# AIIMS B.Sc Nursing Chemistry

## Sample Paper – 4

**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 mass of $2$ moles of water ($H_2O$, molar mass $= 18$ g mol$^{-1}$) is:

(A) $36$ g

(B) $18$ g

(C) $9$ g

(D) $72$ g

**Q2.** The number of moles present in $3.011 \times 10^{23}$ molecules of a gas (Avogadro number $= 6.022 \times 10^{23}$) is:

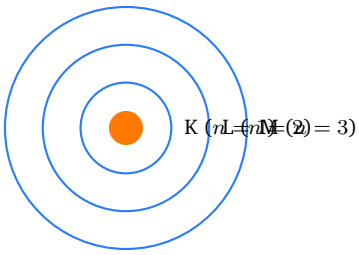
(A) $1$ mol

(B) $0.5$ mol

(C) $2$ mol

(D) $0.25$ mol

**Q3.** The third shell ($n = 3$, the M shell) of an atom is shown below. The number of subshells (s, p, d, ...) it contains, equal to $n$, is:




(A) $1$

(B) $2$

(C) $3$

(D) $4$

**Q4.** The filling of the $2p$ subshell of a carbon atom ($Z = 6$, configuration $1s^2\,2s^2\,2p^2$) is shown. The number of unpaired electrons in a ground-state carbon atom is:

$2p^2$

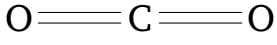
(A) $0$

(B) $1$

(C) $3$

(D) $2$

**Q5.** The bonding in a carbon dioxide molecule is shown below, with carbon doubly bonded to two oxygen atoms and no lone pair on carbon. The shape of the $CO_2$ molecule is:




$180^\circ$ bond angle

(A) linear

(B) bent

(C) trigonal planar

(D) tetrahedral

**Q6.** The hybridization of the nitrogen atom in an ammonia molecule ($NH_3$, which has three N–H bonds and one lone pair on nitrogen) is:

(A) $sp$

(B) $sp^3$

(C) $sp^2$

(D) $sp^3d$

**Q7.** Which one of the following molecules shows intermolecular hydrogen bonding?

(A) $CH_4$

(B) $H_2S$

(C) HF

(D) $CO_2$

**Q8.** On moving from left to right across a period in the periodic table, the metallic character of the elements generally:

(A) increases

(B) remains constant

(C) first decreases, then increases

(D) decreases

**Q9.** Among the elements oxygen (O), nitrogen (N), carbon (C), and boron (B), the most electronegative element is:

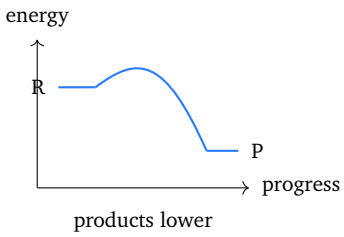
(A) Oxygen

(B) Nitrogen

(C) Carbon

(D) Boron

**Q10.** The energy profiles of two reactions are sketched below. An exothermic reaction is one in which heat is released. Which choice correctly describes an exothermic reaction?




(A) $\Delta H$ is positive; heat is absorbed

(B) $\Delta H$ is negative; heat is released

(C) $\Delta H$ is zero; no heat change

(D) products are higher in energy than reactants

**Q11.** A chemical reaction is spontaneous at *all* temperatures (since $\Delta G = \Delta H - T\Delta S$ stays negative) when:

(A) $\Delta H > 0$ and $\Delta S < 0$

(B) $\Delta H > 0$ and $\Delta S > 0$

(C) $\Delta H < 0$ and $\Delta S > 0$

(D) $\Delta H < 0$ and $\Delta S < 0$

**Q12.** For a reaction at equilibrium, $\mathrm{A} + \mathrm{B} \rightleftharpoons \mathrm{C} + \mathrm{D}$, if some of the product C is continuously removed from the system, the equilibrium shifts:

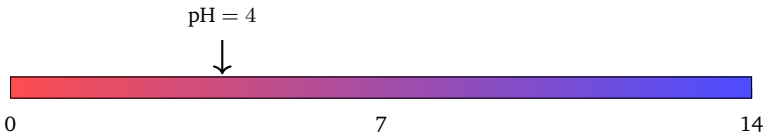
(A) towards the reactants (backward)

(B) no shift occurs

(C) towards both sides equally

(D) towards the products (forward)

**Q13.** At $25^\circ$C a solution has a pH of $4$. Using $\mathrm{pH} + \mathrm{pOH} = 14$, the pOH of this solution is:




(A) 10

(B) 4

(C) 7

(D) 14

**Q14.** A Lewis acid is a species that accepts a pair of electrons. Which one of the following acts as a Lewis acid?

(A) $NH_3$

(B) $BF_3$

(C) $H_2O$

(D) $OH^-$

**Q15.** The oxidation number of nitrogen in nitric acid ($HNO_3$) is:

(A) $+3$

(B) $-3$

(C) $+5$

(D) $+4$

**Q16.** In the redox change shown, the species that brings about reduction of the other reactant is the reducing agent. A reducing agent is best defined as a substance that:

$$\underset{\text{X acts as reducing agent}}{X} \xrightarrow{\text{loses } e^- \text{ (oxidised)}} X^+$$

(A) gains electrons and is itself reduced

(B) neither gains nor loses electrons

(C) only donates protons

(D) loses electrons and is itself oxidised

**Q17.** The quantity of charge required to deposit $0.5$ mole of silver from a silver salt solution ($Ag^{+} + e^{-} \rightarrow Ag$) is (Faraday constant $= 96500$ C $mol^{-1}$):

(A) $48250$ C

(B) $96500$ C

(C) $193000$ C

(D) $24125$ C

**Q18.** The mole fraction of a component A in a solution is defined as:

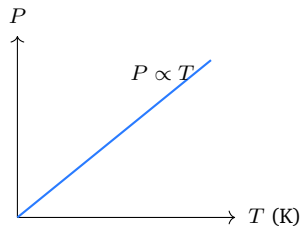
(A) moles of A per litre of solution

(B) moles of A divided by the total moles of all components

(C) moles of A per kilogram of solvent

(D) mass of A divided by total mass of solution

**Q19.** The graph below shows how the pressure of a fixed mass of gas at constant volume varies with absolute temperature. The straight line through the origin illustrates Gay-Lussac's law, which states that pressure is:




(A) inversely proportional to absolute temperature

(B) independent of temperature

(C) directly proportional to absolute temperature

(D) proportional to the square of temperature

**Q20.** The oxidation state of nickel in the complex $[Ni(CO)_4]$ (each carbon monoxide ligand is a neutral molecule) is:

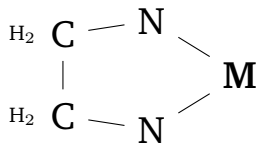
(A) $+4$

(B) $+2$

(C) $+1$

(D) $0$

**Q21.** A bidentate ligand binds the central metal through two donor atoms. Ethylenediamine (en, $H_2N–CH_2–CH_2–NH_2$) is shown chelating a metal M below. Which of the following is a bidentate ligand?




(A) ethylenediamine (en)

(B) ammonia ($NH_3$)

(C) chloride ($Cl^-$)

(D) water ($H_2O$)

**Q22.** Which one of the following elements is a transition metal?

(A) Calcium

(B) Iron

(C) Sodium

(D) Aluminium

**Q23.** Functional isomers have the same molecular formula but different functional groups. Which of the following pair are functional isomers?

(A) ethane and propane

(B) $n$-butane and isobutane

(C) ethanol and dimethyl ether

(D) ethene and ethyne

**Q24.** The IUPAC name of the simplest hydrocarbon, $CH_4$, is:

(A) methanol

(B) ethane

(C) methanal

(D) methane

**Q25.** The general molecular formula of the saturated aliphatic alcohol homologous series is:

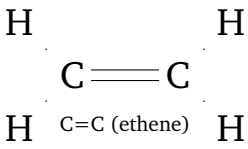
(A) $C_nH_{2n+1}OH$

(B) $C_nH_{2n}O$

(C) $C_nH_{2n}$

(D) $C_nH_{2n-1}OH$

**Q26.** An unsaturated hydrocarbon contains at least one carbon–carbon double or triple bond. Which of the following is an unsaturated hydrocarbon?




(A) methane ($CH_4$)

(B) ethane ($C_2H_6$)

(C) propane ($C_3H_8$)

(D) ethene ($C_2H_4$)

**Q27.** The total number of sigma ($\sigma$) bonds present in one molecule of ethane ($C_2H_6$, which has six C–H bonds and one C–C bond) is:

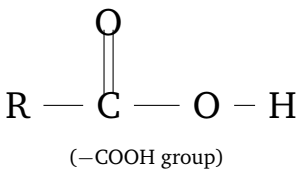
(A) 6

(B) 7

(C) 8

(D) 5

**Q28.** The functional group shown below, a carbon bearing both a double-bonded oxygen and a hydroxyl group on the same carbon, is characteristic of which class of compounds?




(A) alcohols

(B) aldehydes

(C) carboxylic acids

(D) ketones

**Q29.** Which one of the following carbohydrates is a polysaccharide?

(A) starch

(B) glucose

(C) fructose

(D) sucrose

**Q30.** A compound containing the amino group ($-NH_2$), such as an amine, generally behaves as a:

(A) strong acid

(B) neutral substance

(C) base

(D) salt

## Detailed Solutions

### Q1.

**Solution**

**Concept — The mole and molar mass:** The mole is the SI unit for the amount of a substance. One mole of any substance has a mass equal to its molar mass expressed in grams and contains Avogadro's number ($6.022 \times 10^{23}$) of particles. To convert a number of moles into a mass we use $m = n \times M$, where $n$ is the number of moles and $M$ is the molar mass. The molar mass of water is the sum of the atomic masses: $2\times$H $(2 \times 1)+$ O $(16) = 18$ g mol$^{-1}$.

**Given:** $n = 2$ mol; molar mass $M(\mathrm{H_2O}) = 18$ g mol$^{-1}$.

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

**Step 2 — Substitute the data with units:** $m = 2$ mol $\times$ 18 g mol$^{-1}$.

**Step 3 — Evaluate:** the moles cancel, leaving $m = 36$ g.

**Why each other option is wrong:**

- (B) 18 g is the mass of only 1 mole of water, not 2 moles.
- (C) 9 g is half a mole ($0.5 \times 18$), four times too small.
- (D) 72 g would be 4 moles of water ($4 \times 18$).

**Key point:** Mass and moles are linked by $m = n \times M$. Always compute the molar mass first by summing the atomic masses; a common slip is to forget one of the two hydrogen atoms.

**Final Answer:** 36 g $\Rightarrow$ A

**Answer: (A)** Go Back to Q1

### Q2.

**Solution**

**Concept — Moles from a number of particles:** Avogadro's number, $N_A = 6.022 \times 10^{23}$, is the number of particles in exactly one mole. To convert a number of molecules into moles, divide by $N_A$: $n = \dfrac{\text{number of molecules}}{N_A}$. This is simply the reverse of multiplying moles by $N_A$ to get particles.

**Given:** number of molecules $= 3.011 \times 10^{23}$; $N_A = 6.022 \times 10^{23}$.

**Step 1 — Write the formula:** $n = \dfrac{\text{number of molecules}}{N_A}$.

**Step 2 — Substitute:** $n = \dfrac{3.011 \times 10^{23}}{6.022 \times 10^{23}}$.

**Step 3 — Evaluate:** the powers of ten cancel and $3.011/6.022 = 0.5$, so $n = 0.5$ mol.

**Why each other option is wrong:**

- (A) $1$ mol would need $6.022 \times 10^{23}$ molecules, twice the given number.
- (C) $2$ mol would need $1.2 \times 10^{24}$ molecules, four times too many.
- (D) $0.25$ mol corresponds to $1.5 \times 10^{23}$ molecules, half the given number.

**Key point:** Notice that $3.011 \times 10^{23}$ is exactly half of $6.022 \times 10^{23}$, so the answer is $0.5$ mol without a calculator. Dividing a particle count by $N_A$ always gives moles.

**Final Answer:** $0.5$ mol $\Rightarrow$ B

**Answer: (B)** Go Back to Q2

## Q3.

### Solution

**Concept — Subshells within a shell:** Each principal shell (labelled by the principal quantum number $n$) is made up of subshells described by the azimuthal quantum number $\ell$, which runs from $0$ up to $(n-1)$. The values $\ell = 0, 1, 2, 3, \ldots$ correspond to the s, p, d, f subshells. Therefore the number of subshells in a shell is exactly equal to $n$.

**Step 1 — Identify $n$:** the M shell is the third shell, so $n = 3$.

**Step 2 — List the allowed values of $\ell$:** $\ell = 0, 1, 2$, i.e. three values.

**Step 3 — Name the subshells:** $\ell = 0$ is $3s$, $\ell = 1$ is $3p$, $\ell = 2$ is $3d$, giving $3$ subshells in total.

**Why each other option is wrong:**

- (A) $1$ is the number of subshells in the K shell ($n = 1$: only $1s$).
- (B) $2$ is the number of subshells in the L shell ($n = 2$: $2s$ and $2p$).
- (D) $4$ would be the N shell ($n = 4$: $4s, 4p, 4d, 4f$), not the M shell.

**Key point:** The number of subshells in a shell equals $n$, while the maximum number of electrons it can hold equals $2n^2$. Do not confuse the two: for $n = 3$, there

are $3$ subshells but a capacity of $18$ electrons.

**Final Answer:** $3$ subshells $\Rightarrow$ C

**Answer: (C)** Go Back to Q3

## Q4.

### Solution

**Concept — Filling orbitals (aufbau, Pauli, Hund):** Electrons fill orbitals by three rules: the aufbau principle (lowest-energy orbitals first), Pauli's exclusion principle (at most two electrons per orbital, with opposite spins), and Hund's rule of maximum multiplicity (degenerate orbitals are filled singly with parallel spins before any pairing). Carbon has $Z = 6$, so its ground-state configuration is $1s^2\,2s^2\,2p^2$.

**Step 1 — Write the configuration:** $1s^2\,2s^2\,2p^2$. The $1s$ and $2s$ subshells are completely filled (all paired); the deciding part is the $2p^2$.

**Step 2 — Distribute the two $2p$ electrons among the three $2p$ orbitals using Hund's rule:** the two electrons go into two separate orbitals with parallel spins, giving $\uparrow|\uparrow|$ , rather than pairing in one orbital.

**Step 3 — Count the unpaired electrons:** two orbitals each hold a single (unpaired) electron, so there are $2$ unpaired electrons.

**Why each other option is wrong:**

- (A) $0$ would require the two electrons to pair in one orbital, which violates Hund's rule.
- (B) $1$ ignores that both $2p$ electrons remain unpaired.
- (C) $3$ is impossible with only two electrons in the $2p$ subshell.

**Key point:** By Hund's rule, electrons spread out singly into degenerate orbitals before pairing. A $p^2$ configuration therefore gives $2$ unpaired electrons, which is why carbon is paramagnetic in this state.

**Final Answer:** $2$ unpaired electrons $\Rightarrow$ D

**Answer: (D)** Go Back to Q4

## Q5.

**Solution**

**Concept — VSEPR shape of $CO_2$:** The shape of a molecule is set by the number of electron domains (bonding regions and lone pairs) around the central atom, which spread out to be as far apart as possible. A double bond counts as a single domain. In $CO_2$ the central carbon forms two double bonds to the two oxygen atoms and has no lone pair, giving just two electron domains.

**Step 1 — Count the electron domains on carbon:** two C=O bonds and zero lone pairs $= 2$ domains.

**Step 2 — Arrange the domains:** two domains point in opposite directions to be as far apart as possible, i.e. at $180°$.

**Step 3 — State the shape:** the molecule is linear, O=C=O, with a bond angle of $180°$.

**Why each other option is wrong:**

- (B) Bent ($\approx 104.5°$) requires lone pairs on the central atom, as in $H_2O$; carbon here has none.
- (C) Trigonal planar needs three domains, as in $BF_3$.
- (D) Tetrahedral needs four domains, as in $CH_4$.

**Key point:** Two electron domains always give a linear shape. Although each C=O bond is polar, the linear symmetry makes the two dipoles cancel, so $CO_2$ is non-polar overall.

**Final Answer:** Linear $\Rightarrow$ A

**Answer: (A)** Go Back to Q5

## Q6.

**Solution**

**Concept — Hybridisation from electron domains:** The hybridisation of a central atom equals its number of electron domains, that is the number of sigma bonds plus the number of lone pairs. Two domains give $sp$, three give $sp^2$, four give $sp^3$. In ammonia, $NH_3$, nitrogen has five valence electrons: three are used in the three N–H bonds and the remaining two form one lone pair.

**Step 1 — Count the domains on nitrogen:** three N–H sigma bonds plus one lone pair $= 4$ electron domains.

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

**Step 3 — Relate to the shape:** one of the four $sp^3$ orbitals holds the lone pair, so the molecular shape is trigonal pyramidal with a bond angle of about $107°$, though the hybridisation is still $sp^3$.

**Why each other option is wrong:**

- (A) $sp$ (two domains) describes a linear atom such as carbon in $CO_2$.
- (C) $sp^2$ (three domains) describes a trigonal planar atom, ignoring the lone pair on nitrogen.
- (D) $sp^3d$ (five domains) requires $d$ orbitals and does not apply to nitrogen.

**Key point:** Count lone pairs as domains when finding hybridisation. Nitrogen in $NH_3$ has $3$ bonds $+1$ lone pair $= 4$ domains $= sp^3$.

**Final Answer:** $sp^3 \Rightarrow$ B

**Answer: (B)** Go Back to Q6

## Q7.

### Solution

**Concept — Hydrogen bonding:** A hydrogen bond is a strong dipole attraction that forms when a hydrogen atom is bonded directly to a small, highly electronegative atom, namely fluorine (F), oxygen (O), or nitrogen (N). The H then carries a large partial positive charge and is attracted to a lone pair on an F, O, or N atom of a neighbouring molecule.

**Step 1 — Look for an H bonded to F, O, or N:** in HF the hydrogen is bonded to fluorine, the most electronegative element, so it can hydrogen bond.

**Step 2 — Check the other options:** $CH_4$ has only C–H bonds (carbon is not electronegative enough); $H_2S$ has H bonded to sulfur (S is too large and not electronegative enough); $CO_2$ has no H atoms at all.

**Step 3 — Conclusion:** only HF satisfies the requirement and shows strong intermolecular hydrogen bonding.

**Why each other option is wrong:**

- (A) $CH_4$: C–H bonds are nearly non-polar, so no hydrogen bonding.
- (B) $H_2S$: sulfur is not electronegative enough, so $H_2S$ shows only weak dipole forces (this is why it is a gas while water is a liquid).

- (D) $CO_2$: it has no hydrogen atoms, so hydrogen bonding is impossible.

**Key point:** Hydrogen bonding needs H attached to F, O, or N. The mnemonic "FON" captures the only three partners that give true hydrogen bonds.

**Final Answer:** HF $\Rightarrow$ C

**Answer: (C)** Go Back to Q7

## Q8.

### Solution

**Concept — Metallic character across a period:** Metallic character is the tendency of an atom to lose electrons and form positive ions. It depends on how loosely the outer electrons are held. 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 become harder to lose.

**Step 1 — Identify the trend in electron loss:** as the effective nuclear charge increases across a period, ionisation energy rises and electrons are held more firmly.

**Step 2 — Relate to metallic character:** the harder it is to remove an electron, the less metallic (more non-metallic) the element becomes.

**Step 3 — State the trend:** metallic character therefore decreases from left to right across a period. (Down a group it increases, since the outer electrons are farther from the nucleus.)

**Why each other option is wrong:**

- (A) Increasing is the trend *down a group*, not across a period.
- (B) Remaining constant ignores the steadily rising nuclear charge.
- (C) There is no "decrease then increase" reversal within a normal period.

**Key point:** Metals are on the left, non-metals on the right. So across a period metallic character falls; down a group it rises. This is the opposite of the electronegativity trend.

**Final Answer:** Decreases $\Rightarrow$ D

**Answer: (D)** Go Back to Q8

## Q9.

**Solution**

**Concept — Electronegativity trend:** Electronegativity is the tendency of an atom to attract a shared bonding pair of electrons toward itself. On the Pauling scale it increases across a period from left to right and decreases down a group, so within the same period the element farthest to the right is the most electronegative.

**Step 1 — Place the elements:** boron, carbon, nitrogen, and oxygen all lie in the second period, in that order from left to right (groups 13, 14, 15, 16).

**Step 2 — Apply the period trend:** since electronegativity rises across a period, oxygen, being farthest to the right of the four, has the highest value. The rough order is O $(3.5)$ > N $(3.0)$ > C $(2.5)$ > B $(2.0)$.

**Step 3 — Conclusion:** oxygen is the most electronegative among the four.

**Why each other option is wrong:**

- (B) Nitrogen lies to the left of oxygen, so it is slightly less electronegative.
- (C) Carbon is farther left still, with a lower value.
- (D) Boron is the leftmost of the four and therefore the least electronegative.

**Key point:** Fluorine is the most electronegative element overall, but among the choices O, N, C, B (all in period 2), oxygen wins because it sits farthest to the right.

**Final Answer:** Oxygen $\Rightarrow$ $\boxed{\text{A}}$

**Answer: (A)** Go Back to Q9

## Q10.

**Solution**

**Concept — Exothermic reactions and $\Delta H$:** The enthalpy change is $\Delta H = H_{\text{products}} - H_{\text{reactants}}$. An exothermic reaction releases heat to the surroundings, which means the products store *less* energy than the reactants. On an energy profile the products lie below the reactants, and the energy difference is given out as heat, so $\Delta H$ is negative.

**Step 1 — Read the profile:** the products (P) are drawn lower than the reactants (R); the system falls to a lower energy.

**Step 2 — Apply the definition:** since $H_{\text{products}} < H_{\text{reactants}}$, $\Delta H$ = (lower) − (higher) is a negative quantity.

**Step 3 — Interpret:** a negative $\Delta H$ means heat is released to the surroundings, which is exactly what defines an exothermic reaction.

**Why each other option is wrong:**

- (A) A positive $\Delta H$ with heat absorbed describes an *endothermic* reaction, the opposite case.
- (C) Zero $\Delta H$ would mean no net energy change, which is not an exothermic reaction.
- (D) Products higher than reactants again describes an endothermic reaction.

**Key point:** Exothermic $\rightarrow$ heat released $\rightarrow$ $\Delta H$ negative $\rightarrow$ products lower on the diagram. Combustion and neutralisation are everyday exothermic examples.

**Final Answer:** $\Delta H$ negative; heat released $\Rightarrow$ $\boxed{\text{B}}$

**Answer: (B)** Go Back to Q10

## Q11.

### Solution

**Concept — Spontaneity and $\Delta G$:** A reaction is spontaneous when its Gibbs free energy change is negative: $\Delta G = \Delta H - T\Delta S$, where $\Delta H$ is the enthalpy change, $T$ the absolute (always positive) temperature, and $\Delta S$ the entropy change. To make $\Delta G$ negative at *every* temperature, both terms must work in the favourable direction.

**Step 1 — Examine each term:** for $\Delta G$ to be negative we want $\Delta H$ small (or negative) and $-T\Delta S$ negative. The term $-T\Delta S$ is negative only when $\Delta S$ is positive (because $T > 0$).

**Step 2 — Combine the requirements:** if $\Delta H < 0$ (favourable, makes $\Delta G$ more negative) and $\Delta S > 0$ (so $-T\Delta S < 0$, also favourable), then $\Delta G = (\text{negative}) - (\text{positive}) = \text{negative}$ at all $T$.

**Step 3 — Conclusion:** the reaction is spontaneous at all temperatures when $\Delta H < 0$ and $\Delta S > 0$.

**Why each other option is wrong:**

- (A) $\Delta H > 0$, $\Delta S < 0$ gives $\Delta G$ positive at all $T$: never spontaneous.
- (B) $\Delta H > 0$, $\Delta S > 0$ is spontaneous only at *high* temperatures, not all.
- (D) $\Delta H < 0$, $\Delta S < 0$ is spontaneous only at *low* temperatures, not all.

**Key point:** Only the combination $\Delta H < 0$ *and* $\Delta S > 0$ keeps $\Delta G$ negative for

every temperature. The other sign combinations are temperature dependent or never spontaneous.

**Final Answer:** $\Delta H < 0$ and $\Delta S > 0 \Rightarrow$ C

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

## Q12.

### Solution

**Concept — Le Chatelier's principle (concentration):** If a system at equilibrium is disturbed, it shifts in the direction that partly opposes the disturbance. Removing a product lowers the concentration on the product side, so the system responds by making more product to restore the balance, i.e. it shifts in the forward direction.

**Step 1 — Identify the disturbance:** product C is continuously removed, decreasing [C].

**Step 2 — Apply the principle:** to oppose the drop in product concentration, the equilibrium shifts to produce more C (and D), which is the forward (rightward) direction.

**Step 3 — Conclusion:** the equilibrium shifts towards the products. Continual removal of a product is in fact a common laboratory trick used to drive a reaction to completion.

**Why each other option is wrong:**

- (A) A backward shift toward reactants would follow from *adding* more product, not removing it.
- (B) "No shift" would only apply if the disturbance were inert (for example adding a catalyst or an inert gas at constant volume).
- (C) An equilibrium does not respond by moving "toward both sides equally."

**Key point:** Removing a product pulls the equilibrium forward; adding a product pushes it backward. This is how removing water or a gas can drive reactions to completion.

**Final Answer:** Towards the products (forward) $\Rightarrow$ D

**Answer: (D)** Go Back to Q12

## Q13.

### Solution

**Concept — Relation between pH and pOH:** At $25°$C, the pH and pOH of any aqueous solution are linked by $\text{pH} + \text{pOH} = 14$. This comes from the ionic product of water, $K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$, taken as a negative logarithm. Knowing one of the two immediately gives the other by subtraction from $14$.

**Given:** $\text{pH} = 4$.

**Step 1 — Write the relation:** $\text{pH} + \text{pOH} = 14$.

**Step 2 — Rearrange for pOH:** $\text{pOH} = 14 - \text{pH}$.

**Step 3 — Substitute and evaluate:** $\text{pOH} = 14 - 4 = 10$.

**Why each other option is wrong:**

- (B) $4$ simply repeats the given pH; pH and pOH are equal only at the neutral point (both $7$).
- (C) $7$ is the pOH of a neutral solution ($\text{pH} = 7$), not this acidic one.
- (D) $14$ would correspond to a pH of $0$, a far more acidic solution.

**Key point:** Always use $\text{pH} + \text{pOH} = 14$ at $25°$C. A pH of $4$ (acidic) gives a high pOH of $10$, consistent with a low hydroxide concentration.

**Final Answer:** $\text{pOH} = 10 \Rightarrow \boxed{\text{A}}$

**Answer: (A)** Go Back to Q13

## Q14.

### Solution

**Concept — Lewis acids and bases:** In the Lewis theory, an acid is an electron-pair *acceptor* and a base is an electron-pair *donor*. A Lewis acid usually has an incomplete octet or an empty orbital that can receive a lone pair. Boron trifluoride, $BF_3$, is the classic example: boron has only six electrons around it and an empty $2p$ orbital ready to accept a pair.

**Step 1 — Look for the electron-pair acceptor:** $BF_3$ has an electron-deficient boron (only $6$ electrons), so it can accept a lone pair and act as a Lewis acid.

**Step 2 — Check the others as donors:** $NH_3$, $H_2O$, and $OH^-$ all carry lone pairs that they can donate, making them Lewis *bases*, not acids.

**Step 3 — Conclusion:** only $BF_3$ behaves as a Lewis acid. In fact $BF_3$ reacts with

$NH_3$ to form the adduct $F_3B{\leftarrow}NH_3$, the acid accepting the lone pair from the base.

**Why each other option is wrong:**

- (A) $NH_3$ donates the lone pair on nitrogen, so it is a Lewis base.
- (C) $H_2O$ donates a lone pair on oxygen, so it is a Lewis base.
- (D) $OH^-$ is electron rich and donates a lone pair, a strong Lewis base.

**Key point:** Lewis acid = electron-pair acceptor (often electron deficient, like $BF_3$ or $AlCl_3$); Lewis base = electron-pair donor (has a lone pair).

**Final Answer:** $BF_3 \Rightarrow$ B

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

## Q15.

### Solution

**Concept — Oxidation number rules:** In a neutral compound the oxidation numbers of all atoms sum to zero. Hydrogen is usually $+1$ and oxygen usually $-2$. The unknown oxidation number is then found algebraically by balancing the sum.

**Given:** $HNO_3$, a neutral molecule.

**Step 1 — Assign the known values:** H $= +1$; each of the three oxygens is $-2$, giving $3 \times (-2) = -6$. Let the oxidation number of nitrogen be $x$.

**Step 2 — Set the sum to zero:** $(+1) + x + (-6) = 0$, i.e. $1 + x - 6 = 0$.

**Step 3 — Solve:** $x - 5 = 0 \Rightarrow x = +5$.

**Why each other option is wrong:**

- (A) $+3$ is the oxidation number of N in nitrous acid, $HNO_2$.
- (B) $-3$ is N in ammonia, $NH_3$.
- (D) $+4$ is N in nitrogen dioxide, $NO_2$.

**Key point:** Nitrogen in $HNO_3$ is in its highest oxidation state, $+5$, which is why nitric acid is a strong oxidising agent. Always make the oxidation numbers sum to zero for a neutral molecule.

**Final Answer:** Oxidation number of N $= +5 \Rightarrow$ C

**Answer: (C)** Go Back to Q15

## Q16.

**Solution**

**Concept — Reducing agents in redox reactions:** In a redox reaction, the reducing agent is the substance that *reduces* another species by giving electrons to it. In doing so the reducing agent itself *loses* electrons and is therefore *oxidised*. Oxidation is loss of electrons and reduction is gain (the mnemonic OIL RIG: Oxidation Is Loss, Reduction Is Gain).

**Step 1 — Recall what a reducing agent does:** it supplies electrons to the other reactant, causing that reactant to be reduced.

**Step 2 — Determine what happens to the reducing agent itself:** by donating electrons it loses them, so its own oxidation number increases, meaning it is oxidised (as the figure shows, $X \rightarrow X^{+} + e^{-}$).

**Step 3 — Conclusion:** a reducing agent is a substance that loses electrons and is itself oxidised.

**Why each other option is wrong:**

- (A) Gaining electrons and being reduced describes an *oxidising* agent, the opposite role.
- (B) A species that neither gains nor loses electrons takes no part in the redox change.
- (C) Donating protons (not electrons) describes a Bronsted acid, not a reducing agent.

**Key point:** Reducing agent = electron donor = gets oxidised; oxidising agent = electron acceptor = gets reduced. The agent always undergoes the change opposite to what it causes.

**Final Answer:** Loses electrons and is itself oxidised $\Rightarrow$ $\boxed{\text{D}}$

**Answer: (D)** Go Back to Q16

## Q17.

**Solution**

**Concept — Faraday's law of electrolysis:** The charge needed to deposit a substance during electrolysis depends on the number of electrons in the half-reaction. The deposition of $1$ mole of a metal that needs $z$ electrons requires $z$ faradays of charge, where $1$ faraday $= 96500$ C. The total charge is $Q = n \times z \times F$, with $n$ the moles deposited.

**Given:** $n = 0.5$ mol of Ag; half-reaction $Ag^{+} + e^{-} \rightarrow$ Ag so $z = 1$; $F = 96500$ C mol$^{-1}$.

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

**Step 2 — Substitute the values:** $Q = 0.5 \times 1 \times 96500$ C.

**Step 3 — Evaluate:** $Q = 48250$ C.

**Why each other option is wrong:**

- (B) $96500$ C is the charge for $1$ full mole of silver, twice the amount asked.
- (C) $193000$ C would deposit $2$ moles of silver.
- (D) $24125$ C corresponds to $0.25$ mole, half the requested amount.

**Key point:** Silver needs only one electron per ion ($z = 1$), so $1$ mole needs $1$ faraday. For $0.5$ mole the charge is simply half a faraday, $48250$ C.

**Final Answer:** $48250$ C $\Rightarrow$ A

**Answer: (A)** Go Back to Q17

## Q18.

**Solution**

**Concept — Mole fraction:** The mole fraction of a component is a way of expressing concentration in terms of the relative *number of moles*. For a component A it is the number of moles of A divided by the total number of moles of all the components present: $x_A = \dfrac{n_A}{n_A + n_B + \cdots}$. It is a pure ratio with no units, and the mole fractions of all components in a mixture add up to $1$.

**Step 1 — Recall the defining ratio:** mole fraction of A $= \dfrac{\text{moles of A}}{\text{total moles of all components}}$.

**Step 2 — Note its properties:** it is dimensionless, always lies between $0$ and $1$, and $x_A + x_B + \cdots = 1$.

**Step 3 — Match to the options:** only the choice "moles of A divided by the total moles of all components" expresses this definition correctly.

**Why each other option is wrong:**

- (A) Moles of A per litre of solution is the definition of *molarity*, not mole fraction.
- (C) Moles of A per kilogram of solvent is the definition of *molality*.
- (D) Mass of A divided by total mass is the *mass fraction*, which uses masses rather than moles.

**Key point:** Mole fraction uses moles in both numerator and denominator and is unitless. Do not confuse it with molarity (moles per litre) or molality (moles per kg of solvent).

**Final Answer:** Moles of A $\div$ total moles $\Rightarrow$ $\boxed{\text{B}}$

**Answer: (B)** Go Back to Q18

## Q19.

### Solution

**Concept — Gay-Lussac's law:** For a fixed mass of gas kept at *constant volume*, the pressure is directly proportional to the absolute (Kelvin) temperature: $P \propto T$, or equivalently $\frac{P}{T} =$ constant. This means a graph of $P$ against $T$ (in kelvin) is a straight line passing through the origin, exactly as shown.

**Step 1 — State the law:** at constant volume, $P \propto T$, so $\frac{P_1}{T_1} = \frac{P_2}{T_2}$.

**Step 2 — Interpret the graph:** a straight line through the origin means the two quantities are directly proportional; doubling the absolute temperature doubles the pressure.

**Step 3 — Conclusion:** pressure is directly proportional to absolute temperature at constant volume.

**Why each other option is wrong:**

- (A) Inverse proportionality ($P \propto 1/T$) would give a curve falling toward the axes, not a rising straight line; it also describes no real gas law for $P$ and $T$.
- (B) "Independent of temperature" contradicts the sloping line in the graph.
- (D) Proportionality to $T^2$ would give a curve bending upward, not a straight line.

**Key point:** Gay-Lussac's law ($P \propto T$ at constant $V$) must use *absolute* (Kelvin) temperature; using Celsius would not give a line through the origin. This law explains why a sealed can heats to a higher pressure.

**Final Answer:** Directly proportional to absolute temperature $\Rightarrow$ C

**Answer: (C)** Go Back to Q19

## Q20.

### Solution

**Concept — Oxidation state in a complex:** The oxidation state of the central metal in a coordination compound is found by setting the sum of the metal's oxidation state and the charges of all the ligands equal to the overall charge of the complex. Neutral ligands contribute zero charge.

**Given:** $[Ni(CO)_4]$, an overall neutral complex; CO is a neutral ligand (charge $0$).

**Step 1 — Note the ligand charge:** each carbon monoxide molecule is neutral, so four of them contribute $4 \times 0 = 0$.

**Step 2 — Balance the charges:** let the oxidation state of Ni be $x$. Then $x+4(0) = 0$ (the complex is neutral).

**Step 3 — Solve:** $x = 0$, so nickel is in the zero oxidation state.

**Why each other option is wrong:**

- (A) $+4$ would require each CO to carry a $-1$ charge, which it does not.
- (B) $+2$ is the common oxidation state of nickel in salts such as $NiCl_2$, but not in this neutral carbonyl.
- (C) $+1$ does not balance the neutral complex with neutral ligands.

**Key point:** Because CO is a neutral ligand, the metal in a neutral carbonyl such as $[Ni(CO)_4]$ has an oxidation state of $0$. Always check the ligand charge first.

**Final Answer:** Oxidation state of Ni $= 0 \Rightarrow$ D

**Answer: (D)** Go Back to Q20

## Q21.

**Solution**

**Concept — Denticity of ligands:** The denticity of a ligand is the number of donor atoms through which it binds to the central metal. A monodentate ligand binds through one atom; a bidentate ligand binds through two donor atoms, forming a ring (chelate) with the metal. Ethylenediamine ($H_2N$–$CH_2$–$CH_2$–$NH_2$) has two nitrogen atoms, each carrying a lone pair, so it can attach to the metal at two points.

**Step 1 — Identify the donor atoms in ethylenediamine:** the two terminal $-NH_2$ groups each donate a lone pair from nitrogen, giving two points of attachment.

**Step 2 — Classify it:** two donor atoms $\Rightarrow$ bidentate; it wraps around the metal to form a five-membered chelate ring, as drawn.

**Step 3 — Compare the others:** $NH_3$, $Cl^-$, and $H_2O$ each bind through only one donor atom, so they are monodentate.

**Why each other option is wrong:**

- (B) $NH_3$ donates one lone pair from a single nitrogen, so it is monodentate.
- (C) $Cl^-$ binds through the single chloride ion, so it is monodentate.
- (D) $H_2O$ donates one lone pair from oxygen, so it is monodentate.

**Key point:** Count the donor atoms to find denticity. Ethylenediamine (en) is the standard bidentate ligand, forming stable chelate rings.

**Final Answer:** Ethylenediamine (en) $\Rightarrow$ $\boxed{\text{A}}$

**Answer: (A)** Go Back to Q21

## Q22.

**Solution**

**Concept — Transition metals:** A transition metal is a d-block element whose atom or one of its common ions has a partly filled d subshell. They occupy groups 3 to 12 of the periodic table and are known for variable oxidation states, coloured ions, and catalytic behaviour. Iron ($Z = 26$, configuration $[\text{Ar}]3d^6\,4s^2$) is a classic transition metal.

**Step 1 — Test each element's block:** iron lies in the d block with a partly filled $3d$ subshell, so it is a transition metal.

**Step 2 — Eliminate the others:** calcium ($[\text{Ar}]4s^2$) is an s-block (group 2) metal;

sodium ($[Ne]3s^1$) is an s-block (group 1) metal; aluminium ($[Ne]3s^2\,3p^1$) is a p-block (group 13) metal.

**Step 3 — Conclusion:** only iron is a transition (d-block) metal.

**Why each other option is wrong:**

- (A) Calcium is an alkaline-earth (s-block) metal with no partly filled d subshell.
- (C) Sodium is an alkali (s-block) metal.
- (D) Aluminium is a p-block metal, not a transition element.

**Key point:** Transition metals are d-block elements (groups 3 to 12) with partly filled d orbitals. Iron, copper, and zinc lie here, whereas Na, Ca, and Al do not.

**Final Answer:** Iron $\Rightarrow$ $\boxed{\text{B}}$

**Answer: (B)** Go Back to Q22

## Q23.

### Solution

**Concept — Functional isomerism:** Isomers are different compounds with the same molecular formula. *Functional* isomers have the same molecular formula but contain different functional groups, so they belong to different classes of compound. The classic example is ethanol and dimethyl ether, both $C_2H_6O$.

**Step 1 — Compare the molecular formulas:** ethanol ($CH_3CH_2OH$) and dimethyl ether ($CH_3OCH_3$) both have the formula $C_2H_6O$.

**Step 2 — Compare the functional groups:** ethanol contains an alcohol group ($-OH$), while dimethyl ether contains an ether linkage ($-O-$). Same formula, different functional groups, so they are functional isomers.

**Step 3 — Conclusion:** ethanol and dimethyl ether are functional isomers.

**Why each other option is wrong:**

- (A) Ethane ($C_2H_6$) and propane ($C_3H_8$) have different molecular formulas, so they are not isomers at all.
- (B) $n$-butane and isobutane are *chain* isomers (same group, different skeleton), not functional isomers.
- (D) Ethene ($C_2H_4$) and ethyne ($C_2H_2$) have different molecular formulas, so they are not isomers.

**Key point:** Functional isomers share a molecular formula but differ in functional group (e.g. alcohol vs ether). Same formula but different skeleton is chain isomerism instead.

**Final Answer:** Ethanol and dimethyl ether $\Rightarrow$ C

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

## Q24.

### Solution

**Concept — IUPAC naming of alkanes:** In the IUPAC system, a saturated hydrocarbon (alkane) is named by a root that indicates the number of carbon atoms followed by the suffix *-ane*. One carbon gives the root "meth-", so a one-carbon alkane is "methane". $CH_4$ has a single carbon bonded to four hydrogens.

**Step 1 — Count the carbon atoms:** $CH_4$ has exactly one carbon atom, so the root is "meth-".

**Step 2 — Add the alkane suffix:** since all four bonds are single C–H bonds, the compound is a saturated alkane, taking the suffix "-ane".

**Step 3 — Combine:** "meth-" + "-ane" = methane.

**Why each other option is wrong:**

- (A) Methanol is $CH_3OH$, which contains an extra oxygen ($-OH$ group).
- (B) Ethane is $C_2H_6$, a two-carbon alkane, not one carbon.
- (C) Methanal is HCHO, an aldehyde with a C=O group, not $CH_4$.

**Key point:** The root tells the carbon count (meth $= 1$, eth $= 2$, prop $= 3$, but $= 4$) and the suffix tells the family (-ane for alkanes). $CH_4$ is methane, the simplest alkane.

**Final Answer:** Methane $\Rightarrow$ D

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

## Q25.

**Solution**

**Concept — General formula of alcohols:** A homologous series is a family of compounds with the same functional group and a fixed general formula, each member differing from the next by a $CH_2$ unit. Saturated aliphatic alcohols are formed by replacing one hydrogen of an alkane ($C_nH_{2n+2}$) with a hydroxyl group ($-OH$), giving the general formula $C_nH_{2n+1}OH$.

**Step 1 — Start from the alkane formula:** an alkane is $C_nH_{2n+2}$.

**Step 2 — Replace one H with $-OH$:** removing one H and adding $-OH$ converts $C_nH_{2n+2}$ into $C_nH_{2n+1}OH$ (the alkyl group $C_nH_{2n+1}$ bonded to $-OH$).

**Step 3 — Check with an example:** for $n = 2$, $C_2H_5OH$ is ethanol, which fits perfectly.

**Why each other option is wrong:**

- (B) $C_nH_{2n}O$ is a general formula that also covers ethers and aldehydes, so it is not specific to alcohols.
- (C) $C_nH_{2n}$ is the formula of alkenes (or cycloalkanes), with no oxygen.
- (D) $C_nH_{2n-1}OH$ corresponds to an unsaturated alcohol, not the saturated series.

**Key point:** A saturated alcohol is an alkyl group plus $-OH$: $C_nH_{2n+1}OH$. Test it with methanol ($CH_3OH$, $n = 1$) and ethanol ($C_2H_5OH$, $n = 2$).

**Final Answer:** $C_nH_{2n+1}OH \Rightarrow$ A

**Answer: (A)** Go Back to Q25

## Q26.

**Solution**

**Concept — Saturated vs unsaturated hydrocarbons:** A saturated hydrocarbon contains only carbon–carbon single bonds (alkanes), whereas an unsaturated hydrocarbon contains at least one carbon–carbon double bond (alkenes) or triple bond (alkynes). The presence of a multiple bond is what makes a hydrocarbon unsaturated.

**Step 1 — Examine each option for multiple bonds:** methane ($CH_4$), ethane ($C_2H_6$), and propane ($C_3H_8$) contain only single bonds, so they are saturated alkanes.

**Step 2 — Identify the unsaturated one:** ethene ($C_2H_4$) contains a carbon–carbon double bond (C=C), as drawn, so it is unsaturated.

**Step 3 — Conclusion:** ethene is the unsaturated hydrocarbon among the choices.

**Why each other option is wrong:**

- (A) Methane has only C–H single bonds: saturated.
- (B) Ethane ($CH_3$–$CH_3$) has a single C–C bond: saturated.
- (C) Propane has only single bonds: saturated.

**Key point:** Look for a C=C or C≡C bond to spot an unsaturated hydrocarbon. Alkenes and alkynes are unsaturated; alkanes (single bonds only) are saturated.

**Final Answer:** Ethene ($C_2H_4$) $\Rightarrow$ D

**Answer: (D)** Go Back to Q26

## Q27.

### Solution

**Concept — Counting sigma bonds:** A sigma ($\sigma$) bond is a single bond formed by the head-on overlap of orbitals; every single covalent bond is a sigma bond. To count the sigma bonds in a molecule, simply count every bond line in its structure (a double bond has one $\sigma$ and one $\pi$, but ethane has no multiple bonds). Ethane is $CH_3$–$CH_3$.

**Given:** ethane $C_2H_6$, with six C–H bonds and one C–C bond.

**Step 1 — Count the C–H sigma bonds:** each carbon is bonded to three hydrogens, so $2 \times 3 = 6$ C–H bonds, all sigma.

**Step 2 — Add the C–C sigma bond:** the two carbons are joined by one single bond, which is one more sigma bond.

**Step 3 — Total:** $6 + 1 = 7$ sigma bonds.

**Why each other option is wrong:**

- (A) $6$ counts only the C–H bonds and forgets the central C–C bond.
- (C) $8$ over-counts, perhaps by adding an imaginary extra bond.
- (D) $5$ under-counts the C–H bonds.

**Key point:** Every single bond is one sigma bond, so just count the lines: $6$ C–H plus $1$ C–C gives $7$ sigma bonds in ethane (and it has no $\pi$ bonds at all).

**Final Answer:** 7 sigma bonds $\Rightarrow$ B

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

## Q28.

### Solution

**Concept — The carboxyl group:** A functional group is the reactive part of a molecule that defines its class. The group shown, a carbon double-bonded to one oxygen (a carbonyl, C=O) and single-bonded to a hydroxyl group (−OH) on the *same* carbon, is the carboxyl group, written −COOH. It is the defining group of carboxylic acids.

**Step 1 — Read the structure:** the carbon carries a C=O and an −OH together, plus one bond to the rest of the molecule (R), i.e. R–COOH.

**Step 2 — Match the group to its class:** a carbon bearing both C=O and −OH on the same carbon is the carboxyl group, characteristic of carboxylic acids such as acetic acid ($CH_3COOH$).

**Step 3 — Conclusion:** the −COOH group belongs to carboxylic acids.

**Why each other option is wrong:**

- (A) Alcohols have only the −OH group (no carbonyl).
- (B) Aldehydes have −CHO (a carbonyl with an H, but no extra −OH).
- (D) Ketones have a carbonyl bonded to two carbons (>C=O), with no −OH.

**Key point:** −COOH = carbonyl + hydroxyl on one carbon = carboxylic acid. The releasable −OH proton is what makes these compounds acidic.

**Final Answer:** Carboxylic acids $\Rightarrow$ C

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

## Q29.

### Solution

**Concept — Classes of carbohydrates:** Carbohydrates are classified by how many sugar units they contain. Monosaccharides are single sugar units (such as glucose and fructose); disaccharides are two units joined together (such as sucrose); polysaccharides are long polymers of many monosaccharide units. Starch is a polysaccharide made of many glucose units.

**Step 1 — Classify each option:** glucose and fructose are monosaccharides (single units); sucrose is a disaccharide (glucose + fructose); starch is a polysaccharide.

**Step 2 — Identify the polysaccharide:** starch is a polymer of a large number of glucose molecules joined by glycosidic bonds, so it is a polysaccharide.

**Step 3 — Conclusion:** among the choices, starch is the polysaccharide.

**Why each other option is wrong:**

- (B) Glucose is a single sugar unit, i.e. a monosaccharide.
- (C) Fructose is also a monosaccharide.
- (D) Sucrose (table sugar) is a disaccharide of glucose and fructose, not a polysaccharide.

**Key point:** Polysaccharides (starch, cellulose, glycogen) are long polymers of many monosaccharide units. Glucose and fructose are single units; sucrose is just two units joined.

**Final Answer:** Starch $\Rightarrow$ $\boxed{\text{A}}$

**Answer: (A)** Go Back to Q29

## Q30.

### Solution

**Concept — Basicity of amines:** An amine contains the amino group ($-NH_2$), in which the nitrogen atom carries a lone pair of electrons. Because this lone pair can be donated to a proton, an amine accepts $H^+$ and therefore behaves as a base (both in the Bronsted sense as a proton acceptor and in the Lewis sense as an electron-pair donor).

**Step 1 — Locate the reactive feature:** the nitrogen of the $-NH_2$ group has a lone pair available for donation.

**Step 2 — Show the basic behaviour:** this lone pair accepts a proton, for example $R–NH_2 + H^+ \rightarrow R–NH_3^+$, forming a substituted ammonium ion.

**Step 3 — Conclusion:** a compound containing the $-NH_2$ group behaves as a base.

**Why each other option is wrong:**

- (A) A strong acid would *donate* protons; amines accept them instead, so this is wrong.

- (B) A neutral substance neither accepts nor donates protons, but the $-NH_2$ lone pair makes amines distinctly basic.
- (D) A salt is an ionic compound from an acid–base reaction; an amine itself is not a salt.

**Key point:** The lone pair on nitrogen makes the $-NH_2$ group a proton acceptor, so amines are bases. Contrast this with $-COOH$ (carboxylic acids), which donate protons and are acidic.

**Final Answer:** Base $\Rightarrow$ $\boxed{C}$

**Answer: (C)** Go Back to Q30

# Answer Key

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