excess

Working of Plants

⇒ Sulphur burner:

sulphur

Sulphur di oxide is produced by burning sulphur

⇒ Purifying unit:

- Dust remover

Dust is removed either by blowing steam or

- Scrubber or washer

gases with water

Soluble impurities are removed by washing

- Drier

A spray of conc. H_2SO_4 is used for drying of gases

- Arsenic purifier

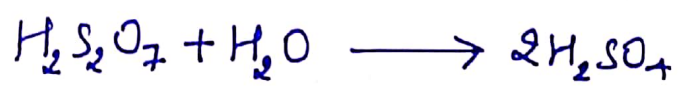
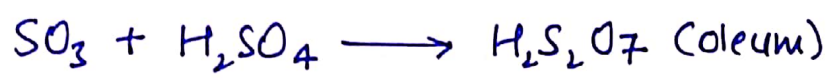
Gelatinous $Fe(OH)_3$ is used to absorb impurities

Catalytic Converter

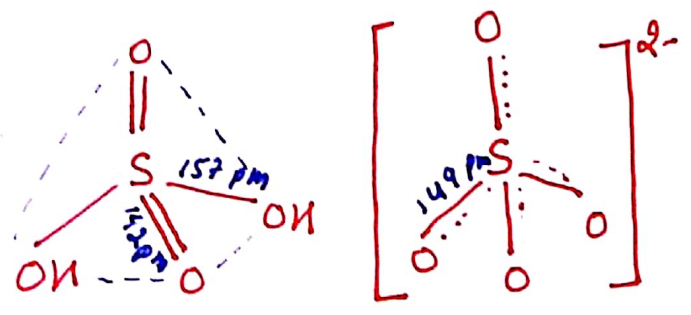
The pure gases coming from the testing box are preheated to 720 K in preheater. The gases are then passed over a catalyst consisting V_2O_5 .

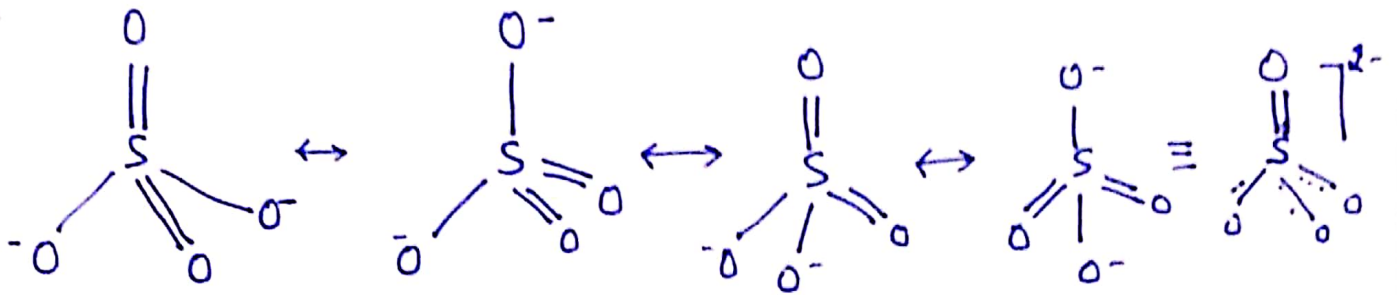
Absorption Tower

SO_3 from the catalytic converter is introduced at the base of absorption tower from top of which H_2SO_4 is showered



Str. of H_2SO_4

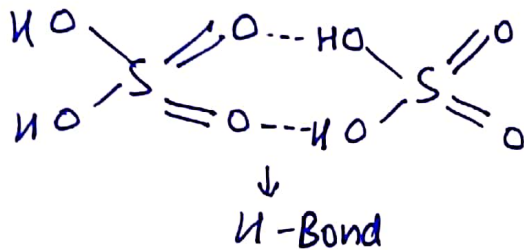




Properties

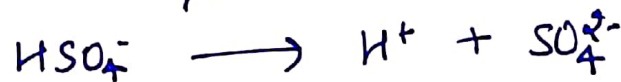
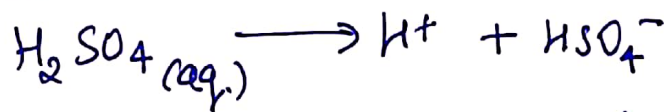
- Pure sulphuric acid is a colourless, dense, viscous liquid (specific gravity of 1.84 at 298K). It is known as oil of vitriol.
- It freezes at 283 K and boils at 611 K

The different molecules are held together by H-bonding

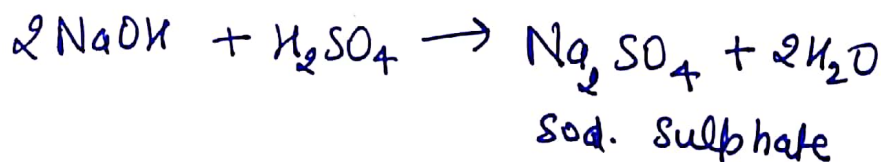
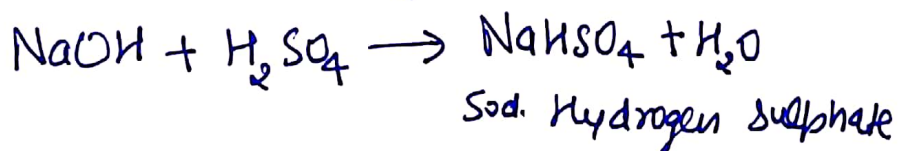


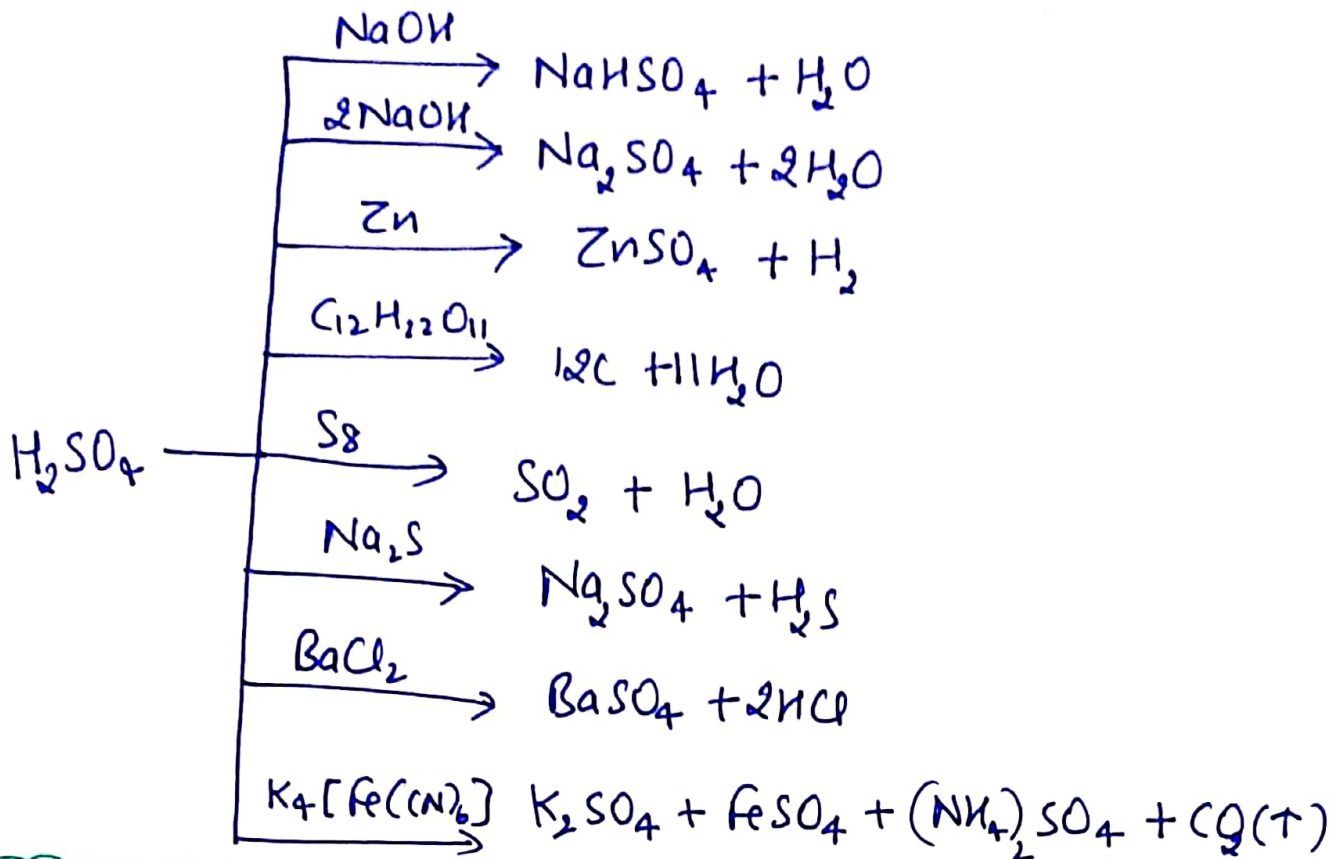
- It dissolves in water with the evolution of a large quantity of heat
- It is highly corrosive and burns on the skin.

Acidic Character



it forms two types of salt





Uses of H_2SO_4

- in the manufacture of HCl , HNO_3 , H_3PO_4 , sulphates and bisulphates.
- in the manufacture of fertilizers like ammonium sulphate.
- In the manufacture of dyes, drugs, paints, detergents.
- Crude petroleum is treated with H_2SO_4 to remove unwanted sulphur.
- In the manufacture of explosive.
- In lead storage batteries as electrolyte.

Group-17 Elements

F - Fluorine	- 9	$2s^2 2p^5$
Cl - Chlorine	- 17	$3s^2 3p^5$
Br - Bromine	- 35	$4s^2 4p^5$
I - Iodine	- 53	$5s^2 5p^5$
At - Astatine	- 85	$6s^2 6p^5$

General Ele. Conf. $\rightarrow ns^2 p^5$

Valence electron $\rightarrow 7$

Also called Halogens
means salt forming

Occurrence

- F is present as insoluble fluorides like fluor spar (CaF_2), Cryolite (Na_3AlF_6) and in small quantities in soil, river water, plants, bones as well as teeth of animals.
- Cl, Br, I are present as chloride, bromide, iodide of sodium, potassium, calcium, magnesium etc. The dried up seas contain sodium chloride and carnalite ($KCl \cdot MgCl_2 \cdot 6H_2O$)
- Iodine is also present in sea weeds (0.5% iodine) and Chile salt petre contains upto 0.2% of sodium iodate

Atomic and ionic radii

Halogens have the smallest atomic radii in their respective periods

i.e. F has smallest in 2nd period

Cl has smallest in 3rd period

Noble gases have larger atomic radii as compared to corresponding elements of G-17

Reason

This is because we measure covalent radius for halogens which is even smaller than their actual radius and we measure Van der Waal radius for noble gases (being atomic in nature) which is larger than their actual radius.

Covalent radius in F-F



Van der Waal radius in Neon



As we move down the group Atomic and ionic radii go on increasing because of increase in number of shells and decrease in effective nuclear charge

Ionization enthalpy

- They have very-very high I.E because of very small size and high effective nuclear charge
- As we move down the gp. their ionization enthalpy goes on decreasing because atomic size increases and effective nuclear charge decreases down the group
- I_2 crystals shine like metals and can even form +ve ions because of low I.E

Electron Gain Enthalpy ($\Delta_{eg} H$)

Halogens have highest electron gain enthalpy in their respective periods because they have small size and high nuclear charge

As we move down the group, electron gain enthalpy decreases

Fluorine has less electron gain enthalpy than chlorine because it has very small size and incoming e^- feels repulsion from valence e^- . So some energy is absorbed to overcome the repulsive forces. And net energy released decreases

Electronegativity

These are the most electronegative elements in their respective periods. Down the group electronegativity decreases. Thus the fluorine is the most electronegative element in the periodic table

Melting and Boiling Point

Increase regularly down the group because of increase in size and nuclear charge resulting in greater van der Waal forces of attraction. Thus F_2 and Cl_2 gas, Br_2 is liquid while I_2 is solid.

Bond Dissociation Enthalpy

Bond dissociation enthalpy decreases from Cl_2 to I_2 . This is due to increase in the size of halogen atom on moving down the group from chlorine to iodine. Bond dissociation enthalpy of fluorine is smaller than chlorine or bromine. This is due to small size of fluorine which results large $e^- - e^-$ repulsion among the lone pairs in F_2 molecule. Thus, the bond dissociation enthalpy decrease in the order $Cl_2 > Br_2 > F_2 > I_2$

Colour

All halogens are coloured. This is due to absorption of radiation in visible region which results in the excitation of outer e^- to higher energy level while the remaining light is transmitted

Oxidation State

Fluorine always show only -1 oxidation state in its ionic as well as covalent compound because it has highest electronegativity and absence of vacant d-orbital.

- Chlorine, bromine and iodine show $-1, +1, +3, +5$ & -7 oxidation state. They show -1 o.s because they can easily gain one e^- to complete their octet whereas +ve oxidation states are due to the presence of vacant d-orbital.

Ground State

	$\uparrow\downarrow$	$\uparrow\downarrow \uparrow\downarrow \uparrow$	$\square \square \square \square$	$-1, +1$
1 st excited state	$\uparrow\downarrow$	$\uparrow\downarrow \uparrow \uparrow$	$\uparrow \square \square \square$	$+3$
2 nd excited state	$\uparrow\downarrow$	$\uparrow \uparrow \uparrow$	$\uparrow \uparrow \square \square$	$+5$
3 rd excited state	\uparrow	$\uparrow \uparrow \uparrow$	$\uparrow \uparrow \uparrow \square$	$+7$

Nature of Bonds

- They have seven valence e^- . So they can easily gain one e^- and form uninegative ions. So, they can form ionic bonds with metal cations eg NaCl, KI, CaCl₂, NaBr etc.
- They can also form covalent bonds with non metals by sharing one e^- as well as with less electropositive metals in higher oxidation states eg F₂, Cl₂, HCl, PCl₅, SF₆ etc.

Oxidising Nature

They have high positive values of reduction potential ($E^0 = +ve$) and behave as strong oxidizing agent because they have high electron gain enthalpy, small bond enthalpies and high enthalpy of hydration



⇒ Reactivity towards Hydrogen (Formation of Hydrides)

directly combine with hydrogen and form their respective hydrogen halides. These are not called hydrides but are called halides because halogens is more electronegative than hydrogen

HF → Hydrogen fluoride HCl → Hydrogen chloride

HBr → Hydrogen bromide HI → Hydrogen iodide

Characteristics of Hydrogen Halides

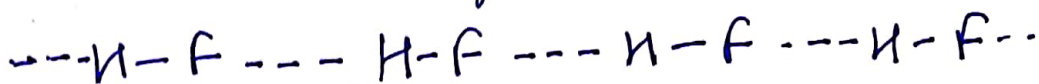
⇒ Solubility

All these are soluble in water and form hydrohalic acid.



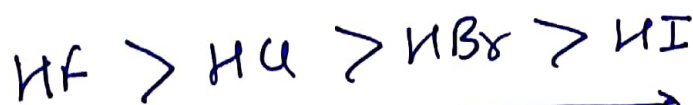
⇒ Physical State

except HF all others are gases at room temperature. HF is liquid because of presence of intermolecular hydrogen bonding



⇒ Thermal Stability

Their thermal stability decreases down the group because as the size of halogen increases, H-X bond length increases & so, bond dissociation energy decreases.



Decreasing order bond dissociation energy

Decreasing order of thermal stability

Reducing character

Their reducing character increases down the group because H-X bond length increases and bond dissociation enthalpy decreases. So the ease of supplying hydrogen becomes more.

- HI behaves as a very strong reducing agent in organic reactions

Acidic Strength

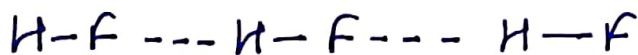
All these behave as acids. Their acidic strength increases down the group due to decrease in their H-X bond dissociation energy



Inc. order of acidic strength \rightarrow

Melting and Boiling Point

HF has the highest b.pt because of presence of intermolecular Hydrogen Bond.



The b.pt increases from HCl to HI down the group because molecular size increases, surface area also increases hence magnitude of Van der Waal's forces of attraction increases and melting and boiling point increases.

Anomalous Behaviour of Fluorine

Fluorine has many different properties from rest of the family members because of

- very small size
- high ionization enthalpy
- high electronegativity
- absence of vacant d-orbitals

Anomalous Behaviour

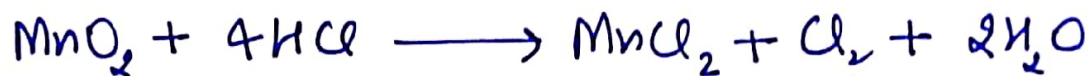
- It has very high +ve reduction potential and so it behaves as a strong oxidising agent
- It forms ionic fluorides
- HF is a liquid with high boiling point than HCl due to Hydrogen Bonding
- Most of the reactions of fluorine are exothermic
- It forms only one oxo acid (HOF)
- It has low F-F bond dissociation energy and so, it is highly reactive
- Shows only -1 oxidation state.
- does not form polyhalide ion like I_3^- , Br_3^- , I_5^- etc

Chlorine

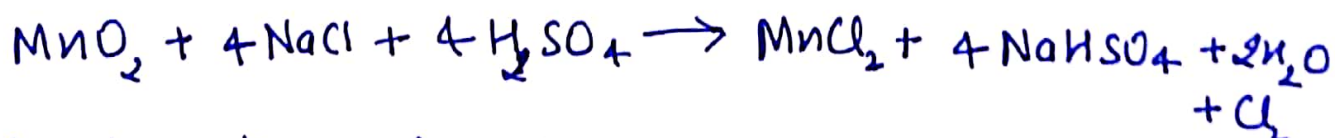
- discovered by Scheele by the action of HCl on MnO_2
- Davy suggested the name chlorine on account of its colour (chloros \rightarrow Greenish yellow)

Preparation

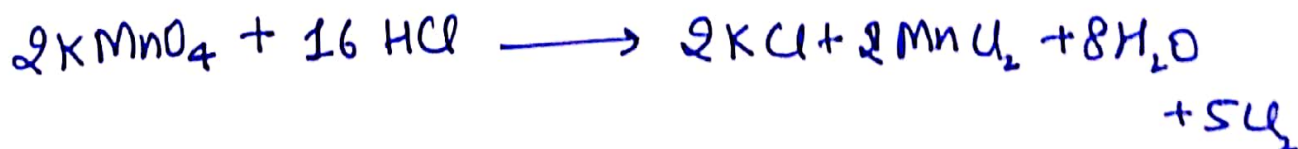
- By heating MnO_2 with HCl



- However a mixture of common salt and conc. H_2SO_4 is used in place of HCl



- By the action of HCl on $KMnO_4$

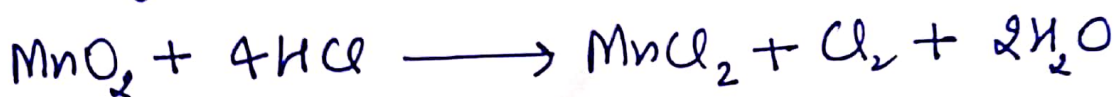


Chlorine

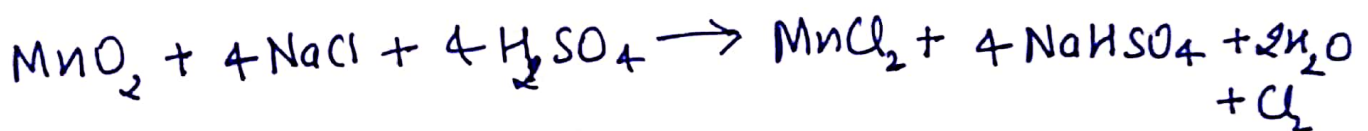
- discovered by Scheele by the action of HCl on MnO_2
- Davy suggested the name chlorine on account of its colour (chloros \rightarrow Greenish yellow)

Preparation

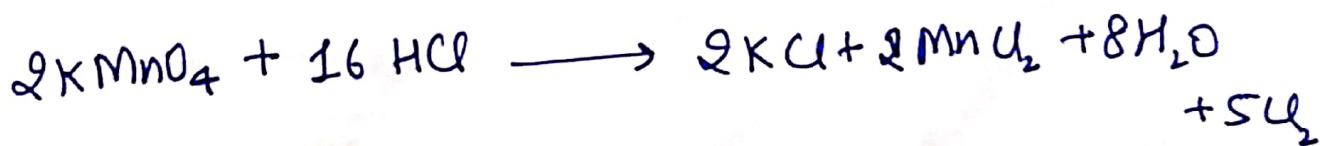
- By heating MnO_2 with HCl



- However a mixture of common salt and conc. H_2SO_4 is used in place of HCl



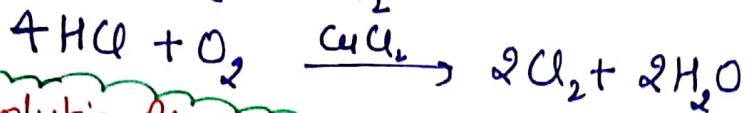
- By the action of HCl on $KMnO_4$



Industrial Preparation

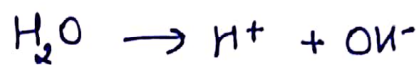
Deacon's Process

HCl is oxidised in atmospheric oxygen in presence of CuCl_2

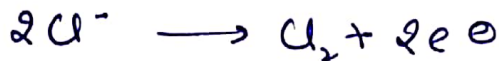


Electrolytic Process

When electrolysis of conc. brine solution is done, then Cl_2 gas is liberated at anode



At Anode



At Cathode



Properties

- It is greenish yellow with pungent odour
- 2-5 times heavier than air
- It is soluble in water
- reacts with metal and non metal to form chlorides



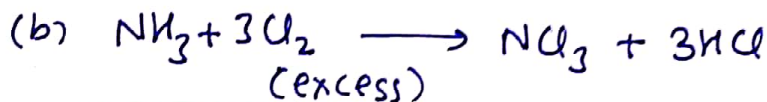
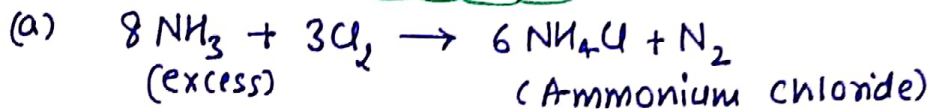
Uses

- It is used for bleaching wood pulp, cotton and textiles
- it is used in extraction of gold and platinum
- it is used in manufacture of CCl_4 , DDT, CuCl_2 & refrigerants
- it is used in sterilizing drinking water
- it is used in preparation of poisonous gases like
 - phosgene (COCl_2)
 - mustard gas $\text{C}_6\text{H}_5\text{N}(\text{CH}_2\text{CH}_2\text{SCH}_2\text{CH}_2\text{Cl})$
 - tear gas (CCl_3NO_2)

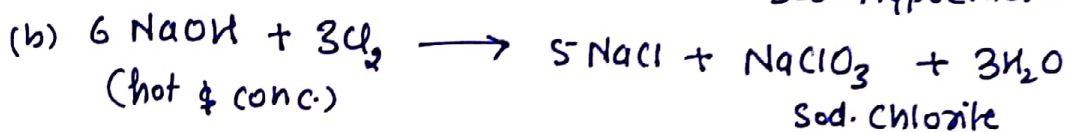
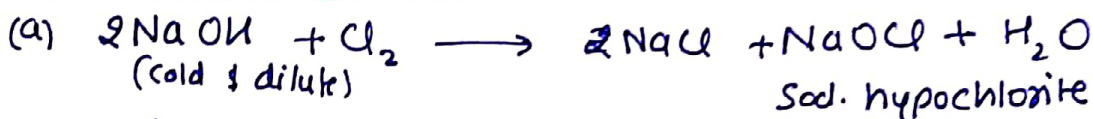
- It reacts with compounds having hydrogen



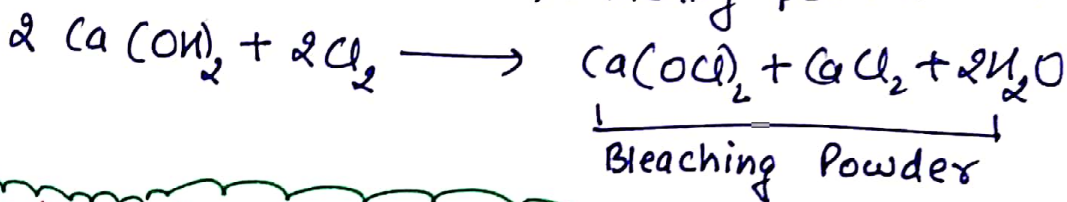
- Reaction with Ammonia



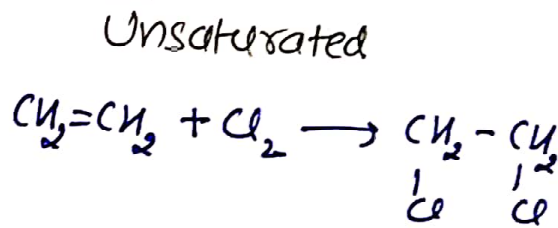
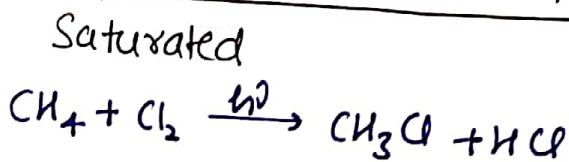
- Reaction with Alkali



(c) with dry slaked lime, bleaching powder is formed



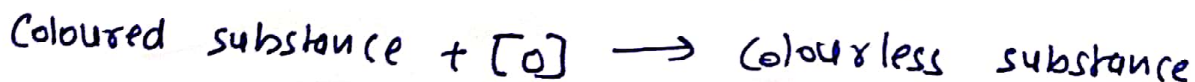
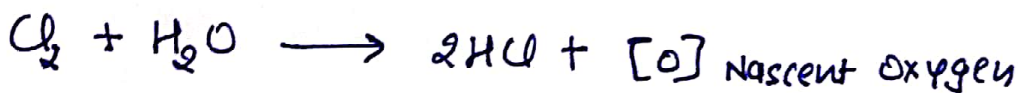
Reaction with Hydrocarbons



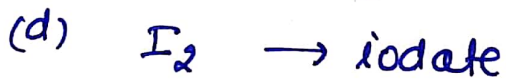
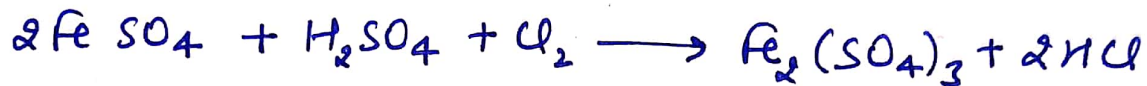
Oxidising and Bleaching Properties

Chlorine on standing loses its yellow colour due to the formation of HCl and HOCl. The HOCl formed gives nascent oxygen which is responsible for oxidizing as well as bleaching properties of chlorine

Bleaching Action



Oxidising Properties



Inter Halogen Compounds

The compounds of one halogen with other halogens are called interhalogen compounds. Among these compounds, the central atom is less electronegative than the surrounding halogens. They form compounds in +1, +3, +5 and +7 oxidation state.

They can be given general composition as XX_3 , XX_5 , XX_7 .

X → halogen of larger size

X' → halogen of smaller size

X is more electronegative than X'

Note

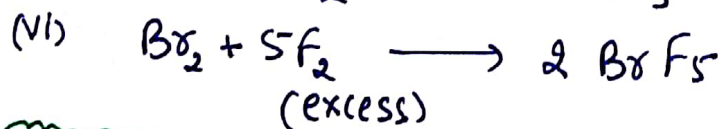
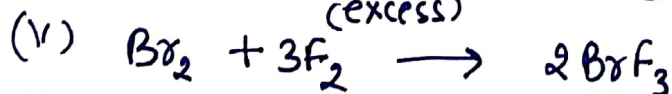
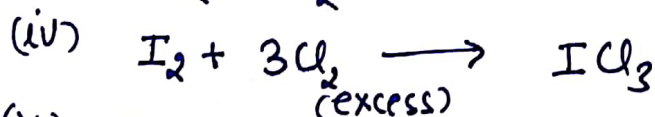
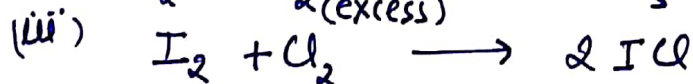
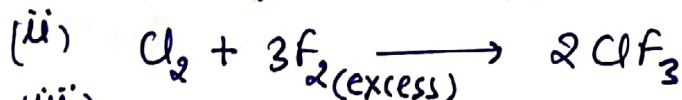
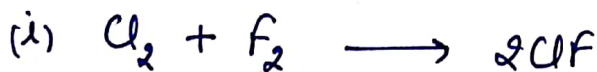
As the ratio b/w X and X' increases, the number of atoms per molecule increases.

e.g. radius ratio of iodine to fluorine is maximum

So, it forms IF_7

Preparation

Prepared by direct combination of Halogens



Properties

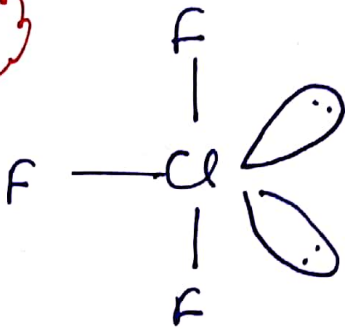
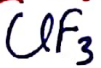
- All these molecules are covalent & diamagnetic
- They are volatile solids or liquids except ClF which is a gas
- Their m.p.t & B.P.T are little higher than the halogens

• Interhalogen compounds are more reactive than halogens (except fluorine) because X-X' bond in interhalogens is weaker than X-X bond in halogen except F-F bond. The overlapping of orbitals of X-X' is ineffective due to large difference in their sizes. So, X-X' are weaker and thus are more reactive than halogens.

• These easily hydrolyse as



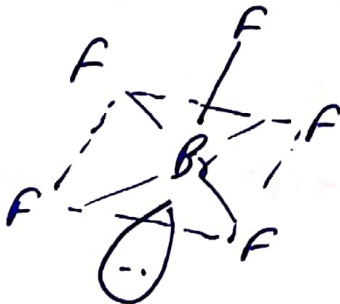
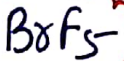
Str. of Interhalogen Compounds



sp³d - Hybridisation

Geometry - Trigonal Bipyramidal

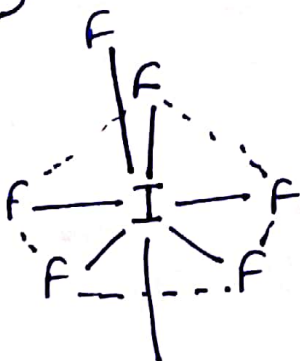
Shape - T-shape



Hybridisation - sp³d²

Geometry - octahedral

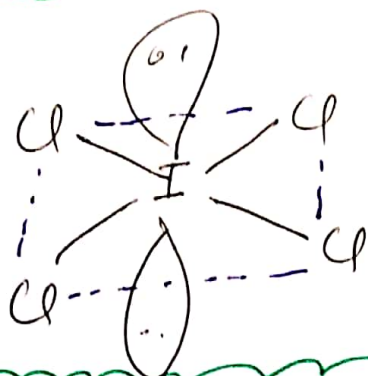
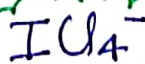
Shape - Square pyramidal



Hybridisation - sp³d³

Geometry - Pentagonal Bipyramidal

Shape -



Hybridisation $\leftarrow sp^3d^2$

shape \leftarrow Octahedral

Geometry \leftarrow Octahedral

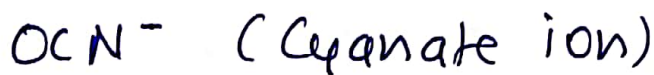
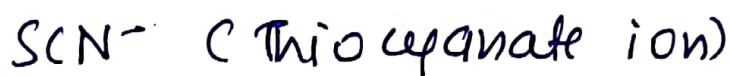
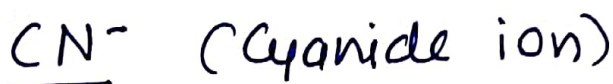
Uses of Interhalogen Compounds

- used as non-aqueous solvents
- used as fluorinating agents.

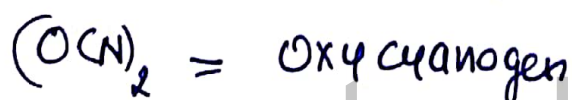
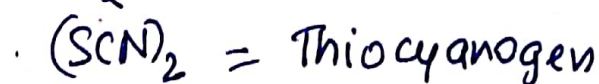
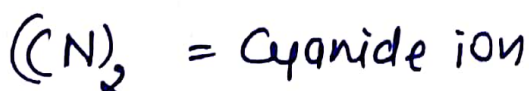


Pseudohalide Ions and Pseudo Halogens

Some ions having two or more atoms in which at least one atom is nitrogen and having properties similar to halide ions.



The dimers of pseudohalide ions behave as halogens and thus are called pseudo halogen



Group-18 (Noble Gases)

He - Helium	-	2	-	1s ²
Ne - Neon	-	10	-	2s ² 2p ⁶
Ar - Argon	-	18	-	3s ² 3p ⁶
Kr - Krypton	-	36	-	4s ² 4p ⁶
Xe - Xenon	-	54	-	5s ² 5p ⁶
Rn - Radon	-	86	-	6s ² 6p ⁶

Also known as

Noble Gases, inert gases
rare gases

General Conf. - ns² np⁶
(except He)

Atomic and Physical Properties

- **Atomic Radii**
Atomic radii increases down the group with increase in atomic number.
Their atomic radii is higher than ~~not~~ halogens because noble gases ~~do~~ exist in monoatomic form due to stable configuration as a result their van der Waal radii is calculated which is larger than atomic radii.
- **Ionisation enthalpy**
They have the highest I.E in their respective periods because of fully filled orbitals and hence greater stability. Down the

group, it decreases due to increase in atomic size

Electron Gain enthalpy

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

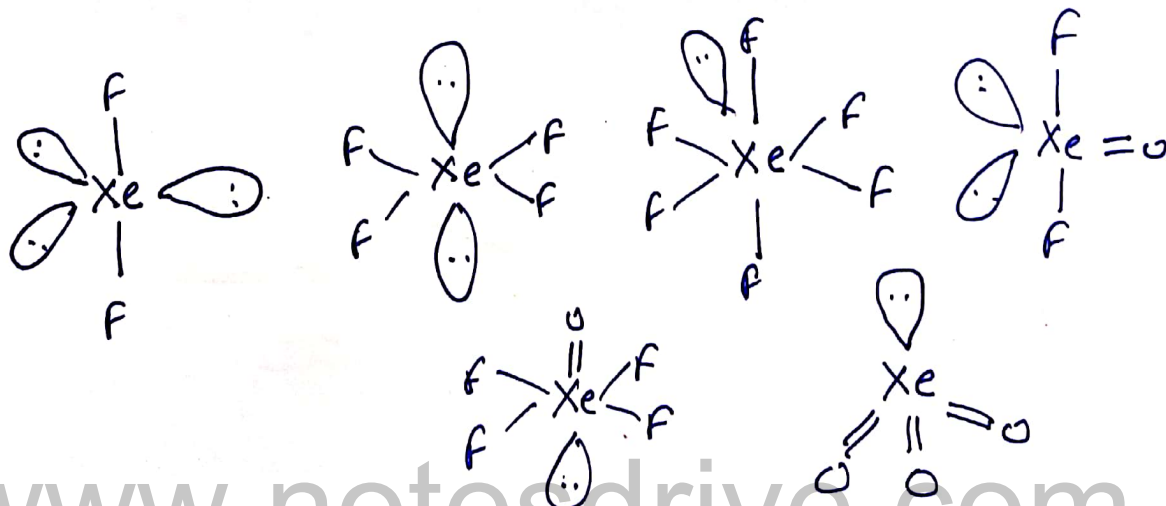
Melting and Boiling Point

They have low m.p.t and b.p.t because weak van der waal forces are present b/w their atoms. Down the group m.p.t and b.p.t increases ~~and~~ because of increase in atomic size and hence increase of van der waal forces. He has the lowest B.P.t of any known substance

Liquefaction

Because of weak van der waal's forces present in them, it is difficult to liquefy them. Down the group, the ease of liquefaction increases because of increase of at size and hence increase of van der waal's forces

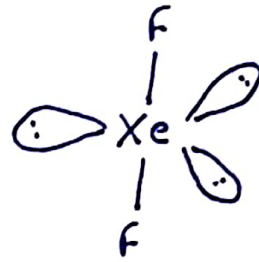
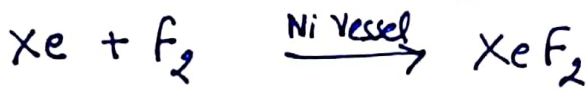
Structure of Xenon Compounds



Compounds of Xenon

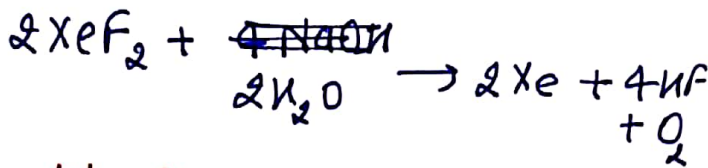
A. Xenon fluorides

1.) Xenon difluoride, XeF_2

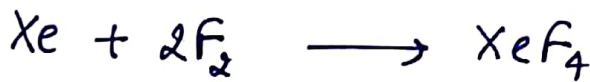


linear str.
 sp^3d hybridisation
 trigonal bipyramidal geometry

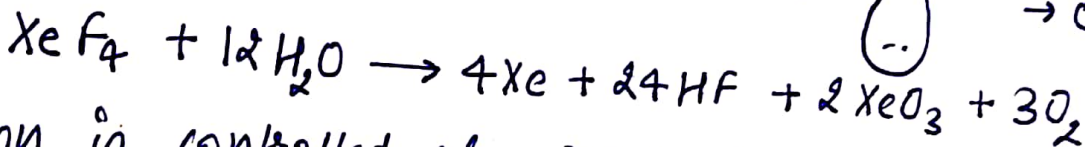
Reaction with water



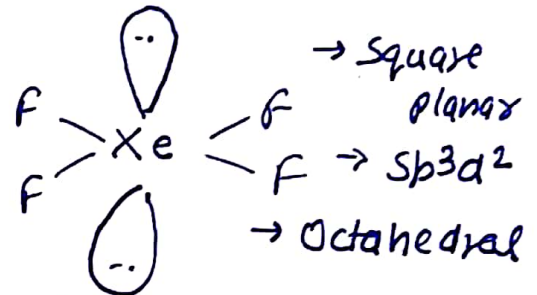
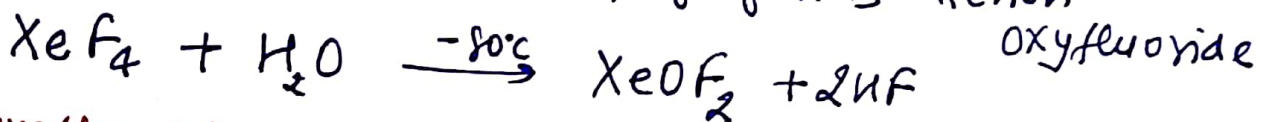
2.) Xenon tetrafluoride, XeF_4



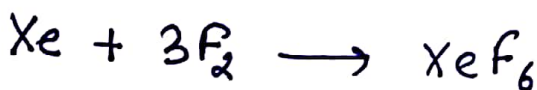
Reaction with water



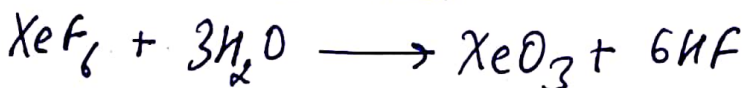
reaction is controlled at $-80^\circ C$, it forms xenon



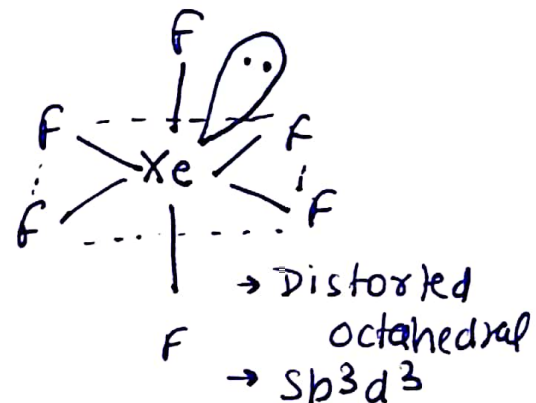
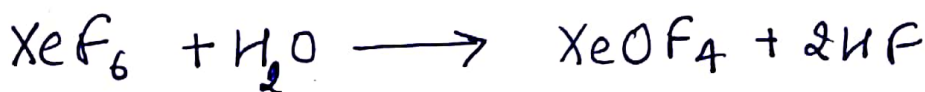
3.) Xenon hexafluoride, XeF_6



Reaction with water



in Ni vessel



Oxoacids of Halogens

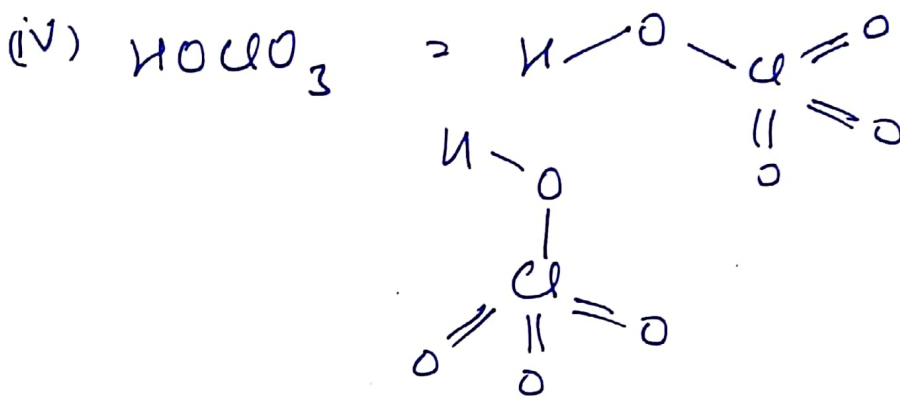
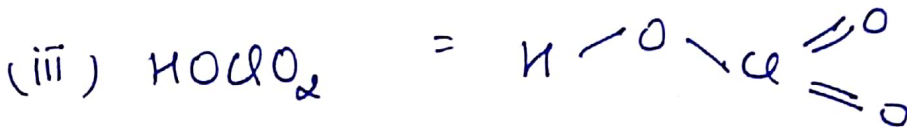
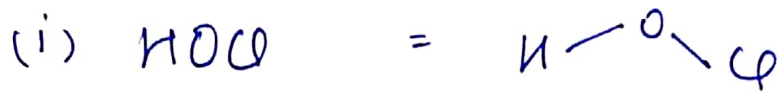
Fluorine	Chlorine	Bromine	Iodine
HFO Hypofluorous Acid	HClO Hypochlorous Acid	HBrO Hypobromous Acid	HIO Hypoiodous Acid
	HClO_2 Chlorous Acid	—	—
	HClO_3 Chloric Acid	HBrO_3 Bromic Acid	HIO_3 Iodic Acid
	HClO_4 Perchloric Acid	HBrO_4 Perbromic Acid	HIO_4 Periodic Acid

Fluorine only form one oxoacid (HOF)

because fluorine has small size and high electronegativity

Other halogens form large no. of oxoacids in +1, +3, +5 and +7 o.s

Str.

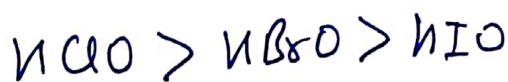


Acidic strength

(i) For same halogen, acidic strength increases with increase in oxidation state of halogen



(ii) Acidic strength of same type of acid of diff. halogen

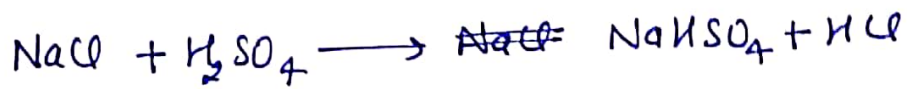


As we move down the group, E.N of halogens decreases and acidic strength also decreases.

Hydrogen Chloride

Preparation

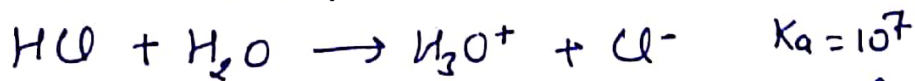
It is prepared in lab by heating sodium chloride with conc. sulphuric acid



HCl formed is dried by passing through conc. H_2SO_4

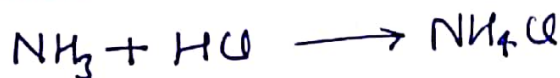
Properties

- It is colourless pungent smelling gas
- It is extremely soluble in water & ionize as

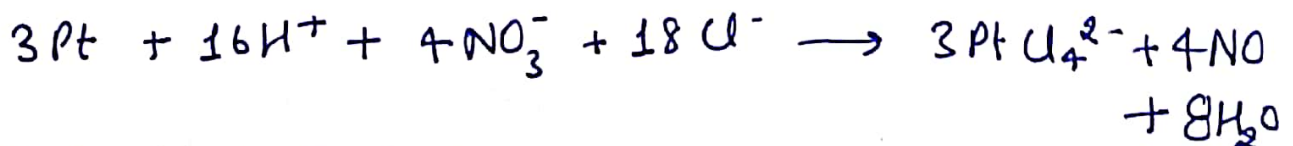
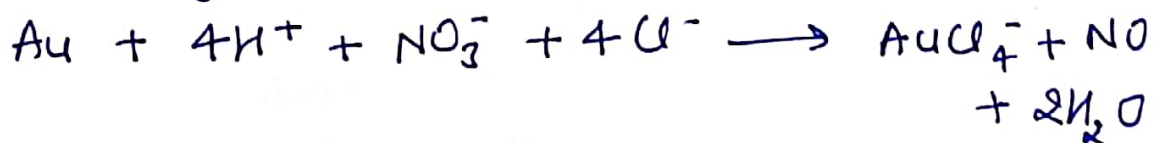


It is a very strong acid. So its K_a value is very High.

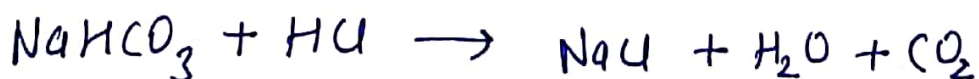
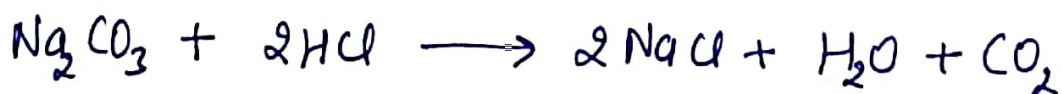
Reaction with Ammonia



- Conc. HCl & conc. HNO_3 when mixed in 3:1 ratio then it is called aqua regia. which is used for dissolving noble metals like Gold and platinum



- Hydrochloric Acid decomposes salts of weaker acids



Uses

- It is used in the manufacturing of chlorine, NH_4Cl and glucose
- It is used for extracting glue from bones and purifying bone black
- It is used in medicine and as a lab. reagent