CHEMICAL EQUILIBRIUM
Chapter 16

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

conc. of products

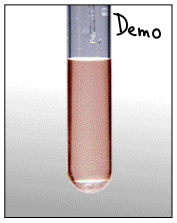
conc. of reactants

equilibrium constant

CONSTANT (at a given T)

Properties of an Equilibrium

Equilibrium systems are
- DYNAMIC (in constant motion)
- REVERSIBLE
- can be approached from either direction



Pink to blue

$Co(H_2O)_6Cl_2$ ---> $Co(H_2O)_4Cl_2$ + 2 H_2O

Blue to pink

$Co(H_2O)_4Cl_2$ + 2 H_2O ---> $Co(H_2O)_6Cl_2$

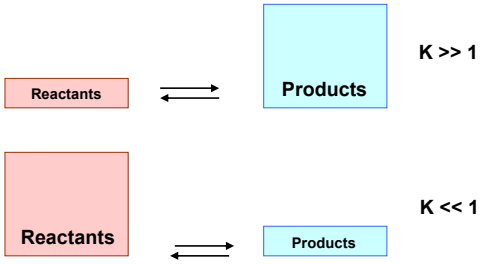
Heterogeneous equilibrium

Solids and liquids *NEVER* appear in equilibrium expressions.

$$S(s) + O_2(g) \leftrightharpoons SO_2(g)$$

$$K = \frac{[SO_2]}{[O_2]}$$

Product or Reactant Favored Reactions



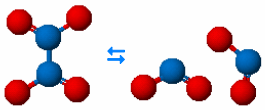
Today's Topics

- Temperature effect on equilibrium
- Equilibrium Constant expressions

Book: The Last Sorcerers (just for fun)
by Richard Morris
The path from Alchemy To The Periodic Table

Nitrogen Dioxide Equilibrium

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



Nitrogen Dioxide Equilibrium
$N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = 0.0059 \text{ at } 298\text{ K}$$

If initial concentration of N_2O_4 is 0.50 M, what are the equilibrium concentrations?

Step 1. Set up an ICE table

	$[N_2O_4]$	$[NO_2]$
Initial	0.50	0
Change	-x	2x
Equilib	0.50 - x	2x

Nitrogen Dioxide Equilibrium
$N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$

Step 2. Substitute into K_c expression and solve.

$$K_c = 0.0059 = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(2x)^2}{(0.50 - x)}$$

Rearrange: $0.0059\ (0.50 - x) = 4x^2$

$0.0029 - 0.0059x = 4x^2$

$4x^2 + 0.0059x - 0.0029 = 0$

This is a **QUADRATIC EQUATION**

$ax^2 + bx + c = 0$

$a = 4 \quad b = 0.0059 \quad c = -0.0029$

Nitrogen Dioxide Equilibrium
$N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$

Solve the quadratic equation for x.

$ax^2 + bx + c = 0$

$a = 4 \quad b = 0.0059 \quad c = -0.0029$

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

$$x = \frac{-0.0059 \pm \sqrt{(0.0059)^2 - 4(4)(-0.0029)}}{2(4)}$$

$x = -0.00074 \pm 1/8(0.046)^{1/2} = -0.00074 \pm 0.027$

Nitrogen Dioxide Equilibrium
$N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$

$$x = \frac{-0.0059 \pm \sqrt{(0.0059)^2 - 4(4)(-0.0029)}}{2(4)}$$

$x = -0.00074 \pm 1/8(0.046)^{1/2} = -0.00074 \pm 0.027$

$x = 0.026$ or -0.028

But a negative value is not reasonable.

Conclusion: *x = 0.026 M*

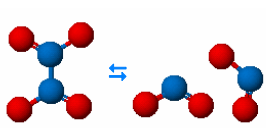
$[N_2O_4] = 0.50 - x = 0.47$ M

$[NO_2] = 2x = 0.052$ M

DEMO

Nitrogen Dioxide Equilibrium

$N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$



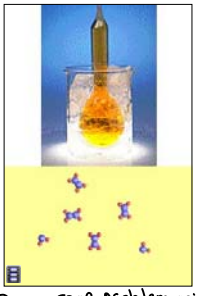
colorless — brown

Temperature Effects on Equilibrium

N_2O_4 + heat $\leftrightharpoons 2\ NO_2$

$\Delta H^o = +57.2$ kJ

endothermic reaction



$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

K_c (273 K) = 0.00077

K_c (298 K) = 0.0059

PRS → same problem using a ≠ K_c

$[NO_2]$=0.052 M

PRS

- **If initial concentration of N_2O_4 is 0.50 M, what is the equilibrium concentration of NO_2?** K_c(273 K)=0.00077

  $N_2O_4(g) \leftrightharpoons 2\ NO_2(g)$

- **Answer: 0.020 M**

Writing and Manipulating K Expressions

1. Adding equations for reactions

$S(s) + O_2(g) \leftrightharpoons SO_2(g)$ $\quad K_1 = \frac{[SO_2]}{[O_2]}$

$+$

$SO_2(g) + 1/2\ O_2(g) \leftrightharpoons SO_3(g)$ $\quad K_2 = \frac{[SO_3]}{[SO_2][O_2]^{1/2}}$

Net equation:

$S(s) + 3/2\ O_2(g) \leftrightharpoons SO_3(g)$ $\quad K_{net} = \frac{[SO_3]}{[O_2]^{3/2}}$

$$K_{net} = K_1 \cdot K_2$$

Adding two balanced equations **multiplies** their Ks

Writing and Manipulating K Expressions

2. Changing coefficients

$S(s) + 3/2\ O_2(g) \leftrightharpoons SO_3(g)$ $\quad K = \frac{[SO_3]}{[O_2]^{3/2}}$

$\times 2$

$2\ S(s) + 3\ O_2(g) \leftrightharpoons 2\ SO_3(g)$ $\quad K_{new} = \frac{[SO_3]^2}{[O_2]^3}$

$$K_{new} = \frac{[SO_3]^2}{[O_2]^3} = (K_{old})^2$$

Doubling the balanced equation **squares** the K

Writing and Manipulating K Expressions

3. Changing direction

$S(s) + O_2(g) \leftrightharpoons SO_2(g)$ $\quad K = \frac{[SO_2]}{[O_2]}$

$SO_2(g) \leftrightharpoons S(s) + O_2(g)$ $\quad K_{new} = \frac{[O_2]}{[SO_2]}$

$$K_{new} = \frac{[O_2]}{[SO_2]} = \frac{1}{K_{old}}$$

Reversing the balanced equation **inverts** the K