

$$14.36(d) \Delta H^\circ = \left(-200.7 \frac{\text{kJ}}{\text{mol}} \times 2\right) + \left(0 \frac{\text{kJ}}{\text{mol}} \times 3\right) - \left[-393.5 \frac{\text{kJ}}{\text{mol}} \times 2 + 241.8 \frac{\text{kJ}}{\text{mol}}\right]$$

$$= 1352.8 \text{ kJ endothermic}$$

$$\Delta S^\circ = \left(239.7 \frac{\text{J}}{\text{mol K}} \times 2\right) + \left(205.0 \frac{\text{J}}{\text{mol K}} \times 3\right) - \left[\left(213.6 \frac{\text{J}}{\text{mol K}} \times 2\right) + \left(188.7 \frac{\text{J}}{\text{mol K}} \times 4\right)\right]$$

$$= -87.6 \text{ J K}^{-1}$$

With $\Delta H^\circ > 0$ and $\Delta S^\circ < 0$, the process is non-spontaneous at all T.

$$14.40 (c) \Delta G^\circ = \left(-379.5 \frac{\text{kJ}}{\text{mol}}\right) + \left(-394.4 \frac{\text{kJ}}{\text{mol}}\right) - \left(-851.0 \frac{\text{kJ}}{\text{mol}}\right)$$

$$= 77.1 \text{ kJ/mol so nonspontaneous}$$

(b) The equilibrium pressure of CO_2 will be higher if the reaction is endothermic, and therefore the K_{eq} will increase with increasing temp.

$$\Delta H^\circ = \left(-425.6 \frac{\text{kJ}}{\text{mol}}\right) + \left(-393.5 \frac{\text{kJ}}{\text{mol}}\right) - \left(-950.8 \frac{\text{kJ}}{\text{mol}}\right)$$

$$= 131.7 \text{ kJ/mol}$$

$$14.45 \Delta G^\circ = -RT \ln K$$

$$= -(0.0083145 \text{ kJ mol}^{-1} \text{ K}^{-1})(298.15 \text{ K}) \ln(1.8 \times 10^{-1})$$

$$= 27 \text{ kJ/mol}$$

14.44 (a) $\Delta G = 0$ for any system at equilibrium

$$(b) \Delta G^\circ = \left(-371.1 \frac{\text{kJ}}{\text{mol}} \times 2\right) - \left(-300.1 \frac{\text{kJ}}{\text{mol}} \times 2\right) - \left(0 \frac{\text{kJ}}{\text{mol}}\right) = -142.0 \text{ kJ/mol}$$

$$(d) Q = \frac{P_{\text{SO}_2}^2}{P_{\text{SO}_2} P_{\text{O}_2}} \quad \Delta G = \Delta G^\circ + RT \ln Q$$

increasing P_{O_2} leads to a decrease in Q and a decrease in $\ln Q$ thus a decrease in ΔG .

(c) ΔG° does not change.