AIIMS B.Sc Nursing Chemistry

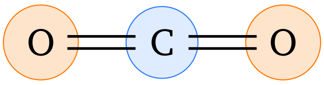
Sample Paper – 7

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. One molecule of carbon dioxide (CO_2) is made of one carbon atom and two oxygen atoms, as shown. The number of moles of atoms present in 1 mole of CO_2 molecules is:



(A) 1 mol
(B) 2 mol
(C) 3 mol
(D) 6 mol

Q2. At standard temperature and pressure (STP), 1 mole of any gas occupies 22.4 L. The number of moles present in 11.2 L of oxygen gas (O_2) at STP is:

(A) 1 mol

(B) 0.5 mol

(C) 2 mol

(D) 0.25 mol

Q3. A p subshell consists of three orbitals (p_x, p_y, p_z), and each orbital can hold at most two electrons. The maximum number of electrons that a p subshell can accommodate is:

(A) 2

(B) 10

(C) 14

(D) 6

Q4. A chlorine atom has the electronic configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$. The number of valence (outermost shell) electrons in a chlorine atom is:

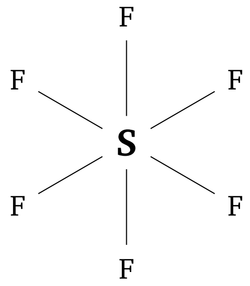
(A) 7

(B) 5

(C) 8

(D) 1

Q5. In sulfur hexafluoride (SF_6) the central sulfur atom is bonded to six fluorine atoms arranged symmetrically around it, as shown. The shape of the SF_6 molecule is:



(A) tetrahedral

(B) octahedral

(C) trigonal bipyramidal

(D) square planar

Q6. In sulfur hexafluoride (SF_6), the central sulfur atom forms six equivalent σ bonds to fluorine with no lone pair. The hybridization of sulfur in SF_6 is:

(A) sp^3

(B) sp^3d

(C) sp^3d^2

(D) sp^2

Q7. The ionic character of a bond increases with the electronegativity difference between the two bonded atoms. Among the following, the bond with the highest ionic character is:

(A) H–F

(B) H–Cl

(C) H–Br

(D) H–I

Q8. When a neutral atom loses one or more electrons to form a cation, the resulting ion compared with the parent atom is:

(A) larger in size

(B) the same size

(C) first larger, then smaller

(D) smaller in size

Q9. The elements of group 2 of the periodic table (the alkaline earth metals such as Mg and Ca) all have the same number of valence electrons. That number is:

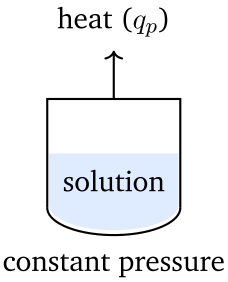
(A) 1

(B) 2

(C) 8

(D) 3

Q10. Heat is exchanged in a reaction carried out in an open vessel at constant pressure, as suggested by the apparatus below. The heat measured under these conditions is equal to the change in:



(A) enthalpy (ΔH)

(B) internal energy (ΔU)

(C) entropy (ΔS)

(D) free energy (ΔG)

Q11. According to the second law of thermodynamics, for any spontaneous (natural) process the total entropy of the universe ($\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$):

(A) decreases

(B) remains constant

(C) increases

(D) becomes zero

Q12. For a gaseous reaction at temperature T, with Δn the change in the number of moles of gas (products minus reactants) and R the gas constant, the relation between K_p and K_c is:

(A) $K_p = K_c$

(B) $K_p = K_c\,(RT)^{-\Delta n}$

(C) $K_p = K_c/RT$

(D) $K_p = K_c\,(RT)^{\Delta n}$

Q13. A strong acid is one that ionises almost completely in aqueous solution. Which of the following is a strong acid?

(A) HCl

(B) CH_3COOH

(C) H_2CO_3

(D) HF

Q14. An amphoteric substance can react with *both* acids and bases. Which one of the following is amphoteric?

(A) NaOH

(B) Al_2O_3

(C) MgO

(D) CO_2

Q15. Taking the oxidation number of oxygen as -2, the oxidation number of sulfur in sulfur dioxide (SO_2) is:

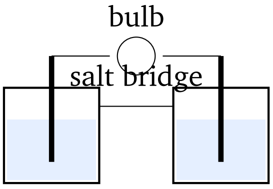
(A) $+2$

(B) $+6$

(C) $+4$

(D) -2

Q16. A galvanic (voltaic) cell, sketched below with two electrodes, a salt bridge, and a light bulb, operates on its own. It converts:



(A) chemical energy into electrical energy

(B) electrical energy into chemical energy

(C) heat energy into chemical energy

(D) light energy into chemical energy

Q17. One faraday ($1\,\mathrm{F} = 96500$ C) is the charge carried by one mole of electrons. The number of electrons contained in one faraday of charge is therefore:

(A) 96500

(B) 1.6×10^{-19}

(C) 3.011×10^{23}

(D) 6.022×10^{23}

Q18. The molar mass of sodium hydroxide (NaOH) is 40 g mol^{-1}. The mass of NaOH required to prepare 1 L of a 0.1 M aqueous solution is:

(A) 40 g

(B) 4 g

(C) 0.4 g

(D) 0.1 g

Q19. For a mixture of non-reacting gases, Dalton's law of partial pressures states that the total pressure exerted by the mixture is equal to:

(A) the product of the individual partial pressures

(B) the partial pressure of the heaviest gas

(C) the sum of the partial pressures of the individual gases

(D) the average of the partial pressures

Q20. In the complex ion $[\mathrm{Ag(NH_3)_2}]^+$, the ammonia ligands are neutral. The oxidation state of silver (Ag) in this complex is:

(A) $+1$
(B) $+2$
(C) 0
(D) $+3$

Q21. A monodentate ligand binds to the central metal ion through only *one* donor atom. Which of the following is a monodentate ligand?

(A) ethylenediamine (en)
(B) ammonia (NH_3)
(C) oxalate ($C_2O_4^{2-}$)
(D) EDTA

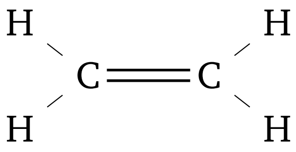
Q22. A metalloid shows properties intermediate between metals and non-metals. Which one of the following elements is a metalloid?

(A) silicon (Si)
(B) sodium (Na)
(C) sulfur (S)
(D) calcium (Ca)

Q23. Geometrical (cis–trans) isomerism arises in compounds in which rotation about a bond is restricted. The essential condition for geometrical isomerism about a carbon–carbon double bond is:

(A) free rotation about the double bond
(B) the presence of a triple bond
(C) at least one carbon bearing two identical groups
(D) restricted rotation, with each doubly bonded carbon carrying two different groups

Q24. The structure shown below, with two carbon atoms joined by a double bond and each carrying two hydrogen atoms ($CH_2{=}CH_2$), has the IUPAC name:



(A) ethane

(B) ethanol

(C) ethene

(D) ethyne

Q25. Carboxylic acids contain the carboxyl group $-\text{COOH}$. The general molecular formula of a saturated aliphatic carboxylic acid (such as HCOOH, CH_3COOH) is:

(A) C_nH_{2n+2}

(B) $C_nH_{2n}O_2$

(C) $C_nH_{2n}O$

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

Q26. One of the hydrocarbons drawn below contains a carbon–carbon triple bond. Which compound is it?

$H_3C\text{–}CH_3$	$H_2C{=}CH_2$	$HC{\equiv}CH$
(I)	(II)	(III)

(A) ethane (I)

(B) ethene (II)

(C) methane

(D) ethyne (III)

Q27. In a methane molecule (CH_4) the carbon atom is sp^3 hybridised with four equivalent C–H bonds pointing to the corners of a regular tetrahedron. The H–C–H bond angle in methane is:

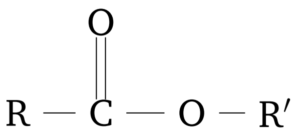
(A) 109.5°

(B) 90°

(C) 120°

(D) 180°

Q28. The functional group shown below, a carbon double-bonded to one oxygen and single-bonded to another oxygen that links to a carbon chain ($-COO-$), is characteristic of which class of compounds?



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

($-COO-$ linkage)

(A) aldehydes

(B) carboxylic acids

(C) esters

(D) ethers

Q29. Fats and oils (triglycerides) are formed by the esterification of glycerol with three molecules of:

(A) amino acids

(B) glucose

(C) nucleotides

(D) fatty acids

Q30. A compound containing the amino group ($-NH_2$), such as an amine, has a lone pair of electrons on the nitrogen that can accept a proton. Such a compound therefore behaves as a:

(A) strong acid

(B) base

(C) neutral salt

(D) oxidising agent

Detailed Solutions

Q1.

Solution

Concept — Moles of atoms inside a compound: One mole of a molecular compound contains Avogadro's number of *molecules*, but each molecule is itself made of several atoms. To find the moles of atoms, multiply the moles of molecules by the number of atoms in one molecule. The number of atoms in CO_2 is read directly from its formula: one carbon (1) plus two oxygens (2) equals 3 atoms per molecule.

Given: 1 mole of CO_2 molecules; atoms per molecule $= 1\,(\text{C}) + 2\,(\text{O}) = 3$.

Step 1 — Write the relation: moles of atoms $=$ (moles of molecules) $\times$ (atoms per molecule).

Step 2 — Substitute: $= 1$ mol $\times\ 3 = 3$ mol of atoms.

Step 3 — Interpret: this comprises 1 mol of C atoms and 2 mol of O atoms, totalling 3 mol of atoms.

Why each other option is wrong:

- (A) 1 mol counts only the molecules (or only the carbon atoms), ignoring the two oxygens.
- (B) 2 mol counts only the oxygen atoms, leaving out the carbon.
- (D) 6 mol would be the moles of atoms in 2 mol of CO_2, double the amount asked.

Key point: Distinguish moles of *molecules* from moles of *atoms*. Multiply by the atom count in the formula: 1 mol $CO_2 \rightarrow 3$ mol atoms.

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

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

Q2.

Solution

Concept — Molar volume at STP: At standard temperature and pressure, one mole of *any* ideal gas occupies a fixed volume of 22.4 L (the molar volume). The number of moles in a measured gas volume is found from $n = \dfrac{V}{V_m}$, where V is the gas volume and $V_m = 22.4$ L mol^{-1} at STP. This relation is independent of the identity of the gas.

Given: $V = 11.2$ L of O_2 at STP; $V_m = 22.4$ L mol^{-1}.

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

Step 2 — Substitute with units: $n = \dfrac{11.2\text{ L}}{22.4\text{ L mol}^{-1}}$.

Step 3 — Evaluate: the litres cancel, giving $n = 0.5$ mol.

Why each other option is wrong:

- (A) 1 mol would occupy the full 22.4 L, not 11.2 L.
- (C) 2 mol would occupy 44.8 L.
- (D) 0.25 mol would occupy only 5.6 L.

Key point: 22.4 L per mole at STP applies to any gas. Half that volume (11.2 L) is exactly half a mole, regardless of whether the gas is O_2, N_2, or CO_2.

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

Answer: (B) Go Back to Q2

Q3.

Solution

Concept — Capacity of a subshell: A subshell is a set of orbitals of the same type. The number of orbitals in a subshell is fixed: s has 1, p has 3, d has 5, and f has 7. By Pauli's exclusion principle each orbital holds at most two electrons (of opposite spin). The maximum electrons a subshell can hold is therefore $2 \times$ (number of orbitals).

Given: a p subshell, which contains 3 orbitals (p_x, p_y, p_z).

Step 1 — Count the orbitals: a p subshell has 3 orbitals.

Step 2 — Multiply by two electrons per orbital: $3 \times 2 = 6$.

Step 3 — Conclusion: a p subshell holds a maximum of 6 electrons.

Why each other option is wrong:

- (A) 2 is the capacity of an s subshell (1 orbital $\times 2$).
- (B) 10 is the capacity of a d subshell (5 orbitals $\times 2$).
- (C) 14 is the capacity of an f subshell (7 orbitals $\times 2$).

Key point: Memorise the orbital counts (s:1, p:3, d:5, f:7) and double them for the electron capacity: s:2, p:6, d:10, f:14.

Final Answer: 6 electrons $\Rightarrow \boxed{\text{D}}$

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

Q4.

Solution

Concept — Valence electrons: The valence electrons are the electrons in the outermost (highest n) shell of an atom; they determine the atom's chemical behaviour. To count them, write the electronic configuration and add up the electrons in the shell with the largest principal quantum number.

Given: chlorine, $Z = 17$, configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$.

Step 1 — Identify the outermost shell: the largest n present is $n = 3$.

Step 2 — Add the electrons in that shell: $3s^2 + 3p^5 = 2 + 5 = 7$.

Step 3 — Conclusion: chlorine has 7 valence electrons, consistent with its position in group 17.

Why each other option is wrong:

- (B) 5 would be a group-15 element (such as nitrogen or phosphorus), not chlorine.
- (C) 8 is a complete octet, the configuration of a noble gas such as argon.
- (D) 1 would be a group-1 (alkali metal) element such as sodium.

Key point: For a main-group element, the number of valence electrons equals its group number (group 17 $\rightarrow$ 7 valence electrons). Chlorine needs just one more electron to reach a stable octet, which is why it is so reactive.

Final Answer: 7 valence electrons $\Rightarrow \boxed{\text{A}}$

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

Q5.

Solution

Concept — VSEPR shape of SF_6: VSEPR theory predicts molecular shape from the number of electron domains around the central atom. Sulfur in SF_6 forms six σ bonds and has no lone pair, giving six electron domains. Six domains repel each other most when placed at the corners of a regular octahedron (four in a plane, one above, one below), all at 90°.

Step 1 — Count electron domains on sulfur: six S–F bonds, zero lone pairs $= 6$ domains.

Step 2 — Assign the electron geometry: six domains $\Rightarrow$ octahedral arrangement.

Step 3 — Shape: since all six positions hold bonding pairs (no lone pairs), the molecular shape is also octahedral, with 90° F–S–F angles.

Why each other option is wrong:

- (A) Tetrahedral is for four electron domains, as in CH_4.
- (C) Trigonal bipyramidal is for five domains, as in PCl_5.
- (D) Square planar arises from six domains with *two* lone pairs (as in XeF_4), not from six bonds.

Key point: Six bonding pairs and no lone pairs $\rightarrow$ a perfectly symmetric octahedron. This high symmetry also makes SF_6 non-polar despite its polar S–F bonds.

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

Answer: (B) Go Back to Q5

Q6.

Solution

Concept — Hybridisation from electron domains: The hybridisation of a central atom equals the total number of σ bonds plus lone pairs (its electron domains). The mapping is: $2 \rightarrow sp$, $3 \rightarrow sp^2$, $4 \rightarrow sp^3$, $5 \rightarrow sp^3d$, $6 \rightarrow sp^3d^2$. Sulfur in SF_6 has six σ bonds and no lone pair.

Step 1 — Count the domains on sulfur: six S–F σ bonds $+ \ 0$ lone pairs $= 6$ electron domains.

Step 2 — Map to a hybridisation: six domains $\Rightarrow sp^3d^2$ (one s, three p, and two d orbitals mix to give six equivalent hybrid orbitals).

Step 3 — Geometry check: the six sp^3d^2 orbitals point to the corners of an octahedron, matching the shape found in Q5.

Why each other option is wrong:

- (A) sp^3 (four domains) describes a tetrahedral centre, not six bonds.
- (B) sp^3d (five domains) describes a trigonal bipyramid, as in PCl_5.
- (D) sp^2 (three domains) describes a trigonal planar centre.

Key point: Six bonds demand six hybrid orbitals, which requires two d orbitals: sp^3d^2. The hybridisation and the octahedral shape go hand in hand.

Final Answer: $sp^3d^2 \Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q6

Q7.

Solution

Concept — Ionic character and electronegativity difference: The percentage ionic character of a covalent bond grows as the electronegativity difference ($\Delta\chi$) between the two bonded atoms grows. Hydrogen has a fixed electronegativity, so for the hydrogen halides the ionic character depends on the halogen. Down group 17, electronegativity falls (F $>$ Cl $>$ Br $>$ I), so $\Delta\chi$ between H and the halogen is largest for fluorine.

Step 1 — List the electronegativities (Pauling): F ≈ 4.0, Cl ≈ 3.0, Br ≈ 2.8, I ≈ 2.5; H ≈ 2.1.

Step 2 — Compute $\Delta\chi$ for each bond: H–F ≈ 1.9; H–Cl ≈ 0.9; H–Br ≈ 0.7; H–I ≈ 0.4.

Step 3 — Pick the largest: H–F has the greatest $\Delta\chi$, so it has the highest ionic character.

Why each other option is wrong:

- (B) H–Cl has a smaller $\Delta\chi$ than H–F, so less ionic character.
- (C) H–Br is even less polar than H–Cl.
- (D) H–I has the smallest $\Delta\chi$ and is the most covalent of the four.

Key point: Larger electronegativity difference → more ionic character. Since fluorine is the most electronegative element, H–F is the most ionic of the hydrogen halides.

Final Answer: H–F $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q7

Q8.

Solution

Concept — Size of a cation vs its parent atom: A cation forms when an atom loses one or more electrons. Removing electrons (often emptying the outermost shell entirely) reduces electron–electron repulsion, while the unchanged nuclear charge now pulls on fewer electrons. Both effects shrink the ion, so a cation is always *smaller* than its parent atom.

Step 1 — Identify the change: the atom loses electrons; the number of protons stays the same.

Step 2 — Effect on size: the higher effective nuclear charge per electron, and often the loss of an entire outer shell, pulls the remaining electrons closer.

Step 3 — Conclusion: the cation is smaller than the parent atom (for example, $r_{\text{Na}^+} \approx 95$ pm is much less than $r_{\text{Na}} \approx 186$ pm).

Why each other option is wrong:

- (A) Larger size describes an *anion*, which gains electrons, not a cation.
- (B) The size cannot remain the same once electrons are removed.
- (C) There is no "first larger then smaller" behaviour for simple cation formation.

Key point: Cation $<$ parent atom $<$ anion. Losing electrons shrinks the species; gaining electrons enlarges it.

Final Answer: Smaller in size $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q8

Q9.

Solution

Concept — Valence electrons and group number: For main-group (representative) elements, the number of valence electrons equals the group number in the older system (groups 1, 2, 13–18). Group 2 elements, the alkaline earth metals (Be, Mg, Ca, Sr, Ba), all end in an ns^2 configuration, giving exactly two valence electrons.

Step 1 — Recall a group-2 configuration: e.g. magnesium is $[\text{Ne}]\,3s^2$, calcium is $[\text{Ar}]\,4s^2$.

Step 2 — Count the outermost electrons: the outer shell contains ns^2, i.e. 2 electrons.

Step 3 — Conclusion: every group-2 element has 2 valence electrons, which it readily loses to form a $+2$ ion.

Why each other option is wrong:

- (A) 1 valence electron is characteristic of group 1 (the alkali metals).
- (C) 8 valence electrons is a complete octet, found in the noble gases (group 18).
- (D) 3 valence electrons belongs to group 13 (such as boron or aluminium).

Key point: Group number gives valence electrons for main-group elements: group $1 \to 1$, group $2 \to 2$. The two ns^2 electrons explain why group-2 metals form $+2$ ions.

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

Answer: (B) Go Back to Q9

Q10.

Solution

Concept — Heat at constant pressure equals ΔH: Enthalpy (H) is defined so that the heat exchanged by a system at constant pressure equals the change in enthalpy: $q_p = \Delta H$. This is why most laboratory reactions, carried out in open vessels exposed to constant atmospheric pressure, directly measure ΔH. (By contrast, heat at constant *volume* equals ΔU, the internal energy change.)

Step 1 — Identify the condition: the reaction occurs in an open vessel, so the pressure is constant (atmospheric).

Step 2 — Apply the definition: at constant pressure $q_p = \Delta H$.

Step 3 — Conclusion: the measured heat equals the enthalpy change, ΔH.

Why each other option is wrong:

- (B) Internal energy change ΔU equals the heat at constant *volume* (q_v), not constant pressure.
- (C) Entropy ΔS measures disorder, not heat at constant pressure.
- (D) Free energy ΔG combines enthalpy and entropy ($\Delta G = \Delta H - T\Delta S$); it is not the directly measured heat.

Key point: Constant pressure → $q_p = \Delta H$; constant volume → $q_v = \Delta U$. Open-vessel reactions measure ΔH.

Final Answer: Enthalpy (ΔH) $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q10

Q11.

Solution

Concept — Second law and entropy of the universe: The second law of thermodynamics states that for any spontaneous (natural, irreversible) process the total entropy of the universe increases: $\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$. Entropy is a measure of disorder, and natural changes always drive the universe toward greater disorder. A process at equilibrium has $\Delta S_{\text{universe}} = 0$.

Step 1 — State the criterion: spontaneity $\Leftrightarrow \Delta S_{\text{universe}} > 0$.

Step 2 — Apply it: since the process is spontaneous, the total entropy of the universe must increase.

Step 3 — Note: the system's own entropy may decrease (e.g. freezing water), but only if the surroundings gain even more entropy, so the *total* still increases.

Why each other option is wrong:

- (A) A decrease in total entropy would make the process non-spontaneous (the reverse would be spontaneous).
- (B) Constant total entropy ($\Delta S_{\text{universe}} = 0$) describes equilibrium, not a spontaneous change.
- (D) Zero is the same equilibrium condition, not a spontaneous one.

Key point: Spontaneous → $\Delta S_{\text{universe}} > 0$; equilibrium → $\Delta S_{\text{universe}} = 0$. Always judge spontaneity by the entropy of the *universe*, not the system alone.

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

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

Q12.

Solution

Concept — Relation between K_p and K_c: For a gaseous equilibrium, the equilibrium constant in terms of partial pressures (K_p) and in terms of molar concentrations (K_c) are linked through the ideal-gas relation $p = (n/V)RT = [\,]\,RT$. Carrying this through the equilibrium expression gives $K_p = K_c\,(RT)^{\Delta n}$, where $\Delta n =$ (moles of gaseous products) $-$ (moles of gaseous reactants).

Step 1 — Start from $p_i = [i]\,RT$ for each gas: substituting into K_p replaces every concentration by a pressure.

Step 2 — Collect the RT factors: each product contributes $+1$ and each reactant -1 to the exponent, summing to Δn.

Step 3 — Result: $K_p = K_c\,(RT)^{\Delta n}$.

Why each other option is wrong:

- (A) $K_p = K_c$ holds only in the special case $\Delta n = 0$, not in general.
- (B) $K_p = K_c\,(RT)^{-\Delta n}$ has the wrong sign on the exponent.
- (C) $K_p = K_c/RT$ assumes $\Delta n = -1$ for every reaction, which is not general.

Key point: $K_p = K_c\,(RT)^{\Delta n}$. When the number of gas moles is unchanged ($\Delta n = 0$), the $(RT)^0 = 1$ factor makes $K_p = K_c$.

Final Answer: $K_p = K_c\,(RT)^{\Delta n} \Rightarrow \boxed{\text{D}}$

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

Q13.

Solution

Concept — Strong vs weak acids: A strong acid ionises almost completely in water, releasing nearly all of its protons; a weak acid ionises only partially. The common strong acids are HCl, HBr, HI, HNO_3, H_2SO_4, and $HClO_4$. Most other acids, including organic acids and carbonic acid, are weak.

Step 1 — Test each option: HCl is a classic strong acid (fully ionised); CH_3COOH (acetic acid) is a weak acid; H_2CO_3 (carbonic acid) is weak; HF is weak (the strong H–F bond resists ionisation).

Step 2 — Select the strong acid: only HCl qualifies.

Step 3 — Conclusion: HCl is the strong acid.

Why each other option is wrong:

- (B) CH_3COOH is a weak acid; it ionises only slightly ($K_a \approx 1.8 \times 10^{-5}$).
- (C) H_2CO_3 is a weak diprotic acid.
- (D) HF is the only hydrogen halide that is weak, owing to its very strong bond.

Key point: Memorise the short list of strong acids (HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$). Note that HF, despite being a halide, is *weak*.

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

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

Q14.

Solution

Concept — Amphoteric substances: An amphoteric substance reacts with both acids and bases. Aluminium oxide, Al_2O_3, is the standard example: it dissolves in acids to form aluminium salts and in alkalis to form aluminates. Such behaviour is typical of the oxides of metals near the metal–non-metal border (Al, Zn, Pb, Sn).

Step 1 — React Al_2O_3 with an acid: $Al_2O_3 + 6\,HCl \rightarrow 2\,AlCl_3 + 3\,H_2O$ (behaves as a base).

Step 2 — React Al_2O_3 with a base: $Al_2O_3 + 2\,NaOH \rightarrow 2\,NaAlO_2 + H_2O$ (behaves as an acid).

Step 3 — Conclusion: since it reacts both ways, Al_2O_3 is amphoteric.

Why each other option is wrong:

- (A) NaOH is a strong base only; it does not behave as an acid.
- (C) MgO is a basic oxide (reacts with acids but not with alkalis).
- (D) CO_2 is an acidic oxide (forms carbonic acid with water; reacts with bases but not acids).

Key point: Amphoteric oxides (Al_2O_3, ZnO, PbO, SnO) react with *both* acids and bases, unlike purely basic (MgO) or purely acidic (CO_2) oxides.

Final Answer: $Al_2O_3 \Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q14

Q15.

Solution

Concept — Oxidation number rules: In a neutral molecule the oxidation numbers of all atoms sum to zero. Oxygen is normally assigned -2. The oxidation number of the other element is then found algebraically from this balance.

Given: SO_2, a neutral molecule; oxygen $= -2$ each.

Step 1 — Assign the known value: the two oxygens contribute $2 \times (-2) = -4$. Let the oxidation number of S be x.

Step 2 — Set the sum to zero: $x + (-4) = 0$.

Step 3 — Solve: $x = +4$.

Why each other option is wrong:

- (A) $+2$ would require the oxygens to total only -2, i.e. a single oxygen.
- (B) $+6$ is the oxidation state of sulfur in SO_3 or H_2SO_4, not SO_2.
- (D) -2 is sulfur's state in sulfides such as H_2S, not in an oxide.

Key point: Balance the oxidation numbers to zero for a neutral compound. With oxygen fixed at -2, sulfur in SO_2 must be $+4$.

Final Answer: Oxidation number of S $= +4 \Rightarrow \boxed{\text{C}}$

Answer: (C) Go Back to Q15

Q16.

Solution

Concept — Energy conversion in a galvanic cell: A galvanic (voltaic) cell harnesses a spontaneous redox reaction to drive electrons through an external wire. The chemical energy released by the reaction is converted into electrical energy, which can light a bulb or run a device. (An electrolytic cell does the reverse: it uses electrical energy to force a non-spontaneous chemical change.)

Step 1 — Identify the driving force: a spontaneous redox reaction takes place in the cell (oxidation at the anode, reduction at the cathode).

Step 2 — Identify the output: the electron flow through the external circuit constitutes an electric current that does work (lights the bulb).

Step 3 — Conclusion: the cell converts chemical energy into electrical energy.

Why each other option is wrong:

- (B) Electrical $\to$ chemical is the conversion in an *electrolytic* cell, not a galvanic one.
- (C) Heat $\to$ chemical is not how a galvanic cell operates.
- (D) Light $\to$ chemical describes a photochemical process, not a galvanic cell.

Key point: Galvanic cell: chemical $\to$ electrical (spontaneous). Electrolytic cell: electrical $\to$ chemical (non-spontaneous). They are exact opposites.

Final Answer: Chemical energy into electrical energy $\Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q16

Q17.

Solution

Concept — One faraday and Avogadro's number: One faraday ($1\,\text{F} = 96500$ C) is defined as the charge carried by one mole of electrons. Since one mole of any particle contains Avogadro's number of particles, one faraday contains exactly $N_A = 6.022 \times 10^{23}$ electrons. (Cross-check: $N_A \times e = 6.022 \times 10^{23} \times 1.6 \times 10^{-19}$ C ≈ 96500 C.)

Given: $1\,\text{F} = 96500$ C = charge of 1 mole of electrons.

Step 1 — Recall the definition: 1 faraday = charge on 1 mole of electrons.

Step 2 — Convert moles to a count: 1 mole $= N_A = 6.022 \times 10^{23}$ electrons.

Step 3 — Conclusion: one faraday contains 6.022×10^{23} electrons.

Why each other option is wrong:

- (A) 96500 is the charge in coulombs, not a number of electrons.
- (B) 1.6×10^{-19} C is the charge of a *single* electron, not the count in a faraday.
- (C) 3.011×10^{23} is half of Avogadro's number (half a mole of electrons), so half a faraday.

Key point: 1 faraday $=$ 1 mole of electrons $=$ 6.022×10^{23} electrons $=$ 96500 C. The three statements are just different ways of describing the same quantity.

Final Answer: 6.022×10^{23} electrons $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q17

Q18.

Solution

Concept — Molarity and mass of solute: Molarity is moles of solute per litre of solution, $M = \frac{n}{V}$. To find the mass needed, first get the moles required ($n = M \times V$), then convert to mass using $m = n \times$ (molar mass).

Given: $M = 0.1$ M; $V = 1$ L; molar mass of NaOH $= 40$ g mol^{-1}.

Step 1 — Find the moles of NaOH: $n = M \times V = 0.1$ mol L^{-1} $\times$ 1 L $= 0.1$ mol.

Step 2 — Convert moles to mass: $m = n \times M_{\text{molar}} = 0.1$ mol $\times$ 40 g mol^{-1}.

Step 3 — Evaluate: $m = 4$ g.

Why each other option is wrong:

- (A) 40 g is the mass for a full 1 M solution (1 mol), ten times too much.
- (C) 0.4 g corresponds to 0.01 M, ten times too dilute.
- (D) 0.1 g confuses the number of moles (0.1) with the mass in grams.

Key point: Mass $= M \times V \times$ molar mass. Here $0.1 \times 1 \times 40 = 4$ g of NaOH.

Final Answer: 4 g $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q18

Q19.

Solution

Concept — Dalton's law of partial pressures: In a mixture of gases that do not react, each gas exerts the same pressure it would if it alone occupied the whole container; this is its partial pressure. Dalton's law states that the total pressure of the mixture is the *sum* of these partial pressures: $P_{\text{total}} = p_1 + p_2 + p_3 + \dots$

Step 1 — State the law: $P_{\text{total}} = \sum p_i$, the sum of the partial pressures.

Step 2 — Reasoning: each gas independently contributes molecular collisions with the walls, so the pressures simply add.

Step 3 — Conclusion: the total pressure equals the sum of the individual partial pressures.

Why each other option is wrong:

- (A) Pressures *add*; they are not multiplied.
- (B) The total is not just the heaviest gas's pressure; every component contributes.
- (D) It is the sum, not the average, of the partial pressures.

Key point: $P_{\text{total}} = p_1 + p_2 + \dots$, and each $p_i = x_i P_{\text{total}}$ where x_i is the mole fraction. The partial pressures of all components add up to the total.

Final Answer: Sum of the partial pressures $\Rightarrow \boxed{\text{C}}$

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

Q20.

Solution

Concept — Oxidation state of a metal in a complex: In a coordination compound, the oxidation state of the central metal is found by setting the sum of (metal oxidation state + total ligand charges) equal to the overall charge of the complex ion. Neutral ligands such as NH_3 contribute zero charge.

Given: $[Ag(NH_3)_2]^+$; each NH_3 is neutral; overall charge $= +1$.

Step 1 — Sum the ligand charges: $2 \times NH_3 = 2 \times 0 = 0$.

Step 2 — Set up the balance: let Ag $= x$; then $x + 0 = +1$.

Step 3 — Solve: $x = +1$, so silver is in the $+1$ oxidation state.

Why each other option is wrong:

- (B) $+2$ would require the ligands to carry a total of -1 charge, but ammonia is neutral.
- (C) 0 would make the complex neutral, contradicting the $+1$ charge shown.
- (D) $+3$ is the common state of metals such as Co or Cr in complexes, not Ag here.

Key point: Neutral ligands (NH_3, H_2O, CO) add nothing to the charge balance, so the metal's oxidation state equals the overall charge of the complex. Here Ag is $+1$.

Final Answer: $+1 \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q20

Q21.

Solution

Concept — Denticity of a ligand: The denticity of a ligand is the number of donor atoms it uses to bind the central metal. A monodentate ligand binds through one donor atom; bidentate uses two, polydentate uses several. Ammonia (NH_3) donates the single lone pair on its nitrogen, so it is monodentate.

Step 1 — Examine each ligand: NH_3 has one donor (N); ethylenediamine (en) has two donor nitrogens; oxalate has two donor oxygens; EDTA has six donor atoms.

Step 2 — Select the one with a single donor atom: only NH_3 binds through one atom.

Step 3 — Conclusion: ammonia is the monodentate ligand.

Why each other option is wrong:

- (A) Ethylenediamine is *bidentate* (two N donors).
- (C) Oxalate is *bidentate* (two O donors).
- (D) EDTA is *hexadentate* (six donor atoms).

Key point: Count donor atoms to classify a ligand. Single donor = monodentate (NH_3, Cl^-, CN^-, H_2O); multiple donors = chelating (en, oxalate, EDTA).

Final Answer: Ammonia (NH_3) $\Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q21

Q22.

Solution

Concept — Metalloids: Metalloids (semi-metals) lie along the zig-zag staircase that separates metals from non-metals in the periodic table. They show intermediate properties, behaving partly like metals and partly like non-metals; many are semiconductors. The common metalloids are B, Si, Ge, As, Sb, Te (and sometimes Po).

Step 1 — Classify each option: silicon (Si) is a metalloid; sodium (Na) is a metal; sulfur (S) is a non-metal; calcium (Ca) is a metal.

Step 2 — Select the metalloid: only silicon qualifies; it is the classic semiconductor used in electronics.

Step 3 — Conclusion: silicon is the metalloid.

Why each other option is wrong:

- (B) Sodium is a soft, highly reactive *metal* (group 1).
- (C) Sulfur is a *non-metal* (group 16).
- (D) Calcium is an alkaline-earth *metal* (group 2).

Key point: Metalloids (B, Si, Ge, As, Sb, Te) border the metal/non-metal staircase and are intermediate in behaviour. Silicon's semiconducting nature underlies all modern electronics.

Final Answer: Silicon (Si) $\Rightarrow \boxed{\text{A}}$

$\boxed{\text{Answer: (A)}}$ **Go Back to Q22**

Q23.

Solution

Concept — Condition for geometrical isomerism: Geometrical (cis–trans) isomerism arises when rotation about a bond is restricted (as about a C=C double bond or in a ring) *and* each of the two doubly bonded carbons carries two *different* groups. The restricted rotation locks the groups in fixed relative positions (same side = cis, opposite sides = trans), and the difference in groups makes the two arrangements distinguishable.

Step 1 — Require restricted rotation: a C=C double bond (or a ring) prevents free rotation, which a single bond would allow.

Step 2 — Require two different groups on each doubly bonded carbon: if either carbon bore two identical groups, swapping them would give the same molecule, so no isomers would exist.

Step 3 — Combine: both conditions together (restricted rotation + two different groups on each carbon) are needed for cis–trans isomerism.

Why each other option is wrong:

- (A) Free rotation (as about a single bond) would let the groups interconvert, destroying any isomerism.
- (B) A triple bond is linear, so there are no fixed "sides" for cis/trans to differ.
- (C) If a carbon bears two *identical* groups, no geometrical isomers are possible; the groups must be different.

Key point: Cis–trans isomerism needs both restricted rotation *and* two different substituents on each doubly bonded carbon. For example, $CHCl{=}CHCl$ shows it, but $CH_2{=}CHCl$ does not.

Final Answer: Restricted rotation, with each doubly bonded carbon carrying two different groups $\Rightarrow \boxed{D}$

Answer: (D) Go Back to Q23

Q24.

Solution

Concept — IUPAC naming of $CH_2{=}CH_2$: IUPAC names are built from a root indicating the carbon count and a suffix indicating the bonding. Two carbons give the root "eth-"; a carbon–carbon double bond gives the suffix "-ene". So a two-carbon molecule with one C=C double bond is named "ethene" (the common name is ethylene).

Step 1 — Count the carbons: the molecule has 2 carbon atoms → root "eth-".

Step 2 — Identify the bond type: the carbons are joined by a double bond → suffix "-ene".

Step 3 — Combine: eth- + -ene = ethene.

Why each other option is wrong:

- (A) Ethane (C_2H_6) has a single C–C bond (suffix "-ane"), not a double bond.
- (B) Ethanol (C_2H_5OH) contains an $-OH$ group; the molecule shown has no oxygen.
- (D) Ethyne (C_2H_2) has a triple bond (suffix "-yne"), not a double bond.

Key point: Suffix tells the bond: "-ane" (single), "-ene" (double), "-yne" (triple). Two carbons with a double bond = ethene.

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

Answer: (C) Go Back to Q24

Q25.

Solution

Concept — General formula of carboxylic acids: A saturated aliphatic monocarboxylic acid is an alkyl group attached to a $-$COOH group. Building from the smallest members, HCOOH (CH_2O_2, $n = 1$) and CH_3COOH ($C_2H_4O_2$, $n = 2$), the series fits the general formula $C_nH_{2n}O_2$.

Step 1 — Test $n = 1$ (formic acid, HCOOH): $C_1H_2O_2$ matches $C_nH_{2n}O_2$ with $n = 1$.

Step 2 — Test $n = 2$ (acetic acid, CH_3COOH): $C_2H_4O_2$ matches with $n = 2$.

Step 3 — Conclusion: the general formula is $C_nH_{2n}O_2$.

Why each other option is wrong:

- (A) C_nH_{2n+2} is the formula of *alkanes*; it contains no oxygen.
- (C) $C_nH_{2n}O$ has only one oxygen, matching aldehydes/ketones, not the two oxygens of $-$COOH.
- (D) $C_nH_{2n+1}OH$ is the formula of *alcohols*, not carboxylic acids.

Key point: The $-$COOH group contributes two oxygens, so carboxylic acids follow $C_nH_{2n}O_2$. Check it against acetic acid ($C_2H_4O_2$).

Final Answer: $C_nH_{2n}O_2 \Rightarrow \boxed{\text{B}}$

Answer: (B) Go Back to Q25

Q26.

Solution

Concept — Identifying a triple bond: The bond type between carbons distinguishes the hydrocarbon families: alkanes have only single bonds (C–C), alkenes have a double bond (C=C), and alkynes have a triple bond (C≡C). Reading the structures, the compound drawn as HC≡CH contains the carbon–carbon triple bond. This is ethyne (acetylene), C_2H_2.

Step 1 — Examine each structure: (I) $H_3C–CH_3$ has a single bond (ethane); (II) $H_2C=CH_2$ has a double bond (ethene); (III) HC≡CH has a triple bond.

Step 2 — Select the triple bond: structure (III), HC≡CH, is the only one with a C≡C triple bond.

Step 3 — Name it: HC≡CH is ethyne (acetylene).

Why each other option is wrong:

- (A) Ethane (I) has a single C–C bond.
- (B) Ethene (II) has a C=C double bond, not a triple bond.
- (C) Methane has only one carbon, so it cannot contain a C–C bond at all.

Key point: Single → alkane, double → alkene, triple → alkyne. The triple bond (≡) appears only in ethyne (III).

Final Answer: Ethyne (III) $\Rightarrow \boxed{\text{D}}$

Answer: (D) Go Back to Q26

Q27.

Solution

Concept — Bond angle in methane: In CH_4 the carbon is sp^3 hybridised, forming four equivalent bonds with no lone pairs. Four electron domains repel each other most effectively when directed to the corners of a regular tetrahedron, which fixes the H–C–H angle at 109.5° (the tetrahedral angle).

Step 1 — Determine the hybridisation: carbon forms four σ bonds and has no lone pair → four domains → sp^3.

Step 2 — Geometry of four domains: they arrange tetrahedrally, giving the characteristic 109.5° angle.

Step 3 — Conclusion: the H–C–H bond angle in methane is 109.5°.

Why each other option is wrong:

- (B) 90° is the angle in an octahedral or square-planar geometry, not tetrahedral.
- (C) 120° is the trigonal-planar (sp^2) angle, as in ethene.
- (D) 180° is the linear (sp) angle, as in CO_2 or ethyne.

Key point: Four equal bonds with no lone pairs $\rightarrow$ tetrahedral $\rightarrow 109.5^\circ$. This regular tetrahedral angle is a defining feature of sp^3 carbon.

Final Answer: $109.5^\circ \Rightarrow \boxed{\text{A}}$

Answer: (A) Go Back to Q27

Q28.

Solution

Concept — The ester functional group: An ester contains the group $-$COO$-$ (written R–CO–O–R′): a carbonyl carbon (C=O) whose carbon is also bonded to a second oxygen that links to another carbon chain. Esters are formed when a carboxylic acid reacts with an alcohol (esterification), losing water.

Step 1 — Read the structure: a carbon is double-bonded to one oxygen (carbonyl) and single-bonded to a second oxygen that connects to R′, i.e. R–CO–O–R′.

Step 2 — Match the group to its class: this $-$COO$-$ linkage between two carbon groups is the ester group, as in ethyl acetate ($CH_3COOC_2H_5$).

Step 3 — Conclusion: the group is characteristic of esters.

Why each other option is wrong:

- (A) Aldehydes carry $-$CHO (carbonyl with an H), with no second oxygen.
- (B) Carboxylic acids carry $-$COOH, where the second oxygen bears a *hydrogen* ($-$OH), not a carbon chain.
- (D) Ethers carry C–O–C with *no* carbonyl (no C=O), unlike the ester group.

Key point: Ester $= -$COO$-$ (carbonyl $+$ a bridging oxygen to a carbon). Replacing the R′ with H gives a carboxylic acid; removing the carbonyl oxygen gives an ether.

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

Answer: (C) Go Back to Q28

Q29.

Solution

Concept — Fats as esters of glycerol: Fats and oils (triglycerides) are esters formed between one molecule of glycerol (a trihydric alcohol with three $-$OH groups) and three molecules of long-chain fatty acids ($-$COOH). Each $-$OH of glycerol esterifies with one fatty-acid $-$COOH, releasing water.

Step 1 — Identify the alcohol part: glycerol provides the three $-$OH groups.

Step 2 — Identify the acid part: three fatty-acid molecules supply the $-$COOH groups that esterify with glycerol.

Step 3 — Conclusion: fats are esters of glycerol and fatty acids (triglycerides).

Why each other option is wrong:

- (A) Amino acids are the monomers of *proteins*, not fats.
- (B) Glucose is the monomer of *carbohydrates* (starch, cellulose).
- (C) Nucleotides build *nucleic acids* (DNA, RNA).

Key point: Fat (triglyceride) $=$ glycerol $+$ 3 fatty acids, joined by ester linkages. This is why the hydrolysis of fats by alkali (saponification) gives glycerol and fatty-acid salts (soaps).

Final Answer: Fatty acids $\Rightarrow \boxed{\text{D}}$

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

Q30.

Solution

Concept — Basicity of the amino group: The nitrogen of an amino group ($-\mathrm{NH_2}$) carries a lone pair of electrons. By the Lewis and Bronsted–Lowry definitions, a species that donates an electron pair (or accepts a proton) is a base. The lone pair on nitrogen can accept a proton ($\mathrm{H^+}$) to form $-\mathrm{NH_3^+}$, so amines behave as bases.

Step 1 — Locate the lone pair: the nitrogen of $-\mathrm{NH_2}$ has one non-bonding lone pair.

Step 2 — Show proton acceptance: $\mathrm{R{-}NH_2 + H^+ \rightarrow R{-}NH_3^+}$, where the lone pair forms a bond to the incoming proton.

Step 3 — Conclusion: accepting a proton is base behaviour, so an $-\mathrm{NH_2}$ compound is a base.

Why each other option is wrong:

- (A) A strong acid *donates* protons; amines accept them, the opposite behaviour.
- (C) A neutral salt has no free lone pair available to accept protons in this way.
- (D) An oxidising agent gains electrons in redox; this is unrelated to the proton-accepting basicity of $-NH_2$.

Key point: The lone pair on the nitrogen of $-NH_2$ makes amines basic (proton acceptors). Compare $-COOH$, which donates protons and is acidic.

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

Answer: (B) Go Back to Q30

Answer Key

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