AIIMS B.Sc Nursing Chemistry

Sample Paper – 10

Duration: 36 Minutes **Maximum Marks: 30**

Instructions

- This paper contains **30 Multiple Choice Questions (single correct answer)**, modelled on the Chemistry section of the **AIIMS B.Sc Nursing** entrance.
- Each correct answer carries **+1 mark**. $\frac{1}{3}$ **mark is deducted** for every wrong answer, and an unattempted question gets **0 marks**.
- Only **one** option is correct. The paper covers physical, inorganic, and organic chemistry.
- Personal calculators, log tables, mobile phones, and other electronic gadgets are strictly prohibited.

Q1. The number of moles present in 4 g of sodium hydroxide (NaOH, molar mass $= 40$ g mol^{-1}) is:

(A) 1 mol

(B) 0.1 mol

(C) 0.4 mol

(D) 10 mol

Q2. The figure shows that 1 mole of carbon dioxide (CO_2) has a mass of 44 g. The mass of 0.5 mole of CO_2 is:

$$\underset{44\text{ g}}{\boxed{1\text{ mol}}} \Rightarrow \underset{?\text{ g}}{\boxed{0.5\text{ mol}}}$$

(A) 88 g

(B) 44 g

(C) 22 g

(D) 11 g

Q3. Which scientist's atomic model first proposed that electrons revolve around the nucleus only in fixed, definite orbits (shells) of fixed energy?

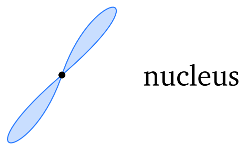
(A) Niels Bohr

(B) J. J. Thomson

(C) Ernest Rutherford

(D) John Dalton

Q4. The boundary surface of one type of atomic orbital is shown below, with two lobes on opposite sides of the nucleus. This dumbbell shape is that of a:



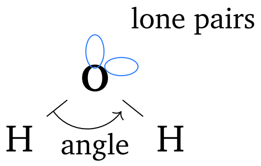
(A) s orbital

(B) f orbital

(C) d orbital

(D) p orbital

Q5. The bent shape of a water molecule (H_2O), with two lone pairs on oxygen, is shown. The H–O–H bond angle is approximately:



(A) 90°

(B) 120°

(C) 104.5°

(D) 180°

Q6. In an ethane molecule (C_2H_6), each carbon atom is bonded to three hydrogens and to the other carbon by single bonds. The hybridization of each carbon atom is:

(A) sp^3

(B) sp^2

(C) sp

(D) sp^3d

Q7. Among the following bonds, which one shows the *least* ionic character (i.e. is the most covalent)?

(A) Na–Cl

(B) H–Cl

(C) K–F

(D) C–I

Q8. On moving down a group of the periodic table, the first ionization energy of the elements generally:

(A) increases

(B) decreases

(C) remains unchanged

(D) first decreases, then increases

Q9. The figure highlights the last vertical column of the periodic table. The noble gases are the elements of:

. . . groups across a period . . .

(A) group 1

(B) group 17

(C) group 18

(D) group 16

Q10. The standard enthalpy of combustion ($\Delta_c H^\circ$) of a fuel, such as methane burning in oxygen, is always:

(A) negative

(B) positive

(C) zero

(D) sometimes positive, sometimes negative

Q11. The thermodynamic relation $\Delta G^\circ = -RT \ln K$ connects the standard free energy change of a reaction with its:

(A) activation energy

(B) enthalpy of formation

(C) molar heat capacity

(D) equilibrium constant

Q12. For the gaseous equilibrium $\mathrm{PCl_5}(g) \rightleftharpoons \mathrm{PCl_3}(g) + \mathrm{Cl_2}(g)$, decreasing the pressure on the system shifts the equilibrium toward the side having:

(A) fewer moles of gas

(B) more moles of gas

(C) equal moles on both sides

(D) no shift at all

Q13. A typical acidic buffer solution, such as one made from acetic acid and sodium acetate, consists of a weak acid together with its:

(A) strong acid

(B) strong base

(C) salt (conjugate base)

(D) another weak acid

Q14. According to the Bronsted–Lowry theory, the conjugate base of sulfuric acid (H_2SO_4) is the species formed when it loses one proton, namely:

(A) HSO_4^-

(B) SO_4^{2-}

(C) $H_3SO_4^+$

(D) H_2SO_3

Q15. The oxidation number of iron in iron(III) oxide, Fe_2O_3 (each oxygen $=$ -2), is:

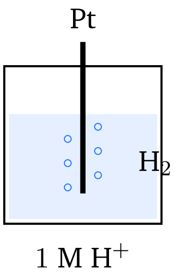
(A) $+2$

(B) $+6$

(C) 0

(D) $+3$

Q16. The standard hydrogen electrode (SHE), shown below with H_2 gas at 1 bar bubbling over a platinum electrode in 1 M H^+, is the reference electrode. Its standard electrode potential is assigned a value of:



(A) $+1.00$ V

(B) 0.00 V

(C) -1.00 V

(D) $+0.76$ V

Q17. The number of faradays of charge required to deposit 1 mole of a trivalent metal ($M^{3+} + 3e^- \rightarrow M$) during electrolysis is:

(A) 1 faraday

(B) 2 faradays

(C) 3 faradays

(D) 6 faradays

Q18. The number of moles of solute present in 250 mL of a 0.2 M solution is:

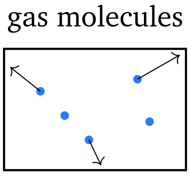
(A) 0.05 mol

(B) 0.2 mol

(C) 0.5 mol

(D) 0.025 mol

Q19. According to the kinetic theory of gases, the pressure exerted by a gas on the walls of its container arises from:



(A) the attraction between molecules

(B) the volume of the molecules themselves

(C) chemical reaction with the walls

(D) collisions of the molecules with the walls

Q20. The oxidation state of manganese in the complex ion $[Mn(H_2O)_6]^{2+}$ (water is a neutral ligand) is:

(A) $+1$

(B) $+2$

(C) $+4$

(D) $+6$

Q21. The total number of ions produced in aqueous solution by one formula unit of potassium ferrocyanide, $K_4[Fe(CN)_6]$, is:

(A) 3

(B) 4

(C) 5

(D) 11

Q22. The chalcogens (the oxygen family, which includes O, S, Se and Te) make up which group of the periodic table?

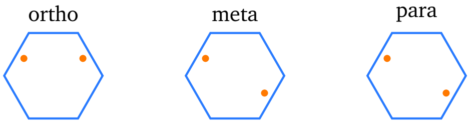
(A) group 16

(B) group 15

(C) group 17

(D) group 14

Q23. Ortho-, meta- and para-xylene (dimethylbenzene) differ only in the positions of the two CH_3 groups on the benzene ring, as shown. They are examples of:



(A) chain isomerism

(B) functional isomerism

(C) geometrical isomerism

(D) position isomerism

Q24. The IUPAC name of formaldehyde, HCHO, is:

(A) methanol

(B) methanal

(C) ethanal

(D) methanoic acid

Q25. Two consecutive members of a homologous series always differ in their molecular formula by:

(A) CH_4

(B) C_2H_4

(C) CH_2

(D) CH_3

Q26. When methane (CH_4) reacts with chlorine in the presence of sunlight to give chloromethane and HCl, the type of reaction taking place is:

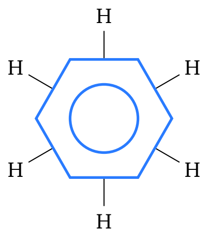
(A) substitution

(B) addition

(C) elimination

(D) polymerisation

Q27. The Kekule structure of benzene (C_6H_6) is shown, with six C–H bonds and a ring of six C–C bonds. The total number of sigma (σ) bonds in a benzene molecule is:



(A) 6

(B) 9

(C) 15

(D) 12

Q28. In an aldehyde such as ethanal (CH_3CHO), the carbonyl group ($>C=O$) is located at the:

(A) middle of the carbon chain

(B) end (terminal carbon) of the chain

(C) nitrogen atom

(D) oxygen atom only

Q29. Starch and cellulose are natural polymers (polysaccharides) made up of many repeating units of:

(A) amino acids

(B) fatty acids

(C) glucose

(D) nucleotides

Q30. The reaction in which a carboxylic acid combines with an alcohol (in the presence of an acid catalyst) to form an ester and water is called:

(A) esterification

(B) saponification

(C) neutralisation

(D) fermentation

Detailed Solutions

Q1.

Solution

Concept — The mole and molar mass: The mole is the SI unit for the amount of a substance. The number of moles connects a measured mass to a count of particles through $n = \dfrac{\text{given mass}}{\text{molar mass}}$, where the molar mass is the mass of one mole expressed in grams. For NaOH the molar mass is the sum of the atomic masses: Na $(23)+$ O $(16)+$ H $(1) = 40$ g mol^{-1}.

Given: mass $= 4$ g; molar mass $M(\text{NaOH}) = 40$ g mol^{-1}.

Step 1 — Write the formula: $n = \dfrac{m}{M}$.

Step 2 — Substitute the data with units: $n = \dfrac{4 \text{ g}}{40 \text{ g mol}^{-1}}$.

Step 3 — Evaluate: the grams cancel, leaving $n = 0.1$ mol.

Why each other option is wrong:

- (A) 1 mol would correspond to a full 40 g of NaOH, not 4 g.
- (C) 0.4 mol comes from dividing 4 by 10 instead of by 40.
- (D) 10 mol inverts the ratio $(40 \div 4)$ instead of $4 \div 40$.

Key point: Always work out the molar mass first by summing atomic masses, then divide the given mass by it. Here 4 g is one-tenth of 40 g, so the answer is 0.1 mol.

Final Answer: 0.1 mol $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q1

Q2.

Solution

Concept — Mass from moles: The mass of a sample is obtained from the number of moles by rearranging $n = \dfrac{m}{M}$ into $m = n \times M$, where n is the number of moles and M is the molar mass in g mol^{-1}. The molar mass of CO_2 is C $(12) + 2\times$O $(2 \times 16) = 44$ g mol^{-1}.

Given: $n = 0.5$ mol; $M(CO_2) = 44$ g mol^{-1}.

Step 1 — Write the formula: $m = n \times M$.

Step 2 — Substitute with units: $m = 0.5 \text{ mol} \times 44 \text{ g mol}^{-1}$.

Step 3 — Evaluate: $m = 22$ g.

Why each other option is wrong:

- (A) 88 g would be the mass of 2 moles (2×44).
- (B) 44 g is the mass of a full 1 mole, not half a mole.
- (D) 11 g corresponds to 0.25 mole, half of the correct amount.

Key point: Half a mole has half the molar mass: $0.5 \times 44 = 22$ g. Multiply moles by molar mass to get mass, and divide mass by molar mass to get moles.

Final Answer: 22 g $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q2

Q3.

Solution

Concept — Bohr's model of the atom: Niels Bohr (1913) was the first to propose that electrons revolve around the nucleus only in certain fixed, allowed circular paths called orbits or shells, each of a definite energy. An electron in such an orbit does not radiate energy; it only absorbs or emits energy when it jumps between orbits. This idea of quantised, fixed orbits is the hallmark of Bohr's model.

Step 1 — Recall what each model proposed: Dalton treated atoms as indivisible solid spheres; Thomson pictured electrons embedded in a positive sphere (plum pudding); Rutherford showed a small dense nucleus but did not fix the electron paths.

Step 2 — Identify the fixed-orbit idea: only Bohr introduced definite energy orbits (shells) in which electrons are allowed to move.

Step 3 — Conclusion: the fixed-orbit (shell) model is Bohr's model.

Why each other option is wrong:

- (B) Thomson's model had electrons spread through a positive sphere, with no orbits.
- (C) Rutherford's nuclear model placed the electrons outside the nucleus but could not explain why they did not spiral in, and gave no fixed orbits.
- (D) Dalton's model treated the atom as a featureless indivisible particle.

Key point: Bohr = fixed orbits / shells of definite energy. Rutherford = nucleus discovered; Thomson = plum pudding; Dalton = indivisible sphere.

Final Answer: Niels Bohr $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q3

Q4.

Solution

Concept — Shapes of atomic orbitals: An orbital is the region around the nucleus where an electron is most likely to be found. Each type of orbital has a characteristic shape: an s orbital is spherical, while a p orbital has two lobes on opposite sides of the nucleus, giving the familiar dumbbell (or figure-of-eight) shape. The figure clearly shows two lobes meeting at the nucleus.

Step 1 — Read the figure: two lobes lie on opposite sides of a central point (the nucleus), with a node at the nucleus between them.

Step 2 — Match to the orbital type: the two-lobed dumbbell shape is the defining shape of a p orbital; the three p orbitals (p_x, p_y, p_z) point along the three axes.

Step 3 — Conclusion: the dumbbell-shaped orbital shown is a p orbital.

Why each other option is wrong:

- (A) An s orbital is spherical, with no directional lobes.
- (B) An f orbital has a complex, many-lobed shape, far more intricate than two lobes.
- (C) A d orbital usually has four lobes (a clover-leaf shape), not two.

Key point: s is spherical; p is dumbbell (two lobes); d is clover-leaf (four lobes, with one exception). The simple two-lobed picture always means a p orbital.

Final Answer: p orbital $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q4

Q5.

Solution

Concept — VSEPR and the shape of water: VSEPR theory predicts molecular shape from the number of electron domains around the central atom. In H_2O, oxygen has 6 valence electrons: two form O–H bonds and four remain as two lone pairs, giving four electron domains (sp^3 hybridisation) arranged tetrahedrally. The two lone pairs repel more strongly than bonding pairs, so they squeeze the two O–H bonds closer together.

Step 1 — Find the ideal angle: four electron domains give an ideal tetrahedral angle of $109.5°$.

Step 2 — Apply lone-pair repulsion: the two lone pairs push the bonding pairs together, reducing the angle below $109.5°$.

Step 3 — State the value: the H–O–H angle in water is about $104.5°$.

Why each other option is wrong:

- (A) $90°$ is far too small; it would suggest no hybridisation, which is not the case.
- (B) $120°$ is the angle for a trigonal planar (sp^2) molecule with no lone pairs.
- (D) $180°$ is the angle of a linear molecule such as CO_2.

Key point: Two lone pairs on oxygen bend water and compress the angle from $109.5°$ down to about $104.5°$. (For comparison, NH_3 with one lone pair is about $107°$.)

Final Answer: $104.5° \Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q5

Q6.

Solution

Concept — Hybridisation from electron domains: The hybridisation of a carbon atom equals its number of σ bonds plus lone pairs (its number of electron domains). In ethane, $CH_3{-}CH_3$, each carbon is joined by single bonds to three hydrogens and to the other carbon, and it has no lone pairs.

Step 1 — Count the σ bonds on one carbon: three C–H bonds + one C–C bond $= 4\ \sigma$ bonds, with no lone pairs, i.e. four electron domains.

Step 2 — Assign the hybridisation: four domains $\Rightarrow sp^3$. The four sp^3 orbitals point toward the corners of a tetrahedron.

Step 3 — Geometry: each carbon is tetrahedral, with bond angles of about $109.5°$, and there are no π bonds in ethane.

Why each other option is wrong:

- (B) sp^2 (three domains) is for a carbon with a double bond, as in ethene.
- (C) sp (two domains) is for a triple bond, as in ethyne.

- (D) sp^3d (five domains) cannot occur for carbon, which has no available d orbitals.

Key point: All single bonds $\rightarrow$ four σ bonds $\rightarrow$ sp^3 $\rightarrow$ tetrahedral. Ethane (sp^3), ethene (sp^2), and ethyne (sp) form the classic comparison set.

Final Answer: $sp^3 \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q6

Q7.

Solution

Concept — Ionic versus covalent character: The ionic character of a bond increases with the electronegativity *difference* between the two atoms. A large difference (a metal joined to a very electronegative non-metal) gives a strongly ionic bond, while a small difference gives a mostly covalent bond. So the bond with the *smallest* electronegativity difference has the least ionic character.

Step 1 — Estimate the electronegativity differences: Na–Cl ($3.0 - 0.9 \approx 2.1$), H–Cl ($3.0 - 2.1 \approx 0.9$), K–F ($4.0 - 0.8 \approx 3.2$), C–I ($2.5 - 2.5 \approx 0.0$).

Step 2 — Compare: C and I have almost equal electronegativities, so the C–I bond has the smallest difference and is the most covalent.

Step 3 — Conclusion: the C–I bond shows the least ionic character.

Why each other option is wrong:

- (A) Na–Cl has a large difference (≈ 2.1) and is strongly ionic.
- (B) H–Cl is polar covalent but still has a bigger difference than C–I.
- (C) K–F has the largest difference of all (≈ 3.2), so it is the most ionic, the opposite of what is asked.

Key point: Least ionic $=$ smallest electronegativity difference $=$ most covalent. Two atoms of similar electronegativity (like C and I) share electrons most evenly.

Final Answer: C–I $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q7

Q8.

Solution

Concept — Ionization energy down a group: Ionization energy is the minimum energy needed to remove the most loosely held electron from a gaseous atom. It depends on how strongly the nucleus holds the outermost electron. Down a group, each successive element has its valence electron in a higher, larger shell that is farther from the nucleus and better shielded by inner electrons.

Step 1 — Identify what changes down a group: the number of inner shells increases, so the outer electron is farther out and more shielded from the nuclear charge.

Step 2 — Net effect: the weaker hold on the outer electron makes it easier to remove.

Step 3 — State the trend: the first ionization energy therefore decreases down a group. (Across a period, by contrast, it generally increases.)

Why each other option is wrong:

- (A) Increasing is the trend *across a period*, not down a group.
- (C) Remaining unchanged ignores the increasing size and shielding down a group.
- (D) There is no "decrease then increase" reversal in the normal group trend.

Key point: Down a group: larger atoms, more shielding → ionization energy decreases. Across a period: rising nuclear charge → ionization energy increases.

Final Answer: Decreases $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q8

Q9.

Solution

Concept — Position of the noble gases: The periodic table is divided into 18 vertical groups. The noble gases (helium, neon, argon, krypton, xenon, radon) form the last column on the far right, group 18. They have completely filled valence shells (a stable octet, or a duplet for He), which makes them very unreactive.

Step 1 — Read the figure: the highlighted column is the last one on the right, labelled group 18.

Step 2 — Recall the group's identity: the far-right group 18 contains the noble (inert) gases.

Step 3 — Conclusion: the noble gases belong to group 18.

Why each other option is wrong:

- (A) Group 1 is the alkali metals (Li, Na, K . . .), highly reactive metals.
- (B) Group 17 is the halogens (F, Cl, Br . . .), reactive non-metals.
- (D) Group 16 is the chalcogens (the oxygen family), not the noble gases.

Key point: The noble gases sit at the extreme right (group 18) with full valence shells, which explains their lack of reactivity. Group 17 (halogens) lies just to their left.

Final Answer: group 18 $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q9

Q10.

Solution

Concept — Enthalpy of combustion: The standard enthalpy of combustion is the heat change when one mole of a substance burns completely in oxygen under standard conditions. Combustion is always an exothermic process: it releases heat to the surroundings. By the sign convention, heat released means a *negative* enthalpy change, so $\Delta_c H^\circ$ is always negative.

Step 1 — Recall the nature of combustion: burning a fuel gives out heat (and usually light), so energy flows from the system to the surroundings.

Step 2 — Apply the sign convention: for an exothermic change, $\Delta H = H_{\text{products}} - H_{\text{reactants}} < 0$.

Step 3 — Conclusion: the standard enthalpy of combustion of any fuel is always negative.

Why each other option is wrong:

- (B) Positive ΔH would mean combustion absorbs heat, which contradicts the fact that fuels release heat.
- (C) Zero would mean no energy change, but combustion clearly gives out energy.
- (D) "Sometimes positive" is wrong: complete combustion is exothermic by definition, so the value is always negative.

Key point: Combustion is always exothermic, so $\Delta_c H^\circ$ is always negative. This is exactly why fuels are used as energy sources.

Final Answer: Negative $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q10

Q11.

Solution

Concept — Free energy and the equilibrium constant: The relation $\Delta G^\circ = -RT\ln K$ links the standard Gibbs free energy change of a reaction to its equilibrium constant K, where R is the gas constant and T is the absolute temperature. It shows that the position of equilibrium is governed by thermodynamics: a large negative ΔG° corresponds to a large K (products favoured).

Step 1 — Identify the symbols: in $\Delta G^\circ = -RT\ln K$, the only reaction-specific term besides ΔG° is K.

Step 2 — Interpret K: K is the equilibrium constant, the ratio of product to reactant activities at equilibrium.

Step 3 — Conclusion: the equation connects free energy with the equilibrium constant.

Why each other option is wrong:

- (A) Activation energy belongs to kinetics (the Arrhenius equation), not to this thermodynamic relation.
- (B) Enthalpy of formation appears in ΔH calculations, not directly in $\Delta G^\circ = -RT\ln K$.
- (C) Molar heat capacity relates heat to temperature change, not free energy to equilibrium.

Key point: $\Delta G^\circ = -RT\ln K$ is the bridge between thermodynamics and equilibrium: negative $\Delta G^\circ \Rightarrow K > 1$ (forward reaction favoured).

Final Answer: equilibrium constant $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q11

Q12.

Solution

Concept — Le Chatelier's principle (decreasing pressure): If the pressure on a gaseous equilibrium is decreased (the volume is increased), the system shifts to partly oppose this by producing *more* gas molecules. So lowering the pressure favours the side with the greater number of moles of gas.

Step 1 — Count moles of gas on each side of $PCl_5 \rightleftharpoons PCl_3 + Cl_2$: the left side has 1 mole of gas, the right side has $1 + 1 = 2$ moles of gas.

Step 2 — Apply the rule: decreasing the pressure favours the side with more gas moles, which is the right side $(2 > 1)$.

Step 3 — Conclusion: the equilibrium shifts toward the side with more moles of gas, i.e. toward the products $PCl_3 + Cl_2$.

Why each other option is wrong:

- (A) Shifting toward fewer gas moles is the response to *increasing* the pressure, not decreasing it.
- (C) The two sides do not have equal moles (1 versus 2), so this does not apply.
- (D) A pressure change does cause a shift whenever the moles of gas differ on the two sides.

Key point: Lower pressure → shift to the side with *more* gas moles; higher pressure → shift to the side with *fewer* gas moles. Only the difference in gas moles matters.

Final Answer: more moles of gas $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q12

Q13.

Solution

Concept — Buffer solutions: A buffer resists changes in pH when small amounts of acid or base are added. An *acidic* buffer is made from a weak acid together with the salt of that acid (which provides its conjugate base). For example, acetic acid (CH_3COOH) plus sodium acetate (CH_3COONa): the weak acid neutralises added base, while the conjugate base (acetate ion) neutralises added acid.

Step 1 — Recall the two components of an acidic buffer: a weak acid and its salt (its conjugate base).

Step 2 — See how it works: the weak acid mops up added OH^-, and the conjugate base mops up added H^+, keeping the pH nearly constant.

Step 3 — Conclusion: the partner of the weak acid in the buffer is its salt (conjugate base).

Why each other option is wrong:

- (A) A strong acid is fully ionised and cannot act as the conjugate-base reservoir a buffer needs.
- (B) A strong base would react completely with the weak acid, not form a buffer with it.
- (D) Two unrelated weak acids do not provide the matched conjugate pair required.

Key point: Acidic buffer = weak acid + its salt (conjugate base); basic buffer = weak base + its salt (conjugate acid). The matched conjugate pair is what gives the buffering action.

Final Answer: salt (conjugate base) $\Rightarrow \boxed{\text{C}}$

$\boxed{\textbf{Answer: (C)}}$ Go Back to Q13

Q14.

Solution

Concept — Conjugate base: In Bronsted–Lowry theory, an acid donates a proton (H^+); the species left behind after donating one proton is its conjugate base. To find the conjugate base of sulfuric acid, remove exactly one H^+ from H_2SO_4.

Step 1 — Remove one proton: $H_2SO_4 \rightarrow H^+ + HSO_4^-$.

Step 2 — Identify what remains: the species left is the hydrogen sulfate (bisulfate) ion, HSO_4^-.

Step 3 — Conclusion: the conjugate base of H_2SO_4 is HSO_4^-.

Why each other option is wrong:

- (B) SO_4^{2-} is formed by removing *two* protons; it is the conjugate base of HSO_4^-, not of H_2SO_4.
- (C) $H_3SO_4^+$ would be the conjugate *acid* (formed by adding a proton).
- (D) H_2SO_3 is sulfurous acid, a different compound with one fewer oxygen.

Key point: Conjugate base = acid minus one H^+. Each proton removed gives the next conjugate base: $H_2SO_4 \rightarrow HSO_4^- \rightarrow SO_4^{2-}$.

Final Answer: $HSO_4^- \Rightarrow \boxed{A}$

Answer: (A) Go Back to Q14

Q15.

Solution

Concept — Oxidation number rules: In a neutral compound the oxidation numbers of all atoms sum to zero. Oxygen is normally -2. The unknown oxidation number of the metal is then found algebraically so that the totals balance.

Given: Fe_2O_3, a neutral compound; O $= -2$.

Step 1 — Assign the known values: three oxygens give $3 \times (-2) = -6$. Let the oxidation number of each Fe be x; there are two Fe atoms, contributing $2x$.

Step 2 — Set the sum to zero: $2x + (-6) = 0$, i.e. $2x - 6 = 0$.

Step 3 — Solve: $2x = 6 \Rightarrow x = +3$.

Why each other option is wrong:

- (A) $+2$ is the oxidation number of Fe in FeO (iron(II) oxide), not in Fe_2O_3.
- (B) $+6$ would come from forgetting that there are *two* iron atoms sharing the charge.
- (C) 0 is the oxidation number of iron metal in its elemental form, not in a compound.

Key point: The name iron(III) oxide already tells you Fe is $+3$. Setting $2x - 6 = 0$ confirms $x = +3$. Always remember to use *both* iron atoms in the balance.

Final Answer: $+3 \Rightarrow \boxed{D}$

Answer: (D) Go Back to Q15

Q16.

Solution

Concept — The standard hydrogen electrode (SHE): Electrode potentials cannot be measured in isolation, so a reference is chosen. The standard hydrogen electrode, with H_2 gas at 1 bar bubbling over a platinum electrode dipped in 1 M H^+ at 298 K, is taken as the universal reference and is assigned a standard potential of exactly 0.00 V. All other standard electrode potentials are measured relative to it.

Step 1 — Recall the convention: the SHE potential is fixed by definition at 0.00 V.

Step 2 — Use as a reference: a metal that is more easily oxidised than hydrogen gets a negative E°, and one less easily oxidised gets a positive E°.

Step 3 — Conclusion: the standard potential of the SHE is 0.00 V.

Why each other option is wrong:

- (A) $+1.00$ V is not assigned; the SHE is the zero of the scale, not a positive value.
- (C) -1.00 V is likewise not the reference value.
- (D) $+0.76$ V is (in magnitude) the standard potential of the zinc electrode relative to the SHE, not the SHE itself.

Key point: The SHE is the agreed zero (0.00 V) against which all other electrode potentials are measured. Negative E° means more reactive than hydrogen; positive means less reactive.

Final Answer: $0.00 \text{ V} \Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q16

Q17.

Solution

Concept — Faraday's laws of electrolysis: One faraday (≈ 96500 C) is the charge carried by one mole of electrons. To deposit one mole of a metal ion M^{n+}, exactly n moles of electrons are needed, because the half-reaction is $M^{n+} + n\,e^- \rightarrow M$. So the number of faradays equals the charge n of the metal ion.

Given: the metal is trivalent, so $n = 3$ ($M^{3+} + 3e^- \rightarrow M$).

Step 1 — Read the half-reaction: depositing one M^{3+} ion needs 3 electrons.

Step 2 — Scale to one mole: depositing one mole of M needs 3 moles of electrons, i.e. 3 faradays.

Step 3 — Conclusion: 3 faradays are required.

Why each other option is wrong:

- (A) 1 faraday deposits one mole of a *monovalent* metal such as Ag^+.
- (B) 2 faradays deposit one mole of a *divalent* metal such as Cu^{2+}.
- (D) 6 faradays would deposit one mole of a hexavalent ion, not a trivalent one.

Key point: The number of faradays needed to deposit one mole of a metal equals the charge on the ion: 1 for M^+, 2 for M^{2+}, 3 for M^{3+}.

Final Answer: 3 faradays $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q17

Q18.

Solution

Concept — Moles from molarity and volume: Molarity is moles of solute per litre of solution: $M = \frac{n}{V}$, where V is in litres. Rearranging gives the number of moles as $n = M \times V$. The volume must be converted from millilitres to litres before substituting.

Given: molarity $M = 0.2$ M; volume $= 250$ mL $= 0.250$ L.

Step 1 — Write the formula: $n = M \times V$.

Step 2 — Substitute with units: $n = 0.2 \text{ mol L}^{-1} \times 0.250 \text{ L}$.

Step 3 — Evaluate: $n = 0.05$ mol.

Why each other option is wrong:

- (B) 0.2 mol would be the amount in a full litre, not in 250 mL.
- (C) 0.5 mol ignores the molarity and treats the volume incorrectly.
- (D) 0.025 mol comes from using 125 mL or halving the result by mistake.

Key point: Always convert mL to L first (250 mL $= 0.25$ L), then multiply by molarity. Here a quarter-litre of 0.2 M solution holds $0.2 \times 0.25 = 0.05$ mol.

Final Answer: 0.05 mol $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q18

Q19.

Solution

Concept — Kinetic theory and gas pressure: The kinetic theory pictures a gas as a large number of tiny molecules in constant, rapid, random motion. These molecules continually strike the walls of the container. Each collision exerts a small force on the wall; the combined effect of countless collisions per second, spread over the wall area, is what we measure as the gas pressure.

Step 1 — Identify the source of pressure: pressure is force per unit area, and the force on the walls comes from molecules hitting them.

Step 2 — Relate to molecular motion: faster or more frequent collisions (higher temperature or more molecules) mean a greater force and hence higher pressure.

Step 3 — Conclusion: gas pressure arises from collisions of the molecules with the container walls.

Why each other option is wrong:

- (A) The kinetic theory of an ideal gas assumes *no* attractive forces between molecules.
- (B) The molecules themselves are assumed to have negligible volume; pressure is not due to their size.
- (C) An inert gas does not chemically react with the walls; the pressure is purely from physical collisions.

Key point: Gas pressure = the result of innumerable elastic collisions of molecules with the walls. More frequent or more energetic collisions give a higher pressure.

Final Answer: collisions of the molecules with the walls $\Rightarrow \boxed{D}$

Answer: (D) Go Back to Q19

Q20.

Solution

Concept — Oxidation state in a complex ion: In a coordination complex, the overall charge equals the oxidation state of the central metal plus the sum of the ligand charges. Neutral ligands (such as H_2O) contribute 0 to the charge balance, so the metal's oxidation state equals the charge on the complex ion.

Given: $[\mathrm{Mn(H_2O)_6}]^{2+}$; each H_2O ligand is neutral (0).

Step 1 — Set up the charge balance: let the oxidation state of Mn be x. Then $x + 6 \times (0) = +2$.

Step 2 — Simplify: $x + 0 = +2$.

Step 3 — Solve: $x = +2$.

Why each other option is wrong:

- (A) $+1$ would make the overall charge $+1$, not the given $+2$.
- (C) $+4$ would require the complex to carry a $+4$ charge.
- (D) $+6$ is the oxidation state of Mn in the manganate ion, not in this aqua complex.

Key point: For neutral ligands like water or ammonia, the metal's oxidation state simply equals the charge on the complex ion. Here that charge is $+2$, so Mn is $+2$.

Final Answer: $+2 \Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q20

Q21.

Solution

Concept — Dissociation of a complex salt: A coordination compound dissociates in water into the ions that lie *outside* the square brackets and the single complex ion *inside* the brackets. The complex ion stays intact as one unit; only the counter-ions are released separately. So we count the outer ions plus the one complex ion.

Given: $\mathrm{K_4[Fe(CN)_6]}$.

Step 1 — Identify the parts: four potassium ions (K^+) lie outside the brackets, and the ferrocyanide complex $[\mathrm{Fe(CN)_6}]^{4-}$ is one ion inside the brackets.

Step 2 — Write the dissociation: $\mathrm{K_4[Fe(CN)_6]} \rightarrow 4\,\mathrm{K^+} + [\mathrm{Fe(CN)_6}]^{4-}$.

Step 3 — Count the ions: 4 potassium ions $+1$ complex ion $= 5$ ions in total.

Why each other option is wrong:

- (A) 3 and (B) 4 undercount; they forget either the complex ion or one of the potassium ions.
- (D) 11 wrongly breaks the complex ion apart into Fe^{2+} and six CN^-, but the CN ligands stay bound inside the complex and are not released.

Key point: The complex ion within the brackets does not split into its ligands in solution; it dissociates as a single ion. So $K_4[Fe(CN)_6]$ gives $4 + 1 = 5$ ions.

Final Answer: 5 ions $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q21

Q22.

Solution

Concept — The chalcogens: The chalcogens are the elements of the oxygen family: oxygen, sulfur, selenium, tellurium and polonium. They occupy group 16 of the periodic table and have six valence electrons (configuration $ns^2\,np^4$), so they tend to gain two electrons or form two covalent bonds.

Step 1 — Recall the family: the oxygen family (chalcogens) is the column headed by oxygen.

Step 2 — Locate the group: this column is group 16 in the modern periodic table.

Step 3 — Conclusion: the chalcogens belong to group 16.

Why each other option is wrong:

- (B) Group 15 is the nitrogen family (pnictogens), with five valence electrons.
- (C) Group 17 is the halogens, with seven valence electrons.
- (D) Group 14 is the carbon family, with four valence electrons.

Key point: Group $16 =$ chalcogens (O, S, Se, Te), six valence electrons. Remember the order: 14 carbon family, 15 nitrogen family, 16 oxygen family (chalcogens), 17 halogens, 18 noble gases.

Final Answer: group $16 \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q22

Q23.

Solution

Concept — Position isomerism: Isomers have the same molecular formula but different structures. Position isomers have the same carbon skeleton and the same functional groups (or substituents), but those substituents are attached at *different positions*. Ortho-, meta- and para-xylene all have two methyl groups on a benzene ring; only the positions of those methyls differ.

Step 1 — Compare the three xylenes: each is dimethylbenzene (C_8H_{10}); the difference lies only in where the two CH_3 groups sit (1,2 for ortho, 1,3 for meta, 1,4 for para).

Step 2 — Match to the type of isomerism: the same groups in different positions on the same skeleton is the definition of position isomerism.

Step 3 — Conclusion: the xylenes show position isomerism.

Why each other option is wrong:

- (A) Chain isomerism needs a different carbon *skeleton*, but all three xylenes share the same benzene ring skeleton.
- (B) Functional isomerism needs different functional groups; here the groups are identical methyls.
- (C) Geometrical isomerism arises from restricted rotation about a double bond or ring (cis/trans), which is not what distinguishes the xylenes.

Key point: Same skeleton, same groups, different *positions* = position isomerism. The ortho/meta/para labels are the classic example.

Final Answer: position isomerism $\Rightarrow \boxed{\text{D}}$

Answer: (D) **Go Back to Q23**

Q24.

Solution

Concept — IUPAC naming of aldehydes: An aldehyde is named by taking the alkane stem for the number of carbons and replacing the final "-e" with the suffix "-al". Formaldehyde, HCHO, has a single carbon bearing the $-$CHO group, so its stem is "meth-" (one carbon) and its IUPAC name is methanal.

Step 1 — Count the carbons: HCHO has one carbon atom, so the stem is "meth".

Step 2 — Add the aldehyde suffix: the $-$CHO functional group gives the suffix

"-al".

Step 3 — Combine: meth + anal = methanal.

Why each other option is wrong:

- (A) Methanol (CH_3OH) is the one-carbon *alcohol* (−ol suffix), not an aldehyde.
- (C) Ethanal (CH_3CHO) has *two* carbons, so its stem is "eth", not "meth".
- (D) Methanoic acid (HCOOH) is the one-carbon carboxylic acid, with a −COOH group, not −CHO.

Key point: Aldehyde → suffix "-al"; alcohol → "-ol"; acid → "-oic acid". One carbon with −CHO is methanal.

Final Answer: methanal $\Rightarrow \boxed{B}$

Answer: (B) Go Back to Q24

Q25.

Solution

Concept — Homologous series: A homologous series is a family of organic compounds with the same general formula and similar chemical properties, in which each member differs from the next by a constant unit. That constant difference is one CH_2 group (a relative mass difference of 14).

Step 1 — Recall the rule: consecutive members of any homologous series differ by CH_2.

Step 2 — Check with the alkanes: methane CH_4, ethane C_2H_6, propane C_3H_8 – each gains one C and two H, i.e. a CH_2 unit, on moving to the next member.

Step 3 — Conclusion: the difference between two consecutive homologues is CH_2.

Why each other option is wrong:

- (A) CH_4 is the formula of methane itself, not the repeating difference.
- (B) C_2H_4 is twice the correct difference (two CH_2 units).
- (D) CH_3 is a methyl group, not the CH_2 difference between successive members.

Key point: Successive homologues always differ by a CH_2 unit (mass 14). This is why members of a series show a smooth gradation in physical properties.

Final Answer: $CH_2 \Rightarrow \boxed{C}$

Answer: (C) Go Back to Q25

Q26.

Solution

Concept — Reactions of alkanes: Alkanes are saturated hydrocarbons with only single bonds, so they cannot undergo addition. Instead, in the presence of sunlight (UV light), they react with halogens such as chlorine by *substitution*: a hydrogen atom of the alkane is replaced by a chlorine atom, with HCl as a by-product. This proceeds by a free-radical chain mechanism.

Step 1 — Look at the reaction: $CH_4 + Cl_2 \xrightarrow{\text{sunlight}} CH_3Cl + HCl$.

Step 2 — Identify what happens: one H of methane is replaced by Cl; an atom is swapped for another, which is substitution (not addition or elimination).

Step 3 — Conclusion: the reaction is a substitution reaction (free-radical halogenation).

Why each other option is wrong:

- (B) Addition occurs across double or triple bonds (alkenes/alkynes); alkanes are saturated and cannot add.
- (C) Elimination removes a small molecule to *form* a double bond; that is not happening here.
- (D) Polymerisation joins many monomer units into a long chain, which is unrelated to chlorination of methane.

Key point: Saturated alkanes undergo free-radical *substitution* with halogens in sunlight; unsaturated alkenes/alkynes undergo addition. The clue "alkane + halogen + sunlight" always points to substitution.

Final Answer: substitution $\Rightarrow \boxed{A}$

Answer: (A) Go Back to Q26

Q27.

Solution

Concept — Sigma bonds in benzene: A σ bond is a single, head-on overlap bond; every single bond is a σ bond, and each double bond contains one σ bond plus one π bond. Benzene, C_6H_6, is a ring of six carbon atoms, each carbon bonded to one hydrogen and to its two ring neighbours.

Step 1 — Count the C–H σ bonds: each of the six carbons has one C–H bond, giving 6 C–H σ bonds.

Step 2 — Count the C–C σ bonds: the six carbons form a closed ring of six C–C bonds; each contributes one σ bond (the extra π electrons are delocalised but do not add to the σ count), giving 6 C–C σ bonds.

Step 3 — Add them up: 6 (C–H) $+6$ (C–C) $= 12$ σ bonds in total.

Why each other option is wrong:

- (A) 6 counts only the C–C ring bonds and ignores the six C–H bonds.
- (B) 9 does not match any correct count of the bonds in benzene.
- (C) 15 wrongly adds the three π bonds to the 12 σ bonds, but π bonds are not σ bonds.

Key point: Benzene has 6 C–C $+6$ C–H $= 12$ σ bonds, plus 3 delocalised π bonds. Only the σ bonds are counted here, giving 12.

Final Answer: $12 \Rightarrow \boxed{\text{D}}$

Answer: (D) **Go Back to Q27**

Q28.

Solution

Concept — Position of the carbonyl in an aldehyde: The carbonyl group is a carbon double-bonded to oxygen ($>\text{C} = \text{O}$). In an aldehyde the carbonyl carbon also carries at least one hydrogen, which means it must sit at the *end* of the carbon chain (a terminal carbon). In a ketone, by contrast, the carbonyl carbon is bonded to two other carbons and therefore lies in the middle of the chain.

Step 1 — Recall the aldehyde group: the $-$CHO group has the carbonyl carbon bonded to an H, so it can only be at the chain end.

Step 2 — Check with ethanal: in CH_3CHO the C=O is on the last carbon, the terminal one.

Step 3 — Conclusion: in an aldehyde the carbonyl group is at the end (terminal carbon) of the chain.

Why each other option is wrong:

- (A) A carbonyl in the middle of the chain (between two carbons) is the feature of a *ketone*, not an aldehyde.
- (C) Nitrogen is not part of the carbonyl group at all.
- (D) Oxygen is double-bonded to the carbon, but the group is located on a carbon atom, not on oxygen alone.

Key point: Aldehyde → carbonyl at a terminal carbon (with an H attached, $-$CHO); ketone → carbonyl on an internal carbon (between two carbons). The chain-end position is what distinguishes an aldehyde.

Final Answer: end (terminal carbon) of the chain $\Rightarrow \boxed{\text{B}}$

Answer: (B) **Go Back to Q28**

Q29.

Solution

Concept — Polysaccharides as glucose polymers: Starch and cellulose are polysaccharides, a class of carbohydrates built from many small sugar units joined together. In both cases the repeating monomer unit is glucose: starch is a polymer of α-glucose, while cellulose is a polymer of β-glucose. The two differ only in the type of glycosidic linkage between the glucose units.

Step 1 — Identify the monomer: the building block of both starch and cellulose is the monosaccharide glucose.

Step 2 — Note the difference: starch uses α-1,4 linkages (digestible by humans), cellulose uses β-1,4 linkages (not digestible by humans), but both are made of glucose.

Step 3 — Conclusion: starch and cellulose are natural polymers of glucose.

Why each other option is wrong:

- (A) Amino acids are the monomers of *proteins*, not of carbohydrates.
- (B) Fatty acids (with glycerol) build up *fats and lipids*, not polysaccharides.
- (D) Nucleotides are the monomers of *nucleic acids* (DNA and RNA).

Key point: Both starch and cellulose are glucose polymers; they differ only in the linkage (α versus β). Match the polymer to its monomer: polysaccharides $\leftarrow$ glucose, proteins $\leftarrow$ amino acids, nucleic acids $\leftarrow$ nucleotides.

Final Answer: glucose $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q29

Q30.

Solution

Concept — Esterification: Esterification is the reaction in which a carboxylic acid reacts with an alcohol, usually in the presence of a small amount of concentrated sulfuric acid as catalyst, to form an ester together with water. The general reaction is R–COOH + R′–OH $\rightleftharpoons$ R–COO–R′ + H_2O.

Step 1 — Identify the reactants and products: acid + alcohol $\rightarrow$ ester + water.

Step 2 — Name the process: the formation of an ester from an acid and an alcohol is called esterification.

Step 3 — Conclusion: the named reaction is esterification.

Why each other option is wrong:

- (B) Saponification is the *reverse*-type reaction: the alkaline hydrolysis of an ester (or fat) to give a soap and an alcohol.
- (C) Neutralisation is the reaction of an acid with a base to give a salt and water, not an ester.
- (D) Fermentation is the enzymatic breakdown of sugars (e.g. to ethanol), not the joining of an acid and an alcohol.

Key point: Acid + alcohol $\xrightarrow{H^+}$ ester + water is esterification (a condensation). Its reverse, the alkaline hydrolysis of an ester, is saponification.

Final Answer: esterification $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q30

Answer Key

Q	Ans	Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	B	2	C	3	A	4	D	5	C
6	A	7	D	8	B	9	C	10	A
11	D	12	B	13	C	14	A	15	D
16	B	17	C	18	A	19	D	20	B
21	C	22	A	23	D	24	B	25	C
26	A	27	D	28	B	29	C	30	A