CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 5: Chemical Thermodynamics – Free PDF Download

Question: Choose the correct answer. A thermodynamic state function is a quantity (i) used to determine heat changes (ii) whose value is independent of path (iii) used to determine pressure volume work (iv) whose value depends on temperature only.

Contents

Chapter 5 of NCERT Class 11 Chemistry, titled “Thermodynamics” (Unit 5 in the current 2026-27 reprint of Chemistry Part-I), builds on the idea of energy changes in chemical and physical processes, introducing state functions, the first law, enthalpy, Hess’s law, entropy and Gibbs energy so that students can predict whether a reaction is spontaneous.

NCERT Exercise Solutions

Question 5.1

Question: Choose the correct answer. A thermodynamic state function is a quantity (i) used to determine heat changes (ii) whose value is independent of path (iii) used to determine pressure volume work (iv) whose value depends on temperature only.

Solution: The correct answer is (ii) whose value is independent of path. A state function depends only on the initial and final states of the system and not on the path taken to reach that state — internal energy (U), enthalpy (H), entropy (S) and Gibbs energy (G) are all state functions, whereas heat (q) and work (w) are path functions.

Question 5.2

Question: For the process to occur under adiabatic conditions, the correct condition is: (i) $\Delta T = 0$ (ii) $\Delta p = 0$ (iii) $q = 0$ (iv) $w = 0$

Solution: The correct answer is (iii) $q = 0$. An adiabatic process is one in which no heat (q) is exchanged between the system and the surroundings, i.e. $q = 0$. (Note: options (i)–(iv) as printed in the textbook list (iv) as “$w = 0$”; the defining condition of an adiabatic process is $q = 0$.)

Question 5.3

Question: The enthalpies of all elements in their standard states are: (i) unity (ii) zero (iii) < 0 (iv) different for each element

Solution: The correct answer is (ii) zero. By convention, the standard enthalpy of formation of an element in its most stable (reference) state at 298 K and 1 bar pressure is taken as zero, since it forms the baseline from which enthalpies of formation of compounds are measured.

Question 5.4

Question: ΔU° of combustion of methane is $-$ X kJ mol^{-1}. The value of ΔH° is (i) $= \Delta U^\circ$ (ii) $> \Delta U^\circ$ (iii) $< \Delta U^\circ$ (iv) $= 0$

Solution: The correct answer is (iii) $< \Delta U^\circ$. For the combustion of methane, $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, the change in moles of gas is $\Delta n_g = 1 - (1 + 2) = -2$ (water is formed as a liquid, so it is not counted as a gas). Using $\Delta H^\circ = \Delta U^\circ + \Delta n_g RT$, since Δn_g is negative, $\Delta H^\circ = \Delta U^\circ - 2RT$, which is more negative than ΔU°. Given $\Delta U^\circ = -X$, ΔH° is therefore less than ΔU°.

Question 5.5

Question: The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are, -890.3 kJ mol^{-1} -393.5 kJ mol^{-1}, and -285.8 kJ mol^{-1} respectively. Enthalpy of formation of $CH_4(g)$ will be (i) -74.8 kJ mol^{-1} (ii) -52.27 kJ mol^{-1} (iii) $+74.8$ kJ mol^{-1} (iv) $+52.26$ kJ mol^{-1}.

Solution: The correct answer is (i) −74.8 kJ mol^{-1}. The formation reaction is C(graphite) + $2H_2(g) \rightarrow CH_4(g)$. By Hess's law, $\Delta_f H(CH_4) = \Delta_c H(C) + 2\Delta_c H(H_2) - \Delta_c H(CH_4) = (-393.5) + 2(-285.8) - (-890.3) = -393.5 - 571.6 + 890.3 = -74.8$ kJ mol^{-1}.

Question 5.6

Question: A reaction, A + B → C + D + q is found to have a positive entropy change. The reaction will be (i) possible at high temperature (ii) possible only at low temperature (iii) not possible at any temperature (iv) possible at any temperature

Solution: The correct answer is (iv) possible at any temperature. Since heat (q) is released, the reaction is exothermic, so ΔH is negative; it is also given that ΔS is positive. From ΔG = ΔH − TΔS, with ΔH negative and −TΔS also negative (since ΔS is positive), ΔG is negative at every temperature, so the reaction is spontaneous at any temperature.

Question 5.7

Question: In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process?

Solution: By the first law of thermodynamics, ΔU = q + w. Heat absorbed by the system, q = +701 J. Since work is done *by* the system (on the surroundings), w = −394 J. Therefore ΔU = 701 J + (−394 J) = 307 J.

Question 5.8

Question: The reaction of cyanamide, $NH_2CN(s)$, with dioxygen was carried out in a bomb calorimeter, and ΔU was found to be −742.7 kJ mol^{-1} at 298 K. Calculate enthalpy change for the reaction at 298 K.
$NH_2CN(g) + 3/2\ O_2(g) \rightarrow N_2(g) + CO_2(g) + H_2O(l)$

Solution: In a bomb calorimeter, volume is constant so the heat measured equals ΔU. To convert to ΔH, use $\Delta H = \Delta U + \Delta n_g RT$. Moles of gas: products = $N_2 + CO_2 = 1 + 1 = 2$ (H_2O is liquid, not counted); reactants = $O_2 = 3/2$. So $\Delta n_g = 2 - 1.5 = 0.5$. Using $R = 8.314 \times 10^{-3}$ kJ mol^{-1} K^{-1} and T = 298 K: $\Delta n_g RT = 0.5 \times 8.314 \times 10^{-3} \times 298 = 1.239$ kJ mol^{-1}. Therefore ΔH = −742.7 + 1.239 = −741.5 kJ mol^{-1} (approximately).

Question 5.9

Question: Calculate the number of kJ of heat necessary to raise the temperature of 60.0 g of aluminium from 35°C to 55°C. Molar heat capacity of Al is 24 J mol^{-1} K^{-1}.

Solution: Molar mass of Al ≈ 27 g mol^{-1}, so moles of Al = 60.0/27 = 2.22 mol. Rise in temperature, ΔT = 55 − 35 = 20°C = 20 K. Heat required, $q = n \times C_m \times \Delta T = 2.22$ mol × 24 J mol^{-1} K^{-1} × 20 K = 1066.7 J ≈ 1.07 kJ.

Question 5.10

Question: Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0°C to ice at −10.0°C. $\Delta_{fus}H$ = 6.03 kJ mol^{-1} at 0°C. $C_p[H_2O(l)] = 75.3$ J mol^{-1} K^{-1}, $C_p[H_2O(s)] = 36.8$ J mol^{-1} K^{-1}.

Solution: This is done in three steps using Hess's law:
Step 1 – cool liquid water from 10°C to 0°C: $\Delta H_1 = C_p(l) \times \Delta T = 75.3$ J mol^{-1} K^{-1} × (0 − 10) K = −753 J = −0.753 kJ.

Step 2 – freeze water to ice at 0 °C: this is the reverse of fusion, so $\Delta H_2 = -\Delta_{fus}H = -6.03$ kJ.
Step 3 – cool ice from 0 °C to −10 °C: $\Delta H_3 = C_p(s) \times \Delta T = 36.8\ J\ mol^{-1}\ K^{-1} \times (-10 - 0)\ K = -368\ J = -0.368$ kJ.
Total enthalpy change = $\Delta H_1 + \Delta H_2 + \Delta H_3 = -0.753 - 6.03 - 0.368 = -7.15\ kJ\ mol^{-1}$ (approximately).

Question 5.11

Question: Enthalpy of combustion of carbon to CO_2 is −393.5 kJ mol^{-1}. Calculate the heat released upon formation of 35.2 g of CO_2 from carbon and dioxygen gas.

Solution: Molar mass of CO_2 = 44 g mol^{-1}, so moles of CO_2 formed = 35.2/44 = 0.8 mol. Since 1 mol of CO_2 formation releases 393.5 kJ, heat released for 0.8 mol = 0.8 × 393.5 = 314.8 kJ.

Question 5.12

Question: Enthalpies of formation of CO(g), CO_2(g), N_2O(g) and N_2O_4(g) are −110, − 393, 81 and 9.7 kJ mol^{-1} respectively. Find the value of $\Delta_r H$ for the reaction: $N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$

Solution: $\Delta_r H = [\Delta_f H(N_2O) + 3\Delta_f H(CO_2)] - [\Delta_f H(N_2O_4) + 3\Delta_f H(CO)] = [81 + 3(-393)] - [9.7 + 3(-110)] = [81 - 1179] - [9.7 - 330] = (-1098) - (-320.3) = -777.7\ kJ\ mol^{-1}$.

Question 5.13

Question: Given $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$; $\Delta_r H° = -92.4\ kJ\ mol^{-1}$. What is the standard enthalpy of formation of NH_3 gas?

Solution: The given reaction produces 2 mol of NH_3, so the enthalpy of formation (per mole) is half of $\Delta_r H°$: $\Delta_f H°(NH_3) = -92.4/2 = -46.2\ kJ\ mol^{-1}$.

Question 5.14

Question: Calculate the standard enthalpy of formation of CH_3OH(l) from the following data: $CH_3OH(l) + 3/2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta_r H° = -726\ kJ\ mol^{-1}$. C(graphite) + $O_2(g) \rightarrow CO_2(g)$; $\Delta_c H° = -393\ kJ\ mol^{-1}$. $H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(l)$; $\Delta_f H° = -286\ kJ\ mol^{-1}$.

Solution: The target reaction is C(graphite) + $2H_2(g) + 1/2\ O_2(g) \rightarrow CH_3OH(l)$. By Hess's law: $\Delta_f H°(CH_3OH) = \Delta_c H°(C) + 2\Delta_f H°(H_2O) - \Delta_r H°$(combustion of methanol) = (−393) + 2(−286) − (−726) = −393 − 572 + 726 = −239 kJ mol^{-1}.

Question 5.15

Question: Calculate the enthalpy change for the process $CCl_4(g) \rightarrow C(g) + 4\ Cl(g)$ and calculate bond enthalpy of C–Cl in CCl_4(g). $\Delta_{vap}H°(CCl_4) = 30.5\ kJ\ mol^{-1}$. $\Delta_f H°(CCl_4) = -135.5\ kJ\ mol^{-1}$. $\Delta_a H°(C) = 715.0$ kJ mol^{-1}, where $\Delta_a H°$ is enthalpy of atomisation. $\Delta_a H°(Cl_2) = 242\ kJ\ mol^{-1}$.

Solution: Combine the given steps by Hess's law: reverse of vaporisation, $CCl_4(g) \rightarrow CCl_4(l)$, $\Delta H = -30.5$ kJ; reverse of formation, $CCl_4(l) \rightarrow$ C(graphite) + $2Cl_2(g)$, $\Delta H = +135.5$ kJ; atomisation of carbon, C(graphite) → C(g), $\Delta H = +715.0$ kJ; atomisation of chlorine (2 mol Cl_2), $2Cl_2(g) \rightarrow 4Cl(g)$, $\Delta H = 2 \times 242 = +484$ kJ. Adding these: $\Delta H = -30.5 + 135.5 + 715.0 + 484 = 1304\ kJ\ mol^{-1}$. This enthalpy corresponds to breaking 4 C–Cl bonds, so the average bond enthalpy of C–Cl = 1304/4 = 326 kJ mol^{-1}.

Question 5.16

Question: For an isolated system, $\Delta U = 0$, what will be ΔS?

Solution: In an isolated system there is no exchange of matter or energy with the surroundings. For any spontaneous process to occur in such a system, the entropy must increase, so ΔS will be positive ($\Delta S > 0$).

Question 5.17

Question: For the reaction at 298 K, $2A + B \rightarrow C$, $\Delta H = 400$ kJ mol^{-1} and $\Delta S = 0.2$ kJ K^{-1} mol^{-1}. At what temperature will the reaction become spontaneous considering ΔH and ΔS to be constant over the temperature range?

Solution: A reaction becomes spontaneous when $\Delta G = \Delta H - T\Delta S < 0$, i.e. when $T > \Delta H/\Delta S$. The boundary (equilibrium) temperature is $T = \Delta H/\Delta S = 400$ kJ mol^{-1} / 0.2 kJ K^{-1} mol^{-1} = 2000 K. So the reaction becomes spontaneous at temperatures above 2000 K.

Question 5.18

Question: For the reaction, $2\,Cl(g) \rightarrow Cl_2(g)$, what are the signs of ΔH and ΔS?

Solution: Formation of a Cl–Cl bond from two free chlorine atoms releases energy, so ΔH is negative. Also, two moles of gaseous atoms combine to give one mole of gaseous molecules, which decreases the randomness/disorder of the system, so ΔS is negative.

Question 5.19

Question: For the reaction $2\,A(g) + B(g) \rightarrow 2D(g)$, $\Delta U^\circ = -10.5$ kJ and $\Delta S^\circ = -44.1$ J K^{-1}. Calculate ΔG° for the reaction, and predict whether the reaction may occur spontaneously.

Solution: First find ΔH° from ΔU°. $\Delta n_g = 2$ (product D) – (2 + 1) (reactants A and B) = –1. At T = 298 K, $\Delta H^\circ = \Delta U^\circ + \Delta n_g RT = -10.5$ kJ + $(-1)(8.314 \times 10^{-3}$ kJ mol^{-1} $K^{-1})(298$ K) = –10.5 – 2.478 = –12.98 kJ. Now $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -12.98$ kJ – (298 K)$(-44.1 \times 10^{-3}$ kJ $K^{-1})$ = –12.98 + 13.14 = +0.16 kJ (approximately). Since ΔG° is positive (though close to zero), the reaction is not spontaneous under standard conditions at 298 K.

Question 5.20

Question: The equilibrium constant for a reaction is 10. What will be the value of ΔG°? R = 8.314 JK^{-1} mol^{-1}, T = 300 K.

Solution: Using $\Delta G^\circ = -RT \ln K = -2.303\ RT \log K$: $\Delta G^\circ = -2.303 \times 8.314$ J K^{-1} $mol^{-1} \times 300$ K $\times \log(10) = -2.303 \times 8.314 \times 300 \times 1 = -5744.6$ J $mol^{-1} \approx -5.74$ kJ mol^{-1}.

Question 5.21

Question: Comment on the thermodynamic stability of NO(g), given $1/2\ N_2(g) + 1/2\ O_2(g) \rightarrow NO(g)$; $\Delta_r H^\circ = 90$ kJ mol^{-1}. $NO(g) + 1/2\ O_2(g) \rightarrow NO_2(g)$; $\Delta_r H^\circ = -74$ kJ mol^{-1}.

Solution: The formation of NO(g) from N_2 and O_2 is endothermic ($\Delta_r H^\circ = +90$ kJ mol^{-1}), which means NO(g) is at a higher energy level than the elements it is formed from and is therefore thermodynamically unstable with respect to decomposition back into N_2 and O_2. Additionally, the oxidation of NO to NO_2 is exothermic ($\Delta_r H^\circ = -74$ kJ mol^{-1}), showing that NO is also unstable with respect to further oxidation and

readily combines with more oxygen to form the more stable NO_2.

Question 5.22

Question: Calculate the entropy change in surroundings when 1.00 mol of $H_2O(l)$ is formed under standard conditions. $\Delta_f H^\circ = -286$ kJ mol^{-1}.

Solution: Since the reaction is exothermic, heat equal to 286 kJ is released to the surroundings, i.e. $q_{surroundings} = +286$ kJ = 286000 J (at standard temperature, T = 298 K). Entropy change of the surroundings, $\Delta S_{surr} = -\Delta H_{system}/T$ = 286000 J / 298 K = 959.7 J K^{-1} mol^{-1} (approximately).

Notes and Extra Questions

This chapter lays the conceptual foundation of chemical thermodynamics for Class 11: the distinction between system, surroundings, state functions and path functions; the first law of thermodynamics ($\Delta U = q + w$) and its application to enthalpy changes at constant pressure; the use of Hess's law to calculate reaction enthalpies indirectly (formation, combustion, atomisation and bond enthalpies); and finally entropy and Gibbs energy as criteria for predicting the spontaneity of a process. Numerical questions in this chapter regularly combine two or more of these ideas (for example, converting ΔU to ΔH using $\Delta n_g RT$, or using both ΔH and ΔS to compute ΔG), so students should practise unit consistency (J vs kJ) and sign conventions carefully, since these are the most common sources of error.