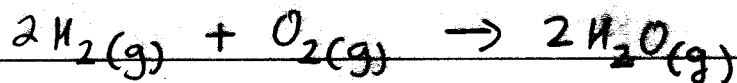


Thermodynamics of the Simplest Combustion Reaction (25°C)



$$\begin{aligned}\Delta_r H^\circ &= 2\Delta_f H^\circ(\text{H}_2\text{O}(\text{g})) - 2\Delta_f H^\circ(\text{H}_2(\text{g})) - \Delta_f H^\circ(\text{O}_2(\text{g})) \\ &= 2(-241.826 \text{ kJ mol}^{-1})\end{aligned}$$

$$\Delta_r H^\circ = -483.652 \text{ kJ mol}^{-1}$$

$$\Delta_r S^\circ = 2S^\circ(\text{H}_2\text{O}(\text{g})) - 2S^\circ(\text{H}_2(\text{g})) - S^\circ(\text{O}_2(\text{g}))$$

$$= [2(188.72) - 2(130.6) - 205.0] \text{ J mol}^{-1} \text{ K}^{-1}$$

$$\Delta_r S^\circ = -88.76 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ = -RT \ln K$$

$$\ln K = -\frac{\Delta_r H^\circ}{RT} + \frac{\Delta_r S^\circ}{R} = \frac{1}{R} \left(-\frac{\Delta_r H^\circ}{T} + \Delta_r S^\circ \right)$$

$$K = \exp \left\{ \left(\frac{\text{mol K}}{8.315 \text{ J}} \right) \left(+ \frac{483652 \text{ J}}{\text{mol}} \left(\frac{1}{298.15 \text{ K}} \right) - \frac{88.76 \text{ J}}{\text{mol K}} \right) \right\}$$

$$K = 1 \times 10^{80}$$