NCERT Exercise Questions and Answers

1. 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?

Ans: Since heat is absorbed, it is given positive sign, i.e., $q = 701\ J$

Since work is done by the system, it is given negative sign i.e., $w = -394\ J$

$\therefore\ \Delta U = q + w = 701 - 394 = +307\ J$

2. The reaction of cyanamide, $NH_2CN(s)$ with O_2 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.

$$NH_2CN\,(s) + \frac{3}{2}O_2\,(g) \longrightarrow N_2\,(g) + CO_2\,(g) + H_2O\,(l)$$

Ans: $\Delta n\,(g) = (1+1) - \frac{3}{2} = +\frac{1}{2}$

$\Delta U = -742.7\ kJ\ mol^{-1}$, $R = 8.314 \times 10^{-3}\ kJ\ mol^{-1}\ K^{-1}$

$\Delta H = \Delta U + \Delta n\,RT = -742.7 + \frac{1}{2} \times 8.314 \times 10^{-3} \times 298 = \mathbf{-741.5\ kJ}$

3. Calculate the number of kJ of heat necessary to raise the temperatures of 60 g of Al from 35°C to 55°C. Molar heat capacity of Al is $24\ J\ mol^{-1}\ K^{-1}$

Ans: Atomic weight of $Al = 27$

$\therefore$ No of moles of $Al = \frac{60}{27}\ mol$

$$q = nC\Delta T = \frac{60}{27} \times 24\,(55-35)\,J = 1066.7\ J = 1.0667\ kJ = 1.07\ kJ$$

4. Enthalpy of combustion of C 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 oxygen gas.

Ans:

$C(s) + O_2\,(g) \longrightarrow CO_2\,(g) \quad \Delta H = -393.5\ kJ\ mol^{-1}$

Heat released in the formation of 1 mol of CO_2 (i.e., 44 g CO_2) = 393.5 kJ

$\therefore$ Heat released in the formation of 35.2 g CO_2 = $\frac{393.5}{44} \times 35.2 = \mathbf{314.8\ kJ}$

or ΔH for the formation of 35.2 g CO_2 = $\mathbf{-314.8\ kJ}$

5. Ethalpies 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 $\Delta_r H$ from the reaction

$$N_2O_4\,(g) + 3CO\,(g) \longrightarrow N_2O\,(g) + 3CO_2\,(g)$$

Ans:

$$\Delta_r H = \left[\Delta_f H\,(N_2O) + 3\Delta_f H\,(CO_2)\right] - \left[\Delta_f H\,(N_2O_4) + 3\Delta_f H\,(CO)\right]$$
$$= [81 + 3 \times (-393)] - [9.7 + 3 \times (-110)] = \mathbf{-777.7\ kJ}$$

6. Given that

$$N_2\,(g) + 3H_2\,(g) \longrightarrow 2NH_3\,(g) \qquad \Delta_r H^\theta = -92.4\ kJ$$

What is the standard enthalpy of formation of NH_3 gas?

Ans: ΔH^0 for the formation of 2 moles of NH_3 from its elements = $-92.4\ kJ$

$\therefore$ ΔH^0 for the formation of 1 mole of NH_3 from its elements (i.e., standard enthalpy of formation of NH_3 $= \dfrac{-92.4}{2} = \mathbf{-46.2\ kJ\,mol^{-1}}$

7. Calculate the standard enthalpy of formation of $CH_3OH\,(l)$ from the following data

$$CH_3OH\,(l) + \frac{3}{2}O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(l) \qquad \Delta H^\theta = -726\ kJ \quad \text{---- (1)}$$

$$C\,(s) + O_2\,(g) \longrightarrow CO_2\,(g) \qquad \Delta H^\theta = -393\ kJ \quad \text{---- (2)}$$

$$H_2\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow H_2O\,(l) \qquad \Delta H^\theta = -286\ kJ \quad \text{---- (3)}$$

Ans: The required equation (i.e., equation representing the formation of 1 mol of CH_3OH from the constituent elements is

$$C\,(s) + 2H_2\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow CH_3OH\,(l) \qquad \Delta H^\theta = ?$$

Multiplying equation (3) by 2, we get

$$2H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(l) \qquad \Delta H^\theta = -572\ kJ \quad \text{---- (4)}$$

Equations (2) + (4) – (1) gives

$$C\,(s) + 2H_2\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow CH_3OH\,(l) \qquad \Delta H^\theta = -393 - 572 - (-726)$$

$$= \mathbf{-239\ kJ}$$

8. Calculate the enthalpy change for the process $CCl_4\,(g) \longrightarrow C\,(g) + 4Cl\,(g)$ and calculate the bond enthalpy of $C - Cl$ bond in CCl_4 from the following data.

$$\Delta_{vap} H^\theta\,(CCl_4) = 30.5\ kJ\,mol^{-1}$$

$\Delta_f H^\theta (CCl_4) = -135.5 \; kJ\,mol^{-1}$

$\Delta_a H^\theta (C) = 715.0 \; kJ\,mol^{-1}$

$\Delta_a H^\theta (Cl_2) = 242 \; kJ\,mol^{-1}$

Ans: The given data may be written in the form of thermochemical equations as follows

$CCl_4\,(l) \longrightarrow CCl_4\,(g) \qquad \Delta H^\theta = 30.5 \; kJ\,mol^{-1}$ —— (1)

$C\,(s) + 2\,Cl_2\,(g) \longrightarrow CCl_4\,(l) \qquad \Delta H^\theta = -135.5 \; kJ\,mol^{-1}$ —— (2)

$C(s) \longrightarrow C\,(g) \qquad \Delta H^\theta = 715.0 \; kJ\,mol^{-1}$ —— (3)

$Cl_2\,(g) \longrightarrow 2Cl\,(g) \qquad \Delta H^\theta = 242.0 \; kJ\,mol^{-1}$ —— (4)

Multiplying equation (4) by 2, we get

$2Cl_2\,(g) \longrightarrow 4Cl\,(g) \qquad \Delta H^\theta = 484.0 \; kJ\,mol^{-1}$ —— (5)

Eqn (3) + Eqn (5) – Eqn (1) – Eqn (2) gives the required equation

$\therefore\ \Delta H$ for the reaction $= 715.0 + 484.0 - 30.5 - (-135.5) = \mathbf{1304\ kJ}$

Bond enthalpy of $C - Cl$ in $CCl_4 = \dfrac{1304}{4} = 326 \; kJ\,mol^{-1}$

9. For an isolated system, $\Delta U = 0$. What will be ΔS?

Ans: Consider an example of two gases contained separately in two bulbs connected by a stop-cock and isolated from the surroundings. On opening the stop-cock, the two gases mix up. i.e., $\Delta S > 0$.

10. For the reaction $2A + B \longrightarrow C$ at 298 K, $\Delta H = 400 \; kJ\,mol^{-1}$ 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.

Ans: At equilibrium, $\Delta G = 0$

$\therefore\ \Delta H - T\Delta S = 0 \quad$ or $\quad \Delta H = T\Delta S$

$$\therefore\ T = \frac{\Delta H}{\Delta S} = \frac{400}{0.2} = 2000\,K$$

i.e. the reaction will be at equilibrium at 2000 K. Since ΔH and ΔS are positive, ΔG will be negative when $T\Delta S > \Delta H$. This takes place at higher temperature. Therefore, the reaction will be spontaneous (i.e., ΔG is negative) at temperature above 2000 K.

11. For the reaction, $2Cl\,(g) \longrightarrow Cl_2\,(g)$, what are the signs of ΔH and ΔS?

Ans: Since the reaction involves formation of bonds, energy is released. So ΔH is negative. Since chlorine atoms combine together to form Cl_2 molecules, randomness decreases. i.e., ΔS is negative.

12. For the reaction $2A\,(g) + B\,(g) \longrightarrow 2D\,(g)$, $\Delta U^\theta = -10.5\,kJ$ and $\Delta S^\theta = -44.1\,JK^{-1}$. Calculate ΔG^θ for the reaction and predict whether the reaction may occur spontaneously.

Ans:

Δn for the reaction $= 2-(2+1) = -1$

$$\therefore\ \Delta H^\theta = \Delta U^\theta + \Delta n RT = -10.5 + (-1)\times 8.314\times 10^{-3}\,kJ \times 298 = -12.98\,kJ$$

$$\Delta G^\theta = \Delta H^\theta - T\Delta S = -12.98 - 298\,(-44.1\times 10^{-3}\,kJK^{-1}) = +0.16\,kJ$$

Since ΔG^θ is positive, the reaction will not occur spontaneously at 298 K.

(*Note:* Since standard values of ΔU and ΔS are given, the temperature is taken as 298 K)

13. The equilibrium constant for a reaction is 10. What will be the value of ΔG^θ at 300 K

Ans:

$$\Delta G^\theta = -2.303\,RT\log K = -2.303\times 8.314\times 300\times \log 10$$
$$= -2.303\times 8.314\times 300\times 1$$
$$= -5744.14\,J = -57.44\,kJ$$

14. Comment on the thermodynamic stability of $NO\,(g)$ from the following data

$$\frac{1}{2}N_2\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow NO\,(g) \qquad \Delta H^\theta = 90\,kJ\,mol^{-1}$$

$$NO\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow NO_2\,(g) \qquad \Delta H^\theta = -74\,kJ\,mol^{-1}$$

Ans: Since energy is absorbed in the formation of NO, it is not very stable. Since energy is released in the second reaction, NO will be easily converted to the more stable NO_2

15. Calculate the entropy change in the surroundings when 1 mole of $H_2O\,(l)$ is formed under standard conditions. $\Delta_f H^\theta = -286\,kJ\,mol^{-1}$.

Ans: The equation representing the enthalpy of formation of $H_2O\,(l)$ is

$$H_2\,(g) + \frac{1}{2}O_2\,(g) \longrightarrow H_2O\,(l) \qquad \Delta_f H^\theta = -286\,kJ\,mol^{-1}$$

i.e., when 1 mole of $H_2O\,(l)$ is formed 286 kJ of heat is evolved. This heat is absorbed by the surroundings.

$$\therefore\ q_{\text{surroundings}} = +286\,kJ\,mol^{-1}$$

$$\Delta S_{\text{(surroundings)}} = \frac{q}{T} = \frac{286\,kJ\,mol^{-1}}{298} = 0.9597\,kJ\,K^{-1}\,mol^{-1} = 959.7\,JK^{-1}\,mol^{-1}$$