AIIMS B.Sc Nursing Chemistry

Sample Paper – 2

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 8 g of oxygen gas (O_2, molar mass $= 32$ g mol^{-1}) is:

(A) 0.125 mol

(B) 0.25 mol

(C) 0.5 mol

(D) 1 mol

Q2. The number of atoms present in 0.5 mol of oxygen gas (O_2) is (Avogadro number $= 6.022 \times 10^{23}$):

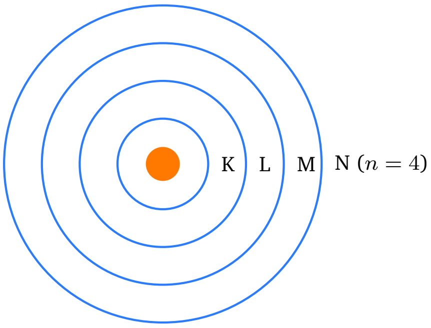
(A) 3.011×10^{23}

(B) 1.505×10^{23}

(C) 6.022×10^{23}

(D) 1.2×10^{24}

Q3. The shells of an atom are shown below. The maximum number of electrons that the N shell ($n = 4$) can hold, given by $2n^2$, is:



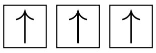
(A) 8

(B) 16

(C) 18

(D) 32

Q4. The filling of the $2p$ subshell of a nitrogen atom ($Z = 7$, configuration $1s^2\,2s^2\,2p^3$) is shown. The number of unpaired electrons in a ground-state nitrogen atom is:



$2p^3$

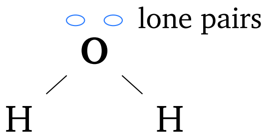
(A) 1

(B) 2

(C) 3

(D) 0

Q5. The structure of a water (H_2O) molecule, with two lone pairs on the oxygen atom, is shown. The shape of the molecule is:



bond angle ≈ 104.5°

(A) bent (angular)
(B) linear
(C) trigonal planar
(D) tetrahedral

Q6. The hybridization of the central carbon atom in a methane molecule (CH_4, which has four single C–H bonds) is:

(A) sp
(B) sp^3
(C) sp^2
(D) sp^3d

Q7. Which one of the following compounds contains an ionic bond?

(A) CH_4
(B) CO_2
(C) HCl
(D) NaCl

Q8. On moving from left to right across a period in the periodic table, the first ionization energy of the elements generally:

(A) decreases
(B) remains constant
(C) increases
(D) first decreases, then increases

Q9. Among the following period-3 elements, which one is the most metallic (and has the largest atomic size)?

(A) Sodium (Na)
(B) Magnesium (Mg)

(C) Sulphur (S)

(D) Chlorine (Cl)

Q10. Which one of the following changes is endothermic (absorbs heat, so ΔH is positive)?

(A) Combustion of methane

(B) Melting of ice

(C) Neutralisation of an acid by a base

(D) Condensation of steam to water

Q11. A chemical reaction at constant temperature and pressure is spontaneous when the Gibbs free energy change ΔG is:

(A) equal to zero

(B) greater than zero (positive)

(C) equal to ΔH

(D) less than zero (negative)

Q12. For the exothermic equilibrium $\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g) + \text{heat}$, lowering the temperature shifts the equilibrium:

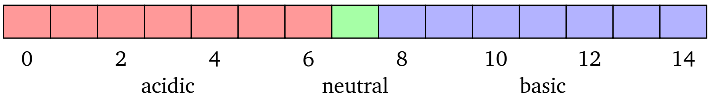
(A) towards the reactants

(B) no shift occurs

(C) towards both sides equally

(D) towards the products (ammonia)

Q13. Using the pH scale shown, the pH of a 0.001 M solution of hydrochloric acid (a strong acid, fully ionized) is:



(A) 1

(B) 2

(C) 11

(D) 3

Q14. According to the Bronsted–Lowry theory, the conjugate acid of ammonia (NH_3) is:

(A) NH_2^-

(B) N^{3-}

(C) NH_4^+

(D) NH_3

Q15. The oxidation number of sulphur in sulphuric acid (H_2SO_4) is:

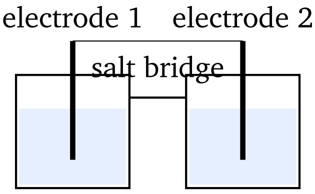
(A) $+2$

(B) $+6$

(C) $+4$

(D) -2

Q16. A galvanic (voltaic) cell with two electrodes joined by a salt bridge is shown. In such a cell, reduction takes place at the:



(A) cathode

(B) anode

(C) salt bridge

(D) electrolyte

Q17. The quantity of charge required to deposit 1 mole of copper from a copper salt solution ($Cu^{2+} + 2e^- \rightarrow Cu$) is (Faraday constant $= 96500$ C mol^{-1}):

(A) 96500 C

(B) 48250 C

(C) 193000 C

(D) 9650 C

Q18. The molarity of a solution that contains 1 mol of solute dissolved in enough water to make 500 mL of solution is:

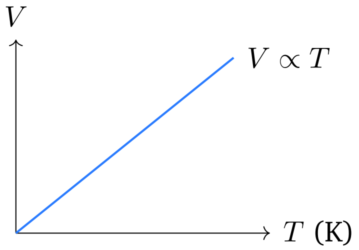
(A) 1 M

(B) 2 M

(C) 0.5 M

(D) 4 M

Q19. The graph shows that, at constant pressure, the volume of a fixed mass of an ideal gas is directly proportional to its absolute temperature (Charles's law). If the absolute (Kelvin) temperature is doubled at constant pressure, the volume becomes:



(A) doubled

(B) halved

(C) unchanged

(D) four times

Q20. The oxidation state of copper in the complex ion $[Cu(NH_3)_4]^{2+}$ (each NH_3 ligand is neutral) is:

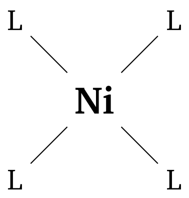
(A) 0

(B) $+1$

(C) $+3$

(D) $+2$

Q21. In the complex $[Ni(CO)_4]$ the four CO ligands (L) are arranged around the central nickel atom as shown. The coordination number of nickel in this complex is:



(A) 2

(B) 6

(C) 4

(D) 3

Q22. Which one of the following elements is an alkali metal (group 1)?

(A) Calcium

(B) Potassium

(C) Aluminium

(D) Iron

Q23. The number of structural (chain) isomers of pentane, C_5H_{12}, is:

(A) 2

(B) 3

(C) 4

(D) 5

Q24. The IUPAC name of the compound CH_3–CHO is:

(A) ethanal

(B) methanal

(C) ethanol

(D) ethanoic acid

Q25. The general molecular formula of the alkane homologous series is:

(A) C_nH_{2n}

(B) C_nH_{2n-2}

(C) C_nH_n

(D) C_nH_{2n+2}

Q26. Which one of the following is a saturated hydrocarbon (containing only single carbon–carbon bonds)?

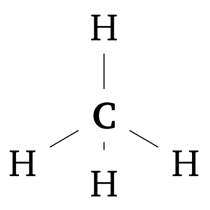
(A) Ethene (C_2H_4)

(B) Ethyne (C_2H_2)

(C) Ethane (C_2H_6)

(D) Benzene (C_6H_6)

Q27. The structure of a methane molecule (CH_4) is shown. The number of C–H bonds in one methane molecule is:



(A) 4

(B) 3

(C) 2

(D) 5

Q28. The functional group shown below, an oxygen bonded to a hydrogen and attached to a carbon chain (written $-OH$), is characteristic of which class of compounds?

$$R - O - H$$

($-OH$ group)

(A) carboxylic acids
(B) alcohols
(C) aldehydes
(D) amines

Q29. The building-block monomer units of polysaccharides (carbohydrates such as starch and cellulose) are:

(A) monosaccharides (e.g. glucose)
(B) amino acids
(C) nucleotides
(D) fatty acids

Q30. Which one of the following functional groups makes an organic compound acidic (able to donate an H^+)?

(A) $-OH$ (hydroxyl)
(B) $-CHO$ (aldehyde)
(C) $-NH_2$ (amino)
(D) $-COOH$ (carboxyl)

Detailed Solutions

Q1.

Solution

Concept — The mole and molar mass: The mole is the SI unit for the amount of a substance, and one mole contains Avogadro's number (6.022×10^{23}) of particles with a mass equal to the molar mass in grams. A measured mass is converted to a count of moles using $n = \dfrac{\text{given mass}}{\text{molar mass}}$. For oxygen gas the molar mass is that of the O_2 molecule, $2 \times 16 = 32$ g mol^{-1}.

Given: mass $= 8$ g; molar mass $M(O_2) = 32$ g mol^{-1}.

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

Step 2 — Substitute the data with units: $n = \dfrac{8\text{ g}}{32\text{ g mol}^{-1}}$.

Step 3 — Evaluate: the grams cancel, leaving $n = 0.25$ mol.

Why each other option is wrong:

- (A) 0.125 mol would need a molar mass of 64 g mol^{-1}, twice the true value of O_2.
- (C) 0.5 mol corresponds to $0.5 \times 32 = 16$ g of O_2, not 8 g.
- (D) 1 mol is the full molar mass, 32 g.

Key point: Use the molar mass of the molecule O_2 (32), not of a single O atom (16). A common slip is dividing by 16 and getting 0.5 mol.

Final Answer: 0.25 mol $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q1

Q2.

Solution

Concept — Atoms versus molecules: Avogadro's number, $N_A = 6.022 \times 10^{23}$, is the count of particles in one mole. The number of *molecules* is $n \times N_A$, but the number of *atoms* also depends on how many atoms each molecule contains. Oxygen gas is diatomic (O_2), so each molecule contains two atoms; the atom count is therefore $n \times N_A \times 2$.

Given: $n = 0.5$ mol of O_2; $N_A = 6.022 \times 10^{23}$; atoms per molecule $= 2$.

Step 1 — Number of molecules: $0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$ molecules.

Step 2 — Convert to atoms: multiply by 2 atoms per O_2 molecule: $3.011 \times 10^{23} \times 2$.

Step 3 — Evaluate: $= 6.022 \times 10^{23}$ atoms.

Why each other option is wrong:

- (A) 3.011×10^{23} is the number of *molecules*, not atoms; it forgets the factor of 2.
- (B) 1.505×10^{23} corresponds to 0.25 mol of molecules, half the given amount.
- (D) 1.2×10^{24} is twice too large; it would be the atom count of 1 mol of O_2.

Key point: For a diatomic gas, atoms = molecules $\times 2$. The phrase "number of atoms" is the trap: half a mole of O_2 has 3.011×10^{23} molecules but 6.022×10^{23} atoms.

Final Answer: 6.022×10^{23} atoms $\Rightarrow \boxed{\text{C}}$

$\boxed{\textbf{Answer: (C)}}$ **Go Back to Q2**

Q3.

Solution

Concept — Capacity of a shell: Electrons occupy shells labelled by the principal quantum number n ($n = 1$ K, $n = 2$ L, $n = 3$ M, $n = 4$ N). The maximum number of electrons a shell can hold is $2n^2$, which comes from n^2 orbitals per shell times two electrons per orbital (Pauli's exclusion principle).

Step 1 — Identify n: the N shell is the fourth shell, so $n = 4$.

Step 2 — Apply $2n^2$: $2 \times (4)^2 = 2 \times 16$.

Step 3 — Evaluate: $= 32$ electrons.

Why each other option is wrong:

- (A) 8 is the capacity of the L shell ($n = 2$): $2 \times 2^2 = 8$.
- (B) 16 comes from incorrectly using n^2 rather than $2n^2$.
- (C) 18 is the M shell ($n = 3$): $2 \times 3^2 = 18$.

Key point: The shell capacities run K$= 2$, L$= 8$, M$= 18$, N$= 32$. Always square n first, then double; using n^2 alone (giving 16) is the usual mistake.

Final Answer: 32 electrons $\Rightarrow \boxed{\text{D}}$

$\boxed{\text{Answer: (D)}}$ Go Back to Q3

Q4.

Solution

Concept — Hund's rule and unpaired electrons: Electrons fill degenerate orbitals singly, with parallel spins, before any pairing begins (Hund's rule of maximum multiplicity). Nitrogen has $Z = 7$, giving the ground-state configuration $1s^2\,2s^2\,2p^3$. The $1s$ and $2s$ subshells are full and paired; the unpaired electrons come from the $2p^3$.

Step 1 — Write the configuration: $1s^2\,2s^2\,2p^3$.

Step 2 — Distribute the three $2p$ electrons among the three $2p$ orbitals using Hund's rule: one electron goes into each orbital with the same spin, giving $\uparrow|\uparrow|\uparrow$, with no pairing.

Step 3 — Count the unpaired electrons: all three $2p$ electrons are unpaired, so there are 3 unpaired electrons.

Why each other option is wrong:

- (A) 1 would require two of the three $2p$ electrons to pair, which violates Hund's rule.
- (B) 2 is the count for an oxygen ($2p^4$) atom, not nitrogen.
- (D) 0 would mean a fully paired subshell, impossible with an odd $2p^3$.

Key point: A half-filled $2p^3$ subshell has the maximum possible unpaired electrons (three), which makes it extra stable. These three unpaired electrons make the nitrogen atom strongly paramagnetic.

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

$\boxed{\text{Answer: (C)}}$ Go Back to Q4

Q5.

Solution

Concept — VSEPR shape of H_2O: VSEPR theory predicts shape from the electron domains around the central atom: bonding pairs and lone pairs spread out as far as possible. Oxygen in water has 6 valence electrons: two form O–H bonds and the other four form two lone pairs, giving four electron domains in total (sp^3 hybridisation).

Step 1 — Electron geometry: four electron domains arrange themselves tetrahedrally.

Step 2 — Account for the lone pairs: two of the four positions are lone pairs, which are not "seen" in the shape but push hard on the two O–H bonds.

Step 3 — Molecular shape: only two bonds remain, so the visible shape is bent (angular). The two lone pairs squeeze the bonds, reducing the angle from 109.5° to about 104.5°.

Why each other option is wrong:

- (B) Linear (180°) needs two domains and no lone pairs, as in CO_2.
- (C) Trigonal planar needs three bonding domains and no lone pairs, as in BF_3.
- (D) Tetrahedral is the *electron* geometry; the shape is tetrahedral only when all four domains are bonds, as in CH_4.

Key point: Two bonds plus two lone pairs always give a bent shape. The bent shape (and the unequal sharing of electrons) is what makes water a polar molecule.

Final Answer: Bent (angular) $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q5

Q6.

Solution

Concept — Hybridisation from σ bonds: A carbon atom's hybridisation equals its number of σ bonds plus lone pairs (its number of electron domains). In methane, CH_4, the central carbon forms four single C–H bonds and has no lone pairs, so it has four σ bonds.

Step 1 — Count the electron domains on carbon: four C–H σ bonds and zero lone pairs $= 4$ domains.

Step 2 — Assign the hybridisation: four domains $\Rightarrow sp^3$. One s orbital mixes with three p orbitals to give four equivalent sp^3 hybrid orbitals.

Step 3 — Geometry: the four sp^3 orbitals point to the corners of a tetrahedron, so methane is tetrahedral with H–C–H angles of $109.5°$.

Why each other option is wrong:

- (A) sp (two domains) is for a triple bond or two double bonds, as in ethyne.
- (C) sp^2 (three domains) is for a double bond, as in ethene.
- (D) sp^3d (five domains) cannot occur for carbon, which has no available d orbitals.

Key point: Four single bonds always mean sp^3 and a tetrahedral $109.5°$ shape. Methane is the textbook example of sp^3 carbon.

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

Answer: (B) Go Back to Q6

Q7.

Solution

Concept — Ionic versus covalent bonding: An ionic bond forms by the complete transfer of electrons from a metal to a non-metal, producing oppositely charged ions held together by electrostatic attraction. A covalent bond forms by the sharing of electrons between non-metals. So the compound with a metal–non-metal pair and a large electronegativity difference is the ionic one.

Step 1 — Classify each option: CH_4 (C and H, both non-metals) is covalent; CO_2 (C and O, non-metals) is covalent; HCl (H and Cl, non-metals) is covalent (a polar covalent molecule).

Step 2 — Examine NaCl: sodium is a metal (group 1) and chlorine a non-metal (group 17). Na loses one electron to become Na^+ and Cl gains it to become Cl^-; the ions attract electrostatically.

Step 3 — Conclusion: NaCl contains an ionic bond.

Why each other option is wrong:

- (A) CH_4 shares electrons between carbon and hydrogen (covalent).
- (B) CO_2 has covalent C=O double bonds.

- (C) HCl is a polar *covalent* molecule, not ionic, because both atoms are non-metals.

Key point: Metal + non-metal → usually ionic (electron transfer); non-metal + non-metal → covalent (electron sharing). HCl is polar but still covalent.

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

Answer: (D) Go Back to Q7

Q8.

Solution

Concept — Ionization energy across a period: The first ionization energy is the minimum energy needed to remove the outermost electron from a gaseous atom. Across a period (left to right), the nuclear charge rises while electrons are added to the same shell, so the outer electrons are held more tightly and are harder to remove.

Step 1 — Identify what changes across a period: the principal shell stays the same, but the effective nuclear charge increases with each successive element.

Step 2 — Net effect: a stronger nuclear pull on the outer electrons means more energy is required to pull one away.

Step 3 — State the trend: the first ionization energy therefore increases from left to right across a period. (Down a *group* it decreases, because the outer electron is farther from the nucleus.)

Why each other option is wrong:

- (A) Decreasing is the trend *down a group*, not across a period.
- (B) Remaining constant ignores the steadily rising nuclear charge.
- (D) There is no overall "decrease then increase" across a period (small dips at group 13 and 16 do not reverse the broad upward trend).

Key point: Ionization energy increases across a period and decreases down a group, the exact opposite of the atomic-radius trend. Smaller, more tightly held atoms have higher ionization energies.

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

Answer: (C) Go Back to Q8

Q9.

Solution

Concept — Metallic character and atomic size: Metallic character is the tendency of an atom to lose electrons. It increases down a group and decreases across a period (left to right), so the most metallic element in a period is the one farthest to the left, which is also the largest atom (the atomic radius decreases across a period).

Step 1 — Place the options in period 3: sodium (Na) is group 1, magnesium (Mg) group 2, sulphur (S) group 16, and chlorine (Cl) group 17 (left to right in that order).

Step 2 — Apply the trend: the farther left, the more metallic and the larger the atom. Sodium sits farthest left of the four.

Step 3 — Conclusion: sodium is the most metallic and has the largest atomic size among these.

Why each other option is wrong:

- (B) Magnesium is metallic but lies to the right of sodium, so it is smaller and less metallic.
- (C) Sulphur is a non-metal, well to the right.
- (D) Chlorine is a non-metal at the far right, the smallest and least metallic of the four.

Key point: Across a period, metallic character and atomic size both fall from left to right. The leftmost element (here Na) is therefore the most metallic and the largest.

Final Answer: Sodium (Na) $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q9

Q10.

Solution

Concept — Endothermic versus exothermic: In an endothermic change the system *absorbs* heat from the surroundings, so ΔH is positive; in an exothermic change the system *releases* heat and ΔH is negative. Processes that break attractions or increase disorder by adding energy (such as melting) are endothermic.

Step 1 — Test melting of ice: converting solid ice to liquid water requires heat to overcome the hydrogen bonds holding the solid together. Heat is absorbed, so $\Delta H > 0$ (endothermic).

Step 2 — Compare the other choices: combustion of methane releases a great deal of heat (exothermic); neutralisation of an acid by a base releases heat (exothermic); condensation of steam to water releases heat as bonds form (exothermic).

Step 3 — Conclusion: only melting of ice absorbs heat, so it is the endothermic change.

Why each other option is wrong:

- (A) Combustion of methane is strongly exothermic (ΔH negative).
- (C) Acid–base neutralisation gives out heat (exothermic).
- (D) Condensation releases the latent heat of vaporisation (exothermic).

Key point: Melting, boiling, and evaporation *absorb* heat (endothermic, $\Delta H > 0$); freezing, condensation, combustion, and neutralisation *release* heat (exothermic, $\Delta H < 0$).

Final Answer: Melting of ice $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q10

Q11.

Solution

Concept — Gibbs free energy and spontaneity: The Gibbs free energy change, $\Delta G = \Delta H - T\Delta S$, is the criterion for spontaneity at constant temperature and pressure. A process is spontaneous (feasible without outside help) when $\Delta G < 0$, at equilibrium when $\Delta G = 0$, and non-spontaneous when $\Delta G > 0$.

Step 1 — State the rule: spontaneity requires the free energy of the system to decrease, i.e. ΔG must be negative.

Step 2 — Interpret the sign: a negative ΔG means the products are favoured and the reaction can proceed on its own.

Step 3 — Conclusion: the reaction is spontaneous when $\Delta G < 0$ (less than zero).

Why each other option is wrong:

- (A) $\Delta G = 0$ is the condition for *equilibrium*, not spontaneity.
- (B) $\Delta G > 0$ (positive) describes a non-spontaneous reaction.

- (C) $\Delta G = \Delta H$ holds only at absolute zero or when $\Delta S = 0$; it is not a spontaneity criterion.

Key point: Remember the sign convention: negative $\Delta G \Rightarrow$ spontaneous; zero $\Rightarrow$ equilibrium; positive $\Rightarrow$ non-spontaneous. Both ΔH and $T\Delta S$ feed into ΔG, but it is the *sign of* ΔG that decides feasibility.

Final Answer: Less than zero (negative) $\Rightarrow \boxed{\text{D}}$

Answer: (D) **Go Back to Q11**

Q12.

Solution

Concept — Le Chatelier's principle (temperature): If a system at equilibrium is disturbed, it shifts to partly oppose the change. For temperature, treat heat as a reactant or product. In an exothermic reaction heat is released (a "product"), so lowering the temperature removes heat and the equilibrium shifts in the direction that produces more heat, i.e. the forward (exothermic) direction.

Step 1 — Identify where heat appears: the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3 + \text{heat}$ gives out heat in the forward direction, so heat is effectively a product.

Step 2 — Apply the rule: lowering the temperature is like removing a product (heat); the system responds by making more heat, which means making more ammonia.

Step 3 — Conclusion: the equilibrium shifts toward the products (forward direction), increasing the yield of ammonia. This is why the Haber process favours ammonia at lower temperatures (though a moderate temperature is used for an acceptable rate).

Why each other option is wrong:

- (A) A shift toward the reactants would follow from *raising* the temperature of an exothermic reaction.
- (B) "No shift" would only apply if the reaction had $\Delta H = 0$, which is not the case here.
- (C) An equilibrium does not move "toward both sides equally."

Key point: For an exothermic reaction, cooling shifts the equilibrium *forward* (more product); heating shifts it backward. The opposite holds for endothermic reactions.

Final Answer: Towards the products (ammonia) ⇒ $\boxed{\text{D}}$

Answer: (D) Go Back to Q12

Q13.

Solution

Concept — pH of a strong acid: pH is defined by $\text{pH} = -\log_{10}[\text{H}^+]$, where $[\text{H}^+]$ is the molar concentration of hydrogen ions. HCl is a strong acid, so it ionises completely ($\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$) and gives one H^+ per molecule; thus $[\text{H}^+]$ equals the acid molarity.

Given: $[\text{HCl}] = 0.001\ \text{M} = 10^{-3}\ \text{M}$.

Step 1 — Find $[\text{H}^+]$: complete ionisation gives $[\text{H}^+] = 10^{-3}\ \text{M}$.

Step 2 — Apply the pH formula: $\text{pH} = -\log_{10}(10^{-3})$.

Step 3 — Evaluate: $\log_{10}(10^{-3}) = -3$, so $\text{pH} = -(-3) = 3$.

Why each other option is wrong:

- (A) $\text{pH} = 1$ corresponds to $[\text{H}^+] = 0.1$ M, a hundred times more concentrated.
- (B) $\text{pH} = 2$ corresponds to 0.01 M, ten times more concentrated.
- (C) $\text{pH} = 11$ is the pOH of this solution ($\text{pH} + \text{pOH} = 14$) and describes a basic solution, not this acid.

Key point: For a strong monoprotic acid, $\text{pH} = -\log(\text{molarity})$. The scale shows pH 3 in the acidic (red) region, below 7, as expected for an acid.

Final Answer: $\text{pH} = 3 \Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q13

Q14.

Solution

Concept — Conjugate acid–base pairs: In the Bronsted–Lowry theory, an acid donates a proton (H^+) and a base accepts one. When a base *gains* a proton, the result is its *conjugate acid*. So the conjugate acid of ammonia is found by adding one H^+ to NH_3.

Step 1 — Add one proton to ammonia: $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$.

Step 2 — Identify the product: the ammonium ion, NH_4^+, carries a $+1$ charge from the extra proton.

Step 3 — Conclusion: NH_4^+ is the conjugate acid of ammonia.

Why each other option is wrong:

- (A) NH_2^- (the amide ion) is the conjugate *base* of ammonia, formed by removing an H^+.
- (B) N^{3-} (nitride) would require removing all three hydrogens, which is not a conjugate-acid step.
- (D) NH_3 is the original species, not its conjugate acid.

Key point: Conjugate acid = original + one H^+; conjugate base = original − one H^+. Ammonia gains a proton to give ammonium (NH_4^+).

Final Answer: $NH_4^+ \Rightarrow \boxed{\text{C}}$

Answer: (C) **Go Back to Q14**

Q15.

Solution

Concept — Oxidation number rules: In a neutral compound the oxidation numbers of all atoms sum to zero. Hydrogen (bonded to a non-metal) is $+1$ and oxygen is usually -2. The oxidation number of the remaining element is then found algebraically.

Given: H_2SO_4, a neutral compound.

Step 1 — Assign the known values: two hydrogens give $2 \times (+1) = +2$; four oxygens give $4 \times (-2) = -8$. Let the oxidation number of sulphur be x.

Step 2 — Set the sum to zero: $(+2) + x + (-8) = 0$, i.e. $2 + x - 8 = 0$.

Step 3 — Solve: $x - 6 = 0 \Rightarrow x = +6$.

Why each other option is wrong:

- (A) $+2$ would not balance the -8 from the four oxygens.
- (C) $+4$ is the oxidation number of S in SO_2 or sulphurous acid (H_2SO_3), not in H_2SO_4.
- (D) -2 is S in sulphides such as H_2S, the opposite extreme.

Key point: Set the oxidation numbers to sum to zero for a neutral molecule.

Sulphur reaches its highest oxidation state, $+6$, in sulphuric acid and the sulphate ion (SO_4^{2-}).

Final Answer: Oxidation number of S $= +6 \Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q15

Q16.

Solution

Concept — Electrodes in a galvanic cell: In any electrochemical cell, oxidation always occurs at the *anode* and reduction always occurs at the *cathode*. A memory aid is "red-cat" (reduction at the cathode) and "an-ox" (oxidation at the anode). In a galvanic cell the cathode is the positive electrode.

Step 1 — Recall the definitions: reduction is the gain of electrons; the electrode where electrons are gained by the species in solution is the cathode.

Step 2 — Match to the question: the question asks where reduction takes place, which is by definition the cathode.

Step 3 — Conclusion: reduction occurs at the cathode.

Why each other option is wrong:

- (B) The anode is where *oxidation* (loss of electrons) occurs, the opposite of reduction.
- (C) The salt bridge completes the circuit and keeps the solutions neutral; no electrode reaction occurs there.
- (D) The electrolyte is the conducting solution, not an electrode.

Key point: "Red-Cat, An-Ox": Reduction → Cathode, Oxidation → Anode. This holds for both galvanic and electrolytic cells, even though the sign of each electrode differs between the two.

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

Answer: (A) Go Back to Q16

Q17.

Solution

Concept — Faraday's first law: The charge needed to deposit one mole of a metal at an electrode is $Q = n \times F$, where n is the number of moles of electrons required per mole of metal and $F = 96500$ C mol^{-1} is the Faraday constant. The value of n comes from the charge on the metal ion.

Given: $\mathrm{Cu}^{2+} + 2e^- \rightarrow \mathrm{Cu}$, so $n = 2$ moles of electrons per mole of Cu; $F = 96500$ C mol^{-1}.

Step 1 — Write the formula: $Q = n \times F$.

Step 2 — Substitute: $Q = 2 \times 96500$ C.

Step 3 — Evaluate: $Q = 193000$ C.

Why each other option is wrong:

- (A) 96500 C is the charge for *one* mole of electrons, which deposits only 0.5 mol of Cu.
- (B) 48250 C is half a Faraday, enough for just 0.25 mol of Cu.
- (D) 9650 C is one-tenth of a Faraday, far too small.

Key point: For a $+2$ ion you need *two* faradays per mole of metal. Contrast silver (Ag^+, one electron, 96500 C) with copper (Cu^{2+}, two electrons, 193000 C).

Final Answer: 193000 C $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q17

Q18.

Solution

Concept — Molarity: Molarity (M) is the number of moles of solute per litre of solution: $M = \dfrac{\text{moles of solute}}{\text{volume of solution in litres}}$. The volume must be converted from millilitres to litres before substituting.

Given: moles of solute $= 1$ mol; volume $= 500$ mL $= 0.5$ L.

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

Step 2 — Convert the volume: $500 \text{ mL} = \dfrac{500}{1000} \text{ L} = 0.5 \text{ L}$.

Step 3 — Substitute and evaluate: $M = \dfrac{1 \text{ mol}}{0.5 \text{ L}} = 2 \text{ mol L}^{-1} = 2 \text{ M}$.

Why each other option is wrong:

- (A) 1 M would need 1 mol in a full litre, not in 500 mL.
- (C) 0.5 M comes from dividing the moles by the volume in the wrong sense ($0.5 \div 1$ instead of $1 \div 0.5$).
- (D) 4 M would require 1 mol in 250 mL.

Key point: Always convert mL to L first. Dissolving the same moles in a smaller volume gives a *higher* molarity, so 1 mol in half a litre is 2 M.

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

Answer: (B) Go Back to Q18

Q19.

Solution

Concept — Charles's law: At constant pressure, the volume of a fixed mass of an ideal gas is directly proportional to its absolute (Kelvin) temperature: $V \propto T$, or $\frac{V_1}{T_1} = \frac{V_2}{T_2}$. The straight line through the origin in the graph shows this direct proportionality, so doubling T doubles V.

Given: pressure constant; new temperature $T_2 = 2T_1$.

Step 1 — Write Charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$.

Step 2 — Substitute $T_2 = 2T_1$: $V_2 = V_1 \times \frac{T_2}{T_1} = V_1 \times \frac{2T_1}{T_1}$.

Step 3 — Evaluate: $V_2 = 2V_1$, so the volume is doubled.

Why each other option is wrong:

- (B) Halved would happen only if the temperature were halved, not doubled.
- (C) Unchanged would require constant temperature (which is not the case).
- (D) Four times would need the temperature to be quadrupled, since $V \propto T$ (not T^2).

Key point: Charles's law uses *absolute* (Kelvin) temperature; the volume is directly proportional to T, so doubling T exactly doubles V. Always convert any Celsius temperature to Kelvin before applying the law.

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

Answer: (A) Go Back to Q19

Q20.

Solution

Concept — Oxidation state in a complex ion: The overall charge of a complex ion equals the oxidation state of the central metal plus the sum of the ligand charges. Rearranging, the metal's oxidation state = (overall charge) − (total ligand charge). Neutral ligands contribute zero charge.

Given: $[Cu(NH_3)_4]^{2+}$; each NH_3 is neutral (charge 0); overall charge $= +2$. Let the oxidation state of Cu be x.

Step 1 — Set up the charge balance: $x + 4 \times (0) = +2$.

Step 2 — Simplify: $x + 0 = +2$.

Step 3 — Solve: $x = +2$.

Why each other option is wrong:

- (A) 0 would make the whole ion neutral, contradicting the $2+$ charge.
- (B) $+1$ would give an overall charge of $+1$, not $+2$.
- (C) $+3$ would give an overall charge of $+3$, too high.

Key point: With neutral ligands like NH_3 or H_2O, the metal's oxidation state simply equals the charge written on the complex. Watch for charged ligands (e.g. CN^-, Cl^-), which change the balance.

Final Answer: Oxidation state of Cu $= +2 \Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q20

Q21.

Solution

Concept — Coordination number: The coordination number of the central metal in a complex is the number of ligand atoms directly bonded (donating lone pairs) to it. Each CO ligand is monodentate, bonding to the metal through one carbon atom, so each CO contributes one to the coordination number.

Step 1 — Count the ligands: $[Ni(CO)_4]$ has four CO ligands around the nickel.

Step 2 — Count the donor atoms: each CO is monodentate (one donor atom), so the donor atoms number $4 \times 1 = 4$.

Step 3 — Conclusion: the coordination number of nickel is 4, and the geometry is tetrahedral.

Why each other option is wrong:

- (A) 2 would describe a linear two-coordinate complex such as $[Ag(NH_3)_2]^+$.
- (B) 6 is an octahedral coordination number, as in $[Co(NH_3)_6]^{3+}$, not here.
- (D) 3 does not match the four CO ligands shown.

Key point: Coordination number $=$ number of donor atoms bonded to the metal, not the charge of the complex. Four monodentate ligands give a coordination number of 4 (tetrahedral for $Ni(CO)_4$).

Final Answer: $4 \Rightarrow \boxed{C}$

Answer: (C) Go Back to Q21

Q22.

Solution

Concept — Alkali metals: The alkali metals are the elements of group 1 of the periodic table (lithium, sodium, potassium, rubidium, caesium, francium). Each has a single electron in its outermost shell, which it loses easily to form a $+1$ ion, making these metals highly reactive.

Step 1 — Recall the group positions: potassium (K) is in group 1; calcium (Ca) is in group 2 (an alkaline-earth metal); aluminium (Al) is in group 13; iron (Fe) is a transition metal (group 8).

Step 2 — Match the definition: only potassium lies in group 1 with one valence electron.

Step 3 — Conclusion: potassium is the alkali metal.

Why each other option is wrong:

- (A) Calcium is an *alkaline-earth* metal (group 2), with two valence electrons.
- (C) Aluminium is a group-13 metal with three valence electrons.
- (D) Iron is a transition metal, not an alkali metal.

Key point: Alkali metals $=$ group 1 (one valence electron, $+1$ ions). Do not confuse them with the alkaline-earth metals of group 2. Among the options only potassium qualifies.

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

Answer: (B) Go Back to Q22

Q23.

Solution

Concept — Structural (chain) isomers: Structural isomers are compounds with the same molecular formula but different arrangements of atoms. For alkanes, chain isomers differ in how the carbon skeleton is branched. Pentane, C_5H_{12}, can be arranged in three distinct ways.

Step 1 — Draw the straight chain: n-pentane, $CH_3–CH_2–CH_2–CH_2–CH_3$ (one structure).

Step 2 — Make one branch: isopentane (2-methylbutane), a four-carbon chain with a methyl branch on C-2 (second structure).

Step 3 — Make two branches on one carbon: neopentane (2,2-dimethylpropane), a three-carbon chain with two methyl groups on the central carbon (third structure). No further distinct skeletons exist, so there are 3 isomers.

Why each other option is wrong:

- (A) 2 misses the doubly branched neopentane.
- (C) 4 over-counts; rearrangements beyond the three above just repeat the same skeletons.
- (D) 5 is far too many for C_5H_{12}.

Key point: The isomer counts to memorise are: butane (C_4H_{10}) has 2, pentane (C_5H_{12}) has 3, and hexane (C_6H_{14}) has 5. Pentane's three are n-pentane, isopentane, and neopentane.

Final Answer: 3 isomers $\Rightarrow \boxed{\text{B}}$

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

Q24.

Solution

Concept — IUPAC naming of aldehydes: An aldehyde is named by taking the longest carbon chain that includes the $-$CHO carbon and replacing the final "-e" of the parent alkane with the suffix "-al". The compound $CH_3–CHO$ has two carbons, so the parent is ethane.

Step 1 — Count the carbons: $CH_3–CHO$ has two carbon atoms, giving the parent "ethane".

Step 2 — Identify the functional group: the $-$CHO group makes it an aldehyde, so the suffix is "-al".

Step 3 — Build the name: ethane $\rightarrow$ ethan- $+$ -al $=$ ethanal.

Why each other option is wrong:

- (B) Methanal (HCHO) has only one carbon, but the compound here has two.
- (C) Ethanol is the *alcohol* CH_3CH_2OH (the "-ol" suffix is for an $-$OH group, not an aldehyde).
- (D) Ethanoic acid is the *carboxylic acid* CH_3COOH ($-$COOH group), not the aldehyde.

Key point: The "-al" suffix marks an aldehyde, "-ol" an alcohol, and "-oic acid" a carboxylic acid. Two carbons plus $-$CHO gives ethanal (common name acetaldehyde).

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

Answer: (A) **Go Back to Q24**

Q25.

Solution

Concept — General formula of alkanes: A homologous series is a family of compounds with the same general formula, each member differing from the next by a CH_2 unit. Alkanes are saturated hydrocarbons (only single bonds), and their general formula is C_nH_{2n+2}, where n is the number of carbon atoms.

Step 1 — Recall the saturated structure: each carbon forms four single bonds; a straight chain of n carbons leaves the maximum number of hydrogens.

Step 2 — Check with examples: methane $n = 1 \rightarrow C_1H_4$; ethane $n = 2 \rightarrow C_2H_6$; propane $n = 3 \rightarrow C_3H_8$. Each fits C_nH_{2n+2}.

Step 3 — Conclusion: the general formula of the alkane series is C_nH_{2n+2}.

Why each other option is wrong:

- (A) C_nH_{2n} is the formula of *alkenes* (one double bond) or cyclic alkanes.
- (B) C_nH_{2n-2} is the formula of *alkynes* (one triple bond).
- (C) C_nH_n does not correspond to any common saturated series and fails the methane check.

Key point: Alkanes C_nH_{2n+2}, alkenes C_nH_{2n}, alkynes C_nH_{2n-2}. Each degree of unsaturation (double or triple bond, or ring) removes two hydrogens.

Final Answer: $C_nH_{2n+2} \Rightarrow \boxed{D}$

Answer: (D) Go Back to Q25

Q26.

Solution

Concept — Saturated hydrocarbons: A saturated hydrocarbon contains only single carbon–carbon bonds, so every carbon holds the maximum number of hydrogens. These are the alkanes (C_nH_{2n+2}). Hydrocarbons with double bonds (alkenes), triple bonds (alkynes), or aromatic rings are unsaturated.

Step 1 — Classify each option: ethene (C_2H_4) has a C=C double bond (unsaturated); ethyne (C_2H_2) has a C≡C triple bond (unsaturated); benzene (C_6H_6) has a ring of delocalised electrons (aromatic, unsaturated).

Step 2 — Examine ethane: ethane (C_2H_6) has only a single C–C bond and follows C_nH_{2n+2} (with $n = 2$), so it is saturated.

Step 3 — Conclusion: ethane is the saturated hydrocarbon.

Why each other option is wrong:

- (A) Ethene has a double bond, so it is unsaturated.
- (B) Ethyne has a triple bond, so it is unsaturated.
- (D) Benzene's aromatic ring counts as unsaturated (it readily undergoes substitution but is not saturated).

Key point: "Saturated" means only single bonds (alkanes). The "-ane" ending signals saturation, while "-ene" (double bond) and "-yne" (triple bond) signal unsaturation.

Final Answer: Ethane (C_2H_6) $\Rightarrow \boxed{C}$

Answer: (C) Go Back to Q26

Q27.

Solution

Concept — Bonding in methane: Carbon has four valence electrons and forms four covalent bonds. In methane, CH_4, each of carbon's four bonds goes to a separate hydrogen atom, so the molecule contains four identical C–H single bonds arranged tetrahedrally.

Step 1 — Count carbon's bonds: carbon needs four bonds to complete its octet, and all four are used to bond hydrogens.

Step 2 — Match to hydrogens: there are four hydrogen atoms in CH_4, each joined to the central carbon by one single bond.

Step 3 — Conclusion: the methane molecule contains 4 C–H bonds.

Why each other option is wrong:

- (B) 3 would leave one of carbon's four valences unused, which cannot happen in stable CH_4.
- (C) 2 matches the formula CH_2 (a reactive carbene), not methane.
- (D) 5 exceeds carbon's four available bonds, which is impossible.

Key point: Carbon is tetravalent: it forms exactly four bonds. In CH_4 these are four C–H single bonds, giving a symmetric tetrahedral molecule.

Final Answer: 4 C–H bonds $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q27

Q28.

Solution

Concept — The hydroxyl group: A functional group is the reactive part of a molecule that defines its class. The group shown, an oxygen bonded to a hydrogen and joined to a carbon (written $-$OH, the hydroxyl group), is the defining group of *alcohols*.

Step 1 — Read the structure: the carbon chain (R) is attached to an oxygen, which is in turn attached to a hydrogen, i.e. R–OH.

Step 2 — Match the group to its class: a carbon chain carrying an $-$OH group is an alcohol, for example ethanol (CH_3CH_2OH) or methanol (CH_3OH).

Step 3 — Conclusion: the $-$OH group is characteristic of alcohols.

Why each other option is wrong:

- (A) Carboxylic acids have $-$COOH (a carbonyl *and* a hydroxyl on the same carbon), not a lone $-$OH.
- (C) Aldehydes have the $-$CHO group (a carbonyl with a hydrogen), not a hydroxyl.
- (D) Amines have the $-NH_2$ group, which contains nitrogen, not oxygen.

Key point: A lone $-$OH on a carbon chain means an alcohol. When the $-$OH sits on a carbonyl carbon it becomes part of $-$COOH (a carboxylic acid), which behaves very differently.

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

Answer: (B)

Go Back to Q28

Q29.

Solution

Concept — Monomers of carbohydrates: Large biomolecules are polymers built from small repeating monomer units. Polysaccharides (complex carbohydrates such as starch, glycogen, and cellulose) are polymers of monosaccharides, the simple sugars, of which glucose is the most common.

Step 1 — Identify the monomer of polysaccharides: the building blocks are monosaccharides (single sugar units), e.g. glucose.

Step 2 — How they link: many monosaccharide units join through glycosidic bonds, each linkage releasing a molecule of water (condensation), to form a long polysaccharide chain.

Step 3 — Conclusion: polysaccharides are made of monosaccharide (glucose) monomers.

Why each other option is wrong:

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

Key point: Match each biomolecule to its monomer: polysaccharides $\leftarrow$ monosaccharides (glucose), proteins $\leftarrow$ amino acids, nucleic acids $\leftarrow$ nucleotides, fats $\leftarrow$ fatty acids $+$ glycerol. Carbohydrates are built from simple sugars.

Final Answer: Monosaccharides (glucose) $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q29

Q30.

Solution

Concept — The acidic functional group: An organic compound is acidic if it can donate a proton (H^+). The carboxyl group, $-COOH$, is a carbonyl ($C=O$) and a hydroxyl ($-OH$) on the *same* carbon. Its O–H bond readily releases an H^+, and the resulting carboxylate ion ($-COO^-$) is stabilised by resonance, which is why carboxylic acids are the standard acidic organic compounds.

Step 1 — Recall the carboxyl group: $-COOH$, as in acetic acid (CH_3COOH) and formic acid ($HCOOH$).

Step 2 — Confirm its acidity: the $-OH$ of the carboxyl donates a proton to give a resonance-stabilised carboxylate ion, making the compound a weak acid.

Step 3 — Conclusion: the $-COOH$ group makes a compound acidic.

Why each other option is wrong:

- (A) $-OH$ (hydroxyl, in alcohols) holds its proton tightly and is essentially neutral, not acidic.
- (B) $-CHO$ (aldehyde) is a carbonyl group; it is not acidic in the proton-donating sense.
- (C) $-NH_2$ (amino, in amines) *accepts* a proton, so it is basic, the opposite of acidic.

Key point: $-COOH$ = carbonyl + hydroxyl on one carbon = acidic (proton donor). Compare neutral alcohols ($-OH$) and basic amines ($-NH_2$); only the carboxyl group is reliably acidic.

Final Answer: $-COOH$ (carboxyl) $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q30

Answer Key

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