

NCERT Exercise

Question 1:

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 1:

A thermodynamic state function is a quantity whose value is independent of a path. Functions like p , V , T etc. depend only on the state of a system and not on the path. Hence, alternative (ii) is correct.

Question 2:

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 2:

A system is said to be under adiabatic conditions if there is no exchange of heat between the system and its surroundings. Hence, under adiabatic conditions, $q = 0$. Therefore, alternative (iii) is correct.

Question 3:

The enthalpies of all elements in their standard states are:

- (i) unity
- (ii) zero
- (iii) < 0
- (iv) different for each element

Solution 3:

The enthalpy of all elements in their standard state is zero. Therefore, alternative (ii) is correct.

Question 4:

ΔU^θ of combustion of methane is $-X \text{ kJ mol}^{-1}$. The value of ΔH^θ is

- (i) $= \Delta U^\theta$
- (ii) $> \Delta U^\theta$
- (iii) $< \Delta U^\theta$
- (iv) $= 0$

Solution 4:

Since $\Delta H^\theta = \Delta U^\theta + \Delta n_g RT$ and $\Delta U^\theta = -X \text{ kJ mol}^{-1}$,

$$\Delta H^\theta = (-X) + \Delta n_g RT.$$

$$\Rightarrow \Delta H^\theta < \Delta U^\theta$$

Therefore, alternative (iii) is correct.

Question 5:

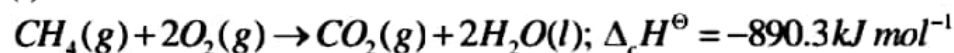
The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are, $-890.3 \text{ kJ mol}^{-1}$, $-393.5 \text{ kJ mol}^{-1}$, and $-285.8 \text{ kJ mol}^{-1}$ respectively. Enthalpy of formation of $\text{CH}_4(\text{g})$ will be

- (i) $-74.8 \text{ kJ mol}^{-1}$
- (ii) $-52.27 \text{ kJ mol}^{-1}$
- (iii) $+74.8 \text{ kJ mol}^{-1}$
- (iv) $+52.26 \text{ kJ mol}^{-1}$

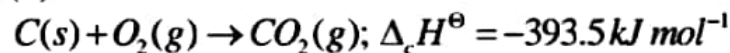
Solution 5:

According to the question,

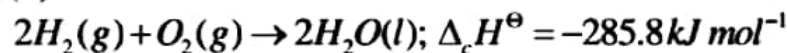
(i)



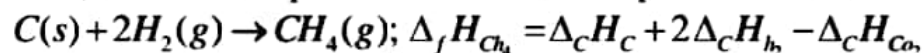
(ii)



(ii)



Thus, the desired equation is the one that represents the formation of $\text{CH}_4(\text{g})$ i.e.,



$$= (-393.5) + 2*(-285.8) - (-890.3) = -74.8 \text{ kJ mol}^{-1}$$

\therefore Enthalpy of formation of $\text{CH}_4(\text{g}) = -74.8 \text{ kJ mol}^{-1}$

Hence, alternative (i) is correct

Question 6:

A reaction, $A + B \rightarrow C + D + q$ 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 6:

For a reaction to be spontaneous, ΔG should be negative.

$$\Delta G = \Delta H - T\Delta S$$

According to the question, for the given reaction,

$$\Delta S = \text{positive}$$

$$\Delta H = \text{negative (since heat is evolved)} \Rightarrow \Delta G = \text{negative}$$

Therefore, the reaction is spontaneous at any temperature.

Hence, alternative (iv) is correct.

Question 7:

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 7:

According to the first law of thermodynamics,

$$\Delta U = q + W \text{ (i)}$$

Where,

ΔU = change in internal energy for a process

q = heat

W = work

Given,

$q = +701 \text{ J}$ (Since heat is absorbed)

$W = -394 \text{ J}$ (Since work is done by the system)

Substituting the values in expression (i), we get

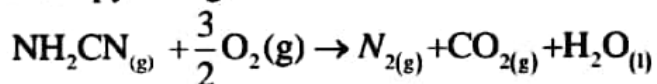
$$\Delta U = 701 \text{ J} + (-394 \text{ J})$$

$$\Delta U = 307 \text{ J}$$

Hence, the change in internal energy for the given process is 307 J.

Question 8:

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



Solution 8:

Enthalpy change for a reaction (ΔH) is given by the expression,

$$\Delta H = \Delta U + \Delta n_g RT$$

Where,

ΔU = change in internal energy

Δn_g = change in number of moles

For the given reaction,

$$\Delta n_g = \Sigma n_g (\text{products}) - \Sigma n_g (\text{reactants})$$

$$= (2 - 1.5) \text{ moles}$$

$$\Sigma n_g = +0.5 \text{ moles}$$

And, $\Delta U = -742.7 \text{ kJ mol}^{-1}$

$T = 298 \text{ K}$

$R = 8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1}$

Substituting the values in the expression of ΔH :

$$\Delta H = (-742.7 \text{ kJ mol}^{-1}) + (+0.5 \text{ mol}) (298 \text{ K}) 8.314 \times 10^{-3} \text{ kJ mol}^{-1} \text{ K}^{-1}$$

$$= -742.7 + 1.2$$

$$\Delta H = -741.5 \text{ kJ mol}^{-1}$$

Question 9:

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

Solution 9:

From the expression of heat (q),

$$q = m \cdot c \cdot \Delta T$$

Where,

c = molar heat capacity

m = mass of substance

ΔT = change in temperature

Substituting the values in the expression of

$$q = \left(\frac{60}{27} \text{ mol} \right) (24 \text{ J mol}^{-1} \text{ K}^{-1}) (20 \text{ K})$$

$$q = 1066.7 \text{ J}$$

$$q = 1.07 \text{ kJ}$$

Question 10:

Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0°C to ice at -10.0°C. $\Delta_{\text{fus}} H = 6.03 \text{ kJ mol}^{-1}$ at 0°C.

$$C_p [\text{H}_2\text{O}(l)] = 75.3 \text{ J mol}^{-1} \text{ K}^{-1};$$

$$C_p [\text{H}_2\text{O}(s)] = 36.8 \text{ J mol}^{-1} \text{ K}^{-1}.$$

Solution 10:

Total enthalpy change involved in the transformation is the sum of the following changes:

(a) Energy change involved in the transformation of 1 mol of water at 10°C to 1 mol of water at 0°C.

(b) Energy change involved in the transformation of 1 mol of water at 0°C to 1 mol of ice at 0°C.

(c) Energy change involved in the transformation of 1 mol of ice at 0°C to 1 mol of ice at -10°C.

$$\begin{aligned} \text{Total } \Delta H &= C_p [\text{H}_2\text{O}(l)] \Delta T + \Delta H_{\text{freezing}} + C_p [\text{H}_2\text{O}(s)] \Delta T \\ &= (75.3 \text{ J mol}^{-1} \text{ K}^{-1}) (0 - 10) \text{ K} + (-6.03 \times 10^3 \text{ J mol}^{-1}) + (36.8 \text{ J mol}^{-1} \text{ K}^{-1}) (-10 - 0) \text{ K} \\ &= -753 \text{ J mol}^{-1} - 6030 \text{ J mol}^{-1} - 368 \text{ J mol}^{-1} \\ &= -7151 \text{ J mol}^{-1} \\ &= -7.151 \text{ kJ mol}^{-1} \end{aligned}$$

Hence, the enthalpy change involved in the transformation is -7.151 kJ mol⁻¹

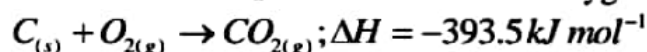
Question 11:

Enthalpy of combustion of carbon to carbon dioxide is -393.5 kJ mol⁻¹

Calculate the heat released upon formation of 35.2 g of CO₂ from carbon and dioxygen gas.

Solution 11:

Formation of CO₂ from carbon and dioxygen gas can be represented as



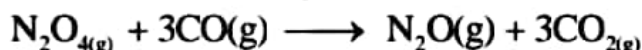
(1 mole = 44g)

Heat released in the formation of 44 g of CO₂ = 393.5 kJ mol⁻¹

Heat released in the formation of 35.2 g of $\text{CO}_2 = (393.5 \text{ KJ}) \times (35.2 \text{ g}) / (44 \text{ g}) = 314.8 \text{ KJ}$

Question 12:

Enthalpies of formation of $\text{CO}(g)$, $\text{CO}_2(g)$, $\text{N}_2\text{O}(g)$ and $\text{N}_2\text{O}_4(g)$ are -110 , -393 , 81 kJ and 9.7 kJ mol^{-1} respectively. Find the value of $\Delta_r H$ for the reaction:



Solution 12:

$\Delta_r H$ for a reaction is defined as the difference between $\Delta_f H$ value of products and $\Delta_f H$ value of reactants.

$$\Delta_r H = \sum \Delta_f H (\text{product}) - \sum \Delta_f H (\text{reactant})$$

For the given reaction,



$$\Delta_r H = [\{\Delta_f H (\text{N}_2\text{O}) + 3\Delta_f H (\text{CO}_2)\} - \{\Delta_f H (\text{N}_2\text{O}_4) + 3\Delta_f H (\text{CO})\}]$$

Substituting the values of $\Delta_f H$ for N_2O , CO_2 , N_2O_4 , and CO from the question, we get:

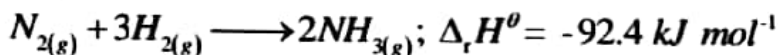
$$\Delta_r H = [81 \text{ kJ mol}^{-1} + 3(-393) \text{ KJ mol}^{-1}] - [9.7 \text{ kJ mol}^{-1} + 3(-110) \text{ KJ mol}^{-1}]$$

$$\Delta_r H = -777.7 \text{ kJ mol}^{-1}$$

Hence, the value of $\Delta_r H$ for the reaction is $-777.7 \text{ kJ mol}^{-1}$.

Question 13:

Given

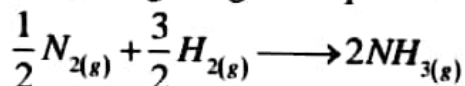


What is the standard enthalpy of formation of NH_3 gas?

Solution 13:

Standard enthalpy of formation of a compound is the change in enthalpy that takes place during the formation of 1 mole of a substance in its standard form from its constituent elements in their standard state.

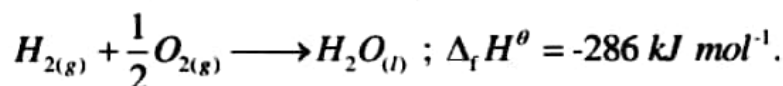
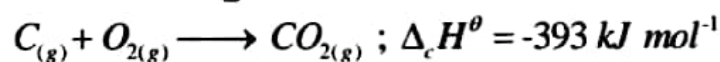
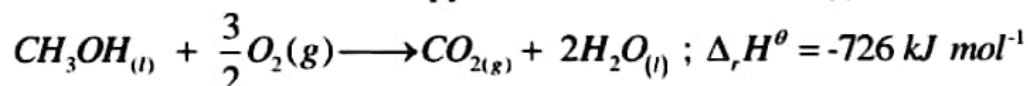
Re-writing the given equation for 1 mole of $\text{NH}_3(g)$,



$$\begin{aligned}
 \therefore \text{Standard enthalpy of formation of } \text{NH}_3(\text{g}) & \\
 = \frac{1}{2} \Delta_r H^\theta & \\
 = \frac{1}{2} (-92.4 \text{ kJ mol}^{-1}) & \\
 = -46.2 \text{ kJ mol}^{-1} &
 \end{aligned}$$

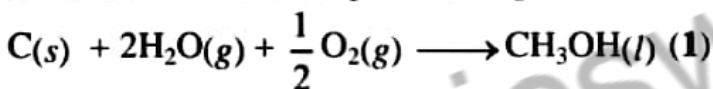
Question 14:

Calculate the standard enthalpy of formation of $\text{CH}_3\text{OH}(\text{l})$ from the following data:



Solution 14:

The reaction that takes place during the formation of $\text{CH}_3\text{OH}(\text{l})$ can be written as:



The reaction (1) can be obtained from the given reactions by following the algebraic calculations as:

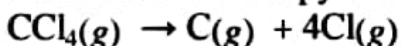
Equation (ii) + 2 × equation (iii) - equation (i)

$$\begin{aligned}
 \Delta_f H^\theta [\text{CH}_3\text{OH}(\text{l})] &= \Delta_c H^\theta + 2\Delta_f H^\theta [\text{H}_2\text{O}(\text{l})] - \Delta_r H^\theta \\
 &= (-393 \text{ kJ mol}^{-1}) + 2(-286 \text{ kJ mol}^{-1}) - (-726 \text{ kJ mol}^{-1}) \\
 &= (-393 - 572 + 726) \text{ kJ mol}^{-1}
 \end{aligned}$$

$$\therefore \Delta_f H^\theta [\text{CH}_3\text{OH}(\text{l})] = -239 \text{ kJ mol}^{-1}$$

Question 15:

Calculate the enthalpy change for the process



and calculate bond enthalpy of C-Cl in $\text{CCl}_4(\text{g})$.

$$\Delta_{\text{vap}} H^\theta (\text{CCl}_4) = 30.5 \text{ kJ mol}^{-1}.$$

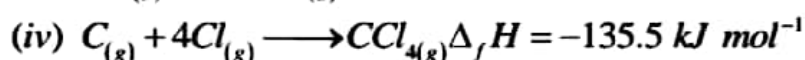
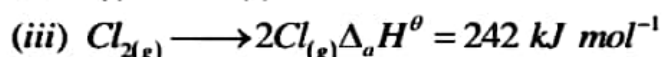
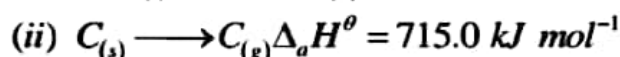
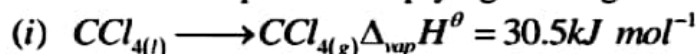
$$\Delta_f H^\theta (\text{CCl}_4) = -135.5 \text{ kJ mol}^{-1}.$$

$$\Delta_a H^\theta (\text{C}) = 715.0 \text{ kJ mol}^{-1}, \text{ where } \Delta_a H^\theta \text{ is enthalpy of atomisation}$$

$$\Delta_a H^\theta (\text{Cl}_2) = 242 \text{ kJ mol}^{-1}$$

Solution 15:

The chemical equations implying to the given values of enthalpies are:



Enthalpy change for the given process $\text{CCl}_{4(g)} \longrightarrow \text{C}_{(g)} + 4\text{Cl}_{(g)}$ can be calculated

using the following algebraic calculations as:

Equation (ii) + 2 × Equation (iii) - Equation (i) - Equation (iv)

$$\Delta H = \Delta_a H^\theta (\text{C}) + 2\Delta_a H^\theta (\text{Cl}_2) - \Delta_{\text{vap}} H^\theta - \Delta_f H$$

$$= (715.0 \text{ kJ mol}^{-1}) + 2(242 \text{ kJ mol}^{-1}) - (30.5 \text{ kJ mol}^{-1}) - (-135.5 \text{ kJ mol}^{-1})$$

$$\therefore \Delta H = 1304 \text{ kJ mol}^{-1}$$

Bond enthalpy of C-Cl bond in $\text{CCl}_4(g)$

$$= \frac{1304}{4} \text{ kJ mol}^{-1}$$

$$= 326 \text{ kJ mol}^{-1}$$

Question 16:

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

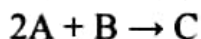
Solution 16:

ΔS will be positive i.e., greater than zero.

Since $\Delta U = 0$, ΔS will be positive and the reaction will be spontaneous.

Question 17:

For the reaction at 298 K,



$$\Delta H = 400 \text{ kJ mol}^{-1} \text{ and } \Delta S = 0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}$$

At what temperature will the reaction become spontaneous considering ΔH and ΔS to be constant over the temperature range?

Solution 17:

From the expression,

$$\Delta G = \Delta H - T\Delta S$$

Assuming the reaction at equilibrium, ΔT for the reaction would be:

$$T = (\Delta H - \Delta G) \frac{1}{\Delta S}$$

$$= \frac{\Delta H}{\Delta S}$$

($\Delta G = 0$ at equilibrium)

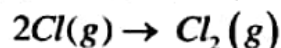
$$= \frac{400 \text{ kJ mol}^{-1}}{0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}}$$

$$T = 2000 \text{ K}$$

For the reaction to be spontaneous, ΔG must be negative. Hence, for the given reaction to be spontaneous, T should be greater than 2000 K.

Question 18:

For the reaction,



what are the signs of ΔH and ΔS ?

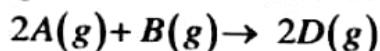
Solution 18:

ΔH and ΔS are negative

The given reaction represents the formation of chlorine molecule from chlorine atoms. Here, bond formation is taking place. Therefore, energy is being released. Hence, ΔH is negative.

Also, two moles of atoms have more randomness than one mole of a molecule. Since spontaneity is decreased, ΔS is negative for the given reaction.

Question 19: For the reaction $2\text{A}(g) + \text{B}(g) \rightarrow 2\text{D}(g)$

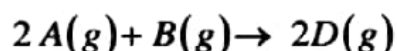


$$\Delta U^\theta = -10.5 \text{ kJ and } \Delta S^\theta = -44.1 \text{ JK}^{-1}.$$

Calculate ΔG^θ for the reaction, and predict whether there action may occur spontaneously.

Solution 19:

For the given reaction,



$$\Delta n_g = 2 - (3)$$

$$= -1 \text{ mole}$$

Substituting the value of ΔU^θ in the expression of ΔH :

$$\Delta H^\theta = \Delta U^\theta + \Delta n_g RT$$

$$= (-10.5 \text{ kJ}) - (-1) (8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}) (298 \text{ K})$$

$$= -10.5 \text{ kJ} - 2.48 \text{ kJ}$$

$$\Delta H^\theta = -12.98 \text{ kJ}$$

Substituting the values of ΔH^θ and ΔS^θ in the expression of ΔG^θ :

$$\Delta G^\theta = \Delta H^\theta - T\Delta S^\theta$$

$$= -12.98 \text{ kJ} - (298 \text{ K}) (-44.1 \text{ J K}^{-1})$$

$$= -12.98 \text{ kJ} + 13.14 \text{ kJ}$$

$$\Delta G^\theta = +0.16 \text{ kJ}$$

Since ΔG^θ for the reaction is positive, the reaction will not occur spontaneously.

Question 20:

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

Solution 20:

From the expression,

$$\Delta G^\theta = -2.303 RT \log K_{eq}$$

ΔG^θ for the reaction,

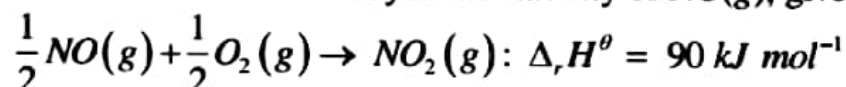
$$= (2.303) (8.314 \text{ JK}^{-1} \text{ mol}^{-1}) (300 \text{ K}) \log 10$$

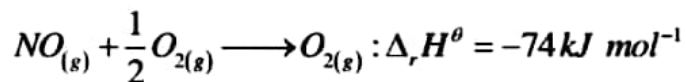
$$= -5744.14 \text{ J mol}^{-1}$$

$$= -5.744 \text{ kJ mol}^{-1}$$

Question 21:

Comment on the thermodynamic stability of $\text{NO}(g)$, given





Solution 21:

The positive value of $\Delta_r H$ indicates that heat is absorbed during the formation of $NO(g)$. This means that $NO(g)$ has higher energy than the reactants (N_2 and O_2).

Hence, $NO(g)$ is unstable.

The negative value of $\Delta_r H$ indicates that heat is evolved during the formation of $NO_2(g)$ from $NO(g)$ and $O_2(g)$. The product, $NO_2(g)$ is stabilized with minimum energy.

Hence, unstable $NO(g)$ changes to unstable $NO_2(g)$.

Question 22:

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

Solution 22:

It is given that 286 kJ mol^{-1} of heat is evolved on the formation of 1 mol of $H_2O(l)$. Thus, an equal amount of heat will be absorbed by the surroundings.

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

$$\text{Entropy change } (\Delta S_{surr}) \text{ for the surroundings} = \frac{q_{surr}}{T}$$

$$= \frac{286 \text{ kJ mol}^{-1}}{298 \text{ K}}$$

$$\therefore \Delta S_{surr} = 959.73 \text{ J mol}^{-1} \text{ K}^{-1}$$
