CLASS 11 CHEMISTRY EXTRA QUESTIONS

Class 11 Chemistry Chapter 4 Chemical Bonding and Molecular Structure – Extra Questions with Answers

Extra practice questions for Class 11 Chemistry Chapter 4 (Chemical Bonding and Molecular Structure), beyond the textbook. These Class 11 Chemistry Chapter 4 important questions are handy for last-minute exam practice.

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Extra practice questions for Class 11 Chemistry Chapter 4 (Chemical Bonding and Molecular Structure), beyond the textbook. These Class 11 Chemistry Chapter 4 important questions are handy for last-minute exam practice.

Very Short Answer Questions (1 mark)

Q1. What type of bond is formed by electron transfer?
Ans: Ionic (electrovalent) bond.

Q2. What is the shape of a methane (CH_4) molecule?
Ans: Tetrahedral.

Q3. What is the hybridization of carbon in CH_4?
Ans: sp^3.

Q4. What is a coordinate (dative) bond?
Ans: A covalent bond where both shared electrons come from the same atom.

Q5. Name the theory that explains bonding using bonding and antibonding molecular orbitals.
Ans: Molecular orbital theory.

Short Answer Questions (2–3 marks)

Q6. Draw the Lewis structure of water (H_2O) and state its molecular shape.
Ans: Oxygen has 2 bond pairs (to H atoms) and 2 lone pairs; shape is bent/angular (due to lone pair repulsion) with bond angle ≈104.5°.

Q7. Explain why NH_3 has a pyramidal shape rather than a perfect tetrahedral shape, despite nitrogen being sp^3 hybridized.
Ans: Nitrogen in NH_3 has 3 bond pairs and 1 lone pair. The lone pair exerts greater repulsion than bond pairs, pushing the three N-H bonds closer together, distorting the shape from ideal tetrahedral (109.5°) to pyramidal (≈107°).

Q8. Calculate the bond order of N_2 using molecular orbital theory (10 bonding, 4 antibonding electrons).
Ans: Bond order=½(10−4)=3 (triple bond), consistent with N_2's known triple bond.

Higher-Order Thinking / Application Questions

Q9. Using molecular orbital theory, explain why O_2 is paramagnetic (has unpaired electrons) even though its Lewis structure suggests all electrons should be paired, and describe how this illustrates a limitation of the simple Lewis/valence bond approach.
Ans: The Lewis structure of O_2 shows a double bond with all electrons apparently paired, predicting O_2 should be diamagnetic. However, molecular orbital theory reveals that when filling the molecular orbitals of O_2 in order of increasing energy, the last two electrons occupy two degenerate (equal energy) π* antibonding orbitals singly (following Hund's rule), rather than pairing up in one orbital. This gives O_2 two unpaired electrons, correctly predicting its observed paramagnetism (attraction to a magnetic field), something the simple Lewis dot structure approach completely fails to predict, demonstrating that molecular orbital theory provides a more accurate picture of electron distribution in molecules like O_2 where subtle orbital-filling effects matter.

Q10. Explain, using the concept of hydrogen bonding, why water has an unusually high boiling point (100°C) compared to H_2S (boiling point around –60°C), even though sulfur is larger and might be expected to have stronger van der Waals forces.
Ans: Oxygen is significantly more electronegative than sulfur and much smaller in size, which allows water molecules to form strong hydrogen bonds between the highly polarized O-H bonds of one molecule and the lone pairs on the oxygen of a neighbouring molecule. These hydrogen bonds, though individually weaker than covalent bonds, are considerably stronger than ordinary van der Waals (dispersion) forces and create an extensive, dynamic hydrogen-bonded network throughout liquid water, requiring significantly more thermal energy to break in order to vaporize the liquid. H_2S, by contrast, involves the less electronegative and larger sulfur atom, which cannot form effective hydrogen bonds (S-H bonds are much less polarized), so H_2S molecules are held together only by weaker van der Waals forces, resulting in a much lower boiling point despite sulfur's larger size (which would otherwise be expected to increase van der Waals attraction).

Class 11 Chemistry Chapter 4 – Solutions and Notes

For complete step-by-step answers and a quick summary, check the Class 11 Chemistry Chapter 4 Solutions and Class 11 Chemistry Chapter 4 Revision Notes.

See also: Chapter 1 | Chapter 2 | Chapter 3 | Chapter 4

Practice more: Chapter 1 | Chapter 2 | Chapter 3

Quick revision: Chapter 1 | Chapter 2 | Chapter 3