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

	N_2O_4 (*g*)	$2NO_2$ (*g*)
Init:	**0.670 *M***	**0.000 *M***
equil:	**0.643 *M***	**0.0547 *M***

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

The NO_2 - N_2O_4 system at 25 ° C

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

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

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

$Q_c = K_c$ 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

Calculating Equilibrium Concentrations

Practice Exercise

Calculate the concentrations at equilibrium.

$K_c = 54.3$ at 430°C

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

	$H_2(g)$	$I_2(g)$	$2HI(g)$
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

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

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

Taking the square root of both sides

$$2x = 3.685 - 7.37x$$

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

Taking the square root of both sides

$$9.37x = 3.685$$

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

Taking the square root of both sides

$$x = 0.393$$

At equilibrium, the concentration are

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

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

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

Practice Exercise

What is the [H^+] of a 1.00M acetic acid solution. K = **1.8×10^{-5}**

$$HOAc\,(aq) \rightleftharpoons H^+(aq) + AcO^-(aq)$$

	HOAc	H^+	AcO^-
Init:	1.00 *M*	0.00*M*	0.00 *M*
final:	1 - x	+ x	+x

$$K_a = \frac{[H^+]\ \ [A^-]}{[HOAc]}$$

$$\frac{(x)^2}{(1.00 - x)} = 1.8 \times 10^{-5}$$

assume x is small compared to 1.00M and solve

If x is 5% or less than the original quantity our assumption is valid

$$\frac{(x)^2}{(1.00 - \cancel{x})} = 1.8 \times 10^{-5}$$

Small K

Approximation $1.00 - x \cong 1.00$

$$\frac{x^2}{1.00} = 1.8 \times 10^{-5}$$

$$100\% \times \frac{0.0042}{1.00} = .42\%$$

assumption is valid

$[H^+] = x = 0.0042\ M$

Practice Exercise

$[H_2] = 1.000\ M$

$[F_2] = 2.000\ M$

$$H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$$

$K_c = 115$ at $430^{\circ}C$

Calculate the equilibrium concentrations of all species present.

Practice Exercise

$$H_2(g) + F_2(g) \rightleftharpoons 2HF\ (g)$$

	$H_2(g)$	$F_2(g)$	$2HF\ (g)$
Init:	1.000 *M*	2.000 *M*	0.000 *M*
final:	-x	-x	+ 2x

$$115 = \frac{[HF]^2}{[H_2]\,[F_2]}$$

$$115 = \frac{(2x)^2}{(1.000-x)(2.000-x)}$$

assume x is small compared to 1.000 and 2.000 and solve

If x is 5% or less than the original quantity our assumption is valid

Practice Exercise

$$H_2(g) + F_2(g) \rightleftharpoons 2HF\ (g)$$

	$H_2(g)$	$F_2(g)$	$2HF\ (g)$
Init:	1.000 *M*	2.000 *M*	0.000 *M*
final:	-x	-x	+ 2x

$$115 = \frac{[HF]^2}{[H_2]\,[F_2]}$$

$$115 = \frac{(2x)^2}{(1.000-x)(2.000-x)}$$

$$115 = \frac{(2x)^2}{(1.000)\,(2.000)}$$

$$x = 7.58$$

$$\frac{7.58}{2.000} \times 100 = 379\%$$

Practice Exercise

$$H_2(g) + F_2(g) \rightleftharpoons 2HF\ (g)$$

Init:	1.000 *M*	2.000 *M*	0.000 *M*
final:	-x	-x	+ 2x

$$115 = \frac{(2x)^2}{(1.000-x)(2.000-x)}$$

$$115\,(1.000-x)\,(2.000-x) = 4x^2$$

$$115(x^2) - 3.000(115)x + 2.000(115) = 4x^2$$

$$111x^2 - 345x + 230 = 0$$

Practice Exercise

Quadratic equation

$$ax^2 + bx + c = 0$$

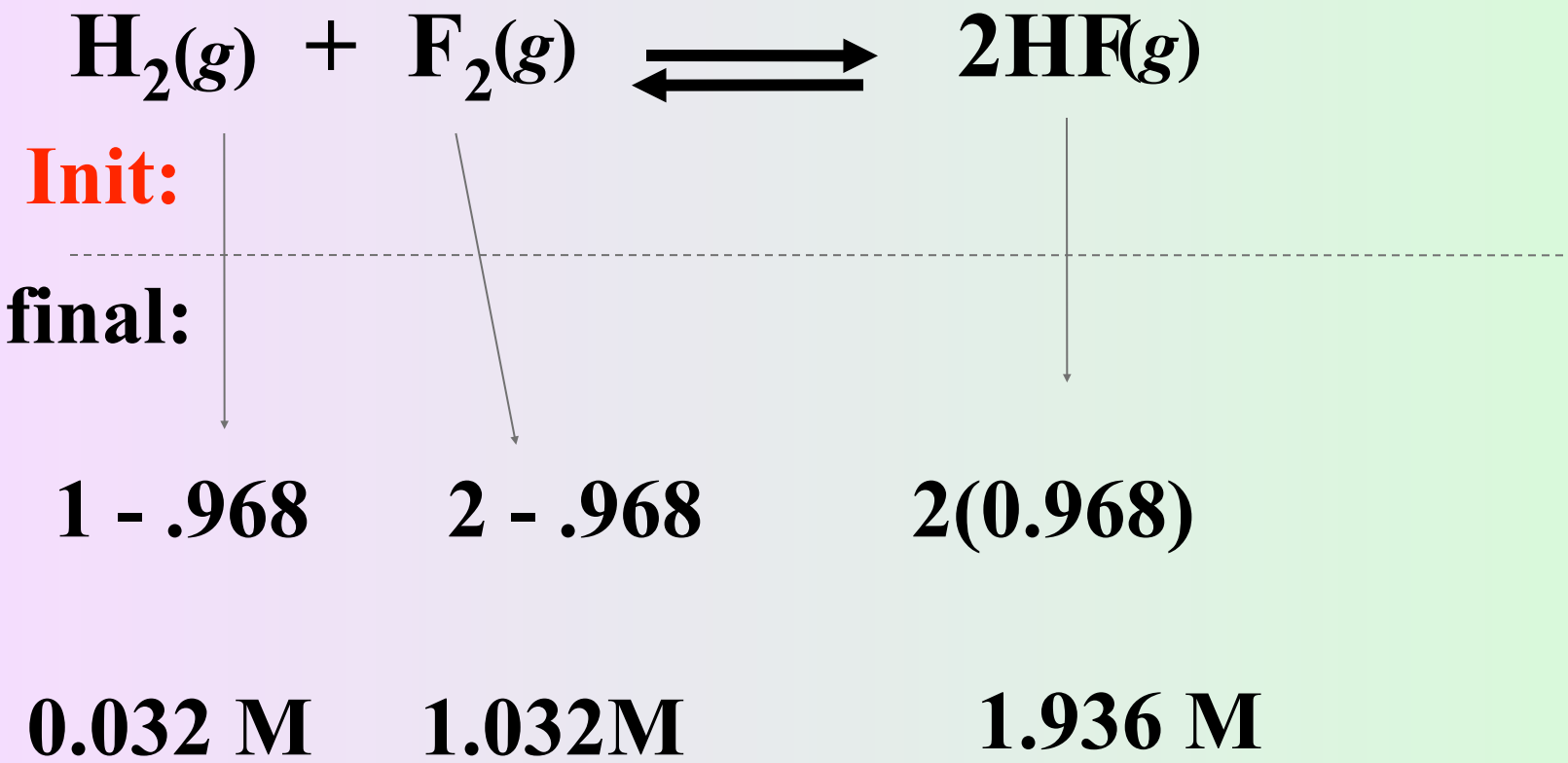
$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$111x^2 - 345x + 230 = 0$$

$$x = \frac{345 \pm \sqrt{-345^2 - 4(111)(230)}}{2(111)}$$

~~$x = 2.14$~~ $x = 0.968$

Practice Exercise



Practice Exercise

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

	$H_2(g)$	$I_2(g)$	$2HI$ (g)
Init:	0.00623 *M*	0.00414*M*	0.0224 *M*

How will the reaction shift ?

Calculate the equilibrium concentrations of all species present.

Practice Exercise

$$H_2(g) + I_2(g) \rightleftharpoons 2HI \quad (g) \qquad K = 54.3$$

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

Practice Exercise

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

	$H_2(g)$	$I_2(g)$	$2HI$ (g)
Init:	0.00623 *M*	0.00414*M*	0.0224 *M*
final:	0.00623 - x	0.00414*M* -x	0.0224 + 2x

$$54.3 = \frac{[HI]^2}{[H_2]\,[I_2]}$$

$$54.3 = \frac{(0.0224 + 2x)^2}{(0.00623 - x)\,(0.00414 - x)}$$

Practice Exercise

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

	H_2	I_2	$2HI$
Init:	0.00623 *M*	0.00414*M*	0.0224 *M*
final:	0.00623 - x	0.00414*M* -x	0.0224 + 2x

$$54.3 = \frac{(0.0224 + 2x)^2}{(0.00623-x)(0.00414-x)}$$

$$54.3\,(2.58 \times 10^{-5} - 0.0104x + x^2) =$$

$$5.02 \times 10^{-4} + 0.0896x + 4x^2$$

$$50.3\,x^2 - 0.654\,x + 8.98 \times 10^{-4} = 0$$

Practice Exercise

Quadratic equation

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$50.3\ x^2 - 0.654\ x + 8.98 \times 10^{-4} = 0$$

$$x = \frac{.654 \pm \sqrt{(-0.654)^2 - 4(50.3)(8.98 \times 10^{-4})}}{2(50.3)}$$

$x = 0.0114$

$x = 0.00156$

Practice Exercise

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

At equilibrium, the concentration are

$$[H_2] = (0.00623 - 0.00156)\ M = 0.00467\ M$$

$$[I_2] = (0.00414 - 0.00156)\ M = 0.00258\ M$$

$$[HI] = [0.0224 + 2(0.00156)]\ M = 0.0255\ M$$

Practice Exercise

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

	H_2	I_2	$2HI$
Init:	0.00 *M*	0.00 *M*	0.04 *M*
final:	+ x	+ x	0.04 - 2x

$$54.3 = \frac{[HI]^2}{[H_2]\,[I_2]}$$

$$54.3 = \frac{(0.040 - 2x)^2}{(x)(x)}$$

assume 2x is small compared to 0.04 and solve

If x is 5% or less than the original quantity our assumption is valid

Practice Exercise

[HI] = 0.040M

K_c = 54.3 at 430°C

Calculate the equilibrium concentrations of all species present.

Practice Exercise

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

	H_2	I_2	$2HI$
Init:	0.00 *M*	0.00 *M*	0.04 *M*
final:	+ x	+ x	0.04 - 2x

$$54.3 = \frac{[HI]^2}{[H_2]\,[I_2]}$$

$$54.3 = \frac{(0.040 - 2x)^2}{(x)(x)}$$

assume 2x is small compared to 0.04 and solve

If x is 5% or less than the original quantity our assumption is valid

Practice Exercise

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

	H_2	I_2	$2HI$
Init:	0.00 *M*	0.00 *M*	0.04 *M*
final:	+ x	+ x	0.04 - 2x

$$54.3 = \frac{[HI]^2}{[H_2]\,[I_2]}$$

$$54.3 = \frac{(0.04 - 2x)^2}{(x)(x)}$$

$$54.3 = \frac{(0.04)^2}{(x)(x)}$$

$$x = .0054$$

$$2x = .0108$$

$$\frac{.0108}{.04} \times 100 = 27\%$$

Practice Exercise

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

	H_2	I_2	$2HI$
Init:	0.00 *M*	0.00 *M*	0.040 *M*
final:	+ x	+ x	0.040 - 2x

$$54.3 = \frac{(0.040 - 2x)^2}{(x)(x)}$$

$$50.3\ x^2 + 0.16\ x - 0.0016 = 0$$

$$54.3\ x^2 = (0.040 - 2x)\ (0.040 - 2x)$$

$$54.3\ x^2 = 4x^2 - 0.16\ x + 0.0016$$

Practice Exercise

Quadratic equation

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$50.3\,x^2 + 0.16\,x - 0.0016 = 0$$

$$x = \frac{-.16 \pm \sqrt{.16^2 - 4(50.3)(-.0016)}}{2(50.3)}$$

$$x = 0.0043$$

Practice Exercise