Equilibrium: Let's Get Balanced!

Dynamic Equilibrium

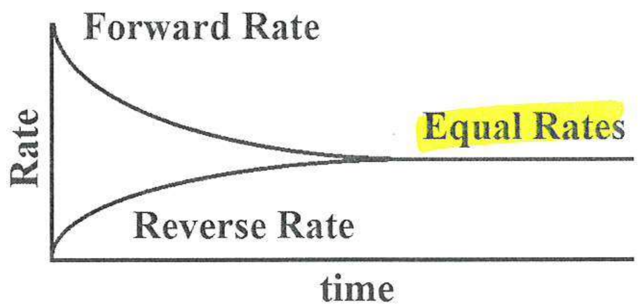
1. Many chemical reactions are at equilibrium, which is indicated by double arrows (⇋).
2. At equilibrium, both reactants and products are present.

Forward and backward reactions **rates** are equal.

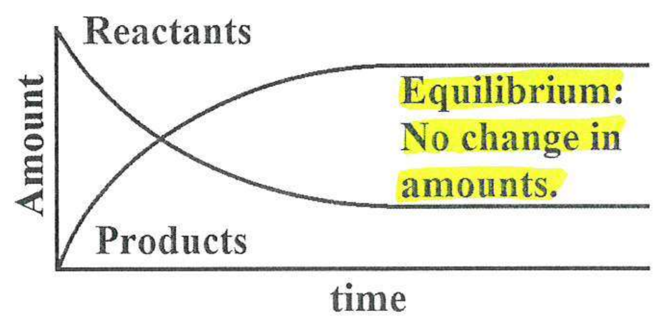
This means that the forward and backward reactions never stop!

- But… concentrations of reactants and products are not (usually) equal.
- Instead, the concentrations of the reactants and products are constant.

Rate of Reaction vs Time



[Reactants], [Products] vs Time



The Equilibrium Constant, K (the Law of Mass Action): relates the concentrations of reactants and products **at equilibrium** at a given temperature.

For the general reaction: $aA + bB \leftrightharpoons cC + dD$

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

a) Each concentration is raised to the power of its stoichiometric coefficient in the balanced equation.

b) **Note:** Pure solids (s) and pure liquids (l) (e.g. water) are NOT placed into the expression.

c) There are NO units for equilibrium constant K (they cancel out).

Let's Practice! Write the K_c expressions for the following reactions:

1. $C_6H_{12}(g) + 9\,O_2(g) \rightleftharpoons 6\,CO_2(g) + 6\,\cancel{H_2O(l)}$

$$K_c = \frac{[CO_2]^6}{[C_6H_{12}][O_2]^9}$$

2. $\cancel{Ni(s)} + 4\,CO(g) \rightleftharpoons Ni(CO)_4(g)$

$$K_c = \frac{[Ni(CO)_4]}{[CO]^4}$$

3. For $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$, at equilibrium $[N_2] = 1.50$ M, $[H_2] = 2.00$ M, and $[NH_3] = 0.010$ M.

a. Write the equilibrium constant expression.

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

b. Solve for K.

$$K = \frac{(0.010)^2}{(1.50)(2.00)} = \boxed{8.3 \times 10^{-6}}$$

c. Is the forward or reverse reaction favored, and how do you know?] Answer after next notes, oops! ☺

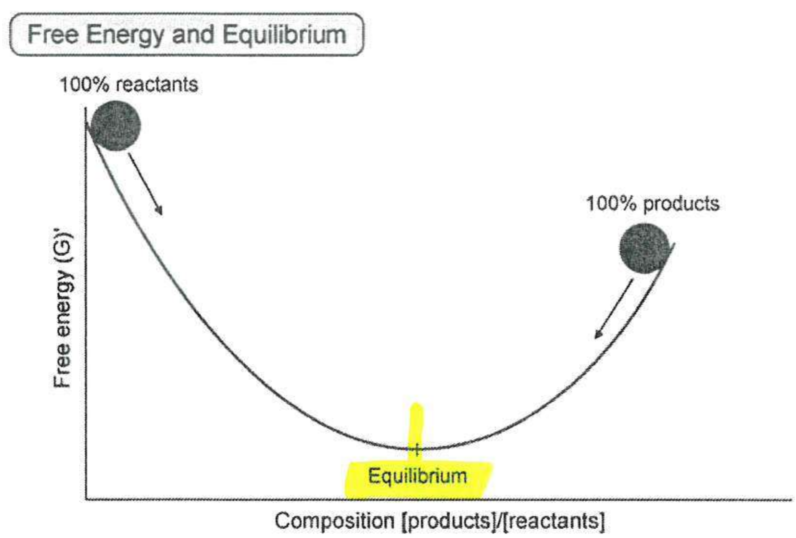
Reverse rxn is favored, b/c K=8.3E-6, which is much less than 1, so there are more reactants than products at equilibrium.

Important Information about K, the equilibrium constant

→ Regardless of initial conditions, <u>at a given temperature</u> a reaction will reach equilibrium with the <u>same</u> ratio of products to reactants

→ **<u>Equilibrium is temperature dependent</u>!!!**

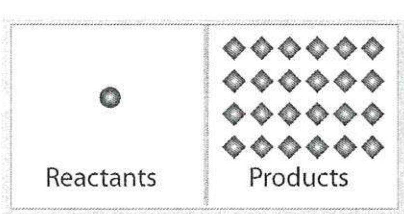
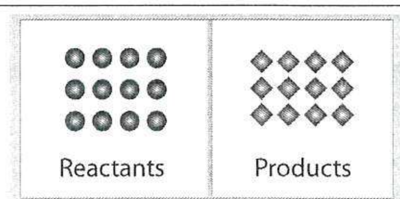
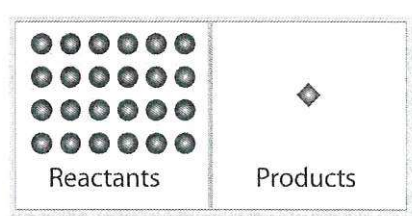
Change the <u>temperature</u>, change the <u>ratio</u> of products to reactants (K).



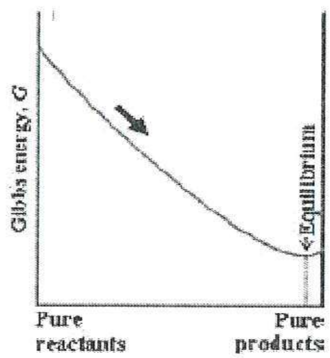
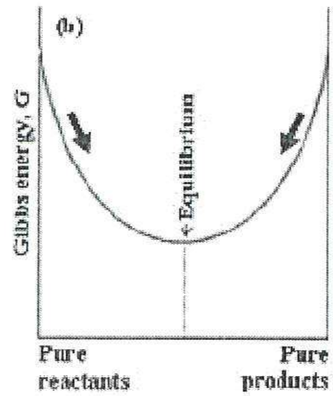
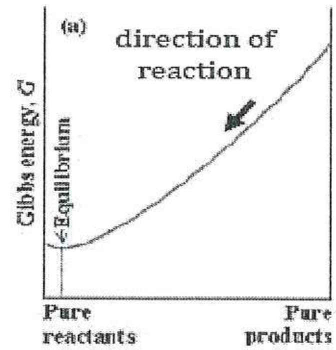
What does *K* mean?

The equilibrium constant, K, tells you:

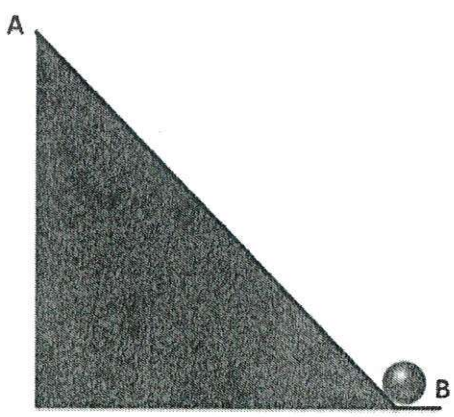
- the ratio of products to reactants when a given reaction reaches its <u>lowest</u> free energy state and "stops".

Large *K* (*K* >> 1)	**Intermediate *K*** (*K* ≈ 1)	**Small *K*** (*K* << 1)
Product-favored	Neither	Reactant-Favored
 Mostly products at equilibrium	 Significant amounts of reactants and products at equilibrium	 Mostly reactants at equilibrium

What does *K* mean?

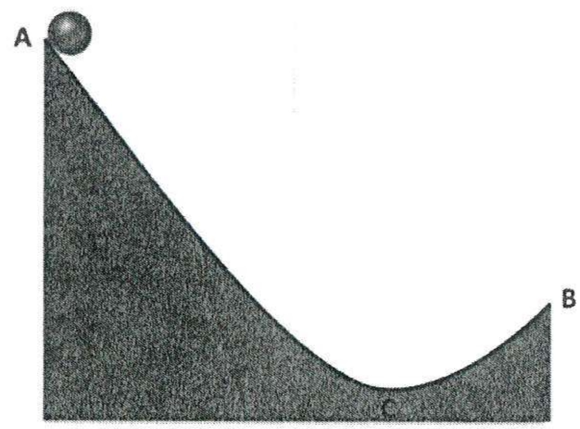
Large *K* ($K >> 1$)	Intermediate *K* ($K \approx 1$)	Small *K* ($K << 1$)
Product-favored	Neither	Reactant-Favored
Forward reaction is thermodynamically favorable 		Reverse reaction is thermodynamically favorable 

Completion Reactions vs Equilibrium Reactions



Complete Reaction: Complete consumption of reactants.

$$A \rightarrow B$$



Equilibrium: Reaction achieves equilibrium by reaching the lowest state of energy.

$$A \rightleftharpoons B$$

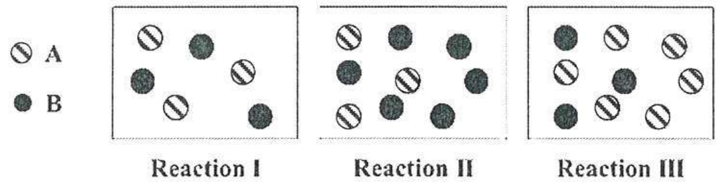
Completion reactions:	Equilibrium reactions:
• Non-reversible • 100% of reactants convert to products • Example: combustion	• Reversible • Reaction will occur until lowest energy state is reached • Example: weak acid dissolution

You can think of K as measuring how close a reaction will go towards 100% completion:

- High K = reaction came pretty close to completely turning reactants into products
- Low K = reaction did NOT come close to turning all reactants into products, mostly reactants just hung around

Let's Practice!

1. The following boxes represent reactions of A ⇋ B at equilibrium.



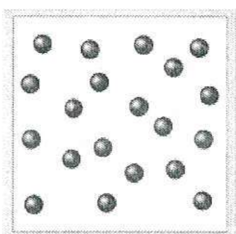
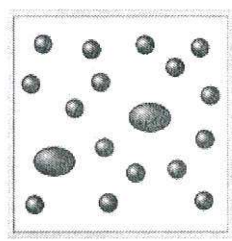
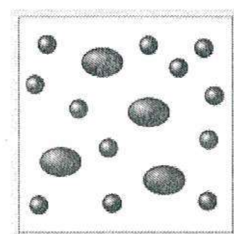
For which reaction shown above is K smallest?

a. Reaction I
b. Reaction II
c. Reaction III
d. Cannot be determined.

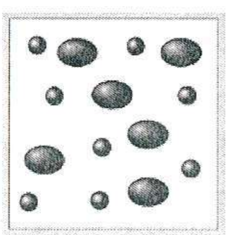
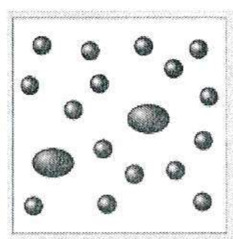
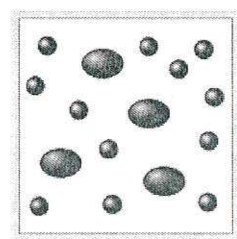
Consider the following reversible reaction to answer #2–3.

$$A \leftrightarrow B$$

2. Which of the following equilibrium systems has the largest value of K?

a.  b.  c. 

3. Which of the following equilibrium systems has the smallest value of K?

b.  b.  c. 

The Reaction Quotient, Q: The ratio of products to reactants at <u>current conditions</u> (querrent quonditions ;)

Calculate Q if: you need to know the answer to the question, "Is the system at equilibrium?"

A: The answer can be yes or no!

For the general reaction: aA + bB ⇋ cC + dD

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

- Q has the appearance of K but the concentrations are not *necessarily* at equilibrium.
- K is constant (at constant temperature), but Q can change!