CHEMISTRY (CODE - 043)

CHAPTER TEST - STEP-BY-STEP SOLUTIONS AND MARKING SCHEME

Class XI (2026-27) | Chapter 5: Thermodynamics

Max. Marks: 35

SECTION A (12 × 1 = 12 marks)

Each correct answer carries 1 mark.

Q1. Ans: C. U and H [1]

A state function depends only on the initial and final states. U and H are state functions; q and w depend on the path.

Q2. Ans: A. open, closed, isolated [1]

X exchanges both matter and energy → open. Y exchanges energy only (lid, no matter) → closed. Z exchanges neither → isolated.

Q3. Ans: B. −2.49 kJ [1]

$\Delta n_g = 2 - 3 = -1$. $\Delta H - \Delta U = \Delta n_g RT = (-1)(8.314\ \text{J K}^{-1}\ \text{mol}^{-1})(300\ \text{K}) = -2494\ \text{J} \approx$ **−2.49 kJ**.

Q4. Ans: D. +6 L atm [1]

$w = -p_{ex}(V_f - V_i) = -2\ \text{atm} \times (2 - 5)\ \text{L} =$ **+6 L atm**. The sign is positive because work is done *on* the gas (area of the shaded rectangle = 2 × 3).

Q5. Ans: C. ½H_2(g) + ½Br_2(l) → HBr(g) [1]

$\Delta_f H°$ refers to the formation of **one mole** of the compound from its elements in their reference states (most stable form at 298 K, 1 bar). Bromine is a liquid at 298 K, so Br_2(l) is the reference state.

Q6. Ans: B. S; P [1]

$\Delta G = \Delta H - T\Delta S$. S (ΔH +, ΔS −): ΔG is positive at all T → non-spontaneous at all temperatures. P (ΔH −, ΔS +): ΔG is negative at all T → spontaneous at all temperatures.

Q7. Ans: A. −283.0 kJ mol^{-1} [1]

By Hess's law $\Delta H_3 = \Delta H_1 + \Delta H_2$, so $\Delta H_2 = -393.5 - (-110.5) =$ **−283.0 kJ mol^{-1}**.

Q8. Ans: B. 400 K [1]

Spontaneous when ΔG < 0, i.e. $T\Delta S > \Delta H \rightarrow T > \Delta H/\Delta S = 40 / 0.10 =$ **400 K**.

Q9. Ans: D. +300 J [1]

q = +500 J (absorbed); w = −200 J (work done by the system). $\Delta U = q + w = 500 - 200 =$ **+300 J**.

Q10. Ans: A. Both A and R are true, and R is the correct explanation of A. [1]

Because H is a state function, ΔH depends only on the initial and final states and not on the path (Hess's law).

Q11. Ans: B. Both A and R are true, and R is not the correct explanation of A. [1]

Both are true. ΔH_{vap} is positive because heat is absorbed to overcome the hydrogen bonds between water molecules; the increase in entropy does not explain the sign of the enthalpy change.

Q12. Ans: C. A is true but R is false. [1]

In free expansion $p_{ex} = 0$ (not infinite), so $w = -p_{ex}\Delta V = 0$; q = 0 for an ideal gas, hence ΔU = 0. A is true, R is false.

SECTION B (4 × 2 = 8 marks)

Q13. [2]

(i) P = steel vessel (bomb); Q = stirrer. [½] The apparatus measures **ΔU**: the bomb is sealed, so the volume is constant (ΔV = 0, w = 0) and $q_V = \Delta U$. [½]

(ii) Heat evolved = C × ΔT = 20.0 kJ K^{-1} × 2.0 K = 40.0 kJ for 1.2 g = 0.1 mol of C. [½]

ΔU = −40.0 kJ / 0.1 mol = −400 kJ mol^{-1}. For C(s) + O_2(g) → CO_2(g), $\Delta n_g = 0$, so **ΔH = ΔU = −400 kJ mol^{-1}**. [½]

Q14. [2]

(i) 78 g of benzene = 1 mol. Heat required = 1 mol × 30.8 kJ mol^{-1} = **30.8 kJ**. [1]

(ii) Water molecules are held together by strong **hydrogen bonds**, whereas acetone has only weaker dipole–dipole interactions. More heat is needed to overcome the stronger intermolecular forces in water (40.79 vs 29.1 kJ mol^{-1}). [1]

Q15. [2]

(a) Increases – liquid → gas, molecules become far more disordered. [½]

(b) Increases – a solid gives a solid and a gas; the number of gaseous molecules increases from 0 to 1. [½]

(c) Decreases – 4 mol of gas change to 2 mol of gas, so the disorder falls. [½]

(d) Decreases – liquid → ordered crystalline solid. [½]

Q16. [2]

(i) P: reactants are at a higher enthalpy than products → **exothermic**, ΔH is **negative**. Q: products are at a higher enthalpy → **endothermic**, ΔH is **positive**. [1]

(ii) Yes, it can be spontaneous. ΔG = ΔH − TΔS; if ΔS is positive and TΔS is large enough (usually at high temperature) to outweigh the positive ΔH, then ΔG is negative. [1]

OR

$\Delta n_g = 1 - 0 = 1$; ΔH = ΔU + Δn_gRT, so ΔU = ΔH − Δn_gRT. [1]

ΔU = 40.7 − (1)(8.314 × 10^{-3} kJ K^{-1} mol^{-1})(373 K) = 40.7 − 3.10 = **37.6 kJ mol^{-1}** [1]

SECTION C (2 × 3 = 6 marks)

Q17. [3]

(i) Bonds broken: 4 C–H and 2 O=O. Bonds formed: 2 C=O (in CO_2) and 4 O–H (in $2H_2O$). [1]

(ii) $\Delta_r H$ = Σ(bond enthalpies of reactants) − Σ(bond enthalpies of products). [½]

Reactants: 4 × 414 + 2 × 498 = 1656 + 996 = 2652 kJ. Products: 2 × 741 + 4 × 464 = 1482 + 1856 = 3338 kJ.

$\Delta_r H$ = 2652 − 3338 = **−686 kJ mol^{-1}** [½]

(iii) Mean (average) bond enthalpies are used – the actual enthalpy of a bond varies slightly from one compound to another – and the relation holds exactly only when all species are in the gaseous state. [1]

Q18. [3]

(i) If a reaction takes place in several steps, its standard reaction enthalpy is the sum of the standard enthalpies of the intermediate reactions into which the overall reaction may be divided at the same temperature. [1]

(ii) Add (1) and (3), and reverse (2): [1]

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ (1)

$H_2 + ½O_2 \rightarrow H_2O$ (3)

$2CO_2 + 3H_2O \rightarrow C_2H_6 + 7/2\ O_2$ (−2)

Sum: C_2H_4(g) + H_2(g) → C_2H_6(g)

(iii) $\Delta_r H$ = (−1411) + (−286) + (+1560) = **−137 kJ mol^{-1}**. [½] ΔH is negative → **exothermic**. [½]

SECTION D (1 × 4 = 4 marks)

Q19. [4]

(i) The standard enthalpy change when **one mole** of a compound is formed from its elements in their most stable states of aggregation (reference states) at 1 bar and the specified temperature. [1]

(ii) $\Delta_r H^\circ = \Sigma\, \Delta_f H^\circ(\text{products}) - \Sigma\, \Delta_f H^\circ(\text{reactants}) = [-635.1 + (-393.5)] - [-1206.9] =$ **+178.3 kJ mol^{-1}** [1]

(iii) $\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ = 178.3 - (298 \times 0.160) = 178.3 - 47.7 =$ **+130.6 kJ mol^{-1}**. $\Delta_r G^\circ$ is positive, so the decomposition is **non-spontaneous** at 298 K. [1]

(iv) Spontaneous when $\Delta_r G^\circ < 0$, i.e. $T > \Delta_r H^\circ/\Delta_r S^\circ = 178300 / 160 =$ **≈ 1114 K** (about 841 °C). [1]

SECTION E (1 × 5 = 5 marks)

Q20. [5]

(i) P – enthalpy of sublimation (atomisation) of M; Q – ionisation enthalpy of M; R – ½ × bond dissociation enthalpy of X_2 (enthalpy of atomisation of X); S – electron gain enthalpy of X. [½ each = 2]

(ii) The sum of enthalpy changes round the cycle is zero (Hess's law): [½]

$\Delta_f H = P + Q + R + S - \Delta_{lattice} H \rightarrow -400 = 100 + 500 + 120 - 350 - \Delta_{lattice} H = 370 - \Delta_{lattice} H$

$\Delta_{lattice} H = 370 + 400 =$ **+770 kJ mol^{-1}** [1]

(iii) $\Delta_{sol} H = \Delta_{lattice} H + \Delta_{hyd} H = +770 + (-760) =$ **+10 kJ mol^{-1}** [1]

(iv) Increase – the dissolution is endothermic ($\Delta_{sol} H$ positive), so a rise in temperature favours it. [½]

OR

(i) $\Delta G = \Delta H - T\Delta S$. The process is spontaneous if $\Delta G < 0$ (non-spontaneous if $\Delta G > 0$; $\Delta G = 0$ at equilibrium). [1]

(ii) $\Delta G = -92 - (298)(-0.200) = -92 + 59.6 =$ **−32.4 kJ mol^{-1}**; negative, so **spontaneous** at 298 K. [1]

(iii) $\Delta G = 0$ at $T = \Delta H/\Delta S = (-92)/(-0.200) =$ **460 K**. Above 460 K, ΔG becomes positive → non-spontaneous. [1]

(iv) $\Delta G = -92 - (600)(-0.200) = -92 + 120 =$ **+28 kJ mol^{-1}** → non-spontaneous at 600 K (consistent with (iii)). [1]

(v) $\Delta_r G^\circ = -RT \ln K$ (= $-2.303\, RT \log K$). A large negative $\Delta_r G^\circ$ means K is much greater than 1: the reaction goes nearly to completion. [1]

***** END OF SOLUTIONS *****