화학 General Chemistry
034.020-005

2018 Spring Semester

Tue/Thr 9:30~10:45
Building 028-302

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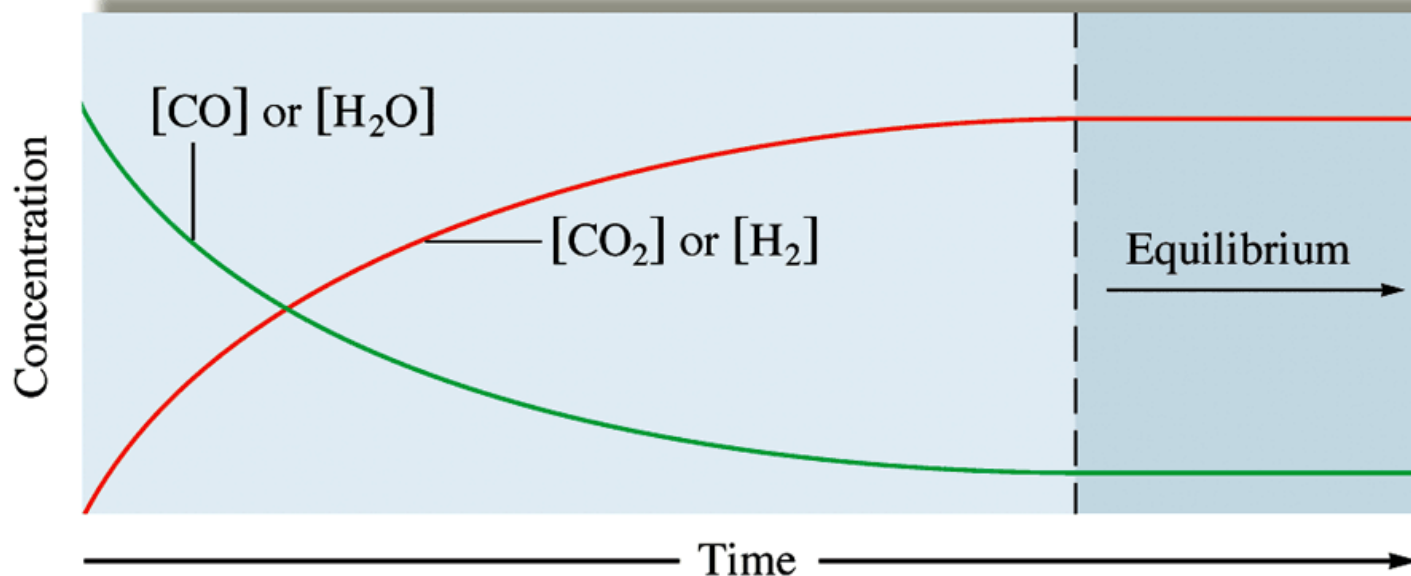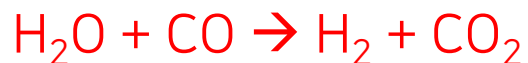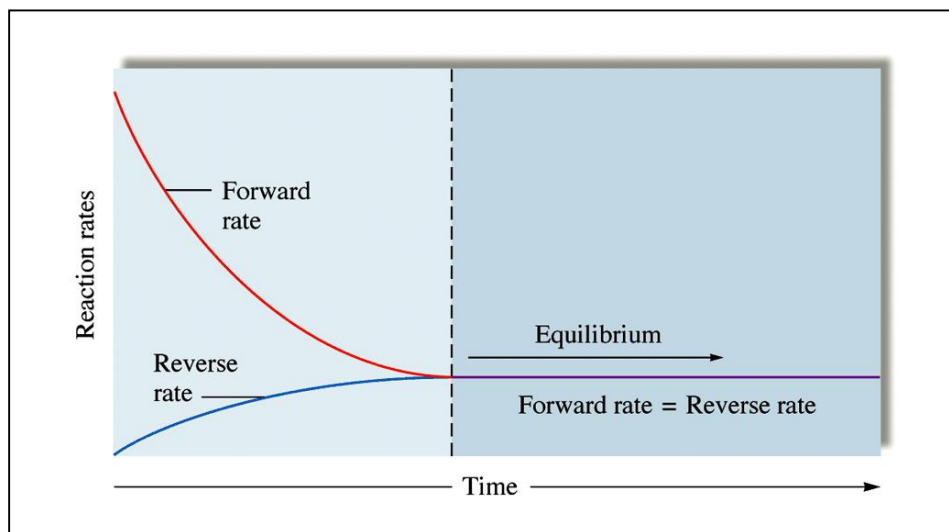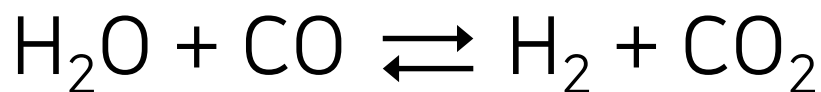
10.1. The Reversibility of Reactions

Chemical Equilibrium

• The state where the concentrations of all reactants and products remain constant with time.

For the rxn:
$H_2O + CO \rightarrow H_2 + CO_2$

At equilibrium, two reactions are occurring at the same rate

$H_2O + CO \rightarrow H_2 + CO_2$ k_f (Forward Reaction)

$H_2O + CO \leftarrow H_2 + CO_2$ k_b (Back Reaction)

$H_2O + CO \rightleftharpoons H_2 + CO_2$
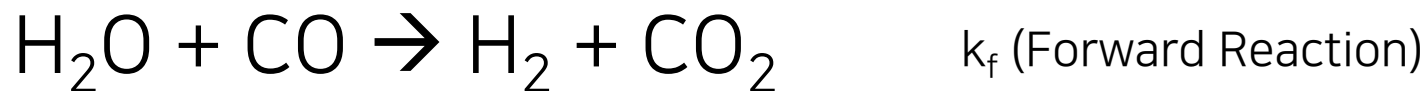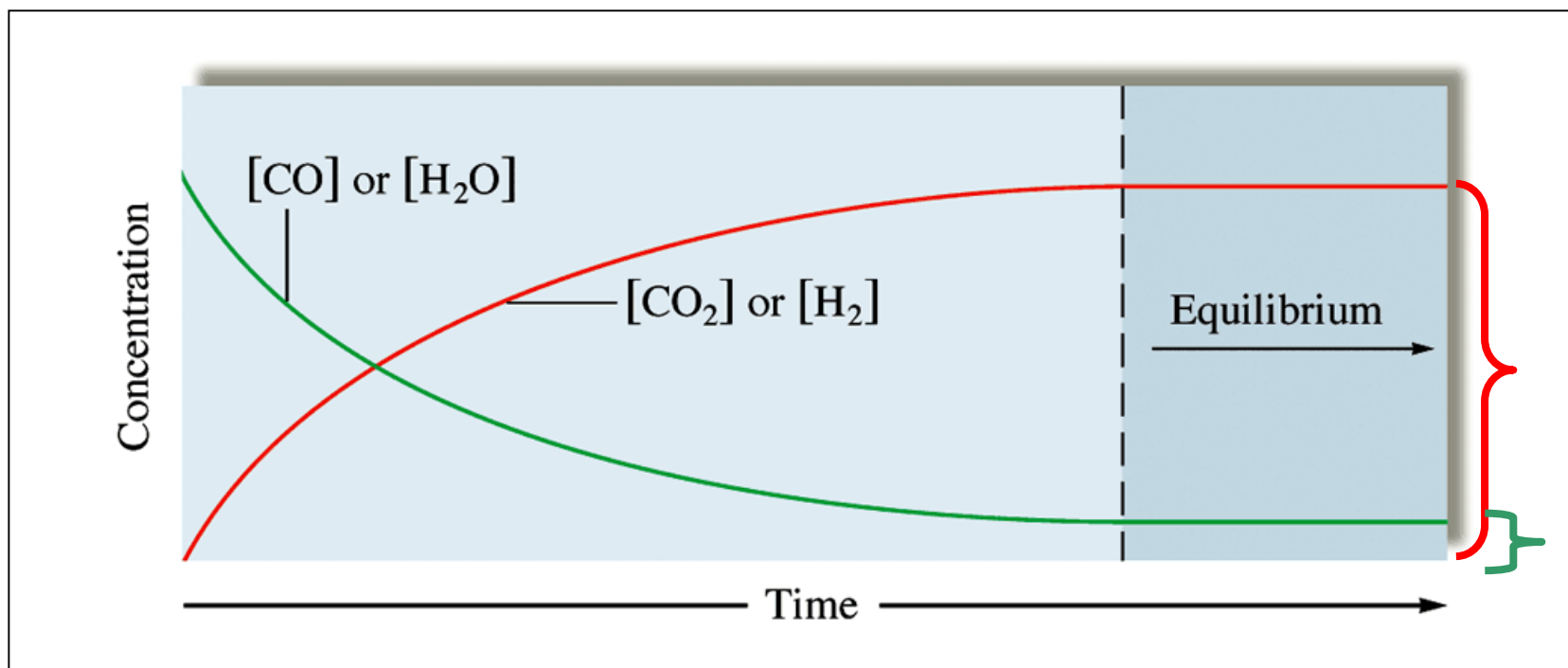
How much CO_2 and CO do we have
at equilibrium?

What determines the equilibrium concentration?

Chemical equilibrium:

(a) Time ⟶

(b) Time ⟶
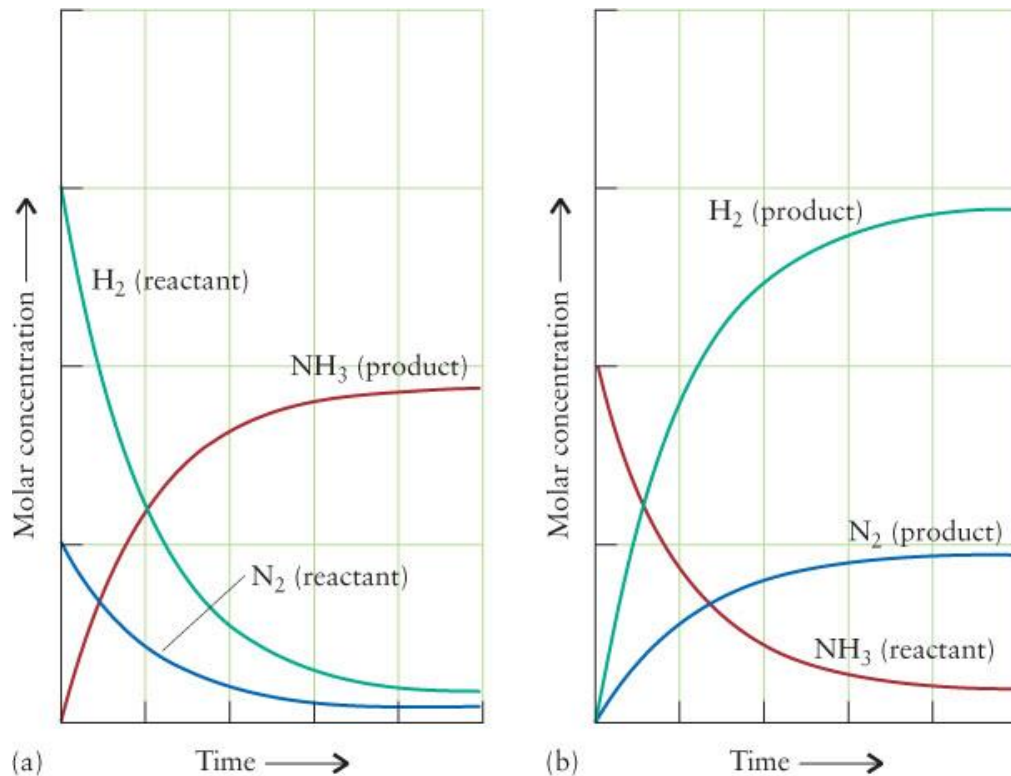
- **= concentration** of reactant and product remain **constant.**

- Forward and backward reactions are occurring **at the same time**

Forward:

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

Backward:

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

Dynamic Equilibrium:

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$
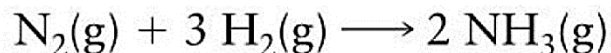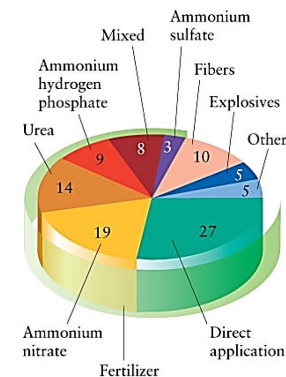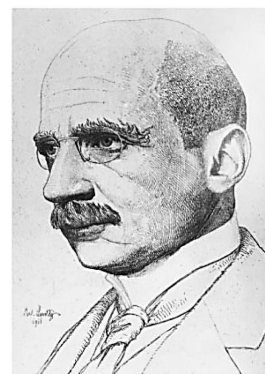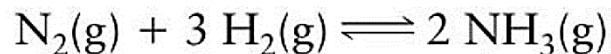
Haber-Bosch process

10.2. Equilibrium and the Law of Mass Action

$$j\text{A} + k\text{B} \rightarrow l\text{C} + m\text{D},$$
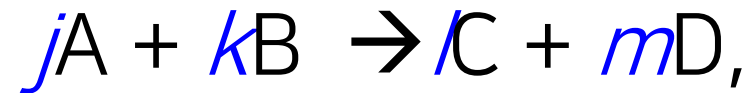
K : equilibrium constant

K is usually written without any unit

Note:

- The K is a constant for a given **temperature** (T)
- However, the equilibrium concentrations will not always be the same.

Ex)
A + B → C:
$$K = [C]/([A][B]) = 10$$

Equilibrium concentrations can be any of the following:

[C] = 10, [A] = 1, [B] = 1 ; K = 10/ (1 x 1) = 10
[C] = 20, [A] = 2, [B] = 1 ; K = 20 / (2 X 1) = 10

Equilibrium Expressions Involving Pressures

- Two different ways to define equilibrium constants:

$$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3 (g)$$
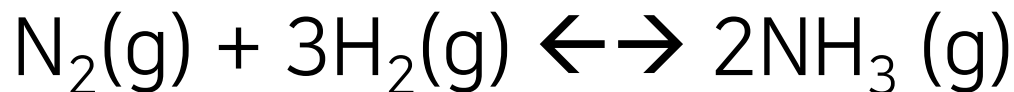
a. Use **concentration**:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Note:
PV = nRT,
P= (n/V)RT = [A]RT
[A] = P/(RT)

b. Use **Partial Pressure**:

$$K_p = \frac{P_{NH3}^2}{P_{N_2} P_{H_2}^3} = \frac{([NH_3]RT)^2}{([N_2]RT)([H_2]RT)^3} = K_c(RT)^{-2}$$

Thermodynamic description of the equilibrium state:

- For ideal gas: aA + bB → cC + dD

$$G^o : \text{Free energy at standard condition}$$
$$G = G^o + RT \ln\left(P / P^0\right)$$
$$G^o : \text{Free energy at standard condition (1 bar)}$$

$$\Delta G_{rxn} = cG_f(\text{C}) + dG_f(\text{D}) - aG_f(\text{A}) - bG_f(\text{B})$$

$$= \left(cG_f^o(\text{C}) + dG_f^o(\text{D}) - aG_f^o(\text{A}) - bG_f^o(\text{B})\right)$$

$$+ RT\left(c\ln\left(P_C / P^0\right) + d\ln\left(P_D / P^0\right) - a\ln\left(P_A / P^0\right) - b\ln\left(P_B / P^0\right)\right)$$

$$= \Delta G_{rxn}^o + RT \ln\left(\frac{(P_C / P^0)^c (P_D / P^0)^d}{(P_A / P^0)^a (P_B / P^0)^b}\right) = \Delta G_{rxn}^o + RT \ln\left(Q\right)$$

$Q = K$ and $\Delta G_{rxn} = 0$

Therefore,

$$\Delta G_{rxn}^o + RT \ln(K) = 0, \qquad \Delta G_{rxn}^o = -RT \ln(K),$$

$$K = \exp\left(-\frac{\Delta G_{rxn}^o}{RT}\right) \qquad K = \left(\frac{(P_{C,eq} / P^0)^c (P_{D,eq} / P^0)^d}{(P_{A,eq} / P^0)^a (P_{B,eq} / P^0)^b}\right) = \left(\frac{(P_{C,eq})^c (P_{D,eq})^d}{(P_{A,eq})^a (P_{B,eq})^b}\right)$$

Law of Mass Action

$$2 \, SO_2 \, (g) + O_2 \, (g) \rightleftharpoons 2 \, SO_3 \, (g)$$

P_{SO_2} (bar)	P_{O_2} (bar)	P_{SO_3} (bar)	K^*
0.660	0.390	0.0840	0.0415
0.0380	0.220	0.00360	0.0409
0.110	0.110	0.00750	0.0423
0.950	0.880	0.180	0.0408
1.44	1.98	0.410	0.0409

$$\boldsymbol{K} = \frac{(P_{SO3})^2}{(P_{SO2})^2 \, P_{O2}} \approx 0.0413 \; \text{(const)}$$

Law of mass action

Reaction	T (K)*	K	K_c†
$H_2(g) + Cl_2(g) \rightleftharpoons 2\ HCl(g)$	300	4.0×10^{31}	4.0×10^{31}
	500	4.0×10^{18}	4.0×10^{18}
	1000	5.1×10^8	5.1×10^8
$H_2(g) + Br_2(g) \rightleftharpoons 2\ HBr(g)$	300	1.9×10^{17}	1.9×10^{17}
	500	1.3×10^{10}	1.3×10^{10}
	1000	3.8×10^4	3.8×10^4
$H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$	298	794	794
	500	160	160
	700	54	54
$2\ BrCl(g) \rightleftharpoons Br_2(g) + Cl_2(g)$	300	377	377
	500	32	32
	1000	5	5
$2\ HD(g) \rightleftharpoons H_2(g) + D_2(g)$	100	0.52	0.52
	500	0.28	0.28
	1000	0.26	0.26
$F_2(g) \rightleftharpoons 2\ F(g)$	500	3.0×10^{-11}	7.3×10^{-13}
	1000	1.0×10^{-2}	1.2×10^{-4}
	1200	0.27	2.7×10^{-3}
$Cl_2(g) \rightleftharpoons 2\ Cl(g)$	1000	1.0×10^{-5}	1.2×10^{-7}
	1200	1.7×10^{-3}	1.7×10^{-5}
$Br_2(g) \rightleftharpoons 2\ Br(g)$	1000	3.4×10^{-5}	4.1×10^{-7}
	1200	1.7×10^{-3}	1.7×10^{-5}
$I_2(g) \rightleftharpoons 2\ I(g)$	800	2.1×10^{-3}	3.1×10^{-5}
	1000	0.26	3.1×10^{-3}
	1200	6.8	6.8×10^{-2}
$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$	298	6.8×10^5	4.2×10^8
	400	41	4.5×10^4
	500	3.6×10^{-2}	62
$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$	298	4.0×10^{24}	9.9×10^{25}
	500	2.5×10^{10}	1.0×10^{12}
	700	3.0×10^4	1.7×10^6
$N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$	298	0.15	6.1×10^{-3}
	400	47.9	1.44
	500	1.7×10^3	41
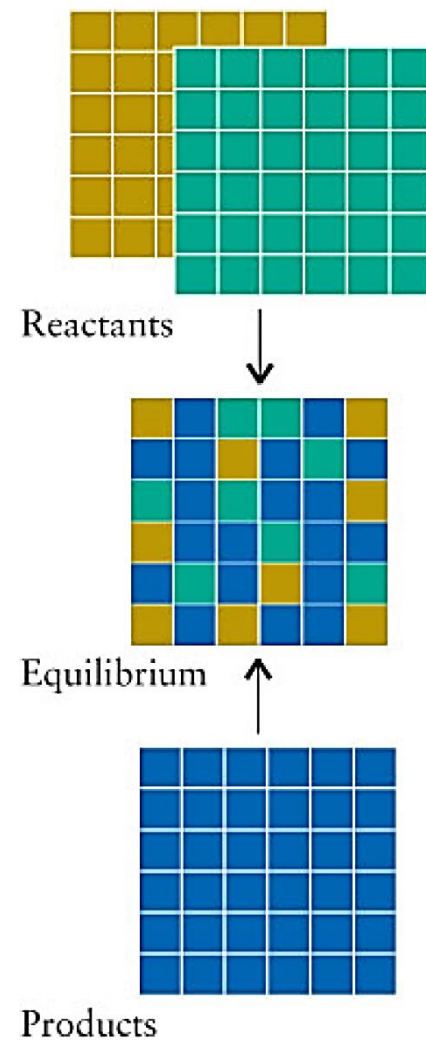
Reactants

Equilibrium

Products

- Homogeneous equil. : reactants and products all in the same phase

- Heterogeneous equil. : reactants and products with different phase

Examples

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

$$Ca(OH)_2\,(s) \leftrightarrow Ca^{2+}\,(aq) + 2\,OH^-\,(aq)$$

$$Ni\,(s) + 4\,CO\,(g) \leftrightarrow Ni(CO)_4\,(g)$$

$$2\,AgNO_3(aq) + 2\,NaOH\,(aq) \leftrightarrow Ag_2O\,(s) + 2\,NaNO_3\,(aq) + H_2O\,(l)$$

$$CH_3COOH\,(aq) + H_2O\,(l) \leftrightarrow CH_3COO^-\,(aq) + H_3O^+\,(aq)$$
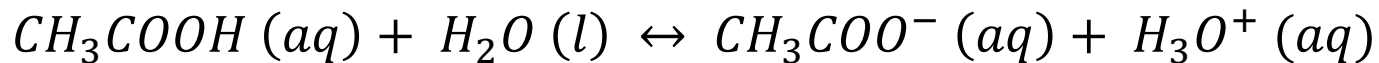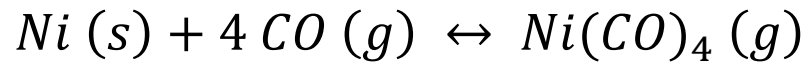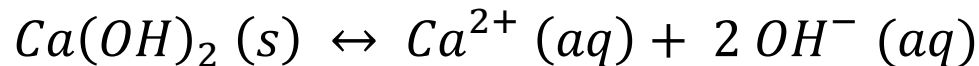
Heterogeneous Equilibria

- **Homogeneous** equil. : reactants and products all in the same phase

- **Heterogeneous** equil. : reactants and products with different phase

$$CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$$

The position of a heterogeneous equilibrium does not depend on the amounts of *pure solids or liquids* present, as long as there exist some amount of the solid and liquid in the equilibrium mixture.

→ *Pure solid and pure liquid concentration do not appear in the equilibrium equation !*
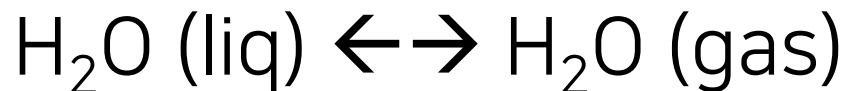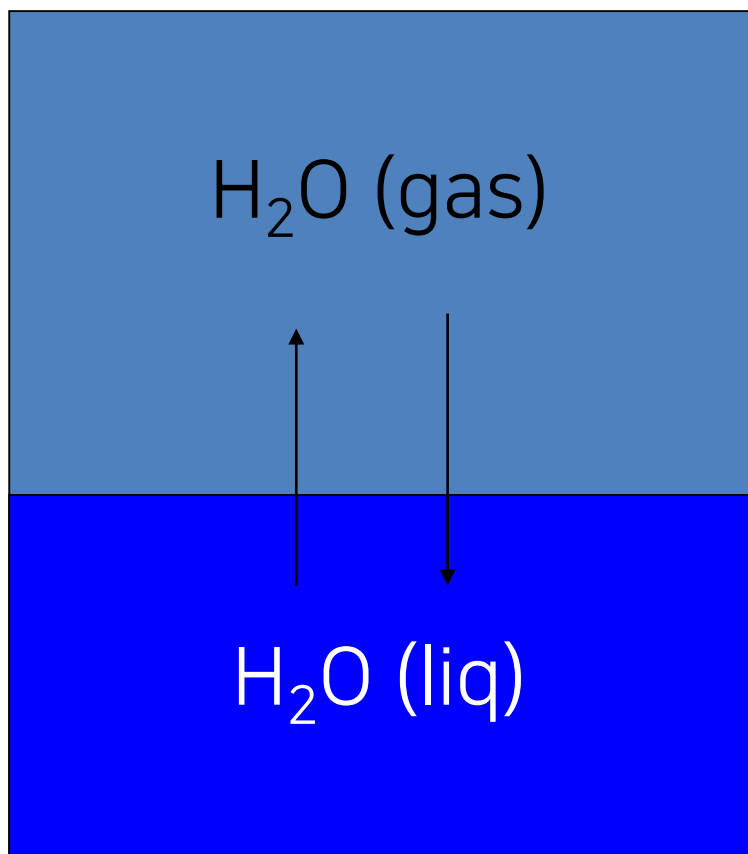
$$K = [CO_2(g)] \times 1/(1) = constant.$$

Same pressure of $CO_2(g)$
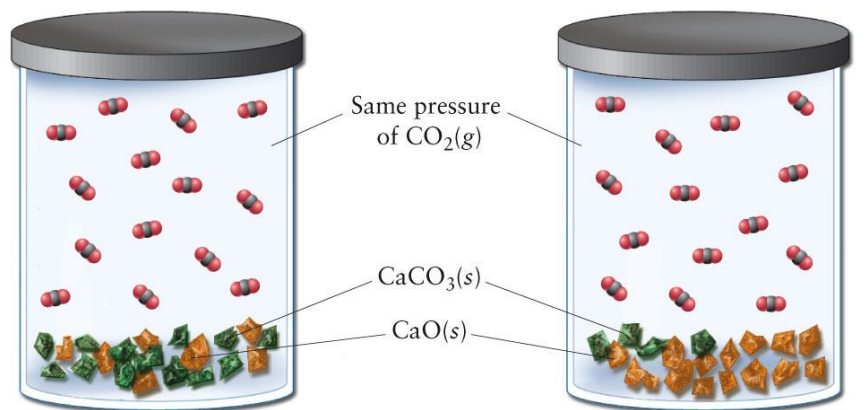
$CaCO_3(s)$

$CaO(s)$

Much $CaCO_3(s)$, little $CaO(s)$ Little $CaCO_3(s)$, much $CaO(s)$

Why do we omit pure liquids and solids in equilibrium expression?

- Densities of pure liquids and solids are constant for a given temperature.

- Therefore, "concentration" of pure liq. and solids are constant.
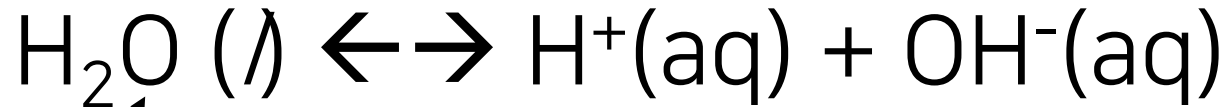
An example of heterogeneous equilibrium

H_2O (liq) $\longleftrightarrow$ H_2O (gas)

H_2O (gas)

H_2O (liq)

$K' = [H_2O(gas)] / [H_2O(liq)]$

Constant!!
55 M for pure water.

Therefore, we write:

$K = [H_2O(gas)]$

Heterogeneous Equilibria: other examples

$$H_2O\ (l) \longleftrightarrow H^+(aq) + OH^-(aq)$$

Pure liquid

$$K = [H^+][OH^-] = 10^{-14}$$

For neutral water , $[H^+] = [OH^-] = x$

$$K = [H^+][OH^-] = x^2 = 10^{-14}$$

$$Therefore,\ [H^+] = [OH^-] = 10^{-7}\ M$$

10.3. The Thermodynamic Origin of K
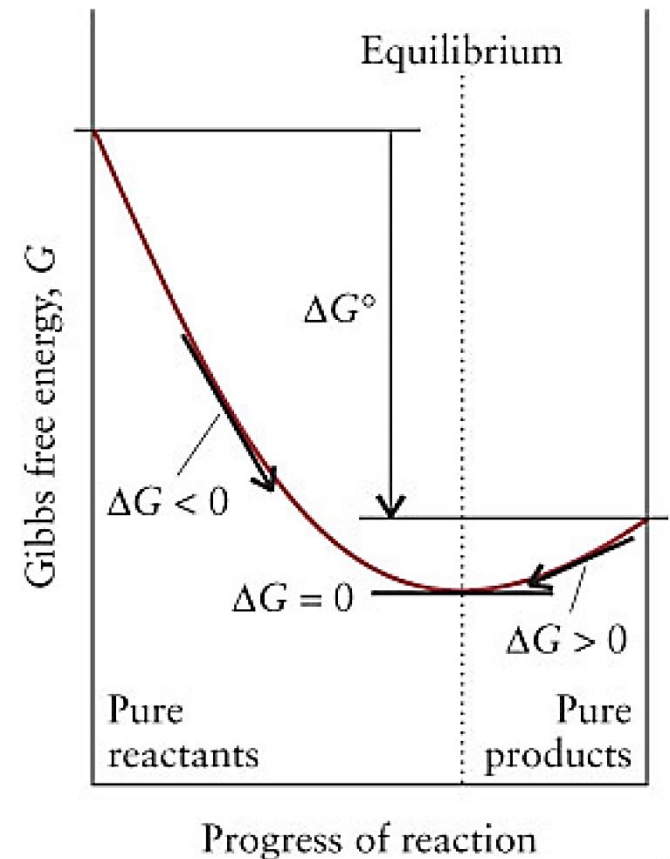
$G = G (T, P, x_i)$

x_i = mole fraction of component i

→ At constant T and P, it depends only on the extent of reaction, ξ

$$\left(\frac{\partial G}{\partial \xi}\right)_{T,P} = 0 \quad \text{at equilibrium}$$

Extent of reaction, /saɪ/

- Regardless of the initial composition of a reaction mixture, the composition tends to adjust until the activities give rise to the characteristic value of K for the reaction and temperature.

Equilibrium

Gibbs free energy, G

$\Delta G°$

$\Delta G < 0$

$\Delta G = 0$

$\Delta G > 0$

Pure reactants

Pure products

Progress of reaction

$$\Delta G_r{}^o = -RT \ln K$$

$$\Delta G_r = \sum_{products} n\, G - \sum_{reactants} n\, G$$

$$\Delta G_r{}^o = \sum_{products} n\, G^o - \sum_{reactants} n\, G^o$$

$G(A) = G^o(A) + RT \ln\left(\frac{P_A}{P^o}\right)$ for ideal gas

P^o: 1 bar standard pressure
c^o : 1 M standard molar conc,

$G(A) = G^o(A) + RT \ln\left(\frac{[A]}{c^o}\right)$ for solute in a dilute (ideal) solution

$G(A) = G^o(A) + RT \ln(1) = G^o(A)$ for pure solids and liquids

$G(A) = G^o(A) + RT \ln a_A$ for any substance in any state

For a reaction, $\alpha A + \beta B \leftrightarrow \gamma C + \delta D$,

$$\therefore \Delta G_r = \Delta G_r{}^o + RT \ln \frac{(a_C)^\gamma (a_D)^\delta}{(a_A)^\alpha (a_B)^\beta}$$

$$\rightarrow 0 = \Delta G_r{}^o + RT \ln \frac{(a_C)^\gamma (a_D)^\delta}{(a_A)^\alpha (a_B)^\beta} = \Delta G_r{}^o + RT \ln K \text{ at equilibrium}$$

$$\therefore \boldsymbol{\Delta G_r{}^o = -RT \ln K}$$

- The reaction Quotient Q defined as

$$\Delta G_r = \Delta G_r{}^o + RT \ln \frac{(a_C)^\gamma (a_D)^\delta}{(a_A)^\alpha (a_B)^\beta} = \Delta G_r{}^o + RT \ln \boldsymbol{Q}$$

Q: reaction quotient

Example

For the reaction 2 SO_2(g) + O_2 (g) → 2 SO_3(g), $\Delta G_r{}^o = -141.74 \ kJ \cdot mol^{-1}$ at 298 K. (a) what is the Gibbs free energy of reaction when the partial pressure of each gas is 100 bar, (b) what is the direction of spontaneous reaction ?

G (A)> 0 or <0??

$$G (A) = G^o(A) + RT \ln(K)$$

- Additionally,

$$\Delta G_r^o = \Delta H_r^o - T \Delta S_r^o$$

$$\Delta G_r^o = -RT \ln K$$

$$K = \exp(-\Delta G_r^o / RT) = \exp\left(\frac{-\Delta H_r^o}{RT} + \frac{\Delta S_r^o}{R}\right)$$

An highly **endothermic** reaction ($\Delta H_r^o \gg 0$) is likely to have **K < 1**.

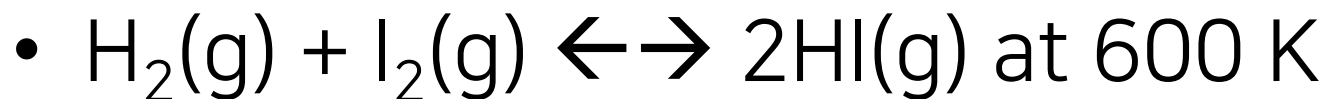An highly **exothermic** reaction ($\Delta H_r^o \ll 0$) is likely to have **K > 1**.

Example

For the reaction ½ H_2(g) + ½ I_2(g) → HI (g), $\Delta G_r^o = +1.7 \; kJ \cdot mol^{-1}$ at 298 K. Calculate K.

$$\Delta G_r^o = -RT \ln K$$

Applications of the Equilibrium Constant

- Equilibrium constant

→ What are the equilibrium concentrations?
→ Which direction does the reaction want to go?

Determination of Equilibrium Concentrations

- $H_2(g) + I_2(g) \leftarrow\rightarrow 2HI(g)$ at 600 K

Initial concentrations ➔ P_{H2} = 1.98 atm, P_{I2} = 1.710 atm

Equilibrium constant ➔ $K_p = P(HI)^2 / \{ P(H_2) \times P(I_2) \} = 92.6$

What are the equilibrium concentrations (or pressures) of the products and reactants?

	H_2	I_2	HI
Before the reaction	1.98	1.71	0
After the reaction	1.98-x	1.71-x	2 x

$$PCl_5(g) \longleftrightarrow PCl_3(g) + Cl_2(g)$$

1 Liter,
Inititally
0.298 mol, $PCl_3(g)$
8.70 x 10^{-3} mol $PCl_5(g)$

1 Liter,
Finally
2.00 x 10^{-3} mol $Cl_2(g)$

$[PCl_5]$, $[PCl_3]$, and K =?

Exer Calculating Equilibrium Concentrations

Carbon monoxide reacts with steam to produce carbon dioxide and hydrogen. At 700 K, the equilibrium constant is 5.10.

Calculate the equilibrium concentrations of all species if 1.000 mol of each component is mixed in a 1.000 liter flask.

Some Trick for complicated situations

Consider:

$2NOCl \longleftrightarrow 2NO + Cl_2$, $K = 1.6 \times 10^{-5}$
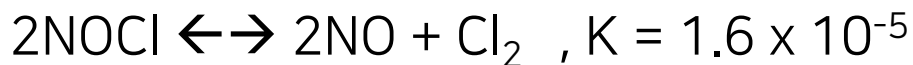
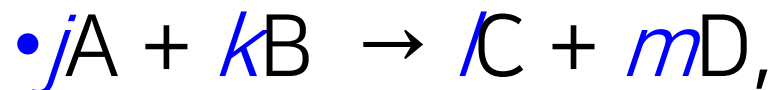|	[NOCl]	[NO]	[Cl$_2$]
Initially	0.5	0	0

(a) equilibrium constant is a very small number.
(b) In this case, x is also a very small number.

Applications of the Equilibrium Constant

- Equilibrium constant

→ What are the equilibrium concentration?
→ Which direction does the reaction want to go?

Reaction Quotient, Q
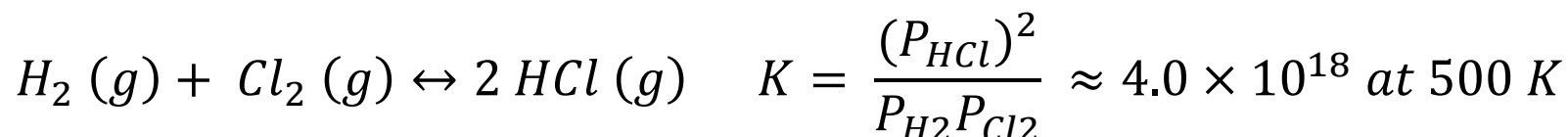
- For

  - jA + kB → lC + mD,

$$Q = \frac{[C]_0^l [D]_0^m}{[A]_0^j [B]_0^k}$$

$[X]_0$ = *Initial concentrations* of reagents and products.
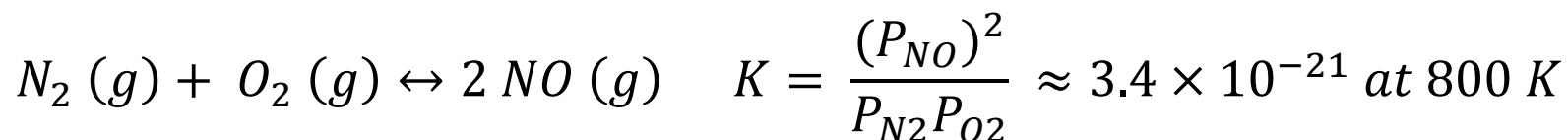
$$K = \frac{[C]_{eq}^l [D]_{eq}^m}{[A]_{eq}^j [B]_{eq}^k}$$ Equilibrium constant = Concentrations at equilibrium

10.4. The Extent of Reaction

- Large K

$$H_2\ (g) +\ Cl_2\ (g) \leftrightarrow 2\ HCl\ (g) \quad K = \frac{(P_{HCl})^2}{P_{H2}P_{Cl2}} \approx 4.0 \times 10^{18}\ at\ 500\ K$$

- Small K

$$N_2\ (g) +\ O_2\ (g) \leftrightarrow 2\ NO\ (g) \quad K = \frac{(P_{NO})^2}{P_{N2}P_{O2}} \approx 3.4 \times 10^{-21}\ at\ 800\ K$$
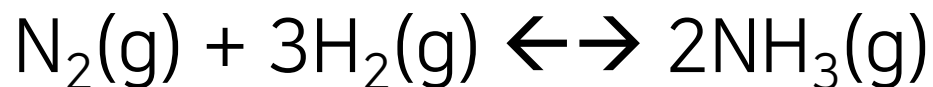
- Guidelines (not a rule)

If $K > 10^3$, products are favored.

If $10^{-3} < K < 10^3$, neither reactants nor products are strongly favored.

If $K < 10^{-3}$, reactants are favored.

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

$$Q = \frac{[NH_3]_0^2}{[N_2]_0[H_2]_0^3}$$

$[X]_0$ = Initital Concentrations

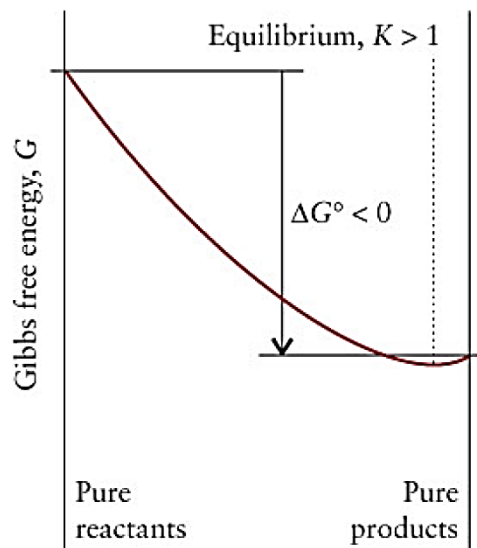| $[NH_3]_0$	$[N_2]_0$	$[H_2]_0$	Q	Q vs K (6.0 X10^{-2})
1.0×10^{-3}	1.0×10^{-5}	2.0×10^{-3}	1.3×10^{7}	
2.0×10^{-4}	1.5×10^{-5}	3.54×10^{-1}	6×10^{-2}	
1.0×10^{-4}	5.0	1.00×10^{-2}	2×10^{-3}	

10.5. The Direction of Reaction

- For a reaction, $\alpha A + \beta B \leftrightarrow \gamma C + \delta D$,

$$K = \left(\frac{(a_C)^\gamma (a_D)^\delta}{(a_A)^\alpha (a_B)^\beta}\right)_{equilibrium} \quad \text{and} \quad Q = \left(\frac{(a_C)^\gamma (a_D)^\delta}{(a_A)^\alpha (a_B)^\beta}\right)_{not\ at\ equilibrium}$$

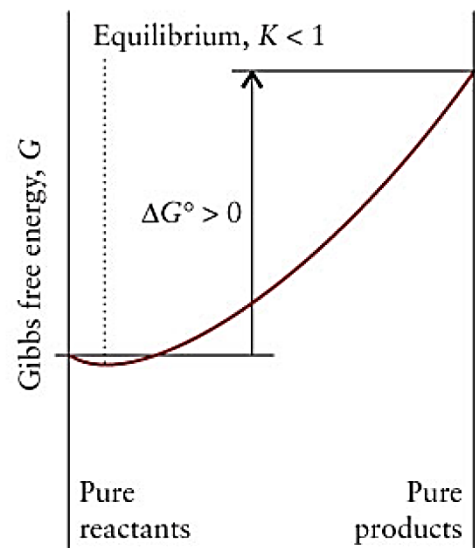If Q < K, the reaction proceeds in the forward direction ($\rightarrow$).

If Q > K, the reaction proceeds in the reverse direction ($\leftarrow$).

If Q = K, dynamic equilibrium ($\leftrightarrow$).

Equilibrium Calculation
10.6. The Alternative Form of K
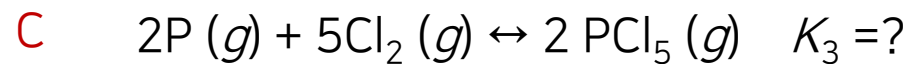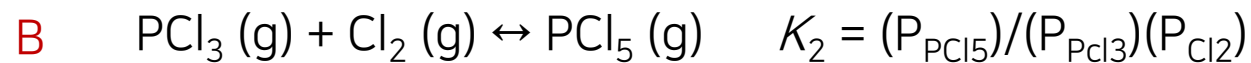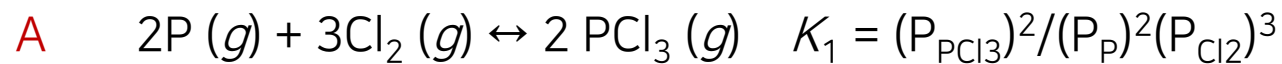
$$\alpha \, A + \beta \, B \leftrightarrow \gamma \, C + \delta \, D \qquad K_1 = \left(\frac{(a_C)^\gamma \, (a_D)^\delta}{(a_A)^\alpha \, (a_B)^\beta} \right)_{equil}$$

$$\gamma \, C + \delta \, D \leftrightarrow \alpha \, A + \beta \, B \qquad K_2 = \left(\frac{(a_A)^\alpha \, (a_B)^\beta}{(a_C)^\gamma \, (a_D)^\delta} \right)_{equil} = K_1^{-1}$$

$$n\alpha \, A + n\beta \, B \leftrightarrow n\gamma \, C + n\delta \, D \qquad K_3 = \left(\frac{(a_C)^{n\gamma} \, (a_D)^{n\delta}}{(a_A)^{n\alpha} \, (a_B)^{n\beta}} \right)_{equil} = K_1^{\,n}$$

$$\Delta G_r^{\,o} = \sum_{products} n \, G^o - \sum_{reactants} n \, G^o$$

$$\Delta G_r^{\,o} = -RT \, \ln K$$

A $2P\ (g) + 3Cl_2\ (g) \leftrightarrow 2\ PCl_3\ (g)$ $K_1 = (P_{PCl3})^2/(P_P)^2(P_{Cl2})^3$

B $PCl_3\ (g) + Cl_2\ (g) \leftrightarrow PCl_5\ (g)$ $K_2 = (P_{PCl5})/(P_{Pcl3})(P_{Cl2})$

C $2P\ (g) + 5Cl_2\ (g) \leftrightarrow 2\ PCl_5\ (g)$ $K_3 = ?$

10.6. Le Chatelier's Principle

Henry Le Chatelier
1850-1936 (France)

If **a change** is imposed on a system at equilibrium, *the position of the equilibrium will shift in a direction that tends to reduce that change.*

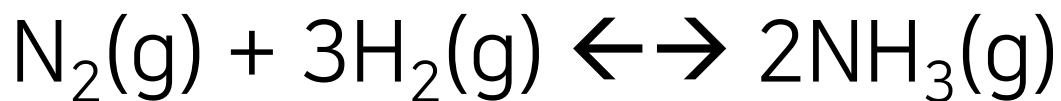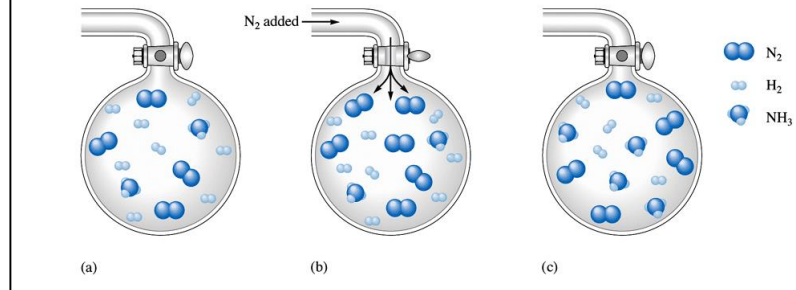** a change*
= pressure, temperature, and concentration

Note: this is not a new principle or discovery.
→Le Chatelier's principle is already contained in the equilibrium expression. Le Chatelier's principle laid a foundation for Chemical equilibrium theory, actually.

Le Chatelier's Principle-1: Change in Concentration

$$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g)$$

Initial Equilibrium	0.399 M	1.197 M	0.202 M	$K = 1.7 \times 10^{-2}$

Add 1.0 M

Le Chatelier's principle:

Concentration of N_2 is too high → FORWARD REACTION (→) in order to DECREASE the N_2 concentration

|			
Final Equilibrium	1.348 M	1.044 M	0.304 M	$K = 1.7 \times 10^{-2}$

Result :
 * Increase in NH_3 concentration   * Decrease in H_2 concentration

Le Chatelier's Principle-2: Change in Pressure

1. Add or remove a gaseous reactant or product (already talked about it).
2. Add an inert gas (one not involved in the reaction).
   → inert gas increase the total pressure, but has no effect on equilibrium (no change in volume).
3. Change the volume of container.
 Container volume reduction
→ Try to decrease its own volume by reducing the total number of gas-phase molecules.

$$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g)$$
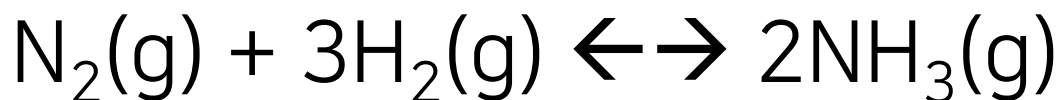
4 molecules   2 molecules

Decrease in volume
→ wants to decrease the number of molecules.
→ EQUILIBRIUM SHIFTS TO RIGHT.

Increase in volume
→ wants to INCREASE the number of molecules.
→ EQUILIBRIUM SHIFTS TO LEFT

Using Le Chatelier's principle II

| | NO. of gas-phase molecules | | |
	LEFT	RIGHT	Decrease in volume leads to···
$P_4(s) + 6Cl_2 (g) \leftrightarrow 4PCl_3(l)$	6	0	
$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$	2	1	
$PCl_3(g) + 3NH_3(g) \leftrightarrow P(NH_2)_3(g) + 3HCl(g)$	4	4	

Le Chatelier's Principle-3: Change in Temperature

- Note: changes in pressure and volume change the equilibrium. However, these won't change the equilibrium constant.

- Change in temperature changes the equilibrium constant.

- Equilibrium constant is a function of temperature.

The Response of Equil. to Changes in Conditions
10.7. Adding and Removing Reagents

- Le Chatelier's Principle

When a stress is applied to a system in dynamic equilibrium, the equilibrium tens to adjust to minimize the effect of the stress. Experimentally discovered, however, thermodynamically expected too.
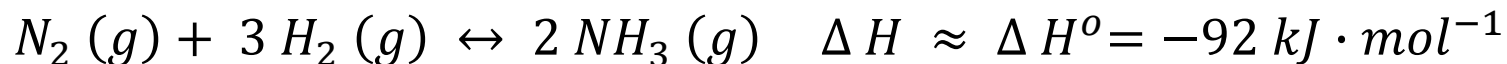
$$H_2\,(g) +\, Cl_2\,(g) \leftrightarrow 2\,HCl\,(g) \qquad Q = \frac{(P_{HCl})^2}{P_{H2}P_{Cl2}}$$
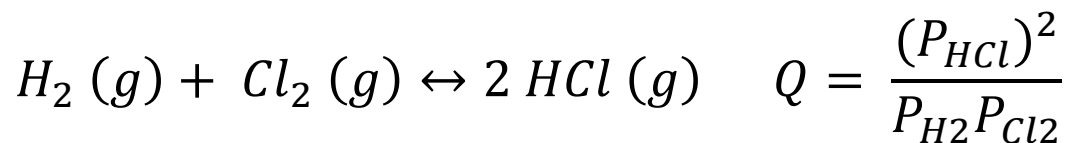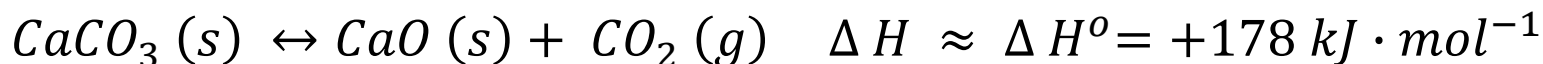
Adding H_2 results in Q < K. The system proceeds forward.

Adding HCl results in Q > K. The system proceed backward.

10.8. Temperature and Equilibrium

- For exothermic reaction, K ?

$$N_2\ (g) + 3\ H_2\ (g)\ \leftrightarrow\ 2\ NH_3\ (g)\quad \Delta H\ \approx\ \Delta H^o = -92\ kJ \cdot mol^{-1}$$

-For endothermic reaction, K ?

$$CaCO_3\ (s)\ \leftrightarrow CaO\ (s) +\ CO_2\ (g)\quad \Delta H\ \approx\ \Delta H^o = +178\ kJ \cdot mol^{-1}$$

- Thermodynamically,

$$K = \exp(-\Delta G_r{}^o / RT) = \exp\left(\frac{-\Delta H_r{}^o}{RT} + \frac{\Delta S_r{}^o}{R}\right)$$

- van't Hoff Equation

$$\ln\frac{K_1}{K_2}\ \approx\ -\ \frac{\Delta H_r{}^o}{R}\left\{\frac{1}{T_1} - \frac{1}{T_2}\right\}$$

Example The K for the synthesis of ammonia is 6.8×10^5 at 298 K.
Predict its value at 400 K. (compare with the experimental value: 41).

10.9. Catalysts and Haber's Achievement

- Catalyst (and activation energy)

→ Rates of forward and reverse reactions are increased.

→ Dynamic equilibrium is unaffected.

→ K is not changed.

→ The identity of the catalyst is unchanged.

- Haber's achievement
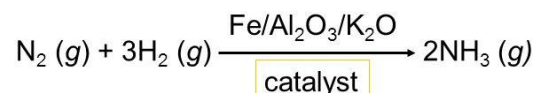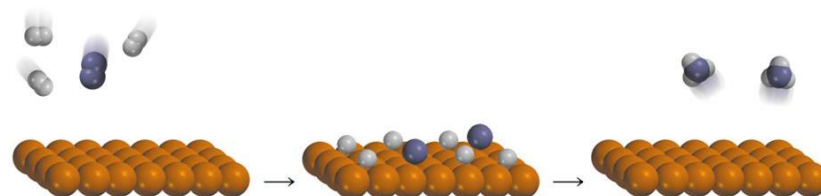
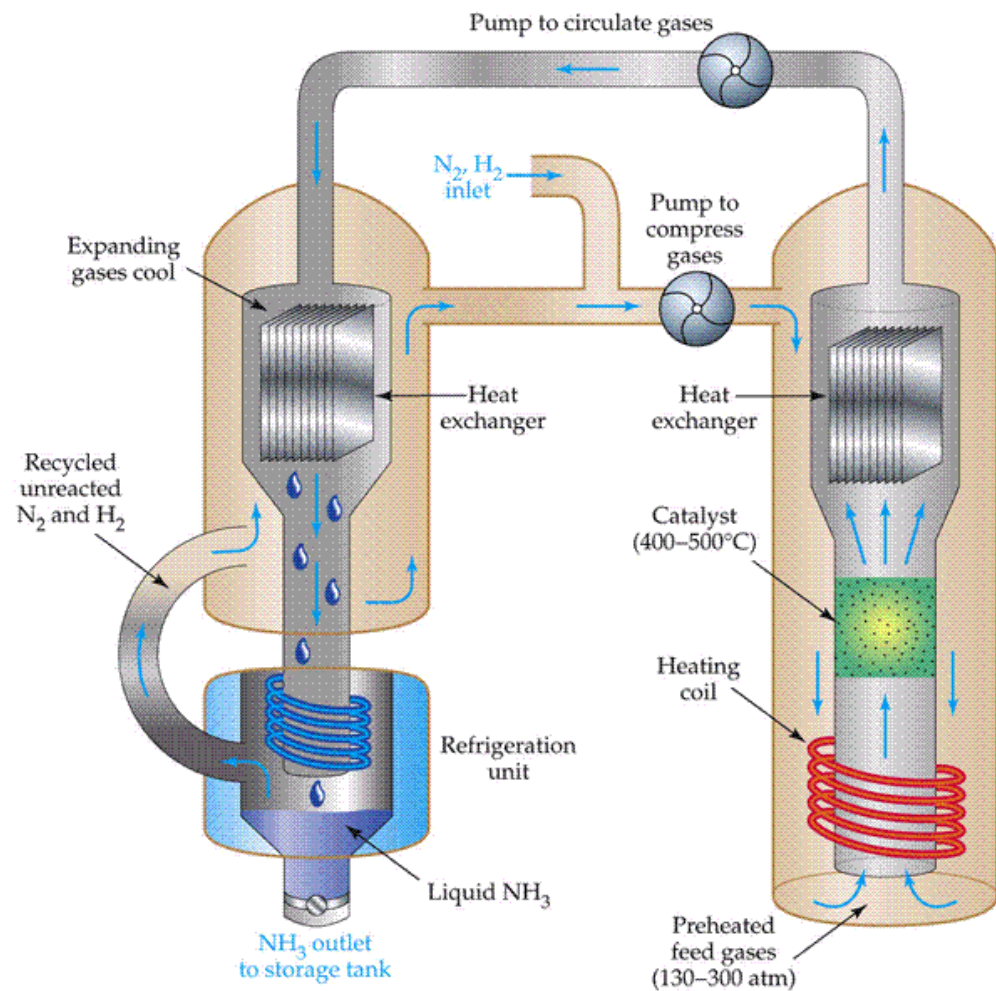$$N_2\,(g) +\ 3\,H_2\,(g) \leftrightarrow 2\,NH_3\,(g) \quad \Delta H_r^{\,o} = -92\,kJ \cdot mol^{-1}$$
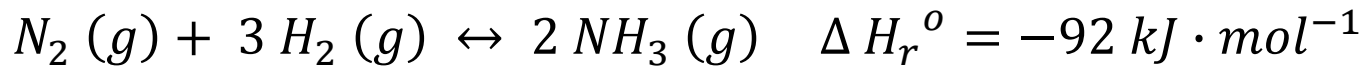
$$\Delta G_r^{\,o} = -32.90\,kJ \cdot mol^{-1} \quad K = 5.80\ \times\ 10^5$$

→ Although thermodynamically favorable, the reaction is very slow at room temperature and pressure.

→ Compression of the gas mixture favors product NH_3.

→ Removal of NH_3 as it is formed encourages more to be formed.

→ Increasing T is required to increase the reaction rate, but it adversely affects the position of the equilibrium

→ Catalyst.
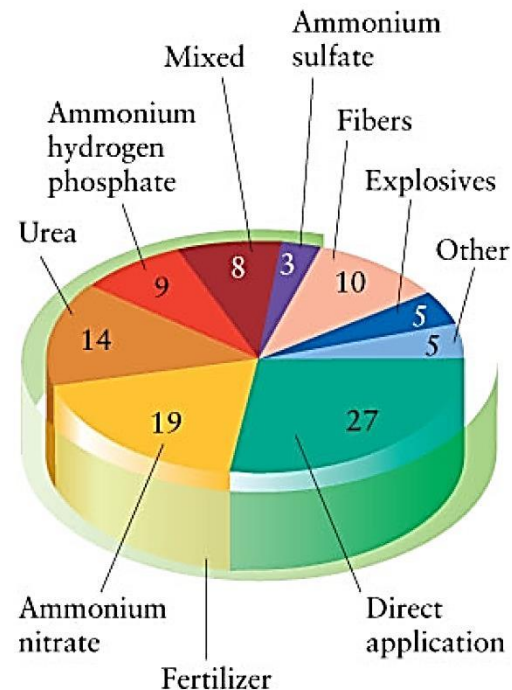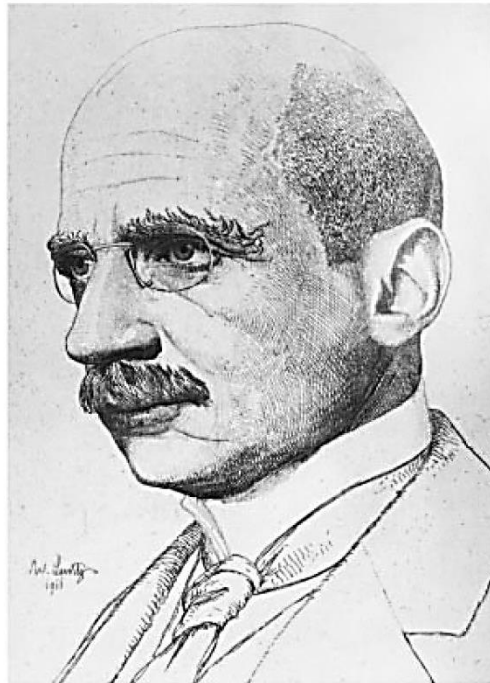
The most popular catalysts are based on iron promoted with K_2O, CaO, SiO_2, and Al_2O_3. Initially, Os and U catalysts were used.

$$N_2\,(g) + 3H_2\,(g) \xrightarrow[\text{catalyst}]{\text{Fe/Al}_2\text{O}_3\text{/K}_2\text{O}} 2NH_3\,(g)$$

Pump to circulate gases

N_2, H_2 inlet

Pump to compress gases

Expanding gases cool

Heat exchanger

Heat exchanger

Recycled unreacted N_2 and H_2

Catalyst (400–500°C)

Heating coil

Refrigeration unit

Liquid NH_3

NH_3 outlet to storage tank
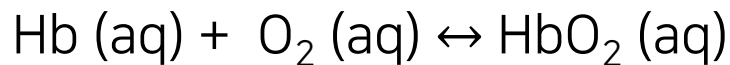
Preheated feed gases (130–300 atm)

10.10. The Impact on Biology: Homeostasis

- Homeostasis

Mechanism similar to chemical equilibrium, which governed by Le Chatelier's Principle

Hemoglobin: Oxygen carrier

Hb (aq) + O_2 (aq) ↔ HbO_2 (aq)

Myoglobin: Oxygen storage protein

CO Poisoning