AP Chemistry
Ms. Ye

Name ___________________________________
Date ________________________Block ______

Equilibrium Constant, K

- K tells where the equilibrium lies
- How likely (to what extent) the reaction is to occur

Reaction Quotient, Q

- Snapshot of reaction at some point (where you are right NOW)
- Same calculation as K ($\frac{[products]}{[reactants]}$)
- Given K, compare Q and K to determine which direction the reaction will shift.
- 3 possible situations
  - If **Q = K**, the reaction is____
  - If **Q < K,** the reaction will

    ____
    - (more reactants than should be present at equilibrium)
  - If **Q > K**, the reaction will

    ____
    - (More products than should be present at equilibrium)

Example:

$$\mathbf{2\ NO_{2\,(g)} \leftrightarrow N_2O_{4\,(g)} \qquad K_p = 11\ @\ 100°C}$$

1) 0.200 moles of NO_2 are mixed with 0.200 moles of N_2O_4 in a 4.00 L flask at 100°C. Is the system at equilibrium? If not, determine which direction the reaction must shift to reach equilibrium.

2) At 448°C, the equilibrium constant K_c for the reaction $\mathbf{H_2(\mathit{g}) + I_2(\mathit{g}) \leftrightarrow 2\ HI(\mathit{g})}$ is 50.5. Predict in which direction the reaction proceeds to reach equilibrium if we start with 2.0×10^{-2} mol of HI, 1.0×10^{-2} mol of H_2, and 3.0×10^{-2} mol of I_2 in a 2.00-L container.

Manipulating Equilibrium Constants

1. The equilibrium constant of a reaction in the ***reverse*** direction is the _________________________________ of the equilibrium constant of the reaction in the forward direction.

$A + B \rightleftharpoons C + D$ $\quad$ $K_{eq} = K_1$

$C + D \rightleftharpoons A + B$ $\quad$ $K_{eq} =$ _____

2. The equilibrium constant of a reaction that has been ***multiplied*** by a number is equal to the original equilibrium constant _____________________________ equal to the number

$A + B \rightleftharpoons C + D$ $\quad$ $K_{eq} = K_1$

$nA + nB \rightleftharpoons nC + nD$ $\quad$ $K_{eq} =$ ___

3. The equilibrium constant for a net reaction made up of ***two or more reactions*** is the ______________________ of the equilibrium constants for the individual reactions.

$A + B \rightleftharpoons C + D$ $\quad$ $K_{eq} = K_1$

$C + F \rightleftharpoons G + A$ $\quad$ $K_{eq} = K_2$

$B + F \rightleftharpoons D + G$ $\quad$ $K_{eq} =$ __________

Practice:

1. The equilibrium constant of nitrogen and oxygen reacting to form nitrogen monoxide is $K_c = 1 \times 10^{-30}$ at 25° C. Using this information write the equilibrium expression and calculate K_c for the reverse reaction.

2. If K = 0.145 for $A_2 + 2B \rightleftharpoons 2\,AB$, write the equilibrium expression and calculate K_c for $AB \rightleftharpoons B + 1/2A_2$

3. Determine the overall equilibrium expression and K_c value for the following equilibria:

$2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$ $\quad$ $K_c = 0.014$

$\underline{Br_2(g) + Cl_2(g) \rightleftharpoons 2BrCl(g)}$ $\quad$ $K_c = 7.2$

$2NOBr(g) + Cl_2(g) \rightleftharpoons 2NO + 2BrCl$

Calculating Equilibrium Constants- ICE Tables

Often we do not know the equilibrium concentrations of all species in an equilibrium mixture. If we know the equilibrium concentration of at least one species, however, we can generally use the stoichiometry of the reaction to deduce the equilibrium concentrations of the others. A useful tool is a (R)ICE Table:

- **R**- Reaction
- **I** – initial concentrations/pressures
- **C** – change in concentration/pressure (using molar ratios)
- **E** – equilibrium concentrations/pressures (I - C)

Example:

1. A closed system initially contains 1.00×10^{-3} M H_2 and 2.00×10^{-3} M I_2 at constant temperature. Calculate Kc for the reaction if 1.87×10^{-3} M HI is present at equilibrium

$$H_2 + I_2 \rightleftharpoons 2HI$$

	H_2	I_2	$2HI$
I			
C			
E			

2. Sulfur trioxide decomposes at high temperatures. Initially the vessel is charged at 1000K with $SO_3(g)$ at a partial pressure of 0.500atm. At equilibrium the $SO_3(g)$ partial pressure is 0.200atm. Calculate the value of Kp at 1000K.

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

	$2SO_3(g)$	$2SO_2(g)$	$O_2(g)$
I			
C			
E			

3. Gaseous NOCl decomposes to form the gases NO and Cl_2. At 35°C, the equilibrium constant is 1.6×10^{-5}. In an experiment where 1.0 mol NOCl is placed in a 2.0 L flask, what are the equilibrium concentrations?

$2\ NOCl(g) \leftrightarrows 2\ NO(g) + Cl_2(g)$ $\quad K_c = 1.6 \times 10^{-5}$

	2 NOCl(g)	2 NO(g)	Cl_2 (g)
I			
C			
E			

4. 0.600 moles of NO and 0.750 moles of O_2 are placed in an empty 2.00 L flask. The system is allowed to establish equilibrium. What will be the equilibrium concentration of each species in the flask?

$2\ NO(g) + O_2(g) \leftrightarrows 2\ NO_2(g)$ $\quad K_c = 5.00 \times 10^{-6}$

	2 NO(g)	O_2(g)	2 NO_2(g)
I			
C			
E			

Q, K, and ICE Tables Problem Set

1. At 100°C the equilibrium constant for the reaction $COCl_{2\,(g)} \leftrightarrow CO_{(g)} + Cl_{2\,(g)}$ has the value $K_c = 2.19 \times 10^{-10}$. Are the following mixtures of $COCl_2$, CO, and Cl_2 at 100°C at equilibrium? If not, indicate the direction that the reaction must proceed to achieve equilibrium:
   a. $[COCl_2] = 2.00 \times 10^{-3}$ M, $[CO] = 3.3 \times 10^{-6}$ M, $[Cl_2] = 6.62 \times 10^{-6}$ M

   b. $[COCl_2] = 4.50 \times 10^{-2}$ M, $[CO] = 1.1 \times 10^{-7}$ M, $[Cl_2] = 2.25 \times 10^{-6}$ M

   c. $[COCl_2] = 0.0100$ M, $[CO] = [Cl_2] = 1.48 \times 10^{-6}$ M

2. At a particular temperature, a 2.00-L flask at equilibrium contains 2.80×10^{-4} moles of N_2, 2.50×10^{-5} moles of O_2, and 2.00×10^{-2} moles of N_2O based on: $\mathbf{N_{2\,(g)} + O_{2\,(g)} \leftrightarrow 2\,N_2O_{(g)}}$. In a different trial of the same reaction, $[N_2] = 2.00 \times 10^{-4}$ M, $[N_2O] = 0.200$ M, and $[O_2] = 0.00245$ M, does this represent a system at equilibrium? If not, what directional shift must take place to establish equilibrium?

3. Consider the following equilibrium, for which $K_p = 0.0752$ at 480°C:

$$2\,Cl_{2\,(g)} + 2\,H_2O_{(g)} \rightleftharpoons 4\,HCl_{(g)} + O_{2\,(g)}$$

a. What is the value of K_p for the reaction $4\,HCl_{(g)} + O_{2\,(g)} \rightleftharpoons 2\,Cl_{2\,(g)} + 2\,H_2O_{(g)}$?

b. What is the value of K_p for the reaction $Cl_{2\,(g)} + H_2O_{(g)} \rightleftharpoons 2\,HCl_{(g)} + \frac{1}{2}\,O_{2\,(g)}$?

4. Calculate a value for the equilibrium constant for the reaction $O_{2\,(g)} + O_{(g)} \rightleftharpoons O_{3\,(g)}$ given:

$NO_{2\,(g)} \rightleftharpoons NO_{(g)} + O_{(g)}$ $\quad K = 6.8 \times 10^{-49}$

$O_{3\,(g)} + NO_{(g)} \rightleftharpoons NO_{2\,(g)} + O_{2\,(g)}$ $\quad K = 5.8 \times 10^{-34}$

5. Consider the equilibrium: $N_2\,(g) + O_2\,(g) + Br_2\,(g) \rightleftharpoons 2\,NOBr\,(g)$

Calculate the equilibrium constant Kp for this reaction, given the following information:

$2\,NO\,(g) + Br_2\,(g) \rightleftharpoons 2\,NOBr\,(g)$ $\quad Kc = 2.0$

$2\,NO\,(g) \rightleftharpoons N_2\,(g) + O_2\,(g)$ $\quad Kc = 2.1 \times 10^{30}$

6. Given the following information,

$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$ $\quad K_c = 6.8\times10^{-4}$

$H_2C_2O_4(aq) \rightleftharpoons 2H^+(aq) + C_2O_4^-(aq)$ $\quad K_c = 3.8\times10^{-6}$

determine the K_c for the overall reaction:

$2HF(aq) + C_2O_4^-\,(aq) \rightleftharpoons 2F^-(aq) + H_2C_2O_4(aq)$

7. A mixture of 0.10 mol of NO, 0.050 mol of H_2, and 0.10 mol of H_2O is placed into a 1.0-L vessel at 300 K. At equilibrium, [NO] = 0.062 M. Determine the equilibrium concentrations of H_2, N_2, and H_2O, then determine K_c.

$$\mathbf{2\ NO(g) + 2\ H_2\ (g) \leftrightarrows N_2(g) + 2\ H_2O(g)}$$

8. Cl_2 gas undergoes homolytic cleavage into chlorine atoms at 1100°C. K_p at 1100°C for this process is 1.13×10^{-4}. If a sample with an initial Cl_2 gas pressure of 0.500 atm was allowed to reach equilibrium, what is the total pressure in the flask?

$$\mathbf{Cl_2(g) \leftrightarrows 2\ Cl(g)}$$

9. A flask is charged with 1.500 atm of $N_2O_4(g)$ and 1.00 atm $NO_2(g)$ at 25°C. After equilibrium is reached, the partial pressure of NO_2 is 0.512 atm. **$N_2O_4(g) \leftrightarrows 2\ NO_2(g)$**
(a) What is the equilibrium partial pressure of N_2O_4?
(b) Calculate K_p for this reaction.