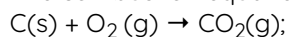




Question 11. Enthalpy of combustion of carbon to carbon dioxide is $-393.5 \text{ J mol}^{-1}$. Calculate the heat released upon formation of 35.2 g of CO_2 from carbon and oxygen gas.

Answer:

The combustion equation is:



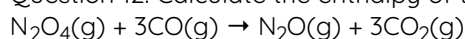
$$\Delta_c H = -393.5 \text{ KJ mol}^{-1}$$

Heat released in the formation of 44g of $\text{CO}_2 = 393.5 \text{ kJ}$

Heat released in the formation of 35.2 g of CO_2

$$= (393.5 \text{ KJ}) \times (35.2\text{g}) / (44\text{g}) = 314.8 \text{ kJ}$$

Question 12. Calculate the enthalpy of the reaction:



Given that; $\Delta_f H^\circ \text{CO}(\text{g}) = -110 \text{ kJ mol}^{-1}$;

$$\Delta_f H^\circ \text{CO}_2(\text{g}) = -393 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ \text{N}_2\text{O}(\text{g}) = 81 \text{ kJ mol}^{-1};$$

$$\Delta_f H^\circ \text{N}_2\text{O}_4(\text{g}) = 9.7 \text{ kJ mol}^{-1}$$

Answer:

$$\text{Enthalpy of reaction } (\Delta_r H) = [81 + 3(-393)] - [9.7 + 3(-110)]$$

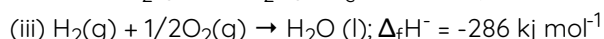
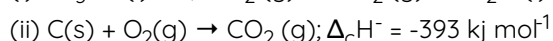
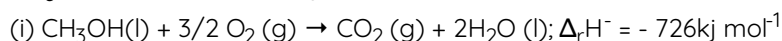
$$= [81 - 1179] - [9.7 - 330] = -778 \text{ kJ mol}^{-1}$$

Question 13. Given : $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$; $\Delta_r H^\circ = -92.4 \text{ kJ mol}^{-1}$

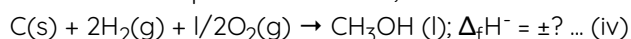
¹ What is the standard enthalpy of formation of NH_3 gas?

$$\text{Answer: } \Delta_f H^\circ \text{NH}_3(\text{g}) = - (92.4) / 2 = -46.2 \text{ kJ mol}^{-1}$$

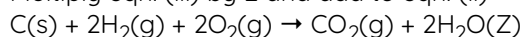
Question 14. Calculate the standard enthalpy of formation of CH_3OH . from the following data:



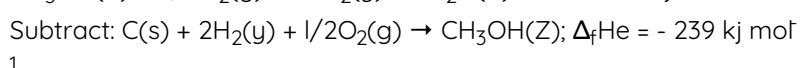
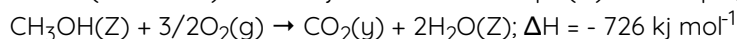
Answer: The equation we aim at;



Multiply eqn. (iii) by 2 and add to eqn. (ii)

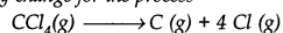


$$\Delta H = - (393 + 522) = -965 \text{ kJ mol}^{-1} \text{ Subtract eqn. (iv) from eqn. (i)}$$



Question 15.

Calculate the enthalpy change for the process



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

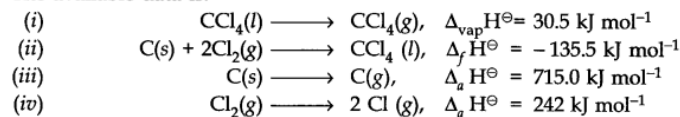
$$\text{Given: } \Delta_{\text{vap}} H^\circ (\text{CCl}_4) = 30.5 \text{ kJ mol}^{-1}; \Delta_f H^\circ (\text{CCl}_4) = -135.5 \text{ kJ mol}^{-1}$$

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

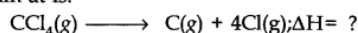
$$\Delta_a H^\circ (\text{Cl}_2) = 242 \text{ kJ mol}^{-1}.$$

Answer:

The available data is:



The equation we aim at is:



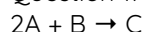
$$\begin{aligned}
 \text{Eqn. (iii)} + 2 \times \text{Eqn. (iv)} - \text{Eqn. (ii)} & \text{ gives the required equation with} \\
 \Delta H & = 715.0 + 2(242) - 30.5 - (-135.5) \text{ kJ mol}^{-1} \\
 & = 1304 \text{ kJ mol}^{-1}
 \end{aligned}$$

$$\text{Bond enthalpy of C-Cl in CCl}_4 \text{ (average value)} = \frac{1304}{4} = 326 \text{ kJ mol}^{-1}.$$

Question 16. For an isolated system $\Delta U = 0$; what will be ΔS ?

Answer: Change in internal energy (ΔU) for an isolated system is zero for it does not exchange any energy with the surroundings. But entropy tends to increase in case of spontaneous reaction. Therefore, $\Delta S > 0$ or positive.

Question 17. For a 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?

Answer:

As per the Gibbs Helmholtz equation:

$$\Delta G = \Delta H - T\Delta S \text{ For } \Delta G = 0;$$

$$\Delta H = T\Delta S \text{ or } T = \Delta H / \Delta S$$

$$T = (400 \text{ KJ mol}^{-1}) / (0.2 \text{ KJ K}^{-1} \text{ mol}^{-1}) = 2000 \text{ K}$$

Thus, reaction will be in a state of equilibrium at 2000 K and will be spontaneous above this temperature.

Question 18. For the reaction; $2\text{Cl(g)} \rightarrow \text{Cl}_2(\text{g})$; what will be the signs of ΔH and ΔS ?

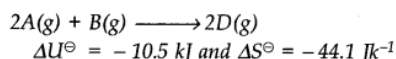
Answer:

ΔH : negative (-ve) because energy is released in bond formation.

ΔS : negative (-ve) because entropy decreases when atoms combine to form molecules.

Question 19.

For the reaction



$$\Delta U^\ominus = -10.5 \text{ kJ and } \Delta S^\ominus = -44.1 \text{ J K}^{-1}$$

Calculate ΔG^\ominus for the reaction, and predict whether the reaction may occur spontaneously.

Answer:

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

$$\Delta U^\ominus = -10.5 \text{ kJ}; \Delta n_g = 2 - 3 = -1 \text{ mol}$$

$$R = 8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}; T = 298 \text{ K}$$

$$\therefore \Delta H^\ominus = (-10.5 \text{ kJ}) + [(-1 \text{ mol}) \times (8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K})]$$

$$= -10.5 \text{ kJ} - 2.478 \text{ kJ} = -12.978 \text{ kJ}$$

According to Gibbs Helmholtz equation:

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

$$\Delta G^\ominus = (-12.978 \text{ kJ}) - (298 \text{ K}) \times (-0.0441 \text{ kJ K}^{-1})$$

$$= -12.978 + 13.112 = -12.978 + 13.112 = 0.164 \text{ kJ}$$

Since ΔG^\ominus is positive, the reaction is non-spontaneous in nature.

Question 20.

The equilibrium constant for the reaction is 10. Calculate the value of ΔG^\ominus ; Given

$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}; T = 300 \text{ K}.$$

Answer:

$$\Delta G^\ominus = -RT \ln K = -2.303 RT \log K.$$

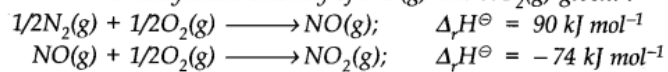
$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}; T = 300 \text{ K}; K = 10$$

$$\Delta G^\ominus = -2.303 \times 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times (300 \text{ K}) \times \log 10$$

$$= -5527 \text{ J mol}^{-1} = -5.527 \text{ kJ mol}^{-1}.$$

Question 21.

Comment on the thermodynamic stability of NO(g) and NO₂(g) given: :



Answer:

For NO (g) ; $\Delta_r H^\ominus = + \text{ve}$: Unstable in nature

For NO₂ (g) ; $\Delta_r H^\ominus = - \text{ve}$: Stable in nature ,

Question 22.

Calculate the entropy change in surroundings when 1.0 mol of H₂O(l) is formed under standard conditions. Given $\Delta H^\ominus = -286 \text{ kJ mol}^{-1}$.

Answer:

$$q_{\text{rev}} = (-\Delta_r H^\ominus) = -286 \text{ kJ mol}^{-1} = 286000 \text{ J mol}^{-1}$$

$$\Delta S_{(\text{Surroundings})} = \frac{q_{\text{rev}}}{T} = \frac{(286000 \text{ J mol}^{-1})}{298 \text{ K}} = 959 \text{ J K}^{-1} \text{ mol}^{-1}.$$

***** END *****