

AL HUDA SCIENCE ACADEMY

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Guess Paper of Chemistry Class 11th

اللہ سب کو سچ کہنے کی ہمت عطا فرمائیں۔ آمین

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11th Class Chemistry Important Short Questions

Question 2

1. Determine the number of protons, neutrons, and electrons in an atom with $Z = 17$ and $A = 35$.
2. Explain why the 4s subshell is filled before the 3d subshell in potassium.
3. What information about an electron can be obtained from the Principal and Azimuthal quantum numbers?
4. Define the Pauli Exclusion Principle and give an example.
5. What does the $(n + 1)$ rule represent in electronic configuration?
6. Why are orbitals of the same subshell called degenerate?
7. Differentiate between an orbit and an orbital.
8. How many electrons can be accommodated in the M-shell? Explain using quantum numbers.
9. Why does the f-subshell have 7 orbitals?
10. Describe the significance of the magnetic quantum number.
11. Explain why the atomic number is more fundamental than the mass number.
12. Compare the mass of a neutron and a proton in amu.
13. What happens to a neutron when passed through an electric field?
14. Why is the deflection of electrons greater than that of protons in an electric field?
15. Draw the orbital box diagram for phosphorus ($Z = 15$) adhering to Hund's rule.
16. Write the electronic configuration of Sodium (Na).
17. Define the azimuthal quantum number (l).
18. How many values of " l " are possible for $n = 3$?
19. What is the maximum number of orbitals in the p-subshell?
20. Distinguish between atomic emission and atomic absorption spectrum.
21. What is 1st ionization energy? Give its standard unit.
22. Explain why sulfur has a lower first ionization energy than phosphorus.
23. The ionization energy of Be is higher than that of B. Justify this exception.
24. Why does ionization energy decrease down a group?
25. Explain why nitrogen has a higher ionization energy than oxygen.
26. Define electronegativity and name the scale used to measure it.

27. How does atomic size affect electronegativity?
28. Why do noble gases have positive 1st electron affinities?
29. Explain why the 2nd electron affinity of oxygen is positive (+844 kJ/mol).
30. Why does fluorine have lower electron affinity than chlorine despite its smaller size?
31. What is the shielding effect and how does it influence atomic radius?
32. Define ionic radius and compare the size of Na and Na^+ .
33. Why are anions always larger than their parent neutral atoms?
34. Explain why Mg^{2+} is smaller than Na^+ .
35. What happens to atomic radius across a period from left to right?
36. Identify semi-metals in groups 14, 15, and 16.
37. Illustrate how metallic character varies in group 14.
38. Why do non-metals have higher electronegativity than metals?
39. Explain the effect of nuclear charge on ionic size.
40. Why is the 3rd ionization energy of magnesium much higher than its 2nd?

Period 3 Elements (Oxides & Chlorides) – Q.NO. 41–60

41. Classify NaCl , MgCl_2 , and PCl_5 as acidic, basic, or neutral.
42. Why are the oxides of Na and Mg more ionic than the oxides of N and P?
43. Describe the reaction of Na_2O with water and identify the product's nature.
44. Why is AlCl_3 considered an acidic halide while NaCl is neutral?
45. Predict whether PCl_5 and NCl_5 would be acidic or basic.
46. Why is SiO_2 considered an acidic oxide despite being a metalloid compound?
47. Give the reaction of ZnO with both HCl and NaOH to prove its amphoteric nature.
48. What causes the difference in oxidation numbers between SO_2 and SO_3 ?
49. Why is sodium kept under kerosene oil while magnesium is not?
50. Explain why magnesium reacts slowly with cold water but rapidly with steam.
51. What is hydration in the context of neutral chlorides like NaCl ?
52. Compare the pH behavior of aqueous solutions of MgCl_2 and AlCl_3 .
53. Why do sulfur and phosphorus show variable oxidation numbers?
54. Describe the nature of oxides formed by Period 3 non-metals.

55. Write balanced equations for the reaction of Mg with oxygen.
56. Explain the term "hydrolysis" in the context of acidic chlorides.
57. Why is the oxidation number of 3rd-period elements in oxides always positive?
58. Compare the reactivity of Na and Mg toward chlorine.
59. What is the important function of the "stair-step line" on the periodic table?
60. How does electronic configuration help determine the block of an element?

Reaction Kinetics – Q.NO. 61–90

61. Define the rate of a reaction and give its units.
62. Differentiate between average rate and instantaneous rate.
63. Why does the instantaneous rate of a reaction decrease with time?
64. Define "Activation Energy" and its role in a chemical reaction.
65. Explain why wood burns more rapidly in pure oxygen than in air.
66. How does temperature affect the collision frequency of molecules?
67. What is the Boltzmann distribution curve?
68. Explain why a 10°C rise in temperature approximately doubles the reaction rate.
69. Define a catalyst and explain how it affects activation energy.
70. Differentiate between homogeneous and heterogeneous catalysis.
71. What is the "Order of Reaction"? How is it experimentally determined?
72. Define the specific rate constant (k).
73. Why is the sum of coefficients in a balanced equation not always equal to the order?
74. Give an example of a zero-order reaction.
75. Define molecularity and distinguish it from the order of reaction.
76. What is the role of proper molecular orientation in collision theory?
77. How does the surface area of a solid reactant affect the rate of reaction?
78. Why is the rate of reaction highest at the beginning?
79. Give an example of a biological catalyst (enzyme).
80. What is a photochemical reaction? How does light affect its rate?
81. Calculate the overall order for: $\text{rate} = k[\text{NO}]^2[\text{NH}_3]^0$.
82. Why do chemists need to know the rate constant for a reaction?

83. How does the nature of reactants (ionic vs. covalent) affect the rate?
84. Describe the effect of a catalyst on the "alternative path" of a reaction.
85. What is meant by "effective collisions"?
86. State the units of the rate constant for a first-order reaction.
87. How can color change be used to monitor the progress of a reaction?
88. Define "Fast Reactions" with an example.
89. Why is pressure important for the rate of gaseous reactions?
90. Explain why a catalyst is not consumed in a reaction.

Chemical Equilibrium – Q.NO. 91–120

91. Define a reversible reaction with a suitable example.
92. What are the macroscopic characteristics of chemical equilibrium?
93. Why does a reversible reaction never go to completion?
94. Explain the term "Dynamic Equilibrium."
95. Define the Equilibrium Constant (K_c).
96. Write the expression for $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$.
97. Why does a catalyst have no effect on the value of K_c ?
98. State Le-Chatelier's Principle.
99. Explain the effect of increasing the concentration of reactants on equilibrium.
100. How does a change in pressure affect equilibrium if the number of moles of gas is equal on both sides?
101. Why is temperature the only factor that changes the value of K_c ?
102. In an exothermic reaction, in which direction does the equilibrium shift if heat is added?
103. Describe the effect of pressure on the synthesis of Ammonia (Haber's Process).
104. Why are pure solids and liquids excluded from the K_c expression?
105. Explain the soda water bottle equilibrium: $CO_2(g) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$.
106. What happens to the equilibrium in a soda bottle when the cap is removed?
107. Differentiate between homogeneous and heterogeneous equilibria.
108. What does a very large value of K_c indicate about the extent of a reaction?
109. How is the equilibrium constant related to the rates of forward and backward reactions?

110. Explain why a closed system is necessary for establishing equilibrium.
111. If $Q_c < K_c$, what does it tell you about the concentrations of reactants and products?
112. Define the "Position of Equilibrium."
113. How does the removal of a product affect the yield in a reversible reaction?
114. Explain the effect of decreasing volume on the reaction: $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$.
115. Why does equilibrium shift to the left in an exothermic reaction when heated?
116. Discuss the industrial importance of Le-Chatelier's Principle.
117. What is the response of an endothermic reaction to a decrease in temperature?
118. In the reaction $\text{BiCl}_3 + \text{H}_2\text{O} \rightleftharpoons \text{BiOCl} + 2\text{HCl}$, what happens when HCl is added?
119. Does a catalyst change the position of equilibrium? Explain.
120. Explain why K_c has no units in some reactions.

Halogens – Q.NO. 121–150

121. Why are halogens called "salt formers"?
122. Explain the trend of volatility in halogens from fluorine to iodine.
123. Why is fluorine the most powerful oxidizing agent?
124. Describe the role of London dispersion forces in the physical states of halogens.
125. Why is the F-F bond weaker than the Cl-Cl bond?
126. Explain why HF is a weaker acid than HCl in aqueous solution.
127. What is a disproportionation reaction? Give the reaction of chlorine with water.
128. Why does the reactivity of halogens with hydrogen decrease down the group?
129. Name the halogen used as an antiseptic and explain how it works.
130. Describe a chemical test to distinguish between Bromide and Iodide ions.
131. Why is HOCl a more effective disinfectant than OCl^- ?
132. Write the ionic equation for the displacement of Iodine by Bromine.
133. Which halogen is the most volatile? Explain based on intermolecular forces.
134. What are the primary active species in the chlorination of water?
135. Why is chlorine used in the purification of drinking water?
136. Explain the color change when chlorine gas is passed through a KI solution.
137. How does the electronegativity of halogens vary down the group?

138. Why can't iodine displace bromine from its salt?
139. Write the oxidation state of Chlorine in HClO and HClO_3 .
140. What is the physical state of Iodine at room temperature? Why?
141. Explain how bond length affects bond strength in halogens.
142. What are hydrogen halides? Name the most stable one.
143. Write the conditions required for the reaction of Iodine with Hydrogen.
144. Why is fluorine gas extremely reactive even in the dark?
145. Compare the oxidizing power of Cl_2 and Br_2 .
146. How is HCl formed in the lab? Write the reaction.
147. What is the environmental concern regarding the chlorination of water?
148. Explain the ionization of halogen acids in water.
149. Why does iodine have a higher melting point than other halogens?
150. Describe the reactivity of halogens with cold and hot NaOH .

Periodic Trends – Q.NO. 151–170

151. Why do some elements show variable oxidation numbers?
152. Define the Modern Periodic Law.
153. What does the period number of an element indicate about its shells?
154. Explain why lithium, sodium, and potassium have similar chemical properties.
155. Why are noble gases placed at the end of each period?
156. How can ionization energy data predict the number of valence electrons?
157. Differentiate between paired and unpaired electrons.
158. What is the significance of arrow direction in orbital diagrams?
159. Why does electron pairing occur only after a subshell is half-filled?
160. Explain the role of surface area in solid-gas reactions.
161. What is an enzyme? Give an example of its catalytic action.
162. How does pressure affect the synthesis of SO_3 in the Contact Process?
163. Define "Macroscopic Events" in a chemical reaction.
164. Why is a large jump seen in the successive ionization energies of magnesium?
165. What is the relationship between group number and valence electrons?

166. Explain the term "Effective Nuclear Charge" (Z_{eff}).
167. Why do atoms become smaller from left to right in a period?
168. Define the "Spin-Pair Repulsion" effect.
169. Why is the second electron affinity of oxygen positive?
170. How does the shielding effect influence the electronegativity of elements in a group?

Reactivity of Elements – Q.NO. 171–180

171. Why are elements on the left side of the periodic table considered metals?
172. Compare the reactivity of Na and Mg with oxygen.
173. Why does sodium form both an oxide and a peroxide?
174. Write the formula of the chloride of a group 15 element.
175. Differentiate between acidic and basic oxides with reactions.
176. Explain why AlCl_3 is acidic while NaCl is neutral.
177. How is the oxidation number of a third-period element generally related to its group?
178. Why is the atomic number more fundamental than the mass number?
179. Find the number of neutrons in $^{31}\text{P}_{15}$.
180. Explain why the Cl^- ion has more electrons than protons.

Atomic Structure Review – Q.NO. 181–190

181. Define the term "Orbit."
182. How do the values of " l " determine the type of subshell (s, p, d, f)?
183. Why are p_x , p_y , and p_z orbitals called degenerate?
184. State Hund's Rule of maximum multiplicity.
185. Explain the importance of the rate law in chemistry.
186. Why is molecularity always a whole number?
187. How does the removal of a product affect the direction of a reversible reaction?
188. What is the response of an equilibrium system to a decrease in temperature in an exothermic reaction?
189. Why is K_c constant at a given temperature?
190. Describe the effect of volume increase on the reaction: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$.

Miscellaneous Short Questions – Q.NO. 191–200

191. Why does a catalyst speed up both forward and reverse reactions equally?
 192. Compare the bond strengths of Cl-Cl and Br-Br.
 193. What is observed when chlorine reacts with KI solution?
 194. Explain why HF is more stable than other hydrogen halides.
 195. How does HClO act on microorganisms to kill them?
 196. Write the electronic configuration of an element in Period 4, Group 2.
 197. Why does metallic character increase down a group?
 198. Explain why nitrogen has a higher ionization energy than oxygen.
 199. What is the significance of the "rate-determining step" in a reaction?
 200. Define "Amphoteric Oxides" with the example of Aluminum Oxide (Al_2O_3).
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Question 3

VSEPR & Hybridization – Q.NO. 1–20

1. By counting electron pairs around the central atom, explain why xenon trioxide (XeO_3) has a pyramidal shape.
2. Explain the difference between the formation of σ and π bonds.
3. Show how the central carbon atom in propanone forms σ and π bonds through hybridization.
4. Predict the shapes of sulfate (SO_4^{2-}), borate (BH_4^-), and tri-iodide (I_3^-) ions according to VSEPR.
5. Sketch the hybrid orbitals and bond formation in PCl_3 , SiCl_4 , and NH_4^+ .
6. Draw the orbital structures of the CO_2 molecule in terms of Valence Bond Theory (VBT).
7. Can you explain why CO has a dipole moment but CO_2 does not?
8. Do you think that individual bonds in CCl_4 are polar? Explain in terms of electronegativity.
9. Draw the Lewis structures for IF_3 and IF_5 and predict their geometry.
10. HI is a strong acid and a robust reducing agent, whereas HF is a weaker acid. Explain.
11. Differentiate between s-s overlap and s-p overlap.
12. Why are covalent bonds directional in nature?
13. Explain the bonding in Cl_2 molecule using orbital overlap.
14. How does the π bond form in O_2 ?
15. Why does H_2S have a bent shape instead of linear?

16. Define sp^3 hybridization and give the example of methane (CH_4).
17. Differentiate between sp and sp^3 hybridization in terms of bond angle and geometry.
18. Why is the energy of hybrid orbitals lower than that of unhybridized orbitals?
19. What type of hybridization is present in $BeCl_2$?
20. Why does BF_3 have a trigonal planar shape?

Thermochemistry – Q.NO. 21–40

21. Define Bond Order.
22. Explain why the enthalpy of hydration is always an exothermic process for gaseous ions.
23. Define standard enthalpy of atomization with an example.
24. The enthalpy of solution can be positive or negative. Explain what this indicates.
25. Calculate ΔH° for the formation of methane using enthalpies of combustion of C, H_2 , and CH_4 .
26. State Hess's Law and explain its advantage.
27. Define Bond Dissociation Energy.
28. Why is bond energy always positive?
29. Define Specific Heat Capacity. What is its value for water?
30. Explain the meaning of each symbol in the formula $q = mc\Delta T$.
31. Define Lattice Energy.
32. Define Hydration Energy.
33. Why is lattice breaking endothermic while hydration is exothermic?
34. What is a Born-Haber cycle?
35. Is electron affinity exothermic or endothermic? Explain.
36. Standard enthalpy of combustion of ethanol?
37. Why are bond energies different in different compounds for the same bond?
38. What happens to energy when a chemical bond is formed?
39. Explain the Joule-Thomson effect.
40. Why can gases be compressed easily?

Intermolecular Forces & States of Matter – Q.NO. 41–70

41. Which forces are present among the molecules of CCl_4 and SiF_4 ?

42. Differentiate between instantaneous dipole-induced dipole (id-id) and permanent dipole (pd-pd) forces.
43. Can the CHF_3 molecule make a hydrogen bond? Explain.
44. Show a hydrogen bond between two molecules of ethanol.
45. Explain why the boiling point difference decreases as the size of the alcohol molecules increases.
46. Why is the viscosity of honey higher than water?
47. Which is more viscous: glycerine or hexane? Why?
48. Why does evaporation get faster at higher temperatures?
49. Why is food cooking difficult in areas with high altitudes?
50. Explain why food cooks faster in a pressure cooker.
51. Why is the boiling point of water (100°C) higher than that of ethanol (78°C)?
52. What is meant by the "habit of a crystal"?
53. Name the properties of liquid crystals in which they resemble solids.
54. Which property of liquid crystals makes them useful in temperature sensing devices?
55. Why do solids not undergo translatory motion?
56. Why does iodine exist as a solid at room temperature?
57. Why does HF have a higher boiling point than HCl?
58. Why do boiling points of halogens increase down the group?
59. Why is HF less acidic than HCl despite stronger bonding?
60. Define Surface Tension and Viscosity.
61. Why can insects like mosquitoes walk on the surface of water?
62. What happens to the surface tension of water when temperature increases?
63. Why does vapour pressure increase with temperature?
64. Explain why vapour pressure is independent of container size.
65. Why are liquid crystals always anisotropic?
66. How are liquid crystals used in breast cancer diagnosis?
67. Define Crystal Lattice and Unit Cell.
68. Why do crystalline solids have sharp melting points?
69. Explain how urea impurity affects the habit of an NaCl crystal.

70. Why is water more effective in temperature regulation than other liquids?

Thermochemistry (Continued) – Q.NO. 71–90

71. Differentiate between exothermic and endothermic reactions with examples.
72. What do you understand by the enthalpy of a system?
73. Distinguish clearly between standard enthalpy of reaction and standard enthalpy of formation.
74. Define and give one example of: Standard enthalpy of solution and hydration.
75. Explain why the lattice enthalpy of an ionic compound is typically a large negative value.
76. Explain why the enthalpy of hydration is always an exothermic process for gaseous ions.
77. Define standard enthalpy of atomization with an example.
78. The enthalpy of solution can be positive or negative. Explain what this indicates.
79. Draw energy profile diagrams for the combustion of methane and decomposition of calcium carbonate.
80. Calculate ΔH° for the formation of methane using enthalpies of combustion of C, H_2 , and CH_4 .
81. State Hess's Law and explain its advantage.
82. Calculate the enthalpy change for the Haber process using bond energies ($N \equiv N$, $H-H$, $N-H$).
83. Define Bond Dissociation Energy.
84. Why is bond energy always positive?
85. Define Specific Heat Capacity. What is its value for water?
86. Explain the meaning of each symbol in the formula $q = mc\Delta T$.
87. Why is insulation used in a calorimeter?
88. Define the calorie content of food in terms of energetics.
89. Give the combustion equation of glucose with its ΔH value.
90. On which fundamental law is Hess's Law based?

Born-Haber Cycle & Bond Energies – Q.NO. 91–100

91. Why is Hess's Law important for reactions that are too slow or have side reactions?
92. Define Lattice Energy.
93. Define Hydration Energy.
94. Why is lattice breaking endothermic while hydration is exothermic?
95. What is a Born-Haber cycle?

96. Name the steps involved in the Born-Haber cycle of NaCl.
97. Is electron affinity exothermic or endothermic? Explain.
98. What is the standard enthalpy of combustion of ethanol?
99. Why are bond energies different in different compounds for the same bond?
100. What happens to energy when a chemical bond is formed?

Organic Chemistry – Q.NO. 1–25

1. Differentiate between Aliphatic and Aromatic hydrocarbons.
2. Differentiate between Homolytic and Heterolytic Fission.
3. Define Electrophile and Nucleophile with examples.
4. Explain why alkanes do not undergo addition reactions.
5. Explain why 2-bromopropane is the major product when propene reacts with HBr.
6. Explain how inductive effects from alkyl groups stabilize carbocations.
7. Write down structural formulas for 2-methylbutane and 2,2-dimethylpropane.
8. Give two differences between molecules of cyclopentane and pentane.
9. Why do branched alkanes have lower boiling points than straight-chain alkanes?
10. Explain why alkanes have a tetrahedral shape.
11. Explain why alkanes are highly stable (paraffins).
12. Explain the mechanism of halogenation of methane.
13. Give the dehydrohalogenation reaction of bromopropane.
14. How does an alkane differ from an alkene in terms of stability?
15. Why is the order of stability of carbocations $3^\circ > 2^\circ > 1^\circ$?
16. Compare the relative rates of addition to alkenes for HCl, HBr, and HI.
17. Explain how Markovnikov's rule is applied in the addition of HBr to 2-pentene.
18. What is a free radical?
19. Why are alkenes called olefins?
20. Describe geometrical isomerism (cis-trans) in alkenes.
21. How can you distinguish between an alkane and an alkene using bromine water?
22. Which bond is weaker in a double bond: sigma or pi?
23. State the bond angle in ethene (120°).

24. Why is the boiling point of CH_3OH (65°C) much higher than CH_3CH_3 (-89°C)?
25. Why does CCl_4 have higher viscosity than CHCl_3 but less than ethanol?

Miscellaneous Short Questions (States of Matter & Hess's Law) – Q.NO. 26–30

26. Why does evaporation continue even at room temperature?
27. Why do we feel cool near the bank of a river?
28. What does the peak in an energy profile diagram represent?
29. What happens to the energy of a system when a bond is broken?
30. What is meant by "indirect route" in Hess's Law?

Periodic Trends Review – Q.NO. 31–50

31. Define Modern Periodic Law.
32. Why does atomic radius decrease across a period?
33. Why does atomic radius increase down a group?
34. Define Ionization Energy and give its trend in the periodic table.
35. Why does shielding effect reduce ionization energy?
36. Explain why the second ionization energy of Calcium is higher than the first.
37. Why do noble gases have the highest ionization energies?
38. Why is the first electron affinity of Oxygen negative while the second is positive?
39. Differentiate between Electron Affinity and Electronegativity.
40. Explain the trend of Electronegativity across a period.
41. Why is the ionic radius of a cation smaller than its parent atom?
42. Why is the ionic radius of an anion larger than its parent atom?
43. Compare the reactivity of Sodium and Magnesium with water.
44. Classify Na_2O , Al_2O_3 , and SO_2 as acidic, basic, or amphoteric.
45. Why does AlCl_3 behave as an acidic chloride in water?
46. What happens when sodium burns in excess oxygen?
47. Explain the variation of metallic character in the periodic table.
48. Define Amphoteric Oxides with a chemical equation.
49. What is the oxidation state of Phosphorus in PCl_3 and PCl_5 ?
50. Why do chlorides of Group 1 and 2 elements form neutral solutions?

Atomic Structure (Detailed) – Q.NO. 51–70

51. Define Atomic Number and Nucleon Number.
52. Describe the behavior of protons, neutrons, and electrons in an electric field.
53. What are isotopes? Give the neutron count for Cl-35 and Cl-37.
54. Explain the Principal Quantum Number (n).
55. What does the Azimuthal Quantum Number (l) determine?
56. Describe the Magnetic Quantum Number (m).
57. State the Pauli Exclusion Principle.
58. State Hund's Rule of Maximum Multiplicity.
59. Explain the Aufbau Principle.
60. Sketch the shape of a p-orbital.
61. Why is the electronic configuration of Chromium (Cr) anomalous?
62. Why is the electronic configuration of Copper (Cu) anomalous?
63. What is a free radical? Give an example.
64. How are n-type semiconductors formed?
65. How are p-type semiconductors formed?
66. Define Exothermic and Endothermic reactions with sign of ΔH .
67. State Hess's Law of Constant Heat Summation.
68. Define Standard Enthalpy of Formation (ΔH_f°).
69. Define Standard Enthalpy of Combustion (ΔH_c°).
70. Define Enthalpy of Atomization (ΔH_{at}°).

Thermochemistry & States of Matter – Q.NO. 71–100

71. Define Standard Enthalpy of Neutralization.
72. Define Lattice Energy.
73. What is the Born-Haber Cycle?
74. Why is bond breaking an endothermic process?
75. Why is heat of neutralization of strong acid and strong base constant?
76. Define Hydrogen Bonding.
77. Why does ice have a lower density than liquid water?

78. What are Liquid Crystals?
79. Define London Dispersion Forces.
80. Why does evaporation cause cooling?
81. State the Law of Mass Action.
82. Define Chemical Equilibrium.
83. Write the Equilibrium Constant (K_c) expression for $aA + bB \rightleftharpoons cC + dD$.
84. Differentiate between Homogeneous and Heterogeneous Equilibrium.
85. What is the relationship between K_p and K_c ?
86. How does temperature affect the equilibrium constant (K_c)?
87. Why is chemical equilibrium dynamic?
88. What is Le Chatelier's Principle?
89. Effect of pressure on the synthesis of Ammonia ($N_2 + 3H_2 \rightleftharpoons 2NH_3$).
90. Does a catalyst affect the equilibrium constant?
91. What are "families" in the periodic table? Name two.
92. Explain the term "Blocks" in the periodic table.
93. What is the role of liquid crystals in medicine?
94. Why are solids incompressible?
95. Define Unit Cell.
96. What is hydration energy?
97. Why is the second electron affinity of Oxygen endothermic?
98. What is the shape of an ethene molecule?
99. Define Isomerism.
100. What is the general formula of Alkanes and Alkenes?

Question 4

Mole Concept & Stoichiometry – Q.NO. 1–20

1. How is the concept of mole derived from Avogadro's number (N_A)?
2. Define: (a) Molar mass (b) Molar volume (c) Molar concentration.
3. 39 g of Potassium (K) and 56 g of Iron (Fe) have equal number of atoms in them. Justify.

4. 4 g of Helium (He), 17 g of Ammonia (NH₃), and 64 g of Sulfur Dioxide (SO₂) occupy 22.414 dm³ at STP despite different sizes. Explain.
5. Differentiate theoretical and actual yields.
6. What are the factors mostly responsible for the low yield of products in chemical reactions?
7. Calculate the molar mass of Potassium Permanganate (KMnO₄).
8. How many molecules are present in 1.75 g of Hydrogen Peroxide (H₂O₂)?
9. How many atoms are present in a 15 g Gold (Au) ring?
10. Determine the volume of 2.5 moles of Chlorine molecules (Cl₂) at STP.
11. Calculate the molar mass of a gas which has a density of 1.97 g/dm³ at STP.
12. Calculate the molar concentration of a solution containing 7.9 g of KMnO₄ dissolved in 1 dm³ (Molar mass = 158).
13. When 3.3 moles of Nitrogen react with Hydrogen, how many moles of Hydrogen are consumed to form Ammonia?
14. Define Stoichiometry. Mention its two basic assumptions.
15. Calculate the mass of Aluminum (Al) needed to react with 32.0 g of Iron(III) Oxide (Fe₂O₃) to produce Iron.
16. What is the mass of 10⁻³ mol of Magnesium Sulfate (MgSO₄)?
17. How many molecules are in 1 × 10⁻⁶ g of isopentyl acetate (C₇H₁₄O₂)?
18. Calculate the volume of CO₂ produced at STP when 4.5 dm³ of Methane (CH₄) is burnt.
19. Why does 1 mole of Uranium weigh more than 1 mole of Hydrogen?
20. Define oxidation and reduction in terms of electron transfer with examples.

Redox Reactions & Electrochemistry – Q.NO. 21–50

21. Define oxidation and reduction in terms of change in oxidation number.
22. Determine the Oxidation Number of: (i) Cr in K₂Cr₂O₇ (ii) N in NO₃⁻.
23. Identify the oxidized and reduced species in: Fe₂O₃ + 3CO → 2Fe + 3CO₂.
24. Define a Disproportionation reaction with an example involving Chlorine.
25. Balance the equation: MnO₂ + HCl → MnCl₂ + Cl₂ + H₂O.
26. Balance the equation: K₂Cr₂O₇ + FeSO₄ + H₂SO₄ → Cr₂(SO₄)₃ + Fe₂(SO₄)₃ + K₂SO₄ + H₂O.
27. Explain the construction and function of a Salt Bridge in a galvanic cell.
28. During electrolysis of aqueous NaCl, why is H₂ liberated at the cathode?

29. Define Standard Hydrogen Electrode (SHE). What is its potential?
30. How and why is an electrical double layer formed?
31. Why is the electrode potential of Cu called reduction potential?
32. Calculate the standard cell potential (E°_{cell}) for a Ni-Co cell.
33. What is the effect of variation in ion concentration on the standard electrode potential?
34. Arrange Ag, Cr, and Fe in increasing order of their reducing powers based on E° values.
35. Describe the redox changes during the electrolysis of molten CuBr_2 .
36. What happens if the salt bridge is removed from a galvanic cell?
37. Why is DC used instead of AC in electrolysis?
38. Define Electrolysis. How is it used to refine metals?
39. Calculate the voltage of a cell having iron and copper electrodes ($E^\circ_{\text{Fe}} = -0.44\text{V}$, $E^\circ_{\text{Cu}} = +0.34\text{V}$).
40. Explain why the electrode potential is measured relatively using SHE.
41. Describe the movement of ions through the salt bridge.
42. (Skipped in numbering)
43. Calculate the mass of 1.5 moles of $\text{Ca}(\text{OH})_2$.
44. Determine the number of molecules in 1.75 g of water.
45. How many atoms of Carbon are in 1 mole of isopentyl acetate?
46. Define Molar Volume. What is its value at STP?
47. State Avogadro's Law in terms of gas volumes.
48. Calculate the mass of sodium hypochlorite produced from 2.25 moles of chlorine.
49. Balance: $\text{MnO}_2 + \text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + \text{H}_2\text{O}$.
50. Calculate the Ox. No. of Chromium in $\text{Cr}_2\text{O}_7^{2-}$.

Nitrogen & Its Compounds – Q.NO. 51–66

51. List two reasons for the chemical inertness of N_2 gas.
52. Why is ammonia (NH_3) considered a weak base?
53. Describe the laboratory preparation of ammonia gas.
54. Why is N_2 gas used in food packaging?
55. Both CO and N_2 have triple bonds. Why is CO more reactive?
56. Why shouldn't a farmer treat a field with ammonium fertilizer and lime at the same time?

57. Draw the structures of N_2O , NO , and NO_2 and explain their bonding.
58. Explain why sulfur is unreactive at room temperature.
59. Determine the oxidation state of S in: H_2SO_4 and $\text{S}_2\text{O}_3^{2-}$.
60. Explain the involvement of d-orbitals in the variable oxidation states of sulfur.
61. What is the role of sulfur in the vulcanization of rubber?
62. Give the reaction of Sulfur with NaOH .
63. What is the function of SO_2 and sulfite salts in food preservation?
64. Explain the oxidation of SO_2 to SO_3 .
65. Why does sulfur show more oxidation states than oxygen?
66. Compare the reactions of sulfur with metals and non-metals.

Environmental Chemistry – Q.NO. 67–100

67. What does PAN stand for? Give its general formula.
68. Describe the role of a catalyst in a catalytic converter.
69. Write the conversion reaction carried out by Nitrobacter.
70. Explain why denitrification occurs in anaerobic conditions.
71. Name the four major layers of the atmosphere and their approximate heights.
72. Why is the Troposphere the most important layer for life?
73. Explain why the Stratosphere is warmer than the Mesosphere.
74. Give the equations for the formation and depletion of ozone in the stratosphere.
75. Identify three major natural sources of air pollutants.
76. Differentiate between Classical and Photochemical smog.
77. Name the major Greenhouse Gases (GHGs). How do they cause global warming?
78. What are the main chemical processes involved in the formation of Acid Rain?
79. How do volatile organic compounds (VOCs) affect air quality?
80. What is the impact of PAN on human health and plants?
81. How does deforestation impact air quality and CO_2 levels?
82. Why is CO called a "silent killer"?
83. Explain the principle of Winkler's method to determine Dissolved Oxygen (DO).
84. Define BOD. How is it carried out?

85. Why are BOD and DO inversely related in polluted water?
86. What are the anthropogenic sources of NO_x?
87. Explain how NO₂ contributes to the brown color of smog.
88. What is the role of sunlight in photochemical smog formation?
89. How do particulates (Particulate Matter) cause lung damage?
90. Write a balanced equation for the formation of sulfuric acid from atmospheric SO₃.
91. Explain the "Greenhouse Effect" in terms of heat absorption.
92. What are the effects of SO₂ on plants and aquatic life?
93. What is the common name of N₂O?
94. Which nitrogen oxide is paramagnetic?
95. Which nitrogen oxide is used in rocket propellants?
96. List two natural sources of NO_x.
97. Define Anthropogenic sources of pollution.
98. What gives the brown color to urban smog?
99. Explain the role of sunlight in smog formation.
100. Name the metals used in a catalytic converter.

Nitrification & Sulfur – Q.NO. 101–116

101. Define Nitrification and name the bacteria involved.
102. Write the conversion reaction of NH₄⁺ to NO₂⁻.
103. What is the atomic number and physical state of sulfur?
104. Define Catenation in sulfur.
105. Which is the most stable molecular form of sulfur?
106. Compare oxidation states of sulfur in SO₂ and H₂SO₄.
107. Explain the stability of the +6 oxidation state in SO₃.
108. What happens when SO₂ is dissolved in water?
109. Why is SCl₂ unstable in moist air?
110. How is sulfur used in the bleaching of paper?
111. Describe the role of sulfur in making dyes and explosives.
112. How is sulfuric acid used in the battery industry?

- 113. What is the physical appearance of sulfur?
- 114. Explain the role of sulfur in the vulcanization of rubber.
- 115. Which atmospheric layer contains the ozone layer?
- 116. Why is the Mesosphere the coldest layer?

Ozone & Air Pollution – Q.NO. 117–144

- 117. Describe how UV radiation contributes to ozone formation.
- 118. Name two primary air pollutants.
- 119. Give an example of a secondary pollutant.
- 120. How are secondary pollutants formed?
- 121. Why are aerosols harmful?
- 122. Why is fossil fuel combustion a major cause of air pollution?
- 123. How does agriculture release ammonia into the air?
- 124. Compare industrial and domestic sources of air pollutants.
- 125. What is the impact of urbanization on air quality?
- 126. What is global warming?
- 127. Which gas has the highest heat-absorbing potential?
- 128. How do particulates cause cardiovascular issues?
- 129. What is the impact of PAN on plant life?
- 130. Explain the role of NO_2 in acid rain.
- 131. How can smog formation be reduced in cities?
- 132. Define the Greenhouse Effect.
- 133. Name one heavy metal pollutant affecting human health.
- 134. Why are motor vehicles considered mobile pollution sources?
- 135. Calculate moles of HCl neutralized by 2.1 g of baking soda (NaHCO_3).
- 136. What volume of H_2 at STP is produced from 7.0 g of Iron and H_2SO_4 ?
- 137. Differentiate between Actual Yield and Percentage Yield.
- 138. Why must electrons lost equal electrons gained in balancing redox reactions?
- 139. Define Oxidation Number.
- 140. Find the Oxidation Number of Mn in KMnO_4 .

141. Why is the mole used by chemists as a unit?
142. How is molar mass related to the periodic table?
143. Do 1 mole of glucose and 1 mole of water contain equal molecules?
144. What happens to a cation during electrolysis?

11th Class Chemistry Important Long Questions

Question NO. 5

1. Explain the four quantum numbers with their names, symbols, possible values, and significance. Also write their formulae where applicable.
2. Draw the shapes of s, p, and d-orbitals. Justify these by keeping in view the azimuthal and magnetic quantum numbers.
3. Describe Aufbau Principle, Pauli Exclusion Principle, and Hund's Rule with diagrams.
4. Discuss the variation of first ionization energies across a period and down a group with reasons.
5. Explain the Bohr and Schrodinger models of the atom. How did the concept of energy levels change with quantum mechanics?
6. Explain the method to determine the number of protons, neutrons, and electrons in atoms and ions, with the help of suitable examples.
7. Hydrogen bonding is present in H_2O , NH_3 , HF , $(\text{CH}_3)_2\text{CO}$ and CHCl_3 molecules. Sketch structures and discuss briefly.
8. Compare the hydrogen bonding in water, ammonia, and hydrogen fluoride. How does this bonding affect their physical states and boiling points?
9. What are London dispersion forces? Give examples, and discuss the factors affecting these forces.
10. Discuss the structural changes when water turns into ice. Justify the empty spaces in its crystals as compared to H_2O at 4°C and lower density of ice.
11. Discuss the factors that affect the boiling point of liquids with examples. Explain how intermolecular forces influence the boiling points of water and ethanol.
12. What are liquid crystals? How are they different from solids and liquids? Explain with their properties and uses.
13. Compare crystalline and amorphous solids.
14. Describe the following properties of crystalline solids: (i) Geometrical shape (ii) Melting point (iii) Cleavage plane (iv) Habit of a crystal.
15. Explain the Ideal Gas Equation ($PV = nRT$).

Question NO. 6

1. What are the postulates of VSEPR model? Discuss the structures of the following species with reference to this theory: (i) CH_4 (ii) NH_3 (iii) H_2O^+ (iv) PCl_3 (v) SO_2 (vi) SF_6 .
 2. Explain the orbital hybridization for CH_4 , NH_3 , BF_3 and BeCl_2 .
 3. Describe the molecular orbital diagram of oxygen (O_2). Explain its bond order and magnetic behavior.
 4. Draw and explain the molecular orbital diagram for nitrogen (N_2). What type of bonding exists?
 5. Differentiate between Sigma and Pi Bonds in detail.
 6. Differentiate between sp , sp^2 , and sp^3 hybridization.
 7. Explain the postulates of collision theory. Also describe the terms "effective collision" and "activation energy" with the help of diagrams or examples.
 8. Explain how a catalyst works. Support your answer with an energy profile diagram.
 9. Discuss in detail the five major factors that affect the rate of a chemical reaction with examples.
 10. Explain the difference between order of reaction and molecularity with examples.
 11. Define rate law, rate constant, and order of reaction. Derive units of k for first and second-order reactions.
 12. Relate the order of a reaction to the rate law for the reaction. How do you distinguish between zero order, first order and second order reaction?
 13. Calculate the reaction rate if the concentration of A is 0.5 M, the concentration of B is 0.2 and the rate constant k is 4.0 M^{-1} . Given the rate law for a reaction: $\text{Rate} = k[\text{A}][\text{B}]^2$.
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Question NO. 7

1. Define and explain the law of mass action and derive the expression for the equilibrium constant (K_c) for a general reaction: $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$.
2. State Le-Chatelier's Principle. Explain its application with respect to changes in concentration, temperature, pressure, and use of a catalyst.
3. Synthesis of ammonia by Haber's process is an exothermic reaction. What should be the possible effect of change of temperature, pressure, and concentration at the equilibrium stage?
4. The change of volume or pressure for the following reactions only changes the equilibrium position. Discuss how the direction changes for: (i) $2\text{H}_2 + \text{O}_2 \rightleftharpoons 2\text{H}_2\text{O}$ (ii) $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$.
5. Ethanol reacts with ethanoic acid to form ethyl ethanoate and water.
6. Propanoic reacts with hydrogen cyanide. At equilibrium, the concentration of the product is $0.0233 \text{ mol dm}^{-3}$. Calculate K_c given initial concentrations were $0.0500 \text{ mol dm}^{-3}$.

7. Define pH and pOH. Derive their mathematical expressions and prove that $\text{pH} + \text{pOH} = 14$ at 25°C .
 8. Define ionization constant (K_a) of acids. Derive the expression for K_a of a weak acid and explain its significance in distinguishing strong vs. weak acids.
 9. What is a buffer solution? Explain the working mechanism of acidic ($\text{CH}_3\text{COOH}/\text{CH}_3\text{COONa}$) and basic buffers with examples.
 10. What is the common ion effect? Explain with examples how the addition of a common ion affects the ionization of weak acids using Le Chatelier's Principle.
 11. The ionic product of water (K_w) is 1.0×10^{-14} at 25°C . If $[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$, calculate $[\text{OH}^-]$, pH, and pOH.
 12. Explain the use of Solubility Product (K_{sp}) in predicting precipitation vs. solubility.
 13. Calculate the solubility of a sparingly soluble salt lead (II) iodide (PbI_2) in water. It has $K_{sp} = 1.4 \times 10^{-8}$.
 14. $\text{Ca}(\text{OH})_2$ is a sparingly soluble compound. Its solubility product is 6.5×10^{-6} . Calculate the solubility of $\text{Ca}(\text{OH})_2$.
 15. A buffer solution has a pH of 5.0. It is made from a weak acid HA with a pK_a of 4.8. What is the ratio of $[\text{A}^-]$ to $[\text{HA}]$?
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Question NO. 8

1. Describe the halogenation of methane using the free radical chain mechanism. Include initiation, propagation, and termination steps.
2. Describe the mechanism of electrophilic addition of hydrogen halides to alkenes. Discuss Markovnikov's Rule in the context of hydrogen halide addition.
3. Describe the important reactions of alkenes including hydrogenation, halogenation, hydration, and ozonolysis with equations.
4. Describe the following methods for the preparation of alkenes: (i) Dehydrohalogenation of alkyl halides (ii) Dehydration of alcohols.
5. Explain the structure and bonding in ethene. How does it affect the reactivity of alkenes?
6. What is geometrical isomerism in alkenes? Illustrate your answer with appropriate structural examples.
7. Explain the characteristics, height ranges, and temperature variations of the four main layers of the atmosphere.
8. Write short notes on the following: (i) CFCs and ozone layer depletion (ii) Greenhouse effect and global warming.

9. How does fossil fuel burning cause acid rain? Discuss in detail with chemical reactions.
 10. Discuss the effects of major air pollutants (CO , SO_2 , NO , O_3 , PAN, particulates) on human health and the environment.
 11. Define primary and secondary pollutants. Explain their sources and harmful effects with examples.
 12. Explain the major natural and anthropogenic (man-made) sources of air pollution with examples.
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Question NO. 9

1. Describe the industrial and laboratory preparation of ammonia. Also give chemical equations and tests for its identification.
2. How oxides of nitrogen (NO_x) cause the formation of photochemical smog and PAN? Give its mechanism.
3. Describe the processes of nitrification and denitrification with the help of relevant chemical equations and bacterial names.
4. Explain the Bronsted-Lowry basic character of ammonia. How is it converted into NH_4^+ ion and what is the geometry of this ion?
5. Discuss the structures, oxidation states, physical properties, and uses of at least three oxides of nitrogen.
6. Describe the different oxidation states shown by sulfur and explain their stability with reference to compounds like H_2S , SO_2 , SO_3 , and S_8 .
7. Discuss the important chemical reactions of sulfur with oxygen, hydrogen, chlorine, water, and metals. Include balanced equations.
8. Describe the uses of sulfur and its compounds in industry, agriculture, and medicine.
9. Explain the trend in oxidizing power of halogens down group 17. Support your answer with reactions and ionic equations.
10. Describe the reactions of halogens with hydrogen. Compare their reactivity trend and write balanced chemical equations.
11. Describe how chlorine purifies drinking water. Include chemical reactions and the role of HClO and ClO^- .
12. Describe the displacement reactions of halogens with halide ions and relate them to oxidizing strength.