

AIIMS B.Sc Nursing Chemistry

Sample Paper – 9

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. Glucose has the molecular formula $C_6H_{12}O_6$. Using atomic masses $C = 12$, $H = 1$, $O = 16$, the molar mass of glucose is:

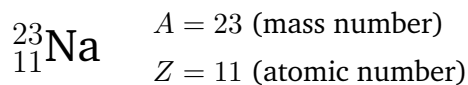
- (A) 90 g mol^{-1}
- (B) 162 g mol^{-1}
- (C) 180 g mol^{-1}
- (D) 342 g mol^{-1}

Q2. The number of moles present in 9 g of water (H_2O , molar mass = 18 g mol^{-1}) is:

- (A) 0.5 mol
- (B) 1 mol
- (C) 2 mol
- (D) 0.25 mol

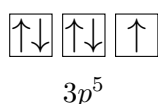


Q3. A sodium atom is represented by the nuclide symbol shown below. The number of neutrons in this atom (neutrons = $A - Z$) is:



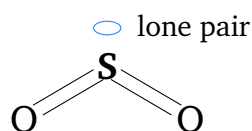
- (A) 11
- (B) 23
- (C) 34
- (D) 12

Q4. The filling of the $3p$ subshell of a chlorine atom ($Z = 17$, configuration ending in $3p^5$) is shown below. The number of unpaired electrons in a ground-state chlorine atom is:



- (A) 0
- (B) 1
- (C) 2
- (D) 3

Q5. The sulfur atom in sulfur dioxide (SO_2) carries one lone pair, as shown. The shape of the SO_2 molecule is:



- (A) linear
- (B) trigonal planar
- (C) bent (angular)



(D) trigonal pyramidal

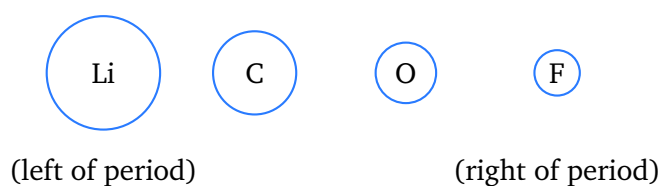
Q6. Each carbon atom in diamond is bonded to four other carbon atoms by single bonds in a tetrahedral arrangement. The hybridization of such a saturated carbon atom is:

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

Q7. Among the hydrogen halides below, the bond with the greatest polarity (largest electronegativity difference between the two atoms) is:

- (A) H–I
- (B) H–Br
- (C) H–Cl
- (D) H–F

Q8. The relative sizes of four period-2 atoms are sketched below (noble gas excluded). Which atom has the smallest atomic radius?



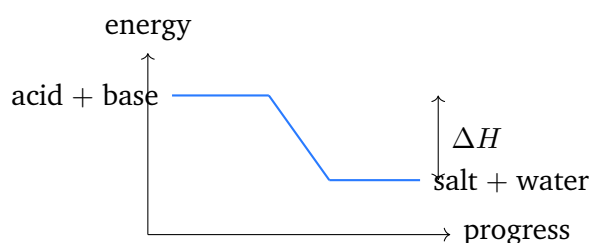
- (A) Lithium (Li)
- (B) Fluorine (F)
- (C) Carbon (C)
- (D) Oxygen (O)

Q9. Excluding helium, every noble gas (group 18) has a completely filled outermost shell. The number of valence electrons in such a noble gas atom is:



- (A) 2
- (B) 4
- (C) 8
- (D) 18

Q10. The energy diagram for the neutralization of a strong acid by a strong base is shown, with the products lower in energy than the reactants. The enthalpy of neutralization of a strong acid and a strong base is always:



- (A) negative (heat released)
 - (B) positive (heat absorbed)
 - (C) zero
 - (D) sometimes positive, sometimes negative
- Q11.** According to the Gibbs free-energy criterion ($\Delta G = \Delta H - T\Delta S$), a chemical reaction at constant temperature and pressure is **non-spontaneous** when the value of ΔG is:
- (A) negative
 - (B) zero
 - (C) negative or zero
 - (D) positive
- Q12.** When a reversible reaction has reached a state of dynamic equilibrium, the rate of the forward reaction and the rate of the backward reaction are:
- (A) both zero



- (B) equal
- (C) forward greater than backward
- (D) backward greater than forward

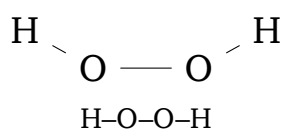
Q13. Hydrochloric acid is a strong acid that ionizes completely in water. The pH of a 0.1 M HCl solution ($\text{pH} = -\log_{10}[\text{H}^+]$) is:

- (A) 0
- (B) 2
- (C) 1
- (D) 13

Q14. Strong acids such as HCl, HNO₃, and H₂SO₄ differ from weak acids in that, when dissolved in water, they ionize:

- (A) almost completely
- (B) only to a very small extent
- (C) not at all
- (D) exactly halfway

Q15. The structure of hydrogen peroxide (H₂O₂), which contains an oxygen–oxygen single bond, is shown below. The oxidation number of each oxygen atom in H₂O₂ is:



- (A) –2
- (B) 0
- (C) +1
- (D) –1



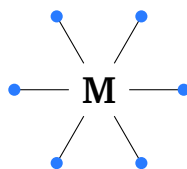
- Q16.** In a galvanic (voltaic) cell, the two half-cells are connected by a salt bridge. The main function of the salt bridge is to:
- (A) supply electrons to the external wire
 - (B) complete the circuit and maintain electrical neutrality of the half-cells
 - (C) act as the oxidising agent
 - (D) increase the temperature of the cell
- Q17.** Faraday's first law of electrolysis states that the mass of a substance deposited (or liberated) at an electrode during electrolysis is directly proportional to the:
- (A) temperature of the electrolyte
 - (B) volume of the electrolyte
 - (C) quantity of charge passed through the cell
 - (D) area of the electrode
- Q18.** The number of moles of solute present in 2 L of a 0.5 M solution (moles = molarity \times volume in litres) is:
- (A) 1 mol
 - (B) 0.25 mol
 - (C) 2.5 mol
 - (D) 4 mol
- Q19.** According to the kinetic theory of gases, when the temperature of a fixed sample of gas is raised, the average speed of its molecules:
- (A) decreases
 - (B) remains unchanged
 - (C) first decreases, then increases
 - (D) increases



Q20. In the complex ion $[\text{Zn}(\text{OH})_4]^{2-}$, each hydroxide ligand (OH) carries a charge of -1 . The oxidation state of zinc in this complex is:

- (A) +4
- (B) +2
- (C) 0
- (D) -2

Q21. A hexadentate ligand wraps around a central metal ion (M) as shown, binding through several donor atoms (shown as dots). The number of donor atoms through which a hexadentate ligand binds to the metal is:



one ligand, six donor atoms

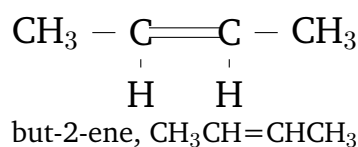
- (A) 2
- (B) 4
- (C) 6
- (D) 1

Q22. An amphoteric oxide reacts with both acids and bases. Which one of the following oxides is amphoteric?

- (A) Al_2O_3
- (B) Na_2O
- (C) CO_2
- (D) MgO

Q23. Geometrical (cis-trans) isomerism arises from restricted rotation about a $\text{C}=\text{C}$ double bond when each double-bonded carbon carries two different groups. Which compound, sketched below, shows cis-trans isomerism?





- (A) ethene ($\text{CH}_2=\text{CH}_2$)
- (B) propene ($\text{CH}_3\text{CH}=\text{CH}_2$)
- (C) but-1-ene ($\text{CH}_2=\text{CHCH}_2\text{CH}_3$)
- (D) but-2-ene ($\text{CH}_3\text{CH}=\text{CHCH}_3$)

Q24. The IUPAC name of the compound $\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{OH}$ is:

- (A) ethanol
- (B) propan-1-ol
- (C) propan-2-ol
- (D) propanal

Q25. A saturated hydrocarbon (an alkane) is one in which the carbon atoms are joined using only:

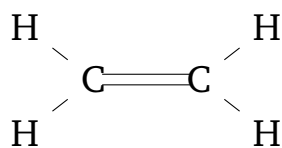
- (A) double bonds
- (B) triple bonds
- (C) single bonds
- (D) alternate single and double bonds

Q26. Because of the electron-rich $\text{C}=\text{C}$ double bond, the reaction most characteristic of alkenes (for example the reaction of ethene with bromine) is:

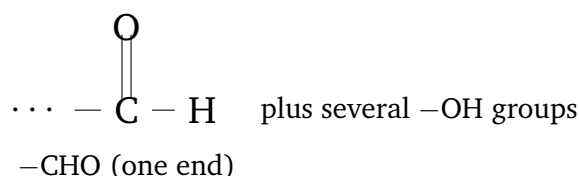
- (A) addition
- (B) substitution
- (C) neutralization
- (D) precipitation



Q27. The planar structure of ethene (C_2H_4), with a $C=C$ double bond and $H-C-H$ angles of about 120° , is shown. The hybridization of each carbon atom in ethene is:



- (A) sp^3
 (B) sp
 (C) sp^3d
 (D) sp^2
- Q28.** The mild (controlled) oxidation of a primary alcohol, such as ethanol (CH_3CH_2OH), first gives a compound containing the $-CHO$ group. This product is:
- (A) a ketone
 (B) an aldehyde
 (C) an ether
 (D) an ester
- Q29.** The open-chain structure of glucose is shown in part below. Glucose (an aldohexose) contains which two functional groups?



- (A) ketone and hydroxyl
 (B) carboxylic acid and amino
 (C) aldehyde and hydroxyl
 (D) aldehyde and carboxylic acid



- Q30.** An amine contains the $-\text{NH}_2$ functional group, in which the nitrogen atom carries a lone pair of electrons. The general chemical character of an amine is that it acts as a:
- (A) base
 - (B) strong acid
 - (C) neutral, unreactive group
 - (D) powerful oxidising agent



Detailed Solutions

Q1.

Solution

Concept — Molar mass from a molecular formula: The molar mass of a compound is the sum of the atomic masses of all the atoms in one formula unit, expressed in grams per mole. You read the subscripts in the formula, multiply each element's atomic mass by the number of its atoms, and add the contributions. For glucose, $C_6H_{12}O_6$, there are 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.

Given: formula $C_6H_{12}O_6$; atomic masses C = 12, H = 1, O = 16 $g\ mol^{-1}$.

Step 1 — Carbon contribution: $6 \times 12 = 72\ g\ mol^{-1}$.

Step 2 — Hydrogen contribution: $12 \times 1 = 12\ g\ mol^{-1}$.

Step 3 — Oxygen contribution and total: $6 \times 16 = 96\ g\ mol^{-1}$; adding all parts, $72 + 12 + 96 = 180\ g\ mol^{-1}$.

Why each other option is wrong:

- (A) 90 is exactly half of the true value, the result of halving every subscript.
- (B) 162 is the molar mass of an *anhydroglucose* unit ($C_6H_{10}O_5$) found in starch, not free glucose.
- (D) 342 is the molar mass of sucrose ($C_{12}H_{22}O_{11}$), a disaccharide, not glucose.

Key point: Always multiply each atomic mass by its subscript before adding. For glucose the tidy result is $180\ g\ mol^{-1}$, a value worth remembering.

Final Answer: $180\ g\ mol^{-1} \Rightarrow$ C

Answer: (C) [Go Back to Q1](#)

Q2.

Solution

Concept — Moles from mass: The number of moles of a substance links its measured mass to a count of particles through the relation $n = \frac{m}{M}$, where m is the given mass and M is the molar mass. One mole of water weighs 18 g, because $M(H_2O) = 2 \times 1 + 16 = 18\ g\ mol^{-1}$.

Given: mass $m = 9\ g$; molar mass $M(H_2O) = 18\ g\ mol^{-1}$.



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

Step 2 — Substitute the data with units: $n = \frac{9 \text{ g}}{18 \text{ g mol}^{-1}}$.

Step 3 — Evaluate: the grams cancel, leaving $n = 0.5 \text{ mol}$.

Why each other option is wrong:

- (B) 1 mol would need the full 18 g of water, not 9 g.
- (C) 2 mol would correspond to 36 g.
- (D) 0.25 mol would be only 4.5 g.

Key point: 9 g is half of the molar mass 18 g mol^{-1} , so the amount is half a mole. Recognising that 9 g is "half of 18" lets you read the answer almost by inspection.

Final Answer: $0.5 \text{ mol} \Rightarrow$

Answer: (A) [Go Back to Q2](#)

Q3.

Solution

Concept — Counting neutrons in a nuclide: A nuclide symbol is written ${}^A_Z\text{X}$, where Z (the atomic number) is the number of protons and A (the mass number) is the total number of protons plus neutrons. The number of neutrons is therefore found by subtraction: $\text{neutrons} = A - Z$.

Given: for ${}^{23}_{11}\text{Na}$, mass number $A = 23$, atomic number $Z = 11$.

Step 1 — Write the relation: $\text{neutrons} = A - Z$.

Step 2 — Substitute the values: $\text{neutrons} = 23 - 11$.

Step 3 — Evaluate: $= 12 \text{ neutrons}$.

Why each other option is wrong:

- (A) 11 is the number of *protons* (Z), not neutrons.
- (B) 23 is the mass number A (protons + neutrons together).
- (C) 34 wrongly *adds* A and Z instead of subtracting.

Key point: Protons = Z , neutrons = $A - Z$, and mass number $A = \text{protons} + \text{neutrons}$. For sodium-23 this gives 11 protons and 12 neutrons.

Final Answer: 12 neutrons \Rightarrow



Answer: (D) [Go Back to Q3](#)

Q4.

Solution

Concept — Unpaired electrons in a p^5 subshell: Electrons fill the three orbitals of a p subshell following Hund's rule (each orbital gets one electron before any pairing) and the Pauli principle (at most two electrons per orbital, with opposite spins). Chlorine ($Z = 17$) has the configuration $1s^2 2s^2 2p^6 3s^2 3p^5$, so the deciding subshell is the $3p^5$.

Step 1 — Distribute five electrons over three $3p$ orbitals: first place one in each of the three orbitals (three electrons), then the remaining two pair up two of those orbitals, giving $\uparrow\downarrow\uparrow\downarrow\uparrow$.

Step 2 — Identify the unpaired orbital: two orbitals are now full (paired), and exactly one orbital holds a single electron.

Step 3 — Count the unpaired electrons: there is 1 unpaired electron.

Why each other option is wrong:

- (A) 0 would require all five electrons to be paired, impossible with an odd number.
- (C) 2 and (D) 3 overcount: only one orbital is left half-filled in a p^5 system.

Key point: A p^5 configuration always has exactly one unpaired electron (the same as p^1). This single unpaired electron is why a chlorine atom is one electron short of the stable octet and forms Cl^- so readily.

Final Answer: 1 unpaired electron \Rightarrow

Answer: (B) [Go Back to Q4](#)

Q5.

Solution

Concept — VSEPR shape of SO_2 : The shape of a molecule is set by the number of electron domains (bonding regions plus lone pairs) around the central atom; lone pairs then distort the visible shape. In SO_2 the central sulfur has two bonding regions (to the two oxygens) and one lone pair, giving three electron domains.

Step 1 — Electron geometry: three electron domains arrange themselves in a trigonal-planar pattern at about 120° .



Step 2 — Account for the lone pair: one of the three positions is occupied by the lone pair, which is not "seen" as an atom but still pushes the two S–O bonds together.

Step 3 — Molecular shape: with only two atoms attached and one lone pair on the central atom, the molecule is bent (angular), with an O–S–O angle of about 119° .

Why each other option is wrong:

- (A) Linear needs two domains and no lone pair on the central atom, as in CO_2 .
- (B) Trigonal planar describes the *electron* geometry; the actual shape is bent because one domain is a lone pair.
- (D) Trigonal pyramidal needs four electron domains (three bonds + one lone pair), as in NH_3 .

Key point: Two bonds plus one lone pair on the central atom always give a bent shape (like SO_2 , O_3 , and H_2O). The lone pair on sulfur is what bends the molecule away from a straight line, making SO_2 polar.

Final Answer: Bent (angular) \Rightarrow C

Answer: (C) [Go Back to Q5](#)

Q6.

Solution

Concept — Hybridisation of a saturated carbon: The hybridisation of a carbon atom equals its number of σ bonds plus lone pairs, that is, its number of electron domains. A saturated carbon forms four single (σ) bonds and has no lone pairs, so it has four electron domains. Four domains correspond to sp^3 hybridisation, in which one s and three p orbitals mix to give four equivalent orbitals pointing to the corners of a tetrahedron.

Step 1 — Count the bonds on a diamond carbon: each carbon is joined to four other carbon atoms by single bonds, i.e. four σ bonds and no lone pairs (four electron domains).

Step 2 — Assign the hybridisation: four electron domains $\Rightarrow sp^3$.

Step 3 — Geometry: the four sp^3 orbitals point toward the corners of a tetrahedron with bond angles of 109.5° , giving diamond its rigid three-dimensional network.



Why each other option is wrong:

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

Key point: A fully saturated carbon (four single bonds) is always sp^3 and tetrahedral, whether in diamond, methane, or any alkane.

Final Answer: $sp^3 \Rightarrow$

Answer: (A) [Go Back to Q6](#)

Q7.

Solution

Concept — Bond polarity and electronegativity: The polarity of a covalent bond depends on the difference in electronegativity between the two bonded atoms: the larger the difference, the more polar the bond. In the hydrogen halides H-X, hydrogen is common to all, so the polarity is decided by the halogen. Electronegativity decreases down group 17 ($F > Cl > Br > I$), so the electronegativity difference with hydrogen is largest for fluorine.

Step 1 — List approximate electronegativities of the halogens: $F \approx 4.0$, $Cl \approx 3.0$, $Br \approx 2.8$, $I \approx 2.5$ (hydrogen ≈ 2.1).

Step 2 — Compute the differences from hydrogen: $H-F \approx 1.9$, $H-Cl \approx 0.9$, $H-Br \approx 0.7$, $H-I \approx 0.4$.

Step 3 — Pick the largest: the biggest difference is for H-F, so the H-F bond is the most polar.

Why each other option is wrong:

- (A) H-I has the smallest electronegativity difference and is the least polar.
- (B) H-Br is only slightly more polar than H-I, still well below H-F.
- (C) H-Cl is more polar than H-Br and H-I, but its difference is about half that of H-F.

Key point: Polarity follows the electronegativity gap. Since fluorine is the most electronegative element, H-F is the most polar of the hydrogen halides, which is also why HF molecules hydrogen-bond strongly.



Final Answer: H-F \Rightarrow

Answer: (D) [Go Back to Q7](#)

Q8.

Solution

Concept — Atomic radius across a period: Across a period (left to right), electrons are added to the same outermost shell while protons are added to the nucleus. The rising effective nuclear charge pulls that shell inward, so the atomic radius decreases steadily across the period. Therefore the atom on the far right (just before the noble gas) is the smallest.

Step 1 — Place the options in period 2: the order left to right is Li, Be, B, C, N, O, F (then the noble gas Ne). Of the choices, Li is farthest left and F is farthest right.

Step 2 — Apply the trend: as we move from Li to F the nuclear charge rises from +3 to +9 while electrons stay in the second shell, so the radius shrinks $\text{Li} > \text{C} > \text{O} > \text{F}$.

Step 3 — Conclusion: fluorine, being the rightmost element listed (excluding the noble gas), has the smallest atomic radius.

Why each other option is wrong:

- (A) Lithium is the leftmost period-2 element, so it has the *largest* radius of the four.
- (C) Carbon lies in the middle of the period, larger than oxygen and fluorine.
- (D) Oxygen is smaller than carbon but still larger than fluorine, which sits one place to its right.

Key point: The atomic radius decreases across a period because of the rising effective nuclear charge. Excluding the noble gas, the last element of the period (here fluorine) is the smallest.

Final Answer: Fluorine (F) \Rightarrow

Answer: (B) [Go Back to Q8](#)



Q9.

Solution

Concept — Valence electrons of the noble gases: Valence electrons are the electrons in the outermost shell of an atom. The noble gases of group 18 are chemically inert because their outermost shell is completely filled. Apart from helium (which has only the $1s$ shell with 2 electrons), every other noble gas has the stable octet configuration $ns^2 np^6$, that is, eight electrons in the outer shell.

Step 1 — Recall the outer-shell configuration: for neon, argon, krypton, etc., the outermost shell is $ns^2 np^6$.

Step 2 — Count the outer electrons: ns^2 contributes 2 and np^6 contributes 6, giving $2 + 6 = 8$ valence electrons.

Step 3 — Conclusion: a noble gas (other than helium) has 8 valence electrons, a complete octet.

Why each other option is wrong:

- (A) 2 is the number for *helium* alone (its $1s^2$ shell), which is the exception excluded here.
- (B) 4 is the number of valence electrons in a group-14 element such as carbon, not a noble gas.
- (D) 18 is the *group number* of the noble gases, not their valence-electron count.

Key point: The stability of the noble gases comes from a filled outer shell. Except for helium (2 electrons), this means a complete octet of 8 valence electrons, the configuration other atoms aim for in bonding.

Final Answer: 8 valence electrons \Rightarrow C

Answer: (C) [Go Back to Q9](#)

Q10.

Solution

Concept — Enthalpy of neutralization: Neutralization is the reaction of an acid with a base to form a salt and water. For a strong acid and a strong base, both are fully ionized, so the only net chemical change is $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. This combination releases energy (it forms a stable O–H bond in water), so heat is given out to the surroundings.



Step 1 — Identify the net reaction: for any strong acid + strong base, the underlying process is always $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$, which releases about 57 kJ per mole of water formed.

Step 2 — Read the energy diagram: the products (salt + water) lie below the reactants, so the system loses energy as it goes from reactants to products.

Step 3 — Assign the sign: because heat is released, the reaction is exothermic and the enthalpy of neutralization is negative ($\Delta H \approx -57 \text{ kJ mol}^{-1}$).

Why each other option is wrong:

- (B) A positive ΔH would mean heat is absorbed (endothermic), which contradicts both the diagram and experiment.
- (C) Zero would mean no energy change, but neutralization clearly liberates heat.
- (D) "Sometimes positive" is wrong for strong acid + strong base, where the value is always close to -57 kJ mol^{-1} .

Key point: For a strong acid and a strong base the enthalpy of neutralization is always negative and nearly constant ($\approx -57 \text{ kJ mol}^{-1}$), because the same $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ reaction underlies them all.

Final Answer: Negative (heat released) \Rightarrow

[Go Back to Q10](#)

Q11.

Solution

Concept — Gibbs free energy and spontaneity: The sign of the Gibbs free-energy change $\Delta G = \Delta H - T\Delta S$ tells us whether a process can occur on its own at constant temperature and pressure. The rule is: $\Delta G < 0$ means the reaction is spontaneous (feasible), $\Delta G = 0$ means the system is at equilibrium, and $\Delta G > 0$ means the reaction is non-spontaneous (it will not proceed in the forward direction without outside help).

Step 1 — State the spontaneity criterion: a reaction is spontaneous only when ΔG is negative.

Step 2 — Identify the non-spontaneous case: the opposite of a negative value is a positive value, so a positive ΔG means the forward reaction is non-spontaneous.

Step 3 — Conclusion: a reaction is non-spontaneous when ΔG is positive.



Why each other option is wrong:

- (A) A negative ΔG marks a *spontaneous* reaction, the exact opposite.
- (B) $\Delta G = 0$ corresponds to a system already at equilibrium, not a non-spontaneous one.
- (C) "Negative or zero" describes feasible or equilibrium states, neither of which is non-spontaneous.

Key point: Remember the sign chart: $\Delta G < 0$ spontaneous, $\Delta G = 0$ equilibrium, $\Delta G > 0$ non-spontaneous. A positive ΔG flags a reaction that will not go forward by itself.

Final Answer: Positive \Rightarrow

Answer: (D) [Go Back to Q11](#)

Q12.

Solution

Concept — Dynamic equilibrium: A reversible reaction reaches equilibrium when the forward and backward reactions are still happening but at the same rate. This is called *dynamic* equilibrium: molecules keep converting in both directions, but because the two rates are equal, the concentrations of reactants and products stay constant with time.

Step 1 — Recall the definition: at equilibrium, rate of forward reaction = rate of backward reaction.

Step 2 — Note that the rates are not zero: both reactions continue actively; only their *net* effect is zero, which keeps the amounts of each species unchanged.

Step 3 — Conclusion: at equilibrium the two rates are equal.

Why each other option is wrong:

- (A) "Both zero" describes a reaction that has stopped, not a dynamic equilibrium where both directions are still active.
- (C) If the forward rate were greater, the reaction would still be moving forward, i.e. not yet at equilibrium.
- (D) If the backward rate were greater, the reaction would be moving in reverse, again not at equilibrium.

Key point: Equilibrium is "dynamic, not dead": the forward and backward rates are equal and non-zero, so concentrations remain constant even though molecules



are continually interconverting.

Final Answer: Equal \Rightarrow B

Answer: (B) [Go Back to Q12](#)

Q13.

Solution

Concept — pH of a strong acid: The pH measures acidity through $\text{pH} = -\log_{10}[\text{H}^+]$, where $[\text{H}^+]$ is the molar concentration of hydrogen ions. HCl is a strong, monoprotic acid, so it ionizes completely in water ($\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$) and provides one H^+ per molecule. The hydrogen-ion concentration therefore equals the acid concentration.

Given: $[\text{HCl}] = 0.1 \text{ M} = 10^{-1} \text{ M}$.

Step 1 — Find $[\text{H}^+]$: because HCl ionizes fully, $[\text{H}^+] = 10^{-1} \text{ M}$.

Step 2 — Apply the pH formula: $\text{pH} = -\log_{10}(10^{-1})$.

Step 3 — Evaluate: $\log_{10}(10^{-1}) = -1$, so $\text{pH} = -(-1) = 1$.

Why each other option is wrong:

- (A) $\text{pH} = 0$ corresponds to $[\text{H}^+] = 1 \text{ M}$, ten times more concentrated.
- (B) $\text{pH} = 2$ corresponds to 0.01 M , ten times more dilute.
- (D) $\text{pH} = 13$ is the pOH of this solution and would describe a strongly basic solution, not this acid.

Key point: For a strong monoprotic acid, $[\text{H}^+]$ equals the molarity, so the pH is just $-\log$ of the concentration. A pH of 1 is strongly acidic, as expected for 0.1 M HCl.

Final Answer: $\text{pH} = 1 \Rightarrow$ C

Answer: (C) [Go Back to Q13](#)



Q14.

Solution

Concept — Strong versus weak acids: The strength of an acid refers to the extent to which it ionizes (dissociates into ions) in water, not to how concentrated it is. A *strong* acid ionizes almost completely, so nearly every acid molecule donates its proton to water. A *weak* acid ionizes only partially, leaving most molecules undissociated in an equilibrium with their ions.

Step 1 — Recall the definition of a strong acid: essentially all of the acid converts to H^+ (or H_3O^+) and the corresponding anion, e.g. $HCl \rightarrow H^+ + Cl^-$ goes effectively to completion.

Step 2 — Compare with a weak acid: a weak acid such as acetic acid sets up an equilibrium, $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$, with only a small fraction ionized.

Step 3 — Conclusion: strong acids ionize almost completely in water.

Why each other option is wrong:

- (B) "Only to a very small extent" describes a *weak* acid, not a strong one.
- (C) "Not at all" would mean the substance is not an acid at all.
- (D) "Exactly halfway" is not a defining property of any acid class; the degree of ionization is not fixed at 50%.

Key point: Acid strength is about the degree of ionization, not the concentration. Strong acids (HCl , HNO_3 , H_2SO_4) are nearly fully ionized in water; weak acids are only slightly ionized.

Final Answer: Almost completely \Rightarrow

Answer: (A) [Go Back to Q14](#)

Q15.

Solution

Concept — Oxidation number of oxygen in peroxides: Oxidation numbers in a neutral compound add up to zero, with hydrogen usually +1. Oxygen is usually -2 , but in *peroxides* (which contain an $O-O$ single bond) each oxygen has an oxidation number of -1 . Hydrogen peroxide, $H-O-O-H$, is the simplest peroxide, so its oxygen atoms are an exception to the usual -2 rule.

Given: H_2O_2 , a neutral molecule with two H (+1 each) and two O (unknown).

Step 1 — Assign the known values: the two hydrogens contribute $2 \times (+1) = +2$.



Let the oxidation number of each oxygen be x .

Step 2 — Set the sum to zero: $(+2) + 2x = 0$.

Step 3 — Solve: $2x = -2 \Rightarrow x = -1$. So each oxygen in H_2O_2 has an oxidation number of -1 .

Why each other option is wrong:

- (A) -2 is the usual value of oxygen in ordinary oxides and in water, but not in a peroxide.
- (B) 0 is the oxidation number of oxygen in the free element O_2 , not in a compound.
- (C) $+1$ would apply only in compounds with fluorine (e.g. O_2F_2), where oxygen is bonded to the more electronegative fluorine.

Key point: The O–O bond in a peroxide is the giveaway: it makes each oxygen -1 instead of the usual -2 . This intermediate state is why H_2O_2 can act as both an oxidising and a reducing agent.

Final Answer: $-1 \Rightarrow$ D

Answer: (D) [Go Back to Q15](#)

Q16.

Solution

Concept — Role of the salt bridge: In a galvanic cell, oxidation in one half-cell releases positive ions into that solution, while reduction in the other half-cell removes positive ions. Without a connection, charge would build up and the cell would stop. The salt bridge (a tube of inert electrolyte such as KCl in a gel) allows ions to flow between the half-cells, completing the internal circuit and keeping both solutions electrically neutral.

Step 1 — See the problem the bridge solves: as the cell runs, the anode solution tends to become positively charged (loses electrons, gains cations) and the cathode solution negatively charged.

Step 2 — How the bridge responds: ions migrate through the bridge (anions toward the anode, cations toward the cathode) to cancel these charge build-ups.

Step 3 — Conclusion: the salt bridge completes the circuit and maintains electrical neutrality, allowing the current to keep flowing.

Why each other option is wrong:



- (A) Electrons flow through the external metal wire, not through the salt bridge; only ions move through the bridge.
- (C) The salt bridge contains an inert electrolyte and takes no part in oxidation or reduction.
- (D) The salt bridge does not heat the cell; it has no thermal role.

Key point: The salt bridge carries the internal ionic current and prevents charge build-up, keeping both half-cells neutral. Remove it and the cell voltage collapses almost at once.

Final Answer: Complete the circuit and maintain electrical neutrality \Rightarrow

Answer: [Go Back to Q16](#)

Q17.

Solution

Concept — Faraday's first law of electrolysis: Faraday's first law states that the mass m of a substance deposited or liberated at an electrode is directly proportional to the quantity of electric charge Q passed through the electrolyte: $m \propto Q$, or $m = Z Q$, where Z is the electrochemical equivalent. Since $Q = I t$ (current \times time), more charge means more product.

Step 1 — State the law: $m \propto Q$.

Step 2 — Express charge in measurable terms: $Q = I t$, so doubling either the current or the time (and hence the charge) doubles the mass deposited.

Step 3 — Conclusion: the mass deposited is proportional to the quantity of charge passed through the cell.

Why each other option is wrong:

- (A) Temperature has only a minor, indirect effect; it is not what the first law links the mass to.
- (B) The volume of electrolyte does not set the mass deposited; the charge does.
- (D) The electrode area affects the current density but not the total mass for a given charge.

Key point: Faraday's first law: mass \propto charge ($m = Z Q = Z I t$). The second law adds that, for the same charge, the masses of different substances are in the ratio of their equivalent masses.



Final Answer: Quantity of charge passed \Rightarrow

Answer: (C) [Go Back to Q17](#)

Q18.

Solution

Concept — Moles from molarity and volume: Molarity (M) is defined as the number of moles of solute per litre of solution: $M = \frac{n}{V}$, where V is the volume in litres. Rearranging gives the number of moles directly: $n = M \times V$.

Given: molarity $M = 0.5 \text{ mol L}^{-1}$; volume $V = 2 \text{ L}$.

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

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

Step 3 — Evaluate: the litres cancel, leaving $n = 1 \text{ mol}$.

Why each other option is wrong:

- (B) 0.25 mol comes from dividing by the volume instead of multiplying.
- (C) 2.5 mol does not follow from 0.5×2 by any correct operation.
- (D) 4 mol would require dividing the volume by the molarity ($2 \div 0.5$), which is dimensionally wrong.

Key point: Moles = molarity \times volume in litres. Keep the volume in litres (not millilitres) so the units cancel cleanly; here $0.5 \times 2 = 1 \text{ mol}$.

Final Answer: 1 mol \Rightarrow

Answer: (A) [Go Back to Q18](#)

Q19.

Solution

Concept — Temperature and molecular speed: In the kinetic theory of gases, the absolute temperature is a direct measure of the average kinetic energy of the molecules: $\overline{KE} = \frac{3}{2}kT$. Since kinetic energy is $\frac{1}{2}mv^2$, a higher temperature means a larger average kinetic energy and therefore a larger average molecular speed (the speed rises with \sqrt{T}).

Step 1 — Link temperature to kinetic energy: the average kinetic energy is proportional to the absolute temperature T .



Step 2 — Link kinetic energy to speed: because $KE = \frac{1}{2}mv^2$, a greater average kinetic energy corresponds to a greater average speed (root-mean-square speed $\propto \sqrt{T}$).

Step 3 — Conclusion: raising the temperature increases the average speed of the gas molecules.

Why each other option is wrong:

- (A) Decreasing would require cooling the gas, the opposite of raising the temperature.
- (B) "Remains unchanged" ignores the direct dependence of molecular speed on temperature.
- (C) There is no "decrease then increase"; the speed rises smoothly and monotonically with temperature.

Key point: Higher temperature \rightarrow higher average kinetic energy \rightarrow higher average molecular speed ($v_{\text{rms}} \propto \sqrt{T}$). This is why warm gases diffuse and effuse faster than cold ones.

Final Answer: Increases \Rightarrow D

Answer: (D) [Go Back to Q19](#)

Q20.

Solution

Concept — Oxidation state of a metal in a complex ion: In a complex ion, the sum of the oxidation state of the central metal and the charges of all the ligands equals the overall charge on the ion. You assign each ligand its known charge, then solve for the metal. The hydroxide ligand, OH^- , carries a charge of -1 .

Given: complex ion $[\text{Zn}(\text{OH})_4]^{2-}$; four OH^- ligands; overall charge -2 .

Step 1 — Total ligand charge: four hydroxides give $4 \times (-1) = -4$. Let the oxidation state of Zn be x .

Step 2 — Set the sum equal to the ion charge: $x + (-4) = -2$.

Step 3 — Solve: $x = -2 + 4 = +2$. So zinc is in the $+2$ oxidation state.

Why each other option is wrong:

- (A) $+4$ would require the overall charge to be 0 , not -2 .



- (C) 0 (free metal) is impossible inside a charged complex with four anionic ligands.
- (D) -2 confuses the metal's oxidation state with the overall charge of the ion.

Key point: For a complex ion, metal oxidation state + sum of ligand charges = overall charge. Zinc almost always shows the $+2$ state, consistent with the result here.

Final Answer: $+2 \Rightarrow$

Answer: (B) [Go Back to Q20](#)

Q21.

Solution

Concept — Denticity of a ligand: The denticity of a ligand is the number of donor atoms it uses to bind to a central metal ion. A monodentate ligand binds through one atom, a bidentate through two, and so on. The prefix "hexa-" means six, so a *hexadentate* ligand binds through six donor atoms, often wrapping right around the metal.

Step 1 — Interpret the prefix: "hexadentate" = "hexa" (six) + "dentate" (toothed/binding), so it has six points of attachment.

Step 2 — Match to the figure: the ligand in the diagram reaches the metal at six donor atoms (the six dots), occupying six coordination positions.

Step 3 — Conclusion: a hexadentate ligand binds through 6 donor atoms. The classic example is EDTA, which uses two nitrogen and four oxygen donor atoms.

Why each other option is wrong:

- (A) 2 donor atoms describe a *bidentate* ligand (e.g. ethylenediamine).
- (B) 4 donor atoms describe a *tetradentate* ligand.
- (D) 1 donor atom describes a *monodentate* ligand (e.g. NH_3 or Cl^-).

Key point: The "-dentate" prefix gives the count of donor atoms directly: mono (1), bi (2), tetra (4), hexa (6). A hexadentate ligand such as EDTA can fill all six positions of an octahedral metal by itself.

Final Answer: 6 donor atoms \Rightarrow

Answer: (C) [Go Back to Q21](#)



Q22.

Solution

Concept — Amphoteric oxides: Oxides are classified by their acid–base behaviour. Basic oxides (of metals such as Na, Mg) react with acids; acidic oxides (of non-metals such as CO_2 , SO_2) react with bases. An *amphoteric* oxide can react with *both* acids and bases, behaving as a base toward acids and as an acid toward bases. Aluminium oxide, Al_2O_3 , is a classic amphoteric oxide.

Step 1 — Test Al_2O_3 with an acid: $\text{Al}_2\text{O}_3 + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2\text{O}$ (acts as a base).

Step 2 — Test Al_2O_3 with a base: $\text{Al}_2\text{O}_3 + 2\text{NaOH} + 3\text{H}_2\text{O} \rightarrow 2\text{Na}[\text{Al}(\text{OH})_4]$ (acts as an acid).

Step 3 — Conclusion: because it reacts with both, Al_2O_3 is amphoteric.

Why each other option is wrong:

- (B) Na_2O is a strongly *basic* oxide; it reacts with acids but not with bases.
- (C) CO_2 is an *acidic* oxide; it reacts with bases (forming carbonates) but not with acids.
- (D) MgO is a *basic* oxide; it reacts with acids only.

Key point: Amphoteric oxides (such as Al_2O_3 , ZnO , and PbO) react with both acids and bases. They typically come from elements on the metal/non-metal borderline.

Final Answer: $\text{Al}_2\text{O}_3 \Rightarrow \boxed{\text{A}}$

Answer: (A) [Go Back to Q22](#)

Q23.

Solution

Concept — Conditions for geometrical (cis–trans) isomerism: Geometrical isomerism arises around a $\text{C}=\text{C}$ double bond because the double bond prevents free rotation. For the isomerism to exist, *each* of the two double-bonded carbons must carry two *different* groups. If either carbon bears two identical groups, the cis and trans forms become the same molecule, and no geometrical isomerism is possible.

Step 1 — Examine but-2-ene, $\text{CH}_3\text{CH}=\text{CHCH}_3$: each double-bonded carbon carries one CH_3 and one H, that is, two different groups, so cis (both CH_3 on the same side) and trans (on opposite sides) forms exist.

Step 2 — Check the other options: ethene has two H's on each carbon, propene



has two H's on its terminal CH_2 carbon, and but-1-ene also has a terminal $=\text{CH}_2$ with two identical H's. In each case one carbon bears two identical groups, so no cis-trans isomers form.

Step 3 — Conclusion: only but-2-ene satisfies the requirement and shows geometrical isomerism.

Why each other option is wrong:

- (A) Ethene ($\text{CH}_2=\text{CH}_2$) has two identical H's on each carbon, so no cis-trans forms.
- (B) Propene has a terminal $=\text{CH}_2$ carbon bearing two identical H's.
- (C) But-1-ene also has a terminal $=\text{CH}_2$ group, so one carbon has two identical substituents.

Key point: For geometrical isomerism, look for a $\text{C}=\text{C}$ with each carbon carrying two different groups. A terminal double bond ($=\text{CH}_2$) never shows cis-trans isomerism, which is why but-2-ene works but but-1-ene does not.

Final Answer: but-2-ene \Rightarrow D

Answer: (D) [Go Back to Q23](#)

Q24.

Solution

Concept — IUPAC naming of alcohols: To name an alcohol, find the longest carbon chain containing the $-\text{OH}$ group, name it after the parent alkane, replace the final "-e" with "-ol", and add a locant number showing the position of the $-\text{OH}$. The compound $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ has three carbons in a row with the $-\text{OH}$ on an end carbon.

Step 1 — Count the carbons: three carbons in the chain \Rightarrow the parent is propane.

Step 2 — Locate the $-\text{OH}$: it is on carbon 1 (an end carbon), so the locant is 1.

Step 3 — Assemble the name: propane \rightarrow propan-1-ol. Hence the IUPAC name is propan-1-ol.

Why each other option is wrong:

- (A) Ethanol has only two carbons ($\text{CH}_3\text{CH}_2\text{OH}$), not three.
- (C) Propan-2-ol has the $-\text{OH}$ on the *middle* carbon ($\text{CH}_3\text{CHOHCH}_3$), not the end.



- (D) Propanal is an *aldehyde* ($\text{CH}_3\text{CH}_2\text{CHO}$), a different functional group, not an alcohol.

Key point: Number the chain so the $-\text{OH}$ gets the lowest locant, and use the "-ol" suffix. A three-carbon chain with $-\text{OH}$ on an end carbon is propan-1-ol.

Final Answer: propan-1-ol \Rightarrow

Answer: (B) [Go Back to Q24](#)

Q25.

Solution

Concept — Saturated hydrocarbons: A hydrocarbon is "saturated" when it holds the maximum possible number of hydrogen atoms, which happens only when all carbon-carbon bonds are single bonds. Such compounds are the alkanes, with the general formula $\text{C}_n\text{H}_{2n+2}$. The presence of any double or triple bond would make the molecule unsaturated.

Step 1 — Recall the meaning of "saturated": every carbon is bonded to as many hydrogens as possible, which requires only single C-C and C-H bonds.

Step 2 — Identify the bond type: single bonds only, with no π bonds present.

Step 3 — Conclusion: a saturated hydrocarbon (an alkane) contains only single bonds.

Why each other option is wrong:

- (A) Double bonds make a hydrocarbon *unsaturated* (an alkene).
- (B) Triple bonds make a hydrocarbon *unsaturated* (an alkyne).
- (D) Alternate single and double bonds describe a conjugated/aromatic system, which is also unsaturated.

Key point: "Saturated" = single bonds only = alkanes ($\text{C}_n\text{H}_{2n+2}$). Any double or triple bond means the molecule is unsaturated and can take up more hydrogen by addition.

Final Answer: Single bonds \Rightarrow

Answer: (C) [Go Back to Q25](#)



Q26.

Solution

Concept — Characteristic reactions of alkenes: Alkenes contain a C=C double bond, which is electron-rich because of its π electrons. This π bond is relatively weak and easily attacked, so alkenes typically undergo *addition* reactions: the double bond opens and two new atoms or groups add across it, converting the unsaturated alkene into a saturated product.

Step 1 — Recall the nature of the C=C bond: the π bond is exposed and reactive toward electrophiles.

Step 2 — See what happens in a typical reaction: for example $\text{CH}_2=\text{CH}_2 + \text{Br}_2 \rightarrow \text{CH}_2\text{Br}-\text{CH}_2\text{Br}$; the double bond opens and a bromine atom adds to each carbon.

Step 3 — Conclusion: the reaction characteristic of alkenes is addition.

Why each other option is wrong:

- (B) Substitution is characteristic of *saturated* alkanes (and of aromatic rings), where one atom replaces another, not of alkenes.
- (C) Neutralization is an acid–base reaction, not a reaction of a carbon–carbon double bond.
- (D) Precipitation involves forming an insoluble solid from solution and is unrelated to the alkene double bond.

Key point: Alkenes "add," alkanes "substitute." The reactive π bond of an alkene undergoes addition (with H_2 , X_2 , HX , water, etc.), turning the double bond into a single bond.

Final Answer: Addition \Rightarrow

[Go Back to Q26](#)

Q27.

Solution

Concept — Hybridisation in ethene: A carbon atom's hybridisation equals its number of σ bonds plus lone pairs (its number of electron domains). A double bond counts as *one* σ bond, since the extra bond is a π bond that does not change the hybridisation. In ethene, $\text{CH}_2=\text{CH}_2$, each carbon is bonded to two hydrogens and to the other carbon.

Step 1 — Count the σ bonds on one carbon: two C–H σ bonds plus one C–C



σ bond (from the double bond) = 3 σ bonds, with no lone pairs (three electron domains).

Step 2 — Assign the hybridisation: three domains $\Rightarrow sp^2$. The three sp^2 orbitals lie in a plane at 120° , and the leftover unhybridised p orbital forms the π bond.

Step 3 — Geometry: each carbon is trigonal planar, and the whole ethene molecule is flat, as shown in the figure.

Why each other option is wrong:

- (A) sp^3 (four domains) is for a carbon with four single bonds, as in ethane.
- (B) sp (two domains) is for a triple bond, as in ethyne.
- (C) sp^3d (five domains) cannot occur for carbon, which has no available d orbitals.

Key point: Count a double bond as one σ bond for hybridisation. sp^2 always goes with a planar 120° geometry and exactly one π bond, which is why ethene is flat.

Final Answer: $sp^2 \Rightarrow$ D

Answer: (D) [Go Back to Q27](#)

Q28.

Solution

Concept — Oxidation of a primary alcohol: A primary alcohol has its $-OH$ on a carbon bonded to only one other carbon ($R-CH_2OH$). On controlled (mild) oxidation, this carbon loses two hydrogens and forms a carbon–oxygen double bond, giving an aldehyde ($R-CHO$). Stronger or prolonged oxidation would go further and convert the aldehyde into a carboxylic acid.

Step 1 — Identify the starting class: a primary alcohol, e.g. ethanol CH_3CH_2OH .

Step 2 — Apply mild oxidation: $CH_3CH_2OH \xrightarrow{[O]} CH_3CHO + H_2O$; the $-CH_2OH$ group becomes a $-CHO$ group.

Step 3 — Conclusion: the first product, containing the $-CHO$ group, is an aldehyde (here ethanal).

Why each other option is wrong:

- (A) A ketone is formed by oxidising a *secondary* alcohol, not a primary one.
- (C) An ether ($R-O-R$) is not an oxidation product of an alcohol.



- (D) An ester forms by reacting an alcohol with a carboxylic acid (esterification), not by mild oxidation.

Key point: Primary alcohol $\xrightarrow{\text{mild [O]}}$ aldehyde $\xrightarrow{\text{further [O]}}$ carboxylic acid; secondary alcohol $\xrightarrow{\text{[O]}}$ ketone. Mild oxidation stops at the aldehyde stage.

Final Answer: An aldehyde \Rightarrow

Answer: [Go Back to Q28](#)

Q29.

Solution

Concept — Functional groups in glucose: Glucose is an aldohexose, a six-carbon monosaccharide. In its open-chain form it has one aldehyde group ($-\text{CHO}$) at carbon 1 and several hydroxyl groups ($-\text{OH}$) on the remaining carbons. The "aldo-" part of "aldohexose" signals the aldehyde, and the many $-\text{OH}$ groups make it highly soluble in water.

Step 1 — Locate the aldehyde group: carbon 1 bears a $-\text{CHO}$ group, which is why glucose gives a positive test with Tollens' and Fehling's reagents (reducing sugar).

Step 2 — Locate the hydroxyl groups: carbons 2 to 6 carry $-\text{OH}$ groups (five of them in all), giving glucose its sugar-like, water-soluble character.

Step 3 — Conclusion: glucose contains an aldehyde group and several hydroxyl groups, so the two functional groups are aldehyde and hydroxyl.

Why each other option is wrong:

- (A) Glucose has an *aldehyde*, not a ketone (fructose is the ketone-containing sugar).
- (B) Glucose contains neither a carboxylic acid nor an amino group.
- (D) Glucose has hydroxyl groups, not a carboxylic acid group, alongside its aldehyde.

Key point: Glucose = one aldehyde ($-\text{CHO}$) + several hydroxyls ($-\text{OH}$). The aldehyde makes it a reducing sugar, while the many $-\text{OH}$ groups make it freely soluble in water.

Final Answer: Aldehyde and hydroxyl \Rightarrow

Answer: [Go Back to Q29](#)



Q30.

Solution

Concept — Basic character of amines: An amine contains the $-\text{NH}_2$ group, in which the nitrogen atom holds a lone pair of electrons. This lone pair can be donated to (accept) a proton, so amines act as Bronsted–Lowry bases (and as Lewis bases). When an amine reacts with water or an acid, the nitrogen lone pair grabs an H^+ , forming a substituted ammonium ion.

Step 1 — Identify the reactive feature: the lone pair on the nitrogen of $-\text{NH}_2$ is available to bond with a proton.

Step 2 — Show the base behaviour: for example $\text{R}-\text{NH}_2 + \text{H}^+ \rightarrow \text{R}-\text{NH}_3^+$; the amine accepts a proton, the defining action of a base.

Step 3 — Conclusion: the amine functional group is basic in character.

Why each other option is wrong:

- (B) Amines do not donate protons readily, so they are not strong acids; their character is the opposite.
- (C) Amines are not unreactive; the nitrogen lone pair makes them clearly basic and nucleophilic.
- (D) Amines are not powerful oxidising agents; that property belongs to species like KMnO_4 or $\text{K}_2\text{Cr}_2\text{O}_7$.

Key point: The lone pair on nitrogen is the key: it lets amines accept protons, so the $-\text{NH}_2$ group is basic. Compare carboxylic acids ($-\text{COOH}$, acidic) with amines ($-\text{NH}_2$, basic).

Final Answer: Base \Rightarrow

Answer: (A) [Go Back to Q30](#)



Answer Key

Q	Ans	Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	C	2	A	3	D	4	B	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

