

**FIRST YEAR**

**CHEMISTRY**

**Revised Nov, 2022**

*Short Questions and  
Answers*



## ***CHEMISTRY (XI)***

### ***Chapter 1***

#### ***Basic Concepts***

#### ***Short Questions***

***1. Define atom. Give examples.***

***Ans:*** Atom is the smallest particle of an element, which can take part in a chemical reaction. For example, He and Ne, etc. have atoms, which have independent existence while atoms of hydrogen, nitrogen and oxygen do not exist independently.

***2. What are the main points of Dalton's atomic theory?***

***Ans:*** In 1808, an English school teacher, John Dalton, recognized that the law of conservation of matter and the law of definite proportions could be explained by the existence of atoms. He developed an atomic theory the main postulate of which is that all matter is composed of atoms of different elements which differ in their properties.

***3. Define molecule. Give examples.***

***Ans:*** A molecule is the smallest particle of a pure substance which can exist independently. For example, He, Cl<sub>2</sub>, O<sub>3</sub>, P<sub>4</sub>, S<sub>8</sub>, HCl, NH<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>.

***4. Define macromolecules. Give an example.***

***Ans:*** Some molecules are so big that they are called macromolecules. Haemoglobin is such a macromolecule found in blood. It helps to carry oxygen from our lungs to all parts of our body. Each molecule of haemoglobin is made up of nearly 10,000 atoms and it is 68,000 times heavier than a hydrogen atom.

**5. Define ion. Give examples.**

**Ans:** Ions are those species which carry either positive or negative charge. For example,  $\text{Cl}^-$ ,  $\text{Na}^+$ ,  $\text{F}^-$ ,  $\text{K}^+$ .

**6. Define cation. Give an example.**

**Ans:** The positively charged species are called cation. For example, the most common positive ions are formed by the metal atoms such as  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ .

**7. Define anion. Give an example.**

**Ans:** The negatively charged species are called anion. For example, there are many examples of negative ions which consist of group of atoms like  $\text{OH}^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{SO}_4^{2-}$ ,  $\text{PO}_4^{3-}$ ,  $\text{MnO}_4^{1-}$ ,  $\text{Cr}_2\text{O}_7^{2-}$  etc

**8. What are molecular ions? How are they formed? Give their use.**

**Ans:** A molecule may lose or gain an electron to form a molecular ion, e.g.,  $\text{CH}_4^+$ ,  $\text{CO}^+$ ,  $\text{N}_2^+$   
Cationic molecular ions are more abundant than anionic ones. These ions can be generated by passing high energy electron beam or  $\alpha$ -particles or X-rays through a gas. The breakdown of molecular ions obtained from the natural products can give important information about their structure.

**9. Define relative atomic mass.**

**Ans:** Relative atomic mass is the mass of an atom of an element as compared to the mass of an atom of carbon taken as 12. Examples are:

Element	Relative Atomic Mass (amu)	Element	Relative Atomic Mass (amu)
H	1.008	Cl	35.453
O	15.9994	Cu	63.546
Ne	20.1797	U	238.0289

**10. Define isotopes. Give examples.**

**Ans:** Atoms of the same element having same atomic number but different mass number. Such atoms of an element are called isotopes.

**Examples:** H-1, H-2, H-3 called as protium, deuterium and tritium, C-12, C-13, C-14 and O-16, O-17, O-18

**11. The atomic masses may be in fraction. Why?**

**Ans:** The atomic masses are in fraction when elements have isotopes as the average of atomic masses of isotopes will definitely be in fraction. For example,

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100} = 20.18 \text{ Answer}$$

The atomic mass of some of the monoisotopic elements is also in fraction.

**12. Why the isotopes have the same chemical properties but different physical properties?**

**Ans:** The isotopes of the same element have same chemical properties as they have same electronic configuration and chemical properties depend upon the number of electrons. They have different physical properties as they have different number of neutrons in the nucleus and physical properties depend upon the nucleus.

**13. No individual neon atom in the sample of the element has a mass of 20.18 amu. Why?**

**Ans:** Neon has three isotopes of atomic masses 20, 21 and 22 with relative abundances as 90.92%, 0.26% and 8.82%

The relative atomic mass of neon comes out to be 20.18 amu. So, 20.18 amu is the average atomic mass of all the three isotopes of neon.

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100} = 20.18 \text{ Answer}$$

**14. Comment on how atomic number and atomic mass define stability of isotopes.**

**Ans:** The elements of odd atomic number almost never possess more than two stable isotopes. The elements of even atomic number usually have larger number of isotopes and isotopes whose mass numbers are multiples of four are particularly abundant. For example,  $^{16}\text{O}$ ,  $^{24}\text{Mg}$ ,  $^{28}\text{Si}$ ,  $^{40}\text{Ca}$  and  $^{56}\text{Fe}$  form nearly 50% of the earth's crust. Out of 280 isotopes that occur in nature, 154 have even mass number and even atomic number.

**15. What are mono-isotopic elements?**

**Ans:** The elements with only one isotope are called mono-isotopic elements. Examples, arsenic, fluorine, iodine and gold.

**16. Define mole. Give examples.**

**Ans:** When the atomic mass of an element, molecular mass of a compound, formula mass of an ionic compound and ionic mass of ionic specie is expressed in grams then it is called mole.

OR

The atomic mass, molecular mass, formula mass or ionic mass of the substance expressed in gram is called molar mass of the substance.

**Examples:** 1 gram atom of hydrogen = 1.008 g

1 gram molecule of water = 18 g

1 gram formula of NaCl = 58.5 g

1 gram ion of  $\text{OH}^-$  = 17 g

**17. Define mass spectrometry.**

**Ans:** Mass spectrometer is an instrument which is used to measure the exact masses of different isotopes of an element.

**18. Mention the types of mass spectrometer.**

**Ans:** Following are the types of mass spectrometer:

**Aston's mass spectrograph**

Designed to identify the isotopes of an element on the basis of their atomic masses.

**Dempster's mass spectrometer**

Designed for the identification of elements which were available in the solid state.

**19. Name the steps involved in the working of mass spectrometer.**

**Ans:** Following are the steps involved:

1. The substance whose analysis for the separation of isotopes is required is converted into the vapour state. The pressure of these vapours is kept very low, that is,  $10^{-6}$  to  $10^{-7}$  torr.
2. These vapours are allowed to enter the ionization chamber where fast moving electrons are thrown upon them. The atoms of isotopic element present in the form of vapours are ionized.
3. When a potential difference (E) of 500-2000 volts is applied between perforated accelerating plates, then these positive ions are strongly attracted towards the negative plate. In this way, the ions are accelerated.
4. These ions are then allowed to pass through a strong magnetic field of strength (H), which will separate them on the basis of their (m/e) values. Actually, the magnetic field makes the ions to move in a circular path. The ions of definite m/e value will move in the form of groups one after the other and fall on the electrometer.

**20. Write functions of  $Mg(ClO_4)_2$  and  $KOH$  in combustion analysis.**

**Ans: Function of  $Mg(ClO_4)_2$ :**-Magnesium perchlorate acts as a dehydrating agents so it absorbs water during combustion analysis.

**Function of  $KOH$ :**-Potassium hydroxide has the ability to absorb carbon dioxide so it is used to absorb  $CO_2$  produced during combustion analysis.

**21. Give the mathematical relationship used to calculate mass to charge ratio.**

**Ans:** The mathematical relationship for (m/e) is

$$m/e = H^2 r^2 / 2E$$

Where H is the strength of magnetic field, E is the strength of electrical field, r is the radius of circular path.

**22. What does the mathematical relationship express in mass spectrometry?**

**Ans:** The mathematical relationship for (m/e) is:

$$m/e = H^2 r^2 / 2E$$

Where H is the strength of magnetic field, E is the strength of electrical field, r is the radius of circular path. If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/e will fall at a different place as compared to the first place.

**23. What happens if strength of the electric field is increased in mass spectrometry?**

**Ans:** If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/e will fall at a different place as compared to the first place. This can also be done by changing the magnetic field.

**24. What is the function of an electrometer?**

**Ans:** Electrometer is also called an ion collector and develops the electrical current. The strength of the current thus measured gives the relative abundance of ions of a definite  $m/e$  value. Similarly, the ions of other isotopes having different masses are made to fall on the collector and the current strength is measured. The current strength in each case gives the relative abundance of each of the isotopes. The same experiment is performed with C-12 isotope and the current strength is compared.

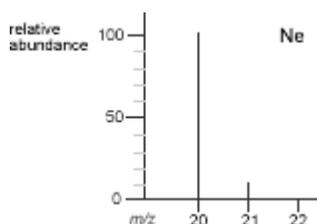
**25. Mention the methods which can be used for the separation of isotopes.**

**Ans:** Following are the methods used for the separation of isotopes:

1. Gaseous diffusion
2. Thermal diffusion
3. Distillation
4. Ultracentrifuge
5. Electromagnetic separation
6. Laser separation

**26. What is mass spectrum?**

**Ans:** The mass spectrum is a graph showing the relative abundance of isotopes plotted against the mass number.



**27. Mention the formula to find percentage of element.**

**Ans:** Following is the formula to calculate percentage of an element:

$$\text{Percentage of an element} = \frac{\text{Mass of the element in the compound}}{\text{mass of the compound}} \times 100$$

**28. How can the percentage of an element be determined by formula mass?**

**Ans:** Percentage composition of a compound can also be determined theoretically if we know the formula mass of the compound. The following equation can be used for this purpose:

$$\text{Percentage of an element} = \frac{\text{Mass of the element in one mole of the compound}}{\text{Formula mass of the compound}} \times 100$$

**29. Define empirical formula.**

**Ans:** It is the simplest formula that gives the small whole number ratio between the atoms of different elements present in a compound. In an empirical formula of a compound,  $A_xB_y$ , there are  $x$  atoms of an element A and  $y$  atoms of an element B. The empirical formula of glucose ( $C_6H_{12}O_6$ ) is  $CH_2O$  and that of benzene ( $C_6H_6$ ) is  $CH$ .

**30. Write the steps involved in the determination of empirical formula?**

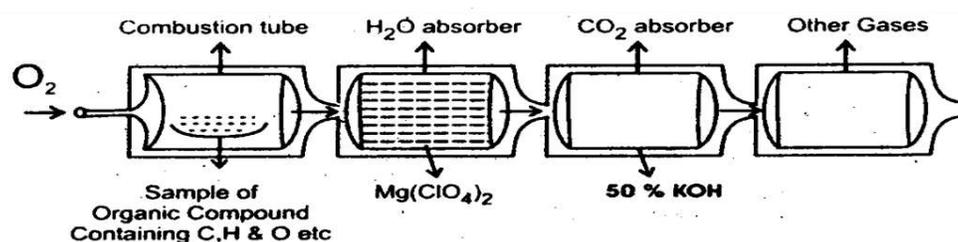
**Ans:** Empirical formula of a compound can be calculated following the steps mentioned below:

1. Determination of the percentage composition.
2. Finding the number of gram atoms of each element. For this purpose divide the mass of each element (% of an element) by its atomic mass.
3. Determination of the atomic ratio of each element. To get this, divide the number of moles of each element (gram atoms) by the smallest number of moles.
4. If the atomic ratio is simple whole number, it gives the empirical formula, otherwise multiply with a suitable digit to get the whole number atomic ratio.

**31. Mention the steps involved in combustion analysis.**

**Ans:** Following are the steps involved in combustion analysis:

1. A weighed sample of the organic compound is placed in the combustion tube. This combustion tube is fitted in a furnace. Oxygen is supplied to burn the compound.
2. Hydrogen is converted to H<sub>2</sub>O and carbon is converted to CO<sub>2</sub>.
3. These gases are absorbed in Mg(ClO<sub>4</sub>)<sub>2</sub> and 50% KOH, respectively.
4. The difference in the masses of these absorbers gives us the amounts of H<sub>2</sub>O and CO<sub>2</sub> produced.
5. The amount of oxygen is determined by the method of difference.



**32. Mention the formulae used in combustion analysis.**

**Ans:** Following formulae are used in combustion analysis:

$$\% \text{ of carbon} = \frac{\text{Mass of CO}_2}{\text{Mass of organic compound}} \times \frac{12.00}{44.00} \times 100$$

$$\% \text{ of hydrogen} = \frac{\text{Mass of H}_2\text{O}}{\text{Mass of organic compound}} \times \frac{2.016}{18} \times 100$$

The percentage of oxygen is obtained by the method of difference.

$$\% \text{ of oxygen} = 100 - (\% \text{ of carbon} + \% \text{ of hydrogen}) .$$

**33. Define molecular formula.**

**Ans:** That formula of a substance which is based on the actual molecule is called molecular formula. It gives the total number of atoms of different elements present in the molecule of a compound. For example, molecular formula of benzene is C<sub>6</sub>H<sub>6</sub> while that of glucose is C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>.

**34. Give the formula relating empirical formula with molecular formula.**

**Ans:** Following is the formula relating empirical and molecular formula:

$$\text{Molecular formula} = n (\text{Empirical formula})$$

Where 'n' is a simple integer.

**35. Mention the compounds having same empirical and molecular formula.**

**Ans:** Those compounds whose empirical and molecular formulae are the same are numerous. For example, H<sub>2</sub>O, CO<sub>2</sub>, NH<sub>3</sub> and C<sub>12</sub>H<sub>22</sub>O<sub>11</sub> have the same empirical and molecular formulas. Their simple multiple 'n' is unity.

**36. What the value of 'n' signifies or tells in a molecular and empirical formula?**

**Ans:** Actually the value of 'n' is the ratio of molecular mass and empirical formula mass of a substance.

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

**37. Define gram atom. Give examples.**

**Ans:** The atomic mass of the element expressed in grams is called one gram atom. It is also called one gram mole or simply a mole of that element.

$$\text{Number of gram atoms or moles of an element} = \frac{\text{Mass of an element in grams}}{\text{Molar mass of an element}}$$

For example,

1 gram atom of hydrogen = 1.008 g

1 gram atom of carbon = 12.000 g

1 gram atom of uranium = 238.0 g

**38. Define gram molecule. Give examples.**

**Ans:** The molecular mass of a substance expressed in grams is called gram molecule or gram mole or simply the mole of a substance.

$$\text{Number of gram molecules or moles of a molecular substance} = \frac{\text{Mass of molecular substance in grams}}{\text{Molar mass of the substance}}$$

For example,

1 gram molecule of water = 18.0 g

1 gram molecule of H<sub>2</sub>SO<sub>4</sub> = 98.0 g

1 gram molecule of sucrose = 342.0 g

**39. Define gram formula mass. Give examples.**

**Ans:** The formula unit mass of an ionic compound expressed in grams is called gram formula of the substance.

$$\text{Number of gram formulas or moles of a substance} = \frac{\text{Mass of the ionic substance in grams}}{\text{Formula mass of the ionic substance}}$$

For example,

1 gram formula of NaCl = 58.50 g

1 gram formula of Na<sub>2</sub>CO<sub>3</sub> = 106 g

1 gram formula of AgNO<sub>3</sub> = 170g

**40. Define gram ion. Give examples.**

**Ans:** Ionic mass of an ionic specie expressed in grams is called one gram ion or one mole of ions.

$$\text{Number of gram ions or moles of an species} = \frac{\text{Mass of the ionic species in grams}}{\text{Formula mass of the ionic species}}$$

For example,

$$1 \text{ g ion of OH}^- = 17 \text{ g}$$

$$1 \text{ g ion of SO}_4^{2-} = 96 \text{ g}$$

$$1 \text{ g ion of CO}_3^{2-} = 60 \text{ g}$$

**41. Define Avogadro's number. Give examples.**

**Ans:** Avogadro's number is the number of atoms, molecules and ions in one gram atom of an element, one gram molecule of a compound and one gram ion of a substance, respectively.

**Examples**

1.008 g of hydrogen = 1 mole of hydrogen =  $6.02 \times 10^{23}$  atoms of H

18 g of  $\text{H}_2\text{O}$  = 1 mole of water =  $6.02 \times 10^{23}$  molecules of water

96 g of  $\text{SO}_4^{2-}$  = 1 mole of  $\text{SO}_4^{2-}$  =  $6.02 \times 10^{23}$  ions of  $\text{SO}_4^{2-}$

**42. What is the number of covalent bonds in 8 g of  $\text{CH}_4$ ?**

**Ans:**

1 mole of  $\text{CH}_4$  = 16 g

So,

0.5 mole of  $\text{CH}_4$  = 8 g

No. of molecules =  $0.5 \times 6.02 \times 10^{23}$   
=  $3.01 \times 10^{23}$  molecules

Each molecule has four bonds. So,

The total number of bonds =  $4 \times 3.01 \times 10^{23}$   
=  $12.04 \times 10^{23}$   
=  $1.204 \times 10^{24}$

**43. How Avogadro's number relates to the masses of chemical substances?**

**Ans:** Avogadro's number relates to the atoms, molecules or ions in one gram mole of an element, compound or ion. One gram mole of the substance is the atomic mass, molecular mass or ionic mass expressed in grams. It means that the number of the species is related with the masses of the

species. For example, 23 grams of sodium and 238 grams of uranium have equal number of atoms in them.

$$1 \text{ mole of Na} = 23 \text{ g} = 6.02 \times 10^{23} \text{ atoms}$$

$$1 \text{ mole of U} = 238 \text{ g} = 6.02 \times 10^{23} \text{ atoms}$$

**44. 180 grams of glucose and 342 grams of sucrose have same number of molecules but different number of atoms present in them. Justify it.**

**Ans:**

$$180 \text{ grams of glucose} = 1 \text{ mole}$$

$$342 \text{ grams of sucrose} = 1 \text{ mole}$$

One mole of every substance has equal number of molecules i.e.  $6.02 \times 10^{23}$

One molecule of  $C_6H_{12}O_6$  has number of atoms = 24

$$= 24 \times 6.02 \times 10^{23}$$

$$= \mathbf{1.44 \times 10^{25}}$$

One molecule of  $C_{12}H_{22}O_{11}$  has number of atoms = 45

$$= 45 \times 6.02 \times 10^{23}$$

$$= \mathbf{2.70 \times 10^{25}}$$

**45. Mg atom is twice heavier than that of carbon atom. How?**

**Ans:** The atomic mass of magnesium is 24 g/mol and that of carbon is 12 g/mol. That is why, Mg atom is twice heavier than carbon atom.

**46. Calculate the mass in grams of  $10^{-3}$  moles of water.**

**Ans:** The formula applied is:

$$\text{Mass of water} = \text{No. of moles} \times \text{molar mass}$$

$$= 10^{-3} \times 18$$

$$= 1.8 \times 10^{-2}$$

$$= \mathbf{0.018 \text{ g}}$$

**47. Two grams of  $H_2$ , 16 g of  $CH_4$  and 44 g of  $CO_2$  occupy separately the volumes of 22.414  $dm^3$  at STP although the sizes and masses of molecules of three gases are very different from each other. Give reason.**

**Ans:** One mole of any ideal gas at STP occupies a volume of 22.414  $dm^3$ . The distance between the gas molecules is 300 times greater than the diameter, therefore, the sizes and masses of different gases do not affect the volume. Hence,

$$2 \text{ g of } H_2 = 1 \text{ mole} = 6.02 \times 10^{23} \text{ molecules } 22.4 \text{ dm}^3 \text{ at STP}$$

$$16 \text{ g of } CH_4 = 1 \text{ mole} = 6.02 \times 10^{23} \text{ molecules } 22.4 \text{ dm}^3 \text{ at STP}$$

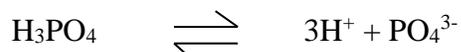
$$44 \text{ g of } CO_2 = 1 \text{ mole} = 6.02 \times 10^{23} \text{ molecules } 22.4 \text{ dm}^3 \text{ at STP}$$

So, they occupy a volume of 22.414  $dm^3$  which is called molar volume ( $V_m$ ).

**48. What is the number of  $H^+$  ions in 9.8 g of  $H_3PO_4$ ?**

**Ans:** 1 mole of  $H_3PO_4 = 98 \text{ g}$

$$0.1 \text{ mole of } H_3PO_4 = 9.8 \text{ g}$$



1 molecule of  $\text{H}_3\text{PO}_4$  has  $3\text{H}^+$ . So, the number of  $\text{H}^+$  ions in 0.1 moles of  $\text{H}_3\text{PO}_4$  is  $1/10^{\text{th}}$  of Avogadro's number multiplied with 3 i.e.  $3 \times 6.02 \times 10^{22} = 1.806 \times 10^{23}$

**49. One mole of  $\text{H}_2\text{SO}_4$  should completely react with 2 moles of  $\text{NaOH}$ . Justify it.**



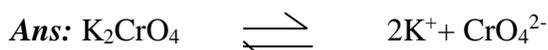
Apply Avogadro's number concept:

1 mole of  $\text{H}_2\text{SO}_4$  generate  $\text{H}^+$  ions =  $2 \times 6.02 \times 10^{23} \text{H}^+$  ions

2 moles of  $\text{NaOH}$  generate  $\text{OH}^-$  ions =  $2 \times 6.02 \times 10^{23} \text{H}^+$  ions

It is clear from the above calculations that the number of  $\text{H}^+$  and  $\text{OH}^-$  ions formed are same although the number of moles of  $\text{H}_2\text{SO}_4$  and  $\text{NaOH}$  are different that is why 1 mole of  $\text{H}_2\text{SO}_4$  reacts completely with 2 moles of  $\text{NaOH}$ .

50. One mg of  $K_2CrO_4$  has thrice the number of ions than the number of formula units when ionized in water. Justify.

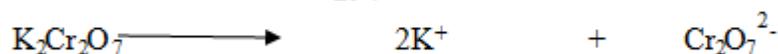


Mass of  $K_2Cr_2O_7 = 1 \text{ mg} = 0.001 \text{ g}$

Molar mass of  $K_2Cr_2O_7 = 294 \text{ gmol}^{-1}$

$$\text{Number of formula units of } K_2Cr_2O_7 = \frac{\text{Mass}}{\text{Molar mass}} \times N_A$$

$$= \frac{0.001 \times 6.02 \times 10^{23}}{294} = 2.04 \times 10^{18}$$



$$2.04 \times 10^{18}$$

$$2 \times 2.04 \times 10^{18}$$

$$2.04 \times 10^{18} = 6.12 \times 10^{18} \text{ ions}$$

Hence, it is justified that total number of ions ( $6.12 \times 10^{18}$  ions) is thrice the number of formula units ionized ( $2.04 \times 10^{18}$ )

51.  $N_2$  and  $CO$  have same number of electrons, protons and neutrons. Justify.

Ans: According to calculations for  $N_2$

Number of electrons in  $N_2 = 7 + 7 = 14$

Number of protons in  $N_2 = 7 + 7 = 14$

Number of neutrons in  $N_2 = 14 - 7 = 7 + 7 = 14$

According to calculations for  $C$  and  $O$

Number of electrons in C = 6

Number of electrons in O = 8

Total number of electrons =  $6+8 = 14$

Number of protons in C = 6

Number of protons in O = 8

Total number of protons =  $6+8 = 14$

Number of neutrons in C = 6

Number of neutrons in O = 8

Total number of neutrons =  $6 + 8 = 14$

Hence,  $N_2$  and CO have same number of electrons, protons and neutrons

**52. Calculate the mass in kg of  $2.6 \times 10^{23}$  molecules of  $SO_2$ .**

**Ans:**

$6.02 \times 10^{23}$  molecules of  $SO_2$  have mass = 64 g

One molecule of  $SO_2$  has mass = 64 g

$6.02 \times 10^{23}$  molecules of  $SO_2$  have mass =  $64/ 6.02 \times 10^{23}$

$2.6 \times 10^{23}$  molecules of  $SO_2$  have mass =  $64 \times 2.6 \times 10^{23}/6.02 \times 10^{23}$

$$= 64 \times 2.6/6.02 = 27.641 \text{ g}$$

**53. One mole of  $H_2O$  has two moles of bonds, three moles of atoms, ten moles of electrons and twenty-eight moles of the total fundamental particles present in it. Justify.**

**Ans:** The formula of water is:



- i. Two moles of covalent bonds
- ii. Total three moles of atoms
- iii. 10 moles of electrons because 8 moles of electrons are contributed by one mole of oxygen and 2 moles of electrons are contributed by 2 moles of hydrogen.
- iv. Number of particles in one mole of oxygen =  $8P + 8n + 8e = 24$  mole particles  
No. of particles in two moles of H atoms = 4 mole particles  
Total no. of particles =  $24 + 4 = 28$  moles

**54. Define stoichiometry.**

**Ans:** Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

**55. What are the basic assumptions of stoichiometry?**

**Ans:** Following are the basic assumptions of stoichiometry:

1. All the reactants are completely converted into products
2. No side reaction occurs
3. Law of conservation of mass and law of definite proportions are obeyed

**56. Which relationships can be studied in stoichiometry?**

**Ans:** Following relationships can be studied in stoichiometry:

### 1) Mass-mass Relationship

If we are given the mass of one substance, we can calculate the mass of the other substances involved in the chemical reaction.

### 2) Mass-mole Relationship or Mole-mass Relationship

If we are given the mass of one substance, we can calculate the moles of other substance and vice versa.

### 3) Mass-volume Relationship

If we are given the mass of one substance we can calculate the volume of the other substances and vice-versa.

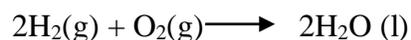
### 4) Mole-mole Relationship

If we are given the moles of one substance we can calculate the moles of the other substances and vice versa.

#### *57. What is a limiting reactant? How it helps to control the reaction?*

**Ans:** The limiting reactant is a reactant that controls the amount of the product formed in a chemical reaction due to its smaller amount. Whenever it is consumed then the further formation of the product stops although excess amount of the other reagent is still in the reaction flask. The unavailability of limiting reactant on consumption stops product formation. This is how it controls the reaction.

For example, consider the reaction between hydrogen and oxygen to form water.



When we take 2 moles of hydrogen (4g) and allow it to react with 2 moles of oxygen (64g), then we will get only 2 moles (36 g) of water. Actually, we will get 2 moles (36g) of water because 2 moles (4g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation. Since less hydrogen is present as compared to oxygen, so hydrogen is a limiting reactant. If we would have reacted 4 moles (8g) of hydrogen with 2 moles (64 g) of oxygen, we would have obtained 4 moles (72 g) of water.

***58. How can the efficiency of a chemical reaction be expressed? OR Why a chemist is interested in calculating efficiency of a reaction?***

***Ans:*** The efficiency of a chemical reaction can be expressed by the percentage yield of the chemical reaction. Percentage yield depends upon the ratio of actual yield and theoretical yield.

$$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Efficiency of a reaction is also called percentage yield.

A chemist is interested in calculating efficiency of a reaction so that the results come in front of him in terms of percentage which tells about the success of a reaction.

**59. Why the experimental yield is mostly less than the theoretical yield?**

**Ans:** Following are the reasons for actual yield to be less than theoretical yield:

1. A practically inexperienced worker has many shortcomings and cannot get the expected yield.
2. The processes like filtration, separation by distillation, separation by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the actual yield.
3. Some of the reactants might take part in a competing side reaction and reduce the amount of the desired product.

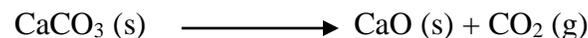
So in most of the reactions the actual yield is less than the theoretical yield.

**60. Distinguish between actual yield and theoretical yield.**

**Ans:** The amount of the product that is obtained in real in a chemical reaction is called actual yield.

The amount of the product calculated from balanced chemical equation is called theoretical yield.

For example,



56 g of CaO is produced according to the balanced chemical equation so it is the theoretical yield.

In real suppose 40 g of CaO is produced so it is the actual yield.

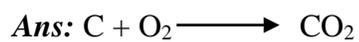
**61. Concept of limiting reactant is not applicable to the reversible reactions. Explain it.**

**Ans:** In case of reversible reactions we cannot understand the concept of limiting reactant as certain amounts of reactants are left unreacted at equilibrium stage. The idea of limiting reactant becomes clear when out of the two reactants one reactant gets consumed completely.

**62. The reaction of combustion in atmosphere consumes O<sub>2</sub> which is in excess. What is limiting reactant?**

**Ans:** In the process of combustion oxygen (O<sub>2</sub>) is always in excess and is left in the atmosphere after complete burning of material. Hence, burning material is limiting reactant.

**63. 11 g of carbon is reacted with 32 g of oxygen to give CO<sub>2</sub>. Which is the limiting reactant?**



According to the balanced equation,

12 g of C reacts with 32 g O<sub>2</sub> to give 44 g CO<sub>2</sub>. Therefore, oxygen is in excess and carbon is the limiting reactant.

**64. Give examples of limiting reactant in daily life.**

**Ans:** Following are the examples of yield in daily life:

1. Burning of wood in which oxygen is in excess and wood is the limiting reactant.
2. The concept of limiting reactant is analogous to the relationship between the number of “kababs” and the “slices” to prepare “sandwiches”. If we have 30 “kababs” and five breads “having 58 slices”, then we can only prepare 29 “sandwiches”. One “kabab” will be extra (excess reactant) and “slices” will be the limiting reactant. It is a practical problem that we cannot purchase exactly sixty “slices” for 30 “kababs” to prepare 30 “sandwiches”.
3. Rusting of iron in which iron is the limiting reactant and moisture and air are in excess.

**65. Mention steps to identify limiting reactant.**

**Ans:** To identify a limiting reactant, the following three steps are performed:

1. Calculate the number of moles from the given amount of reactant.
2. Find out the number of moles of product with the help of a balanced chemical equation.
3. Identify the reactant which produces the least amount of product as limiting reactant.

**66. Why isotopes have same chemical but different physical properties?**

**Ans:** Chemical properties depend upon atomic numbers i.e. arrangement of electrons. Isotopes have the same electronic configuration and as chemical properties depend upon the number of electrons present in outermost shell so isotopes have same chemical properties. Physical properties depend upon atomic masses i.e. number of protons and neutrons so isotopes have different physical properties.

**67. Why oxygen cannot be determined directly in combustion analysis?**

**Ans:** During combustion analysis, an excess of oxygen is provided to make sure that the organic compound must be burnt to produce  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . At the end some of the oxygen is

obtained as extra amount and is not surely given out by the organic compound only. So oxygen cannot be determined directly in combustion analysis. We have to subtract total percentage of C and H from 100.

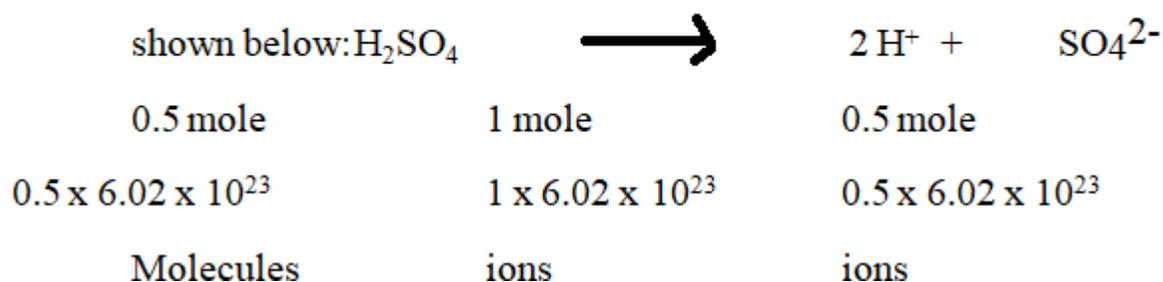
$$\% \text{ age of O} = 100 - (\% \text{ age of C} + \% \text{ age of H})$$

**68. How 4.9 g of  $\text{H}_2\text{SO}_4$  when completely ionized in water have equal number of +ve and -ve charges but the number of positively charged ions is twice the number of negatively charged ions?**

*Ans:*

$$\text{Number of moles of } \text{H}_2\text{SO}_4 = 4.9/98 = 0.5 \text{ mole}$$

0.5 mole of  $\text{H}_2\text{SO}_4$  yields 1 mole of  $\text{H}^+$  ions and 0.5 mole of  $\text{SO}_4^{2-}$  ions as



The relationship shows that total positive charges are equal to total negative charges because each  $\text{SO}_4^{2-}$  ion has -2 charge and two  $\text{H}^+$  have also +2 charge. However, the above relationship shows that number of positive ions are twice as compared to negative ions.

**69. 23g of sodium and 238g of uranium have equal number of atoms in them. Justify.**

**Ans:**

It is justified as follows:

Given mass of Na	=23 g
Atomic mass of Na	=23g/mol
No. of moles of Na	=mass of Na/atomic mass of Na
	=23/23 = 1mole

So, 1 mole of Na contains  $=6.02 \times 10^{23}$  atoms

Given mass of U	=238g
Atomic mass of U	=238g/mol
No. of moles of U	=mass of Na/atomic mass of Na
	=238/238 = 1mole

So, 1 mole of U contains  $=6.02 \times 10^{23}$  atoms

**70. Differentiate between empirical and molecular formula.**

**Ans:**

**Empirical Formula:** - It is the simplest formula that gives the small whole number ratio between the atoms of different elements present in a compound. For example, the empirical formula of glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>) is CH<sub>2</sub>O and that of benzene (C<sub>6</sub>H<sub>6</sub>) is CH.

**Molecular Formula:** - The formula of a substance which is based on the actual molecule is called molecular formula. It gives the total number of atoms of different elements present in the molecule of a compound. For example, the molecular formula of benzene is C<sub>6</sub>H<sub>6</sub> while the molecular formula of glucose is C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>.

**71. A compound may have same molecular and empirical formula. Justify.**

**Ans:** Some compounds have the molecules in which elements are already present in the simplest whole number ratio. So such compounds have the same empirical and molecular formula. For example, empirical and molecular formula for water is H<sub>2</sub>O and for carbon dioxide is CO<sub>2</sub>.

Molecular formula is related with empirical formula by the following relationship:

$$\text{Molecular formula} = n \times \text{Empirical formula}$$

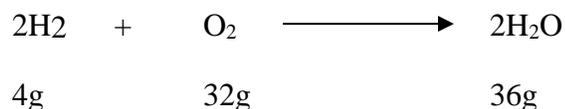
Where 'n' is a simple integer.

The value of the 'n' is the ratio of the molecular mass and empirical formula mass of a substance.

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

**72. Law of conservation of mass must be considered during stoichiometric calculations. How? OR How is law of conservation of mass obeyed during stoichiometric calculations?**

**Ans:** Law of conservation of mass must be obeyed while doing stoichiometric calculations. For example, in the following reaction of hydrogen with oxygen to form water, the mass of reactants must be equal to the mass of the products; for this purpose, the balanced chemical equation is used for stoichiometric calculations.



Here 4 gram of hydrogen reacts with 32 g of oxygen to form 36 gram of water, so law of conservation of mass is obeyed.

**73. Calculate the number of water molecules in 10 g of ice.**

**Ans:**

**Given data:**

Mass of ice (water)=10g  
Molar mass of ice=18g/mol  
No.of molecules of water=?

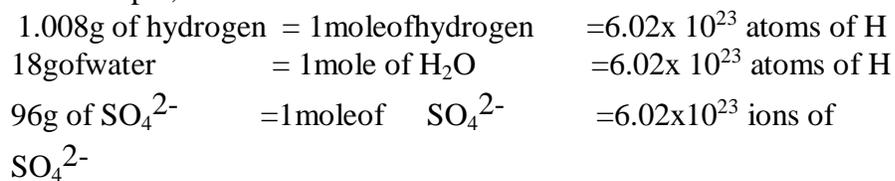
**Solution:**

$$\begin{aligned} n &= \frac{\text{Mass of ice}}{\text{molar mass of ice}} \times \text{NA} \\ n &= \frac{10}{18} \times 6.02 \times 10^{23} \\ &= 3.31 \times 10^{23} \text{ molecules} \end{aligned}$$

**74. What is Avogadro's number? Give equation to relate the Avogadro's number and mass of element.**

**Ans:** Avogadro's number is the number of atoms, molecules and ions in one gram atom of an element, one gram molecule of a compound and one gram ion of a substance, respectively.

For example,



**Equation:**

$$\text{Number of atoms of an element} = \frac{\text{Mass of the element}}{\text{Atomic mass}} \times N_A$$

**75. Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e, Ag-107 and Ag-109. Given the following mass spectrometric data, calculate the average atomic mass of silver.**

Isotopes	Mass(amu)	Percentage abundance
$^{107}\text{Ag}$	106.90509	51.84
$^{109}\text{Ag}$	108.90476	48.16

**Solution**

$$\text{Average Atomic Mass of Ag} = \frac{(\text{Mass of } ^{107}\text{Ag} \times \% \text{ abundance}) + (\text{Mass of } ^{109}\text{Ag} \times \% \text{ abundance})}{100}$$

$$\text{Average Atomic Mass of Ag} = \frac{(106.90509 \times 51.84) + (108.90476 \times 48.16)}{100}$$

$$\text{Average Atomic Mass of Ag} = 107.83 \text{ amu}$$

**76. Calculate the following quantities:**

(a) Mass in grams of 2.74 moles of  $\text{KMnO}_4$

**Solution:**



**Given Information:**

$$\text{Moles of KMnO}_4 = 2.74 \text{ mol}$$

**Requirement:**

Mass in grams of  $\text{KMnO}_4 = ?$

**Formula Applied:**

$$n = \frac{\text{Given mass}}{\text{Molar mass}}$$

**Calculation and Result:**

$$\begin{aligned} \text{Mass of } \text{KMnO}_4 &= \text{No. of moles} \times \text{Molar mass} \\ &= 2.74 \times 158 \\ &= \mathbf{432.9 \text{ g}} \end{aligned}$$

(b) Moles of O atoms in 9.00g of  $\text{Mg}(\text{NO}_3)_2$

**Solution:**  $\text{Mg}(\text{NO}_3)_2$  ( $M = 24 + 28 + 96 = 148 \text{ g mol}^{-1}$ )

**Given information:**

Mass of  $\text{Mg}(\text{NO}_3)_2 = 9 \text{ g}$

Requirement:

Moles of oxygen atoms = ?

**Formula Applied:**

$$n = \frac{\text{Given mass}}{\text{Molar mass}}$$

**Calculation and Result:**

$$\begin{aligned} \text{Moles of } \text{Mg}(\text{NO}_3)_2 &= \frac{9 \text{ g}}{148 \text{ g mol}^{-1}} \\ &= 0.061 \text{ mol} \end{aligned}$$

$$\begin{aligned}
 1 \text{ mole of Mg(NO}_3)_2 \text{ contains oxygen} &= 6 \text{ mol} \\
 0.061 \text{ moles of Mg(NO}_3)_2 \text{ contains oxygen} &= 6 \times 0.061 \\
 &= \mathbf{0.36 \text{ mol}}
 \end{aligned}$$

**(c) Number of O atoms in 10.037 g of CuSO<sub>4</sub>. 5H<sub>2</sub>O**

**Solution:**  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  ( $M=63.5 + 32 + 64 + 90=249.5 \text{ g mol}^{-1}$ )

**Given Information:**

$$\text{Mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 10.037\text{g}$$

**Requirement:**

$$\text{No. of Oxygen atoms} = ?$$

**Formula Applied:**

$$\text{Number of formula units} = \frac{\text{mass} \times \text{Avogadro Number}}{\text{Molar mass}}$$

**Calculation and Result:**

$$\text{Number of formula units of CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{10.037}{249.5} \times 6.02 \times 10^{23}$$

$$= 2.42 \times 10^{22}$$

$$1 \text{ formula unit of CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ contains} = 9 \text{ oxygen atoms}$$

$$2.41 \times 10^{22} \text{ formula units of CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ contains} = 9 \times 2.42 \times 10^{22} \text{ oxygen atoms}$$

$$= 2.17 \times 10^{23} \text{ oxygen atoms}$$

**(d) Mass in grams of 5.136 moles of silver carbonate**

**Solution:**

$$\text{Ag}_2\text{CO}_3 \text{ } M=2(107.87)+ 12+ 48 = 275.74\text{g mol}^{-1}$$

**Given Information:**

$$\text{Moles of Silver Carbonate} = 5.136 \text{ mol}$$

**Requirement:**

Mass in grams of Silver Carbonate = ?

**Formula Applied:**

$$n = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$\text{Mass} = \text{moles} \times \text{molar mass}$$

**Calculation and Result:**

$$\begin{aligned} \text{Mass of Ag}_2\text{CO}_3 &= \text{No. of moles} \times \text{molar mass} \\ &= 5.136 \times 275.74 \\ &= \mathbf{1416.2 \text{ g}} \end{aligned}$$

(e) Mass in grams of  $2.78 \times 10^{21}$  molecules of  $\text{CrO}_2\text{Cl}_2$

$$\text{CrO}_2\text{Cl}_2 \text{ (M)} = 52 + 32 + 71 = 155 \text{ gmol}^{-1}$$

**Given Information:**

$$\text{Molecules of CrCO}_2\text{Cl}_2 = 2.78 \times 10^{21}$$

**Formula Applied:**

$$\text{Number of particles} = \frac{\text{Mass}}{\text{Molar mass}} \times N_A$$

$$\text{Mass} = \frac{\text{Number of particles}}{N_A} \times \text{Molar mass}$$

**Calculation and Result:**

$$\begin{aligned} \text{Mass of CrO}_2\text{Cl}_2 &= \frac{2.78 \times 10^{21} \times 155}{6.02 \times 10^{23}} \\ &= \mathbf{0.715 \text{ g}} \end{aligned}$$

**(f) Number of moles and formula units in 100g of KClO<sub>3</sub>**

**Solution:**  $\text{KClO}_3$

$$M = 39 + 35.5 + 48 = 122.5 \text{ gmol}^{-1}$$

**Given Information:**

$$\text{Mass of KClO}_3 = 100 \text{ g}$$

**Requirement:**

$$\text{Number of moles of KClO}_3 = ?$$

$$\text{Number of Formula units of KClO}_3 = ?$$

**Formula Applied:**

$$n = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$\text{Number of particles} = \text{number of moles} \times N_A$$

**Calculation and Result:****i. Calculation of Number of moles:**

$$\text{Moles of KClO}_3 = \frac{100}{122.5} = 0.816 \text{ mol}$$

**ii. Calculation of Number of formula units:**

$$\begin{aligned} \text{Number of formula units of KClO}_3 &= 0.816 \times 6.02 \times 10^{23} \\ &= 4.91 \times 10^{23} \end{aligned}$$

**(g) Number of K<sup>+</sup>ions, ClO<sup>3-</sup>ion, Cl atoms and O atoms in it**KClO<sub>3</sub>:K<sup>+</sup>ions

1 : 1

$$4.91 \times 10^{23} : 4.91 \times 10^{23}$$

$$\text{Number of K}^+\text{ions} = 4.91 \times 10^{23}$$

KClO<sub>3</sub>: ClO<sup>3-</sup>

1 : 1

$$4.91 \times 10^{23} : 4.91 \times 10^{23}$$

$$\text{Number of ClO}^{3-}\text{ions} = 4.91 \times 10^{23}$$


---

KClO<sub>3</sub>: Chlorine atoms

1 : 1

$$4.91 \times 10^{23} : 4.91 \times 10^{23}$$

$$\text{Number of Chlorine atoms} = 4.91 \times 10^{23}$$


---

KClO<sub>3</sub>: Oxygen atoms

1 : 3

$$4.9 \times 10^{23} : 3 \times 4.9 \times 10^{23}$$

$$: 1.47 \times 10^{24}$$

Number of Oxygen atoms =  $1.47 \times 10^{24}$

**(h) Moles of Cl atoms in 0.822 g C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub>**

**Solution:** C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub> (M = 24 + 4 + 71 = 99 g mol<sup>-1</sup>)

**Given Information:**

Mass of C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub> = 0.822g

**Requirement:**

Moles of Cl atoms = ?

**Calculation and Result:**

99g of C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub> contains moles of Cl = 2

1g of C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub> contains moles of Cl =  $\frac{2}{99}$

0.822 g C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub> contains moles of Cl =  $\frac{2}{99} \times 0.822$   
= **0.017 mol**

**(i) Mass in kilogram of  $2.6 \times 10^{20}$  molecules of SO<sub>2</sub>**

**Solution:** SO<sub>2</sub> (M = 32 + 32 = 64 g mol<sup>-1</sup>)

**Given Information:**

Molecules of SO<sub>2</sub> =  $2.6 \times 10^{20}$

**Requirement:**

Mass in kilogram of SO<sub>2</sub> = ?

**Formula Applied:**

Number of particles =  $\frac{\text{mass}}{\text{molar mass}} \times N_A$

Mass =  $\frac{\text{Number of particles} \times \text{molar mass}}{N_A}$

**Calculation and Result:**

Mass =  $\frac{2.6 \times 10^{20} \times 64}{N_A}$

=  $\frac{6.02 \times 10^{23}}{27.64 \times 10^{-3} \text{ g}}$

=  $27.64 \times 10^{-3} \times 10^{-3} \text{ g}$

$$\begin{aligned} &= 27.64 \times 10^{-6} \text{ g} \\ &= 2.764 \times 10^{-5} \text{ g} \end{aligned}$$

*77. In each pair, choose the larger of the indicated quantity, or state if the samples are equal*

(a) **Individual particles: 0.4 mole of oxygen molecules or 0.4 moles of oxygen atoms.**

**Solution:**

**Given Information:**

Moles of oxygen molecules = 0.4 moles  
Moles of oxygen atoms = 0.4 moles

**Requirement:**

Individual particles (larger in number) = ?

**Calculation and Result:**

0.4 moles of  $O_2$  or 0.4 mol of oxygen atoms are equimolar, so **both have equal number of individual particles.**

$$\begin{aligned} \text{The number of particles} &= \text{Moles} \times N_A \\ &= 0.4 \times 6.02 \times 10^{23} \\ &= 2.4 \times 10^{23} \end{aligned}$$

(b) **Mass: 0.4 moles of ozone molecules or 0.4 moles of oxygen atoms**

**Solution:**

**Given Information:**

Moles of  $O_3$  (ozone) molecules = 0.4

Moles of oxygen atoms = 0.4

**Requirement:**

Larger mass = ?

**Calculation and Result:**

Mass of Ozone( $O_3$ )molecules = moles  $\times$  mol.Mass

$$= 0.4 \times 48 = 19.2 \text{ g}$$

Mass of oxygen atoms = moles  $\times$  molar

$$= \text{mass } 0.4 \times 16$$

$$= 6.4 \text{ g}$$

**Ozone has larger mass**

(c) **Mass: 0.6 moles of C<sub>2</sub>H<sub>4</sub> or 0.6 mole of I<sub>2</sub>**

**Solution:**

**Given Information:**

Moles of C<sub>2</sub>H<sub>4</sub> = 0.6

Moles of I<sub>2</sub> = 0.6

**Requirement:**

Larger mass = ?

**Calculation and Result:**

Mass of C<sub>2</sub>H<sub>4</sub> = moles × mol. Mass

Mass of I<sub>2</sub> = 0.6 × 28 = **16.8g**

= Moles × mol. mass

= 0.6 × 254 = **152.4 g**

**Iodine has larger mass**

(d) **Individual particles: 4.0g N<sub>2</sub>O<sub>4</sub> and 3.3g SO<sub>2</sub>**

**Solution:**

**Given Information:**

Mass of N<sub>2</sub>O<sub>4</sub> = 4g

$$\text{Mass of SO}_2 = 3.3\text{g}$$

**Requirement:**

$$\text{Individual particles (in larger number)} = ?$$

**Calculation and Result:**

$$\text{Mass of N}_2\text{O}_4 = 4\text{g}$$

$$\text{Moles of N}_2\text{O}_4 = \frac{4}{92}$$

$$= 0.043 \text{ mol}$$

$$1 \text{ mole N}_2\text{O}_4 \text{ contains molecules} = 6.02 \times 10^{23}$$

$$0.043 \text{ mole N}_2\text{O}_4 \text{ contains molecules} = 6.02 \times 10^{23} \times 0.043$$

$$= 2.6 \times 10^{22} \text{ molecules}$$

$$\text{Mass of SO}_2 = 3.3\text{g}$$

$$= \frac{3.3}{64}$$

$$= 0.051 \text{ (moles)}$$

$$1 \text{ moles SO}_2 \text{ contains molecules} = 6.02 \times 10^{23}$$

$$0.051 \text{ mole SO}_2 \text{ contains molecules} = 6.02 \times 10^{23} \times 0.051$$

$$= 3.1 \times 10^{22}$$

**SO<sub>2</sub> contains larger number of particles**

**(e) Total ions: 2.3 mol of NaClO<sub>3</sub> or 2.0 mole of MgCl<sub>2</sub>**

**Given information**

$$\text{Moles of NaClO}_3 = 2.3$$

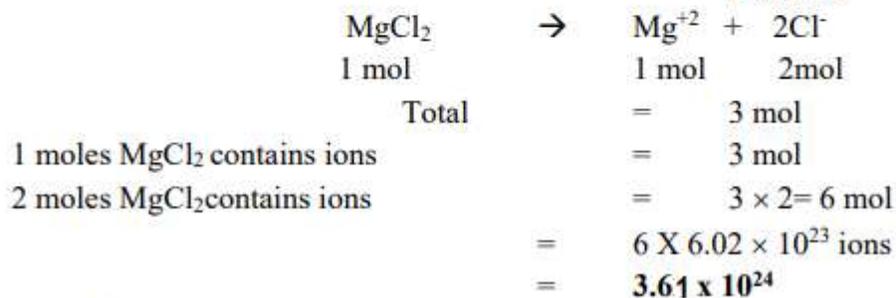
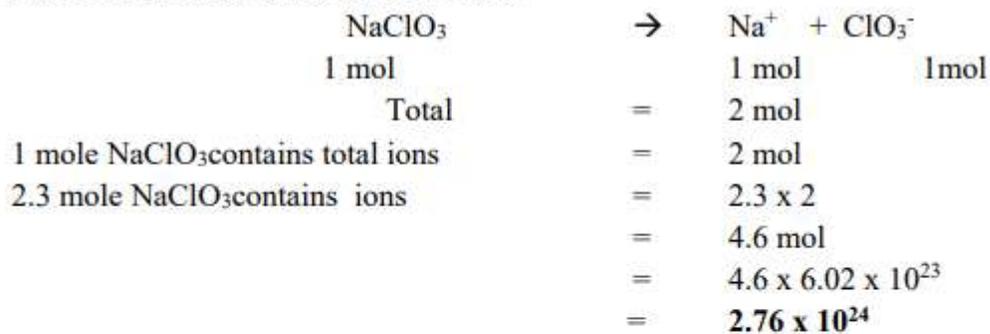
$$\text{Moles of MgCl}_2 = 2.0$$

**Requirement**

$$\text{Number of total ions (in larger number)} = ?$$

**Calculation and Result**

Number of ions in 2.3 moles of  $\text{NaClO}_3$



**$\text{MgCl}_2$  contains larger number of ions**

**(f) Molecules: 11.0 g of  $\text{H}_2\text{O}$  or 11.0 g of  $\text{H}_2\text{O}_2$**

**Given Information:**

Mass of  $\text{H}_2\text{O}$  = 11.0 g

Mass of  $\text{H}_2\text{O}_2$  = 11.0 g

**Requirement:**

No. of Molecules (in larger number) = ?

**Calculation and Result:**

Number of molecules in 11g  $\text{H}_2\text{O}$  = ?

Mass of  $\text{H}_2\text{O}$  = 11g

$$\begin{aligned}
 \text{Moles of H}_2\text{O} &= \frac{11}{18} \\
 &= 0.61 \text{ mol} \\
 \text{Number of molecules of H}_2\text{O} &= \text{No. of moles} \times N_A \\
 &= 0.61 \times 6.02 \times 10^{23} \text{ molecules} \\
 &= \mathbf{3.67 \times 10^{23} \text{ molecules}}
 \end{aligned}$$

Number of molecules in 11g H<sub>2</sub>O<sub>2</sub>

$$\begin{aligned}
 \text{Mass of H}_2\text{O}_2 &= 11 \text{ g} \\
 \text{Moles of H}_2\text{O}_2 &= \frac{11}{34} = 0.32 \text{ mol} \\
 \text{Number of molecules of H}_2\text{O}_2 &= \text{No. of moles} \times N_A \\
 &= 0.32 \times 6.02 \times 10^{23} \\
 &= \mathbf{1.92 \times 10^{23} \text{ molecules}}
 \end{aligned}$$

**H<sub>2</sub>O contains larger number of molecules**

**(g) Na<sup>+</sup> ion: 0.500 moles of NaBr or 0.0145 kg NaCl**

**Given Information:**

$$\begin{aligned}
 \text{Moles of NaBr} &= 0.500 \text{ moles} \\
 \text{Mass of NaCl} &= 0.0145 \text{ kg} \\
 &= 14.5 \text{ g}
 \end{aligned}$$

**Requirement:**

$$\text{Number of Na}^+ \text{ ions} = ?$$

**Calculation and Result:**

$$\begin{aligned}
 1 \text{ moles NaBr contains Na}^+ \text{ ions} &= 6.02 \times 10^{23} \\
 0.5 \text{ mole NaBr contains Na}^+ \text{ ions} &= 6.02 \times 10^{23} \times 0.5 \\
 &= 3.01 \times 10^{23} \text{ Na}^+ \text{ ions} \\
 \text{Mole of NaCl} &= \frac{14.5}{58.5} = 0.248 \text{ moles}
 \end{aligned}$$

$$\begin{aligned}
 1 \text{ mole NaCl contains Na}^+ \text{ ions} &= 6.02 \times 10^{23} \\
 0.248 \text{ mole NaCl contains Na}^+ \text{ ions} &= 0.248 \times 6.02 \times 10^{23} \\
 &= 1.49 \times 10^{23} \text{ Na}^+ \text{ ions}
 \end{aligned}$$

**0.5 moles NaBr contains larger number of Na<sup>+</sup> ions**

(h) Mass:  $6.02 \times 10^{23}$  atoms of U<sup>235</sup> or  $6.02 \times 10^{23}$  atoms of U<sup>238</sup>

**Given Information:**

$$\begin{aligned}
 \text{Atoms of U}^{235} &= 6.02 \times 10^{23} \\
 \text{Atoms of U}^{238} &= 6.02 \times 10^{23}
 \end{aligned}$$

**Requirement:**

$$\text{Mass (larger)} = ?$$

**Calculation and Result:**

Mass of  $6.02 \times 10^{23}$  atoms of U<sup>235</sup>

$$6.02 \times 10^{23} \text{ atoms of U}^{235} = 1 \text{ mole} = 235 \text{ g (molar mass)}$$

$$6.02 \times 10^{23} \text{ atoms of U}^{238} = 1 \text{ mole} = 238 \text{ g (molar mass)}$$

**U<sup>238</sup> has larger mass**

**78. Calculate the percentage of nitrogen in the four important fertilizers**

(i)  $\text{NH}_3$       (ii)  $\text{NH}_2\text{CONH}_2$       (iii)  $(\text{NH}_4)_2\text{SO}_4$       (iv)  $\text{NH}_4\text{NO}_3$

**Solution:**

**Required:**

$$\% \text{ age of Nitrogen} = ?$$

**Formula Applied:**

$$\% \text{age of Nitrogen} = \frac{\text{Mass of Nitrogen in fertilizer}}{\text{Molar mass}} \times 100$$

**(i) NH<sub>3</sub> (Ammonia)**

$$\text{Molar mass of NH}_3 = 14 + 3 = 17\text{gmol}^{-1}$$

$$\begin{aligned} \text{\%age of Nitrogen} &= \frac{\text{Mass of Nitrogen} \times 100}{\text{Molar mass}} \\ &= \frac{14}{17} \times 100 = \mathbf{82.35\%} \end{aligned}$$

**(ii) Urea (NH<sub>2</sub>CONH<sub>2</sub>)**

$$\text{Molar mass of NH}_2\text{CONH}_2 = 60\text{gmol}^{-1}$$

$$\mathbf{\%age\ of\ N = 28/60 \times 100 = 46.67\%}$$

**(iii) (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub> (Ammonium Sulphate)**

$$\text{Molar mass of (NH}_4)_2\text{SO}_4 = 132\text{gmol}^{-1}$$

$$\mathbf{\%age\ of\ N = 28/132 \times 100 = 21.21\%}$$

**(iv) NH<sub>4</sub>NO<sub>3</sub> (Ammonium Nitrate)**

$$\text{Molar mass of NH}_4\text{NO}_3 = 80\text{gmol}^{-1}$$

$$\mathbf{\%age\ of\ N = \frac{28}{80} \times 100 = 35\%}$$

79. Calculate the percentage of nitrogen and phosphorus in each of the following:

- (i)  $\text{NH}_4\text{H}_2\text{PO}_4$                       (ii)  $(\text{NH}_4)_2\text{HPO}_4$                       (iii)  $(\text{NH}_4)_3\text{PO}_4$

Solution:

Required:

$$\% \text{age of Nitrogen \& Phosphorous} = ?$$

- (i)  $\text{NH}_4\text{H}_2\text{PO}_4$  Ammonium hydrogen phosphate

$$\text{Molar mass of } \text{NH}_4\text{H}_2\text{PO}_4 = 115 \text{ gmol}^{-1}$$

$$\% \text{ age of N} = \frac{14}{115} \times 100 = 12.17\%$$

$$\% \text{ age of P} = \frac{31}{115} \times 100 = 26.96\%$$

- (ii)  $(\text{NH}_4)_2\text{HPO}_4$  (Diammonium Hydrogen Phosphate)

$$\text{Molar mass of } (\text{NH}_4)_2\text{HPO}_4 = 132 \text{ g mol}^{-1}$$

$$\% \text{ age of N} = \frac{28}{132} \times 100 = 21.21\%$$

$$\% \text{ age of P} = \frac{31}{132} \times 100 = 23.48\%$$

- (iii)  $(\text{NH}_4)_3\text{PO}_4$  (Ammonium Phosphate)

$$\begin{aligned} \text{Molar mass of } (\text{NH}_4)_3\text{PO}_4 &= 3(14 + 4) + 31 + 4(16) \\ &= 3(18) + 31 + 64 \\ &= 54 + 31 + 64 = 149 \text{ g mol}^{-1} \end{aligned}$$

$$\begin{aligned} \% \text{ age of N} &= \frac{42}{149} \times 100 \\ &= 28.19\% \end{aligned}$$

$$\begin{aligned}\% \text{ age of P} &= \frac{\text{Mass of P}}{\text{Molar mass}} \times 100 \\ &= \frac{31}{149} \times 100 \\ &= \mathbf{20.8 \%}\end{aligned}$$

## ***CHEMISTRY (XI)***

### ***CHAPTER 2 (EXPERIMENTAL TECHNIQUES IN CHEMISTRY)***

#### ***Short Questions***

***1. Define analytical chemistry.***

**Ans:** Analytical chemistry is a branch of chemistry that deals with the complete chemical characterization of a chemical compound.

***2. Define qualitative and quantitative analysis.***

**Ans:**

**Qualitative Analysis:** In qualitative analysis, a chemist is only concerned with the detection or identification of elements present in a compound.

Example: Determination of radicals present in a salt (Salt Analysis)

**Quantitative Analysis:** In quantitative analysis, a chemist is also concerned with the exact amount of elements present in the compound.

Example: Determination of concentration of a solution.

***3. Mention the steps involved in quantitative determination of a sample.***

**Ans:** A complete quantitative determination generally consists of four major steps:

- (i) Obtaining a sample for analysis
- (ii) Separation of the desired constituent
- (iii) Measurement and calculation of results
- (iv) Drawing conclusion from the analysis

***4. Name the various experimental techniques used for the purification of the substances.***

**Ans:** Purification techniques are as follows:

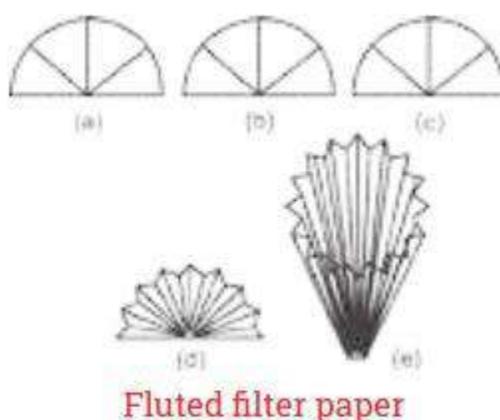
- (i) Filtration (ii) Crystallization (iii) Sublimation (iv) Solvent Extraction (v) Chromatography

***5. Define filtration. Mention various filter media used.***

**Ans:** The process of filtration is used to separate insoluble particles from liquids. It can be performed with several types of filter media. Nature of the precipitate and other factors dictate which filter medium must be used. The most convenient ways of filtration are either through a filter paper or through a filter crucible.

**6. What is the purpose of fluted filter paper? How can it be obtained?**

**Ans:** The rate of filtration through conical funnel can be considerably increased using a fluted filter paper. For preparation of such a paper ordinary filter paper is folded in such a way that a fan like arrangement with alternate elevations and depressions at various folds is obtained.

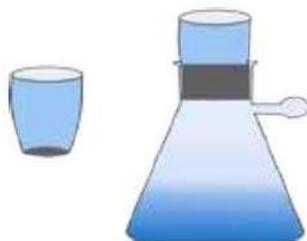


**7. How does the rate of filtration increase by using fluted filter paper?**

**Ans:** In fluted filter paper, filter paper is folded in such a way that a fan like arrangement with alternate elevations and depressions at various folds is obtained. In such a way, contact area or surface area of filter paper is increased due to which rate of filtration is also increased.

**8. Tell about construction and working of Gooch crucible.**

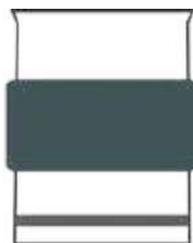
**Ans:** It is made of porcelain having a perforated bottom which is covered with paper pulp or a filter paper cut to its size. Quick filtration can be done by placing the Gooch crucible in a suction filtering apparatus. It is useful for the filtration of precipitates, which need to be ignited at high temperature. If its perforations are covered with asbestos mat then it may be used to filter solutions that react with paper e.g. concentrated HCl and  $\text{KMnO}_4$  solutions.



Gooch Crucible with filtering apparatus

**9. Describe construction and working of sintered glass crucible.**

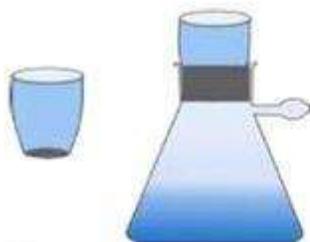
**Ans:** Sintered glass crucible is a glass crucible with a porous glass disc sealed into the bottom. It is very convenient to use because no preparation is needed as with the Gooch crucible.



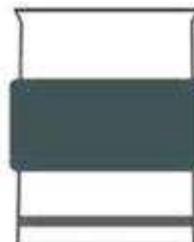
Sintered glass Crucible

**10. What is the difference between Gooch crucible and Sintered glass crucible?**

**Ans:** In Gooch crucible, a filter paper or asbestos mat is needed to cover the perforations of crucible. Whereas in Sintered glass crucible, porous glass is sealed with the bottom and it does not require any filter paper or asbestos mat.



Gooch Crucible with filtering apparatus



Sintered glass Crucible

**11. Why concentrated  $KMnO_4$  and  $HCl$  solutions can't be filtered by Gooch crucible?**

**Ans:**  $KMnO_4$  and  $HCl$  are not filtered by Gooch crucible because both these chemicals can react with filter paper used in Gooch crucible.

**12. Define crystallization.**

**Ans:** Crystallization is the removal of a solid from solution by increasing its concentration above the saturation point in such a manner that the excess solid separates out in the form of crystals.

**13. Write the names of major steps of crystallization.**

**Ans:** Names of the major steps of crystallization are as follows:

- (i) Choice of Solvent
- (ii) Preparation of Saturated Solution
- (iii) Filtration
- (iv) Cooling
- (v) Collecting the Crystals
- (vi) Drying the Crystals
- (vii) Decolourization

**14. Why there is a need to crystallize the crude product?**

**Ans:** Crude product prepared in laboratory may contain soluble and insoluble impurities in it. The insoluble impurities of the product are removed by filtration process but the soluble impurities can be removed by crystallization. So, crude product is crystallized to remove impurities.

**15. Desiccator is the safest method of drying the crystals. Explain.**

**Ans:** Desiccator is slow but safest method of drying, because in this method, crystals preserve their shape and identity. Whereas, if we use other methods of drying, crystals can be crushed or contaminated. Crystals are placed in a vacuum desiccator for several hours. Drying agents that can be used in desiccator are calcium chloride ( $\text{CaCl}_2$ ), silica gel or phosphorous pentoxide ( $\text{P}_2\text{O}_5$ ) etc.

**16. How crystals are dried by safest and reliable method?**

**Ans:** Safest and reliable method for drying of crystals is using vacuum desiccator. Desiccator is slow but safest method of drying, because in this method, crystals preserve their shape and identity. Whereas, if we use other methods of drying, crystals can be crushed or contaminated. Crystals are placed in a vacuum desiccator for several hours. Drying agents that can be used in desiccator are calcium chloride ( $\text{CaCl}_2$ ), silica gel or phosphorous pentoxide ( $\text{P}_2\text{O}_5$ ) etc.

**17. Name the commonly used solvents in crystallization.**

**Ans:** The solvents which are mostly used for crystallization are:

Water

Rectified spirit (95% ethanol)

Absolute ethanol

Diethyl ether

Acetone

Chloroform

Carbon tetrachloride

Acetic acid

Petroleum ether

**18. How decolourization of undesirable colours from crystals is carried?**

**Ans:** Sometimes during the preparation of a crude substance, the coloring matter or resinous products affect the appearance of product and it may appear colored. Such impurities are conveniently removed by boiling the substance in the solvent with the sufficient quantity of finely powdered animal charcoal and then filtering the hot solution. The colored impurities are adsorbed by animal charcoal and the pure decolorized substance crystallizes out from the filtrate on cooling.

**19. What are the characteristics of a good solvent/ideal solvent?**

**Ans:** An ideal solvent should have the following features.

- i. It should dissolve a large amount of the substance at its boiling point and only a small amount at the room temperature.
- ii. It should not react chemically with the solute.
- iii. It should either not dissolve the impurities or the impurities should not crystallize from it along with the solute.

- iv. On cooling it should deposit well-formed crystals of the pure compound.
- v. It should be inexpensive.
- vi. It should be safe to use and should be easily removable.

**20. *How crystallized substances are dried?***

**Ans:** Crystallized substances can be dried by following processes:

- (i) Pressing between several folds of filter paper
- (ii) Drying in an Oven
- (iii) Placing in vacuum desiccator

**21. *Define sublimation. Give examples.***

**Ans:** It is a process in which a solid, when heated, vaporizes directly without passing through the liquid phase and these vapours can be condensed to form the solid again. It is frequently used to purify a solid. Examples of such solids are ammonium chloride, iodine, naphthalene, benzoic acid, etc.

**22. *What type of substances can be purified by sublimation?***

**Ans:** Solid substances can be purified by sublimation. Examples of such solids are ammonium chloride, iodine, naphthalene, benzoic acid etc.

**23. *Define sublimand and sublimate.***

**Ans:** Sublimand is the impure solid substance to be sublimed.

Sublimate is the pure solid substance which is obtained after sublimation of impure solid substance.

**24. *How naphthalene can be purified by the process of sublimation?***

**Ans:** To carry out the process, the substance is taken in a watch glass covered with an inverted funnel. The substance is then heated slowly over a sand-bath and the funnel is cooled with wet cotton. The pure solid deposits on the inner side of the funnel.

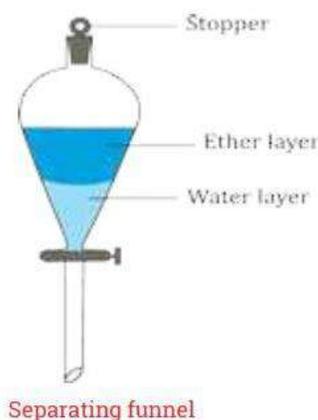


**25. What is solvent extraction?**

**Ans:** Solvent extraction is an important technique in chemical analysis. According to this technique, a solute can be separated from a solution by shaking the solution with a solvent in which the solute is more soluble and the added solvent does not mix with the solution. Usually it is done by placing the solution and the second liquid into a separating funnel. The funnel is stoppered and the two liquids are shaken together.

**26. What is ether extraction?**

**Ans:** Ether extraction is used to separate the products of organic synthesis from water. In a typical organic synthesis, the aqueous solution containing the organic product is shaken up with ether in a separating funnel and allowed to separate. The inorganic impurities remain in aqueous phase whereas the organic compound goes to the ether layer. The ether layer is separated and the organic product is obtained by evaporating the ether.



27. *Why repeated extractions are more efficient than single extraction in solvent extraction?*

OR

*In solvent extraction technique, why repeated extraction using small portions of solvents are more efficient than using a single extraction but larger volume of solvent?*

**Ans:** In solvent extraction technique, repeated extraction using small portions of solvents are more efficient than using a single extraction than larger volume of solvent because organic solvents (such as CCl<sub>4</sub>, CHCl<sub>3</sub>, ether etc.) can absorb approximately half of amount of solute from aqueous solvent. Thus no complete extraction can be done in a single extraction.

28. *What is distribution law or partition law?*

**Ans:** Partition law states that a solute distributes itself between two immiscible liquids in a constant ratio of concentrations irrespective of the amount of solute added.

$$\text{Distribution coefficient (KD)} = \frac{\text{Concentration of solute in organic phase}}{\text{Concentration of solute in aqueous phase}}$$

29. *Iodine is more soluble in water in the presence of KI. Give reason.*

**Ans:** For the dissolution of I<sub>2</sub> in water a dilute solution of KI should be prepared. The molecule I<sub>2</sub> combines with I<sup>-</sup> of KI to generate I<sub>3</sub><sup>-</sup> ion. The ion I<sub>3</sub><sup>-</sup> is soluble in water. I<sub>2</sub> is not soluble in water.



30. *What is stationary phase and mobile phase?*

**Ans:** The stationary phase may be a solid or a liquid supported as a thin film on the surface of an inert solid. The mobile phase flowing over the surface of the stationary phase may be a gas or a liquid.

31. *What is distribution coefficient? To which technique is it applicable?*

**Ans:** Distribution coefficient is the ratio of the amounts of solute dissolved in the immiscible liquids at equilibrium.

$$\text{Distribution coefficient (KD)} = \frac{\text{Concentration of solute in organic phase}}{\text{Concentration of solute in aqueous phase}}$$

This is applicable to solvent extraction and is also applied to chromatography

$$K = \frac{\text{Concentration of the component in the moving phase}}{\text{Concentration of the component in the stationary phase}}$$

**32. What is the effect of distribution coefficient on distribution of solute?**

**Ans:** The component of a mixture with a small value of K mostly remains in the stationary phase as the moving phase flows over it. The component with a greater value of K remains largely dissolved in the mobile phase and passes over the stationary phase quickly.

**33. What is adsorption chromatography and partition chromatography?**

*Or*

**What are the types of chromatography?**

**Ans: Adsorption Chromatography:** Chromatography in which the stationary phase is a solid, is classified as adsorption chromatography. In this type, a substance leaves the mobile phase to become adsorbed on the surface of the solid phase.

**Example:** Thin layer chromatography or column chromatography

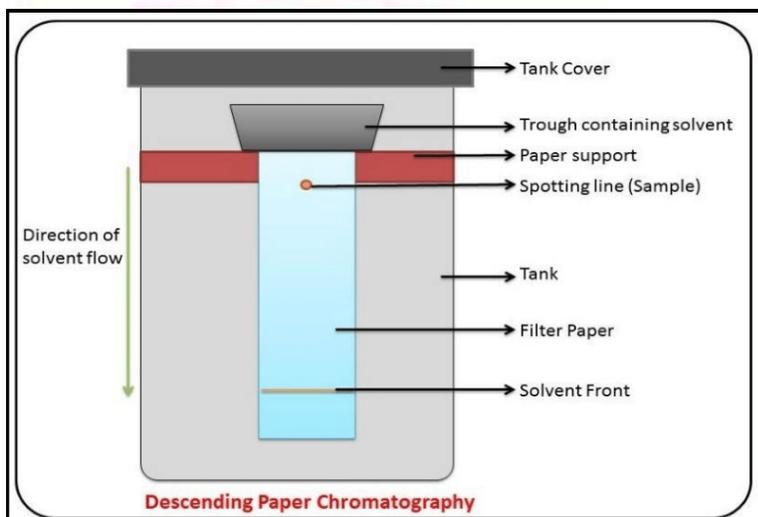
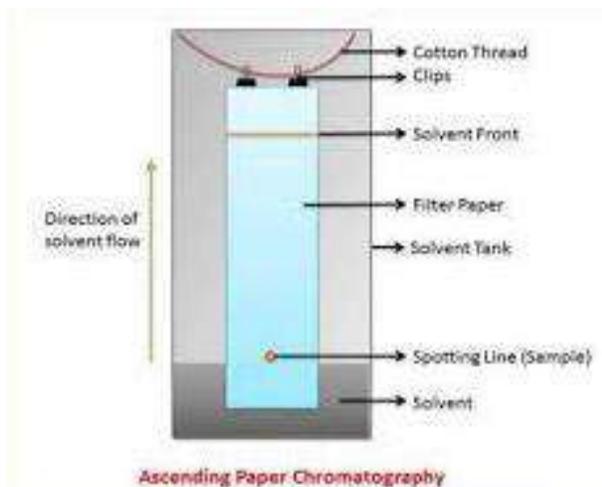
**Partition Chromatography:** Chromatography in which the stationary phase is a liquid, is called partition chromatography. In this type, the substances being separated are distributed throughout both the stationary and mobile phases.

**Example:** Paper chromatography

**34. What are the types of paper chromatography?**

**Ans:** There are three common ways of carrying out paper chromatography namely:

- (i) ascending
- (ii) descending
- (iii) radial/circular



35. *What is  $R_f$ ? Why it has no units?*

**Ans:**  $R_f$  stands for retention factor or retardation factor.

$$R_f = \frac{\text{Distance travelled by a component from the original spot (cm)}}{\text{Distance travelled by solvent from the original spot (cm)}}$$

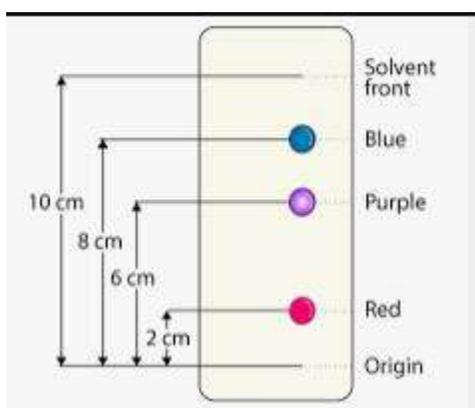
As  $R_f$  is the ratio between two similar quantities it has no units.

**36. What are the uses of chromatography? OR Give any two uses of chromatography.**

**Ans:** The techniques of chromatography are very useful in organic synthesis for separation, isolation and purification of the products. They are equally important in qualitative and quantitative analyses and for determination of the purity of a substance.

**37. What is chromatogram?**

**Ans:** The final filter paper having all the separated components on it is called a chromatogram.



**38. A water insoluble organic compound aspirin is prepared by the reaction of salicylic acid with a mixture of acetic acid and acetic anhydride. How will you separate the product from the reaction mixture?**

**Ans:** The product of the reaction mixture is insoluble in water. It can be separated easily by filtration. The reaction mixture is added to the cold water. Aspirin crystallizes out when dissolved

in cold water while other products are water soluble. So aspirin can be separated from the reaction mixture by filtration.

***CHEMISTRY (XI)***  
***Chapter 3***  
***Gases***  
***Short Questions***

***1. Name the states of matter. Which state is the simplest one?***

**Ans:** Matter exists in four states i.e., solid, liquid, gas and plasma. The simplest form of matter is the gaseous state.

***2. What are the properties of gases? (Write any four properties as an answer to short question)***

**Ans:** Following are the properties of gases:

1. Gases don't have a definite volume and occupy all the available space. The volume of a gas is the volume of the container.
2. They don't have a definite shape and take the shape of the container just like liquids.
3. Due to low densities of gases, as compared to those of liquids and solids, the gases bubble through liquids and tend to rise up.
4. Gases can diffuse and effuse. This property is negligible in solids but operates in liquids as well.
5. Gases can be compressed by applying a pressure because there are large empty spaces between their molecules.
6. Gases can expand on heating or by increasing the available volume. Liquids and solids, on the other hand, do not show an appreciable increase in volume when they are heated.
7. When sudden expansion of gases occurs cooling takes place. It is called Joule-Thomson effect.
8. Molecules of gases are in a constant state of random motion. They can exert a certain pressure on the walls of the container and this pressure is due to the number of collisions.
9. The intermolecular forces in gases are very weak.

***3. What are the properties of liquids? (Write any four properties as an answer to short question)***

**Ans:** Following are the properties of liquids:

1. Liquids don't have a definite shape but have a definite volume. Unlike solids they adopt the shape of the container.
2. Molecules of liquids are in a constant state of motion. The evaporation and diffusion of liquid molecules is due to this motion.
3. The densities of liquids are much greater than those of gases but are close to those of solids.
4. The spaces among the molecules of liquids are negligible just like solids.
5. The intermolecular attractive forces in liquids are intermediate between gases and solids. The melting and boiling points of gases, liquids and solids depend upon the strength of such forces.
6. Molecules of liquids possess kinetic energy due to their motion. Liquids can be converted into solids on cooling i.e., by decreasing their kinetic energy. Molecules of liquids collide among themselves and exchange energy but those of solids cannot do so.

#### ***4. What are the properties of solids?***

**Ans:** Following are the properties of solids:

1. The particles present in solid substances are very close to each other and they are tightly packed. Due to this reason solids are non-compressible and they cannot diffuse into each other.
2. There are strong attractive forces in solids which hold the particles together firmly and for this reason solids have definite shape and volume.
3. The solid particles possess only vibrational motion.

#### ***5. What are the various units of pressure?***

**Ans:** The S.I. unit of pressure is expressed in  $\text{Nm}^{-2}$ . One atmospheric pressure i.e 760 torr is equal to  $101325 \text{ Nm}^{-2}$ .  $1 \text{ pascal} = 1 \text{ Nm}^{-2}$ . So,  $760 \text{ torr} = 101325 \text{ Pa} = 101.325 \text{ kilopascals}$  (kpa is another unit of pressure) The unit pounds per square inch (psi) is used most commonly in engineering work, and  $1 \text{ atm} = 760 \text{ torr} = 14.7 \text{ pounds inch}^{-2}$ . The unit millibar is commonly used by meteorologists.

**6. Define one atmospheric pressure.**

**Ans:** The pressure of air that can support 760 mmHg column at sea level, is called one atmosphere. It is the force exerted by 760mm or 76cm long column of mercury on an area of  $1\text{cm}^2$  at  $0^\circ\text{C}$ . It is the average pressure of atmosphere at sea level  $1\text{mmHg}=1\text{torr}$ .

**7. Define gas laws.**

**Ans:** The relationships between volume of a given amount of gas and the prevailing conditions of temperature and pressure are called the gas laws. Different scientists, like Boyle, Charles, Graham and Dalton have given their laws relating to the properties of gases.

**8. State Boyle's law.**

**Ans:** The volume of a given mass of a gas at constant temperature is inversely proportional to the pressure applied to the gas.

$V \propto 1/P$  (when the temperature and number of moles are constant)

or  $V = k/p$

$PV = k$  (when T and n are constant) (1)

'k' is proportionality constant. The value of k is different for the different amounts of the same gas.

According to the equation (1), Boyle's law can also be defined as

"The product of pressure and volume of a fixed amount of a gas at constant temperature is a constant quantity"

So  $P_1V_1 = k$  and  $P_2V_2 = k$

Hence  $P_1V_1 = P_2V_2$

$P_1V_1$  are the initial values of pressure and volume while  $P_2V_2$  are the final values of pressure and volume.

**9. State Charles's law.**

**Ans:** The volume of the given mass of a gas is directly proportional to the absolute temperature

when the pressure is kept constant.

**$V \propto T$  (when pressure and number of moles are constant)**

$$V = kT$$

$$V/T = k$$

If the temperature is changed from  $T_1$  to  $T_2$  and volume changes from  $V_1$  to  $V_2$ , then

$$V_1/T_1 = k \text{ and } V_2/T_2 = k$$

$$V_1/T_1 = V_2/T_2$$

The ratio of volume to temperature remains constant for same amount of gas at same pressure.

**10. Give quantitative definition of Charles's law. Throw some light on the factor 1/273 in Charles's law.**

**Ans:** At constant pressure, the volume of the given mass of a gas increases or decreases by 1/273 of its original volume at 0 °C for every 1 °C rise or fall in temperature respectively.

$$V_t = V_o(1 + \frac{t}{273}) \dots\dots\dots (3)$$

Where  $V_t$  = volume of gas at temperature T

$V_o$  = Volume of gas at 0°C

t = Temperature on centigrade or celsius scale

**11. Why Celsius scale of temperature does not justify Charles's law.**

**Ans:** The increase in temperature from 10 °C to 100 °C, increases the volume from 566 cm<sup>3</sup> to 746 cm<sup>3</sup>. Applying Charles's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{566}{10} \neq \frac{746}{100}$$

The two sides of equation are not equal. So, Charles's law is not being obeyed when temperature is measured on the Celsius scale.

**12. Why Kelvin scale of temperature justifies Charles's law?**

**Ans:** Charles's law is obeyed when the temperature is taken on the Kelvin scale. For example, at 283 K (10 °C) the volume is 566 cm<sup>3</sup>, while at 373 K (100 °C) the volume is 746 cm<sup>3</sup>.

According to Charles's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} = K$$

$$\frac{566}{283} = \frac{746}{373} = 2 = K$$

**13. Tell about Celsius scale of temperature.**

**Ans:** It has a zero mark for the temperature of ice at one atmospheric pressure. The mark 100 °C indicates the temperature of boiling water at 1 atmospheric pressure. The space between these temperature marks is divided into 100 equal parts and each part is 1 °C.

**14. Tell about Fahrenheit scale of temperature.**

**Ans:** The melting point of ice at 1 atmospheric pressure has a mark 32°F and that of boiling water is 212 °F. The space between these temperature marks is divided into 180 equal parts and each part is 1 °F.

**15. Tell about absolute or Kelvin scale of temperature.**

**Ans:** The melting point of ice at 1 atmospheric pressure is 273K. The water boils at 373K or more precisely at 373.16 K.

$$K = ^\circ C + 273.16$$

**16. Calculate the value of ideal gas constant according to STP.**

**Ans:** The volume of one mole of an ideal gas at STP (one atmospheric pressure and 273.16 K) is 22.414 dm<sup>3</sup>.

Putting these values in the general gas equation will give the value of R.

$$R = \frac{PV}{nT}$$

Putting their values, alongwith units

$$R = \frac{1 \text{ atm} \times 22.414 \text{ dm}^3}{1 \text{ mole} \times 273.16 \text{ K}}$$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

When the pressure is in atmospheres, volume in dm<sup>3</sup> then the value of R used should be 0.0821 dm<sup>3</sup> atm K<sup>-1</sup> mol<sup>-1</sup>

### **Physical meaning**

The physical meanings of this value is that, if we have one mole of an ideal gas at 273.16 K and one atmospheric pressure and its temperature is increased by 1 K, then it will absorb 0.0821 dm<sup>3</sup>-atm of energy, dm<sup>3</sup>-atm is the unit of energy in this situation. Hence, the value of R is a universal parameter for all the gases. It tells us that the Avogadro's number of molecules of all the ideal gases have the same demand of energy.

### ***17. Calculate the value of gas constant in SI units.***

**Ans:** Using SI units of pressure, volume and temperature in the general equation, the value of R is calculated as follows. The SI units of pressure are Nm<sup>-2</sup> and of volume are m<sup>3</sup>. By using Avogadro's principle

$$1 \text{ atm} = 760 \text{ torr} = 101325 \text{ Nm}^{-2}$$

$$1 \text{ m}^3 = 1000 \text{ dm}^3$$

$$n = 1 \text{ mole}$$

$$T = 273.16 \text{ K}$$

$$P = 1 \text{ atm} = 101325 \text{ Nm}^{-2}$$

$$V = 22.414 \text{ dm}^3 = 0.022414 \text{ m}^3$$

Putting their values, alongwith units.

$$R = \frac{PV}{nT} = \frac{101325 \text{ N m}^{-2} \times 0.022414 \text{ m}^3}{1 \text{ mol} \times 273.16 \text{ K}}$$

$$R = 8.3143 \text{ Nm K}^{-1} \text{ mol}^{-1} = 8.3143 \text{ J K}^{-1} \text{ mol}^{-1} \text{ (1 Nm = 1J)}$$

Since 1cal. = 4.18 J

$$\text{so } R = \frac{8.3143}{4.18} = 1.989 \text{ cal K}^{-1} \text{ mol}^{-1}$$

### 18. Calculate density of an ideal gas.

**Ans:** For calculating the density of an ideal gas, we substitute the value of number of moles (n) of the gas in terms of the mass (m), and the molar mass (M) of the gas.

$$n = \frac{m}{M}$$

$$PV = \frac{m}{M} RT$$

The equation is another form of general gas equation that may be employed to calculate the mass of a gas whose P, T, V and molar mass are known.

$$PM = \frac{m}{V} RT$$

$$PM = d RT \quad \left(d = \frac{m}{V}\right)$$

$$d = \frac{PM}{RT}$$

Hence, the density of an ideal gas is directly proportional to its molar mass. Greater the pressure on the gas, closer will be the molecules and greater the density. Higher temperature makes the gases

to expand hence density falls with the increase in temperature. With the help of equation one can calculate the relative molar mass (M) of an ideal gas if its temperature, pressure and density are known.

**19. State Avogadro's law.**

**Ans:** Equal volumes of all the ideal gases at the same temperature and pressure contain equal number of molecules.

$$V \propto n$$

$$V = kn$$

For example,

22.414dm<sup>3</sup> of an ideal gas at STP = 1 mole of gas = 6.02 x 10<sup>23</sup> molecules

**20. State Dalton's law of partial pressures.**

**Ans:** The total pressure exerted by a mixture of non-reacting gases is equal to the sum of their individual partial pressures. Let the gases are designated as 1,2,3 and their partial pressures are p<sub>1</sub>, p<sub>2</sub>, p<sub>3</sub>. The total pressure (P) of the mixture of gases is given by:

$$P_t = p_1 + p_2 + p_3$$

**21. How Dalton's law of partial pressures is applicable to the collection of gases under water?**

**Ans:** Some gases are collected over water in the laboratory. The gas during collection gathers water vapours and becomes moist. The pressure exerted by this moist gas is, therefore, the sum of the partial pressures of the dry gas and that of water vapours.

$$P_{\text{moist}} = P_{\text{dry}} + P_{\text{wvap}}$$

$$P_{\text{moist}} = P_{\text{dry}} + \text{aqueous tension}$$

$$P_{\text{dry}} = P_{\text{moist}} - \text{aqueous tension}$$

The partial pressure exerted by the water vapours is called aqueous tension.

**22. How Dalton's law of partial pressure finds application in breathing process?**

**Ans:** Dalton's law finds its applications during the process of respiration. The process of respiration depends upon the difference in partial pressures. When animals inhale air then oxygen moves into lungs as the partial pressure of oxygen in the air is 159 torr, while the partial pressure of oxygen in the lungs is 116 torr. CO<sub>2</sub> produced during respiration moves out in the opposite direction, as its partial pressure is more in the lungs than that in air.

**23. Why pilots feel uncomfortable breathing at high altitudes?**

**Ans:** At higher altitudes, the pilots feel uncomfortable breathing because the partial pressure of oxygen in the un-pressurized cabin is low, as compared to 159 torr, where one feels comfortable breathing.

**24. Why deep sea divers take oxygen mixed with an inert gas?**

**Ans:** Deep sea divers take oxygen mixed with an inert gas say He and adjust the partial pressure of oxygen according to the requirement. Actually, in sea after every 100 feet depth, the diver experiences approximately 3 atm pressure, so normal air cannot be breathed in depth of sea. Moreover, the pressure of  $N_2$  increases in depth of sea and it diffuses in the blood.

**25. Mention main postulates of kinetic molecular theory.**

**Ans:** Following are the fundamental postulates of this kinetic theory of gases:

1. Every gas consists of a large number of very small particles called molecules. Gases like He, Ne, Ar have mono-atomic molecules.
2. The molecules of a gas move haphazardly, colliding among themselves and with the walls of the container and change their directions.
3. The pressure exerted by a gas is due to the collisions of its molecules with the walls of a container. The collisions among the molecules are perfectly elastic.
4. The molecules of a gas are widely separated from one another and there are sufficient empty spaces among them.

5. The molecules of a gas have no forces of attraction for each other.
6. The actual volume of molecules of a gas is negligible as compared to the volume of the gas.
7. The motion imparted to the molecules by gravity is negligible as compared to the effect of the continued collisions between them.
8. The average kinetic energy of the gas molecules varies directly as the absolute temperature of the gas.

**26. What is mean square velocity?**

**Ans:** All the molecules of a gas under the given conditions don't have the same velocities. Rather different velocities are distributed among the molecules. According to Maxwell's law of distribution of velocities if there are  $n_1$  molecules with velocity  $c_1$ ,  $n_2$  molecules with velocity  $c_2$ , and so on then,

$$\overline{c^2} = \frac{c_1^2 + c_2^2 + c_3^2 + \dots}{n_1 + n_2 + n_3 + \dots}$$

In this reference  $n_1+n_2+n_3+\dots=N$

**27. What is root mean square velocity?**

**Ans:** When we take the square root of mean square velocity then it is called root mean square velocity

$$C_{rms} = \sqrt{\frac{3RT}{M}}$$

Where,  $C_{rms}$  = root mean square velocity

M = molar mass of the gas

T = temperature

**28. Prove Charles's law according to KMT/Prove Boyle's law according to KMT/Prove**

**Avogadro's law according to KMT/Prove Graham's law according to KMT. (Any one can be**

*asked)*

**Ans:**

### **Boyle's Law**

(a) Boyle's Law

According to one of the postulates of kinetic theory of gases, the kinetic energy is directly proportional to the absolute temperature of the gas. The kinetic energy of N molecules is

$$\begin{aligned} & \frac{1}{2} mN\overline{c^2} \\ & \frac{1}{2} mN\overline{c^2} \propto T \\ & \frac{1}{2} mN\overline{c^2} = kT \dots\dots(1) \end{aligned}$$

Where k is the proportionality constant. According to the kinetic equation of gases:

$$PV = \frac{1}{3} mN\overline{c^2}$$

Multiplying and dividing 2 on right hand side

$$PV = \frac{2}{3} \left( \frac{1}{2} mN\overline{c^2} \right) \dots\dots(2)$$

Putting equation (1) in equation (2)

$$PV = \frac{2}{3} kT \dots\dots(3)$$

If the temperature (T) is constant then right hand side of equation (3)  $\frac{2}{3} kT$  is constant. Let that constant be k.

So,  $PV = k$  (which is Boyle's law)

Hence at constant temperature and number of moles the product PV is a constant quantity.

### **(b) Charles's Law**

Consider the following equation:

$$PV = \frac{2}{3} kT$$

$$V = \frac{2}{3P} kT$$

At constant pressure,

Therefore,

$$k'' = \frac{2}{3P} k$$

$$V = k''T$$

$V/T = k''$  (which is Charles's Law)

### (c) Avogadro's Law

Consider two gases 1 and 2 at the same pressure  $P$  and having the same volume  $V$ . Their number of molecules are  $N_1$  and  $N_2$ , masses of molecules are  $m_1$  and  $m_2$  and mean square velocities are  $\overline{c_1^2}$  and  $\overline{c_2^2}$  respectively.

Their kinetic equations can be written as follows:

$$PV = \frac{1}{3} m_1 N_1 \overline{c_1^2} \quad \text{for gas(1)}$$

$$PV = \frac{1}{3} m_2 N_2 \overline{c_2^2} \quad \text{for gas(2)}$$

$$\text{Equalizing } \frac{1}{3} m_1 N_1 \overline{c_1^2} = \frac{1}{3} m_2 N_2 \overline{c_2^2}$$

$$\text{Hence, } m_1 N_1 \overline{c_1^2} = m_2 N_2 \overline{c_2^2} \quad (1)$$

When the temperature of both gases is the same, their mean kinetic energies per molecule will also be same, so

$$\frac{1}{2} m_1 \overline{c_1^2} = \frac{1}{2} m_2 \overline{c_2^2}$$

$$m_1 \overline{c_1^2} = m_2 \overline{c_2^2} \quad (2)$$

Dividing eq (1) by (2)

$$N_1 = N_2$$

Hence equal volumes of all the gases at the same temperature and pressure contain equal number of molecules which is Avogadro's law.

**(d) Graham's Law of Diffusion**

$$PV = \frac{1}{3} mN\overline{c^2} \quad (1)$$

Applying the kinetic equation

$$PV = \frac{1}{3} mN_A\overline{c^2} \quad (1)$$

If we take one mole of a gas having Avogadro's number of molecules ( $N=N_A$ ) then the equation (1) can be written as:

$$PV = \frac{1}{3} M\overline{c^2} \quad (M=mN_A) \quad (2)$$

If we take one mole of a gas having Avogadro's number of molecules ( $N=N_A$ ) then the equation (1) can be written as:

$$PV = \frac{1}{3} M\overline{c^2} \quad (M=mN_A) \quad (3)$$

$$\overline{c^2} = \frac{3PV}{M}$$

Where M is the molecular mass of the gas

Taking square root

$$\begin{aligned} \sqrt{\overline{c^2}} &= \sqrt{\frac{3PV}{M}} \\ \sqrt{\overline{c^2}} &= \sqrt{\frac{3P}{M/V}} = \sqrt{\frac{3P}{d}} \quad \left(\frac{M}{V} = d\right) \end{aligned}$$

'V' is the molar volume of gas at given conditions. Since the root mean square velocity of the gas is proportional to the rate of diffusion of the gas.

$$\begin{aligned} \sqrt{\overline{c^2}} &\propto r \\ r &\propto \sqrt{\frac{3P}{d}} \end{aligned}$$

At constant pressure

$$r \propto \sqrt{\frac{1}{d}}$$

Which is Graham's law.

**29. Define plasma.**

**Ans:** Plasma is the fourth state of matter. Plasma is a distinct state of matter containing a significant number of electrically charged particles a number sufficient to affect its electrical properties and behaviour.

**30. How is plasma formed?**

**Ans:** An electron may gain enough energy to escape its atom. This atom loses one electron and develops a net positive charge. It becomes an ion. In a sufficiently heated gas, ionization happens many times, creating clouds of free electrons and ions. All the atoms are not necessarily ionized and some of them may remain completely intact with no net charge. This ionized gas mixture, consisting of ions, electrons and neutral atoms is called plasma.

**31. Differentiate between natural and artificial plasma.**

**Ans:** Artificial plasma can be created by ionization of a gas as in neon signs. Plasma at low temperatures is hard to maintain because outside a vacuum low temperature plasma reacts rapidly with any molecule it encounters. This aspect makes this material, both very useful and hard to use. Natural plasma exists only at very high temperatures, or low temperature vacuums.

Natural plasma, on the other hand, does not breakdown or react rapidly, but is extremely hot (over 20,000°C minimum). Its energy is so high that it vaporizes any material it touches.

**32. What are the characteristics of plasma?**

**Ans:** Following are the characteristics of plasma:

1. A plasma must have sufficient number of charged particles so as a whole it exhibits a collective response to electric and magnetic fields. The motion of the particles in the plasma generate fields

and electric currents from within plasma density. It refers to the density of the charged particles.

This complex set of interactions makes plasma a unique, fascinating, and complex state of matter.

2. Although plasma includes electrons and ions and conducts electricity, it is macroscopically neutral. In measurable quantities the number of electrons and ions are equal.

**33. *Where is plasma found?***

**Ans:** Entire universe is almost of plasma. It existed before any other forms of matter came into being. Plasmas are found in everything from the sun to quarks, the smallest particles in the universe. Plasma is the most abundant form of matter in the universe. It is the stuff of stars. A majority of the matter in inner-stellar space is plasma. All the stars that shine are all plasma. The sun is a 1.5 million kilometer ball of plasma, heated by nuclear fusion.

On earth it only occurs in a few limited places, like lightning bolts, flames, auroras, and fluorescent lights. When an electric current is passed through neon gas, it produces both plasma and light.

**34. *What are the applications of plasma? (Mention any four applications as an answer to short question)***

**Ans:** Following are the applications of plasma:

1. A fluorescent light bulb is not like regular light bulbs. Inside the long tube is a gas. When the light is turned on, electricity flows through the tube. This electricity acts as that special energy and charges up the gas. This charging and exciting of the atoms creates glowing plasma inside the bulb.

2. Neon signs are glass tubes filled with gas. When they are turned on then the electricity flows through the tube. The electricity charges the gas, possibly neon, and creates plasma inside the tube. The plasma glows with a special colour depending on what kind of gas is inside.

3. They find applications such as plasma processing of semiconductors, sterilization of some medical products, lamps, lasers, diamond coated films, high power microwave sources and pulsed power switches.

4. They also provide the foundation for important potential applications such as the generation of electrical energy from fusion pollution control and removal of hazardous chemicals.
5. Plasma light up our offices and homes, make our computers and electronic equipment work.
6. They drive lasers and particle accelerators, help to clean up the environment, pasteurize foods and make tools corrosion-resistant.

**35. *Why is the Boyle's law applicable only to the ideal gases?***

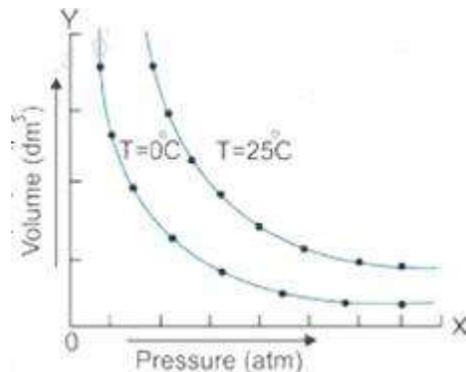
**Ans:** Boyle's law is applicable only to the ideal gases because in ideal gases there are no forces of attraction among the gas molecules.

**36. *When a gas obeys the Boyle's law, the isotherms for the gas can be plotted. How is it true?***

**Ans:** Isotherms are the graphs between pressure and volume at constant temperature and number of moles. This condition is fulfilled by Boyle's law. The word isotherm means "same temperature". They are curves. At higher temperature the curves go away from the axis.

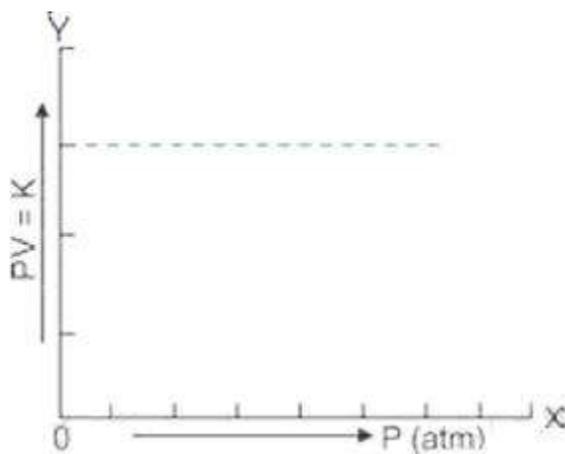
**37. *What are isotherms? What happens to the positions of isotherms when they are plotted at high temperature for a particular gas?***

**Ans:** Isotherms are the graphs between pressure and volume by keeping the temperature constant. Plot a graph at two different temperatures and keeping them constant at different pressure and volume and plot the isotherm. It goes away from both the axes. The reason is that at higher temperature, the volume of the gas has increased. Similarly, if we increase the temperature further, make it constant and plot another isotherm, it further goes away from the axis.



**38. The plot of  $PV$  versus  $P$  is a straight line. Justify.**

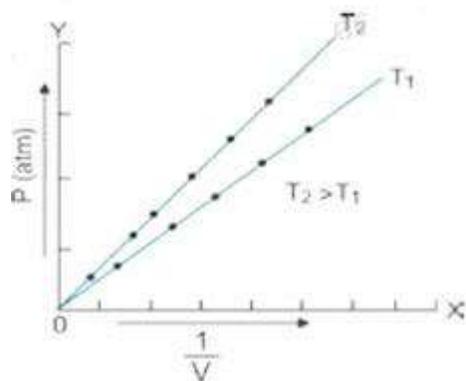
**Ans:** Plot a graph between pressure on x-axis and the product  $PV$  on Y-axis. A straight line parallel to the pressure axis is obtained. This straight line indicates that 'k' is a constant quantity. At higher constant temperature, the volume increase and value of product  $PV$  should increase due to increase of volume at same pressure, but  $PV$  remains constant at this new temperature and a straight line parallel to the pressure axis is obtained. This type of straight line will help us to understand the non-ideal behaviour of gases. Boyle's law is applicable only to ideal gases.



**39. Greater the temperature closer the straight line of  $P$  versus  $1/V$  to the pressure axis, Justify.**

**Ans:** If a graph is plotted between  $1/V$  on x-axis and the pressure  $P$  on the y-axis then a straight line is obtained. This shows that the pressure and inverse of volume are directly proportional to

each other. This straight line will meet at the origin which means that when the pressure is very close to zero, then the volume is so high that its inverse is very close to zero. By increasing the temperature of the same gas from  $T_1$  to  $T_2$  and keeping it constant, one can vary pressure and volume. The graph of this data between  $P$  and  $1/V$  will give another straight line. This straight line at  $T_2$  will be closer to the pressure-axis.

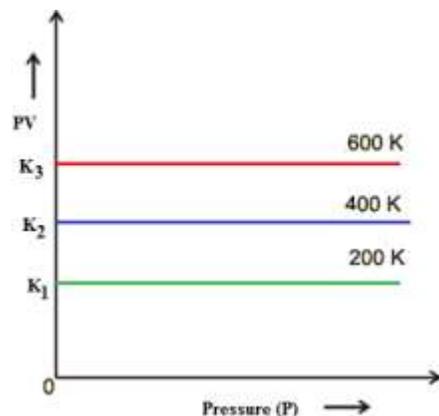


**40. What is absolute zero? What happens to real gases while approaching it?**

**Ans:** It is the lowest possible temperature which a gas would attain while remaining in the gaseous state. Its value is  $-273.16^{\circ}\text{C}$  or  $0\text{ K}$ . It is a hypothetical and unattainable temperature. In fact all gases especially real gases are converted to liquid before reaching this temperature.

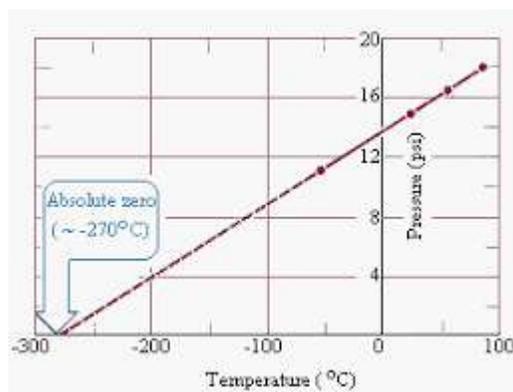
**41. The plot of  $PV$  vs  $P$  is a straight line at constant temperature and with a fixed number of moles of an ideal gas. Justify.**

**Ans:** The plot of  $PV$  vs  $P$  is a straight line at constant temperature and with a fixed number of moles of an ideal gas showing that 'k' is a constant quantity. At higher constant temperature, the volume increases and value of product  $PV$  should increase due to increase of volume at same pressure, but  $PV$  remains constant at this new temperature and a straight line parallel to the pressure axis is obtained.



**42. Justify that volume of gas becomes theoretically zero at  $-273\text{ }^{\circ}\text{C}$ .**

**Ans:** If we plot a graph between temperature on x-axis and the volume of one mole of an ideal gas on y-axis, we get a straight line which cuts the temperature axis at  $-273.16\text{ }^{\circ}\text{C}$ . This can be possible only if we extrapolate the graph up to  $-273.16\text{ }^{\circ}\text{C}$ . This temperature is the lowest possible temperature which would have been achieved if the substance remains in the gaseous state. Actually, all the gases are converted into liquids even before reaching this temperature.



**43. Why lighter gases diffuse more rapidly than heavier gases?**

**Ans:** Lighter gases diffuse more rapidly than heavier gases following Graham's law of diffusion or effusion. In fact, lighter gases have greater velocities and thus greater rates of diffusion.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

$r_1$  = rate of diffusion of gas 1

$r_2$  = rate of diffusion of gas 2

$M_1$  = Molar mass of gas 1

$M_2$  = Molar mass of gas 2

**44. Calculate the density of methane at**

**STP.**

**Ans: Given data:-**

Temperature of gas = 0°C or

273 K

Pressure of gas = 1 atm

Molecular mass of CH<sub>4</sub> = 16

g mol<sup>-1</sup> Density of CH<sub>4</sub> = ?

**Solution**

$$d = PM/RT$$

$$= 1 \times 16 / 0.0821 \times 273$$

$$= 0.7138 \text{ g dm}^{-3}$$

**45. Calculate number of molecules and number of atoms in 20 cm<sup>3</sup> of CH<sub>4</sub> at 0°C and**

**700mm of Hg.**

**Ans:**

Volume of CH<sub>4</sub> = 20 cm<sup>3</sup>

Temperature = 0°C or 273K

Pressure = 700 mm of Hg = 0.921 atm

No. of moles of CH<sub>4</sub> = ?

Molecules of CH<sub>4</sub> = ?

Atoms of CH<sub>4</sub> = ?

**Solution**

$$PV = nRT$$

$$n = PV/RT$$

$$n = 0.921 \times 20 / 0.082 \times 273$$

$$n = 18.42 / 22.386$$

$$n = 0.823 \text{ mol}$$

$$\text{No. of molecules of CH}_4 = 0.823 \times 6.02 \times 10^{23} = 4.95 \times 10^{23} \text{ molecules}$$

$$\text{No. of atoms} = 5 \times 4.95 \times 10^{23} = 24.77 \times 10^{23} = 2.477 \times 10^{24} \text{ atoms}$$

**46. State Joule-Thomson Effect. Write its application OR Define Joule-Thomson Effect.**

**Ans:** When a highly compressed gas is allowed to expand into the region of low pressure it gets cooled. Nitrogen and oxygen are liquefied on industrial scale by Linde's method of liquefaction which uses the phenomenon of Joule-Thomson effect.

**47. Hydrogen and Helium are ideal at room temperature but  $SO_2$  and  $Cl_2$  are non-ideal.**

**Ans:** Gases are non-ideal at high pressure and low temperature because then intermolecular forces become stronger. In helium and hydrogen, already there are weaker Van der Waal's forces because these are non-polar and their particle size is very small so they behave ideally at room temperature. On the other hand,  $SO_2$  and  $Cl_2$  are either polar ( $SO_2$ ) or having bigger molecules so there are strong intermolecular forces in them which make them non-ideal at room temperature.

**48. Some of the postulates of Kinetic Molecular Theory are faulty. Justify OR Write down two faulty assumptions of KMT of gases.**

**Ans:**

Following are the faulty assumptions of KMT:

1. There are no forces of attraction among the molecules of a gas.
2. The actual volume of gas molecules is negligible as compared to the volume of the gas.

**49. Derive expression for the molecular mass of the gas using general gas equation.**

**Ans:** According to the general gas equation:

$$PV = nRT$$

But  $n = m/M$ , putting in above equation

$$PV = \frac{m}{M}RT$$

$$M = \frac{mRT}{PV}$$

**50. Prove that  $p_A = P_t \cdot x_A$**

**Ans:** Let us suppose that we have a mixture of gas A and gas B. This mixture is enclosed in a container having volume (V). The total pressure is one atm. The number of moles of the gases A and B are  $n_A$  and  $n_B$  respectively. If they are maintained at temperature T, then

$P_t V = n_t RT$  (equation for the mixture of gases)

$p_A V = n_A RT$  (equation for gas A)

$P_B V = n_B RT$

(equation for gas B)

$p_A V / P_t V = n_A RT / n_t RT$

$p_A / P_t = n_A / n_t$

$p_A = n_A / n_t$

$P_t$

Hence  $p_A = x_A P_t$  ( $x_A$  is a mole fraction of gas A)

**51. Define critical temperature ( $T_c$ ).**

**Ans:** The highest temperature at which a substance can exist as a liquid, is called its critical temperature ( $T_c$ ). For example,  $O_2$  has a critical temperature 154.4 K (-118.75 °C).

**52.  $SO_2$  is comparatively non-ideal at 273K but behave ideally at 373K.**

**Ans:** Low temperature decreases the kinetic energies of molecules resulting in more intermolecular forces. At 273K intermolecular forces are more and that creates non-ideality but 373K breaks intermolecular forces and creates ideality.

**53. Rate of diffusion of ammonia is more than that of HCl. Why?**

**Ans:** According to Graham's law, lighter gases diffuse more rapidly than heavier gases. As ammonia is lighter (17 g/mol) than HCl (36.5 g/mol) its rate of diffusion is more than HCl.

**54. Pressure of ammonia gas at given conditions is less as calculated by Vander Waal equation than that calculated by general gas equation. Why?**

**Ans:** Van der Waal's equation tells about intermolecular forces which decrease the observed pressure of a gas. So, pressure of NH<sub>3</sub> gas calculated with this equation shall be mathematically lesser as compared to the value of pressure calculated from ideal gas equation:

$$PV = nRT$$

Vander Waal's equation is as follows:

$$(P + n^2a/V^2) (V - nb) = nRT$$

**55. What is physical significance of Vander Waal's constants 'a' and 'b'?**

**Ans:**

'a' is the attraction per unit volume and is called co-efficient of attraction for one mole of a gas. Its value depends directly upon the strength of intermolecular forces among gas particles. 'b' is effective volume or excluded or incompressible volume per mole. Its value depends on the size of the gas molecules.

**56. Justify that  $1\text{ cm}^3$  of  $\text{H}_2$  and  $1\text{ cm}^3$  of  $\text{CH}_4$  at STP will have same number of molecules, when one molecule of  $\text{CH}_4$  is 8 times heavier than that hydrogen.**

**Ans:** According to Avogadro's law  $1\text{ cm}^3$  of  $\text{H}_2$  and  $1\text{ cm}^3$  of  $\text{CH}_4$  at STP will have same number of molecules and it does not depend on mass of molecules. The distances between the gas molecules is 300 times greater than the diameter so sizes and masses do not affect the volume and number of molecules are equal.

$22414\text{ cm}^3$  of 1 mole gas contains molecules =  $6.02 \times 10^{23}$

$1\text{ cm}^3$  of 1 mole gas contains molecules =  $6.02 \times 10^{23} / 22414 = 2.68 \times 10^{19}$

Molar mass of  $\text{H}_2$  is  $2\text{ g mol}^{-1}$  while that of  $\text{CH}_4$  is  $16\text{ g mol}^{-1}$  which shows that the molecule of methane is 8 times heavier than that of hydrogen ( $\text{H}_2$ ).

**57. Do you think that the volume of any quantity of gas becomes zero at  $-273.16\text{ }^\circ\text{C}$ ? Is it not against the law of conservation of mass? How do you deduce the idea of absolute zero from this information?**

**Ans:** The volume of any quantity of a gas does not become zero at  $-273.16\text{ }^\circ\text{C}$  because gases are liquefied before reaching this temperature. In fact,  $-273.16\text{ }^\circ\text{C}$  or zero Kelvin is the lowest possible temperature which has not been achieved by any gas remaining in the gaseous state. It is against the law of conservation of mass. Practically we cannot attain this temperature.

This 0 K is called absolute zero.

**58. Why do we feel comfortable in expressing the densities of gases in units of  $\text{g dm}^{-3}$  rather than  $\text{g cm}^{-3}$ , a unit which is used to express the densities of liquids and solids?**

**Ans:** We feel comfortable in expressing the densities of gases in the units of  $\text{g dm}^{-3}$  rather than  $\text{g cm}^{-3}$  because the gases have high volumes as compared to liquids and solids. So larger units of volume must be preferred. For example,  $0.00071 \text{ g/cm}^3$  is the density of methane at  $0^\circ\text{C}$ . It can also be expressed as  $0.71 \text{ g/dm}^3$  which is a more appropriate value.

**59. Do you think that 1 mole of  $H_2$  and 1 mole of  $NH_3$  at  $0^\circ C$  and 1 atm pressure will have Avogadro's number of particles?**

**Ans:** It is true because Avogadro's law states that one mole of any substance has  $6.02 \times 10^{23}$  particles in it. So, one mole of any gas shall also follow the same principle.

1 mole of  $H_2 = 6.02 \times 10^{23}$  molecules of  $H_2$

1 mole of  $NH_3 = 6.02 \times 10^{23}$  molecules of  $NH_3$

**60. Dalton's law of partial pressures is only obeyed by those gases which don't have attractive forces among their molecules. Explain it.**

**Ans:** If there are attractive forces among the gas molecules then their pressures against the walls of the container may reduce and observed total pressure of the gaseous mixture may not be equal to expected total pressure. In such cases Dalton's law is not obeyed.

**61. Define pressure.**

**Ans:** The force applied per unit area is called pressure. Its SI units are  $Nm^{-2}$ .

**62. Define diffusion and effusion. Give examples.**

**Ans:**

The spontaneous intermingling of molecules of one gas with another at a given temperature and pressure is called diffusion. For example, the spreading of fragrance of a rose or a scent is due to diffusion.

The effusion of a gas is its movement through an extremely small opening into a region of low pressure. For example, the escape of gas molecules one by one through a punctured tyre.

**63. Under which conditions gases behave ideally?**

**Ans:**

(i) Gases are ideal at low pressure and non-ideal at high pressure.

(ii) Gases are ideal at high temperature and non-ideal at low temperature.

***CHEMISTRY (XI)***  
***Chapter 4***  
***Liquids and Solids***  
***Short Questions***

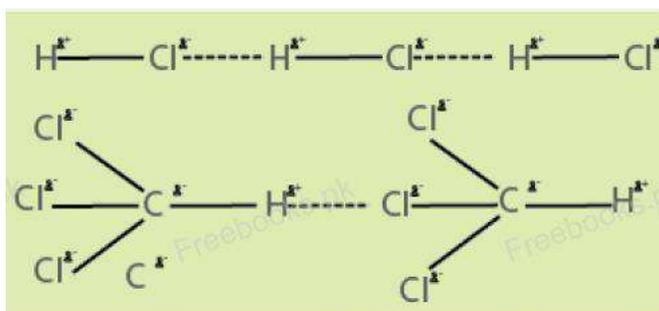
***1. Describe various types of intermolecular forces.***

***Ans:*** The forces holding the molecules together are called intermolecular forces. They are as follows:

1. Dipole-dipole forces
2. Ion-dipole forces
3. Dipole-induced dipole forces
4. Instantaneous dipole-induced dipole forces or London dispersion forces

***2. Define intramolecular forces.***

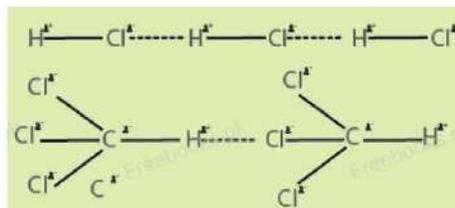
***Ans:*** The forces present between atoms of a molecule are called intramolecular forces. For example, the two bonds in molecule of water are intramolecular forces.



The solid lines are intramolecular forces, whereas, the dotted lines are intermolecular forces.

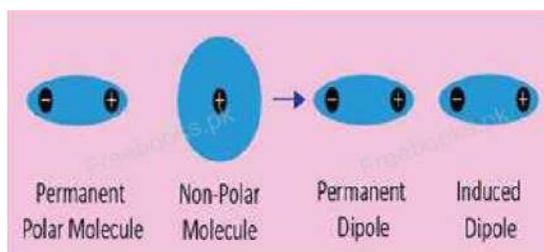
***3. What are dipole-dipole forces? Give an example.***

**Ans:** When the positive end of one molecule attracts the negative end of the other molecule then these electrostatic forces of attraction are called dipole-dipole forces. For example, dipole-dipole forces are present in HCl and CHCl<sub>3</sub>



**4. What are dipole-induced dipole forces or Debye forces?**

**Ans:** Sometimes we have a mixture of substances containing polar and non-polar molecules. The positive end of the polar molecule attracts the mobile electrons of the nearby non-polar molecule. In this way polarity is induced in non-polar molecule, and both molecules become dipoles. These forces are called dipole-induced dipole forces or Debye forces.



**5. What are instantaneous dipole-induced dipole forces or London dispersion forces?**

**Ans:** The momentary force of attraction created between instantaneous dipole and the induced dipole is called instantaneous dipole-induced dipole interaction or London force.

London forces are present in all types of molecules whether polar or non-polar but they are very significant for non-polar molecules like Cl<sub>2</sub>, H<sub>2</sub> and noble gases (helium, neon, etc.)

**6. Define polarizability.**

**Ans:** Polarizability is the quantitative measurement of the extent to which the electronic cloud can be polarized or distorted.

### 7. Why the boiling point of halogens increase down the group?

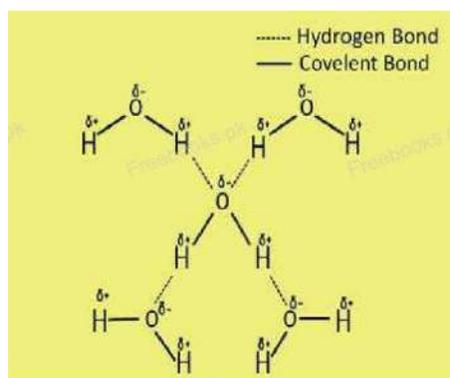
**Ans:** The increased distortion of electronic cloud creates stronger London forces and hence the boiling points are increased down the group VII-A. All the halogens are non-polar diatomic molecules, but there is a big difference in their physical states at room temperature. Fluorine is a gas and boils at  $-188.1\text{ }^{\circ}\text{C}$ , while iodine is a solid at room temperature which boils at  $+184.4\text{ }^{\circ}\text{C}$ . The polarizability of iodine molecule is much greater than that of fluorine.

### 8. How the number of atoms affect London forces?

**Ans:** Greater the number of atoms in a molecule, greater is its polarizability. For example, in saturated hydrocarbons the hydrocarbons have chain of carbon atoms linked with hydrogen atoms. Compare the length of the chain for  $\text{C}_2\text{H}_6$  and  $\text{C}_6\text{H}_{14}$ . They have the boiling points  $-88.6\text{ }^{\circ}\text{C}$  and  $68.7\text{ }^{\circ}\text{C}$ , respectively. This means that the molecule with a large chain length experiences stronger attractive forces. The reason is that longer molecules have more places along its length where they can be attracted to other molecules. With the increasing molecular mass of hydrocarbons they change from gaseous to liquid and then finally become solids.

### 9. Define hydrogen bonding.

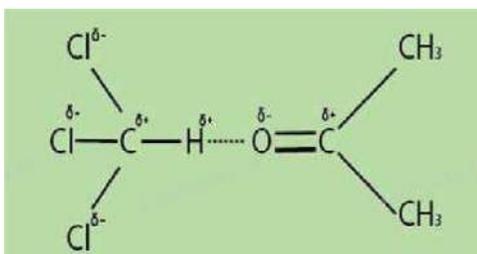
**Ans:** The electrostatic force of attraction between a highly electronegative atom and partial positively charged hydrogen atom. For example, hydrogen bonding is present in water.



The electronegative atoms responsible for creating hydrogen bonding are fluorine, oxygen, nitrogen and rarely chlorine. The strength of hydrogen bond is generally twenty times less than that of a covalent bond.

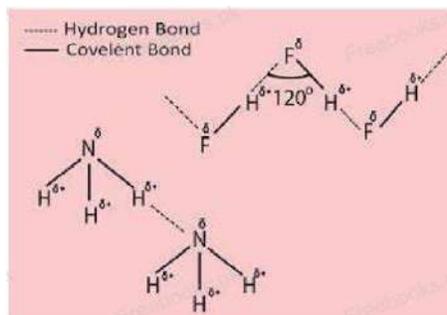
**10. Show hydrogen bonding between acetone and chloroform.**

**Ans:** Following is the structure showing hydrogen bonding between acetone and chloroform



**11. Show hydrogen bonding in HF and NH<sub>3</sub>**

**Ans:** Following is the structure showing hydrogen bonding in HF and NH<sub>3</sub>



**12. Why HF is a weak acid?**

**Ans:** Low acidic strength of HF molecule as compared to HCl, HBr and HI is due to strong hydrogen bonding, because the partial positively charged hydrogen is entrapped between two highly electronegative atoms.

**13. The hydrides of group IV-A have low boiling points as compared to hydrides of group V-A, VI-A and VII-A. Justify.**

**Ans:** The hydrides of group IV-A have low boiling points as compared to hydrides of group V-A, VI-A and VII-A because these elements are least electronegative.  $\text{CH}_4$  has the lowest boiling point because it is a very small molecule and its polarizability is the least.

**14. Why water is liquid but  $\text{H}_2\text{S}$  and  $\text{H}_2\text{Se}$  are gases?**

**Ans:** Oxygen has enhanced electronegative character due to which water shows hydrogen bonding and it is liquid. There is no hydrogen bonding in  $\text{H}_2\text{S}$  and  $\text{H}_2\text{Se}$  as S and Se are not electronegative due to which they are gases.

**15. Why HF has lower boiling point than  $\text{H}_2\text{O}$ ?**

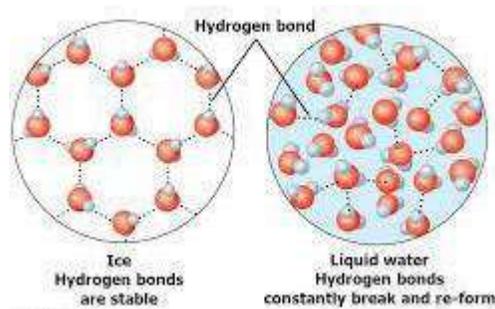
**Ans:** Fluorine atom can make only one hydrogen bond with electropositive hydrogen of a neighboring molecule. Water can form two hydrogen bonds per molecule, as it has two hydrogen atoms and two lone pairs on oxygen atom.

**16. Why ice floats on water?**

**Ans:** When water freezes it occupies 9% more space and its density decreases. The result is that ice floats on water.

**17. Ice occupies more space than water. Give reason.**

**Ans:** The molecules of water have tetrahedral structure. When the temperature of water is decreased and ice is formed then the molecules become more regular and this regularity extends throughout the whole structure and empty spaces are created in the structure. The structure of ice is just like that of a diamond and that is why when water freezes into ice it occupies 9% more space and its density is decreased with increase in volume. Hence, ice occupies more space than water.



**18. Why fish survive in winter in frozen ponds?**

**Ans:** The lower density of ice than liquid water at 0°C causes water in ponds and lakes to freeze from surface to the downward direction. Water attains the temperature of 4°C by the fall of temperature in the surrounding. As the outer atmosphere becomes further cold, the water at the surface becomes less dense. This less dense water below 4 °C stays on the top of slightly warm water underneath. A stage reaches when it freezes. This layer of ice insulates the water underneath for further heat loss. Fish and plants survive under this blanket of ice for months.

**19. How the size of electronic cloud affects London forces?**

**Ans:** When the size of the atom or molecule is large then the dispersion becomes easy and these forces become more prominent. The elements of the zero group in the periodic table are all mono-atomic gases. They don't make covalent bonds with other atoms because their outermost shells are complete.

**20. Define liquid crystals. Give an example.**

**Ans:** A liquid crystalline state exists between two temperatures i.e. melting temperature and clearing temperature. Example is cholesteryl benzoate.

**21. Mention the types of liquid crystals.**

**Ans:** Depending upon the nature of ordering, liquid crystals can be divided into nematic, smectic and cholesteric.

**22. How liquid crystal acts as temperature sensor?**

**Ans:** Like solid crystals, liquid crystals can diffract light. When one of the wavelengths of white light is reflected from a liquid crystal it appears coloured. As the temperature changes, the distances between the layers of the molecules of liquid crystals changes. Therefore, the colour of the reflected light changes accordingly. Thus, liquid crystals can be used as temperature sensors.

**23. Mention the biological applications of liquid crystals.**

**Ans:** Liquid crystalline substances are used to locate the veins, arteries, infections and tumors. The reason is that these parts of the body are warmer than the surrounding tissues. Specialists can use the techniques of skin thermography to detect blockages in veins and arteries. When a layer of liquid crystal is painted on the surface of the breast, a tumor shows up as a hot area which is coloured blue. This technique has been successful in the early diagnosis of breast cancer.

**24. What are the types of solid?**

**Ans:** Solids can be classified on the basis of the regular arrangements of constituent atoms, ions or molecules. There are two types of solids in this respect:

1. Crystalline solids
2. Amorphous solids

**25. Define crystalline solids.**

**Ans:** Those solids in which atoms, ions or molecules are arranged in a definite three dimensional pattern are called crystalline solids. This recurring regular geometrical pattern of structure extends three dimensionally. Example, sodium chloride crystal.

**26. Define amorphous solids.**

**Ans:** Amorphous substances are those whose constituent atoms, ions, or molecules do not possess a regular orderly arrangement. The best examples are glass, plastics, rubber, glue, etc. These substances have solid state properties and virtually complete maintenance of shape and volume. But they do not have an ordered crystalline state.

**27. What are crystallites?**

**Ans:** Long range regularity does not exist in amorphous solids but they can possess small regions of orderly arrangements. These crystalline parts of otherwise amorphous solids are known as crystallites.

**28. Define crystal lattice.**

**Ans:** A crystal lattice is an array of points representing atoms, ions or molecules of a crystal, arranged at different sites in three dimensional space.

**29. Define unit cell.**

**Ans:** The smallest part of the crystal lattice has all the characteristic features of the entire crystal and is called a unit cell.

**30. What are unit cell dimensions or crystallographic elements?**

**Ans:** There are three unit cell lengths  $a$ ,  $b$ ,  $c$  and three unit cell angles  $\alpha$ ,  $\beta$  and  $\gamma$ . These six parameters of the unit cell are called unit cell dimensions or crystallographic elements.

**31. Define crystal system.**

**Ans:** A crystal system may be identified by the dimensions of its unit cell along its three edges or axes,  $a$ ,  $b$ ,  $c$  and three angles between the axes  $\alpha$ ,  $\beta$ ,  $\gamma$ . There are seven crystal systems.

**32. How the cleansing action of soaps and detergents is due to hydrogen bonding?**

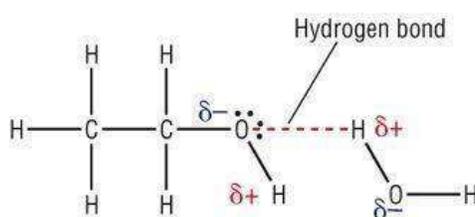
**Ans:** Soaps and detergents perform the cleansing action because the polar part of their molecules are water soluble due to hydrogen-bonding and the non-polar parts remain outside water, because they are alkyl or benzyl portions and are insoluble in water.

**33. Justify that structure of ice is just like that of diamond.**

**Ans:** Oxygen in water molecule is  $sp^3$  hybridized just like the carbon atom in diamond and each atom of carbon in diamond is at the center of tetrahedron just like the oxygen of water molecule in ice.

**34. Water and ethanol can mix easily in all proportions. Justify.**

**Ans:** Water ( $H_2O$ ) and ethanol ( $C_2H_5OH$ ) can mix easily in all proportions due to hydrogen bonding as both have OH group in them.



**35. Lower alcohols are water soluble but corresponding alkanes are insoluble in water.**

**Why?**

**Ans:** In alcohols  $-OH$  group is present due to which lower alcohols are soluble in water due to hydrogen bonding but alkanes do not have any functional group due to which they are not soluble in water due to lack of hydrogen bonding.

**36. Why intramolecular forces are stronger than intermolecular forces?**

**Ans:** Intramolecular forces are stronger than intermolecular forces because they are formed by sharing of electrons, whereas, intermolecular forces do not involve sharing of electrons.

**37. Lower alcohols are soluble into water but hydrocarbons are not. Give reason.**

**Ans:** Lower alcohols are soluble in water as they can form hydrogen bonds with water.

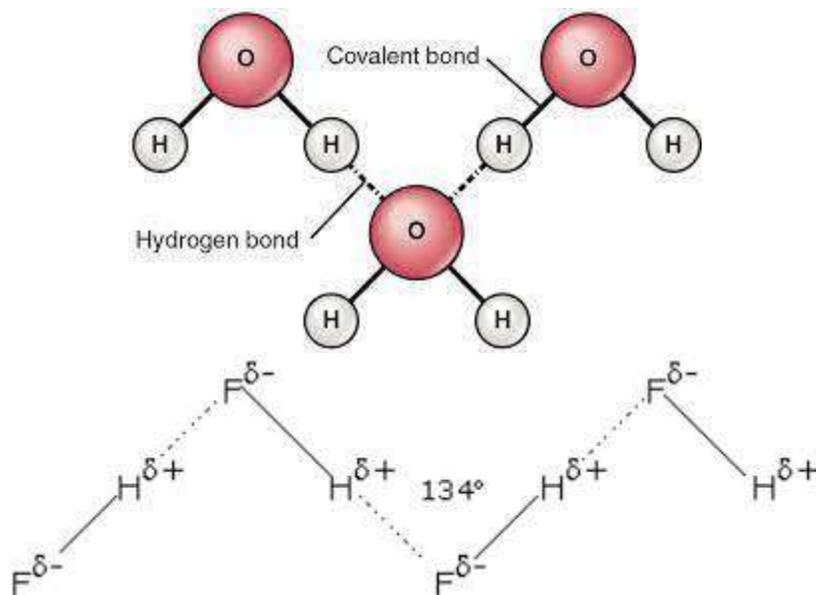
Hydrocarbons are not soluble in water because they are non- polar compounds and there is no chance of hydrogen bonding between water and hydrocarbon molecules.

**38. Write a brief note on solubility of hydrogen bonded molecules.**

**Ans:** Water is the best example of hydrogen bonded molecules. Similarly, ethyl alcohol also has the tendency to form hydrogen bonds. So, ethyl alcohol can dissolve in water because both can form hydrogen bonds with each other. Similarly, carboxylic acids are also soluble in water if their sizes are small. Hydrocarbons are not soluble in water because they are non- polar compounds and there is no chance of hydrogen bonding between water and hydrocarbon molecules.

**39. Why boiling point of water is greater than HF?**

**Ans:** HF and water both are liquid at room temperature as both have strong hydrogen bonds among their molecules. However, water has two hydrogen bonds per molecule and HF has one hydrogen bond per molecule. Thus, water has stronger forces of attractions between its molecules and has higher boiling point than HF.



**40. Earthen ware vessels keep water cool. Explain.**

**Ans:** Earthen ware vessels are porous in nature. Water molecules come out from the pores and evaporate. These molecules of water need energy to overcome their intermolecular forces of attraction. They get this energy from other molecules of water and get evaporated. This evaporation process lowers the energy of water molecules left behind. Thus, water has low temperature in earthen ware vessels and is kept cool.

**41. Why evaporation causes cooling?**

**Ans:** Evaporation of a liquid causes cooling because high energy molecules escape and change into vapours during evaporation. So the temperature of liquid falls. To compensate this heat loss, heat flows from surrounding to the region of lower temperature. This causes the temperature of surroundings to decrease. Hence, evaporation causes cooling.

**42. Explain why evaporation takes place at all temperatures?**

**Ans:** The molecules whose kinetic energies are greater than the average kinetic energies of the molecules, escape from the surface of the liquid. If temperature is increased, rate of evaporation

also increases. Thus, evaporation takes place at all temperatures and only the rate of evaporation differs with temperature change.

**43. Why boiling point of water is different at Murree Hills and at Mount Everest?**

**Ans:** When vapor pressure of a liquid becomes equal to the external pressure then the liquid boils, so when external pressure is changed, boiling point will also be changed. Therefore, water boils at 98 °C at Murree hills due to external pressure of 700 torr while at the top of Mount Everest, water boils at only 690 °C at 323 torr.

**44. Why different liquids evaporate at different rates even at the same temperature?**

**Ans:** There are many factors which control the rate of evaporation of the liquids even at the same temperature. These factors include attractive forces among molecules of the liquid and surface area of the liquid. If attractive forces are weak, the rate of evaporation is faster, e.g. gasoline having weaker forces of attraction than water evaporates much faster than water. Similarly, if surface area is increased, then more molecules are able to escape and liquid evaporates more quickly.

**45. Vacuum distillation can be used to avoid decomposition of sensitive liquids. Explain.**

**Ans:** The decomposition of many compounds can be avoided by vacuum distillation. For example glycerin boils at 290 °C at 760 torr pressure but it decomposes at this temperature. Hence, glycerin cannot be distilled at this temperature. Under vacuum, the boiling temperature of glycerin decreases to 210 °C at 50 torr. It is distilled at this temperature without decomposition and hence can be purified easily.

**46. Heat of sublimation of iodine is very high. Justify it.**

**Ans:** In the solid state the molecules of iodine align in the form of layer lattice with I-I bond distance 271.5 pm. Thus, heat of sublimation of iodine is very high.

**47. Ionic solids do not conduct electricity in solid state. Give reason.**

**Ans:** Ionic solids do not conduct electricity in solid state because on account of electrostatic force existing between them, the cations and anions remain tightly held together and hence occupy fixed positions. Ionic crystals conduct electricity when they are in the molten state.

**48. Write down applications of liquid crystals (Write any two as an answer to short question)**

**Ans:** Following are the two applications of liquid crystals:

(i) Like solid crystals, liquid crystals can diffract light. When one of the wavelengths of white light is reflected from a liquid crystal it appears coloured. As the temperature changes, the distances between the layers of the molecules of liquid crystals change. Therefore, the colour of the reflected light changes accordingly. Thus, liquid crystals can be used as temperature sensors.

(ii) Liquid crystals are used to find the point of potential failure in electrical circuits. Room thermometers also contain liquid crystals with a suitable temperature range. As the temperature changes, figures show up in different colours.

(iii) Liquid crystalline substances are used to locate the veins, arteries, infections and tumors. The reason is that these parts of the body are warmer than the surrounding tissues. Specialists can use the techniques of skin thermography to detect blockages in veins and arteries. When a layer of liquid crystal is painted on the surface of the breast, a tumor shows up as a hot area which is coloured blue. This technique has been successful in the early diagnosis of breast cancer.

(iv) Liquid crystals are used in the display of electrical devices such as digital watches, calculators and laptop computers. These devices operate due to the fact that temperature, pressure and electro-magnetic fields easily affect the weak bonds, which hold molecules together in liquid crystals.

(v) In chromatographic separations, liquid crystals are used as solvents.

(vi) Oscilloscopic and TV displays also use liquid crystal screens.

**49. Define isomorphism and polymorphism with examples.**

**Ans:**

**Isomorphism** is the phenomenon in which two different substances exist in the same crystalline form. These different substances are called isomorphs of each other. A crystalline form is independent of the chemical nature of the atoms and depends only on the number of atoms and their way of combinations. For example,  $\text{NaNO}_3$  and  $\text{KNO}_3$  have rhombohedral crystals, whereas, Cu and Ag have cubic crystals.

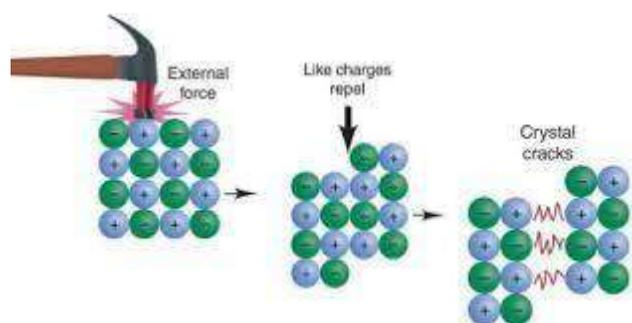
**Polymorphism** is a phenomenon in which a compound exists in more than one crystalline forms and such compounds are called polymorphic and these forms are called polymorphs of each other. For example,  $\text{AgNO}_3$  has rhombohedral and orthorhombic crystalline forms.

**50. Define polymorphism. Give an example.**

**Ans:** Polymorphism is a phenomenon in which a compound exists in more than one crystalline forms and such compounds are called polymorphic and these forms are called polymorphs of each other. For example,  $\text{AgNO}_3$  has rhombohedral and orthorhombic crystalline forms.

**51. Why ionic solids are highly brittle?**

**Ans:** Ionic crystals are highly brittle because ionic solids are composed of parallel layers which contain cations and anions in alternate positions so that the opposite ions in the various parallel layers lie over each other. When an external force is applied, one layer of the ions slides a bit over the other layer along a plane. In this way, the like ions come in front of each other and begin to repel. So, the application of a little external force develops repulsion between two layers causing brittleness.



**52. Why heat of sublimation of iodine is very high?**

**Ans:** In the solid state the molecules of iodine align in the form of layer lattice with I-I bond distance 271.5 pm. Thus, heat of sublimation of iodine is very high.

**53. Define transition temperature with an example.**

**Ans:** It is the temperature at which two crystalline forms of the same substance can co-exist in equilibrium with each other. At this temperature, one crystalline form of a substance changes to another. Above and below this temperature, only one form exists. For example, transition temperature of tin is  $13.2\text{ }^{\circ}\text{C}$  at which Grey tin having cubic crystals and White tin having tetragonal crystals co-exist. Similarly, transition temperature of Sulphur rhombic and monoclinic crystals is  $95.5\text{ }^{\circ}\text{C}$ .

**54. Cleavage is an anisotropic behaviour. Explain it.**

**Ans:** Whenever the crystalline solids are broken they do so along definite planes. This is called cleavage. Cleavage is an isotropic behavior because solids show cleavage depending upon direction and solids have orderly arrangements of their particles in different directions.

**55. How the liquid crystals help in the detection of blockages in veins and arteries. OR How are liquid crystals used to locate veins, arteries, infections and tumors?**

**Ans:** Liquid crystals are used to locate veins, arteries, infections and tumors. The reason is that these parts of the body are warmer than the surrounding tissues. Specialists can use the technique of skin thermography to detect blockages in veins and arteries. When a layer of liquid crystals is painted on the surface of the breast, a tumor shows up as a hot area which is coloured blue. This technique has been successful in the early diagnosis of breast cancer.

**56. What is relationship between polymorphism and allotropy?**

**Ans:**

**Polymorphism:** Polymorphism is a phenomenon in which a compound exists in more than one crystalline form and such compounds are called polymorphic and these forms are called polymorphs of each other. For example,  $\text{AgNO}_3$  has rhombohedral and orthorhombic crystalline forms.

**Allotropy:** The existence of an element in more than one crystalline form is called allotropy and these forms are called allotropes of each other. For example, carbon has allotropic forms of diamond (cubic) and graphite (hexagonal).

### **Relationship between Polymorphism and Allotropy:**

The relationship between both phenomena i.e polymorphism and allotropy is that both describe the existence of the substances in more than one crystalline form. However, polymorphism tells about compounds and allotropy describes about elements.

*57. Transition temperature is the term used for elements as well as compounds. Explain.*

*Ans:* It is the temperature at which two crystalline forms of the same substance can co-exist in equilibrium with each other. At this temperature, one crystalline form of a substance changes to another. Above and below this temperature, only one form exists. For example, transition temperature of tin is  $13.2\text{ }^{\circ}\text{C}$  at which Grey tin having cubic crystals and White tin having tetragonal crystals co-exist. Similarly, transition temperature of Sulphur rhombic and monoclinic crystals is  $95.5\text{ }^{\circ}\text{C}$ . There are many compounds which show transition temperature. For example,  $\text{KNO}_3$  has transition temperature of  $32.38\text{ }^{\circ}\text{C}$  having orthorhombic and rhombohedral crystals. Thus transition temperature is the term used for elements as well as compounds.

**58. Define transition temperature. Give two examples.**

**Ans:** It is the temperature at which two crystalline forms of the same substance can co-exist in equilibrium with each other. At this temperature, one crystalline form of a substance changes to another. Above and below this temperature, only one form exists. For example, transition temperature of tin is  $13.2\text{ }^{\circ}\text{C}$  at which Grey tin having cubic crystals and White tin having tetragonal crystals co-exist. Similarly, transition temperature of sulphur rhombic and monoclinic crystals is  $95.5\text{ }^{\circ}\text{C}$ .

**59. The vapor pressure of diethyl ether is higher than water at same temperature?**

**Ans:** The forces of attraction are weaker in diethyl ether, whereas, water has strong hydrogen bonding between its molecules. Thus, diethyl ether evaporates more quickly than water and its vapor pressure is higher than water.

**60. Give properties of molecular solids.**

**Ans:** Following are the four properties of molecular solids:

1. Molecular solids have weak forces of attraction among their molecules.
2. They are soft and easily compressible.
3. They are mostly volatile and have low melting and boiling points.
4. They are bad conductors of electricity.

**61. Define molar heat of fusion and molar heat of vaporization.**

**Ans:**

**Molar Heat of Fusion:** It is the amount of heat absorbed by one mole of a solid when it melts into liquid form at its melting point. The pressure during the change is kept constant.

**Molar Heat of Vaporization:** It is the amount of heat absorbed when one mole of a liquid is changed into vapors at its boiling point. The pressure during the change is kept constant.

**62. Describe that heat of sublimation is greater than heat of vaporization.**

**Ans:** The heat of sublimation is greater than heat of vaporization because vaporization is a single stage phase change whereas sublimation is a double stage phase change as solid changes to vapors.

**63. Why the electrical conductivity of metals decrease by increasing temperature?**

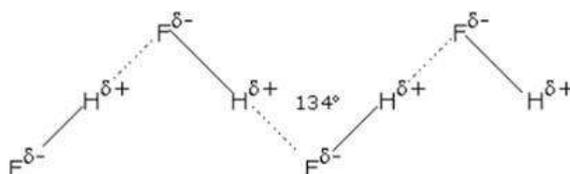
**Ans:** Metals are good conductors of electricity. Sometimes, the electrical conductivity of metals decreases with increase in temperature. The reason is that with the increase in temperature the positive metal ions also begin to oscillate and the motion hinders the free movement of mobile electrons between the positive ions. This hindrance decreases the electrical conductivity.

**64. What is meant by dynamic equilibrium? Give an example.**

**Ans:** Whenever a change of state occurs, the system moves towards the condition of dynamic equilibrium. Dynamic equilibrium is a situation when two opposing changes occur at equal rates. For example, at 0 °C, solid water (ice) exists in dynamic equilibrium with liquid water.

**65. HF is weaker acid than HCl. Why?**

**Ans:** HF has strong hydrogen bonding among its molecules and its H atoms are trapped between F atoms. Due to this HF does not release its proton  $H^+$  easily as compared to HCl which donates its proton  $H^+$  easily. Thus, HF is a weaker acid than HCl.

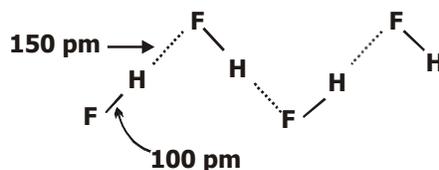


**66. Diamond is hard and an electrical insulator. Give reason?**

**Ans:** There is  $sp^3$ -  $sp^3$  overlapping of the carbon atoms in diamond and it has tetrahedral structure with 1.54 °A bond lengths most suitable for effective packing of atoms, thus, diamond is hard. As there are no free electrons present in any of the carbon atoms which could conduct electricity, therefore, diamond is an electrical insulator.

**67. In the hydrogen bonded structure of HF, which is stronger bond? The shorter covalent bond or the longer hydrogen bond between different molecules.**

**Ans:** A covalent bond is formed by the mutual sharing of electrons. Hence, it is true chemical bond and very strong, Whereas, hydrogen bond is not true chemical bond. It is formed due to polarity of H – F molecules.



Therefore, hydrogen bond is weaker. It just adds to the strength of bonding. Real strength is because of covalent bond.

**68. In a very cold winter the fish in garden ponds owe their lives to hydrogen bonding?**

**Ans:** When the temperature of surrounding drops to 0 °C, ice is formed on the surface of water only. This layer of ice insulates the water underneath from further cooling. Fish and other aquatic life can live under this layer of ice easily. Moreover fresh air containing oxygen easily dissolves in water by passing through air spaces of ice.

**69. What is the origin of intermolecular forces in water?**

**Ans: Origin of Dipole-Dipole Forces:**

Due to greater electronegativity of oxygen atoms, the shared electron pair is closer to oxygen atom creating partial –ve charge on it and corresponding partial +ve charge on covalently bonded hydrogen atoms. Water molecules become polar and dipole-dipole forces start operating.

**Origin of Hydrogen bonding:**

Smaller size of hydrogen atom and presence of two lone pairs on oxygen atom makes possible the origin of hydrogen bonding in which lone pairs of oxygen are donated to hydrogen atoms of nearby molecules.

**70. Boiling needs a constant supply of heat. Justify.**

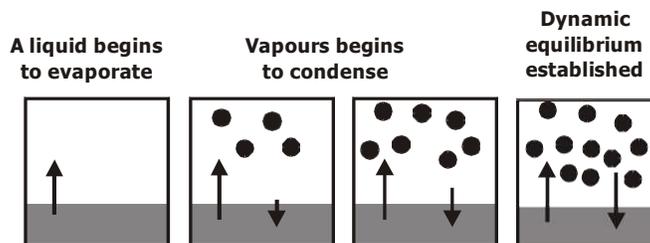
**Ans:** At boiling point intermolecular forces break abruptly. Molecules get freedom, convert to bubbles and come out of the liquid. In order to continue this process, continuous heating is required for continuous boiling.

**71. Earthenware vessels keep water cool. Justify.**

**Ans:** Earthenware vessels have very small pores in them. Evaporation of high energy molecules of water takes place through these pores as a result of which temperature of liquid falls. Earthen ware vessel is made of clay which keeps that water cool and does not absorb heat energy from surrounding.

**72. Dynamic equilibrium is established during evaporation of liquid in a closed vessel at constant temperature.**

**Ans:** When the molecules of a liquid leave the open surface, they are mixed up with air above the liquid. This process is called evaporation. But if we close the system the molecules of liquid start gathering above the surface. These molecules do not only collide with the walls of the container, but also with the surface of the liquid. There are chances that these molecules are recaptured by the surface of liquid. This process is called condensation. The two-processes i.e., evaporation and condensation continue till a stage reaches when the rate of evaporation becomes equal to the rate of condensation. This is called the state of dynamic equilibrium.



**73. Crystals of salts fracture easily, but metals are deformed under stress without fracturing. Explain the difference.**

**Ans:** Crystals of salt fracture easily because salt crystals are mostly ionic in nature and when some stress is applied then same ions come in front of each other. Hence, forces of repulsion are developed between same ions and crystals are broken down showing brittleness. On the other hand, metals have metallic bonding with each other which is very strong and does not allow the metal to break and only deformation occurs.

**74. Name the two forms of metal packing.**

**Ans:** Following are the two forms of metal packing:

- i. Cubic close packing
- ii. Hexagonal close packing

**75. What is the difference between cubic close packing and hexagonal close packing?**

**Ans:** Cubic close packing is also called face-centered cubic cell. Closest packed means that the atoms are packed together as closely as possible. In hexagonal close packing layers of spheres are packed so that spheres in alternating layers overlie one another.

**76. What happens to conduction of metals with temperature?**

**Ans:** As the temperature increases the conduction of metals decreases as the positive spheres start vibrating about their fixed position, thereby, preventing the flow of free moving electrons.

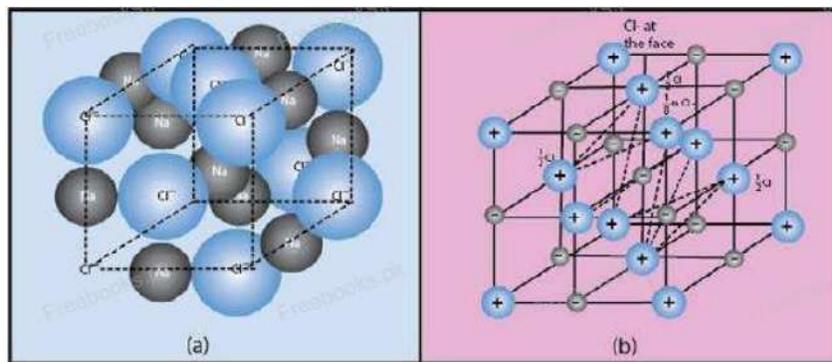
**77. Define lattice energy. Give an example.**

**Ans:** The lattice energy is the energy released when one mole of the ionic crystal is formed from the gaseous ions. It is also defined as the energy required to break one mole of solid into isolated ions in the gas phase. It is expressed in kJ/mole. For example, lattice energy of sodium chloride is -787 kJ/mol.

**78. Explain structure of sodium chloride.**

**Ans:** In the structure of NaCl each  $\text{Na}^+$  ion is surrounded by six chloride ions. The distance between two nearest ions of the same kind i.e.,  $\text{Cl}^-$  ions is  $5.63 \text{ \AA}$ . So the distance between two adjacent ions of different kind is  $5.63/2 = 2.815 \text{ \AA}$ .

The location of  $\text{Na}^+$  and  $\text{Cl}^-$  is such that each  $\text{Na}^+$  is surrounded by six  $\text{Cl}^-$  placed at the corners of a regular octahedron. So the coordination number of each  $\text{Na}^+$  is six. Similarly, each  $\text{Cl}^-$  is also surrounded by six  $\text{Na}^+$  ions. There are eight  $\text{Cl}^-$  at the corners of the cube, and each is being shared amongst eight cubes.  $1/8^{\text{th}}$  part of each  $\text{Cl}^-$  ion is considered for this unit cell. So, one complete  $\text{Cl}^-$  is contributed by eight corners. Similarly, six chloride ions are present at the face centres and each is being shared between two cells. Thus, per unit cell there are  $8/8 + 6/2 = 4 \text{ Cl}^-$  ions. The presence of 4  $\text{Na}^+$  can be justified if you take a unit cell having 8 $\text{Na}^+$  at eight corners and 6 $\text{Na}^+$  at faces. So, there is equal number of  $\text{Na}^+$  ions and four NaCl units are present per unit cell.



**79. Define crystal system.**

**Ans:** A crystal system may be identified by the dimensions of its unit cell along its three edges or axes,  $a$ ,  $b$ ,  $c$  and three angles between the axes  $\alpha$ ,  $\beta$ ,  $\gamma$ . There are seven crystal systems.

**80. Explain cubic, tetragonal, orthorhombic, monoclinic, hexagonal, rhombohedral and triclinic system.**

**Ans:**

Sr. No	Crystal system	Axes	Angles	Examples
1.	cubic	$a = b = c$	$\alpha = \beta = \gamma = 90^\circ$	Fe, Cu, Ag, Au, NaCl, NaBr, Diamond
2.	Tetragonal	$a = b \neq c$	$\alpha = \beta = \gamma = 90^\circ$	Sn, $\text{SnO}_2$ , $\text{MnO}_2$ , $\text{NH}_4\text{Br}$
3.	Orthorhombic	$a \neq b \neq c$	$\alpha = \beta = \gamma = 90^\circ$	Idoine, Rhombic, Sulphur, $\text{BaSO}_4$ , $\text{K}_2\text{SO}_4$
4.	Monoclinic	$a \neq b \neq c$	$\alpha = \gamma = 90^\circ, \beta \neq 90^\circ$	Sugar, Sulphur, Borax, $\text{NaSO}_4 \cdot 10\text{H}_2\text{O}$
5.	Hexagonal	$a = b \neq c$	$\hat{a} = \hat{a} = 90^\circ, \hat{a} = 120^\circ$	Graphite, ZnO, CdS, Ice, Zn, Cd
6.	Rhombohedral or Trignol	$a = b = c$	$\alpha = \beta \neq \gamma = 90^\circ \text{ and } 120^\circ$	Bi, $\text{Al}_2\text{O}_3$ , $\text{NaNO}_3$ , $\text{KNO}_3$
7.	Triclinic	$a \neq b \neq c$	$\alpha \neq \beta \neq \gamma \neq 90^\circ$	$\text{H}_3\text{BO}_3$ , $\text{K}_2\text{Cr}_2\text{O}_7$ , $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

(Use this table to write about each type in your own words)

**81. Define enthalpy change.**

**Ans:** If a physical or a chemical change takes place at constant pressure, then the heat change during this process is also called enthalpy change. This is denoted by  $\Delta H$ .

**82. Define molar heat of fusion.**

**Ans:** It is the amount of heat absorbed by one mole of a solid when it melts into liquid form at its melting point. The pressure, during the change is kept one atmosphere. It is denoted by  $\Delta H_f$ .

**83. Define molar heat of vaporization.**

**Ans:** It is the amount of heat absorbed when one mole of a liquid is changed into vapours at its boiling point. The pressure, during the change is kept one atmosphere. It is denoted by  $\Delta H_v$ .

**84. Define molar heat of sublimation.**

**Ans:** It is the amount of heat absorbed when one mole of a solid sublimates to give one mole of vapours at a particular temperature and one atmospheric pressure. It is denoted by  $\Delta H_s$ .

**85. Define molar heat of vaporization.**

**Ans:** The amount of heat required to vapourize one mole of a liquid at its boiling point is called its molar heat of vapourization. The molar heat of vapourization of water is  $40.6 \text{ kJmol}^{-1}$ .

**86. Name the method used to determine vapour pressure. Mention the mathematical expression applied.**

**Ans:** The method which is used to determine vapour pressure is called manometric method.

Following mathematical equation is used to calculate vapour pressure:

- ∴  $P = P_a + \Delta h$
- P = Vapour pressure of the liquid at one atm pressure.
- P = Atmospheric pressure.
- $\Delta h$  = Difference in the heights of the mercury levels in the two limbs of the manometer, giving us the vapour pressure of liquid.

**87. Mention the applications of hydrogen bonding.**

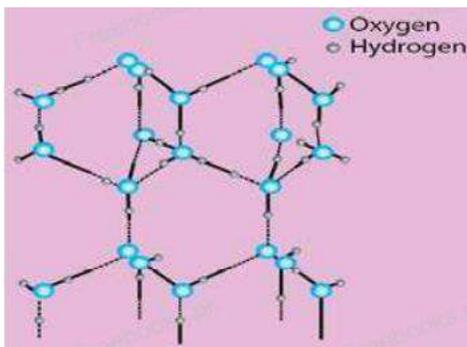
**Ans:** Following are the applications of hydrogen bonding:

1. Thermodynamic properties of covalent hydrides
2. Solubility of hydrogen bonded molecules
3. Structure of ice
4. Cleansing action of soaps and detergents
5. Hydrogen bonding in biological compounds and food materials

**89. Explain structure of ice.**

**Ans:** The molecules of water have tetrahedral structure. Two lone pairs of electrons on oxygen atom occupy two corners of the tetrahedron. In the liquid state, water molecules are extensively associated with each other and these associations break and are reformed because the molecules of water are mobile. When the temperature of water is decreased and ice is formed then the molecules become more regular and this regularity extends throughout the whole structure. Empty spaces are created in the structure. That is why when water freezes it occupies 9% more space and its density decreases. The result is that ice floats on water. The structure of ice is just like that of a diamond because each atom of carbon in diamond is at the center of tetrahedron just

like the oxygen of water molecule in ice.



***90. Tell about hydrogen bonding in DNA helical structure and proteins.***

***Ans:*** Fibres like those found in the hair, silk and muscles consist of long chains of amino acids. These long chains are coiled about one another into a spiral. This spiral is called a helix. Such a helix may either be right handed or left handed. In the case of right handed helix the groups like  $>NH$  and  $>C=O$  are vertically adjacent to one another and they are linked together by hydrogen bonds. These H-bonds link one spiral to the other. X-ray analysis has shown that on the average there are 27 amino acid units for each turn of the helix. Deoxyribonucleic acid (DNA) has two spiral chains. These are coiled about each other on a common axis. In this way, they give a double helix. This is 18-20 Å in diameter. They are linked together by H-bonding between their sub-units.

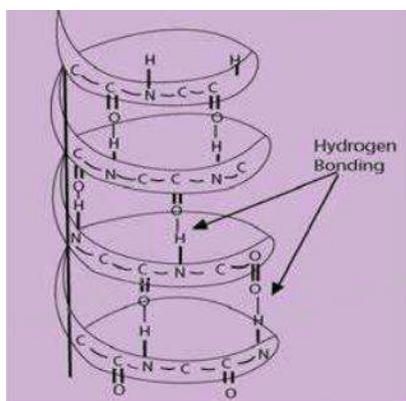


Fig (4.9 a) Hydrogen bonding

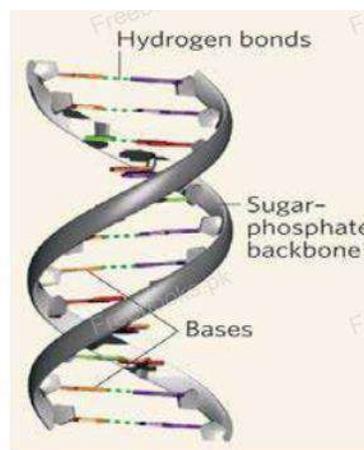


Fig (4.9 b) Hydrogen bonding in DNA double helix

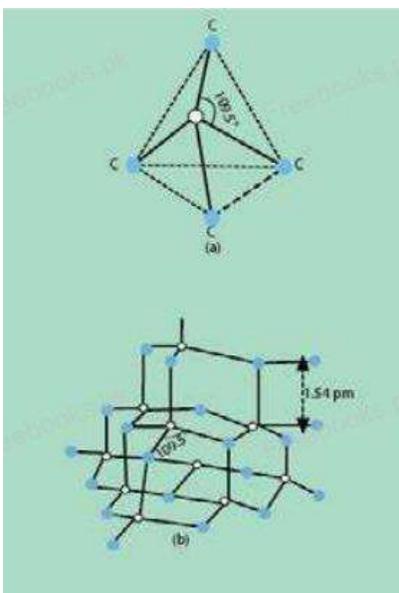
**91. Tell about structure of iodine.**

**Ans:** In the solid state the molecules of iodine align in the form of layer lattice. I -I bond distance is 271.5 pm and it is appreciably longer than in gaseous iodine (266.6 pm). As expected from its structure, iodine is a poor conductor of electricity. (draw structure from PTB)

**92. Tell about structure of diamond.**

**Ans:** Diamond is one of the allotropic modifications of carbon. It is best understood by taking into consideration the number of electrons in the outermost shell of carbon, which are four. The four atomic orbitals (one 2s and three 2p) undergo  $sp^3$  hybridization to give four  $sp^3$  hybridized orbitals. They are directed in space along the four corners of a tetrahedron. This is the unit cell of diamond and a large number of such unit cells undergo  $sp^3$  - $sp^3$  overlapping to form a huge structure. Each carbon atom is linked with four other carbon atoms. The bonds between carbon atoms are covalent which run through the crystal in three-dimensions. All the bond angles are  $109.5^\circ$  and the bond lengths are 154 pm.

The whole lattice is, therefore, continuous and because of the continuity of C-C covalent bonding, the entire diamond crystal behaves as a huge or giant three-dimensional carbon molecule. This is also called a macromolecule. The overall structure of diamond looks face centred-cubic.



**93. Ionic crystals don't conduct electricity in the solid state. Justify.**

**Ans:** Current is only carried by moving charged particles. In solid state of the ionic solid, all the ions are held together strongly by strong ionic bonding so they are unable to move and carry electrical current.

**94. In the closest packing of atoms of metals, only 74% space is occupied.**

**Justify.**

**Ans:** Cubic close packing is also called ABCABC packing. Although it is close packing even vacant spaces, also called interstices, are left behind which never fill up and occupy 26% empty space in a crystal and only 74% is occupied by atoms.

**95. The number of positive ions surrounding the negative ions in the ionic crystal lattice depends upon the sizes of the two ions.**

**Ans:** The crystal structure of ionic compounds depends upon the coordination number of cations.  $\text{Na}^+$  ion has smaller size and its coordination number is six because in its neighbor, six  $\text{Cl}^-$  ions are present.  $\text{Cs}^+$  ion has bigger size and its coordination number is eight. In the neighborhood of  $\text{Cs}^+$  eight  $\text{Cl}^-$  ions are present.  $\text{NaCl}$  has face-centered cubic structure while  $\text{CsCl}$  has body centered cubic structure

**96. The crystals showing isomorphism mostly have the same atomic ratios.**

**Justify.**

**Ans:** If different substances have same crystalline shape they are called isomorphous substances. Isomorphous substances have same atomic ratio. The crystalline form is independent of chemical nature and depends upon the atomic ratio.

Substance	Crystalline	Atomic ratio
$\text{NaNO}_3, \text{CaCO}_3$	Rhombohedral	1 : 1 : 3
$\text{K}_2\text{SO}_4, \text{K}_2\text{CrO}_4$	Orthorhombic	2 : 1 : 4

**97. Cleavage of the crystals is itself anisotropic behavior. Justify.**

**Ans:** The crystals showing different intensities of a given property in different directions are called anisotropic. This phenomenon is called 'Anisotropy'. Cleavage of crystals is an anisotropic property. Crystals have specific cleavage planes and can be divided only in those planes. Example: Mica sheets can be cleaved only parallel to length of sheet and not in any other direction.

**98. Amorphous solid like glass is also called super cooled liquid. Justify.**

**Ans:** An amorphous solid like glass has random structure like a liquid which has disarranged structure. Amorphous solid (glass) is formed initially in liquid state afterwards it freezes

and becomes hard and rigid but microscopically it looks like just as liquid because of random arrangement of particles. Some of the particles are also in the state of motion, whereas, particles of normal solids can't move. That is why, it is also called super cooled liquid.

**99. The vapour pressures of solids are far less than those of liquids. Justify.**

**Ans:** The vapor pressures of solids are lesser than those of liquids because of relatively stronger intermolecular forces among their particles. Due to stronger forces, molecules of solids can't escape easily from the solid.

**100. Iodine dissolves readily in tetrachloromethane. Justify.**

**Ans:** Iodine is more soluble in  $\text{CCl}_4$  because both are non-polar in nature. Non-polar solutes are more soluble in non-polar solvent.

**101. Sodium chloride and Caesium chloride have different structures. Justify.**

**Ans:** The crystal structure of ionic compounds depends upon the coordination number of cations.  $\text{Na}^+$  ion has smaller size and its coordination number is six because in its neighbor, six  $\text{Cl}^-$  ions are present.  $\text{Cs}^+$  ion has bigger size and its coordination number is eight. In the neighborhood of  $\text{Cs}^+$  eight  $\text{Cl}^-$  ions are present.  $\text{NaCl}$  has face-centered cubic structure while  $\text{CsCl}$  has body centered cubic structure.

**102. Diamond is hard and electrical insulator. Justify.**

**Ans:** Diamond is an example of covalent solids based on carbon atoms. It is hard because of close packing of  $\text{sp}^3$  hybrid carbon atoms which are tightly connected to each other with covalent bonds and form a three dimensional macromolecule. It is non-conductor (insulator) because all the valence electrons of each carbon take part in the bond formation and no one is free to move in the crystal.

**103. Sodium is softer than copper, but both are very good electrical conductors.**

**Justify.**

**Ans:** Both sodium and copper have large number of mobile electrons and hence are good conductors. The atomic number (number of protons in nucleus) of copper is 29 and that of sodium is 11. The strength of metallic bonds in Copper is greater due to greater nuclear charge as compared with sodium. So copper is harder than sodium.

**104. Define the following:**

**Ans:**

### **Crystalline Solids**

Those solids in which atoms, ions or molecules are arranged in a definite three dimensional pattern are called crystalline solids. This recurring regular geometrical pattern of structure extends three dimensionally.

### **Amorphous Solids**

The word amorphous means shapeless. Amorphous substances are those whose constituent atoms, ions, or molecules do not possess a regular orderly arrangement. The best examples are glass, plastics, rubber, glue, etc

### **Cleavage Planes**

Whenever the crystalline solids are broken they do so along definite planes. These planes are called the cleavage planes and they are inclined to one another at a particular angle for a given crystalline solid. The value of this angle varies from one solid to another solid.

### **Anisotropy**

Some of the crystals show variation in physical properties depending upon the direction. Such properties are called anisotropic properties and the phenomenon is referred to as anisotropy. The

physical properties of crystalline solids like refractive index, coefficient of thermal expansion, electrical and thermal conductivities are sometimes anisotropic in nature for some crystals.

### **Symmetry**

The repetition of faces, angles or edges when a crystal is rotated by  $360^\circ$  along its axis is called symmetry. This an important property of the crystal and there are various types of symmetry elements found in crystals like, center of symmetry, plane of symmetry and axis of symmetry, etc.

### **Habit of a Crystal**

The shape of a crystal in which it usually grows is called habit of a crystal. Crystals are usually obtained by cooling the saturated solution or by slow cooling of the liquid substance. These are formed by growing in various directions. If the conditions for growing a crystal are maintained, then the shape of the crystal always remains the same. If the conditions are changed the shape of the crystal may change. For example, a cubic crystal of NaCl becomes needle like when 10% urea is present in its solution as an impurity.

#### ***105. Mention the properties of metallic solids.***

**Ans:** Following are the properties of metallic solids:

1. Metals are good conductor of electricity.
2. The electrical conductivity of metals decrease with the increase in temperature.
3. The thermal conductivity is another property associated with metallic solids.
4. Whenever the metals are freshly cut, most of them possess metallic luster which means that they have a shining surface.
5. Metals are malleable and ductile whenever stress is applied on them.

#### **106. Why do metals have metallic luster?**

**Ans:** When light falls on the metallic surface, the incident light collides with the mobile electrons and they are excited. These electrons when de-excited give off some energy in the form of light. This light appears to be reflected from the surface of the metal which gives a shining look and so metals have a metallic luster.

## ***CHEMISTRY (XI)***

### ***Chapter 5 Atomic Structure Short Questions***

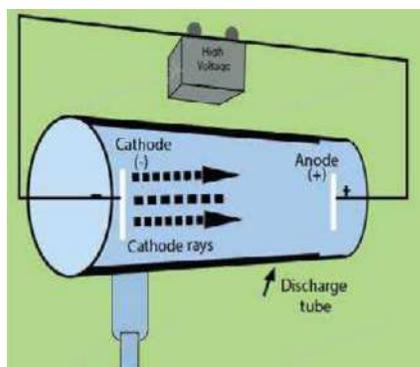
1. *Name the fundamental particles of atom.*

*Ans:* The fundamental particles of atom are electron, proton and neutron.

2. *How cathode rays (electrons) were discovered?*

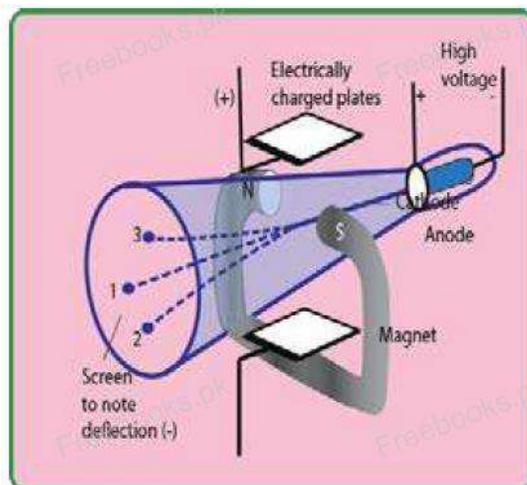
*Ans:* A gas discharge tube is fitted with two metallic electrodes acting as cathode and anode. The tube is filled with a gas, air or vapours of a substance at any desired pressure. The electrodes are connected to a source of high voltage. The tube is attached to a vacuum pump by means of a small side tube so that the conduction of electricity may be studied at any value of low pressure

It is observed that current does not flow through the gas at ordinary pressure even at high voltage of 5000 volts. When the pressure inside the tube is reduced and a high voltage of 5000-10000 volts is applied, then an electric discharge takes place through the gas producing a uniform glow inside the tube. When the pressure is reduced further to about 0.01 torr, the original glow disappears. Some rays are produced which create fluorescence on the glass wall opposite to the cathode. These rays are called cathode rays.



**3. How was it proved that cathode rays are negatively charged particles? OR How the bending of the cathode rays in the electric and magnetic field prove that they are negatively charged?**

**Ans:** Cathode rays are negatively charged. In 1895, J. Perrin showed that when the cathode rays passed between the poles of the magnet, the path of the negatively charged particles was curved downward to point 2 by the magnetic field. In 1897, J. Thomson established their electric charge by the application of electric field, the cathode ray particles were deflected upwards (towards the positive plate) to point 3. Thomson found that by carefully controlling the charge on the plates when the plates and the magnet were both around the tube, he could make the cathode rays strike the tube at point 1 again. He was able to cancel the effect of the magnetic field by applying an electric field that tended to bend the path of the cathode rays in the opposite direction.

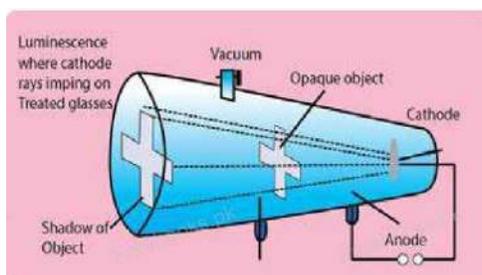


**4. What type of fluorescence is shown by cathode rays?**

**Ans:** They produce a greenish fluorescence on striking the walls of the glass tube. These rays also produce fluorescence in rare earths and minerals. When placed in the path of these rays, alumina glows red and tin stone yellow.

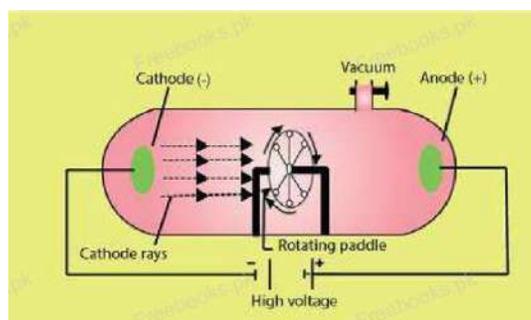
**5. Which observations prove that cathode rays cast shadow?**

**Ans:** Cathode rays cast a shadow when an opaque object is placed in their path. This proves that they travel in straight line perpendicular to the surface of cathode.



**6. Which observations prove that cathode rays have momentum? Or Prove that cathode rays have momentum.**

**Ans:** These rays can drive a small paddle wheel placed in their path. This shows that these rays possess momentum. From this observation, it is inferred that cathode rays are not rays but material particles having a definite mass and velocity.



**7. Why is it necessary to decrease the pressure in the discharge tube to get the cathode rays?**

**Ans:** The pressure in discharge tube is decreased to allow the cathode to move freely from cathode towards opposite wall. Hence, possibility of collisions between rays and the gas molecules are minimized.

**8. Whichever gas is used in the discharge tube the nature of the cathode rays remains the same. Why?**

**Ans:** All the gases have atoms and molecules having electrons in the outermost shell. Due to high voltage these electrons become free. They are repelled by the cathode and attracted towards the anode. That is why, they are called cathode rays as they are always electrons and nature of electrons is same no matter which gas is used in the discharge tube experiment.

**9. Why  $e/m$  value of the cathode rays is just equal to that of electron?**

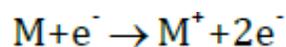
**Ans:** Cathode rays are actually electrons and the nature of electrons remains the same irrespective of the substance used in the discharge tube. So,  $e/m$  value of cathode rays is also called  $e/m$  value of electron.

**10. Why the positive rays are also called canal rays? Give reason.**

**Ans:** Positive rays after passing through the perforated cathode produce a glow on the wall opposite to the anode. Since these rays pass through the canals or the holes of cathode, they are called canal rays.

**11. What is the reason for the production of positive rays?**

**Ans:** These positive rays are produced, when high speed cathode rays (electrons) strike the molecules of a gas enclosed in the discharge tube. They knock out electrons from the gas molecules and positive ions are produced, which start moving towards the cathode.



**12. The  $e/m$  values of positive rays for different gases are different, but those for cathode rays the  $e/m$  values are same. Justify it.**

**Ans:** The nature of particles of positive rays depends upon the nature of the gas as every gas has its own atomic mass. As the mass is different so  $e/m$  values of positive rays for different gases are different. Cathode rays are simply electrons, so the  $e/m$  value of cathode rays remains the same no matter which gas is used in the experiment.

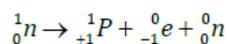
**13. The  $e/m$  value of positive rays obtained from the hydrogen gas is 1836 times less than that of cathode rays. Justify it.**

**Ans:** The  $e/m$  value for the positive rays is always smaller than that of electrons and depends upon the nature of the gas used in the discharge tube. Heavier the gas, smaller is the  $e/m$  value. When hydrogen gas is used in the discharge tube, the  $e/m$  value is found to be the maximum in comparison to any other gas because the value of 'm' is the lowest for the positive particle obtained from the hydrogen gas. Hence, the positive particle obtained from hydrogen gas is the lightest among all the positive particles. The mass of a proton is 1836 times more than that of an electron.

**14. Which particles are formed by the decay of free neutron?**

**Ans:**

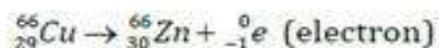
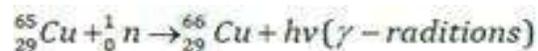
Free neutron decays into a proton  ${}^1_1P$  with the emission of an electron  ${}^{-1}_0e$  and a neutrino  ${}^0_0n$ .



**15. How the slow neutrons prove to be more effective than fast neutrons?**

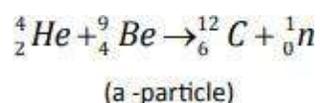
**Ans:** When the travelling neutrons have energy below 1 eV then they are called slow neutrons.

They are more effective in fission processes. Slow neutrons hit the copper metal and gamma radiations are emitted giving a new isotope of copper i.e. Cu (66) which is again radioactive.

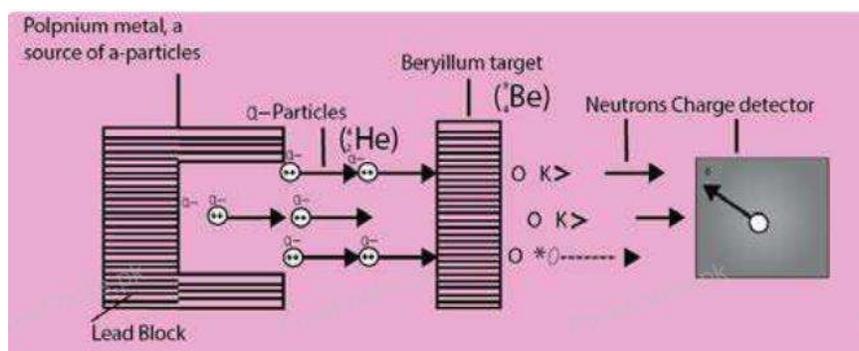


### 16. How neutron was discovered?

**Ans:** Chadwick discovered neutron in 1932 and was awarded Nobel Prize in Physics in 1935. A stream of  $\alpha$ -particles produced from a polonium source was directed at beryllium ( ${}^9_4\text{Be}$ ) target. It was noticed that some penetrating radiations were produced. These radiations were called neutrons because the charge detector showed them to be neutral.



Actually  $\alpha$ -particles and the nuclei of Beryllium are re-arranged and extra neutron is emitted.



### 17. Give properties of Neutron.

**Ans:** Some properties of neutron are as follows:

1. Neutrons cannot ionize gases.
2. Neutrons are highly penetrating particles.
3. They can expel high speed protons from paraffin, water, paper and cellulose.
4. When neutrons travel with an energy of 1.2 Mev (Mega electron volt  $10^6$ ), they are called fast neutrons but with energy below 1ev they are called slow neutrons.

**18. How did Rutherford's model of an atom proved the existence of nucleus of the atom?**

**Ans:** A beam of  $\alpha$ -particles was directed onto a gold foil of 0.00004 cm thickness as target through a pin-hole in lead plate. A photographic plate or a screen coated with zinc sulphide was used as a detector. Whenever, an  $\alpha$ -particle struck the screen, flash of light was produced at that point. It was observed that most of the particles went through the foil undeflected. Some were deflected at fairly large angles, and a few were deflected backward. Rutherford proposed that the rebounding particles must have collided with the central heavy portion of the atom which he called as nucleus.

**19. What are the defects of Rutherford's model of atom?**

**Ans:** Following are the defects in Rutherford's model of atom:

- (i) It is based on the laws of the motion and the gravitation. These laws are not for the charged bodies like electrons.
- (ii) The revolving electron should emit the energy continuously and by producing a spiral path it should ultimately fall into the nucleus. But actually the atoms don't do that.
- (iii) The atom should give continuous spectra but actually line spectra are obtained.

**20. What does Planck's quantum theory say? Mention main points.**

**Ans:** According to Planck's theory, energy travels in a discontinuous manner and it is composed of large number of tiny discrete units called quanta. The main points of his theory are:

- (i) Energy is not emitted or absorbed continuously. Rather, it is emitted or absorbed in a discontinuous manner and in the form of wave packets. Each wave packet or quantum is associated with a definite amount of energy. In case of light, the quantum of energy is often called photon.
- (ii) The amount of energy associated with a quantum of radiation is proportional to the frequency of the radiation.

$$E \propto \nu$$

$$E = h\nu$$

Where 'h' is a constant known as Planck's constant and its value is  $6.626 \times 10^{-34}$  Js. Frequency ( $\nu$ ) has the units of cycles  $s^{-1}$  or Hz. (1 Hz = 1 cycle  $s^{-1}$ )

**21. Define frequency.**

**Ans:** Frequency is the number of waves passing through a point per second. The SI unit of frequency is hertz (Hz).

**22. Define wavelength.**

**Ans:** Wavelength is the distance between the two adjacent crests or troughs and expressed in  $\text{\AA}$ , nm or pm. (1  $\text{\AA}$  =  $10^{-10}$  m, 1 nm =  $10^{-9}$  m, 1 pm =  $10^{-12}$  m)

**23. Define wave number.**

**Ans:** Wave number ( $\bar{\nu}$ ) is the number of waves per unit length and is reciprocal to wavelength. It is usually measured in units of  $m^{-1}$  or  $cm^{-1}$ .

$$\bar{\nu} = 1/\lambda$$

*24. How the energy of the photon can be calculated from the measurement of the frequency, wavelength or wave number of the photon?*

*Ans:* According to Planck's theory, the energy depends upon the frequency, wavelength and wave number of photons as in the equations

$$E = h\nu$$

$$E = hc/\lambda$$

$$\bar{\nu} = 1/\lambda$$

$$E = hc\bar{\nu}$$

We need two constants for this i.e., 'c' and 'h'

**25. How do you justify that the distances between adjacent orbits of Hydrogen atom go on increasing from lower to the higher orbits?**

**Ans:** From equation of radius for hydrogen atom

$$r = 0.529 \text{ \AA} \times n^2$$

where  $n = 1 \quad r_1 = 0.529 \times (1)^2 = 0.529 \text{ \AA}$

$$n = 2 \quad r_2 = 0.529 \times (2)^2 = 2.11 \text{ \AA}$$

$$n = 3 \quad r_3 = 0.529 \times (3)^2 = 4.75 \text{ \AA}$$

$$n = 4 \quad r_4 = 0.529 \times (4)^2 = 8.4 \text{ \AA}$$

From the above data it is clear that distance gaps go on increasing from lower to higher orbits. Therefore,

$$r_2 - r_1 < r_3 - r_2 < r_4 - r_3 < \dots \dots \dots \text{SO ON.}$$

The second orbit is four times away from the nucleus than first orbit, third orbit is nine times away and similarly fourth orbit is sixteen times away.

26. How do you prove that the energy associated with the electron which is revolving around the nucleus of H-atom is negative?

The work done is the potential energy of electron, so

*Ans:* Work done =  $E_{\text{potential}} = -\frac{Ze^2}{4\pi\epsilon_0 r}$

The total energy (E) of the electron is:

So,  $E = E_{\text{kinetic}} + E_{\text{potential}}$

$$= \frac{1}{2} mv^2 - \frac{Ze^2}{4\pi\epsilon_0 r} \quad \dots\dots\dots (1)$$

Now, we want to eliminate the factor of **velocity from** equation (1). So, from equation (2) substitute the value of  $mv^2$

Since  $mv^2 = \frac{Ze^2}{4\pi\epsilon_0 r} \quad \dots\dots\dots (2)$

$$E = \frac{Ze^2}{8\pi\epsilon_0 r} - \frac{Ze^2}{4\pi\epsilon_0 r}$$

Simplifying it,  $E = -\frac{Ze^2}{8\pi\epsilon_0 r} \quad \dots\dots\dots (3)$

27. How do you come to know that the velocities of electrons in higher orbits are less than those in lower orbits of hydrogen atom?

*Ans:* According to Bohr's proposals, the centrifugal force of the electron is equal to the force of attraction between nucleus and electron.

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\epsilon_0 r^2}$$

Rearranging it,

$$r = \frac{Ze^2}{4\pi \epsilon_0 mv^2}$$

According to equation, the radius of a moving electron is inversely proportional to the square of its velocity. Electron should move faster nearer to the nucleus in an orbit of smaller radius. It also tells that if hydrogen atom has many possible orbits, then the promotion of electron to higher orbits makes it move with less velocity.

**28. Energy of an electron is inversely proportional to 'n<sup>2</sup>' but energy of higher orbits is always greater than those of the lower orbits. Why?**

**Ans:** The formula for the energy of an electron revolving in any orbit is given by the equation:

$$E_n = -2.178 \times 10^{-18} \left[ \frac{1}{n^2} \right] \text{J}$$

This equation gives the energy associated with electron in the n<sup>th</sup> orbit of hydrogen atom. Its negative value shows that greater the value of n greater the energy of electron.

**29. The energy difference between adjacent levels goes on decreasing sharply. Why?**

**Ans:** Substituting the values of n as 1, 2, 3, 4, 5, etc. in equation, we get the energy associated with an electron revolving in 1<sup>st</sup>, 2<sup>nd</sup>, 3<sup>rd</sup>, 4<sup>th</sup> and 5<sup>th</sup> orbits of H-atoms.

$$E_1 = -\frac{1313.31}{1^2} = -1313.31 \text{kJmol}^{-1}$$

$$E_2 = -\frac{1313.31}{2^2} = -328.32 \text{kJmol}^{-1}$$

$$E_3 = -\frac{1313.31}{3^2} = -145.92 \text{kJmol}^{-1}$$

$$E_4 = -\frac{1313.31}{4^2} = -82.08 \text{ kJmol}^{-1}$$

$$E_5 = -\frac{1313.31}{5^2} = -52.53 \text{ kJmol}^{-1}$$

$$E_\infty = -\frac{1313.31}{\infty^2} = 0 \text{ kJmol}^{-1} \text{ (electron is free from the nucleus)}$$

$$E_2 - E_1 = (-328.32) - (-1313.31) = 984.99 \text{ kJmol}^{-1}$$

$$E_3 - E_2 = (-145.92) - (-328.32) = 182.40 \text{ kJmol}^{-1}$$

$$E_4 - E_3 = (-82.08) - (-145.92) = 63.84 \text{ kJmol}^{-1}$$

The data obtained above proves that

$$E_2 - E_1 > E_3 - E_2 > E_4 - E_3$$

Energy difference between lower orbits is greater and energy difference between higher orbits is lower.

### ***30. How Bohr's model of H-atom can help us to justify the ionization potential of H-atom?***

**Ans:** Substituting the values of n as 1, 2, 3, 4, 5, etc. in equation we get the energy associated with an electron revolving in various orbits of hydrogen atom.

$$E_1 = -\frac{1313.31}{1^2} = -1313.31 \text{ kJmol}^{-1}$$

$$E_\infty = -\frac{1313.31}{\infty^2} = 0 \text{ kJmol}^{-1} \text{ (electron is free from the nucleus)}$$

$$E_2 - E_1 = (-328.32) - (-1313.31) = 984.99 \text{ kJmol}^{-1}$$

The energy difference between first and infinite levels of energy is calculated as:

$$E_\infty - E_1 = 0 - (-1313.31) = 1313.31 \text{ kJmol}^{-1}$$

1313.31 kJmol<sup>-1</sup> is the ionization energy of hydrogen. This value is the same as determined experimentally.

### ***31. What are the defects of Bohr's model of atom?***

**Ans:** Following are the defects of Bohr's model of atom:

1. Bohr's theory can successfully explain the origin of the spectrum of H-atom and ions like  $\text{He}^{+1}$ ,  $\text{Li}^{+2}$  and  $\text{Be}^{+3}$ , etc. These are all one electron systems. But this theory is not able to explain the origin of the spectrum of multi-electrons or poly-electrons system like He, Li and Be etc.
2. When the spectrum of hydrogen gas is observed by means of a high resolving power spectrometer, the individual spectral lines are replaced by several very fine lines. The  $\text{H}\alpha$  line in the Balmer series is found to consist of five - component lines. This is called fine structure or multiple structure. The appearance of several lines in a single line suggests that only one quantum number is not sufficient to explain the origin of various spectral lines.
3. Bohr suggested circular orbits of electrons around the nucleus of hydrogen atom, but researches have shown that the motion of electron is not in a single plane, but takes place in three dimensional space. The atomic model is not flat.
4. When the excited atoms of hydrogen (which give an emission line spectrum) are placed in a magnetic field, its spectral lines are further split up into closely spaced lines. This type of splitting of spectral lines is called Zeeman effect. So, if the source which is producing the Na - spectrum is placed in a weak magnetic field, it causes the splitting of two lines of Na into component lines. Similarly, when the excited hydrogen atoms are placed in an electrical field, then similar splitting of spectral lines takes place which is called "Stark effect". Bohr's theory does not explain either Zeeman or Stark effect.

### ***32. What is Zeeman effect and Stark effect?***

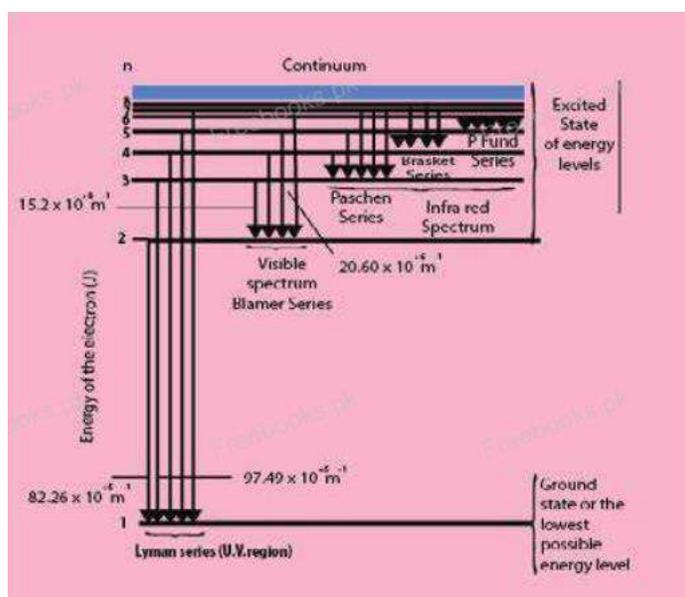
**Ans:** When the excited atoms of hydrogen (which give an emission line spectrum) are placed in a magnetic field, its spectral lines are further split up into closely spaced lines. This type of splitting

of spectral lines is called Zeeman effect. Similarly, when the excited hydrogen atoms are placed in an electrical field, then similar splitting of spectral lines takes place which is called “Stark effect”.

**33. What are the names of the spectral series in hydrogen spectrum?**

**Ans:** These spectral lines of hydrogen spectrum can be classified into five groups called spectral series. These series are named after their discoverers as:

- (i) Lyman series (U.V region)
- (ii) Balmer series (visible region)
- (iii) Paschen series (LR region)
- (iv) Brackett series (I.R region)
- (v) Pfund series (I.R region)



**34. Define spectrum.**

**Ans:** A visual display or dispersion of the components of white light, when it is passed through a

prism is called a spectrum.

**35. What is continuous spectrum?**

**Ans:** In this type of spectrum, the boundary line between the colours cannot be marked. The colours diffuse into each other. One colour merges into another without any dark space. The best example of continuous spectrum is rainbow. It is obtained from the light emitted by the sun or incandescent (electric light) solids. It is the characteristic of matter in bulk.

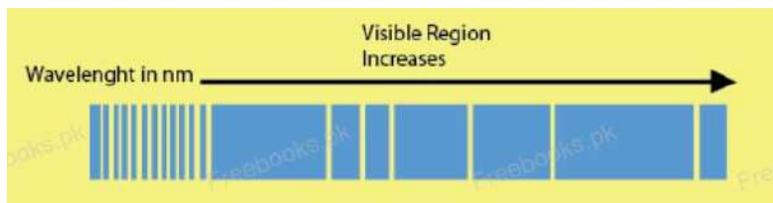
**36. What is line or atomic spectrum?**

**Ans:** When an element or its compound is volatilized on a flame and the light emitted is seen through a spectrometer, we see distinct lines separated by dark spaces. This type of spectrum is called line spectrum or atomic spectrum.

**37. What are the features/characteristics of line spectrum?**

**Ans:** Following are the features of line spectrum:

1. This is characteristic of an atom.
2. The number of lines and the distance between them depends upon the element volatilized. For example, line spectrum of sodium contains two yellow-coloured lines separated by a definite distance. Similarly, the spectrum of hydrogen consists of a number of lines of different colours having different distances from each other. It has also been observed that distances between the lines for the hydrogen spectrum decrease with the decrease in wavelength and the spectrum becomes continuous after a certain value of wavelength.
3. Atomic spectrum can also be observed when elements in gaseous state are heated at high temperature or subjected to an electric discharge.



### Atomic spectrum of hydrogen

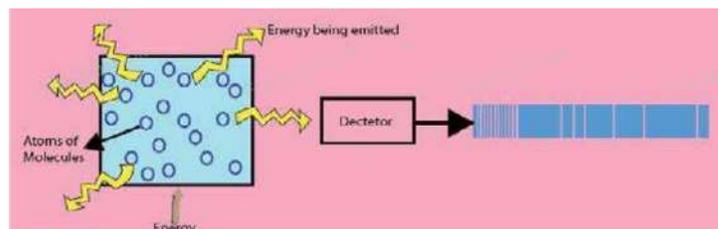
**38. Name the two types of atomic spectrum.**

**Ans:** There are two types of atomic spectrum:

- (i) Atomic emission spectrum
- (ii) Atomic absorption spectrum

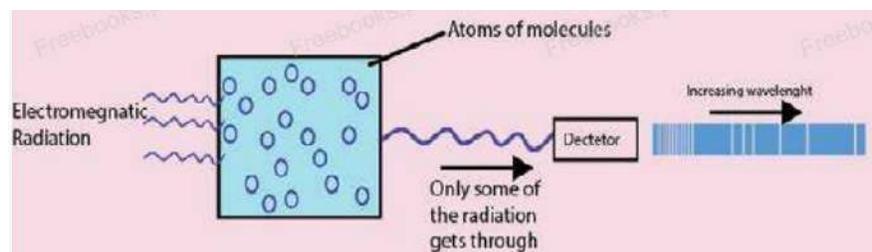
**39. What is atomic emission spectrum?**

**Ans:** When solids are volatilized or elements in their gaseous states are heated to high temperature or subjected to an electrical discharge, radiations of certain wavelengths are emitted. The spectrum of this radiation contained bright lines against a dark background. This is called atomic emission spectrum.



**40. What is atomic absorption spectrum?**

**Ans:** When a beam of white light is passed through a gaseous sample of an element, the element absorbs certain wavelengths while the rest of wavelengths pass through it. The spectrum of this radiation is called an atomic absorption spectrum. The wavelengths of the radiation that have been absorbed by the element appear as dark lines and the background is bright.



41. How wave number of photons is calculated for Lyman series?

First line  $n_1 = 1$  (lower orbit),  $n_2 = 2$  (higher orbit)

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{1^2} - \frac{1}{2^2} \right] = 82.26 \times 10^5 \text{ m}^{-1}$$

Second line  $n_1 = 1$   $n_2 = 3$

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{1^2} - \frac{1}{3^2} \right] = 97.49 \times 10^5 \text{ m}^{-1}$$

Limiting line  $n_1 = 1$   $n_2 = \infty$

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{1^2} - \frac{1}{\infty^2} \right] = 109.678 \times 10^5 \text{ m}^{-1}$$

Ans:

42. How wave number of photons is calculated for Balmer series?

Ans:

First line  $n_1 = 2$ ,  $n_2 = 3$

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{2^2} - \frac{1}{3^2} \right] = 15.234 \times 10^5 \text{ m}^{-1}$$

Second line  $n_1 = 2$   $n_2 = 4$

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{2^2} - \frac{1}{4^2} \right] = 20.566 \times 10^5 \text{ m}^{-1}$$

Third line  $n_1 = 2$   $n_2 = 5$

$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{2^2} - \frac{1}{5^2} \right] = 23.00 \times 10^5 \text{ m}^{-1}$$

Limiting line  $n_1 = 2$   $n_2 = \infty$

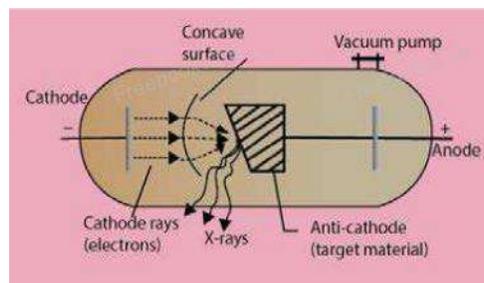
$$\bar{\nu} = 1.09678 \times 10^7 \left[ \frac{1}{2^2} - \frac{1}{\infty^2} \right] = 27.421 \times 10^5 \text{ m}^{-1}$$

#### 43. How are the X-ray series generated?

**Ans:** The X-rays are passed through a slit in platinum plate and then emerged through aluminum window. This is thrown on a crystal of  $\text{K}_4[\text{Fe}(\text{CN})_6]$ , which analyzes the X-ray beam. The rays are diffracted from the crystal and are obtained in the form of line spectrum of X-rays. This is allowed to fall on photographic plate. This line spectrum is the characteristic of target material used. This characteristic X-rays spectrum has discrete spectral lines. These are grouped into K-series, L-series and M-series, etc. Each series has various lines as  $\text{K}_\alpha$ ,  $\text{K}_\beta$ ,  $\text{L}_\alpha$ ,  $\text{L}_\beta$ ,  $\text{M}_\alpha$ ,  $\text{M}_\beta$  etc.

#### 44. On what factors the wavelength of X-rays depends?

**Ans:** The wavelength of X-rays produced depends upon the nature of the target metal. Every metal has its own characteristic X-rays.



#### 45. Detail out the work of Moseley on X-rays.

**Ans:** A systematic and comprehensive study of X-rays was undertaken by Moseley in 1913-1914. His research covered a range of wavelengths 0.04 - 8 Å. He employed thirty-eight different elements from Aluminium to Gold as target in X-rays tube. Moseley was able to draw the following important conclusions from a detailed analysis of the spectral lines which he obtained.

- (i) The spectral lines could be classified into two distinct groups. One of shorter wavelengths are identified by K-series and the other of comparatively longer wavelengths are identified by L-series.
- (ii) If the target element is of higher atomic number the wavelength of X-rays becomes shorter.
- (iii) A very simple relationship was found between the frequency ( $\nu$ ) of a particular line of X-rays and the atomic number  $Z$  of the element emitting it.

$$\sqrt{\nu} = a(Z-b)$$

Here 'a' and 'b' are the constants characteristic of the metal under consideration. This linear equation is known as Moseley's Law. 'a' is proportionality constant and 'b' is called screening constant of the metals.

#### 46. State Moseley's law.

**Ans:** This law states that the frequency of a spectral line in X-ray spectrum varies as the square of atomic number of an element emitting it. This law convinces us that it is the atomic number and not the atomic mass of the element which determines its characteristic properties, both physical

and chemical. If value of  $\sqrt{\nu}$  for K-series are plotted against Z, then a straight line is obtained.

$$\sqrt{\nu} = a(Z-b)$$

**47. Write about importance of Moseley's law.**

**Ans:** Following are the points of importance of Moseley's law:

- (i) Moseley arranged K and Ar, Ni and Co in a proper way in Mendeleev's periodic table.
- (ii) This law has led to the discovery of many new elements like Tc(43), Pr(59), Rh(45).
- (iii) The atomic numbers of rare earths have been determined by this law.

**48. What is meant by dual particle nature of matter?**

**Ans:** According to de-Broglie, all matter particles in motion have a dual character. It means that electrons, protons, neutrons, atoms and molecules possess the characteristics of both the material particle and a wave. This is called wave-particle duality in matter. de-Broglie derived a mathematical equation which relates the wavelength ( $\lambda$ ) of the electron to the momentum of electron.

$$\lambda = \frac{h}{mv}$$

where

$\lambda$  = de-Broglie's wavelength

m = mass of the particle

v = velocity of electron

According to this equation, the wavelength associated with an electron is inversely proportional to its momentum (mv). This equation is called de-Broglie's equation.

**49. Prove that macroscopic bodies don't have waves.**

**Ans:** If we imagine a proton moving in a straight line with the velocity of  $2.188 \times 10^6 \text{ ms}^{-1}$  its wavelength will be 1836 times smaller than that of electron. Similarly, an  $\alpha$ -particle moving with the same velocity should have a wavelength 7344 times smaller as compared to that of electron. Now, consider a stone of mass one gram moving with a velocity of  $10 \text{ ms}^{-1}$  then its wavelength will be:

$$\lambda = \frac{6.626 \times 10^{-34} \text{ Js}}{10^{-3} \text{ kg} \times 10 \text{ ms}^{-1}}$$

$$= 6.626 \times 10^{-30} \text{ m}$$

This wavelength is so small, that it cannot be measured by any conceivable method. It means that heavy material particles have waves associated with them, but they cannot be captured and we say that the macroscopic bodies don't have the waves.

**50. Prove that microscopic bodies can have waves in measurable range.**

**Ans:** Consider an electron which is moving with a velocity of  $2.188 \times 10^6 \text{ ms}^{-1}$  in the first orbit of Bohr's model of hydrogen atom. Then, wavelength associated with it can be calculated as:

$$h = 6.626 \times 10^{-34} \text{ Js}$$

$$m_e = 9.108 \times 10^{-31} \text{ kg}$$

$$\lambda = \frac{6.626 \times 10^{-34} \text{ Js}}{9.108 \times 10^{-31} \text{ kg} \times 2.188 \times 10^6 \text{ ms}^{-1}}$$

$$9.108 \times 10^{-31} \text{ kg} \times 2.188 \times 10^6 \text{ ms}^{-1}$$

$$\text{Since } (J = \text{kg m}^2 \text{ s}^{-2})$$

$$\lambda = 0.33 \text{ nm } (10^9 \text{ m} = 1 \text{ nm})$$

$$\lambda = 0.33 \text{ nm}$$

This value of wavelength ( $\lambda$ ) of electron while moving in the first orbit of H-atom is comparable to the wavelength of X-rays and can be measured.

**51. What is line spectrum of hydrogen and how it differs from continuous spectrum?**

**Ans:** Hydrogen spectrum is an important example of atomic spectrum. The spectrum of hydrogen consists of a number of lines of different colours having different distances from each other. It has also been observed that distances between the lines for the hydrogen spectrum decrease with the decrease in wavelength and the spectrum becomes continuous after a certain value of wavelength

**52. How Bohr's model justifies the H-spectra?**

**Ans:** Lyman and Balmer discovered spectra of hydrogen in 1887. They did not know the reasons for their line. In 1913, Bohr's equation of wave number gives us the values of wave numbers as were measured by Lyman, Balmer, Paschen, Brackett and Pfund. So, Bohr's model can explain the spectrum of hydrogen.

**53. Heisenberg's uncertainty principle has no relation with Bohr's atomic model. Justify it.**

**Ans:** Since the electron has wave nature having elliptical paths so the simultaneous determination of position and momentum is not possible. But Bohr's model does not accommodate the wave nature of electron. He proposes fixed orbits. It means that Bohr's model is very simple as compared to Heisenberg's uncertainty principle.

**54. State Heisenberg's uncertainty principle.**

**Ans:** It states that certainty in the determination of momentum introduces uncertainty in the determination of position and vice versa.

Suppose that  $\Delta x$  is the uncertainty in the measurement of the position and  $\Delta p$  is the uncertainty in the measurement of momentum of an electron, then the mathematical expression is:

$$\Delta x \Delta p \geq \frac{h}{4\pi}$$

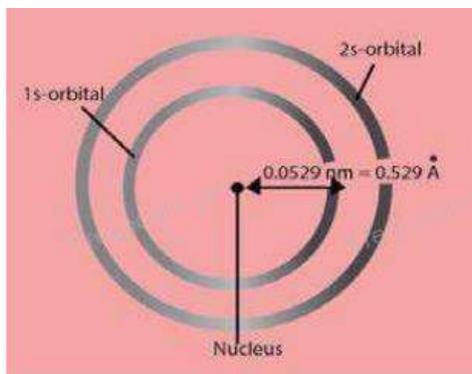
This relationship is called uncertainty principle. This equation shows that if  $\Delta x$  is small then  $\Delta p$  will be large and vice versa. So, if one quantity is measured accurately then the other becomes less accurate. It is applicable to microscopic bodies.

**55. How Compton Effect helps to study uncertainty principle?**

**Ans:** Compton's effect helps us understand the uncertainty principle. Suppose, we need to determine the position of an electron. Visible light cannot help us because the wavelength of visible light is millions time large as compared to the diameter of electron. We have to use X-rays which have very short wavelength as compared to that of visible light. When this photon of X-rays strikes an electron, the momentum of electron is expected to change. Uncertainty of momentum will appear due to change of velocity of electron. Smaller the wavelength of X-rays greater will be the energy of the photon. Hence, the collision of X-rays with electron will bring about the greater uncertainty in momentum. So, an effort to determine the exact position of electron has rendered its momentum uncertain. When we use the photons of longer wavelength to avoid the change of momentum, the determination of the position of electron becomes impossible.

**56. Define orbital.**

**Ans:** The volume of space in which there is 95% chance of finding an electron is called atomic orbital. The term orbital should not be confused with the term orbit as used in the Bohr's theory. The orbital can be regarded as a spread of charge surrounding the nucleus. This is often called the "electron cloud".



**57. Define quantum numbers.**

**Ans:** Quantum numbers are the sets of numerical values which give the acceptable solutions to Schrodinger wave equation for hydrogen atom. An electron in an atom is completely described by its four quantum numbers:

- (1) Principal quantum number ( $n$ )
- (2) Azimuthal quantum number ( $l$ )
- (3) Magnetic quantum number ( $m$ )
- (4) Spin quantum number ( $s$ )

**58. What is the function of principal quantum number?**

**Ans:** Principal quantum number gives information about:

1. Energy of electron
2. Distance of electron from the nucleus
3. Radius of orbit

**59. What is the function of azimuthal quantum number?**

**Ans:** Azimuthal quantum number ( $l$ ) gives us information about the orbitals present in an orbit.

$l=0$  s-orbital

$l=1$  p-orbital

$l=2$  d-orbital

$l=3$  f-orbital

**60. Calculate the number of electrons in s, p, d and f sub-shells from the formula.**

**Ans:** The formula is  $2(2l+1)$

For s-orbital  $2(2 \times 0 + 1) = 2$  electrons

For p-orbital  $2(2 \times 1 + 1) = 6$  electrons

For d-orbital  $2(2 \times 2 + 1) = 10$  electrons

For f-orbital  $2(2 \times 3 + 1) = 14$  electrons

**61. The magnetic quantum number gives orientation of orbital in space. Justify it.**

**Ans:** The value of  $m$  gives us the information of degeneracy of orbitals in space. It tells us the number of different ways in which a given s, p, d or f-subshell can be arranged along x, y and z-axes in the presence of a magnetic field. Thus, different values of 'm' for a given value of 'l', represent the total number of different space orientations for a subshell. In case of s-subshell  $l = 0$ , so,  $m = 0$ . It implies that s-subshell of any energy level has only one space orientation and can be arranged in space only in one way along x, y and z-axes. So, s-subshell is not sub-divided into any other orbital. The shape of s' orbital is such that the probability of finding the electron in all the directions from the nucleus is the same. It is spherical and symmetrical orbital.

**62. Define Hund's rule.**

**Ans:** This rule is concerned with the distribution of electrons in degenerate orbitals. These orbitals may be atomic or molecular. It states that:

If degenerate orbitals are available and more than one electrons are to be placed in them then place them in separate orbitals with the same spin rather than putting them in the same orbital with opposite spins.

**63. State Pauli's exclusion principle.**

**Ans:** According to Pauli's exclusion principle:

No two electrons in an atom can have the same set of four quantum numbers.

**64. What is  $n+l$  rule?**

**Ans:** To determine the energy position of an electron we take the sum of principal and azimuthal quantum number. Greater the value of  $n+l$  means greater the energy of electron.

**65. What is the function of spin quantum number?**

**Ans:** An electron while moving in an orbital around the nucleus also rotates or spins about its own axis either in a clockwise or anti-clockwise direction. This is also called self-rotation. Clockwise motion is shown as  $+1/2$  and anti-clockwise as  $-1/2$ . This is explained by spin quantum number.

**66. Calculate mass of an electron when  $e/m=1.7588 \times 10^{11} \text{Ckg}^{-1}$  OR Calculate mass of an electron from its  $e/m$  value.**

**Ans:**

$$\text{Charge of electron} = 1.602 \times 10^{-19} \text{ C}$$

$$e/m \text{ of electron} = 1.7588 \times 10^{11} \text{ Ckg}^{-1}$$

$$\text{Mass of electron} = \frac{1.602 \times 10^{-19} \text{ C}}{1.7588 \times 10^{11} \text{ Ckg}^{-1}}$$

$$\text{Mass of electron} = 9.1095 \times 10^{-31} \text{ Kg}$$

**67. How do you come to know that velocities of electrons in higher orbits are less than those in lower orbits of hydrogen atom?**

**Ans:** According to Bohr's proposal, the centrifugal force of the electron is equal to the force of attraction between nucleus and electron.

$$mv^2 = \frac{Ze^2}{4\pi\epsilon_0 r}$$

Rearranging

$$r = \frac{Ze^2}{4\pi\epsilon_0 mv^2}$$

The factors are  $Z, 4, \pi, \epsilon_0$  and  $m$  are constant, so

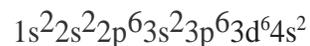
$$r \propto \frac{1}{v^2}$$

According to this equation, radius and velocities are inverse to each other. Greater the velocity of the moving electron smaller is the radius.

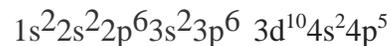
**68. Write down electronic configuration of Fe (26) and Br (35) (Electronic configuration of the complete list at the end of chapter should be done)**

**Ans:**

Fe (26)



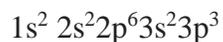
Br(35)



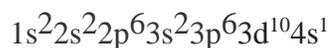
**69. Write down electronic configuration of P(15) and Cu(29).**

**Ans:**

P(15)



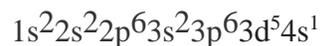
Cu (29)



**70. Write down electronic configuration of Cr (24).**

**Ans:**

Cr (24)



**71. Calculate the number of electrons in s, p, d and f-sub-shells from the formula and write separately.**

**Ans:** Following is the formula for calculating the number of electrons in s,p,d,f  $2(2l+1)$

For s-sub-shell =  $2(2l+1) = 2(2(0)+1) = 2(0+1) = 2(1) = 2$  electrons

f or p-subshell =  $2(2l+1) = 2(2(1)+1) = 2(2+1) = 2(3) = 6$  electrons

f or d-subshell =  $2(2l+1) = 2(2(2)+1) = 2(4+1) = 2(5) = 10$  electrons

f or f-subshell =  $2(2l+1) = 2(2(3)+1) = 2(6+1) = 2(7) = 14$  electrons

**72. The e/m value for positive rays obtained from hydrogen gas is 1836 times less than that of cathode rays. Justify it.**

**Ans:**  $e/m$  ratio depends directly on 'e' and inversely on 'm'.  $e/m$  ratio of positive rays obtained from hydrogen has less value than that of cathode rays, because the mass of electron is 1836 times smaller than the mass of proton ( $H^+$ ) positive particle.

**73. Hydrogen atom and  $He^+$  are mono-electronic systems, but the size of the  $He^+$  is much smaller than H-atom why?**

**Ans:** It is because the atomic number (number of protons) is higher in Helium and lesser in hydrogen. Radius or size of atom or ion depends inversely on its atomic number.

According to Bohr's atomic model,

$$r = 0.529 \text{ \AA} \frac{n^2}{z}$$

For H-atom,	For $He^+$ ,
$r = 0.529 \text{ \AA} \frac{1^2}{1}$	$r = 0.529 \text{ \AA} \frac{1^2}{2}$
$r = 0.529 \text{ \AA}$	$r = 0.26 \text{ \AA}$

**74. Do you think that the size of  $Li^{2+}$  is even smaller than  $He^+$ ? Justify with calculations.**

**Ans:**

The size of  $Li^{2+}$  is even smaller than  $He^+$  because  $Li^{2+}$  has three protons in the nucleus. It has only one electron. Therefore, nucleus will have greater force of attraction for electron and pull it towards itself.

According to Bohr's atomic model,

$$r = 0.529 \text{ \AA} \frac{n^2}{z}$$

For $\text{Li}^{2+}$ ,	For $\text{He}^+$ ,
$r = 0.529 \text{ \AA} \frac{1^2}{3}$	$r = 0.529 \text{ \AA} \frac{1^2}{2}$
$r = 0.176 \text{ \AA}$	$r = 0.26 \text{ \AA}$

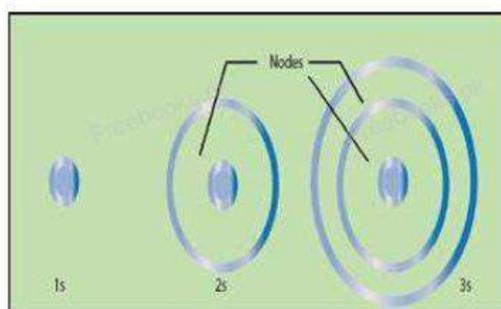
**75. How does the Bohr's model justify the Moseley's equation?**

**Ans:** According to Bohr's atomic model, energy is released in the form of radiations when electrons are de-excited from higher energy levels to lower ones. During the production of X-Rays the same process occurs in the ionized anode's atom, as a result of which X-Rays are produced. Electrons of higher shells fall into K or L or M-shell and produce K-Series or L-Series or M-Series X-Rays.

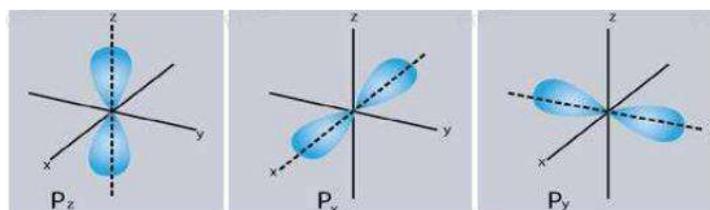
**76. Draw the shapes of s, p and d sub-shells.**

**Ans:**

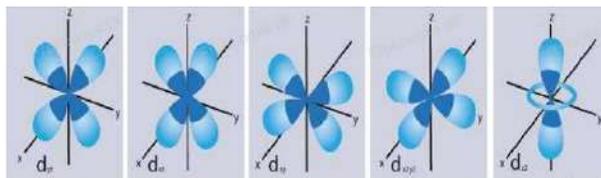
**s-subshell**



**p-sub-shell**



***d-sub-shell***



***77. Define node.***

***Ans:*** The probability of finding the electron is zero between two orbitals.

This place is called nodal plane or nodal surface.

## ***CHEMISTRY (XI)***

### ***Chapter 6***

### ***Chemical Bonding***

### ***Short Questions***

***1. What is octet rule? Give two examples of compounds which deviate from it.***

***Ans:*** The tendency of an atom to attain a maximum of eight electrons in the valance shell is called octet rule.

There are many compounds in which atoms have no eight electrons (octet) in the valance shell after chemical combination. For example,  $\text{BF}_3$ ,  $\text{BCl}_3$ ,  $\text{AlCl}_3$ ,  $\text{SF}_6$ ,  $\text{PF}_5$

***2. Define chemical bond. Give examples.***

***Ans:*** A chemical bond is the force which holds together two or more atoms or ions to form a large variety of compounds e.g. ionic bond, covalent bond.

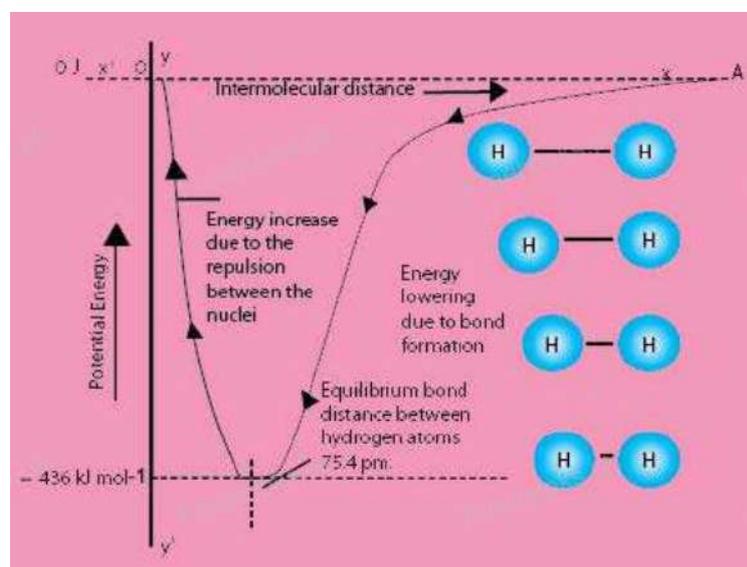
***3. Bond distance is the compromise distance. Justify.***

***Ans:*** When two atoms approach each other then there are two forces that operate i.e. attractive forces and repulsive forces. Attractive forces are between the nucleus of one atom and electrons of the other atom and repulsive forces are between the nuclei and electrons of both the atoms. There is competition between the attractive and repulsive forces till a point is reached where attractive forces dominate and repulsive forces are minimized. This is the point where bond is formed so bond distance is called the compromise distance between two atoms making up a molecule.

**4. The compromise distance between the two bonded hydrogen atoms is 75.4 pm.**

**Justify.**

**Ans:** When two atoms come close to each other, they have attraction as well as repulsion between them. These bonded atoms stay at the distance, where the attraction is maximum. This distance is also called bond length. These bonded atoms cannot fly apart from each other in ordinary conditions. In  $H_2$  molecule compromised distance between the two hydrogen atoms is 75.4pm, which is also called equilibrium bond distance or bond length of  $H_2$  molecule.



**5. Define atomic radius.**

**Ans:** The atomic radius means the average distance between the nucleus of the atom and its outermost electronic shell.

**6. Why atomic radii cannot be determined accurately?**

**Ans:** The radius of an atom cannot be determined precisely due to the following reasons:

- (i) There is no sharp boundary of an atom. The probability of finding an electron never becomes exactly zero even at large distances from the nucleus.
- (ii) The electronic probability distribution is affected by neighbouring atoms. For this reason, the size of an atom may change from one compound to another.

**7. Give trend of atomic radii in the periodic table.**

**Ans:** Atomic radii decrease from left to right in the periodic table because the effective nuclear charge increases from left to right and shielding effect remains constant. It increases from top to bottom in the periodic table because the effective nuclear charge decreases and shielding effect increases.

**8. Define ionic radius.**

**Ans:** The ionic radius of an ion is the average distance between nucleus of ion and the outer shell. It is measured in pm or nm etc.

**9. Why the radius of  $Cl^-$  ion increases from 99 pm to 181 pm OR Why the ionic radius is greater than atomic radius?**

**Ans:** The radius of anion is larger than its original atom. The increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valance shell electrons. We can say that effective nuclear charge decreases by the increase of one electron in the valance shell. This causes expansion of the shells. For example, size of chloride ion  $Cl^-$  is greater than Cl. Size of chlorine atom Cl is 99 pm while size of  $Cl^-$  is 181 pm.

**10. No bond in chemistry is 100% ionic. Justify.**

**Ans:** For a compound to be 100% ionic the electronegativity difference between the elements must be 4.0. So, no bond in chemistry is 100% ionic in nature. Highest ionic characters are present in CsF because Cesium is least electronegative and fluorine is most electronegative element. Difference of electronegativity is  $4 - 0.7 = 3.3$ . Cesium fluoride is 92% ionic and has 8% covalent character in it. NaCl is 72% ionic and 28% covalent.

**11. Define ionization energy. Give an example.**

**Ans:** The ionization energy of an element is the minimum energy required to remove an electron from the outermost orbit of an isolated gaseous atom to form a positive ion. The process is called ionization, e.g.



**12. Give the trend of ionization energy in the periodic table.**

**Ans:** Ionization energy increases from left to right in the periodic table because the atomic size decreases, shielding effect remains constant and effective nuclear charge increases. It decreases from top to bottom because the atomic size increases, shielding effect increases and effective nuclear charge decreases.

**13. Mention the factors affecting ionization energy.**

**Ans:** It is observed that the ionization energies of atoms depend upon the following factors:

- (i) Atomic radius
- (ii) Nuclear charge or proton number of the atom
- (iii) Shielding effect of inner electrons
- (iv) Nature of orbital

**14. Ionization energy is an index to the metallic character. Justify.**

**Ans:** Ionization energy is an index to the metallic character. The elements having low ionization energies are good metals and those having high ionization energies are non-metals. Those with intermediate values are mostly metalloids.

**15. Define electron affinity. Give an example.**

**Ans:** The electron affinity of an atom is the energy released or absorbed when one or more electrons are added to an empty or partially filled orbital of an isolated gaseous atom in its valence energy level to form negative ion. e.g.,



**16. Give the trend of electron affinity in the periodic table.**

**Ans:** Electron affinities increase from left to right in period of the periodic table as the effective nuclear charge increases, shielding effect remains constant and size of atom decreases. There is more attraction of nucleus for outermost shell electrons, so electron affinity increases from left to right. Electron affinity decreases from top to bottom as the effective nuclear charge decreases, shielding effect increases and size of atom increases. There will be less attraction of nucleus for outermost shell electrons, so electron affinity decreases from top to bottom.

**17. Define electronegativity. Give an example.**

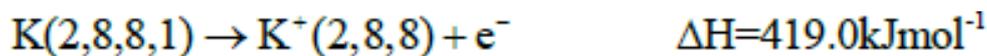
**Ans:** The tendency of an atom to attract a shared electron pair towards itself is called its electronegativity. For example, electronegativity of Hydrogen is 2.1 and that of Fluorine is 4.0.

**18. Give the trend of electronegativity in the periodic table.**

**Ans:** Electronegativity increases from left to right in the periodic table because atomic size decreases, shielding effect remains constant, effective nuclear charge increases. It decreases from top to bottom because atomic size increases, shielding effect increases and effective nuclear charge decreases.

**19. Define ionic bond. Give an example.**

**Ans:** According to the Lewis theory, ionic bond is formed by the complete transfer of one or more electrons from an atom with low ionization energy to another atom with high electron affinity. For example:



**20. Why is the radius of cation smaller than its parent atom?**

**Ans:** Size of cation is smaller than its parent atom because of the imbalance of electron-proton ratio which can be further explained as:

- i. No. of protons are greater than no. of electrons so nuclear charge increases
- ii. Nucleus hold increases on the remaining electrons
- iii. In some cases number of shells also decreases

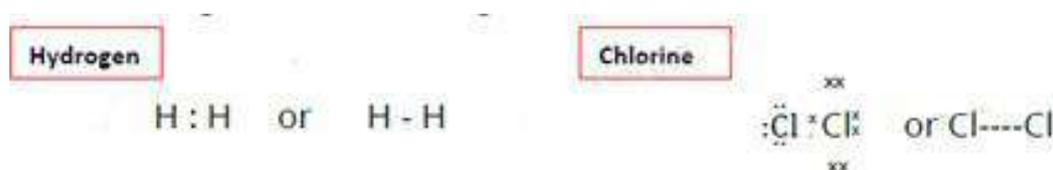
The radius of  $\text{Na}^+$  ion is smaller than Na atom. The radius of  $\text{Na}^+$  ion is 95 pm while Na is 186 pm. By losing one electron, the effective nuclear charge increases and shells shrink to smaller size.

**21. How electronegativity helps us to understand nature of bonds?**

**Ans:** Criteria of electronegativity also helps us to understand the nature of bond. In order to decide the % of ionic nature in a compound, it is better to note the difference of electronegativity between the bonded atoms. If the difference is 1.7 or more then the bond is said to be ionic. Keeping this aspect in view, NaCl has 72% ionic character. CsF has 92% ionic character and calculations tell us that there is no bond with 100% ionic character.

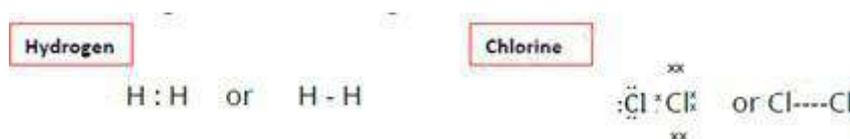
**22. Define covalent bond. Give an example.**

**Ans:** According to Lewis and Kossel, a covalent bond is formed by the mutual sharing of electrons between two atoms. For example:



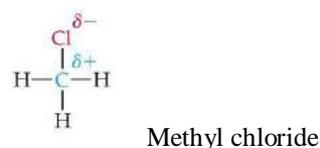
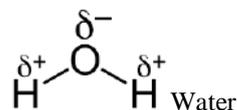
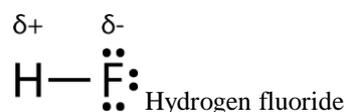
**23. Define non-polar covalent bond. Give an example.**

**Ans:** In such bonds, the bonding electron pairs are equally shared. For example, in H<sub>2</sub> or Cl<sub>2</sub> molecules, the two electrons forming the covalent bond are equally shared by the two identical atoms having same electronegativities. For example:



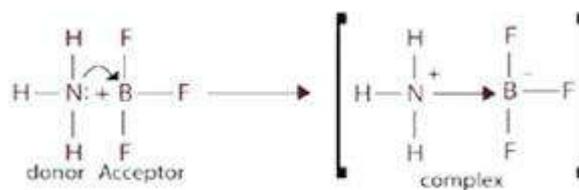
**24. Define polar covalent bond. Give an example.**

**Ans:** When two different atoms are joined by a covalent bond, the electron pair will be displaced towards the more electronegative atom and the pair shall go away from less electronegative atom. In this one atom gets partial negative charge and other partial positive charge. Such bond is considered polar. For example,



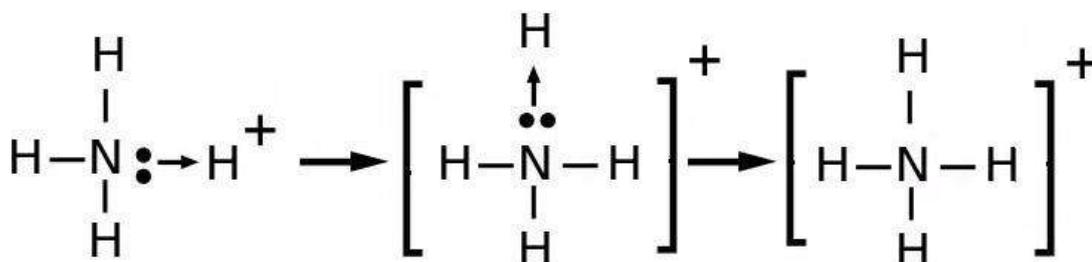
**25. Define coordinate covalent bond. Give an example.**

**Ans:** A coordinate covalent bond is formed between two atoms when the shared pair of electrons is donated by one of the bonded atoms. For example, bond formation between NH<sub>3</sub> and BF<sub>3</sub>.



**26. Distinction between covalent and coordinate covalent bond vanishes after bond formation. Justify.**

**Ans:** Coordinate covalent bond is just like a single bond. For example, this bond is formed by the empty orbital of  $H^+$  and completely filled orbital of Nitrogen. According to Lewis concept, nitrogen donates a pair of electron to  $H^+$ , So  $NH_3$  is a donor while  $H^+$  is acceptor. Usually, this bond is donated by an arrow. Arrow head is from donor towards acceptor.



Experimentally all four N-H bonds are identical in bond length and bond strength. Therefore, ammonium ion is written as  $NH_4^+$  and it is that after bond formation the distinction between coordinate covalent bond and covalent bond vanishes.

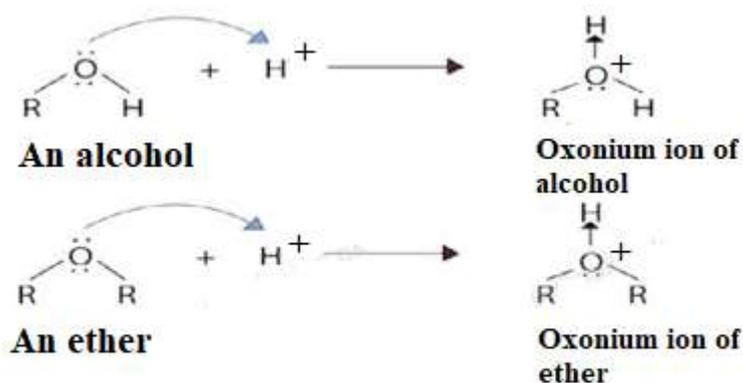
**27. Differentiate between covalent bond and coordinate covalent bond.**

**Ans:**

Covalent Bond	Coordinate Covalent Bond
The bond formed by the mutual sharing of electrons between two atoms is called covalent bonds.	The bond formed when the shared pair of electrons is donated by one of the bonded atoms is called coordinate covalent bond.
Example: CH <sub>4</sub> , H <sub>2</sub> O	Example: NH <sub>4</sub> <sup>+</sup> , BF <sub>4</sub> <sup>-</sup>
Two atoms donate an equal share of electrons.	One atom called donor donates a pair of electrons and one atom called acceptor accepts the pair of electron.

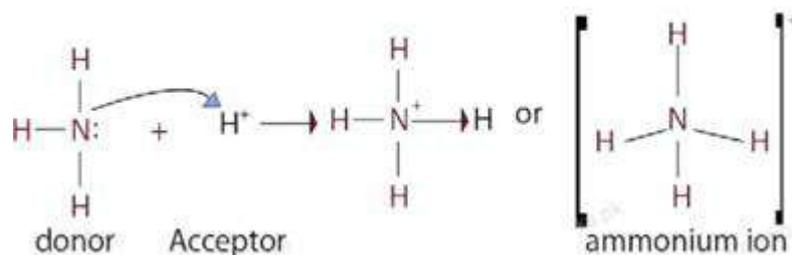
**28. What are oxonium ions? Give an example.**

**Ans:** All the alcohols and ethers offer their lone pairs to H<sup>+</sup> just like water to give coordinate covalent bonds. The ions so produced with positively charged oxygen atom are called oxonium ions.



**29. How NH<sub>4</sub><sup>+</sup> ion is formed?**

**Ans:** Ammonia donates its electron pair to H<sup>+</sup> ion to give NH<sub>4</sub><sup>+</sup> ion. All the four bonds behave alike in NH<sub>4</sub><sup>+</sup> ion.



**30. What is the basic assumption of VSEPR theory?**

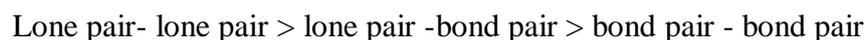
**Ans:** The valence electron pairs (lone pairs and the bond pairs) are arranged around the central atom to remain at a maximum distance apart to keep repulsions at a minimum.

**31. Why the non-bonding electron pairs occupy more space than the bonding electrons?**

**Ans:** A bonding electron pair is attracted by both nuclei of atoms while non-bonding by only one nucleus. Because a lone pair experiences less nuclear attraction, its electronic charge is spread out more in space than that for bonding pair. As a result, the non-bonding electron pairs exert greater repulsive forces on bonding electron pairs and thus tend to compress the bond pairs.

**32. Mention the magnitude of repulsions between the electron pairs in a given molecule.**

**Ans:** The magnitude of repulsions between the electron pairs of a molecule decreases in the following order:



These repulsions are called Van der Waal's repulsions.

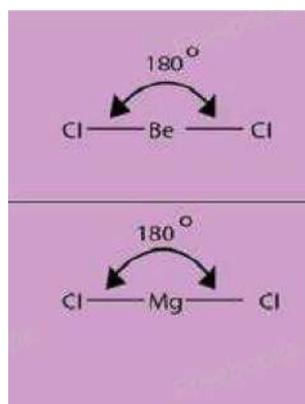
**33. Why multiple bonds are considered as single bond in VSEPR theory?**

**Ans:** The two electron pairs of a double bond and three electron pairs of a triple bond, contain a higher electronic charge density. Therefore, they occupy more space than one electron pair of a single bond but behave like a single electron pair in determining the geometry of the molecule. This is because they tend to occupy the same region between the two nuclei like a single bond.

**34. Explain  $AB_2$  type molecules according to VSEPR theory.**

**Ans:** In such molecules two electron pairs around the central atom are arranged at an angle of  $180^\circ$  in order to minimize repulsions between them. Thus, they form a linear geometry.

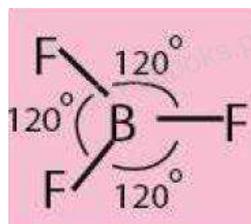
Beryllium chloride is a typical linear molecule which contains two electron pairs.  $MgCl_2$ ,  $CaCl_2$ ,  $SrCl_2$ ,  $CdCl_2$  and  $HgCl_2$  are also linear molecules.



**35. Explain  $AB_3$  type molecules with no lone pairs.**

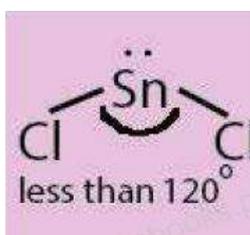
**Ans:** In such molecules, central atom contains three bonding electron pairs, which are arranged at maximum distance apart at a mutual angle of  $120^\circ$ , giving a triangular planar geometry. The boron atom in  $BH_3$  is surrounded by three charge clouds, which remain farthest apart in one plane, each pointing towards the corners of an equilateral triangle. Thus,  $BH_3$  molecules has a trigonal planar geometry, with each H- B-H bond angles of  $120^\circ$ .

Similar geometries are of hydrides of group III-A ( $\text{AlH}_3$ ,  $\text{GaH}_3$ ,  $\text{InH}_3$  and  $\text{TlH}_3$ ) and their halides ( $\text{BF}_3$ ,  $\text{AlCl}_3$ , etc.)



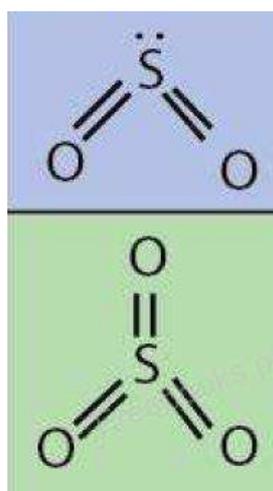
**36. Explain  $\text{AB}_3$  type molecules with one lone pair and two bond pairs.**

**Ans:** In  $\text{SnCl}_2$ , one of the corners of the triangle is occupied by a lone pair, giving rise to a distorted triangular structure in vapor phase.



**37. Explain  $\text{AB}_3$  type molecules with multiple bonds.**

**Ans:** In  $\text{SO}_2$ , one corner of triangle is occupied by a lone pair and two corners each by  $\text{S}=\text{O}$  double bond, while in  $\text{SO}_3$  all three regions, each are occupied by  $\text{S}=\text{O}$  bonds. This structure of  $\text{SO}_3$  is perfectly triangular.

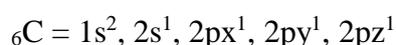


**38. Explain AB<sub>4</sub> type molecules with no lone pairs.**

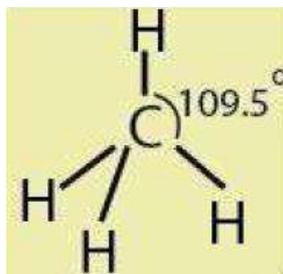
**Ans:** The charge clouds due to four electron pairs avoid their electrostatic repulsions by drifting apart to maintain a mutual bond angle of 109.5°. These molecules have shape of a regular tetrahedron.

**Example**

Each of the four valence electrons of carbon pair up with one electron of hydrogen in methane.



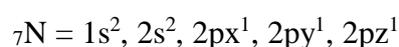
The four electron pairs are directed from the center towards the corners of a regular tetrahedron, with each apex representing a hydrogen nucleus. The arrangement permits a non-planar arrangement of electron pairs. Each H-C-H bond is perfectly 109.5°. Molecules of similar geometry are SiH<sub>4</sub>, GeH<sub>4</sub>, CCl<sub>4</sub>.

**39. Explain AB<sub>3</sub> type molecules with one lone pair and three bond pairs.**

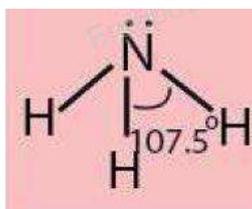
**Ans:** The charge cloud of lone pair electrons (non-bonding electrons) spreads out more than that of bonding electrons.

Large lone pair charge cloud tends to compress the bond angles in rest of the molecules.

Ammonia, NH<sub>3</sub> is a typical example.



The non-bonding electron in 2s orbital takes up more space and exerts a strong repulsive force on the bonding electron pairs. Consequently, to avoid a larger repulsion, the bonding electron pairs move closer that reduces the ideal bond angle from  $109.5^\circ$  to  $107.5^\circ$ . The resultant molecule has triangular pyramidal geometry instead of tetrahedral.



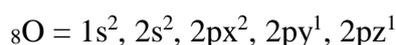
**40. Why the bond angle is reduced to  $102^\circ$  in  $NF_3$ ?**

**Ans:** Substitution of hydrogen in  $NH_3$  with electronegative atoms like F or Cl further reduces the bond angle. In  $NF_3$ , the strong polarity of N-F bond pulls the lone pair of N atom closer to its nucleus. It exerts a stronger repulsion over bonding electrons. The angle further shrinks to  $102^\circ$ . Moreover, the bond pairs N-F bonds are closer to F atoms than N atoms. The increased distances in these bond pairs make their repulsions less operative.

**41. Explain  $AB_4$  type molecules with two lone pairs and two bond pairs.**

**Ans:** Presence of two lone pairs, introduces three types of repulsion i.e., lone pair-lone pair, lone pair-bond pair and bond pair-bond pair repulsion.

For example: water ( $H_2O$ ), a triatomic molecule is expected to be an  $AB_2$  type linear molecule like  $BeCl_2$  and  $CO_2$ . But experimental evidence confirms a bent or angular geometry. VSEPR theory explain geometry of water molecule.



Two of the corners of a tetrahedron are occupied by each of the two lone pairs and remaining by bond pairs. By lone pair-lone pair and lone pair-bond pair repulsions the bond angle is further reduced to  $104.5^\circ$ .

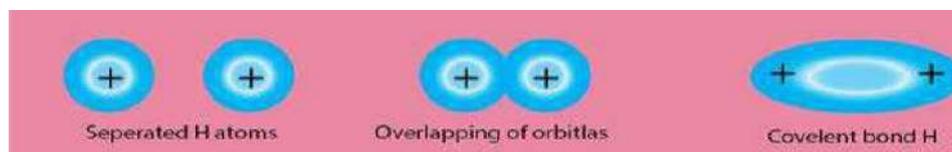
**42. Write two points of Valence bond theory.**

**Ans:** According to valence bond theory,

- i. The partially filled atomic orbitals overlap to form bonds but the individual character of atomic orbitals is retained.
- ii. Greater the overlap, stronger will be the bond formed.

**43. How hydrogen molecule is formed according to VBT?**

**Ans:** As the two atoms of hydrogen approach each other, their 1s orbitals overlap, thereby giving the H-H bond. The electron density becomes concentrated between the two nuclei. The bond is called a sigma ( $\sigma$ ) bond.

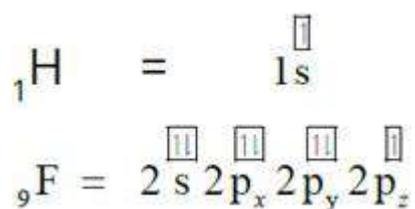


**44. Define sigma bond.**

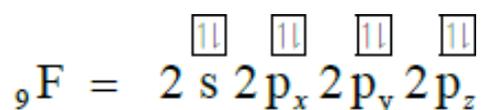
**Ans:** A single bond is formed when two partially filled atomic orbitals overlap in such a way that the probability of finding the electron is maximum around the line joining the two nuclei.

**45. How hydrogen fluoride molecule is formed according to VBT?**

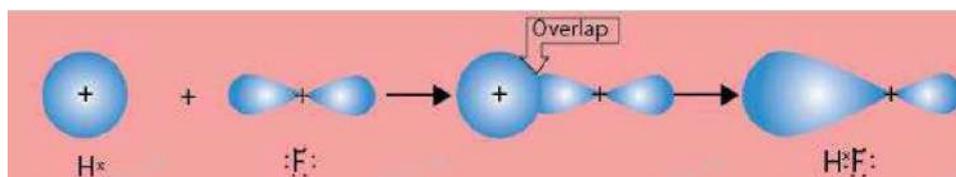
**Ans:** The H-F bond is formed by the pairing of electrons - one from hydrogen and one from fluorine.



The fluorine atom completes its 2p sub-shell by acquiring a share of an electron from hydrogen as shown below.

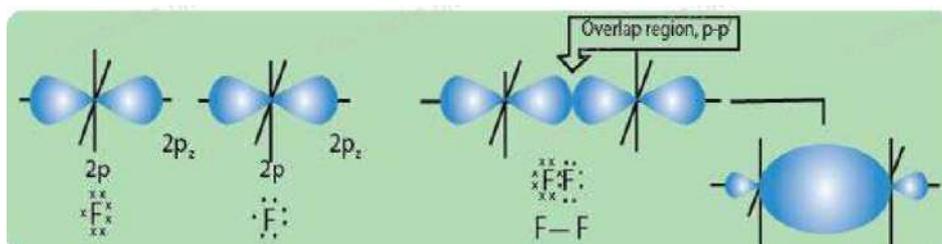


There are then two orbitals plus two electrons whose spins can adjust so they are paired.



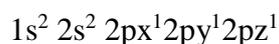
**46. How fluorine molecule is formed according to VBT?**

**Ans:** The sigma bond in the fluorine molecule, F<sub>2</sub> is formed by the overlap of half-filled 2p<sub>z</sub> orbital on each fluorine atom.

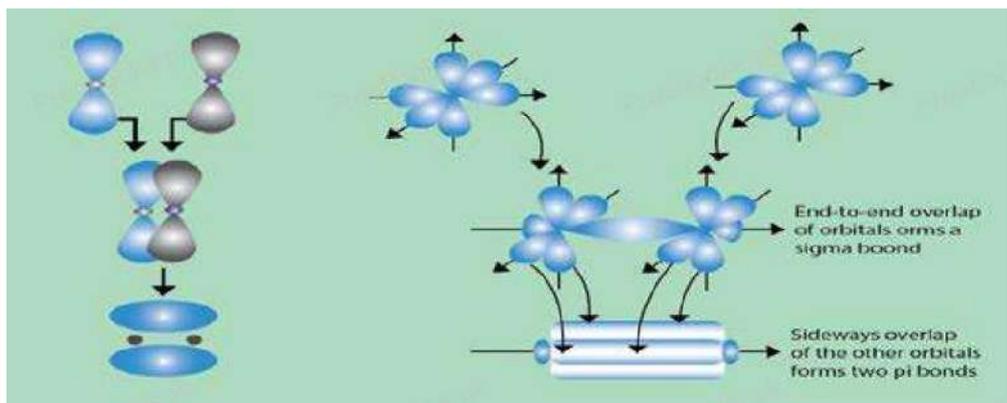


**47. How nitrogen molecule is formed according to VBT?**

**Ans:** Nitrogen atoms have the following electronic configuration

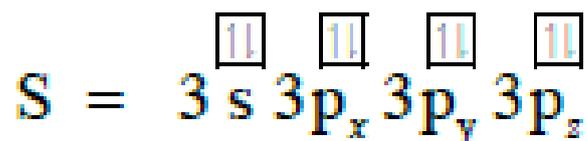


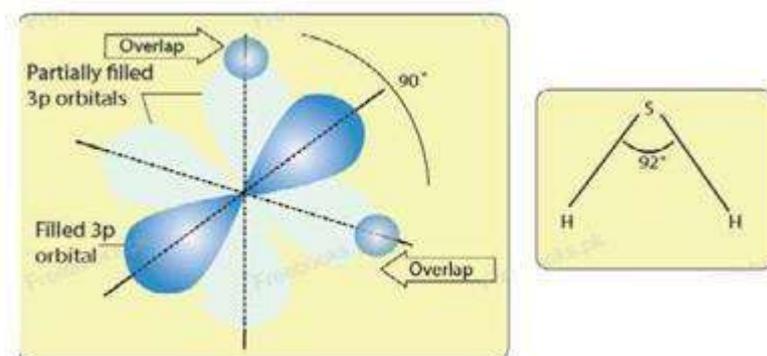
One end-to-end p orbital overlap occurs resulting in a sigma bond, the other two p orbitals are aligned parallel to the corresponding orbitals in the other atom giving two pi bonds.



**48. How H<sub>2</sub>S molecule is formed according to VBT?**

**Ans:** H<sub>2</sub>S is a non-linear molecule, and the bond angle between the two H-S bonds is about 92°. Each two 3p orbitals of sulphur containing one electron can overlap with the 1s orbitals of hydrogen atoms.





**49. Define hybridization.**

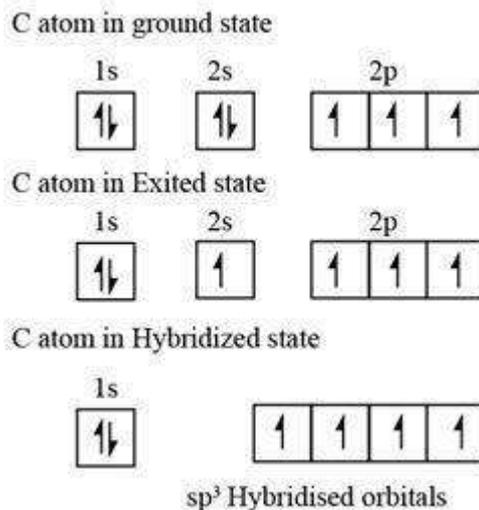
**Ans:** According to this, atomic orbitals differing slightly in energy intermix to form new orbitals of exactly same energies which are called hybrid atomic orbitals. They differ from the parent atomic orbitals in shape and possess specific geometry.

**50. Define  $sp^3$  hybridization. Give examples.**

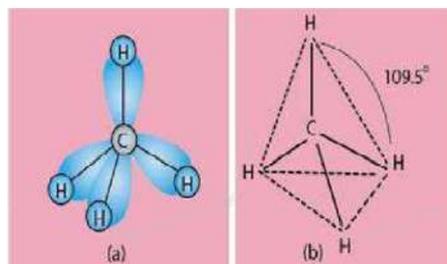
**Ans:** In  $sp^3$  hybridization, one s and three p atomic orbitals of an atom intermix to form four equivalent orbitals called  $sp^3$  hybrid atomic orbitals. Examples  $CH_4$ ,  $NH_3$  and  $H_2O$ .

**51. Discuss structure of methane according to  $sp^3$  (Similar question may be asked for ammonia and water)**

**Ans:** Electronic configuration of  ${}_6C$ , its electronic excitation and hybridization is given as follows:



The four equivalent hybrid orbitals are directed towards the four corners of a regular tetrahedron. The hybrid orbitals are oriented in space in such a manner that the angle between them is  $109.5^\circ$ . They are six in number. Methane molecule is formed by the overlap of  $sp^3$  hybrid orbitals of carbon with 1s orbitals of four hydrogen atoms separately to form four sigma bonds. The tetrahedral structure of  $CH_4$  has four faces, four corners and six edges.



**52. The bond angles of  $H_2O$  and  $NH_3$  are not  $109.5^\circ$  like that of  $CH_4$  although Oxygen and Nitrogen atoms are  $sp^3$  hybridized. Why?**

**Ans:** According to VSEPR theory, lone pairs occupy more space than bond pairs and cause more repulsions.

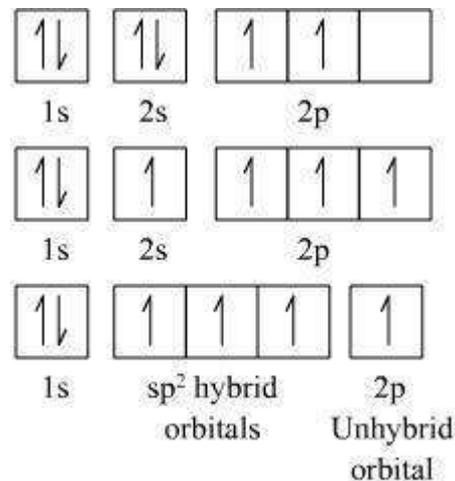
- H<sub>2</sub>O has two lone pairs, so it repels the bond pairs much more and makes bond angle shorter till 104.5°.
- NH<sub>3</sub> has one lone pair that repels the three bond pairs so the bond angle between hydrogen atoms of ammonia is 107.5°.
- CH<sub>4</sub> molecule has no lone pair, and each bond pair repels each other with equal force and bond angle between two adjacent hydrogen atoms becomes 109.5°.

**53. Define  $sp^2$  hybridization. Give examples.**

**Ans:** In  $sp^2$  hybridization, one 's' and two 'p' atomic orbitals of an atom intermix to give three orbitals called  $sp^2$  hybrid orbitals. Example, BF<sub>3</sub> and Ethene.

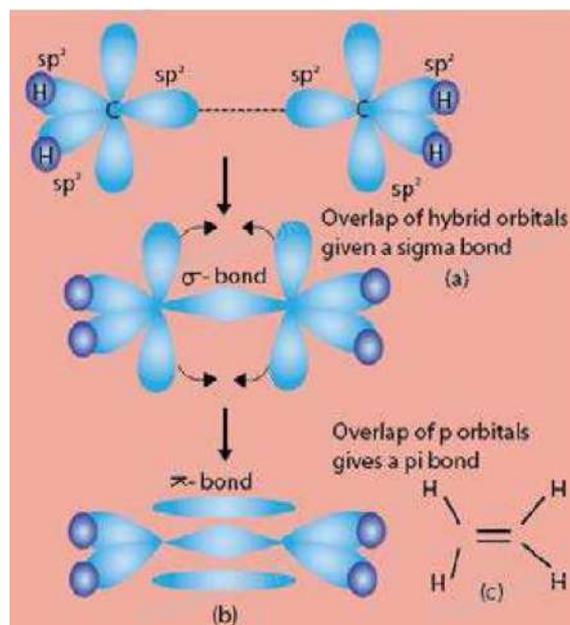
**54. Explain structure of ethene (ethylene) according to hybridization.**

**Ans:** Electronic configuration of <sub>6</sub>C is



Each carbon atom undergoes  $sp^2$  hybridization to form three hybrid orbitals which are coplanar and are oriented at an angle of 120°.

One of the p-orbitals does not take part in hybridization. Each carbon atom undergoes  $sp^2$ -s overlaps with two hydrogen atoms and  $sp^2$ - $sp^2$  overlap between themselves to form sigma bonds. The partially filled p-orbitals undergo sideways overlap to form a  $\pi$ -bond.

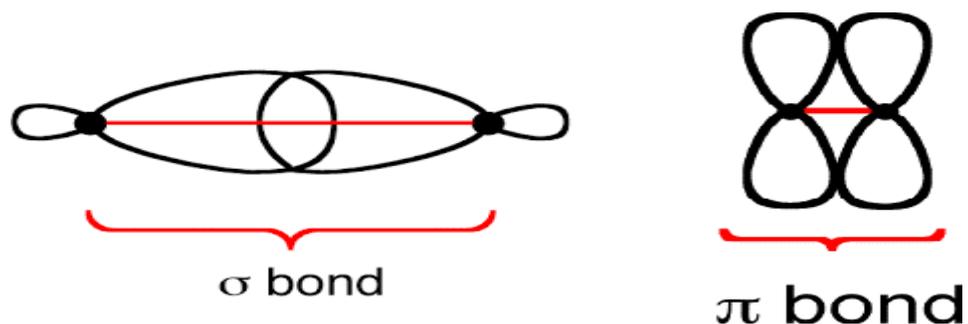


**55. Define pi bond.**

**Ans:** A pi-bond is formed by the sideways overlap of two half-filled co-planar p-orbitals in such a way that the probability of finding the electron is above or below the line joining the two nuclei.

**56.  $\pi$  bonds are more diffused than  $\sigma$  bonds. Why?**

**Ans:** In sigma bonded atoms, the electron density is between two nuclei or at bond axis. In pi bond, electronic cloud lies above and below the line joining the nuclei, so pi bonds are more diffused (spread) than sigma bonds.



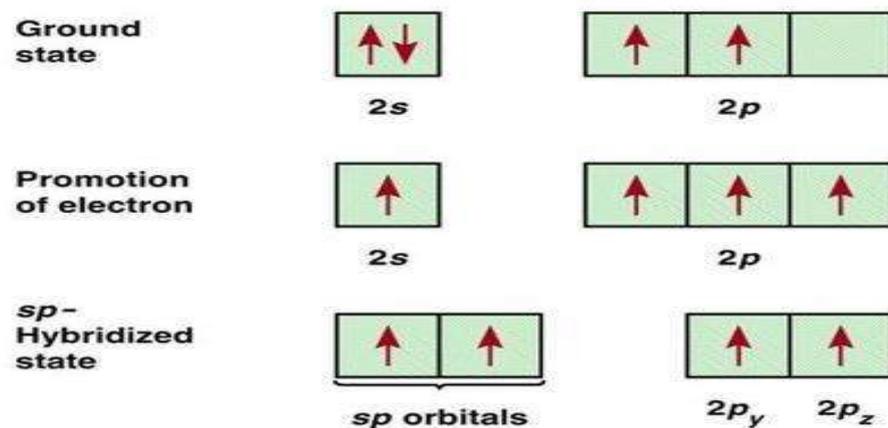
**57. Define sp hybridization.**

**Ans:** In sp hybridization, one 's' and one 'p' orbitals intermix to form two sp-hybrid orbital.

Examples,  $\text{BeCl}_2$  and ethyne.

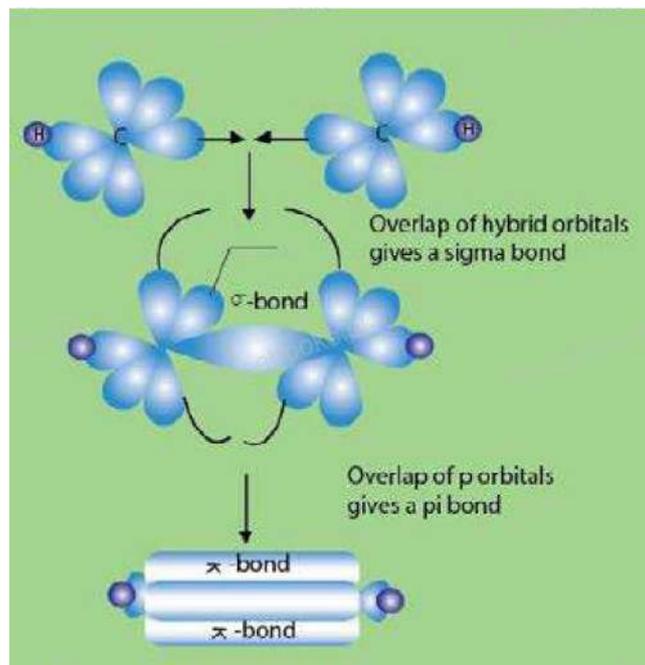
**58. Explain structure of ethyne according to hybridization.**

**Ans:** Electronic configuration of  ${}_6\text{C}$  is



Each carbon atom undergoes sp-s overlap with one hydrogen atom and sp-sp overlap with other carbon atom. Each carbon atom is left with two unhybridized p orbitals perpendicular to the plane of sp hybrid orbitals. The sideways p overlap between the p-orbitals

results in the formation of two pi bonds. Ethyne molecule contains one  $\sigma$  and two pi bonds between the two carbon atoms and each carbon atom is bonded with one H atom through  $\sigma$  bond.



**59. What is the difference between VBT and MOT?**

**Ans:** Following are the points of difference between VBT and MOT:

1. VBT considers overlap of valence shells whereas MOT considers atoms as a whole to participate in bonding.
2. Simplest diagrammatic representation of no bond formation was given by MOT not VBT.
3. MOT very well explains the paramagnetism and diamagnetism phenomena in molecules whereas VBT fails to explain it.
4. Simplest formula to calculate bond order was given by MOT not VBT.

**60. Define bond order. Give an example.**

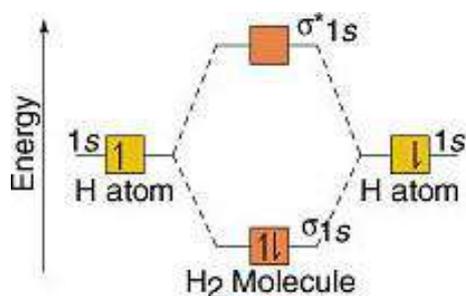
**Ans:** The number of bonds formed between two atoms after the atomic orbitals overlap is called the bond order and is taken as half of the difference between the number of bonding

electrons and anti-bonding electrons. The number of bonds formed between H-atoms in hydrogen molecule may be calculated as follows:

Number of electrons in the bonding orbitals = 2

Number of electrons in the anti-bonding orbitals = 2

$$\text{Bond order} = \frac{2-0}{2} = 1$$

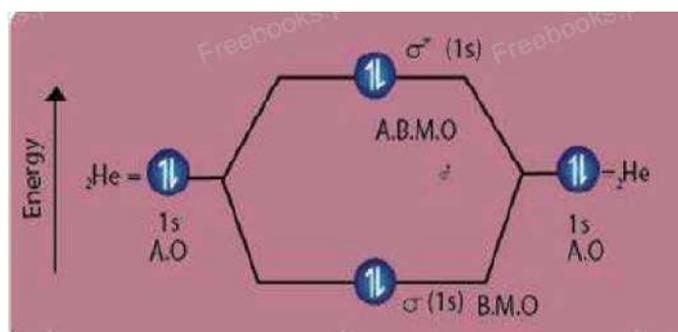


**61. Prove that no bond is formed in helium.**

**Ans:** The electronic configuration of He is 1s<sup>2</sup>. The 1s orbitals of He-atoms combine to form one bonding σ (1s) and one anti-bonding σ\* (1s) orbitals.

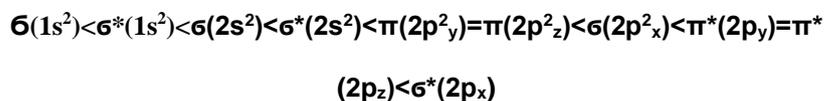
Each He-atom contributes two electrons. Two electrons enter bonding molecular orbital σ(1s) and the remaining two go to antibonding σ\* (1s) molecular orbital. The bond order for He<sub>2</sub> is zero i.e.

$$\frac{2-2}{2} = 0$$



**62. Explain structure of nitrogen according to MOT.**

**Ans:** Electronic configuration of N<sub>2</sub> molecule is:

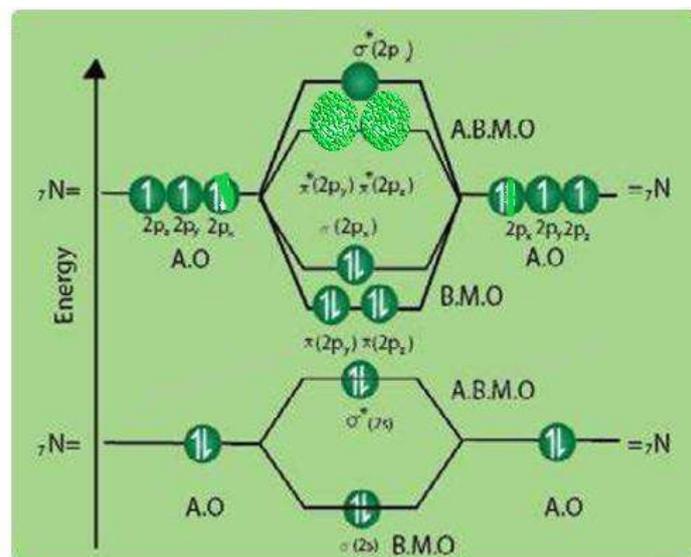


Six electrons enter into three outermost bonding orbitals while no electrons enter into anti-bonding orbitals.

Thus, the bond order in N<sub>2</sub> molecule is

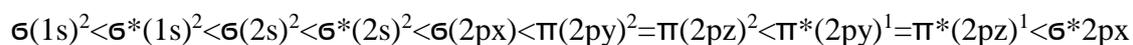
$$\frac{6 - 0}{2} = \frac{6}{2} = 3,$$

which corresponds to the triple bond consisting of one sigma and two pi bonds. The bond dissociation energy of N<sub>2</sub> is very high, i.e. 941 kJ mol<sup>-1</sup>.



**63. Explain structure of oxygen according to MOT.**

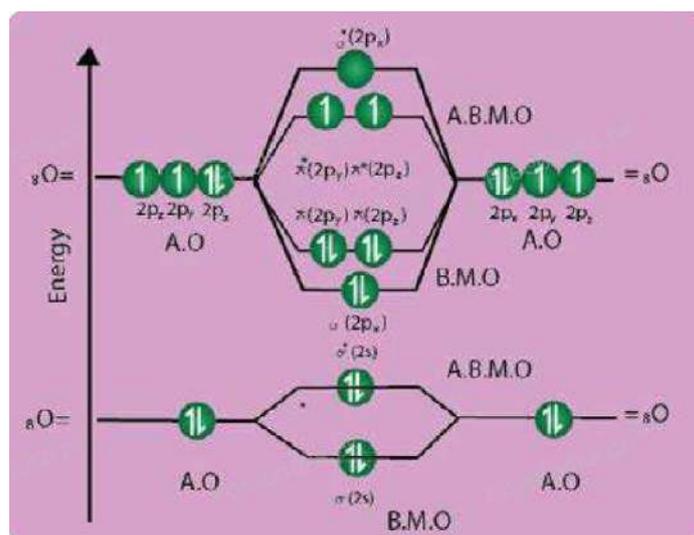
**Ans:** The electronic configuration of O<sub>2</sub> is:



The bond order of O<sub>2</sub> is:

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

This corresponds to a double bond. There are two unpaired electrons in anti-bonding molecular orbitals showing paramagnetic properties of oxygen. Liquid O<sub>2</sub> is attracted towards the magnet.



**64. Differentiate between hybrid orbital and molecular orbital?**

*Ans:*

Hybrid orbital	Molecular orbital
Hybrid orbitals are formed by the interactions of atomic orbitals in the same atom.	Molecular orbitals are formed by the interactions of atomic orbitals of two different atoms.
They have same energies	They may have different energies

**65. Why the energy of anti-bonding molecular orbital is higher than corresponding bonding molecular orbital?**

**Ans:** Anti-bonding orbitals are higher in energy because there is less electron density between the two nuclei. An anti-bonding orbital is formed when two atoms approach each other and the overlap between atomic orbitals results in destructive interference. Thus, when the electrons in an anti-bonding orbital spend less time between the two nuclei, they are at a higher energy level.

**66. Differentiate between atomic orbital and molecular orbital.**

**Ans:**

Atomic orbital	Molecular orbital
Atomic orbital is the region, having the highest probability of finding an electron in an atom.	Molecular orbital is the region, having the highest probability of finding an electron in a molecule.
Formed by the electronic cloud around the atom.	Formed by the fusion of atomic orbitals that have nearly the same energy.

**67. Define Dipole moment and give its S.I units.**

**Ans:** The product of electric charge (q) and the distance between the positive and negative centers (r) is called dipole moment.

$$\mu = q \times r$$

The common unit of dipole moment is Debye (D). The S.I units of dipole moment is coulomb meter (Cm).  $1D = 3.336 \times 10^{-30} \text{ Cm}$

**68. Define bond length. Give an example.**

**Ans:** The distance between the nuclei of two atoms forming a covalent bond is called the bond length. For example, Cl-Cl bond length is 198 pm.

**69. Define bond energy. Give an example.**

**Ans:** The bond energy is the average amount of energy required to break all bonds of a particular type in one mole of the substance. For example, C-C bond energy is 348 kJ/mol.

**70. Why the abnormality of bond length and bond strength in HI is less prominent than that of HCl.**

**Ans:** Electronegativity of Cl is more than that of iodine that is why the difference of electronegativity between the HCl is more than of HI between the bonded atoms. The decrease in polarity from the HCl to HI indicates the increase in trend of equal sharing of electrons due to decreasing electronegativity between the bonded atoms. Therefore, the bond length and bond strength of HCl is more prominent in its abnormality than that of HI.

**71. Why the dipole moment of CO<sub>2</sub> is Zero but that of SO<sub>2</sub> is 1.61 D?**

**Ans:** The dipole moment of CO<sub>2</sub> is zero, as it has a linear structure where the dipoles being equal and opposite cancel out each other's effect.

SO<sub>2</sub> is also a triatomic molecule but it has a lone pair of electron due to which it has angular structure. Polarity is not cancelled out so it has a dipole moment of 1.61 D.

**72. Why the dipole moment of SO<sub>2</sub> is 1.61D but that of SO<sub>3</sub> is zero?**

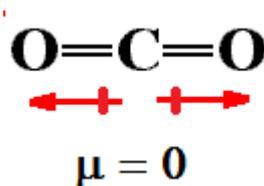
**Ans:** SO<sub>2</sub> is a triatomic molecule but it has a lone pair of electrons, due to which it has angular structure. Polarity is not cancelled out, so it has a dipole moment of 1.61 D.

SO<sub>3</sub> has a dipole moment of 0. SO<sub>3</sub>, a trigonal molecule, has no dipole moment because the bond dipoles cancel each other. This is because the negative and positive centers are both

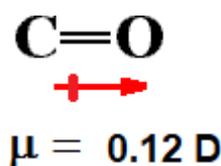
located in the same plane and therefore there is no permanent dipole therefore the molecule is non-polar.

**73. Why the dipole moment of CO<sub>2</sub> is Zero but that of CO is 0.12 D?**

**Ans:** The dipole moment of CO<sub>2</sub> is zero, as it has a linear structure where the dipoles being equal and opposite, cancel out each other's effect.



CO has a permanent pole, hence it shows a dipole moment of 0.12 D.



**74. Why BF<sub>3</sub> is non-polar but SO<sub>2</sub> polar?**

**Ans:** BF<sub>3</sub> has a symmetrical triangular planar molecule, having zero dipole moment so it is a non-polar molecule. On the other hand, SO<sub>2</sub> is also a triatomic molecule but it has a lone pair of electrons, due to which it has angular structure. Polarity is not cancelled out, so it has a dipole moment of 1.61 D and is polar in nature.

**75. How the percentage of ionic character of covalent bond is determined by Dipole moment?**

**Ans:** From the experimentally determined dipole moments, the percentage ionic character in a bond can be calculated.

$$\% \text{age ionic character} = \mu_{\text{observed}} / \mu_{\text{ionic}} \times 100$$

**Example**

The observed dipole moment of HF is 1.90D. The distance between the charges is  $0.917 \times 10^{-10}$  m. (unit positive charge =  $1.6022 \times 10^{-19}$  C)

$$\begin{aligned}\mu_{\text{ionic}} &= q \times r \\ &= 1.6022 \times 10^{-19} \text{ C} \times 0.917 \times 10^{-10} \text{ m} \\ &= 1.469 \times 10^{-29} \text{ Cm} = 4.4 \text{ D} \quad (1\text{D} = 3.336 \times 10^{-30} \text{ Cm})\end{aligned}$$

$$\begin{aligned}\% \text{ ionic character} &= \mu_{\text{observed}} / \mu_{\text{ionic}} \times 100 \\ &= 1.90\text{D} / 4.4\text{D} \times 100 \\ &= 43.2\%\end{aligned}$$

**76. Why the melting points, boiling points, heats of sublimation and heats of vaporization of electrovalent compounds are higher as compared with those of covalent compounds?**

**Ans:** The melting points, boiling points, heats of sublimation and heats of vaporization of electrovalent compounds are usually higher as compared to covalent compounds. In ionic compound, strong inter-ionic forces are present. Large amount of energy is required to break these forces. Ionic compounds are solid in nature while covalent compounds are usually found in solids, liquids and gases. Covalent compounds have less attractive forces between molecules and have less melting points, boiling points, heats of sublimation and heats of vaporization. For example, melting point of an ionic compound NaCl is  $801^\circ\text{C}$  while melting point of a covalent compound  $\text{H}_2\text{O}$  is  $0^\circ\text{C}$ .

**77. Why it is impossible for  $\text{CH}_4$  to make a coordinate covalent bond with  $\text{H}^+$  ion while water and ammonia can do so?**

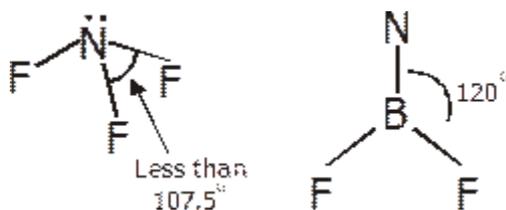
**Ans:** A coordinate covalent bond is formed between two atoms, when the shared pair of electrons is donated by one of the bonded atoms, called donor and accepted by the other bonded atom, called the acceptor.

$H^+$  ion acts as an electron pair acceptor. Ammonia and water both have lone pair of electrons, which they can donate to  $H^+$  ion to form coordinate covalent bond.

Methane  $CH_4$  has only bond pairs, and no lone pair is available so it cannot form coordinate covalent bond.

**78. The molecules  $NF_3$  and  $BF_3$  all have molecular formula of the type  $XF_3$ . But they have different structural formulas. Keeping in view VSEPR theory, sketch the shape of each molecule and explain the origin of differing in shapes.**

**Ans:** The difference in their shapes or geometries is due to a lone pair which is present on nitrogen in  $NF_3$  but not on boron in  $BF_3$ . This lone pair repels the three N-F bonds in  $NF_3$  downwards and shape becomes trigonal pyramidal, whereas, the shape of  $BF_3$  remains triangular planar because of absence of such repulsions.



**79. Why the lone pairs of electrons occupy more space than bond pairs?**

**Ans:** A lone pair is attracted by only one nucleus, on the other hand, a bonding electron pair is attracted by both nuclei of atoms. Because a lone pair experiences less nuclear attraction, its electronic charge is spread out more in space than that of bond pair. As a result, electron pairs are more influential and exert greater repulsive forces on bond pairs, thus tend to compress the bond pairs occupying more space.

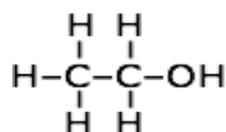
**80. On what factors strength of bond depends?**

**Ans:** The strength of bond depends on

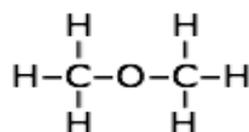
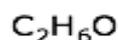
- i. Electronegativity
- ii. Size of atom
- iii. Bond length

**81. Why ionic compounds do not show the phenomenon of isomerism, but covalent compounds do?**

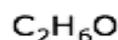
**Ans:** Since ionic bond is non-directional, ionic compounds do not exhibit the phenomenon of isomerism. Covalent bonds are rigid and directional as overlapping of orbitals occurs along definite axes. This leads to the possibility of a variety of isomerism. For example, structural isomerism is shown by the compound,  $C_2H_6O$ .



ethanol



dimethyl ether



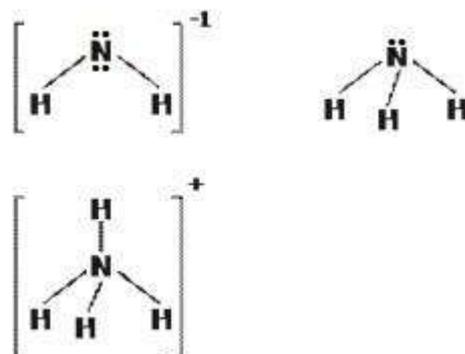
**82. How the type of bonding affects the solubility of compounds.**

**Ans:** Mostly ionic compounds are soluble in water but insoluble in non-aqueous solvents. When a crystal of ionic substance is placed in water, the polar water molecules detach the cation and anion from the crystal lattice by their electrostatic attraction and salt is dissolved in water. If the attraction of the water molecules cannot overcome the attraction between the ions then dissolution does not occur. For the same reasons, non-polar solvents like benzene and hexane do not dissolve ionic compounds. Covalent compounds dissolve easily in non-polar organic solvents like benzene, ether etc. Here the attractive forces of solvent molecules are enough for overcoming the intermolecular forces of attraction.

Compounds like glucose, sucrose, urea etc. dissolve in water due to hydrogen bonding.

**83. The species  $\text{NH}_2^-$ ,  $\text{NH}_3$ ,  $\text{NH}_4^+$ , have bond angles of  $105^\circ$ ,  $107.5^\circ$  and  $109.5^\circ$  respectively. Justify these values by drawing their structures.**

**Ans:** This is due to presence of lone pairs on central atom of these species (except  $\text{NH}_4^+$ ). They repel the bond pairs, thereby, decreasing the bond angles. Perfect tetrahedral bond angle is shown by  $\text{NH}_4^+$  only.



**84. The linear geometry of  $\text{BeCl}_2$  suggests that central Be atom is  $sp$ - hybridized. What type of hybridization a central atom undergoes, when the atoms bonded to it are located at the corners of (a) an equilateral triangle (b) a regular tetrahedron and (c) triangular bipyramidal?**

**Ans: (a)**

In case of equilateral triangle, the central atom should be  $sp^2$  hybridized e.g. Boron in  $\text{BF}_3$ .

**(b)**

In case of regular tetrahedron, the central atom should be  $sp^3$  hybridized e.g. Carbon in  $\text{CH}_4$ .

**(c)**

In case of Triangular bipyramidal, the central atom should be  $dsp^3$  hybridized  
e.g. Phosphorous in  $PCl_5$ .

**85. How does molecular orbital theory explain the paramagnetic character of  $O_2$ ,  $O_2^{2+}$  and  $O_2^{2-}$  species?**

**Ans:**

**$O_2$ :**

$O_2$  will be paramagnetic because it has 2 unpaired electrons in anti-bonding molecular orbitals.

**$O_2^{2+}$ :**

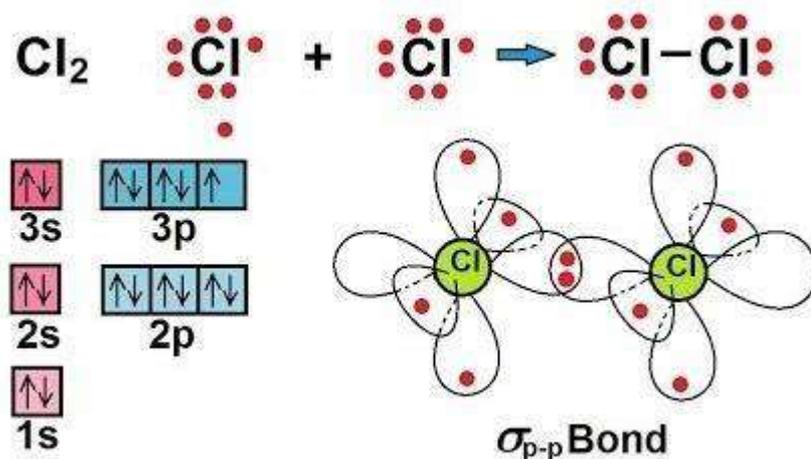
$O_2^{2+}$  will be diamagnetic because it has no unpaired electrons.  $O_2$  after losing two electrons becomes  $O_2^{2+}$

**$O_2^{2-}$ :**

$O_2^{2-}$  will be again diamagnetic because it has no unpaired electrons.  $O_2$  after gaining two electrons becomes  $O_2^{2-}$ .

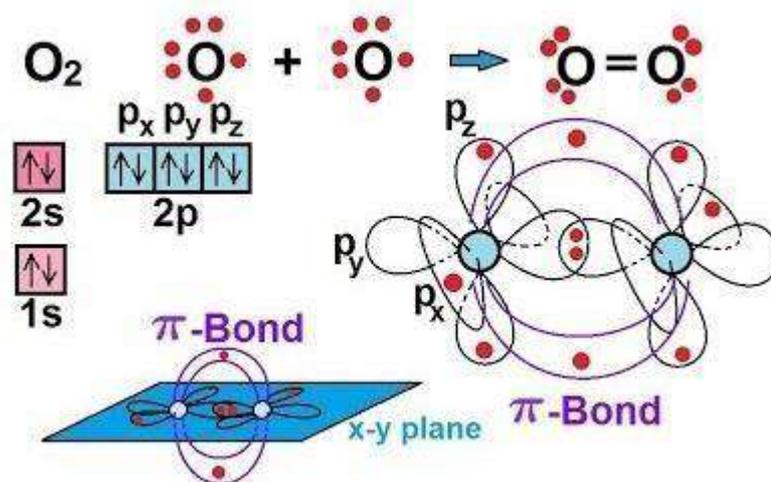
**86. How the bonding in  $Cl_2$  is explained according to valence bond theory?**

**Ans:** In  $Cl_2$  formation, two  $3p_x$  orbitals of two chlorine atoms overlap with each other to make a single covalent bond.



87. How bonding in  $\text{O}_2$  is explained according to valence bond theory?

**Ans:** In  $\text{O}_2$  formation, two  $2p_x$  orbitals of two oxygen atoms overlap with each other to make a single sigma covalent bond and later on two  $2p_y$  orbitals and two  $2p_z$  orbitals engage in the formation of two pi bonds.



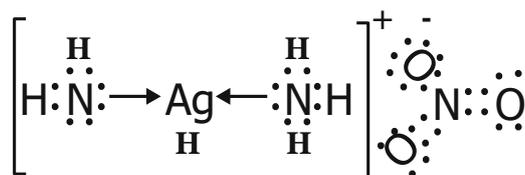
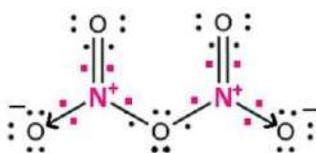
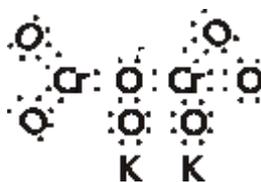
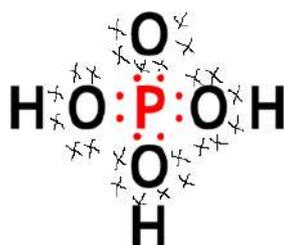
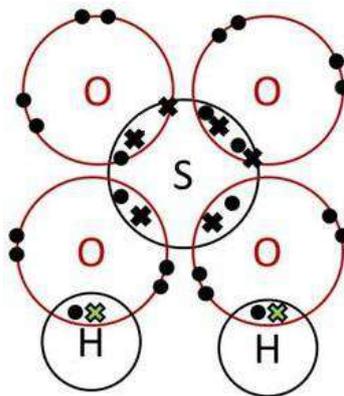
88. Write the Lewis structures for the following compounds:

- (i)  $\text{HCN}$  (ii)  $\text{CCl}_4$  (iii)  $\text{CS}_2$  (iv)  $\text{H}_3\text{N} \rightarrow \text{AlF}_3$  (v)  $\text{NH}_4\text{OH}$  (vi)  $\text{H}_2\text{SO}_4$  (vii)  $\text{H}_3\text{PO}_4$  (viii)  $\text{K}_2\text{Cr}_2\text{O}_7$  (ix)  $\text{N}_2\text{O}_5$  (x)  $\text{Ag}(\text{NH}_3)_2\text{NO}_3$

**Ans:**

$\text{HCN}$





89. How do you compare the bond strength of:

- (i) Polar and non-polar molecules      (ii)  $\sigma$  and  $\pi$  bonds?

Ans: (i)

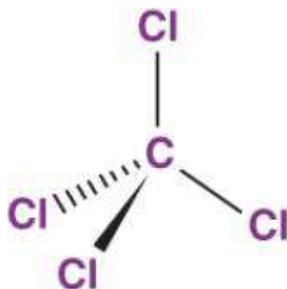
The polar bond has more strength as compared with non-polar bond because of attractions between partial positive and negative ends of the molecule.

(ii)

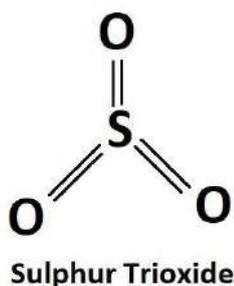
Sigma bond is formed by the linear overlap of half-filled orbitals of atoms, thus, it is less diffused (less spread) and stronger than pi bond which is formed by parallel overlap of half-filled orbitals above and below the plane of molecule making it more diffused and weak.

90. Which of the following molecules will be polar or non-polar, sketch the structures and justify your answer? (i)  $\text{CCl}_4$  (ii)  $\text{SO}_3$  (iii)  $\text{NF}_3$  (iv)  $\text{SO}_2$

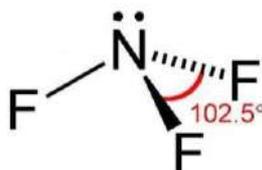
Ans:  $\text{CCl}_4$ : non-polar due to zero dipole moment resulting in tetrahedral structure.



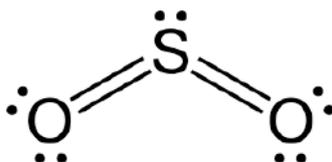
$\text{SO}_3$ : non-polar due to zero dipole moment resulting in triangular planar structure.



$\text{NF}_3$  : polar due to dipole moment resulting in trigonal pyramidal structure.



$\text{SO}_2$  shall be polar due to resultant dipole moment (1.61D) of its angular structure.



### 91. How Pauling calculated inter-ionic radii?

**Ans:** Let us consider, the positive and negative ions, which are held together by electrostatic forces of attraction in a crystal lattice.

$r^+$  and  $r^-$  are the values of radii of cation and anion, respectively. The interionic distance 'R' will be:

$$R = r^+ + r^-$$

Pauling was able to determine the distance between  $\text{K}^+$  and  $\text{Cl}^-$  ions in potassium chloride crystal and found that it was equal to the sum of the radii of the two ions.

$$R = 133\text{pm} + 181\text{ pm} = 314\text{ pm}$$

Thus, the ionic radius appeared to be an additive property. Pauling extended this concept to other  $\text{K}^+$  salts and calculated the radii of other ions from the relationship:

$$r^- = R - r^+$$

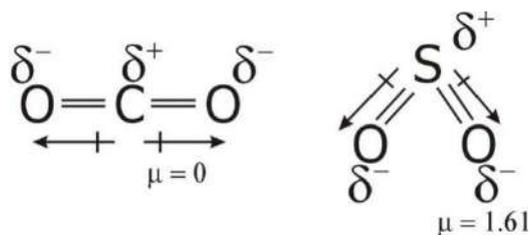
Similarly, the ionic radii of different cations can also be determined.

### 92. How the covalent radius of bonded atoms can be calculated?

**Ans:** The covalent radius of an atom can be used to determine the covalent radius of another atom. For example, the experimentally determined bond length of C-Cl in  $\text{CH}_3\text{Cl}$  is 176.7 pm. The covalent radius of Cl-atom is 99.4 pm. Covalent radius of C-atom can be calculated by subtracting this value from C-Cl bond length. So, the covalent radius of C-atom =  $176.7 - 99.4 = 77.3$  pm.

**93. The dipole moments of  $\text{CO}_2$  and  $\text{CS}_2$  are zero but that of  $\text{SO}_2$  is 1.61D.**

**Ans:** The dipole moments of  $\text{CO}_2$  and  $\text{CS}_2$  are zero because of their linear structures in which resultant dipole moment is zero due to cancellation of dipole moment on both sides. In  $\text{SO}_2$ , a lone pair is present on sulphur introducing repulsions on bond pairs and molecular geometry becomes angular which will have a resultant dipole moment of 1.61 D.



## ***CHEMISTRY (XI)***

### ***Chapter 7***

### ***Thermochemistry***

### ***Short Questions***

***1. Define thermochemistry.***

***Ans:*** The study of heat changes accompanying a chemical reaction is known as thermochemistry.

***2. What is thermochemical equation? Give two examples.***

***Ans:*** A balanced chemical equation which shows not only the reactants and products but also the amount of heat energy absorbed or released is known as a thermochemical equation.



***3. What are thermochemical reactions? Give their types.***

***Ans:*** The chemical reactions which involve heat changes are called as thermodynamic reactions. Most chemical reactions are thermochemical reactions because energy of reactants is never equal to that of products and difference is either released to surrounding or absorbed from surrounding. There are two types of such reactions:

- ✓ Exothermic reactions: Exo means out and therm means heat. It is a reaction in which heat energy is released.
- ✓ Endothermic reactions: Endo means in and therm means heat. It is a reaction in which heat energy is absorbed.

**4. What information thermochemical equations convey?**

**Ans:** The important information that can be drawn from thermochemical equations is:

- Nature of reaction is predicted, whether it is endothermic or exothermic.
- Exact amount of net heat released or absorbed in the reaction.

**5. What is spontaneous process? Give two examples.**

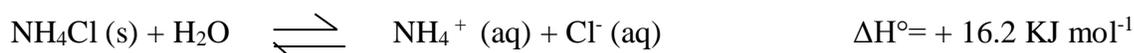
**Ans:** The process which takes place on its own without any outside assistance and moves from a non-equilibrium state towards an equilibrium state is termed as spontaneous process.

Example:

1. Evaporation of water



2. Dilution of  $\text{NH}_4\text{Cl}$  in water



**6. What is the difference between a spontaneous and a non-spontaneous reaction?**

**Ans:**

**Spontaneous reaction**

A process which takes place on its own without any outside assistance and moves from a non-equilibrium state towards an equilibrium state is termed as spontaneous process or natural process. It is unidirectional, irreversible and a real process. Some examples of spontaneous processes are given below:

- (i) Water flows from higher level to the lower level. The flow cannot be reversed without some external aid.

(ii) Neutralization of a strong acid with a strong base is a spontaneous acid-base reaction.



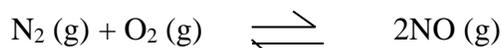
### Non-spontaneous reaction

Non-spontaneous process is the reverse of the spontaneous process. It does not take place on its own and does not occur in nature. Reversible processes constitute a limiting case between spontaneous and non-spontaneous processes. Some non-spontaneous processes can be made to take place by supplying energy to the system from external source. Some examples of non-spontaneous processes are given below:

(i) Pumping of water uphill.

(ii) Transfer of heat from cold interior part of the refrigerator to the hot surroundings.

(iii) When nitrogen reacts with oxygen, nitric oxide is formed. This reaction takes place by the absorbance of heat. Although,  $\text{N}_2$  and  $\text{O}_2$  are present in air, but they do not react chemically at ordinary conditions. The reaction takes place when the energy is provided by lightening.

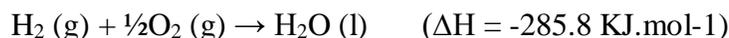


**7. What are endothermic and exothermic reactions? Give examples.**

**Ans:**

### Exothermic Reactions

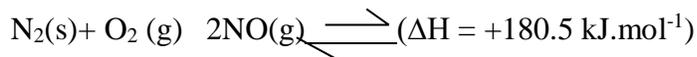
In this reaction reactants are at higher energy in the beginning of a reaction and products obtained are at lower energy after reaction. The difference in the energy of products and reactants appears in the form of heat released by the system in the surroundings.



### Endothermic Reactions

In this reaction reactants are at lower energy before reaction starts and products obtained are at

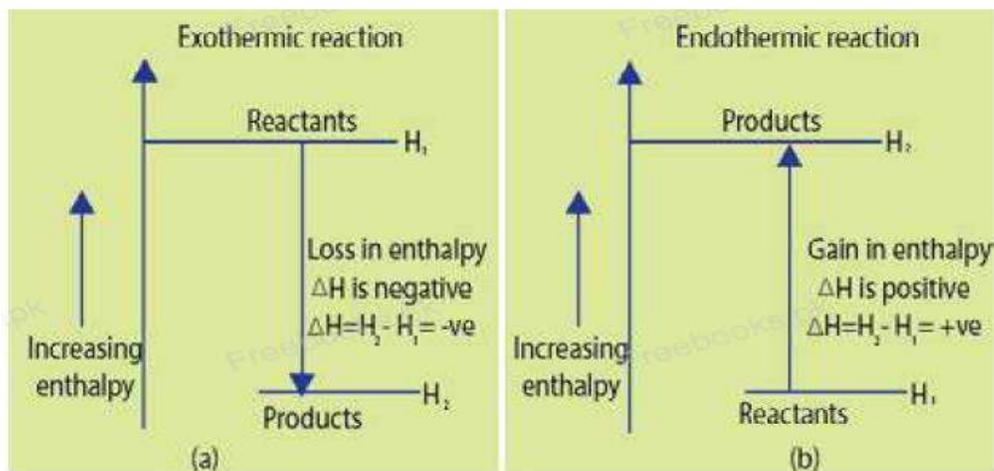
higher energy after reaction stops. The difference in the energy of products and reactants appears in the form of heat absorbed by the system from the surroundings.



### 8. Differentiate between endothermic and exothermic reactions.

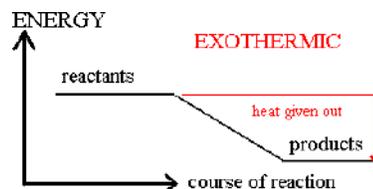
**Ans:** Following are the differences between exothermic and endothermic reactions:

Sr No	Exothermic reactions	Endothermic reactions
1	In this reaction reactants are at higher energy in the beginning of a reaction and products obtained are at lower energy after reaction. The difference in the energy of products and reactants appears in the form of heat released by the system in the surroundings.	In this reaction reactants are at lower energy before reaction starts and products obtained are at higher energy after reaction stops. The difference in the energy of products and reactants appears in the form of heat absorbed by the system from the surroundings.
2	The reactions in which heat energy is evolved	The reactions in which heat energy is absorbed
3	The enthalpy change is shown with a negative sign	The enthalpy change is shown with a positive sign
4	Example $\text{C}(\text{s}) + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) \quad \Delta H = -393.7 \text{ kJ/mol}$	Example $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{NO}(\text{g}) \quad \Delta H = +180.51 \text{ kJ/mol}$



**9. Why heat energy is released in exothermic reactions?**

**Ans:** In exothermic reactions the energy content of products is lesser than that of reactants. Therefore, during reaction heat is released from system to surroundings.



**10. Define system, surrounding, boundary, state and state function.**

**Ans: System**

The term system is used for anything (materials) under test in the laboratory or under consideration in the classroom for the purpose of argument.

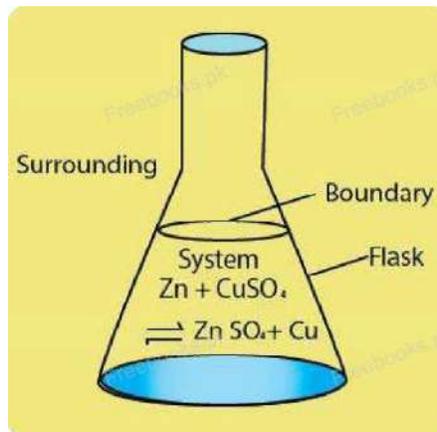
**Surrounding**

The remaining portion of the universe other than the system is known as its surroundings.

**Boundary**

The real or imaginary surface separating the system from the surroundings is called the boundary.

**Example**



**Diagram for system, surrounding and boundary**

### **State**

The state of a system is the condition of a system.

There are two states

1. Initial state
2. Final state

**Example:** Initial temperature  $T_1$  and final temperature  $T_2$

### **State Function**

A state function is a macroscopic property of a system which has some definite values for initial and final states and which is independent of the path adopted to bring about a change. By convention, we use capital letters as symbols for a state function, e.g. pressure (P), temperature (T), volume (V), internal energy (E) and enthalpy (H) are all state functions.

#### ***11. Define internal energy.***

**Ans:** The total of all the possible kinds of energies of the system i.e kinetic energy and potential energy is called its internal energy, E.

### **Mathematically**

$$E = \Sigma \text{ kinetic energy} + \Sigma \text{ potential energy}$$

**12. Define heat and work.**

**Ans:**

**Heat**

Heat is not a property of a system. It is therefore not a state function. It is defined as the quantity of energy that flows across the boundary of a system during a change in its state due to the difference in temperature between the system and the surroundings. Heat evolved or absorbed by the system is represented by a symbol **q**.

The sign of **q** is positive when heat is absorbed by the system from surroundings and it is negative when heat is absorbed by the surroundings from the system.

**Work**

Work is also a form in which energy is transferred from one system to another. It is defined as the product of force and distance i.e.

$$W = F \times S$$

Work is measured in **Joules in SI units**.

Work is not a state function. The sign of **W** is positive when work is done on the system and it is negative when work is done by the system.

**13. State first law of thermodynamics.**

**Ans:** The first law of thermodynamics, also called the law of conservation of energy, states that energy can neither be created nor destroyed, but can be changed from one form to another. In other words, a system cannot destroy or create energy. However, it can exchange energy with its surroundings in the form of heat and work.

**14. Prove that  $\Delta E = q_v$** 

**Ans:** Consider a gas enclosed in a cylinder having a frictionless piston. When a quantity of heat  $q$  is supplied to the system, its internal energy  $E_1$  changes to  $E_2$  and piston moves upwards. The change in internal energy  $\Delta E$  is given by the following equation.

$$\Delta E = E_2 - E_1 = q + w$$

$$\Delta E = q + w$$

In this equation  $q$  represents the amount of heat absorbed by the system and  $w$  is the work done by the system in moving the piston up. If  $w$  is pressure-volume work then the above expression assumes the following form:

$$\Delta E = q - P\Delta V$$

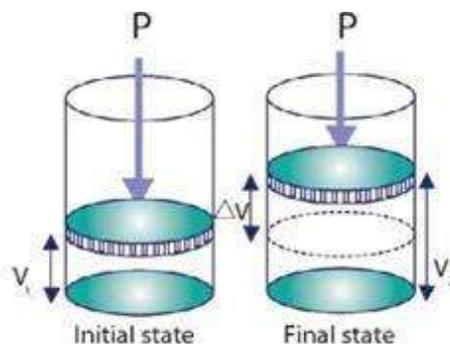
(1)

When the piston is kept in its original position or the volume of the gas is not allowed to change then  $\Delta V = 0$  will take the following form:

$$\Delta E = q_v$$

(2)

This eq (2) shows that a change in internal energy of a system at constant volume is equal to heat absorbed by the system ( $q_v$ ).



**15. Define enthalpy.**

**Ans:** To take account of increase in internal energy and accompanying work done by the gas, there is a property of the system called enthalpy or heat content. It is represented by H. In general, enthalpy is equal to the internal energy E plus the product of pressure and volume (PV).

$$H = E + PV$$

Enthalpy is a state function. It is measured in joules.

**16. Prove that  $\Delta H = q_p$** 

**Ans:** Enthalpy is equal to the internal energy, E plus the product of pressure and volume (PV).

$$H = E + PV$$

Enthalpy is a state function. It is measured in joules. It is not possible, to measure the enthalpy of a system in a given state. However, change in enthalpy ( $\Delta H$ ) can be measured for a change in the state of system. A change in enthalpy of a system can be written as:

$$\Delta H = \Delta E + \Delta(PV)$$

$$\text{or } \Delta H = \Delta E + V\Delta P + P\Delta V$$

Since the gas is kept at constant pressure = 0

Hence

$$\Delta H = \Delta E + P\Delta V \dots \dots \dots (1)$$

**17. In case of liquids and solids, the changes in state do not cause significant volume change**

**i.e.  $\Delta V = 0$ . For such process,  $\Delta H$  and  $\Delta E$  are approximately the same i.e.  $\Delta H \approx \Delta E$**

**Ans:** According to first law of thermodynamics:

$$\Delta E = q + w$$

If w is pressure - volume work done by the system, then:

$$w = - P\Delta V$$

$$\text{So } \Delta E = q - P\Delta V$$

Putting the value of  $\Delta E$  in equation (1) we get:

$$\Delta H = q - P\Delta V + P\Delta V$$

$$\Delta H = q$$

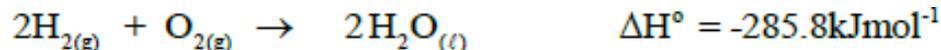
Since the pressure is constant, therefore,

$$\Delta H = q_p$$

This shows that change in enthalpy is equal to heat of reaction at constant pressure.

**18. Define enthalpy of a reaction. Give an example.**

**Ans:** The standard enthalpy of a reaction  $\Delta H^\circ$  is the enthalpy change which occurs when the certain number of moles of reactants as indicated by the balanced chemical equation, react together completely to give the products under standard conditions, i.e. 25 °C (298K) and one atmosphere pressure. All the reactants and products must be in their standard physical states. Its units are  $\text{kJmol}^{-1}$ .



**19. Define enthalpy of formation. Give an example.**

**Ans:** The standard enthalpy of formation of a compound is the amount of heat absorbed or evolved when one mole of the compound is formed from its elements. It is denoted by  $\Delta H^\circ_f$ . All the substances involved are in their standard physical states and the reaction is carried out under standard conditions i.e. at 25 °C (298 K) and one atm. pressure. Its units are  $\text{kJ mol}^{-1}$ .



**20. Define enthalpy of atomization. Give an example.**

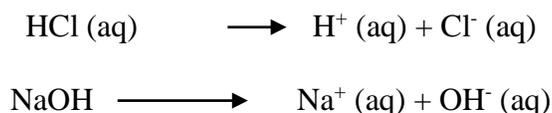
**Ans:** The standard enthalpy of atomization of an element is defined as the amount of heat absorbed when one mole of gaseous atoms is formed from the element under standard

conditions. It is denoted by  $\Delta H^\circ_{\text{at}}$ . For example, the standard enthalpy of atomization of hydrogen is given below:



**21. Define enthalpy of neutralization. Which instrument is used to measure it?**

**Ans:** The standard enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions  $[\text{H}^+]$  from an acid react with one mole of hydroxide ions  $[\text{OH}^-]$  from a base to form one mole of water. For example, the enthalpy of neutralization of sodium hydroxide by hydrochloric acid is  $-57.4 \text{ kJ mol}^{-1}$ . A strong acid HCl and a strong base, NaOH, ionize completely in dilute solutions as follows:

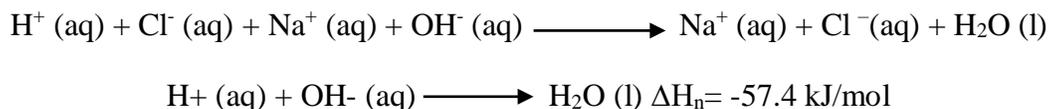


It is measured by using glass calorimeter. Formula applied is

$$q = m \times s \times \Delta T$$

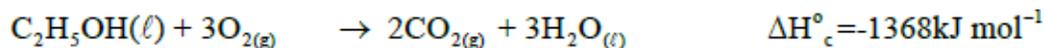
**22. The enthalpy of neutralization is merely the heat of formation of one mole of liquid water from its ionic components. Justify.**

**Ans:** When strong acid strong base solutions are mixed together during the process of neutralization, the only change that actually occurs is the formation of water molecules leaving the sodium ions and the chloride ions as free ions in solution. Thus, the enthalpy of neutralization is merely the heat of formation of one mole of liquid water from its ionic components.



**23. Define enthalpy of combustion. Give an example.**

**Ans:** The standard enthalpy of combustion of the substance is the amount of heat evolved when one mole of the substance is completely burnt in excess of oxygen under standard conditions. It is denoted by  $\Delta H^{\circ}_c$ .

**24. Define enthalpy of solution. Give an example.**

**Ans:** The standard enthalpy of a solution is the amount of heat absorbed or evolved when one mole of a substance is dissolved in so much solvent that further dilution results in no detectable heat change. For example, enthalpy of solution ( $\Delta H^{\circ}_{\text{sol}}$ ) of ammonium chloride is  $+16.2 \text{kJmol}^{-1}$  and that of sodium carbonate is  $-25.0 \text{kJmol}^{-1}$ .

**25. What is the purpose of bomb calorimeter?**

**Ans:** A bomb calorimeter is usually used for the accurate determination of the enthalpy of combustion for food, fuel and other compounds.

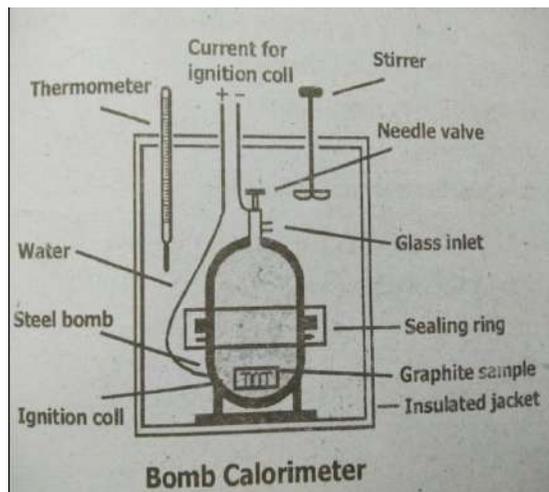
**26. Define heat capacity.**

**Ans:** The heat capacity 'c' of a body or a system is defined as the quantity of heat required to change its temperature by 1 kelvin.

$$q = c \times \Delta T$$

**27. Draw a labeled diagram of Bomb Calorimeter.**

**Ans:**



**28. State Hess's law. Give an example.**

**Ans:** If a chemical change takes place by several different routes, the overall energy change is the same, regardless of the route by which the chemical change occurs, provided the initial and final conditions are the same.

**Example:** Formation of CO<sub>2</sub> is explained in direct and indirect ways below,

**Direct route:**



**Indirect route:**



According to Hess's law

$$\Delta H = \Delta H_1 + \Delta H_2$$

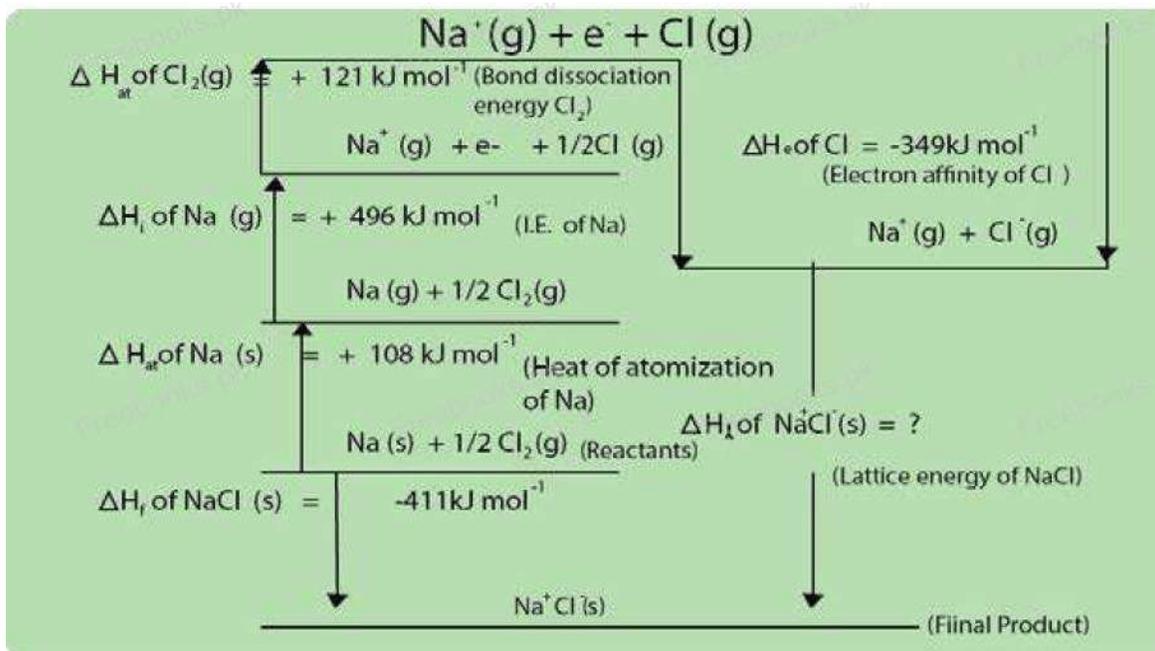
$$-393.7 = -283.7 + (-110)$$

$$-393.7 = -393.7$$

**29. State Born Haber cycle.**

**Ans:** It states that energy change in a cyclic process is always zero. It enables us to calculate the lattice energies of binary ionic compounds such as  $M^+X^-$ .

**30. Draw Born Haber cycle for sodium chloride.**



**Ans:**

**Lattice energy of NaCl= -787 kJ/mol**

**31. Define lattice energy. Give an example.**

**Ans:** The lattice energy of an ionic crystal is the enthalpy of formation of one mole of the ionic compound from gaseous ions under standard conditions.



**32. Why it is necessary to mention the physical states of reactants and products in the thermochemical equation?**

**Ans:** Thermochemical equation tells about heat change and heat change is always accompanied

by a physical change. For example,



The two values of formation of water are due to different physical states of water. Therefore, it is necessary to mention physical states of reactants and products in a thermochemical equation.

**33. *Burning of a candle is a spontaneous process. Justify it.***

**Ans:** There are two types of spontaneous processes:

- Those which start and proceed on their own.
- Those which are once started (by giving some energy) then proceed on their own.

Burning of candle is the example of second type of reactions which are once started with some amount of energy but afterwards proceed and complete on their own

**34. Differentiate between Law of conservation of energy and Hess's Law.**

**Ans:** The law of conservation of energy is a physical law that states energy cannot be created or destroyed but can be changed from one form to another.

Hess's law is defined as if a chemical change takes place by several different routes the overall energy change is the same regardless of the route by which the chemical change occurs, provided the initial and final conditions are the same.

**35. Is it true that  $\Delta H$  and  $\Delta E$  have the same values for the reaction taking place in solution state?**

**Ans:** In case of solids and liquids the change in volume is negligibly small. Therefore, in equation

$\Delta H = \Delta E + P\Delta V$  the factor becomes  $\Delta V = 0$  and we are left with

$$\Delta E \approx \Delta H$$

Hence, it is true that  $\Delta H$  and  $\Delta E$  have the same values for the reactions taking place in solution state.

**36. Mention the various units of energy.**

**Ans:** The various units of energy are Joule, KiloJoule and Calorie.

**37. Differentiate between Internal energy and enthalpy.**

**Ans:**

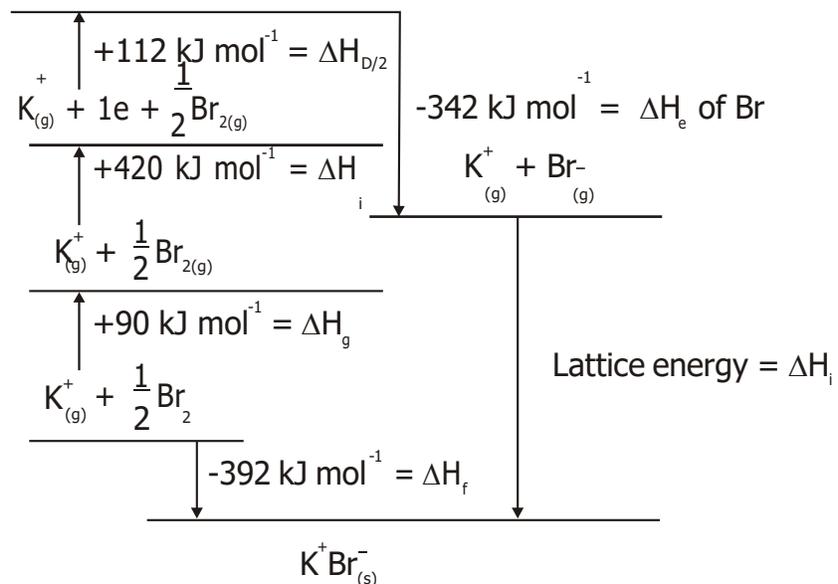
<b>Internal Energy(E)</b>	<b>Enthalpy(H)</b>
The total energy of a system including all forms of kinetic and potential energies is known as Internal Energy	The sum of internal energy change and work done is known as Enthalpy or Heat content.  <b>OR</b> The heat absorbed by system at constant pressure is known as Enthalpy.
Mathematically it is  $E = \text{Total kinetic energy} + \text{total potential energy}$	Mathematically it is  $H = E + PV$

**38. Is it true that a non-spontaneous process never happens in the universe? Explain it.**

**Ans:** No, it is not true that non-spontaneous process never happens in the universe. A non-spontaneous process occurs but it needs energy to take place. For example, water cannot flow from the ground to the top of the hill on its own. However, pumping of water uphill can be made possible with the help of devices like motor pumps, which consume electrical energy.

**39. Draw a complete fully labeled Born-Haber Cycle for the formation of Potassium bromide.**

**Ans:**



Requirement:	$\Delta H_l = ?$
$\text{K}^+_{(g)} + \text{Br}^-_{(g)} \rightarrow \text{KBr}_{(g)}$	Lattice energy of KBr

### Calculation and Result:

According to Born-Haber cycle,

$$\Delta H_s + \Delta H_i + \Delta H_s + \Delta H_e + \Delta H_l = \Delta H_f$$

Putting the values

$$+90 + 420 + 112 - 342 + \Delta H_l = -392$$

$$\Delta H_l = -392 - 90 - 420 - 112 + 392$$

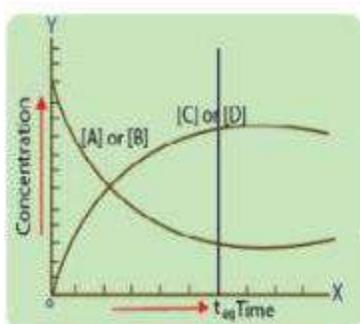
$$= -1014 + 392$$

$$\Delta H_l = -672 \text{ KJ mol}^{-1}$$

**CHEMISTRY (XI)**  
**Chapter 8**  
**Chemical Equilibrium**  
**Short Questions**

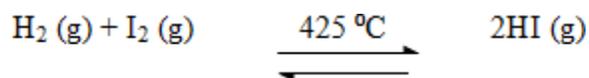
**1. What is meant by the stage of chemical equilibrium?**

**Ans:** In reversible chemical reactions two opposing reactions occur. A stage reaches for the reaction when the rates of two opposing reactions are equal. This stage is called stage of chemical equilibrium.



**2. Define reversible reaction. Give an example.**

**Ans:** A reversible reaction is a reaction in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously. For example,



**3. Define irreversible reaction. Give an example.**

**Ans:** An irreversible reaction is a reaction that proceeds in one direction only. The products do not react together to form the reactants. For example,

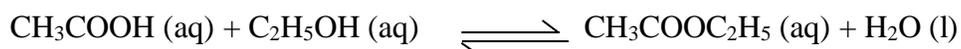


**4. Justify that chemical equilibrium is dynamic in nature.**

**Ans:** In reversible chemical reactions the molecules of reactants collide and convert into products. At the same time the molecules of the products are converting into reactants. When two opposing forces maintain the equal rates then equilibrium is there and that is dynamic equilibrium in nature.

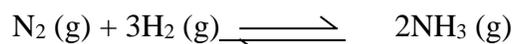
**5. Why the equilibrium constant value has its units for some of the reversible reactions but has no units for some other reactions?**

**Ans:** If the number of moles of reactants and products are equal in a reversible balanced equation then the units are cancelled, and the value of  $K_c$  has no units.



$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} \frac{[\text{moles dm}^{-3}][\text{moles dm}^{-3}]}{[\text{moles dm}^{-3}][\text{moles dm}^{-3}]} \text{ no units}$$

If the number of moles of reactants and products are unequal, then  $K_c$  has net units.



$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \frac{[\text{moles dm}^{-3}]^2}{[\text{moles dm}^{-3}][\text{moles dm}^{-3}]^3} \text{ moles}^{-2}\text{dm}^{-6}$$

**6. When four types of chemical equilibrium constants for a reaction become equal?**

**How  $K_p$  and  $K_c$  are related?**

**Ans:** The expressions of equilibrium constants depend upon the concentration units used. Mostly the concentrations are expressed in  $\text{mole dm}^{-3}$ . Let us consider the following reversible reaction.



$$K_c = \frac{C^c D^d}{A^a B^b}$$

The square brackets represent the concentration of species in moles  $\text{dm}^{-3}$ . Anyhow, the capital C is also used for molar concentrations.

If the reactants A, B, and the products C, D of the reaction under consideration are ideal gases, then molar concentration of each gas is proportional to its partial pressure. When the concentrations are expressed in terms of partial pressures, the expression of  $K_p$  is

$$K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

Here  $P_A$ ,  $P_B$ ,  $P_C$  and  $P_D$  are partial pressures of A, B, C, D respectively at equilibrium position. As long as the number of moles of products and reactants, which are in the gaseous state, are equal, the values of  $K_c$  and  $K_p$  remain the same. Otherwise, the following relationship between  $K_p$  and  $K_c$  can be derived by using Dalton's law of partial pressures.

$$K_p = K_c (RT)^{\Delta n}$$

### 7. How equilibrium constant helps to determine direction of reaction?

$$K_c = \frac{[\text{Products}]}{[\text{Reactants}]}$$

**Ans:**

The direction of a chemical reaction at any particular time can be predicted by means of [products] / [reactants] ratio, calculated before the reaction attains equilibrium. The value of [product] / [reactants] ratio leads to one of the following three possibilities.

- (a) The ratio is less than  $K_c$ . This implies that more of the product is required to attain the equilibrium, therefore, the reaction will proceed in the forward direction.
- (b) The ratio is greater than  $K_c$ . It means that the reverse reaction will occur to attain the equilibrium.
- (c) When the ratio is equal to  $K_c$ , then the reaction is at equilibrium.

***8. How the value of  $K_c$  helps to determine extent of reaction?***

- Ans:*** (a) If the equilibrium constant is very large, this indicates that the reaction is almost complete.
- (b) If the value of  $K_c$  is small, it reflects that the reaction does not proceed appreciably in the forward direction.
- (c) If the value of  $K_c$  is very small, this shows a very little forward reaction.

***9. State Le-Chatelier's principle.***

***Ans:*** This principle states that if a stress is applied to a system at equilibrium, the system acts in such a way so as to nullify, as far as possible, the effect of that stress.

***10. What happens to the direction of a reversible reaction when the ratio of the concentrations is less than actual  $K_c$ ?***

***Ans:*** When the ratio of the concentrations for a reversible reaction is less than  $K_c$  then it means that the reaction is not at equilibrium stage. It must go to the forward direction to attain the actual value of  $K_c$ .

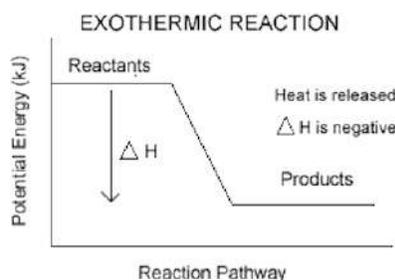
**11. The change of volume disturbs the equilibrium position for some of the gaseous phase reactions but not the equilibrium constant. Why?**

**Ans:** Those gaseous phase reversible reactions which happen with changing number of moles are affected by the change of volume at equilibrium stage. Their equilibrium position is disturbed but equilibrium constant is not changed.



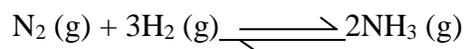
**12. Why are the exothermic reactions favored to the forward direction by cooling, but reverse is true for endothermic reactions?**

**Ans:** The amount of heat which is being evolved by the exothermic reaction is taken up by that body which has a cooling effect. So, the reactions move to that direction where there is less energy. The endothermic reactions require extra energy to take place. If the system is cooled it will go to that direction where there is a less energy and that is the backward direction of reaction



**13. The change of temperature disturbs the equilibrium position and the equilibrium constant of reaction. Justify.**

**Ans:** All the reversible reactions are disturbed by changing their equilibrium position and equilibrium constant by disturbing the temperature. Actually, change of temperature changes the energy contents of reactants and products.

**14. What will be effect of change in pressure on  $NH_3$  and  $SO_3$  synthesis?****Ans:**

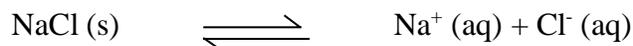
This is a gaseous reaction having less number of moles of products. So this reaction happens with the decrease of volume. The increase of pressure will shift the equilibrium position of reaction to the forward direction and greater amount of ammonia will be produced. Equilibrium constant does not change. In  $2SO_2 + O_2 \rightleftharpoons 2SO_3$  same principle is applicable.

**15. How does a catalyst affect a reversible reaction?**

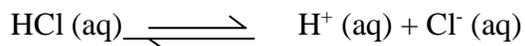
**Ans:** A catalyst increases the rate of reaction by lowering the activation energy and giving new path to the reaction. The value of equilibrium constant as well as equilibrium position remains the same. It just increases the rate of both forward and reverse reactions so that the equilibrium is reached earlier.

**16. How NaCl can be purified by common ion effect?**

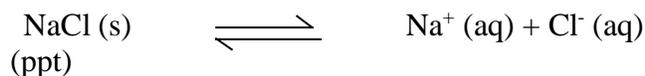
**Ans:** The impure sample of NaCl is dissolved in water to prepare the saturated solution.



If HCl gas is passed in saturated solution of NaCl the  $Cl^-$  ions are generated in excess in the solution

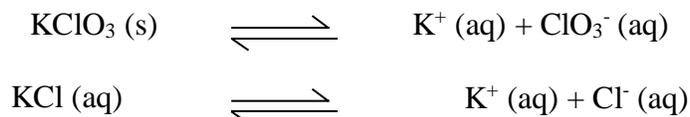


Due to excess of  $Cl^-$  the ionization of NaCl is suppressed and NaCl settles down in the form of precipitate



**17. Define common ion effect. Give an example.**

**Ans:** The suppression of ionization of a weak electrolyte by adding a common ion from outside is called common ion effect. For example the solubility of a less soluble salt  $\text{KClO}_3$  in water is suppressed by the addition of a more soluble salt  $\text{KCl}$  by common ion effect.  $\text{K}^+$  is a common ion. The ionization of  $\text{KClO}_3$  is suppressed and it settles down as precipitate.

**18. What is the effect of common ion on solubility?**

**Ans:** The presence of a common ion decreases the solubility of a slightly soluble ionic compound. In order to explain it, consider a saturated solution of  $\text{PbSO}_4$ , which is a sparingly soluble ionic salt.

$$K_c = \frac{[\text{Pb}^{2+}][\text{SO}_4^{2-}]}{[\text{PbSO}_4]}$$

**19. Why the ionic product of water increases with the increase of temperature?**

**Ans:** The value of  $K_w$  is  $0.11 \times 10^{-14}$  at  $0^\circ \text{C}$  and is  $10^{-14}$  at  $25^\circ \text{C}$ . The value increases approximately 10 times when the temperature changes from  $0^\circ \text{C}$  to  $25^\circ \text{C}$ . The reason is that the increase of temperature increases the kinetic energy of the water and possibility of bond breakage increases.

**20. What is percentage ionization of acid?**

**Ans:** We can calculate the percentage ionization of weak acid and the formula is as follows:

$$\% \text{ionization} = \frac{\text{Amount of acid ionized}}{\text{Amount of acid initially available}} \times 100$$

**21. Define pH and pOH. Give mathematical expression.**

**Ans:** pH is negative log of hydrogen ion concentration.

$$\text{pH} = -\log [\text{H}^+]$$

pOH is negative log of hydroxide ion concentration

$$\text{pOH} = -\log [\text{OH}^-]$$

**22. Prove that  $\text{pK}_w = 14$  at  $25^\circ \text{C}$**

**Ans:**

$$\text{pH} = -\log [\text{H}^+]$$

and

$$\text{pOH} = -\log [\text{OH}^-]$$

For neutral water,

$$\text{pH} = -\log 10^{-7} = 7$$

$$\text{pOH} = -\log 10^{-7} = 7$$

when

pH = 7, → solution is neutral

pH < 7, → solution is acidic

pH > 7, → solution is basic

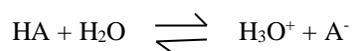
If we take the negative log of  $K_w$ , then it is called  $\text{pK}_w$ .

$$\begin{aligned} \text{pK}_w &= -\log K_w \\ &= -\log 10^{-14} \end{aligned}$$

$$\text{pK}_w = 14 \log 10$$

Since ( $\log 10 = 1$ )

$$\text{pK}_w = 14 \times 1 = 14 \text{ (at } 25^\circ \text{C)}$$

**23. Calculate ionization constant of acid.****Ans:**

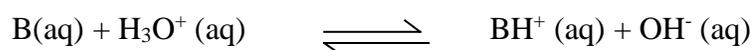
$$K_c = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]}$$

$$K_c[\text{H}_2\text{O}] = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Let  $K_c[\text{H}_2\text{O}] = K_a$

$K_a$  is another constant

Hence  $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$

**24. Calculate ionization constant of base.****Ans:** Let the base is represented by B. Then

$$K_c = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}][\text{H}_2\text{O}]}$$

Since, the concentration of  $\text{H}_2\text{O}$  constant, being in large excess

So,  $K_c[\text{H}_2\text{O}] = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$

Put  $K_c[\text{H}_2\text{O}] = K_b$

Hence  $K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$

**25. Define  $pK_b$  and  $pK_a$ .**

**Ans:** Negative log of  $K_a$  is called  $pK_a$  and negative log of  $K_b$  is called  $pK_b$

$$pK_a = -\log K_a$$

$$pK_b = -\log K_b$$

**26. Define buffers.**

**Ans:** Those solutions, which resist the change in their pH when a small amount of an acid or a base is added to them, are called buffer solutions. They have a specific constant value of pH and their pH values do not change on dilution and on keeping for a long time. Buffer solutions are mostly prepared by mixing two substances.

**27. What are the types of buffers?**

**Ans:** There are two types of buffers:

- (i) By mixing a weak acid and a salt of it with a strong base. Such solutions give acidic buffers with pH less than 7. Mixture of acetic acid and sodium acetate is one of the best examples of such a buffer.
- (ii) By mixing a weak base and a salt of it with a strong acid. Such solutions will give basic buffers with pH more than 7. Mixture of  $NH_4OH$  and  $NH_4Cl$  is one of the best examples of such a basic buffer.

**28. Why do we need buffer solutions? OR What are the applications of buffers?**

**Ans:** It is a common experience that the pH of the human blood is maintained at pH 7.35. If it goes to 7.00 or 8.00, a person may die. Sometimes one wants to study a reaction under conditions that would suffer any associated change in the pH of the reaction mixture. So, by suitable choice of the solutes, a chemist can ensure that a solution will not experience more than

a very small change in pH, even if a small amount of a strong acid or a strong base is added. Buffers are important in many areas of chemistry and allied sciences like molecular biology, microbiology, cell biology, soil sciences, nutrition and the clinical analysis.

**29. Define buffer capacity.**

**Ans:** The buffer capacity of a solution is the capability of a buffer to resist the change of pH.

**30. What is Henderson equation?**

**Ans:** For acidic buffer Henderson equation is

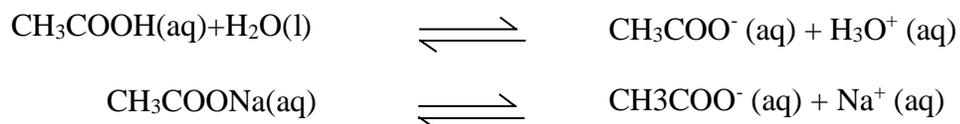
$$\text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

For basic buffer Henderson equation is

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

**31. How acidic buffer works?**

**Ans:** Let us take the example of an acidic buffer consisting of  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COONa}$ .  $\text{CH}_3\text{COOH}$ , being a weak electrolyte undergoes very little dissociation. When  $\text{CH}_3\text{COONa}$  is added to  $\text{CH}_3\text{COOH}$  solution, then the dissociation of  $\text{CH}_3\text{COOH}$  is suppressed, due to common ion effect of  $\text{CH}_3\text{COO}^-$



If one goes on adding  $\text{CH}_3\text{COONa}$  in  $\text{CH}_3\text{COOH}$  solution, then the added concentrations of  $\text{CH}_3\text{COO}^-$  decrease the dissociation of  $\text{CH}_3\text{COOH}$  and the pH of solution increases. Greater the concentration of acetic acid as compared to  $\text{CH}_3\text{COONa}$ , lesser is the pH of solution.

When an acid or  $\text{H}_3\text{O}^+$  ions are added to this buffer, they will react with  $\text{CH}_3\text{COO}^-$  to give back acetic acid and hence the pH of the solution will almost remain unchanged. The reason is that  $\text{CH}_3\text{COOH}$  being a weak acid will prefer to remain undissociated.

**32. How a basic buffer works?**

**Ans:** The buffer solution consisting of  $\text{NH}_4\text{Cl}$  and  $\text{NH}_4\text{OH}$ , can resist the change of pH and pOH, when acid or base is added from outside. When a base or  $\text{OH}^-$  ions are added in it, they will react with  $\text{H}_3\text{O}^+$  to give back  $\text{H}_2\text{O}$  and the pH of the solution again will remain almost unchanged.

**33. The solubility of glucose in water is increased by increasing the temperature.**

**Explain.**

**Ans:** The solubility of glucose in water involves an endothermic process. The solution has temperature lower than original temperature of solvent. Therefore, according to Le-Chatelier's principle, an increase in temperature will increase the solubility of glucose in solution.

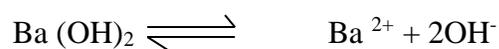
**34. Calculate the pH of  $10^{-4} \text{ mol dm}^{-3}$  solution of HCl.**

**Ans:** Mathematically

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log [10^{-4}] \\ &= -(-4)\log [10] && \text{as } \log 10 = 1 \\ &= 4 \end{aligned}$$

**35. Calculate the pH of  $10^{-4} \text{ mol dm}^{-3}$  solution of  $\text{Ba}(\text{OH})_2$ .**

**Ans:**



$$[\text{OH}^-] = 2 \times 10^{-4} \text{ mol.dm}^{-3}$$

$$\begin{aligned} \text{pOH} &= -\log 2 \times 10^{-4} \\ &= 3.69 \end{aligned}$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ &= 14 - 3.69 \\ &= 10.31 \end{aligned}$$

**36. Define solubility product. Give an example.**

**Ans:** The solubility product is the product of the concentrations of ions raised to an exponent equal to the co-efficient of the balanced equation.



$$\text{PbCl}_2 \quad K_{\text{sp}} = [\text{Pb}^{2+} (\text{aq})][\text{Cl}^- (\text{aq})]^2$$

**37. Mention the applications of solubility product.**

**Ans:** Following are the applications of solubility product:

- i. Determination of solubility from  $K_{\text{sp}}$
- ii. Determination of  $K_{\text{sp}}$  from solubility

**38. How  $K_{\text{sp}}$  is determined from solubility?**

**Ans:** From the solubility of the compounds, we can calculate  $K_{\text{sp}}$  of the salt. The solubility for most of the compounds are given in terms of the grams of the solute per 100 g of water. Since the quantity of solute is very very small, so 100 g of water solution is 100 mL of solution. The reason is that the density of water is very close to unity. Hence, we get the concentration in moles  $\text{dm}^{-3}$ . The number of moles of solute  $\text{dm}^{-3}$  of the solution is calculated by dividing the

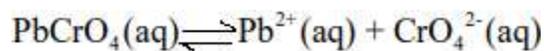
mass of solute by its molar mass. Then by using the balanced equation, we find the molarity of each ion and then find  $K_{sp}$ .

**39. How solubility is determined from  $K_{sp}$ ?**

**Ans:** For the determination of solubility from  $K_{sp}$  we need the formula of the compound and  $K_{sp}$  value. Then the unknown molar solubility  $S$  is calculated and the concentration of the ions are determined.

**40. What is the effect of common ion on solubility?**

**Ans:** The presence of a common ion decreases the solubility of a slightly soluble ionic compound. In order to explain it, consider a saturated solution of  $PbCrO_4$ , which is a sparingly soluble ionic salt.



Now add  $Na_2CrO_4$  which is a soluble salt.  $CrO_4^{2-}$  is the common ion. It combines with  $Pb^{2+}$  to form more insoluble  $PbCrO_4$ . So, equilibrium is shifted to the left to keep  $K_{sp}$  constant.

**41. Write two applications of equilibrium constant?**

**Ans:** Equilibrium constant of reversible reaction is a very informative parameter. It can be used to determine:

- i. Direction of reaction before the reversible reaction attains equilibrium.
- ii. Extent of reaction in forward and reverse side.

**42. Derive solubility product expression for  $Ag_2CrO_4$ ,  $PbCl_2$ ,  $AgCl$ ?**

**Ans:** The solubility product is the product of the concentrations of ions raised to exponent equal to the co-efficient of the balanced equation.

$$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$$

$$K_{sp} = [Pb^{2+}][Cl^-]^2$$

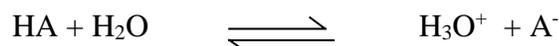
$$K_{sp} = [Ag^+][Cl^-]$$

**43. How change in volume disturbs the equilibrium position for some of the gas phase reactions but not the equilibrium constant?**

**Ans:** The change in volume disturbs those reactions in which number of moles of reactants and products are different. According to Le-Chatelier's principle, if volume of equilibrium system is decreased at equilibrium position the reaction will move in the direction of decreased number of moles and vice versa. New equilibrium position will be established but the value of  $K_c$  ultimately remains constant because it is only permanently affected by temperature change only.

**44. Define Lowry-Bronsted concept of acids and bases.**

**Ans:** According to this concept, acids are those species which donate the proton or have a tendency to donate and bases are those species which accept the proton or have a tendency to accept the proton.

**45. Prove by equations that what happens when  $Na_2CrO_4$  is added to saturated solution of  $PbCrO_4$ .**

**Ans:**





The presence of a common ion decreases the solubility of a slightly soluble ionic compound.  $\text{CrO}_4^{2-}$  is a common ion, it combines with  $\text{Pb}^{2+}$  to form more insoluble  $\text{PbCrO}_4$ . So equilibrium is shifted to the left to keep  $K_{\text{sp}}$  constant.

**46. Why solid ice at 0°C can be melted by applying pressure without supply of heat from outside.**

**Ans:** When pressure is applied to the broken pieces of ice at 0°C then according to Le-Chatelier's principle the ice moves to that direction where its volume should decrease i.e., towards liquid water. Actually, ice occupies 9% more volume than liquid water.

**47. Why do the rates of forward reactions slow down when a reversible reaction approaches the equilibrium stage?**

**Ans:** Since rate of forward reaction is directly proportional to molar concentration of reactants, with the passage of time the concentration of reactants decreases and rate of forward reaction decreases also.

**48. Reversible reaction attains the position of equilibrium which is dynamic in nature and not static. Explain it.**

**Ans:** At equilibrium state the reaction is not stopped. Only the rate of forward reaction becomes equal to the rate of reverse reaction. Since reaction is in progress in both the directions, therefore, equilibrium is dynamic in nature not static one.

**49. Mention the properties of chemical equilibrium.**

**Ans:** The properties of chemical equilibrium are:

- i. All reactions cease at equilibrium so that the system becomes stationary.
- ii. The forward and reverse reactions are taking place simultaneously at exactly the same rate.

**50. What is the justification for the increase of ionic product with temperature?**

**Ans:** The value of ionic product of water ( $K_w$ ) increases almost 75 times when temperature is increased from 0 °C to 100 °C. The increase in  $K_w$  is not regular. This occurs because ionization of water is an endothermic process and with the increase in temperature bonds become weaker and ionization increases.

**51. The change of volume disturbs the equilibrium position for some of the gaseous phase reactions but not the equilibrium constant.**

**Ans:** The change in volume disturbs those reactions in which number of moles of reactants and products are different. According to Le-Chatelier's principle if volume of equilibrium mixture is decreased at equilibrium point then the reaction will move in the direction of decreased number of moles and vice versa. New equilibrium position will be established but the value of  $K_c$  ultimately remains constant because it is temporarily affected by volume change.

**52. When a graph is plotted between time on x-axis and the concentrations of reactants and products on y-axis for a reversible reaction, the curves become parallel to time axis at a certain stage.**

**(a) At what stage the curves become parallel?**

**(b) Before the curves become parallel, the steepness of curves falls? Give reasons.**

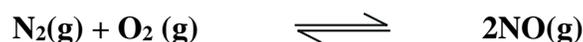
*(c) The rate of decreases of concentrations of any of the reactants and rate of increase of concentrations of any of the products may or may not be equal, for various types of reactions, before the equilibrium time. Explain it.*

**Ans:**

(a) The curves become parallel at the stage of chemical equilibrium.

(b) The steepness of curves falls before the curves become parallel because rate of reaction decreases at that stage.

(c) The rate of decrease of concentrations of any of the reactants and rate of increase of concentrations of any of the products may or may not be equal, for various types of reactions, before the equilibrium time. This is true for some reactions, when concentration of products is different from those of reactants. For example,



**53. Decide the comparative magnitudes of  $K_c$  and  $K_p$ , for the following reversible reactions.**

**(i) Ammonia synthesis**

**(ii) Dissociation of  $\text{PCl}_5$**

**Ans: (i) Ammonia synthesis:**



$$\text{Since } K_p = K_c (\text{RT})^{\Delta n}$$

$$\Delta n = \text{number of moles of products} - \text{no of moles of reactants}$$

$$\Delta n = 2 - 4 = -2$$

As value of  $\Delta n$  is negative so  $K_p$  will be lesser than  $K_c$  for this reaction.

(ii) **Dissociation of  $\text{PCl}_5$**



Since  $K_p = K_c (RT)^{\Delta n}$

$\Delta n$  = number of moles of products – no of moles of reactants

In this equation

$$\Delta n = +1$$

As value of  $\Delta n$  is positive so  $K_p$  will be greater than  $K_c$  for this reaction.

**CHEMISTRY (XI)**  
**Chapter 9**  
**Solutions**  
**Short Questions**

**1. Define phase. Give an example.**

**Ans:** Every sample of matter with uniform properties and a fixed composition is called a phase. For example, water at room temperature and normal pressure exists as a single liquid phase, that is, all the properties of water are uniform throughout this liquid phase.

**2. Define solution. Give an example.**

**Ans:** A solution, on average, is a homogeneous mixture of two or more kinds of different molecular or ionic substances. For example, sugar added in water makes a solution.

**3. Define solvent and solute.**

**Ans:** The substance which is present in large quantity is called a solvent and the other component in small quantity is called a solute. For example, in the mixture of sugar and water sugar is solute and water is solvent.

**4. Define concentration of solution.**

**Ans:** For a given solution, the amount of solute dissolved in a unit volume of solution (or a unit amount of solvent) is termed as the concentration of the solution.

**5. Define dilute solution and concentrated solution.**

**Ans:** Solutions containing relatively lower concentrations of solute are called dilute solutions, whereas, those containing relatively higher concentrations of solutes are called concentrated solutions.

**6. Mention the concentration units of solution.**

**Ans:** Following are the concentration units of solution:

- a. Percentage weight/weight
- b. Percentage weight/volume
- c. Percentage volume/weight
- d. Percentage volume/volume
- e. Molarity (M)
- f. Molality (m)
- g. Mole Fraction (x)
- h. Parts per million (ppm)

**7. Mention the units showing percentage composition.**

**Ans:** Following are the units showing percentage composition:

- a. Percentage weight/weight
- b. Percentage weight/volume
- c. Percentage volume/weight
- d. Percentage volume/volume

**8. Mention the various types of solution (Any type can come for short question)**

**Ans:**

State of Solute	State of Solvent	Example
Gas	Gas	Air
Gas	Liquid	O <sub>2</sub> in water, CO <sub>2</sub> in water.
Gas	Solid	H <sub>2</sub> adsorbed by palladium
Liquid	Gas	Mist, fog, clouds, liquid air pollutants.
Liquid	Liquid	Alcohol in water, milk, benzene in toluene.
Liquid	Solid	Mercury in silver, butter, cheese.
Solid	Liquid	Sugar in water, jellies, paints.
Solid	Gas	Dust particles in smoke.
Solid	Solid	Metal alloys pearls, opals, carbon in iron (steel).

(Use the table to write in your own words)

**9. What is percentage weight/weight? Give an example.**

**Ans:** It is the weight of a solute dissolved per 100 parts by weight of solution. 5% w/w sugar solution will contain 5 g of sugar dissolved in 100 g of solution in water. This solution contains 95 g of water.

$$\% \text{ by weight} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

**10. What is percentage weight/volume? Give an example.**

**Ans:** It is the weight of a solute dissolved per 100 parts by volume of solution. 10 g of glucose dissolved in 100 cm<sup>3</sup> of solution is 10% w/v solution of glucose. The quantity of the solvent is not exactly known. In such solutions, the total volume of the solution is under consideration.

**11. What is percentage volume/weight? Give an example.**

**Ans:** It is the number of cm<sup>3</sup> of a solute dissolved per 100 g of the solution. If we dissolve 10 cm<sup>3</sup> of alcohol in water and the total weight of the solution is 100 g then it is 10% v/w solution of alcohol in water. In such type of solutions, we don't know the total volume of the solution.

**12. What is percentage volume/volume? Give an example.**

**Ans:** It is the volume of a solute dissolved per 100 cm<sup>3</sup> of the solution. This unit of concentration is best applicable to the solutions of liquids in liquids. A 12 % alcohol beverage is 12 cm<sup>3</sup> of alcohol per 100 cm<sup>3</sup> of solution. In such solutions, the total volume of the solution may not be necessarily equal to the sum of volumes of solute and the solvent.

**13. Define molarity. Give an example.**

**Ans:** Molarity is the number of moles of solute dissolved per dm<sup>3</sup> of the solution. To prepare one molar solution of glucose in water, we take 180 g of glucose and add sufficient water to make the total volume 1 dm<sup>3</sup> (litre) in a measuring flask.

Mathematical relation

$$\text{Molarity(M)} = \frac{\text{Number of moles of solute}}{\text{Volume of solution (dm}^3\text{)}}$$

**14. Define molality. Give an example.**

**Ans:** Molality is the number of moles of solute in 1000 g (1 kg) of the solvent. In order to prepare molal solutions, we don't have to take any flask. 180 g of glucose when dissolved in 1000 g of water gives one molal solution of glucose.

$$\text{Molality(m)} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$$

**15. Define mole fraction (x). Give an example.**

**Ans:** The mole fraction of any component in a mixture is the ratio of the number of moles of it to the total number of moles of all the components present.

$$X_A = \frac{n_A}{n_A + n_B + n_C}$$

$$X_B = \frac{n_B}{n_A + n_B + n_C}$$

$$X_C = \frac{n_C}{n_A + n_B + n_C}$$

The sum of the mole fractions of all the components of a solution must be equal to one. In the case of mixture of gases, one can determine the mole fraction from the partial pressure data of the mixture. Hence,

$$X_A = \frac{P_A}{P_A + P_B + P_C}, X_B = \frac{P_B}{P_A + P_B + P_C}, X_C = \frac{P_C}{P_A + P_B + P_C}$$

Where  $p_a$ ,  $p_b$ ,  $p_c$  are the partial pressures of various gases in the mixture. Generally, we can say that

$$\text{Mole fraction of any gas} = \frac{\text{Partial pressure of that gas}}{\text{Total pressure of the mixture of gases}}$$

#### 16. Define parts per million (ppm).

**Ans:** It is defined as the number of parts (by weight or volume) of a solute per million parts (by weight or volume) of the solution. This unit is used for very low concentrations of solutions, e.g. to express the impurities of substances in water.

$$\text{Parts per million (ppm)} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

**17. Tell about the solutions of solids in liquids.**

**Ans:** When a solid comes in contact with a suitable liquid, it dissolves forming a solution i.e. a homogeneous mixture. This process of dissolution can be explained in terms of attraction between the particles of a solute and that of a solvent. The molecules or ions in solids are arranged in such a regular pattern that the inter-molecular or inter-ionic forces are at a maximum.

**18. Justify that the solubility principle is “like dissolves like”.**

**Ans:** The inter-ionic forces of attraction are very strong in ionic solids so equally strong polar solvents are needed to dissolve them. Such solids cannot be dissolved by moderately polar solvents e.g. acetone. A moderately polar solvent, fails to dissolve sodium chloride, which is an ionic solid. Thus the solubility principle is “like dissolves like”.

**19. What are completely miscible liquids?**

**Ans:** Liquids which mix completely with each other are called completely miscible liquids. For example, liquids like alcohol and water or alcohol and ether mix in all proportions.

**20. What are partially miscible liquids?**

**Ans:** A large number of liquids are known which dissolve into one another up to a limited extent. For example, ether  $\text{C}_2\text{H}_5\text{-O-C}_2\text{H}_5$  dissolves water to the extent of about 1.2 % and water dissolves ether up to the extent of about 6.5%.

**21. What are conjugate solutions?**

**Ans:** Each liquid layer when is a saturated solution of the other liquid then such solutions are called conjugate solutions. The mutual solubility of these conjugate solutions is affected by temperature changes. Typical examples of such systems are:

- a. Phenol-water system
- b. Triethylamine-water system
- c. Nicotine-water system

**22. What is critical solution temperature or upper consolute temperature?**

**Ans:** The temperature of 65.9 °C at which two conjugate solutions merge into one another is called critical solution temperature or upper consolute temperature.

**23. How a water-phenol system works?**

**Ans:** If equal volumes of water and phenol are mixed together they show partial miscibility. It has been observed that around room temperature phenol will dissolve in a lot of water giving us the upper layer and water will dissolve in a lot of phenol giving us the lower layer. At 25 °C the upper layer is 5% solution of phenol in water and the lower layer is 30% water in phenol. These two solutions are conjugate solutions to each other. The lower layer has a greater density due to greater percentage of phenol. Water acts as a solute in the lower layer while phenol is a solute in the upper layer. When the temperature of water-phenol system is increased, the compositions of both layers change. Water starts travelling from upper to the lower layer and phenol travels from lower to the upper layer. When the temperature of this system approaches 65.9 °C, a homogeneous mixture of two components is obtained. This homogeneous mixture contains 34% phenol and 66% water.

**24. What are immiscible liquids?**

**Ans:** Those liquids which do not dissolve into each other in any proportion are immiscible.

Examples:

(i) Water and benzene ( $\text{H}_2\text{O} + \text{C}_6\text{H}_6$ )

(ii) Water and carbon disulphide ( $\text{H}_2\text{O} + \text{CS}_2$ )

**25. What is the difference between ideal and non-ideal solutions?**

**Ans:** Following points are the differences between ideal and non-ideal solution:

- i. If the forces of interactions between the molecules of different components are same as when they were in the pure state, they are ideal solutions, otherwise non-ideal.
- ii. If the volume of solution is not equal to the sum of the individual volumes of the components, the solution is non-ideal.
- iii. Ideal solutions have zero enthalpy change as their heat of solution.
- iv. If the solutions obey Raoult's law then they are ideal. This is one of the best criterion for checking the ideality of a solution

**26. State Raoult's law.**

**Ans:** The vapour pressure of a solvent above a solution is equal to the product of the vapour pressure of pure solvent and the mole fraction of solvent in solution.

$$p = p^\circ x_1$$

**27. Define solubility. Give an example.**

**Ans:** The solubility is defined as the concentration of the solute in the solution when it is in equilibrium with the solid substance at a particular temperature. For example, solubility is expressed in terms of number of grams of solute in 1000g of solvent. At a particular temperature,

saturated solution of NaCl in water at 0°C contains 37.5g of NaCl in 100 g of water. Similarly the solubility of CuSO<sub>4</sub> in water at 0°C is 14.3g/100g, while at 100°C it is 75.4g/100g.

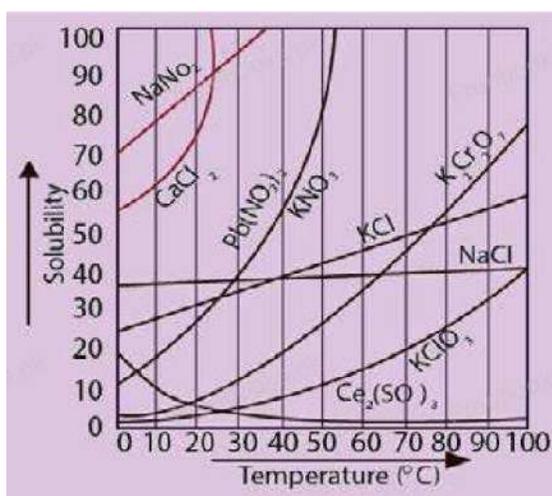
**28. Define solubility curves.**

**Ans:** A graphical representation between temperature and solubility of solution is called solubility curves. There are two types of solubility curves.

- (a) Continuous solubility curves
- (b) Discontinuous solubility curves

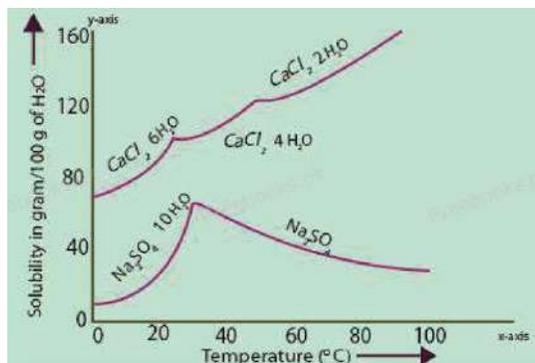
**29. What are continuous solubility curves?**

**Ans:** Continuous solubility curves don't show sharp breaks anywhere. KClO<sub>3</sub>, K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>, Pb(NO<sub>3</sub>)<sub>2</sub> and CaCl<sub>2</sub> are showing continuous solubility curves. The solubility curves of KCl, NaCl and NaNO<sub>3</sub> give the straight lines. NaCl shows a very small change of solubility from 0°C to 100°C increase of temperature. Ce<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> shows the exceptional behaviour whose solubility decreases with the increase in temperature and becomes constant from 40°C onwards. It shows continuous solubility curve.



**30. What are discontinuous solubility curves?**

**Ans:** Sometimes, the solubility curves show sudden changes of solubilities and these curves are called discontinuous solubility curves. The best examples in this reference are  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ ,  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ . Actually, these curves are combination of two or more solubility curves. At the break a new solid phase appears and another solubility curve of that new phase begins. It is the number of molecules of water crystallization which changes and hence solubility changes.



### 31. What is fractional crystallization? How is it carried out?

**Ans:** Fractional crystallization is a technique for the separation of impurities from the chemical products. By using the method, the impure solute is dissolved in a hot solvent in which the desired solute is less soluble than impurities. As the hot solution is cooled, the desired solute being comparatively less soluble, separates out first from the mixture, leaving impurities behind. In this way, pure desired product crystallizes out from the solution.

### 32. Define colligative properties.

**Ans:** The colligative properties are the properties of solution that depend on the number of solute and solvent molecules or ions. Following are colligative properties of dilute solution:

- (i) Lowering of vapour pressure
- (ii) Elevation of boiling point
- (iii) Depression of freezing point

(iv) Osmotic pressure

**33. Why molal solution is dilute as compared to molar solution?**

**Ans:** Molal solution is dilute as compared to molar solution because in molal solution 1 kg of solvent is used whereas in molar solution 1 litre of solvent is used. As 1 kg solvent is more than 1 litre solvent so molal solution is dilute as compared to molar solution.

**34. What is ebullioscopic constant?**

**Ans:** The elevation in boiling point of 1 molal solution of non-volatile non-electrolyte solute in water as solvent is  $0.52^{\circ}\text{C}$ . It is also called molal boiling point elevation constant.

$$\Delta T_b = K_b m$$

When  $m=1$

$$\Delta T_b = K_b$$

**Unit**

The unit of  $K_b$  is  $^{\circ}\text{C Kg mol}^{-1}$

**Example**

Dissolve 6g of urea in 500g of  $\text{H}_2\text{O}$  or 18g of glucose in 500g of  $\text{H}_2\text{O}$  both give 0.2 molal solution and both have same elevation of boiling points i.e.  $0.1^{\circ}\text{C}$ .

**35. What is cryoscopic constant?**

**Ans:** The depression in freezing point of 1 molal solution of non-volatile non-electrolyte solute in water as solvent is  $1.86^{\circ}\text{C}$ . It is called molal freezing point depression constant or cryoscopic constant.

**36. What are the conditions to observe colligative properties?**

**Ans:** To observe the colligative properties following condition should be fulfilled by the solutions:

- (i) Solution should be dilute
- (ii) Solute should be non-volatile
- (iii) Solute should be non-electrolyte

**37. What are the parts in assembly of Landsberger's apparatus?**

**Ans:** The apparatus consists of four major parts.

- (a) An inner tube with a hole in its side. This tube is graduated.
- (b) A boiling flask which sends the solvent vapours into the graduated tube through a rosehead.
- (c) An outer tube, which receives hot solvent vapours coming from the side hole of the inner tube.
- (d) A thermometer which can read up to 0.01K.

**38. Why ethylene glycol is used as an anti-freeze in car radiators?**

**Ans:** Ethylene glycol is used as an antifreeze in the radiator of an automobile. It is not only completely miscible with water but has a very low vapour pressure and non-volatile in character. When mixed with water, it lowers the freezing point as well as raises the boiling point. During winter it protects a car by preventing the liquid in the radiator from freezing. In hot summer, the antifreeze solution also protects the radiator from boiling over.

**39. Why the relative lowering of vapour pressure is independent of temperature?**

**Ans:** The expression of relative lowering of vapour pressure is as follows:

$$\Delta p/p^{\circ} = x_2$$

In this expression temperature is not involved so it is independent of temperature.

**40. *Non-ideal solutions do not obey the Raoult's law. Why?***

**Ans:** The molecules of components in non-ideal solutions have forces of attractions for each other. The values of vapour pressures of individual components are not proportional to their mole fractions as in Raoult's law.

**41. *Boiling points of solvents increase due to the presence of solutes. Why?***

**Ans:** Solute particles hinder the escape of solvent particles so more energy is required by the solvent particles to leave the surface of solvent. That is why in the presence of solute boiling point increases.

**42. *Why Beckmann thermometer is used to note the depression of freezing point?***

**Ans:** Beckmann thermometer can read up to 0.01 K. The elevation of boiling point and depression in freezing point are very small quantities so this small temperature change can be read by Beckmann thermometer.

**43. *Why the freezing points are depressed due to presence of solutes?***

**Ans:** The lowering of vapour pressure compels the solutions to freeze at those temperatures which are below the freezing point of pure solvent. The reason is that the vapour pressure temperature curve meets the solid phase of pure solvent at lower temperature than the pure solvent.

**44. *Calculate the molarity of a solution containing 9g of glucose in 250 cm<sup>3</sup> of solution.***

**Ans:**

18 g of glucose in 1000 mL of solution is 0.1 molar.

**Calculation**

1M solution of glucose contains 180 g of glucose in 1 L water

250 cm<sup>3</sup> of solution =  $250/1000 = 0.25 \text{ dm}^3$

$$\text{Molarity (M)} = \frac{\text{Mass}}{\text{Molar Mass}} \times \frac{1}{\text{volume of soln. in dm}^3}$$

$$= \frac{9}{180} \times \frac{1}{0.25}$$

$$= 0.2 \text{ M}$$

$$= 2 \times 10^{-1} \text{ M}$$

**45. How molality is independent of temperature but molarity depends on temperature?**

**Ans:** In molal solutions the mass of the solvent and that of the solute are also fixed. The masses of the substances are not temperature dependent. In molar solutions we have the volumes of solutions. Volume of a liquid is temperature dependent. So, the molality is not influenced by temperature but molarity does change.

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent in Kg}}$$

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in dm}^3}$$

**46. One molal solution of glucose is dilute as compared to one molar solution of glucose. Justify it.**

**Ans:** In one molal solution of glucose, 180 g of glucose is dissolved in 1000 g of water. In one molar solution of glucose, 180 g of glucose is added in water to make total volume of solution as

1000 cm<sup>3</sup>. As 1000g (1 kg) of solvent is greater than 1000 mL of solvent so molar solution is concentrated and molal solution is dilute.

**47. One molal solution of urea is dilute as compared to one molar solution of urea. Justify it.**

**Ans:** In one molal solution of urea, 60 g of urea is dissolved in 1000 g of water. In one molar solution of urea, 60 g of urea is added in water to make total volume of solution as 1000 cm<sup>3</sup>. So the volume of water in molar solution is less than that in molal solution. Hence, molar solution is concentrated and molal solution is dilute.

**48. Relative lowering of vapor pressure is independent of temperature. Justify it.**

**Ans:** The relative lowering of vapour pressure and mole fraction of solute are related as:

$$\frac{\Delta P}{p_0} = x_2$$

Vapour pressure and lowering of vapour pressure depends upon temperature. So, when the temperature of a solution is increased both the factors  $\Delta P$  and  $P^0$  increase in such a way that the ratio remains the same.

**49. Justify that boiling point of solvents increase due to presence of non-volatile solutes.**

**Ans:** The surface of the solution has molecules of solute as well. They do not allow the solvent to leave the surface as rapidly as in pure solvent. To boil the solutions, we have to increase the temperature of solutions in comparison to pure solvents. So the boiling points of solutions are higher than pure solvents.

**50. Depression of freezing point is a colligative property. Justify it.**

**Ans:** Depression of freezing point is a colligative property as colligative properties depend upon number of particles. Suppose there is 6g of urea, 18g of glucose and 34.2g of sucrose and they are dissolved separately in 1kg of water. This will produce 0.1 molal solution of each substance. In these three solutions, the vapour pressure of water will be lowered. The reason is that the molecules of a solute present upon the surface of a solution decrease the evaporating capability. The lowering of vapour pressure in all these solutions will be same at a given temperature as the number of particles of the solute in the solutions is equal. We have added  $1/10^{\text{th}}$  of Avogadro's number of particles. The lowering of vapour pressure depends upon the number of particles and not upon their molar mass and structures. The freezing points will be depressed for these solutions and the value of depression in these three cases is  $0.186^{\circ}\text{C}$ .

**51. Why NaCl and KNO<sub>3</sub> are used to lower the melting points of ice?**

**Ans:** NaCl and KNO<sub>3</sub> are electrolytes and are sufficiently soluble in water. They double the number of particles after dissociation in water. In this way they can manage to decrease the freezing point of water to a greater extent as compared to a non-electrolyte.

**52. Why is the vapor pressure of a solution lesser than vapor pressure of pure solvent?**

**Ans:** The particles of the solute are distributed throughout the bulk of the solution and some of the particles of the solute are also present on the surface of the solution. The number of

molecules of the solvent per unit area on the surface of the solution becomes less. Hence, the evaporating tendency of the solvent decreases and vapour pressure of solution becomes less.

**53. Calculate the percentage by weight of NaCl if 2g of it is dissolved in 20g of water.**

**Ans:**

$$\% \text{ by weight} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100$$

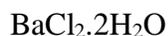
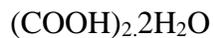
$$\% \text{ by weight} = \frac{2}{20} \times 100$$

$$= 10\%$$

**54. What is meant by water of crystallization? Give an example.**

**Ans:** The number of water molecules which combine with compounds as they are crystallized from aqueous solutions are called water of crystallization or water of hydration.

**Examples**



**55. What are zeotropic and azeotropic mixtures?**

**Ans: Zeotropic mixtures**

Such liquid mixtures which distil with a change in composition are called zeotropic mixtures.

**Example**

methyl alcohol-water solution.

**Azeotropic mixtures**

Azeotropic mixtures are those which boil at constant temperature and distil over without change in composition at any temperature like a pure chemical compound.

***Example***

ethanol-water mixture.

***56. Differentiate between hydration and hydrolysis.***

***Ans:***

***Hydration***

Hydration is the process in which water molecules surround and interact with solute ions or molecules.

***Example***

When a salt is added to water it is dissolved because of hydration process.

***Hydrolysis***

When a salt is dissolved in water, it dissociates into cations and anions. These ions may react with water and the resulting solution may be acidic, basic or neutral, depending upon the extent of hydrolysis.

***Example***

When NaCl is dissolved in water, the resulting solution is neutral because in solution the concentration of  $H^+$  and  $OH^-$  ions is equal to  $10^{-7}$  M.

***57. What are the names of four major parts of apparatus used in Landsberger's method for elevation of boiling point?***

***Ans:*** The four major parts of Landsberger's apparatus are:

1. An inner tube with a hole in its side. This tube is graduated.

2. A boiling flask which sends the solvent vapours into the graduated tube through a rosehead.
3. An outer tube, which receives hot solvent vapours coming from the side hole of the inner tube.
4. A thermometer which can read up to 0.01K.

**58. Why the solubility of glucose into water increases by increasing temperature?**

**Ans:** The solubility of glucose into water increases by increasing temperature because the space between the water molecules increases due to increase in kinetic energy and more glucose molecules are accommodated in between the spaces.

**59. Give two statements of Raoult's law.**

**Ans:**

**Statement 1**

The vapour pressure of a solvent above a solution is equal to the product of vapour pressure of pure solvent and the mole fraction of the solvent in solution.

**Relation**

$$p \propto x_1$$

$$p = p^{\circ} x_1 \text{ ----- (1)}$$

Where  $P^{\circ}$  is vapour pressure of pure solvent.  $P$  is the vapour pressure of solvent in solution or of pure solution and  $x_1$  is the mole fraction of solvent.

**Statement 2**

The lowering of vapour pressure of a solvent is directly proportional to the mole fraction of solute.

**Relation**

$$\Delta p = p^{\circ} x_2$$

$$\Delta p/p^{\circ} = x_2$$

$\Delta p$

$p^{\circ}$  is called relative lowering of vapour pressure and it is more important than lowering of vapour pressure.

**60. Aqueous solution of  $\text{CuSO}_4$  is acidic in nature. Justify it.**

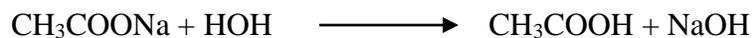
**Ans:**



As sulphuric acid is a strong acid and copper hydroxide is a weak base so the solution is overall acidic in nature.

**61. Aqueous solution of  $\text{CH}_3\text{COONa}$  is basic in nature.**

**Ans:**



As acetic acid is a weak acid and sodium hydroxide is a strong base so the solution is overall basic in nature.

**62. Define and explain the following with one example in each case.**

**i. A homogeneous phase:**

**Definition:**

Every sample of matter with uniform properties and fixed composition is called a phase.

**Example:**

Water at room temperature and normal pressure exists as a single liquid phase. Its all properties are uniform throughout this liquid phase.

**ii. A concentrated solution:**

**Definition:**

A solution which has relatively larger concentration of solute is called concentrated solution.

**Example:**

10 g  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  in 5 mL water

**iii. A solution of solid in a solid:**

**Definition:**

When solid solute is dissolved in solid solvent such solution is said to be solution of solid in solid.

**Example:**

Metal alloys like Brass, Bronze etc.

**iv. Upper consolute temperature:**

**Definition:**

The temperature at which two conjugate solutions merge into one another is called critical solution temperature or upper consolute temperature.

**Example:**

Upper consolute temperature of phenol water system is  $65.9^\circ\text{C}$ .

**v. A non-ideal solution:**

***Definition:***

Solutions which do not follow Raoult's law are called non-ideal solutions. They show deviations from Raoult's Law due to difference in their molecular structures i.e. size, shape and intermolecular forces.

***Example:***

Solution of 95.5 % Ethanol and 4.5 % water (B.P.=78.1°C)

**vi. *Zeotropic solutions:***

***Definition:***

The solutions which follow Raoult's law are called ideal solutions or Zeotropic mixtures

***Example:***

The solution of 50% ethanol and water

**vii. *Heat of hydration***

***Definition:***

The heat energy released or absorbed when one mole ions are dissolved into water is called heat of hydration.

***Example:***

Heat of hydration of  $H^+$  ions is  $-1075 \text{ kJ mol}^{-1}$

**viii. *Water of crystallization***

***Definition:***

Those water molecules which combine with substances as they are crystallized from aqueous solutions are called water molecules of crystallization or water of hydration.

Some familiar examples are as follows:

***Example:***

$(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$  (Oxalic acid)

**ix. *Azeotropic solution***

**Definition:**

Solution which does not follow Raoult's law is called non-ideal solution or azeotropic solution. It shows deviations from Raoult's law.

**Example:**

Solution of 95.5 % Ethanol and 4.5 % water (B.P.=78.1°C)

**x. Conjugate solution:****Definition:**

The solution of two partially miscible liquids is called conjugate solution.

**Example:**

Solution of phenol and water

**63. The sum of mole fractions of all the components is always equal to unity for any solution. Justify.**

**Ans:** Consider a solution which consists of two components A and B. Their mole fraction  $x_A$  and  $x_B$  are given as:

$$x_A = \frac{n_A}{n_A + n_B} \qquad x_B = \frac{n_B}{n_A + n_B}$$

$$x_A + x_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B}$$

$$x_A + x_B = \frac{n_A + n_B}{n_A + n_B}$$

$$x_A + x_B = 1$$

Hence it is proved that sum of mole fractions for any solution is always unity.

**64. 100g of 98%  $H_2SO_4$  has a volume of  $54.34 \text{ cm}^3$  of  $H_2SO_4$ . (Density =  $1.84 \text{ g cm}^{-3}$ )?**

**Ans:** Mass of 98%  $H_2SO_4$  = 100g

Density of 98%  $\text{H}_2\text{SO}_4 = 1.84 \text{ g cm}^{-3}$

Volume of 98%  $\text{H}_2\text{SO}_4 = ?$

Now using the formula:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Or Volume = mass / density

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

$$= \frac{100}{1.84}$$

$$= 54.34 \text{ cm}^3$$

Hence, 100 g of 98 %  $\text{H}_2\text{SO}_4$  has a volume of  $54.34 \text{ cm}^3$  of  $\text{H}_2\text{SO}_4$  because its density is  $1.84 \text{ gcm}^{-3}$ .

**65. Colligative properties are obeyed when the solute is non-electrolyte, and also when the solutions are dilute. Justify.**

**Ans:** Yes, colligative properties are only obeyed by the solution when the solute is non-electrolyte. If it would be electrolyte then it will ionize in the solution and number of solute particles will be changed means increased and solution will not remain dilute.

When amount of solute is very high in solution the force of attraction between solute and solvent particles becomes very strong creating non-ideal solution so it should be dilute.

**66. The solutions showing positive and negative deviations cannot be fractionally distilled at their specific compositions. Explain it.**

**Ans:** Many solutions do not behave ideally. Such solutions are also called Azeotropic mixtures. They show positive or negative deviations from Raoult's Law due to differences in their

molecular structures i.e. size, shape and intermolecular forces. Such solutions boil at constant temperature and distil over without change in composition at any temperature like a pure chemical compound. It means solute cannot be separated from solvent as both do not maintain their original boiling points in the solution. Solute and solvent start boiling at the same temperature.

***Examples:***

- i. Solution of 95.5% ethanol and 4.5 % water
- ii. Solution of 20.24% HCl and 79.76% water

***67. Explain the effect of temperature on phenol-water system.(You may consult how a phenol-water system works for answer)***

***Ans:*** At 25°C the upper layer is 5% solution of phenol in water and the lower layer is 30% water in phenol. Water acts as a solute in the lower layer while phenol is a solute in the upper layer. When the temperature of water-phenol system is increased, the compositions of both layers change. Water starts travelling from upper solution to lower solution and phenol from lower to upper layer. When temperature of this system approaches 65.9°C, a single homogeneous mixture (solution) will be obtained. This homogeneous mixture contains 34% (25.75 ml) phenol and 66% (50ml) water. The temperature of 65.9°C at which two conjugate solutions merge into one another is called critical solution temperature or upper consolute temperature.

***68. What is the physical significance of  $K_b$  and  $K_f$  of solvents? OR Define ebullioscopic constant and cryoscopic constant with an example.***

***Ans: Physical significance of  $K_b$***

The elevation in boiling point of one molal solution is called molal boiling point constant or ebullioscopic constant ( $K_b$ ).

The unit of  $K_b$  is °C Kg.mol<sup>-1</sup>

Value of  $K_b$  indicates strength of attractive forces which solvent particles can develop with solute particles.

***Physical significance of  $K_f$***

The depression in freezing point of one molal solution is called molal freezing point constant or cryoscopic constant ( $K_f$ ). Its significance and units are also the same as that of  $K_b$ .

***69. Boiling points of the solvents increase due to the presence of solutes.***

***How?***

***Ans:*** A substance boils at that temperature when its vapour pressure equalizes external pressure. The vapour pressure of solvent decreases due to the addition of non-volatile solute. Therefore, we have to supply more (high) temperature in order to equalize the vapour pressure of solvent to external pressure in order to boil it. In this way boiling point of solution will be greater than boiling point of solvent.

***70. In summer the antifreeze solutions protect the liquid of the radiator from boiling over.***

***Justify.***

***Ans:*** By adding some antifreeze into water for example ethylene glycol, the boiling point of water shows elevation. This mixture is used in the radiators of automobiles. In hot summer, antifreeze protects the water of radiator from boiling at 100 °C. In winter ethylene glycol protects the freezing of the radiator water at 0 °C.

***71. Freezing points are depressed due to the presence of solutes. Justify.***

***Ans:*** Presence of non-volatile and non-electrolyte solutes decreases freezing point of solvent because forces of attraction between solvent molecules become weak and vapour pressure of solution is reduced.

***Example:***

The freezing point of pure water is  $0^{\circ}\text{C}$  and that of solution containing 1 mole solute is -  
 $1.86^{\circ}\text{C}$ .

**72. The boiling point of one molal urea solution is  $100.52^{\circ}\text{C}$  but the boiling point of two molal urea solution is less than  $101.04^{\circ}\text{C}$ .**

**Ans:** Two molal urea solution contains twice the number of solute particle ( $2 \times 6.02 \times 10^{23}$ ) as compared to one molal urea solution. If a solution has large number of solute particles then some of its particles will not be associated with solvent molecules and hence will not contribute in decrease of vapour pressure of solvent. Therefore, the lowering of vapour pressure will not be in accordance to Raoult's law. So the boiling point of two molal urea solution is lesser than  $101.14^{\circ}\text{C}$  which is expected from two molal solution.

**73. The total volume of the solution by mixing  $100\text{ cm}^3$  of water with  $100\text{ cm}^3$  of alcohol may not be equal to  $200\text{ cm}^3$ . Justify it.**

**Ans:** If the solution formed by mixing  $100\text{ cm}^3$  of water with  $100\text{ cm}^3$  of alcohol is ideal the total volume will be  $200\text{ cm}^3$  otherwise not. As alcohol makes hydrogen bonding with water it is likely to be less than  $200\text{ cm}^3$ .

**74.  $100\text{ g}$  of  $98\%$   $\text{H}_2\text{SO}_4$  has a volume of  $54.34\text{ cm}^3$  of  $\text{H}_2\text{SO}_4$ .**

**Ans:**

Mass of  $98\%$   $\text{H}_2\text{SO}_4 = 100\text{g}$

Density of  $98\%$   $\text{H}_2\text{SO}_4 = 1.84\text{ g cm}^{-3}$

Volume of  $98\%$   $\text{H}_2\text{SO}_4 = ?$

Now using the formula:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Or Volume = mass / density

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

$$= \frac{100}{1.84}$$

$$= 54.34 \text{ cm}^3$$

Hence, 100 g of 98 % H<sub>2</sub>SO<sub>4</sub> has a volume of 54.34 cm<sup>3</sup> of H<sub>2</sub>SO<sub>4</sub> because its density is 1.84 gcm<sup>-3</sup>.

**75. Non-ideal solutions do not obey the Raoult's law. Justify.**

**Ans:** Non-ideal solutions have strong attractive forces among the particles means solute and solvent molecules. As a result the vapor pressures of both are changed from the expected values. Such solutions do not obey the Raoult's law.

**76. The concentration in terms of molality is independent of temperature but molarity depends upon temperature.**

**Ans:** In case of molality the concentration of solvent is expressed in terms of mass. The mass of a substance is not affected by the change in temperature. Hence, molality is independent of temperature.

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent in Kg}}$$

In case of molarity the concentration of solution is expressed in terms of volume which changes with the change of temperature. Hence, molarity is affected by temperature.

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in dm}^3}$$

Moreover, molal solution is dilute as compared to molar so change in temperature does not affect it.

## ***CHEMISTRY (XI)***

### ***Chapter 10***

#### ***Electrochemistry***

#### ***Short Questions***

***1. What is Electrochemistry?***

**Ans:** Electrochemistry is concerned with the conversion of electrical energy into chemical energy in electrolytic cells as well as the conversion of chemical energy into electrical energy as in galvanic or voltaic cells.

***2. Define oxidation number or state.***

**Ans:** It is the apparent charge on an atom of an element in a molecule or an ion. It may be positive or negative or zero. For example,  $\text{Na}^{+1}$ ,  $\text{S}^{-2}$ ,  $\text{K}^{+1}$

***3. State rules for assigning oxidation number (Mention any four as an answer to short question).***

**Ans:** (i) The oxidation number of all elements in the free state is zero. This is often shown as a zero written on the symbol. For example,  $\text{H}_2^0$ ,  $\text{Na}^0$ ,  $\text{Mg}^0$

(ii) The oxidation number of an ion, consisting of a single element, is the same as the charge on the ion. For example, the oxidation number of  $\text{K}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Al}^{3+}$ ,  $\text{Br}^-$ ,  $\text{S}^{2-}$  are + 1, +2, +3, -1, -2, respectively.

(iii) The oxidation number of hydrogen in all its compounds except metal hydrides is +1. In metal hydrides it is -1. ( $\text{Na}^+\text{H}^-$ ,  $\text{Mg}^{2+} \text{H}_2^{(-1)2}$ )

(iv) The oxidation number of oxygen in all its compounds except in peroxides,  $\text{OF}_2$  and in super oxides is -2. It is -1 in peroxides +2 in  $\text{OF}_2$  and -1/2 in super oxides.

(v) In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.

(vi) In ions, the algebraic sum of oxidation number equals the charge on the ion.

(vii) In any substance the more electronegative atom has the negative oxidation number.

**4. Calculate the oxidation no. of Mn in  $KMnO_4$ .**

**Ans:**  $KMnO_4$

$$+1 + Mn + (-2)4 = 0$$

$$Mn + 1 - 8 = 0$$

$$Mn = +7$$

**5. Calculate the oxidation no. of Mn in  $Na_2MnO_4$ .**

**Ans:**  $Na_2MnO_4$

$$(+1)2 + Mn + (-2)4 = 0$$

$$+2 + Mn - 8 = 0$$

$$Mn - 6 = 0$$

$$Mn = +6$$

**6. Calculate the oxidation no. of S in  $Cr_2(SO_4)_3$  and  $SO_4^{2-}$ .**

**Ans:**  $Cr_2(SO_4)_3$

$$(+3)2 + (S)3 + (-2)12 = 0$$

$$+6 + (S)3 - 24 = 0$$

$$(S)3 - 18 = 0$$

$$S = +6$$

**$SO_4^{2-}$**

$$S + (-2)4 = -2$$

$$S - 8 = -2$$

$$S = -2 + 8$$

$$S = +6$$

**7. Calculate the oxidation no. of Cr in  $CrCl_3$ .**

**Ans:**  $CrCl_3$

$$Cr + (-1)3 = 0$$

$$Cr - 3 = 0$$

$$Cr = +3$$

**8. Calculate the oxidation no. of P in  $HPO_3$ .**

***Ans:***  $\text{HPO}_3$

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

$$+1 + P - 6 = 0$$

$$P - 5 = 0$$

$$P = +5$$

9. Calculate the oxidation no. of the elements underlined in the following compounds.

Ans:



Oxidation number of O = -2

Oxidation number of Ca = +2

Oxidation number of Cl = x

$$\begin{aligned} +2 + 2x + 6(-2) &= 0 \\ 2x &= +12 - 2 = +10 \\ x &= +5 \end{aligned}$$



Oxidation number of O = -2

Oxidation number of Na = +1

Oxidation number of C = x

$$\begin{aligned} 2(+1) + (x) + 3(-2) &= 0 \\ 2 + x - 6 &= 0 \\ x &= +6 - 2 = +4 \end{aligned}$$



Oxidation number of O = -2

Oxidation number of Na = +1

Oxidation number of P = x = ?

$$\begin{aligned} 3(+1) + x + 4(-2) &= 0 \\ +3 + x - 8 &= 0 \\ x &= +8 - 3 = +5 \end{aligned}$$



Oxidation number of O = -2

Oxidation number of H = +1

Oxidation number of N = x = ?

$$\begin{aligned} (+1) + x + 3(-2) &= 0 \\ +1 + x - 6 &= 0 \\ x &= +6 - 1 = +5 \end{aligned}$$



Oxidation number of O = -2

Oxidation number of Cr = +3

Oxidation number of S = x = ?

$$\begin{aligned} 2(+3) + 3x + 12(-2) &= 0 \\ +6 + 3x - 24 &= 0 \\ 3x &= +2 - 6 = +18 \\ x &= +6 \end{aligned}$$

(vi) **HPO<sub>3</sub>**

Oxidation number of O = -2  
Oxidation number of H = +1  
Oxidation number of P = x = ?

$$\begin{aligned} +1 + x + 3(-2) &= 0 \\ +1 + x - 6 &= 0 \\ x &= +6 - 1 = +5 \end{aligned}$$

(vii) **K<sub>2</sub>MnO<sub>4</sub>**

Oxidation number of O = -2  
Oxidation number of K = +1  
Oxidation number of Mn = x = ?

$$\begin{aligned} 2(+1) + x + 4(-2) &= 0 \\ 2 + x - 8 &= 0 \\ x &= +8 - 2 = +6 \end{aligned}$$

**10. Define metallic conduction.**

**Ans:** Most metals are conductors of electricity because of the relatively free movement of their electrons throughout the metallic lattice. This electronic conduction is simply called metallic conduction.

**11. Define ionization. Give an example.**

**Ans:** Ionization is the process in which ionic compounds when fused or dissolved in water split up into charged particles called ions.



**12. What is an electrolytic cell?**

**Ans:** The movement of ionic charges through the liquid brought by the application of electricity is called electrolytic conduction and the apparatus used is known as electrolytic cell.

**13. Differentiate between electrolytic and galvanic cell.**

**Ans:**

Electrolytic Cell	Voltaic cell
<ol style="list-style-type: none"> <li>1. The electrochemical cell in which electrical energy is converted into chemical energy is called Electrolytic cell.</li> <li>2. In this cell, non- spontaneous reaction occurs.</li> <li>3. Electric current is used to drive the non-spontaneous oxidation reduction reaction.</li> <li>4. Electrolysis takes place in this cell.</li> </ol> <p><b>Example:</b> Down's cell, Nelson's cell</p>	<ol style="list-style-type: none"> <li>1. The electrochemical cell in which chemical energy is converted into electrical energy is called Voltaic cell.</li> <li>2. In this cell, spontaneous reaction occurs.</li> <li>3. Electric current is produced due to spontaneous reaction.</li> <li>4. Electric conduction takes place in this cell.</li> </ol> <p><b>Example:</b> Daniel's cell, Fuel cells</p>

**14. Define electrolysis OR Define electrolysis. Give one example.**

**Ans:** The electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis.

OR

Electro comes from electricity and lysis means breakdown, so electrolysis is the breakdown of salts by passing electric current. Moreover, the electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis. The process is carried out in an electrolytic cell.

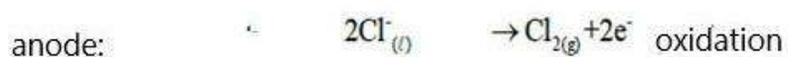
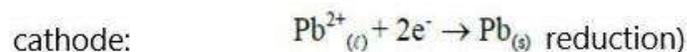
**Example:** Electrolysis of Fused salts



<p>2. In this case, conductance decreases with increase in temperature.</p> <p>3. No chemical reaction occurs during conduction.</p> <p>4. Chemical composition of metal is not changed during conduction and no new substances are produced.</p> <p><b>Example:</b> All metals are conductors. e.g. Fe, Pb etc.</p>	<p>2. In this case, conductance increase with increase in temperature</p> <p>3. Redox reactions occurs take place during conduction</p> <p>4. Since chemical reactions occur, therefore new substances are produced</p> <p><b>Example:</b> Molten salt e.g. NaCl (l) or their aqueous solutions, acids, bases etc.</p>
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**17. How electrolysis takes place in fused salts?**

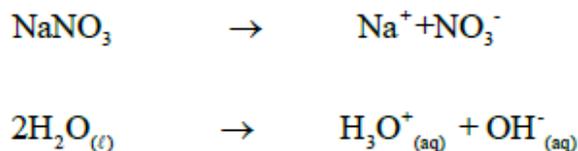
**Ans:** When a fused salt is electrolyzed, the metal ions called cations arrive at the cathode which being negatively charged supply electrons to them and thus discharge the cations. The anions move towards the anode, give up their electrons and are thus discharged. In the case of fused lead chloride, the equations for electrode processes are given as under:



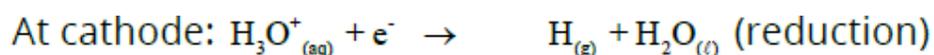
So, oxidation happens at anode and reduction at the cathode. Similarly, for fused NaCl and fused PbBr<sub>2</sub> the electrolytes are decomposed during electrolysis. Fused Pb and Na are deposited at cathode and Cl<sub>2</sub> (g) and Br<sub>2</sub> (l) at anode. Electrons flow through the external circuit from anode to cathode. The electric current is conducted through the cell by the ions and through the external circuit by the electrons.

**18. How electrolysis takes place in aqueous solution of salts? OR Explain electrolysis of aqueous solution of sodium nitrate.**

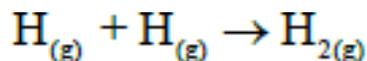
**Ans:** While electrolyzing aqueous sodium nitrate ( $\text{NaNO}_3$ ) solution, sodium ions present are not discharged at the cathode. A small concentration of hydronium and hydroxyl ions arises from the dissociation of water:



Hydronium ions accept electrons from the cathode to form hydrogen atoms:

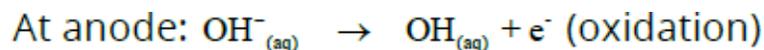


Subsequently, hydrogen atoms combine rapidly to form hydrogen molecules at the cathode.

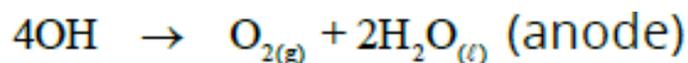


So,  $\text{H}_2$  gas evolves at the cathode.

At the anode, both nitrate and hydroxide ions are present. Hydroxide ions are easier to discharge than nitrate ions. Nitrate ions remain in solution while the electrode reaction is:



The OH groups combine to give  $\text{O}_2$  gas as follows.



So,  $\text{O}_2$  gas evolves at the anode.

**19. Give the chemistry of electrolysis of aqueous solution of sodium chloride.**

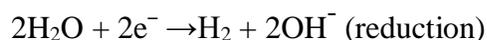
**Ans:** Caustic Soda is obtained on commercial scale by the electrolysis of concentrated aqueous solution of sodium chloride using Titanium anode and Mercury or Steel cathode. This electrolysis is carried out in Nelson's cell or Castner- Kellner cell or Hg- cell.



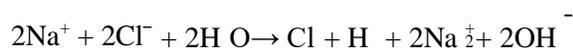
At Anode:



At Cathode:

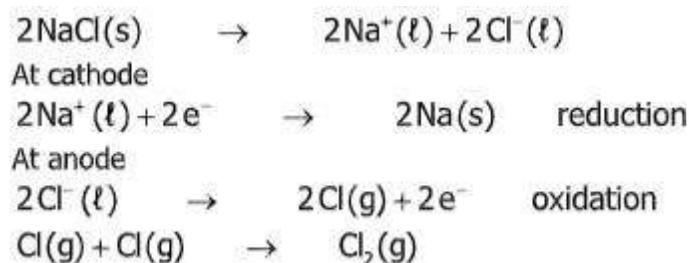


Overall reaction: (combining Na<sup>+</sup> ions)



**20. Explain extraction of sodium by Down's cell.**

**Ans:** Extraction of sodium by the electrolysis of fused sodium chloride is carried out in Down's cell. In this case, molten sodium chloride is electrolyzed between iron cathode and graphite anode. The cell is planted to get sodium metal commercially. Chlorine is obtained as a by-product.

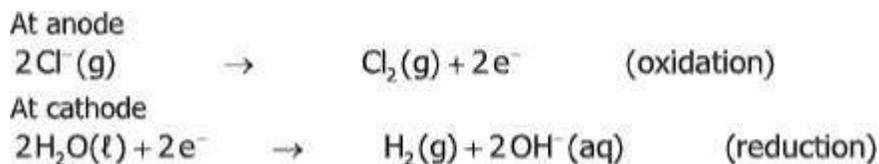


By adding the two reactions at anode and cathode, the overall reaction is

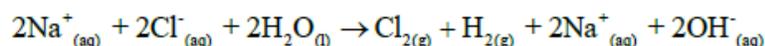


**21. How the process of electrolysis is applied in the production of caustic soda?**

**Ans:** Caustic soda is obtained on industrial scale by the electrolysis of concentrated aqueous solution of sodium chloride using titanium anode and mercury or steel cathode. This electrolysis is carried out in Nelson cell and Castner- Kellner cell or Hg- cell.



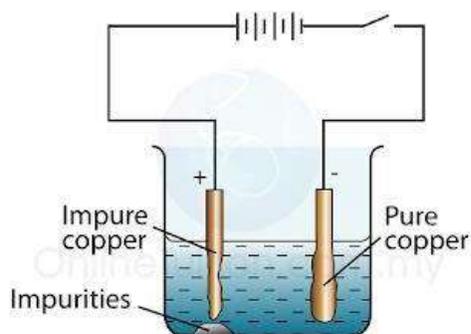
By combining, the electrode reactions and including  $\text{Na}^+$  ions, the overall reaction is:



Here, chlorine and hydrogen are obtained as by-products and  $\text{Na}^+$  is not discharged at cathode.

### 22. How impure copper is purified by the process of electrolysis?

**Ans:** Electrolytic cell can also be used for the purification of copper. Impure copper is made the anode and a thin sheet of pure copper is made the cathode. Copper sulphate solution is used as an electrolyte. The atoms of Cu from impure Cu- anode are converted to  $\text{Cu}^{2+}$  ions and migrate to cathode which is made up of pure Cu. In this way, Cu anode is purified. Impurities are left at anode.



### 23. What is anodized aluminium?

**Ans:** Anodized aluminium is prepared by making it an anode in an electrolytic cell containing sulphuric acid or chromic acid, which coats a thin layer of oxide on it. The aluminium oxide layer resists attack for corrosive agents. The freshly anodized aluminium is hydrated and can absorb dyes.

**24. What is the function of salt bridge?**

**Ans:** The purpose of the salt bridge is to prevent any net charge accumulation in either beaker by allowing negative ions to leave the right beaker, diffuse through the bridge and enter the left beaker. If this diffusional exchange of ions does not occur, the net charge accumulating in the beakers would immediately stop the flow of electrons through the external circuit and the oxidation-reduction reaction would stop.

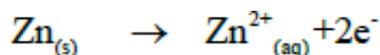
**25. A salt bridge maintains the electrical neutrality in the cell. Explain.**

**Ans:** In galvanic cell the salt bridge connects the two half cells together and maintains the electrical neutrality in both the half cells by transferring certain ions from one half cell to the other and thus preventing the accumulation of ions in either of the half cell. If salt bridge is not connected the cell will stop working very soon.

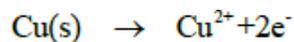
**26. What is a voltaic cell?**

**Ans:** A voltaic cell is one in which chemical energy is converted to electrical energy. It is a spontaneous cell.

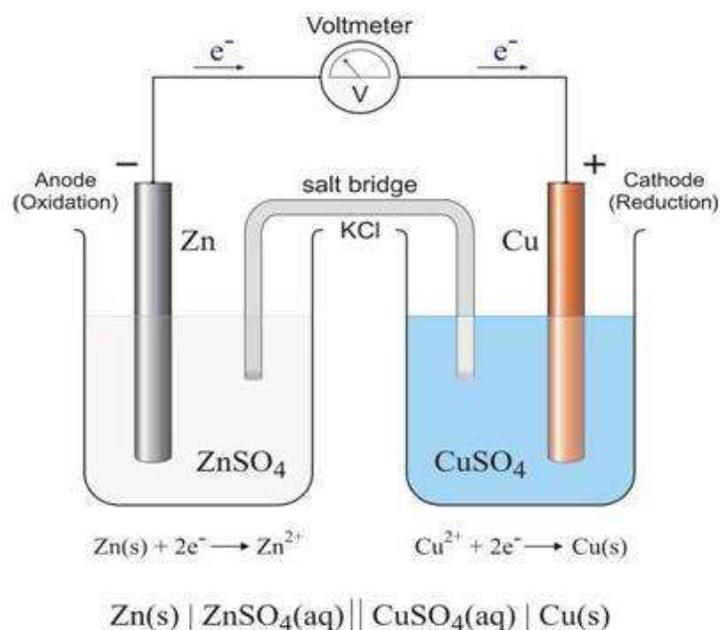
A voltaic or a galvanic cell consists of two half-cells that are electrically connected. Each half cell is a portion of the total cell in which a half reaction takes place. The left half cell consists of a strip of zinc metal dipped in 1.0 M solution of zinc sulphate giving the following equilibrium:



The right half-cell is a copper metal strip that dips into 1.0 M copper sulphate solution and the equilibrium here is represented as follows:



These half-cells are connected electrically by a salt bridge. Oxidation takes place at Zinc and



reduction takes place at Copper.

### 27. Is voltaic cell a reversible cell?

**Ans:** If the external circuit is replaced by a source of electricity that opposes the voltaic cell, the electrode reactions can be reversed. Now, the external source pushes the electrons in the opposite direction and supplies energy or work to the cell so that the reverse non-spontaneous reaction occurs. Such, a cell is called a reversible cell. For the zinc-copper cell, the half-cell reactions are reversed to give.



and the overall reaction being reversed, becomes



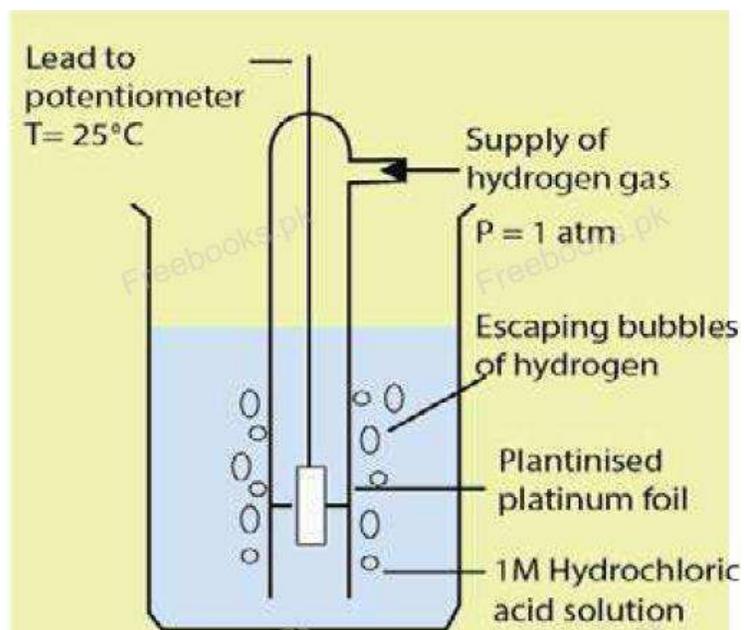
**28. Define electrode potential OR What is standard electrode potential?**

**Ans:** The potential set up when an electrode is in contact with one molar solution of its own ions at 298 K is known as standard electrode potential or standard reduction potential of the element. It is represented as  $E^{\circ}$ .

**29. What is SHE?**

**Ans:** It stands for standard hydrogen electrode which is used as a standard. It consists of a piece of platinum foil, which is coated electrolytically with finely divided platinum black, to give a large surface area and suspended in one molar solution of HCl.

Pure hydrogen gas at one atmosphere pressure is continuously bubbled into 1M HCl solution. The platinum acts as an electrical conductor and also facilitates the attainment of equilibrium between the gas and its ions in solution. The potential of this electrode is arbitrarily taken as zero.



**30. The equilibrium is set up between metal atoms of electrode and ions of metal in a cell.**

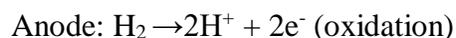
**How?**

**Ans:** The electrode potential can be explained on the basis of equilibrium which is established between atoms of metal deposited on the surface of metal through reduction process and the ions released by metal atoms into the solution due to oxidation of metal atoms. Whichever process is greater or faster determines whether the metal becomes positively or negatively charged. At equilibrium both the processes occur at the same rate and no further change occurs in the potential difference.

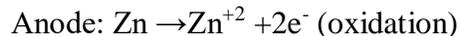
**31. SHE acts as anode when connected with Cu but acts as cathode when connected with Zn.**

**Justify your answer with equations.**

**Ans:** The reduction potential of Copper is +0.34V i.e a positive value, so it brings about reduction and act as cathode and SHE as anode.



The oxidation potential of Zn is +0.76V i.e a positive value so Zn brings about oxidation and act as anode while SHE act as cathode.



**32. A porous plate or a salt bridge is not required in Lead acid storage battery. Justify.**

**Ans:** A porous plate or a salt bridge is not required in a Lead storage battery because all the cells are dipped in the same electrolyte (**30% H<sub>2</sub>SO<sub>4</sub> solution/ d= 1.25gcm<sup>-3</sup>**). Salt bridge usually separates the two electrodes in which different electrolytes are used.

**33. Write down reactions taking place at the electrodes during discharging of Nickel Cadmium cell.**

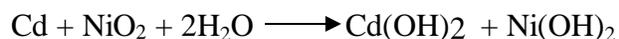
**Ans:** At Anode:



(oxidation) At Cathode:



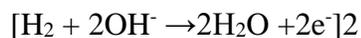
Net reaction:



**34. Give chemical reactions taking place at anode and cathode of Fuel cell.**

**Ans:**

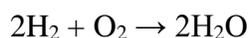
At Anode:



At Cathode:



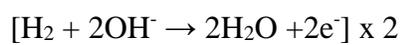
Overall reaction:



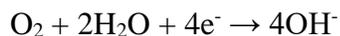
**35. How Fuel cells produce electricity?**

**Ans:** In fuel cells  $\text{H}_2$  is oxidized and  $\text{O}_2$  is reduced. The electrolyte of the cell is aqueous KOH solution. Electrodes of the cell are made of porous carbon impregnated with platinum. The following reactions take place at the respective electrodes to produce electricity.

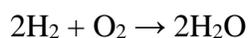
At Anode:



At Cathode:



Overall reaction:



**36. Write two advantages of Fuel cells.**

**Ans:** Advantages of fuel cells are as follows.

- Fuel cells are light, portable and produce electricity and pure water during space flights.
- These are environment friendly, efficient and convert about 75% fuel bond energy into electrical energy.

**37. Write recharging of lead accumulator battery.**

**Ans:** During recharging the lead accumulator battery is connected to an external battery through the electrodes. As a result, the electrode half reactions are reversed as follows:

At Anode:  $\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{-2}$  (reduction)

At Cathode:  $\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{-2} + 2\text{e}^-$  (oxidation)

Overall reaction:  $2\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{Pb} + \text{PbO}_2 + 4\text{H}^+ + 2\text{SO}_4^{-2}$

Both the density of the acid and voltage of the battery are restored.

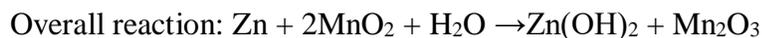
**38. Lead accumulator is a chargeable battery. Justify.**

**Ans:** Lead accumulator is a chargeable battery and is based on secondary cells. The battery is chargeable the reactions of which can be reversed by applying external electrical source. During the process of discharging, the density of  $\text{H}_2\text{SO}_4$  decreases from  $1.25 \text{ g cm}^{-3}$  to  $1.15 \text{ g cm}^{-3}$  and volts drop down from 12V. After recharging the density of sulphuric acid again increases up to  $1.25 \text{ g cm}^{-3}$ , electrodes regain original physical states and cell restarts providing E.M.F of 12V.

**39. What is alkaline battery?**

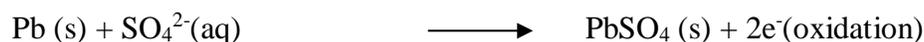
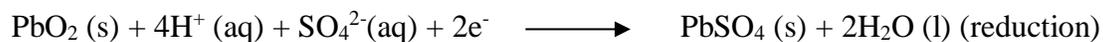
**Ans:** A dry alkaline battery is that which uses an alkali i.e KOH as an electrolyte. Zinc rod serves as anode and manganese dioxide as cathode. The battery is enclosed in a steel container. The voltage of the cell is 1.5V. The electrode reactions are as follows:

Anode:  $\text{Zn} + 2\text{OH}^- \rightarrow \text{Zn(OH)}_2 + 2\text{e}^-$



**40. Write reactions of discharging of lead accumulator battery.**

**Ans:** At the anode the lead atoms release two electrons each to be oxidized to  $\text{Pb}^{2+}$  ions, which combine with  $\text{SO}_4^{2-}$  ions present in the electrolyte and get deposited on the anode as  $\text{PbSO}_4$ .



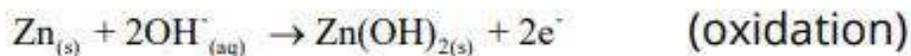
At the cathode the electrons from the anode are accepted by  $\text{PbO}_2$  and hydrogen ions from the electrolyte then undergo a redox reaction to produce lead ions and water as follows: The  $\text{Pb}^{2+}$  ions then combine with the  $\text{SO}_4^{2-}$  ions and they both deposit at the cathode as  $\text{PbSO}_4$ . When both electrodes are completely covered with  $\text{PbSO}_4$  deposits, the cell will cease to discharge any more current until it is recharged. The overall reaction is:



**41. Explain the working of silver oxide battery.**

**Ans:** The cathode is of silver oxide,  $\text{Ag}_2\text{O}$  and the anode is of zinc metal. The following reactions occur in a basic electrolyte.

At the anode



At the cathode



The overall reaction is:



The voltage of silver oxide battery is about 1.5 V

**42. Define electrochemical series.**

**Ans:** When elements are arranged in the order of their standard, electrode potentials on the hydrogen scale the resulting list is known as electrochemical series.

**43. How electrochemical series predict feasibility of a reaction?**

**Ans:** Electrochemical series help to predict whether a particular reaction will take place or not.

For example,  $\text{Cu}^{2+}$  (aq) can oxidize solid zinc but  $\text{Zn}^{2+}$  (aq) cannot oxidize solid copper. Standard

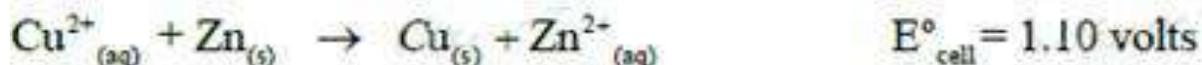


reduction potential values of copper and zinc can explain this.

Since zinc is being oxidized so the reverse reaction will be considered.



The overall reaction will thus be



The overall positive value for the reaction potential suggests that the process is energetically feasible. If the sum of  $E^\circ$  values of the two half-cell reactions is negative, then the reaction will not be feasible.

**44. The standard oxidation potential of Zinc is 0.76V and its reduction potential is -0.76.**

**Why?**

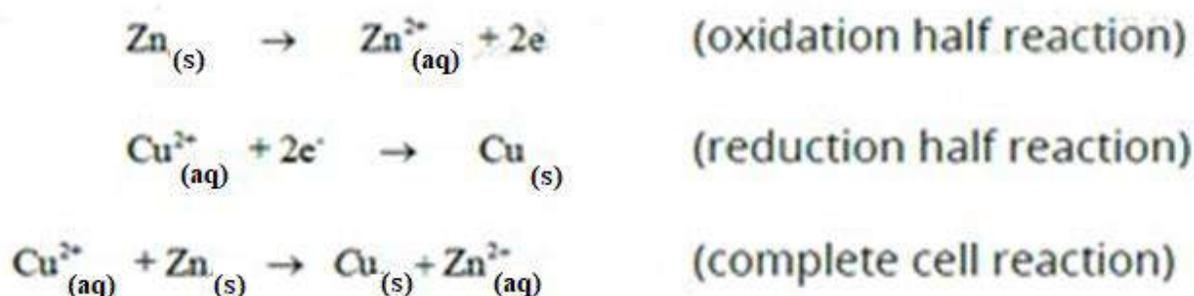
**Ans:** When Zn electrode is connected to S.H.E, Zn gives electrons to Hydrogen electrode and gets oxidized:



The standard oxidation potential of Zn is +0.76V as it is a spontaneous process. Its reduction will be non-spontaneous process with reduction potential of -0.76V.

**45. How electrochemical series helps to calculate voltage or emf?**

**Ans:** In a galvanic cell, the electrode occupying a higher position in the electrochemical series, will act as anode and oxidation takes place on it. Similarly, the electrode occupying the lower position in the series will act as a cathode and reduction will take place on it. The half-cell reactions are:



The oxidation potential of Zn is positive. The reduction potential of  $\text{Cu}^{2+}$  is also positive. The cell voltage or emf of the cell is given by

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{oxi}} + E^{\circ}_{\text{red}}$$

$$E^{\circ}_{\text{cell}} = 0.76 + 0.34 = 1.10 \text{ volts}$$

The cell voltage or emf measures the force with which electrons move in the external circuit and therefore measures the tendency of the cell reaction to take place. Galvanic cells, thus, give quantitative measure of the relative tendency of the various reactions to occur.

**46. How relative tendency of elements for redox reactions can be compared by using electrochemical series?**

**Ans:** The value of the reduction potential of a metal or a nonmetal tells us the tendency to lose electrons and act as a reducing agent. It also gives the information about the tendency of a species to gain electrons and act as an oxidizing agent. Greater the value of standard reduction potential of a given species, greater is its tendency to accept electrons to undergo reduction and hence to act as an oxidizing agent. For example, ions like  $\text{Au}^{3+}$ ,  $\text{Pt}^{2+}$ ,  $\text{Hg}^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Cu}^{2+}$  and the nonmetals elements like  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$  which lie below the SHE, have a strong tendency to gain electrons and undergo reduction. The series tell us that strong oxidizing agents like  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$ , etc. have a large positive value of standard reduction potentials, while strong reducing agents have large negative values like Li, K, Ca, Na, etc. which lie above SHE.

**47. How can relative chemical reactivity of metals be compared by using electrochemical series?**

**Ans:** Greater the value of standard reduction potential of a metal, smaller is its tendency to lose electrons to change into a positive ion and hence lower will be its reactivity. For example, metals like Li, Na, K and Rb are highly reactive. Coinage metals, Cu, Ag, and Au are the least reactive because they have positive reduction potentials. Similarly, metals like Pb, Sn, Ni, Co and Cd which

are very close to SHE react very slowly with steam to liberate hydrogen gas, while the metals like Fe, Cr, Zn, Mn, Al and Mg which have more negative reduction potentials react with steam to produce the metallic oxides and hydrogen gas.

**48. How reaction of metals with dilute acids can be compared by electrochemical series?**

**Ans:** Greater the value of standard reduction potential of a metal, lesser is its tendency to lose electrons to form metal ions and so weaker is its tendency to displace  $H^+$  ions from acids as  $H_2$  gas. For example, metals like Au, Pt, Ag and Cu which have sufficiently high positive values of reduction potentials, do not liberate hydrogen gas from acids. While, metals like Zn, Mg and Ca which are close to the top of the series and have very low reduction potentials, liberate hydrogen gas, when they react with acids.

**49. Na and K can displace hydrogen from acids but Pt, Pd and Cu cannot.**

**Ans:** The elements whose reduction potential are negative show oxidation and are considered as oxidizable elements. On the other hand, the elements with positive reduction potential exhibit reduction and fall under the category of reducible elements. Therefore, Na and K have negative reduction potential so they can react with dilute acids like HCl and  $H_2SO_4$  releasing  $H_2$  gas. Pt, Pd and Cu cannot react with dilute acids.

**50. How displacement of one metal by another can be studied by electrochemical series?**

**Ans:** One metal will displace another metal from the aqueous solution of its salt if it lies above in the electrochemical series. For example, Fe can displace Cu from  $CuSO_4$ , Zn does not displace Mg from solution of  $MgSO_4$ .

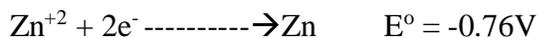
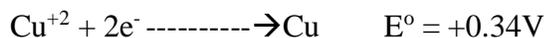
**51. Give two applications of electrochemical series.**

**Ans:**

Following are the two applications of electrochemical series:

**1. Prediction of feasibility of a chemical reaction**

When we look at the electrochemical series, it is easy to predict whether a particular reaction will take place or not. For example  $\text{Cu}^{+2}$  ion can oxidize solid Zinc but  $\text{Zn}^{+2}$  ion cannot oxidize solid copper. Standard reduction potential values of Copper and Zinc are shown below.

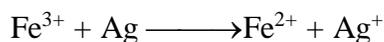


## 2 .Displacement of one metal by another from its solution

Metal will displace another metal from the aqueous solution of its salt, if it lies above the electrochemical series. For example Fe can displace Cu from  $\text{CuSO}_4$  but Zn does not displace Mg from solution of  $\text{MgSO}_4$ .

**52. Is the reaction  $\text{Fe}^{3+} + \text{Ag} \longrightarrow \text{Fe}^{2+} + \text{Ag}^{+}$  spontaneous? If not, write spontaneous reaction involving these species.**

**Ans:**



Fe is reduced and Ag is oxidized. Hence,  $\text{Fe}^{+3}$  will act as cathode while  $\text{Ag}^{\circ}$  as anode.

Emf of the cell can be calculated as:

$$E^{\circ}\text{cell} = E^{\circ}(\text{ox}) + E^{\circ}(\text{red})$$

$$E^{\circ}\text{cell} = - 0.7994 + (- 0.44)$$

$$E^{\circ}\text{cell} = - 0.7994 - 0.44$$

$$E^{\circ}\text{cell} = - 1.2394$$

emf of cell is negative so the cell-reaction is non-spontaneous.

On reversing the electrodes the cell-reaction becomes spontaneous



**53. What is the difference between single electrode potential and standard electrode potential?**

***Give its importance.***

***Ans:***

**Single Electrode potential:**

The potential set up when an electrode is in contact with one molar solution of its own ions is known as standard electrode potential or standard reduction potential of the element. It is represented as E.

**Standard electrode potential:**

The potential set up when an electrode is in contact with one molar solution of its own ions at 298 K is known as standard electrode potential or standard redox potential of the element. It is represented as  $E^{\circ}$ .

**Importance:**

1. The standard electrode potential value of an element indicates the potential generated when an element is reduced or oxidized.
2. Total Emf of the cell can also be calculated by standard electrode potential of the elements working as electrode.

## ***CHEMISTRY (XI)***

### ***Chapter 11***

#### ***Reaction Kinetics***

##### ***Short Questions***

***1. Define reaction kinetics.***

***Ans:*** The studies concerned with rates of chemical reactions and the factors that affect the rates of chemical reactions constitute the subject matter of reaction kinetics.

***2. Define rate of reaction.***

***Ans:*** The rate of a reaction is defined as the change in concentration of a reactant or a product divided by the time taken for the change.

$$\text{Rate of reaction} = \frac{\text{change in concentration of the substance}}{\text{time taken for the change}}$$

It has the units of  $\text{mol.dm}^{-3}\text{s}^{-1}$

***3. Define average rate of reaction and instantaneous rate of reaction.***

***Ans:*** The rate at any one instant during the interval is called the instantaneous rate. The rate of reaction between two specific time intervals is called the average rate of reaction.

$$\text{Rate of reaction} = \frac{dx}{dt}$$

***4. Define specific rate constant or velocity constant.***

***Ans:*** The specific rate constant of a chemical reaction is the rate of reaction when the

concentrations of the reactants are unity.

$$\text{If } [A] = 1 \text{ mol dm}^{-3} \text{ and } [B] = 1 \text{ mol dm}^{-3}$$

$$\text{Rate of reaction} = k \times 1^a \times 1^b = k$$

**5. Define order of reaction. Give an example.**

**Ans:** The order of reaction is given by the sum of all the exponents to which the concentrations in the rate equation are raised.

**OR**

The order of reaction may also be defined as the number of reacting molecules, whose concentrations alter as a result of the chemical change. For example, decomposition of nitrogen pentoxide involves the following equation.



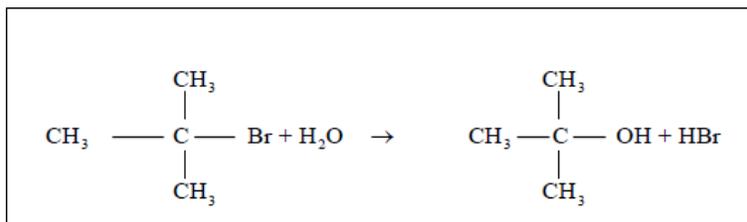
The experimentally determined rate equation for this reaction is as follows:

$$\text{Rate} = k[\text{N}_2\text{O}_5]$$

This equation suggests that the reaction is first order with respect to the concentration of  $\text{N}_2\text{O}_5$ .

**6. Hydrolysis of tertiary butyl bromide is a pseudo first order reaction. Justify.**

**Ans:** Hydrolysis of tertiary butyl bromide



The rate equation determined experimentally for this reaction is

$$\text{Rate} = k[(\text{CH}_3)_3\text{CBr}]$$

The rate of reaction remains effectively independent of the concentration of water because, being a solvent, it is present in very large excess. Such types of reactions have been named as pseudo first order reactions.

**7. Define half-life period. Give an example.**

**Ans:** Half-life period of a reaction is the time required to convert 50% of the reactants into products. For example, the half-life period for the decomposition of  $\text{N}_2\text{O}_5$  at  $45^\circ\text{C}$  is 24 minutes.

**8. How radioactive decay is a first order reaction?**

**Ans:** The disintegration of radioactive  ${}_{92}\text{U}^{235}$  has a half-life of  $7.1 \times 10^8$  or 710 million years. If one kilogram sample disintegrates, then 0.5 kg of it is converted to daughter elements in 710 million years. Out of 0.5 kg of  ${}_{92}\text{U}^{235}$ , 0.25kg disintegrates in the next 710 million years. So, the half-life period for the disintegration of a radioactive substance is independent of the amount of that substance and it is a first order reaction.

**9. Half-life period of first order reaction is independent of the initial concentration of reactant. Justify.**

**Ans:** If we decompose  $0.10 \text{ mole dm}^{-3}$  of  $\text{N}_2\text{O}_5$  at  $45^\circ\text{C}$ , then after 24 minutes  $0.05 \text{ mole dm}^{-3}$  of  $\text{N}_2\text{O}_5$  will be left behind. Similarly after 48 minutes  $0.025(25\%) \text{ mole dm}^{-3}$  of  $\text{N}_2\text{O}_5$  will remain unreacted and after 72 minutes (3 half times)  $0.0125 (12.5\%) \text{ mole dm}^{-3}$  of  $\text{N}_2\text{O}_5$ , will remain unreacted. The disintegration of radioactive  ${}_{92}\text{U}^{235}$  has a half-life of  $7.1 \times 10^8$  or 710 million years. If one kilogram sample disintegrates, then 0.5 kg of it is converted to daughter elements in 710 million years. Out of 0.5 kg of  ${}_{92}\text{U}^{235}$ , 0.25kg disintegrates in the next 710 million years. So, the half-life period for the disintegration of a radioactive substance is independent of the amount of that substance and it is a first order reaction.

**10. What is the relation between half-life period and order of reaction?**

**Ans:**

$$[t_{1/2}]_1 \propto \frac{1}{a^0}, \text{ since } [t_{1/2}]_1 = \frac{0.693}{k}$$

$$[t_{1/2}]_2 \propto \frac{1}{a^1}, \text{ since } [t_{1/2}]_2 = \frac{1}{k_a}$$

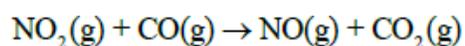
$$[t_{1/2}]_3 \propto \frac{1}{a^2}, \text{ since } [t_{1/2}]_3 = \frac{1.5}{ka^2}$$

In general for the reaction of  $n^{\text{th}}$  order:

$$[t_{1/2}]_n \propto \frac{1}{a^{n-1}}$$

**11. What is rate determining step?**

**Ans:** The slowest step of a reaction is called the rate determining step. Let us consider the following reaction:

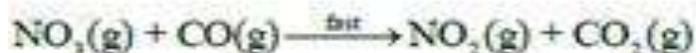


The rate equation of the reaction is found to be:

$$\text{Rate} = k[\text{NO}_2]^2$$

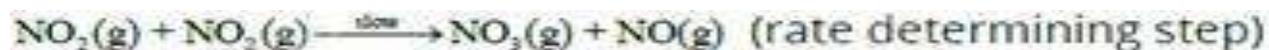
This equation shows that the rate of reaction is independent of the concentration of carbon monoxide. In other words the equation tells us that reaction involves more than one steps and

two molecules of  $\text{NO}_2$  are involved in the rate determining step. The proposed mechanism for this reaction is as follows.



**12. Define reaction intermediate. Give an example.**

**Ans:** The reaction intermediate has a temporary existence and it is unstable relative to the reactants and the products. For example:



$\text{NO}_3$  is the reaction intermediate.

**13. Name the physical methods used to determine rate of reaction.**

**Ans:** Following are the physical methods used to determine rate of reaction:

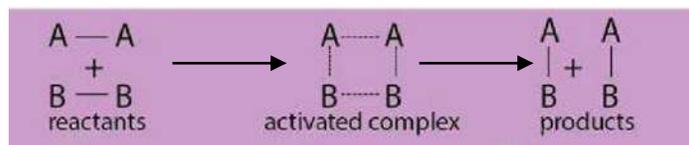
1. Spectrometry
2. Electrical conductivity method
3. Dilatometric method
4. Refractometric method
5. Optical Rotation method

**14. Define activation energy.**

**Ans:** The minimum amount of energy, required for an effective collision is called activation energy ( $E_a$ ).

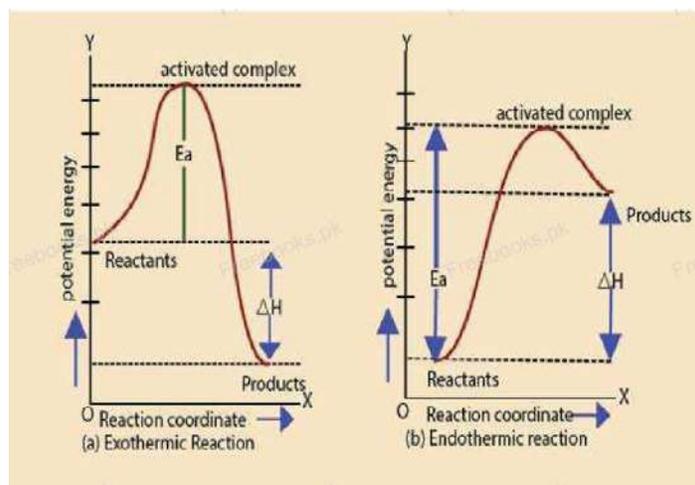
**15. Define activated complex. Give an example.**

**Ans:** Activated complex is an unstable combination of all the atoms involved in the reaction for which the energy is maximum. It is a short lived species and decomposes into the products immediately. It has a transient existence that is why it is also called a transition state.



**16. Draw potential energy diagram for exothermic and endothermic reaction.**

**Ans:** The potential energy diagram can be used to study the heat evolved or absorbed during the reaction. For exothermic reactions, the products are at a lower energy level than the reactants and the decrease in potential energy appears as increase in kinetic energy of the products. For endothermic reactions, the products are at higher energy level than the reactants and for such reactions a continuous source of energy is needed to complete the reaction.



**17. Mention methods to find order of reaction.**

**Ans:** Following are the methods to find order of reaction:

(i) Method of hit and trial

- (ii) Graphical method
- (iii) Differential method
- (iv) Half-life method
- (v) Method of large excess

**18. Explain half-life method.**

**Ans:** Half-life of a reaction is inversely proportional to the initial concentration of reactants raised to the power one less than the order of reaction.

Therefore, 
$$(t_{1/2})_n \propto \frac{1}{a^{n-1}}$$

Let us perform a reaction twice by taking two different initial concentrations 'a<sub>1</sub>' and 'a<sub>2</sub>' and their half-life periods are found to be t<sub>1</sub> and t<sub>2</sub> respectively.

$$t_1 \propto \frac{1}{a_1^{n-1}} \quad \text{and} \quad t_2 \propto \frac{1}{a_2^{n-1}}$$

Dividing the two relations: 
$$\frac{t_1}{t_2} = \left[ \frac{a_2}{a_1} \right]^{n-1}$$

Taking log on both sides: 
$$\log \frac{t_1}{t_2} = (n-1) \log \left[ \frac{a_2}{a_1} \right]$$

$$n-1 = \frac{\log \left[ \frac{t_1}{t_2} \right]}{\log \left[ \frac{a_2}{a_1} \right]}$$

Rearranging

$$n = 1 + \frac{\log \left[ \frac{t_1}{t_2} \right]}{\log \left[ \frac{a_2}{a_1} \right]}$$

So, if we know the two initial concentrations and two half-life values we can calculate the order of reaction (n).

**19. What is meant by method of large excess? Explain.**

**Ans:** In this method, one of the reactants is taken in a very small amount as compared to the rest of the reactants. The active masses of the substances in large excess remain constant throughout. That substance taken in small amount controls the rate and the order is noted with respect to that. The reason is that a small change in concentration of a substance taken in very small amount affects the value of rate more appreciably. The hydrolysis of ethyl acetate as mentioned earlier shows that water being in large excess does not determine the order. In this way, the reaction is repeated by taking rest of the substances in small amounts one by one and overall order is calculated.

**20. Mention factors affecting rate of reaction.**

**Ans:** Following are the factors affecting rate of reaction:

1. Nature of reactants
2. Concentration of reactants
3. Surface area
4. Light
5. Effect of temperature on rate of reaction
6. Arrhenius equation

**21. How activation energy, temperature and rate of reaction are related?**

**Ans:** Arrhenius equation explains the effect of temperature on the rate constant of a reaction. The rate constant 'k' for many simple reactions is found to vary with temperature. According to Arrhenius:

$$k = Ae^{-E_a/RT}$$

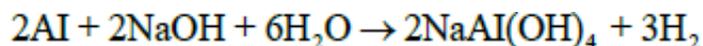
'k' is exponentially related to activation energy ( $E_a$ ) and temperature (T). R is general gas constant and e is the base of natural logarithm. The equation shows that the increase in temperature, increases the rate constant and the reactions of high activation energy have low 'k' values.

**22. What is the effect of nature of reactants on rate of reaction?**

**Ans:** The rate of reaction depends upon the nature of reacting substances. The chemical reactivity of the substances is controlled by the electronic arrangements in their outermost orbitals. The elements of I-A group have one electron in their outermost s-orbital. They react with water more swiftly than those of II-A group elements having two electrons in their outermost s-orbital. Similarly, the neutralization and double decomposition reactions are very fast as compared to those reactions in which bonds are rearranged. Oxidation-reduction reactions involve the transfer of electrons and are slower than ionic reactions.

**23. What is the effect of surface area on rate of reaction?**

**Ans:** The increased surface area of reactants increases the possibilities of atoms and molecules of reactants to come in contact with each other and the rates enhance. For example, Al foil reacts with NaOH moderately when warmed, but powdered Al reacts rapidly with cold NaOH and  $H_2$  is evolved with frothing.



**24. What is the effect of light on rate of reaction?**

**Ans:** Light consists of photons having definite amount of energies depending upon their frequencies. When the reactants are irradiated, this energy becomes available to them and rates

of reactions are enhanced. The reaction of  $\text{CH}_4$  and  $\text{Cl}_2$  requires light. The reaction between  $\text{H}_2$  and  $\text{Cl}_2$  at ordinary pressure is negligible in darkness, slow in daylight, but explosive in sunlight. Similarly, light is vital in photosynthesis, and the rate is influenced by light.

**25. What is the effect of concentration of reactants on rate of reaction?**

**Ans:** An increase in the concentrations of the reactants will result in the corresponding increase in the reaction rate, while a decrease in the concentrations will have a reverse effect. For example

1. Combustion that occurs slowly in air (21% oxygen) will occur more rapidly in pure oxygen.
2. Limestone reacts with different concentrations of hydrochloric acid at different rates. In the case of a gaseous reactant, its concentration can be increased by increasing its pressure. Therefore, a mixture of  $\text{H}_2$  and  $\text{Cl}_2$  will react twice as fast if the partial pressure of  $\text{H}_2$  or  $\text{Cl}_2$  is increased from 0.5 to 1.0 atmosphere in the presence of excess of the other component.

**26. How the effect of concentration of reactants on rate of reaction can be expressed mathematically?**

**Ans:**

The effect of change in concentration on the rate of a chemical reaction can be nicely understood from the following gaseous reaction.



$$\text{Rate} \propto [\text{H}_2]$$

The overall rate equation of reaction is

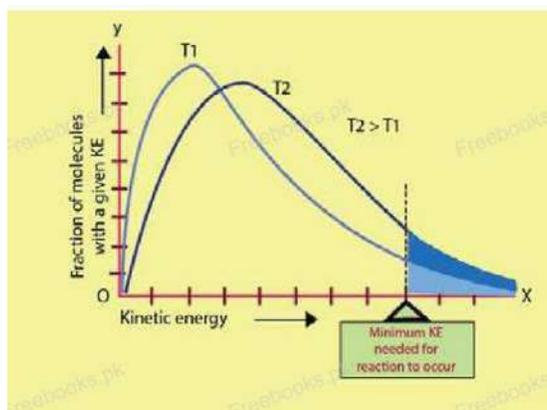
$$\text{Rate} \propto [\text{H}_2][\text{NO}]^2$$

$$\text{Rate} = k[\text{H}_2]^1[\text{NO}]^2$$

Hence, the reaction is a third order one.

**27. What is the effect of temperature on rate of reaction?**

**Ans:** For a collision to be effective the molecules must possess the activation energy and they must also be properly oriented. All the molecules of a reactant do not possess the same energy at a particular temperature. A fraction of total molecules will have energy more than the average energy. This fraction of molecules is indicated as shaded area. As the temperature increases, the number of molecules in this fraction also increases. There happens a wider distribution of velocities. The curve at higher temperature T2 has flattened. It shows that molecules having higher energies have increased and those with less energies have decreased. So, the number of effective collisions increases and hence the rate increases.



**28. What happens to the rate of chemical reaction with the passage of time?**

**Ans:** According to law of mass action, rate of reaction is directly proportional to concentration of

reactants. In all reactions, concentration of reactants decreases with the passage of time. With the decreases in concentration of reactants, rates also show continuous decrease from start till completion of reaction. This justifies that the rate of reaction is an ever changing parameter.

**29. Define with example second order reaction?**

**Ans:** Second order reaction: - When the sum of all the exponents to which the concentrations in the rate equation are raised, is equal to two, then the order of reaction is 2 and it is called a second order reaction.

**Example**

Oxidation of nitric oxide with ozone has shown to be first order with respect to NO and first order with respect to O<sub>3</sub>. The sum of the individual orders gives the overall order of reaction as two.



$$\text{Rate} = k [\text{NO}][\text{O}_3]$$

**30. How does a catalyst affect a reversible reaction? OR Equilibrium constant is not changed in the presence of a catalyst.**

**Ans:** A catalyst cannot affect the equilibrium constant of a reaction but it helps the equilibrium to be reached earlier. The rates of forward and backward steps are increased equally. Even a small amount of the catalyst proves to be sufficient.

**31. Does the physical state of a catalyst changes at the end of reaction?**

**Ans:** A catalyst remains unchanged in mass and chemical composition at the end of reaction. It may not remain in the same physical state.  $\text{MnO}_2$  is added as a catalyst for the decomposition of  $\text{KClO}_3$  in the form of granules. It is converted to fine powder at the end of reaction. It has been found in many cases that the shining surfaces of the solid catalyst become dull.

**32. A very small amount of catalyst may prove sufficient for a reaction. Justify.**

**Ans:** Sometimes we need traces of a metal catalyst to catalyze reactions. For example, 1 mg of fine platinum powder can convert  $2.5 \text{ dm}^3$  of  $\text{H}_2$  and  $1.25 \text{ dm}^3$  of  $\text{O}_2$  to water.

**33. How enthalpy change of a reaction and energy of activation are distinguished?**

**Ans:**

**Activation Energy:**

The minimum amount of energy in addition to average kinetic energy which is just sufficient to convert the reactants into products is called activation energy.

**Enthalpy Change of a Reaction:**

The enthalpy change occurs when the certain number of moles of reactants as indicated by the balanced chemical equation react together completely to give the products under standard conditions, i.e.  $25^\circ\text{C}$  (298 K) and one atmosphere pressure.

**34. Define and give an example of the process of activation of a catalyst.**

**Ans:**

Such substances which promote the activity of a catalyst are called promoters or activators. It is also called “catalyst of a catalyst”.

For example, hydrogenation of vegetable oils is accelerated by nickel. The catalyst activity of nickel can be increased by using copper and tellurium.

**35. Define catalyst. Give an example.**

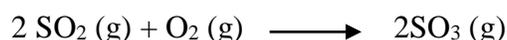
**Ans:** A catalyst is defined as a substance which alters the rate of a chemical reaction, but remains chemically unchanged at the end of the reaction. A catalyst is often present in a very small proportion. For example, the reaction between  $H_2$  and  $O_2$  to form water is very slow at ordinary temperature, but proceeds more rapidly in the presence of platinum. Platinum acts as a catalyst. Similarly,  $KClO_3$  decomposes much more rapidly in the presence of a small amount of  $MnO_2$ .  $HCl$  is oxidized to  $Cl_2$  in the presence of  $CuCl_2$ .

**36. Define catalysis.**

**Ans:** The process which takes place in the presence of a catalyst is called catalysis.

**37. Define homogeneous catalysis. Give an example.**

**Ans:** In this process, the catalyst and the reactants are in the same phase and the reacting system is homogeneous throughout. The catalyst is distributed uniformly throughout the system. For example, the formation of  $SO_3(g)$  from  $SO_2(g)$  and  $O_2(g)$  in the lead chamber process for the manufacture of sulphuric acid, needs  $NO(g)$  as a catalyst. Both the reactants and the catalyst are gases.

**38. Define heterogeneous catalysis. Give an example.**

**Ans:** In some systems the catalyst and the reactants are in different phases. Mostly, the catalysts are in the solid phase while the reactants are in the gaseous or liquid phase. For example oxidation of ammonia to  $NO$  in the presence of platinum gauze helps us to manufacture  $HNO_3$ .



39. Differentiate between Homogeneous and heterogeneous catalysis. Give two examples.

Ans:

Homogeneous catalysis	Heterogeneous catalysis
<p>In this process, the catalyst and the reactants are in the same phase and the reacting system is homogeneous throughout.</p>	<p>In such system, the catalyst and the reactants are in different phases. Mostly, the catalysts are in solid phase, while the reactants are in the gaseous or liquid phase.</p>
<p><b>Example:</b> The formation of SO<sub>3</sub> (g) from SO<sub>2</sub>(g) and O<sub>2</sub> (g) in the lead chamber process needs NO (g) as a catalyst. Both the reactants and the catalyst are gases.</p> $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \xrightleftharpoons{\text{NO}(\text{g})} 2\text{SO}_3(\text{g})$ <p>SO<sub>3</sub> is used for the manufacture of sulphuric acid (H<sub>2</sub>SO<sub>4</sub>).</p>	<p><b>Example:</b> Oxidation of ammonia to NO in the presence of solid platinum gauze is an example of heterogeneous catalysis. Reactants are gases while catalyst is solid.</p> $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \xrightleftharpoons{\text{Pt}(\text{s})} 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$ <p>NO is used for the preparation of Nitric acid (HNO<sub>3</sub>).</p>

**40. What is catalytic poisoning? Give two examples.**

**Ans:** Catalytic poisoning happens due to presence of trace amounts of foreign substances which render them ineffective. Such substances are called poisons. For example, in the presence of Co as an impurity with hydrogen the catalytic activity of catalyst is decreased in the Haber's process for the manufacture of  $\text{NH}_3$ . The manufacture of  $\text{H}_2\text{SO}_4$  in the contact process needs platinum as catalyst. The traces of arsenic present as impurities in the reacting gases makes platinum ineffective.

**42. Define autocatalyst. Give an example.**

**Ans:**

In some of the reactions, a product formed acts as a catalyst. This phenomenon is called autocatalysis. For example when copper is allowed to react with nitric acid, the reaction is slow in the beginning. It gains the speed gradually and finally becomes very fast. This is due to the formation of nitrous acid during the reaction which accelerates the process.

**43. What are enzymes? How they act as catalysts?**

**Ans:**

**Enzymes:**

Enzymes are defined as catalyst of biological systems (animal and plants). They are either purely made of proteins or contain some non-protein part also for their help. Enzyme increase the rate of reaction within the body of living organisms by decreasing the activation energy of reaction. For example, urease catalyzes the hydrolysis of urea and Lipase catalyzes the hydrolysis of fats.

**44. What is meant by negative catalysis?**

**Ans:** When the rate of reaction is retarded by adding a substance, then it is said to be a negative catalyst or inhibitor. For example, tetraethyl lead is added to petrol, because it saves the petrol from pre-ignition.

**45. Mention the characteristics of enzyme catalysis. (Mention any two as answer to short question)**

**Ans:** Following are the characteristics of enzymes:

- (i) Enzymes are the most efficient catalysts known and they lower the energy of activation of a reaction.
- (ii) Enzymes catalysis is highly specific, for example, urease catalyses the hydrolysis of urea only and it cannot hydrolyze any other amide even methyl urea.
- (iii) Enzyme catalytic reactions have the maximum rates at an optimum temperature.
- (iv) The pH of the system also controls the rates of the enzyme catalyzed reaction and the rate passes through a maximum at a particular pH, known as an optimum pH. The activity of enzyme catalyst is inhibited by a poison.
- (v) The catalytic activity of enzymes is greatly enhanced by the presence of a co-enzyme or activator.

**46. Mention the characteristics of a catalyst (Mention any two as answer to short question)**

**Ans:** Following are the characteristics of a catalyst:

1. A catalyst remains unchanged in mass and chemical composition at the end of reaction. It may not remain in the same physical state.  $\text{MnO}_2$  is added as a catalyst for the decomposition

of  $\text{KClO}_3$  in the form of granules. It is converted to fine powder at the end of reaction. It has been found in many cases that the shining surfaces of the solid catalyst become dull.

2. Sometimes, we need a trace of a metal catalyst to affect very large amount of reactants. For example, 1 mg of fine platinum powder can convert  $2.5 \text{ dm}^3$  of  $\text{H}_2$  and  $1.25 \text{ dm}^3$  of  $\text{O}_2$  to water. Dry  $\text{HCl}$  and  $\text{NH}_3$  don't combine, but in the presence of trace of moisture, they give dense white fumes of  $\text{NH}_4\text{Cl}$ . Thousands of  $\text{dm}^3$  of  $\text{H}_2\text{O}_2$ , can be decomposed in the presence of 1 g of colloidal platinum.

3. A catalyst is more affective, when it is present in a finely divided form. For example, a lump of platinum will have much less catalytic activity than colloidal platinum. In the hydrogenation of vegetable oils finely divided nickel is used.

4. A catalyst cannot affect the equilibrium constant of a reaction but it helps the equilibrium to be established earlier. The rates of forward and backward steps are increased equally.

5. A catalyst cannot start a reaction, which is not thermodynamically feasible. It is now considered that a catalyst can initiate a reaction. The mechanism of a catalyzed reaction is different from that of an uncatalyzed reaction.

***47. Enzymes are specific in action. Justify.***

***Ans:***

Enzymes function in the lock and key mechanism. A specific enzyme can combine with a specific substrate having complementary structure. This is called lock-key mechanism of enzyme action.

**48. Differentiate between rate and rate constant of reaction?****Ans:**

<b>Rate</b>	<b>Rate constant</b>
<p>The rate of a reaction is defined as the change in concentration of a reactant or a product divided by the time taken for the change.</p> <p>Rate of reaction</p> $= \frac{\text{Change in concentration of the substance}}{\text{time taken for the change}}$	<p>The ratio between rate of reaction and concentration of reactants is called rate constant.</p> <p>‘The specific rate constant of a chemical reaction is the rate of reaction when the concentrations of the reactants are unity’.</p>
<p><b>Calculation of rate:</b></p> <p>It is calculated as follows</p> <p>Rate of reaction = <math>\frac{\text{Change in concentration of reactant or product}}{\text{time taken for the change}}</math></p>	<p><b>Calculation of rate constant(k):</b></p> <p>It is calculated from rate equation of a reaction.</p>

**49. The unit of rate constant of a second order reaction is  $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  but, but the unit of rate of reaction is  $\text{mol dm}^{-3} \text{s}^{-1}$ .**

**Ans:**

$$\text{Rate} = \frac{\Delta[\text{C}]}{\Delta t} = \frac{\text{mole dm}^{-3}}{\text{s}} = \text{mol dm}^{-3} \text{s}^{-1}$$

In case of second order reaction is

$$\text{Rate} = k[A][B]$$

$$k = \frac{\text{Rate}}{[A][B]} = \frac{\text{mol dm}^{-3} \text{ mol}}{\text{mol dm}^{-3} \text{ mol dm}^{-3}}$$

$$k = \text{dm}^3 \text{ mol}^{-1} \text{s}^{-1}$$

**50. Rate of chemical reaction is an every changing parameter and decreases under the given conditions.**

**Ans:** According to the Law of mass action rate of reaction is directly proportional to concentration of reactants. In all reactions, concentration of reactants decreases with the passage of time. With the decrease in concentration of reactants rates shall also show continuous decrease from start till completion of reaction. This justifies the rate of reaction is an ever changing parameter.

**51. The sum of the coefficient of a balanced chemical equation is not necessarily important to give the order of a reaction.**

**Ans:** Rate of reaction is an experimentally determined quantity. It cannot be predicated from balanced chemical equation. The order of reaction is determined from experimentally obtained rate equation. Through balanced chemical equation we can only calculate the molecularity of reaction.

**Example:**

Consider the decomposition of  $N_2O_5$ .



Apparently it seems to be 2<sup>nd</sup> order reaction but experimentally determined rate equation is:

$$\text{Rate} = k [N_2O_5]$$

It indicates that it is first order reaction.

**52. Explain that the half-life method for measurement of the order of a reaction can help us to measure the order of even those reactions which have a fractional order.**

**Ans:** Following equation is used to calculate order (n) of any reaction including fractional order.

$$n-1 = \frac{\log \left[ \frac{t_1}{t_2} \right]}{\log \left[ \frac{a_2}{a_1} \right]}$$

$$n = 1 + \frac{\log \left[ \frac{t_1}{t_2} \right]}{\log \left[ \frac{a_2}{a_1} \right]}$$

The values of following parameters must be known,  $a_1$  (concentration of reactant in the first experiment),  $a_2$  (concentration of reactant in the second experiment),  $t_1$  (half-life of reaction in the first experiment) and  $t_2$  (half-life of reaction in the second experiment).

**53. A finely divided catalyst may prove more effective. Justify.**

**Ans:** A catalyst is more effective when it is present in a finely divided form. For example a lump of platinum will have much less catalytic activity than colloidal platinum. In the hydrogenation of vegetable oils finely divided nickel is used.

**54. 50% of a hypothetical first order reaction completes in one hour. The remaining 50% needs more than one hour to complete.**

**Ans:** First order reaction is independent of concentration of reactant. Suppose initial amount of reactant is  $1 \text{ mol dm}^{-3}$  and half-life of reaction is 1 hour.

When time according to half-life (1Hour) shall pass the concentration of reactant will be  $0.5 \text{ mol dm}^{-3}$ . After next hour i.e. after 2 hours the concentration of reactant shall decrease to  $0.25 \text{ mol dm}^{-3}$ .

$\text{dm}^{-3}$ . This will continue till completion of reaction.



# Prayer Before an Examination

Dear Lord, as I take this exam, I thank you that my value

is not based on my performance, but on your great love for me.

Come into my heart so that we can walk through this time together.

Help me, not only with this test, but the many tests of life that are sure to come my way.

As I take this exam, bring back to my mind everything I studied

and be gracious with what I have overlooked.

Help me to remain focused and calm, confident in the facts and in my ability,

and firm in the knowledge that no matter what happens today you are there with me.

- A Salesian Prayer

**GOOD**

**LUCK**

**DEAR**

**STUDENTS**

