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Chapter 1: Basic Concepts

Q 1: The diameter of atoms are of the order:
Ⓐ 2 m Ⓑ 0.2 nm Ⓒ 0.2 m Ⓓ 0.2 um

Q 2: 1a.m.u is equal to:
Ⓐ 1.661×10^{-27}kg Ⓑ 1.661×10^{-24} g Ⓒ 1.661×10^{-21}mg Ⓓ All of these

Q 3: Masses of atoms ranges from:
Ⓐ 10^{-27} - 10^{-25} kg Ⓑ 10^{-24}g-10^{-22}g Ⓒ 10^{-21}mg-10^{-19}mg Ⓓ All of these

Q 4: Nickel has isotopes:
Ⓐ 3 Ⓑ 5 Ⓒ 9 Ⓓ 11

Q 5: Total No. of naturally occurring isotopes is:
Ⓐ 240 Ⓑ 40 Ⓒ 280 Ⓓ 154

Q 6: 27g of Al will react completely with how much mass of O_2 to produce Al_2O_7
Ⓐ 8g Ⓑ 16 g Ⓒ 32g Ⓓ 24g

Q 7: In combustion analysis H_2O vapors are absorbed by:
Ⓐ Mg $(ClO_2)_2$ Ⓑ Mg $(ClO_3)_2$ Ⓒ Mg $(ClO_4)_2$ Ⓓ 50% KOH

Q 8: The number of CO_2 which contains 8.0 g Oxygen:
Ⓐ 0.25 Ⓑ 0.50 Ⓒ 0.75 Ⓓ 1.0

Q 9: Largest number of molecules is in:
Ⓐ 3.6g of H_2O Ⓑ 4.8g of C_2H_5OH Ⓒ 2.8g of CO Ⓓ 5.8g of N_2O_5

Q10: Tin has isotopes:
Ⓐ 7 Ⓑ 9 Ⓒ 11 Ⓓ 5

Q11: How many isotopes are present in Palladium?
Ⓐ 4 Ⓑ 5 Ⓒ 6 Ⓓ 7

Q12: Volume occupied by 1.4 g of N_2 at S.T.P is:
Ⓐ 22.4 dm^3 Ⓑ 22.44 dm^3 Ⓒ 1.12 dm^3 Ⓓ 112.0 Cm^3

Q13: One mole of SO_2 contains:
Ⓐ 6.02×10^{23} atoms of oxygen Ⓑ 18.1×10^{23} molecules of SO_2
Ⓒ 6.02×10^{23} atoms of Sulphur Ⓓ 4 g atoms of SO_2

Q14: Molecular Formula = n (empirical formula). Value of n for Sugar is:
Ⓐ 0 Ⓑ 1 Ⓒ 2 Ⓓ 0.5

Q15: 1a.m.u is equal to:
Ⓐ 1.661×10^{-27}kg Ⓑ 1.661×10^{-24}ug Ⓒ 1.661×10^{-21}ng Ⓓ All of these

ANSWERS:

1	B	9	A
2	D	10	C
3	A	11	C
4	B	12	C
5	C	13	C
6	D	14	B
7	C	15	A
8	A		

Short Questions

1	What are isotopes? Why they have same chemical but different physical properties?
2	What is Avogadro's number? Give equation to relate the Avogadro's number and mass of an element.
3	How N, and CO have same number of electrons, protons and neutrons?
4	Mg atom is twice heavier than that of carbon atom, comment
5	No individual atom of neon in the sample has a mass of 20.18 a.m.u. Give reason.
6	Atomic masses are in fractions. Justify.
7	Write down the limitations of a chemical reaction.
8	Define isotopes. Why they have same chemical properties but different physical properties?
9	Define mass spectrum. Which type of information we can get from it?
10	What are monoisotopic elements? Give name and symbol of such an element.
11	Define limiting reactant. Amount of product is controlled by limiting reactant. Why?
12	How limiting reactant is identified? Discuss Steps to determine Limiting Reactant.
13	Law of conservation of mass has to be obeyed during stoichiometric calculations. Justify it.
14	Write down assumptions of stoichiometry.
15	Many chemical reactions take place in our surrounding involves limiting reactants justify.
16	Actual yield is usually less than the theoretical yield. Give reasons. OR Why theoretical yield is greater than actual yield?
17	How the efficiency of a chemical reaction is determined? OR Why we calculate percentage yield?
18	How many molecules of water are there is 10g of ice?
19	Why 23 g of Na and 238 g of uranium have equal number of atoms in them?
20	One mole of H_2SO_4 should completely react with two moles of NaOH. How does Avogadro's Number help us too explain it?

Long Questions

1	What is a Mass Spectrometer? How it is used to determine the percentage abundance and atomic masses of elements.
2	Define combustion analysis. How percentage composition of each element in an organic compound is determined?

3	Define stoichiometry. Give its assumptions. Mentions two laws which help us to perform stoichiometric calculations.
4	Define limiting reactant. Explain with two examples.
5	Define yield. Give its types. Explain why actual yield is lesser than theoretical yield?
6	Write down four steps to determine empirical formula of a compound.

Chapter 2: Experimental Techniques in Chemistry

Q 1: The substance used for decolourization of undesirable color in a crystalline substance:

(A) H_2SO_4 (B) Silica gel (C) $NaNO_3$ (D) Animal Charcoal

Q 2: Which substance does not undergo sublimation?

(A) $KMaO_4$ (B) Naphthalene (C) NH_4Cl (D) Iodine

Q 3: The most common solvent used in solvent extraction is:

(A) Acetone (B) Ethanol (C) Rectified Spirit (D) Diethyl ether

Q 4: Solvent extraction method is particularly used technique for separation when the product to be separated is:

(A) Non-volatile or thermally unstable (B) Volatile or thermally stable
(C) Non-volatile or thermally stable (D) Volatile and thermally unstable

Q 5: Solvent extraction is an equilibrium process and it is controlled by:

(A) Law of mass action (B) The amount of solvent used
(C) Distribution law (D) The amount of solute

Q 6: The chromatography in which the stationary phase is a solid:

(A) Gas liquid chromatography (B) Partition chromatography
(C) Adsorption chromatography (D) Paper chromatography

Q 7: The comparative rates at which solutes move in chromatography depends on:

(A) The size of the paper (B) R_f values of solutes
(C) Temperature of the experiment (D) Size of chromatographic tank used

Answers:

1	D	5	C
2	C	6	C
3	D	7	B
4	D		

Short Questions

1	Mention only steps involved in complete quantitative determination.
2	What is the difference between Gooch's crucible and Sintered glass crucible?
3	Why there is a need to crystallize a crude product?
4	Write down the characteristics of the solvent selected for crystallization of a compound.
5	Define sublimation with two examples.
6	How does the rate of filtration increase by using fluted filter paper?
7	How are crystals decolorized?

8	How are crystals dried?
9	Define Ether Extraction.
10	Define Distribution law (or Partition law) and how it is helpful in solvent extraction?
11	What do you mean by solvent extraction? Which law controls it?
12	Differentiate between Stationary phase and mobile phase?
13	Define Chromatography? Discuss its uses.
14	What is Retardation Factor (R_f)? why it has no units?
15	Differentiate between Adsorption and Partition chromatography with examples.

Note : No long Question included from this chapter according to the scheme 2023

Chapter 3: Gases

Q 1: One torr is equal to:
- Ⓐ One Pascal
- Ⓑ One mm of Hg
- Ⓒ 76 cm of Hg
- Ⓓ One atmosphere

Q 2: The absolute zero is:
- Ⓐ Attainable
- Ⓑ May be attainable
- Ⓒ Unattainable
- Ⓓ May not be attainable

Q 3: Normal human body temperature is?
- Ⓐ 37^oC
- Ⓑ 98.6^oC
- Ⓒ 37^oF
- Ⓓ 273K

Q 4: Which has same number of molecules at STP?
- Ⓐ 11.2 dm^3 of O_2 and 32 g of O_2
- Ⓑ 44g of CO_2 and 1 dm^3 of CO_2
- Ⓒ 28g of N_2 and 5.6 dm^3 of O_2
- Ⓓ 280 ml of CO_2 and 280 cm^3 of N_2O

Q 5: Partial pressure of O_2 in the lungs is:
- Ⓐ 760 torr
- Ⓑ 320 torr
- Ⓒ 116 torr
- Ⓓ 159 torr

Q 6: The exothermic process is:
- Ⓐ Evaporation
- Ⓑ Sublimation
- Ⓒ Respiration
- Ⓓ Boiling

Q 7: Smell of cooking gas during leakage from a gas cylinder is due to the property of gases?
- Ⓐ Diffusion
- Ⓑ Evaporation
- Ⓒ Osmosis
- Ⓓ All of these

Q 8: If absolute temperature of a gas is doubled and pressure is reduced to one half, the volume of the gas will
- Ⓐ Remains unchanged
- Ⓑ Increases four times
- Ⓒ Reduces to $\frac{1}{4}$
- Ⓓ Be doubled

Q 9: Which gas will diffuse more rapidly among the following?
- Ⓐ N_2
- Ⓑ H_2
- Ⓒ NH_3
- Ⓓ CO

Q10: How should the conditions be changed to prevent the volume of a given gas from expanding when its mass is increased?
- Ⓐ Temperature is lowered and pressure is increased.
- Ⓑ Temperature is increased and pressure is lowered.
- Ⓒ Temperature and pressure both are lowered.
- Ⓓ Temperature and pressure both are increased.

Q11: Gases deviate from ideal behavior at high pressure. Which of the following is correct for non-ideality?
- Ⓐ At high pressure, the gas molecules move only in one direction.

Ⓑ At high pressure, the collisions between the gas molecules are increased manifolds.
Ⓒ At high pressure, the volume of the gas becomes insignificant.
Ⓓ At high pressure, the intermolecular attractions become significant.

Q 12: Plasma are found everywhere from the sun to..............

Ⓐ Atoms Ⓑ Molecules Ⓒ Electrons Ⓓ Quarks

Q 13: Deviation of gas from ideal behavior is maximum at:

Ⓐ -10°C and 5 atm Ⓑ -10°C and 2 atm
Ⓒ 400 °C and 2 atm Ⓓ 0°C and 2 atm

Q 14: A real gas obeying van der Waal's equation will resemble an ideal gas if:

Ⓐ Both 'a' and 'b' are large Ⓑ Both 'a' and 'b' are small
Ⓒ a' is small and 'b' is large Ⓓ 'a' is large and 'b' is small

Answers:

1	Ⓒ	8	Ⓑ
2	Ⓒ	9	Ⓑ
3	Ⓐ	10	Ⓐ
4	Ⓓ	11	Ⓓ
5	Ⓒ	12	Ⓓ
6	Ⓒ	13	Ⓐ
7	Ⓐ	14	Ⓑ

Short Questions

1	Define atmospheric pressure. Give its two units.
2	What is the Quantitative definition of Charles's Law OR Throw some light on factor 1/273 in Charles's Law.
3	State Avogadro's law of gases. Give two examples.
4	What do you mean by absolute zero temperature of gases? Give its value. OR What is absolute zero? What happens to the real gases while approaching it?
5	Justify that volume of gas becomes theoretically zero at -273 °C.
6	Convert (i) 37^oC into °F (ii) -40^oC into °F
7	Derive expression of density of gas with the help of general gas equation. OR Prove that d=PM/RT.
8	Calculate the value of 'R' gas constant in S1 units.
9	Derive the value of 'R' when the pressure is measured in atmosphere and volume in dm.
10	State Dalton's law of partial pressure. Give its mathematical form
11	Give two application of Dalton's law of partial pressure OR Apply Dalton's law of partial pressure to determine the Partial pressure of a dry gas?
12	Why deep sea divers take oxygen mixed with an inert gas like He? OR Regular air cannot be used in diver's tank. Give reason Why Pilots feel uncomfortable at high altitude?
13	Differentiate between diffusion and effusion.
14	SO_2 is comparatively non-ideal at 273K but behave ideally at 327°C. Give reason.
15	Hydrogen and Helium are ideal at room temperature, but SO_2 and Cl_2 are non-Ideal.
16	Define the Joule-Thomson effect.

17	Water vapors do not behave ideally at 273K. Why?
18	Define Critical temperature and Pressure. Give one example of each.
19	Derive Charles's law from kinetic molecular theory of gases.
20	Why gases show non-Ideal behavior at low temperature and high pressure?
21	What is plasma? Write down any two applications (Uses) of plasma.
22	Write down two characteristics of plasma. OR Why Plasma is Neutral?
23	Differentiate between Natural and Artificial plasma.

Long Questions

1	A gas having volume of $10dm^3$ is enclosed in a vessel at 0°C and the pressure is 2.5 atm. This gas is allowed to expand until the new pressure is 2 atm. What will be the new volume of this gas, if temperature maintained at 273 K.
2	Calculate the mass of $1dm^3$ of NH_3 gas at 30 °C and 1000 mm Hg pressure, considering that NH_3 is behaving ideally.
3	250 cm^3 of a sample of hydrogen effuses four times as rapidly as 250 cm^3 of an unknown gas. Calculate the molar mass of the unknown gas.
4	Calculate the density of $CH_{4(g)}$ at 0°C and 1 atm pressure. What will happen to the density if (a) temperature is increased to 27°C (b) the pressure is increased to 2 atm at 0°C
5	There is a mixture of hydrogen, helium, and methane occupying a vessel of volume 13 dm^3 at 37°C and pressure l atm. The masses of H_2 and He are 0.8 g and 0.12 g respectively. Calculate the partial pressure in torr of each gas in the mixture.
6	Working at a vacuum line, a chemist isolated a gas in a weighing bulb with a volume of 255cm, at a temperature of 25°C and under a pressure in the bulb of 10.0 torr. The gas weighed 12.1 mg. What is the molecular mass of this gas?
7	The relative densities of two gases A and B are 1:1.5. Find out the volume of B which will diffuse in the same time in which 150dm^3 of A will diffuse?

Chapter 4: Liquids and Solids

Q 1: Dipole-dipole interactions are present in:

Ⓐ Atoms of He gas Ⓑ Molecules of CCl_4 Ⓒ Molecules of solid iodine Ⓓ Molecules of NH_3

Q 2: Forces which are present between iron and water molecules are:

Ⓐ Dipole-induced dipole forces Ⓑ Dipole-dipole forces
Ⓒ Ion-dipole forces Ⓓ London dispersion forces

Q 3: London dispersion forces are significant for:

Ⓐ Polar molecules Ⓑ ionic solids Ⓒ metals Ⓓ non-polar molecules

Q 4: NH_3 shows a maximum among the hydrides of VA group elements due to:

Ⓐ Very small size of the Nitrogen Ⓑ Lone pair of electrons present on nitrogen
Ⓒ Enhanced electronegative character of N Ⓓ Pyramidal structure of NH_3

Q 5: The boiling point of glycerin at 1 atmospheric pressure is:

Ⓐ 210 °C Ⓑ 270 °C Ⓒ 290 °C Ⓓ 300 °C

Q 6: Boiling point of water and ethanol is:

Ⓐ Equal Ⓑ Different Ⓒ 98 °C and 70°C Ⓓ 100°C and 90° C

Q 7: Vapour pressure of a substance does not depend upon:

Ⓐ Temperature Ⓑ Intermolecular forces
Ⓒ Surface Area Ⓓ Physical state of substance

Q8: Amorphous solids:

Ⓐ Have sharp melting points Ⓑ Undergo clean cleavage when cut with knife
Ⓒ Have perfect arrangement of atom Ⓓ Can possess small portions of orderly arrangement

Q9: Glass may begin to crystallize by a process called:

Ⓐ Super cooling Ⓑ Sublimation Ⓒ Crystallization Ⓓ Annealing

Q10: Allotropy is the property of:

Ⓐ Compound Ⓑ Element Ⓒ Atom Ⓓ Mixture

Q11: The branch of science which deals with structure of crystals is called:

Ⓐ Anisotropy Ⓑ Isomorphy Ⓒ Crystallography Ⓓ Stoichiometry

Q12: In an orthorhombic crystal, the unit cell dimensions are:

Ⓐ $a = b \neq c \quad \alpha = \beta = \gamma = 90°$ Ⓑ $a \neq b \neq c \quad \alpha = \beta = \gamma = 90°$
Ⓒ $a \neq b \neq c \quad \alpha = \beta \neq \gamma = 90°$ Ⓓ $a \neq b \neq c \quad \alpha = \beta = \gamma \neq 90°$

Q13: The example of hexagonal structure is:

Ⓐ Sulphur Ⓑ NaCl Ⓒ Graphite Ⓓ Diamond

Q14: The no. of Cl ions per unit cell of NaCl:

Ⓐ 6 Ⓑ 4 Ⓒ 2 Ⓓ 8

Q15: There are .... Bravis lattices:

Ⓐ 7 Ⓑ 10 Ⓒ 14 Ⓓ 17

Q16: The molecules of CO_2:

Ⓐ Ionic crystals Ⓑ Metallic Crystals Ⓒ Covalent crystals Ⓓ Any type of crystal

Q17: The number of carbon atoms in 22.0 g of CO_2 are:

Ⓐ 3.01×10^{23} Ⓑ 6.02×10^{23} Ⓒ 3.01×10^{22} Ⓓ 6.02×10^{22}

Answers:

1	Ⓓ	10	Ⓑ
2	Ⓑ	11	Ⓒ
3	Ⓓ	12	Ⓑ
4	Ⓒ	13	Ⓒ
5	Ⓒ	14	Ⓑ
6	Ⓑ	15	Ⓒ
7	Ⓒ	16	Ⓒ
8	Ⓓ	17	Ⓐ
9	Ⓓ		

Short Questions

1	What are dipole-dipole forces? How they affect the thermodynamic properties of substances?
2	Lower alcohols are soluble in water, but hydrocarbons are insoluble. Why? OR Why ethyl alcohol is soluble in water?
3	What are "dipole-induced dipole forces? OR What are Debye forces?
4	Explain cleansing Action of detergents and soap on the basis of H-Bonding.
5	How does polarizability effect the strength of London forces?
6	Why HF has a less acidic strength than HCl, HBr, HI?
7	Why ice floats over the surface of water? OR Ice has less density than Liquid water. Why?
8	Water is a liquid at room temperature, while H_2S is a gas. Why?
9	Define vapor pressure. Explain the two factors which affect the vapor pressure of liquid.
10	Earthenware vessels keep water cool. Justify.
11	One feels the sense of cooling under the fan after bath. Describe it.
12	Evaporation causes cooling. Give reason.
13	Why vacuum distillation can be used to avoid decomposition of a sensitive liquid?
14	Why are the vapor pressures of solids far less than those of liquids?
15	Why boiling point of H_2O is different at Murree hills and at Mount Everest?
16	Diamond is hard and electrically insulator, Give reason.
17	Cleavage of crystals is itself an anisotropic behavior. Why?
18	Define Isotropy and Anisotropy.
19	Why sodium chloride and cesium chloride have different structures?
20	Sodium is softer than copper, but both are very good electrical conductor. Why?
21	Ionic crystals are highly brittle. How?
22	Why electrical conductivity of metals decreases with increase in temperature? Why?
23	What are: (i) Symmetry (ii) Habit of crystal
24	Write four properties of molecular solids.
25	Distinguish between isomorphism and polymorphism.
26	How liquid crystals act as temperature sensors?
27	Freshly cut metals show the property of metallic luster. Justify.
28	Define transition temperature. Give two examples.

Long Questions

1	Define vapour pressure. How vapour pressure of a liquid is determined by manometric method?
2	Define liquid crystal. Give its types. Give their use in daily life.
3	What are molecular solids? Give examples and explain their properties.
4	Explain electron pool theory of gases. Also give three properties of metallic solids.
5	Define hydrogen bonding. Explain hydrogen bonding in biological molecules.
6	Discuss the factors affecting London dispersion forces.

Chapter 5: Atomic Structures

Q 1: In the ground state of an atom, the electron is present:

- Ⓐ In the nucleus
- Ⓑ In the second shell
- Ⓒ Nearest to the nucleus
- Ⓓ Farthest from the nucleus

Q 2: The e/m value for the positive rays is the maximum for the gas:

- Ⓐ Hydrogen
- Ⓑ Helium
- Ⓒ Oxygen
- Ⓓ Nitrogen

Q 3: The nature of positive rays /anode rays depends on:

- Ⓐ Hydrogen
- Ⓑ The nature of the discharge tube
- Ⓒ The nature of residual gas
- Ⓓ All of these

Q 4: Free neutron decays into a proton with the emission of an electron and a:

- Ⓐ Positron
- Ⓑ Neutrino
- Ⓒ Beta particle
- Ⓓ Helium nucleus

Q 5: The velocity of photon is:

- Ⓐ Independent of its wavelength
- Ⓑ Depends on its source
- Ⓒ Equal to square of its amplitude
- Ⓓ Independent on its wavelength

Q 6: According to Bohr's atomic mode, radius of second orbit of hydrogen atom is:

- Ⓐ 0.529 A
- Ⓑ 5.0 A
- Ⓒ 2.116A
- Ⓓ 4.0 A

Q 7: Negative charge on cathode rays was established by:

- Ⓐ William Crook
- Ⓑ J. Perin
- Ⓒ JJ. Thomson
- Ⓓ Hittrof

Q 8: What is 'X' in the given reaction? $\boxed{He + Be \longrightarrow C + X}$

- Ⓐ Electron
- Ⓑ Proton
- Ⓒ Neutron
- Ⓓ Gamma Rays

Q 9: Which of the following series lie in ultra-violet region?

- Ⓐ Lyman
- Ⓑ Balmer
- Ⓒ Paschen
- Ⓓ Brackett

Q 10: The limiting line of the Balmer series lies in:

- Ⓐ U.V region
- Ⓑ Visible region
- Ⓒ I.R region
- Ⓓ None of these

Q 11: Splitting of spectral lines when excited atoms are subjected to strong electric field is called:

- Ⓐ Zeeman effect
- Ⓑ Stark effect
- Ⓒ Photoelectric effect
- Ⓓ Compton's effect

Q12: All the d-orbitals have,

- Ⓐ Spherical shape
- Ⓑ Dumbbell shape
- Ⓒ Four lobe shape
- Ⓓ None of above

Q13: Quantum number values for 3d orbitals are:

- Ⓐ n=3, l=1
- Ⓑ n=3, l=2
- Ⓒ n=3, l=0
- Ⓓ n=2, l=2

Q14: When 6d orbital is complete, the next entering electron goes into:

- Ⓐ 7f orbital
- Ⓑ 7s obital
- Ⓒ 7p orbital
- Ⓓ 7d orbital

Q15: How many unpaired electrons are present in an atom of configuration: $1s^2$, $2s^2$, 2p?

- Ⓐ 4
- Ⓑ 0
- Ⓒ 2
- Ⓓ 3

1	Ⓒ	4	Ⓑ	7	Ⓒ	10	Ⓑ	13	Ⓑ
2	Ⓐ	5	Ⓓ	8	Ⓒ	11	Ⓑ	14	Ⓑ
3	Ⓒ	6	Ⓒ	9	Ⓐ	12	Ⓒ	15	Ⓒ

Short Questions	
1	Give two defects of the Rutherford Model.
2	Differentiate between slow and fast neutrons with examples.
3	Why it is necessary to decrease the pressure in the discharge tube?
4	Whichever gas in the discharge tube nature of cathode rays remains the same. Justify. OR Why is e/m value of the cathode rays just equal to that of electron?
5	e/m value of is same for cathode rays (electrons) but different for positive rays?
6	Why positive rays are also called 'Canal rays"? How Positive rays are produced?
7	Calculate Mass of an electron when m/e = $1.758 \times 10^{11} C\ kg^{-1}$
8	Justify that cathode rays are material partials OR cathode rays have momentum.
9	Define Zeeman's effect and Stark effect.
10	How neutrons were discovered by Chadwick? Give the equation of nuclear reaction involved.
11	State the Heisenberg's Uncertainty Principle and give its mathematical form.
12	What is the difference between Orbit and Orbital? Draw the shape of the P orbital.
13	What is the difference between continuous spectrum and line spectrum?
14	Differentiate between Atomic Emission and atomic Absorption spectrum.
15	What particles are formed by the decay of free neutron? Give equation?
16	Why the potential energy of revolving electron is with negative sign?
17	How the dual nature of an electron was got verified?
18	State Moseley's law. Give its importance.
19	Derive de-Broglie equation.
20	Draw the shapes of d-orbitals.
21	What is the difference between orbit and orbital?
22	Hund's rule and Pauli's exclusion principle.
23	Describe (n+1) rule for the distribution of electron.

Long Questions	
1	What are quantum numbers? Discuss Azimuthal and Magnetic quantum numbers.
2	Give defects of Bohr's atomic model.
3	Write down Millikan's oil drop method for the measurement of charge on electron.
4	Derive an expression to calculate of radius of revolving electron in orbit by Bohr's model of atom
5	Write down main postulates of Bohr's atomic model.
6	What are X-rays? Discuss its origin. Explain Moseley study.
7	Write down four properties of cathode rays.

Chapter 6: Chemical Bonding

Q 1: The decrease in atomic radius is small when travel from left to right in Transition Elements. It is due to:

(A) Valence electrons (B) Number of shells
(C) Nuclear charge (D) Intervening electrons

Q 2: The elements having low ionization energy are:

(A) Non-metals (B) Metals (C) Semi-metals (D) Metalloids

Q 3: The value of the third ionization energy of Mg is:

(A) 1450 kJ mol^{-1} (B) 7730 kJ mol^{-1} (C) 7850 kJ mol^{-1} (D) 1890 kJ mol^{-1}

Q 4: Ionic, Covalent and Coordinate Covalent bond is present in:

(A) SO_2 (B) C_2H_5 (C) NH_4Cl (D) H_2O

Q 5: VSEPR theory was proposed by?

(A) Nyholm and Gillespie (B) Kossel (C) Lewis (D) Sidwick & Powell

Q 6: Bond angle between two H-S-H bond is:

(A) 104.5° (B) 107.5° (C) 92° (D) 95°

Q 7: Which of the following is AB, type molecule with two lone pairs?

(A) $BeCl_2$ (B) CH_4 (C) BF_4 (D) H_2S

Q 8: The hybridization in ammonia molecule is:

(A) dsp^2 (B) Sp^2 (C) Sp^3 (D) sp

Q 9: Total number of sigma bonds in Ethyne (C_2H_2)/Acetylene:

(A) Five (B) Three (C) Two (D) Four

Q10: The number of bonds in nitrogen molecules is:

(A) One sigma(σ) and one Pi(π) bond (B) One sigma(σ) and two Pi(π) bonds
(C) Three sigma (σ) bonds only (D) Two sigma(σ) and one Pi(π) bond

Q11: Which has unpaired electrons in anti-bonding molecular orbitals?

(A) N_2^{-2} (B) O_2^{-2} (C) B_2 (D) F_2

Q12: The paramagnetic property of oxygen is well-explained on the basis of:

(A) VSEPR-theory (B) VB-theory (C) MOT-theory (D) None of these

Q13: The bond order of O_2^{-2}:

(A) One (B) Two (C) Three (D) Four

Q14: Which of the hydrogen halide has highest percentage of ionic character?

(A) HCl (B) HBr (C) HF (D) HI

Answers:

1	D	4	C	7	D	10	B	13	A
2	B	5	D	8	C	11	A	14	C
3	B	6	C	9	B	12	C		

Short Questions

1	What is the octet rule? Why it is not universal?
2	Bond distance is the compromise distance between two atoms, Justify.
3	Name the four factors affecting ionization energies.
4	Why second ionization energy is higher than first ionization energy?
5	Why is the size of anion greater than the size of parent atom?
6	Cationic radius is always smaller than the size of the parent atom. Why?
7	Describe the variation of electron affinity along periods and groups in periodic table?
8	lonization energy is an index to the metallic nature of element Justify.
9	Why sizes of the atoms cannot be measure preciously?
10	No bond in chemistry is 100% ionic. Justify it
11	Sigma bond is stronger than pi-bond. Why?
12	Draw the hybridization diagram of H_2O , CH_4 , NH_3
13	Why lone pair occupies more space then bond pair of electrons.
14	What is the basic assumption of VSEPR theory?
15	Why are pi-bonds more diffused than sigma bonds? OR Differentiate between sigma and pi bond?
16	How nature of bond can be determined by electronegativity values?
17	Differentiate between Bonding and Antibonding molecular orbitals.
18	What is bond order? Give an example.
19	Why does Helium not exist in the form of He2??
20	Why is MOT superior to VBT?
21	How dipole moment is helpful to determine the molecular structure?
22	Define dipole moment and write the units of it.

Long Questions

1	Explain the bonding in O_2 or N_2 according to molecular orbital theory and explain its paramagnetic property.
2	Give the postulates of VSEPR theory. Explain the structure of ammonia on the basis of this theory.
3	Define dipole moment. Give its units. How is it used to determine the geometry of molecules?
4	Define ionization energy. Give example. Discuss its trend in groups and periods.
5	Describe sp^2 hybridization to explain the structure of ethene.
6	Describe sp hybridization to explain the structure of ethyne.
7	Define sp^3 hybridization. Discuss the structure of water and methane on its basis.

Chapter 7: Thermochemistry

Q1: If an endothermic reaction is allowed to take place very rapidly in the air, the temperature of the surrounding air:

Ⓐ Remains constant Ⓑ Increases Ⓒ Decreases Ⓓ Remains unchanged

Q2: For a given process, the heat changes at constant pressure (q_p) and at constant volume (q_v) are related to each other as:

Ⓐ $q_p = q_v$ Ⓑ $q_p > q_v$ Ⓒ $q_p < q_v$ Ⓓ $q_p = q_v/2$

Q3: Which of the following is not a state function?

(A) Enthalpy (B) Temperature (C) Heat (D) Pressure

Q4: The units of Heat Capacity are:

(A) $kjK^{-1}mol^{-1}$ (B) $kjK^{-1}g^{-1}$ (C) kjK^{-1} (D) $kjK^{-1}mol^{-2}$

Q5: The change in heat energy of a chemical reaction at constant temperature and pressure is called:

(A) Enthalpy change (B) Bond Energy (C) Internal energy (D) Heat of sublimation

Q6: The pressure of oxygen inside of bomb calorimeter:

(A) 100 atm (B) 50 atm (C) 125 atm (D) 20 atm

Q7: The enthalpy of atomization of hydrogen is:

(A) 180 kj mol^{-1} (B) 218 kj mol^{-1} (C) -1368 kj mol^{-1} (D) -57.4 kj mol^{-1}

Answers:

1	B	4	C	7	B
2	B	5	A		
3	C	6	D		

Short Questions

1	What is a thermochemical equation? Give two examples.
2	Differentiate between exothermic reaction and endothermic reaction.
3	Differentiate between spontaneous and non-spontaneous reactions with examples.
4	The burning if candle is a spontaneous process. Explain.
5	Define Enthalpy of Neutralization and Enthalpy of Combustion.
6	Define Enthalpy of Solution. Give examples.
7	Define Enthalpy of atomization with an example.
8	Define Hess's law of constant heat summation with one example.
9	Define System, Surrounding and Boundary.
10	What is State Function? Give two examples.
11	Why is it necessary to mention the physical states of reactants and products in a thermochemical equation of a reaction
12	What is difference between heat and temperature? Write a mathematical relationship between these two parameters?
13	State first law of thermodynamics.
14	State Born Haber Cycle.
15	Differentiate between internal energy and enthalpy.
16	Prove that AE = qv
17	Define heat and work.

Long Questions

1	What is Born Haber Cycle? How it used to calculate the Lattice energy of NaCI?
2	How enthalpy of food is determined by bomb calorimeter?
3	Explain glass calorimeter is used to determine the enthalpy of a reaction?
4	State and explain with two example, the Hess's law of constant heat summation
5	State first law of thermodynamics. Prove that: AE = q_v and AH = q_p

6	Define the terms: system, surroundings, boundary, and state function.
7	Differentiate between spontaneous and non-spontaneous reactions.

Chapter 8: Chemical Equilibrium

Q1: Molar concentration is called:

Ⓐ Active mass Ⓑ Weight Ⓒ Mass Ⓓ None of theses

Q2: For which system does the equilibrium Ke have units of (conc^{-1})?

Ⓐ $N_2 + 3H_2 \rightarrow 2NH_3$ Ⓑ $H_2 + I_2 \rightarrow 2HI$
Ⓒ $2NO_2 \rightarrow N_2H_4$ Ⓓ $2HF \rightarrow H_2 + F_2$

Q3: Equilibrium constant for gaseous equilibrium is represented by:

Ⓐ K_o Ⓑ K_g Ⓖ K_v Ⓓ K_p

Q4: When K_c value is small, the equilibrium position is:

Ⓐ Towards left Ⓑ Towards right Ⓖ Remains unchanged Ⓓ None of these

Q5: Which statement about the following equilibrium is correct: $2SO_2 + O_2 \rightarrow 2SO_4$ $\Delta H = -183 kjmol^{-1}$. The yield of SO_3 will be maximum if:

Ⓐ Both temperature and pressure are reduced Ⓑ Temperature is increased and pressure is kept constant
Ⓒ Both temperature and pressure are increased Ⓓ Temperature is reduced and pressure is increased

Q6: The unit of ionic product of water:

Ⓐ mol^{-1} dm^3 Ⓑ mol^{-2} dm^3 Ⓖ mol^{-2} dm^5 Ⓓ mol^{-2} dm^4

Q7: The ionic product of water will increase if:

Ⓐ H^+ ions are added Ⓑ OH^- sons are added
Ⓒ Temperature is increased Ⓓ H^+ and OH^- ions are added in equal amount

Q8: The law of mass action was given by:

Ⓐ Vant's Hoff Ⓑ Bondeinstin Ⓖ Guldberg and Waage Ⓓ Berthelot

Q9: When 50% of reactants in a reversible reaction are converted into a product, the value of equilibrium constant Ke is:

Ⓐ 2 Ⓑ 1 Ⓖ 3 Ⓓ 4

Q10: The term pH was introduced by:

Ⓐ Henderson Ⓑ Sorenson Ⓖ Goldsmith Ⓓ Thomson

Q11: The pH of 10^{-2} moles dm^{-3} of an aqueous solution of NaOH is:

Ⓐ 2 Ⓑ 12 Ⓖ 3 Ⓓ 10

Q12: Which relationship is correct about the strength of an acid with the strength of its conjugated base?

Ⓐ $K_a \propto \frac{1}{K_a}$ Ⓑ $K_a \propto K_b$ Ⓖ $\overline{Ka} \propto K_b$ Ⓓ None of these

Q13: Sum of pK_a and pK_b is equal to:

Ⓐ 14 Ⓑ 7 Ⓖ 0 Ⓓ 1

Q14: Ionization of hydrogen sulphide gas is suppressed by:

Ⓐ KCI Ⓑ NaCI Ⓖ HCI Ⓓ NH_4CI

Q15: The K_p value for $PbSO_4$ is 1.8×10^4. The maximum concentration of Pb^{-2} ions is:

Ⓐ 1.8×10^{-4} Ⓑ 1×10^{-4} Ⓖ 1.34×10^{-4} Ⓓ 1.69×10^{-4}

Q16: When a small amount of acid or a base is added to a buffer solution, its pH value will change:
Ⓐ Drastically Ⓑ A little Ⓒ Rapidly Ⓓ Not at all

Q17: If NaOH is added to a solution of CH_2COOH, then
Ⓐ pH of solution decreases Ⓑ H^+ ions concentration decreases
Ⓒ CH_3COO^- ions concentration increases Ⓓ OH^- ions concentration increases

Q18: POH of a solution is 4. The $[H^+]$ ion concentration of the solution is:
Ⓐ 10^{-10} mol dm^{-3} Ⓑ 4 mol dm^{-3} Ⓒ 0.4 mol dm^{-3} Ⓓ 4×10^4 mol dm^{-3}

1	A	6	D	11	B	16	D
2	C	7	C	12	A	17	A
3	D	8	C	13	A	18	A
4	A	9	B	14	C		
5	D	10	B	15	C		

Short Questions

1	How do the Buffer Act?
2	Differentiate between Reversible and Irreversible reactions.
3	How values of Ke of a reaction help to predict the direction of a reversible reaction?
4	How value of equilibrium constant (Kc) helps to predict extent of a reaction?
5	Define pKa and pKb.
6	Define Law of Mass Action and Equilibrium constant (Kc).
7	Calculate the pH of 10^{-4} mole dm^{-3} of $Ba(OH)_2$.
8	Calculate the pH of 10^{-4} mole dm^3 of HCl.
9	Define pH and pOH with mathematical expressions.What is the sum of pH and pOH at 25°C?
10	What is ionic product of water? Write its value at 25°C.
11	What is the Common Ion Effect? How is NaCl purified by common ion effect? OR What is the effect of common ion on solubility?
12	What are buffer solutions? Why do we need buffers in daily life?
13	Write down Henderson's equation for acidic and basic buffer.
14	How does a catalyst affect the equilibrium position?
15	State Le Chatelier's principle.
16	Give optimum conditions for synthesis of Ammonia gas by Haber's process.
17	Define buffer capacity. Write down Henderson's equation for acidic buffers.
18	Explain that a mixture of NH_4OH and NH_4Cl gives us the basic buffer.
19	What is the solubility product? Derive solubility product expression for $Ag2CrO_4$

Long Questions

1	Describe the Beckmann method to determine the freezing point depression.
2	Explain the Landsberger method to calculate the elevation of boiling point.
3	Explain the Raoult's law when both components in the solution and volatile.
4	How lowering of vapour pressure as colligative property is used to calculate molar mass of solute.
5	Discuss phenol water system and explain the term upper consulate temperature.
6	Define hydration and hydrolysis with examples.

Chapter 9: Solutions

Q1: Molarity of pure water is:

Ⓐ 1 Ⓑ 18 Ⓒ 55.5 Ⓓ 6

Q2: A solution of glucose is 10%. The volume in which 1 g mole of it is dissolved will be:

Ⓐ 1 dm^3 Ⓑ 1.8 dm^3 Ⓒ 200 dm^3 Ⓓ 900 dm^3

Q3: The consulate temperature of water-aniline system is:

Ⓐ 69.5 °C Ⓑ 64.5 °C Ⓒ 167 °C Ⓓ 49.1 °C

Q4: If 2.0 g of NaCl is dissolved in 20 g of water, the percentage by weight of NaCl:

Ⓐ 0.99% Ⓑ 10% Ⓒ 9.09% Ⓓ 0.90%

Q5: Which one is not equation of Raoult's law:

Ⓐ $\Delta p = px$ Ⓑ $\Delta p = x_2$ Ⓒ $PV = nRT$ Ⓓ $PV = RT$

Q6: Which of the following pair of liquids is practically immiscible?

Ⓐ Water and phenol Ⓑ Water and aniline
Ⓒ Water and benzene Ⓓ Water and Nicotine

Q7: In azeotropic mixture showing positive deviation from Raoult's law, the volume of the mixture is:

Ⓐ Slightly more than the total volume of the components
Ⓑ Equal to total volume of components
Ⓒ Slightly less than the total volume of the components
Ⓓ None of these

Q8: Which of the following solutions has the highest boiling point?

Ⓐ 5.58% solution of sodium chloride Ⓑ 18% solution of glucose
Ⓒ 6% solution of urea Ⓓ All have the same boiling point

Q9: Freezing point of Equi-molal aqueous solutions will be minimum for:

Ⓐ Glucose Ⓑ Fructose Ⓒ NH_4Cl Ⓓ Urea

Q10: Which one of the following substances when dissolved in water give acidic solution?

Ⓐ NaCl Ⓑ Na_2SO_4 Ⓒ NH_4Cl Ⓓ NH_3COONH_4

Q11: The molal boiling point constant depends upon:

Ⓐ Nature of solvent Ⓑ Nature of solute
Ⓒ Temperature of Solution Ⓓ Vapor pressure of Solution

Q12: Chemical used to protect a car by preventing the liquid in the radiator from freezing is:

Ⓐ Phenol Ⓑ Ethylene Glycol Ⓒ KNO_3 Ⓓ Methanol

Q13: 18 g glucose in dissolved in 90 g of water. The relative lowering of vapor pressure is equal to:

Ⓐ 1/5 Ⓑ 5.1 Ⓒ 1/51 Ⓓ 6

Q14: The molal boiling point constant is the ratio of the elevation in boiling point to:

Ⓐ Molality Ⓑ Molarity Ⓒ Mole fraction of solvent Ⓓ Mole fraction of solute

1	Ⓒ	4	Ⓒ	7	Ⓐ	10	Ⓒ	13	Ⓒ
2	Ⓑ	5	Ⓒ	8	Ⓐ	11	Ⓐ	14	Ⓐ
3	Ⓒ	6	Ⓒ	9	Ⓒ	12	Ⓑ		

Short Questions

1	What is parts per million? Write its mathematical expression.
2	Define molarity and molality and write its formula.
3	How the molality is independent of temperature, but molarity depends upon temperature?
4	Define mole fraction. Justify that sum of mole fractions is always equal to unity.
5	Give three statements of Raoult's law.
6	Non-ideal solutions so not obey Raoult's Law. Give reason.
7	Differentiate between ideal and non-ideal solutions.
8	Define conjugate solution with one example.
9	What is fractional crystallization?
10	Why freezing points of solvents are depressed due to presence of solutes?
11	Define molal boiling point (Ebullioscopic) and molal freezing point (Cryoscopic) constant giving example.
12	Why NaCl and KNO_3 are used to lower melting point of ice?
13	Define colligative properties. Name them.
14	Write down conditions which should be fulfilled to observe colligative properties.
15	Differentiate between zeotropic and azeotropic mixtures.
16	Define solubility and solubility curves. Name the two types of solubility curves. OR What are Continuous and discontinuous solubility curves?
17	Define hydration energy of ions.
18	Define water of crystallization with on example.
19	What are hydrates? How are they formed? Give two examples.

Long Questions

1	The solubility of PbF_2 at 25 °C is 0.64 g dm^{-3}. Calculate Ksp of PbF.
2	$Ca(OH)_2$ is a sparingly soluble compound. Its solubility product is 6. 5 x 10^{-6}. Calculate the solubility of $Ca(OH)_2$
3	The solubility of CaF_2 in water at 25°C is found to be 2.05 x 10^{-4} $moldm^{-3}$. What is the value of Kap at this temperature?
4	What is the percentage ionization of acetic acid in a solution in which 0.1 moles of it has been dissolved per dm^3 of the solution?
5	Calculate the pH of a buffer solution in which 0.11 molar CH_3COONa and 0.09 molar acetic acid solutions are present. K_a for CH_3COOH is 1.85 x 10^{-5}.

Chapter 10: Electrochemistry

Q1: Oxidation no. of oxygen is -1 in:

Ⓐ Na_2O Ⓑ F_2O Ⓒ Na_2O_2 Ⓓ Cl_2O

Q2: Sulphur has highest oxidation state in:

Ⓐ SO_2 Ⓑ H_2SO_3 Ⓒ H_2S Ⓓ H_2SO_4

Q3: In conversion of Br^2 to BrO^{-1} oxidation state of bromine changes from:

Ⓐ 1 to 5 Ⓑ 0 to 5 Ⓖ 0 to -1 Ⓓ 2 to -3

Q4: In the reaction $2Fe + 3Cl_2 \rightarrow 2FeCl_3$

Ⓐ Iron is reduced Ⓑ Iron is oxidized
Ⓒ Chlorine is oxidized Ⓓ Neither of the elements is oxidized

Q5: Which one of the following cells is used for the extraction of Na metal?

(A) Nelson's Cell (B) Galvanic Cell (C) Down's Cell (D) All of these cells

Q6: The cathodic reaction in the electrolysis of dil. H_2SO_4 with Fe electrode is:

(A) Reduction (B) Oxidation
(C) Both oxidation and reduction (D) Neither oxidation or reduction

Q7: Cathode in NICAD:

(A) Ag_2O (B) NiO_2 (C) Cd (D) Zn

Answers:

1	C	3	C	5	C	7	B
2	D	4	B	6	A		

Short Questions

1	Define electrochemistry.
2	Define electrolytic cell. Give example.
3	Differentiate between Galvanic and Electrolytic cell.
4	Voltaic cell is a reversible cell. Justify it.
5	Define oxidation state with two examples. Rules for assigning Oxidation state
6	Calculate oxidation number of 'S' in H_2SO_4, $Cr_2(SO_4)_3$ SO_4^{2-} and 'Mn' in $KMnO_4$ Na_2MnO_4
7	What is the difference between metallic and electrolytic conduction?
8	Differentiate between primary and secondary cells with examples.
9	A porous plate or salt bridge is not required in lead storage cell. Give reason.
10	Write down the difference between ionization and electrolysis.
11	Impure copper can be purified by electrolytic process. Give reason.
12	What is Anodized Aluminium? Give its advantages.
13	What is Hall-Beroult process?
14	What is function of salt bridge? OR How salt bridge maintains the electrical neutrality in the cell.
15	Write down the reactions for the electrolysis of fused sodium chloride.
16	Why Na and K can displace hydrogen from acids, but Pt, Pd, and Cu cannot?
17	How electrochemical series helps to predict the feasibility of a chemical reaction?
18	Define electrode potential and standard reduction potential.
19	Describe the construction of SHE.
20	Write chemical reactions taking place in NICAD cell.
21	Write down the reaction at anode and cathode of silver oxide battery.
22	Write down chemical reactions taking place in alkaline battery.
23	Explain the electrolysis of fused $PbCl_2$.

Long Questions

1	Explain the construction and working of fuel cells. Give their advantages.
2	What is standard hydrogen electrode (SHE). How it is used to measure the electrode potential of zinc.

3	Give two industrial applications of electrolysis. OR How can you prepare sodium metal and caustic soda electrolysis OR Describe the electrolysis of molten (fused) NaCl and aqueous solution of NaCl.
4	Explain charging and discharging of lead accumulator.
5	Define electrochemical series. Give its two applications.
6	Explain construction and working of galvanic cell. Discuss why salt bridge is necessary in this cell.

Chapter 11: Reaction Kinetics

Q1: The rate of reaction is:

Ⓐ Increases as the reaction proceeds Ⓑ Decreases as the reaction proceeds
Ⓒ Remains the same as reaction proceeds Ⓓ May increase or decrease as reaction proceeds

Q2: The true representative for unit of rate constant for 1" order reaction:

Ⓐ Sec^{-1} Ⓑ mol dm^3 s^{-1} Ⓒ Mol^2 dm^3 s^{-1} Ⓓ mol dm^3 s

Q3: Photosynthesis is:

Ⓐ First order Ⓑ Second order Ⓒ Zero order Ⓓ Third order

Q4: If 75% of any given amount of radioactive element disintegrates in 60 minutes, the half-life of radioactive is:

Ⓐ 20 minutes Ⓑ 30 minutes Ⓒ 45 minutes Ⓓ 25 minutes

Q5: NO catalyzes the oxidation of Sulphur dioxide to Sulphur trioxide. This is an example of:

Ⓐ Homogeneous Catalysis Ⓑ Heterogeneous Catalysis
Ⓒ Neutralization Reaction Ⓓ None of these

Q6: The catalyst used for the reaction, $HCOOH \rightarrow H_2 + CO_2$

Ⓐ Copper Ⓑ Alumina Ⓒ Silica Ⓓ Iron

Q7: The energy of activated complex is:

Ⓐ Greater than reactants and products Ⓑ Less than the reactants and products
Ⓒ Equal to the products Ⓓ Equal to the reactants

Answers:

1	Ⓑ	3	Ⓒ	5	Ⓐ	7	Ⓐ
2	Ⓐ	4	Ⓑ	6	Ⓐ		

Short Questions	
1	Differentiate between rate of chemical reaction and rate constant.
2	Differentiate between instantaneous and average rate of reaction.
3	Define Specific Rate Constant or Velocity Constant.
4	Rate of a reaction is an ever-changing parameter. Justify
5	Define Zero and Pseudo 1 Order reactions with example.
6	Photochemical reactions are usually zero-order reactions. Justify.
7	Define half-life period. How is it used to determine the order of reaction?

8	The radioactive decay is always a first-order reaction. Explain.
9	Define order of reaction with an example.
10	Name the four methods to determine the order of reaction.
11	What is rate determining step? Give a suitable example.
12	How surface area affects the rate of reaction? Give one example.
13	Define energy of activation and activated complex.
14	What are reaction intermediates? Give one example.
15	Give two characteristics of enzyme catalysis.
16	Define Autocatalysis with example.
17	Differentiate between homogeneous and heterogeneous catalysis. Give example.
18	Define catalytic poisoning and give an example. OR what is negative catalysis?
19	A finely divided catalyst may prove effective. Give reason.
20	A catalyst is specific in its action. Justify.

Long Questions

1	Give four characteristics of a catalyst.
2	Define enzyme catalysis. Give its four characteristics.
3	How does Arrhenius equation help us to calculate energy of activation of a reaction?
4	What is order of reaction? Name five methods to calculate order of a reaction. Explain two methods to find the order of a reaction.
5	Explain the effect of temperature on rate of reaction.
6	How light and surface area affects rate of a chemical reaction.
7	Discuss chemical method to determine the rate of a chemical reaction.