Class 11 Chemistry · Chapter 4: Chemical Bonding and Molecular Structure

Answer key and marking scheme · 20 marks · about 40 minutes · based on NCERT

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1. In which of the following molecules does the central atom have fewer than eight electrons around it? [1]

(a) PCl_3 (b) H_2S
(c) $AlCl_3$ (d) SiF_4

Answer: (c) Al has only three valence electrons; after three shared pairs with Cl it holds six electrons, an incomplete octet. P in PCl_3, S in H_2S and Si in SiF_4 each reach an octet.

2. In which species are there twelve electrons around the central atom? [1]

(a) H_2SO_4 (b) SCl_2
(c) NF_3 (d) CCl_4

Answer: (a) In H_2SO_4 sulphur, a third-period element, has an expanded octet of 12 electrons. S in SCl_2, N in NF_3 and C in CCl_4 have eight.

3. The total number of valence electrons to be placed in the Lewis structure of the sulphate ion, SO_4^{2-}, is [1]

(a) 30 (b) 40
(c) 31 (d) 32

Answer: (d) S gives 6 and each of the four O atoms gives 6: 6 + 24 = 30; the two negative charges add two electrons, so 32.

4. The enthalpies needed to remove the first, second and third hydrogen atoms from $NH_3(g)$ are 449, 385 and 339 kJ mol^{-1}. The average N–H bond enthalpy in NH_3 is [1]

(a) 449 kJ mol^{-1} (b) 391 kJ mol^{-1}
(c) 1173 kJ mol^{-1} (d) 339 kJ mol^{-1}

Answer: (b) The three N–H bonds break with different enthalpies because the chemical environment changes after each step, so the mean is used: (449 + 385 + 339)/3 = 1173/3 = 391 kJ mol^{-1}.

5. The number of σ bonds and π bonds in a molecule of acrylonitrile, $CH_2{=}CH{-}C{\equiv}N$, is [1]

(a) 3 σ and 6 π (b) 6 σ and 2 π
(c) 6 σ and 3 π (d) 7 σ and 3 π

Answer: (c) Every bond between two atoms has one σ bond: three C–H, C=C, C–C and C≡N give 6 σ. The double bond adds one π and the triple bond two π: 3 π.

6. For KBr, the ionization enthalpy of K(g) is 419 kJ mol^{-1}, the electron gain enthalpy of Br(g) is −325 kJ mol^{-1} and the enthalpy of lattice formation is −671 kJ mol^{-1}. Show that making the gaseous ions costs energy, and explain why solid KBr still forms. [2]

$K(g) \rightarrow K^+(g)$ needs +419 and $Br(g) \rightarrow Br^-(g)$ releases 325: net 419 − 325 = +94 kJ mol^{-1}, so forming the gaseous ions costs energy. Lattice formation releases 671 kJ mol^{-1}, which more than compensates for the 94 kJ mol^{-1}, so the solid is stabilised by its lattice enthalpy.

- net +94 kJ mol^{-1} to form the gaseous ions (1)
- lattice enthalpy released (671) more than compensates, so KBr(s) is stable (1)

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7. Taking z as the internuclear axis, state whether each pair of atomic orbitals can combine to form molecular orbitals, with a reason: (i) 1s of one hydrogen atom with 2s of another; (ii) 2pz of one oxygen atom with 2px of the other. [2]

(i) No: atomic orbitals combine only if their energies are the same or nearly the same, and the energy of 2s is appreciably higher than that of 1s. (ii) No: although their energies are equal, 2pz and 2px do not have the same symmetry about the molecular (z) axis.

- (i) no, energies too different (1)
- (ii) no, different symmetry about the z axis (1)

8. Using H = ½(V + M − C + A), find the hybridisation of the central atom and the number of lone pairs on it in (i) ICl_2^- and (ii) SiF_6^{2-}. [2]

(i) ICl_2^-: V = 7, M = 2, A = 1; H = ½(10) = 5, sp^3d; lone pairs = 5 − 2 = 3. (ii) SiF_6^{2-}: V = 4, M = 6, A = 2; H = ½(12) = 6, sp^3d^2; lone pairs = 6 − 6 = 0.

- ICl_2^-: H = 5, sp^3d, 3 lone pairs (1)
- SiF_6^{2-}: H = 6, sp^3d^2, no lone pair (1)

9. Using the molecular orbital energy order for B_2, C_2 and N_2, find the bond orders of C_2 and C_2^{2-} and state the magnetic behaviour of each. Adding two electrons to O_2 to give O_2^{2-} lowers the bond order from 2 to 1; explain why adding two electrons to C_2 has the opposite effect. [3]

C_2 (12 e): KK $\sigma 2s^2\ \sigma^*2s^2\ (\pi 2p_x^2 = \pi 2p_y^2)$; counting σ1s and σ*1s, $N_b = 8$, $N_a = 4$, b.o. = ½(8 − 4) = 2; all electrons paired, diamagnetic. C_2^{2-} (14 e, like N_2): the two extra electrons fill σ2pz, which lies above π2p in this order: $N_b = 10$, $N_a = 4$, b.o. = ½(10 − 4) = 3; all paired, diamagnetic. In C_2^{2-} the added electrons enter a bonding orbital (σ2pz), so N_b rises and the bond order rises; in O_2^{2-} they enter the antibonding π*2p orbitals, so N_a rises and the bond order falls.

- bond orders 2 (C_2) and 3 (C_2^{2-}), worked (1)
- both diamagnetic (1)
- C_2^{2-}: electrons enter bonding σ2pz; O_2^{2-}: antibonding π* (1)

10. Describe the bonding in the linear HCN molecule in terms of hybridisation: name the hybridisation of carbon, the orbitals that overlap to form each σ bond, and how the π bonds arise. [3]

Carbon is sp hybridised; its two sp hybrids point in opposite directions, giving the linear shape. One sp hybrid overlaps axially with the 1s orbital of H (σ C–H); the other overlaps axially with an orbital of N (σ C–N). The two unhybridised 2p orbitals of carbon, at right angles to each other and to the axis, overlap sidewise with two p orbitals of N to form two π bonds, so C≡N is one σ and two π.

- carbon sp, linear (1)
- σ C–H from sp–1s and σ C–N by axial overlap (1)
- two π from sidewise overlap of unhybridised p orbitals (1)

11. State three conditions or features of hybridisation, and correct the statement: 'Only half-filled orbitals can take part in hybridisation.' [3]

Only orbitals of the valence shell are hybridised; the orbitals mixed must have almost equal energy; the number of hybrid orbitals equals the number of atomic orbitals mixed, and the hybrids are equivalent in energy and shape (promotion of an electron is not essential). Correction: filled orbitals of the valence shell can also take part, as the lone-pair orbitals of N in NH_3 do.

- any two correct conditions or features (1)
- a third correct condition or feature (1)
- correction: filled valence orbitals can also hybridise (1)