



## Answer Key & Explanations: 90 Marks MCQs - Set 1

(From "90 marks MCQs\_NEET\_Practice Set\_1.pdf")

### Part I: 4-Marks Questions

1. A) 90 g

- **Explanation:** Molar mass of glucose ( $C_6H_{12}O_6$ ) =  $(6 \times 12) + (12 \times 1) + (6 \times 16) = 72 + 12 + 96 = 180$  g/mol. Mass = Number of moles  $\times$  Molar mass =  $0.5 \text{ mol} \times 180 \text{ g/mol} = 90$  g.

2. A)  $CH_2Cl$

- **Explanation:** Assume 100 g of the compound. | Element | Percentage | Atomic Mass | Moles (Percentage/Atomic Mass) | Mole Ratio (Divide by smallest) | Simplest Whole Number Ratio || :----- | :----- | :----- | :-----  
----- | :----- | :----- | | C | 24.27 | 12 |  
 $24.27/12 \approx 2.02$  |  $2.02/2.01 \approx 1$  | 1 | | H | 4.07 | 1 |  $4.07/1 = 4.07$  |  $4.07/2.01 \approx 2$  | 2 | |  
Cl | 71.65 | 35.5 |  $71.65/35.5 \approx 2.01$  |  $2.01/2.01 \approx 1$  | 1 | The empirical formula is  $CH_2Cl$ .

3. A)  $CaCO_3$

- **Explanation:** Balanced equation:  $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$  Molar mass of  $CaCO_3 = 40 + 12 + 3(16) = 100$  g/mol. Molar mass of  $HCl = 1 + 35.5 = 36.5$  g/mol. Moles of  $CaCO_3 = 100 \text{ g/mol} / 20 \text{ g} = 0.2$  mol. Moles of  $HCl = 36.5 \text{ g/mol} / 20 \text{ g} \approx 0.548$  mol. From the balanced equation, 1 mole of  $CaCO_3$  reacts with 2 moles of  $HCl$ . To react completely with 0.2 mol of  $CaCO_3$ , the required moles of  $HCl = 0.2 \times 2 = 0.4$  mol. Since we have 0.548 mol of  $HCl$  (which is more than the 0.4 mol required),  $HCl$  is in excess. Therefore,  $CaCO_3$  is the limiting reagent.





4. **B) 0.2 M**

- **Explanation:** Molar mass of  $\text{H}_2\text{SO}_4 = 2(1)+32+4(16) = 2+32+64=98 \text{ g/mol}$ .  
 Number of moles of  $\text{H}_2\text{SO}_4 = \frac{\text{Molar mass}}{\text{Given mass}} = \frac{98 \text{ g/mol}}{4.9 \text{ g}} = 0.05 \text{ mol}$ .  
 Volume of solution = 250 mL = 0.250 L. Molarity (M) =  $\frac{\text{Volume of solution (L)}}{\text{Moles of solute}} = \frac{0.250 \text{ L}}{0.05 \text{ mol}} = 0.2 \text{ M}$ .

5. **A)  $7 \times 6.022 \times 10^{23}$  atoms**

- **Explanation:** Molar mass of  $\text{H}_2\text{SO}_4 = 98 \text{ g/mol}$ . Given mass = 98 g. Number of moles of  $\text{H}_2\text{SO}_4 = \frac{98 \text{ g}}{98 \text{ g/mol}} = 1 \text{ mol}$ . One molecule of  $\text{H}_2\text{SO}_4$  contains  $2(\text{H})+1(\text{S})+4(\text{O})=7$  atoms. Total number of atoms = Moles of molecules  $\times$  Avogadro's Number  $\times$  Number of atoms per molecule =  $1 \text{ mol} \times 6.022 \times 10^{23} \text{ molecules/mol} \times 7 \text{ atoms/molecule} = 7 \times 6.022 \times 10^{23} \text{ atoms}$ .

6. **B) Balmer**

- **Explanation:** In the hydrogen spectrum, electron transitions that end in the  $n=2$  energy level (i.e., from  $n=3, 4, 5, \dots$  to  $n=2$ ) produce lines in the **Balmer series**, which falls within the visible region of the electromagnetic spectrum.

7. **A)  $3.3 \times 10^{-10} \text{ m}$**

- **Explanation:** Use the de Broglie wavelength equation:  $\lambda = h/mv$  Given:  
 $h = 6.626 \times 10^{-34} \text{ J s}$   $m = 9.1 \times 10^{-31} \text{ kg}$  (mass of electron)  $v = 2.2 \times 10^6 \text{ m/s}$   
 $\lambda = \frac{(9.1 \times 10^{-31} \text{ kg}) \times (2.2 \times 10^6 \text{ m/s})}{6.626 \times 10^{-34} \text{ J s}} = \frac{20.02 \times 10^{-25}}{6.626 \times 10^{-34}} = 0.3309 \times 10^{-9} \text{ m} = 3.309 \times 10^{-10} \text{ m}$ .

8. **A) -3.4 eV**

- **Explanation:** The energy of an electron in the  $n$ th Bohr orbit for a hydrogen atom is given by  $E_n = E_1/n^2$ , where  $E_1$  is the ground state energy (-13.6 eV). For the second orbit ( $n=2$ ):  $E_2 = -13.6 \text{ eV}/(2)^2 = -13.6 \text{ eV}/4 = -3.4 \text{ eV}$ .

9. **C) 14**





- **Explanation:** The maximum number of electrons that can be present in a subshell for a given azimuthal quantum number ( $l$ ) is calculated using the formula  $2(2l+1)$ . Given:  $l=3$  (which corresponds to an f-subshell). Maximum electrons  $= 2 \times (2 \times 3 + 1) = 2 \times (6 + 1) = 2 \times 7 = 14$  electrons.

10. A)  $Z^2/n^2$

- **Explanation:** The energy of an electron in the  $n$ th Bohr orbit for a hydrogen-like species (an atom or ion with only one electron) is given by the formula  $E_n = -R_H n^2 Z^2$ , where  $R_H$  is the Rydberg constant,  $Z$  is the atomic number, and  $n$  is the principal quantum number. Thus, the energy is directly proportional to  $Z^2/n^2$ .

11. A)  $H_2O < NH_3 < CH_4$

- **Explanation:** The bond angles are primarily influenced by the number of lone pairs on the central atom and the resulting electron pair repulsions (VSEPR theory).
  - $CH_4$  (Methane): Tetrahedral geometry,  $sp^3$  hybridized, no lone pairs. Ideal bond angle =  $109.5^\circ$ .
  - $NH_3$  (Ammonia): Pyramidal geometry,  $sp^3$  hybridized, one lone pair. Lone pair-bond pair repulsion compresses the bond angle to  $\approx 107^\circ$ .
  - $H_2O$  (Water): Bent (V-shaped) geometry,  $sp^3$  hybridized, two lone pairs. Two lone pairs exert stronger repulsion, compressing the bond angle further to  $\approx 104.5^\circ$ . Therefore, the increasing order of bond angles is  $H_2O < NH_3 < CH_4$ .

12. C) See-saw

- **Explanation:** For  $SF_4$ : Sulfur (S) is the central atom. Valence electrons of S = 6. Four Fluorine (F) atoms form 4 single bonds, using 4 electrons from S. Remaining electrons on S =  $6 - 4 = 2$  electrons, which form 1 lone pair. Steric





number (Number of bond pairs + Number of lone pairs) =  $4+1=5$ . A steric number of 5 corresponds to a trigonal bipyramidal electron geometry. With one lone pair in an equatorial position, the molecular geometry becomes **see-saw**.

13. C)  $sp^3d^2$

- **Explanation:** For  $XeF_4$ : Xenon (Xe) is the central atom. Valence electrons of Xe = 8. Four Fluorine (F) atoms form 4 single bonds, using 4 electrons from Xe. Remaining electrons on Xe =  $8-4=4$  electrons, which form 2 lone pairs. Steric number (Number of bond pairs + Number of lone pairs) =  $4+2=6$ . A steric number of 6 corresponds to  $sp^3d^2$  hybridization. (The electron geometry is octahedral; the molecular geometry is square planar).

14. C)  $CO_2$

- **Explanation:** A molecule has a zero dipole moment if it is symmetrical and the individual bond dipoles cancel each other out.
  - $CO_2$  (Carbon Dioxide): Linear  $O=C=O$  molecule. The two  $C=O$  bond dipoles are equal in magnitude and opposite in direction, canceling each other out. Net dipole moment = 0.
  - $NH_3$  (Ammonia): Pyramidal geometry with a lone pair. Has a net dipole moment.
  - $H_2O$  (Water): Bent geometry with two lone pairs. Has a net dipole moment.
  - $SO_2$  (Sulfur Dioxide): Bent geometry with one lone pair. Has a net dipole moment.

15. C) 1

- **Explanation:** To find the bond order of  $O_2^{2-}$  using Molecular Orbital Theory (MOT): Total electrons in  $O_2^{2-}$  = (8 from O) + (8 from O) + 2 (for 2- charge)





= 18 electrons. MO configuration (for >14 electrons):  $\sigma_{1s}^2 \sigma_{1s}^{*2} \sigma_{2s}^2 \sigma_{2s}^{*2} \sigma_{2p_z}^2 (\pi_{2p_x}^2, \pi_{2p_y}^2) (\pi_{2p_x}^{*1}, \pi_{2p_y}^{*1})$  Number of electrons in bonding orbitals (Nb) = 2 ( $\sigma_{1s}$ ) + 2 ( $\sigma_{2s}$ ) + 2 ( $\sigma_{2p_z}$ ) + 4 ( $\pi_{2p}$ ) = 10. Number of electrons in antibonding orbitals (Na) = 2 ( $\sigma_{1s}^*$ ) + 2 ( $\sigma_{2s}^*$ ) + 2 ( $\pi_{2p}^*$ ) = 6. Bond Order =  $\frac{1}{2}(Nb - Na) = \frac{1}{2}(10 - 6) = 2$ . Wait, I've re-calculated the bond order for  $O_2^{2-}$  and it's 2. The given answer is C) 1. Let's re-check the MO configuration for 18 electrons.  $\sigma_{1s}^2 \sigma_{1s}^{*2} \sigma_{2s}^2 \sigma_{2s}^{*2} \sigma_{2p_z}^2 \pi_{2p_x}^2 \pi_{2p_y}^2 \pi_{2p_x}^{*2} \pi_{2p_y}^{*2}$  Nb = 2 + 2 + 2 + 4 = 10. Na = 2 + 2 + 4 = 8. Bond Order =  $\frac{1}{2}(10 - 8) = 1$ . My previous calculation for  $O_2^{2-}$  was correct after all, the one I had in the latest immersive was incorrect. The bond order is indeed 1.

## Part II: 2-Marks Questions

### 16. C) Density

- **Explanation:** Intensive properties are independent of the amount of substance (e.g., temperature, density, boiling point). Extensive properties depend on the amount of substance (e.g., mass, volume, energy).

### 17. B) 4

- **Explanation:** In 0.002040, the leading zeros (0.00) are not significant. The non-zero digits (2, 4) are significant. The zero between non-zero digits (0) is significant. The trailing zero (the final 0) is significant because it is after the decimal point. Thus, there are 4 significant figures.

### 18. C) Molecular Formula = Empirical Formula $\times$ n

- **Explanation:** The molecular formula represents the actual number of atoms in a molecule, while the empirical formula gives the simplest whole-number ratio.





The molecular formula is always an integral multiple (n) of the empirical formula.

**19. C) Law of Conservation of Mass**

- **Explanation:** Balancing a chemical equation ensures that the total number of atoms of each element on the reactant side is equal to the total number of atoms of each element on the product side. This is consistent with the Law of Conservation of Mass, which states that matter cannot be created or destroyed.

**20. C) 3**

- **Explanation:** Orthophosphoric acid ( $\text{H}_3\text{PO}_4$ ) has three hydrogen atoms directly bonded to oxygen atoms, which are acidic and replaceable. Hence, its basicity is 3.

**21. B) 4s**

- **Explanation:** According to the Aufbau principle, electrons fill orbitals in order of increasing energy, which can be determined by the  $(n+l)$  rule. For 3p:  $n=3$ ,  $l=1$ ;  $n+l = 4$  For 4s:  $n=4$ ,  $l=0$ ;  $n+l = 4$  When two orbitals have the same  $(n+l)$  value, the one with the lower 'n' value is filled first. Both 3p and 4s have  $(n+l)=4$ , but 4s has a lower 'n' value (4) than 3d ( $n=3$ ,  $l=2$ ,  $n+l=5$ ). Therefore, after 3p, 4s is filled before 3d.

**22. B) Position and momentum**

- **Explanation:** Heisenberg's Uncertainty Principle states that it is fundamentally impossible to simultaneously know both the exact position and exact momentum (or velocity) of a particle. The more accurately one is known, the less accurately the other can be determined.

**23. C) Photoelectric effect**





- **Explanation:** The photoelectric effect, where electrons are ejected from a metal surface when light of a certain frequency shines on it, can only be explained by considering light as behaving like particles (photons), not just waves.

24. C) 2

- **Explanation:** The number of radial nodes (or spherical nodes) in an orbital is given by the formula  $(n-l-1)$ . For a 3s orbital:  $n=3$ ,  $l=0$ . Number of radial nodes  $= 3-0-1=2$ .

25. C) Same number of neutrons but different atomic numbers.

- **Explanation:** Isotones are atoms of different elements (different atomic numbers and thus different numbers of protons) that have the same number of neutrons.

26. B) Ionic bond

- **Explanation:** An ionic bond is formed by the complete transfer of one or more electrons from one atom (typically a metal) to another atom (typically a non-metal), resulting in oppositely charged ions that are held together by electrostatic attraction.

27. C) Valence Shell Electron Pair Repulsion (VSEPR) Theory

- **Explanation:** VSEPR theory predicts the shapes of molecules based on the idea that electron pairs (both bonding and lone pairs) in the valence shell of a central atom repel each other and arrange themselves to minimize this repulsion, leading to specific geometries.

28. D)  $N_2^{2-}$

- **Explanation:** Bond length is inversely proportional to bond order. The species with the lowest bond order will have the longest bond length. Let's determine the bond order for each:







- $\text{N}_2$ : Total electrons = 14. Bond Order = 3 (triple bond).
- $\text{N}_2^+$ : Total electrons = 13. Bond Order = 2.5.
- $\text{N}_2^-$ : Total electrons = 15. Bond Order = 2.5.
- $\text{N}_2^{2-}$ : Total electrons = 16. Bond Order = 2. Comparing bond orders:  $\text{N}_2 (3) > \text{N}_2^+ (2.5) = \text{N}_2^- (2.5) > \text{N}_2^{2-} (2)$ .  $\text{N}_2^{2-}$  has the lowest bond order (2), hence it will have the maximum bond length.

### 29. C) HF

- **Explanation:** Hydrogen bonding occurs when hydrogen is directly bonded to a highly electronegative atom: Fluorine (F), Oxygen (O), or Nitrogen (N). Among the given options, only HF has hydrogen directly bonded to fluorine, allowing for hydrogen bonding.

### 30. A) The delocalization of electrons.

- **Explanation:** Resonance is a concept used to describe the bonding in molecules or ions where a single Lewis structure is insufficient. It involves the delocalization of electrons over multiple atoms, leading to a more stable and accurate representation of the bonding.

