

using $\Delta G^\circ = -RT \ln K$ & $\Delta G = \Delta G^\circ + RT \ln Q$

The equilibrium constant for the formation of NO from nitrogen and oxygen gas at ~~1500 K~~ is 1.0×10^{-5} . Answer the following questions.

~~25.0°C~~
25.0°C

- a. What is the ΔG° for this reaction?

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

- b. If the initial concentration of nitrogen and oxygen is 0.100 M and the concentration of NO is 0.0100 M what is ΔG at ~~1500 K~~ 25.0°C?

$$[N_2] = 0.1 \quad [O_2] = 0.1$$

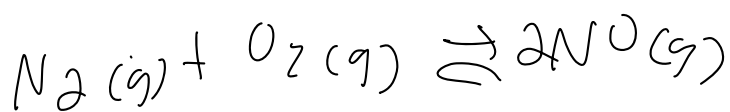
- c. Based on your answer from b, which direction will the reaction shift to establish equilibrium?

a) $\Delta G^\circ = -RT \ln K$

$$\Delta G^\circ = -\left(8.314 \frac{J}{mol \cdot K}\right)(298 K) \ln(1.0 \times 10^{-5})$$

$\Delta G^\circ = 28,524 \text{ J}$
 or
 28.5 kJ
 reactant favored
 @ eq. $\Delta G^\circ > 0$ and $K < 1$

b) $\Delta G = \Delta G^\circ + RT \ln Q$



$$Q = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$

$$Q = \frac{(0.01)^2}{(0.10)(0.10)}$$

$$Q = 0.01$$

$$\Delta G = 28.5 \text{ kJ} + (8.314 \text{ J mol}^{-1} \text{ K}^{-1})(298 \text{ K}) \ln(0.01)$$

$$\Delta G = 28.5 \text{ kJ} + -11.4 \text{ kJ}$$

$$K = 1.0 \times 10^{-5}$$

$$Q = 0.01$$

$$Q > K$$

$$\Delta G = +17.1 \text{ kJ}$$