MCQ 50 marks

1. **B. 21σ, 5π, 8 pairs**
Explanation: Aspirin $C_9H_8O_4$ has 9 + 8 + 4 = 21 atoms in total and 1 ring.
σ bonds = 21 − 1 + 1 = 21
π bonds = 3 (benzene ring) + 2 (C = O) = 5
Lone pairs = 2 × 2 (C = O) + 2 (−OH) + 2 (−O−) = 8

2. **B. BrF_5 and XeF_4**
Explanation: BrF_5: steric number = 5 + 1 = 6, hybridization sp^3d^2, one lone pair gives a square pyramidal shape
XeF_4: steric number = 4 + 2 = 6, hybridization sp^3d^2, two lone pairs give a square planar shape

3. **A. 3 bridging bonds and 6 double bonds**
Explanation: In cyclic $(SO_3)_3$, 3 S atoms and 3 bridging O atoms alternate to form a six-membered ring.
Bridging bonds (S − O − S) = 3
Each S carries 2 terminal S = O, so double bonds = 3 × 2 = 6

4. **B. sp^3d^2 (2 lone pairs) and sp^3d (3 lone pairs)**
Explanation: $[ICl_4]^-$: electrons on I = 7 + 1 = 8, used in 4 bonds = 4, remaining 4 electrons = 2 lone pairs; steric number = 6, hybridization sp^3d^2
$[ICl_2]^-$: electrons on I = 8, used in 2 bonds = 2 , remaining 6 electrons = 3 lone pairs; steric number = 5, hybridization sp^3d

5. **C. 173°**
Explanation: In SF_4, the equatorial lone pair repels the axial S − F bonds and pushes them downward.
So the axial angle contracts from 180° to about 173°.

6. **B. Rapid axial–equatorial ligand exchange**
Explanation: In Berry pseudorotation, the trigonal bipyramidal molecule passes through a square pyramidal intermediate. Two axial and two equatorial ligands exchange positions, and no bond is broken.

7. **D. 3c − 4e bonding model**
Explanation: In the F − Xe − F unit, one 5p orbital of Xe and the 2p orbitals of two F atoms form 3 molecular orbitals.
4 electrons occupy the bonding and non-bonding orbitals among them.
Therefore, the 3c − 4e model needs no d-orbitals.

8. **C. +1**
Explanation: Formal charge = V − N − $\frac{B}{2}$ (where V = valence electrons, N = non-bonding electrons, B = bonding electrons)
Here, N has 4 bonds (B = 8) and no lone pair (N = 0)
Therefore, formal charge = 5 − 0 − $\frac{8}{2}$ = +1

9. **C. $BF_3 < BCl_3 < BBr_3 < BI_3$**
Explanation: The halogen lone pair is donated into the vacant 2p orbital of B through pπ − pπ back-bonding.
From F to I, the size mismatch increases, so back-bonding decreases and Lewis acidity increases.
Therefore, $BF_3 < BCl_3 < BBr_3 < BI_3$

10. **C. Effective 2pπ − 2pπ back-bonding**
Explanation: B and F both belong to the second period, so their 2p orbitals are of nearly the same size.
So 2pπ − 2pπ back-bonding from the lone pair of F into the vacant 2p orbital of B is most effective.
Therefore, BF_3 has the lowest Lewis acidity.

11. **D. One σ and two π bonds**
Explanation: In C_2^{2-} ($[C \equiv C]^{2-}$), the two carbons are joined by a triple bond.
A triple bond consists of 1 σ and 2 π bonds

12. **D. MgO**
Explanation: In MgO, the ions are Mg^{2+} and O^{2-} with small radii.
So the lattice energy $\propto \frac{z_+z_-}{r_0}$ is very high.

Therefore, MgO has the highest melting point (2852°C).

13. **D.** Al_2O_3
Explanation: Values of $|z_+z_-|$: MgO = 2 × 2 = 4, LiF = 1 × 1 = 1, Al_2O_3 = 3 × 2 = 6
Al_2O_3 has the highest charge product and small ionic radii.
Therefore, Al_2O_3 has the most negative lattice enthalpy.

14. **B.** Ag_2CO_3
Explanation: Because of its pseudo-noble-gas ($18e^-$) configuration, Ag^+ has high polarising power and strongly distorts CO_3^{2-}.
Therefore, Ag_2CO_3 decomposes into Ag_2O and CO_2 on slight heating.

15. **D.** sp^2 **and about** 116.8°
Explanation: In O_3, the steric number of the central O is 2 + 1 = 3, so it is sp^2 hybridized.
Repulsion from the one lone pair reduces the bond angle from 120° to about 116.8°.

16. **A.** $[F - H - F]^-$
Explanation: In $[F - H - F]^-$, F is extremely electronegative, so the hydrogen bond is the strongest.
The proton lies exactly midway between the two F atoms and the bond is symmetric.

17. **B. Bond order = 1, diamagnetic**
Explanation: Total electrons in O_2^{2-} = 18; bonding orbitals hold N_b = 10 and antibonding orbitals hold N_a = 8
Bond order = $\frac{10-8}{2}$ = 1
All electrons are paired, so the ion is diamagnetic.

18. **D.** NO^+
Explanation: A higher bond order means a shorter bond length.
The bond order of NO^+ is 3, the highest.
Therefore, NO^+ has the shortest bond length.

19. **C. Bonding** π_{2p_x} **and** π_{2p_y}
Explanation: B_2 has 10 electrons in total; by Hund's rule the last 2 electrons go one each into the degenerate π_{2p_x} and π_{2p_y} orbitals.
Therefore, the 2 unpaired electrons reside in the bonding π_{2p_x} and π_{2p_y} orbitals.

20. **A.** C − C > C = C > C ≡ C
Explanation: As bond order increases, bond length decreases.
Bond orders: C − C = 1, C = C = 2, C ≡ C = 3
Therefore, the length order is C − C > C = C > C ≡ C

21. **C.** C − F > C − Cl > C − Br > C − I
Explanation: From F to I, the halogen size and the C − X bond length increase, so orbital overlap decreases.
Therefore, the bond dissociation enthalpy order is C − F > C − Cl > C − Br > C − I

22. **D. Lone pair repulsion due to the small size of fluorine**
Explanation: Because F is very small, the F − F bond length is short.
So the lone pairs on the two F atoms repel each other strongly and the bond weakens.
Therefore, the dissociation enthalpy of F_2 is lower than that of Cl_2.

23. **D. Banana bond or bent bond**
Explanation: In cyclopropane, the C − C − C angle is about 60°, far below the ideal sp^3 angle (109.5°).
So the orbitals cannot overlap fully along the internuclear axis and a bent (banana) bond forms.

24. **C.** sp^3**, trigonal pyramidal**
Explanation: In $(CH_3)_2SO$, S has 3 bond pairs (2 S − C and 1 S − O) and 1 lone pair.
Steric number = 3 + 1 = 4, so hybridization is sp^3
Because of the one lone pair, the shape is trigonal pyramidal.

25. **D.** CO_3^{2-} **and** NO_3^-
Explanation: Total electrons in CO_3^{2-} = 6 + 24 + 2 = 32
Total electrons in NO_3^- = 7 + 24 + 1 = 32
Both are trigonal planar (sp^2), so they are isoelectronic and isostructural.

26. **A. Network structure of** SiO_2
Explanation: Because C is small, it forms effective pπ − pπ double bonds with O, so CO_2 consists of discrete molecules.
Because Si is large, it cannot form effective

$p\pi - p\pi$ double bonds; it forms a three-dimensional network through 4 Si − O single bonds.
So SiO_2 is a high-melting solid.

27. C. Ligand-to-metal charge transfer
Explanation: In $Cr_2O_7^{2-}$, the oxidation number of Cr is +6 (d^0), so d − d transitions are impossible.
Visible light is absorbed in the $O^{2-} \rightarrow Cr^{6+}$ charge transfer (LMCT), giving the orange colour.

28. C. Diamond
Explanation: In diamond, each C is joined to four C atoms by sp^3 single bonds; there is no π bond or resonance.
So the C − C bond length stays at the pure single-bond length (0.154 nm).

29. A. $AlCl_3 > MgCl_2 > NaCl$
Explanation: From Na^+ to Mg^{2+} to Al^{3+}, the charge increases and the size decreases, so ionic potential and polarising power increase. Therefore, covalent character decreases in the order $AlCl_3 > MgCl_2 > NaCl$

30. B. The in-layer C–C bond in graphite is shorter and stronger than the C–C bond in diamond
Explanation: In diamond each C is sp^3 hybridized and the C–C bond length is about 0.154 nm. In a graphite layer each C is sp^2 hybridized and, due to π-electron delocalization, the bond length is about 0.142 nm; hence this bond is shorter and stronger.

31. A. Orbital interaction between the lone pair on N and the Si centres
Explanation: In the conventional bonding model, the lone pair on N interacts with suitable orbitals on the Si centres. This reduces the availability of the lone pair, makes the geometry around N nearly planar, and makes $(SiH_3)_3N$ less basic than NH_3.

32. A. More p-character is directed towards the more electronegative substituent
Explanation: According to Bent's rule, the central atom uses hybrid orbitals with more p-character towards more electronegative substituents and more s-character towards more electropositive substituents.

33. D. $[SeCl_6]^{2-}$
Explanation: Although the central Se in $[SeCl_6]^{2-}$ has a lone pair, it behaves as stereochemically inactive. The six Cl atoms therefore keep an almost regular octahedral arrangement, with no significant geometric distortion.

34. D. Pentagonal bipyramidal
Explanation: In IF_7 the central I has seven bond pairs and no lone pair. By VSEPR theory, seven electron domains adopt a pentagonal bipyramidal arrangement, so the molecular geometry is pentagonal bipyramidal.

35. C. NF_3
Explanation: In the conventional textbook model, the central atoms of BF_3, H_3BO_3 and CCl_2 have a vacant p orbital, so $p\pi - p\pi$ back-bonding from lone pairs on the neighbouring atoms is possible. In NF_3, however, N already has a lone pair and no suitable vacant orbital, so this type of conventional back-bonding is not expected.

36. B. XeF_2
Explanation: In XeF_2 the central Xe has 2 bond pairs and 3 lone pairs.
Steric number = 2 + 3 = 5, so the electron-pair geometry is trigonal bipyramidal.
The three lone pairs occupy the equatorial positions, so the molecular geometry is linear.

37. B. sp, $\mu \equiv 0$
Explanation: OCS has the structure O = C = S, and the central C has two electron domains, so C is sp hybridized. Although the molecule is linear, the C=O and C=S bond dipoles are unequal, so the opposing dipoles do not cancel completely and $\mu \equiv 0$.

38. C. LiCl
Explanation: Li^+ has the smallest ionic radius, hence the highest charge density and the largest hydration enthalpy magnitude. LiCl therefore has a greater tendency to form hydrates than the other alkali-metal chlorides, e.g. $LiCl \cdot 2H_2O$.

39. **A. Ligand-to-metal charge transfer (LMCT)**
Explanation: In MnO_4^- the oxidation state of Mn is +7, i.e. Mn(VII) is d^0, so a d − d transition is impossible. Charge transfer from the oxygen ligands to vacant Mn-centred orbitals absorbs visible light, giving the intense purple colour.

40. **B. Stabilization of the conjugate base by intramolecular hydrogen bonding**
Explanation: In salicylic acid, the ortho −OH group can form an intramolecular hydrogen bond with the carboxylate part of the conjugate base. This gives the conjugate base extra stabilization and increases the acidity.

41. **C.** 0.00
Explanation: A terminal iodine has 1 bond (2 bonding electrons) and 3 lone pairs (6 electrons)
$FC = 7 - 6 - \frac{2}{2} = 0$
The central iodine has 2 bonds and 3 lone pairs: $FC = 7 - 6 - \frac{4}{2} = -1$
Total charge = $0 + (-1) + 0 = -1$
Hence the formal charge on each terminal iodine is 0

42. **A. H_2O forms a more extensive three-dimensional hydrogen-bond network**
Explanation: One H_2O molecule can take part in up to four hydrogen bonds through its two O–H bonds and two lone pairs, creating an extensive three-dimensional intermolecular network. Although each hydrogen bond in HF is stronger, its network is far less extensive, so water boils at a higher temperature.

43. **D. Strong −I effect of the CCl_3- group**
Explanation: The strong −I effect of three electronegative Cl atoms increases the electrophilicity of the carbonyl carbon. Nucleophilic addition of H_2O therefore becomes easy, and the gem-diol (hydrate) form of chloral is comparatively stable.

44. **C. Two degenerate π bonding orbitals**
Explanation: Simplified valence MO configuration of C_2:
$(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2$
The bonding effects of σ2s and σ^*2s cancel each other
Bond order = $\frac{(2+4)-2}{2} = 2$
Hence the main bonding contribution comes from the two degenerate π2p bonding orbitals

45. **C. It acts as a Lewis acid by accepting OH^- from water**
Explanation: The B centre of $B(OH)_3$ is electron-deficient, so it behaves as a Lewis acid:
$B(OH)_3 + 2H_2O \rightleftharpoons [B(OH)_4]^- + H_3O^+$
The formation of H_3O^+ makes the solution acidic.

46. **A. Tungsten (W)**
Explanation: Owing to its electronic structure and strong cohesive metallic bonding, tungsten has an extremely high melting point of about 3422°C, the highest among the given metals.

47. **A. Formation of solvated electrons**
Explanation: An alkali metal loses an electron in liquid ammonia, forming a solvated metal cation and a solvated electron:
$M + (x+y)NH_3 \rightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-$
The solvated electron absorbs visible light, so the dilute solution is blue.
This electron is unpaired, so the solution is paramagnetic.

48. **D. sp^3, tetrahedral**
Explanation: In $POCl_3$ the central P is bonded to four atoms and has no lone pair, so the steric number is 4. In the conventional hybridization model P is sp^3 hybridized and the geometry is tetrahedral. Since the P=O and P–Cl bonds are not identical, the actual bond angles may deviate slightly from the ideal tetrahedral angle.

49. **D. Each axial bond experiences 90° repulsion from three equatorial bond pairs**
Explanation: In the trigonal bipyramidal structure of PCl_5, each axial bond faces 90° repulsion from three equatorial bond pairs, so repulsion on the axial bonds is greater.

Hence $d_{axial} > d_{equatorial}$, and the axial P–Cl bonds are comparatively longer and weaker.

50. B. AgI

Explanation: Among the silver halides, AgI has the lowest K_{sp}, so it is the least soluble in water. Also, I^- is more polarizable than F^-, Cl^- and Br^-, so AgI has comparatively more covalent character.