What Does the Equilibrium Constant Tell Us?

A large value for K_{eq} > 1

Tells us the reaction is spontaneous. (product favored)

A small value for K_{eq} < 1

Tells us the reaction is non-spontaneous. (reactant favored)

The NO_2 - N_2O_4 system at 25 ° C

Init: 0.670 *M* 0.000 *M*

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

equil: 0.643 *M* 0.0547 *M*

$$\frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.0547)^2}{0.643} = 4.65 \times 10^{-3}$$

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$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

equil: 0.643 *M* 0.0547 *M*

the reaction is non-spontaneous. (reactant favored)

$K_{eq} = 4.65 \times 10^{-3}$

Predicting the Direction of a Reaction

The reaction quotient (*Q*) is a useful tool to analyze what must happen in order for a system to reach equilibrium

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$$Q_c = \frac{[\text{products}]^{x}}{[\text{reactants}]^{y}}$$

$Q_c = K_c$ reaction is at equilibrium

$Q_c > K_c$ products revert to reactants to achieve equilibrium

$Q_c < K_c$ reactants form products to achieve equilibrium

Practice Exercise

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$$

(a) How will system shift when

$[N_2] = 0.0711$ M

$[H_2] = 9.17 \times 10^{-3}$ M

$[NH_3] = 1.83 \times 10^{-4}$ M

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$$

$$K_c = 1.2$$

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

$$Q = \frac{(1.83 \times 10^{-4})^2}{(0.0711)\;(9.17 \times 10^{-3})^3}$$

$$= 0.611$$

if Q is smaller than K the concentration of the products increases

Practice Exercise

$$H_2(g) + I_2(g) \underset{}{\overset{K = 54.3}{\rightleftharpoons}} 2HI(g)$$

Init: 0.00623 *M* 0.00414*M* 0.0224 *M*

How will the reaction shift ?

Practice Exercise

$$H_2(g) + I_2(g) \underset{}{\overset{K = 54.3}{\rightleftharpoons}} 2HI(g)$$

Init: 0.00623 *M* 0.00414*M* 0.0224 *M*

$$Q = \frac{[HI]^2}{[H_2][I_2]}$$

$$Q = \frac{(0.0224)^2}{(0.00623)(0.00414)}$$

$= 19.45$

if Q is smaller than K the concentration of the products increases

Calculating Equilibrium Concentrations

Practice Exercise

Calculate the concentrations at equilibrium.

$K_c = 54.3$ at 430^oC

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

	$H_2(g)$	$I_2(g)$	$2HI(g)$
Init:	$0.50\ M$	$0.50\ M$	$0.00\ M$
final:	$-\ x$	$-\ x$	$+2x$
	$0.50 - x$	$0.50 - x$	$2x$

$$\frac{[HI]^2}{[H_2]\ [I_2]} = \frac{(2x)^2}{(0.50 - x)\ (0.50 - x)} = 54.3$$

$$\frac{[HI]^2}{[H_2]\;[I_2]} = \frac{(2x)^2}{(0.50 - x)\;(0.50 - x)} = 54.3$$

Taking the square root of both sides

$$\frac{2x}{0.50 - x} = 7.37$$

$$\boxed{x = 0.393}$$

At equilibrium, the concentration are

$[H_2] = (\,0.50 - 0.393\,)\,M = 0.107\ M$

$[I_2] = (\,0.50 - 0.393\,)\,M = 0.107\ M$

$[HI] = (\,2 \times 0.393\,)\,M = 0.786\ M$

Calculating equilibrium concentration can be difficult.

So we will focus on the easiest situation to solve for.

The solubility of insoluble ionic compounds.

K_{sp}