Let's Practice!

1. At 1000 K, the value of K_p for the reaction $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$ is 0.338. Predict which direction the reaction will proceed toward equilibrium if the initial partial pressures are: P_{SO3} = 0.16 atm, P_{SO2} = 0.41 atm, and P_{O2} = 2.5 atm.

$$Q_p = \frac{(P_{SO_2})^2(P_{O_2})}{(P_{SO_3})^2} = \frac{(0.41)^2(2.5)}{(0.16)^2} = 16$$

$K_p < Q_p$, $0.338 < 16$ — The rxn will shift left to decrease $P_{products}$ and reach equilibrium.

2. For the synthesis of ammonia at 500°C, $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$, the equilibrium constant is 6.0×10^{-2}. Predict the direction in which direction the system will shift to reach equilibrium if $[NH_3]_{initial} = 1.0 \times 10^{-4}$ M, $[N_2]_{initial} = 5.0$ M, and $[H_2]_{initial} = 1.0 \times 10^{-2}$ M at 500°C.

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(1.0E-4)^2}{(5.0)(1.0E-2)^3} = 0.0020 \ (2.0E-3)$$

$K_c > Q_c$, $6.0E-2 > 2.0E-3$ — The rxn will shift right to make more products and achieve equilibrium.

3. The equilibrium partial pressure of Br_2 is 4.00 atm and that of NOBr is 8.00 atm.

$$2\ NO(g) + Br_2(g) \leftrightarrow 2\ NOBr(g) \qquad K_p = 100$$

Using the equation above, determine the equilibrium partial pressure of nitrogen monoxide, NO, at equilibrium.

a. 0.400 atm b. 0.566 atm c. 1.77 atm d. 5.83 atm

$$K_p = \frac{(P_{NOBr})^2}{(P_{NO})^2(P_{Br_2})} = \frac{(8.00)^2}{(P_{NO})^2(4.00)} = 100 \Rightarrow P_{NO} = \sqrt{\frac{(8.00)^2}{(100)(4)}} = \frac{\sqrt{8^2}}{\sqrt{100} \times \sqrt{4}} = \frac{8}{10 \times 2} = \frac{8}{20} = \frac{4}{10}$$

4. The value of the equilibrium constant, K_c, at 25°C is 8.1 for the following reaction: $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$. What must happen for the reaction to reach equilibrium if the initial concentrations of all three species was 2.0 M?

a. The rate of the forward reaction would increase, and $[SO_3]$ would decrease.
b. The rate of the reverse reaction would increase, and $[SO_2]$ would decrease.
c. Both the rate of the forward and reverse reactions would increase, and the value for the equilibrium constant would also increase.
d. No change would occur in either the rate of reaction or the concentrations of any of the species.

$$Q_c = \frac{[SO_2]^2[O_2]}{[SO_3]^2} = \frac{(2)^2(2)}{(2)^2} = 2$$

$K > Q$, $8.1 > 2$

5. For the reaction $2\ X(g) + Y(g) \leftrightarrow 2\ Z(g)$, $K_c = 4.0 \times 10^4$. Determine the value of the equilibrium constant, K_c, for the following reaction:

$$2\ Z(g) \leftrightarrow 2\ X(g) + Y(g) \qquad K_c = ??$$

a. 2.5×10^{-5} b. 2.5×10^{-4} c. 4.0×10^{-5} d. 4.0×10^{-4}

reverse! $K' = \frac{1}{K} = \frac{1}{4E4} = \frac{1}{4}E-4 = 0.25E-4 = 2.5E-5$