Question 11. Enthalpy of combustion of carbon to carbon dioxide is - 393.5 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:

$C(s) + O_2(g) \rightarrow CO_2(g);$

AcH = - 393.5 KJ mol^{-1}

Heat released in the formation of 44g of CO_2 = 393.5 kj

Heat released in the formation of 35.2 g of CO_2

= (393.5 KJ) x (35.2g)/(44g) = 314.8 kj

Question 12. Calculate the enthalpy of the reaction:

$N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$

Given that; $\Delta_f H^- CO(g)$ = - 110 kj mot^{-1};

$\Delta_f HCO_2(g)$ = - 393 kj mol^{-1}

$\Delta_f HN_2O(g)$ = 81 kj mot^{-1};

$\Delta_f N_2O_4(g)$ = 9.7 kj mol^{-1}

Answer:

Enthalpy of reaction (Δ_r,H) = [81 + 3 (- 393)] - [9.7 + 3 (- 110)]

= [81 - 1179] - [9.7 - 330] = - 778 kj mol^{-1}

Question 13. Given : $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$; $\Delta_{r\ H^-}$ = -92.4 kj mot^{-1} What is the standard enthalpy of formation of NH_3 gas?

Answer: ΔH^- NH_3 (g) = - (92.4)/2 = - 46.2 kj mol^{1}

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

(i) $CH_3OH(l) + 3/2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta_r H^-$ = - 726kj mol^{-1}

(ii) $C(s) + O_2(g) \rightarrow CO_2(g)$; $\Delta_c H^-$ = -393 kj mol^{-1}

(iii) $H_2(g) + 1/2O_2(g) \rightarrow H_2O(l)$; $\Delta_f H^-$ = -286 kj mol^{-1}

Answer: The equation we aim at;

$C(s) + 2H_2(g) + l/2O_2(g) \rightarrow CH_3OH(l)$; $\Delta_f H^-$ = ±? ... (iv)

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

$C(s) + 2H_2(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(Z)$

ΔH = - (393 + 522) = - 965 kj moH Subtract eqn. (iv) from eqn. (i)

$CH_3OH(Z) + 3/2O_2(g) \rightarrow CO_2(y) + 2H_2O(Z)$; ΔH = - 726 kj mol^{-1}

Subtract: $C(s) + 2H_2(y) + l/2O_2(g) \rightarrow CH_3OH(Z)$; $\Delta_f He$ = - 239 kj mol^{-1}

Question 15.

Calculate the enthalpy change for the process

$$CCl_4(g) \longrightarrow C(g) + 4\,Cl(g)$$

and calculate bond enthalpy of C – Cl in $CCl_4(g)$

Given: $\Delta_{vap} H^\ominus (CCl_4)$ = 30.5 kJ mol^{-1}; $\Delta_f H^\ominus(CCl_4)$ = – 135.5 kJ mol^{-1}

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

$\Delta_a H^\ominus (Cl_2)$ = 242 kJ mol^{-1}.

Answer:

The available data is:

(i) $CCl_4(l) \longrightarrow CCl_4(g)$, $\Delta_{vap}H^\ominus = 30.5$ kJ mol^{-1}

(ii) $C(s) + 2Cl_2(g) \longrightarrow CCl_4(l)$, $\Delta_f H^\ominus = -135.5$ kJ mol^{-1}

(iii) $C(s) \longrightarrow C(g)$, $\Delta_a H^\ominus = 715.0$ kJ mol^{-1}

(iv) $Cl_2(g) \longrightarrow 2Cl(g)$, $\Delta_a H^\ominus = 242$ kJ mol^{-1}

The equation we aim at is:

$$CCl_4(g) \longrightarrow C(g) + 4Cl(g); \Delta H = ?$$

Eqn. (iii) + 2 × Eqn. (iv) – Eqn. (i) – Eqn. (ii) gives the required equation with

$$\Delta H = 715.0 + 2(242) - 30.5 - (-135.5) \text{ kJ mol}^{-1}$$
$$= 1304 \text{ kJ mol}^{-1}$$

Bond enthalpy of C–Cl in CCl_4 (average value) = $\frac{1304}{4}$ = **326 kJ mol⁻¹.**

Question 16.For an isolated system ΔU = 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, ΔS > 0 or positive.

Question 17. For a reaction at 298 K
2A + B → C
ΔH = 400 kj mol^{-1} and ΔS = 0.2 kj Kr^{-1} 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:
ΔG = ΔH- TΔS For ΔG=0 ;
ΔH=TΔS or T=ΔH/ΔS
T = (400 KJ mol^{-1})/(0.2 KJ K^{-1} mol^{-1}) = 2000 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; 2Cl(g) → Cl_2(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

$$2A(g) + B(g) \longrightarrow 2D(g)$$
$$\Delta U^\ominus = -10.5 \text{ kJ and } \Delta S^\ominus = -44.1 \text{ JK}^{-1}$$

Calculate $\Delta G^\ominus$ for the reaction, and predict whether the reaction may occur spontaneously.

Answer:

$$\Delta H^\ominus = \Delta U^\ominus + \Delta^{ng} RT$$
$$\Delta U^\ominus = -10.5 \text{ kJ}; \quad \Delta^{ng} = 2 - 3 = -1 \text{ mol}$$
$$R = 8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}; T = 298 \text{ K}$$
$$\therefore \quad \Delta H^\ominus = (-10.5 \text{ kJ}) + [(-1 \text{ mol}) \times (8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1}) \times (298\text{K})]$$
$$= -10.5 \text{ kJ} - 2.478 \text{ kJ} = -12.978 \text{ kJ}$$

According to Gibbs Helmholtz equation:

$$\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$$
$$\Delta G^\ominus = (-12.978 \text{ kJ}) - (298 \text{ K}) \times (-0.0441 \text{ kJ K}^{-1})$$
$$= -12.978 + 13.112 = -12.978 + 13.142 = \mathbf{0.164 \text{ kJ}}$$

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

Answer:

$$\Delta G^\ominus = -RT \ln K = -2.303 RT \log K.$$
$$R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}; T = 300 \text{ K}; K = 10$$
$$\Delta G^\ominus = -2.303 \times 8.314 \text{ JK}^{-1} \text{ mol}^{-1} \times (300 \text{ K}) \times \log 10$$
$$= -5527 \text{ J mol}^{-1} = \mathbf{-5.527 \text{ kJ mol}^{-1}}.$$

Question 21.

Comment on the thermodynamic stability of NO(g) and $NO_2(g)$ given: :

$$1/2N_2(g) + 1/2O_2(g) \longrightarrow NO(g); \quad \Delta_r H^\ominus = 90 \text{ kJ mol}^{-1}$$
$$NO(g) + 1/2O_2(g) \longrightarrow NO_2(g); \quad \Delta_r H^\ominus = -74 \text{ kJ mol}^{-1}$$

Answer:

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

For NO_2 (g) ; $\Delta_r H^\ominus$ = – ve : Stable in nature

Question 22.

Calculate the entropy change in surroundings when 1.0 mol of $H_2O(l)$ is formed under standard conditions. Given $\Delta H^\ominus = -286$ kJ mol^{-1}.

Answer:

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

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

*********** END **********