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 [redacted] is 1.0×10^{-5} . Answer the following questions.

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 [redacted] 25.0° C?

$[N_2] = 0.1$ $[O_2] = 0.1$

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

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

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

$\Delta G^\circ = 28,524\,J$ or $28.5\,kJ$

reactant favored @ eq. $\Delta G^\circ > 0$ and $K << 1$

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

$N_2(g) + O_2(g) \rightleftarrows 2NO(g)$

$Q = \frac{[NO]^2}{[N_2][O_2]}$

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

$Q = 0.01$

$\Delta G = 28.5\,kJ + \left(8.314\,\frac{J}{mol \cdot K}\right)(298\,K) \ln(0.01)$

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

$\Delta G = +17.1\,kJ$

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

$Q = 0.01$

$Q > K$