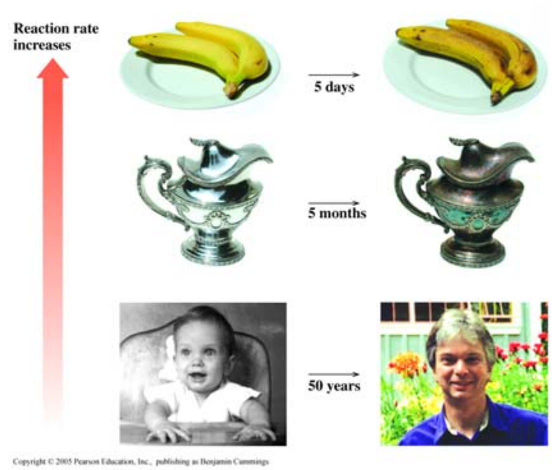
Chapter 13

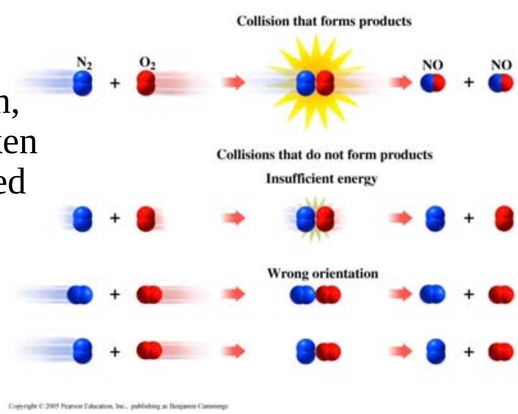
Chemical Equilibrium



Collision Theory

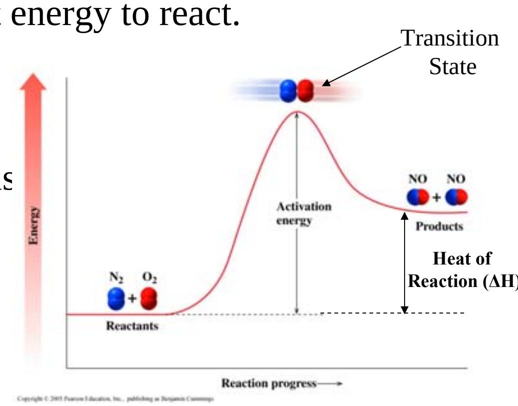
- The ***collision theory*** states that a reaction only takes place when molecules collide with the proper orientation and energy.

- In a successful collision, existing bonds are broken as new bonds are formed and the reactants are transformed into products.



Energy Barriers in Reactions

- For a chemical reaction to occur, the reactants must collide with sufficient energy to react.
- The amount of energy required to break the bonds between the atoms in the reactants is called the ***Activation Energy (E_{act})***
- The difference in the energy of the reactants and the products is called ***heat of reaction***, **ΔH**



This plot is called a "Reaction Profile"

Reaction Rates

- **Reaction Rate:** *The change in the concentration of a reactant or a product with time* (M/s).

Reactant $\rightarrow$ Products

aA $\rightarrow$ bB

$$\text{Rate} = -\frac{\Delta[A]}{\Delta t} \qquad \text{Rate} = \frac{\Delta[B]}{\Delta t}$$

Negative sign for the Reactants
Positive sign for the Products

Factors that Affect the Reaction Rate

- Table 13.1 Here

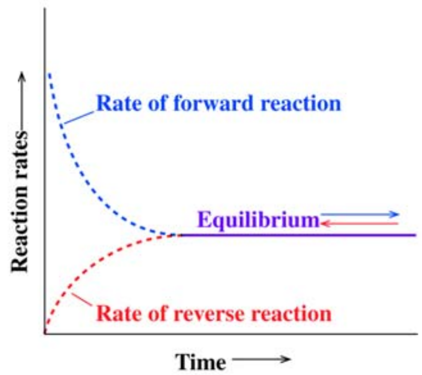
Chemical Equilibrium

- Most reactions are ongoing, reversible processes; preceding in both the forward direction to give products and in the reverse direction to give the original reactants.
- We indicate an equilibrium reaction with a double arrow:

$$\text{Reactants} \underset{\textit{reverse reaction}}{\overset{\textit{forward reaction}}{\rightleftharpoons}} \text{Products}$$

Chemical Equilibrium and Reaction Rates

- In an equilibrium reaction, initially the rate of the forward reaction is very fast.
- As more products are formed, the rate of the reverse reaction begins to speed up.
- When the rates of the forward and reverse reactions are the same, the system is at ***equilibrium***.



Equilibrium Constant Expression

- Consider the following general reaction:

$$a\,\mathrm{A} + b\,\mathrm{B} \leftrightarrows c\,\mathrm{C} + d\,\mathrm{D}$$

- Mathematically, we express the ***equilibrium constant expression*** as follows:

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

- The constant, K_{eq}, is the ***general equilibrium constant***.
- The value of K_{eq} varies with temperature. So a given value of K_{eq} is valid only for a specific temperature.

Homogeneous Equilibria

- A ***<u>homogeneous equilibrium</u>*** is a reaction where all of the products and reactants are in the physical same state.
- What is the equilibrium constant expression for the homogeneous equilibrium?

$$2\ SO_2(g) + O_2(g) \leftrightarrows 2\ SO_3(g)$$

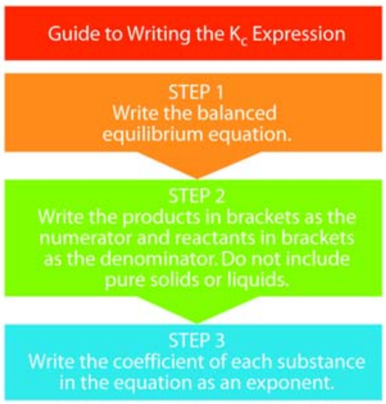
Heterogeneous Equilibria

- A ***<u>heterogeneous equilibrium</u>*** is a reaction where one of the substances is in a different physical state.

$$C(s) + H_2O(g) \leftrightarrows CO(g) + H_2(g)$$

- What is the Equilibrium expression?
- The concentrations of liquids and solids **<u>do not change</u>**, and they are therefore **<u>omitted</u>** from equilibrium constant expressions

Equilibrium Constant Expressions



What are the equilibrium constant expressions for the following reactions?

$NH_4NO_3(s) \leftrightarrows N_2O(g) + 2\ H_2O(g)$

$CO(g) + H_2(g) \leftrightarrows CH_3OH\ (l)$

$HCl(aq) + O_2(g) \leftrightarrows H_2O(l) + Cl_2(g)$

Using Equilibrium Constants

- The value of K_{eq} can indicate whether products or reactants are favored in a reaction.

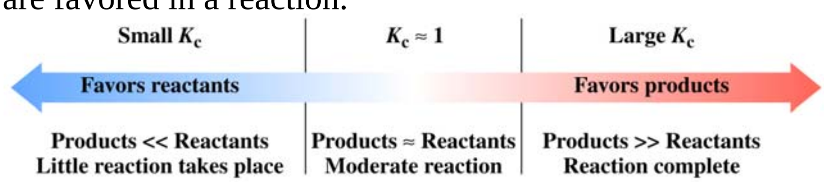


Table 12.3 Examples of Reactions with Large and Small K_c Values

Reactants Favors		Products	K_c	Equilibrium
$2CO(g) + O_2(g)$	$\rightleftharpoons$	$2CO_2(g)$	2×10^{11}	
$2H_2(g) + S_2(g)$	$\rightleftharpoons$	$2H_2S(g)$	1.1×10^{7}	
$N_2(g) + 3H_2(g)$	$\rightleftharpoons$	$2NH_3(g)$	1.6×10^{2}	
$PCl_5(g)$	$\rightleftharpoons$	$PCl_3(g) + Cl_2(g)$	1.2×10^{-2}	
$N_2(g) + O_2(g)$	$\rightleftharpoons$	$2NO(g)$	2×10^{-9}	
$COCl_2(g)$	$\rightleftharpoons$	$CO(g) + Cl_2(g)$	2.2×10^{-10}	