

# N13 - Equilibrium

**Target:** I can calculate  $K$ , convert  $K$  when changes to a reaction are made, and can use the size of  $K$  to determine product or reactant favored.

# **N13 - Equilibrium**

**K – Equilibrium Constant  
(Capitalized!)**

# Equilibrium Constant

Even though the concentrations of reactants and products are not equal at equilibrium, there is a relationship between them.

**Law of Mass Action** or also **Equilibrium Expression**

The relationship between the chemical equation and the concentrations of reactants and products.

# Equilibrium Constant

For the general equation  $aA + bB \rightarrow cC + dD$

**The law of mass action** gives the relationship below.

- The lowercase letters represent the coefficients of the balanced chemical equation.
- Always products over reactants

**$K$**  is called the **equilibrium constant**.

- Unitless

**Law of Mass Action**

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Products

Reactants

# Writing Equilibrium Constant Expressions

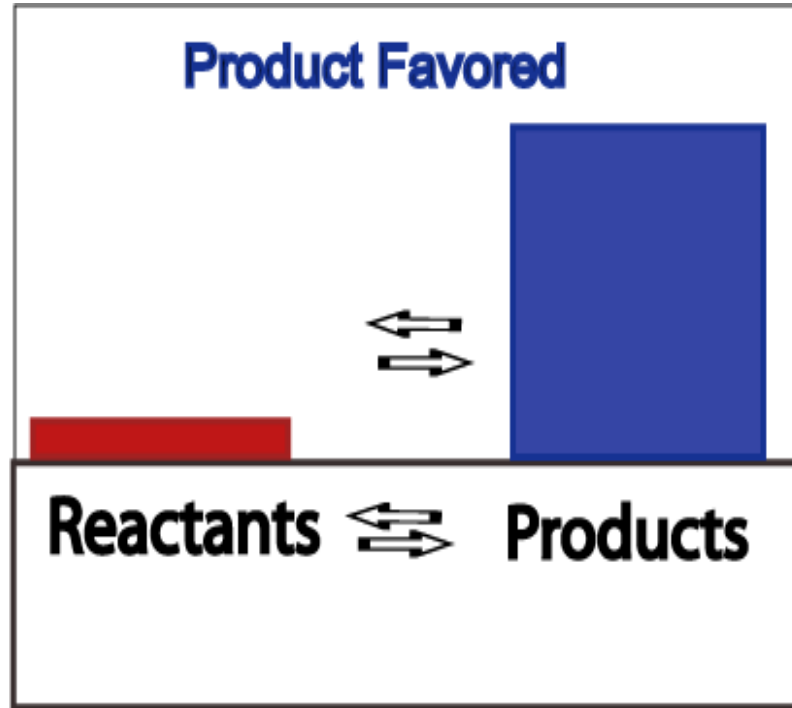


The equilibrium constant expression is:

$$K = \frac{[\text{NO}_2]^4 [\text{O}_2]}{[\text{N}_2\text{O}_5]^2}$$

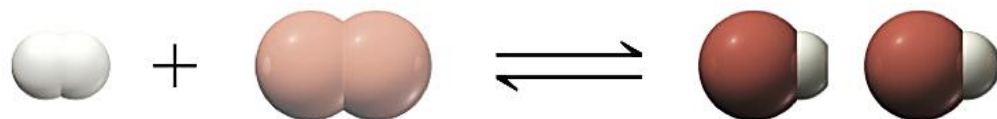
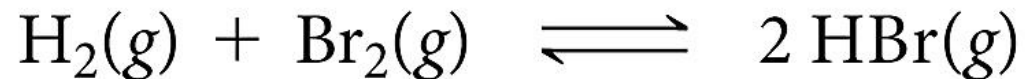
# Product Favored Equilibrium

Large values for  $K$  signify the reaction is **product favored**



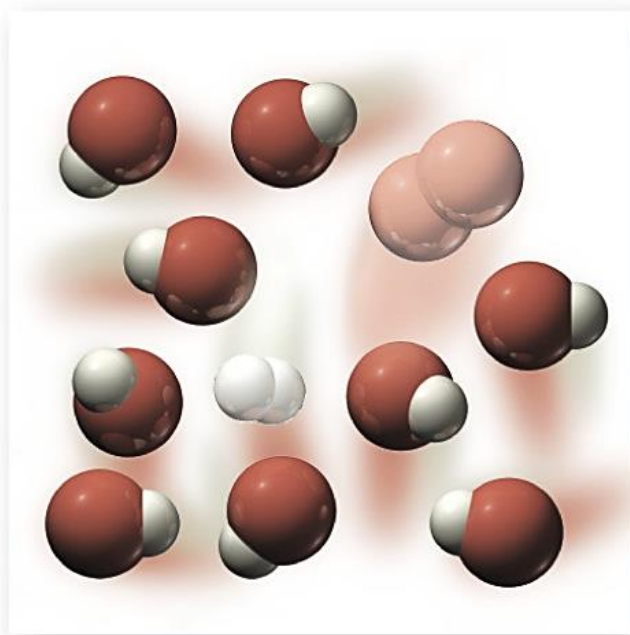
When equilibrium is achieved, most reactant  
has been converted to product

# A Large Equilibrium Constant



**PRODUCT  
FAVORED**

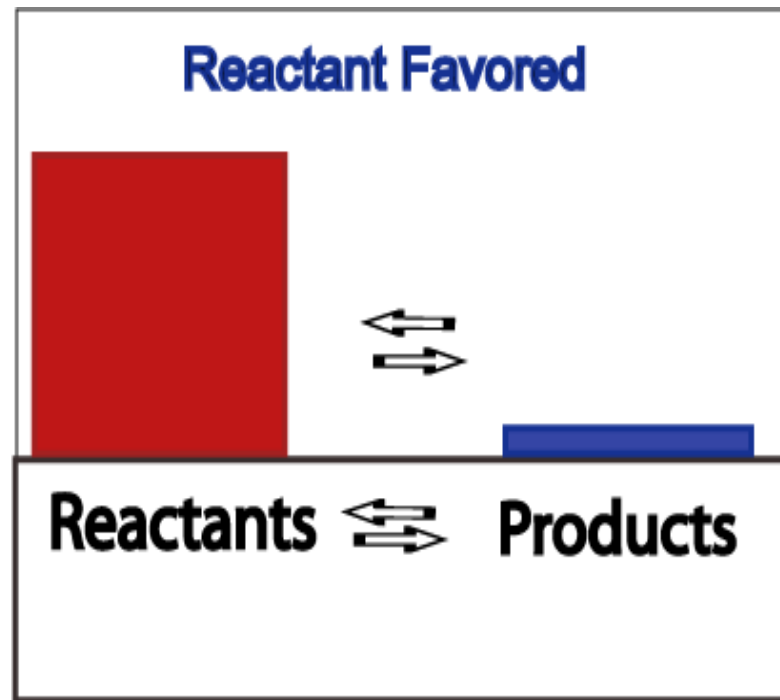
**Lots of products =  
Large numerator =  
Large K value**



$$K = \frac{[\text{HBr}]^2}{[\text{H}_2][\text{Br}_2]} = \text{large number}$$

# Reactant Favored Equilibrium

Small values for  $K$  signify the reaction is **reactant favored**



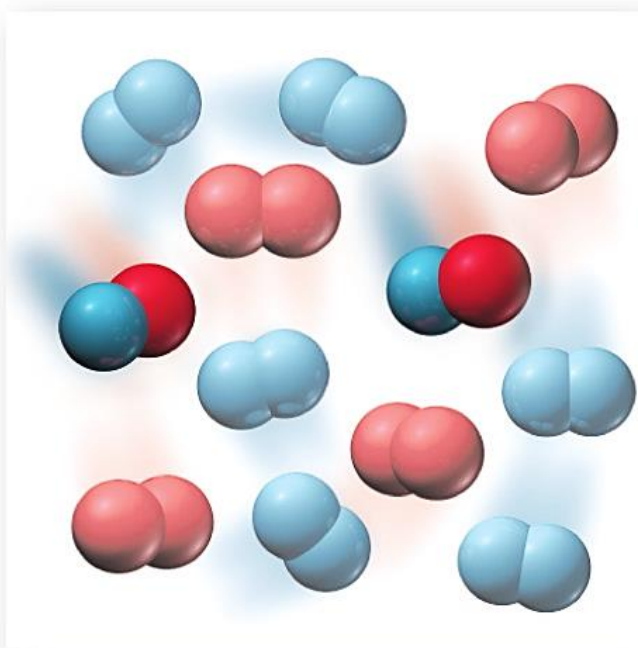
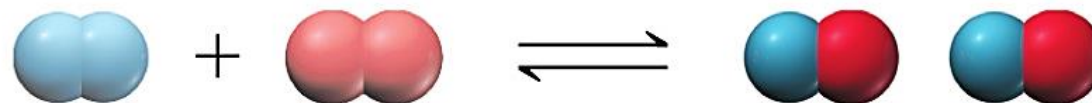
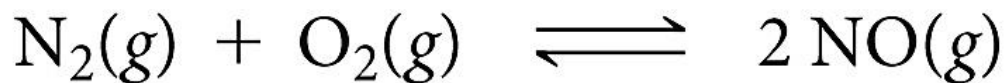
When equilibrium is achieved, very little reactant  
has been converted to product



# A Small Equilibrium Constant

Reactant FAVORED

Lots of reactants =  
Large denominator =  
Small K value



$$K = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = \text{small number}$$

# Relationships between $K$ and Chemical Equations

- When the reaction is written backward, the equilibrium constant is inverted.



the equilibrium constant expression is:

$$K_{\text{forward}} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$



the equilibrium constant expression is:

$$K_{\text{backward}} = \frac{[A]^a \times [B]^b}{[C]^c \times [D]^d}$$

$$K_{\text{backwards}} = \frac{1}{K_{\text{forward}}}$$

# Relationships between $K$ and Chemical Equations

- When the coefficients of an equation are multiplied by a factor, the equilibrium constant is raised to that factor.



the equilibrium constant expression is:

$$K_{\text{original}} = \frac{[C]^c}{[A]^a \times [B]^b}$$

$$K_{\text{new}} = K_{\text{original}}^n$$



the equilibrium constant expression is:

$$\begin{aligned} K_{\text{new}} &= \frac{[C]^{2c}}{[A]^{2a} \times [B]^{2b}} \\ &= \left( \frac{[C]^c}{[A]^a \times [B]^b} \right)^2 \end{aligned}$$

# Relationships between $K$ and Chemical Equations

- When you add equations to get a new equation, the equilibrium constant of the new equation is the product of the equilibrium constants of the old equations.



the equilibrium constant expressions are:

$$K_1 = \frac{[B]^b}{[A]^a} \quad K_2 = \frac{[C]^c}{[B]^b}$$

$$K_{new} = K_1 \times K_2$$



the equilibrium constant is:

$$K_{new} = \frac{[C]^c}{[A]^a} \\ = \frac{\cancel{[B]^b}}{[A]^a} \times \frac{[C]^c}{\cancel{[B]^b}}$$

# Equilibrium Constants for Rxns Involving Gases

- The [ ]s of a gas in a mixture is proportional to its partial pressure.
- Therefore, K can be expressed as the ratio of the partial pressures of the gases.



the equilibrium constant expressions are:

$$K_c = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

# K<sub>c</sub> and K<sub>p</sub>

- $K_p$ , the partial pressures are always in atm.
- $K_p$  and  $K_c$  are not necessarily the same because of the difference in units.

$$K_p = K_c \text{ when } \Delta n = 0$$

$$K_p = K_c \times (RT)^{\Delta n}$$

$\Delta n$  is the difference between the number of moles of reactants and moles of products.

# Deriving the Relationship Between Kc and Kp

$$[A] = \frac{n_A}{V}, n_A = \text{moles of A}, V = \text{volume of gas}$$

$$P_A V = n_A RT, \text{ from the Ideal Gas Law}$$

$$\text{substituting } P_A = \frac{n_A}{V} RT = [A] RT$$

$$[A] = \frac{P_A}{RT}$$

# Deriving the Relationship Between Kc and Kp

$$[X] = \frac{P_X}{RT} \quad aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g) \quad K_c = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

substituting

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

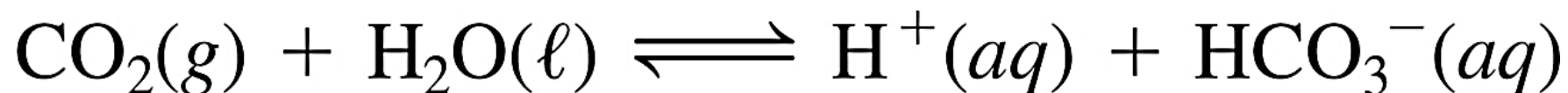
$$K_c = \frac{\left(\frac{P_C}{RT}\right)^c \times \left(\frac{P_D}{RT}\right)^d}{\left(\frac{P_A}{RT}\right)^a \times \left(\frac{P_B}{RT}\right)^b} = \frac{P_C^c P_D^d \left(\frac{1}{RT}\right)^{c+d}}{P_A^a P_B^b \left(\frac{1}{RT}\right)^{a+b}} = K_p \left(\frac{1}{RT}\right)^{(c+d)-(a+b)}$$

rearranging  $K_p = K_c (RT)^{(c+d)-(a+b)} = K_c (RT)^{\Delta n}$



# Heterogeneous Equilibria

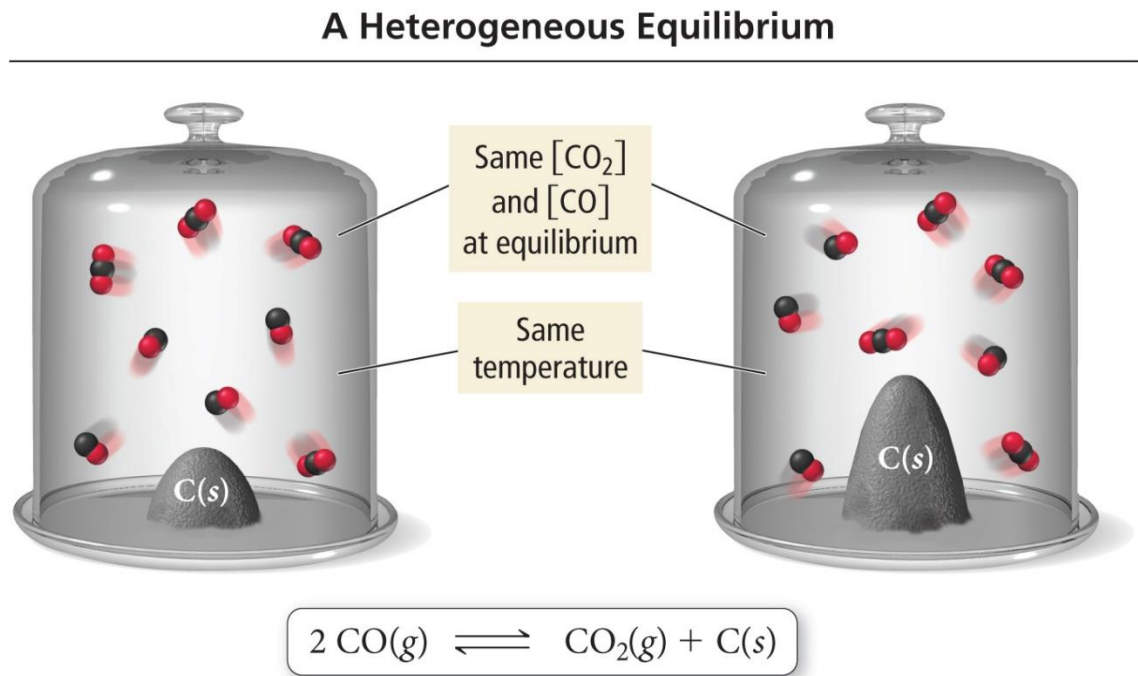
- [ ]s of pure solids and pure liquids do not change during the course of a reaction.
- Because their [ ]s don't change, solids and liquids are not included in the equilibrium constant expression.



**the equilibrium constant expression is as follows:**

$$K_c = \frac{[\text{H}^+][\text{HCO}_3^-]}{[\text{CO}_2]}$$

# Heterogeneous Equilibria



The amount of C is different, but the concentrations of CO and  $\text{CO}_2$  remain the same. Therefore, the amount of C has no effect on the position of equilibrium.

# **YouTube Link to Presentation**

<https://youtu.be/JbbaquQpHsw>