CHAPTER 1

(BASIC CONCEPTS)

Short Questions :

1. **Define molecular ion. How is it formed?**

Ans: Definition: - The ion which is produced when a molecule loses or gains one or more than one electrons, is called molecular ion, e.g. CH_4^+, CO^+, N_2^+. Cationic molecular ions are more abundant than anionic molecular ions.

Generation of molecular ions: - These ions can be generated by passing high energy electron beam or α-particles or X-rays through a gas.

2. **Define molecular ion, write its uses.**

Ans: Definition: - The ion which is produced when a molecule loses or gains one or more than one electrons, is called molecular ion, e.g. CH_4^+, CO^+, N_2^+.

Uses: - The breakdown of molecular ions obtained from the natural products can give important information about their structure.

3. **What are isotopes? Why they have same chemical but different physical properties?**

Ans: Isotopes: - The isotopes are different kind of atoms of the same element having same atomic number but different masses. For example, carbon has three isotopes, i.e. $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$.

Chemical properties depend upon atomic number so isotopes have same chemical properties. Physical properties depend upon atomic masses so isotopes have different physical properties.

4. **Explain mathematical relationship of m/e of an ion in mass spectrometry.**

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

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

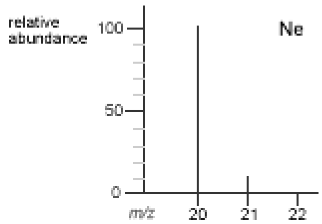
Where H is the strength of magnetic field, E is the strength of electrical field, r is the radius of circular path. If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/e will fall at a different place as compared to the first place.

5. **What is mass spectrum?**

Ans: Mass Spectrum: -In modern spectrographs, each ion strikes a detector, the ionic current is amplified and is fed to the recorder. The recorder makes a graph showing the relative abundance of isotopes plotted against the mass number. This graph is known as mass spectrum.

Example

Following is the mass spectrum of Neon



6. Write functions of Mg $(ClO_4)^2$ and KOH in combustion analysis.

Ans: Function of Mg $(ClO_4)^2$: - Magnesium perchlorate acts as dehydrating agent 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.

7. How does no individual neon atom in the sample of the element has mass 20.18 amu?

Ans: Neon has three isotopes i.e. $^{20}_{10}Ne$, $^{20}_{10}Ne$ and $^{20}_{10}Ne$ with different relative abundance of 90.92%, 0.26% and 8.82% respectively. Average atomic mass of Ne is calculated as:

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

So 20.18 is average atomic mass as no individual neon atom in the sample has a mass of 20.18 amu.

8. Why oxygen cannot be determined directly in combustion analysis?

Ans: During combustion analysis, an excess of oxygen is provided to make sure that all the organic compound must be burnt to produce CO_2 and H_2O. At the end some of 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.

% age of O = 100 – (% age of C + % age of H)

9. Magnesium atom is twice heavier than carbon atom. Comment.

Ans: Magnesium atom (24 amu) is twice heavier than Carbon atom (12amu) because number of fundamental sub-atomic particles is double in Magnesium as compared with Carbon.

Sub-atomic particles	**Carbon**	**Magnesium**
Protons	6	12
Neutrons	6	12
Electrons	6	12
Total particles	18	36
Atomic mass (amu)	12	24

10. How one mg of K_2CrO_4 has thrice the number of ions than the number of formula units when ionized.

Ans: It can be justified as following:

Mass of $K_2Cr_2O_7$ = 1 mg = 0.001 g

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

Number of formula units of $K_2Cr_2O_7$ = $\frac{\text{mass x Na}}{\text{Molar mass}}$

$$= \frac{0.001 \times 06.02 \times 10^3}{294} = 2.04 \times 10^{18}$$

$K_2Cr_2O_7$	$\longrightarrow$	$2K^+$	+	$Cr_2O_7^{2-}$
2.04×10^{18}		$2 \times 2.04 \times 10^{18}$		2.04×10^{18}

Hence it is justified that total number of ions is thrice the number of formula units ionized

11. How 4.9 g of H_2SO_4 when completely ionized in water have equal number of +ve and –ve charges but the number of positively charged ions are twice the number of negativity charged ions.

Ans: Number of moles of H_2SO_4 = 4.9/98 = 0.5 mole

0.5 mole of H_2SO_4 yields 1 mole of H^+ ions and 0.5 mole of SO_4^{2-} ions as shown below:

H_2SO_4	$\longrightarrow$	$2H^+$	+	SO_4^{2-}
0.5 mole		1 mole		0.5 mole
$0.5 \times 6.02 \times 10^{23}$		$1 \times 6.02 \times 10^{23}$		$0.5 \times 6.02 \times 10^{23}$
Molecules		ions		ions

The relationship shows that total positive charges are equal to total negative charges because each SO_4^{2-} ion has -2 charge and two H^+ have also +2 charge. However, the above relationship shows that number of positive ions are twice as compared to negative ions.

12. 23 g of sodium and 39 g of potassium have equal number of atoms in them. Justify.

Ans: It is justified as follows:

	Given mass of Na	= 23 g
	Atomic mass of Na	= 23 g/mol
	No. of moles of Na	= mass of Na/atomic mass of Na
		=23/23 = 1 mole
So,	1 mole of Na contains	= 6.02 x 1023 atoms
	Given mass of U	= 238 g
	Atomic mass of U	= 238 g/mol
	No. of moles of U	= mass of Na/atomic mass of Na
		=238/238 = 1 mole
So,	1 mole of U contains	= 6.02 x 1023 atoms

13. What are molecular ions?

Ans: Definition: - The ion which is produced when a molecule loses or gains one or more than one electrons, is called molecular ion, e.g. CH_4^+, CO^+, N_2^+. Cationic molecular ions are more abundant than anionic molecular ions.

Generation of molecular ions: - These ions can be generated by passing high energy electron beam or α-particles or X-rays through a gas.

14. **Define isotopes**.

Ans: Isotopes: - The isotopes are different kind of atoms of the same element having same atomic number but different masses. For example, carbon has three isotopes, i.e. $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$.

Chemical properties depend upon atomic number so isotopes have same chemical properties. Physical properties depend upon atomic masses so isotopes have different physical properties.

15. 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 ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) 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_6H_6 while the molecular formula of glucose is $C_6H_{12}O_6$.

16. A compound may have same molecular and empirical formula, Justify.

Ans: Empirical formula 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 ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) is CH.

Molecular 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_6H_6 while the molecular formula of glucose is $C_6H_{12}O_6$.

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_2O and for carbon dioxide is CO_2.

17. Define molecular formula. How is it related with empirical formula?

Ans: 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_6H_6 while the molecular formula of glucose is $C_6H_{12}O_6$.

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

Molecular formula = n (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.

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

18. 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.

$$\underset{4g}{2H_2} + \underset{32g}{O_2} \longrightarrow \underset{36g}{2H_2O}$$

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.

19. Calculate the number of water molecules in 10 g of ice.

Ans: Given data:

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

Solution:

$$n = \frac{\text{Mass of ice}}{\text{molar mass of ice}} \times N_A$$

$$n = \frac{10}{18} \times 6.02 \times 10^{23}$$

$$= 3.31 \times 10^{23} \text{ molecules}$$

20. Define empirical formula and molecular formula with examples.

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 ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) 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_6H_6 while the molecular formula of glucose is $C_6H_{12}O_6$.

21. Give assumptions of stoichiometry.

Ans: Assumption of stoichiometry

There are two assumptions of stoichiometry:

1. All the reactants are completely converted into the products.
2. No side reaction occurs.

22. Calculate the mass in Kg of 2.6×10^{20} molecules of SO_2.

Ans: Given data:

Molecules of SO_2 = 2.6×10^{20}
Molar mass of SO_2 = 32 + 32 = 64 g/mol
Mass of SO_2 = ?

Solution:

$$\text{mass of } SO_2 = \frac{\text{Molecules of } SO_2}{N_A} \times \text{molar mass of } SO_2$$

$$\text{mass of } SO_2 = \frac{2.6\times10^{20}}{6.02 \times 10^{23}} \times 64$$

$$= 2.76 \times 10^{-4}\ \text{g} = 2.76 \times 10^{-7}\ \text{Kg}$$

23. 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,

1.008 g of hydrogen	= 1 mole of hydrogen	= 6.02×10^{23} atoms of H
18 g of water	= 1 mole of H_2O	= 6.02×10^{23} atoms of H
96 g of SO_4^{2-}	= 1 mole of SO_4^{2-}	= 6.02×10^{23} ions of SO_4^{2-}

Equation:

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

24. One mole of H2SO4 should completely react with two moles of NaOH. How does Avogadro's number help to explain it?

Ans: According to balanced chemical equation

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

1 mole 2 moles 1 moles 2 moles

Apply Avagadro's number concept:

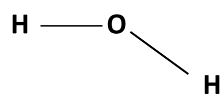
1 mole of H2SO$_4$ generate H+ ions = $2 \times 6.02 \times 10^{23}$ H^+ ions

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

It is clear from the above calculations that the number of H^+ and OH^- ions formed are same although the number of moles of H_2SO_4 and NaOH are different that is why 1 mole of H_2SO_4 reacts completely with 2 moles of NaOH.

25. One mole of H2O has 2 moles of bonds, 3 moles of atoms, 10 moles of electrons and 28 moles of total fundamental particles preset in it.

Ans: Following is the structure of water:



It is clear from this structure that one molecule of water has two covalent bonds, two hydrogen atoms and one oxygen atom. So one mole of water shall contain:

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 re contributed by 2 moles of hydrogen.
iv. Number of particles in one moles of oxygen = 8 P + 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

26. How N2 and CO have same number of electrons, protons and neutrons.

Ans: Both N2 and CO have same number of electrons, protons and neutrons as it is clear from the following explanation.

For N2

Number of electrons = 7 + 7 = 14
Number of protons = 7 + 7 = 14
Number of neutrons = 7 + 7 = 14

For CO

Number of electrons or protons or neutron in C = 6
Number of electrons or protons or neutron in O = 8
Number of electrons or protons or neutron in CO = 6 + 8 = 14

27. Calculate the number of moles of oxygen atoms in 9 g of $Mg(NO_3)_2$.

Ans: Given data:

Mass of Mg $(NO_3)_2$ = 9.00 g

Molar mass of $Mg(NO_3)_2$ = 24 + 2 (14 + 3 x 16) = 148 $gmol^{-1}$

Number of moles of O atoms= ?

Solution:

Number of moles of $Mg(NO_3)_2$ = Mass of $Mg(NO_3)_2$ / Molar mass of $Mg(NO_3)_2$

= 9/148 = 0.06 mol

1 mole of $Mg(NO_3)_2$ contains moles of O atoms = 6 mol

0.06 moles of $Mg(NO_3)_2$ contain moles of O atoms = 0.06 x 6 = 0.36 moles of O atoms

28. Calculate the mass in grams of 2.74 moles of $KMnO_4$ (At.wt. K=39 amu, Mn 55 amu, O=16 amu)

Ans: Given data:

Number of moles of KMnO4 = ?

Molar mass of $KMnO_4$ = 39 + 55 + 64 = 158 $gmol^{-1}$

Solution

Mass of $KMNO_4$ = Moles of $KMnO_4$ x Molar mass of $KMnO_4$

= 2.74 x 158 = 432.92 g

29.Why do 2 g of H_2, 16g of CH_4, 44g of CO_2 occupy separately the volume of 22.414 dm^3 although the sizes and masses of molecules of three gases are very different from each other?

Ans: 2 g of H_2 = 1 mole = 6.02×10^{23} molecules 22.4 dm^3 at STP

16 g of CH_4 = 1 mole = 6.02×10^{23} molecules 22.4 dm^3 at STP

44 g of CO_2 = 1 mole = 6.02×10^{23} molecules 22.4 dm^3 at STP

According to Avogadro's law, equal number of molecules of all gases occupy same volumes at same temperature and pressure. Since H_2, CH_4 and CO_2 have same number of molecules that is why these occupy same volume.

30. Define limiting reactant. Give an example.

Ans: Limiting Reactant: - A reactant which is consumed earlier due to its lesser quantity and gives less amount of product in a chemical reaction is called a limiting reactant.

Example: - The burning of a piece of paper in the atmosphere is common example in which paper is limiting reactant and atmospheric oxygen is in excess.

31. How do many chemical reactions taking place in our surrounding involve limiting reactants?

Ans: There are many reactions taking place in our surrounding which involve limiting reactant. Some of them are given here:

1. Burning of paper where paper is a limiting reactant.

2. Rusting of iron where iron is a limiting reactant.

32. Define actual yield. Write formula for the calculation of % age yield.

Ans: Actual Yield:- The amount of the products obtained in a chemical reaction is known as actual yield.

Formula

The formula to calculate, percentage yield is as follows:

$$\%\ \text{age yield} = \frac{\text{Actual yield}}{\text{Theoretical Yield}} \times \text{molar mass of } SO_2$$

33. Why theoretical yield is greater than actual yield?

Ans: Following are some of the reasons for this issue:

1. A practically inexperienced worker has many shortcomings and cannot get the expected yield.
2. Leakage of gaseous products due to effervescence and bubbling
3. Impurities present in reactants
4. Side reactions leading to formation of side products (un-expected products)
5. The processes like filtration, separation by distillation, separation by a separating funnel, washing, drying and crystallization if not properly carried out, decreases the actual yield.

34. Why we calculate %age yield?

Ans: A chemist is usually interested in the efficiency of a reaction. The efficiency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage yield.

$$\%\ \text{age yield} = \frac{\text{Actual yield}}{\text{Theoretical Yield}} \times \text{molar mass of } SO_2$$

CHEMISTRY (XI) CHAPTER 2 (EXPERIMENTAL TECHNIQUES IN CHEMISTRY)

Short Questions:

1. Define analytical chemistry.

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

2. Define qualitative and quantitative analysis.

In Qualitative analysis, a chemist is only concern with the detection or identification of elements present in a compound. Where as in quantitative analysis, a chemist is also concern with the exact amount of elements present in the compound.

3. Name the various experimental techniques used for the purification of the substances.

Purification techniques are as follows;

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

4. In solvent extraction technique, repeated extractions using small portions of solvent are more

efficient than using a single extraction but with larger volume of solvent. Comment.

The benefit of repeated extractions using small portion of solvent is that, we can extract almost all solute. Where as some traces are left if we use single extraction with large volume.

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

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

6. What is difference between Gooch crucible and Sintered glass crucible?

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

7. Write the names of major steps of crystallization.

Steps 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) De-colorization

8. Desiccator is the safest method of drying the crystals. Explain.

Desiccator is slow but safest method of drying, because in this method, crystals preserve their shape and identity. Where as 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 $CaCl_2$, Silica gel or Phosphorous Pentoxide.

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

Crude products contain many contaminations that need to be removed. So, the crystallization process is used for this purpose.

10. How crystallized substances are dried?

Crystallized substances can be dried by following processes;

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

11. How crystals are dried by safest and reliable method?

Safest and reliable method for drying of crystals is using vacuumed desiccator.

12. Write four properties of a good solvent.

Four properties of a good solvent are as follows:

(i) It should not react chemically with the solute
(ii) It should be inexpensive
(iii) It should be safe to use and easily removable
(iv) On cooling, it should deposit well-formed crystals of pure compound

13. Write the names of eight solvents used for the crystallization.

Names of solvents are as follows:

(i) Water
(ii) Rectified Spirit (95% Ethanol)
(iii) Absolute Alcohol
(iv) Diethyl Ether
(v) Acetone
(vi) Chloroform
(vii) Carbon Tetrachloride
(viii) Acetic Acid

14. Define sublimation and partition law.

Process in which a solid, on heated, directly converts into vapors without passing through liquid phase, is called Sublimation.

Partition law states that a solute distributes itself between two immiscible liquids in a constant ratio of concentration, irrespective of amount of solute added.

15. Give the importance of sublimation.

Impure solid substances can be pure through sublimation.

16. What type of substances can be purified by sublimation?

Solid substances can be purified by sublimation. Examples of such solids are Ammonium Chloride, Iodine, Naphthalene, Benzoic acid etc.

17. Define Sublimand and sublimate.

Sublimand is the impure solid substance to be sublime.

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

18. What is solvent extraction?

Solvent extraction is a technique to separate the solute from the solution by shaking the solution with a solvent in which solute is more soluble, and the added solvent does not mix with the solution.

19. Define Rf value. Give its unit.

For the separation of two component, when we use the technique Chromatography, a term Retardation factor is used which depends upon the distribution coefficient of each component. Retardation factor is denoted as Rf value. Since it is a ratio of distance travelled by component from original spot to the distance travelled by solvent from original spot, so it has no unit.

20. Give two applications of paper chromatography.

Paper chromatography is used to separate the colored mixtures like pigments.

It is also used to separate and identify the organic and inorganic compounds from a mixture.

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

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 of filter paper is increased due to which rate of filtration is also increased.

22. Define chromatography. Give its two uses.

Chromatography is originating from Greek word "Khromatos" means color writing. It is mainly used for separation and purification.

23. Differentiate between partition and adsorption chromatography.

Chromatography in which stationary phase is a liquid, is call partition chromatography. Whereas chromatography in which stationary phase is a solid, is called adsorption chromatography.

24. Write main uses of chromatography.

Main uses of Chromatography are separation, purification and identification of mixtures.

CHAPTER 3: (GASES)

Short Questions:

1. Define pressure. Give its different units.

Ans: Pressure: The force applied per unit area is called pressure.

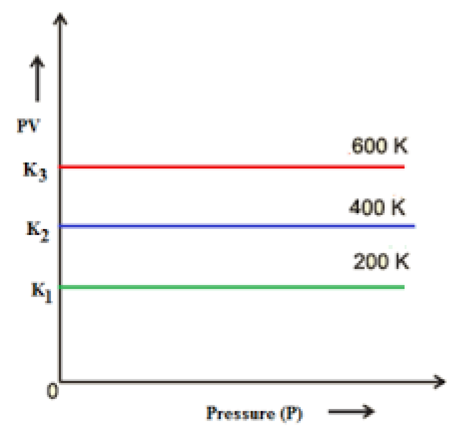
Units: Its different units are atmosphere, mm Hg, torr, Nm^{-2} and Pascal.

2. Write down the value of atmospheric pressure in four different units.

Ans: 1 atm = 760 mmHg = 760 torr = 101325 Nm^{-2}

3. 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.



4. Explain Boyle's law with the help of KMT.

Ans:- According to one of the postulates of kinetic molecular theory of gases, the kinetic energy is directly proportional to the absolute temperature of the gas.

The kinetic energy of N molecules = $1/2\ mN\bar{c}^2$

So $$1/2\ mN\bar{c}^2 \propto T$$

$$1/2\ mN\bar{c}^2 = kT \qquad \text{eq (1)}$$

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

$$PV = 1/3\ mN\bar{c}^2$$

Multiplying and dividing by 2 on right hand side

$$PV = 2/3\ (1/2\ mN\bar{c}^2) \qquad \text{eq (2)}$$

Putting equation 1 into equation 2

$$PV = 2/3\ kT$$

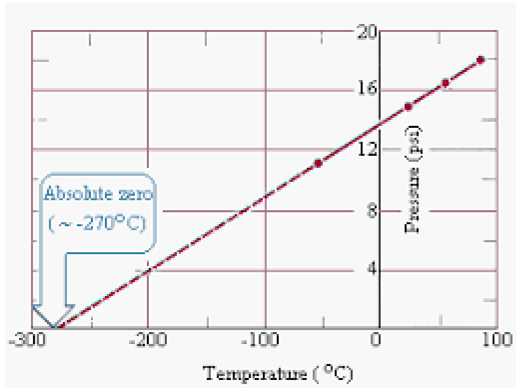
If the temperature T is constant then right hand side of the above equation i.e., 2/3kT is constant. Let that constant be k′. So,

$$PV = k'$$

Which is Boyle's law. Hence at constant temperature and at a fixed number of moles, the product PV is a constant quantity.

5. Justify that volume of gas becomes theoretically zero at -273°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 astraight line which cuts the temperature axis at -273.16^{0}C. This can be possible only if we extrapolate the graph upto -273.16^{0}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 above this temperature.



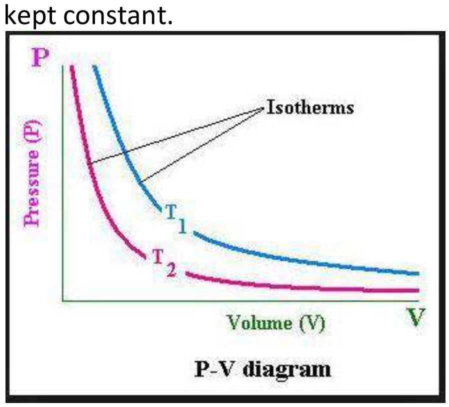
6. What do you mean by absolute zero temperature of gases?

Ans: The hypothetical temperature at which the volume of a given mass of an ideal gas becomes zero is called Absolute zero. Its value is -273.16 ^{0}C or zero K.

7. What are isotherms?

Ans: 'Iso' means same and 'therm' means heat. When a graph is plotted between two parameters keeping the temperature constant, the curve is called an isotherm.

For example, for graphical explanation of Boyle's law, if a graph is plotted between pressure on x-axis (abscissa) and volume on the y-axis (ordinate), then a curve is obtained as shown in the figure. This is isotherm because temperature is kept constant.



8. Why lighter gases diffuse more rapidly than heavier gases?

Ans: The diffusion is the process of gradual mixing of molecules of one gas with molecules of another gas. Lighter gases diffuse more rapidly than heavier gases following Graham's law of diffusion or effusion which stated that

"Rate of diffusion or effusion of a gas in inversely proportional to the square root of its molar mass of density".

In fact, lighter gases have greater velocities and thus greater rates of diffusion.

9. Calculate the density of methane at STP.

Ans: Given data:-

Temperature of gas = 0^oC or 273 K

Pressure of gas = 1 atm

Molecular mass of CH_4 = 16 $gmol^{-1}$

Density of CH_4 = ?

Solution

$$d = PM/RT$$
$$= 1 \times 16/ 0.0821 \times 273$$
$$= 0.7138 \text{ g dm}^{-3}$$

10. State Avogadro's Law.

Ans: It is defined as "equal volumes of all ideal gases at the same temperature and pressure contain equal number of molecules." It means

22.414dm^3 of an ideal gas at STP = 1 mole of gas = 6.02×10^{23} molecules

11. Calculate number of molecules and number of atoms in 20 cm^3 of CH_4 at 0^oC and 700mm of Hg.

Ans:- Given data:-

Volume of CH_4 = 20 cm^3

Temperature = 0^oC or 273 K

Pressure = 700 mm of Hg = 0.921 atm

No. of moles of CH_4 = ?

Molecules of CH_4 = ?

Atoms of CH_4 = ?

Solution

$$PV = nRT$$
$$n = PV/RT$$
$$n = 0.921 \times 20/ 0.082 \times 273$$
$$n = 18.42 / 22.386$$
$$n = 0.823 \text{ mol}$$

No. of molecules of CH_4 = $0.823 \times 6.02\ 10^{23} = 4.95 \times 10^{23}$ molecules

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

12. State Joule-Thomson Effect. Write its application.

OR Define Joule-Thomson Effect.

Ans: Joule- Thomson Effect: When a highly compressed gas is allowed to expand into the region of low pressure, it gets cooled.

Application: N_2 and O_2 are liquefied on industrial scale by Linde's method of liquefaction which is the practical application of Joule- Thomson effect.

13. 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 under these conditions, 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 sized is very small, so they behave ideally at room temperature.

Whereas SO_2 and Cl_2 are either polar (SO_2) or having bigger molecules (both SO_2 and Cl_2), so there are stronger intermolecular forces in them which make them non-ideal at room temperature.

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

Ans: The faulty assumptions of KMT are:

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.

15. Calculate the value of R in S.I units.

Ans:- Value of "R" in SI units

$n = 1$ mole
$T = 273.17$ K
$P = 1$ atm $= 101325\ Nm^{-2}$
$V = 22.414\ dm^3 = 0.022414\ m^3$

Putting their values along with units.

$R = PV/nT$
$= 101325\ Nm^{-2} \times 0.022414\ m^3 / 1\ mol \times 273.16\ K$
$= 8.3143\ Nm\ K^{-1}\ mol^{-1}$
$= 8.3143\ J\ K^{-1}\ mol^{-1}$

16. Calculate the value of R in units' atm.dm^3.k^{-1}mol^{-1}.

Ans : Value of "R" in SI units

$n = 1$ mole
$T = 273.17$ K
$P = 1$ atm
$V = 22.414\ dm^3$

Putting their values along with units.

$R = PV/nT$
$= 1\ atm \times 22.414\ dm^3 / 1\ mol \times 273.16\ K$
$= 0.0821\ atm\ dm^3\ K^{-1}\ mol^{-1}$

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

Ans: Relationship for molecular mass of a gas

According to the general gas equation:

$PV = nRT$

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

$PV = m\ RT / M$
$M = mRT/PV$

18. Derive expression for the density of the gas using general gas equation.

Ans: Relationship for density of a gas

According to the general gas equation:

$PV = nRT$

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

$PV = m\ RT / M$√

And $d = m/V$

So above equation becomes

$$m/V = PM/RT$$

or $$d = PM/RT$$

19. Give four fundamental postulates of KMT of gases.

Ans: 1. Every gas consists of a large number of very small particles called molecules. Gases like He, Ne, Ar have monoatomic 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 thw walls of the container. The collisions among the molecules are perfectly elastic.

4. The average kinetic energy of the gas molecules varies directly as the absolute temperature of the gas.

20. Derive Graham's law of diffusion in the light of KMT of gases.

Ans: Applying the kinetic equation

$$Pv = 1/3\ mN\bar{c}^2$$

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

$$Pv = 1/3\ mN_A\bar{c}^2$$

Or $$PV = 1/3\ M\bar{c}^2 \quad (M = mN_A)$$

Where M is the molecular mass of the gas

Or $$\bar{c}^2 = 3PV/M$$

Taking square root

$$\sqrt{\bar{c}^2} = \sqrt{3PV/M}$$

$$\sqrt{\bar{c}^2} = \sqrt{3P/M/V} = \sqrt{3P/d} \quad (M/V = d)$$

'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.

$$\sqrt{\bar{c}^2}\ \alpha\ r$$

So $$r\ \alpha\ \sqrt{3P/d}$$

At constant pressure

$$r\ \alpha\ \sqrt{1/d}$$

which is Graham's law of diffusion.

21. Prove that $P_A=Pt.X_A$

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

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

$P_AV = n_A RT$ (equation for gas A)

$P_BV = n_B RT$ (equation for gas B)

$P_AV/P_tV = n_A RT/ n_t RT$

$P_A / P_t = n_A / n_t$

$P_A = n_A / n_t\ P_t$

Hence $P_A = X_A P_t$ (X_A is a mole fraction of gas A)

22. Why regular air can't be used in diver's tanks?

Ans: Regular air can't be used in diver's tanks because in sea, after every 100 feet depth, the diver experiences approximately 3 atm pressure. Moreover, the pressure of N_2 increases in depth of sea and it diffuses in the blood.

23. Calculate fraction of total pressure exerted by Oxygen when equal masses of CH_4 and O_2 are mixed into an empty container at 25₀C.

24. What do you mean by critical temperature of gases?

Ans: The highest temperature at which a gas can exist as a liquid is called its critical temperature (T_C). For example, the critical temperature of oxygen is 154.4K (-118.75^{0}C)

25. H_2 and He behave ideally while Cl_2 and SO_2 do not. Why?

Ans: Gases are non- ideal at high pressure and low temperature because under these conditions, 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 sized is very small, so they behave ideally at room temperature.

Whereas SO_2 and Cl_2 are either polar (SO_2) or having bigger molecules (both SO_2 and Cl_2), so there are stronger intermolecular forces in them which make them non-ideal at room temperature.

26. SO_2 is comparatively non-ideal at 273K but behave ideally at 373K.

Ans: Low temperature decreases the kinetic energies of molecules and will be responsible for strengthening of intermolecular forces. A gas with intermolecular forces is always non- ideal. 273K generates intermolecular forces and creates non- ideality but 373K breaks intermolecular forces and creates ideality.

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

Ans: The diffusion is the process of gradual mixing of molecules of one gas with molecules of another gas. Lighter gases diffuse more rapidly than heavier gases following Graham's law of diffusion. As ammonia is lighter than HCl, thus its rate of diffusion is greater than HCl.

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

Ans: Vander Waal's equation accounts for intermolecular forces which decrease the observed pressure of a gas. So pressure of NH3 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:

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

29. Where do natural and artificial plasma exist?

Ans: Natural Plasma: Entire universe is almost in a state of plasma. Plasmas are found in everything from the sun to quarks. The sun is 1.5 million kilometer ball of plasma. All the stars that shine are plasma. On earth it occurs in a few limited places, like lightning bolts, flames and auroras.

Artificial Plasma: Fluorescent light bulbs, neon signs, plasma processing of semiconductors, sterilization of some medical equipment, lamps, lasers, diamond coated films, high power microwave sources and pulsed power switches are examples of artificial plasma.

30. Write two characteristics of plasma.

Ans: 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 generates fields and electric currents from within plasma density of 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.

31. Define Plasma. Give its one application.
Ans: Plasma is called as fourth state of matter. A gaseous mixture containing positive ions, electrons and neutral atoms is called as plasma.
Neon signs are glass tubes filled with gas. When they are turned on, the electricity flows through the tube. The gas gets charged and plasma is produced inside the tube. The plasma glows with a special color depending on what kind of gas is inside.

32. Give two applications of Plasma.
Ans: 1. Neon signs are glass tubes filled with gas. When they are turned on, the electricity flows through the tube. The gas gets charged and plasma is produced inside the tube. The plasma glows with a special color depending on what kind of gas is inside.
2. Plasmas drive lasers and particle accelerators.

33. 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. Greater the value of 'a', stronger are the attractive forces and greater is the non- ideal behavior of the gas.
'b' is effective volume or excluded or incompressible volume per mole. Its value depends on the size of the gas molecules. Greater is the size of gas molecules, greater the value of 'b' and greater is the non-ideal behavior.

34. What are the units of Vander Waal's constants 'a' and 'b'.
Ans: Units of constant 'a': Common units of a = atm dm^6 mol^{-2}
SI units of a = Nm^{+4} mol^{-2}

35. What are Neon advertisement signs?
Ans: Neon signs are glass tubes filled with gas. When they are turned on, the electricity flows through the tube. The gas gets charged and plasma is produced inside the tube. The plasma glows with a special color depending on what kind of gas is inside.

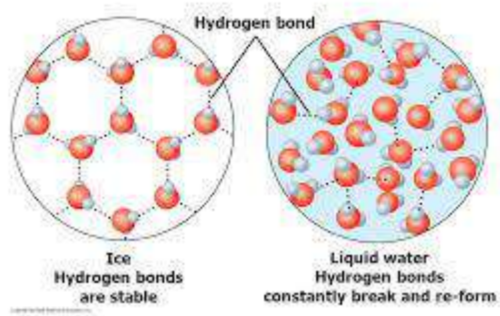
36. How is fluorescent light bulb different from ordinary light bulb?
Ans: A fluorescent light bulb is not like regular light bulbs. Inside the long tube is the gas. When the light is turned on, electricity flows through the tube. This electricity acts as special energy and charges up the gas. This charging and excitation of the atoms create a glowing plasma inside the tube.

CHEMISTRY CHAPTER 4:
(LIQUIDS AND SOLIDS)
Short Questions:

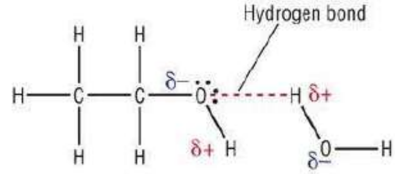
1. What are dipole-dipole forces of attraction? Explain with examples.
Ans: The positive end of one molecule attracts the negative end of the other molecule and these electrostatic forces of attraction are called dipole- dipole forces of attraction.
For example, in case of HCl molecules, both atoms differ in electronegativity. Chlorine being more electronegative develops the partial negative charge and hydrogen develops the partial negative charge. So they have dipole- dipole attractive forces among their molecules. Similarly,
Chloroform $CHCl_3$ molecules have such attractive forces among their molecules.

2. What are 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 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.

3. 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.



4. Water and ethanol can mix in all proportions. Give reason.
Ans: Water and ethanol can mix in all proportions because both can form hydrogen bonds with each other.

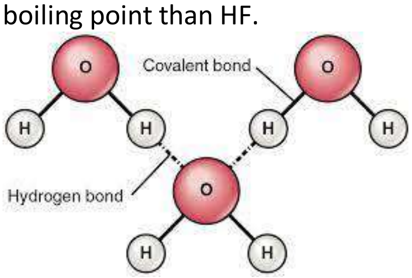


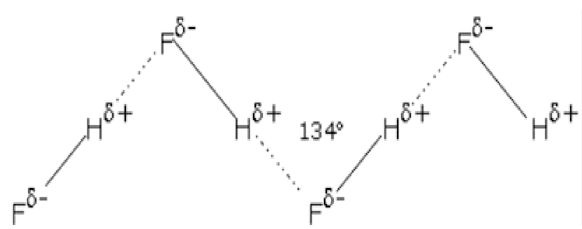
5. Lower alcohols are soluble into water but hydrocarbons not. Give reason.

Ans: Lower alcohols are soluble in water as they can form hydrogen bonds with each other. 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.

6. Write a brief note on solubility oh 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.

7. 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.





8. 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 cool.

9. 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.

10. 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.

11. Why vapor pressure increases with temperature?
Ans: At high temperature, the molecules having high energy increase and so capability of molecules to leave the surface increases. It causes the increase of vapor pressure.

12. 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 98^0C at Murree hills due to external pressure of 700 torr while at the top of Mount Everest, water boils at only 69^0C at 323 torr.

13. 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.

14. 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 290^0C at 760 torr pressure but decomposes at this temperature. Hence, glycerin cannot be distilled at this temperature. Under vacuum, the boiling temperature of glycerin decreases to 210^0C at 50 torr. It is distilled at this temperature without decomposition and hence can be purified easily.

15. 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.

16. 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.

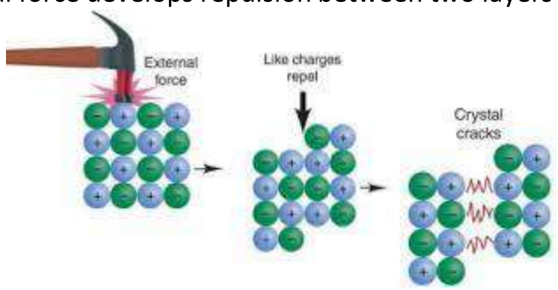
17. Write down two applications of liquid crystals.
Ans: 1. Liquid crystals are used to find the point of potential failure in electric circuits. Room thermometers also contain liquid crystals with a suitable temperature range. As the temperature changes, figures show up in different colours.
2. Liquid crystals are used in the display of electrical devices such as digital watches, calculators and laptop computers.

18. Define isomorphism and polymorphism with examples.
Ans: Isomorphism: 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 the number of atoms and their way of combinations. For example, $NaNO_3$ and KNO_3 have rhombohedral crystals whereas Cu and Ag have cubic crystals.
Polymorphism: 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, $AgNO_3$ has rhombohedral and orthorhombic crystalline forms.

19. 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, $AgNO_3$ has rhombohedral and orthorhombic crystalline forms.

20. 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, like ions come in front of each other and hence begin to repel. So, the application of a little external force develops repulsion between two layers causing brittleness.



21. 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.

22. Define transition temperature with 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 on form exists. For example, transition temperature of tin is 13.2^0C 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 95.5^0C.

23. 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.

24. 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.

25. 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, $AgNO_3$ 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 phenomenon 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.

26. What is Isomorphism? Give example?

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 the number of atoms and their way of combinations. For example, $NaNO_3$ and KNO_3 have rhombohedral crystals whereas Cu and Ag have cubic crystals.

27. 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 on form exists. For example, transition temperature of tin is 13.2^0C 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 95.5^0C. There are many compounds which show transition temperature. For example, KNO_3 has transition temperature of 32.380C having orthorhombic and rhombohedral crystals. Thus transition temperature is the term used for elements as well as compounds.

28. 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 on form exists. For example, transition temperature of tin is 13.2^0C 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 95.5^0C.

29. The vapor pressure of diethyl ethyl 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.

30. Give four properties of molecular solids.

Ans: 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.

31. 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.

32. 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.

33. Why ice floats over the surface of water?
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. Thus, ice floats over the surface of water as water is more denser than ice.

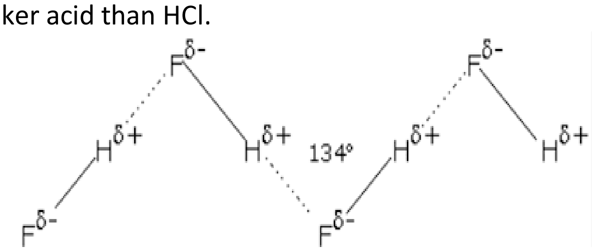
34. Define allotropy. Give its one example.
Ans: The existence of an element in more than one crystalline form is called allotropy. For example, Sulphur has rhombic and monoclinic crystalline forms.

35. Write two properties of molecular solids.
Ans: 1. Molecular solids have weak forces of attraction among their molecules.
2. They are soft and easily compressible.

36. Why the electrical conductivity of metals decrease by increasing temperature?
Ans: Metals are good conductors of electricity. Sometimes, they 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.

37. 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 0^0C, solid water (ice) exists in dynamic equilibrium with liquid water.

38. 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 it proton H^+ easily. Thus, HF is a weaker acid than HCl.



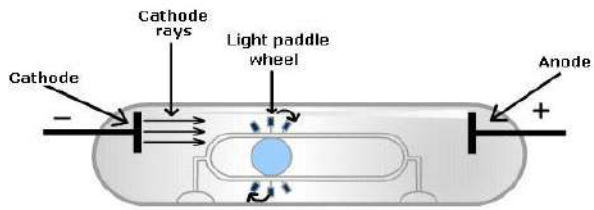
39. Diamond is hard and an electrical insulator. Give reason?
Ans: There is SP^3- SP^3 overlapping of the carbon atoms in diamond and it has tetrahedral structure with 1.54 A^0 bond lengths most suitable for effective packing of atoms, thus diamond is hard. As there are no

free electrons present in any of the carbon atoms which could conduct electricity, therefore, diamond is an electrical insulator.

CHAPTER 5 (Atomic structure)

Short Question and Answers

1. **Cathode rays are material in nature justify it.**
   These rays drive a small paddle wheel placed in their path. This shows that they possess momentum. From this observation, it is inferred that cathode rays are not rays but, material particles having a definite mass and velocity.



2. **Why it is important (necessary) to decrease the pressure in the discharge tube to get the cathode rays?**
   The pressure in discharge tube is decreased to allow the cathode rays to move freely from one electrode to the other. In this way, the possibility of collisions between rays and the gas molecules are minimized. **OR**

   The current does not flow through the gas at ordinary pressure even at high voltage about 500 volts. However when the pressure inside the tube is decreased, the gas in the tube begins to conduct electricity at low pressure. Therefore it is necessary to decrease the pressure in the discharge tube to get the cathode rays.

3. W**hy e/m of cathode rays is just equal to that of electron?**
   A cathode ray consists of beam of electrons, so cathode rays are actually electrons. Therefore e/m value of cathode ray is just equal to that of electron.

4. **Give properties of positive rays.**
   Following are the properties of positive rays:
   - These rays travel in a straight in a direction opposite to the cathode rays.
   - They produce flashes on ZnS plate.
   - There are deflected by an electric as well as magnetic field.
   - They are positively charged.
5. **Give reason for the production of positive rays.**
   The positive rays are produced when high speed cathode rays (electrons) strike the molecules of the gas enclosed in a discharge tube. They knock out electrons from the gas molecules and positive ions are produced, which start moving towards cathode in the form of positive rays.

$$M + e^- \longrightarrow M^+ + 2e^-$$

6. **Give any two properties of Neutron.**

   Following are the two properties of neutron.

   - Neutrons cannot ionize gases.
   - Neutrons are highly penetrating particles

7. **Write the nuclear reaction for the decay of neutron.**

   Free neutron decays into a proton ($^1{}_+P$) with the emission of an electron (e) and a neutrino (n)

$${}^1_0n \longrightarrow {}^1_{+0}P + {}^0_{-1}e + {}^0_0n$$

8. **How neutron was discovered by Chadwick. Give nuclear equation involved.**

   When a stream of α–particles from a polonium source is directed at beryllium target, penetrating radiations are produced, which are called neutrons because the charge detector showed them to be neutral particles.

$${}^4_2He + {}^9_4Be \longrightarrow {}^{12}_6C + {}^1_0n$$

   (α- particles)

9. **Calculate mass of an electron when e/m=1.7588 x 1011C.Kg-1 OR Calculate mass of an electron from its e/m value.**

   The value of charge on electron is 1.602×10^{-19} , while e/m of electron is 1.7588×10^{11} So,

$$e/m = 1.6022 \times 10^{-19}\ C / \text{Mass of electron} = 1.7588 \times 10^{11}\ Ckg^{-1}$$

$$\text{Mass of electron} = 1.6022 \times 10^{-19}\ C / 1.7588 \times 10^{11}\ Ckg^{-1}$$

   Rearranging,

$$\text{Mass of electron} = 9.1095 \times 10^{-31}\ Kg$$

10. **What is Moseley's law?**

   Mosely's law states that the frequency of 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.Mathematically ,

$$\sqrt{v} = a(Z - b)$$

11. **How do you come to know that velocities of electrons in higher orbits are less than those in lower orbits of hydrogen atom?**

   According to Bohr's proposal, the centrifugal force of the electron is equal to the force of attraction between nucleus and electron.

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_o r^2}$$

   Rearranging it,

$$r = \frac{Ze^2}{4\pi\varepsilon_o mv^2}$$

The factors are Z, 4, π, Ɛ∘ and m are constant, so r α $\frac{1}{V^2}$

According to this equation, radius and velocities are inverses to each other. Greater the velocity of the moving electron, smaller the radius.

12. Energy of an electron is inversely proportional to n2 but energies of higher orbits are always greater than those of the lower orbits. Justify it.

The formula for the energy of an electron revolving in any orbit is given by the equation.

$$E = -\ 2.18\times10^{-18}\ j\ (\frac{1}{n^2})$$

Greater the value of 'n' greater the value of energy because energy is negative inverse of n. It becomes more and less negative. The value of energy approaches to zero when n= ∞

13. **What is Plank's quantum theory? OR Write down any two postulates of Plank's Quantum theory.**

According to this theory, energy travels in a discontinue manner and it is composed of large number of tiny discrete units called quanta.

The main postulates of this theory are:

1. 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.
2. The amount of energy associated with a quantum of radiation is proportional to the frequency (ν) of the radiation

$$E \propto \nu$$
$$E = h\nu$$

14. State Heisenberg's uncertainty principle and give its equation.

Heisenberg showed that it is impossible to determine simultaneously both the position and momentum of an electron. Suppose that ∆x is the uncertainty in the measurement of the position and ∆p is the uncertainty in the measurement of momentum of an electron.

$$\Delta x\ \Delta p \geq \frac{h}{4\pi}$$

This relationship is called uncertainty principle.

15. Write down electronic configuration of Fe (26) and Br (35).

Fe (26)

$1s^2 2s^2 2p^6 3s^2 3p^6\ 3d^6\ 4s^2$

Br (35)

$1s^2 2s^2 2p^6 3s^2 3p^6\ 3d^{10}\ 4s^2\ 4p^5$

16. Write down electronic configuration of P (15) and Cu (29).

P (15)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Cu (29)

$1s^2 2s^2 2p^6 3s^2 3p^6\ 3d^{10}\ 4s^1$

17. Write down electronic configuration of Cr (24).

Cr (24)

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5\ 4s^1$

18. What is difference between continuous spectrum and line spectrum?

Continuous Spectrum	Line Spectrum
The Spectrum in which the boundary line between the colors cannot be marked and the colors diffuse into each other.	When an element or its compounds is volatized on a flames the light emitted is seen through a spectrometer we see distinct lines separated by dark spaces such a spectrum is called line spectrum
It is a characteristic of matter in bulk	It is the characteristic of an atom
E.g. Hydrogen Spectrum	E.g. Rainbow

19. Write two points in importance of Moseley's law.

Following are the two points in importance of Moseley's law:

- The atomic number of rare earths have been determined by this law.
- Mosely arranged Potassium (K), Argon (Ar), Nickel (Ni) and Cobalt (Co) in a proper way in Mendeleev's Periodic table

20. What are the defects (draw backs) in the Rutherford's atomic model?

Following are the defects of Rutherford Model.

- Rutherford's planet like picture was defective and unsatisfactory because the moving electron must be accelerated towards the nucleus.
- The radius of the orbiting electron should become smaller and smaller and the electron should fall into the nucleus. Thus, an atomic structure as proposed by Rutherford would collapse.
- If the electron radiates energy continuously, we should get continuous spectrum but, line spectrum was observed.

21. **Justify that distance gaps between different orbitals go on increasing from the lower to higher orbit.**

According to Bohr's Postulates, radius of revolving electron in nth orbital around the nucleus in hydrogen atom.

$$r = \frac{n^2h^2\varepsilon_\circ}{Ze^2\pi m}$$

We know that: $r=0.529 \quad x[n^2]$

When n=2 $r_2=0.529 \quad x4=2.11$

When n=3 $r_3=0.529 \quad x9=4.75$

When n=4 $r_4=0.529 \quad x16=8.4$

When n=5 $r_5=0.529 \quad x25=13.22$

Distance between orbits are:

$r_2 - r_1 = (2.11 - 0.529) = 1.581$

$r_3 - r_2 = (4.75 - 2.11) = 2.64$

$r_4 - r_3 = (8.4 - 4.75) = 3.65$

$r_5 - r_4 = (13.22 - 8.4) = 4.82$

From the data of radius difference, it is clear that the distance gaps between different orbits go on increasing from the lower to the higher orbits.

22. Define Zeeman's effect and Stark's effect. OR What is Zeeman's effect?

Zeeman's effect: The splitting of spectral lines in the presence of strong magnetic field is called Zeeman's effect.

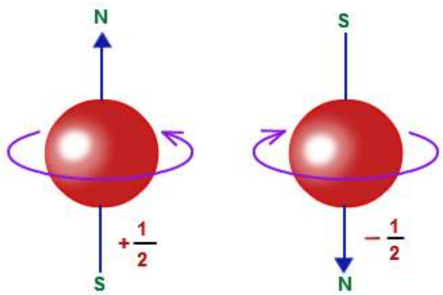
Stark's effect: The splitting of spectral lines in the presence of strong electric field is called Stark's effect

23. **Differentiate between line spectrum and continuous spectrum.**

Continuous Spectrum	Line Spectrum
The Spectrum in which the boundary line between the colors cannot be marked and the colors diffuse into each other.	When an element or its compound is volatized on a flame the light emitted is seen through a spectrometer we see distinct lines separated by dark spaces such a spectrum is called line spectrum
It is a characteristic of matter in bulk	It is the characteristic of an atom
E.g. Hydrogen Spectrum	E.g. Rainbow

24. State spin Quantum number briefly.

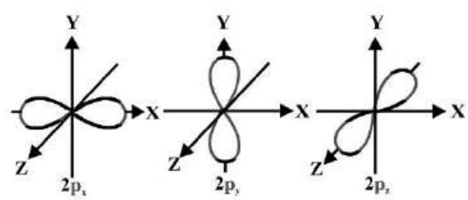
A quantum number that describes the spin about its own axis (also called self-rotation) either clockwise or anti-clockwise to satisfy the magnetic moment. It is represented by “s”. This spinning is responsible for duplet line structure in the spectrum.



25. What is orbital? Discuss the shape of p-orbitals.

The volume of space around the nucleus in which there is 95% chance of finding an electron is called orbital .It is called “atomic orbital” or “electron orbital”.

There are three values of magnetic quantum number for p-subshell. So, there are three orientations of P-subshell in space. All the 3p-orbitals namely p_x, p_y, p_z have dumb-bell shapes.



26. Calculate the number of electrons in s, p, d and f- sub shells from the formula and write separately.

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

For p-subshell = (2l +1) = 2(2(1) +1) = 2(2+1) = 2(3) = 6 electrons

For d-subshell= 2(2l +1) = 2(2(2) +1) = 2(4+1) = 2(5) = 10 electrons

For f-subshell= 2(2l +1) = 2(2(3) +1) = 2(6+1) = 2(7) = 14 electrons

27. **State Pauli’s exclusion principle.**

This principle can be stated as follow:

It is impossible for two electrons residing in the same orbital of a ploy-electron atom to have the same values of four quantum numbers. **Or**

Two electrons in the same orbital should have opposite spins ↑↓

28. State Pauli’s exclusion principle and Hund’s rule.

Pauli's exclusion principle

This principle can be stated as follow:

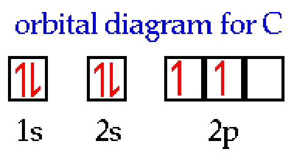
It is impossible for two electrons residing in the same orbital of a ploy-electron atom to have the same values of four quantum numbers. **Or**

Two electrons in the same orbital should have opposite spins

Hund's rule:

If degenerate orbitals are available and more than one electron are to be placed in them, they should be placed in separate orbitals with same spin rather than putting them in the orbital with opposite spin

For Example: The two electrons in 2p subshells of carbon will be distributed as follows:

$_6$C = 

29. What is Hund's rule?

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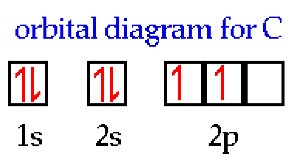
$_6$C = 

The three orbitals of 2p subshell are degenerate.

CHEMISTRY CHAPTER 6
(CHEMICAL BONDING)
Short Questions:

1. **What is octet rule? Give two examples of compounds which deviate from it.**

<u>Octet Rule:</u>
The tendency of an atom to attain a maximum of eight electrons in the valance shell is called octet rule”.
<u>Deviation from octet rule:</u>
There are many compounds in which atoms have no eight electrons (octet) in the valance shell after chemical combination. For example, BF_3, BCl_3, $AlCl_3$, SF_6, PF_5

2. **75.4 pm is the compromised distance between the bonded hydrogen atoms. Justify.**

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 75.4 pm, which is also called equilibrium bond distance or bond length of H_2 molecule.

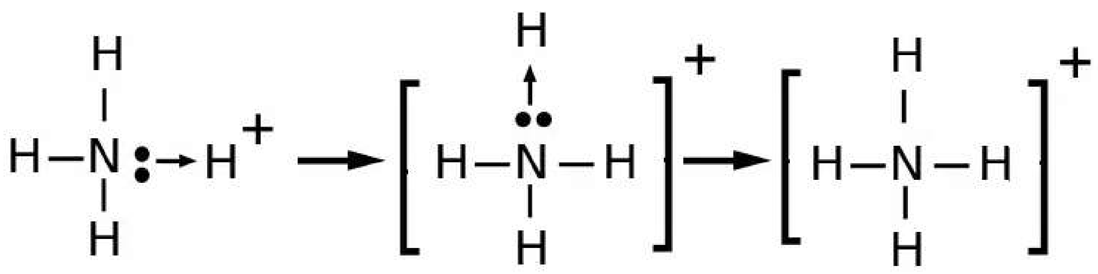
3. **Define ionic bond with an example.**

<u>Ionic bond:</u>
The electrostatic force of attraction between oppositely charged ions is called ionic bond.
<u>Examples:</u>
Na^+Cl^-, K^+Br^-

4. **Distinction between covalent and coordinate covalent bond vanishes after bond formation.**

<u>Justification:</u>
Coordinate covalent bond is just like a single bond. According to molecular orbital theory, this bond is formed by the empty orbital of H^+ and completely filled orbital of Nitrogen. According to Lewis concept, nitrogen donate 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.



Above structure gives impression that one N-H bond is different from other three N-H bonds. However, it has been shown experimentally that 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.

5. Why the radius of Cl^- ion increases from 99 pm to 181 pm.

OR

Why the ionic radius is greater than atomic radius?

<u>Ionic radius of anion:</u>
The radius of anion is larger than its original atom. 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.
<u>Reason:</u>
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.

6. State electronegativity and electron affinity.

<u>Electronegativity:</u>
The measure of the ability of an atom in a molecule to attract the shared pair of electron towards itself is called the electronegativity.
<u>Example:</u> Electronegativity of Fluorine is 4.
<u>Electron affinity:</u>
The amount of energy evolved and absorbed when an electron is added to partially filled orbital of an isolated atom to form a negative ion is called the electron affinity.

Example:

$$Cl_{(g)} + e^- \rightarrow Cl^-_{(g)} \quad \Delta E = -349\ kJ/mol$$

7. Why is the radius of a cation smaller than its parent atom?

Ionic radius of Cation:
Size of cation is smaller than from its parent atom because:

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 no. of shells also decreases

Example:
The radius of Na^+ ion is smaller than Na atom. The radius of 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.

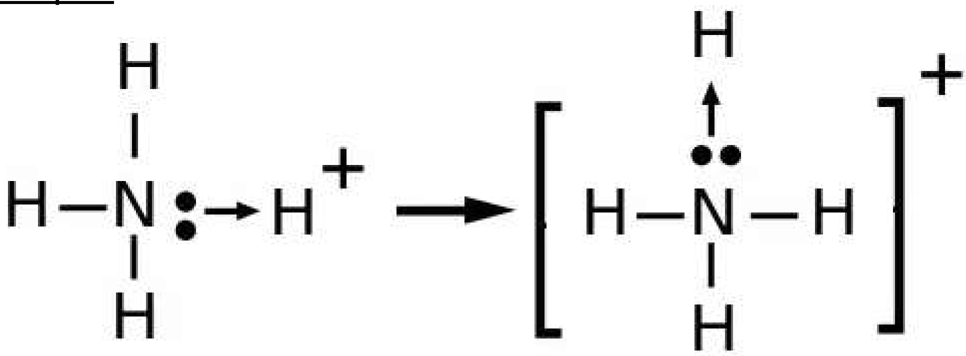
8. Differentiate between covalent bond and coordinate covalent bond.

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_4, H_2O	Example: NH_4^+, BF_4^-
Two atoms donate an equal share of electrons.	One atom called donor donates a pair of electron and one atom called acceptor accepts the pair of electron.
They can be polar or non-polar depending upon the electronegativity of atoms.	They are always polar.

9. Define coordinate covalent bond and give an example.

Coordinate covalent bond:
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.

Example:



In formation of NH_4^+ ion, NH_3 acts as electron pair donor and H^+ act as electron pair acceptor.

10. Why the energy of anti-bonding molecular orbital is higher than corresponding bonding molecular orbital?

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.
The destructive interference causes a node to form, which decreases the electron density between the two atoms. This raises both the kinetic and potential energy of any electrons in the resulting anti-bonding orbital.
Thus, when the electrons in an anti-bonding orbital spend less time between the two nuclei, they are at a higher energy level.

11. How does ionization energy vary in periodic table?

Variation of Ionization Energy in Periodic Table:
Along the Period:
In the period the ionization energy increases from left to right in a period, due to increase in proton number and decrease in atomic radius.
Down the group:
The ionization energy decreases from top to bottom in a group with the increase in atomic radius and increase in shielding effect.

12. Ionization energy is an index to the metallic nature of an element. Justify.

<u>Index to metallic character:</u>
The ionization energy is an index to metallic character, as metallic character is associated with tendency of atom to loose electrons. The elements which have

i. Low ionization energies are metals
ii. High ionization energies are non-metals
iii. Intermediate ionization energies are metalloids

13. Define Bond order. Calculate bond order of hydrogen molecule.

OR

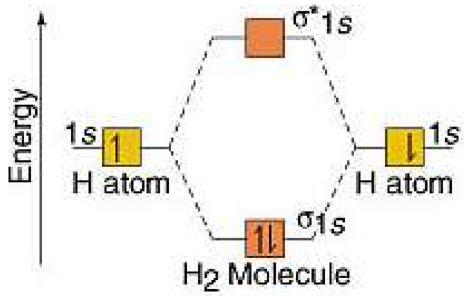
What is Bond order? Give an example.

<u>Bond Order:</u>
The no. of bonds formed between two atoms after the atomic orbitals overlap, is called the bond order.
OR
Half of the difference between the number of bonding electrons and anti-bonding electrons is called the bond order.

$$Bond\ order = \frac{No.of\ electrons\ in\ B.M.O - No.of\ electrons\ in\ A.B.M.O}{2}$$

Bond order of Hydrogen molecule:

$$Bond\ Order = \frac{2-0}{2} = 1$$

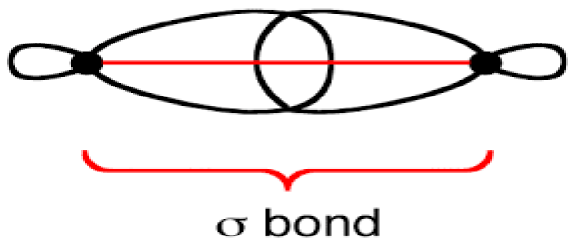


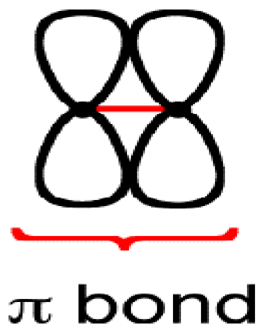
14. Differentiate between atomic orbital and molecular orbital.

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.
Single nucleus effects the electronic cloud.	More than one nuclei effects the electronic cloud.
Formed by the electronic cloud around the atom.	Formed by the fusion of atomic orbitals that have nearly the same energy.

15. π bonds are more diffused than Ϭ bonds. Why?

Sigma bond is formed by the overlapping of partially filled orbitals between two atoms. In sigma bonded atoms, the electron density is between two nuclei or at bond axis. In pi bond, electronic cloud is expanded at more region than sigma bond, so pi bond or pi electrons are more diffused than sigma electrons.





16. Define electronegativity. Give its trend in the periodic table.

Definition:
"The measure of the ability of an atom in a molecule to attract the shared pair of electron toward itself is called its electronegativity".
Trend of Electronegativity in Periodic Table:
The electronegativity values increases from left to right in the periodic table while it decreases from top to bottom in groups. Thus the most electronegative atoms are the non-metals at the right portion of periodic table.

17. How the nature of a chemical bond is predicted with the help of electronegativity values of two bonded atoms?

Elements which show greater electronegativity difference than 1.7 from ionic bonds. A difference of 1.7 units shows roughly equal contribution of ionic and covalent bonds.
The bond between Na (E.N = 0.9) and Cl (E.N = 3.0) in NaCl is ionic because E.N difference is 2.1. The bond in HCl shows covalent character because electronegativity is 0.9 (E.N of H = 2.1 and Cl = 3.0).

18. Why is no bond in chemistry 100% ionic?

Usually cations and anions of ionic bond are considered as hard spheres with some charge. But when cations and anions approaches to each other, distortion of electronic clouds of cations and anions take place. This process is called polarizability. Due to polarization of cations and anions, sharing of electrons also take place upto some extent. 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 have 8% covalent character in it. NaCl is 72% ionic and 28% covalent.

19. The bond angles of H_2O and NH_3 are not 109.5° like that of CH_4 although Oxygen and Nitrogen atoms are sp^3 hybridized. Why?

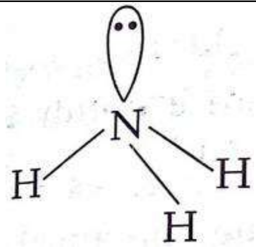
According to VSEPR theory, lone pairs occupy more space than bond pairs and causes more repulsions.

- H_2O has two lone pairs, so it repels the bond pairs much more and makes bond angle shorter of 104.5°.

- NH_3 has one lone pair that repels the three bond pair but not much effectively and strongly as two lone pairs of water repel the bond pair so the bond angle between hydrogen atoms of ammonia is 107.5° greater than that of water.

- CH_4 molecule also has sp^3 hybridization but it 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°.

20. State the geometry of ammonia molecule on the basis of VSEPR theory.

Total electron pairs	4	**Shape of NH_3 Molecule**
Bond pairs	3	
Lone pairs	1	
Arrangement of electron pairs	Tetrahedral	
Molecular geometry	Tetragonal pyramidal	
Bond angle	107.5°	

21. Define Dipole moment and give its S.I units.

<u>Definition:</u>
The product of electric charge (q) and the distance between the positive and negative centers (r) is called dipole moment.

$$\mu = q \times r$$

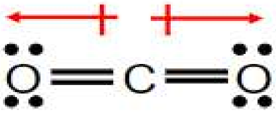
<u>Units:</u>
The common unit of dipole moment is Debye (D). The S.I units of dipole moment is coulomb meter (Cm). 1D = 3.336×10^{-30} Cm

22. Why the abnormality of bond length and bond strength in HI is less prominent than that of HCl.

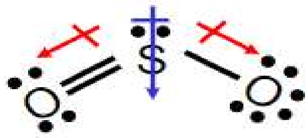
<u>Abnormality of bond length and bond strength:</u>
Electronegativity of Cl is more than that of iodine that is why the difference of electronegativity between the HCl is more than of Hl between the bonded atoms. The decrease in polarity from the HCl to Hl indicates the increase in trend of equal sharing 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 Hl.

23. Why the dipole moment of CO_2 is Zero but that of SO_2 is 1.61 D?

The dipole moment of CO_2 is zero, as it has a linear structure where the dipoles being equal and opposite, cancel out each other's effect.
SO_2 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.



$\mu = 0$



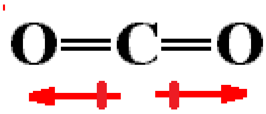
$\mu = 1.6$ D

24. Why the dipole moment of SO_2 is 1.61D but that of SO_3 is zero?

<u>Dipole moment of SO_2</u>
SO_2 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.
<u>Dipole moment of SO_3</u>
SO_3 has a dipole moment of 0. SO_3, 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 as the dipoles in the bond cancel each other out, therefore the molecule is non polar.

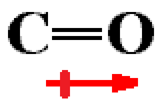
25. Why the dipole moment of CO_2 is Zero but that of CO is 0.12 D?

<u>Dipole moment of CO_2</u>
The dipole moment of CO_2 is zero, as it has a linear structure where the dipoles being equal and opposite, cancel out each other's effect.



$\mu = 0$

<u>Dipole moment of CO</u>
CO has a permanent pole, hence it shows a dipole moment of 0.12 D.



$\mu = 0.12$ D

26. Why BF_3 is non-polar but SO_2 polar?

BF3 has a symmetrical triangular planar molecule, having zero dipole moment so it is a non-polar molecule.
On the other hand, SO_2 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, and is polar in nature.

27. How the percentage of ionic character of covalent bond is determined by Dipole moment?

Percentage ionic character:
From the experimentally determined dipole moments, the percentage ionic character in a bond can be calculated.

$$\%\text{age ionic character} = \mu_{observed}/\mu_{ionic} \times 100$$

Example
The observed dipole moment of HF is 1.90D. The distance between the charges is 0.917×10^{-10} m. (unit positive charge = 1.6022×10^{-19} C).

$$\mu_{ionic} = q \times r$$
$$= 1.6022 \times 10^{-19}\ \text{C} \times 0.917 \times 10^{-10}\ \text{m}$$
$$= 1.469 \times 10^{-29}\ \text{Cm} = 4.4\ \text{D} \qquad (1\text{D} = 3.336 \times 10^{-30}\ \text{Cm})$$

$$\%\text{age ionic character} = \mu_{observed}/\mu_{ionic} \times 100$$
$$= 1.90\text{D}/4.4\text{D} \times 100$$
$$= 43.2\%$$

28. 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?

The melting points, boiling points, heats of sublimation and heats of vaporization of electrovalent compounds are usually higher as compared with those of covalent compounds. In ionic compound, strong interionic 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° while melting point of a covalent compound H_2O is 0°C.

29. Write two points of Valence bond theory.

According to valance bond theory,

i. The partially filled atomic orbitals overlap to form bonds but the individual character of atomic orbitals are retained.
ii. Greater the overlap, stronger will be the bond formed.

30. Why ionization energy decreases down the group although nuclear charge increases. Explain.

Ionization energy decreases down the group:

In groups the ionization energy decreases inspite of the increase in nuclear charge. This is due to successive addition of electronic shells as a result of which the valance electrons are placed at a larger distance from the nucleus.
As the force of attraction between the nucleus and the outer electron decreases with the increase in distance, the electrons can be removed more easily or with less energy.
Moreover, the force of attraction also decreases due to shielding effect of the intervening electrons.

31. Why it is impossible for CH_4 to make a coordinate covalent bond with H^+ ion while water and ammonia can do so?

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 act as the 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.

32. Why the lone pairs of electrons occupy more space than bond pairs?

Lone pairs occupy more space:
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.

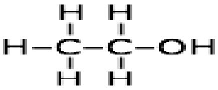
33. On what factors strength of bond depends?

The strength of bond depends on

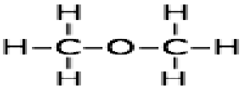
i. Electronegativity
ii. Size of atom
iii. Bond length

34. Why ionic compounds do not show the phenomenon of isomerism but covalent compounds do?

Since ionic bond is non-directional, ionic compounds do not exhibit the phenomenon of isomerism. Covalent bonds are rigid and directional. This leads to the possibility of a variety of isomerism. For example, structural isomerism is shown by the compound, C_2H_6O.



ethanol
C_2H_6O



dimethyl ether
C_2H_6O

35. How the type of bonding affects the solubility of compounds.

<u>Solubility of ionic compounds:</u>
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. Thus the ions are freed from the crystal lattice by hydration and salt is dissolved in water. Many ionic compounds do not dissolve in water as the attraction of the water molecules cannot overcome the attraction between the ions. For the same reasons, non-polar solvents like benzene and hexane do not dissolve ionic compounds.

<u>Solubility of covalent compounds:</u>
In general, 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. Mostly covalent compounds are insoluble in water. However, some of them like glucose, sucrose, urea etc. dissolve in water due to hydrogen bonding.

36. Why is it difficult to measure the correct atomic radius?

Atomic radius cannot measure precisely because

i. There is no sharp boundary of an atom, the probability of finding an electron never becomes exactly zero even at large distance from nucleus.
ii. The electronic probability distribution is affected by neighboring atoms.

Hence size of an atom is changed from compound to compound.

37. Differentiate between hybrid orbital and molecular orbital?

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.
Formed in the same atom	Formed between two atoms
e.g. sp^3 hybrid orbital	e.g. σ (1s)

Long Questions

1. Define bond energy. Discuss two facts which affect it.
2. Write the main postulates of VSEPR theory and explain the structure of Ammonia on the basis of this theory.
3. Define hybridization process and explain the structure of ethyne on the basis of it
4. Explain the structure of CH_4 on the basis of hybridization.
5. Explain the molecular orbital structure of following molecules on the basis of MOT. N_2 and O_2 molecule.
6. Describe the bonding in O_2 according to Molecular orbital theory and explain its paramagnetic property. Explain important points of Molecular orbital theory and draw structure of Nitrogen (N_2) molecule according to this theory.
7. Explain paramagnetic behaviour of O_2 on the basis of MOT and prove that MOT is superior to other theories. How does MOT explain the paramagnetic nature of O_2 molecule? Also calculate its bond order.
8. Define electron affinity. Name the factors affecting it. How does it vary in the periodic table?
9. Discuss the valence bond theory. How the sigma and pi bonds are formed by the overlapping of different orbitals?
10. Define dipole moment. Give its units. How is it used to determine the geometry of molecule? Give an example.

CHEMISTRY CHAPTER 7 (Thermochemistry)

Short Questions:

1. **What is thermochemical equation? Give two examples.**

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 thermochemical equation.

Examples: $C(s) + O_2(g) \rightarrow CO_2(g)$ ($\Delta H = -393.5\ KJ.mol^{-1}$)

$N_2(s) + O_2(g) \rightarrow 2NO(g)$ ($\Delta H = +180.5\ KJ.mol^{-1}$)

2. What is thermochemical equation? What information does it convey?

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 thermochemical equation.

Examples: $C(s) + O_2(g) \rightarrow CO_2(g)$ ($\Delta H = -393.5\ KJ.mol^{-1}$)

$N_2(s) + O_2(g) \rightarrow 2NO(g)$ ($\Delta H = +180.5\ KJ.mol^{-1}$)

The important information that can be drawn from thermochemical equations are:

- Nature of reaction is predicted, whether it is endothermic or exothermic.
- Exact amount of net heat released or absorbed in the reaction.

3. **What are thermochemical reactions? Give their types.**

The chemical reactions which involve heat changes are called as thermodynamic reactions. Actually 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. Why it is necessary to mention the physical states of reactants and products in the thermochemical equation?

To explain anomalies in enthalpies of formation of a compound, it is necessary to mention physical states of reactants and products.

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ ($\Delta H = -285.8\ KJ.mol^{-1}$)

$H_2 (g) + ½O_2 (g) \rightarrow H_2O (g)$ $(\Delta H = -241.8\ KJ.mol^{-1})$

The two values of formation of water are due to different physical states of water.

5. Burning of candle is spontaneous process. Justify it.

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 examples of 2nd type of reactions which are once started with some amount of energy but afterwards proceed and complete on their own.

6. Differentiate between endothermic and exothermic reactions.

Exothermic Reactions	Endothermic 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. Examples are as follows: $C(s) + O_2 (g) \rightarrow CO_2(g)$ $(\Delta H = -393.5\ KJ.mol^{-1})$ $H_2 (g) + ½O_2 (g) \rightarrow H_2O (l)$ $(\Delta H = -285.8\ KJ.mol^{-1})$	In this reaction reactants are at lower energy before reaction start and products obtained are at higher energy after reaction stop. The difference in the energy of products and reactants appears in the form of heat absorbed by the system from the surroundings. Examples are as follows: $N_2(s) + O_2 (g) \rightarrow 2NO(g)$ $(\Delta H = +180.5\ KJ.mol^{-1})$ $C(s) + 2S(s) \rightarrow CS_2(l)$ $(\Delta H = +92.0\ KJ.mol^{-1})$

7. What are endothermic and exothermic reactions? Give examples.

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.

$C(s) + O_2 (g) \rightarrow CO_2(g)$ $(\Delta H = -393.5\ KJ.mol^{-1})$

$H_2 (g) + ½O_2 (g) \rightarrow H_2O (l)$ $(\Delta H = -285.8\ KJ.mol^{-1})$

Endothermic Reactions

In this reaction reactants are at lower energy before reaction start and products obtained are at higher energy after reaction stop. The difference in the energy of products and reactants appears in the form of heat absorbed by the system from the surroundings.

$N_2(s) + O_2 (g) \rightarrow 2NO(g)$ $(\Delta H = +180.5\ KJ.mol^{-1})$

$C(s) + 2S(s) \rightarrow CS_2(l)$ $(\Delta H = +92.0\ KJ.mol^{-1})$

8. What is internal energy of a system?

The total energy of a system including all forms of kinetic and potential energies is known as internal energy**.**

Mathematically

$E = \Sigma$ kinetic energy $+ \Sigma$ potential energy

9. Differentiate between Law of conservation of energy and Hess's Law.

The law of conservation of energy is a physical law that states energy cannot be created or destroyed but may be changed from one form to another. Another way of stating this law of chemistry is to say the total energy of an isolated system remains constant or is conserved within a given frame of reference.

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.

10.Explain the term enthalpy of atomization.

Enthalpy change or amount of heat energy absorbed when an element is split into one mole atoms or a molecule is broken down to produce one mole of gaseous atoms under standard conditions is called standard enthalpy of atomization. It is represented by $\Delta H°_{at}$

Examples:

$\frac{1}{2} H_2 \rightarrow H \qquad \Delta H°_{at} = -218 \text{ KJ mol}^{-1}$

$\frac{1}{2} Cl_2 \rightarrow Cl \qquad \Delta H°_{at} = -121 \text{ KJ mol}^{-1}$

11.Define enthalpy of solution. Give an example.

The standard enthalpy of solution is the amount of heat energy evolved or absorbed when one mole of a substance is dissolved in so much solvent that further dilution result in no detectable heat change. It is represented by $\Delta H°_{sol}$

Examples:

$NH_4Cl\ (s) + H_2O \rightleftharpoons NH_4^+\ (aq) + Cl^-\ (aq) \qquad \Delta H°_{sol} = +16.2 \text{ KJ mol}^{-1}$

$Na_2CO_3\ (s) + H_2O \rightleftharpoons 2Na^+\ (aq) + CO_3^{-2}\ (aq) \qquad \Delta H°_{sol} = -25 \text{ KJ mol}^{-1}$

12.Define Enthalpy of solution and enthalpy of neutralization.

The standard enthalpy of solution is the amount of heat energy evolved or absorbed when one mole of a substance is dissolved in so much solvent that further dilution result in no detectable heat change. It is represented by ΔH°_{sol}

Examples:

$NH_4Cl\ (s) + H_2O \rightleftharpoons NH_4^+\ (aq) + Cl^-\ (aq) \qquad \Delta H^\circ_{sol} = +\ 16.2\ KJ\ mol^{-1}$

The standard enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions [H^+] from an acid, reacts with one mole of hydroxide ions [OH^-] from a base to form one mole of water. It is represented by $\Delta H^\circ_{n.}$

$H^+\ (aq) + Cl^-\ (aq) + Na^+\ (aq) + OH^-\ (aq) \rightleftharpoons Na^+\ (aq) + Cl^-\ (aq) + H_2O(l)$

$\Delta H^\circ_n = -57.4\ KJ\ mol^{-1}$

13.Define enthalpy of neutralization with an example.

The standard enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions [H^+] from an acid, reacts with one mole of hydroxide ions [OH^-] from a base to form one mole of water.

$H^+\ (aq) + Cl^-\ (aq) + Na^+\ (aq) + OH^-\ (aq) \rightleftharpoons Na^+\ (aq) + Cl^-\ (aq) + H_2O(l) \qquad \Delta H^\circ_n = -57.4\ KJ\ mol^{-1}$

14.Define standard enthalpy of formation and give two examples.

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 represented by ΔH°_f.

$Mg(s) + \frac{1}{2}O_2\ (g) \rightarrow MgO\ (s) \qquad (\Delta H^\circ_f = -692\ KJmol^{-1})$

$C(s) + O_2\ (g) \rightarrow CO_2\ (g) \qquad (\Delta H^\circ_f = -393.7\ KJmol^{-1})$

15.What is state function? Give two examples.

A state function is a macroscopic property of the system which has some definite initial and final values and difference can be directly calculated by subtractinginitial state from final value e.g., Enthalpy, Temperature, Presure and volume, Internal energy e.t.c.

$$\Delta T = T_2 - T_1$$

16.State the Hess's Law of constant heat summation.

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 CO2 is explained in direct and indirect ways below,

Direct route:

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ KJ.mol^{-1}

Indirect route:

$CO(g) + ½O_2(g) \rightarrow CO_2(g)$ $\Delta H_1 = -283.7$ KJ mol^{-1}

$C(g) + ½O_2(g) \rightarrow CO(g)$ $\Delta H_2 = -110$ KJ mol^{-1}

According to Hess's law

$\Delta H = \Delta H_1 + \Delta H_2$

$-393.7 = -283.7 + (-110)$

$-393.7 = -393.7$

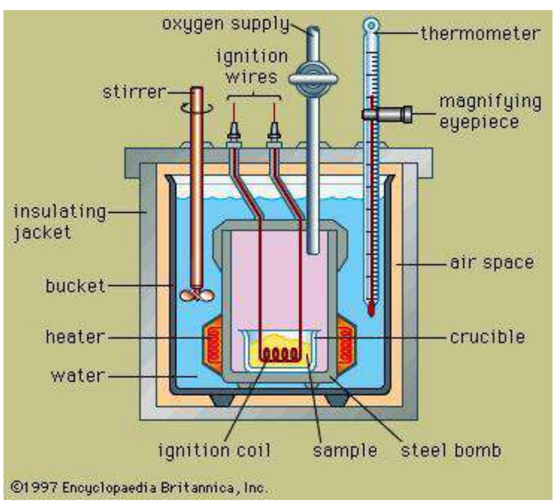
17.Is it true that ΔH and ΔE have the same values for the reaction taking place in solution state?

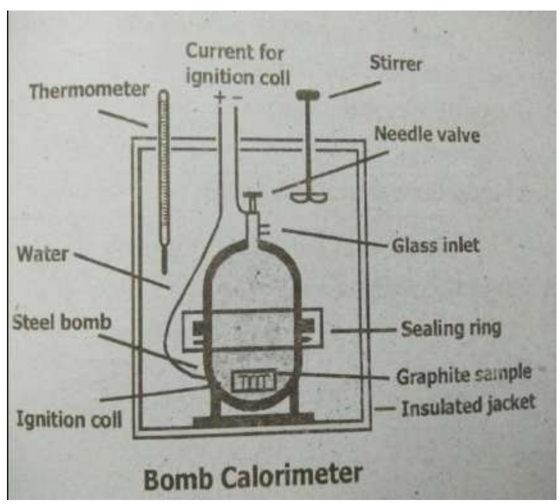
In case of solids and liquids the change in volume is negligibly small. Therefore, in equation $\Delta H = \Delta E + P\Delta V$ the factor becomes $\Delta V = 0$ and we are left with

$\Delta E \approx \Delta H$

Hence, it is true that ΔH and ΔE have the same values for the reactions taking place in solution state.

18.Draw a labeled diagram of Bomb Calorimeter.





19.What is spontaneous process? Give two examples.

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

   $H2O(l) \rightarrow H2O(g)$ $\Delta H = +44$ KJ mol^{-1}

2. Dilution of NH_4Cl in water

NH_4Cl (s) + H_2O $\rightleftharpoons$ NH_4^+ (aq) + Cl^- (aq) $\Delta H° = + 16.2$ KJ mol^{-1}

20.Prove that ΔE= qv

First Law of Thermodynamics states **ΔE = q + w** and is independent of path.

The internal energy (E) is a state function

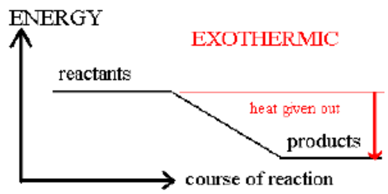
so **ΔE = 0** for a cyclic path

It is impossible to create a machine which, operating cyclically, will yield net work without aborption of heat.

For a constant volume process, **ΔE = qv**

21.Why heat energy is released in exothermic reactions?

In exothermic reactions the energy content of products is lesser than that of reactants. Therefore during reaction heat is released from system to surroundings.



CHEMISTRY CHAPTER 8 (CHEMICAL EQUILIBIUM)

Short Question Answers:

1. **Explain the term reversible reaction and state of equilibrium.**

   The reaction which proceeds in both the directions, forward as well as in reverse are called **reversible reactions.** Such reactions are represented by double headed arrow.

   Examples:

   $H_2(g) + I_2(g) \rightleftharpoons 2HI\ (g)$

   $N2O4\ (g) \rightleftharpoons 2NO_2\ (g)$

   The state at which rate of forward reactions becomes equal to the rate of reverse reaction is called state of equilibrium or chemical equilibrium.

2. **How the direction of a reversible reaction at any instant can be determined by Kc value? Mathematically**

   $K_c = \frac{[\text{Products}]}{[\text{Reactants}]}$ for any reaction

   The value of [products]/ [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.

3. **State Le-Chatelier's principle. And discuss the effect of change in concentration of a product on reversible reaction.**

   Le- Chatelier 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.

   Change in concentration bring two kinds of effects in a reversible reaction

   i. If a substance is added among the reactants, or the substance is removed from the products at equilibrium stage disturbs the equilibrium position and reaction is shifted to the **forward direction.**

   ii. If a substance is added among the products, or the substance is removed from the reactants at equilibrium stage disturbs the equilibrium position and reaction is shifted to the **backward direction.**

4. **How does change of pressure shifts the equilibrium position in the synthesis of ammonia?**

   During the synthesis of ammonia in Haber's process the increase in pressure, decrease the volume of the reaction vessel. Four moles of reactants combine to give two moles of the products. High pressure will shift the equilibrium position to the right to give more and more ammonia.

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

5. **The change of temperature disturbs both the equilibrium position and equilibrium constant of a reaction. Explain it.**
   According to le-Chatelier's principle, increase in temperature will favour the endothermic reactions and decrease in temperature will favour the exothermic reactions. Therefore, change of temperature will disturb equilibrium position.
   The equilibrium constant is temperature dependant, therefore, with the change of temperature, a new equilibrium position will be established and new value of equilibrium constant will be obtained.

6. **How the values of equilibrium constant help to predict the direction of a reversible reaction?**
   Mathematically

   $$K_c = \frac{[\text{Products}]}{[\text{Reactants}]} \qquad \text{for any reaction}$$

   The value of [products]/ [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.

7. **The solubility of glucose in water is increased by increasing the temperature. Explain.**
   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.

8. **Define pH and pOH. OR Define pOH of a solution. Give its mathematical equation.**
   pH is defined as the –log of H^+ ion concentration
   Mathematically

   $$pH = -\log [H^+]$$

   pOH is defined as the –log of OH^- ion concentration
   Mathematically

   $$pH = -\log [OH^-]$$

9. **Calculate the pH of 10^{-4} mol.dm^{-3} solution of HCl.**

Mathematically

$$\begin{aligned} pH &= -\log [H^+] \\ &= -\log [10^{-4}] \\ &= -(-4)\log [10] \qquad \text{as } \log[10] = 1 \\ &= 4 \end{aligned}$$

10. Calculate the pH of 10^{-4} mol.dm^{-3} solution of $Ba(OH)_2$.

$$Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH^-$$

$[OH^-] = 2\times10^{-4}$ mol.dm^{-3}

pH = $-\log 2\times10^{-4}$

11. How do the buffers act? Give example.

Buffer solution show almost constant pH value and are formed either by mixing weak acid and its salt with strong base e.g $CH_3COOH+CH_3COONa$ or by mixing weak base and its salt with strong acid e.g $NH_4OH + NH_4Cl$

Suppose we have acidic buffer solution of acetic acid and sod-acetate. See the ionization of both components

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

When an acid or H_3O^+ is added to this buffer the pH value remains unchanged due to the following reaction.

$$CH_3COONa + H_3O^+ \rightleftharpoons CH_3COOH + NaOH$$

All extra H_3O^+ ions are used tomake weak acid which does not affect buffer's pH.

Similarly when a base or OH^- ions added in it they will react with acid to give back H2O and anion and the pH value remain almost constant.

$$CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$$

12. How does the catalyst affect the equilibrium constant?

.A catalyst does not affect the equilibrium constant of the reaction.

13. How the buffer solutions are prepared?

There are two classes of buffer solutions i.e. acidic buffer and basic buffer

Preparation of acidic buffer:

It is prepared by mixing a weak acid with a salt which (i) produces strong base in water and (ii) provides a common ion.

Example: CH3COOH + CH3CONa

Preparation of basic buffer:

It is prepared by mixing a weak base with a salt which (i) produces strong acid in water and (ii) provides a common ion.

Example: $NH_4OH + NH_4Cl$

14. What do you mean by Buffer capacity?

The amount of external acid or base which a buffer can absorb without showing significant change in its pH is called as buffer capacity. Buffer capacity is the capability of a buffer to resist the changes in its pH. Buffer capacity depends upon the concentrations of its components.

15. Write two applications of equilibrium constant?

Equilibrium constant of reversible reaction is 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

16. Write two uses of buffer solutions.

i. Buffers are used as preservatives for food.
ii. They are used in pharmaceuticals.

17. Give two applications of solubility product.

Solubility product can be used to calculate:

i. Solubility of a compound from value of its solubility product.
ii. Solubility product of a compound from value of its solubility.

18. What is Handerson's equation?

Handerson's equation for acidic buffer is as follow:

$$pH = pKa + \log \frac{[salt]}{[acid]}$$

Handerson's equation for acidic buffer is as follow:

$$pH = pKa + \log \frac{[salt]}{[base]}$$

19. What are buffer solutions? How a basic buffer can be prepared?

The 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.

Preparation of basic buffer:

It is prepared by mixing a weak base with a salt which (i) produces strong acid in water and (ii) provides a common ion.

Example: $NH_4OH + NH_4Cl$

20. Define solubility product. Derive solubility product expression for Ag_2CrO_4?

The solubility product is the product of the concentrations of ions raised to exponent equal to the co-efficient of the balanced equation.

Solubility product expression for Ag_2CrO_4 $K_{sp}= [Ag^+][CrO_4^{-2}]$

21. Define solubility product. Derive solubility product expression for $PbCl_2$?

The solubility product is the product of the concentrations of ions raised to exponent equal to the co-efficient of the balanced equation.

Solubility product expression for $PbCl_2$ $K_{sp}= [Pb^{2+}][Cl^-]^2$

22. How change in volume disturbs the equilibrium position for some of the gas phase reactions but not the equilibrium constant?

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 Kc ultimately remains constant because it is only permanently affected by temperature change only.

23. How does a catalyst affect a reversible reaction?

A catalyst does not affect the equilibrium position of the reaction. It increases the rates of both forward and backward reactions and this reduces the time to attain the state of equilibrium.

24. What is the formula to calculate the percentage ionization of weak acids?

The percentage ionization of weak acid depends upon the extent of dilution of their aqueous solutions. The formula to calculate the percentage ionization of weak acids is as follows:

$$\%\ \text{ionization} = \frac{\text{Amount of acid ionized}}{\text{Amount of acid initially available}}$$

25. Define Lowry-Bronsted concept of acids and bases?

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.

$$\underset{\text{Aicd}}{HA} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\substack{\text{conjugate acid}\\ \text{Of } H_2O}}{H_3O^+} + \underset{\substack{\text{conjugate base}\\ \text{of HA}}}{A}$$

$$\underset{\text{Base}}{A^-} + \underset{\text{acid}}{H_2O} \rightleftharpoons \underset{\text{acid}}{HA} + \underset{\text{base}}{OH^-}$$

26. Prove by equations that what happens when Na_2CrO_4 is added to saturated solution of $PbCrO_4$.

$$PbCrO_4\,(aq) \rightleftharpoons Pb^{2+}(aq) + CrO_4^{2-}\,(aq)$$

$$Na_2CrO_4 \rightleftharpoons Na^{+1} + CrO_4^{2-}$$

The presence of a common ion decreases the solubility of a slightly soluble ionic compound. CrO_4^{2-} is a 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.

27. Why solid ice at 0°C can be melted by applying pressure without supply of heat from outside.

When pressure is applied to the broken pieces of ice 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.

28. Write the relationship of Kp with Kc.

Kc and Kp both are equilibrium constants, Kc is in terms of concentrations of reactants and products in mol. dm^{-3} and Kp is in terms of partial pressures in atm or torr e.t.c.

They are related as follows

$K_p = K_c\,(RT)^{\Delta n}$

Where R= gas constant T= absolute temperature and n is difference in total moles of products and reactants.

29. Define Chemical equilibrium. Give its any two properties.

If a reversible reaction is allowed to continue for a considerable long time without changing the conditions, there is no further change in composition of the reaction mixture. The reaction is said to have attained a state of chemical equilibrium.

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.

Chemistry Chapter 9

Solutions

1. **Define molarity and molality.**

Molarity (M)

It is the number of moles of solute dissolved per dm^3 of solution.

Molarity (M) = Mass of solute/Molar Mass of solute × 1/volume of soln. in dm^3

Example

18 g of glucose in 1000 mL of solution is 0.1 molar.

Molality (m)

It is the number of moles of solute in 1000 grams of the solvent.

Molality (m) = Mass of solute/Molar Mass of solute × 1/volume of solvent in kg

Example

For one molal sucrose solution 342 g of sucrose are dissolved in 1000g of H_2O.

2. **What is molarity? Calculate the molarity of a solution containing 9g of glucose in 250 cm^3 of solution.**

Molarity (M)

It is the number of moles of solute dissolved per dm^3 of solution.

Molarity (M) = Mass of solute/Molar Mass of solute × 1/volume of soln. in dm^3

Example

18 g of glucose in 1000 mL of solution is 0.1 molar.

Molarity Calculation

1M solution of glucose contains 180 g of glucose in 1 L water

250 cm^3 of solution= 250/1000=0.25 dm^3

$$\text{Molarity (M)} = \text{Mass/Molar Mass} \times 1/\text{volume of soln. in } dm^3$$

$$= 9/180 \times 1/0.25$$

$$= 0.2\ M = 2\times10^{-1} M$$

3. **How molality is independent of temperature but molarity depends on temperature?**

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.

4. **One molal solution of glucose is dilute as compared to one molar solution of glucose. Justify it.**

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^3. As 1000g (1 kg) of solvent is greater than 1000 mL of solvent so molar solution is concentrated and molal solution is dilute.

5. **One molal solution of urea is dilute as compared to one molar solution of urea. Justify it.**

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^3. 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.

6. **Define Ebullioscopic constant with example.**

The elevation in boiling point when 1 mole of solute is dissolved to the one kilogram of solvent is called molal boiling point constant or ebullioscopic constant.

Relation

$$\triangle T_b = K_b m$$

When m=1,

$\triangle T_b = K_b$

Unit

The unit of K_b is $^oC\ Kgmol^{-1}$

Example

Dissolve 6g of urea in 500g of H_2O or 18g of glucose in 500g of H_2O both give 0.2 molal solution and both have same elevation of boiling points i.e. 0.1^oC.

7. **Relative lowering of vapor pressure is independent of temperature. Justify it.**

The relative lowering of vapour pressure and mole fraction of solute are related as:

$$\frac{\triangle p}{po} = x_2$$

Vapour pressure and lowering of vapour pressure depend upon temperature. So, when the temperature of a solution is increased both the factors $\triangle P$ and P^o increase in such a way that the ratio remains the same.

8. **Justify that boiling point of solvents increase due to presence of non-volatile solutes.**

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.

9. **Depression of freezing point is a colligative property. Justify it.**

Depression of freezing point is a colligative property as colligative properties depend upon number of particles. This can be further explained with the help of an example. 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. Pure water has certain value of vapour pressure at a given temperature. In these three solutions, the vapour pressures will be lowered. The reason is that the molecules of a solute present upon the surface of a solution decrease the evaporating capability. Apparently it seems that sucrose solution should show the maximum lowering of vapour pressure while urea should have the minimum lowering of vapour pressure. The reality is that the lowering of vapour pressure in all these solutions will be same at a given temperature. Actually the number of particles of the solute in the solutions is equal. We have added $1/10^{th}$ 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 $0.186^{o}C$.

10. What is meant by molality? Give its formula.

Molality (m)

It is the number of moles of solute in 1000 grams of the solvent.

Molality (m) = Mass of solute/Molar Mass of solute × 1/volume of solvent in kg

Example

For one molar sucrose solution 342 g of sucrose are dissolved in 1000g of H_2O.

11. Why NaCl and KNO_3 are used to lower the melting points of ice?

NaCl and KNO_3 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.

12. Why Beckman’s thermometer is used to find the depression in freezing point?

Beckmann thermometer can measure up to $1/20^{th}$ of the degree. The elevation of boiling points and the depressions of freezing points for dilute solutions are very small quantities. Hence, one can measure these very small changes of temperatures.

13. Differentiate between ideal and non-ideal solutions.

An ideal solution obeys Raoult’s law at all concentrations and at all temperatures, whereas, a non-ideal solution does not obey Raoult’s law. In an ideal solution solute-solute interaction, solvent-solvent interaction and solute-solvent interactions are same. In a non-ideal solution the force of attraction between solute and solvent molecules are changed after making the solution. So they show positive and negative deviations from Raoult’s law. In an ideal solution, the total volume of solution is equal to the sum of volumes of all components. In a non-ideal solution, the total volume of solution is not equal to the sum of volumes of all the individual components. In an ideal solution, when the components of solution are mixed, there is no evolution or absorption of heat. In a non-ideal solution, when the components of solution are mixed there is evolution or absorption of heat. Example of ideal solution is benzene-toluene and that of non-ideal solution is acetone-methanol.

14. Why is the vapor pressure of a solution lesser than vapor pressure of pure solvent?

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.

15. Calculate the percentage by weight of NaCl if 2g of it is dissolved in 20g of water.

$$\% \text{ by weight} = \text{Mass of Solute/Mass of Solution} \times 100$$

$$\% \text{ by weight} = 2/20 \times 100$$

$$= 10\%$$

16. What is meant by water of crystallization? Give an example.

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

$(COOH)_2.2H_2O$

$BaCl_2.2H_2O$

$Na_2CO_3.10H_2O$

17. Define Zeotropic mixtures. Give one example.

Such liquid mixtures which distil with a change in composition are called zeotropic mixtures.

Example

methyl alcohol-water solution.

18. What are zeotropic and azeotropic mixtures?

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.

19. Differentiate between molarity and molality.

Molarity (M)

It is the number of moles of solute dissolved per dm^3 of solution.

Molarity (M) = Mass of solute/Molar Mass of solute × 1/volume of soln. in dm^3

Example

18 g of glucose in 1000 mL of solution is 0.1 molar.

Molality (m)

It is the number of moles of solute in 1000 grams of the solvent.

Molality (m) = Mass of solute/Molar Mass of solute × 1/volume of solvent in kg

Example

For one molal sucrose solution 342 g of sucrose are dissolved in 1000g of H_2O.

20. Define upper consulate temperature. Give two examples. Define hydrolysis with example.

Upper Consulate Temperature

Partially miscible liquids make two distinct layers of conjugate solutions. By changing the temperature, the two layers become completely miscible and homogeneous solution is produced. This temperature is called consulate temperature.

Examples

phenol-water system

methanol-cyclohexane system.

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 are equal to 10^{-7} M.

21. Differentiate between hydration and hydrolysis.

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.

22. What are the names of four major parts of apparatus used in Landsberger's method for elevation of boiling point?

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.

23. Why the solubility of glucose into water increases by increasing temperature?

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.

24. Give two statements of Raoult's law.

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

$$P \propto X_1$$

$$P = P^o X_1 \text{----(1)}$$

Where P^o 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

$$\triangle P = P^o X_2$$

$$\frac{\triangle P}{P^o} = X_2$$

$\frac{\triangle P}{P^o}$ is called relative lowering of vapour pressure and it is more important than lowering of vapour pressure.

25. What is fractional crystallization?

The separation of crystalline substances from a mixture on the basis of their different solubilities. In this method, the impure solute is dissolved in a hot solvent in which it is less soluble than the impurities. When this hot solution is cooled the solute being less soluble separates out first from the mixture while impurities remain in the mother liquor. In this way, pure desired product crystallizes out from the solution.

26. Aqueous solution of $CuSO_4$ is acidic in nature. Justify it.

$$CuSO_4 + 2HOH \longrightarrow Cu(OH)_2 + 2H^+ + SO_4^{-2}$$

As sulphuric acid is a strong acid and copper hydroxide is a weak base so the solution is overall acidic in nature.

27. Aqueous solution of CH_3COONa is basic in nature.

$$CH_3COONa + HOH \longrightarrow CH_3COOH + NaOH$$

As acetic acid is a weak acid and sodium hydroxide is a strong base so the solution is overall basic in nature.

Chemistry (Part 1)

Chapter 10. Electrochemistry

Short Questions

Q1.What is Electrochemistry?

Electrochemistry is concerned with the conversion of electrical energy in to chemical energy in electrolytic cells as well as the conversion of chemical energy in to electrical energy in Galvanic or Voltaic cells

Q2.What is the difference between metallic conduction and electrolytic conduction?

Metallic conduction	Electrolytic conduction
1. It takes place due to free electrons.	1. It takes place due to movement of ions.
2. Conductance decreases with increase in temperature.	2. Conductance increases with increase in temperature.
3. No chemical reaction takes place during conduction.	3. Redox reactions take place during conduction.
4. Chemical composition is not changed during conduction, so no new substance is produced.	4. Chemical reactions occur so, new products are produced.
5. Example: All metals are conductors.	5. Example: Molten salt eg. NaCl.

Q3.Differentiate between electrolytic and Galvanic cell.

Electrolytic Cell	Voltaic Cell
1. The electrochemical cell in which electrical energy is converted in to chemical energy is called electrolytic cell.	1.The electrochemical cell in which chemical energy is converted in to electrical energy is called Voltaic cell.
2.In this cell, electric current is used to drive a non spontaneous reaction.	2.In this cell, electric current is produced due to spontaneous reaction.
3. Electrolysis takes place in this cell.	3. Electrical conduction takes place in it.
4. Examples are Down’s cell, Nelson’s cell	4. Examples are Daniel’s cell, Fuel cell

Q4. Explain how impure Copper can be purified by electrolytic process.

In electrolytic cell anode is made of impure copper and cathode is made of pure copper. Aqueous solution of $CuSO_4$ is used as an electrolyte. The atoms of Cu from impure Cu anode are converted to Cu+2 ions, which go to cathode and discharge by accepting electrons and deposit on it as pure copper. Impurities are left at anode.

Q5. The standard oxidation potential of Zinc is 0.76V and its reduction potential is -0.76. Why?

When Zn electrode is connected to S.H.E, Zn gives electrons to Hydrogen electrode and gets oxidized: $Zn \rightarrow Zn^{+2} + 2e^-$ $\qquad E^o = +0.76$

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.

Q6. Give two applications of electrochemical series.

Electrochemical Series:

When elements are arranged in the order of their standard electrode potentials on the Hydrogen scale, the resulting list is called Electrochemical series.

Applications:

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 Cu^{+2} ion can oxidize solid Zinc but Zn^{+2} ion cannot oxidize solid copper. Standard reduction potential values of Copper and Zinc are shown below.

$$Cu^{+2} + 2e^- \rightarrow Cu \qquad E^o = +0.34V$$

$$Zn^{+2} + 2e^- \rightarrow Zn \qquad E^o = -0.76V$$

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 $CuSO_4$ but Zn does not displace Mg from solution of $MgSO_4$.

Q7. A salt bridge maintains the electrical neutrality in the cell. Explain.

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.

Q8. Write down the function of the salt bridge.

Same as above.

Q9. A porous plate or a salt bridge is not required in Lead acid storage battery.

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% H2SO4 solution/ d= 1.25gcm-3).Salt bridge usually separates the two electrodes in which different electrolytes are used.

Q10. Define Electrochemical Series.

When elements are arranged in the order of their standard electrode potentials on the hydrogen scale, the resulting list is called Electrochemical series.

Q11. Write down two functions of salt bridge in a Galvanic cell.

a. Salt bridge brings about the transference of ions between the two electrodes to prevent the accumulation of ions in either of the two half cells.

b. Salt bridge maintains the electrical neutrality of the Galvanic cell.

Q12. Write down reactions taking place at the electrodes during discharging of Nickel Cadmium cell.

<u>At Anode:</u>

$Cd + 2OH^- \rightarrow Cd(OH)_2 + 2e-$ (oxidation)

<u>At Cathode:</u>

$NiO_2 + 2H_2O + 2e- \rightarrow Ni(OH)_2 + 2OH^-$ (reduction)

<u>Net reaction:</u>

$Cd + NiO_2 + 2H_2O \rightarrow Cd(OH)_2 + Ni(OH)_2$

Q13. What is Standard Electrode Potential?

The potential set up when electrode is in contact with one molar solution of its own ions at 298K is called Standard electrode potential eg. standard electrode potential of Zn is 0.76V.

Q14. Give chemical reactions taking place at anode and cathode of Fuel cell.

At Anode:

$[H_2 + 2OH^- \rightarrow 2H_2O + 2e^-]2$

At Cathode:

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Overall reaction:

$2H_2 + O_2 \rightarrow 2H_2O$

Q15. Calculate the oxidation no. of Mn in $KMnO_4$.

$KMnO_4$

$+1 + Mn + (-2)4 = 0$

$Mn + 1 - 8 = 0$

$Mn = +7$

Q16. Calculate the oxidation no. of Mn in Na_2MnO_4.

Na_2MnO_4

$(+1)2 + Mn + (-2)4 = 0$

$+2 + Mn - 8 = 0$

$Mn - 6 = 0$

$Mn = +6$

Q17. Calculate the oxidation no. of S in $Cr_2(SO_4)_3$ and SO_4^{-2}.

$Cr_2(SO_4)_3$

$(+3)2 + (S)3 + (-2)12 = 0$

$+6 + (S)3 - 24 = 0$

$(S)3 - 18 = 0$

$S = +6$

$\underline{SO_4^{-2}}$

$S + (-2)4 = -2$

$S - 8 = -2$

$S = -2 + 8$

$S = +6$

Q18. Calculate the oxidation no. of Cr in $CrCl_3$.

$\underline{CrCl_3}$

$Cr + (-1)3 = 0$

$Cr - 3 = 0$

$Cr = +3$

Q19. Calculate the oxidation no. of P in HPO_3.

$\underline{HPO_3}$

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

$+1 + P - 6 = 0$

$P - 5 = 0$

$P = +5$

Q20. Calculate the oxidation no. of the elements underlined in the following compounds.

$K_2\underline{Mn}O_4$

$(+1)2 + Mn + (-2)4 = 0$

$+2 + Mn - 8 = 0$

$Mn - 6 = 0$

$Mn = +6$

$Ca(\underline{Cl}O_3)_2$

$+2 + (Cl)2 + (-2)6 = 0$

$+2 + (Cl)2 - 12 = 0$

$(Cl)2 - 12 + 2 = 0$

$(Cl)2 - 10 = 0$

$Cl = +5$

Q21. What is the difference between Electrolytic cell and Voltaic cell?

Same as question no. 3

Q22. Voltaic cell is reversible cell .State.

Some Voltaic cells are reversible in which electrode reactions can be reversed by using an external battery. Such cells are rechargeable and are also called secondary cells. eg. lead storage battery.

Q23. How Fuel cells produce electricity?

In fuel cells H_2 is oxidized and 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.

<u>At Anode:</u>

$2H_2 + 4OH^{-1} \rightarrow 4H_2O + 4e^-$

<u>At Cathode:</u>

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

<u>Overall reaction:</u>

$2H_2 + O_2 \rightarrow 2H_2O$

Q24. Write two advantages of Fuel cells.

<u>Advantages of fuel cells are as follows.</u>

a. Fuel cells are light, portable and produce electricity and pure water during space flights.

b. These are environment friendly, efficient and convert about 75% fuel bond energy into electrical energy.

Q25. Give the chemistry of electrolysis of aqueous solution of Sodium Chloride.

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.

$NaCl \leftrightarrow Na^+ + Cl^-$

<u>At Anode:</u>

$2Cl^{-1} \rightarrow Cl_2 + 2e^-$ (oxidation)

<u>At Cathode:</u>

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (reduction)

<u>Overall reaction: (combining Na^+ ions)</u>

$2Na^+ + 2Cl^- + 2H_2O \rightarrow Cl_2 + H_2 + 2Na^+ + 2OH^-$

Q26. What is electrolysis? Give example.

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.

<u>Example: Electrolysis of Fused salts</u>

When a fused salt is electrolyzed the metal ions called cations move to cathode and get discharged by picking up electrons hence undergoing redction. The anions move towards anode and also get discharged by losing electrons hence undergoing oxidation.

For example in case of fused lead chloride, the equations for electrode half reactions are as follows:

At Cathode: $Pb^{+2} + 2e^- \rightarrow Pb$ (reduction)

At Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$ (oxidation)

Q27. Write recharging of lead accumulator battery.

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: $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{-2}$ (reduction)

At Cathode: $PbSO_4 + 2H_2O \rightarrow PbO_2 + 4H^+ + SO_4^{-2} + 2e^-$ (oxidation)

Overall reaction: $2PbSO_4 + 2H_2O \rightarrow Pb + PbO_2 + 4H^+ + 2SO_4^{-2}$

Both the density of the acid and voltage of the battery are restored.

Q28. Lead accumulator is a chargeable battery. Justify.

Same as above

Q29. SHE acts as anode when connected with Cu but act as cathode when connected with Zn. Justify your answer with equations.

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.

Anode: $H_2 \rightarrow 2H^+ + 2e^-$ (oxidation)

Cathode: $Cu^{+2} + 2e^- \rightarrow Cu$ (reduction)

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.

Anode: $Zn \rightarrow Zn^{+2} + 2e^-$ (oxidation)

Cathode: $2H^+ + 2e^- \rightarrow H_2$ (reduction)

Q30. What is Alkaline battery?

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 which is more than that of a comman dry cell. The electrode reactions are as follows:

Anode: $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$

Cathode: $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$

Overall reaction: $Zn + 2MnO_2 + H_2O \rightarrow Zn(OH)_2 + Mn_2O_3$

Q31. Write down electrode reactions of Dry cell.

Electrode reactions are same as above.

Q32. Differentiate between cathode and anode.

Cathode	Anode
1. It is the negative electrode.	1. It is positive electrode.
2. It attracts positive ions and brings about their reduction	2. It attracts negative ions and carry out their oxidation.

Q33. How is Aluminium anodized ?

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.

CHEMISTRY CHAPTER 11

(Reaction Kinetics)

Short Questions:

1. **Define rate of chemical reaction and give its units.**

Ans: Rate of Chemical Reaction: - 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.
Units: -

$$\text{Rate of reaction} = \frac{\text{Mol dm}^{-3}}{\text{seconds}} = \text{mold m}^{-3}\ \text{sec}^{-1}$$

2. **Define specific rate constant. Give equation to support your answer.**

Ans: - The rate constant is called as specific rate constant or velocity constant when concentrations of reactants are unity.

OR

Specific rate constant is the rate of reaction when the concentrations of reactants are unity.

$aA + bB \rightarrow cC + dD$

Rate = $k\,[A]^a\,[B]^b$

As $[A] = [B] = 1\ \text{mol dm}^{-3}$

Rate = k (Specific rate constant)

3. **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 the rate of reaction is an ever changing parameter.

4. **Define order of reaction with the help of an example.**

Ans: - Order of Reaction
Ii is the sum of the exponents of concentration of the reactants on which rate of reaction actually depends.
OR
The order of reaction is given by the sum of all the exponents to which the concentrations in the rate equation are changed.
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.

Example
For the reaction $aA + bB \rightarrow cC + dD$
Rate = $k [A]^a [B]^b$
Order = n = a +b

5. **The radioactive decay is always a first order reaction. Give reason.**
Ans: - Radioactive decay means emission of radioactive rays (alpha, beta and gamma) by a substance. In this process only radioactive substance acts a reactant, and its concentration doesn't effect on the rate of its decay. So it will be a first order reaction.
It is first order reaction. Half-life is independent of initial concentration of reactant.

$$[t_{1/2}] \ \alpha \ 1/a^{n-1}$$

For first order n=1, then

$$[t_{1/2}] \ \alpha \ 1/a^{o}$$

6. **Define with example 2nd 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 nitric oxide with ozone has shown to be first oedrer with respect to NO and first orfer with reaspect to O_3. The sum of the individual orders gives the overall order of reaction as two.
$NO + O_3 \rightarrow NO_2 + O_2$
Rate = $k [NO] [O_3]$

7. **What is specific rate constant or velocity constant?**

Ans: - The rate constant is called as specific rate constant or velocity constant when concentrations of reactants are unity.
OR
Specific rate constant is the rate of reaction when the concentrations of reactants are unity.
$aA + bB \rightarrow cC + dD$
Rate = $k [A]^a [B]^b$
As $[A] = [B] = 1 \text{ mol dm}^{-3}$
Rate = k (Specific rate constant)

8. **What is half-life period? Give example.**
Ans:- Half-life period: - Half-life period of a reaction is the time in which 50% of initial concentration of reactants is converted into products.
For example,
Half-life of decomposition of N_2O_5 at 45°C is 24 minutes. It means if we take 0.1 moldm^{-3} N_2O_5, after 24 minutes 0.05 mol dm^{-3} of N_2O_5 will decompose to products.
9. **How surface area affects the rate of reaction? Give one example.**

Ans: Effect of surface area: -The increased suffice 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.

$$2Al + 2\,NaOH + 6H_2O \rightarrow 2NaAl(OH)_4 + 3H_2$$

10. **Define activation energy and activated complex.**

    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.

    Activated Complex: - Activated complex is an unstable combination of all the atoms involved in the reaction for which the energy id 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.

11. **What do you mean by activation energy of a reaction?**

    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.

12. **How does a catalyst affect a reversible reaction?**

    Ans: A catalyst cannot affect the equilibrium constant of a reaction but it helps the equilibrium to be slashed earlier. The rates of forward and backward steps are increases equally.

13. **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. $25^{o}C$ (298 K) and one atmosphere pressure.

14. **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 promotors 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.

15. **Define homogeneous catalysis. Give two examples.**

    Ans:- Homogeneous catalysis: - 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.

    Example 1

    The formation of SO_3 from SO_2 and O_2 in the lead chamber process for the manufacture of suphuric acid, needs NO as a catalyst. Both the reactants and the catalyst are gases.

$$2SO_{2(g)} + O_{2(g)} \xrightleftharpoons{NO_{(g)}} 2SO_{3(g)}$$

Example 2

Esters are hydrolyzed in the presence of H2SO4, both the reactants and the catalyst are in the solution state.

$$CH_3COOC_2H_{5(aq)} + H2O_{(l)} \xrightleftharpoons[H_2SO_4]{H_3O+} CH_3COOH_{(aq)} + C_2H_5OH_{(aq)}$$

16. What is catalytic poisoning? Give two examples.

Ans: Catalytic poisoning: Catalytic poisoning happens due to presence of trace amounts of foreign substances which render them ineffective. Such substances are called poisons.

Example 1

The presence of Co as an impurity with hydrogen decreases the catalytic activity of catalyst in the Haber's process for the manufacture of NH_3.

Example 2

The manufacture of H_2SO_4 in the contact process needs platinum as catalyst. The traces of arsenic present as impurities in the reacting gases makes platinum ineffective.

17. 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 of contain some non-protein part also for their help.

Enzyme Aaction:- Enzyme increase the rate of reaction within the body of living organisms by decreasing the activation energy of reaction.

For example, Urease catalyses the hydrolysis of urea and Lipase catalyses the hydrolysis of fats.

18. Write down any two characteristics of enzyme catalysis.

Ans:- Characteristics of Enzymes

1. Enzymes are the most efficient catalysts known and they lower the energy of activation of a reaction.
2. Enzymes catalytic reactions have the maximum rates at an optimum temperature and optimum pH.
3.

19. Enzymes are specific in action. Justify.

Ans: Specific action of enzymes:- 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.

20. What is auto catalyst? Give an example.

Ans:- Autocatalysis: - In some reactions, the product formed acts as a catalyst and this phenomenon is called autocatalysis.

For example, the reaction oxalic acid with acidified $KMnO_4$ is slow in the beginning, it becomes fast with with the passage of time due to the formation of Mn^{2+} ion which act as auto-catalyst.

$$2KMnO_4 + 3H_2SO_4 + 5\,H_2C_2O_4 \xrightarrow{Mn^{2+}} K_2SO_4 + 2MnSO_4 + 8H_2O + 10\,CO_2$$

21. What is auto catalysis? Give example to support answer.

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