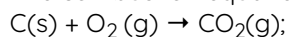




Question 11. Enthalpy of combustion of carbon to carbon dioxide is - 393.5 J mol<sup>-1</sup>. Calculate the heat released upon formation of 35.2 g of CO<sub>2</sub> from carbon and oxygen gas.

Answer:

The combustion equation is:



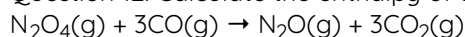
$$\Delta_c H = - 393.5 \text{ KJ mol}^{-1}$$

Heat released in the formation of 44g of CO<sub>2</sub> = 393.5 kJ

Heat released in the formation of 35.2 g of CO<sub>2</sub>

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

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:

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

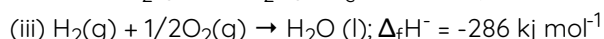
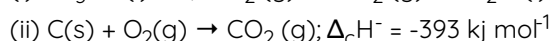
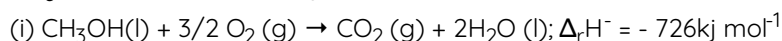
$$= [81 - 1179] - [9.7 - 330] = - 778 \text{ kJ mol}^{-1}$$

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

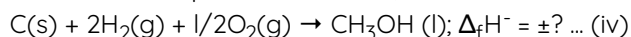
<sup>1</sup> What is the standard enthalpy of formation of NH<sub>3</sub> gas?

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

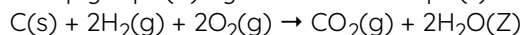
Question 14. Calculate the standard enthalpy of formation of CH<sub>3</sub>OH. from the following data:



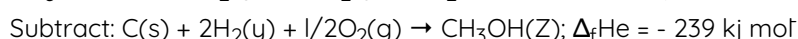
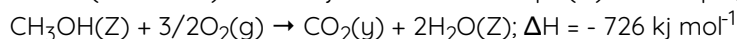
Answer: The equation we aim at;



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



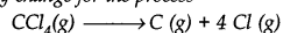
$$\Delta H = - (393 + 522) = - 915 \text{ kJ mol}^{-1} \text{ Subtract eqn. (iv) from eqn. (i)}$$



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Question 15.

Calculate the enthalpy change for the process



and calculate bond enthalpy of C-Cl in CCl<sub>4</sub>(g)

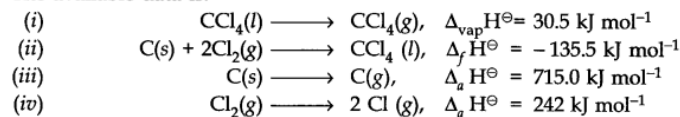
$$\text{Given: } \Delta_{\text{vap}} H^\circ (\text{CCl}_4) = 30.5 \text{ kJ mol}^{-1}; \Delta_f H^\circ (\text{CCl}_4) = - 135.5 \text{ kJ mol}^{-1}$$

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

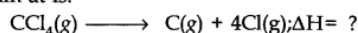
$$\Delta_a H^\circ (\text{Cl}_2) = 242 \text{ kJ mol}^{-1}.$$

Answer:

The available data is:



The equation we aim at is:



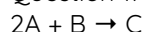
$$\begin{aligned}
 \text{Eqn. (iii)} + 2 \times \text{Eqn. (iv)} - \text{Eqn. (ii)} & \text{ gives the required equation with} \\
 \Delta H &= 715.0 + 2(242) - 30.5 - (-135.5) \text{ kJ mol}^{-1} \\
 &= 1304 \text{ kJ mol}^{-1}
 \end{aligned}$$

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

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

Answer: Change in internal energy ( $\Delta U$ ) 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  $\Delta H$  and  $\Delta S$  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;  $2\text{Cl}(g) \rightarrow \text{Cl}_2(g)$ ; what will be the signs of  $\Delta H$  and  $\Delta S$ ?

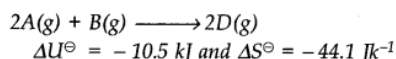
Answer:

$\Delta H$ : negative (-ve) because energy is released in bond formation.

$\Delta S$ : 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  $\Delta G^\ominus$  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  $\Delta G^\ominus$  is positive, the reaction is non-spontaneous in nature.

Question 20.

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

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

Answer:

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

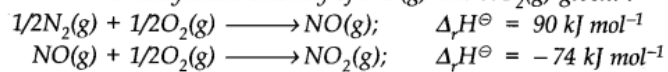
$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}; T = 300 \text{ K}; K = 10$$

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

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

Question 21.

Comment on the thermodynamic stability of NO(g) and NO<sub>2</sub>(g) given: :



Answer:

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

**For NO<sub>2</sub> (g) ;  $\Delta_r H^\ominus = - \text{ve}$  : Stable in nature** ,

Question 22.

Calculate the entropy change in surroundings when 1.0 mol of H<sub>2</sub>O(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 \*\*\*\*\*