

April 14: Chemical Equilibrium

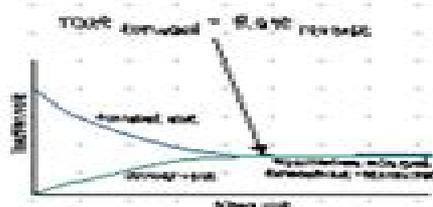
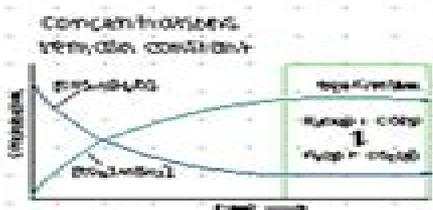
Chemical equilibrium: dynamic process in which concentrations of reactants + products remain constant over time

rate of reaction in forward direction matches rate in reverse direction - forward + reverse rates not zero at equilibrium



dynamic equilibrium: state of chemical reaction in which reactant and product concentrations do not change, but products and reactants are continuously interconverting

Equilibrium



EXAMPLE: The Haber process



- Initially, only N_2 and H_2 collide and react at some rate
- as NH_3 forms, there is less N_2 and H_2 therefore less likely to collide... less likely to react
- equilibrium concentrations are constant, rate of NH_3 dissociation + formation are equal



from experimental data

$$\text{rate}_f = k_f [\text{NO}_2]^2$$

$$\text{rate}_r = k_r [\text{N}_2\text{O}_4]$$

at equilibrium: $k_f [\text{NO}_2]^2 = k_r [\text{N}_2\text{O}_4]$

rearrange: $\frac{k_f}{k_r} = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} = K_{eq}$ (equilibrium constant)

Equilibrium constant

Equilibrium constant expression:

ratio of equilibrium concentrations or partial pressures of products to reactants each term raised to a power equal to the coefficient of that substance in the balanced chemical equation

Equilibrium constant:

value of the ratio of concentrations (or partial pressures) terms in the equilibrium constant expression at a specific temperature

for general overall reaction



equilibrium constant defined as:

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

where [A], [B], etc are equilibrium concentrations

$K \gg 1$ → equilibrium lies to right; PRODUCTS DOMINATE

$K \ll 1$ → equilibrium lies to left; REACTANTS DOMINATE

$K = 0$ → no reaction

$K = \infty$ → irreversible reaction

LAW OF MASS ACTION

equilibrium constant expression of chemical equilibrium has characteristic value @ given temperature

mass action expression:

equivalent to equilibrium constant expression, but applied to reaction mixtures that may or may not be at equilibrium



Experiment	Starting Concentrations	Equilibrium Concentrations	Calculated value of K_c
1	$[\text{H}_2] = 0.2000 \text{ M}$ $[\text{CO}] = 0.2000 \text{ M}$	$[\text{CH}_4] = 0.0400 \text{ M}$ $[\text{H}_2\text{O}] = 0.0400 \text{ M}$ $[\text{H}_2] = 0.1200 \text{ M}$ $[\text{CO}] = 0.1200 \text{ M}$	$K_c = 0.04$
2	$[\text{H}_2] = 0.2000 \text{ M}$ $[\text{CO}] = 0.2000 \text{ M}$	$[\text{CH}_4] = 0.0800 \text{ M}$ $[\text{H}_2\text{O}] = 0.0800 \text{ M}$ $[\text{H}_2] = 0.0800 \text{ M}$ $[\text{CO}] = 0.0800 \text{ M}$	$K_c = 0.04$
3	$[\text{H}_2] = 0.2000 \text{ M}$ $[\text{CO}] = 0.2000 \text{ M}$	$[\text{CH}_4] = 0.0200 \text{ M}$ $[\text{H}_2\text{O}] = 0.0200 \text{ M}$ $[\text{H}_2] = 0.1600 \text{ M}$ $[\text{CO}] = 0.1600 \text{ M}$	$K_c = 0.04$

equilibrium constant expression: $K_c = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{H}_2]^3[\text{CO}]}$

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