Let's Practice!

1. 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}. In which direction will the system shift to reach equilibrium (at 500°C) if $[NH_3]_{initial} = 1.0 \times 10^{-3}$ M, $[N_2]_{initial} = 1.0 \times 10^{-5}$ M, and $[H_2]_{initial} = 2.0 \times 10^{-3}$ M?

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(1.0E\text{-}3)^2}{(1.0E\text{-}5)(2.0E\text{-}3)^3} = 1.3 \times 10^{7} > 6.0 \times 10^{-2}$$

Q K

$K < Q$, so the reaction will need to shift left + make more reactants to reach equilibrium.

2. For the reaction $2\ NO(g) \rightleftharpoons N_2(g) + O_2(g)$, the equilibrium constant $K = 2.4 \times 10^3$ at a certain temperature. The initial concentrations are 0.044 *M* NO, 2.0 *M* N_2, and 0.65 *M* O_2. Is the system at equilibrium? If not, which way will the reaction shift and why?

$$Q = \frac{[N_2][O_2]}{[NO]^2} = \frac{(2.0)(0.65)}{(0.044)^2} = 670 < 2{,}400$$

Q K

$K > Q$, so the system is not at equilibrium, + will need to shift right and make more products to achieve equilibrium.

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 ~~$[SO_2]$~~ $[SO_3]$ 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.

$$Q = \frac{[SO_2]^2[O_2]}{[SO_3]^2} = \frac{(2.0)^2(2.0)}{(2.0)^3} = 1 < 8.1$$

Q K

$K > Q \Rightarrow$ shift right, $\uparrow$ [products]

$[A] = [B] = [C] = 0.6\,mol / 3.1\,L = 0.19$

5. Under equilibrium conditions, 0.60 moles of A, 0.60 moles of B and 0.60 moles of C are present in a 3.1 L solution for the reaction shown below. Determine the value of the equilibrium constant, *K*.

$$2\,A + B \leftrightarrow 2\,C$$

a. $K = 17$ (b.) $K = 5.2$ c. $K = 1.7$ d. $K = 0.52$

$$K = \frac{[C]^2}{[A]^2[B]} = \frac{(0.19)^2}{(0.19)^2(0.19)} = \frac{1}{0.19} = 5.3$$

Use the following information to answer #6-8: A sample of H_2S gas is placed in an evacuated, sealed container and heated until the following decomposition reaction occurs at 1000 K:

$$2\,H_2S(g) \leftrightarrow 2\,H_2(g) + S_2(g) \quad K_c = 86.2$$

6. Which of the following represents the equilibrium constant for this reaction?

(a.) $K_c = \frac{[H_2]^2[S_2]}{[H_2S]^2}$ c. $K_c = \frac{2[H_2][S_2]}{2[H_2S]}$

b. $K_c = \frac{[H_2S]^2}{[H_2]^2[S_2]}$ d. $K_c = \frac{2[H_2S]}{2[H_2][S_2]}$

7. Initially, $[H_2S] = 0.1$ M, $[H_2] = 0.01$ M, and $[S_2] = 0.01$ M. Which way will the reaction shift and why?

(a.) The reaction will shift to make more products because Q < K.

b. The reaction will shift to make more reactants because Q > K. $Q < K$.

c. The reaction will shift to make more products because Q < K. $Q > K$.

d. The reaction will shift to make more reactants because Q > K.

$$Q = \frac{(0.01)^2(0.01)}{(0.1)^2} = 1 \times 10^{-4} < 86.2$$

Q K

$K > Q \Rightarrow$ shift right!

8. If, at a given point in the reaction, the value for the reaction quotient *Q* is determined to be 2.5×10^8, which of the following is occurring?

a. The concentration of the reactant is decreasing while the concentration of the products is increasing.

(b.) The concentration of the reactant is increasing while the concentration of the products is decreasing.

c. The system has passed the equilibrium point, and the concentration of all species involved in the reaction will remain constant.

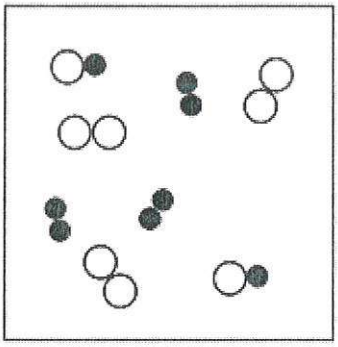
d. The concentrations of all species involved are changing at the same rate.

$Q = \frac{[C]}{[A][B]} = \frac{2}{(3)(3)}$

$= 0.22$

$2.3E-3 < 0.22$

$K \quad Q$



$A_{(g)} = 3$

$B_{(g)} = 3$

$C_{(g)} = 2$

9. The picture above shows the species initially present in a 1.0 L container. The chemical reaction shown below takes place.

$$A + B \leftrightarrow C \qquad K_c = 2.3 \times 10^{-3}$$

Which of the following statements is true?

a. The reaction shifts towards the products to reach equilibrium.

(b.) The reaction shifts towards the reactants to reach equilibrium.

c. The reaction mixture is at equilibrium.

d. The direction of shift cannot be determined from the information given.

10. Consider the following reaction:

$$CH_4(g) + 2\,H_2S(g) \leftrightarrow CS_2(g) + 4\,H_2(g)$$

1.00 M CH_4, 1.00 M CS_2, 2.00 M H_2S and 2.00 M H_2 are mixed in a reaction vessel at 960°C. At this temperature, the reaction will make more methane, CH_4, and more hydrogen sulfide, H_2S. What is a possible K for this reaction?

c. K = 16 c. K = 4

d. K = 8 (d.) K = 1

↳ making more reactants ⇒ $K < Q$

$K < 4$

$Q = \frac{[CS_2][H_2]^4}{[CH_4][H_2S]^2} = \frac{(1)(2)^4}{(1)(2)^2} = 4$