

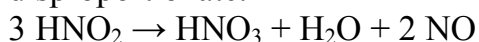
1. **p-Block elements:** Elements belonging to groups 13 to 18 of the periodic table are called p-block elements.
2. **General electronic configuration of p-block elements:** The p-block elements are characterized by the  $ns^2 np^{1-6}$  valence shell electronic configuration.
3. **Representative elements:** Elements belonging to the s and p-blocks in the periodic table are called the representative elements or main group elements.
4. **Inert pair effect:** The tendency of  $ns^2$  electron pair to participate in bond formation decreases with the increase in atomic size. Within a group the higher oxidation state becomes less stable with respect to the lower oxidation state as the atomic number increases. This trend is called 'inert pair effect'. In other words, the energy required to unpair the electrons is more than energy released in the formation of two additional bonds.
5. **Nitrogen family:** The elements of group 15 – nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb) and bismuth (Bi) belong to configuration is  $ns^2 np^3$ .
6. **Oxygen family:** Group 16 of periodic table consists of five elements – oxygen (O), sulphur (S), selenium (Se), tellurium (Te) and polonium (Po). Their general electronic configuration is  $ns^2 np^4$ .
7. **The halogen family:** Group 17 elements, fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At), belong to halogen family. Their general electronic configuration is  $ns^2 np^5$ .
8. **Group 18 elements:** Helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn) are Group 18 elements. They are also called noble gases. Their general electronic configuration is  $ns^2 np^6$  except helium which has electronic configuration  $1s^2$ . They are called noble gases because they show very low chemical reactivity.

### GROUP 15 ELEMENTS

9. **Atomic and ionic radii:** Covalent and ionic radii increase down the group. There is appreciable increase in covalent radii from N to P. There is small increase from As to Bi due to presence of completely filled d or f orbitals in heavy elements.

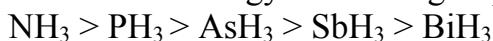
10. **Ionisation energy:** It goes on decreasing down the group due to increase in atomic size. Group 15 elements have higher ionisation energy than group 14 elements due to smaller size of group 15 elements. Group 15 elements have higher ionization energy than group 16 elements because they have stable electronic configuration i.e., half filled p-orbitals.

11. **Allotropy:** All elements of Group 15 except nitrogen show allotropy. Catenation: Nitrogen shows catenation to some extent due to triple bond but phosphorus shows catenation to maximum extent. The tendency to show catenation decreases down the group. 13. Oxidation states: The common oxidation states are +3, +5, –3. The tendency shows –3 oxidation state decreases down the group due to decrease in electronegativity which is due to increase in atomic size. The stability of +5 oxidation state decreases whereas stability of +3 oxidation state increases due to inert pair effect. Nitrogen shows oxidation states from –3 to +5. Nitrogen and phosphorus with oxidation states from +1 to +4 undergo oxidation as well as reduction in acidic medium. This process is called disproportionation.

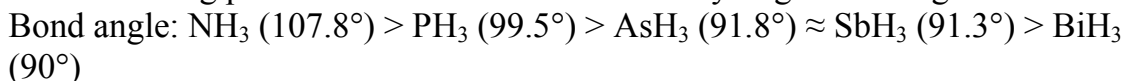


12. **Reactivity towards hydrogen:** All group 15 elements form trihydrides,  $\text{MH}_3$ . Hybridisation -  $\text{sp}^3$

The stability of hydrides decrease down the group due to decrease in bond dissociation energy down the group.



Boiling point increases with increase in size due to increase in van der Waals forces. Boiling point of  $\text{NH}_3$  is more because of hydrogen bonding.



Electronegativity of N is highest. Therefore, the lone pairs will be towards nitrogen and hence more repulsion between bond pairs.

Therefore bond angle is the highest. After nitrogen, the electronegativity decreases down the group.

Basicity decreases as  $\text{NH}_3 > \text{PH}_3 > \text{AsH}_3 > \text{SbH}_3 < \text{BiH}_3$ .

This is because the lone pair of electrons are concentrated more on nitrogen and hence the basicity will be maximum in the case of  $\text{NH}_3$ . It will decrease down the group as the electronegativity decreases down the group.

The reducing power of hydrides increases down the group due to decrease in bond dissociation energy down the group.

13. **Reactivity towards oxygen:** All group 15 elements form trioxides ( $M_2O_3$ ) and pentoxides ( $M_2O_5$ ).

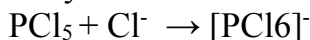
Acidic character of oxides decreases and basicity increases down the group. This is because the size of nitrogen is very small. It has a strong positive field in a very small area. Therefore, it attracts the electrons of water's O-H bond to itself and release  $H^+$  ions easily. As we move down the group, the atomic size increases. Hence, the acidic character of oxides decreases and basicity increases as we move down the group.

14. **Reactivity towards halogen:** Group 15 elements form trihalides and pentahalides.

Trihalides – covalent compounds and become ionic down the group.  $sp^3$  hybridisation, pyramidal shape

Pentahalides -  $sp^3d$  hybridisation, TBP shape

They are Lewis acids because of the presence of vacant d – orbitals.



$PCl_5$  is ionic in solid state and exist as  $[PCl_4]^+ [PCl_6]^-$

In  $PCl_5$ , there are three equatorial bonds and two axial bonds.

The axial bonds are longer than equatorial bonds because of greater repulsion from equatorial bonds.

Nitrogen does not form pentahalides due to absence of d- orbitals.

15. **Reactivity towards metals:** All elements react with metals to form binary compounds in  $-3$  oxidation state.

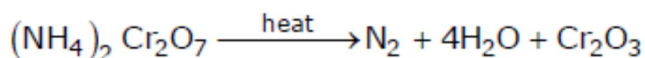
16. **Anomalous behaviour of nitrogen:** The behaviour of nitrogen differs from rest of the elements.

Reason:

- i. It has a small size.
- ii. It does not have d – orbitals
- iii. It has high electronegativity
- iv. It has high ionization enthalpy

17. **Dinitrogen:**

Preparation:



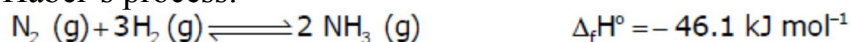
Properties:

It is a colourless, odourless, tasteless and non – toxic gas. It is chemically un-reactive at ordinary temperature due to triple bond in  $N \equiv N$  which has high bond dissociation energy.

18. **Ammonia:** Ammonia molecule is trigonal pyramidal with nitrogen atom at the apex. It has 3 bond pairs and 1 lone pair. N is  $sp^3$  hybridised.

Preparation:

Haber's process:



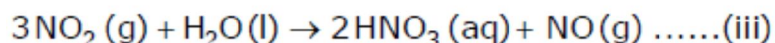
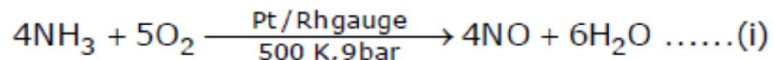
Pressure =  $200 \times 10^5$  Pa

Temperature = 773 K

Catalyst is FeO with small amounts of  $K_2O$  and  $Al_2O_3$

19. **Nitric Acid:**

a. Ostwald Process:



NO thus formed is recycled and the aqueous  $HNO_3$  can be concentrated by distillation upto ~ 68% by mass. Further concentration to 98% can be achieved by dehydration with concentrated  $H_2SO_4$ .

Nitric acid is strong oxidizing agent in the concentrated as well as in the dilute state.

20. **Phosphorus:**

a. It shows the property of catenation to maximum extent due to most stable P – P bond.

b. It has many allotropes, the important ones are :

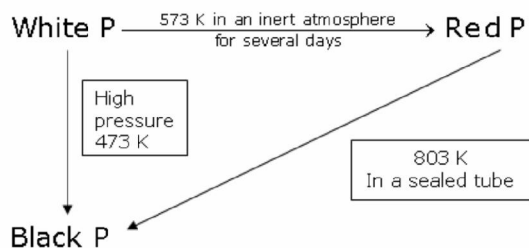
i. White phosphorus

ii. Red phosphorus

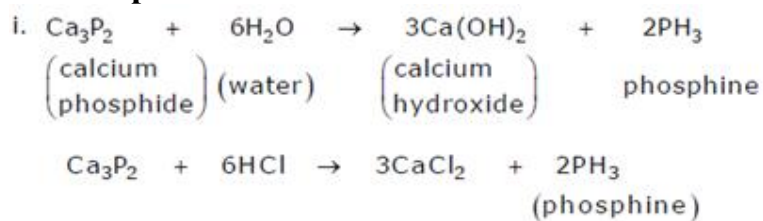
## iii. Black phosphorus

White phosphorus	Red phosphorus	Black phosphorus
Discrete tetrahedral P <sub>4</sub> molecules	Polymeric structure consisting of chains of P <sub>4</sub> units linked together	Exists in two forms - α black P and β black P
Very reactive	Less reactive than white P	Very less reactive
Glow in dark	Does not glow in dark	-
Translucent waxy solid	Has an iron grey luster	Has an opaque monoclinic or rhombohedral crystals
Soluble in CS <sub>2</sub> but insoluble in water	Insoluble in water as well as CS <sub>2</sub>	-
It has low ignition temperature, therefore, kept under water	-	-

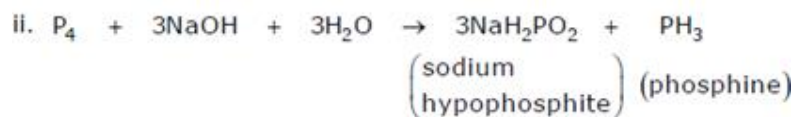
## Preparation:



## 21. Phosphine:



Phosphine is highly poisonous, colourless gas and has a smell of rotten fish.



## 22. Chlorides of Phosphorous:

$\text{PCl}_3$	$\text{PCl}_5$
Colourless oily liquid	Yellowish white powder
$\text{P}_4 + 6\text{Cl}_2 \rightarrow 4\text{PCl}_3$ $\text{P}_4 + 8\text{SOCl}_2 \rightarrow 4\text{PCl}_3 + 4\text{SO}_2 + 2\text{S}_2\text{Cl}_2$	$\text{P}_4 + 10\text{Cl}_2 \rightarrow 4\text{PCl}_5$ $\text{P}_4 + 10\text{SO}_2\text{Cl}_2 \rightarrow 4\text{PCl}_5 + 10\text{SO}_2$
Is hydrolysed in the presence of moisture $\text{PCl}_3 + 3\text{H}_2\text{O} \longrightarrow \text{H}_3\text{PO}_3 + 3\text{HCl}$	$\text{PCl}_5 + \text{H}_2\text{O} \longrightarrow \text{POCl}_3 + 2\text{HCl}$ $\text{POCl}_3 + 3\text{H}_2\text{O} \longrightarrow \text{H}_3\text{PO}_4 + 3\text{HCl}$
$3\text{CH}_3\text{COOH} + \text{PCl}_3 \rightarrow 3\text{CH}_3\text{COCl} + \text{H}_3\text{PO}_3$	$3\text{CH}_3\text{COOH} + \text{PCl}_5 \rightarrow \text{CH}_3\text{COCl} + \text{POCl}_3 + \text{HCl}$
$3\text{C}_2\text{H}_5\text{OH} + \text{PCl}_3 \rightarrow 3\text{C}_2\text{H}_5\text{Cl} + \text{H}_3\text{PO}_3$	$\text{C}_2\text{H}_5\text{OH} + \text{PCl}_5 \rightarrow \text{C}_2\text{H}_5\text{Cl} + \text{POCl}_3 + \text{HCl}$
Pyramidal shape, $\text{sp}^3$ hybridisation	TBP geometry, $\text{sp}^3\text{d}$ hybridisation
-	$2\text{Ag} + \text{PCl}_5 \rightarrow 2\text{AgCl} + \text{PCl}_3$ $\text{Sn} + 2\text{PCl}_5 \rightarrow \text{SnCl}_4 + 2\text{PCl}_3$

## GROUP 16 ELEMENTS

23. **Oxidation states:** They show -2, +2, +4, +6 oxidation states. Oxygen does not show +6 oxidation state due to absence of d – orbitals. Po does not show +6 oxidation state due to inert pair effect.

The stability of -2 oxidation state decreases down the group due to increase in atomic size and decrease in electronegativity.

Oxygen shows -2 oxidation state in general except in  $\text{OF}_2$  and  $\text{O}_2\text{F}_2$

The stability of +6 oxidation state decreases and +4 oxidation state increases due to inert pair effect.

24. **Ionisation enthalpy:** Ionisation enthalpy of elements of group 16 is lower than group 15 due to half filled p-orbitals in group 15 which are more stable. However, ionization enthalpy decreases down the group.

25. **Electron gain enthalpy:** Oxygen has less negative electron gain enthalpy than S because of small size of O. From S to Po electron gain enthalpy becomes less negative to Po because of increase in atomic size.

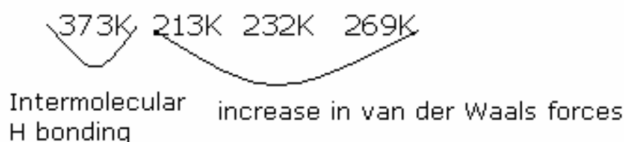
26. **Melting and boiling point:** It increases with increase in atomic number. Oxygen has much lower melting and boiling points than sulphur because oxygen is diatomic ( $O_2$ ) and sulphur is octatomic ( $S_8$ ).

27. **Reactivity with hydrogen:**

All group 16 elements form hydrides.

Bent shape

Bond angle:  $H_2O > H_2S < H_2Se < H_2Te$



Acidic nature:  $H_2O < H_2S < H_2Se < H_2Te$

This is because the H-E bond length increases down the group. Therefore, the bond dissociation enthalpy decreases down the group.

Thermal stability:  $H_2O < H_2S < H_2Se < H_2Te < H_2Po$

This is because the H-E bond length increases down the group. Therefore, the bond dissociation enthalpy decreases down the group.

Reducing character:  $H_2O < H_2S < H_2Se < H_2Te < H_2Po$

This is because the H-E bond length increases down the group. Therefore, the bond dissociation enthalpy decreases down the group.

28. **Reactivity with oxygen:**  $EO_2$  and  $EO_3$  Reducing character of dioxides decreases down the group because oxygen has a strong positive field which attracts the hydroxyl group and removal of  $H^+$  becomes easy.

Acidity also decreases down the group.

$SO_2$  is a gas whereas  $SeO_2$  is solid. This is because  $SeO_2$  has a chain polymeric structure whereas  $SO_2$  forms discrete units.

29. **Reactivity with halogens:**  $EX_2$ ,  $EX_4$  and  $EX_6$  The stability of halides decreases in the order  $F^- > Cl^- > Br^- > I^-$ . This is because E-X bond length increases with increase in size.

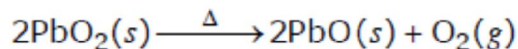
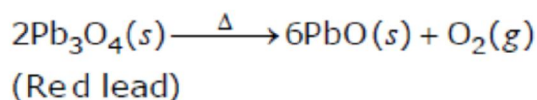
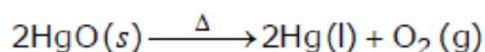
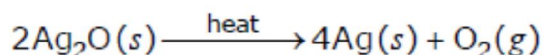
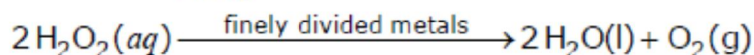
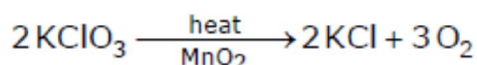
Among hexa halides, fluorides are the most stable because of steric reasons.

Dihalides are  $sp^3$  hybridised, are tetrahedral in shape.

Hexafluorides are only stable halides which are gaseous and have  $sp^3d^2$  hybridisation and octahedral structure.

$H_2O$  is a liquid while  $H_2S$  is a gas. This is because strong hydrogen bonding is present in water. This is due to small size and high electronegativity of O.

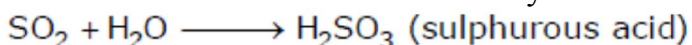
30. **Oxygen:** Preparation:



31. **Oxides:** The compounds of oxygen and other elements are called oxides.

Types of oxides:

a. Acidic oxides: Non-metallic oxides are usually acidic in nature.



b. Basic oxides: Metallic oxides are mostly basic in nature. Basic oxides dissolve in water forming bases e.g.,



c. Amphoteric oxides: They show characteristics of both acidic as well as basic



oxides.  $Al_2O_3 + 6NaOH(aq) + 3H_2O(l) \longrightarrow 2Na_3[Al(OH)_6](aq)$

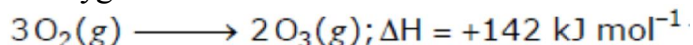


d. Neutral oxides: These oxides are neither acidic nor basic.

Example: CO, NO and N<sub>2</sub>O

32. **Ozone:** Preparation:

i. It is prepared by passing silent electric discharge through pure and dry oxygen  
10 – 15 % oxygen is converted to ozone.



Structure of Ozone: Ozone has angular structure. Both O = O bonds are of equal bond length due to resonance.

33. **Sulphur:**

Sulphur exhibits allotropy:

a. Yellow Rhombic ( $\alpha$  - sulphur):

b. Monoclinic ( $\beta$ - sulphur):



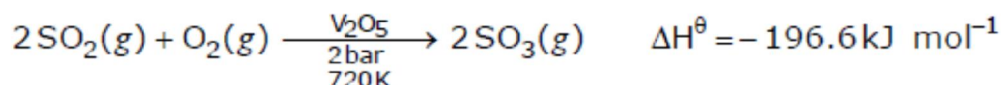
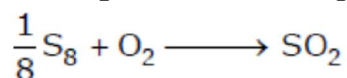
At 369 K both forms are stable. It is called transition temperature.

Both of them have S<sub>8</sub> molecules. The ring is puckered and has a crown shape.

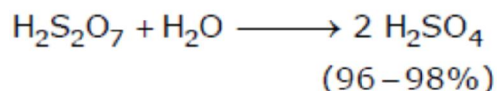
Another allotrope of sulphur – cyclo S<sub>6</sub> ring adopts a chair form.

S<sub>2</sub> is formed at high temperature (~ 1000 K). It is paramagnetic because of 2 unpaired electrons present in anti bonding  $\pi^*$  orbitals like O<sub>2</sub>

34. **Sulphuric acid:** Preparation: By contact process



Exothermic reaction and therefore low temperature and high pressure are favorable



It is dibasic acid or diprotic acid. It is a strong dehydrating agent.

It is a moderately strong oxidizing agent.

**GROUP 17 ELEMENTS**

35. **Atomic and ionic radii:** Halogens have the smallest atomic radii in their respective periods because of maximum effective nuclear charge.
36. **Ionisation enthalpy:** They have very high ionization enthalpy because of small size as compared to other groups.
37. **Electron gain enthalpy:** Halogens have maximum negative electron gain enthalpy because these elements have only one electron less than stable noble gas configuration.  
Electron gain enthalpy becomes less negative down the group because atomic size increases down the group.
38. **Electronegativity:** These elements are highly electronegative and electronegativity decreases down the group. They have high effective nuclear charge.
39. **Bond dissociation enthalpy:**  
  
Bond dissociation enthalpy follows the order  $\text{Cl}_2 > \text{Br}_2 > \text{F}_2 > \text{I}_2$   
This is because as the size increases bond length increases.  
  
Bond dissociation enthalpy of  $\text{Cl}_2$  is more than  $\text{F}_2$  because there are large electronic repulsions of lone pairs present in  $\text{F}_2$ .
40. **Colour:** All halogens are coloured because of absorption of radiations in visible region which results in the excitation of outer electrons to higher energy levels.
41. **Oxidising power:** All halogens are strong oxidizing agents because they have a strong tendency to accept electrons.  
Order of oxidizing power is  $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$
42. **Reactivity with  $\text{H}_2$ :**  
Acidic strength:  $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$   
Stability:  $\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$   
This is because of decrease in bond dissociation enthalpy.  
Boiling point:  $\text{HCl} < \text{HBr} < \text{HI} < \text{HF}$   
HF has strong intermolecular H bonding  
As the size increases, van der Waals forces increases and hence boiling point increases.

% Ionic character:  $\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$

Dipole moment:  $\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$

Electronegativity decreases down the group.

Reducing power:  $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$

43. **Reactivity with metals:** Halogens react with metals to form halides.

Ionic character:  $\text{MF} > \text{MCl} > \text{MBr} > \text{MI}$

Halides in higher oxidation state will be more covalent than the one in the lower oxidation state.

44. **Interhalogen compounds: Reactivity of halogens towards other halogens:**

Binary compounds of two different halogen atoms of general formula  $\text{X X}'_n$  are called interhalogen compounds where  $n = 1, 3, 5,$  or  $7$

These are covalent compounds.

Interhalogen compounds are more reactive than halogens because  $\text{X-X}'$  is a more polar bond than  $\text{X-X}$  bond.

All are diamagnetic.

Their melting point is little higher than halogens.

$\text{XX}'$  ( $\text{ClF}$ ,  $\text{BrF}$ ,  $\text{BrCl}$ ,  $\text{ICl}$ ,  $\text{IBr}$ ,  $\text{IF}$ ) (Linear shape)

$\text{XX}'_3$  ( $\text{ClF}_3$ ,  $\text{BrF}_3$ ,  $\text{IF}_3$ ,  $\text{ICl}_3$ ) (Bent T- shape)

$\text{XX}'_5$  –  $\text{ClF}_5$ ,  $\text{BrF}_5$ ,  $\text{IF}_5$ , (square pyramidal shape)

$\text{XX}'_7$  –  $\text{IF}_7$  (Pentagonal bipyramidal shape)

45. **Oxoacids of halogens:**

Fluorine forms only one oxoacid  $\text{HOF}$  (Fluoric (I) acid or hypofluorous acid) due to high electronegativity.

Acid strength:  $\text{HOCl} < \text{HClO}_2 < \text{HClO}_3 < \text{HClO}_4$

Reason:  $\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$  most stable

Acid strength:  $\text{HOF} > \text{HOCl} > \text{HOBr} > \text{HOI}$

This is because Fluorine is most electronegative.

### GROUP 18 ELEMENTS:

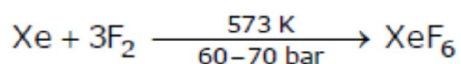
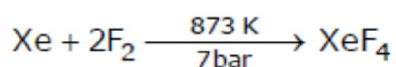
46. **Ionisation enthalpy:** They have very high ionization enthalpy because of completely filled orbitals.

Ionisation enthalpy decreases down the group because of increase in size.

47. **Atomic radii:** Increases down the group because number of shells increases down the group.

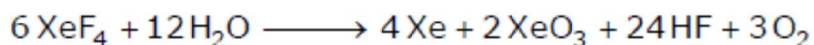
48. **Electron gain enthalpy:** They have large electron gain enthalpy because of stable electronic configuration.
49. **Melting and boiling point:** Low melting and boiling point because only weak dispersion forces are present.
50.  $\text{XeF}_2$  is linear,  $\text{XeF}_4$  is square planar and  $\text{XeF}_6$  is distorted octahedral.  $\text{KrF}_2$  is known but no true compound of He Ne and Ar are known.

51. **Compounds of Xe and F:**



$\text{XeF}_2$ ,  $\text{XeF}_4$  and  $\text{XeF}_6$  are powerful fluorinating agents.

52. **Compounds of Xe and O**



### Conceptual questions and answers

Q1) Why penthalides more covalent than trihalides

Ans) Higher the positive oxidation state of central atom, more will be its polarizing power. Which it turns increases the covalent character of bond formed between the central atom and the other atom

Q2) Why is  $\text{BiH}_3$  the strong reducing agent amongst all the hydrides of group 15 elements

Ans) Because  $\text{BiH}_3$  is the least stable amongst the hydrides of group 15. Bi –H bond is the weakest amongst the hydrides of elements of group 15.

Q3) Why does the reactivity of nitrogen differs from Phosphorous?

Ans) Nitrogen forms  $p\pi - p\pi$  multiple bonds with itself where as phosphorous does not form  $p\pi - p\pi$  multiple bonds. Thus nitrogen exists as a diatomic molecule with a triple bond ( $N \equiv N$ ) between the two atoms, consequently, its bond enthalpy is very high. Therefore nitrogen is inert and unreactive in its elemental state. In contrast, phosphorous forms single bonds and exists as  $P_4$  molecule. Since P-P single bond is much weaker than  $N \equiv N$  triple bond, therefore, phosphorous is much more reactive than nitrogen.

Q4) Why does  $NH_3$  forms hydrogen bond but  $PH_3$  does not?

Ans) The electro negativity of Nitrogen is much higher than that of phosphorous, as a result, N-H bond is quite polar and hence  $NH_3$  undergoes intermolecular H-bonding where as P-H bond is less polar and hence  $PH_3$  does not undergo H – bonding.

Q5)  $PH_3$  has lower boiling point than  $NH_3$  why?

Ans)  $NH_3$  undergoes extensive intermolecular H-bonding and hence it exists as an associated molecule. On the other hand,  $PH_3$  does not undergo H-bonding and thus exists as discrete molecules. Therefore  $PH_3$  has lower boiling point than  $NH_3$

Q6) Why does  $NH_3$  act as a Lewis base?

Nitrogen atom in  $NH_3$  has lone pair of electrons which is available for donation. Therefore, it acts as Lewis base.

Q7) Give reason for anomalous property of Nitrogen

Ans) Nitrogen differs from the rest of the members of group 15 due to its smaller size, high electro negativity, higher ionization enthalpy and non-availability of d-orbitals

Q8) Write the reaction for thermal decomposition of Sodium azide

Ans)  $2NaN_3 \rightarrow 2Na + 3N_2$

Q9) Nitrogen exists as diatomic molecule and phosphorous as  $P_4$  why?

Ans) Nitrogen because of its small size and higher electro negativity forms  $p\pi - p\pi$  multiple bonds with itself. Therefore it exists as a diatomic molecule having a triple bond between two Nitrogen atoms. Phosphorous, on the other hand, due to its large size and lower electro negativity does not form  $p\pi - p\pi$  multiple bonds with itself. Instead it prefers to form P-P single bonds and hence exists as  $P_4$  molecules

Q10) Nitrogen exists in gaseous state whereas Phosphorous in solid state why?

Ans) Refer Q9

Q11) Explain why  $\text{NH}_3$  is basic while  $\text{BiH}_3$  is only feebly basic?

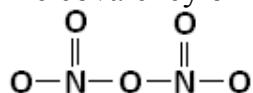
Ans) N in  $\text{NH}_3$  and Bi in  $\text{BiH}_3$  have a lone pair of electrons on the central atoms and hence should behave as Lewis base. Since atomic size of 'N' is much smaller than that of Bi therefore, electron density on the N-atom is much higher than that of Bi atom. Consequently, the tendency of N in  $\text{NH}_3$  to donate its pair of electrons is much higher than  $\text{BiH}_3$

Q12) Why does  $\text{R}_3\text{P}=\text{O}$  exist but  $\text{R}_3\text{N}=\text{O}$  does not ( R = alkyl group)

Ans) Nitrogen due to the absence of d-orbital's cannot form  $p\pi - d\pi$  multiple bonds and hence cannot expand its covalency beyond 4. Therefore N cannot form  $\text{R}_3\text{N}=\text{O}$  in which the covalency of N is 5. In contrast, P due to the presence of d-orbitals forms  $p\pi - d\pi$  multiple bonds and hence can expand its covalency beyond 4. Therefore P forms  $\text{R}_3\text{P}=\text{O}$  in which the covalency of P is 5

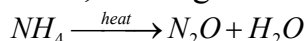
Q13) What is the covalency of Nitrogen in  $\text{N}_2\text{O}_5$

The covalency of nitrogen in  $\text{N}_2\text{O}_5$  is four

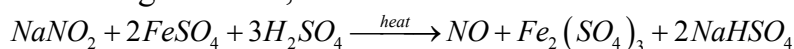


Q14) Name the various oxides of Nitrogen. Mention their oxidation states. Give their preparation

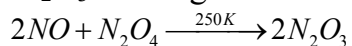
1)  $\text{N}_2\text{O}$  , Dinitrogen oxide or Nitrogen (I)oxide, oxidation State +1



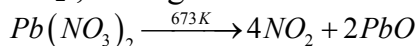
2)  $\text{NO}$  Nitrogen oxide , oxidation state +2



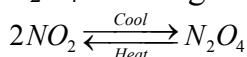
3)  $\text{N}_2\text{O}_3$  Dinitrogen trioxide or Nitrogen (III) oxide, oxidation state +3



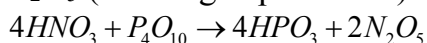
4)  $\text{NO}_2$  , Nitrogen dioxide or Nitrogen (IV) oxide, oxidation state +4



5)  $\text{N}_2\text{O}_4$  Dinitrogen tetroxide or Nitrogen (IV) oxide, oxidation state +5



6)  $\text{N}_2\text{O}_5$  ( Dinitrogen pentoxide) or Nitrogen (V) oxide, oxidation state +6



Q15) Why does  $\text{NO}_2$  dimerise?

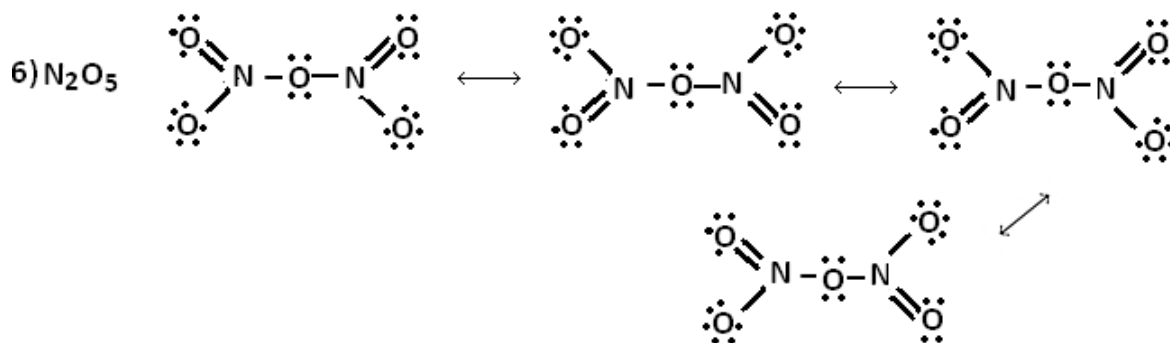
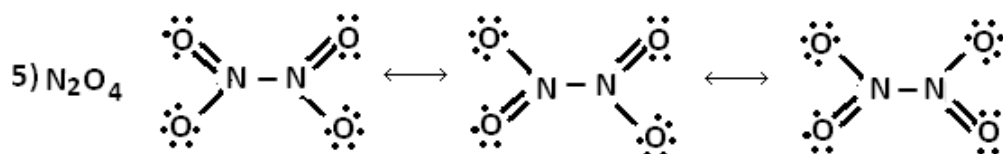
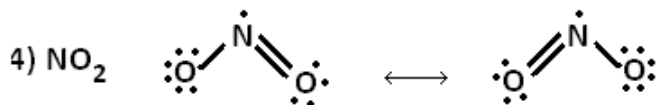
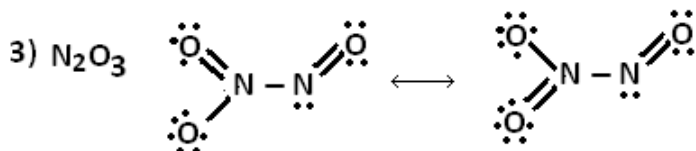
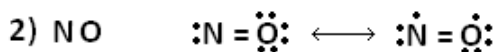
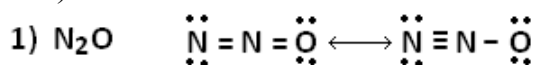
Ans)  $\text{NO}_2$  contains odd number of valence electrons. It behaves as a typical odd molecule. On dimerisation it is converted to stable  $\text{N}_2\text{O}_4$  molecule with even number of electrons

Q16) Why does nitrogen show catenation property less than Phosphorus?

Ans) The property of catenation depends upon the strength of the element – element bond. Since the single N-N bond is weaker than single P-P bond strength, therefore nitrogen shows less catenation properties than phosphorus

Q17) Give resonating structures of following oxides

Ans)



Q18) Metals like Cr, Al do not dissolve in concentrated  $\text{HNO}_3$  why?

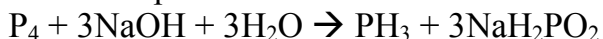
Ans) It is because of the formation of passive film of oxide on the surface

Q19) Bond angle in  $\text{PH}_4^+$  is higher than in  $\text{PH}_3$  why?

Ans) Both are  $\text{sp}^3$  hybridised. In  $\text{PH}_4^+$ , all the four orbitals are bonded where as in  $\text{PH}_3$ , there is a lone pair of electrons on P. Which is responsible for lone pair-bond pair repulsion  $\text{PH}_3$  reducing the bond angle to less than  $109^\circ 28'$

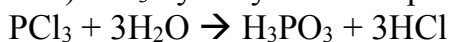
Q20) What happens when white phosphorous is heated with conc. NaOH solution in an inert atmosphere of  $\text{CO}_2$ ?

Ans) White phosphorous reacts with NaOH solution in an inert atmosphere to form Phosphine



Q21) Why does the  $\text{PCl}_3$  fume in moisture?

Ans)  $\text{PCl}_3$  hydrolyses in the presence of moisture giving fumes of HCl

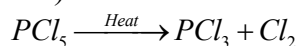


Q22) Are all the five bonds in  $\text{PCl}_5$  molecule equivalent? Justify your answer

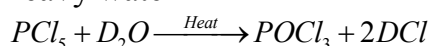
Ans)  $\text{PCl}_5$  has a trigonal bipyramidal structure and the three equatorial P-Cl bonds are equivalent, while two axial bonds are different and longer than equatorial bond

Q23) What happens when  $\text{PCl}_5$  is heated?

Ans) When heated  $\text{PCl}_5$  sublimates but decomposes on strong heating

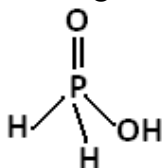


Q24) Write a balanced chemical equation for the hydrolysis reaction of  $\text{PCl}_5$  in heavy water



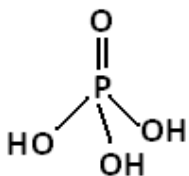
Q25) how do you account for the reducing behavior of  $\text{H}_3\text{PO}_2$ ,

Ans) In  $\text{H}_3\text{PO}_2$ , Two H atoms are bounded directly to P atom which imports reducing character to the acid.



Q26) What is the basicity of  $\text{H}_3\text{PO}_4$ ?

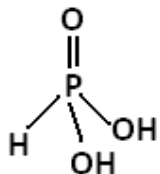
Ans) Since it contains three ionizable H-atoms which are present as OH group, it behaves as a tribasic acid i.e the basicity is three





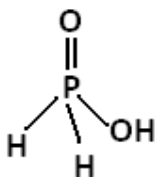
Q27)  $\text{H}_3\text{PO}_3$  is a dibasic acid why?

Ans) Since it contains two ionizable H atoms which are present as OH groups, it behaves as a dibasic acid. i.e. the basicity is two



Q28) What happens when  $\text{H}_3\text{PO}_2$  is a monobasic acid why?

Ans) Since it contains only one ionizable H-atom which is present as OH group it behaves as a monobasic acid i.e basicity is one



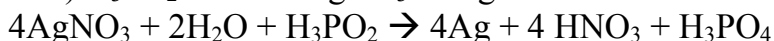
Q29) What happens when  $\text{H}_3\text{PO}_3$  is heated?

Ans) Phosphorous acid on heating disproportionate to give phosphoric acid and phosphine



Q30) Give an example for reducing property of  $\text{H}_3\text{PO}_2$  ?

Ans)  $\text{H}_3\text{PO}_2$  reduces  $\text{AgNO}_3$  to Ag



Q31) NO ( nitric oxide) is paramagnetic in the gaseous state but diamagnetic in the liquid and solid state why?

Ans) NO has an odd number of electrons and hence is paramagnetic in gaseous state, but in the liquid and solid state it exists as dimer and hence diamagnetic in these states

Q32)  $\text{NCl}_3$  gets readily hydrolyzed while  $\text{NF}_3$  does not. Why?

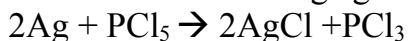
Ans) In  $\text{NCl}_3$ , Cl has vacant d-orbital to accept the lone pair of electrons donated by O-atom of  $\text{H}_2\text{O}$  molecules but in  $\text{NF}_3$ , F does not have d-orbital

Q33)  $\text{PCl}_5$  is ionic in nature in the solid state. Give reason

Ans) In the solid state it exists as  $[\text{PCl}_4]^+ [\text{PCl}_6]^-$ . Therefore  $\text{PCl}_5$  is ionic nature in the solid state

Q34) Can  $\text{PCl}_5$  act as an oxidizing as well as reducing agent? Justify

Ans) The oxidation state of P in  $\text{PCl}_5$  is +5. Since P has five electrons in its valence shell, therefore, it cannot increase its oxidation state beyond +5 by donating electrons, therefore  $\text{PCl}_5$  cannot act as a reducing agent. However, it can decrease its oxidation state from +5 to lower oxidation state, therefore  $\text{PCl}_5$  can act as an oxidizing agent



Q35)  $\text{H}_2\text{S}$  is less acidic than  $\text{H}_2\text{Te}$ . Why?

Ans) H-S bond dissociation energy is higher than that of H-Te bond dissociation energy and hence H-S bond breaks less easily than H-Te bond. Hence  $\text{H}_2\text{S}$  is a weaker acid than  $\text{H}_2\text{Te}$

Q36) Why is  $\text{H}_2\text{O}$  is liquid and  $\text{H}_2\text{S}$  a gas?

Ans) Because of small size and high electronegativity of oxygen, molecules of water are highly associated through hydrogen bonding resulting in its liquid state. In contrast,  $\text{H}_2\text{S}$  does not undergo H-bonding and exists as discrete molecules. Therefore,  $\text{H}_2\text{S}$  is gas at room temperature

Q37) Give reason for anomalous behavior of Oxygen

Ans) Oxygen differs from the rest of members of the group due to its smaller size, high electronegativity and non-availability of d-orbitals

Q38) Which of the following do not reacts with oxygen directly. Zn, Ti, Rt, Fe

Ans) Platinum does not reacts with oxygen directly. It is because the sum of first four ionization enthalpy is very large

Q39) Why is dioxygen is a gas while sulphur is solid

Ans) Due to small size and high electronegativity, oxygen form  $p\pi-p\pi$  multiple bonds, as a result oxygen exists as diatomic molecules ( $\text{O}_2$ ) and has less intermolecular force of attraction. Therefore,  $\text{O}_2$  is a gas at room temperature. Sulphur, on the other hand, because of its bigger size and lower electronegativity, does not form  $p\pi-p\pi$  multiple bonds. Instead, it prefers to form S-S single bond as a result, sulphur exists as octa-atomic ( $\text{S}_8$ ) and has high intermolecular force of attraction. Therefore, sulphur is a solid at room temperature

Q40) What are acidic oxide? Give examples

Ans) An oxide which give a acid with water is known as acidic oxides. E.g  $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{N}_2\text{O}_5$

Q41) What are basic oxides? Give examples

Ans) An oxide which give a base with water is known as basic oxide. E.g.  $N_2O$ ,  $CaO$ ,  $BaO$

Q)42) What are amphoteric oxides? Give example

Ans) Those oxides which shows characteristics of both acidic as well as basic oxides are known as amphoteric oxides . they react with acids as well as bases e.g.  $Al_2O_3$ ,  $ZnO$ ,  $PbO$ ,  $SnO$

Q43) What are neutral Oxides? Give examples

Ans) Oxides which neither react with acid nor with bases are called as neutral oxides. E.g.  $NO$ ,  $CO$ ,  $N_2P$

Q44) High concentration of ozone is dangerously explosive. Explain

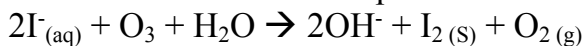
Ans) The decomposition of ozone into oxygen results in the liberation of heat ( $\Delta H$  is negative) and an increase in entropy ( $\Delta S$ ) . these two effects reinforce each other resulting in large negative Gibbs energy change ( $\Delta G$  is negative) for its conversion into oxygen. Therefore, higher concentrations of ozone is dangerously explosive

Q45) Why does the ozone acts as a power oxidizing agent?

Ans) Due to the ease with which it liberates atom of nascent oxygen it acts as powerful oxidizing agent

Q46) How is  $O_3$  estimated quantitatively

Ans) When  $O_3$  is treated with excess of  $KI$  solution buffered with Borate buffer ( $pH = 9.2$ ).  $I_2$  is liberated quantitatively which can be titrated against a standard solution of sodium thiosulphate



Q47) Which aerosols deplete ozone?

Ans) Freons deplete the ozone layer supplying  $Cl$  free radical which converts  $O_3$  to  $O_2$

Q48) Explain how nitrogen oxides deplete ozone layer

An) Nitrogen oxides emitted from the exhaust system of supersonic jet aeroplanes might be slowly depleting the concentration of the ozone in the upper atmosphere



Q48) What is meant by transition temperature of sulphur

Ans) The temperature at which both  $\alpha$  –sulphur and  $\beta$ -sulphur are stable is known as transition temperature. The transition temperature of sulphur is  $369K$  .

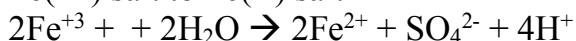
Below 369K  $\alpha$  –sulphur is stable where as above 369K  $\beta$ -sulphur is stable

Q49) Which form of sulphur shows paramagnetic behavior

Ans) In vapour state sulphur partly exists as  $S_2$  molecule which has two unpaired electrons in the antibonding  $\pi^*$  orbitals like  $O_2$  and hence, exhibits paramagnetism

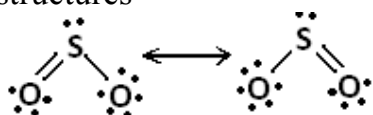
Q50) What happens when  $SO_2$  ( sulphur dioxide ) is passed through an aqueous solution of Fe(III) salt

Ans)  $SO_2$  acts as a reducing agent and hence reduces an aqueous solution of Fe(III) salt to Fe(II) salt



Q51) Comment on the two S-O bonds formed in  $SO_2$  molecule , are the two bonds in this molecule equal

Ans) Both S-O bonds are covalent and have equal strength due to resonating structures



Q52) Why is  $K_{a2} \ll K_{a1}$  for  $H_2SO_4$  in water

Ans)  $H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$  ,  $K_{a1}$

$HSO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$  ,  $K_{a2}$

$K_{a2}$  is less than  $K_{a1}$  because the negatively charged  $HSO_4^-$  ion has much less tendency to donate a proton to  $H_2O$  as compared to neutral  $H_2SO_4$  to donate a proton to water

Q53) Halogens have maximum negative electron gain enthalpy in the respective periods of the periodic table why

Ans) Halogens have the smallest size in the respective periods and therefore high effective nuclear charge. As a consequence they readily accept one electron to acquire noble gas electronic configuration.

Q54) The electron gain enthalpy of fluorine is less than that of chlorine why?

Ans) It is due to small size of fluorine atom. As a result, there is strong interelectronic repulsion in the relatively small 2p orbitals of fluorine and thus the incoming electron does not experience much attraction

Q55) Although electrogain enthalpy is a stronger oxidizing agent than chlorine why?

Ans) It is due to

- (i) Low enthalpy of dissociation of F-F bond
- (ii) High hydration and enthalpy of F<sup>-</sup> ion

Q)56) Give two examples to show the anomalous behavior of fluorine?

Ans) (i) Fluorine shows only an oxidation state of -1 while other halogens.

Show positive oxidation states of +1, +3, +5 and +7 besides oxidation state of -1

- (ii) The bond dissociation enthalpy of F-F bond is lower than that of Cl-Cl bond

Q57) Explain why fluorine forms only one oxoacid, HOF?

Ans) Fluorine is the most electronegative element and can exhibit only -1 oxidation state. Secondly Fluorine does not have d-orbitals in its valence shell and therefore cannot expand its octet. Hence Fluorine forms only one oxoacid

Q58) What are interhalogen compounds? Give examples

Ans) When two different halogens react with each other to form a number of compounds called interhalogen compounds. Their general formula is XX'<sub>n</sub> where X is a less electronegative halogen while X' is a more electronegative halogen (smaller size halogen) and n is its number e.g IF<sub>7</sub>, BrF<sub>5</sub>, IF<sub>5</sub>

Q59) Explain why in spite of nearly same electronegativity oxygen forms hydrogen bonding while chlorine does not

Ans) Oxygen has a smaller size than chlorine. As a result electron density per unit volume on oxygen atoms is much higher than that of on chlorine atoms.

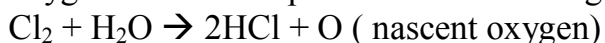
Therefore, oxygen forms H-Bonds while chlorine does not

Q60) Why are halogens coloured?

Ans) All the halogens are coloured. This is due to absorption of radiation in the visible region which results in the excitation of outer electrons to a higher energy level. While the remaining light is transmitted. The colour of the halogens is actually the colour of this transmitted light. By absorbing different quanta of radiation they display different colours

Q61) Give the bleaching action of Cl<sub>2</sub>

Ans) In the presence of moisture or in aqueous solution, Cl<sub>2</sub> liberates nascent oxygen which is responsible for the bleaching action of Cl<sub>2</sub>



The bleaching action of Cl<sub>2</sub> is due to oxidation

Q62) Why is ICl more reactive than I<sub>2</sub>?

Ans) In general, interhalogen compounds are more reactive than halogens due to weaker X-X' bonding than X-X bond thus ICl is more reactive than I<sub>2</sub>

Q63) Why are the elements of Group 18 known as noble gases?

Ans) The elements present in Group 18 have their valence shell orbitals completely filled and therefore, react with a few elements only under certain condition. Therefore, they are known as noble gases

Q64) Noble gases have very low boiling points. Why?

Ans) Noble gases being mono-atomic have no inter-atomic forces except weak dispersion forces and therefore, they are liquefied at very low temperatures. Hence, they have low boiling point.

Q65) Noble gases are least reactive. Why?

Ans) It is due to following reasons

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

Q66) What inspired N. Bartlett for carrying out reaction between Xe and  $PtF_6$ ?

Ans) N. Bartlett observed that  $PtF_6$  reacts with  $O_2$  to yield an ionic solid  $O_2^+PtF_6^-$ . Since the first ionization enthalpy of molecular oxygen ( $1175 \text{ kJmol}^{-1}$ ) is almost identical with that of xenon is ( $1170 \text{ kJmol}^{-1}$ ). Bartlett thought that  $PtF_6$  should also oxidise Xe to  $Xe^+$ . This inspired Bartlett to carry out the reaction between Xe and  $PtF_6$ . He prepared  $Xe^+[PtF_6]^-$  mixing  $PtF_6$  and Xe  
 $Xe + PtF_6 \rightarrow Xe^+ [PtF_6]^-$

Q67) Does the hydrolysis of  $XeF_6$  leads to a redox reaction

Ans) No, the products of hydrolysis are  $XeOF_4$  and  $XeO_2 F_2$  where oxidation states of all elements remain the same as it was in the reacting state

Q68) Why is helium used in diving apparatus?

Ans) Because of its low solubility compared to  $N_2$  in blood a mixture of oxygen and helium is used in diving apparatus

Q69) Why has been difficult to study the chemistry of radon?

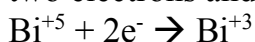
Ans) Radon is radioactive with very short half life which makes the study of chemistry of radon difficult

Q70) With what neutral molecule is  $ClO^-$  iso-electronic . is that molecule is a Lewis base

Ans)  $ClF$  , Yes it is Lewis base due to presence of lone pair of electrons

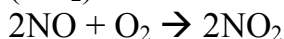
Q71) Bismuth is a strong oxidizing agent in the pentavalent state. Why?

Ans) Due to inter pair effect, +5 oxidation state of Bi is less stable than +3 oxidation state. Therefore, Bismuth in the pentavalent state can easily accept two electrons and this gets reduced to trivalent bismuth



Q72) Nitric oxide becomes brown when released in air. Why?

Ans) Nitric oxide readily combines with  $\text{O}_2$  of the air to form a nitrogen dioxide ( $\text{NO}_2$ ) which has brown colour



Q73)  $\text{H}_2\text{S}$  is less acidic than  $\text{H}_2\text{Te}$ . Why?

Ans) H-S bond dissociation enthalpy is higher than that of H-Te bond dissociation enthalpy and hence H-Te bond breaks easily than H-S bond.

Therefore  $\text{H}_2\text{S}$  is a weaker acid than  $\text{H}_2\text{Te}$

Q74)  $\text{H}_2\text{S}$  acts only as a reducing agent but  $\text{SO}_2$  acts as a reducing agent as well as an oxidizing agent. Why?

Ans) The minimum oxidation state of S is -2 while its maximum oxidation state is +6. In  $\text{SO}_2$ , the oxidation state of S is +4 therefore, it can not only increase its oxidation state by losing electrons but can also decrease its oxidation state by gaining electrons. As a result,  $\text{SO}_2$  acts both as a reducing agent as well as oxidizing agent. On other hand, in  $\text{H}_2\text{S}$ , S has an oxidation state of -2. As a result it can only increase its oxidation state by losing electrons and hence it acts only as a reducing agent

Q75) Why the compounds of Fluorine with oxygen are called fluorides of oxygen (Oxygen fluorides) and not the oxides of fluorine (Fluorine oxide)

Ans) This is because fluorine is more electronegative than oxygen

Q76) Both  $\text{NO}$  and  $\text{ClO}_2$  are odd electron species whereas.  $\text{NO}$  dimerises but  $\text{ClO}_2$  does not. Why?

Ans) In  $\text{NO}$ , the odd electron on N is attracted by only one O-atom but in  $\text{ClO}_2$ , the odd electron on Cl is attracted by two O-atoms. As a result the odd electron on N in  $\text{NO}$  is localized while the odd electron on Cl in  $\text{ClO}_2$  is delocalized.

Consequently  $\text{NO}$  has a tendency to dimerize but  $\text{ClO}_2$  does not.

Q77) Bleaching by Chlorine is permanent while that by  $\text{SO}_2$  is temporary Why?

Ans)  $\text{Cl}_2$  bleaches coloured material by Oxidation and hence bleaching is permanent. On the other hand,  $\text{SO}_2$  bleaches coloured material by reduction and hence is temporary since when the bleached colour material exposed to air, it gets oxidized and the colour is restored.

Q78) Hydrogen iodide is stronger acid than Hydrogen fluoride in aqueous solution. Why?

Ans) Because of larger size of Iodine as compared to fluorine, the bond dissociation enthalpy of H-I bond is much lower than that of H-F bond. As a result, in aqueous solution, H-I bond breaks more easily to form  $H^+$  ion than H-F bond. Thus HI is strong acid than HF in aqueous solution

Q79) Helium and Neon do not forms compounds with fluorine why?

Ans) He and Ne do not contain d-orbitals in their respective valence shells and hence their electrons cannot be promoted to higher energy levels. Therefore, He and Ne do not forms compounds with fluorine

Q80) Neon is generally used for warning signals. Why?

Ans) Neon lights are visible from long distance even in fog and Mist and hence Neon is generally used for warning signals

Q81) Xenon does not form fluorides such as  $XeF_3$  and  $XeF_5$

Ans: As all the filled orbitals of Xe have paired electrons. The promotion of one, two or three electrons from the 5P – filled orbitals to 5d- vacant orbitals will give rise two, four and six half filled orbitals. So Xe can combine with even but not odd number of F atoms. Hence it cannot form  $XeF_3$  and  $XeF_5$

Q82) Out of noble gases, only Xenon is known to form chemical; compounds why?

Ans) Except radon, which is radioactive, Xe has least ionization enthalpy among noble gases and hence it readily forms chemical compounds particularly  $O_2$  and  $F_2$

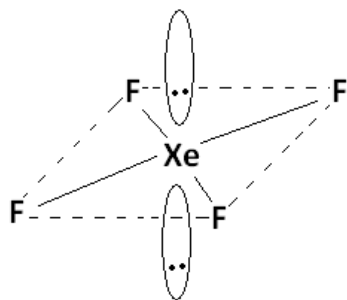
Q83) Knowing the electron gain enthalpy values of  $O_2 \rightarrow O^-$  and  $O \rightarrow O^{2-}$  as  $141 \text{ KJ mol}^{-1}$  and  $700 \text{ KJ mol}^{-1}$  respectively, how can you account for the formation of a large number of oxides having  $O^{2-}$  species and not  $O^-$

Ans) It is because lattice energy of Oxides which overcome second electron gain enthalpy that is oxides are formed with  $O^{2-}$

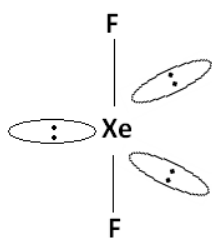
Q84) Give the formula and describe the structure of a noble gas species which is isostructural with (i)  $ICl_4^-$  (ii)  $IBr_2^-$  3)  $BrO_3^-$

Ans (i)  $ICl_4^-$  is isostructural with  $XeF_4$ .  $XeF_4$  has four bonding pair and two lone pairs of electrons. Therefore according to VSEPR theory  $XeF_4$  has square planar structure.

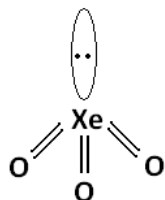




(II)  $\text{IBr}_2^-$  : is isostructural with  $\text{XeF}_2$ .  $\text{XeF}_2$  has two bonding pairs and three lone pairs of electrons. Therefore according to VSEPR theory,  $\text{XeF}_2$  has linear shape.



(III)  $\text{BeO}_3^-$  is isostructural with  $\text{XeO}_3$ ,  $\text{XeO}_3$  has three bonding pairs and one lone pair of electrons. Therefore according to VSEPR theory,  $\text{XeO}_3$  has pyramidal structure



Q85) Deduce the molecular shape of  $\text{BrF}_3$  on the basis of VSEPR theory

Ans) The central atom Br has seven electrons in the valence shell. Three of these will form electron pair with three fluorine atoms leaving behind four electrons. Thus there are three bond pairs and two lone pairs.

According to VSEPR theory, these will occupy the corners of trigonal bipyramid. The two lone pairs will occupy the equatorial position to minimize. Lone pair – lone pair repulsion. Which are greater than the bond pair – bond pair repulsion. In addition the axial Fluorine atom will be bent towards the equatorial fluorine in order to minimize the lone pair – lone pair repulsions. The shape will be that of slightly bent ‘T’

Q86) Which one of  $\text{PCl}_4^-$  and  $\text{PCl}_4^+$  is not likely to exist and why?

Ans)  $\text{PCl}_4^-$  does not exist because octet of ‘P’ is not complete and it is unstable

