# GOVERNMENT HIGHER SECONDARY SCHOOL- KEERIPATTY

## XI- CHEMISTRY PRACTISE QUESTIONS

### Unit-1

1) Define relative atomic mass.

2) What do you understand by the term mole.

3) Define equivalent mass.

4) What do you understand by the term oxidation number.

5) Distinguish between oxidation and reduction.

6) Calculate the molar mass of the following compounds.

i) urea [$CO(NH_2)_2$] ii) acetone [$CH_3COCH_3$] iii) boric acid [$H_3BO_3$] iv) sulphuric acid [$H_2SO_4$]

7) The density of carbon dioxide is equal to 1.965 kgm-3 at 273 K and 1 atm pressure. Calculate the molar mass of $CO_2$.

8) Which contains the greatest number of moles of oxygen atoms

i) 1 mol of ethanol ii) 1 mol of formic acid iii) 1 mol of $H_2O$

9) What is the difference between molecular mass and molar mass? Calculate the molecular mass and molar mass for carbon monoxide.

10) What is the empirical formula of the following ?

i) Fructose ($C_6H_{12}O_6$) found in honey

ii) Caffeine ($C_8H_{10}N_4O_2$) a substance found in tea and coffee.

11) How many moles of ethane is required to produce 44 g of $CO_2$ (g) after combustion.

12) Hydrogen peroxide is an oxidising agent. It oxidises ferrous ion to ferric ion andreduced itself to water. Write a balanced equation.

13) Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38 % hydrogen and rest oxygen its vapour density is 47.

14) A Compound on analysis gave Na = 14.31% S = 9.97% H= 6.22% and O= 69.5% calculate the molecular formula of the compound if all the hydrogen in the compoundis present in combination with oxygen as water of crystallization. (molecular mass of the compound is 322).

15) Balance the following equations by oxidation number method

i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

iii) $Cu + HNO_3 \rightarrow Cu(NO3)_2 + NO_2 + H_2O$

iv) $KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

## Unit-2

16. Which quantum number reveal information about the shape, energy, orientation and size of orbitals?

17. How many orbitals are possible for n =4?

18. How many radial nodes for 2s, 4p, 5d and 4f orbitals exhibit? How many angular nodes?

19. The stabilisation of a half filled d - orbital is more pronounced than that of the p-orbital why?

20. Consider the following electronic arrangements for the $d^5$ configuration

a) b) c)

(i) which of these represents the ground state

(ii) which configuration has the maximum exchange energy.

21. State and explain pauli's exclusion principle.

22. Define orbital? what are the n and l values for 3px and $4dx^2$-$y^2$ electron?

23. Explain briefly the time independent schrodinger wave equation?

24. Determine the values of all the four quantum numbers of the 8th electron in O- atom and 15th electron in Cl atom and the last electron in chromium.

25. For each of the following, give the sub level designation, the allowable m values and the number of orbitals

i) n = 4, l =2, ii) n =5, l = 3 iii) n=7, l=0

26. Give the electronic configuration of $Mn^{2+}$ and $Cr^{3+}$.

27. Describe the Aufbau principle.

28. An atom of an element contains 35 electrons and 45 neutrons. Deduce i) the number of protons ii) the electronic configuration for the element iii) All the four quantum numbers for the last electron.

29. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wave length associated with the electron revolving arround the nucleus.

## Unit-3

30. Define modern periodic law.

31. What are isoelectronic ions? Give examples.

32. What is effective nuclear charge ?

33. Is the definition given below for ionisation enthalpy is correct? "Ionisation enthalpy is defined as the energy required to remove the most loosely bound electron from the valence shell of an atom".

34. Magnesium loses electrons successively to form $Mg^{+}$, $Mg^{2+}$ and $Mg^{3+}$ ions. Which step will have the highest ionisation energy and why?

35. Define electronegativity.

36. How would you explain the fact that the second ionisation potential is always higherthan first ionisation potential?

37. The electronic configuration of atom is one of the important factor which affects the value of ionisation potential and electron gain enthalpy. Explain

38. In what period and group will an element with Z = 118 will be present?

39. Justify that the fifth period of the periodic table should have 18 elements on the basis of quantum numbers.

35. Elements a, b, c and d have the following electronic configurations:

a) $1s^2, 2s^2, 2p^6$ b) $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ d) $1s^2, 2s^2, 2p^1$

Which elements among these will belong to the same group of periodic table

36. Give the general electronic configuration of lanthanides and actinides?

37. Why halogens act as oxidising agents?

38. Mention any two anomalous properties of second period elements.

39. Explain the pauling method for the determination of ionic radius.

40. Explain the periodic trend of ionisation potential.

41. Explain the diagonal relationship.

42. Why the first ionisation enthalpy of sodium is lower that that of magnesium while

its second ionisation enthalpy is higher than that of magnesium?

43. By using paulings method calculate the ionic radii of $K^{+}$ and $Cl^{-}$ ions in the potassium

chloride crystal. Given that $dK^{+}\text{-}Cl^{-} = 3.14\ A^{0}$

44. Explain the following, give appropriate reasons.

(i) Ionisation potential of N is greater than that of O

(ii) First ionisation potential of C-atom is greater than that of B atom, where as the reverse is true is for second ionisation potential.

(iii) The electron affinity values of Be, Mg and noble gases are zero and those of N (0.02 eV) and P (0.80 eV) are very low

(iv) The formation of $F^-$ (g) from F(g) is exothermic while that of $O^{2-}$(g) from O (g) is endothermic.

45. What is screening effect? Briefly give the basis for pauling's scale of electronegativity.

46. State the trends in the variation of electronegativity in group and periods.

**Unit-4**

47. Explain why hydrogen is not placed with the halogen in the periodic table.

48. An the cube at $0^0$ C is placed in some liquid water at $0^0$ C, the ice cube sinks - Why ?

49. Discuss the three types of Covalent hydrides.

50. Predict which of the following hydrides is a gas on a solid (a) HCl (b) NaH. Give your reason.

51. Write the expected formulas for the hydrides of 4th period elements. What is the trend in the formulas? In what way the first two numbers of the series different from the others ?

52. Write chemical equation for the following reactions.

i) reaction of hydrogen with tungsten (VI) oxide $NO_3$ on heating. ii) hydrogen gas and chlorine gas.

53. Complete the following chemical reactions and classify them in to (a) hydrolysis (b) redox (c) hydration reactions.

1) $KMnO_4 + H_2O_2 \rightarrow$ 2) $CrCl_3 + H_2O \rightarrow$ 3) $CaO + H_2O \rightarrow$

54. Hydrogen peroxide can function as an oxidising agent as well as reducing agent. substantiate this statement with suitable examples.

55. Do you think that heavy water can be used for drinking purposes ?

56. What is water-gas shift reaction ?

57. Justify the position of hydrogen in the periodic table ?

58. What are isotopes? Write the names of isotopes of hydrogen.

59. Give the uses of heavy water.

60. Explain the exchange reactions of deuterium.

61. How do you convert parahydrogen into ortho hydrogen ?

62. Mention the uses of deuterium.

63. Explain preparation of hydrogen using electrolysis.

64. A groups metal (A) which is present in common salt reacts with (B) to give compound (C) in which hydrogen is present in –1 oxidation state. (B) on reaction with a gas (C) to give universal solvent (D). The compound (D) on reacts with (A) to give (B), a strong base. Identify A, B, C, D and E. Explain the reactions.

65. An isotope of hydrogen (A) reacts with diatomic molecule of element which occupies group number 16 and period number 2 to give compound (B) is used as a modulator in nuclear reaction. (A) adds on to a compound (C), which has the molecular formula $C_3H_6$ to give (D). Identify A, B, C and D.

66. $NH_3$ has exceptionally high melting point and boiling point as compared to those of the hydrides of the remaining element of group 15 - Explain.

67. Why interstitial hydrides have a lower density than the parent metal.

68. How do you expect the metallic hydrides to be useful for hydrogen storage ?

69. Arrange $NH_3$, $H_2O$ and HF in the order of increasing magnitude of hydrogen bonding and explain the basis for your arrangement.

70. Compare the structures of $H_2O$ and $H_2O_2$.

**Unit-5**

71. Why sodium hydroxide is much more water soluble than chloride?

72. Explain what to meant by efflorescence;

73. Write the chemical equations for the reactions involved in solvay process of preparation of sodium carbonate.

74. An alkali metal (x) forms a hydrated sulphate, $X_2SO_4$. $10H_2O$. Is the metal more likely to be sodium (or) potassium.

75. Write balanced chemical equation for each of the following chemical reactions.

(i) Lithium metal with nitrogen gas (ii) heating solid sodium bicarbonate (iii) Rubidum with oxgen gas

(iv) solid potassium hydroxide with $CO_2$ (v) heating calcium carbonate (vi) heating calcium with oxygen

76. Discuss briefly the similarities between beryllium and aluminium.

77. Give the systematic names for the following

(i) milk of magnesia (ii) lye (iii) lime (iv) Caustic potash (v) washing soda (vi) soda ash (v) trona

78. Substantiate Lithium fluoride has the lowest solubility among group one metal fluorides.

79. Mention the uses of plaster of paris

80. Beryllium halides are Covalent whereas magnesium halides are ionic why?

81. Alkaline earth metal (A), belongs to 3rd period reacts with oxygen and nitrogen to form compound (B) and (C) respectively. It undergo metal displacement reaction with AgNO3 solution to form compound (D).

82. Write balanced chemical equation for the following processes (a) heating calcium in oxygen (b) heating calcium carbonate

(c) evaporating a solution of calcium hydrogen carbonate (d) heating calcium oxide with carbon

83. Explain the important common features of Group 2 elements.

84. Discus the similarities between beryllium and aluminium.

85. Why alkaline earth metals are harder than alkali metals.

86. How is plaster of paris prepared?

87. Give the uses of gypsum.

88. Describe briefly the biological importance of Calcium and magnesium.

89. Which would you expect to have a higher melting point, magnesium oxide or magnesium fluoride? Explain your reasoning.

**Unit-6**

90. State Boyle's law.

91. A balloon filled with air at room temperature and cooled to a much lower temperature can be used as a model for Charle’s law

92. Name two items that can serve as a model for Gay Lusaac’ law and explain.

93. Give the mathematical expression that relates gas volume and moles. Describe in words what the mathematical expression means.

94. What are ideal gases? In what way real gases differ from ideal gases.

95. Can a Van der Waals gas with a=0 be liquefied? explain.

96. Suppose there is a tiny sticky area on the wall of a container of gas. Molecules hitting this area stick there permanently. Is the pressure greater or less than on the ordinary area of walls?

97. Explain the following observations a) Aerated water bottles are kept under water during summer b) Liquid ammonia bottle is cooled before opening the seal c) The tyre of an automobile is inflated to slightly lesser pressure in summer than in winter d) The size of a weather balloon becomes larger and larger as it ascends up into larger altitude

98. Give suitable explanation for the following facts about gases. a) Gases don’t settle at the bottom of a container b) Gases diffuse through all the space available to them and

99. Suggest why there is no hydrogen ($H_2$) in our atmosphere. Why does the moon have no atmosphere?

100. Explain whether a gas approaches ideal behavior or deviates from ideal behaviour if a) it is compressed to a smaller volume at constant temperature. b) the temperature is raised at while keeping the volume constant c) more gas is introduced into the same volume and at the same temperature

101. Which of the following gases would you expect to deviate from ideal behaviour under conditions of low temperature $F_2$, $Cl_2$, or $Br_2$? Explain.

102. Distinguish between diffusion and effusion.

103. Aerosol cans carry clear warning of heating of the can. Why?

104. When the driver of an automobile applies brake, the passengers are pushed toward the front of the car but a helium balloon is pushed toward back of the car. Upon forward acceleration the passengers are pushed toward the front of the car. Why?

105. Would it be easier to drink water with a straw on the top of Mount Everest?

106. Write the Van der Waals equation for a real gas. Explain the correction term for pressure and volume

107. Derive the values of van der Waals equation constants in terms of critical constants.

108. Why do astronauts have to wear protective suits when they are on the surface of moon?

109. When ammonia combines with HCl, $NH_4Cl$ is formed as white dense fumes. Why do more fumes appear near HCl?

**unit-7**

109. Define is Gibb's free energy.

110. Define enthalpy of combustion.

111. Define molar heat capacity. Give its unit.

112. Define the calorific value of food. What is the unit of calorific value?

113. Define enthalpy of neutralization.

114. What is lattice energy?

115. What are state and path functions? Give two examples.

116. Give Kelvin statement of second law of thermodynamics.

117. The equilibrium constant of a reaction is 10, what will be the sign of $\Delta G$? Will this reaction be spontaneous?

118. Enthalpy of neutralization is always a constant when a strong acid is neutralized by a strong base: account for the statement.

119. State the third law of thermodynamics.

120. Write down the Born-Haber cycle for the formation of $CaCl_2$ .

121. Identify the state and path functions out of the following: a)Enthalpy b)Entropy c) Heat d) Temperature e) Work f)Free energy.

122. State the various statements of second law of thermodynamics.

123. What are spontaneous reactions? What are the conditions for the spontaneity of a process?

124. List the characteristics of internal energy.

125. Explain how heat absorbed at constant volume is measured using bomb calorimeter with a neat diagram.

126. Calculate the work involved in expansion and compression process.

127. Derive the relation between ΔH and ΔU for an ideal gas. Explain each term involved in the equation.

128. Suggest and explain an indirect method to calculate lattice enthalpy of sodium chloride crystal.

129. List the characteristics of Gibbs free energy.

**Unit-8**

131. If there is no change in concentration, why is the equilibrium state considered dynamic ?

132. For a given reaction at a particular temperature, the equilibrium constant has constant value. Is the value of Q also constant ? Explain.

133. What the relation between $K_P$ and $K_C$. Give one example for which $K_P$ is equal to $K_C$.

134. For a gaseous homogeneous reaction at equilibrium, number of moles of products are greater than the number of moles of reactants. Is $K_C$ is larger or smaller than $K_P$.

135. When the numerical value of the reaction quotient (Q) is greater than the equilibrium constant (K), in which direction does the reaction proceed to reach equilibrium ?

136. For the reaction, $A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$ ; ΔH is –ve. the following molecular scenes represent different reaction mixture (A – green, B – blue)

137. Deduce the Vant Hoff equation.

**Unit-9**

138. Define (i) molality (ii) Normality

139. What is a vapour pressure of liquid? What is relative lowering of vapour pressure?

140. State and explain Henry's law

141. State Raoult law and obtain expression for lowering of vapour pressure when nonvolatile solute is dissolved in solvent.

142. What is molal depression constant? Does it depend on nature of the solute ?

143. What is osmosis?

144. Define the term 'isotonic solution'.

145. You are provided with a solid 'A' and three solutions of A dissolved in water - one saturated, one unsaturated, and one super saturated. How would you determine which solution is which ?

146. Explain the effect of pressure on the solubility.

**Unit-10**

147. Define the following i) Bond order ii) Hybridisation iii) σ- bond

148. What is a pi bond?

149. In $CH_4$ , $NH_3$ and $H_2O$, the central atom undergoes $sp^3$ hybridisation - yet their bond angles are different. why?

150. Explain $Sp^2$ hybridisation in $BF_3$

151. Draw the M.O diagram for oxygen molecule calculate its bond order and show that $O_2$ is paramagnetic.

152. Draw MO diagram of CO and calculate its bond order.

153. What do you understand by Linear combination of atomic orbitals in MO theory.

154. Discuss the formation of $N_2$ molecule using MO Theory

155. What is dipolment?

156. Linear form of carbondioxide molecule has two polar bonds. yet the molecule has Zero dipolement why?

157. Draw the Lewis structures for the following species. i) $NO^{3-}$ ii) $SO_4^{2-}$ iii) $HNO_3$ iv) $O_3$

158. Explain the bond formation in $BeCl_2$ and $MgCl_2$.

159. Which bond is stronger σ or π? Why?

160. Define bond energy.

161. Hydrogen gas is diatomic where as inert gases are monoatomic – explain on the basis of MO theory.

162. What is Polar Covalent bond? explain with example.

163. Considering x- axis as molecular axis, which out of the following will form a sigma bond.

i) 1s and 2py ii) 2Px and 2Px iii) 2px and 2pz iv) 1s and 2pz

164. Explain resonance with reference to carbonate ion?

165. Explain the bond formation in ethylene and acetylene.

166. What type of hybridisations are possible in the following geometeries? a) octahedral

b) tetrahedral c) square planer.

167. Explain VSEPR theory. Applying this theory to predict the shapes of $IF_7$, and $SF_6$

168. $CO_2$ and $H_2O$ both are triatomic molecule but their dipole moment values are different. Why?

169. Which one of the following has highest bond order? $N_2$, $N^{2+}$ or $N^{2-}$

170. Explain the covalent character in ionic bond.

171. Describe fajan's rule.

**Unit-11**

172. Give the general characteristics of organic compounds?

173. Describe the classification of organic compounds based on their structure.

174. Write a note on homologous series.

175.What is meant by a functional group? Identify the functional group in the following compounds.

(a) acetaldehyde (b) oxalic acid (c) di methyl ether (d) methylamine

176.Give the general formula for the following classes of organic compounds. (a) Aliphatic monohydric alcohol (b) Aliphatic ketones (c) Aliphatic amines.

177.Write the molecular formula of the first six members of homologous series of nitro alkanes.

178. Write the molecular and possible structural formula of the first four members of homologous series of carboxylic acids.

179. Give the structure for the following compound. (i) 3- ethyl - 2 methyl -1-pentene

(ii) 1,3,5- Trimethyl cyclohex - 1 -ene (iii) tertiary butyl iodide

(iv) 3 - Chlorobutanal (v) 3 - Chlorobutanol (vi) 2 - Chloro - 2- methyl propane

(vii) 2,2-dimethyl-1-chloropropane (viii) 3 - methylbut -1- ene (ix) Butan - 2, 2 - diol

(x) Octane - 1,3- diene (xi) 1,5- Dimethylcyclohexane (xii) 2-Chlorobut - 3 - ene

(xiii) 2 - methylbutan - 3 - ol (xiv) acetaldehyde

180. Describe the reactions involved in the detection of nitrogen in an organic compound by Lassaigne method.

181. Give the principle involved in the estimation of halogen in an organic compound by carius method.

182. Give a brief description of the principles of i) Fractional distillation ii) Column Chromatography

183 Explain paper chromatography

184. Explain varions types of constitutional isomerism (structural isomerism) in organic compounds

185. Describe optical isomerism with suitable example.

186. Briefly explain geometrical isomerism in alkene by considering 2- butene as an example.

**Unit-15**

187. Dissolved oxygen in water is responsible for aquatic life. What processes are responsible for the reduction in dissolved oxygen in water?

188. What would happen, if the greenhouse gases were totally missing in the earth's atmosphere?

189. Define smog.

190. Which is considered to be earth's protective umbrella? Why?

191. What are degradable and non-degradable pollutants?

192. From where does ozone come in the photo chemical smog?

193. A person was using water supplied by corporation. Due to shortage of water he started using underground water. He felt laxative effect. What could be the cause?

194. What is green chemistry?

195. Explain how does greenhouse effect cause global warming

196. Mention the standards prescribed by BIS for quality of drinking water

197. How does classical smog differ from photochemical smog?

198. What are particulate pollutants? Explain any three.

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P.ANBARASAN, PGT-CHEMISTRY,GHSS, KEERIPATTY, SALEM DT.