Chemical thermodynamics of materials 8th class pdf

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$$W_{\text{rev.}} = -\int_{V_i}^{V_f} P_{ex}\,dV = -\int_{V_i}^{V_f} (P_{in} \pm dP)\,dV$$

Since $dP \times dV$ is very small we can write,

$$W_{\text{rev.}} = -\int_{V_i}^{V_f} P_{in}\,dV$$

For ideal gas $PV = nRT$

$$P = \frac{nRT}{V}$$

Therefore, at constant temperature,

$$W_{\text{rev.}} = -\int_{V_i}^{V_f} nRT\frac{dV}{V} = -nRT \ln \frac{V_f}{V_i}$$

$$= -2.303\, nRT \log \frac{V_f}{V_i}$$

$$F = \frac{1}{g_c} ma$$

$$32.1740(\text{poundal}) = \frac{1}{g_c} \times 1(\text{lb}_m) \times 32.1740(\text{ft/s}^2)$$

$$1(\text{poundal}) = \frac{1}{g_c} \times 1(\text{lb}_m) \times 1(\text{ft/s}^2)$$

$$g_c = 1(\text{lb}_m \cdot \text{ft/poundal} \cdot \text{s}^2)$$

From the expression,

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

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

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

$= \frac{\Delta H}{\Delta S}$ (ΔG = 0 at equilibrium)

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

T = 2000 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.

Mathematically,

Change in entropy (ΔS) = $S_{\text{final state (product)}} - S_{\text{initial state (reactions)}}$

For reversible and isothermal process,

Change in entropy $\Delta S = \frac{q_{rev}}{T} = \frac{\Delta H}{T}$

When a system is in equilibrium, $\Delta S = 0$

Second law of thermodynamics: The entropy of universe is continuously increasing due to spontaneous processes taking place in it.

$\Delta S_{\text{system}} + \Delta S_{\text{surrounding}} > 0$ *i.e.*, $\Delta S_{\text{total}} > 0$

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