Equilibrium Law and the Equilibrium Constant

Chapter 7.2

- **Equilibrium law** is the mathematical description of a chemical system at equilibrium
- The **equilibrium constant (K)** is the numerical value defining the equilibrium law fora given system

$$aA + bB \rightleftharpoons cC + dD$$

The equation must be correctly balanced

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

K is a unitless quantity

K is constant for a given reaction at a given temperature

Example:

$4NH_{3\ (g)} + 7O_{2\ (g)} \leftrightarrow 4NO_{2\ (g)} + 6H_2O_{\ (g)}$

$$K_{eq} = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$$

Practice

- Write equilibrium law equations for these reactions:

a) $2O_{3\,(g)} \leftrightarrow 3O_{2\,(g)}$

b) $H_{2\,(g)} + F_{2\,(g)} \leftrightarrow 2HF_{(g)}$

Results for three experiments for the reaction:

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)} \quad \text{at } 500\ ^{\circ}C$$

expt	Initial Concentrations			Equilibrium Concentrations			$K = \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$
	$[N_2]$	$[H_2]$	$[NH_3]$	$[N_2]$	$[H_2]$	$[NH_3]$	
I	1.000	1.000	0	0.921	0.763	0.157	0.0602
II	0	0	1.00	0.399	1.197	0.203	0.0602
III	2.00	1.00	3.00	2.59	2.77	1.82	0.0602

$$N_2O_{4(g)} \leftrightarrow 2NO_{2(g)}$$

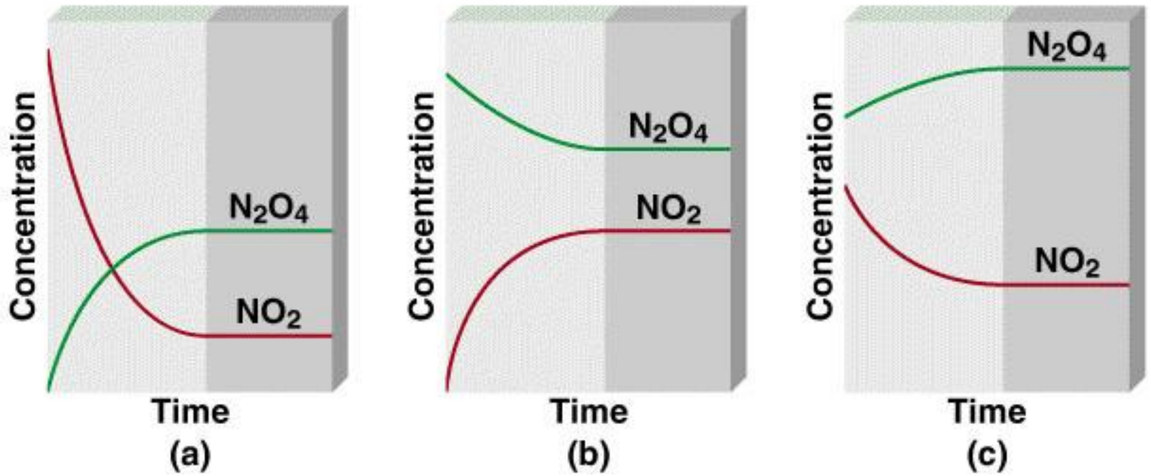


Table 14.1 The NO_2–N_2O_4 System at 25°C

Initial Concentrations (*M*)		Equilibrium Concentrations (*M*)		Ratio of Concentrations at Equilibrium	
$[NO_2]$	$[N_2O_4]$	$[NO_2]$	$[N_2O_4]$	$\frac{[NO_2]}{[N_2O_4]}$	$\frac{[NO_2]^2}{[N_2O_4]}$
0.000	0.670	0.0547	0.643	0.0851	4.65×10^{-3}
0.0500	0.446	0.0457	0.448	0.102	4.66×10^{-3}
0.0300	0.500	0.0475	0.491	0.0967	4.60×10^{-3}
0.0400	0.600	0.0523	0.594	0.0880	4.60×10^{-3}
0.200	0.000	0.0204	0.0898	0.227	4.63×10^{-3}

The Equilibrium Constant Varies with Temperature

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

Table 3 Equilibrium Constant for the Production of Ammonia Gas from Elemental Nitrogen and Hydrogen at Various Temperatures

Temperature (°C)	*K*
25	4.26×10^{8}
300	1.02×10^{-5}
400	8.00×10^{-7}

Practice

$$4SO_{2\,(g)} + O_{2\,(g)} \leftrightarrow 2SO_{3(g)}$$

Experiment 1

Initial	Equilibrium
$[SO_2] = 2.00M$	$[SO_2] = 1.50M$
$[O_2] = 1.50M$	$[O_2] = 1.25M$
$[SO_3] = 3.00M$	$[SO_3] = 3.50M$

Equilibrium constant for Experiment 1 =

Experiment 2

Initial	Equilibrium
$[SO_2] = 0.500M$	$[SO_2] = 0.590M$
$[O_2] = 0.00M$	$[O_2] = 0.045M$
$[SO_2] = 0.350M$	$[SO_3] = 0.260M$

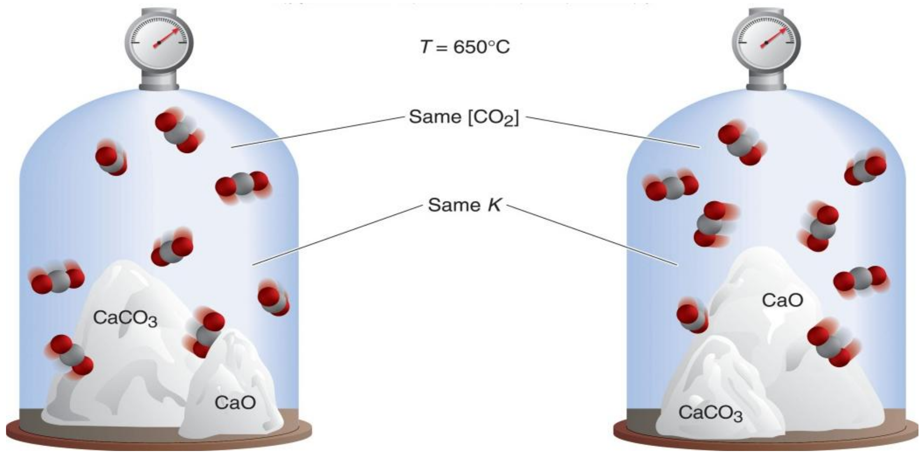
Equilibrium constant for Experiment 2 =

Heterogeneous Equilibria

- A **heterogeneous equilibrium** system is one in which the reactants and products are present in at least two different states, such as gases and solids
- If pure solids or pure liquids are involved in a chemical equilibrium system, their concentrations are not included in the equilibrium law equation for the reaction system

$$CaCO_{3\,(s)} \rightleftharpoons CO_{2\,(g)} + CaO_{(s)}$$

Calcium carbonate ⇌ carbon dioxide + lime



Practice

- Write equilibrium law equations for these reactions:

a) $H_{2\,(g)} + O_{2\,(g)} \rightleftharpoons H_2O_{(l)}$

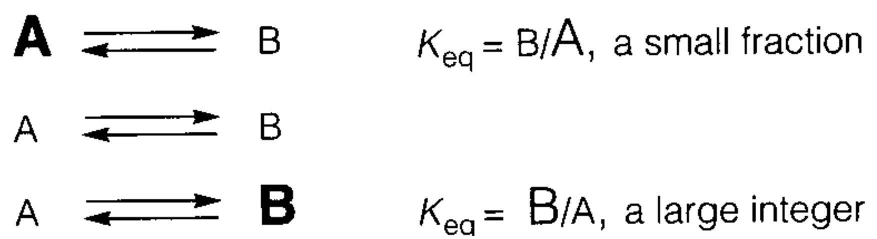
b) $NH_4Cl_{(s)} \rightleftharpoons NH_{3\,(g)} + HCl_{(g)}$

The Magnitude of K

- The magnitude of the equilibrium constant, K, tells us whether the equilibrium position favours products or reactants

$$2A + 5B \rightleftharpoons 3C + 4D$$

$$K = \frac{[\text{Products}]}{[\text{Reactants}]} \qquad K = \frac{[C]^3\,[D]^4}{[A]^2\,[B]^5}$$



If K=1

If K>1

If K<1

$K_{(forward)}$ and $K_{(reverse)}$

$$2O_{3(g)} \rightleftharpoons 3O_{2(g)}$$

HOMEWORK

Required Reading:

p. 429-436

(remember to supplement your notes!)

Questions:

p. 431 #1-3

p. 34 #1

p. 436 #1-6