# CBSE NCERT Solutions for Class 12 Chemistry Chapter 7

***Back of Chapter Questions***

**1.** Why are pentahalides more covalent than trihalides?

**Solution:**

The oxidation state of $+5$ in pentahalides is more as compared to the $+3$ oxidation state in trihalides. Due to the higher positive oxidation state of the central atom in pentahalide state, these atoms will have larger polarizing power than the halogen atom attached to them since the polarizing power is directly proportional to the charge. The central atom in pentahalide state will tend to polarize more the halide ion to which it is attached. In the case of trihalides, the central atom will polarize the halogen atom to a lesser extent as compared to a pentahalide state. More is the polarization, larger will be the covalent character of the bond. Hence due to larger polarization of bond in pentahalide state as compared to trihalide state, the pentahalides are more covalent than trihalides.

**2.** Why is $BiH_3$ the strongest reducing agent amongst all the hydrides of Group 15 elements?

**Solution:**

As we move down the group, the size of the element increases and, therefore, the bond length of the M— H bond increases and its strength decreases(i.e. stability decreases). If the hydride is less stable, it is more reactive and possesses higher reducing strength. The reducing character of hydrides increases on moving from $NH_3$ to $BiH_3$. Hence, $BiH_3$ is the strongest reducing agent among hydrides of Group-15 elements.

**3.** Why is $N_2$ less reactive at room temperature?

**Solution:**

Three electron pairs are shared between two nitrogen atoms to form a dinitrogen $(N_2)$ molecule. The two nitrogen atoms are joined by a strong triple bond$(N \equiv N)$. Since the nitrogen atom is very small in size, therefore the bond length is also quite small. Also, both the nitrogen complete their octet. So the bond dissociation energy is very high. Due to this reason, $N_2$ is very less reactive at room temperature.

**4.** Mention the conditions required to maximise the yield of ammonia.

**Solution:**

For large scale production of ammonia Haber's process is used.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

The conditions required to maximize the yield of ammonia are as follows:

1) In accordance with Le Chatelier's principle, low temperature favours the reaction. It shifts the equilibrium to the right because the reaction is exothermic. This gives a greater yield of ammonia. Therefore a temperature of about 700 K will be optimal for the preparation of ammonia.

2) High pressure on the reaction at equilibrium favours the shift of the equilibrium to the right. It is because the forward reaction proceeds with a decrease in the number of moles. Therefore a pressure of about 200 atm will be optimal for a higher yield of ammonia.

3) A mixture of iron oxide with small amounts of $K_2O$ and $Al_2O_3$ should be used as a catalyst.

**5.** How does ammonia react with a solution of $Cu^{2+}$?

**Solution:**

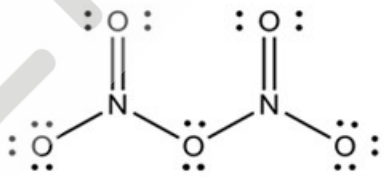
When ammonia reacts with a solution of $Cu^{2+}$, ammonia acts as a Lewis base due to the presence of lone pair of electrons on the nitrogen atom. Ammonia reacts with a solution of $Cu^{2+}$ to from a deep blue coloured complex compound $[Cu(NH_3)_4]^{2+}$.The reaction can be written as follows:

$$\underset{\text{blue}}{Cu^{2+}}(aq)+4NH_3(aq) \rightarrow \underset{\text{deep blue}}{[Cu(NH_3)_4]^{2+}}(aq)$$

**6.** What is the covalence of nitrogen in $N_2O_5$ ?

**Solution:**

Covalence is the number of electron pairs that an atom can share with other atoms. The structure of $N_2O_5$ is shown below:



From the above diagram, we see that the covalence of nitrogen is 4.

**7.** Bond angle in $PH_4^+$ is higher than that in $PH_3$. Why?

**Solution:**

Phosphorus(P) atom in both $PH_3$ and $PH_4^+$ is $sp^3$ hybridized. Due to the absence of lone pair-bond pair repulsion and the presence of four identical bond pair-bond pair interactions, $PH_4^+$ has tetrahedral geometry with a bond angle of $109^o$ 28'. But $PH_3$ has three bond pairs and one lone pair around P. Due to greater lone pair-bond pair repulsion than bond pair-bond pair repulsion, the tetrahedral angle decreases from $109^o$ 28' to $93.6^o$. Therefore, $PH_3$ is pyramidal in shape. Hence bond angle in $PH_4^+$ is higher than that in $PH_3$.

**8.** What happens when white phosphorus is heated with concentrated NaOH solution in an inert atmosphere of $CO_2$?

**Solution:**

When white phosphorus is heated with concentrated NaOH in an inert atmosphere of $CO_2$, the following reaction takes place:

$$P_4 + 3NaOH + 3H_2O \xrightarrow[\Delta]{CO_2 \text{ atmosphere}} \underset{\text{Phosphine}}{PH_3} + \underset{\text{Sodium hypophosphite}}{3NaH_2PO_2}$$

Phosphine and sodium hypophosphite are formed as products. This reaction is an example of a disproportionation reaction in which the oxidation state of phosphorus decreases from 0 in $P_4$ to -3 in $PH_3$, while it increases from 0 in $P_4$ to +1 in $NaH_2PO_2$.

**9.** What happens when $PCl_5$ is heated?

**Solution:**

In $PCl_5$, phosphorus undergoes $sp^3d$ hybridization and has a trigonal bipyramidal structure. It has three equatorial $P-Cl$ bonds and two axial $P-Cl$ bonds which are different. The axial bonds are larger than equatorial bonds. $PCl_5$ is thermally less stable than $PCl_3$.On heating, it sublimes but decomposes on stronger heating into phosphorus trichloride and chlorine.

$$PCl_5 \xrightarrow{\Delta} PCl_3 + Cl_2$$

**10.** Write a balanced equation for the reaction of $PCl_5$ with water.

**Solution:**

$PCl_5$ partially hydrolyses to phosphorus oxychloride.

$$PCl_5 + H_2O \rightarrow POCl_3 + 2HCl$$

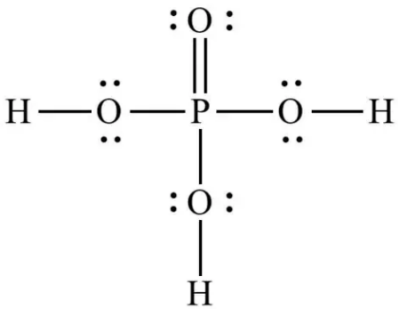
and on complete hydrolysis, the $POCl_3$ gets converted to phosphoric acid.

$$POCl_3 + 3H_2O \rightarrow H_3PO_4 + 3HCl$$

**11.** What is the basicity of $H_3PO_4$?

**Solution:**

The basicity of a compound is defined as the number of acidic hydrogen atoms present in the compound. Acidic hydrogen is bonded to a strongly electronegative element like fluorine or oxygen atom.



From the above figure, we see that one hydrogen atom is attached to each oxygen atom, which makes them acidic. Since the total number of acidic hydrogen is 3, the basicity of $H_3PO_4$ is 3.

**12.** What happens when $H_3PO_3$ is heated?

**Solution:**

Heating of $H_3PO_3$ is a disproportionate reaction. The products formed are $PH_3$ and $H_3PO_4$. Phosphorus has oxidation states of +3,+5 and -3 in orthophosphorous acid ($H_3PO_3$), orthophosphoric acid ($H_3PO_4$) and phosphine ($PH_3$) respectively. Since the oxidation number of the same element is decreasing and increasing during a particular reaction, the reaction is a disproportionation reaction.

$$\underset{\text{orthophosphorous acid}}{\overset{+3}{4H_3PO_3}} \xrightarrow{\Delta} \underset{\text{orthophosphoric acid}}{\overset{+5}{3H_3PO_4}} + \underset{\text{phosphine}}{\overset{-3}{PH_3}}$$

**13.** List the important sources of sulphur.

**Solution:**

Sulphur mainly occurs in the earth's crust in the combined state primarily in the form of sulphates and sulphides.

Sulphates are gypsum($CaSO_4.2H_2O$), Epsom salt $MgSO_4.7H_2O$, Baryte $BaSO_4$

Sulphides are galena(PbS), zinc blende(ZnS), copper pyrites($CuFeS_2$)

Traces of sulphur also occur as hydrogen sulphide in volcanoes. Organic materials such as eggs, proteins, garlic, onion, mustard, hair and wool also contain sulphur.

**14.** Write the order of thermal stability of the hydrides of Group 16 elements.

**Solution:**

On moving down the group, there is a decrease in the bond dissociation enthalpy($H-E$) of hydrides with an increase in atomic size. ($E = O, S, Se, Te, Po$). Hence, the thermal stability of the hydrides decreases down the group. The thermal stability is in the order $H_2O > H_2S > H_2Se > H_2Te > H_2Po$.

**15.** Why is $H_2O$ a liquid and $H_2S$ a gas?

**Solution:**

Weak Van der Waal's force is present in $H_2S$ molecules whereas intermolecular hydrogen bonding is present in water ($H_2O$). This is possible because the oxygen atom has a smaller size and greater electronegativity than sulphur atom. Hence water is a liquid and $H_2S$ is a gas.

**16.** Which of the following does not react with oxygen directly?

Zn, Ti, Pt, Fe

**Solution:**

Pt is a noble metal (inert) and does not react with oxygen directly, whereas Zn, Ti, and Fe are active metals which react quickly with oxygen to form their respective oxides.

**17.** Complete the following reactions:

(i) $C_2H_4 + O_2 \rightarrow$

(ii) $4Al + 3O_2 \rightarrow$

**Solution:**

(i) $\underset{\text{Ethane}}{C_2H_4} + O_2 \rightarrow 2CO_2 + 2H_2O$

Ethane on reaction with oxygen gives carbon dioxide and water.

(ii) $4Al + 3O_2 \rightarrow \underset{\text{alumina}}{2Al_2O_3}$

Aluminium combines with oxygen to form alumina.

**18.** Why does $O_3$ act as a powerful oxidising agent?

**Solution:**

Ozone is not a very stable compound at room temperature. On heating, it decomposes readily to give a molecule of oxygen and nascent oxygen.

$O_3 \xrightarrow{\Delta} O_2 + [O]$

Nascent oxygen, being a free radical, is very reactive. Hence, it acts as a powerful oxidizing agent.

**19.** How is $O_3$ estimated quantitatively?

**Solution:**

Ozone can be estimated quantitatively with the help of potassium iodide. When ozone is made to react with potassium iodide solution buffered with a borate buffer (pH 9.2), iodine is liberated. This liberated iodine is titrated against a standard solution of sodium thiosulphate using starch as an indicator.

$$\underset{\text{Iodide}}{2I^-} + H_2O + \underset{\text{Ozone}}{O_3} \rightarrow 2OH^- + \underset{\text{Iodine}}{I_2} + O_2$$

$$I_2 + 2\underset{\substack{\text{Sodium}\\\text{thiosulphate}}}{Na_2S_2O_3} \rightarrow \underset{\substack{\text{Sodium}\\\text{tetrathionate}}}{Na_2S_4O_6} + NaI$$

**20.** What happens when sulphur dioxide is passed through an aqueous solution of Fe(III) salt?

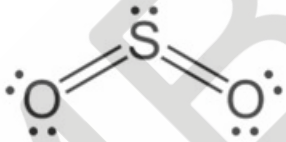
**Solution:**

$SO_2$ acts as a reducing agent when passed through an aqueous solution containing Fe(III) salt. It reduces ferric ions [Fe(III)] to ferrous ions[Fe(II)].

$$2Fe^{3+} + SO_2 + 2H_2O \rightarrow 2Fe^{2+} + SO_4^{2-} + 4H^+$$

**21.** Comment on the nature of two S– O bonds formed in $SO_2$ molecule. Are the two S– O bonds in this molecule equal?

**Solution:**




Steric number of sulphur = number of lone pair + number of bond pair $= 1 + 2 = 3$. So it is $sp^2$ hybridised. Two of these orbitals form sigma bonds with two oxygen atoms and the third orbital contains a lone pair of electrons. The p-orbital and d-orbital contain an unpaired electron each. An electron in the p-orbital forms p π – p π bond with one oxygen atom and the other electron forms p π – d π bond with the other oxygen atom. Therefore $SO_2$ has a bent structure. Also, both the S – O bonds have equal strength due to resonating structures.

**22.** How is the presence of $SO_2$ detected?

**Solution:**

Sulphur dioxide($SO_2$) is a colourless and pungent smelling gas. Acidified potassium permanganate(VII) solution can be used to detect $SO_2$. When $SO_2$ is passed through acidified potassium permanganate(VII) solution it decolourises the solution.

$$5SO_2 + 2MnO_4^- + 2H_2O \rightarrow 5SO_4^{2-} + 4H^+ + 2Mn^{2+}$$

**23.** Mention three areas in which $H_2SO_4$ plays an important role.

**Solution:**

$H_2SO_4$ (Sulphuric acid) plays an important role in the following three areas:

(1) Sulphuric acid is a very important industrial chemical. It is used for the manufacture of hundreds of other compounds and also in many industrial processes. The bulk of sulphuric acid produced is used in the manufacture of fertilisers(e.g., ammonium sulphate, superphosphate).

(2) Sulphuric acid is used for metallurgical applications (e.g., cleansing metals before enamelling, electroplating and galvanising.

(3) Sulphuric acid is widely used as a laboratory reagent.

**24.** Write the conditions to maximise the yield of $H_2SO_4$ by contact process.

**Solution:**

Contact process consists of three steps:

(1) $SO_2$ is produced by burning of sulphide ores.

(2) Conversion of $SO_2$ to $SO_3$ by the reaction with oxygen in the presence of a catalyst $(V_2O_5)$

(3) Absorption of $SO_3$ in $H_2SO_4$ to give Oleum $(H_2S_2O_7)$.

The second step is the key step. It is very crucial to maximise the yield of $H_2SO_4$.

$$2SO_2(g)+O_2(g) \overset{V_2O_5}{\rightarrow} 2SO_3(g) \qquad \Delta_r H^0 = -196.6\ kJmol^{-1}$$

The above reaction is exothermic, reversible and the forward reaction leads to a decrease in volume. So, in accordance with Le Chatelier's principle, low temperature and high pressure are the favourable conditions for maximum yield. But the temperature should not be too low since the rate of reaction will become slow.

The optimum conditions in the contact process to maximise the yield of sulphuric acid in an industrial plant is a pressure of 2 bar and temperature around 720 K.

**25.** Why is $K_{a_2} << K_{a_2}$ for $H_2SO_4$ in water?

**Solution:**

$$H_2SO_4(aq)+H_2O(l) \rightarrow H_3O^+(aq)+HSO_4^-(aq); \quad K_{a_1} > 10 \text{ (very large)}$$

$$HSO_4^-(aq)+H_2O(l) \rightarrow H_3O^+(aq)+SO_4^{2-}(aq); \quad K_{a_2} = 1.2 \times 10^{-2}$$

$$K_{a_1} = \frac{[HSO_4^-][H_3O^+]}{[H_2SO_4]}$$

$$K_{a_2} = \frac{[H_3O^+][SO_4^{2-}]}{[HSO_4^-]}$$

The conjugate base of $H_2SO_4$ is $HSO_4^-$ and it is stable due to resonance. The conjugate base of $HSO_4^-$ $SO_4^{2-}$ is which is a doubly negatively charged ion. so it is less stable than $HSO_4^-$.

**26.** Considering the parameters such as bond dissociation enthalpy, electron gain enthalpy and hydration enthalpy, compare the oxidising power of $F_2$ and $Cl_2$

**Solution:**

The oxidising power of an element is determined by three factors. They are:

(1) Bond dissociation enthalpy

(2) Electron gain enthalpy

(3) Hydration enthalpy

The electron gain enthalpy of chlorine $(-349\ \text{KJ mol}^{-1})$ is more negative than that of fluorine $(-333\ \text{KJ mol}^{-1})$ whereas the bond dissociation energy of fluorine $(-158.8\ \text{KJ mol}^{-1})$ is much lesser than that of chlorine $(-242.6\ \text{KJ mol}^{-1})$. Also, due to the small size of fluorine, the hydration energy of fluorine is much higher than that of chlorine $(-381\ \text{KJ mol}^{-1})$ Therefore, the two factors(1 & 3) compensate more than the less negative electron gain enthalpy of fluorine. Thus, fluorine is a much stronger oxidizing agent than chlorine.

**27.** Give two examples to show the anomalous behaviour of fluorine.

**Solution:**

Two examples to show anomalous behaviour of fluorine:

(1) Fluorine does not tend to form polyhalide ion (This is due to non-availability of d orbitals in valence shell) whereas other halogen form polyhalide ions.

(2) Fluorine shows only one oxidation state ($-1$ due to nonavailability of vacant d-orbitals) while all other halogens show variable oxidation states like $-1, +1, +3, +5, +7$ (due to the presence of vacant d-orbitals).

**28.** Sea is the greatest source of some halogens. Comment.

**Solution:**

Sea water contains chlorides, bromides and iodides of sodium, potassium, magnesium and calcium, but is mainly sodium chloride solution (2.5% by mass). Dried up sea shores deposits contain compounds like sodium chloride and carnallite, $KCl.MgCl_2.6H_2O$. Some forms of marine life contain iodine in their systems; various seaweeds, for example, contain up to 0.5% of iodine and Chile saltpetre contains up to 0.2% of sodium iodate. Thus, the sea is the greatest source of some halogens.

**29.** Give the reason for bleaching action of $Cl_2$.

**Solution:**

Chlorine on reaction with water in the absence of sunlight forms hydrochloric acid and hypochlorous acid(HOCl). The hypochlorous acid is not stable and it decomposes into hydrochloric acid and nascent oxygen.

$Cl_2 + H_2O \rightarrow HCl + HOCl$

$HOCl \rightarrow HCl + \underset{\text{Nascent oxygen}}{[O]}$

This nascent oxygen is responsible for bleaching action.

**30.** Name two poisonous gases which can be prepared from chlorine gas.

**Solution:**

Two poisonous gases that can be prepared from chlorine gas are:

(1) Mustard gas $(ClCH_2CH_2SCH_2CH_2Cl)$

(2) Phosgene $(COCl_2)$

**31.** Why is ICl more reactive than $I_2$?

**Solution:**

The $I-Cl$ bond in ICl is weaker than $I-I$ bond in $I_2$. So, the bond between ICl will break easily, and the I and Cl atoms will be easily available to form another compound as compared to $I-I$ atoms in $I_2$. This shows that the bond dissociation enthalpy of ICl is less than $I_2$. Therefore ICl is more reactive than $I_2$.

**32.** Why is helium used in diving apparatus?

**Solution:**

Air contains a large amount of nitrogen and the solubility of gases in liquids increases with increase in pressure. When sea divers dive deep into the sea, a large amount of nitrogen dissolves in their blood. When they come back to the surface, the solubility of nitrogen decreases and it separates from the blood and forms small air bubbles which lead to a dangerous medical condition called bends. Therefore, the air in oxygen cylinders used for diving is diluted with helium gas. This is done as Helium is sparingly less soluble in blood.

**33.** Balance the following equation:

$XeF_6 + H_2O \rightarrow XeO_2F_2 + HF$

**Solution:**

The balanced chemical equation is:

$XeF_6 + 2H_2O \rightarrow XeO_2F_2 + 4HF$

**34.** Why has it been difficult to study the chemistry of radon?

**Solution:**

Radon is a radioactive substance which has a very short half-life of 3.82 days. Since Radon is a radioactive substance it is dangerous to study as it emits harmful radiations. Also, compounds of radon such as $RaF_2$ have not been isolated yet. They have only been identified and further properties of radon have not been determined. Hence it has been difficult to study the chemistry of radon.

**35.** Discuss the general characteristics of Group 15 elements with reference to their electronic configuration, oxidation state, atomic size, ionisation enthalpy and electronegativity.

**Solution:**

The general characteristics of the group 15 elements with reference to their electronic configuration, oxidation state, atomic size, ionisation enthalpy and electronegativity are as follows:

**(i)** **Electronic configuration:** All the elements of group 15 have 5 electrons in their valence shell. These elements have the general electronic configuration as $ns^2np^3$.

**(ii)** **Oxidation state:** The elements belonging to group 15 generally exhibit $-3$, $+3$ and $+5$ oxidation states. The tendency to exhibit $-3$ oxidation state decreases as we move down the group due to an increase in the size of the atom and the metallic character. Due to the inert pair effect, the stability of $+5$ state decreases and $+3$ state increases as we move down the group.

**(iii)** **Atomic size:** Atomic size of the elements increase as we move down the group due to an increase in the number of orbitals.

**(iv)** **Ionisation enthalpy:** Ionisation enthalpy decreases as we move from top to bottom in the group due to increase in atomic size.

**(v)** **Electronegativity:** Electronegativity decreases as we move from top to bottom in the group due to increase in atomic radius.

**36.** Why does the reactivity of nitrogen differ from phosphorus?

**Solution:**

Nitrogen atom forms a strong triple bond with another nitrogen atom by strong $p\pi - p\pi$ overlap. This triple bond in $N_2$ has a high bond strength which leads to high bond dissociation energy. Phosphorous does not show this property of $p\pi - p\pi$ overlap. Hence, nitrogen is less reactive than phosphorous.

**37.** Discuss the trends in chemical reactivity of group 15 elements.

**Solution:**

The trends in chemical reactivity of group 15 elements are:

(i) **Reactivity towards hydrogen:** The group 15 elements react with hydrogen to form hydrides. They are of the type $EH_3$, where $E = N, P, As, Sb,$ or $Bi$. The stability of these hydrides decreases on moving down the group from $NH_3$ to $BiH_3$.

(ii) **Reactivity towards oxygen:** Group 15 elements form two types of oxides. They are of the type $E_2O_3$ and $E_2O_5$, where $E = N, P, As, Sb,$ or $Bi$. The oxide with the group 15 element in the higher oxidation state is more acidic than those with an element having a lower oxidation state. Down the group, the acidic character of the oxides decreases.

(iii) **Reactivity towards halogens:** The group 15 elements react with halogens to form salts of type: $EX_3$ and $EX_5$, where $E = N, P, As, Sb,$ or $Bi$ and $X = F, Cl, Br,$ or $I$. But nitrogen does not form $NX_5$ since it lacks d −orbital. All trihalides except $NX_3$ are stable.

(iv) **Reactivity towards metals:** The group 15 elements react with metals to form binary compounds of the type $M_3E_2$, where $E = N, P, As, Sb,$ or $Bi$. The metals in these compounds exhibit −3 oxidation state.

**38.** Why does $NH_3$ form hydrogen bond but $PH_3$ does not?

**Solution:**

The electronegativity of nitrogen (3.0) is higher compared to phosphorus (2.1). This means that electrons in $NH_3$ are more attracted towards nitrogen than towards phosphorus in $PH_3$. Therefore, the extent of hydrogen bonding in $PH_3$ is very less as compared to $NH_3$. So $NH_3$ forms hydrogen bond but $PH_3$ does not.

**39.** How is nitrogen prepared in the laboratory? Write the chemical equations of the reactions involved.

**Solution:**

When an aqueous solution of ammonium chloride ($NH_4Cl$) is heated with sodium nitrite ($NaNO_2$), ammonium nitrite ($NH_4NO_2$) is obtained, which is unstable. This ammonium nitrite further breaks down into nitrogen and water.

$$NH_4Cl(aq) + NaNO_2(aq) \rightarrow NH_4NO_2 + NaCl(aq)$$

$$NH_4NO_2 \rightarrow N_2(g) + H_2O(l)$$

Minute amounts of $NO$ and $HNO_3$ are also produced in the reaction. These impurities can be removed by passing nitrogen gas through aqueous sulphuric acid containing potassium dichromate.

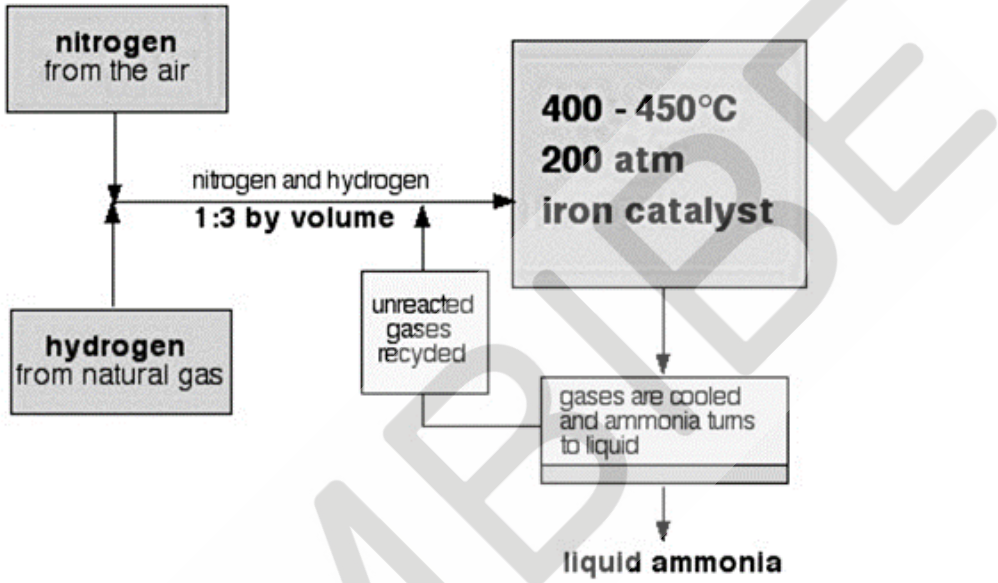
**40.** How is ammonia manufactured industrially?

**Solution:**

Ammonia is manufactured on a large scale industrially by Haber's process. The Haber's process combines nitrogen gas from the air with hydrogen derived from natural gas (methane) in the ratio of 1: 3 into ammonia.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

The reaction is reversible, and the production of ammonia is exothermic in nature.




The optimum conditions for the manufacturing of ammonia are a pressure of 200 atm, the temperature of $400 - 450°C$ and iron oxide catalyst with a small amount of $Al_2O_3$ and $K_2O$ acting as promoters.

**41.** Illustrate how copper metal can give different products on reaction with $HNO_3$.

**Solution:**

Copper metal on reaction with nitric acid ($HNO_3$) gets oxidised. The products depend on the temperature, concentration of the acid and the copper metal undergoing oxidation.

The reaction of copper with dilute and concentrated $HNO_3$ is as shown below:

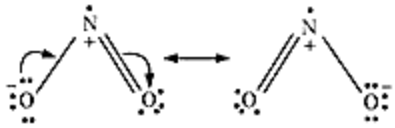
$3Cu + 8HNO_3(dil.) \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

$Cu + 4HNO_3(conc.) \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

**42.** Give the resonating structures of $NO_2$ and $N_2O_5$.

**Solution:**

(1)



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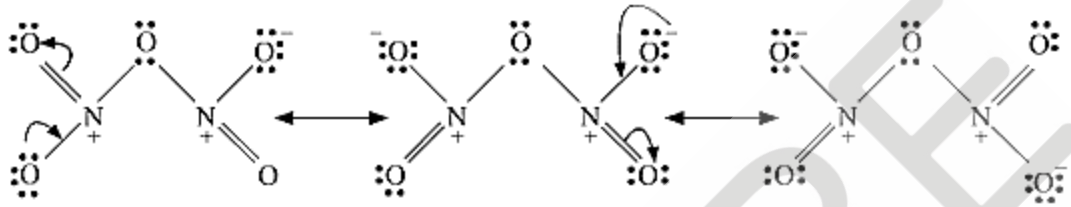


**43.** The HNH angle value is higher than HPH, HAsH and HSbH angles. Why?

[Hint: Can be explained on the basis of $sp^3$ hybridisation in $NH_3$ and only s– p bonding between hydrogen and other elements of the group].

**Solution:**

N has the highest electronegativity among the group 15 elements. Due to this reason, there is high electron density around N. This leads to repulsion between the electron pairs around N atom, causing high HNH angle value. As we move down the group, the electronegativity of the elements decreases and the bond angle also decreases due to lesser repulsion on the cental atom.

**44.** Why does $R_3P = O$ exist but $R_3N = O$ does not (R = alkyl group)?

**Solution:**

Nitrogen lacks a d-orbital to expand its octet. Thus, it cannot have a coordination number greater than four. But, phosphorous has vacant d-orbital and can extend its octet and form $R_3P = O$. Therefore, $R_3P = O$ exist but $R_3N = O$ does not.

**45.** xplain why $NH_3$ is basic while $BiH_3$ is only feebly basic.

**Solution:**

The atomic size order of the group 15 elements is

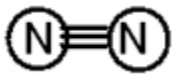
$N < P < As < Sb < Bi$.

Nitrogen has a small size. Due to a small size, there is a very high electron density around Nitrogen. Whereas in case of Bismuth, the size of the atom is greater than nitrogen, and the charge gets distributed over a large area, thereby decreasing the electron density. Therefore, nitrogen atom can easily release electrons as compared to Bismuth. We know that higher the electron donating tendency, higher is the basic strength. Therefore, $NH_3$ is basic while $BiH_3$ is only feebly basic.

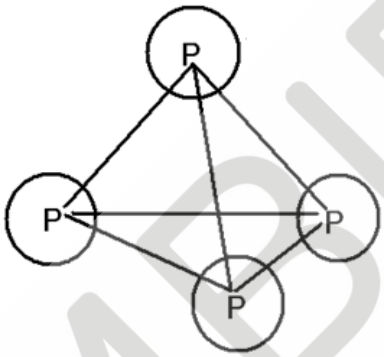
**46.** Nitrogen exists as a diatomic molecule and phosphorus as $P_4$. Why?

**Solution:**

Nitrogen owing to its small size tends to form pπ – pπ multiple bonds with itself. Nitrogen thus forms a very stable diatomic molecule, $N_2$.




When we from top to bottom in a particular group, the tendency to form pπ–pπ bonds decreases due to the larger size of the atoms of elements. With the increase in the size of atoms, the strength of pπ– pπ bonds decreases. Therefore, pπ– pπ bonds are weaker for heavier atoms. Hence, phosphorus (heavier element) exists in the $P_4$ state.




**47.** Write the main differences between the properties of white phosphorus and red phosphorus.

**Solution:**

| White phosphorus | Red phosphorus |
|---|---|
| It is a soft and waxy solid. It has a garlic smell. | It is a hard and crystalline solid, without any smell. |
| It is translucent waxy solid and acquires pale yellow colour when exposed to light. | It is a red solid |
| It is insoluble in water but soluble in carbon disulphide and oils. | It is insoluble in both water as well as carbon disulphide |
| It is highly reactive and undergoes spontaneous combustion in air. | It is relative more stable and less reactive than white phosphorus. |

**48.** Why does nitrogen show catenation properties less than phosphorus?

**Solution:**

Catenation is the property of bonding of atoms of the same element into a series to form a chain. Catenation is shown by phosphorous more than nitrogen. This is due to the reason that the N – N single bond is relatively weaker as compared to the P–P single bond. The nitrogen atom is small. So, there is greater repulsion of

electron density of two nitrogen atoms, thereby weakening the $N-N$ single bond. Phosphorous has a larger atomic size, so it has less repulsion of electron density of two phosphorous atoms. Therefore, the $P-P$ single bond is stronger.

**49.** Give the disproportionate reaction of $H_3PO_3$.

**Solution:**

Orthophosphorus acid ($H_3PO_3$) disproportionates to give orthophosphoric acid ($H_3PO_4$) and phosphine ($PH_3$).

Oxidation State: +3 +5 -3

$$4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$$

**50.** Can $PCl_5$ act as an oxidising as well as a reducing agent? Justify.

**Solution:**

Phosphorous Pentachloride can only act as an oxidising agent. $PCl_5$ cannot act as a reducing agent because in the $PCl_5$, the oxidation state of phosphorus is $+5$, and it is not possible for phosphorus to increase its oxidation state beyond $+5$, whereas it can easily act as an oxidising agent because it can decrease its oxidation state from $+5$ to oxidation state $\leq +3$.

**51.** Justify the placement of O, S, Se, Te and Po in the same group of the periodic table in terms of electronic configuration, oxidation state and hydride formation.

**Solution:**

The elements $O, S, Se, Te$ and $Po$, belong to group 16 and are collectively called chalcogens.

(i) Elements of group 16 have six valence electrons each. The general electronic configuration of these elements is $ns^2\ np^4$, where the n value varies from 2 to 6.

(ii) Oxidation state: As chalcogens have six valence electrons [$ns^2\ np^4$], they should display an oxidation state of $-2$. But only oxygen predominantly shows the oxidation state of $-2$ due to its high electronegativity. It also exhibits the oxidation state of $-1$ [in $H_2O_2$], zero [in $O_2$], and $+2$ [in $OF_2$]. However, the stability of the $-2$ oxidation state decreases on moving down a group because of the decrease in the electronegativity of the elements. The heavier elements of the group show an oxidation state of $+2, +4$, and $+6$ due to the availability of d $-$orbitals.

(iii) Formation of hydrides: Chalcogens form hydrides of formula $H_2E$, where $E = O, S, Se, Te, Po$. Elements oxygen and sulphur also form hydrides of type $H_2E_2$.

**52.** Why is dioxygen a gas but sulphur a solid?

**Solution:**

The oxygen atom is smaller in size as compared to sulphur. Because of its small size, it can effectively form $p\pi - p\pi$ bonds and form $O_2$ molecule. Also, the intermolecular forces in oxygen are weak van der Waal forces. Hence oxygen molecule exists as a gas.

Whereas, sulphur does not form $S_2$ molecule but exists as a puckered structure held together by strong covalent bonds. Hence, it occurs as a solid.

**53.** Knowing the electron gain enthalpy values for $O \rightarrow O^-$ and $O \rightarrow O^{2-}$ as $-141$ and $702\ kJ\ mol^{-1}$ respectively, how can you account for the formation of a large number of oxides having $O^{2-}$ species and not $O^-$?

(**Hint**: Consider lattice energy factor in the formation of compounds).

**Solution:**

Stability of an ionic compound depends on its lattice energy. More the lattice energy of a compound, more stable it will be, and the lattice energy is directly proportional to the charge carried by an ion. When a metal combines with oxygen, the lattice energy of the oxide involving $O^{2-}$ ion is much more than the oxide involving $O^-$ ion. Hence, the oxide having $O^{2-}$ ions are relatively more stable than oxides having $O^-$. Therefore, the formation of $O^{2-}$ is more favorable that the formation of $O^-$.

**54.** Which of the aerosol is responsible for the depletion of ozone?

**Solution:**

Freons or chlorofluorocarbons (CFCs) are the aerosols that are responsible for the depletion of ozone. In the presence of UV rays, the CFCs break down to give chlorine free radicals which combine with the ozone atoms present in the atmosphere to give oxygen atoms.

**55.** Describe the manufacture of $H_2SO_4$ by contact process?

**Solution:**

Sulphuric acid is manufactured on a large scale by the contact process. It involves the following three steps:

**Step (I):** Sulphur or sulphide ores are burnt in air to form $SO_2$.

**Step (II)** $SO_2$ is converted into $SO_3$ in the presence of $V_2O_5$ as a catalyst.

**Step (III):** $SO_3$ produced in the above step is absorbed on $H_2SO_4$ to give $H_2S_2O_7$ (Oleum). This Oleum is then diluted to obtain $H_2SO_4$ of the desired concentration.

The second step is the key step. It is very crucial to maximise the yield of $H_2SO_4$.

$$2SO_2(g)+O_2(g) \overset{V_2O_5}{\rightarrow} 2SO_3(g) \qquad \Delta_r H^0 = -196.6\ kJmol^{-1}$$

The above reaction is exothermic, reversible, and the forward reaction leads to a decrease in volume. So, in accordance with Le Chatelier's principle, low temperature and high pressure are the favourable conditions for maximum yield. But the temperature should not be too low since the rate of reaction will become slow.

The optimum conditions in the contact process to maximise the yield of sulphuric acid in an industrial plant is a pressure of 2 bar and temperature around 720 K.

$$H_2SO_4 + SO_3 \rightarrow \underset{\text{Oleum}}{H_2S_2O_7}$$

$$H_2S_2O_7 + H_2O \rightarrow \underset{\text{Sulphuric acid}}{2H_2SO_4} .$$

**56.** How is $SO_2$ an air pollutant?

**Solution:**

Sulphur dioxide causes harm to the environment in many ways:

1. It combines with water vapour present in the atmosphere to form sulphuric acid. This leads to acid rain. Acid rain damages soil, plants, and buildings, especially those made of marble.
2. $SO_2$ causes irritation in the respiratory tract even when it is present in low concentrations. It also causes throat and eye irritation and can also affect the larynx to cause breathlessness.
3. It is very harmful to plants. Plants exposed to sulphur dioxide for a long time lose colour from their leaves. This condition is known as chlorosis. This happens because the formation of chlorophyll is affected by the presence of sulphur dioxide.

**57.** Why are halogens strong oxidising agents?

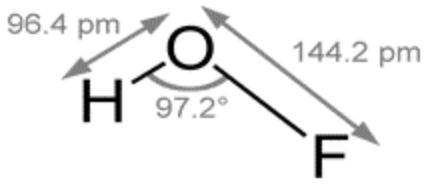
**Solution:**

The general electronic configuration of a halogen is $np^5$, where $n = 2$ to 6. So, a halogen atom needs only one additional electron to complete its octet and to attain the nearest noble gas configuration. Also, halogens are highly electronegative with low dissociation energies and high negative electron gain enthalpies. Therefore, they have a high tendency to gain an electron and they act as strong oxidising agents.

**58.** Explain why fluorine forms only one oxoacid, HOF.

**Solution:**

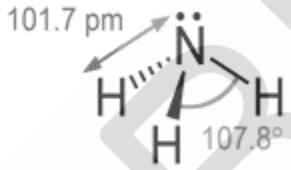
Fluorine forms only one oxoacid, i.e., HOF because of its high electronegativity and small size.



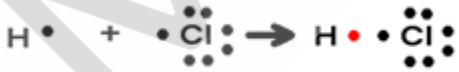

**59.** Explain why inspite of nearly the same Electronegativity, nitrogen forms hydrogen bonding while chlorine does not.

**Solution:**

Both chlorine and nitrogen have almost the same electronegativity values, but chlorine rarely forms hydrogen bonding. This is because, in comparison to chlorine, nitrogen has a smaller size and as a result, a higher electron density per unit volume. Hence nitrogen forms hydrogen bonding more readily.




Whereas, chlorine has a larger atomic size as compared to nitrogen, and has lower electron density per unit volume. Hence chlorine does not readily form hydrogen bonding.




**60.** Write two uses of $ClO_2$.

**Solution:**

Chlorine Dioxide ($ClO_2$) has following uses:

(1) It is used as a bleaching agent in paper pulp and textile industries.

(2) It is used as a disinfectant in sewage and for purification of drinking water.

(3) It is used to control tastes and odours associated with algae and decaying vegetation.

**61.** Why are halogens coloured?

**Solution:**

Most of the halogens are coloured because they absorb radiations which are in the visible region which in turn excite the valence electrons to higher energy levels. As the amount of energy which is required to excite the electrons to a higher level is different for different halogens, each halogen has a different colour.

**62.** Write the reactions of $F_2$ and $Cl_2$ with water.

**Solution:**

When chlorine gas reacts with water it gives Hydrochloric acid and Hypochlorous acid

$Cl_2 + H_2O \rightarrow$ HCl (Hydrochloric Acid) + HOCl (Hypochlorous acid)

Fluorine gas on reaction with water gives Hydrogen ions, Fluorine ions, Oxygen gas and Hydrofluoric acid as products.

$2F_2 + 2H_2O \rightarrow 4H^+ + 4F^- + O_2 + 4HF$

**63.** How can you prepare $Cl_2$ from HCl and HCl from $Cl_2$? Write reactions only.

**Solution:**

(1) $Cl_2$ can be prepared from HCl in presence of $CuCl_2$ by Deacon's process.

$4HCl + O_2 \rightarrow 2Cl_2 + 2H_2O$

(2) HCl can be produced by treating $Cl_2$ with water.

$Cl_2 + H_2O \rightarrow HCl + HOCl$

**64.** What inspired N. Bartlett for carrying out the reaction between Xe and $PtF_6$?

**Solution:**

Neil Bartlett carried out an experiment in which a reaction between oxygen and $PtF_6$ lead to the formation of $O_2^+[PtF_6]^-$ - a red coloured compound

Neil observed that the first ionization energy of both oxygen and xenon were almost same (~1170 kJ/mol). Therefore, he tried to react Xe and $PtF_6$ and was successful to obtain a red coloured compound $Xe^+[PtF_6]^-$.

**65.** What are the oxidation states of phosphorus in the following:

(i) $H_3PO_3$

(ii) $PCl_3$

(iii) $Ca_3P_2$

(iv) $Na_3PO_4$

(v) $POF_3$?

**Solution:**

Let us assume the oxidation state of P be x in all the mentioned compounds:

**Diagram**

(i) $H_3PO_3$

$3 + x + 3(-2) = 0$

$3 + x - 6 = 0$

$x - 3 = 0$

$x = 3$

(ii) $PCl_3$

$x + 3(-1) = 0$

$x - 3 = 0$

$x = 3$

(iii) $Ca_3P_2$

$3(+2) + 2(x) = 0$

$6 + 2x = 0$

$2x = -6$

$x = \frac{-6}{2}$

$x = -3$

(iv) $Na_3PO_4$

$3(+1) + x + 4(-2) = 0$

$3 + x - 8 = 0$

$x - 5 = 0$

$x = 5$

(v) $POF_3$

$x + (-2) + 3(-1) = 0$

$x - 2 - 3 = 0$

$x - 5 = 0$

$x = 5$

**66.** Write balanced equations for the following:

(i) NaCl is heated with sulphuric acid in the presence of $MnO_2$.

(ii) Chlorine gas is passed into a solution of NaI in water.

**Solution:**

(i) $4NaCl + MnO_2 + 4H_2SO_4 \rightarrow MnCl_2 + 4NaHSO_4 + 2H_2O + Cl_2$

Manganese(IV) oxide reacts with sodium chloride and sulfuric acid to produce manganese(II) chloride, chlorine, sodium bisulfate and water.

This reaction takes place at a temperature near $100°C$.

(ii) $Cl_2 + NaI \rightarrow 2NaCl + I_2$

Chlorine reacts with sodium iodide to produce sodium chloride and iodine.

Chlorine - diluted solution.

Sodium iodide - cold solution.

**67.** How are xenon fluorides $XeF_2, XeF_4$ and $XeF_6$ obtained?

**Solution:**

Xe and $F_2$ combine under different conditions to produce $XeF_2, XeF_4, XeF_6$ can be obtained by combining Xe and $F_2$ under different conditions as follows:

| Ratio | Temperature and pressure condition | Reaction |
|---|---|---|
| Excess | At {673 K, 1bar} | $Xe(g) + F_2(g) \rightarrow XeF_2(s)$ |
| 1:5 ratio | At {873 K, 7bar} | $Xe(g) + 2F_2(g) \rightarrow XeF_4(s)$ |
| 1:20 ratio | At {573 K, 60-70bar} | $Xe(g) + 3F_2(g) \rightarrow XeF_6(s)$ |

**68.** With what neutral molecule is $ClO^-$ isoelectronic? Is that molecule a Lewis base?

**Solution:**

$ClO^-$ isoelectronic with ClF as both the compounds contain 26 electrons in all.

Number of electrons in $ClO^-$ : $17 + 8 + 1 = 26$

Number of electrons in ClF : $17 + 9 = 26$

Yes, ClF Molecule is considered as a Lewis base as it accepts electrons from F atom to form $ClF_3$.

**69.** How are $XeO_3$ and $XeOF_4$ prepared?

**Solution:**

(1) $XeO_3$ can be prepared by hydrolysis of $XeF_4$ and $XeF_6$ under controlled pH of the medium in which reaction is taking place. The reaction is:

$6XeF_4 + 12H_2O \rightarrow 4Xe + 2XeO_3 + 24HF + 3O_2$

$XeF_6 + 3H_2O \rightarrow XeO_3 + 6HF$

(2) $XeOF_4$ can be prepared on partial hydrolysis of $XeF_6$ as shown below:

$XeF_6 + H_2O \rightarrow XeOF_4 + 2HF$

**70.** Arrange the following in the order of property indicated for each set:

(i) $F_2, Cl_2, Br_2, I_2$ - increasing bond dissociation enthalpy.

(ii) $HF, HCl, HBr, HI$ - increasing acid strength.

(iii) $NH_3, PH_3, AsH_3, SbH_3, BiH_3$ – increasing base strength.

**Solution:**

(1) Bond dissociation energy usually decreases as we move down the group becuase the atomic size of the element increases.

However, among halogens, the bond dissociation energy of $F_2$ is lower than that of $Cl_2$ due to the small size of the fluorine atom.

Thus increasing order for bond dissociation energy among halogens is as follows:

$I_2 < F_2 < Br_2 < Cl_2$

(2) The bond dissociation energy of $H - X$ molecules (where $X$ is the halogen) decreases with increase in the atomic size. $HI$ is the strongest acid as it loses $H$ atom easily due to weak bonding between $H$ and $I$ atoms.

So Increasing order of acid strength is as follows:

$HF < HCl < HBr < HI$

(3) Basic strength decreases as we move down the group because the size of the atom increases and the electron density of the atom decreases.

So Basic Strength is as follows:

$BiH_3 < SbH_3 < AsH_3 < PH_3 < NH_3$

**71.** Which one of the following does not exist?

(i) $XeOF_4$

(ii) $NeF_2$

(iii) $XeF_2$

(iv) $XeF_6$

**Solution:**

$NeF_2$ does not exist as it would require d-orbital for bonding which Ne lacks as it belongs to period 2. Also, a lot of energy would be used to excite the electrons to a

higher energy level as the valence electrons are very close to the nucleus, thereby experiencing a strong electrostatic attraction. Hence, $NeF_2$ does not exist.

**72.** Give the formula and describe the structure of a noble gas species which is isostructural with:

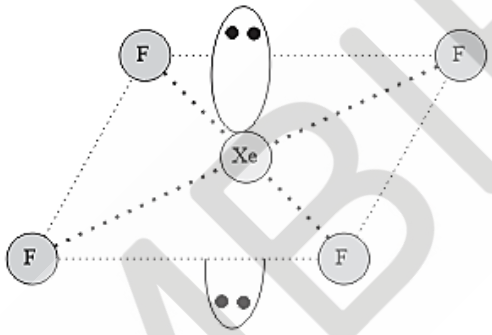
(i) $ICl_4^-$

(ii) $IBr_2^-$

(iii) $BrO_3^-$

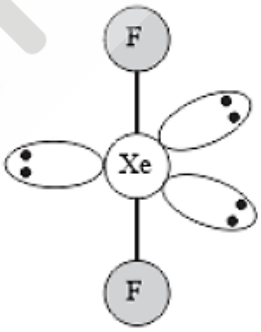
**Solution:**

(i) $XeF_4$ is isoelectronic with $ICl_4^-$. It has square planar geometry, as shown below:



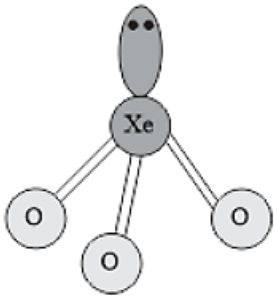

$XeF_4$ : Square planar

(ii) $XeF_2$ is isoelectronic with $IBr_2^-$. And the geometry is linear structure.




$XeF_2$ : Linear

(iii) $XeO_3$ is isoelectronic with $BrO_3^-$. And the geometry is pyramidal structure.




$XeO_3$: Pyramidal structure

**73.** Why do noble gases have comparatively large atomic sizes?

**Solution:**

Noble gases have comparatively large atomic sizes as their shells are completely filled with electrons, and they have a stable configuration due to this completely filled electronic configuration. Electrons tend to be away from each other to reduce the electronic repulsion leading to the larger size of the noble gases.

The atomic radius of the noble gases is determined by van der waal's radius, which is comparatively greater than covalent or ionic radii.

**74.** List the uses of neon and argon gases.

**Solution:**

Uses of Neon gas:

(1) It is used in beacon lights

(2) It is filled in discharge tubes with characteristic colours.

(3) It is mixed with helium to protect electrical appliances from high voltage.

Uses of Argon:

(1) It is used in gas-filled electric lamps due to its inert property.

(2) It is used in laboratories to handle air-sensitive substances.

(3) It is used to provide an inert temperature in a high metallurgical process.

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