

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 $\text{NO}_2 - \text{N}_2\text{O}_4$ system at 25°C

Init: **0.670 M** **0.000 M**



equil: **0.643 M** **0.0547 M**

$$\frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{(0.0547)^2}{0.643} = 4.65 \times 10^{-3}$$

The $\text{NO}_2 - \text{N}_2\text{O}_4$ system at 25°C

Init: **0.670 M** **0.000 M**



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

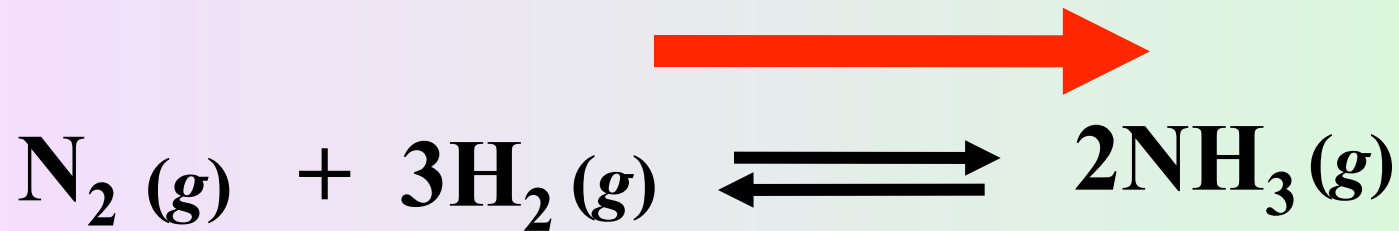


(a) How will system shift when

$$[\text{N}_2] = 0.0711 \text{ M}$$

$$[\text{H}_2] = 9.17 \times 10^{-3} \text{ M}$$

$$[\text{NH}_3] = 1.83 \times 10^{-4} \text{ M}$$



$$K_c = 1.2$$

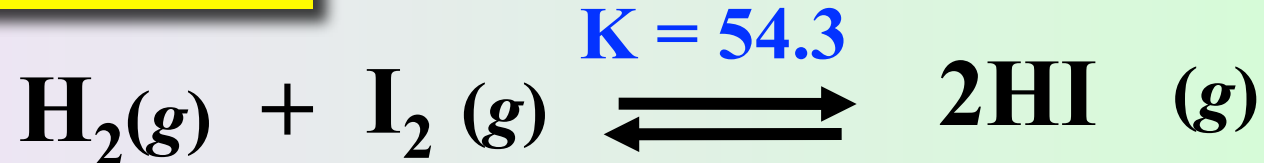
$$Q = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3}$$

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

$$= 0.611$$

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

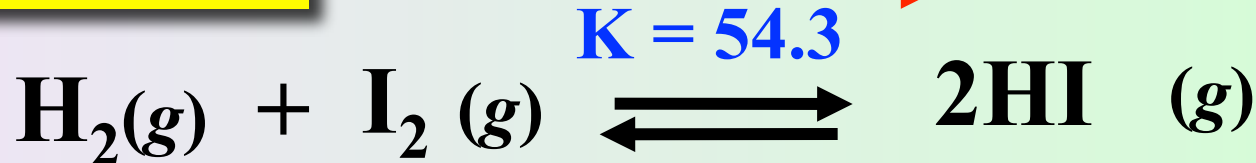
Practice Exercise



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

How will the reaction shift ?

Practice Exercise



Init: 0.00623 M 0.00414M 0.0224 M

$$Q = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

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

$$= 19.45$$

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

Calculating Equilibrium Concentrations

Practice Exercise

Calculate the concentrations at equilibrium.

$$K_c = 54.3 \text{ at } 430^\circ\text{C}$$



Init: **0.50 M** **0.50 M** **0.00 M**

final: **- x** **- x** **+2x**

0.50 - x **0.50 - x** **2x**

$$\frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(2x)^2}{(0.50 - x)(0.50 - x)} = 54.3$$

$$\frac{[\text{HI}]^2}{[\text{H}_2] [\text{I}_2]} = \frac{(2x)^2}{(0.50 - x)(0.50 - x)} = 54.3$$

Taking the square root of both sides

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

$$x = 0.393$$

At equilibrium, the concentration are

$$[\text{H}_2] = (0.50 - 0.393) \text{ M} = 0.107 \text{ M}$$

$$[\text{I}_2] = (0.50 - 0.393) \text{ M} = 0.107 \text{ M}$$

$$[\text{HI}] = (2 \times 0.393) \text{ M} = 0.786 \text{ M}$$

Calculating equilibrium concentration can be difficult.

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

The solubility of insoluble ionic compounds.



