AIIMS B.Sc Nursing Chemistry

Sample Paper – 3

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

Instructions

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

Q1. The number of atoms present in 1 mole of helium gas (He, a monoatomic gas) is (Avogadro number $= 6.022 \times 10^{23}$):

(A) 6.022×10^{23}

(B) 1.2044×10^{24}

(C) 3.011×10^{23}

(D) 6.022×10^{22}

Q2. The number of molecules present in 4 g of methane (CH_4, molar mass $= 16$ g mol^{-1}) is (Avogadro number $= 6.022 \times 10^{23}$):

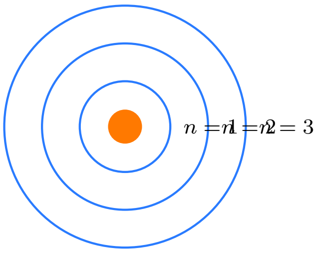
(A) 6.022×10^{23}

(B) 3.011×10^{23}

(C) 1.5055×10^{23}

(D) 2.4088×10^{24}

Q3. The shells of an atom are shown below. The number of orbitals in the third shell ($n = 3$), given by n^2, is:



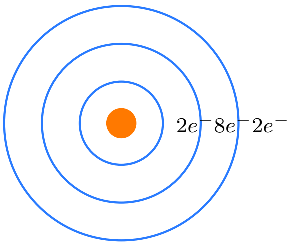
(A) 3

(B) 9

(C) 18

(D) 6

Q4. The electrons of a magnesium atom ($Z = 12$) are distributed in shells as shown. The correct ground-state electronic configuration (K, L, M shells) of magnesium is:



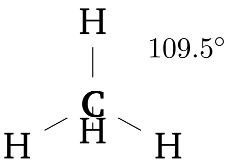
(A) 2, 2, 8

(B) 8, 2, 2

(C) 2, 10

(D) 2, 8, 2

Q5. The structure of a methane (CH_4) molecule, in which carbon forms four equivalent C–H bonds with no lone pairs, is shown. The shape of the molecule is:



(A) tetrahedral

(B) trigonal pyramidal

(C) square planar

(D) trigonal planar

Q6. The hybridization of the central carbon atom in a carbon dioxide molecule (CO_2, structure O=C=O) is:

(A) sp^2

(B) sp

(C) sp^3

(D) sp^3d

Q7. Which one of the following compounds contains a covalent bond?

(A) NaCl

(B) KBr

(C) CH_4

(D) MgO

Q8. On moving from top to bottom down a group of the periodic table, the atomic radius of the elements generally:

(A) decreases

(B) remains constant

(C) first decreases, then increases

(D) increases

Q9. Among the following elements, which one has the lowest first ionization energy?

(A) Caesium (Cs)

(B) Lithium (Li)

(C) Sodium (Na)

(D) Potassium (K)

Q10. The first law of thermodynamics is a statement of the law of conservation of energy. It is expressed as $\Delta U = q + W$, which states that:

(A) energy can be created but not destroyed

(B) energy is neither created nor destroyed, only converted from one form to another

(C) the entropy of the universe always decreases

(D) heat always flows from a colder body to a hotter body

Q11. For a chemical reaction that has reached a state of equilibrium at constant temperature and pressure, the value of the Gibbs free energy change ΔG is:

(A) positive

(B) negative

(C) zero

(D) infinite

Q12. When a catalyst is added to a reversible reaction that is at equilibrium, the position of the equilibrium:

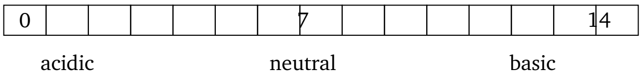
(A) shifts towards the products

(B) shifts towards the reactants

(C) is destroyed permanently

(D) does not change at all

Q13. The pH scale at 25°C is shown below. The pH of a neutral aqueous solution at 25°C is:



(A) 7

(B) 0

(C) 14

(D) 1

Q14. According to the Bronsted–Lowry theory, the conjugate base of hydrochloric acid (HCl) is:

(A) H^+

(B) Cl^-

(C) H_2Cl^+

(D) OH^-

Q15. The oxidation number of chromium in potassium dichromate ($K_2Cr_2O_7$) is:

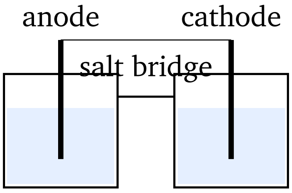
(A) $+3$

(B) $+7$

(C) $+6$

(D) $+2$

Q16. A galvanic cell with two electrodes joined by a salt bridge is shown below. In any redox reaction, the oxidising agent is the species that is itself:



(A) oxidised

(B) evaporated

(C) neutralised

(D) reduced

Q17. One faraday (96500 C) is the magnitude of the electric charge carried by one mole of:

(A) electrons

(B) protons and neutrons together

(C) atoms

(D) neutrons

Q18. Molality of a solution is defined as the number of moles of solute dissolved per:

(A) litre of solution

(B) kilogram of solvent

(C) litre of solvent

(D) kilogram of solution

Q19. For an ideal gas, $PV = nRT$. A gas occupies a volume of $V = 24.6$ L at a pressure of $P = 1$ atm and a temperature of $T = 300$ K (gas constant $R = 0.0821$ L atm $\text{K}^{-1}\text{mol}^{-1}$). The number of moles n of the gas is:

(A) 0.5 mol

(B) 2 mol

(C) 1 mol

(D) 24.6 mol

Q20. The oxidation state of cobalt in the complex ion $[\text{Co}(\text{NH}_3)_6]^{3+}$ (ammonia, NH_3, is a neutral ligand) is:

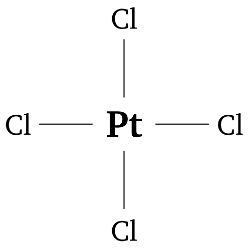
(A) 0

(B) $+2$

(C) $+6$

(D) $+3$

Q21. The complex ion $[PtCl_4]^{2-}$ has the arrangement of four chloride ligands (Cl) around the central platinum shown below. The coordination number of platinum in this complex is:



(A) 4
(B) 2
(C) 6
(D) 3

Q22. Which one of the following elements is a halogen (group 17)?

(A) Neon
(B) Bromine
(C) Calcium
(D) Sulphur

Q23. The number of structural isomers possible for propane, C_3H_8, is:

(A) 3
(B) 4
(C) 2
(D) 1

Q24. The IUPAC name of the compound acetic acid, CH_3COOH, is:

(A) methanoic acid
(B) ethanol
(C) ethanoic acid

(D) propanoic acid

Q25. The general molecular formula of the alkyne homologous series (hydrocarbons containing one carbon–carbon triple bond) is:

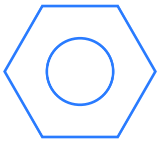
(A) C_nH_{2n-2}

(B) C_nH_{2n}

(C) C_nH_{2n+2}

(D) C_nH_n

Q26. The aromatic hydrocarbon benzene has the ring skeleton shown below, with delocalised electrons represented by the inner circle. The molecular formula of benzene is:



(A) C_6H_{12}

(B) C_6H_6

(C) C_5H_5

(D) C_6H_{14}

Q27. The number of pi (π) bonds present in one molecule of ethyne (acetylene, C_2H_2, which contains a $C\equiv C$ triple bond) is:

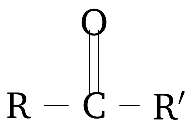
(A) 1

(B) 3

(C) 2

(D) 0

Q28. The functional group shown below, with a carbon double-bonded to oxygen and bonded to two other carbon atoms (R and R′), is characteristic of which class of compounds?



$$R-\overset{\overset{\displaystyle O}{\|}}{C}-R'$$

(>C=O carbonyl)

(A) alcohols

(B) aldehydes

(C) carboxylic acids

(D) ketones

Q29. The building-block monomer units that make up nucleic acids (DNA and RNA) are:

(A) nucleotides

(B) amino acids

(C) monosaccharides

(D) fatty acids

Q30. The acid responsible for the sour taste and smell of vinegar is:

(A) citric acid

(B) formic acid

(C) acetic acid

(D) hydrochloric acid

Detailed Solutions

Q1.

Solution

Concept — The mole and Avogadro's number: One mole of any substance contains Avogadro's number of elementary particles, $N_A = 6.022 \times 10^{23}$. The number of particles equals the number of moles times N_A: number of particles $= n \times N_A$. Helium is a *monoatomic* gas, so each helium particle is a single atom; one molecule of helium is the same as one atom of helium.

Given: $n = 1$ mol of He; $N_A = 6.022 \times 10^{23}$.

Step 1 — Write the formula: number of atoms $= n \times N_A$.

Step 2 — Substitute the data: $= 1 \times 6.022 \times 10^{23}$.

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

Why each other option is wrong:

- (B) 1.2044×10^{24} is $2\,N_A$; this would be right only if each particle had two atoms (a diatomic gas), but helium is monoatomic.
- (C) 3.011×10^{23} is $0.5\,N_A$, the count in only half a mole.
- (D) 6.022×10^{22} is $0.1\,N_A$, ten times too small.

Key point: Because helium is monoatomic, the number of atoms in 1 mole equals exactly N_A. For a diatomic gas like H_2 or O_2, 1 mole would contain N_A molecules but $2N_A$ atoms.

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

Answer: (A) Go Back to Q1

Q2.

Solution

Concept — From mass to number of molecules: To count molecules from a mass, first convert mass to moles using $n = \frac{m}{M}$ (where m is the given mass and M the molar mass), then multiply by Avogadro's number N_A. The molar mass of methane is $M(CH_4) = 12 + 4 \times 1 = 16$ g mol^{-1}.

Given: mass $m = 4$ g; $M(CH_4) = 16$ g mol^{-1}; $N_A = 6.022 \times 10^{23}$.

Step 1 — Find the moles: $n = \dfrac{m}{M} = \dfrac{4\text{ g}}{16\text{ g mol}^{-1}} = 0.25$ mol.

Step 2 — Convert moles to molecules: number of molecules $= n \times N_A = 0.25 \times 6.022 \times 10^{23}$.

Step 3 — Evaluate: $= 1.5055 \times 10^{23}$ molecules.

Why each other option is wrong:

- (A) 6.022×10^{23} is the count in 1 full mole (16 g), not 4 g.
- (B) 3.011×10^{23} is for 0.5 mol (8 g of methane).
- (D) 2.4088×10^{24} is for 4 mol (64 g), wrongly dividing 16 by 4 instead of 4 by 16.

Key point: Always convert mass to moles first, then multiply by N_A. Here 4 g of CH_4 is one-quarter of a mole, so it has one-quarter of N_A molecules.

Final Answer: 1.5055×10^{23} molecules $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q2

Q3.

Solution

Concept — Number of orbitals in a shell: A shell with principal quantum number n contains exactly n^2 orbitals. This follows from the subshells present: for a given n the azimuthal quantum number l runs from 0 to $n-1$, and each subshell l has $2l+1$ orbitals; summing these gives n^2. Each orbital can hold two electrons, so the electron capacity of the shell is $2n^2$.

Given: third shell, $n = 3$.

Step 1 — Write the formula: number of orbitals $= n^2$.

Step 2 — Substitute: $= (3)^2$.

Step 3 — Evaluate (and check by subshells): $= 9$ orbitals. Indeed $n = 3$ has one $3s$ (1), three $3p$ (3), and five $3d$ (5) orbitals, $1 + 3 + 5 = 9$.

Why each other option is wrong:

- (A) 3 counts only the three $3p$ orbitals, ignoring $3s$ and $3d$.
- (C) 18 is the maximum number of *electrons* ($2n^2$), not orbitals.
- (D) 6 has no basis; it is neither n^2 nor $2n^2$.

Key point: Orbitals in a shell $= n^2$; electrons in a shell $= 2n^2$. For $n = 3$ that is 9 orbitals and 18 electrons. Do not confuse the orbital count with the electron count.

Final Answer: 9 orbitals $\Rightarrow \boxed{\text{B}}$

$\boxed{\textbf{Answer: (B)}}$ **Go Back to Q3**

Q4.

Solution

Concept — Bohr–Bury shell filling: Electrons fill shells outward from the nucleus (K, L, M, N $\ldots$), with the maximum capacity of each shell limited to $2n^2$ (K$=2$, L$=8$, M$=18$). For a magnesium atom, $Z=12$, so there are 12 electrons to place, starting from the innermost shell.

Given: magnesium, $Z=12$ (so 12 electrons).

Step 1 — Fill the K shell ($n=1$): maximum 2 electrons $\rightarrow$ K holds 2. (10 electrons remain.)

Step 2 — Fill the L shell ($n=2$): maximum 8 electrons $\rightarrow$ L holds 8. (2 electrons remain.)

Step 3 — Fill the M shell ($n=3$): the last 2 electrons go into M. The distribution is therefore $2, 8, 2$ (equivalently $1s^2\,2s^2\,2p^6\,3s^2$).

Why each other option is wrong:

- (A) $2, 2, 8$ wrongly puts only 2 in the L shell and overfills the M shell out of order.
- (B) $8, 2, 2$ reverses the filling order; the innermost K shell can hold only 2, not 8.
- (C) $2, 10$ overfills the L shell, which can hold at most 8 electrons.

Key point: Fill inner shells first and respect the $2n^2$ limit. Magnesium's outermost shell has 2 electrons, which is why it is a group-2 metal that forms Mg^{2+} ions.

Final Answer: $2, 8, 2 \Rightarrow \boxed{\text{D}}$

$\boxed{\textbf{Answer: (D)}}$ **Go Back to Q4**

Q5.

Solution

Concept — VSEPR shape of CH_4: The shape of a molecule is set by the number of electron domains around the central atom, which spread out as far apart as possible. In methane, carbon has 4 valence electrons, all of which form bonds to four hydrogens, giving four bonding pairs and *no* lone pairs, i.e. four electron domains (sp^3 hybridisation).

Step 1 — Count the electron domains on carbon: four C–H bonds, no lone pairs $\Rightarrow 4$ domains.

Step 2 — Electron geometry: four domains arrange themselves tetrahedrally, pointing to the corners of a regular tetrahedron.

Step 3 — Molecular shape: since all four domains are bonds (no lone pairs to distort it), the molecular shape is also tetrahedral, with H–C–H bond angles of $109.5°$.

Why each other option is wrong:

- (B) Trigonal pyramidal needs three bonds and one lone pair, as in NH_3.
- (C) Square planar would require four bonds plus two lone pairs (AB_4E_2), not the case here.
- (D) Trigonal planar needs only three bonding domains, as in BF_3.

Key point: Four bonding pairs and no lone pairs always give a perfect tetrahedron ($109.5°$). Methane is the textbook example of a tetrahedral molecule.

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

Answer: (A) Go Back to Q5

Q6.

Solution

Concept — Hybridisation from σ bonds and lone pairs: The hybridisation of an atom equals its number of σ bonds plus lone pairs (its number of electron domains). A double bond counts as *one* σ bond, since the second bond is a π bond that does not change the hybridisation. In CO_2, the central carbon is double-bonded to each oxygen, O=C=O.

Step 1 — Count the domains on carbon: two C=O double bonds count as two σ bonds, and carbon has no lone pairs $\Rightarrow 2$ electron domains.

Step 2 — Assign the hybridisation: two domains $\Rightarrow sp$. The two sp orbitals point in opposite directions at $180°$, and the two leftover unhybridised p orbitals form the two π bonds.

Step 3 — Geometry: the molecule is linear, O$=$C$=$O at $180°$.

Why each other option is wrong:

- (A) sp^2 (three domains) describes a carbon with three σ groups, as in ethene or a carbonate ion.
- (C) sp^3 (four domains) is for four single bonds, as in methane.
- (D) sp^3d (five domains) cannot occur for carbon, which has no available d orbitals.

Key point: Count each double bond as one σ bond. Two such bonds on carbon give two domains, hence sp and a linear shape. This is why CO_2 is linear and non-polar.

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

$\boxed{\textbf{Answer: (B)}}$ Go Back to Q6

Q7.

Solution

Concept — Ionic versus covalent bonding: A covalent bond forms by the *sharing* of electron pairs between atoms, typically between two non-metals with a small electronegativity difference. An ionic bond forms by the *transfer* of electrons from a metal to a non-metal, producing oppositely charged ions held by electrostatic attraction.

Step 1 — Classify each compound: NaCl (metal + non-metal), KBr (metal + non-metal), and MgO (metal + non-metal) are all ionic compounds formed by electron transfer.

Step 2 — Examine the remaining option: CH_4 is made of two non-metals (carbon and hydrogen) that share electrons, so its C–H bonds are covalent.

Step 3 — Conclusion: only CH_4 contains a covalent bond.

Why each other option is wrong:

- (A) NaCl is ionic: Na transfers an electron to Cl, forming Na^+ and Cl^-.
- (B) KBr is ionic, formed from K^+ and Br^-.
- (D) MgO is ionic, formed from Mg^{2+} and O^{2-}.

Key point: Metal + non-metal usually gives an ionic bond (electron transfer); non-metal + non-metal gives a covalent bond (electron sharing). CH_4 is purely covalent.

Final Answer: $CH_4 \Rightarrow \boxed{C}$

Answer: (C) Go Back to Q7

Q8.

Solution

Concept — Atomic radius down a group: The atomic radius is roughly the distance from the nucleus to the outermost electron shell. On going down a group, each successive element has its valence electrons in a *new, higher* principal shell, so the outermost electrons are farther from the nucleus. Although the nuclear charge also rises, the added inner shells shield the outer electrons, so the size effect of new shells dominates.

Step 1 — Identify what changes down a group: the number of occupied shells increases by one with each step down (for example Li, Na, K each add a shell).

Step 2 — Net effect: adding shells places the valence electrons progressively farther out, and inner-shell shielding offsets the rising nuclear charge.

Step 3 — State the trend: the atomic radius therefore increases down a group.

Why each other option is wrong:

- (A) Decreasing is the trend *across a period*, not down a group.
- (B) Remaining constant ignores the new shells added down a group.
- (C) There is no "decrease then increase" reversal down a normal group.

Key point: Down a group the radius increases (added shells dominate); across a period it decreases (rising nuclear charge dominates). Keep the two trends separate.

Final Answer: Increases $\Rightarrow \boxed{D}$

Answer: (D) Go Back to Q8

Q9.

Solution

Concept — Trend in ionization energy: Ionization energy is the minimum energy needed to remove the outermost electron from a gaseous atom. It *decreases* down a group (the outer electron is farther from the nucleus and more shielded, so it is easier to remove) and *increases* across a period. The element with the lowest ionization energy is therefore the one lowest in its group among the choices.

Step 1 — Note all options are group-1 alkali metals: Li, Na, K, Cs, listed from top to bottom of group 1.

Step 2 — Apply the group trend: going down group 1, the valence electron lies in an ever higher shell, so it is held more loosely; ionization energy falls: Li $>$ Na $>$ K $>$ Cs.

Step 3 — Conclusion: caesium, being lowest in the group, has its single valence electron held most loosely and so has the lowest ionization energy.

Why each other option is wrong:

- (B) Lithium is at the top of group 1, so it holds its electron most tightly and has the *highest* ionization energy of these four.
- (C) Sodium lies above potassium and caesium, so its ionization energy is higher than theirs.
- (D) Potassium has a lower ionization energy than Na and Li, but caesium (still lower in the group) is lower still.

Key point: Ionization energy decreases down a group. Among these alkali metals the order is Cs $<$ K $<$ Na $<$ Li, so caesium loses its electron most easily.

Final Answer: Caesium (Cs) $\Rightarrow \boxed{\text{A}}$

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

Q10.

Solution

Concept — First law of thermodynamics: The first law is the law of conservation of energy applied to thermodynamic systems. It states that energy can be neither created nor destroyed, only transformed from one form to another or transferred between system and surroundings. Mathematically, $\Delta U = q + W$, where ΔU is the change in internal energy, q the heat added to the system, and W the work done on the system.

Step 1 — Recall the statement: the total energy of an isolated system is constant; any energy gained by the system is lost by the surroundings and vice versa.

Step 2 — Match to the options: the statement that exactly captures conservation of energy is "energy is neither created nor destroyed, only converted from one form to another."

Step 3 — Conclusion: option (B) is the correct statement of the first law.

Why each other option is wrong:

- (A) "Created but not destroyed" violates conservation; energy can be neither created nor destroyed.
- (C) The decrease of entropy is not part of the first law; the second law in fact says the entropy of the universe increases for a spontaneous process.
- (D) Heat flowing from cold to hot on its own contradicts the second law (the Clausius statement), not the first law.

Key point: First law = conservation of energy ($\Delta U = q + W$). Statements about entropy or the direction of heat flow belong to the *second* law, not the first.

Final Answer: Energy is neither created nor destroyed $\Rightarrow \boxed{\text{B}}$

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

Q11.

Solution

Concept — ΔG at equilibrium: The Gibbs free energy change ΔG measures the driving force of a reaction at constant temperature and pressure. A negative ΔG means the forward reaction is spontaneous, a positive ΔG means the reverse is spontaneous, and at equilibrium there is no net driving force in either direction. The relation $\Delta G = \Delta G^{\circ} + RT \ln Q$ becomes, at equilibrium, $\Delta G = 0$ (with $Q = K$).

Step 1 — State the condition for equilibrium: the system has no tendency to move forward or backward, so the free energy is at a minimum with respect to the extent of reaction.

Step 2 — Translate to ΔG: a minimum in free energy means the slope is zero, i.e. $\Delta G = 0$ at equilibrium.

Step 3 — Conclusion: for a reaction at equilibrium, $\Delta G = 0$.

Why each other option is wrong:

- (A) A positive ΔG means the forward reaction is non-spontaneous (it has not yet reached equilibrium).
- (B) A negative ΔG means the forward reaction is still spontaneous and proceeding, so equilibrium is not yet reached.
- (D) ΔG is a finite state-function difference; it never becomes infinite.

Key point: $\Delta G < 0$ (forward spontaneous), $\Delta G > 0$ (reverse spontaneous), $\Delta G = 0$ (equilibrium). At equilibrium the free energy is minimised and ΔG is exactly zero.

Final Answer: $\Delta G = 0 \Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q11

Q12.

Solution

Concept — Effect of a catalyst on equilibrium: A catalyst speeds up a reaction by providing an alternative pathway with a lower activation energy. It lowers the activation energy of the forward and the reverse reactions by the *same* amount, so it speeds both up equally. A catalyst therefore helps the system reach equilibrium faster but does not alter the equilibrium constant or the equilibrium position.

Step 1 — Recall what determines the equilibrium position: the position depends on the relative thermodynamic stabilities of reactants and products (through ΔG° and K), not on the rate of the reaction.

Step 2 — See what a catalyst affects: it changes only the rate (kinetics), lowering the activation barrier equally in both directions, so the forward and reverse rates are increased by the same factor.

Step 3 — Conclusion: since both rates are affected equally, the equilibrium position does not change at all; only the time taken to reach equilibrium is shortened.

Why each other option is wrong:

- (A) A shift towards products would require changing K or the conditions (concentration, pressure, temperature), which a catalyst does not do.
- (B) Likewise, a catalyst does not shift the equilibrium towards the reactants.
- (C) A catalyst is not consumed and certainly does not "destroy" the equilibrium.

Key point: A catalyst changes the rate, not the position, of equilibrium. It speeds up the approach to equilibrium but leaves the equilibrium amounts (and K) unchanged.

Final Answer: Does not change at all $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q12

Q13.

Solution

Concept — pH of a neutral solution: pH is defined as $\text{pH} = -\log_{10}[\text{H}^+]$. In pure water at $25°\text{C}$, the self-ionization gives equal concentrations of hydrogen and hydroxide ions: $[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7}$ M, because the ionic product $K_w = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$. A solution is neutral when $[\text{H}^+] = [\text{OH}^-]$.

Given: neutral aqueous solution at $25°\text{C}$, so $[\text{H}^+] = 1 \times 10^{-7}$ M.

Step 1 — Apply the pH formula: $\text{pH} = -\log_{10}(10^{-7})$.

Step 2 — Evaluate the logarithm: $\log_{10}(10^{-7}) = -7$.

Step 3 — Conclusion: $\text{pH} = -(-7) = 7$.

Why each other option is wrong:

- (B) $\text{pH} = 0$ means $[\text{H}^+] = 1$ M, a strongly acidic solution.
- (C) $\text{pH} = 14$ means $[\text{H}^+] = 10^{-14}$ M, a strongly basic solution.
- (D) $\text{pH} = 1$ means $[\text{H}^+] = 0.1$ M, an acidic solution, not neutral.

Key point: At $25°\text{C}$, neutral means $\text{pH} = 7$ (where $[\text{H}^+] = [\text{OH}^-]$). Below 7 is acidic and above 7 is basic. The neutral value of 7 is specific to $25°\text{C}$.

Final Answer: $\text{pH} = 7 \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q13

Q14.

Solution

Concept — Conjugate acid–base pairs: In Bronsted–Lowry theory an acid is a proton (H^+) donor and a base is a proton acceptor. When an acid loses a proton, the species left behind is its *conjugate base*. To find the conjugate base of HCl, remove one H^+ from it.

Step 1 — Write the ionization: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$.

Step 2 — Identify what remains after losing H^+: the species left is the chloride ion, Cl^-.

Step 3 — Conclusion: the conjugate base of HCl is Cl^-.

Why each other option is wrong:

- (A) H^+ is the proton that is *donated*, not the conjugate base that remains.
- (C) H_2Cl^+ would be the conjugate *acid* of HCl, formed by *adding* a proton (and is not a stable species).
- (D) OH^- is the conjugate base of water, not of HCl.

Key point: Conjugate base = acid minus one H^+. Since HCl is a strong acid, its conjugate base Cl^- is a very weak base (it has essentially no tendency to accept a proton back).

Final Answer: $Cl^- \Rightarrow \boxed{\text{B}}$

$\boxed{\textbf{Answer: (B)}}$ **Go Back to Q14**

Q15.

Solution

Concept — Oxidation number rules: In a neutral compound the oxidation numbers of all atoms sum to zero. A group-1 metal such as potassium is $+1$, and oxygen is usually -2. The unknown (chromium) is then found algebraically, remembering to multiply each oxidation number by the number of those atoms.

Given: $K_2Cr_2O_7$, a neutral compound.

Step 1 — Assign known values: two K at $+1$ give $2 \times (+1) = +2$; seven O at -2 give $7 \times (-2) = -14$. Let each Cr be x, and there are two Cr atoms.

Step 2 — Set the sum to zero: $(+2) + 2x + (-14) = 0$, i.e. $2x - 12 = 0$.

Step 3 — Solve: $2x = 12 \Rightarrow x = +6$.

Why each other option is wrong:

- (A) $+3$ is Cr in Cr^{3+} or Cr_2O_3, not in dichromate.
- (B) $+7$ would not balance the charges; it is the state of Mn in permanganate, not Cr here.
- (D) $+2$ is a rare low oxidation state of chromium, not the value in $K_2Cr_2O_7$.

Key point: Both chromium atoms are equivalent and each is $+6$. This high oxidation state is why dichromate is a strong oxidising agent. Always multiply oxidation numbers by the atom counts before summing.

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

Answer: (C) Go Back to Q15

Q16.

Solution

Concept — Oxidising agent in a redox reaction: In a redox reaction, oxidation is loss of electrons and reduction is gain of electrons. The *oxidising agent* is the species that causes another to be oxidised by accepting its electrons; in accepting electrons, the oxidising agent is *itself reduced*. Likewise the reducing agent is itself oxidised.

Step 1 — Recall the definitions: oxidation = loss of electrons (OIL), reduction = gain of electrons (RIG).

Step 2 — Trace the role of the oxidising agent: it removes electrons from the other reactant (oxidising it). By gaining those electrons, the oxidising agent undergoes reduction.

Step 3 — Conclusion: the oxidising agent is itself reduced during the reaction.

Why each other option is wrong:

- (A) "Oxidised" describes the reducing agent, which loses electrons; the oxidising agent is the opposite.
- (B) Evaporation is a physical change of state, not a redox process.
- (C) Neutralisation is an acid–base reaction, unrelated to the electron-transfer role of an oxidising agent.

Key point: The oxidising agent gains electrons and is reduced; the reducing agent loses electrons and is oxidised. A handy memory aid: the agent does the *opposite* of its name to itself.

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

Answer: (D) Go Back to Q16

Q17.

Solution

Concept — The faraday: One faraday (F) is defined as the magnitude of electric charge carried by one mole of electrons. Numerically $F = N_A \times e = (6.022 \times 10^{23}) \times (1.602 \times 10^{-19}\ \text{C}) \approx 96500\ \text{C mol}^{-1}$, where N_A is Avogadro's number and e is the elementary charge. It is the conversion factor between moles of electrons and charge in electrolysis.

Step 1 — Recall the definition: $1\ F$ is the charge on 1 mole of electrons.

Step 2 — Check the magnitude: $N_A \times e = 6.022 \times 10^{23} \times 1.602 \times 10^{-19}\ \text{C} \approx 96500\ \text{C}$, matching the stated value.

Step 3 — Conclusion: one faraday is the charge carried by one mole of electrons.

Why each other option is wrong:

- (B) The faraday refers to electrons (the charge carriers), not to a mole of protons and neutrons together.
- (C) A mole of neutral atoms carries no net charge, so it cannot define a unit of charge.
- (D) Neutrons are uncharged, so a mole of neutrons carries zero charge.

Key point: $1\ F = 96500\ \text{C} =$ charge of 1 mole of electrons. In electrolysis, the charge needed to deposit 1 mole of a metal ion M^{n+} is n faradays.

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

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

Q18.

Solution

Concept — Molality: Molality (m) is a concentration unit defined as the number of moles of solute dissolved per kilogram of *solvent*: $m = \dfrac{\text{moles of solute}}{\text{mass of solvent in kg}}$. Because it is based on the mass of solvent (which does not change with temperature), molality is independent of temperature, unlike molarity which uses volume.

Step 1 — State the formula: $m = \dfrac{n_{\text{solute}}}{\text{mass of solvent (kg)}}$.

Step 2 — Identify the denominator: the denominator is the mass of the *solvent* in kilograms, not the solution and not a volume.

Step 3 — Conclusion: molality is moles of solute per kilogram of solvent.

Why each other option is wrong:

- (A) Moles of solute per litre of *solution* is the definition of *molarity*, not molality.
- (C) Per litre of solvent is not a standard concentration unit; molality uses mass, not volume, of solvent.
- (D) Per kilogram of *solution* is incorrect; molality uses the mass of solvent only.

Key point: Molality uses mass of *solvent* (kg) in the denominator; molarity uses volume of *solution* (L). Because mass does not change with temperature, molality is temperature-independent.

Final Answer: Kilogram of solvent $\Rightarrow \boxed{\text{B}}$

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

Q19.

Solution

Concept — The ideal gas equation: An ideal gas obeys $PV = nRT$, relating pressure P, volume V, amount n (moles), gas constant R, and absolute temperature T. To find any one variable, rearrange the equation for that variable and substitute, keeping the units consistent (here L, atm, K, and $R = 0.0821$ L atm K^{-1}mol^{-1}).

Given: $P = 1$ atm, $V = 24.6$ L, $T = 300$ K, $R = 0.0821$ L atm K^{-1}mol^{-1}.

Step 1 — Rearrange for n: $n = \dfrac{PV}{RT}$.

Step 2 — Substitute with units: $n = \dfrac{(1\ \text{atm})(24.6\ \text{L})}{(0.0821\ \text{L atm K}^{-1}\text{mol}^{-1})(300\ \text{K})}$.

Step 3 — Evaluate: the denominator is $0.0821 \times 300 = 24.63$ L atm mol^{-1}, so $n = \dfrac{24.6}{24.63} \approx 1$ mol.

Why each other option is wrong:

- (A) 0.5 mol would occupy about 12.3 L under these conditions, half the given volume.
- (B) 2 mol would occupy about 49.2 L, twice the given volume.
- (D) 24.6 mol confuses the volume in litres with the number of moles, ignoring RT entirely.

Key point: At 1 atm and 300 K, the molar volume is about 24.6 L, so 24.6 L corresponds to 1 mole. Always check that P, V, T units match the value of R used.

Final Answer: $n = 1$ mol $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q19

Q20.

Solution

Concept — Oxidation state of the metal in a complex: In a coordination complex, the sum of the oxidation state of the central metal and the charges of all the ligands equals the overall charge of the complex ion. Neutral ligands contribute zero. Ammonia (NH_3) is a *neutral* ligand, so it contributes 0 to the charge balance.

Given: $[Co(NH_3)_6]^{3+}$; six neutral NH_3 ligands; overall charge $+3$.

Step 1 — Set up the charge balance: (oxidation state of Co) $+6 \times (0) = +3$.

Step 2 — Simplify: (oxidation state of Co) $+0 = +3$.

Step 3 — Solve: oxidation state of Co $= +3$.

Why each other option is wrong:

- (A) 0 would require the complex to be neutral, but it carries a $+3$ charge.
- (B) $+2$ would give an overall charge of $+2$ with neutral ligands, not $+3$.
- (C) $+6$ ignores that ammonia is neutral; it would wrongly hold only if the ligands carried $-\frac{1}{2}$ each.

Key point: For neutral ligands like NH_3 and H_2O, the metal oxidation state simply equals the overall charge of the complex ion. Here Co is $+3$.

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

Answer: (D) Go Back to Q20

Q21.

Solution

Concept — Coordination number: The coordination number of the central metal in a complex is the number of donor atoms (ligand atoms) directly bonded to it, i.e. the number of coordinate bonds it forms. It is read directly from the structure by counting the ligands attached to the metal.

Step 1 — Examine the complex $[PtCl_4]^{2-}$: platinum is bonded to four chloride ligands, each donating a lone pair to form one coordinate bond.

Step 2 — Count the donor atoms: there are four Cl ligands arranged around Pt (in a square planar geometry), so four coordinate bonds.

Step 3 — Conclusion: the coordination number of platinum is 4.

Why each other option is wrong:

- (B) 2 would mean only two ligands; the formula clearly shows four chlorides.
- (C) 6 is the coordination number of an octahedral complex such as $[\mathrm{PtCl_6}]^{2-}$, not $[\mathrm{PtCl_4}]^{2-}$.
- (D) 3 does not match the four ligands present.

Key point: The coordination number equals the number of ligand donor atoms bonded to the metal. With four chloride ligands, $[\mathrm{PtCl_4}]^{2-}$ has a coordination number of 4 and adopts a square planar shape.

Final Answer: $4 \Rightarrow \boxed{\mathrm{A}}$

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

Q22.

Solution

Concept — The halogens (group 17): The halogens are the elements of group 17 of the periodic table: fluorine, chlorine, bromine, iodine, and astatine. They each have seven valence electrons ($ns^2\,np^5$) and readily gain one electron to form -1 ions, making them highly reactive non-metals.

Step 1 — Classify each option by group: neon is a noble gas (group 18); calcium is an alkaline-earth metal (group 2); sulphur is a chalcogen (group 16); bromine is in group 17.

Step 2 — Identify the group-17 element: only bromine, with the configuration ending in $4s^2\,4p^5$ (seven valence electrons), belongs to the halogens.

Step 3 — Conclusion: bromine is the halogen.

Why each other option is wrong:

- (A) Neon is a noble gas (group 18), not a halogen.
- (C) Calcium is a group-2 metal, not a halogen.
- (D) Sulphur is in group 16 (the chalcogens), with six valence electrons, not seven.

Key point: Halogens are group 17 (F, Cl, Br, I, At) with seven valence electrons.

Bromine is the only halogen among the choices; it exists as a reddish-brown liquid at room temperature.

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

Answer: (B) Go Back to Q22

Q23.

Solution

Concept — Structural isomers: Structural (constitutional) isomers are compounds with the same molecular formula but different arrangements of atoms (different connectivity). For small alkanes, the number of distinct carbon-skeleton arrangements gives the number of isomers. The first three alkanes (methane, ethane, propane) each have only one possible structure.

Given: propane, C_3H_8.

Step 1 — Try to arrange the three carbons: three carbon atoms can only be joined in a single straight chain, C–C–C. There is no way to branch a three-carbon skeleton (a branch would need at least four carbons).

Step 2 — Count the distinct structures: only one arrangement of C_3H_8 exists.

Step 3 — Conclusion: propane has only 1 structural isomer (itself).

Why each other option is wrong:

- (A) 3 is far too many; branching is impossible for only three carbons.
- (B) 4 is the number of isomers for a larger alkane, not propane.
- (C) 2 is the number of isomers of butane (C_4H_{10}: n-butane and isobutane), not propane.

Key point: Branching first becomes possible at four carbons, so methane, ethane, and propane each have just one structure. Butane is the first alkane with structural isomers (2 of them).

Final Answer: 1 isomer $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q23

Q24.

Solution

Concept — IUPAC naming of carboxylic acids: A carboxylic acid is named by taking the alkane name of the longest carbon chain that includes the $-$COOH carbon, dropping the final "-e", and adding the suffix "-oic acid". The carbon of the $-$COOH group is counted as part of the chain and is given position 1.

Given: CH_3COOH (acetic acid).

Step 1 — Count the carbon atoms in the chain: there are two carbons, the CH_3 carbon and the COOH carbon, so the parent chain is "ethane".

Step 2 — Apply the carboxylic-acid suffix: replace the "-e" of ethane with "-oic acid", giving "ethanoic acid".

Step 3 — Conclusion: the IUPAC name of CH_3COOH is ethanoic acid.

Why each other option is wrong:

- (A) Methanoic acid (HCOOH) has only one carbon, but CH_3COOH has two.
- (B) Ethanol (CH_3CH_2OH) is an alcohol, not an acid; it does not contain the $-$COOH group.
- (D) Propanoic acid (CH_3CH_2COOH) has three carbons, one more than the given compound.

Key point: Count the carbons (including the $-$COOH carbon) to fix the parent chain, then add "-oic acid". Two carbons → ethanoic acid, the common acid in vinegar.

Final Answer: Ethanoic acid $\Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q24

Q25.

Solution

Concept — General formula of a homologous series: Each homologous series of hydrocarbons has a characteristic general formula. Alkanes (all single bonds) are C_nH_{2n+2}; introducing one double bond removes two hydrogens to give alkenes C_nH_{2n}; introducing one triple bond removes two more hydrogens to give alkynes C_nH_{2n-2}.

Step 1 — Start from the alkane formula: C_nH_{2n+2} (the saturated baseline).

Step 2 — Add one triple bond: a triple bond corresponds to two degrees of unsaturation, removing four hydrogens? No, one triple bond removes two H relative to an alkene, which itself has two fewer than an alkane. Net: a single triple bond removes four H from the alkane, giving $C_nH_{2n+2-4} = C_nH_{2n-2}$.

Step 3 — Verify with ethyne: for $n = 2$, $C_2H_{2(2)-2} = C_2H_2$, which is ethyne (acetylene). Correct.

Why each other option is wrong:

- (B) C_nH_{2n} is the alkene (one double bond) general formula.
- (C) C_nH_{2n+2} is the alkane (saturated) general formula.
- (D) C_nH_n does not correspond to any standard hydrocarbon series.

Key point: Alkanes C_nH_{2n+2}, alkenes C_nH_{2n}, alkynes C_nH_{2n-2}. Each step of unsaturation removes two hydrogens; a triple bond is two such steps below the alkane.

Final Answer: $C_nH_{2n-2} \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q25

Q26.

Solution

Concept — The structure of benzene: Benzene is the simplest aromatic hydrocarbon. It is a flat, six-membered ring of carbon atoms, each carbon bonded to one hydrogen, with six delocalised π electrons spread evenly over the ring (shown as the inner circle). Each ring carbon is sp^2 hybridised.

Step 1 — Count the carbons: the hexagonal ring has six carbon atoms, one at each corner.

Step 2 — Count the hydrogens: each ring carbon carries exactly one hydrogen, so there are six hydrogens.

Step 3 — Write the molecular formula: six carbons and six hydrogens give C_6H_6.

Why each other option is wrong:

- (A) C_6H_{12} is cyclohexane, the fully saturated six-carbon ring (no aromatic system), not benzene.
- (C) C_5H_5 wrongly uses a five-membered ring and an odd count.
- (D) C_6H_{14} is hexane, an open-chain saturated alkane, not the aromatic ring.

Key point: Benzene is C_6H_6: a six-carbon aromatic ring with one H per carbon and a delocalised π system. Its high degree of unsaturation (and aromatic stability) distinguishes it from cyclohexane (C_6H_{12}).

Final Answer: $C_6H_6 \Rightarrow \boxed{B}$

Answer: (B) Go Back to Q26

Q27.

Solution

Concept — Sigma and pi bonds in multiple bonds: A single bond is one σ bond. A double bond is one σ bond plus one π bond. A triple bond is one σ bond plus *two* π bonds. To count π bonds in a molecule, identify each multiple bond and add up its π components.

Given: ethyne, C_2H_2, structure H–C≡C–H.

Step 1 — Identify the bonds: two C–H single bonds (each one σ, zero π) and one C≡C triple bond between the two carbons.

Step 2 — Break down the triple bond: a C≡C triple bond consists of one σ bond and two π bonds.

Step 3 — Total the π bonds: the only π bonds are the two in the triple bond, so the molecule has 2 π bonds.

Why each other option is wrong:

- (A) 1 would be the number of π bonds in a *double* bond (as in ethene), not a triple bond.
- (B) 3 counts the triple bond as three π bonds, but one of its three bonds is a σ bond.
- (D) 0 would be the case only for a molecule with single bonds alone, such as ethane.

Key point: Triple bond $= 1\,\sigma + 2\,\pi$. So ethyne has two π bonds (plus three σ bonds: one C–C and two C–H). The C–H bonds contribute no π bonds.

Final Answer: 2 π bonds $\Rightarrow \boxed{C}$

Answer: (C) Go Back to Q27

Q28.

Solution

Concept — The ketone functional group: A functional group is the reactive part of a molecule that defines its class. The carbonyl group (>C=O) bonded to *two* carbon atoms (R–CO–R′) is the defining group of *ketones*. The position of the carbonyl, between two carbons, distinguishes a ketone from an aldehyde (where the carbonyl carbon bears a hydrogen).

Step 1 — Read the structure: the central carbon is double-bonded to oxygen (carbonyl) and single-bonded to two carbon groups, R and R′.

Step 2 — Match the group to its class: a carbonyl flanked by two carbons is characteristic of a ketone, for example propanone (acetone, CH_3COCH_3).

Step 3 — Conclusion: the group shown belongs to ketones.

Why each other option is wrong:

- (A) Alcohols have an $-$OH group, not a carbonyl.
- (B) Aldehydes have a carbonyl bonded to at least one *hydrogen* ($-$CHO), not to two carbons.
- (C) Carboxylic acids have a carbonyl bearing an $-$OH on the same carbon ($-$COOH), an extra oxygen compared with a ketone.

Key point: The carbonyl's neighbours decide the class: two carbons → ketone; one carbon and one hydrogen → aldehyde; an attached $-$OH → carboxylic acid.

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

$\boxed{\textbf{Answer: (D)}}$ Go Back to Q28

Q29.

Solution

Concept — Monomers of nucleic acids: Large biomolecules are polymers built from small repeating monomer units. Nucleic acids (DNA and RNA) are polymers of *nucleotides*. Each nucleotide is itself made of three parts: a nitrogenous base, a pentose sugar, and a phosphate group.

Step 1 — Identify the monomer of nucleic acids: the repeating unit is the nucleotide.

Step 2 — How they link: the phosphate of one nucleotide joins the sugar of the next through phosphodiester bonds, building the long sugar–phosphate backbone of DNA or RNA.

Step 3 — Conclusion: nucleic acids are made of nucleotide monomers.

Why each other option is wrong:

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

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

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

$\boxed{\textbf{Answer: (A)}}$ **Go Back to Q29**

Q30.

Solution

Concept — The acid in vinegar: Vinegar is a dilute aqueous solution (typically 4–8%) of acetic acid (ethanoic acid, CH_3COOH). Acetic acid is a weak carboxylic acid; it is responsible for the characteristic sour taste and pungent smell of vinegar.

Step 1 — Recall the composition of vinegar: the active acid is acetic acid, CH_3COOH, produced by the fermentation of ethanol.

Step 2 — Link the acid to its properties: the partial ionisation of the $-COOH$ group ($CH_3COOH \rightleftharpoons CH_3COO^- + H^+$) makes the solution weakly acidic and sour.

Step 3 — Conclusion: the acid present in vinegar is acetic acid.

Why each other option is wrong:

- (A) Citric acid is found in citrus fruits (lemons, oranges), not in vinegar.
- (B) Formic acid ($HCOOH$) is found in ant stings and nettle stings, not in vinegar.
- (D) Hydrochloric acid is a strong mineral acid present in gastric juice, not the mild acid of vinegar.

Key point: Vinegar $=$ dilute acetic acid (CH_3COOH), whose IUPAC name is ethanoic acid. Knowing the common acids and their sources (citric in citrus, formic in ant stings, lactic in sour milk) is a frequently tested fact.

Final Answer: Acetic acid $\Rightarrow \boxed{\text{C}}$

$\boxed{\textbf{Answer: (C)}}$ Go Back to Q30

Answer Key

Q	Ans	Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	A	2	C	3	B	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	D	24	C	25	A
26	B	27	C	28	D	29	A	30	C