Alpha Order

K -> Q

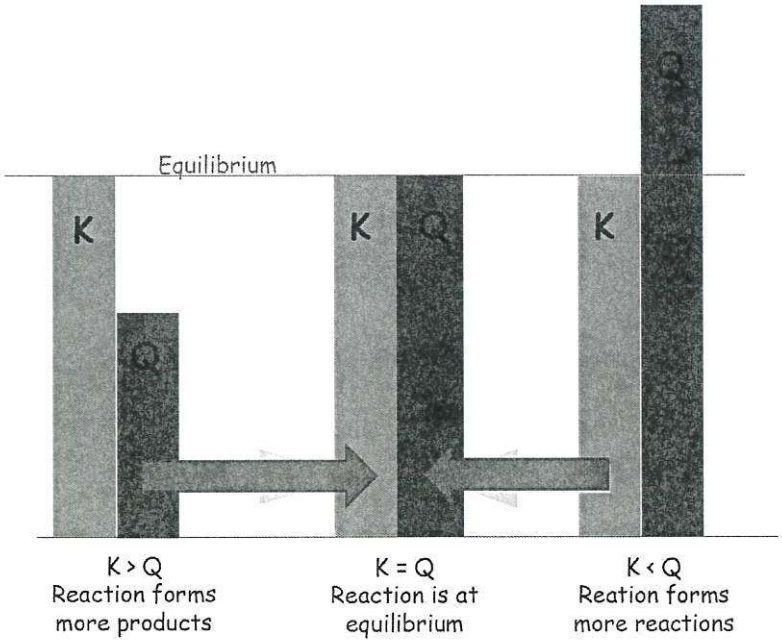
K <- Q

Compare K and Q to see if you're at equilibrium AND how to get there!

1. If **K** ___>___ **Q**, the system is not at equilibrium: forward reaction is favored (shift right, or ___→___) to make Q = K.
2. If **K** ___≈___ **Q**, the system is at equilibrium.
3. If **K** ___<___ **Q**, the system is not at equilibrium: reverse reaction is favored (shift left, or ___←___) to make Q = K.

In Summary

Current conditions	$K > Q$	$K \approx Q$	$K < Q$
change needed for system to reach equilibrium	shift right (make more products)	already at equilibrium	shift left (make more reactants
reaction rates	forward > reverse reaction rate (until equilibrium reached)	forward ≈ reverse reaction rate	forward < reverse reaction rate (until equilibrium reached)



Notes about Language: Talking about equilibrium can be tricky! Here's a quick guide to the terminology used.

Phrases used to describe directions of reaction shift	
If the reaction needs more <u>products</u> to reach equilibrium:	**If the reaction needs more <u>reactants</u> to reach equilibrium:**
1. The reaction will shift right.	1. The reaction will shift left.
2. The forward reaction is occurring more rapidly than the reverse reaction.	2. The reverse reaction is occurring more rapidly than the forward reaction.
3. The reaction will shift to form more products.	3. The reaction will shift to form more reactants.
4. The reaction will proceed to the right.	4. The reaction will proceed to the left.

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]