CHEMISTRY (CODE - 043)

CHAPTER TEST - STEP-BY-STEP SOLUTIONS AND MARKING SCHEME

Class XI (2026-27) | Chapter 4: Chemical Bonding and Molecular Structure

Max. Marks: 35

SECTION A (12 × 1 = 12 marks)

Each correct answer carries 1 mark.

Q1. Ans: B. 24 [1]

C = 4, 3 × O = 18, plus 2 for the 2− charge: 4 + 18 + 2 = **24** electrons.

Q2. Ans: A. $BeCl_2$, H_2O, NH_3 [1]

X: two bond pairs, no lone pair → linear, 180° ($BeCl_2$). Y: two bond pairs and two lone pairs → bent, 104.5° (H_2O). Z: three bond pairs and one lone pair → trigonal pyramidal, 107° (NH_3).

Q3. Ans: B. sp^3d [1]

P in PCl_5 uses one 3s, three 3p and one 3d orbital → five sp^3d hybrid orbitals (trigonal bipyramidal).

Q4. Ans: C. BF_3 [1]

BF_3 is trigonal planar; the three B–F bond dipoles (120° apart) cancel, so μ = 0. H_2O, NH_3 and $CHCl_3$ have a net dipole moment.

Q5. Ans: B. 1.5 [1]

O_2^- has 17 electrons: N_b = 10, N_a = 7 (σ^*1s^2, σ^*2s^2, π^*2p^3). Bond order = ½(10 − 7) = **1.5**.

Q6. Ans: A. (first option) [1]

4 bp + 0 lp → tetrahedral; 3 bp + 1 lp → trigonal pyramidal; 2 bp + 2 lp → bent; 4 bp + 1 lp → see-saw.

Q7. Ans: A. (first option) [1]

Q is the minimum of the curve: attraction balances repulsion at the equilibrium distance of 74 pm (the bond length), and the energy released (435.8 kJ mol^{-1}) is the bond enthalpy. At P the atoms are far apart; at R the repulsion dominates.

Q8. Ans: B. O_2 [1]

O_2 has two unpaired electrons in π^*2p_x and π^*2p_y → paramagnetic. N_2, F_2 and H_2 have all electrons paired.

Q9. Ans: A. 3 σ and 2 π [1]

H–C≡C–H: two C–H σ bonds + one C–C σ bond = 3 σ; the triple bond contains two π bonds → **3 σ and 2 π**.

Q10. Ans: A. Both A and R are true, and R is the correct explanation of A. [1]

BeF_2 is linear; the two equal Be–F bond dipoles act in opposite directions and cancel, so the net μ = 0.

Q11. Ans: B. Both A and R are true, and R is not the correct explanation of A. [1]

Both statements are true, but the pyramidal shape does not explain the reduced angle. The angle is reduced to 107° because the lone pair–bond pair repulsion is greater than bond pair–bond pair repulsion.

Q12. Ans: D. A is false but R is true. [1]

He_2: $(\sigma1s)^2(\sigma^*1s)^2$, bond order = ½(2 − 2) = 0. The molecule is unstable and does not exist, so A is false; R is true.

SECTION B (4 × 2 = 8 marks)

Q13. [2]

(i) X_2 is O_2. Configuration: KK $(\sigma 2s)^2$ $(\sigma^* 2s)^2$ $(\sigma 2p_z)^2$ $(\pi 2p_x^{\,2} = \pi 2p_y^{\,2})$ $(\pi^* 2p_x^{\,1} = \pi^* 2p_y^{\,1})$ [½]

$N_b = 8$, $N_a = 4$ (valence) → bond order = ½(8 − 4) = **2** [½]

(ii) Paramagnetic – two unpaired electrons are present, one each in $\pi^* 2p_x$ and $\pi^* 2p_y$. [1]

Q14. [2]

(i) When a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energy, positions of nuclei, bonding and non-bonding pairs are written; these canonical structures taken together (as their resonance hybrid) describe the molecule. [1]

(ii) Each Lewis structure of O_3 has one O–O single bond (148 pm) and one O=O double bond (121 pm), but the two bonds are found to be equal (128 pm), intermediate between single and double. O_3 is a resonance hybrid of the two canonical structures, in which the bond character is averaged. [1]

Q15. [2]

(i) The dipole moment **decreases** from HF (1.78 D) to HI (0.38 D). Electronegativity of the halogen decreases down the group, so the polarity of the H–X bond (charge separation) decreases. [1]

(ii) H_2 is made of two identical atoms; the shared pair is equally attracted by both nuclei (nonpolar bond), so there is no charge separation and μ = 0. [1]

Q16. [2]

(i) P: p–p end-on (head-on, axial) overlap → **σ bond**; Q: p–p sidewise overlap, axes parallel and perpendicular to the internuclear axis → **π bond**; R: s–s head-on overlap → **σ bond**. [1]

(ii) P (σ bond) is stronger: head-on overlap takes place to a larger extent than the sidewise overlap in Q, and bond strength depends on the extent of overlap. [1]

OR

CH_3 carbon: $\mathbf{sp^3}$; =CH– carbon: $\mathbf{sp^2}$; $=CH_2$ carbon: $\mathbf{sp^2}$. [1]

Bonds: 6 C–H + 2 C–C = 8 **σ** bonds (the C=C contributes one σ); one **π** bond. So 8 σ and 1 π. [1]

SECTION C (2 × 3 = 6 marks)

Q17. [3]

(i) Formal charge = [total valence electrons in the free atom] − [total non-bonding (lone pair) electrons] − ½ [total bonding (shared) electrons]. [1]

(ii) O1: 6 − 2 − ½(6) = **+1**; O2: 6 − 4 − ½(4) = **0**; O3: 6 − 6 − ½(2) = **−1**. [1]

(iii) O1 (central, +1) is joined to O2 by a double bond (O2 with two lone pairs, charge 0) and to O3 by a single bond (O3 with three lone pairs, charge −1): $O2{=}O1^{+}{-}O3^{-}$. [½]

Formal charges help to keep track of valence electrons and to select the lowest-energy Lewis structure (the one with the smallest formal charges). They do not indicate real charge separation. [½]

Q18. [3]

(i) $BeCl_2$: 2 bond pairs, no lone pair → **linear**, 180°. BF_3: 3 bond pairs, no lone pair → **trigonal planar**, 120°. [1]

(ii) N has 4 electron pairs (3 bond pairs + 1 lone pair), so the pairs are tetrahedral, but the lone pair repels the bond pairs more than bond pairs repel each other (lp–bp > bp–bp), so the angle shrinks to 107° and the shape is trigonal pyramidal. [1]

(iii) O in H_2O has **two** lone pairs and two bond pairs. The lp–lp repulsion is greater than lp–bp, which is greater than bp–bp, so the angle is reduced more (104.5°) than in NH_3 with only one lone pair. [1]

SECTION D (1 × 4 = 4 marks)

Q19. [4]

(i) Bond enthalpy is the amount of energy required to break one mole of bonds of a particular type between two atoms in the gaseous state. [½] **N_2** (946 kJ mol^{-1}, bond order 3) has the strongest bond. [½]

(ii) As bond order increases, bond length **decreases** and bond enthalpy **increases**. O_2 (b.o. 2): 121 pm, 498 kJ mol^{-1}; N_2 (b.o. 3): 109 pm, 946 kJ mol^{-1}. [1]

(iii) Average O–H bond enthalpy = (502 + 427) ÷ 2 = **464.5 kJ mol^{-1}** [1]

(iv) After the first O–H bond is broken, the second O–H bond is in a changed chemical environment (it is now in the OH radical), so its enthalpy differs. Hence the mean (average) bond enthalpy is used for polyatomic molecules. [1]

SECTION E (1 × 5 = 5 marks)

Q20. [5]

(i) O_2 (16 electrons): KK $(\sigma 2s)^2$ $(\sigma^* 2s)^2$ $(\sigma 2p_z)^2$ $(\pi 2p_x^{\,2} = \pi 2p_y^{\,2})$ $(\pi^* 2p_x^{\,1} = \pi^* 2p_y^{\,1})$ [1]

(ii) Bond order = ½(N_b − N_a). [2 – ½ for the formula, ½ each for the three values]

O_2: N_b = 10, N_a = 6 → **2**

O_2^+ (15 e^-, one π* electron): N_b = 10, N_a = 5 → **2.5**

O_2^- (17 e^-, three π* electrons): N_b = 10, N_a = 7 → **1.5**

(iii) Stability increases with bond order: **$O_2^+ > O_2 > O_2^-$**. Bond length increases as bond order decreases: **$O_2^- > O_2 > O_2^+$**. [1]

(iv) Paramagnetic – O_2^+ has one unpaired electron in a π* orbital. [1]

OR

(i) P ground state: $3s^2\ 3p^3$ (3d empty). Excited state: one 3s electron is promoted to a 3d orbital → $3s^1\ 3p^3\ 3d^1$ (five unpaired electrons). [1]

One s, three p and one d orbital mix to give **five sp^3d hybrid orbitals**, which overlap with the 3p orbitals of five Cl atoms to form five P–Cl σ bonds. [1]

(ii) Trigonal bipyramidal; the three equatorial P–Cl bonds are at 120° to each other, and the two axial bonds make 90° with the equatorial plane. [1]

(iii) The axial bond pairs suffer more repulsion from the equatorial bond pairs (90° interactions), so the axial bonds are slightly longer and weaker. [1]

(iv) S in SF_6 is **sp^3d^2** hybridised; the shape is **regular octahedral**. [1]

***** END OF SOLUTIONS *****