

## NCERT Exercise Questions and Answers

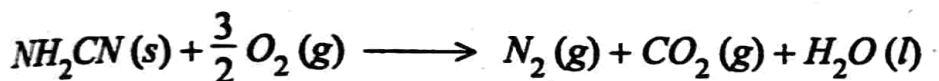
1. In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process?

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

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

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

2. The reaction of cyanamide,  $\text{NH}_2\text{CN}(s)$  with  $\text{O}_2$  was carried out in a bomb calorimeter and  $\Delta U$  was found to be  $-742.7 \text{ kJ mol}^{-1}$  at 298 K. Calculate enthalpy change for the reaction at 298 K.



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

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

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

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

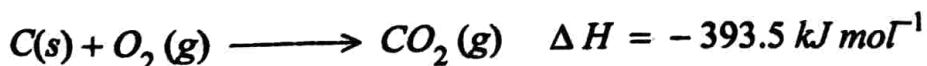
**Ans:** Atomic weight of Al = 27

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

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

4. Enthalpy of combustion of C to  $\text{CO}_2$  is  $-393.5 \text{ kJ mol}^{-1}$ . Calculate the heat released upon formation of 35.2 g of  $\text{CO}_2$  from carbon and oxygen gas.

**Ans:**



Heat released in the formation of 1 mol of  $\text{CO}_2$  (i.e., 44 g  $\text{CO}_2$ ) = 393.5 kJ

$$\therefore \text{Heat released in the formation of 35.2 g } \text{CO}_2 = \frac{393.5}{44} \times 35.2 = 314.8 \text{ kJ}$$

or  $\Delta H$  for the formation of 35.2 g  $\text{CO}_2 = -314.8 \text{ kJ}$

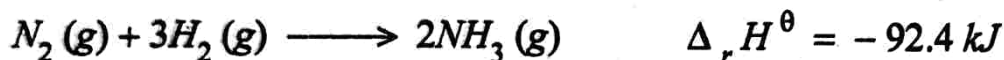
5. Enthalpies of formation of  $CO(g)$ ,  $CO_2(g)$ ,  $N_2O(g)$  and  $N_2O_4(g)$  are  $-110$ ,  $-393$ ,  $81$  and  $9.7 \text{ kJ mol}^{-1}$  respectively. Find the value of  $\Delta_r H$  from the reaction



**Ans:**

$$\begin{aligned} \Delta_r H &= [\Delta_f H(N_2O) + 3\Delta_f H(CO_2)] - [\Delta_f H(N_2O_4) + 3\Delta_f H(CO)] \\ &= [81 + 3 \times (-393)] - [9.7 + 3 \times (-110)] = -777.7 \text{ kJ} \end{aligned}$$

6. Given that



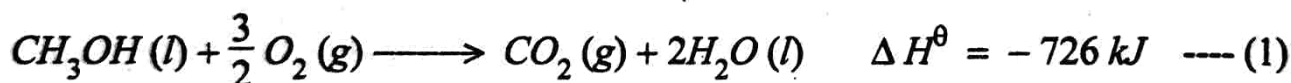
What is the standard enthalpy of formation of  $NH_3$  gas?

**Ans:**  $\Delta H^\theta$  for the formation of 2 moles of  $NH_3$  from its elements =  $-92.4 \text{ kJ}$

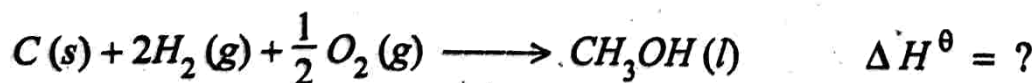
$\therefore \Delta H^\theta$  for the formation of 1 mole of  $NH_3$  from its elements (i.e., standard enthalpy of

$$\text{formation of } NH_3 = \frac{-92.4}{2} = -46.2 \text{ kJ mol}^{-1}$$

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



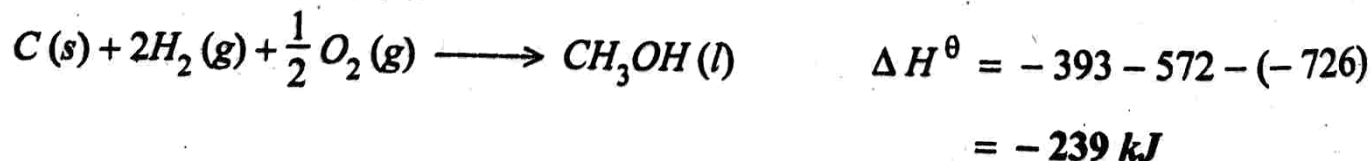
**Ans:** The required equation (i.e., equation representing the formation of 1 mol of  $CH_3OH$  from the constituent elements is



Multiplying equation (3) by 2, we get



Equations (2) + (4) - (1) gives



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.

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

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

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

$$\Delta_a H^\theta (Cl_2) = 242.0 \text{ kJ mol}^{-1}$$

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



Multiplying equation (4) by 2, we get



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

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

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

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

*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,  $\Delta H = 400 \text{ kJ mol}^{-1}$  and  $\Delta S = 0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}$ . At what temperature will the reaction become spontaneous considering  $\Delta H$  and  $\Delta S$  to be constant over the temperature range.

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

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

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

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

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

**Ans:** Since the reaction involves formation of bonds, energy is released. So  $\Delta H$  is negative. Since chlorine atoms combine together to form  $Cl_2$  molecules, randomness decreases. i.e.,  $\Delta S$  is negative.

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

**Ans:**

$$\Delta n \text{ for the reaction} = 2 - (2 + 1) = -1$$

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

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

Since  $\Delta G^\theta$  is positive, the reaction will not occur spontaneously at 298 K.

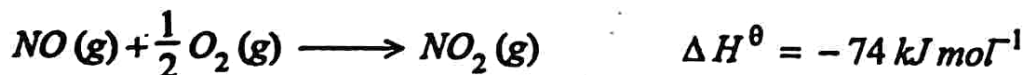
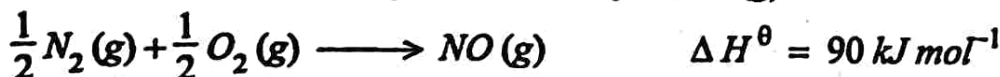
(Note: Since standard values of  $\Delta U$  and  $\Delta S$  are given, the temperature is taken as 298 K)

13. The equilibrium constant for a reaction is 10. What will be the value of  $\Delta G^\theta$  at 300 K

**Ans:**

$$\begin{aligned} \Delta G^\theta &= -2.303 RT \log K = -2.303 \times 8.314 \times 300 \times \log 10 \\ &= -2.303 \times 8.314 \times 300 \times 1 \\ &= -5744.14 \text{ J} = -57.44 \text{ kJ} \end{aligned}$$

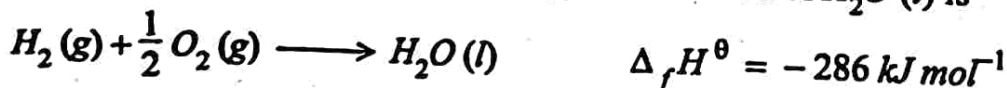
14. Comment on the thermodynamic stability of  $NO(g)$  from the following data



**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.  $\Delta_f H^\theta = -286 \text{ kJ mol}^{-1}$ .

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



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

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

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