

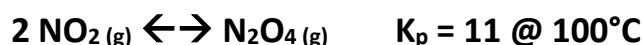
### Equilibrium Constant, K

- 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:



1) 0.200 moles of  $\text{NO}_2$  are mixed with 0.200 moles of  $\text{N}_2\text{O}_4$  in a 4.00 L flask at  $100^\circ\text{C}$ . Is the system at equilibrium? If not, determine which direction the reaction must shift to reach equilibrium.

2) At  $448^\circ\text{C}$ , the equilibrium constant  $K_c$  for the reaction  $\text{H}_2(g) + \text{I}_2(g) \leftrightarrow 2 \text{HI}(g)$  is 50.5. Predict in which direction the reaction proceeds to reach equilibrium if we start with  $2.0 \times 10^{-2}$  mol of HI,  $1.0 \times 10^{-2}$  mol of  $\text{H}_2$ , and  $3.0 \times 10^{-2}$  mol of  $\text{I}_2$  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.



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



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.

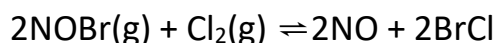
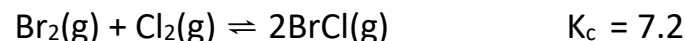
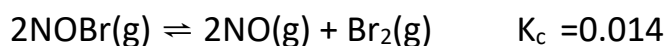


### **Practice:**

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

2. If  $K = 0.145$  for  $A_2 + 2B \rightleftharpoons 2AB$ , 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:



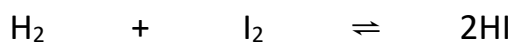
## Calculating Equilibrium Constants- ICE Tables

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)

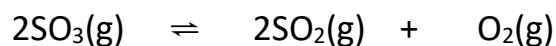
Example:

1. A closed system initially contains  $1.00 \times 10^{-3}$  M  $\text{H}_2$  and  $2.00 \times 10^{-3}$  M  $\text{I}_2$  at constant temperature. Calculate  $K_c$  for the reaction if  $1.87 \times 10^{-3}$  M  $\text{HI}$  is present at equilibrium



I			
C			
E			

2. Sulfur trioxide decomposes at high temperatures. Initially the vessel is charged at 1000K with  $\text{SO}_3(\text{g})$  at a partial pressure of 0.500atm. At equilibrium the  $\text{SO}_3(\text{g})$  partial pressure is 0.200atm. Calculate the value of  $K_p$  at 1000K.



I			
C			
E			

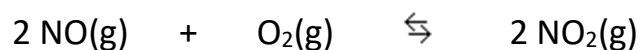
3. Gaseous NOCl decomposes to form the gases NO and Cl<sub>2</sub>. At 35°C, the equilibrium constant is 1.6 x 10<sup>-5</sup>. In an experiment where 1.0 mol NOCl is placed in a 2.0 L flask, what are the equilibrium concentrations?



$$K_c = 1.6 \times 10^{-5}$$

I			
C			
E			

4. 0.600 moles of NO and 0.750 moles of O<sub>2</sub> 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?



$$K_c = 5.00 \times 10^{-6}$$

I			
C			
E			

### Q, K, and ICE Tables Problem Set

1. At 100°C the equilibrium constant for the reaction  $\text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g})$  has the value  $K_c = 2.19 \times 10^{-10}$ . Are the following mixtures of  $\text{COCl}_2$ ,  $\text{CO}$ , and  $\text{Cl}_2$  at 100°C at equilibrium? If not, indicate the direction that the reaction must proceed to achieve equilibrium:

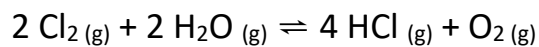
a.  $[\text{COCl}_2] = 2.00 \times 10^{-3} \text{ M}$ ,  $[\text{CO}] = 3.3 \times 10^{-6} \text{ M}$ ,  $[\text{Cl}_2] = 6.62 \times 10^{-6} \text{ M}$

b.  $[\text{COCl}_2] = 4.50 \times 10^{-2} \text{ M}$ ,  $[\text{CO}] = 1.1 \times 10^{-7} \text{ M}$ ,  $[\text{Cl}_2] = 2.25 \times 10^{-6} \text{ M}$

c.  $[\text{COCl}_2] = 0.0100 \text{ M}$ ,  $[\text{CO}] = [\text{Cl}_2] = 1.48 \times 10^{-6} \text{ M}$

2. At a particular temperature, a 2.00-L flask at equilibrium contains  $2.80 \times 10^{-4}$  moles of  $\text{N}_2$ ,  $2.50 \times 10^{-5}$  moles of  $\text{O}_2$ , and  $2.00 \times 10^{-2}$  moles of  $\text{N}_2\text{O}$  based on:  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{N}_2\text{O}(\text{g})$ . In a different trial of the same reaction,  $[\text{N}_2] = 2.00 \times 10^{-4} \text{ M}$ ,  $[\text{N}_2\text{O}] = 0.200 \text{ M}$ , and  $[\text{O}_2] = 0.00245 \text{ M}$ , does this represent a system at equilibrium? If not, what directional shift must take place to establish equilibrium?

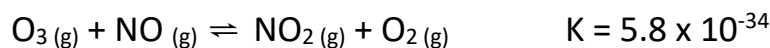
3. Consider the following equilibrium, for which  $K_p = 0.0752$  at  $480^\circ\text{C}$ :



a. What is the value of  $K_p$  for the reaction  $4 \text{HCl}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{Cl}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$ ?

b. What is the value of  $K_p$  for the reaction  $\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons 2 \text{HCl}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g})$ ?

4. Calculate a value for the equilibrium constant for the reaction  $\text{O}_2(\text{g}) + \text{O}(\text{g}) \rightleftharpoons \text{O}_3(\text{g})$  given:

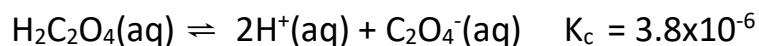
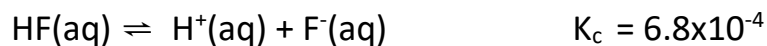


5. Consider the equilibrium:  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2 \text{NOBr}(\text{g})$

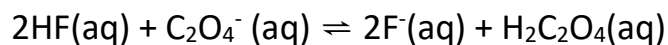
Calculate the equilibrium constant  $K_p$  for this reaction, given the following information:



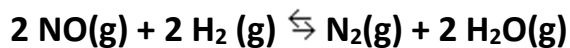
6. Given the following information,



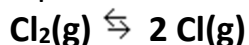
determine the  $K_c$  for the overall reaction:



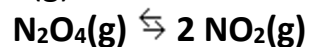
7. A mixture of 0.10 mol of NO, 0.050 mol of H<sub>2</sub>, and 0.10 mol of H<sub>2</sub>O is placed into a 1.0-L vessel at 300 K. At equilibrium, [NO] = 0.062 M. Determine the equilibrium concentrations of H<sub>2</sub>, N<sub>2</sub>, and H<sub>2</sub>O, then determine K<sub>c</sub>.



8. Cl<sub>2</sub> gas undergoes homolytic cleavage into chlorine atoms at 1100°C. K<sub>p</sub> at 1100°C for this process is 1.13 × 10<sup>-4</sup>. If a sample with an initial Cl<sub>2</sub> gas pressure of 0.500 atm was allowed to reach equilibrium, what is the total pressure in the flask?



9. A flask is charged with 1.500 atm of N<sub>2</sub>O<sub>4</sub>(g) and 1.00 atm NO<sub>2</sub>(g) at 25°C. After equilibrium is reached, the partial pressure of NO<sub>2</sub> is 0.512 atm.



- (a) What is the equilibrium partial pressure of N<sub>2</sub>O<sub>4</sub>?  
(b) Calculate K<sub>p</sub> for this reaction.