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CHAPTER : 1
SOME BASIC CONCEPTS OF CHEMISTRY

1 mark questions

1. What is Chemistry?

Ans: It is a Branch of science deals with the study of composition, properties and interaction of matter.

2. What are the basic constituents of matter?

Ans: The basic constituents of matter are atoms and molecules

3. Mention any two life saving drugs

Ans: 1. Cisplatin 2. Taxol

4. Name the drug used in the treatment of AIDS

Ans: Azidothymidine (AZT)

5. Name the drug used in the cancer therapy

Ans: 1. Cisplatin 2. Taxol

6. Which chemical is responsible for depletion of Ozone

Ans: Chloroflurocarbons (CFCs)

7. Mention green house gases

Ans: Methane (CH_4) and carbon dioxide (CO_2)

8. What is matter?

Ans: Matter is anything which has mass and occupies space.

9. Define solid

Ans: Solid is a substance have definite shape and volume

10. Define Liquid

Ans: Liquid is a substance which has definite volume but not definite shape.

11. Define Gas

Ans: Gas is a substance which has neither definite volume nor definite shape.

12. Name the SI unit of density

Ans: S.I unit of density is Kg/m^3 or Kgm^{-3} or gm/cm^3

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13. Define Significant figure

Ans: Significant figures are meaningful digits which are known with certainty.

14. What is meant by dimensional analysis?

Ans: Any calculation involving the use of the dimensions of the different physical quantities involved is called dimensional analysis.

15. What is atomic mass unit?

Ans: Atomic mass unit is defined as a mass exactly equal to $1/12^{th}$ the mass of one carbon 12 atom

16. What is the value of 1 a.m.u ?

Ans: 1 a.m.u = 1.66056×10^{-24} g

17. Define molecular mass

Ans: Molecular mass is the sum of atomic masses of the elements present in a molecule.

18. What is molar mass in gms?
Ans: The mass of one mole of a substance in grams is called molar mass.
19. Write the formula used to calculate the % composition of elements
Ans: Mass percentage of an element =

$$\frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$$

20. What is empirical formula?
Ans: It is a simplest whole number ratio of various atoms present in a compound.
21. What is molecular formula?
Ans: It is a exact formula indicates number of different types of atoms present in a molecule of a compound.
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22. What is a stoichiometry?
Ans: The quantitative relationships between the various reactants and products interms of moles, masses, molecules and volumes is called stoichiometry.
23. What is limiting reagent?
Ans: Out of various reactants in a reaction, a reactant that is completely consumed in a chemical reaction is called limiting reagent.
24. Write the general formula of molarity equation
Ans: Molarity (M) = No. of moles of solute
Volume of solution in litre
25. Define mole fraction
Ans: It is a ratio of no of moles of a particular component to the total number of moles of the solution.
26. Define molality
Ans: Molality is the number of moles of solute present in one Kg of solvent
27. Define mass percentage
Ans: It is defined as the mass of the component in 100 grams of the solution.
28. Define molarity
Ans: It is defined as the number of moles of solute present in one litre of solution.
29. Define normality
Ans: Normality can be defined as gram equivalent mass of the substance present in one dm^3 of the solution.
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30. Define mole
Ans: Mole can be defined as the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 gms of carbon – 12 isotope.
31. Write the equation to give relationship between ^oC and ^oF
Ans: $^oF = 9/5\ (^oC) + 32$

Two marks questions

1. What is the importance of chemistry in daily life ?

Ans: Chemistry plays an important role in meeting human needs for food, health care products and other materials aimed at improving the quality of life.

2. What is homogeneous mixture? Give an example.

Ans: Homogeneous mixture is a mixture in which the components are completely mix with each other and its composition is uniform throughout the solution

Ex: sugar solution, air, sodium chloride solution

3. What is heterogeneous mixture? Give an example.

Ans: It is a mixture in which the components are not completely mix with each other and its composition is not uniform throughout the solution.

Ex: mixture of salt and sugar, grains and pulses

4. Distinguish between pure substance and mixture

Pure Substance Mixture

1. Pure substance is composed of the same kind of particles

1. The composition of mixture is variable

2. The pure substance is homogeneous, irrespective of its origin

2. In mixture each of its components retains its characteristic properties.

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5. Define atom with an example.

Ans: Atom is the smallest particle of an element which may or may not be capable of independent existence.

Ex: Atoms of iron, copper, hydrogen

6. Define molecule with an example

Ans: A molecule is the smallest particle of an element or a compound which can exist freely

Ex: Hydrogen molecule (H_2), water molecule (H_2O)

7. Define physical properties of matter with an example

Ans: The property which can be measured or observed without changing the identity or the composition of the substance is called physical property.

Ex: Colour, odour, melting point and boiling point

8. Define Chemical properties of matter with an example.

Ans: Chemical properties are those in which a chemical change in the substance occurs.

Ex: Acidity, basacity, combustibility

9. Give the units of S.I. systems for the following a) Electric current b) Amount of substance

Ans: a) Electric current – Ampere (A)

b) Amount of substance – Mole (mol)

10. Define mass and weight

Ans: Mass of a substance is the amount of matter present in it and it is constant.

Weight is the force exerted by gravity on an object it changes one place to another due to change in the gravity.

11. Express the following into scientific notation

a) 0.00016 b) 33693.68

Ans: a) 0.00016 = 1.6×10^{-4} b) 33693.68 = 3.369368×10^{4}

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12. Define significant figures

Ans: The total number of digits in a number including the last digit whose value is uncertain is called the number of significant figures.

13. Define precision and accuracy

Ans: Precision refers to the closeness of various measurements for the same quantity. Accuracy is the agreement of a particular value to the true value of the result.

14. How many significant figures are present in the following ?

a) 6.005 b) 6.002×10^{23}

Ans: a) 6.005 = Four because the zeroes between the non zero digits are significant figures

b) 6.022×10^{23} = Four because the exponential term is not considered.

c) 4.01×10^{2} = 3 significant figures

15. State the law of conservation of mass and who proposed it ?

Ans: It states that matter can neither be created nor be destroyed. It was proposed by Antomic Lavoisier.

16. State Law of Definite proportion

Ans: It states that a given compound always contains exactly same proportion of elements by weight.

17. State Law of Multiple proportions

Ans: Law of multiple proportions can be defined as if two elements can combine to form more than one compound the masses of one element that combines with a fixed mass of the other element are in the ratio of small whole numbers.

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18. State Gay Lussac's law of Gaseous volumes

Ans: It can be defined as when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

19. State Aragadro Law

Ans: It can be defined as equal volumes of gases at the same temperature and pressure should contain equal no of molecules

20. What are isotopes? Mention the isotopes of carbon

Ans: Atoms having same atomic number with different mass number are called isotopes. Isotopes of carbon are ^{12}C, ^{13}C, ^{14}C

21. Define Aragadro Number and mention its values
Ans: The number of particles present in one mole of a substance
6.022×10^{23}
22. A piece of copper wire is 2.00 inch long. What is its length in centimeter
Ans: 1 inch = 2.54 cm
Therefore 2 inch = 2.54 x 2 = 5.08 Cms
1
23. A plastic Jug contains 3.5 Ltrs of milk calculate the volume of milk in meter ^{3}
Ans: 1 L = 1000 Cm^3
3.5 L = 1000 x 3.5 Cm^3 = 3500.0 Cm^3
$1m^3$ = 1mx1mx1m
= 10^2cm x 10^2cm x 10^2cm = 10^6 Cm^3
= 3.5 x 1000 x $1m^3$ / 10^6
= $3.5x10^3$ = 3.5 x 10^{-3} m^3
10^6
24. How many seconds are there in 3 days
Ans: 1 day = 24 hours
1 hour = 60 minutes
8
1 minute = 60 seconds
3 day = 3 x 24 Hrs x 60 min x 60 sec
= 259 200 seconds
25. Calculate the molecular mass of the following a) Ethane (C_2H_6)
b) Ammonia (NH_3)
Ans: a) Ethane (C_2H_6)
Ethane = C_2H_6 2 x atomic mass of carbon + 6 x atomic mass of Hydrogen
= 2 x (12.011u) + 6 x (1.008 U)
= 24.022u + 6.048u
= 30.070u
b) Ammonia (NH_3)
Ammonia=(NH_3) = 1x atomic mass of Nitrogen + 3 x atomic mass of Hydrogen
= 1x(14.01u) + 3 (1.008u)
= 14.01u + 3.024u = 17.034u
26. Calculate the formula mass of KCl (Potassium chloride)
Ans: formula mass of potassium chloride = Atomic mass of Potassium + Atomic mass of chlorine
39.10u + 35.5u = 74.60u
27. Calculate the no of molecules present in 2.5 moles of water (H_2O)
Ans: 1 mole of water = 6.022×10^{23}
Therefore 2.5 moles of water = $2.5x6.022x10^{23}$
1
= 15.055×10^{23} molecules
28. Calculate the percent (%) composition of elements in methanol (CH_3OH)
Ans: Molecular formula of methanol CH_3OH
Molecular mass of methanol = 1x12.01+4x1.008+ 1x16.0

= 32.042 gm
Percent composition of carbon = Mass of carbon x 100
Molecular mass of CH_3OH
= 12.01 X 100 = 37.48%
32.042
9
Percent composition of Hydrogen = Mass of Hydrogen x 100
Molecular mass of CH_3OH
= 4.032 x 100 = 12.5%
32.042
Percent composition of Oxygen = 16x100 = 49.93%
32.04
29. Calculate the amount of water in gm produced by the combustion of 8 gms of Ethane
Ans: Reaction $C_2H_6 + 3½ O_2 \_ 2CO_2 + 3H_2O$
(g) (g) (g) (g)
30 gms of ethane = 1 mole = 3 moles of H_2O
8 gms of ethane
1 mole of H_2O = 18 gm of H_2O
3 mole of H_2O = 18x3 = 54 gms of water
30 gms of ethane produced = 54 gms of water
Therefore 8 gms ethane produced = 14.4 gms
8 x 54 = 14.4 gms
30
30. How many moles of methane required to produce 88 gms of CO_2 after combustion
Ans: Combustion reaction
$CH_4 + 2O_2 \_ CO_2 + 2H_2O$
(g) (g) (g) (g)
16 gms of Methane = 1 mole = 44 gms of CO_2
44 gms of CO_2 is produced from 1 mole of methane
88 gms of CO_2 is produced from 88 x 1 = 2 moles
44
31. 25 Kg of N_2 (gm) and 5 Kg of H_2 are mixed to produce Ammonia (NH_3). Calculate the ammonia formed
Ans: Equation $N_2+3H_2 \_ 2NH_3$
(g) (g) (g)
10
= 25 Kg N_2 + 5 Kg H_2 _ NH_3
1 Kg of N_2 = 1000gm N_2
25 Kg of N_2 = 25 x 1000 = 25000/28 = 892.85 mol
1 Kg of H_2 = 1000 gm H_2
5 Kg of H_2 = 5x 1000 = 5000/2.016 = 2480.15 mol
According to above equation 1 mol of $N_{2(g)}$ requires 3 moles of $H_{2(g)}$
Hence 892.85 mole of N_2 and the mols of H_2 required would be
892.85 x 3 mol of H_2 / 1 mol N_2
= 2678.55 mol = 2.67855 x 10_3 mol H_2

But we have 2480.15 mol of H_2. Hence in this process 2678.55
mol of H_2 is required.
3 moles of $H_2(g)$ _ 2 moles of NH_3
2480.15 mole of H_2 _ 1653.43 mole of NH_3 gas is formed
2480.15 x 2 = 1653.43 mol of NH_3
3
32. A solution is prepared by adding 4.00 gm of a substance A to 18 gm of water calculate the mass percent of the solute.
Ans: Mass % of solute (A) = Mass of solute (A) x 100
Mass of solution
= 4 x 100 = 400 = 18.18 %
4+18g H_2O 22gm
Mass % of solute = 18.18%
33. Calculate the molarity of sodium Hydroxide (NaOH) in the solution prepared by dissolving 4 gm in 500ml of the solution.
Ans : Molarity = No of moles of solute
Volume of solution in litre
= Mass of NaOH/molar mass of NaOH
0.5 Litre
= 4/40 = 0.1/0.5 = 1/5 = 0.2 mol / litre
0.5
11
34. The density of 2M solution of sodium chloride (NaCl) 1.13 g mol/litre. Calculate molality of the solution. M= 2 mol per litre.
Ans: Mass of Nacl in 1 litre solution = 2 x 58.5
= 117.0 g
Mass of 1 litre solution = 1000 x 1.13 = 1130 g
Mass of water in solution = 1130 g – 117.0 g = 1013 g
Molality = No of moles of solute
Mass of solvent in Kg
= 2 mol / 1.013 Kg = 1.9743 Molality
4 Marks questions
1. Write any four postulates of Daltons atomic theory
Ans: Dalton published a new system of chemical philosophy in 1808 in which the proposed the following:
1) Matter consists of indivisible atoms
2) All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
3) Compounds are formed when Atoms of different elements combine in a fixed ratio.
4) Chemical reactions involve reorganization of atoms these are neither created nor destroyed in a chemical reaction.
2. A Organic compound contain 57.14% of carbon, 6.16% Hydrogen, 9.52% Nitrogen 27.18% oxygen. Calculate the empirical formula and molecular formula. If its molecular mass is 294.3 gm/mole.
12

Ans:
Element % At mass % at mass Nearest
whole no
Carbon 57.14 12 57.14/12 4.76 4.76/0.68=7
Hydrogen 6.16 1 6.16/1 6.16 6.16/0.68=9.06
Nitrogen 9.52 14 9.52/14 0.68 0.68/0.68=1
Oxygen 27.18 16 27.18/16 1.698 1.698/0.68=2.45
Empirical formula = $C_7H_9N_1O$
Empirical formula mass = $C_7H_9N_1O_{2.5}$
= 12x7 + 1x9 + 1x14 + 2.5x16
= 84+9+14+40
= 147
Molecular formula mass = 294.3
Molecular formula = Empirical formula x n
n = Molecular formula mass = 294.3
Empirical formula 147
n = 2
Therefore Molecular formula = Empirical formula x n
= $(C_7H_9N_1O_2)_2$
= $C_{14}H_{18}N_2O_4$
3. Compound contains 4.07% Hydrogen 24.27% Carbon and 71.65% chlorine. Its molecular mass is 98.96 gm what are its empirical formula and molecular formula?
Ans:
Element Symbol % of
element
At mass of
element
Moles of the
element = %
Simpler molar
mass
Hydrogen H 4.07 1 4.07/1 = 4.07 4.04/2.018=2.01
Carbon C 24.27 12 24.27/12
= 2.022
2.022/2.018
=1.0019
Chlorine Cl 71.65 35.5 71.65/35.5
= 2.018
2.018/2.018 = 1
13
Therefore Empirical formula = H_2C Cl = CH_2Cl
Empirical formula weight = 1x12 + 2x1 + 35.5x1
= 12+2+35.5 = 49.5
Molecular formula = Empirical formula x n
Therefore n = 98.96 = 2
49.5
Therefore molecular formula = $(CH_2Cl)_2 = C_2H_4Cl_2$

4. An organic substance containing carbon Hydrogen and oxygen gave the percentage composition as C=40.687 % H=5.085 % and 0=54.228%
The vapour density of the compound is 59 calculate the molecular formula of the compound.

Ans: Solution = Step 1

Element	Symbol	% of element	At mass of element	Moles of the element = %	Simplest molar ratio	Simplest whole No. multiplied by 2
Carbon	C	40.687	12	40.687/12 = 3.390	3.390/3.389 = 1	2
Hydrogen	H	5.085	1	5.085/1 = 5.085	5.085/3.389=1.5	3
Oxygen	O	54.228	16	54.228/16 = 3.389	3.389/3.389=1	2

Empirical formula $C_2H_3O_2$

Step 2

To calculate the empirical formula mass

Empirical formula $C_2H_3O_2$

Therefore Empirical formula mass 2x12+3x1+2x16 = 59 14

Step 3

To calculate the molecular mass of the compound

The vapour density of the compound = 59

Molecular mass = Vapour density x 2

= 59 x 2 = 118

Step 4

To calculate the value of ‘n’

n = molecular mass = 118 = 2

Empirical formula mass 59

Step 5

Molecular formula = Empirical formula x n

= $C_2H_3O_2$ x 2

= $C_4H_6O_4$

Therefore Molecular formula is $C_4H_6O_4$

CHAPTER 2: STRUCTURE OF ATOM

NUMBER OF TEACHING HOURS: 10 HOURS
MARKS WEIGHTAGE: 09 (1 M, 2M, and 4M)
QUESTIONS CARRYING ONE MARK:

1. Name the person who first proposed the atomic theory of matter on scientific basis.
2. Mention the sub atomic particles.
3. What is the basic rule, regarding the behaviour of charged particles?
4. Under what conditions of pressure and voltage, the electrical discharge through the gases can be observed?
5. How the pressure of gases in the discharge tube can be adjusted?
6. What are cathode rays(cathode ray particles)?
7. Name the phosphorescent material coated inside the discharge tube behind the anode.
8. Give the conditions under which cathode rays travel in straight line.
9. What is the name given to the particles which constitute the cathode rays?
10. Does the nature of cathode ray depend on the nature of gas in the discharge tube or
the electrode material?
11. Name the scientist who was able to determine e/m value of an electron.
12. What is the charge of an electron in coulomb?
13. What is the mass of an electron in terms of kilogram?
14. What are canal rays?

15. Name the fundamental particle of an atom that has highest value for its e/m value.
16. Does the e/m value of canal rays depend on the nature of gas in the discharge tube?
17. Name the gas to be filled in the discharge tube to obtain the smallest and the lightest positive ion.
18. Name the smallest and lightest positive ion obtained when hydrogen gas is subjected to electrical discharge.
19. Who discovered neutron?
20. What is the mass of proton in terms of kilogram?
21. What is the mass of neutron in terms of kilogram?
22. Name the electrically neutral particle obtained by bombarding beryllium with _-particles.
23. Who proposed the spherical shape of an atom?
24. Write one name given to the Thomson model of an atom.
25. What is the important feature of Thomson model of an atom?
26. Name the metal foil used in Rutherford's _-particle scattering experiment.
27. What are orbits?
28. What is atomic number?
29. What are nucleons?
30. What is atomic mass number?
31. What are isotopes?
32. What are isobars?
33. Name the three isotopes of hydrogen.
34. ${}_{Z}X$: what does A and Z represent?
35. How many neutrons are present in ${}_{17}Cl$?
36. How many protons are present in ${}_{6}C$?
37. How many electrons are present in ${}_{11}Na$?
38. Name the fundamental particle of an atom which determines the chemical properties of an element.
39. Write the SI unit of frequency.
40. What is the value of speed of light in vacuum?
41. Define wave number.
42. What is the SI unit of wave number?
43. What is the relation between energy (E) and frequency (_) of an electromagnetic radiation?
44. What is the name given to the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation?
45. What is the wave length range of the visible spectrum?
46. What is the value of Planck's constant?
47. What is black body?
48. What is photoelectric effect?
49. Give an example of a metal which exhibits photoelectric effect.

50. What is threshold frequency?
51. What is emission spectrum?
52. What is spectroscopy?
53. Name an element which is identified by spectroscopic method.
54. What is the value of Rydberg's constant in joule?
55. Name the series of spectral line of hydrogen obtained in visible region.
56. Write the mathematical expression for the calculation of angular momentum of an electron in a given stationary state.
57. Write an expression for the calculation of energy associated with an electron in a given stationary state.
58. Write an expression to calculate the wave number of a spectral line in the hydrogen spectrum.
59. What is the value of the radius of the first stationary state (Bohr orbit)?
60. Give an expression to calculate the radii of the stationary states.
61. Write the relationship between wave length (_), velocity (c) and frequency (_) of a radiation.
62. What is Zeeman effect?
63. What is Stark effect?
64. Write de Broglie equation.
65. Write the mathematical form of Heisenberg's uncertainty principle.
66. What is the implication of uncertainty principle?
67. What does quantum mechanics deal with?
68. What is Schrödinger equation?
69. What is an atomic orbital?
70. What are quantum numbers?
71. What is the significance of (a) principal quantum number(n); (b) azimuthal quantum number (l); (c) magnetic quantum number (m_l); (d) spin quantum number (m_s)?
72. What are the possible values of n (principal quantum number)?
73. What are the possible values of 'l' (azimuthal quantum number) for a given value of 'n'?
74. What are the possible values for m_l (magnetic quantum number) for a given value of 'l'?
75. What are the possible values of m_s (spin quantum number)?
76. What is the total value of m_l (magnetic quantum number) for a given value of 'l'?
77. What is the value of 'l' for:- (a) s- sub shell; (b) p- sub shell; (c) d-sub shell; (d) f-sub shell; (e) g-sub shell; (f) h- sub shell?
78. Write the possible values of magnetic quantum number (m_l) for l= 2.
79. Name the quantum number that specifies the shape of an atomic orbital.
80. Name the quantum number that specifies the size of an atomic orbital.
81. Name the quantum number that designates the orientation of the atomic orbital.
82. What are nodes?
83. Write the total number of nodes for a given value of 'n'.

84. The total number of nodes for 3s orbital is ----------
85. What is the shape of: (a) s-orbital: (b) p- orbital; (c) d-orbital?
86. On the basis of orientation, how the p-orbitals are designated?
87. On the basis of orientation, how the d-orbitals are designated?
88. What is the maximum number of electrons that can be accommodated in: (a) sorbital:
(b) p- orbitals; (c) d-orbitals; (d) f-orbitals?
89. How many number of orbital are possible for f-sub shell?
90. Using s, p, d, notations, write the orbitals having following quantum numbers: (a) n=4, l=0; (b) n=5, l=1; (c) n=3, l= 2.
91. What are degenerate orbitals?
92. What is the ground state of an atom?
93. What is the excited state of an atom?
94. What is effective nuclear charge?
95. How many electrons in an atom may have the following quantum numbers: n=3, l= 0?
96. What is electronic configuration?
97. Write the electronic configuration of the following elements: (a) Cr (Z=24) (b) Cu (Z=29) (c) Ca (Z=20).
98. Write the electronic configuration of the following: (a) Na^{+} (b) Cl^{-} (c) O^{-2}.
99. Between 3d and 4s orbitals which is having higher energy?
100. How many unpaired electrons are present in the following: (a) Na (b) P (c) O?

Two marks Questions

101. State Heisenberg's uncertainty principle.
102. Name the four quantum numbers.
103. Write any two limitations of Bohr's model of an atom.
104. Distinguish between orbit and orbital.
105. Draw the shape of .a) s – orbital. b) p – orbital.
106. State aufbau principle.
107. State Pauli's exclusion principle.
108. State Hund's rule of maximum multiplicity.
109. What is an orbital? Mention different orbitals.
110. Write the schematic diagram to remember sequences of filling atomic orbitals.
111. Name two series of hydrogen spectra which fall in infra red region.
112. Name the series of lines in the emission spectrum of hydrogen.
113. Write de Broglie equation and explain the terms.
114. Write Rydberg's equation and explain the terms.
115. Write two draw backs of Rutherford's model of an atom.
116. State (n+l) rule.
117. Calculate the wave length of the radiation emitted with a frequency of 1,200kHz ($c = 3.0 \times 10^{8} m/s$)
118. Calculate the wave number of radiation having wavelength $5800 A^{0}$.
119. Calculate the energy of one mole of photon of radiation whose frequency is 4

$X10^{12}$ Hz.

120. The threshold frequency ν_0 for a metal is 6.0×10^{13} s^{-1}. Calculate the kinetic energy of an electron when the radiation of frequency $\nu = 1.0 \times 10^{14}$ s^{-1} hits the metal.

121. What will be the wavelength of a ball of mass 0.2 kg moving with velocity of 10 ms^{-1}?

122. Calculate the wave number of the spectral line of shortest wavelength appearing in the Balmer series of hydrogen spectrum (Given $R_H = 1.09 \times 10^7$ m^{-1})

Four Mark Question

123. What are the results drawn from the Cathode ray discharge experiment?

124. What are the Characteristics of Canal Rays?

125. What are the observations made out of Ruther ford's α-ray scattering experiment?

126. What are the conclusions drawn regarding the structure of the atom on the basis of observations in the α-ray scattering experiment?

127. Describe Rutherford's nuclear model of the atom.

128. What are the properties of electromagnetic waves (electromagnetic radiation)?

129. What are the factors that cannot be explained by electromagnetic theory?

130. What are the observations made by Hertz after conducting the photo electric effect experiment?

131. What are the postulates of Bohr's model of hydrogen atom?

132. Write the significance of the four quantum numbers.

Answers

Chapter 2: Structure of atom

1. John Dalton
2. Protons, Neutrons and electrons
3. 'Like charges repel each other and unlike charges attract each other'.
4. At very low pressure and at very high voltage
5. By using vacuum pump
6. The particles moving in the discharge tube from cathode to anode.
7. Zinc sulphide
8. In the absence of electrical or magnetic field.
9. Electrons
10. No. Neither nature of the gas nor electrode material
11. J.J. Thomson.
12. -1.6×10^{-19}C.
13. 9.1094×10^{-31}Kg.
14. The particles carrying positive charge in a discharge tube.
15. Electron.

16. Yes.
17. Hydrogen.
18. Proton.
19. James Chad Wick.
20. 1.672×10^{-27} Kg.
21. 1.675×10^{-27} Kg.
22. Neutron.
23. J. J. Thomson.
24. Plum pudding or raisin pudding or watermelon.
25. Mass of the atom is uniformly distributed over the atom.
26. Gold.
27. The circular path of an electron moving around the nucleus of an atom.
28. No. of Protons present in the nucleus or No. of electrons in a neutral atom.
29. Protons and neutrons present in the Nucleus of an atom.
30. Total No. of Protons and neutrons present in the nucleus of an atom.
31. Atoms of the same element having identical atomic Number but different mass No.
32. Atoms of different elements having same mass No. but different atomic No.
33. Protium, deuterium and Tritium.
34. A= Mass No, Z= Atomic No.
35. 18.
36. 6.
37. 11.
38. Electron
39. Hertz2
40. 3.0×10^{8} m/s.
41. Number of wave lengths per unit length is called wave number (_).
42. m^{-1} or per meter
43. E =h_
44. Proton.
45. 400 nm to 750 nm
46. 6.626×10^{-34} Js.
47. The ideal body which emits and absorbs radiations of all frequency.
48. The ejection of electrons from metal surface when radiation strikes it.
49. Potassium or Caesium or Rubidium.
50. The minimum frequency below which photoelectric effect is not observed.
51. The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.
52. The study of emission or absorption spectra is known as spectroscopy.
53. Rubidium or Caesium, or Thallium or Indium, or Gallium or Scandium or Helium.
54. 2.18×10^{-18} J.

55. Balmer series
56. $m_e v r = n h$

2π

57. $E_n = -R_H \frac{1}{n^2}$

58. $\frac{1}{\lambda} = R_H \left(\frac{1}{n1^2} - \frac{1}{n2^2}\right) cm^{-1}$

59. $a_0 = 52.9$ Pm.
60. $r_n = n^2 a_0$
61. C= $\frac{\lambda}{\nu}$
62. The splitting of spectral lines in the presence of magnetic field.
63. The splitting of spectral lines in an electric field.
64. $\lambda = \frac{h}{mv} = \frac{h}{p}$
65.
66. It rules out existence of definite paths of electrons.
67. It deals with dual behavior of matter.
68. $H \psi = E \psi$
69. It is the region around the nucleus where electron will most probably be found.
70. Quantum Numbers specify the energy, size, shape and orientation of an orbital.
71. a) It determines the size and energy of the orbital.
b) It determines the three dimensional shape of the orbital.
c) It gives the spatial orientation of the orbital.
d) It refers to orientation of the spin of the electron.
72. n=1,2,3 - - - - - ∞.
73. l=0,1,2,3 - - - - (n-1).
74. m=-l, -(l-1), -(l-2)- - - - 0,1 - - - (l-2), (l-1),l
75. + 1 and - 1

2 2

76. (2l+1) values.
77. a) l=0, b) l=1, c) l=2, d) l=3, e) l=4, f) l=5
78. m_l = -2, -1, 0, + 1, + 2.
79. Azimuthal Q.no(l)
80. Principal Q. no (n)
81. Magnetic Q.no (m_l)
82. The region where probability density function reduces to zero.
83. Total no. of nodes = (n-1).
84. No. of nodes = 2
85. a) Spherical b) dumb bell c) double dumb – bell.
86. p_x , p_y and p_z
87. d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$

$_2$ and d_{z^2}
88. a) 2 b) 6 c) 10 d) 14
89. 7
90. a) 4s b) 5p c) 3d
91. Orbital's having the same energy.
92. Atom having electrons occupying lowest energy level.
93. The state of an atom having its electron in higher energy level.
94. The net positive charge experienced by the outer electrons.
95. 2 electrons.
96. The distribution of electrons into orbitals of an atom.
97. a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 3d^5\ 4s^1$ b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 3d^{10}\ 4s^1$
c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
98. a) $1s^2\ 2s^2\ 2p^6$ b) $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6$ c) $1s^2\ 2s^2\ 2p^6$
99. 3d
100. a) 1 b) 3 c) 2

Answers to two marks questions

101. It is impossible to determine simultaneously the exact position and exact momentum of an electron."
102. a. Principal Q. Number (n)
b. Azimuthal Q. Number (l)
c. Magnetic Q. Number (ml)
d. Spin Q. Number (ms)
103. It fails to accounts for the finer details of the hydrogen atom spectrum. It could not explain the ability of atom to form molecules by chemical bonds
104.

Orbit	Orbital
1 It is a circular path around the nucleus in which an electron moves.	1 It is the three dimensional region of space where the probability of finding the electron is maximum
2 Maximum number of electrons that can be accommodated is equal to $2n^2$	2 Maximum number of electrons in an orbital is 2.

105.
S – orbital p – Orbital
106. " In the ground state of the atom orbitals are filled in the order of their increasing energies".
107. " No two electrons in an atom can have the same set of four quantum

numbers".
108. " Pairing of electrons in the orbital's belonging to the same sub shell does not takes place until each orbital belonging to that sub shell has got one electron each."
109. " It is the three dimensional region of space where the probability of finding the electron is maximum."
The different orbitals are s, p, d, and f
110.
111. Bracket series
Paschen series
Pfund series (any two)
112. Lyman series
Balmer Series
Bracket series
Paschen series
Pfund series
113. De Broglie Equation $\lambda = \frac{h}{mv}$
Where m is the mass of the particle?
v is the velocity of the particle
114. $\bar{\nu} = R_H \left(1 - \frac{1}{n_1^2} \frac{1}{n_2^2}\right)$ cm^{-1}
$\bar{\nu}$ - wave number
R_H – Rydberg constant
115. a. This model fails to account for the stability of an atom.
b. This model does not explain the line spectrum of an atom.
Or
This model does not have any specific radius for radius electron orbit.
116. " Lower the value of (n+l) for an orbital , the lower is its energy . If two orbital's have the same value of (n + l) , the orbital with lower value of n will have the lower energy."
117. $\bar{\nu} = \frac{C}{\lambda}$

$\bar{\nu} = 3 \times 10^8$ ms^{-1}
1200×10^3 s^{-1}
$\lambda = 0.0025 \times 10^5$ m
$\lambda = 250$ m
118. Calculation of wave number
$\bar{\nu} = 5800$ $A^0 = 5800 \times 10^{-8}$ cm $= 5800 \times 10^{-10}$ m
$\bar{\nu} = 1 = 1/ 580 \times 10^{-10}$ m

$= 1.724 \times 10^6$ m^{-1}

119. Energy of photon E = h _
h = 6.626 X 10^{-34} JS
_ = 4 X 10 12 Hz (s^{-1})
E = 6.626 X 10^{-34} Js X 4 X 10^{12} S^{-1}
= 26.504 X 10^{-22}
= 2.65 X 10^{-21} J
120. Kinetic Energy = h (_ – $_o$)
= (6.626 X 10^{-34} Js) (1.0x10^{14} s^{-1}– 6.0 X 10^{13} s^{-1})
= (6.626 X 10^{-34} Js) (4.0 X 10^{13} s^{-1})
= 2.65 X 10^{-20}J
121. _ = h
m_
= (6.626 X 10^{-34} Js)/ (0.2 Kg X 10 ms^{-1})
= 3.313 X 10^{-34} m
122.
_ = R_H (1/n_1
^{2} – 1/n_2
^{2})
For Balmer series with shortest wavelength
n_1 = 2 n_2 = _
R_H = 1.09 X 10^7 m^{-1}
_ = 1.09 X 10^7 (1/2^2 – 1/ _ ^{2})
= 1.09 X 10^7 (¼ - 0)
_ = 2.725 X 10^6 m^{-1}
123. .
i. The Cathode rays start form cathode and move towards anode.
ii. Cathode rays are not visible.
iii. In the absence of electrical or Magnetic field, Cathode rays travel in straight lines.
iv. In the presence of electrical or magnetic field, the direction of deflection of cathode rays shows that they contain negatively charged particles.
v. The characteristics of cathode rays do not depend on the material of electrodes and nature of the gas present in the cathode ray tube.
(Any Four Points)
124.
i. Positively charged particles (Canal Rays) depend upon the nature of the gas present in the cathode-ray tube.
ii. The charge to mass ratio of the particles depend on the gas from which they originate.
iii. Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.
iv. In the presence of electrical or magnetic field, the behavior of positively

charged particles is opposite to that observed for cathode rays.
125.
i. Most of the _–particles passed through the gold foil undeflected.
ii. A small fraction of the _–Particles was deflected by small angles.
iii. A very few _-particles bounced back, that is were defected nearly 180^0
126.
i. Most of the space in the atom is empty
ii. The Positive charge of the atom is not spread through out the atom, but concentrated in a very small volume
iii. The volume occupied by the nucleus is negligibly small as compared to the total volume of the atom
127.
i. The positive charge and most of the mass of the atom is concentrated in a small region called nucleus
ii. The nucleus is surrounded by Electrons
iii. The Electrons move around the nucleus in circular paths called orbits.
iv. The Electrons and the nucleus are held together by Electrostatic force of attraction
128.
i. Electrical and Magnetic waves are perpendicular to each other and both are perpendicular to the direction of the propagation of the wave.
ii. Electromagnetic waves do not require medium and they can move in vacuum
iii. There are many types of electromagnetic radiations. They differ from one another in wavelength or frequency
iv. Different kinds of units are used to represent Electromagnetic
v. radiation.
129.
i. The nature of emission of radiation from hot bodies (black body radiation)
ii. Ejection of electrons from metal surface when radiation strikes it (Photo electric effect)
iii. Variation of heat capacity of solids as a function of temperature
iv. Line spectra of atoms with special reference to hydrogen
130.
i. The electrons are ejected from the metal surface as soon as the beam of light strikes the surface
ii. The number of electrons ejected is proportional to the intensity of light
iii. For each metal there is a characteristic minimum frequency called threshold frequency below which photo electric effect is not observed.
iv. The kinetic energies of ejected electrons increase with increase of frequency of light used.
131.
i. The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy and the paths are called orbits

ii. The energy of an electron in the orbit does not change with time
iii. The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by _ E is given by

$$\_ = \frac{\_ E}{h}$$

$$\_ = \frac{E_2 - E_1}{h}$$

iv. The angular momentum of an electron in a given stationary state is integral multiple of h

$$\text{i.e., } m_e V_r = n \frac{h}{2\_}$$

132.
i. Principal quantum number determines energy and size of the orbital
ii. Azimuthal quantum number defines three dimensional shape of the orbital.
iii. Magnetic quantum number gives the information about the spatial orientation of the orbital.
iv. Spin quantum number refers to orientation of the spin of the electron.

UNIT-3
Classification of elements and periodicity in properties
One mark questions:
1. For the triad of elements A, B and C if the atomic weights of A and C are 7 and 39.
Predict the atomic weight of B.

2. “Every eighth element has property similar to the 1st element when placed (arranged) in increasing order of their atomic weight”. Name the law for the above statement.
3. Which property of the element was the basis for the classification of elements by
Mendeleev?
4. State Mendeleev’s periodic law.
5. Name the scientist whose experiment on x-ray spectra of elements led to modern
periodic law.
6. State modern periodic law.
7. What is more fundamental property for an atom of an element according to Moseley?
8. Which quantum number corresponds to the period number in the modern periodic
table?
9. How many elements are there in the 4th period of long form of periodic table?
10. Write the atomic number of the element unniltrium.
11. Give the IUPAC name of the element whose atomic number is 109?
12. Which one of the following subshell is not filled in the 5th period (5s, 5p, 5d, 4d)?
13. In which period does the lanthanoids appear?
14. In which period does the actinide series of elements appear?
15. Name the series of inner transition element found in the 7th period.
16. How many elements are in lanthanide series?
17. Elements of a group have similar chemical properties. Why?
18. The position of helium is in 18th group of p block and not in 2nd group of s block of
long form of periodic table. Justify the statement.
19. To which block of the periodic table do the elements of group-I and II belong?
20. How many groups of elements form p block of the periodic table?
21. What are ‘representative elements’?
22. Which group of elements are called ‘chalcogens’?
23. Write the general outer electronic configuration of d block elements.
24. Which block of elements are more known for exhibiting paramagnetism and catalytic properties?
25. A metal X forms coloured ions, is paramagnetic and is used as a catalyst. Predict
the block to which the metal belongs.
26. Why are d block elements also called transition elements?
27. Write the general outer electronic configuration of f block elements.

28. What are transuranium elements?
29. What is the trend in the metallic character of elements down a group?
30. In bromine molecule (Br_2) the bond distance is 228 pm. What is the covalent radius of bromine?
31. What is covalent radius for a non-metal?
32. Define metallic radius.
33. Arrange the following in the decreasing order of their atomic radius:
K Be N Na
(Z=19) (Z=4) (Z=7) (Z=11)
34. For noble gases covalent radius cannot be defined. Why?
35. Name a group of non-metals for which covalent radius cannot be defined.
36. Size of the cation is smaller than the parent atom. Give reason.
37. Size of the anion is larger than the parent atom. Give reason.
38. Hydride ion (H^-) is larger than hydrogen atom. Give reason.
39. Mention a species that is isoelectronic with neon.
40. Arrange the following in the increasing order of their ionic radii: F^-, N^{-3}, Na^+, O^{-2}.

41. $X^{+}_{(s)} \square\square \rightarrow X^{+2}_{(g)} + e^-$. Correct the equation to represent the 2nd ionisation enthalpy of the element X.
42. Define second ionisation enthalpy.
43. 2nd ionisation enthalpy for an element is higher than the 1st ionisation enthalpy. Give reason.
44. Using the graph, match the ionisation energy of Ar, Na, Ne, Li
with positions P, Q, R and S.
45. Define electron gain enthalpy.
46. Why is ionisation enthalpy always positive?
47. Arrange the below elements X, Y, Z in the increasing order of electron gain enthalpy.
Element Electron gain enthalpy kJ/mol
X
Y
Z

−48

+77

−295

48. What is electronegativity for an atom?
49. “Ability of an atom in a compound to attract the shared electrons to itself”. Which
periodic property does the above quote represent?

50. On the basis of Pauling scale what is the value of electronegativity of fluorine?
51. How is electronegativity related to non-metallic property of an element?
52. Give reason for the anamalous properties of 2^{nd} period elements when compared to
the elements in their respective groups.
53. Mention one property common to all actinoid elements.
54. How are H^- and He related?
55. By what name do we know the 17^{th} group elements?
56. Isoelectronic species do not have the same size. Why?
57. Arrange the following in the increasing order of their metallic character: Cu, K,
Ge, Br.
58. What is the difference between an amphoteric oxide and a neutral oxide?
59. Give an example for a basic oxide.
60. Give an example for a neutral oxide.
61. Name the element that is diagonally related to beryllium.
62. Why do 17^{th} group elements have high negative electron gain enthalpy?

Two Marks questions:

1. Mention one merit and one drawback of Mendeleev's periodic table.
2. What was the name given by Mendeleev to the element if existed and had properties similar to that of the aluminium? What is the present name of the element?
3. What observation made by Moseley showed that atomic number and not atomic mass is more fundamental property of an element?
4. With respect to long form of periodic table what are groups and periods?
5. How many groups and periods are present in the long form of periodic table?
6. How many elements are there in 2^{nd} period? Justify your answer.
7. Which is the i) shortest ii) longest period in the long form of periodic table?
8. What are the subshells filled in i) 2^{nd} period ii) 4^{th} period?
9. Hydrogen is placed separately at the top of the long form periodic table. Justify its
position giving two reasons.
10. The electronic configuration of an element is [Ar] $3d^7\,4s^2$. To which block and period does it belong?
11. Write the electronic configuration of the element with atomic number 118. Predict
the group the element belongs to.
12. Mention any two differences in the properties of metals and non-metals.
13. What are metalloids? Give an example.
14. Determination of size of an atom cannot be precise. Give reason.
15. How does atomic radius vary along a period and down a group in the periodic

table.
16. Explain why atomic size decreases along a period and increases down a group.
17. Given: a) F, Cl, Br b) N, O, F : For which one of these set of elements does the atomic radius increase? Justify your answer.
18. What are isoelectronic species? Give example.
19. Which among the following has i)largest ii)smallest ionic radii:Na^{+},Al^{+3},Mg^{+2}, F^{-}?
20. Define ionisation enthalpy. Give its SI unit.
21. Compare the 1st and 2nd ionisation enthalpies of magnesium with reason.
22. Which group of elements have maximum value for ionisation enthalpy. Give reason.
23. Draw a graph showing the variation of ionisation enthalpy with atomic number for
the elements Li, Be, B and C.
24. How does ionisation enthalpy vary along a period? Give reason.
25. How does ionisation enthalpy vary down a group? Give reason.
26. Ionisation enthalpy of boron is less than that of beryllium. Give reason.
27. Why is ionisation enthalpy of oxygen less than that of nitrogen?
28. Ionisation enthalpy of aluminium is less than that of magnesium. Justify the statement.
29. Give two differences between ionisation enthalpy and electron gain enthalpy for an
element.
30. How does electron gain enthalpy vary along a period? Give reason.
31. How does electron gain enthalpy vary down the group? Explain.
32. Between flourine and chlorine which one of these has low electron gain enthalpy.
Give reason.
33. Which among the following has the most positive and the most negative electron
gain enthalpy: O, Ar, Na, S?
34. How does electronegativity vary along a period? Give reason.
35. How does electronegativity vary down a group? Give reason.
36. Name any two periodic properties that increases along a period.
37. Name any two periodic properties that decreases down a group.
38. Arrange the elements in the following two sets in the increasing order of their electronegativity values: a) [Cl, F, I, Br] b) [S, Cl, Al, Na]
39. Which one of these is possible: $_4BF^-$ or $_6BF^-_3$. Justify your answer. (atomic number
of boron = 5)
40. Why is the chemical reactivity of elements at the two extremes (except noble

gases) of the periodic table very high?
41. Why are the elements placed in the extreme left of the periodic table most metallic?
42. Mention two factors on which ionisation enthalpy depends.
43. Mention one exception (anomaly) each found in the variation of i) ionisation enthalpy in the 2nd period ii) electron gain enthalpy in the 16th period.
44. Classify these into acidic, basic, amphoteric and neutral oxide: CO, Na_2O, Cl_2O_7,
Al_2O_3.

Four Marks:

1. a) What is the valence of 2nd group elements?
b) Using the table given below, evolve the formulae of i) aluminium oxide,
ii) aluminium nitride, iii) aluminium fluoride:

Group No.	13	15	16	17
Element	Al	N	O	F

2. a) When is a cation and an anion is formed from an atom?
b) Arrange H^+, H^-, H in decreasing order of their size.
c) What does the energy absorbed in the reaction represent?

$X_{(g)} \square\square \rightarrow {}_{(g)}X^+ + e^-$

3. What does ionisation enthalpy, electron gain enthalpy and electronegativity measure for an atom? Which one of these is not a measurable quantity?
4. Given:

Element	Electronic configuration
P	$1s^2\ 2s^2$
Q	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
R	$1s^2\ 2s^2\ 2p^6\ 3s^1$
S	$1s^2\ 2s^2\ 2p^4$

i) Arrange P, Q, R, S in decreasing order of their atomic radii.
ii) Which one among these is a chalcogen?
iii) Which one among these has more positive electron gain enthalpy?
iv) Give the formula of the oxide formed from R and S.
5. Mention any two characteristic properties each for s and p block elements.
6. How does metallic and non-metallic character vary along a period and down a group. Give reasons.
7. Given the outer electronic configurations of A and B as .... $3s^2$ and .... $3s^2\ 3p^5$.
a) Locate their position in the periodic table (group and period)

b) Which one of these has i) larger atomic size ii) higher ionisation energy

Scheme of valuation with complete answers

One mark questions:

Qn.No.	**Value points**	**Marks allotted**
1	23	1
2	Law of octaves	1
3	Atomic weight or atomic mass	1
4	Properties of the elements are periodic functions of their atomic weight	1
5	Moseley or Henry Moseley	1
6	Physical and chemical properties or properties of the elements are periodic functions of their atomic numbers	1
7	Atomic number	1
8	Principal quantum number	1
9	18	1
10	103	1
11	Unnilennium	1
12	5d	1
13	6^{th} period or 6	1
14	7^{th} period or 7	1
15	Actinoids or actinide series	1
16	14	1
17	Similar outermost electronic configuration or same number and same distribution of electrons in their outermost orbital	1
18	It has completely filled value shell ($1s^2$) and has properties characteristic of noble gases	1
19	s block	1
20	6	1
21	s and p block	1
22	16^{th} group or 16	1
23	$(n-1)\, d^{1-10}\, ns^{0-2}$	1
24	d – block	1
25	d – block	1
26	They act as a bridge between chemically most active metals of sblock and less active elements of groups 13 and 14.	1
27	$(n-2)\, f^{1-14}\, (n-1)\, d^{0-1}\, ns^2$	1
28	Elements appearing after uranium	1
29	Metallic character increases	1
30	114 pm	1

31 It is one half the distance between two atoms bonded by a single covalent bond 1
32 It is half the internuclear distance separating the metal cores in the metallic crystal 1
33 K > Na > Be > N 1
34 They are monoatomic gases 1
35 Noble gases 1
36 It has fewer electrons and nuclear charge remains same 1
37 It has more electrons and effective nuclear charge decreases 1
38 In H^- due to more number of electrons the effect of nuclear charge is less 1
39 Na^+ (or Mg^{+2}, Al^{+3} or any other) 1
40 $Na^+ < F^- < O^{-2} < N^{-3}$ 1
41 $_{(g)\ (g)}$ $X^+ \square \square \rightarrow X^+ + e^-$ 1
42 It is the energy required to remove the 2^{nd} most loosely bound electron 1
43 It is more difficult to remove an electron from a positively charged ion than from a neutral atom 1
44 P = Li, Q = Ne, R = Na, S = Ar 1
45 It is the enthalpy change that occurs when an electron is added to a neutral gaseous atom 1
46 It is an endothermic process or energy is always absorbed to remove an electron 1
47 Y < X < Z 1
48 It is the ability of an atom in a compound to attract the shared electrons to itself 1
49 Electronegativity 1
50 4 1
51 Electronegativity α non-metallic character 1
52 Same size, large charge / radius ratio, high electronegativity (any two) 1
53 Radioactivity 1
54 Isoelectronic 1
55 Halogens 1
56 They have different nuclear charge 1
57 Br < Ge < Cu < K 1
58 Amphoteric oxide has both acidic and basic character. But a neutral oxide is neither acidic nor basic. 1

59 Na_2O 1
60 CO 1
61 Aluminium 1
62 By gaining one electron they attain noble gas configuration 1

Two mark questions:

Qn.No. Value points Marks allotted

1
i) Position of some elements do not fit in with the scheme of classification or element of lower atomic weight are placed before the element with high atomic weight.
ii) He correctly predicted the existence of few elements.
1
1

2 Eka-aluminium
Gallium
1
1

3
The regularities in the x-ray spectra of elements could be accounted if a plot of ν (ν is the frequency of x-rays emitted) versus atomic number and not atomic mass was done.
or
A graph of ν versus atomic number and not atomic mass could explain the characteristics of x-ray spectra of elements.
1
1
1
1

4 The vertical column of elements are called groups.
The horizontal rows of elements are called periods.
1
1

5 18 groups and 7 periods 1+1

6 8 elements
In the 2^{nd} period (n = 2) the sub-shells filled are 2s and 2p only
1
1

7 1^{st} period
6^{th} period

1
1
8 i) 2s and 2p
ii) 4s, 3d and 4p
1
1
9
i) It has only one electron in 1_{st} orbital and hence could be placed in I group
ii) It can also gain one electron to achieve a noble gas configuration like 17 group elements and hence can be placed in 17 group.
1
1
10 Block – d

Period − 4

1
1
11
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$ or [Rn] $7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$
18 group or noble gas group
1
1
12

Metals	**Non-metals**
1. Have high melting point	Have low melting point.
2. Good conductor of heat and electricity	Bad conductor of heat and electricity
3. Malleable and ductile	Not malleable and ductile

(any two)
1+1
13
Elements that show properties similar to both metals and nonmetals
Silicon (or any other)

1
1
14 i) Atom is very small
ii) Electron cloud of the atom does not have a sharp boundary
1
1
15 It decreases along a period
It increases down a group
1
1
16
i) Along a period the electron is added to same valence shell and effective nuclear charge increases.
ii) Down a group the electron is added to a new shell and effective nuclear charge decreases.
1
1
17
Set a
The elements belong to a group and down the group the atomic size increases.
1
1
18 Species containing same number of electrons.
E.g.: Na^{+}, Mg^{+2}
1
1
19 F^{-}
Al^{+3}
1
1
20
It is the energy required to remove an electron from an isolated gaseous atom in its ground state.
kJ mol^{-1} or kJ
1
1
21
1^{st} ionisation enthalpy of magnesium is less than the 2^{nd}.
More energy is required to remove an electron from Mg^{+} than from Mg

1
1
22 Noble gases or 18 group.
They have a very stable electronic configuration.
1
1
23
Graph-1 Position of elements-1
1+1
24
It increases.
Increase in nuclear charge outweighs (or is more prominent) the shielding effect
1
1
25
It decreases.
Increase in shielding effect outweighs the increase in nuclear charge.
1
1
26
In Beryllium electron is removed from 2s electron which is closer to the nucleus.
In boron the electron is removed from 2p orbital which is far away from the nucleus and is also shielded by 2s electrons.
1
1
27
In nitrogen the three 2p electrons are in different atomic orbitals.
In oxygen, two of the four 2p electrons occupy the same 2p orbital resulting in electron-electron repulsion.
1
1
28
In magnesium the electron has to be removed from 3s orbital which is closer to the nucleus.
In aluminium the electron has to be removed from 3p orbital which is far from the nucleus and is also well shielded by 3s electrons.
1

1
29
Ionisation enthalpy Electron gain enthalpy
Its always positive It may be positive, negative or zero
It measures the hold of the nucleus on its electron.
It measures the affinity of an atom to add an electron.
1
1
30 It increases or it becomes more negative.
Effective nuclear charge increases.
1
1
31 It decreases or it becomes less negative.
Size of the atom increases and effective nuclear charge decreases.
1
1
32
Fluorine
The electron is added to 2p orbital which is small and experiences more repulsion from other electrons.
1
1
33 Ar
O
1
1
34 It increases
Atomic radius decreases
1
1
35 It decreases
Atomic radius increases
1
1
36 Ionisation enthalpy and electronegativity 2
37 Electron gain enthalpy and electronegativity 2
38 I < Br < F < Cl
Na < Al < S < Cl

1
1
39 $[BF_4]^-$
Boron has only 4 valence orbitals.
1
1
40
Towards the left, ionisation energy is the least and atoms have the tendency to lose the electron.
Towards the right, electron gain enthalpy is more negative and hence atoms have tendency to gain an electron.
1
1
41 The elements have low ionisation enthalpy and have high tendency to form cations.
1
1
42 Attraction of electrons towards the nucleus.
Repulsion of electrons from each other.
1
1
43 Beryllium and Boron or oxygen and nitrogen
Oxygen and sulphur
1
1
44 CO- neutral, Na_2O – basic, Al_2O_3 – amphoteric, Cl_2O_7 – acidic 2
Any other suitable answer to be considered as value point.

Four mark questions:

Qn.No. Value points Marks allotted

1
a) 2
b) i) Al_2O_3
ii) AlN
iii) AlF_3
1
1
1
1
2
When an atom loses an electron and when an atom gains an

electron.
$H^- > H > H^+$
1st ionisation enthalpy
2
1
1
3
Ionisation energy measures tendency of an atom to lose an electron.
Electron gain enthalpy measures tendency of an atom to gain an electron.
Electronegativity measures the ability of an atom in a compound to attract shared electrons.
Electronegativity
1
1
1
1
4
Q > R > P > S
S
Q
R_2S
1
1
1
1
5 s block : They are metals and form basic oxides.
p block : They are non-metals and form acidic oxides.
2
2
6
Metallic character increases down a group.
Ionisation energy decreases.
Non-metallic character increases along a period.
Electron gain enthalpy increases.
1
1
1
1
7 A: 2nd group, 3rd period ; B: 17th group, 3rd period

i) A ii) B
2
2

I PUC QUESTION BANK

Chapter: 4

Chemical bonding

1 Mark Questions

1) Who were the first to propose a theory on chemical bonding using electrons?
2) What is a chemical bond?
3) Write Lewis dot symbols for atoms of the following elements
Mg, Na, B, O, N, and Br.
4) Draw the Lewis structures for the following molecules and ions.
H_2, O_2, CO_2, C_2H_4, C_2H_2, HNO_3, CO
5) In the periodic table, the group of highly electronegative elements is _____.

6) In the periodic table, the group of highly electropositive elements is _____.
7) Write the general electronic configuration of noble gases.
8) What type of bond is present in NaCl?
9) Which force holds oppositely charged ions together in an ionic bond?
10) Name a cation that contains cation having two non metallic elements in an ionic compound.
11) How does resonance stabilize a molecule?
12) Give the mathematical expression for dipole moment.
13) Expand VSEPR
14) What is the basis of VSEPR theory?
15) Arrange the repulsive interaction between electron pairs (lp-bp, bp-bp, and lp-lp) in increasing order.
16) What is lone pair of electrons?
17) What is bonded pair of electrons?
18) Among bonded pair of electron and lone pair of electron, which occupy more space in a molecule?
19) How many lone pairs of electrons are present in a molecule of ammonia?
20) How many lone pairs of electrons are present in a molecule of water?
21) How many lone pairs of electrons are present in a molecule of ClF_3?
22) How many lone pairs of electrons are present in a molecule of SF_4?
23) What is hybridization?
24) Write the shape and bond angle of sp hybrid orbitals?
25) Give an example of a molecule having sp hybridization?
26) What is the percentage of s character in sp hybridization?
27) What is the shape and bond angle of sp^2 hybrid orbitals?
28) What is the percentage of s character in sp^3 hybridizations?
29) Give an example for sp^2 and sp^3 hybrid molecules
30) Write the shape and bond angle of sp^3 hybridized orbitals.
31) Write the shape, hybridization and bond angle of a) NH_3 and, b) H_2O
32) How many σ and π bonds are there in a) ethylene b) ethyne?
33) What is the type of hybridization found in PCl_5?
34) What is the shape and hybridization of SF_6?
35) Write the number of axial and equatorial bonds in PCl_5
36) Give example of a molecule showing dsp^2 hybridization.
37) Write the shape of Br F_5 molecule.
38) Arrange the following orbitals in the increasing order of s- character.sp, sp^2, sp3
39) Define hydrogen bond.
40) How is the magnitude of hydrogen bonding in different states of matter?
41) Is the force between two nuclei of hydrogen atoms attractive or repulsive?
42) Which of the force (i.e. attractive or repulsive) is stronger?
43) What is bond enthalpy?
44) The electrons present in valence shells should have ___spins in order to be paired up.
45) Will a bond be stronger when there is more overlap or less overlap of atomic orbitals?
46) Hydrogen molecule is formed due to overlap of ____ orbitals.

47) What is tetrahedral bond angle?
48) What are the types of overlap depending on the sign (phase) of orbital wave function?
49) If the wave functions of participating atomic orbitals are same phase, the overlap is called _____.
50) The energy required for changing the electronic configuration from ground to exited state is provided from where?
51) What is the angle between any two p orbitals?
52) What are the two types of covalent bonds?
53) What is a pi bond?
54) What is the shape of the electron cloud in a pi bond?
55) Can a pi bond exist without a sigma bond?
56) Is there greater overlap in sigma or pi bonds?
57) Define hydrogen bond.

2 Marks

1) Which type of elements is likely to form anions and give the reason for the same?
2) Which factor favours formation of cation? Explain.
3) Among KCl and NaCl, which is more stable? Give reason.
4) Under what conditions the concept of resonance is applied?
5) The dipole moment in BF_3 is zero. Explain
6) The net dipole of NH_3 is greater than that of NF_3. Why?
7) Write the resonance structures of CO_3^{2-} and CO_2.
8) List the outcome of repulsive interaction between electron pairs in a molecule.
9) Account for the shape of the following molecules based on VSEPR theory
i) Water molecule.
ii) SO_2 molecule
iii) Ammonia molecule
iv) SF_4 molecule.
10) What are the causes of formation of hydrogen bond?
11) What are the conditions for hydrogen bonding?
12) Draw the shapes of following hybrid orbitals.
sp, sp^2,sp^3
13) Considering X- axis as the internuclear axis, which out of the following will not form sigma bond? Why?
1s and 2s b) 1s and 2px c) 2py and 2py d) 1s and 2s
14) Draw the shapes of H_2O and C_2H_4 molecule
15) Describe the change in the hybridization of the Al atom in the following reaction.

$AlCl_3 + Cl^- \rightarrow AlCl_4^-$

16) Is there any change in hybridization of B and N as a result of the reaction?

$BF_3 + NH_3 \rightarrow BF_3.NH_3$

17) What are the various forces that arise between the two atoms of hydrogen as they approach each other?
18) Give the ground and exited configuration of carbon for the formation of methane molecule.

19) What is a sigma bond? Mention types.
20) Explain the LCAO method for the formation of molecular orbitals from the atomic orbitals.
21) Distinguish between atomic orbital and molecular orbital.
22) Distinguish between bonding molecular orbital and anti-bonding molecular orbital.
23) Define bond order and give the relationship between bond order and bond length
24) What are the causes of formation of hydrogen bond?
25) What are the conditions for formation of hydrogen bonds?

3 Mark questions

1) Explain the conditions for the combination of atomic orbitals to form molecular orbitals.
2) Write the assumptions of the molecular orbital theory.
3) Explain the diamagnetic behaviour of Hydrogen molecule on the basis of molecular orbital theory.
4) Show the non-existence of helium molecule based on molecular orbital theory.
5) Explain the formation of Lithium molecule on the basis of molecular orbital.
6) Show that oxygen molecule is paramagnetic based on molecular orbital theory.
7) Write the molecular orbital electronic configuration for carbon molecule and comment on its magnetic property
8) Define octet rule. Write its significance and limitations.
9) Explain why crystal structure of NaCl is stable in spite of having sum of electron gain energy and ionization enthalpy as positive. Substantiate with an example
10) Why it is required to take average bond enthalpy in case of polyatomic molecules? Explain with example.
11) Among O_2 and N_2 which is more stable and why?
12) The net dipole moment of NH_3 is greater than that of NF_3. Why?
13) State the postulates of VSEPR theory
14) Indicate the geometry in the following compounds and give one example for each.
i) AB_2
ii) AB_3
iii) AB_4
iv) AB_5
v) AB_6
15) What are the types of hydrogen bonds?
16) What are the salient features of hybridization?
17) With the help of potential energy diagram explain how bonding between two hydrogen atoms takes place.
18) Explain the shape of methane molecule through valence bond theory.
19) Explain the formation of $BeCl_2$ using hybridization.
20) Explain the sp^2 hybridization with the example of BCl_3.
21) Explain the shape of NH_3 molecule using hybridization.
22) Explain the formation of SF_6 using sp^3d^2 hybridization.

1mark Answers:

1). Ans**:** Kossel and Lewis
2). Ans: The attractive force which holds various constituents (atoms or ions etc) together in different chemical species is called a chemical bond

3). **Ans**:
$_{12}Mg=1s^2 2s^2 2p^6 3s^2$
$_{11}Na= 1s^2 2s^2 2p^6 3s^1$
$_{5}B=1s^2 2s^2 2p^1$
$_{8}O = 1s^2 2s^2 2p^6$
$_{7}N =1s^2 2s^2 2p^3$
$_{35}Br =1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
4). **Ans:** Refer to pages 99 and 100
5) **Ans:** Halogens
6) **Ans:** Alkali metals
7) **Ans:** The general electronic configuration of noble gases is $ns^2 np^6$
8) **Ans:** Electrovalent bond
9) **Ans:** Electrostatic force or coulombic force.
10) **Ans:** $\mathbf{NH_4}$
$_{+}$ in NH_4Cl
11) **Ans: Resonance** stabilizes the molecule as the energy of the resonance hybrid is less than the energy of any single canonical structure.

12) **Ans:** ⌈=Q·Ƴ

Dipole moment=charge·distance

13) **Ans:** Valence shell electron pair repulsion
14) **Ans:** Based on repulsive interactions between valence electron pairs in a molecule.
15) **Ans:** Bond pair-bond pair < bond pair-lone pair < lone pair-lone pair.
16) **Ans:** The localized pair of valence electrons over the central atom of a molecule which do not take part in covalent bond formation
. 17) **Ans:** The pair of electron shared between two atoms in a molecule.
18) **Ans:** Lone pair.
19) **Ans:** One lone pair.
20) **Ans:** Two lone pair.
21) **Ans:** Two lone pair.
22) **Ans:** One lone pair
23) **Ans:** The process of intermixing of atomic orbitals of different energies to get same number of new orbitals of equivalent energies is called hybridization
24) **Ans:** Linear shape, bond angle-180^0
25) **Ans:** $BeCl_2$

26) **Ans:** 50%

27) **Ans:** Shape-trigonal planar, bond angle-120^0
28) **Ans:** 25 %
29) **Ans:** BCl_3 - sp^2
CH_4 – sp^3
30) **Ans:** Shape: regular tetrahedron

Bond angle: $109^0\,28'$

31) **Ans:** a) NH_3
Shape: Pyramidal
Hybridization: sp^3

Bond angle: 107^o
b) H_2O
Shape: V shape or angular shape
Hybridization: sp^3
Bond angle: 104.5^o

32) **Ans:** Ethylene σ bond = 5, π bond = 1

Ethyne σ bond = 3, π bond = 2

33) **Ans:** sp^3d hybridization
34) **Ans:** Shape – Octahedral
Hybridization – sp^3d^2
35) **Ans:** Axial bond – 2
Equatorial bond -3
36) **Ans:** $[PtCl_4]^{2-}$
37) **Ans:** Square pyramidal
38) **Ans:** $Sp^3<sp^2<sp$
39) **Ans:** Hydrogen bond is defined as the attractive force which binds hydrogen atom of one molecule with the electronegative atom (F, O or N) of another molecule.
40) **Ans:** It is maximum in solid state and minimum in the gaseous state.
41) **Ans:** Repulsive
42) **Ans:** Attractive
43) **Ans:** The energy released when a bond is formed between two atoms is called bond enthalpy.
44) **Ans:** Opposite
45) **Ans:** More overlap
46) **Ans:** 1s

47) **Ans:** $109^o28'$

48) **Ans:** Positive, negative, zero
49) **Ans:** Positive
50) **Ans:** When atomic orbitals overlap there is release in energy, this energy is used to change the electronic configuration
51) **Ans:** 90^0
52) **Ans:** Sigma and pi
53) **Ans:** It is a type of covalent bond formed due to overlap of atomic orbitals side ways.
54) **Ans:** A pi electron cloud consists of two saucer shaped clouds above and below the plane of participating atoms.
55) **Ans:** No
56) **Ans:** Sigma bonds.
57) **Ans:** Hydrogen bond is defined as the attractive force which binds hydrogen atom of one molecule with the electronegative atom (F, O or N) of same molecule or another molecule.

2 Marks answers

1) **Ans**: Non-metallic elements have high electron gain enthalpy to form anions which is produced in the process

$Cl_{(g)} + e^{-1} \rightarrow Cl^-_{(g)}$ $\Delta H = 348.7 kJmol^{-1}$

2) **Ans**: Metallic elements have low ionization energy which facilitates the easy release of electrons from the metal and formation of cation is

$Na \rightarrow Na^+ + e^-$

3) **Ans**: NaCl is more stable than KCl.
Lattice energy of NaCl (788KJ/mol) greater than lattice energy of KCl (718KJ/mol) because smaller ionic radius of Na^+ (95pm) when compared to K^+ (133pm).

4)**Ans**: According to the concept of resonance, whenever a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energy, positions of nuclei, bonding and nonbonding pairs of electrons are taken as the canonical structures and the hybrid describes the molecule accurately.

5) **Ans:** In BF_3 ⌈ = 0, although the B-F bonds are oriented at an angle of $120^0$to one another. This is because the bond moments give a net sum of zero as the resultant of any two is equal and opposite of third.

6)**Ans:** In case of NH_3 the 0rbital dipole due to lone pair of electron on nitrogen atom , is in the same direction as the resultant dipole moment of N-H bonds, where as in NF_3, it is in the direction opposite to resultant dipole moment of 3 N-F bonds. The orbital dipole decreases the effect of the resultant N-F bond moments which reduces dipole moment of NF_3.

7) **Ans**: For the resonance structures of CO_3^{2-} and CO_2 refer to page no 106(Part-1)

8) **Ans:** a) Deviation in the shape of the molecule
b) Alterations in the bond angle in the molecule.

9) **Ans:** i) Water molecule belongs to type of AB_2E_2
The shape of water molecule should have been tetrahedral if there were all bond pair. But because of the presence of two lone pair, the shape is distorted tetrahedral or angular. The reason is lp-lp repulsion is more than bp-bp repulsion. Thus the angle is reduced to 104.5^0 from 109.5^0.
Refer to page no 112 for structure.
ii) SO_2 molecule
SO_2 molecule belongs to AB_2E
The shape of SO_2 molecule should have been triangular planar but it is found to be bent or V-shaped .This is due to the fact that lp-bp repulsion is much more than bp-bp repulsion. So the angle is reduced to 119.5^0 from 120^0.
iii) Ammonia molecule
Ammonia molecule belong to AB_3E type
The shape of NH_3 molecule should have been tetrahedral if all the electrons were bond pair. But because of the presence of one lp,ther is lp-bp repulsion which is more than bp-bp repulsion and the angle is reduced from 107^0 to 104.5^0.
iv) SF_4 molecule.
SF_4 molecule belongs to AB_4E type.

The shape of SF_4 molecule is distorted tetrahedron or a folded square or a see- saw in which the lp is in an equatorial position where there are only two lp-bp repulsions. This arrangement is more stable than the arrangement where the lp is present at axial position where there are three lp-bp repulsions at 90^0.

10) **Ans:** When there is formation of covalent bond between hydrogen and electronegative elements, the electrons of the covalent bond are shifted towards more electronegative atom as of which the positively charged hydrogen forms a bond with the other electronegative atom .

11) **Ans:** a) Hydrogen atom should be bonded to highly electronegative atoms such as fluorine, oxygen or nitrogen.

b) Size of the electronegative atoms should be small.

12) **Ans:** sp – Refer to page 117,fig 4.10

sp^2 – Refer to page 117, fig 4.11

sp^3 – Refer to page 118, fig 4.12

13) **Ans:** $2p_y$ and $2p_y$ will not form a sigma bond because taking x- axis as the internuclear axis, there will be lateral overlapping between 2py orbitals forming a pi bond.

14) **Ans:** For H_2O refer page 118, fig 4.14 and for C_2H_4 refer page 119, Fig 4.15

15) **Ans:** Electronic configuration of Al = $1s^2 2s^2 2p^6 3s^2 3p^2$

Exited state configuration is $1s^2 2s^2 2p^6 3s^1 3p_x^1 3p_y^1$

Hence hybridization is sp^2

In $AlCl_4^-$ empty 3pz orbital is also involved so the hybridization is sp^3 and shape is tetrahedral.

16) **Ans:** In BF_3, B atom is sp^2 hybridized. In NH_3, N atom is sp^3 hybridized. After the reaction hybridization changes from sp^2 to sp^3.

17) **Ans:** a) Attractive force between nucleus and its own electron.

b) Attractive force between nucleus and electron of other atom

c) Repulsive force between the two nuclei

d) Repulsive force between two electrons.

18) **Ans:** Ground state [He] $2s^2 2p^2$

Excited state [He] $2s^1 2p_x^1 2p_y^1 2p_z^1$

19) **Ans:** This type of covalent bond is formed by the head on overlap of atomic orbitals along the internuclear axis. The sigma bonds result from the following types of overlap

a) s-s b) s-p c) p-p

20) **Ans:** The molecular orbitals are formed by the linear combination of wave functions of the participating atomic orbitals. They may combine either by addition or by

subtraction. Let ψ_A and ψ_B represent the wave functions of the two combining atomic orbitals A and B taking part in chemical combination.

21) **Ans:** Atomic orbitals

1. Atomic orbitals are monocentric.
2. Atomic orbitals have simple shapes like spherical or dumb-bell.

Molecular orbitals

1. Molecular orbitals are polycentric.
2. Molecular orbitals have complex shapes.

22) **Ans:** Bonding molecular orbitals
1. Formed by symmetric combination of atomic orbitals.
2. Has more electron density between the nuclei.
Antibonding molecular orbitals.
1. Formed by asymmetric combination of atomic orbitals.
2. Has less electron density between the nuclei.
23) **Ans:** It is half of the difference between the number of electrons present in bonding molecular orbitals and the number of electrons present in the antibonding molecular orbitals.
Bond order=Number of electrons in B.M.O-Number of electrons in A.B.M.O/2
24) **Ans:** When there is formation of covalent bond between hydrogen and electronegative elements, the electrons of the covalent bond are shifted towards more electronegative atom which makes the hydrogen atom positive and it forms the hydrogen bond with other electronegative atom.
25) **Ans:** The conditions are
a) Hydrogen atom should be bonded to a highly electronegative atom such as F, O or N
b) Size of the electronegative atoms should be small.

3 Mark answers

1) **Ans**: The conditions for the combination of atomic orbitals to form molecular orbitals are
1. The combining atomic orbitals must have same or nearly the same energy.
2. The combining atomic orbitals must have the same symmetry about the molecular axis.
3. The combining atomic orbitals must overlap to the maximum extent.
2) **Ans:** The assumptions of the molecular orbital theory are
1. The molecular orbital is the region in space comprising the nuclei of the combining atoms around which there is maximum probability of finding electron density.
2. The number of molecular orbitals formed is equal to the number of atomic orbitals taking part in the combination.
3. When two atomic orbitals combine according to L.C.A.O. principle, they form two molecular orbitals i.e. bonding and antibonding.
4. The bonding molecular orbital has lower energy (more stability) as compared to antibonding molecular orbital.
3) **Ans:** It is formed by the combination of the two hydrogen atoms.
Each hydrogen atom has one electron in 1s orbital.
Hence there are two electrons in Hydrogen molecule.

The molecular orbital configuration is $\sigma\ 1s^2$

The bond order is calculated as follows
Bond order = $N_b - N_a/2$
= 2-0/2
=1
Hydrogen molecule is diamagnetic since there is no unpaired electron.
4) **Ans:** It is formed by the combination of the two helium atoms.
Each helium atom has two electrons in 1s orbital.
Hence there are four electrons in Helium molecule.

The molecular orbital configuration is $\sigma\ 1s^2, \sigma\ *1s^2$

The bond order is calculated as follows
Bond order = $N_b - N_a/2$
=2-2/2
=0
Helium molecule is unstable and does not exist.
5) **Ans:** It is formed by the combination of the two Lithium atoms.
Each Lithium atom has three electrons and its electronic configuration is $1s^2\ 2s^1$
Hence there are two electrons in Hydrogen molecule.

The molecular orbital configuration is $\sigma\ 1s^2\ \sigma^*1s^2\ \sigma\ 2s^2$

The bond order is calculated as follows
Bond order = $N_b - N_a/2$
= 4-2/2
=1
Lithium molecule is diamagnetic since there is no unpaired electron.
6) **Ans:** It is formed by the combination of the two oxygen atoms.
Each oxygen atom has eight electrons and its electronic configuration is
$1s^2$, $2s^2$, $2p_x$
^{2}, $2p_y$
^{1}, $2p_z$
^{1}.
Hence there are 16 electrons in oxygen molecule.
The molecular orbital configuration is

$\sigma\ 1s^2$, σ^*1s^2, $\sigma\ 2s^2$, σ^*2s^2, $\sigma\ 2pz^2$, $\pi 2px^2 = \pi 2py^2$, $\pi^*2px^1 = \pi^*2py^1$

The bond order is calculated as follows
Bond order = $N_b - N_a/2$
=10-6/2
=2
Oxygen molecule is paramagnetic since there are unpaired electrons.
7) **Ans:** It is formed by the combination of the two carbon atoms.
Each carbon atom has six electrons and its electronic configuration is
$1s^2$, $2s^2$, $2p_x$
^{1}, $2p_y$
^{1}.
Hence there are 12 electrons in carbon molecule.
The molecular orbital configuration is

$\sigma\ 1s^2\ \sigma^*1s^2\ \sigma\ 2s^2\ \sigma^*2s^2\ \pi 2px^2 = \pi 2py^2$

The bond order is calculated as follows
Bond order = $N_b - N_a/2$
= 8-4/2
=2
Carbon molecule is diamagnetic since there no unpaired electrons.
8) **Ans:** Octet rule: Atoms can combine either by transfer of valence electrons from one atom to another (gaining or losing) or by sharing of electron in order to acquire stable gas configuration.

Significance: It helps to explain why different atoms combine with each other to form ionic compounds or covalent compounds.
Limitations:
1) According to octet rule, atoms take part in chemical combination to achieve the configuration of nearest noble gas elements. However some of noble gas elements like Xenon have formed compounds with fluorine and oxygen.
2) The theory does not account for shape of molecule.
9) **Ans:** It is because energy released during the formation of crystal lattice (lattice energy) is greater than the sum of electron gain enthalpy and ionization enthalpy.
Example: In NaCl
Ionization enthalpy

$Na\ (g) \rightarrow Na^{+} + e^{-1}\ \Delta H_I = 495.8 kJmol^{-1}$

Electron gain enthalpy

$Cl\ (g) + e^{-1} \rightarrow Cl^{-}$

$_{(g)}\ \Delta H_E = 348.7 kJmol^{-1}$

Therefore $Na\ (g) + Cl\ (g) \rightarrow Na^{+} + Cl^{-}\ \Delta H = 147.1 kJmol^{-1}$

This is compensated by the lattice energy of NaCl

$Na^{+}\ (g) + Cl^{-}(g) \rightarrow NaCl(s)\ \Delta H = -788\ kJmol^{-1}$

Thus stability of ionic compound is due to enthalpy of lattice formation not just by octet ionic species in gaseous state.
10) **Ans:** In polyatomic molecules the enthalpy required to break the same bond in different chemical environment are different.

Example: $H_2O\ (g) \rightarrow H\ (g) + OH\ (g)\ \Delta H_1 = 502\ kJmol^{-1}$

$OH_{(g)} \rightarrow H_{(g)} + O_{(g)}\ \Delta H_2 = 427\ kJmol^{-1}$

Average bond enthalpy is $502+427/2 = 464.5\ Kjmol^{-1}$
11) **Ans**: Bond order bond length bond dissociation enthalpy

$N_2 \rightarrow N \equiv N$ 3 110nm 945 $kJmol^{-1}$

$O_2 \rightarrow O=O$ 2 121nm 498 $kJmol^{-1}$

Greater the bond order, bond enthalpy (energy required to break the bond) increases, bond length decreases. Hence stability of the molecule increases. Thus N_2 is more stable than O_2.
12) **Ans:** In case of NH_3 the orbital dipole due to lone pair of electrons on N atom , is in the same direction as the resultant dipole moment of N-H bonds, whereas in NF_3 it is in the direction opposite to resultant dipole moment of three N-F bonds. The orbital

dipole decreases the effect of the resultant N-F bond moments which reduces ⌠ of

NF_3.
13) **Ans:** a) The shape of the molecule depends upon the number of valence shell electron pairs (bonded or non bonded) around the central atom.
b) Pairs of electrons in the valence shell repel one another since their electron clouds

are negatively charged.
c) These pairs of electrons tend to occupy such positions in space that minimize repulsion and thus maximize distance between them.
d) The valence shell is taken as a sphere with the electron pairs localizing on the spherical surface at maximum distance from one another.
e) A multiple bond is treated as if it is a single electron pair and the two or three electron pairs of a multiple bond are treated as a single super pair.
f) Where two or more resonance structure can represent a molecule, the VSEPR model is applicable to any such structure.
14) **Ans:** a) AB_2
Linear geometry, Ex: $BeCl_2$
b) AB_3
Trigonal planar geometry, Ex: BF_3
c) AB_4
Tetrahedral geometry, Ex: CH_4
d) AB_5
Trigonal bipyramidal geometry, Ex: PCl_5
e) AB_6
Octahedral geometry, Ex: SF_6.
15) **Ans:** There are two types of hydrogen bonds
Intermolecular hydrogen bond: It is formed between two different molecules of the same or different molecules of the same or different compounds.
Example: HF, H2O, Alcohols
Intramolecular hydrogen bond: It is formed when hydrogen atom is in between the two highly electronegative atoms present within the same molecule.
Example: O-nitrophenol.
16) **Ans:** a) The number of hybrid orbitals is equal to the number of atomic orbitals undergoing hybridization.
b) The hybridized orbitals are always equivalent in energy and shape
c) Hybrid orbitals are more effective in forming stable bonds than pure atomic orbitals
d) Type of hybridization indicates the geometry of the molecule
e) The orbitals in the valence shell of atom are hybridized
f) The orbitals undergoing hybridization should have comparable energy.
g) Not only half filled orbitals but filled orbitals also take part in hybridization
17) **Ans:** When two atoms approach each other, a number of attractive and repulsive forces operate. Experimentally it has been found that attractive forces have greater magnitude and hence the atoms come closer together. A point is reached here net attractive force balances net repulsive force and potential energy reaches a minimum. At this stage the atoms are said to be bonded. Refer fig 4.8, pg: 114
18).**Ans:** In the formation of methane molecule the unpaired electrons in each of the orbital s of carbon pair up with electron in 1s of hydrogen, resulting in formation of 4 C-H bonds. As the angle between p-orbitals is 90^0 the H-C-H bond angle should also be 90^0 and as 1s orbital is spherical the overlap can occur in any direction and direction of fourth C-H bond cannot be ascertained but CH_4 has a tetrahedral structure hence simple atomic overlap cannot explain the tetrahedral shape of methane.

19) **Ans:** The ground state electronic configuration of Be is $1s^2 2s^2$. In the exited state one of the 2s-electrons is promoted to vacant 2p orbital to account for its bivalency. One 2s and one 2p orbitals get hybridized to form two sp hybridized orbitals. These two sp hybrid orbitals are oriented in opposite direction forming an angle of 180^o. Each of the sp hybridized orbital overlaps with the 2p- orbital of chlorine axially and form two Be-Cl sigma bonds. Ref Pg 117 , Fig 4.10 (b).

20).**Ans:** In BCl_3 molecule, the ground state electronic configuration of central boron atom is $1s^2 2s^2 2p^1$. In the exited state, one of the 2s electrons is promoted to vacant 2p orbital as a result boron has three unpaired electrons. These 3 orbitals (one 2s and two 2p) hybridized to form three sp^2 hybrid orbitals. The three hybrid orbitals so formed are oriented in a trigonal planar arrangement and overlap with 2p orbitals of chlorine to form three B – Cl bonds. Therefore, in BCl_3, the geometry is trigonal planar with Cl-B-Cl bond angle of 120^0. Ref pg 117, Fig 4.11.

21) **Ans:** In ammonia, the valence shell electronic configuration of N in the ground state is

$2s^2 2p_x^1 2p_y^1 2p_z^1$

having three unpaired electrons in the sp^3 hybrid orbitals and a lone pair of electrons is present in the fourth one. These 3 hybrid orbitals overlap with 1s orbitals of hydrogen atoms to form 3 N-H sigma bonds. Force of repulsion between a lone pair and a bond pair is more than the force of repulsion between two bond pairs of electrons. The molecules thus get distorted and the bond angle is reduced to 107^0 from 109.5^0. The geometry of such a molecule will be pyramidal and can be explained with sp^3 hybridization. Ref Pg 118, Fig 4.13.

22).**Ans:** In SF_6 the central sulphur atom has the ground state outer electronic configuration of

$3s^2 3p^4$

. In the exited state the available 6 orbitals i.e., 1s, 3p and 2d are singly occupied by electrons. These orbitals hybridise to form 6 new sp^3d^2 hybrid orbitals, which are projected towards the 6 corners of a regular octahedron in SF_6. These 6 sp^3d^2 hybrid orbitals overlap with singly occupied orbitals of Fluorine atoms to form 6 S-F sigma bonds. Thus it has regular octahedron shape. Ref pg121, Fig 4.18

UNIT 5
States of matter
I. Questions carrying one mark

5.1 What are van der Waals forces?
5.2 What type of van der Waals force exists between HCl molecules?
5.3 Between which type of molecules does dipole – induced dipole forces exist?
5.4 Even though HF has lower molecular mass compared to HCl, HF is in the liquid state and HCl is in the gaseous state at room temperature .why?
5.5 Why are liquids & solids are hard to compress?
5.6 State Boyle's law
5.7 What happens to the pressure when the volume of a gas is doubled at constant temperature?
5.8 Give the relationship between pressure and density of a gas
5.9 Mountaineers carry oxygen cylinders along with them . Which gas law will account to the above statement?
5.10 Relate absolute temperature scale or Kelvin temperature to Celsius scale.
5.11 If a volume vs temperature graph is plotted in degree centigrade at constant pressure, at what temperature the line cuts the temperature axis?
5.12 Each line of volume vs pressure graph at constant temperature is known as what?
5.13 Define absolute zero temperature
5.14 Represent Charles law graphically
5.15 What is the volume occupied by the gas at -273.15ºC?
5.16 State Gay – Lussac's law
5.17 Equal volumes of all gases under the same condition of temperature and pressure contain equal number of molecules. Name the gas law for the above statement
5.18 Which law relates temperature and pressure ?
5.19 If the numbers of moles of a gas are doubled by keeping the temperature and pressure constant, what would be the new volume of the gas?
5.20 Relate density and molar mass of a gas molecule
5.21 Give the value of R in SI unit.

5.22 Write the value of molar volume of an ideal gas at STP, when pressure is 1bar
5.23 State Dalton's law of partial pressure
5.24 Write a postulate of kinetic molecular theory of gases that explains the great compressibility of gases
5.25 Define aqueous tension
5.26 Why do the gases expand &occupy the entire space available to them?
5.27 How does a gas exert pressure?
5.28 Collision of the gas molecules does not change average kinetic energy.why?
5.29 Write van der Waals equation for ´n' moles of a gas
5.30Write the expression for compressibility factor Z for one mole of a gas.
5.31Define Boyle temperature
5.32 The value of Van der walls constant 'a' for a gas is zero, what does it signify?
5.33 Define boiling point temperature at a given pressure for a liquid
5.34 Give the value for standard boiling point of water
5.35 Define surface energy
5.36 What would be SI unit of a quantity $PV_2 T_2/n$?
5.37 The magnitude of coefficient of viscosity of liquid at 25°C ,45°C, 19ºC and 57°C are p, q, r, s respectively. Arrange them in the increasing value of co-efficient of viscosity
5.38 What is meant by most probable speed of gas molecule ?
5.39 Water has higher vapour pressure than mercury. Justify
5.40 Name the property of the liquid that measures the resistance to the flow of liquid due to internal friction.
5.41 How does the volume of a gas under given temperature and pressure vary with molar mass?
5.42. The size of the weather balloon becomes larger and larger as it ascends up into higher attitude. Which gas law explains the above phenomenon?
5.43 Arrange U_{rms}, U_{mp}, U_{av} in the increasing value for a gas at a given temperature

II. Questions carrying Two marks

5.1 Name two types of forces which determines the physical state of substances
5.2 Write any two important physical properties of gases.
5.3 How does pressure for a given amount of gas vary with volume at a given temperature? Represent this graphically
5.4 In terms of Charles law, explain why -273.15ºC is the lowest possible temperature of gases.
5.5 Represent Charles law graphically. What is each line in the graph known as?
5.6 Write the Avogadro's law. Write the mathematical representation of the law.
5.7 Using ideal gas equation derive an equation that relates density and molar mass of a gaseous substance .
5.8 Deduce the value of R in SI unit
5.9 State Daltons law of partial pressure. How is partial pressure of a gas is related to mole fraction
5.10 Justify: collisions among the molecules in a gas are perfectly elastic
5.11 Two flasks of equal volumes contains nitrogen and oxygen gases at the same temperature and pressure. Which will have greater number of molecules? Justify the answer by stating the law.
5.12 If the number of moles of a gas is doubled by keeping the temperature and pressure constant, what will happen to the volume? Give reason.
5.13 Which two postulates of the kinetic theory of gases are not applicable for real gases?
5.14 Explain the physical significance of Van der Waals parameters 'a' & 'b'for a real gas.
5.15 Under what condition real gases tend to show ideal gas behaviour?
5.16 Draw Z vs P graph for ideal gas and CO_2 gas.
5.17 At high attitude time required to cook food is more than at sea level. Justify the statement.
5.18 Out of NH_3 and N_2, which will have higher value of 'a'- the van der Waals parameter and why?

5.19 What happens to the compressibility factor for gases like CO_2 at
i) Very high pressure & ordinary temperature
ii) Low pressure & ordinary temperature
5.20 Can Dalton's law of partial pressure be used to calculate pressure of mixture of NH_3 & HCl? Justify the answer .
5.21 Two gases A and B have critical temperature of 250K & 125K respectively. Which one of these can be liquefied easily and why?
5.22 Define surface tension. Write the SI unit of surface tension.
5.23 A drop of liquid assumes spherical shape. why?
5.24 Name the two factors on which the magnitude of surface tension depends?
5.25 A balloon has a volume of 175dm^3 when filled with hydrogen gas at a pressure of 1.0 bar. Calculate the volume of the balloon when the pressure is 0.8bar. Assume that the temperature remains constant.
5.26 A gas cylinder containing cooking gas can withstand a pressure of 14.9bar. The pressure gauge of the cylinder indicates 12bar at 27ºC. Due to sudden fire in the building, the temperature starts rising. At what temperature will the cylinder explode?
5.27 A weather balloon filled with hydrogen at 1 atm & 27ºC has volume equal to 1200dm^3. On ascending it reaches a place where the temperature is -23ºC and pressure is 0.5 atm. What is the volume of the balloon at this temperature?
5.28 Write two factors on which molecular speed of a gas depends?
5.29. The density of a gas is found to be 1.56g /dm^3 at 0.98bar pressure & 65ºC. Calculate the molar mass of the gas (R=0.083 bar dm^3/K/mol).

III.Questions carrying three marks

5.1 On a hot summer day, pressure in the well inflated tyre of an automobile increases considerably and the tyre may burst if the pressure is not adjusted properly. Name the gas law suitable for the above phenomenon .State the law & write the mathematical statement for the above law
5.2 Define an ideal gas. Derive ideal gas equation using gas laws

5.3. Enlist the three postulates of kinetic molecular theory of gases.
5.4. Define the terms
i) Critical temperature for a gas
ii) Aqueous tension
iii) Critical volume
5.5. What is the effect of temperature on -
i) Density ii) Surface tension iii) Vapour pressure of the liquid.
5.6 The mass of 0.5 dm^3 of hydrogen gas at a pressure of 1bar of Hg and at a temperature of 300K was found to be 4 x 10^{-2}g. Calculate the molar mass of hydrogen (R= 0.083 bar dm^3/K/mol).
5.7 2.9g of a gas at 95ºC occupies the same volume as 0.184 g of dihydrogen at 17 ºC at the same pressure . What is the molar mass of the gas?
5.8 Calculate the total pressure of mixture of 8g of oxygen and 4g of hydrogen confined in a vessel of 1dm^3 at 27ºC (R=0.083 bar dm^3/K/mol).
5.9 Calculate volume occupied by 8.8gof CO_2 at 31.1^{0}C&1bar pressure
(R =0.083 bar LK^{-1} mol^{-1}).

Answers

I. Answers for Questions carrying 1 mark Q. NO	Value point	Marks
5.1	Attractive intermolecular forces	1
5.2	Dipole- dipole forces	1
5.3	Polar molecule &non-polar molecule	1
5.4	Due to the presence of hydrogen bonding between the molecules in HF.	1
5.5	Intermolecular forces are very strong or Particles in liquids &solids are very close to each other	1
5.6	At constant temperature, the pressure of fixed amount of a gas varies inversely with its volume	1
5.7	Reduces to half of the initial pressure	1

I PUC – CHEMISTRY
CHAPTER - 06

Thermodynamics

One mark questions

1. Define System.
2. Define surroundings.
3. What is an open system? Give one example.
4. What is closed system? Give one example.
5. What is an isolated system? Give one example.
6. What is an extensive property? Give example.
7. What is an intensive property? Give example.
8. What is isothermal process?
9. What is an adiabatic process?
10. Write an expression for work done during an isothermal reversible expansion of an ideal gas.
11. State I law of Thermodynamics.
12. Write mathematical expression to represent I law.
13. Define Enthalpy.
14. Write mathematical statement of enthalpy.
15. What is an exothermic reaction?
16. Give one example for exothermic reaction.
17. What is an endothermic reaction?
18. Give one example for endothermic reaction.
19. Define heat of a reaction (ΔH_r^o)
20. Define standard enthalpy of formation.
21. State Hess's law of constant heat summation.
22. What is spontaneous process?
23. Give one example for spontaneous reaction.
24. Define entropy.
25. Write the S.I.Unit of entropy.
26. State second law of thermodynamics.
27. Define free energy.
28. Write mathematical statement of free energy.
29. Define standard free energy change for a reaction.
30. Out of Diamond and graphite which one has zero value of standard free energy of formation.
31. What is the change in entropy when sugar is dissolved in water?
32. What happens to entropy when water freezes?
33. Give the mathematical form of Gibbs-Helmholtz equation.
34. What is the state of a chemical reaction when i) $\Delta G = 0$ ii) $\Delta G > 0$ iii) $\Delta G < 0$

35. Mention the sign of ΔH for the formation of explosives.

36. Write the thermochemical equation for combustion of Benzene(l)

37. Write the mathematical statement of II law of thermodynamics.

2

38. How is change in enthalpy related to change in internal energy?

39. What is meant by free expansion?

40. State III Law of thermodynamics.

41. Define reversible process.

42. Define irreversible process.

43. Define internal energy.

44. Give an expression for the work done during an adiabatic process.

45. Define the term state function.

46. What is state of a system?

47. What is standard state of a system?

48. Define Enthalpy of combustion.

49. How is standard free energy change of a reaction is related to equilibrium constant?

50. What is the change in internal energy of a system if 10 joules of heat is supplied to it and 15 joules of work is done by it?

51. Give a chemical reaction for which $\Delta H = \Delta U$.

52. A closed thermos flask containing hot coffee represents what type of a system.

53. What is the relationship between ΔH and ΔU for the reaction.

$PCl_5(g)$ $PCl_3(g) + Cl_2(g)$

54. Choose the correct answer:

a) A thermodynamic state function is a quantity

i) used to determine heat changes

ii) whose value is independent of path

iii) used to determine pressure, volume work

iv) whose value depends on temperature only

b) For the process to occur under adiabatic conditions, the correct condition is,

i) $\Delta T = 0$ ii) $\Delta p = 0$ iii) $q = 0$ iv) $w = 0$

c) The enthalpies of all elements in their standard states are

i) unity ii) zero iii) < 0 iv) different for each element

d) ΔU

Θ

of combustion of methane is –XKJ mol^{-1}. The value of ΔH

Θ is

i) = ΔU

Θ

ii) $>\Delta U$

Θ

iii) <ΔU

Θ

iv) 0

e) The enthalpy of combustion of methane, graphite and dihydrogen at 298K are -890.3 kJ mol^{-1} , - 393.5 k J mol^{-1} and - 285.8 kJ mol^{-1} respectively. Enthalpy of formation of $CH_4(g)$ will be

i) -74.8 k J mol^{-1} ii) -52.27 k J mol^{-1}

iii) +74.8 k J mol^{-1} iv) + 52.26 k J mol^{-1}

f) A reaction , A + B → C + D + q is found to have a positive entropy change.

The reaction will be

i) possible at high temperature

ii) possible only at low temperature

iii) not possible at any temperature

iv) possible at any temperature

Two Marks questions

1. In a process, 701J of heat is absorbed by a system and 394J of work is done by the system. What is the change in internal energy for the process?

2. For the reaction; $2Cl(g) \rightarrow Cl_2(g)$; what will be the signs of ΔH and ΔS?

3

3. Define standard Enthalpy of vapourisation.

4. Write thermochemical equation for the vaporization of Ethanol.

5. Calculate the Enthalpy of vapourisation of Ethanol, given enthalpies of formation of liquid ethanol and gaseous Ethanol as -277.6kJ and -235.4kJ respectively.

6. Define standard enthalpy of fusion or molar enthalpy of fusion.

7. Define standard enthalpy of sublimation.

8. Define enthalpy of atomization ($\Delta_a H^\Theta$)

9. Define bond dissociation enthalpy.

10. For a equilibrium reaction.

A + 2B C ΔH = + 400kJ and ΔS is +200 J/K. Calculate the temperature above which the reaction becomes spontaneous

11. For $Cl_2 \rightarrow 2Cl$. Assign the signs for ΔH and ΔS

(g) (g)

12. Define enthalpy of a solution $\Delta_{sol} H^\Theta$

13. For the reaction

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$ Entropy change and enthalpy change at 298K are -549.4 J|K|mol and -1648 x 10_3 J | mol respectively. Calculate the free energy change for the reaction.

14. For an isolated system $\Delta U = 0$; what will be ΔS.

15. Comment on the thermodynamic stability of NO(g) and NO_2(g) given

½ N_2(g) + ½ O_2(g) → NO(g) ; $\Delta_f H^\ominus$ = 90 kJ mol^{-1}

NO(g) + ½ O_2(g) → NO_2(g) ; $\Delta_f H^\ominus$ = -74 kJ mol^{-1}

16. Given : N_2(g) + 3H_2(g) → 2NH_3(g) ; $\Delta_r H^\ominus$ = -92.4 kJ mol^{-1}
What is the standard enthalpy of formation of NH_3 gas.

17. Explain the spontaineity of Exothermic reactions using Gibb's equation.

18. Explain the spontaineity of endothermic reactions using Gibb's equation.

19. Calculate the entropy change in surroundings when 1.0 mol of $H_2O(l)$ is formed under standard conditions. Given $\Delta H^\ominus$ = -286 kJ mol^{-1}

20. Under what conditions $\Delta H < \Delta U$. Give an example.

21. For an isolated system $\Delta U = 0$; what will be ΔS?

22. For a reaction at 298 K 2A + B → C ΔH = 400 kJ mol^{-1} and ΔS = 0.2 kJ k^{-1} mol^{-1}.

At what temperature will the reaction become spontaneous considering ΔH and ΔS to be constant over the temperature range?

23. Calculate the heat of formation of Carbon-monoxide from the following data.

i) C(s) + O_2(g) → CO_2(g) ; $\Delta_f H$ = -393.5 kJ

ii) CO(g) + ½ O_2(g) → CO_2(g) ; $\Delta_f H$ = -282.8 kJ

24. Calculate the work done when a gas expands at a constant temperature from volume 2 x 10^{-3} m^3 to 4 x 10^{-3} m^3 against a constant pressure of 1.2 x 10^5 Nm^{-2}.

25. Define : specific heat capacity, Molar heat capacity.

26. Standard enthalpy of vapourisation of water at 373 K is 40.66 k J mol^{-1}. Calculate internal energy of vapourisation.

4

3 and 4 marks questions.

1. The reaction of cyanamide, NH_2CN(s) with oxygen was affected in a bomb calorimeter and ΔU was found to be -742.7 kJ mol^{-1} of cyanamide at 298K. Calculate the enthalpy change for the reaction at 298 K.

$NH_2CN(s) + 3/2\ O_2(g) \rightarrow N_2(g) + CO_2(g) + H_2O(l)$

2. Calculate the number of kJ necessary to raise the temperature of 60 g of aluminium from 35^o to 55^o C. Molar heat capacity of Al is 24 J mol^{-1} K^{-1}.

3. Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0^{o}C to ice at -

10.0^{o}C. $\Delta_{fus}H$ = 6.03 kJ mol^{-1} at 0^{o}C. Cp [$H_2O(l)$] = 75.3 J mol^{-1} k^{-1}.

Cp[$H_2O(s)$] = 36.8 J mol^{-1} k^{-1}.

4. Enthalpy of combustion of carbon to carbon-di-oxide is -393.5 k J mol^{-1}. Calculate the heat released upon formation of 35.2g of CO_2 from carbon and oxygen gas.

5. Calculate the enthalpy of the reaction

$N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$

Given that : Δ_fH CO(g) = -11- k J mol^{-1}

Δ_fH CO2(g) = -393 k J mol^{-1}

$\Delta_fHN_2O(g)$ = 81 k J mol^{-1}

Δ_fH $N_2O_4(g)$ = 9.7 k J mol^{-1}

6. Calculate the standard enthalpy of formation of $CH_3OH(l)$ from the following data.

i) $CH_3OH(l)$ +

2

3 $O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta_fH^{\ominus}$ = -726 kJ mol^{-1} → (i)

ii) $C(s) + O_2(g) \rightarrow CO_2(g)$; $\Delta_cH^{\ominus}$ = -393 kJ mol^{-1} → (ii)

iii) $H_2(g)$ + ½ $O_2(g) \rightarrow H_2O(l)$; $\Delta_fH^{\ominus}$ = -286 kJ mol^{-1} →(iii)

7. Calculate the enthalpy change for the process

$CCl_4(g) \rightarrow C(g) + 4\,Cl(g)$ and calculate bond enthalpy of C – Cl in $CCl_4(g)$

Given : $\Delta_{vap}H^{\ominus}(CCl_4)$ = 30.5 kJ mol^{-1}

$\Delta_fH^{\ominus}(CCl_4)$ = -135.5 kJ mol^{-1}

$\Delta_aH^{\ominus}$ (C) = 715.0 kJ mol^{-1} where $\Delta_aH^{\ominus}$ is enthalpy of atomisation

$\Delta_aH^{\ominus}(Cl_2)$ = 242 kJ mol^{-1}

8. For a reaction ; $2A(g) + B(g) \rightarrow 2D(g)$

ΔU_{298} = -10.5 kJ and $\Delta S^{\ominus}$ = -44.1 J k^{-1}.

Calculate ΔU_{298} for the reaction and predict whether the reaction is spontaneous or not.

9. The equilibrium constant for the reaction is 10. Calculate the value of $\Delta G^{\ominus}$; given
R = 8 J k^{-1} mol^{-1} ; T = 300 K

10. Explain the determination of ΔU using bomb caloriemeter.

11. Explain the determination of ΔH using caloriemeter.

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5

Thermodynamics

One mark Answers

1. It is the specific part of the universe in which energy changes are taking place.
2. Rest of the universe which surrounds the system.
3. A system is said to be open if both matter and energy can be exchanged with the surroundings. Example: Water kept in an open beaker.
4. A system is said to be closed if it exchanges only energy with the surroundings. Example. Water kept in a closed container.
5. A system is said to be isolated if it neither exchanges matter nor energy. Example: Coffee taken in a thermos flask.
6. Extensive properties of a system are the properties which depend upon the quantity of the matter present in the system. Example: Volume, internal energy, Enthalpy, heat capacity etc.
7. Intensive properties of a system are the properties which do not depend upon the quantity of the matter present in the system. Example: pressure, temperature, density, specific heat, surface tension etc.
8. A process is said to be isothermal if the temperature of the system remains constant. (dT = 0)
9. A process is said to be adiabatic if no heat exchange between the system and surroundings takes place. (dq = 0)
10. $W = -2.303\ nRT \log \frac{P_2}{P_1}$

$W = -2.303\ nRT \log \frac{V_2}{V_1}$

11. Energy can neither be created nor destroyed, it can be transformed from one form to another.

12. $\Delta U = q + w$, ΔU = change in internal energy
q = heat supplied
w = work done on the system
13. It is the sum of the internal energy and pressure volume energy.
14. H = U + PV H = Enthalpy
U = Internal energy
PV = pressure volume energy

15. A reaction in which heat energy is evolved is an exothermic reaction.

16. $C(s) + O_2(g) \_ CO_2(g)$; $\Delta H = -393.5$ k J.

17. A reaction in which heat energy is absorbed is an endothermic reaction.

18. $N_2(g) + O_2(g) \_ 2\ NO(g)$

19. The change in heat that takes place in a chemical reaction represented by balanced chemical equation.

20. The change in enthalpy that takes place when one mole of compound is formed from its constituent elements, at standard conditions. [298 K, 101.3 k.Pa]

6

21. The heat change that takes place in a chemical reaction is independent of time taken and number of intermediate steps involved.

22. A process that can take place on its own with or without initiation is called spontaneous process.

23. $Zn + H_2SO_4 \_ Zn\ SO_4 + H2 \uparrow$

24. Entropy is a measure of randomness or disorder of a system.

25. Joule / Kelvin / mole or $JK^{-1}\ mol^{-1}$

26. All spontaneous processes are thermodynamically irreversible

27. The amount of energy available for doing useful work under conditions of constant temperature and pressure.

28. G = H – TS. G = free energy
H = Enthalpy
S = Entropy
T = Temperature on Kelvin scale

29. The change in free energy that takes place when the reactants in their standard states are converted into product in their standard states at 298K and 101.3kPa.

30. Graphite.

31. Increases (or) positive

32. Decreases (or) negative

33. $\Delta G = \Delta H - T\Delta S$

34. i) Reaction attains equilibrium
ii) Reaction is non spontaneous
iii) Reaction is spontaneous

35. Positive

36. $C_6H_6(\ l)\ ) + (\ ) \rightarrow$

2
15

$_2 O\ g\ 6\ CO_2(g) + 3\ H_2O(\ l)\ \Delta H = -\ qkJ$

37. $\Delta S > 0$ or ΔS system + ΔS surrounding > 0

38. $\Delta H = \Delta U + \Delta_{(g)} nRT$

39. Expansion of a gas in vacuum is called free expansion.

40. Entropy of perfectly crystalline solid at absolute zero is zero.

41. A process which can be reversed at any instant of time by increasing the opposing force by an infinitesimal amount.
42. A process which is carried out rapidly so that the system does not get a chance to attain equilibrium.
43. The energy possessed by the system due to its nature, chemical composition and thermodynamic state.
44. $W = n C_v(T_2 – T_1)$ for 'n' moles of a gas.
45. The variables like temperature, pressure, volume etc, which define the state of a system are called state functions.
46. Condition of the system expressed by giving definite values for its properties such as temperature, pressure, volume.
47. The state of a system at 298K and 101.3 k.Pa is known as standard state of a system.
48. The change in enthalpy that takes place when one mole of a substance is completely burnt in air or oxygen at a given temperature.

49. $\Delta G^o = - 2.303 RT \log K_P$

50. $\Delta U = q – w$

7

$\Delta U = 10 – 15$

$\Delta U = - 5$ joules

51. $2 HI_{(g)}$ $H_2(g) + I_2(g)$
Or any other suitable example.
52. Isolated system.

53. $\Delta H = \Delta U + RT$

54. a) ii) is the correct answer
b) iii) is the correct answer
c) ii) is the correct answer
d) iii) is the correct answer
Explanation: $CH_4(g) + 2O_2(g)$ _ $CO_2(g) + 2H_2O(l)$

$\Delta_{ng} = 1 – 3 = -2$

$\Delta H^\ominus = \Delta U^\ominus + \Delta_{ng} RT = \Delta U^\ominus - 2RT$

$\therefore \Delta H^\ominus < \Delta U^\ominus$

e) i) is the correct answer
Explanation :

According to available data:

i) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta H^\ominus = -890.3$ k J mol^{-1}

ii) $C(s) + O_2(g) \rightarrow CO_2(g)$; $\Delta_c H^\ominus = -393.5$ k J mol^{-1}

iii) $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$; $\Delta_c H^\ominus = -285.8$ k J mol^{-1}

The equation we aim at

$C(s) + 2H_2(g) \rightarrow CH_4(g)$; $\Delta_f H^\ominus = ?$

Equation (ii) + 2 · equation (iii) – equation (i) and the correct $\Delta_f H^\ominus$ Value is

$= (-393.5) + 2 \cdot (-285.8) – (-890.3) = -74.8$ k J mol^{-1}

f) (iv) is the correct answer.

Two Marks Answers

1. Heat absorbed by the system, q = 701J
Work done by the system, w = -394J

Change in internal energy (ΔU) = q + w = 701 – 394
= 307J

2. ΔH: negative (-ve) because energy is released in bond formation.

ΔS : negative (-ve) because entropy decreases when atoms combine to form molecules.

3. Amount of heat required to vapourise one mole of a liquid at constant temperature and under standard pressure (1 bar) is called its standard enthalpy of vapourisation or molar enthalpy of vapourisation, $\Delta_{vap} H^\ominus$

4. $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$; $\Delta H = +$……..kJ

5. $\Delta H_{vap} = Hp – H_R$
= (-235.4 + 277.6)kJ
= + 42.2kJ
8

6. The enthalpy change that accompanies melting of one mole of a solid substance in standard state is called standard enthalpy of fusion or molar enthalpy of fusion.

$\Delta_{fus} H^\ominus$

7. Standard enthalpy of sublimation, $\Delta_{sub}H^{\ominus}$ is the change in enthalpy when one mole of a solid substance sublimes at a constant temperature and under standard pressure. (1 bar)

8. It is the enthalpy change on breaking one mole of bonds completely to obtain atoms in the gas phase.

9. It is the change in enthalpy when one mole of covalent bonds of the gaseous covalent compound is broken to form products in the gaseous phase.

10. $\Delta G = \Delta H - T\Delta S$

At equilibrium $\Delta G = 0$, $\Delta H = T\Delta S$

$T = \frac{200}{400\ 10^{-3}}$

$= \frac{\Delta H}{\Delta S}$

$T = 2000\ K$

11. ΔH = + ve , ΔS = + ve

12. It is the change in enthalpy when one mole of a substance is dissolved in a specified amount of a solvent.

13. $\Delta G = \Delta H - T\Delta S$

$\Delta G = -1648 \times 10^3 - [(298)(-549.4)]$

$= -1648000 + 163721.2$

$= -1484278.8\ J/mol = -1484.27\ kJ/mol$

14. Change in internal energy (ΔU) for an isolated system is zero because it does not exchange any energy with the surroundings. But entropy tends to increase in case of spontaneous reaction. Therefore, $\Delta S > 0$ or positive.

15. for NO(g) ; $\Delta_f H^{\ominus}$ = +ve : unstable in nature

For $NO_2(g)$; $\Delta_f H^{\ominus}$ = -ve : stable in nature.

16. $\Delta_f H^{\ominus} NH_3(g) = -(92.4) / 2 = -46.2\ kJ\ mol^{-1}$

17. Gibb's equation is $\Delta G = \Delta H - T\Delta S$ _ (1)

For exothermic reaction ΔH is –ve

i) If ΔS is +ve according to equation (1)

ΔG is –ve, reaction is spontaneous at all temperature.

ii) If ΔS is –ve, according to equation (1)

ΔG is –ve at low temperature such that

$T\Delta S < \Delta H$. Reaction is spontaneous.

18. Gibb's equation is $\Delta G = \Delta H - T\Delta S$ _ (1)

For endothermic reaction, ΔH is +ve

i) If ΔS is +ve according to equation (1)

ΔG is –ve, at high temperature

such that $T\Delta S > \Delta H$. Reaction is spontaneous

ii) If ΔS is –ve, according to equation (1)

ΔG is always +ve. Reaction is nonspontaneous at all temperature.

19. $q_{rev} = (-\Delta H^{\ominus}) = -286 \text{ kJ mol}^{-1} = 286000 \text{ J mol}^{-1}$

$\Delta S_{(surroundings)} = \frac{q_{rev}}{T} = \frac{(286000)}{298} = -959 \text{ J k}^{-1} \text{ mol}^{-1}$

20. $\Delta H < \Delta U$; if Δ_{ng} is negative

Example: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

$\Delta_{ng} = 2 - 4 = -2$

21. Change in internal energy (ΔU) for an isolated system is zero because it does not exchange any energy with the surroundings. But entropy tends to increase in case of

spontaneous reaction. Therefore, $\Delta S > 0$ or positive.

22. According to Gibbs – Helmholtz equation;

$\Delta G = \Delta H - T\Delta S$

For $\Delta G = O$; $\Delta H = T\Delta S$ or $T = \frac{\Delta H}{\Delta S}$

$T = \frac{(0.2)}{(400)}$

$= 2000\ k$

Thus, reaction will be in a state of equilibrium at 2000K and will be spontaneous above this temperature.

23. Required equation

$C(s) + ½\ O_2(g) \rightarrow CO(g)$; $\Delta_f H = ?$

Subtract equation (ii) from (i)

$C(s) + O_2(g) \rightarrow CO_2(g)$; $\Delta H = -393.5$ kJ

$CO(g) + ½\ O_2(g) \rightarrow CO_2(g)$; $\Delta H = -282.8$ kJ

(Subtract) $C(s) + ½\ O_2(g) - CO(g) \rightarrow \Delta H = (-393.5 + 282.8)$

Or $C(s) + ½\ O2(g) \rightarrow CO(g)$; $\Delta_f H = -110.7$kJ

24. $W = -P\Delta V$

$P = 1.2 \cdot 10^5\ Nm^{-2}$

$\Delta V = V_2 - V_1 = (4 \cdot 10^{-3}) - (2 \cdot 10^{-3}) - 2 \cdot 10^{-3}\ m^3$ $W = ?$

$= -1.2 \cdot 10^5\ Nm^{-2} \cdot 2 \cdot 10^{-3}\ m^3$

$= -2.4 \cdot 10^2\ Nm$

$= -2.4 \cdot 10^2\ J$ (since Nm = J)

Work done by the system $= -2.4 \cdot 10^2$ J

25. Specific heat capacity : It is defined as the quantity of heat required to raise the temperature of 1 gram of a substance by 1^oC or 1 K.

Molar heat capacity: It is defined as the quantity of heat required to raise the temperature of 1 mol of a substance by one degree celsius or one Kelvin.

26. $H_2O(l)$ $H_2O(g)$

$\Delta n = \eta_{gp} - \eta_{gR}$

$= 1 - 0$

$= 1$

$\Delta H = \Delta U + \Delta\eta_g RT$

$\Delta U = \Delta H - \Delta\eta_g RT$

$= 40.66 - (1) \cdot 8.314 \cdot 10^{-3} \cdot 373$

$= 40.66 - 3.1011$

$= 37.5589$ k J mol^{-1}

10

3 and 4 marks Answers

1. $\Delta U - 742.7$ kJ mol^{-1} ; $\Delta_{ng} = 2 - 3/2 = +$

2

1

mol

R= 8.314 x 10^{-3} kJ k^{-1} mol^{-1} ; T = 298 K

According to the relation, $\Delta H = \Delta U + \Delta_{ng} RT$

ΔH = (-742.7 kJ) + (½ mol) x (8.314 x 10^{-3} kJ k^{-1} mol^{-1}) x 298K

= - 742.7 kJ + 1.239 kJ = -741.5 kJ

2. No. of moles of Al(m) =

(27)

60

$^{-1}$ *gmol*

g

= 2.22 mol

Molar heat capacity (c) = 24 J mol^{-1} k^{-1}

Rise in temperature (ΔT) = 55 – 35 = 20^{o}C = 20 K

Heat evolved (q) = $c \cdot m \cdot T$

= (24 J mol^{-1} k^{-1}) $\cdot$ (2.22 mol) $\cdot$20 K

= 1065.6 J = 1.067 kJ

3. Ans: The change may be represented as

H_2O (l) (10^{o}C) $\xrightarrow{\Delta H}$ H_2O (s) (-10^{o}C)

$\downarrow \Delta H_1$ $\uparrow \Delta H_3$

H_2O(l) (0^{o}C) $\xrightarrow{\Delta}$ $_{H_2}H_2O$ (s) (0^{o}C)

According to Hess's law;

$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$

$\Delta H_1 = 75.3 \text{ J mol}^{-1} \text{ k}^{-1} (10K) = 753 \text{ J mol}^{-1}$

ΔH_2 (solidification) = - 6.03 k J mol^{-1} = -6030 J mol^{-1}
(sign changed)

$\Delta H_3 = 36.8 \text{ J mol}^{-1} \text{ k}^{-1} (-10K) – 36.8 \text{ J mol}^{-1}$

ΔH = (753 – 6030 – 368) J mol^{-1} = -5645 J mol^{-1}
= -5.645 k J mol^{-1}
4. Ans: The combustion equation is

$C(s) + O_2(g) \rightarrow CO_2(g)$; ΔcH = -393.5 k J mol^{-1}
(44g)
Heat released in the formation of 44 g of
CO2 = 393.5 k J
Heat released in the formation of 35.2 g of
CO_2 =
(44)
(393.5) (35.2)
g

kJ · *g*

= 314.8 kJ

5. Enthalpy of reaction (ΔrH)
= [81 + 3(-393)] – [9.7 + 3(-110)]
= (81 – 1179) – (9.7 – 330)
= - 778 k J mol^{-1}
6. The equation we aim at;
11

$C(s) + 2H_2(g) + ½ O_2(g) \rightarrow CH_3OH(l)$; $\Delta_f H^{\ominus} = \pm$?
Multiply equation (iii) by 2 and add to equation (ii)

$C(s) + 2H_2(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

$\Delta H = -(393 + 522) = -965 \text{ kJ mol}^{-1}$
Subtract equation (i) from equation (iv)
$CH_3OH(l)$ +
2

$3\ O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta H = -726 \text{ kJ mol}^{-1}$

Subtract $C(s) + 2H_2(g) + ½ O_2(g) \rightarrow CH_3OH(l)$; $\Delta_f H^{\ominus}$
= -239 kJ mol^{-1}
7. The available data is

i) $CCl_4(l) \rightarrow CCl_4(g)$ $\Delta_{vap}H^\ominus = 30.5$ kJ mol^{-1}

ii) $C(s) + 2Cl_2(g) \rightarrow CCl_4(l)$, $\Delta_fH^\ominus = -135.5$ kJ mol^{-1}

iii) $C(s) \rightarrow C(g)$, $\Delta_aH^\ominus = 715.0$ kJ mol^{-1}

iv) $Cl_2(g) \rightarrow 2Cl(g)$, $\Delta_aH^\ominus = 242$ kJ mol^{-1}

The equation we aim at is

$CCl_4(g) \rightarrow C(g) + 4Cl(g)$; $\Delta H = ?$

Equation (iii) + 2 x equation (iv) – equation (i) – equation (ii) gives the required equation with

$\Delta H = 715.0 + 2(242) – 30.5 – (-135.5)$ kJ mol^{-1}

$= 1304$ kJ mol^{-1}

Bond enthalpy of C-Cl in CCl_4 (average value)

$$= \frac{1304}{4}$$

$= 326$ kJ mol^{-1}

8. $\Delta H^\ominus = \Delta U^\ominus + \Delta_{ng}RT$

$\Delta U^\ominus = -10.5$ kJ ; $\Delta_{ng} = 2 – 3 = -1$ mol

$R = 8.314 \times 10^{-3}$ kJ mol^{-1} ; $T = 298$ K

$\therefore \Delta H^\ominus = (-10.5 \text{ kJ}) + [(-1\text{mol}) \times 8.314 \times 10^{-3} \text{ kJ k}^{-1}\text{mol}^{-1} \times (298\text{K})]$

$= -10.5$ kJ $– 2.478$ kJ $= -12.978$ kJ

According to Gibb's Hemholtz equation

$\Delta G^\ominus = \Delta H^\ominus$

- TΔS

Θ

ΔGΘ = (-12.978 kJ) – (298K) x (-0.0441 kJ k^{-1})

= - 12.978 + 13.142
= - 12.978 + 13.142
= 0.164 kJ

Since ΔG

Θ is positive, the reaction is non spontaneous in nature.

9. ΔG

Θ
= - RT in K = -2.303 RT log K
R = 8.0 J k^{-1} mol^{-1} ; T = 300 K ; k = 10

ΔG

Θ = - 2.303 x (8 J k^{-1} mol^{-1}) x (300K) x log 10
= - 5527 J mol^{-1} = - 5.527 kJ mol^{-1} .
10. Refer page 163 and 164 of the prescribed text book
11. Refer page 164 of the prescribed text book

SUBJECT : CHEMISTRY
CHAPTER-08 : REDOX REACTIONS
QUESTIONS CARRYING ONE MARK:

1. Define ‘oxidation’ in terms of electron transfer.
2. Give the electronic interpretation of ‘reduction’.
3. What is an oxidizing agent (or oxidant)?
4. Which is the most powerful oxidizing agent?
5. What is a reducing agent (or reductant)?
6. Which is the most powerful reducing agent?
7. Complete the following equation: $2Fe^{2+} + 2H^{+} + H_2O_2$ . ..... $+2H_2O$.
8. Define oxidation number. (or oxidation state).
9. Calculate the oxidation number of Cr in Cr_2O_7

$^{2-}$.
10. Calculate the oxidation number of Mn in $KMnO_4$.
11. What is the oxidation number (or oxidation state) of an element?
12. What happens to the oxidation number (O.N.) of an element during oxidation?
13. What happens to the oxidation number of an element during reduction?
14. What is the oxidation state of hydrogen in hydrides?
15. What is the oxidation state of oxygen in peroxides?
16. What is the oxidation state of P_4 ?.
17. What is an electrode?
18. What is electrode potential?
18. What is standard electrode potential?
19. Name the cell obtained by coupling a zinc electrode with a copper electrode.
20. Identify the oxidant in the following reaction: $H_2O_2 + O_3$ $H_2O + 2O_2$
21. What is the oxidation state of oxygen in OF_2?.

QUESTIONS CARRYING TWO MARKS:

1. What is a redox reaction? Give an example.
2. Justify the reaction: $H_2S + Cl_2$ $2HCl + S$ is a redox reaction.
3. Define oxidation and reduction in terms of oxygen and hydrogen. Give one example
for each.
4. What is oxidation number? What is the oxidation number(O.N) of Cl in $KClO_3$?
5. Define oxidation and reduction in terms of oxidation number.
6. How are the oxidizing agent and reducing agents defined in terms if oxidation number?
7. Write separate equations for the oxidation and reduction reactions occurring in the
following redox reaction: $2Fe + 2HCl$ $FeCl_2 + H_2$
8. For $2H_2O_2$ $2H_2O + O_2$
(1) (2) (3)
i) What is the oxidation number of Oxygen in (2)?
ii) What type of Redox reaction is it?
9. Explain whether the following reaction is a redox reaction or not:
$CaCO_3(s)$ $CaO(s) + CO_2(g)$
10 Calculate the oxidation number of: (i) S in H_2SO_4 (ii) P in H_3PO_4.
11. What is a redox couple? Identify the redox couples in the reaction:
$Zn(s) + 2Ag^+(aq)$ $Zn^{2+}(aq) + 2Ag(s)$
12 What is an electrochemical series?
13. What is a spectator ion? Give an example of a reaction involving such an ion.
14. Write the formula for the following compounds represented using Stock notation:
(a) Nickel (II) sulphate (b) Tin (IV) oxide
(c) Thallium (I) sulphate (d) Iron (III) sulphate
15. Using Stock notation, represent the following compounds: Fe_2O_3, CuO, MnO and

MnO_2
16. Calculate the oxidation number of phosphorus in the following species:
(a) HPO_3
$^{2-}$ and (b)PO_4
$^{3-}$
17. Balance the Redox reaction using oxidation number method :
$SO_2 + H_2S$ S + H_2O
18. Assign oxidation number to the underlined elements in each of the following species:(a) NaH_2PO_4 (b) $NaHSO_4$ (c) $H_4P_2O_7$ (d) K_2MnO_4
19. Justify that the following reactions are redox reactions:
(a) CuO(s) + H_2(g) Cu(s) + H_2O(g)
(b) Fe_2O_3(s) + 3CO(g) 2 Fe(s) + $3CO_2$(g)
19. Give an example of a redox combination reaction. Mention the species that undergo
oxidation and reduction.
20. Give an example of a redox decomposition reaction. Mention the species that undergo oxidation and reduction.
21. Give an example of a redox displacement reaction. Mention the species that undergo
oxidation and reduction.
22. Give an example of a redox disproportionation reaction. Mention the species that
undergo oxidation and reduction.
23. F_2 does not undergo disproportionation. Why?
24. What type of redox reactions are the following?
(a) 3Mg(s) + N_2(g) _ Mg_3N_2(s)
(b) $2KClO_3$(s) _ 2KCl(s) + $3O_2$(g)
(c) Cr_2O_3 (s) + 2 Al (s) _ Al_2O_3 (s) + 2Cr(s)
(d) (d) $2NO_2$(g) + $2OH^-$(aq) _ NO_2
$^-$(aq) +NO_3
$^-$ (aq)+H_2O(l)
25. Name the redox indicator used in the titration of
(i). KMnO4 v/s FAS.(or $H_2C_2O_4$).
(ii) $Na_2S_2O_3$ v/s I_2.

QUESTIONS CARRYING THREE MARKS:

1. When blue coloured solution of copper sulphate is stirred with a zinc rod, the blue colour of the solution fades off and the zinc rod is coated with reddish copper metal.
Write the chemical reaction taking place in the above observation and identify the species undergoing oxidation and reduction.
2. A solution of silver nitrate turns blue slowly on stirring with a copper rod which in turn
gets coated with a white deposit of silver. Write a chemical reaction for this observation
and identify the oxidizing and reducing agents in it.

3. Balance the following equations by the oxidation number method.(3marks each)
(i) $Fe^{2+} + H^{+} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$
(ii)$MnO_4^{-}(aq) + I^{-}(aq) \rightarrow MnO_2(s) + I_2(s)$ (in basic medium)
(iii) $MnO_4^{-}(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^{-}(aq)$ (in acidic solution)
(iv) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$ (in acidic solution)
(v) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$ (in acidic solution)
4. Balance the following equations by half reaction method (ion-electron method). (3 marks each)
(a) $MnO_4^{-}(aq) + SO_2(g) \_ Mn^{2+}(aq) + HSO_4^{-}(aq)$ (in acidic solution)
(b) $MnO_4^{-}(aq) + I^{-}(aq) \_ MnO_2(s) + I_2(s)$ (in basic medium)
(c) $H_2O_2(aq) + Fe^{2+}(aq) \_ Fe^{3+}(aq) + H_2O(l)$ (in acidic solution)
(d) $Cr_2O_7^{2-} + SO_2(g) \_ Cr^{3+}(aq) + SO_4^{2-}(aq)$ (in acidic solution)
5. In the reactions given below, identify the species undergoing oxidation and reduction:
(i) $H_2S(g) + Cl_2(g \; 2\,HCl(g) + S(s)$
(ii) $2\,Na(s) + H_2(g) \; 2\,NaH(s)$
(iii) $2Fe(s) + 2HCl(aq) \; FeCl_2(aq) + H_2(g)$
\
6. Justify that the reaction: $2Cu_2O(s) + Cu_2S(s) \; 6Cu(s) + SO_2(g)$ is a redox reaction. Identify the species oxidized/reduced, which acts as an oxidant and which acts
as a reductant.

CHAPTER-08 : REDOX REACTIONS
ANSWERS:
QUESTIONS CARRYING ONE MARK:
1. Loss of electron(s) by any species is called oxidation.
2. Gain of electron(s) by any species is called reduction.
3. An oxidizing agent (or an oxidant) is an acceptor of electron(s).
4. Fluorine (F_2).
5. A reducing agent(or a reductant) is a donor of electron(s).
6. Lithium (Li).
7. $2Fe^{2+} + 2H^{+} + H_2O_2 \; . \; \mathbf{2Fe^{3+}} + 2H_2O$.
8. The term Oxidation number denotes the oxidation state of an element in a compound
ascertained according to a set of rules formulated on the basis that electron pair in a

covalent bond belongs entirely to more electronegative element.
9. Oxidation number of oxygen = -2.
Hence, oxidation number of Cr, (x) in $Cr_2O_7^{2-}$ = 2x + 7x(-2) = 0, x = +6
10. Oxidation number of K = +1, oxygen, O = -2.
Hence, oxidation number of Mn, (x) in $KMnO_4$ = (+1) + x + 4(-2) = 0, x = +7
11. Zero.
12. It increases.
13. It decreases
14. In Hydrides, hydrogen has an oxidation state of -1.
15. In peroxides, oxygen has an oxidation state of -1.
16. Zero.
17. A setup consisting of a metal in contact with its salt solution is called an electrode.
18. The potential attained by a metal in contact with a solution containing its own ions is
called electrode potential.
19. The potential attained by a metal in contact with its salt solution of concentration 1
moldm^{-3} at 298 K.
20. The oxidant is O_3.
21. +2
QUESTIONS CARRYING TWO MARKS:
ANSWERS:
1. A chemical reaction in which both oxidation and reduction are taking place simultaneously is called a redox reaction.
reduction
Ex: $Zn(s) + Cu^{2+} \text{--}\_ Zn^{2+} + Cu$.
Oxidation
2. : $H_2S^{(-2)} + Cl_2^{(0)} \text{-}\_ 2HCl^{(-1)} + S^{(0)}$
The O.N. of S increases from -2 to 0. So it is undergoing oxidation.
The O.N. of Cl_2 decreases from 0 to -1. So it is undergoing reduction.
Therefore it is a redox reaction.
3. **Oxidation**: Addition of oxygen or removal of hydrogen.
Ex: $2\ Mg + O_2 \_\ 2\ MgO$ (Addition of oxygen to Mg)
, $2\ H_2S + O_2 \_\ 2\ S + 2\ H_2O$ (Removal of hydrogen from H_2S)
Reduction: Addition of hydrogen or removal of oxygen.
Ex: $H_2C = CH_2 + H_2 \_\ H_3C - CH_3$. (Addition of hydrogen to ethene)
$2\ HgO \_\ 2\ Hg + O_2$ (Removal of oxygen from HgO)
4. Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that electron pair in a
covalent bond belongs entirely to more electronegative element.

Let the O.N of Cl in $KClO_3$ be x.
O.N. of K = +1, O = -2 . ∴ O.N of Cl in $KClO_3$ = 1+ x + 3(-2) = +5.
5. In terms of oxidation number,
Oxidation: An increase in the oxidation number of an element in a given substance.
Reduction: A decrease in the oxidation number of an element in a given substance.
6. Oxidising agent: A reagent which can increase the oxidation number of an element
in a given substance. These reagents are also called as oxidants.
Reducing agent: A reagent which lowers the oxidation number of an element in a given
substance. These reagents are also called as reductants.
7. $2Fe + 2HCl \_ FeCl_2 + H_2$
Oxidation reaction: $Fe \_ Fe^{2+} + 2e^-$
Reduction reaction: $2HCl + 2e^- \_ H_2$
8. $2\overset{(-1)}{H_2O_2} \_ 2\overset{(-2)}{H_2O} + \overset{(0)}{O_2}$
(1) (2) (3)
(i) The O.N. of oxygen in (2) is -2.
(ii) It is a disproportionation redox reaction (∵ oxygen undergoes both oxidation and reduction.)
9. +2 + 4 –2 +2 –2 +4 –2
CaCO3 (s) _ CaO(s) + CO2(g)
It is not a redox reaction because the oxidation number of no element changes.
10. (i) Let the O.N. of S be ' x'
O.N. of H = +1, O = -2 ∴ O.N. of S in H_2SO_4 = 2(+1) + x + 4(-2) = +6.
(ii) Let the O.N. of P be 'x'.
O.N. of H = +1, O = -2 ∴ O.N. of P in H_3PO_4 = 3(+1)+x+4(-2) = +5.
11. A redox couple is defined as having together the oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction.
The redox couples in the reaction are, Zn^{2+}/ Zn(s) and Ag^+/Ag.
12. A series of electrode potential values arranged in the increasing or decreasing order
constitute an electrochemical series.
13. An ion which is present in a redox reaction, but does not take part in a reaction during
electron transfer is called a spectator ion.
Ex: SO_4^{2-} ion in the reaction: $Zn(s) + CuSO_4(aq) \_ ZnSO_4(aq) + Cu$.
14. (a) Ni(II)SO_4 (b) Sn(IV)O_2
(c) Tl_2(I)SO4 (d) Fe_2(III)$(SO_4)_3$
15. Fe_2O_3 - Fe_2(III)O_3 , CuO – Cu(II)O
MnO - Mn(II)O, MnO_2 – Mn(IV)O_2.
16. (a) Let the O.N of P in HPO_3^{2-} be x.
(+1) + x + 3(-2) = -2
∴ x = +3

(b)Let the O.N of P in PO_4^{3-} be x.
X + 4(-2) = -3
∴ x = +5
17. Step 1: Write skeletal equation with O.N of each element.
4+ 2- 0
$SO_2 + H_2S$ _ S + H_2O.
Step 2: Multiply H_2S by 2 to equalize the oxidation numbers on either side of the equation.
4+ 2 x (2-) 0
SO_2 +2 H_2S _ S + H_2O.
Step 3: Now, balance S atoms on RHS.
$SO_2 + 2H_2S$ _ 3 S + H_2O.
Step 4: Finally balance H and O atoms to get a balanced equation.
$SO_2 + 2H_2S$ _ 3S + $2H_2O$
18. (a) NaH_2PO_4 : O.N. of P = (+1)+2(+1)+x+4(-2); x = +5.
(b) $NaHSO_4$: O.N of S = (+1) + (+1) + x + 4(-2); x = +6
(c) $H_4P_2O_7$: O.N. of P = 4(+1) + 2x + 7(-2) ; x = +5
(d) K_2MnO_4: O.N. of Mn = 2(+1) + X + 4(-2): x = + 7
(Taking O.N. of H=+1, Na = +1, K = +1, O = -2.).
19. Example for Redox combination reaction:
0 0 +4 –2
C(s) + O_2 (g) _ CO_2(g)
In this reaction, the O.N. of ‘C’ increases from 0 to +4. So it is undergoing oxidation.
the O.N. of ‘O’ decreases from 0 to -2. So it is undergoing reduction.
20. Example for Redox decomposition reaction:
-1+1 0 0
2NaH(s) _ 2Na(s) + H_2(g)
In this reaction, the O.N. of ‘Na’ increases from -1 to 0. So it is undergoing oxidation.
the O.N. of ‘H’ decreases from +1 to 0. So it is undergoing reduction.
21. Example for Redox displacement reaction:
+2 +6 -2 0 0 +2 +6 –2
$CuSO_4$(aq) + Zn (s) _ Cu(s) + $ZnSO_4$ (aq)
In this reaction, the O.N. of ‘Zn’ increases from 0 to +2. So it is undergoing oxidation.
the O.N. of ‘Cu’ decreases from +2 to 0. So it is undergoing reduction.
22. Example for Redox disproportionation reaction:
+1 –1 +1 –2 0
$2H_2O_2$ (aq) _ $2H_2O$(l) + O_2(g)
In this reaction, the O.N. of ‘O’ increases from -1 to 0 as well as decreases from -1 to -2.
So oxygen is undergoing both oxidation and reduction(disproportionation).
23. Among halogens, fluorine (F_2) is the most electronegative element; it cannot exhibit any
positive oxidation state. Hence it does not show a disproportionation tendency.
24. (a) $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$ - Redox combination reaction
(b) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ - Redox decomposition reaction

(c) Cr_2O_3 (s) + 2 Al (s) → Al_2O_3 (s) + 2Cr(s) - Redox displacement reaction
(d) $2NO_2(g) + 2OH^-(aq) \rightarrow NO_2^-(aq) + NO_3^-(aq) + H_2O(l)$ - Redox disproportionation reaction.
25. (i) MnO_4^- ion itself act as a self indicator
(ii) Starch.
QUESTIONS CARRYING THREE MARKS:
Answers:
reduction
1. $Zn(s) + Cu^{2+} \_ Zn^{2+} + Cu$.
Oxidation
In this reaction, Zn loses $2e^-$ to Cu and hence is undergoing oxidation; Cu^{2+} is undergoing reduction to Cu.
2. reduction
$Cu(s) + 2Ag^+(aq)\ Cu^{2+}(aq) + 2Ag(s)$
Oxidation
In this reaction, Cu is giving two electrons to Ag+ and so it is a reducing agent. Ag+, in turn, is accepting the electrons from Cu to undergo reduction and so it is an oxidizing agent.

BALANCING EQUATIONS BY OXIDATION NUMBER METHOD

3. (i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$
. Step 1: Write skeletal equation with O.N of each element.
oxidation
2+ 6+ 3+ 3+
$Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$
reduction
Step 2: Multiply Cr^{3+} by 2 and Fe^{2+} and Fe^{3+} by 6 to equalize the oxidation numbers on
either side of the equation.
2+ 1+ 6+ 2x3+ 3+
$6Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 6\ Fe^{3+} + H_2O$
Step 3: Now, balance O atoms on RHS by adding $7H_2O$
$6Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 6Fe^{3+} + 7H_2O$.
Step 4: Finally balance H atoms by adding $14H^+$ on LHS to get a balanced equation as:
$6\ Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 6Fe^{3+} + 7H_2O$
3. (ii) $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$ (in basic medium)
. Step 1: Write skeletal equation with O.N of each element Undergoing change in

oxidation number.
Oxidation
7+ 1- 6+ 0
$MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$
reduction
Step 2: Multiply I^- by 6 and MnO_4^- by 2 to equalize the oxidation numbers on either side of the equation.
2 x (7+) 6 x (1-) 2 x (4+) 0
$2MnO_4^-(aq) + 6I^-(aq) \rightarrow 2MnO_2(s) + 3I_2(s)$
Step 3: Now, add 8 OH^- on RHS to balance –ve charges on either side.
$2MnO_4^-(aq) + 6I^-(aq) \rightarrow 2MnO_2(s) + 3I_2(s) + 8OH^-$
Step 4: Finally balance H and O atoms by adding $4H_2O$ on LHS to get a balanced equation as:
$2MnO_4^-(aq) + 6I^-(aq) + 4H_2O \rightarrow 2MnO_2(s) + 3I_2(s) + 8OH^-$
3. (iii) $MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$ (in acidic solution)
. Step 1: Write skeletal equation with O.N of each element undergoing change in oxidation number.
Oxidation-2e-
7+ 4+ 2+ 6+
$MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$
reduction-5e-
Step 2: Multiply SO_2 by 5 and MnO_4^- by 2 to balance +ve charges on both sides.
2 x (7+) 5 x (4+) 2 x (2+) 5 x (6+)
$2MnO_4^-(aq) + 5SO_2(g) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$
Step 3: Now, add 2H2O and H+ on LHS to balance oxygen atoms
$2\ MnO_4^-(aq) + 5\ SO_2(g) + 2H2O \rightarrow 2Mn^{2+}(aq) + 5\ HSO_4^-(aq)$
Step 4: Finally add H^+ on LHS to get a balanced equation as:
$2\ MnO_4^-$

(aq) + $5SO_2$ (g) +$2H_2O$ +H^+ → $2Mn^{2+}$ (aq) + $5HSO_4$
$^-$ (aq)
3. (iv) H_2O_2 (aq) + Fe^{2+} (aq) →Fe^{3+} (aq) + H_2O (l) (in acidic solution)
. Step 1: Write skeletal equation with O.N of each element undergoing change in oxidation number
Oxidation-1e1-
2+ 3+ 2-
H_2O_2 (aq) + Fe^{2+} (aq) →Fe^{3+} (aq) + H_2O (l))
reduction-2 x 1e-
Step 2: Since the number of charges on both sides are not equal, $2Fe^{2+}$ on LHS and $2Fe^{3+}$ on RHS
2 x (1-) 2x(2+) 2x (3+) (2-)
H_2O_2 (aq) +2 Fe^{2+} (aq) → $2Fe^{3+}$ (aq) + H_2O (l))
Step 3: Now, put $2H_2O$ to balance 'O' atoms.
H_2O_2 (aq) + Fe^{2+} (aq) → Fe^{3+} (aq) +2 H_2O (l))
Step 4: Finally add2 H^+ on LHS to get a balanced equation as:
H_2O_2 (aq) +2 Fe^{2+} (aq) +$2H^+$(aq) → $2Fe^{3+}$ (aq) +2 H_2O (l))
3.(v) $Cr_2O_7^{2-}$ + SO_2(g) → Cr^{3+} (aq) + SO_4
$^{2-}$ (aq) (in acidic solution)
Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number.
Oxidation-2e-
2 x (6+) 4+ 3+ 6+
$Cr_2O_7^{2-}$ + SO_2(g) → Cr^{3+} (aq) + SO_4
$^{2-}$ (aq)
reduction-3e-
Step 2: Multiply SO2 by 3 and Cr3+ by 2 on RHS .
$Cr_2O_7^{2-}$ +3 SO_2(g) → $2Cr^{3+}$ (aq) + 3 SO_4
$^{2-}$ (aq)
Step 3: Balance charges by adding 2H+ on LHS
$Cr_2O_7^{2-}$ +3 SO_2(g) +$2H^+$ → $2Cr^{3+}$ (aq) + 3 SO_4
$^{2-}$ (aq)
Step 4: Finally add H_2O on RHS to get a balanced equation as:
$Cr_2O_7^{2-}$ + SO_2(g) → $2Cr^{3+}$ (aq) + SO_4
$^{2-}$ (aq)+ H_2O(l)

BALANCING EQUATIONS BY ION-ELECTRON METHOD

4. **(a) MnO_4**
$^-$ (aq) + SO_2 (g) _Mn^{2+} (aq) + HSO_4
$^-$ (aq) (in acidic solution)
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.
oxidation-2e-
MnO_4
$^-$ (aq) + SO_2 (g) → Mn^{2+} (aq) + HSO_4-(aq)
reduction-5e-
Step2: Write out oxidation and reduction separately and balance the atoms other than H

and O.
Oxidation half reaction: $SO2 \rightarrow HSO_4^-$
Reduction half reaction: $MnO_4^- \rightarrow Mn^{2+}$
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction
by the extent of oxidation and add.
Oxidation half reaction: $[SO_2 \rightarrow HSO_4^-]$ x5
Reduction half reaction: $[MnO_4^- \rightarrow Mn^{2+}]$ x2
$2MnO_4^- + 5SO_2 \rightarrow 2Mn^{2+} + 5HSO_4^-$
Step4: Add H^+ and $2H_2O$ on LHS to balance H and O atoms in the acid medium to get a
balanced equation.
$2MnO_4^- + 5SO_2 + H^+ + 2H_2O \rightarrow 2Mn^{2+} + 5HSO_4^-$.
4. (b) $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$ (in basic medium)
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.
oxidation-2e-
$MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$
Step2: Write out oxidation and reduction separately and balance the
atoms other than H and O.
Oxidation half reaction: $2I^- \rightarrow I_2$
Reduction half reaction: $MnO_4^- \rightarrow MnO_2$
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction
by the extent of oxidation and add.
Oxidation half reaction: $[2I^- \rightarrow I_2]$ x3
Reduction half reaction: $[MnO_4^- \rightarrow MnO_2]$ x2
$2MnO_4^- + 6I^- \rightarrow MnO_2 + 3I_2$
Step4: Add $4OH^-$ on RHS and $2H_2O$ on LHS to balance H and O atoms in the basic medium to
get a balanced equation.
$2MnO_4^- + 6I^- + 4H_2O \rightarrow 2MnO_2 + 3I_2 + 8OH^-$
4. (c) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$ (in acidic solution)
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.
oxidation-1e-
$H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$
reduction-2e-
Step2: Write out oxidation and reduction separately and balance the atoms other than H
and O.

Oxidation half reaction: $Fe^{2+} \rightarrow Fe^{3+}$
Reduction half reaction: $H_2O_2 \rightarrow H_2O$
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction
by the extent of oxidation and add.
Oxidation half reaction: [$Fe^{2+} \rightarrow Fe^{3+}$] x 2
Reduction half reaction: [$H_2O_2 \rightarrow H_2O$] x 1
$2Fe^{2+}$ + H2O2 → $2Fe^{3+}$ + H2O
Step4: Add 2H+ on LHS and H_2O on RHS to balance H and O atoms in the acid medium to get
a balanced equation.
$2Fe^{2+}$ + H2O2 + $2H^+ \rightarrow 2Fe^{3+}$ +2 H2O.
4.(d) $Cr_2O_7^{2-}$(aq)+ SO_2(g) →Cr^{3+} (aq) + SO_4^{2-} (aq) (in acidic solution)
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.
oxidation-2e-
$Cr_2O_7^{2-}$ + SO_2(g) →Cr^{3+} (aq) + SO_4^{2-} (aq)
reduction-2x3e-
Step2: Write out oxidation and reduction separately and balance the atoms other than H
and O.
Oxidation half reaction: $SO_2 \rightarrow SO_4^{2-}$
Reduction half reaction: $Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction
by the extent of oxidation and add.
Oxidation half reaction: [SO2 → SO_4^{2-}] x6
Reduction half reaction: [$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$] x2
$2Cr_2O_7^{2-}$ +6 $SO_2 \rightarrow 4Cr^{3+}$ + $6SO_4^{2-}$
Step4: Add H^+ and $2H_2O$ on LHS to balance H and O atoms in the acid medium to get a
balanced equation.
$2Cr_2O_7^{2-}$(aq) + 6 SO_2(g) +$4H^+ \rightarrow 4\ Cr^{3+}$ (aq) +6 SO_4^{2-} (aq)+$2H_2O$.
OR, $Cr_2O_7^{2-}$(aq) + 3 SO_2(g) +$2H^+ \rightarrow 2\ Cr^{3+}$ (aq) +$3SO_4^{2-}$ (aq)+H_2O.
4.(e) **MnO_4^- + $C_2O_4^{2-}$ → Mn^{2+} + CO2 in acid medium**
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.
oxidation-2x1e-
MnO_4^- + $C_2O_4^{2-} \rightarrow Mn^{2+}$ + CO_2

reduction-5e-
Step2: Write out oxidation and reduction separately and balance the atoms other than H
and O.
Oxidation half reaction: $C_2O_4^{2-} \rightarrow 2CO_2$
Reduction half reaction: $MnO_4^- \rightarrow Mn^{2+}$
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction
by the extent of oxidation and add.
Oxidation half reaction: $[C_2O_4^{2-} \rightarrow 2CO_2]$ x5
Reduction half reaction: $[MnO_4^- \rightarrow Mn^{2+}]$ x2
$2MnO_4^- + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2$
Step4: Add required number H^+ on LHS and H_2O on RHS to balance H and O atoms in the
acid medium to get a balanced equation.
$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$
5. **(i) H_2S (g) + Cl_2 (g 2 HCl (g) + S (s)**
In this reaction, the species undergoing oxidation is: H_2S
(∵ the O.N. of S in H_2S increases from -2 to 0)
The species undergoing reduction is: Cl_2 (∵ the O.N. of Cl decreases from 0 to -1)
5.(ii) 2 Na (s) + H_2 (g) 2 NaH (s)
In this reaction, the species undergoing oxidation is: Na (∵ the O.N. of Na increases from 0 to +1).
The species undergoing reduction is: H_2 (∵ the O.N. of H_2 decreases from 0 to -1)
5.(iii) 2Fe(s) + 2HCl(aq) $FeCl_2$(aq) + H_2(g)
In this reaction, the species undergoing oxidation is: Fe (∵ the O.N. of Fe increases from 0 to +2)
The species undergoing reduction is: HCl (∵ the O.N. of H in HCl decreases from +1 to 0)
6. The reaction: $2Cu_2O(s) + Cu_2S(s) \rightarrow 6Cu(s) + SO_2(g)$ is a redox reaction because, in Cu_2O, Cu is in +1 oxidation state. It is reduced to Cu in which the oxidation state is 0.
In Cu_2S, S is in -2 oxidation state, which is oxidized to +4 oxidation state in SO_2.
The oxidizing agent(oxidant) is Cu(I) in Cu_2O
The reducing agent(reductant) is sulphur of Cu2S.

1

I P U C QUESTION BANK

SUBJECT:- CHEMISTRY, UNIT – 9 : HYDROGEN

Questions carrying one mark:

1. Which is the most abundant element in the universe?

2. Name the isotope of hydrogen that do not contain neutron.

3. What is the composition of water gas?

4. What is the chemical used in clarke's process to remove the temporary hardness of
water?

5. What volume of oxygen is produced by one litre of "10 volume" H_2O_2 at STP?

6. Name the isotope of hydrogen containing two neutrons.

7. What is the role of heavy water in a nuclear reactor?

8. What is a "syn gas"?

9. Give an example of an ionic hydride.

10. Give an example of a covalent hydride.

11. What is "demineralised water "?

12. Arrange LiH, NaH, and CsH in the increasing order of ionic character.

13. Arrange H_2, D_2, T_2 in the increasing order of their boiling points.

14. Which isotope of hydrogen is radioactive?

15. What causes temporary hardness of water?

16. What causes permanent hardness of water?

17. Why H_2O has higher boiling point compared to H_2S?

18. What are Non-stoichiometric hydrides?

19. Why is H_2O_2 not stored in glass containers?

20. What is "calgon"?

21. What is the chemical name of zeolite used in softening of hard water?

22. Out of ice and water, which has low density?

2

23. What is the bond angle in H_2O molecule?

24. Name the gas liberated by the reaction of zinc with aqueous NaOH solution.

25. Complete the following reaction

$$Mg(HCO_3)_2\,(aq) \xrightarrow[Boil]{} \_\_\_\_\_ + 2\ CO_2.$$

Questions carrying Two / Three marks:

26. Justify the position of hydrogen in the periodic table.

27. Write the names of isotopes of hydrogen. What is the mass ratio of these isotopes?

28. Why does hydrogen occur in a diatomic form rather than in a monoatomic form
under normal conditions?
29. How can the production of dihydrogen obtained from 'coal gasification' be increased?
30. Describe the bulk preparation of dihydrogen by electrolytic method. What is the role
of electrolyte in this process?
31. Complete the following reactions:

i) $H_2(g) + M_mO_o(s) \longrightarrow$

Δ

ii) $CO(g) + H_2(g)$

Δ

$\rightarrow$
Catalyst

iii) $C_3H_8(g) + 3\ H_2O(g)$

Δ

$\rightarrow$
Catalyst

iv) $Zn(s) + NaOH\ (aq) \longrightarrow$
Heat

32. Discuss the consequences of high enthalpy of H-H bond in terms of chemical
reactivity of dihydrogen.
33. What do you understand by (i) electron- deficient (ii) electron-precise (iii) electron-rich compounds of hydrogen? Provide justification with suitable examples.
3
34. How do you expect the metallic hydrides to be useful for hydrogen storage? Explain.
35. How does the atomic hydrogen or oxy hydrogen torch function for cutting and
welding purposes? Explain.
36. Among NH_3, H_2O, and HF, which has highest magnitude of hydrogen bonding and
why?
37. What is "auto-protolysis "of water? Mention its significance.
38. Consider the reaction of water with F_2 and suggest, in terms of oxidation and reduction, which species are oxidized/reduced?
39. Complete the following chemical reactions:

i) PbS (s) + H_2O_2 (aq) _

ii) Mn –

O_4 (aq) + H_2O(aq) _

iii) CaO (s) + H_2O (g) _

iv) $AlCl_3$(s) + H_2O (l) _

v) Ca_3N_2 (s) + H_2O _

40. Write the chemical reactions to show amphoteric nature of water.

41. Discuss the principle and method of softening of hard water by synthetic ion exchange resins.

42. Write chemical reactions to justify that H_2O_2 can function as an oxidizing as well as
reducing agent?

43. What properties of water make it useful as a solvent? What type of compounds can
it (i) dissolve and (ii) hydrolyse?

44. How can saline hydrides remove traces of water from organic compounds?

45. What is the difference between the terms hydrolysis and hydration?

46. Mention any two uses of dihydrogen.

47. Calculate the strength of '10 volume' solution of hydrogen peroxide.

4

48. With an example explain the oxidizing property of water.

49. Explain the reducing property of water with an example.

50. With equations explain the action of washing soda on hard water in removing it's
permanent hardness.

UNIT-9. HYDROGEN. Model answers.

Questions carrying one mark

1. Hydrogen

2. Protium—$_1H_1$

3. A mixture of carbon monoxide and hydrogen or CO + H_2

4. Lime or $Ca(OH)_2$ or Calcium hydroxide.

5. 10 litres.

6. Tritium—$_1T_3$

7. As a moderator or to slow down fast moving neutrons.

8. A mixture of carbon monoxide and hydrogen (CO+H_2)

9. LiH, BeH_2, or MgH_2

10. CH_4, NH_3, H_2O, HF etc.

11. Water which does not contain cations and anions is called demineralised water.

12. LiH < NaH < CsH.

13. H_2 < D_2 < T_2.

14. Tritium.

15. Temporary hardness is due to the presence of bicarbonates of calcium and magnesium in water.
16. Permanent hardness is due to the presence of chlorides and sulphates of calcium
and magnesium.
17. H_2O molecules associate due to intermolecular hydrogen bonding.
18. Non-stoichiometric hydrides are hydrogen deficient compounds formed by the
reaction of dihydrogen with d and f block elements.
19. Alkali metal oxides present in glass catalyse the decomposition of H_2O_2.
20. Sodium hexa metaphosphate or $Na_6P_6O_{18}$.
21. Sodium Aluminium silicate-($NaAlSiO_4$)
5
22. Ice has low density.
23. 104.5^0
24. Hydrogen gas
25. $Mg(OH)_2$

Model answers for questions carrying two or three marks.

26. Resemblance of hydrogen with alkali metals:
i) Hydrogen has $1S_1$ configuration in it's valence shell and has a tendency to lose electron to form H_+ ion.
ii) Similar to alkali metals hydrogen forms oxides, halides, and sulphides.
Resemblance of hydrogen with halogens:
i) Similar to halogens it can gain one electron and form hydride (H_{--}) ion.
ii) It also forms covalent compounds and exists as a diatomic molecule "H_2"
27. I) Protium- $_1H_1$ ii) Deuterium- $_1D_2$ iii) Tritium- $_1T_3$.
28. Hydrogen molecule has a very high bond dissociation enthalpy . It forms a covalent
bond with another H-atom and exists as a diatomic molecule.
29. By reacting CO with steam in presence of iron chromate catalyst.
$CO(g) + H_2O(g)$ $CO_2(g) + H_2(g)$
This is called water gas-shift reaction.
30. Hydrogen is prepared by the electrolysis of acidified water using platinum.
$2H_2O$ $2H_2(g) + O_2(g)$electrodes.
Pure water is not an electrolyte. Addition of electrolyte makes the ions available for
electrolysis.

31. I) $H_2(g) + M_mO_o(s) \xrightarrow[\Delta]{}$
$mM(s) + 0\ H_2O(l)$.
ii)$CO(g) + H_2(g)$

Δ

→
Catalyst
$CH_3OH(l)$
iii)$C_3H_8(g) + 3\ H_2O(g)$
Δ

→
Catalyst
$3CO(g) + 7\ H_2(g)$
673K
Catalyst
Electrolysis
Traces of acid/base
6
iv)$Zn(s) + 2\ NaOH\ (aq) \longrightarrow$
Heat
$Na_2znO_2(aq) + H_2(g)$.

32.Dihydrogen is chemically inert at room temperature due to it's high H-H bond dissociation enthalpy. At high temperature it accomplishes the reactions in which it is
converted to H_+ or $OH-$ions and formation of covalent bond by sharing electrons.

33.I) An electron deficient hydride has less number of electrons for writing it's conventional lewis structure. Example: diborane-B_2H_6.
ii)An electron precise hydride has the required number of electrons to write the conventional lewis structure. Example: methane-CH_4.
iii)Electron rich hydrides have excess of electrons which are present as lone pairs.Example: ammonia-NH_3.

34. Metallic hydrides of Ni, Pd, Ce, Ac etc. have the property of absorption of dihydrogen. The property of absorption of large volume of hydrogen on transition
metals like Pd and Pt is widely used in catalytic reduction / hydrogenation. This property has high potential for hydrogen storage and as source of energy.

35. Atomic hydrogen produced by the dissociation of dihydrogen with the help of an
electric arc is made is made to combine on the surface to be welded which produces a
very high temperature of 4000k.
Oxy-hydrogen flame generates a high temperature used for cutting and welding purposes.

36. Fluorine is the most electronegative atom and H-F bond is highly polar leading to
strong intermolecular bonding in H-F.
37. $H_2O(l) + H_2O(l) \_ H_3O^+(aq) + OH^-(aq)$
Auto-protolysis or self ionization of water is a chemical reaction in which two water
molecules react to produce a hydroxide ion(OH^-) and hydronium ion.(H_3O^+). This indicates the amphoteric nature or acid base behavior of water.
38. $2F_2(g) + 2H_2O(l) \_ 4H^+(aq) + 4\ F^-(aq) + O_2(g)$.
H_2O is oxidized to O_2. (oxidation number of oxygen increases from –2 to 0)
F_2 is reduced to F^-or HF.(oxidation number of fluorine decreases from 0 to –1)
7
39. I) $PbS(s) + 4H_2O_2(l) \_ 4H_2O(l) + PbSO_4$ (Redox reaction)
ii)$2MnO_4^-(aq) + 5\ H_2O_2(aq) + 6\ H^+(aq) \_ 2\ Mn^{2+}(aq) + 5O_2(g) + 8\ H_2O(l)$(Redox reaction)
iii)$CaO(s) + H_2O(l) \_ Ca(OH)_2$ (Hydration reaction)
iv)$AlCl_3(s) + 3\ H_2O(i) \_ Al(OH)_3(aq) + 3HCl(aq)$. (Hydrolysis)
v)$Ca_3N_2(s) + 6\ H_2O\ (l) \_ 3\ Ca(OH)_2(s) + 2\ NH_3(g)$ (Hydrolysis)

40 i) $H_2O(l) + NH_3(aq) \Leftrightarrow NH_4^+(aq) + OH^-(aq)$
In this reaction water acts as Bronsted acid.

ii)$H_2O(l) + H_2S(aq) \Leftrightarrow H_3O^+(aq) + HS^-(aq)$
In this reaction H_2O acts as a Bronsted base. Hence H_2O is amphoteric.
41.Cation exchange resins contain large organic molecules like $R—SO_3H$ which are
water soluble. Ion exchange resin ($R—SO_3H$) is changed to R—Na by treating with
NaCl. The resin exchanges Na^+ ions with Ca^{2+} and Mg^{2+} ions present in hard water
and make it soft. Here R is anion.
$2R—Na(s) + M^{2+}(aq) \_ R_2M(s) + 2Na^+(aq)$
The resin can be regenerated by adding aqueous NaCl solution.
42.H_2O_2 oxidises PbS to $PbSO_4$
$PbS(s) + 4\ H_2O_2(l) \_ PbSO_4(s) + 4\ H_2O(l)$
Example of a reaction for reducing action of H_2O_2 in acidic medium
$2MnO_4^- + 6\ H^+ + 5\ H_2O_2 \_ 2Mn^{2+} + 8\ H_2O + 5O_2$.
43. Water is a universal solvent due to high value of dielectric constant and dipole
moment.

i) water can dissolve many ionic compounds because of ion—dipole interaction. Ii) water can hydrolyse metallic and non-metallic oxides , hydrides ,carbides etc.
8
44. Saline hydrides are ionic and contain H^-ion which react with water liberating H_2
gas.
For example: $NaH + H_2O \_ NaOH + H_2$.
45. Hydrolysis is reaction in which H^+ and OH^-ions of water react with a compound to
form products.Hydration is association of one or more molecules of water to form
hydrated compounds.
46.Dihydrogen is used i) in the synthesis of ammonia ii) in the manufacture of vanaspathi by hydrogenation of oils. Iii) as a rocket fuel iv) in the manufacture of metal hydrides, methanol etc. v) in metallurgical reduction of metal oxides vi) in fuel
cells for generating electrical energy.
47.$2H_2O_2(l) \_ H_2O(l) + O_2(g)$
2x34=68g 22.7 L at STP
22.7 litres of $O_2$is produced from 68 g of H_2O_2 at STP
10 litres of O_2 is produced from 68x10/22.7=29.9g or 30 g of H_2O_2
Strength of H_2O_2 in 10volume solution= 30 g/litre= 3% H_2O_2 solution.
48. Highly electropositive metals like Na reacts with H_2O and reduces H_2O to dihydrogen.
$2Na (s) + 2H_2O(l) \_ 2NaOH (aq) + H_2(g)$
49. Water is oxidized to O_2 during photosynthesis.
$6CO_2(g) + 12 H_2O(i) -\_ C_6H_{12}O_6(aq) + 6H_2O(l) + 6 O_2(g)$.
50.Washing soda converts soluble chlorides and sulphates of Ca/Mg to insoluble
carbonates and reduces the permanent hardness of water.
$MCl_2 + Na_2CO_3 \_ MCO_3 + 2NaCl$
$MSO_4 + Na_2CO_3 \_ MCO_3 + Na_2SO_4$. (where M= Ca/Mg)

Chapter – 11: P-Block Element

One mark questions

1. What are p-block Elements?

p-block elements are those in which the lost electron enters
the p orbital of the outermost shell.

2. Is an aqueous solution of borax is basic or acidic?

Basic.

3. What is the type of hybridization of boran in diborabe?

Sp_3.

4. Which is the thermodynamically most stable form of carbon?

Graphite.

5. Why boric acid is polymeric?

Due to the presence of hydrogen bonds.

6. What happens when

a)Borax is heated strongly

loses water molecule to form sodium metaborate on further
heating boric anhydride is formed.

b) Boric acid is added to water.

It accepts electrons from OH^- ion.

c) Aluminium is treated with dilute NaOH.

Sodium tetrahydroxoaluminate(III) and hydrogen gas is liberated.

d) Boron trifluoride is reacted with ammonia.

Form an adduct and results in complete octet around boron.

7. Classify the following oxides as neutral,acidic ,basic or
amphoteric.
CO,B_2O_3 ,SiO_2,CO_2.Al_2O_3,PbO_2,Tl_2O_3.
Acidic-B_2O_3,SiO_2,CO_2
Basic-Ti_2O_3
Neutral-CO
Amphoteric-Al_2O_3,PbO_2.
8. What are allotropes? Name the allotropes of carbon?
Elements having different physical properties and same chemical properties are known as Allotropes.
Allotropes of carbon – Graphite, Diamond, Fullerene .
9. Write the oxidation states exhibited by elements of group
14?
+2 and +4 .
10. What is the nature of hybridization in diamond and graphite?
SP_3 and SP_2 respectively.
11. write the valence shell electronic configuration of p-block
elements
ns^2 np^{1-6} (except helium)
12. write the chemical composition of borax.
Na_2 B_4 O_7 . $10H_2O$
13. Which is the catalyst used to convert alcohols directly into
gasoline?
ZSM-5
14. Which is the basic structural unit of silicates?

SiO_4^{4-}

15. Which is the repeating unit of organosillicon polymer?

R_2SiO

16. Name the important oxides of carbon?

CO, CO_2

17. Why carbon shows allotropic forms?

Due to property of catenation and Pπ - Pπ bond formation.

18. What happens atomic and ionic radii of p-block elements
along a period?

Decreases.

19. What happens to ionization enthalphy of p-block elements
of down a group?

Decreases.

20. Name the solid metalloid of p-block elements?

Boron.

21. Which is the most abundant metal in the earth crust?

Aluminium.

22. Which is the simplest boron hydride?

Di borane.

23. Which compound is known as inorganic benzene?

Borazine.

24. Who discovered fullerenes?

H.W.Kroto, E.Smalley and R.F.Curl.

25. What type of hybridization does carbon undergo in diamond?

Sp_3.
26. What type of hybridization does carbon undergo in Graphite?
Sp_2.
27. What is water gas ?
$CO+H_2$.
28. What is producer gas?
$CO + N_2$.
29. What is dry ice?
Solid CO_2.
30. What is the type of hybridization of boron in diborane?
Sp_3.
31. Which gas is used in fizzy drinks and fire extinguisher?
CO_2.
32. Why diamond donot conduct electricity?
Because it has no free electrons.
33. Name the allotropic form of carbon whose structure resembles soccer ball?
Fullerene,
34. Why CH_4 is more stable than SiH_4?
Due to small size of carbon.
35. Mention one use of dry ice?
As a refrigerant for ice cream.
36. Write the equation for the decomposition of steam by tin?
$Sn+2H_2O$
Δ
_ SnO_2+2H_2.

Two marks questions:-

37. Write balanced equations for

a) $B_2H_6+O_2 \rightarrow$

$B_2H_6+3O_2 \rightarrow B_2O_3+3H_2O$

b) BF_3+NaH 450 K.

$2BF_3+6NaH$ 450 K B_2H_6+NaF

c) $SiCl_4+H_2O \rightarrow$

$SiCl_4+2H_2O \rightarrow Si(OH)_4 + 4HCl$

d) $C+O_2+N_2$ 1273K

$2C+O_2+4N_2$ 1273K $2CO+4N_2$

38. Diamond is covalent, yet it has high melting point. why?

Diamond has a three dimensional network involving strong c-c
bonds, which are very difficult to break and in turn has high
melting point.

39. Write the uses of boron.

In making bullet proof vest and light composite material for
aircraft

40. Name the allotropes of carbon.

Diamond, graphite, fullerene.

41. Boron is unable to form BF_6^{3-} ion, explain.

Due to non-availability of d-orbitals boron is unable to expand
its octet therefore,maximum covalence of boron exceed 4

42. Write the uses of the aluminium.

In making utensils,aeroplane and transportation industry etc.

43. Why Carbon monoxide is poisonous? Explain.

Because its ability to form a complex with heamoglobin.The COHb complex is more stable than O_2-Hb complex and prevents Hb to bind with oxygen, thus person dies due to lack of oxygen

44. Distinguish between diamond and graphite.

Diamond –each carbon atom is sp3 hybridised and bonded to 4 other carbon atoms. C-C bond length is 154 pm

Graphite-each carbon atom is sp2 hybridised and is bonded to 3 other carbon atoms. C-C bond length is 141.5 pm

45. Why does boron trifluoride behave as a lewis acid?

Octet remains incomplete hence electron deficient and acts as a lewis acid

46. What is the action of borax on heating?

It first loses water molecules to form sodium metaborate. On further heating forms glass like material known as borax bead

47. How is orthoboric acid is prepared from borax?

It can be prepared by acidifying an aqueous solution of borax

$Na_2B_4O_7+2HCl +5H_2O \rightarrow 2Nacl+4H_3BO_3$

48. Why is boric acid considered as a weak acid?

Because it is not able to relese H^+ ions on its own. It receives OH^- ions from water molecule to complete its octet in turn releases H^+ ions.

49. How to prepare diborane on industrial scale?

Boran trifluride reacts with sodium hydride at 450k forms
diborane.

$2BF_3+6NaH \xrightarrow{450K} B_2H_6+6NaF$

50. Write the structure of diborane

Reffer text book.

51. What is catenation?

Carbon atoms have the tendency to link with one another
through covalent bonds to form chains and rings this property
is called catenation.

52. Why boric acid is considered as a weak acid?

Boric acid is considered as a weak acid because it is not able to
release H+ ions on its own. It receives OH- ions from water
molecules to complete its octet and in turn releases H+ ions.

53. Give one method for industrial preparation of CO.

$C+H2O \rightarrow CO+H2$

54. Give one method for laboratory preparation of co.

$HCOOH \rightarrow H2O+CO$

55. Explain why is there a phenomenal decrease in ionization
enthalpy from carbon to silicon?

As carbon is very small in size its ionization energy is high .
however on moving down the group the size of silicon increases
and its ionization enthalpy decreases sharply.
56. How would explain the lower atomic radius of Ga as compared to Al?
Although Ga has one extra shell than Al , atomic radius is less
than Al this is because the presence of additional 10d-electrons
offer poor screening effect for the outer electrons in Ga ,
hence atomic radius of Ga is less than that of Al
57. Why carbon differs from rest of the members of its family.?
Due to its smaller size, higher electronegetivity, higher ionization enthalpy and unavailability of d orbitals.
58. Diamond is covalent yet it has high melting point. Why.?
Diamond has a three – dimensional network involving strong
C-C bonds, which are very difficult to break and in turn has high
melting point.
59. Why graphite is used as a dry lubricant in machines running
at high temperature.?
Graphite cleaves easily between the layers and therefore, it is

very soft and slippery for this reason it is used as dry lubricant.

60. How to obtain carbon dioxide in laboratory?

By the action of dilute HCl on calcium carbonate.

$CaCO_3+2HCl \rightarrow Ca\ Cl_2+CO_2+H_2O$

61. Write the equation for the chemical change during photosynthesis?

$6CO_2+12H_2O \rightarrow C_6H_{12}O_6+6O_2+6H_2O$

62. Write the resonance structures of carbon dioxide?

Reffer text book.

63. What are the crystalline forms of silica? Are they interconvertable?

Quartz,Cristobalite and trydymite.Yes

64. What are silicons.?

Group of organosilicon polymers, which have (R_2SiO) as a
repeating unit.

65. Name the important man made silicates.

Glass and cement

Three/four marks questions

66. Discuss the pattern of variation in the oxidation states of
(1) B to Tl (2) C to Pb

(1)Due to small size of boron, the sum of its first 3 ionization
enthalpies is very high, this prevents it to form +3 ions and
forces it to form only covalent compounds Al shows +3 oxidation state,on moving down the group, due to poor

shielding effect of intervening d and f orbitals and inert pair
effct,the two electrons present in the s-shell are stongly attracted by the nucleus and not participating in bonding .
hence +1 state becomes more stable down the group. In Ga, In
and Tl, both +1 and +3 oxidation states are observed.
(2) the common oxidation state +4 and +2,down the group +2
oxidation statebecomes common.C and Si shows the +4
state.down the group higher oxidation state less stable because
of inert pair effct.
67. Explain the structure of diborane .
The four terminal hydrogen atoms and the boron atoms lie in
one plane .Above and below this plane ,there are two bridging
hydrogen atoms. The four terminal B-H bonds are regular two
centre –two electron bonds while the two bridge(B-H-B)bonds
are different and can be described in terms of three centre-two
electron bonds as in the figure(refer text book).
68. Explain the anomalus behavior of carbon.
Due to its smaller size ,higher electronegetivity,higher

ionization enthalpy and unavailability of d-orbitals only s and p
orbitals are available for bonding so maximum covalence is 4
and due to catenation it shows allotropic forms.
69. How can you explain higher stability of BCl_3 as compared to
$TlCl_3$.?
Outer electronic configuration- ns^2np^1.
Bcl_3 is more stable than $TlCl_3$ because +3 oxidation state of B is
more stable than +3 oxidation state of Tl,+3 state is highly
oxidizing.
70. Explain the structure of diamond ?
It has a crystalline lattice.
Carbon atom undergoes sp^3 hybridasation.
Linked to other 4 carbon atoms using hybridized orbitals in
tetrahedral fashion.
C-C bond length is 154 pm.
Presence of covalent bonds.
71.Discuss the layered structure of graphite?
Vanderwaals forces between layers,distance between layers
340pm,each layer is composed of planar hexagonal rings of
carbon atoms,C-C bond length 141.5 pm-sp^2 hybridisation
3sigma bonds and a pi bond.

72. Explain the structure of fullerenes?
These are cage like molecules contains 26 membered rings and
12 five mebered rings.All carbon atoms are equal and undergoes sp_2 hybridisation. Each carbon atom forms 3 sigma
bonds with other 3 carbon atoms. The remaining electron at
each carbon is delocalized in molecular orbitals. Both cabon –
carbon single and double with distance 143.5pm and 138.3pm
respectively.
73. How to synthesis organo silicon polymers?
Methyl chloride reacts with silicon in the presence of copper at
573 K methyl substituted chlorosilane are formed,hydralisis of
dimethyldichlorosilane followed by condensation polymerization yields organo silicon polymers.
$2CH_3Cl + Si \xrightarrow{Cu/570\ K} (CH_3)_2SiCl_2 \xrightarrow[-2HCl]{+2H_2O} (CH_3)_2Si(OH)_2$
74. A. Diamond is a bad conductor of electricity but graphite is
good conductor,justify.
Due to sp_3 hybridisation in diamond no free electron. In graphite due sp_2 hybridisation there free to conduct electricity.

B. Diamond is hardest substance while graphite is very soft,
give reasons.
It is difficult to break extended covalent bonds in diamonds so
it is hard. Graphite cleaves easily between the layers therefore
it is very soft
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