

Topic 2.4: Free Energy – Solutions



$$\Delta H_{\text{rxn}}^{\circ} = \sum \Delta H_f^{\circ} \text{products} - \sum \Delta H_f^{\circ} \text{reactants}$$

$$\therefore \Delta H_{\text{rxn}}^{\circ} = \left[2\text{mol} (90.2 \text{ kJmol}^{-1}) \right] - \left[(0) \text{ kJ} \right] = 180.4 \text{ kJ} = 180400 \text{ J}$$

$$\Delta S^{\circ} = \sum \Delta S^{\circ} \text{Products} - \sum \Delta S^{\circ} \text{Reactants}$$

$$\therefore \Delta S^{\circ} = \left[2\text{mol} (210.7 \text{ Jmol}^{-1}\text{K}^{-1}) + \right] - \left[1\text{mol} (191.5 \text{ Jmol}^{-1}\text{K}^{-1}) + 1\text{mol} (205.0 \text{ Jmol}^{-1}\text{K}^{-1}) \right] = 24.9 \text{ JK}^{-1}$$

$$\Delta G^{\circ} = \Delta H^{\circ} - T \Delta S^{\circ}$$

$$\therefore \Delta G^{\circ} = 180400 \text{ J} - 298 \text{ K} (24.9) \text{ JK}^{-1} = 172980 \text{ J} = 173 \text{ kJ}$$

As ΔG° is positive, the reaction will not proceed spontaneously at this temperature.



$$\Delta G^{\circ} = \sum \Delta G_f^{\circ} \text{Products} - \sum \Delta G_f^{\circ} \text{Reactants}$$

$$\therefore \Delta G^{\circ} = \left[1\text{mol} (0 \text{ kJmol}^{-1}) + 6\text{mol} (-4.7 \text{ kJmol}^{-1}) + 4\text{mol} (-228.6 \text{ kJmol}^{-1}) \right] \\ - \left[6\text{mol} (-78.9 \text{ kJmol}^{-1}) + 1\text{mol} (-727.8 \text{ kJmol}^{-1}) + 8\text{mol} (0 \text{ kJmol}^{-1}) \right] = 258.6 \text{ kJ}$$



$$\Delta G^{\circ} = \sum \Delta G_f^{\circ} \text{Products} - \sum \Delta G_f^{\circ} \text{Reactants}$$

$$\therefore \Delta G^{\circ} = \left[2(-394.4) + 3(-120.4) \right] \text{ kJmol}^{-1} - \left[(-174.8) + 3(0) \right] \text{ kJmol}^{-1} = -975.2 \text{ kJmol}^{-1}$$

$$\text{mol C}_2\text{H}_5\text{OH} = \frac{1.00 \text{ g}}{46.068 \text{ gmol}^{-1}} = 0.0217 \text{ mol}$$

$$\text{Gibbs free energy per gram} = -975.2 \text{ kJmol}^{-1} \times 0.0217 \text{ mol} = -21.2 \text{ kJ}$$



$$\Delta G^{\circ} = \sum \Delta G_f^{\circ} \text{Products} - \sum \Delta G_f^{\circ} \text{Reactants}$$

$$\therefore \Delta G^{\circ} = \left[6(-394.4) + 6(-120.4) \right] \text{ kJmol}^{-1} - \left[(-917.2) + 6(0) \right] \text{ kJmol}^{-1} = -2171.6 \text{ kJmol}^{-1}$$

$$\text{mol C}_6\text{H}_{12}\text{O}_6 = \frac{1.00 \text{ g}}{180.156 \text{ gmol}^{-1}} = 0.00555 \text{ mol}$$

$$\text{Gibbs free energy per gram} = -2171.6 \text{ kJmol}^{-1} \times 0.00555 \text{ mol} = -12.05 \text{ kJ}$$

1 g of ethanol produces more Gibbs free energy by oxidation than does glucose.

ΔG° has no effect on the rate of the reaction.

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SJ Ball



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