

CHEM103

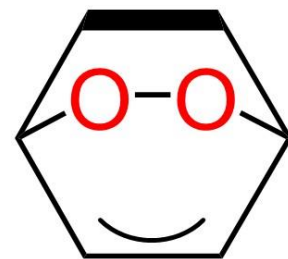
General Chemistry

Chapter 15: Chemical Equilibrium



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Homeworks 13 & 14

Homework 13

Due date: 28th Nov. (Mon)

Homework 14

Due date: 5th Dec. (Mon)



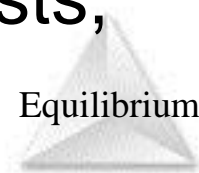
Review on Chapter 14

Rate (average/instantaneous/initial rate), rate law, rate constant, reaction order, half-life

Collision model, activation energy (barrier), transition state (activated complex), reaction coordinate diagrams, Arrhenius equation

Reaction mechanism: elementary process, intermediate, molecularity (unimolecular, bimolecular, termolecular), rate-determining step

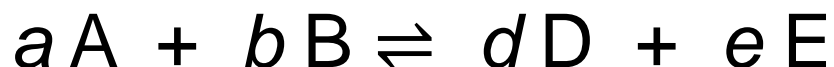
Catalyst: homogeneous/heterogeneous catalysts, enzymes



Equilibrium

Outline of Chapter 15

Equilibrium constant (K),
Homogeneous equilibria, Heterogeneous equilibria



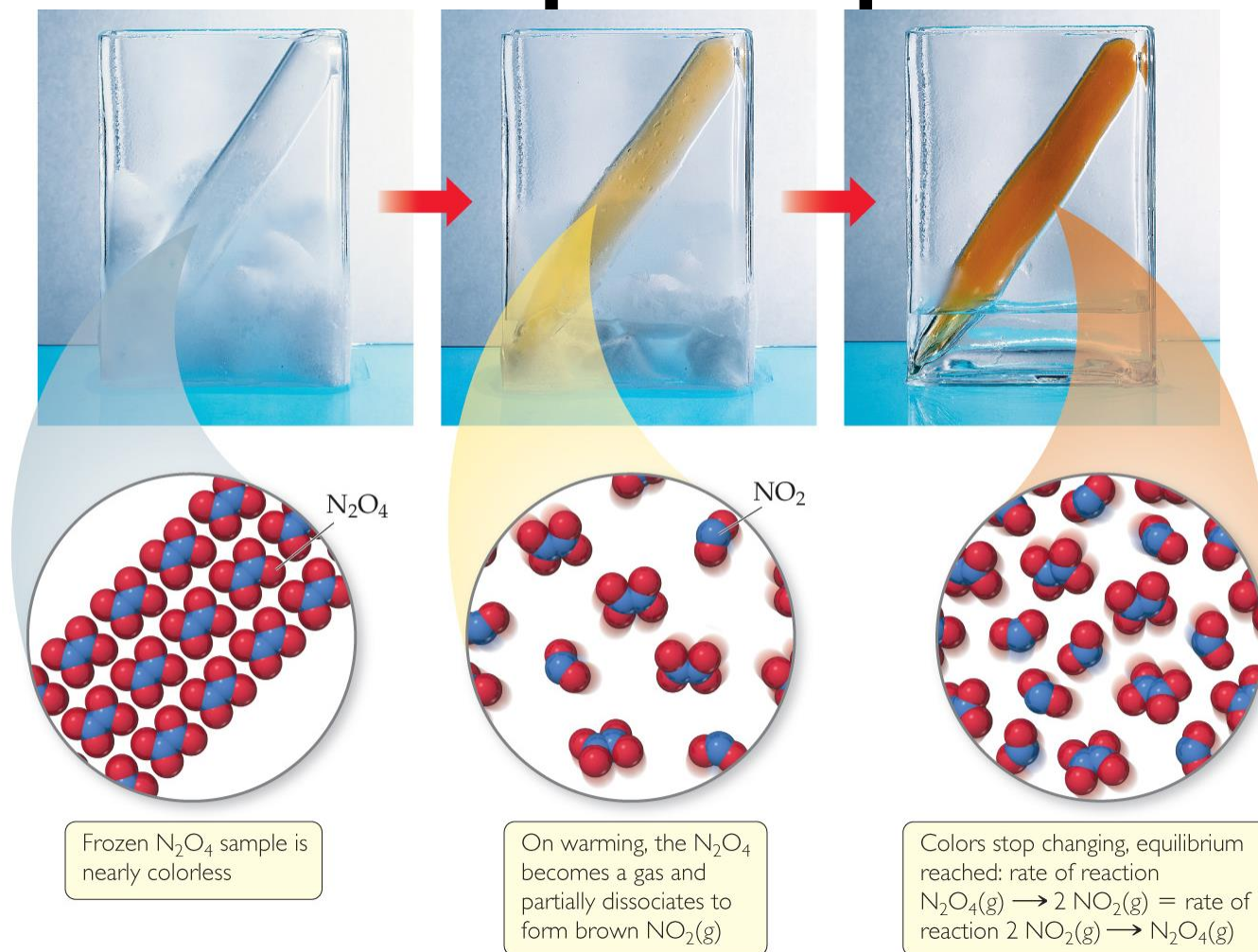
$$K_c = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$

Reaction quotient (Q)

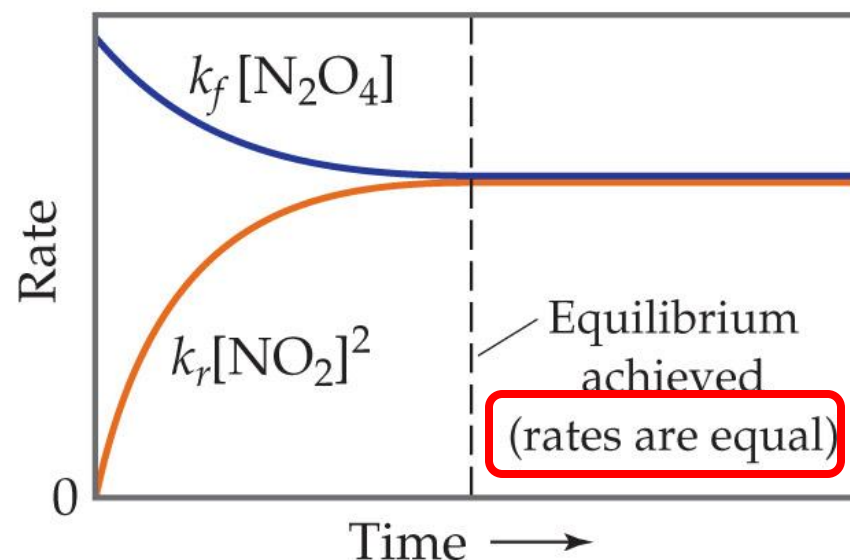
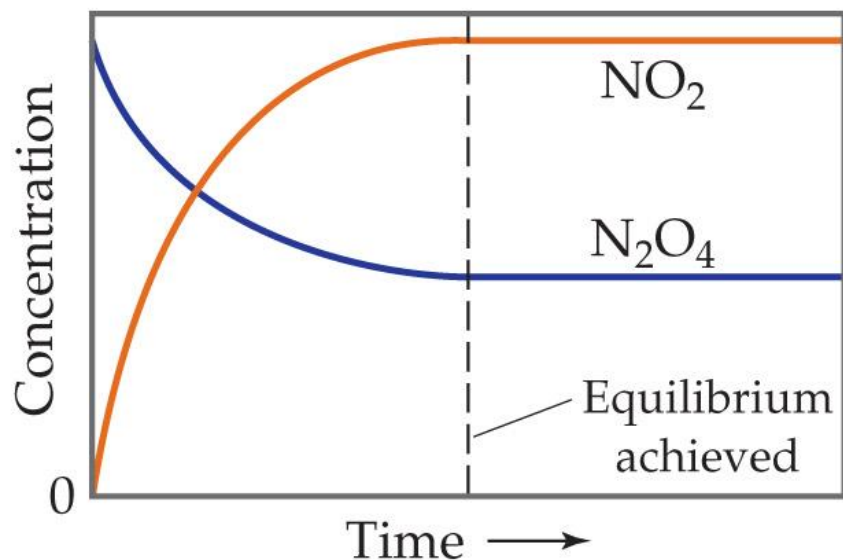
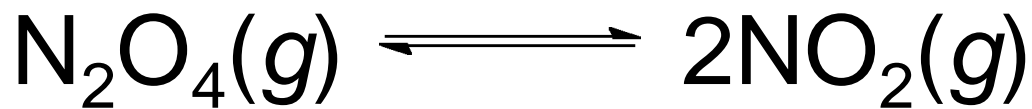
Le Châtelier's Principle (勒夏特列原理)



The Concept of Equilibrium



• **Chemical equilibrium** occurs when a **forward reaction & its reverse reaction** proceed at the **same rate**. Equilibrium is finally reached in the right picture.

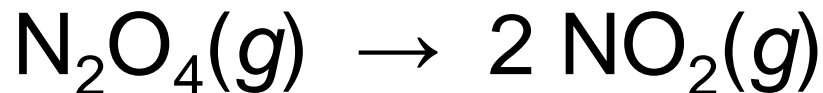


- As a system approaches equilibrium:
 1. both the **forward & reverse reactions** are occurring at the **same rate**.
 2. the **amount/concentration** of each **reactant & product remains constant**.
 3. We write its equation with **a double arrow**.



Comparing Rates

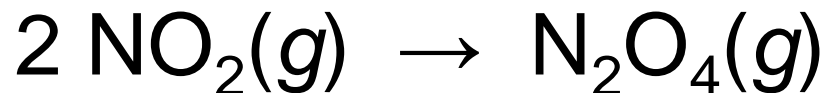
- For the forward reaction:



- The rate law is:

$$\text{Rate}_f = k_f^*[\text{N}_2\text{O}_4]$$

- For the reverse reaction:



- The rate law is:

$$\text{Rate}_r = k_r^*[\text{NO}_2]^2$$



Equilibrium Constant

- At equilibrium:

$$\begin{aligned}\text{Rate}_f &= \text{Rate}_r \\ k_f[\text{N}_2\text{O}_4] &= k_r[\text{NO}_2]^2\end{aligned}$$

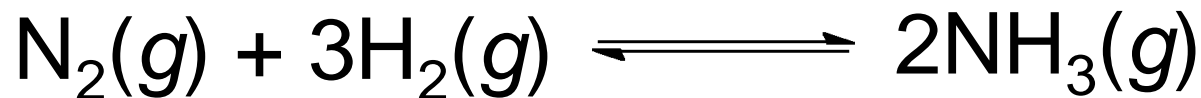
- Rewriting this, it becomes:

$$K_{\text{eq}} = \frac{k_f}{k_r} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

- A ratio of concentration terms equals the **equilibrium constant** (K_{eq}) at **that temperature** (a temperature-dependent thermochemical property).

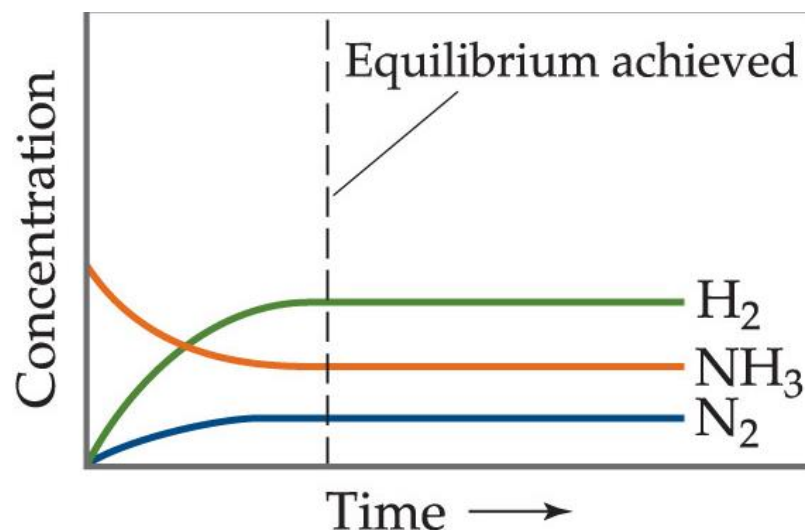
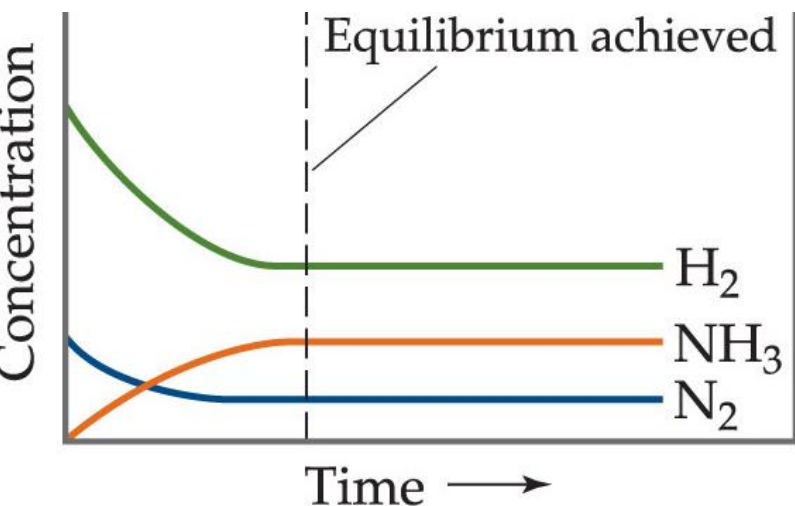


Another Example: The Haber Process



- The **equilibrium constant** depends on **stoichiometry**:

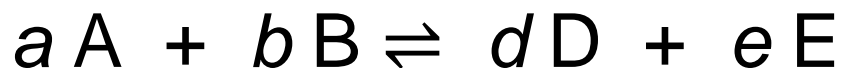
$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$



- Regardless which direction it starts, we can have the **same proportions** of all three substances at equilibrium.

The Equilibrium Constant

- Consider the generalized reaction



- The equilibrium expression for this reaction:

$$K_c = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$

← products
← reactants

- K depends **only on the stoichiometry** of the **overall reaction**, **not its mechanism** (different from rate law).
- As pressure is proportional to concentration for gases in a **closed** system, the equilibrium expression can also be written:

$$K_p = \frac{(P_D)^d (P_E)^e}{(P_A)^a (P_B)^b}$$



Relationship between K_c and K_p

- For gases (the ideal gas Law):

$$PV = nRT$$

- Rearranging,

$$P = (n/V)RT$$

As (n/V) is equal to concentration, thus :

$$K_p = K_c (RT)^{\Delta n}$$

where

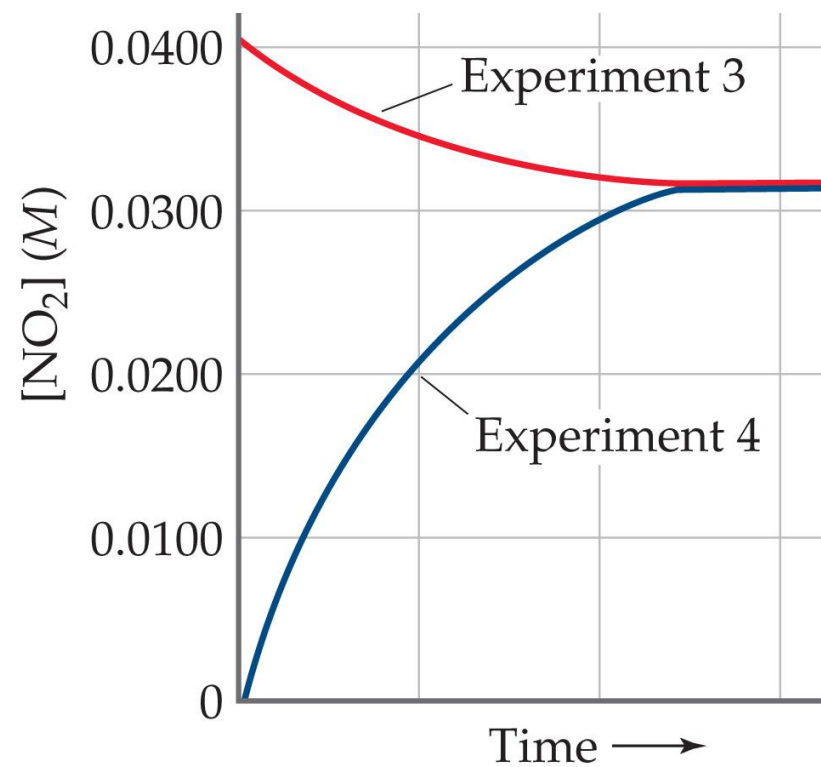
$\Delta n = (\text{moles of gaseous product}) - (\text{moles of gaseous reactant})$

Equilibrium

Table 15.1 Initial and Equilibrium Concentrations of $\text{N}_2\text{O}_4(g)$ and $\text{NO}_2(g)$ at 100°C

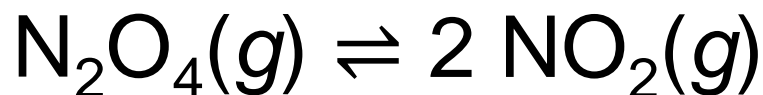
| Experiment | Initial $[\text{N}_2\text{O}_4]$ (M) | Initial $[\text{NO}_2]$ (M) | Equilibrium $[\text{N}_2\text{O}_4]$ (M) | Equilibrium $[\text{NO}_2]$ (M) | K_c |
|------------|--------------------------------------|-----------------------------|--|---------------------------------|-------|
| 1 | 0.0 | 0.0200 | 0.00140 | 0.0172 | 0.211 |
| 2 | 0.0 | 0.0300 | 0.00280 | 0.0243 | 0.211 |
| 3 | 0.0 | 0.0400 | 0.00452 | 0.0310 | 0.213 |
| 4 | 0.0200 | 0.0 | 0.00452 | 0.0310 | 0.213 |

• The **ratio** of $[\text{NO}_2]^2$ to $[\text{N}_2\text{O}_4]$ ($\rightarrow K_c$) **remains constant** at this temperature *no matter what the initial concentrations* of NO_2 and N_2O_4 are.

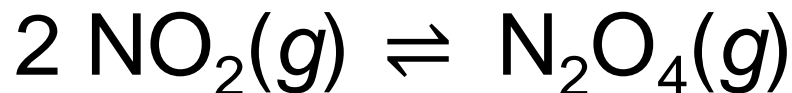


Direction of the Chemical Equation & K

- The equilibrium constant of a reaction in the **reverse reaction** is the **reciprocal** of the equilibrium constant of the **forward** reaction: $K_c = 1/K_{c^*}$



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.212 \text{ at } 100^\circ\text{C}$$

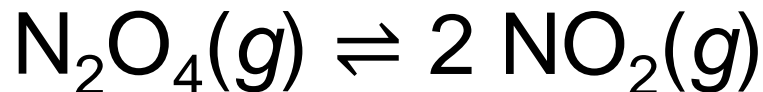


$$K_{c^*} = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} = 4.72 \text{ at } 100^\circ\text{C}$$

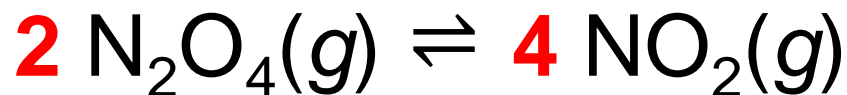
Equilibrium

Stoichiometry & Equilibrium Constants

- When the equation is multiplied by a number (e.g. 2 in the below example), we simply raise the original equilibrium constant to that power. As the **stoichiometry** is **doubled**; the **constant** is the **squared**! *We should always write the balanced chemical reaction!*



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.212 \text{ at } 100^\circ\text{C}$$



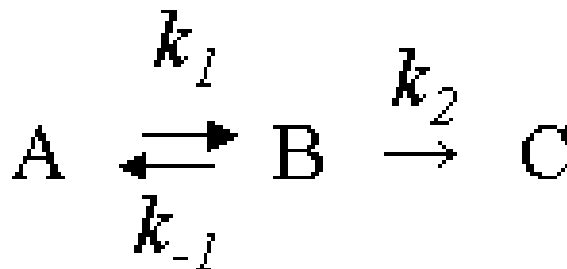
$$K_c = \frac{[\text{NO}_2]^4}{[\text{N}_2\text{O}_4]^2} = (0.212)^2 \text{ at } 100^\circ\text{C}$$

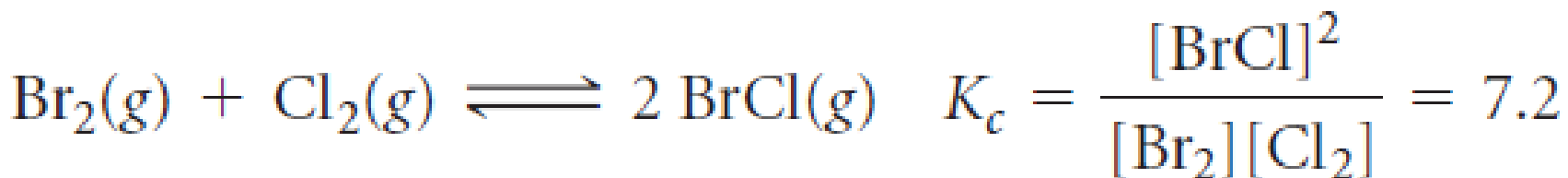
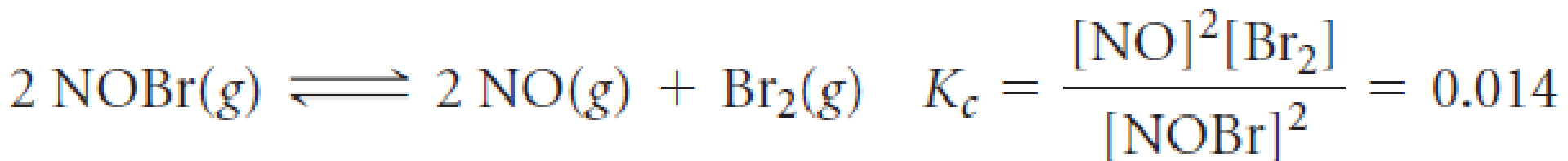


Consecutive Equilibria

- When two or more **consecutive equilibria** occur, the equations can be **combined** to give a single equilibrium.
- The equilibrium constant of the **overall reaction** is the **product** of the constants of the **two reactions/steps** (K_{c1} & K_{c2}):

$$K_c = K_{c1} \times K_{c2}$$





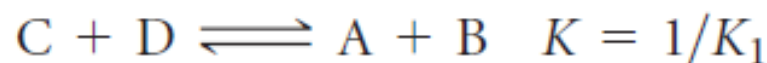
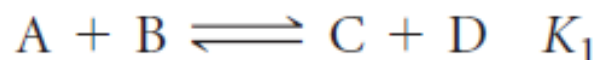
$$K_c = \frac{[\text{NO}]^2[\text{BrCl}]^2}{[\text{NOBr}]^2[\text{Cl}_2]} = \frac{[\text{NO}]^2[\text{Br}_2]}{[\text{NOBr}]^2} \times \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]}$$

$$K_{c3} = (K_{c1})(K_{c2}) = (0.014)(7.2) = 0.10$$

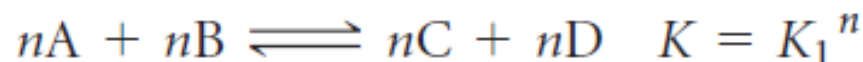
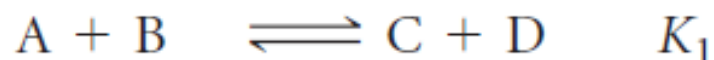
ibrium

To summarize:

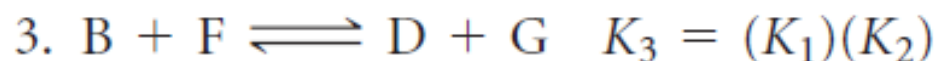
1. The equilibrium constant of a reaction in the *reverse* direction is the *inverse* (or *reciprocal*) of the equilibrium constant of the reaction in the forward direction:



2. The equilibrium constant of a reaction that has been *multiplied* by a number is equal to the original equilibrium constant raised to a *power* equal to that number.



3. The equilibrium constant for a net reaction made up of *two or more reactions* is the *product* of the equilibrium constants for the individual reactions:



Equilibrium Constants and Units

- Equilibrium constants are reported **without units**.
- Equilibrium constants derived from thermodynamic measurements are defined in terms of **activities** (a ; 活度) rather than concentrations or partial pressures.

$$K = \left(\prod_j a_j^{\nu_j} \right)$$

- molalities, by replacing a_j by b_j/b^\ominus , where $b^\ominus = 1 \text{ mol kg}^{-1}$
- molar concentrations, by replacing a_j by $[J]/c^\ominus$, where $c^\ominus = 1 \text{ mol dm}^{-3}$
- partial pressures, by replacing a_i by p_i/p^\ominus , where $p^\ominus = 1 \text{ bar}$

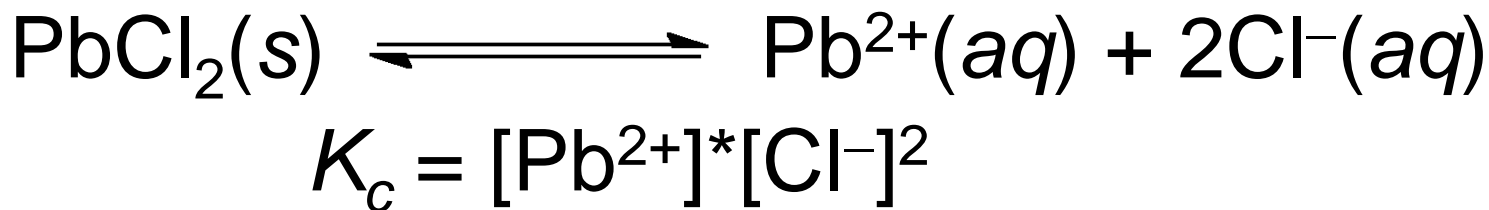
- The **activity** of any substance in an ideal mixture is the **ratio** of the **concentration** or **pressure** of the substance to a **reference concentration or pressure** (1 M or 1 atm). Therefore, activities have **no unit**.

- If the concentration of a substance in an equilibrium mixture is 0.01 M, its activity is $0.01 \text{ M} / 1 \text{ M} = 0.01$.



Homogeneous vs. Heterogeneous

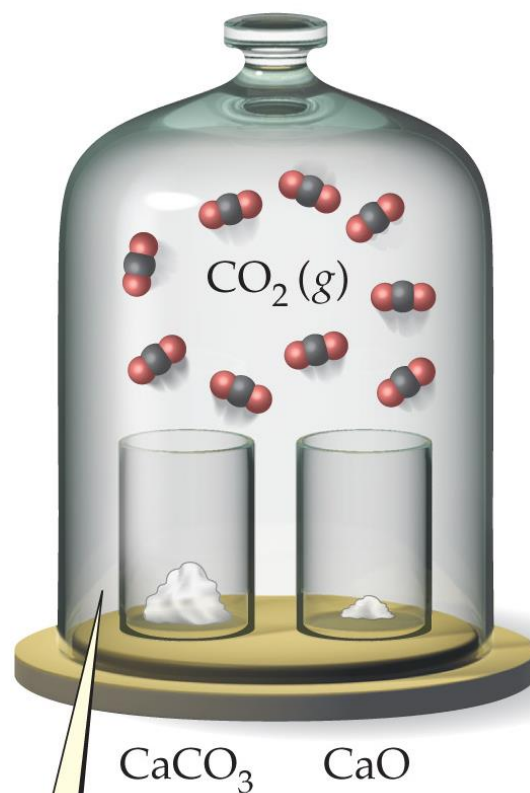
- **Homogeneous equilibria** occur when all reactants and products are in the **same phase**.
- **Heterogeneous equilibria** occur when something in the equilibrium is in a **different phase**.
- The **activity value** used for a **pure solid or liquid** is always **1**.



The Decomposition of CaCO_3 : A Heterogeneous Equilibrium



- $K_c = [\text{CO}_2]$
- $K_p = P_{\text{CO}_2}$
- The amount of CO_2 above the solid remain the same.

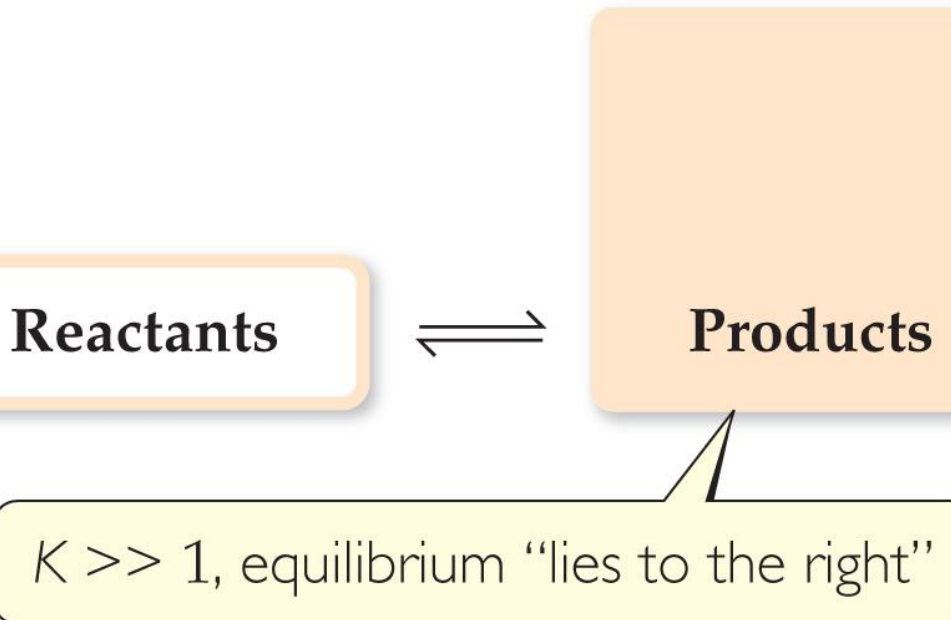


Large amount of CaCO_3 , small amount of CaO , gas pressure P



Small amount of CaCO_3 , large amount of CaO , gas pressure still P

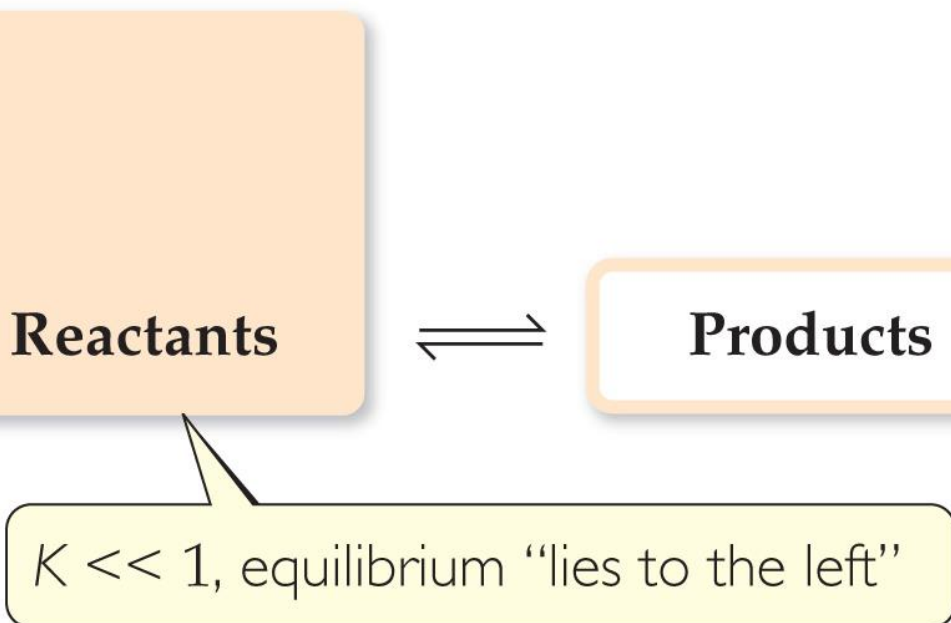
Magnitude of K



$$K_c = \frac{[D]^d[E]^e}{[A]^a[B]^b}$$

← products
← reactants

- If $K \gg 1$, the reaction favors products; **products predominate** at equilibrium.



- If $K \ll 1$, the reaction favors reactants; **reactants predominate** at equilibrium.

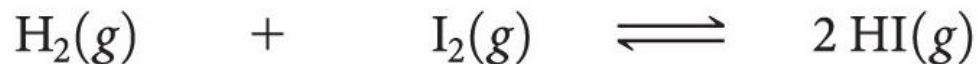
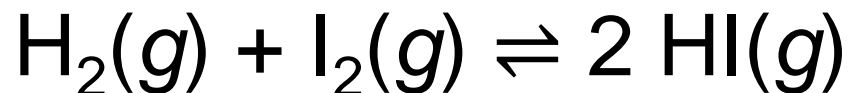
Calculating Equilibrium Constant

- 1) If any initial and equilibrium concentrations are known, **calculate the change**.
- 2) Use the **balanced equation** and **initial concentrations** to find ***change/equilibrium concentration for all reactants and products***.
- 3) Calculate the equilibrium constant using the **equilibrium concentrations**.



An Example

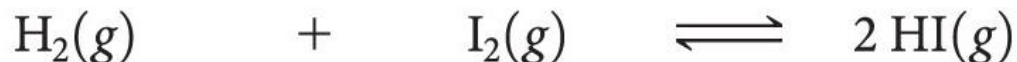
A closed system initially containing $1.000 \times 10^{-3} \text{ M H}_2$ and $2.000 \times 10^{-3} \text{ M I}_2$ at 448°C is allowed to reach equilibrium. Analysis of the equilibrium mixture shows that the concentration of HI is $1.87 \times 10^{-3} \text{ M}$. Calculate K_c at 448°C for the reaction.



| | | | |
|-------------------------------|------------------------|------------------------|-----------------------|
| Initial concentration (M) | 1.000×10^{-3} | 2.000×10^{-3} | 0 |
| Change in concentration (M) | | | |
| Equilibrium concentration (M) | | | 1.87×10^{-3} |

(If any initial and equilibrium concentrations are known, calculate the change.)





| | | | |
|-------------------------------|------------------------|------------------------|------------------------|
| Initial concentration (M) | 1.000×10^{-3} | 2.000×10^{-3} | 0 |
| Change in concentration (M) | | | $+1.87 \times 10^{-3}$ |
| Equilibrium concentration (M) | | | 1.87×10^{-3} |

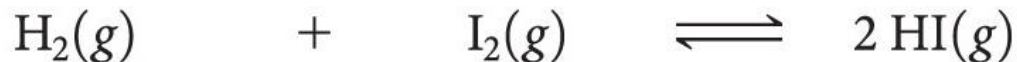
- The change in concentration of HI is $1.87 \times 10^{-3} \text{ M}$.

| | | | |
|-------------------------------|-------------------------|-------------------------|------------------------|
| Initial concentration (M) | 1.000×10^{-3} | 2.000×10^{-3} | 0 |
| Change in concentration (M) | -0.935×10^{-3} | -0.935×10^{-3} | $+1.87 \times 10^{-3}$ |
| Equilibrium concentration (M) | | | 1.87×10^{-3} |

- Stoichiometry tells us $[\text{H}_2]$ and $[\text{I}_2]$ decrease by half.

(Use the balanced equation and initial concentrations to find change/equilibrium concentration for all reactants and products.)





| | | | |
|-------------------------------|-------------------------|-------------------------|------------------------|
| Initial concentration (M) | 1.000×10^{-3} | 2.000×10^{-3} | 0 |
| Change in concentration (M) | -0.935×10^{-3} | -0.935×10^{-3} | $+1.87 \times 10^{-3}$ |
| Equilibrium concentration (M) | 0.065×10^{-3} | 1.065×10^{-3} | 1.87×10^{-3} |

- We can now calculate the equilibrium concentrations of H_2 & I_2 compounds.
- The equilibrium constant becomes:

$$\begin{aligned}
 K_c &= \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \\
 &= \frac{(1.87 \times 10^{-3})^2}{(6.5 \times 10^{-5})(1.065 \times 10^{-3})} \\
 &= 51
 \end{aligned}$$

(Calculate the equilibrium constant using the equilibrium concentrations.)



Is a Mixture in Equilibrium? Which Way Does the Reaction Go?

- We need to calculate the **reaction quotient**, Q .
- Q looks like the equilibrium constant (K), but the **values used** to calculate it are the **current conditions**, not necessarily those for equilibrium.
- To calculate Q , one substitutes the **initial (or current) concentrations of reactants and products** into the equilibrium expression.



$$Q_c = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$



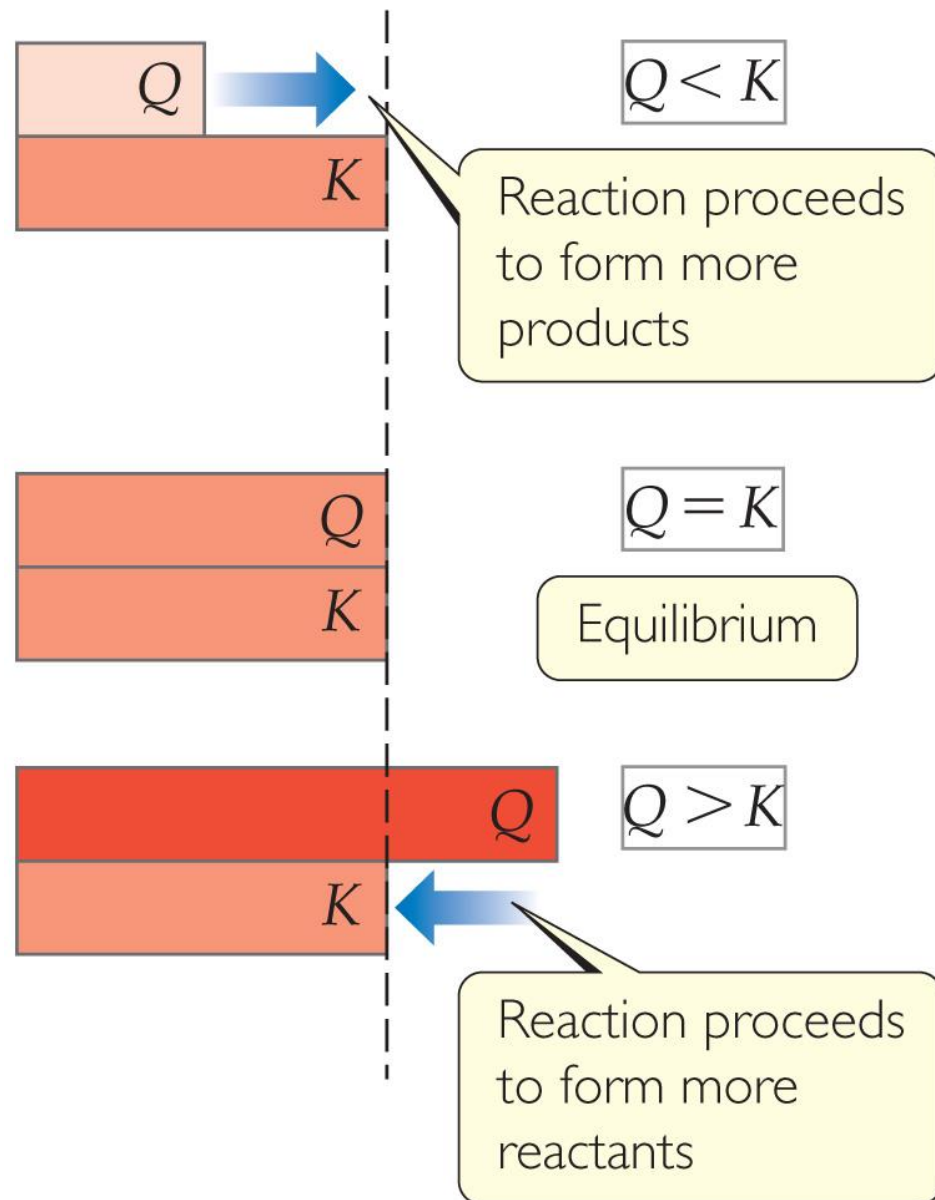
Comparing Q and K

At equilibrium

- If $Q < K$, nature will make the reaction **proceed to products**.

- If $Q = K$, the reaction is **in equilibrium**.

- If $Q > K$, nature will make the reaction **proceed to reactants**.



Calculating Equilibrium Concentrations

- If we know the equilibrium constant, we can *deduce equilibrium concentrations* from the known initial concentrations and changes (based on stoichiometry).

A 1.000 L flask is filled with 1.000 mol of $\text{H}_2(\text{g})$ and 2.000 mol of $\text{I}_2(\text{g})$ at 448 °C. Given a K_c of 50.5, what are the equilibrium concentrations of H_2 , I_2 , and HI ?

- Set up a below table & *set* the “change in concentration” row to be a factor of “*x*” based on the stoichiometry.



| | | | |
|-----------------------------------|-------|-------|-------|
| Initial concentration (M) | 1.000 | 2.000 | 0 |
| Change in concentration (M) | $-x$ | $-x$ | $+2x$ |
| Equilibrium concentration (M) | | | |



| | | | |
|--|-------------|-------------|-------|
| Initial concentration (<i>M</i>) | 1.000 | 2.000 | 0 |
| Change in concentration (<i>M</i>) | $-x$ | $-x$ | $+2x$ |
| Equilibrium concentration (<i>M</i>) | $1.000 - x$ | $2.000 - x$ | $2x$ |

- The equilibrium concentration = the initial concentration minus the “change in concentration”.
- Set up the equilibrium constant expression:

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$



| | | | | | |
|-------------------------------|------------------------|---|------------------------|----------------------|-------------------------|
| | $\text{H}_2(\text{g})$ | + | $\text{I}_2(\text{g})$ | \rightleftharpoons | $2 \text{HI}(\text{g})$ |
| Initial concentration (M) | 1.000 | | 2.000 | | 0 |
| Change in concentration (M) | $-x$ | | $-x$ | | $+2x$ |
| Equilibrium concentration (M) | $1.000 - x$ | | $2.000 - x$ | | $2x$ |

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(2x)^2}{(1.000 - x)(2.000 - x)} = 50.5$$

- Solving the above equation by using the quadratic formula, we get $x = 2.323$ or 0.935 .
- However, x *must* be 0.935 .
- So $[\text{H}_2]_{\text{eq}} = 1.000 - 0.935 = 0.065 \text{ M}$;
 $[\text{I}_2]_{\text{eq}} = 2.000 - 0.935 = 1.065 \text{ M}$;
 $[\text{HI}]_{\text{eq}} = 2(0.935) = 1.87 \text{ M}$



Le Châtelier's Principle (勒夏特列原理)

*“If a system at equilibrium is disturbed by a change in temperature, pressure, or the concentration of one of the components, **the system will shift its equilibrium position so as to counteract the effect of the disturbance.**”*

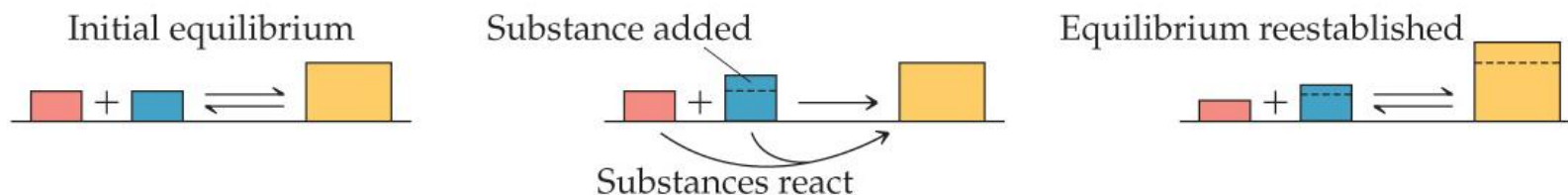


Henri-Louis Le Châtelier (1850–1936)

If a system at equilibrium is disturbed by a change in **concentration, pressure, or temperature**, the system will shift its equilibrium position so as to counter the effect of the disturbance.

Concentration: adding or removing a reactant or product

If a substance is added to a system at equilibrium, the system reacts to consume some of the substance. If a substance is removed from a system, the system reacts to produce more of substance.



Pressure: changing the pressure by changing the volume

At constant temperature, reducing the volume of a gaseous equilibrium mixture causes the system to shift in the direction that reduces the number of moles of gas.



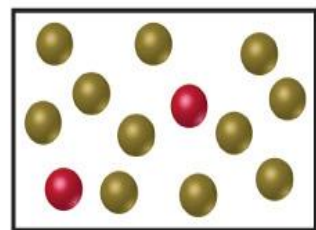
Temperature:

If the temperature of a system at equilibrium is increased, the system reacts as if we added a reactant to an endothermic reaction or a product to an exothermic reaction. The equilibrium shifts in the direction that consumes the "excess reactant," namely heat.



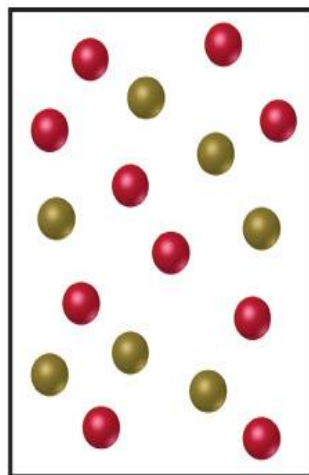
• Use Le Châtelier's Principle qualitatively to **predict shifts in equilibrium** based on **changes in conditions**.

Change in Volume or Pressure



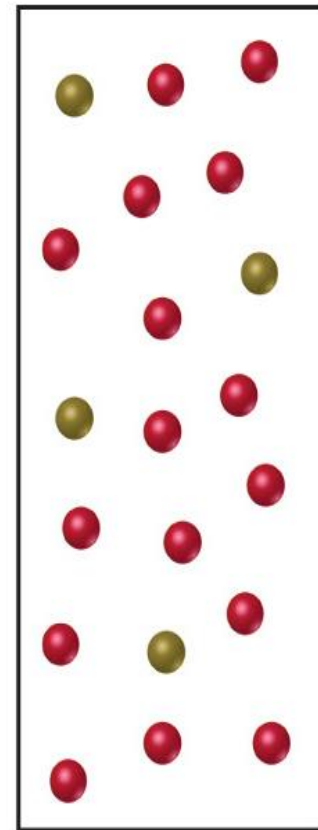
New equilibrium favors products to reduce total moles of gas

Decrease volume,
increase pressure

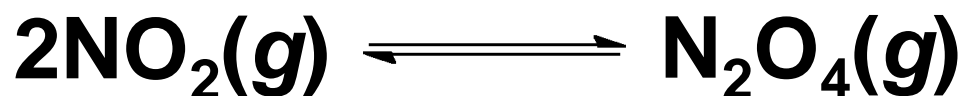


Initial volume

Increase volume,
decrease pressure



New equilibrium favors reactants to increase total moles of gas

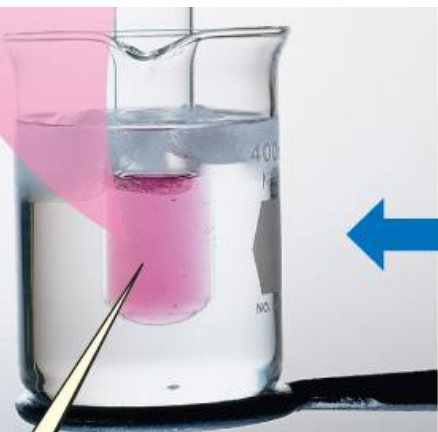
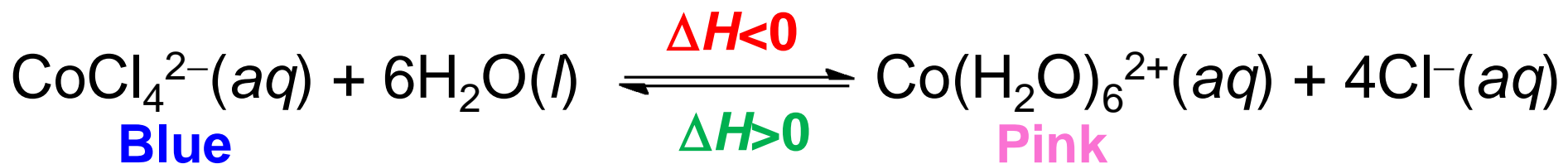


- A change in pressure or volume of gases affects equilibrium: **Higher volume or lower pressure favors** the side of the equation with **more moles** → **more pressure** (and *vice-versa*).

Equilibrium

Change in Temperature

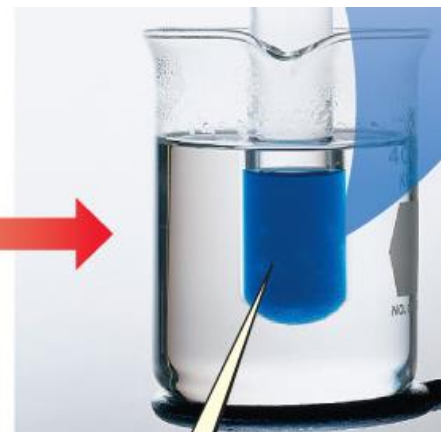
- Does the reaction endothermic or exothermic affect equilibrium?
- Endothermic**: Heat acts *like* a reactant; **adding heat** drives a reaction **toward products**.
- Exothermic**: Heat acts *like* a product; **adding heat** drives a reaction **toward reactants**.



Cool

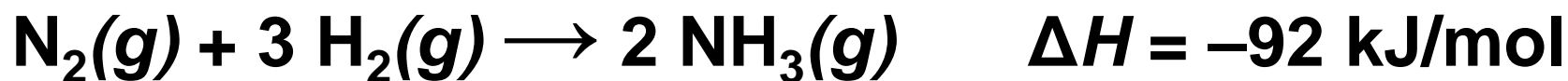


Heat



| System | Change | Result |
|---|--|---|
| $\text{CO}_2 + \text{H}_2 \rightarrow \text{H}_2\text{O}(g) + \text{CO}$ | A drying agent is added to absorb H_2O | Shift to the right. Continuous removal of a product will force any reaction to the |
| $\text{H}_2(g) + \text{I}_2(g) \rightarrow 2\text{HI}(g)$ | Some nitrogen gas is added | No change; N_2 is not a component of this reaction system. |
| $\text{NaCl}(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{Na}_2\text{SO}_4(s) + \text{HCl}(g)$ | Reaction is carried out in an open container | HCl gas can escape from the system, the reaction is forced to the right : the basis for the commercial production of HCl. |
| $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ | Water evaporates from an open container | Continuous removal of water vapor forces the reaction to the right , so equilibrium is never achieved. |
| $\text{HCN}(aq) \rightarrow \text{H}^+(aq) + \text{CN}^-(aq)$ | The solution is diluted | Shift to right ; the product $[\text{H}^+][\text{CN}^-]$ diminishes more rapidly than does $[\text{HCN}]$. |
| $\text{AgCl}(s) \rightarrow \text{Ag}^+(aq) + \text{Cl}^-(aq)$ | Some NaCl is added to the solution | Shift to left due to increase in $[\text{Cl}^-]$: the common ion effect on solubility. |
| $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$ | A catalyst is added to speed up this reaction | No change. Catalysts affect only the rate of a reaction; they have no effect at all on the composition of the equilibrium state. |

The Haber Process-Changes in Temperature



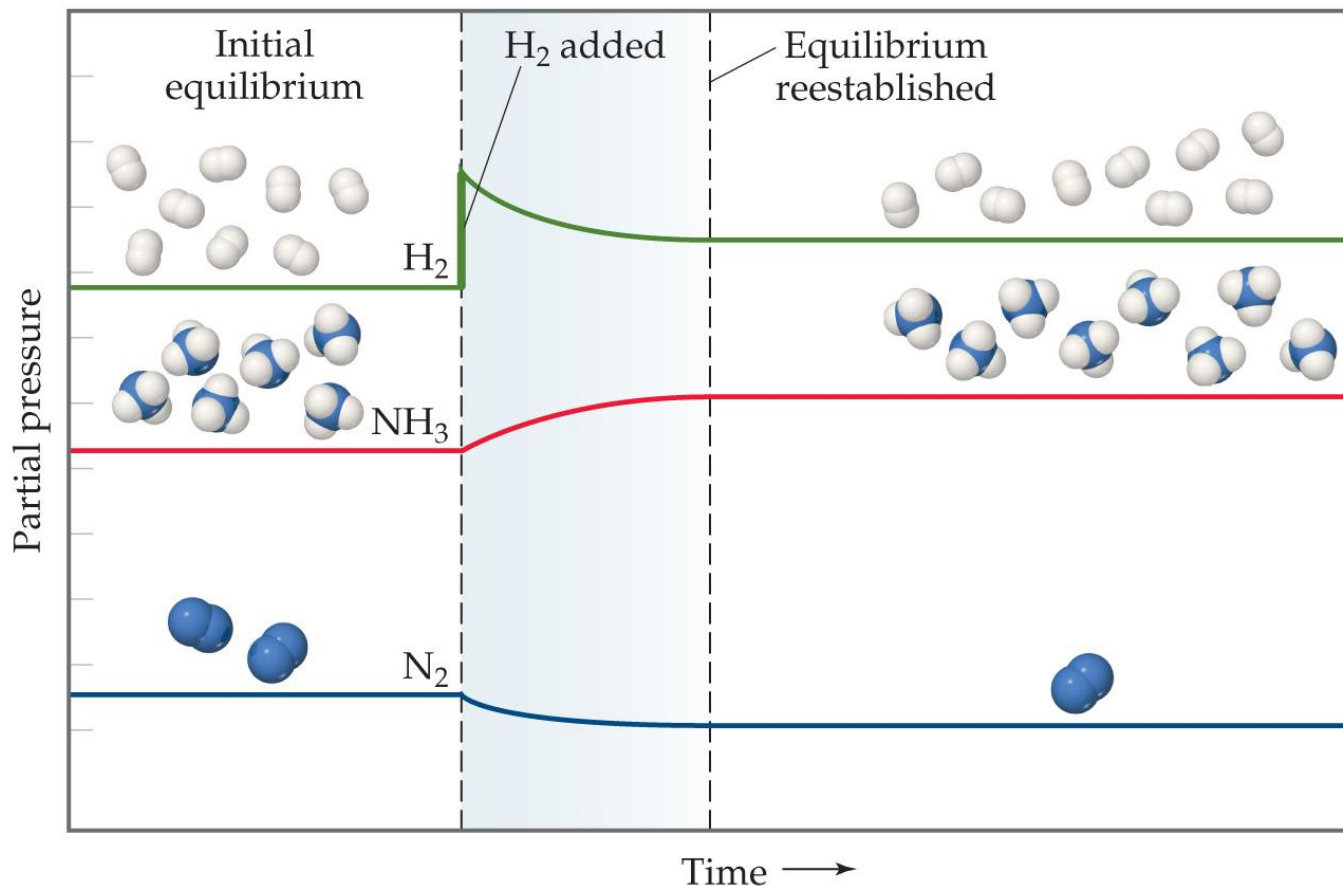
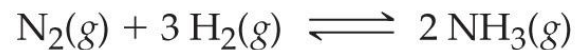
Temperature (°C)

K_p

| | |
|-----|-----------------------|
| 300 | 4.34×10^{-3} |
| 400 | 1.64×10^{-4} |
| 450 | 4.51×10^{-5} |
| 500 | 1.45×10^{-5} |
| 550 | 5.38×10^{-6} |
| 600 | 2.25×10^{-6} |

Equilibrium

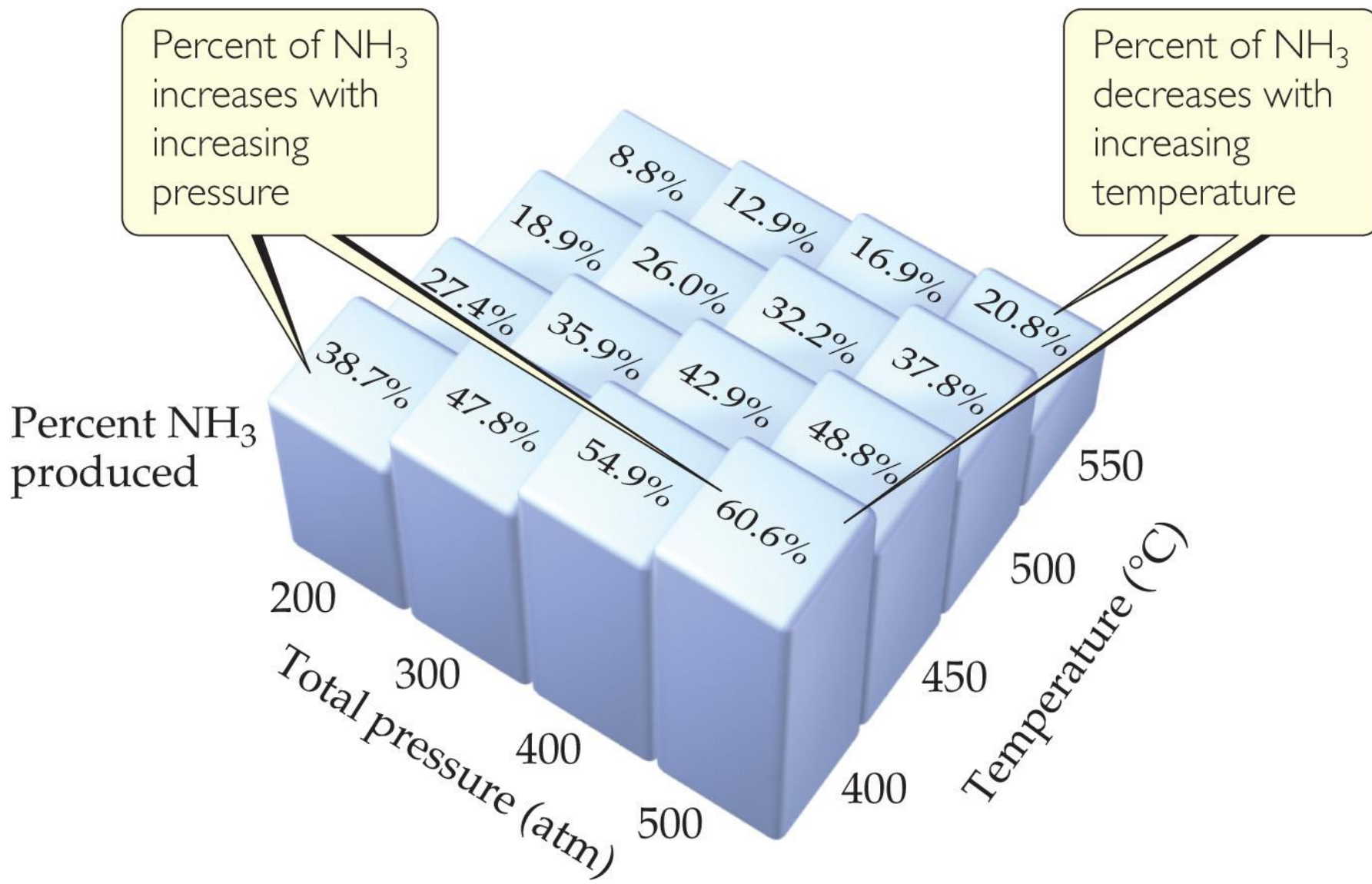
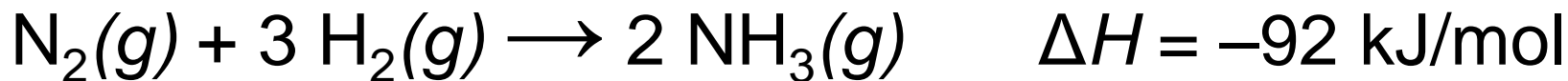
Change in Reactant or Product Concentration

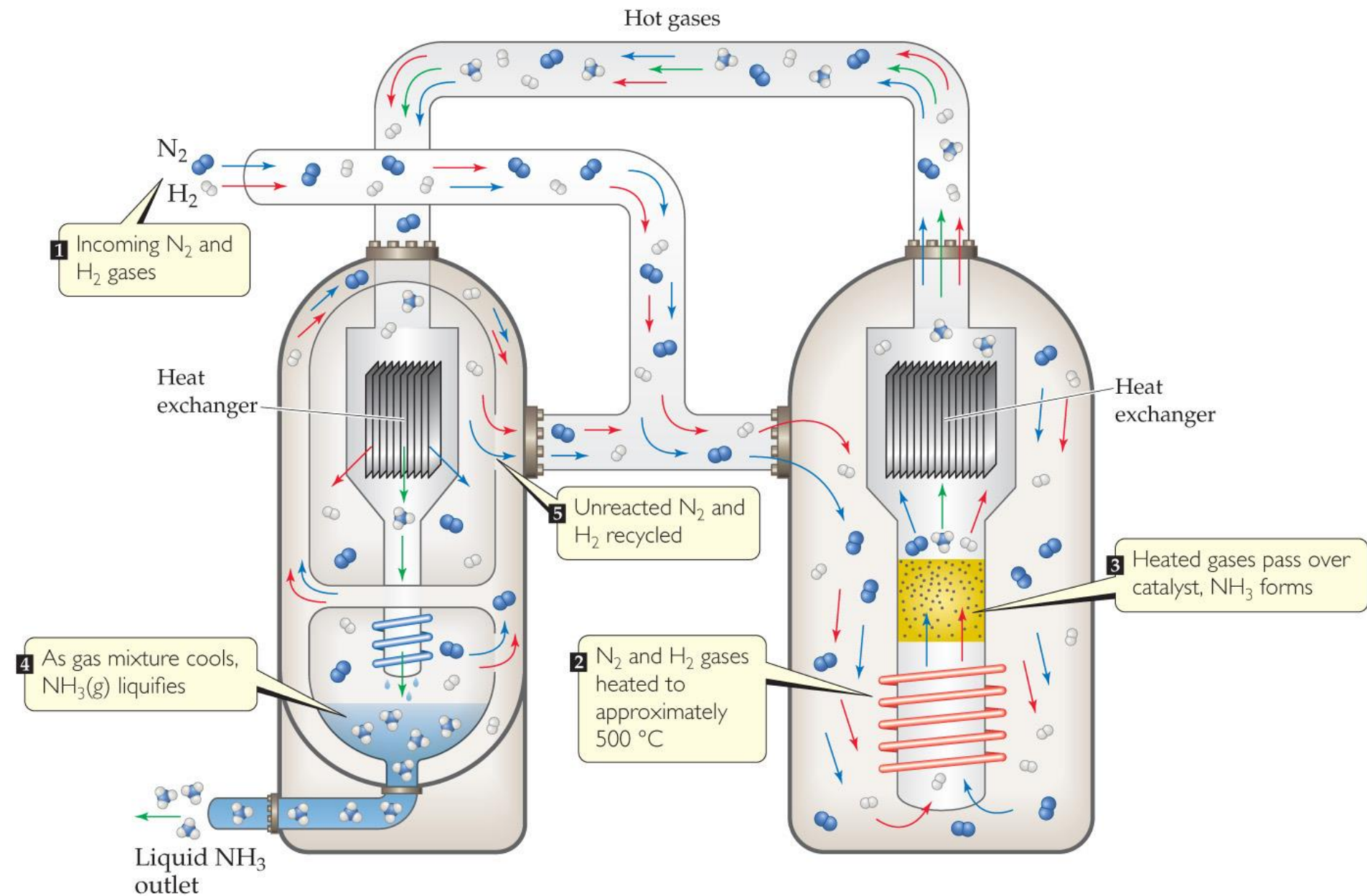


- When adding more reactant, that are used up more.
- If H₂ is added to the system, N₂ will be consumed & the two reagents will form more NH₃.

Equilibrium

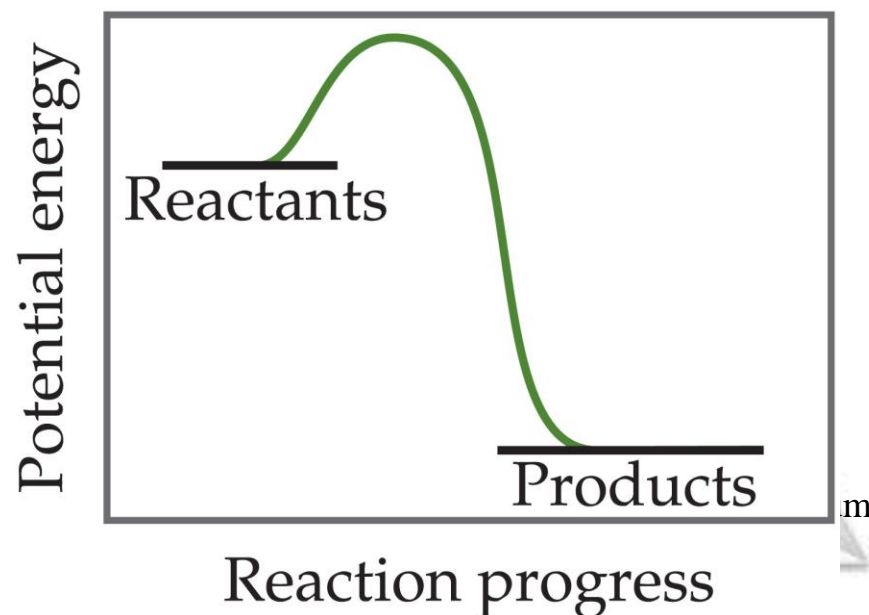
Effect of Temperature & Pressure



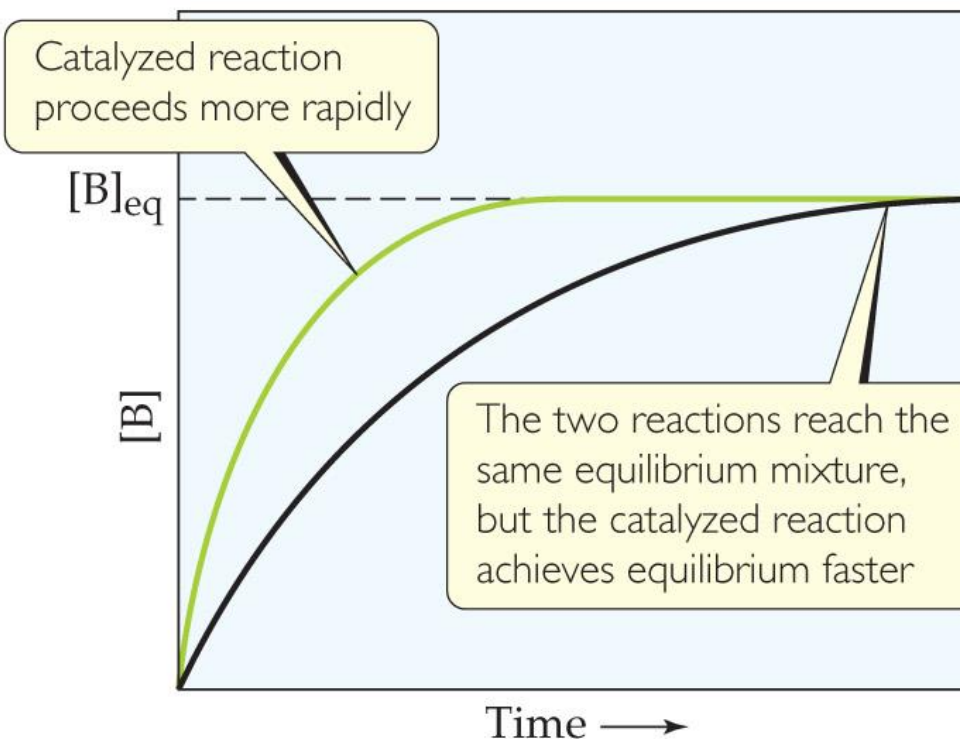
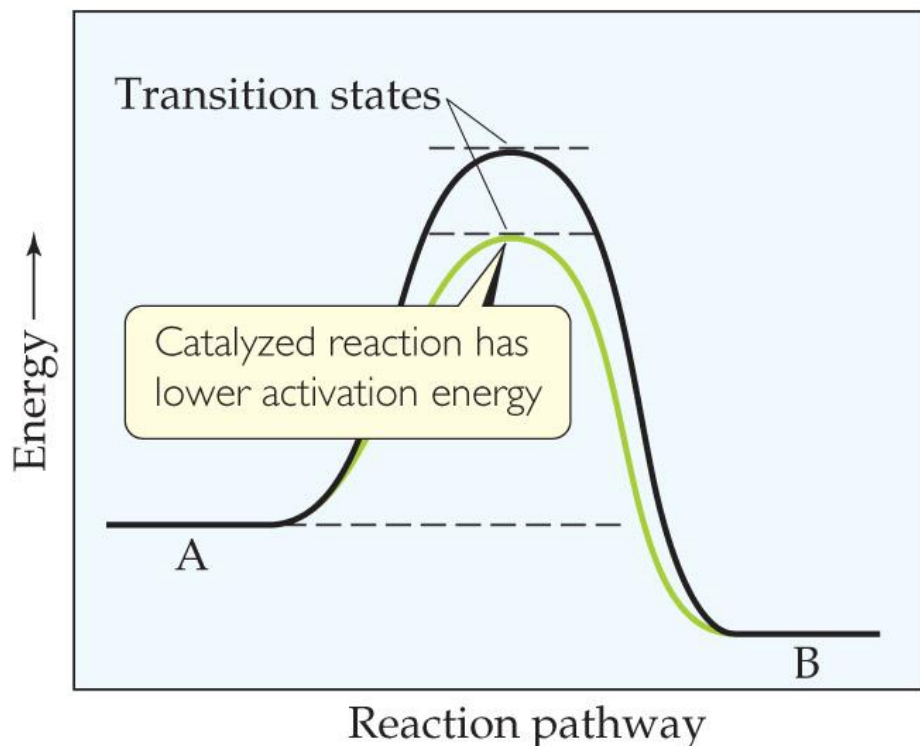


This industrial apparatus helps push the equilibrium to the right by removing NH_3 from the system as a **liquid**.

- The Haber process for producing ammonia from the elements is exothermic.
- One would think that cooling down the reactants would result in more product.
- However, the activation energy for this reaction is high!
- This is the *one* instance where a system in equilibrium can be affected by a catalyst by lowering the activation energy!



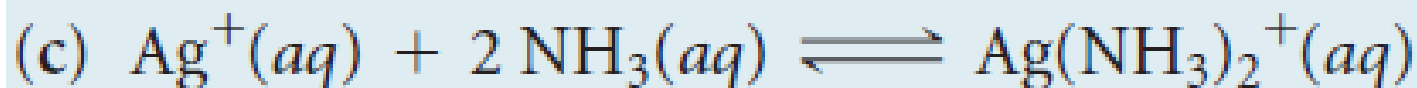
Catalysts



- Catalysts **increase** the **rate** of **both** the **forward & reverse reactions**.
- **Equilibrium** is achieved **faster**, but the equilibrium composition **remains unaltered**.
- Activation energy is lowered, allowing equilibrium to be established at lower temperatures.



Write the equilibrium expression for K_c for the following reactions:



Solve

$$(a) K_c = \frac{[\text{O}_2]^3}{[\text{O}_3]^2}$$

$$(b) K_c = \frac{[\text{NOCl}]^2}{[\text{NO}]^2[\text{Cl}_2]}$$

$$(c) K_c = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2}$$

Write the equilibrium-constant expression for the evaporation of water, $\text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g)$, in terms of partial pressures.

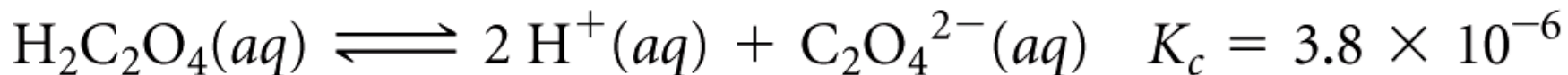
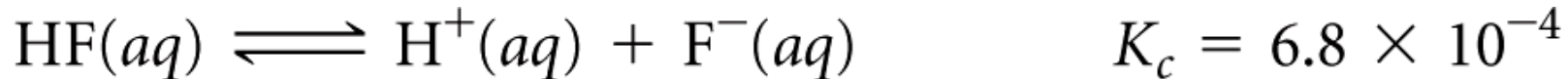
A. $K_p = 1/[\text{H}_2\text{O}(g)]$

B. $K_p = P_{\text{H}_2\text{O}(g)}$

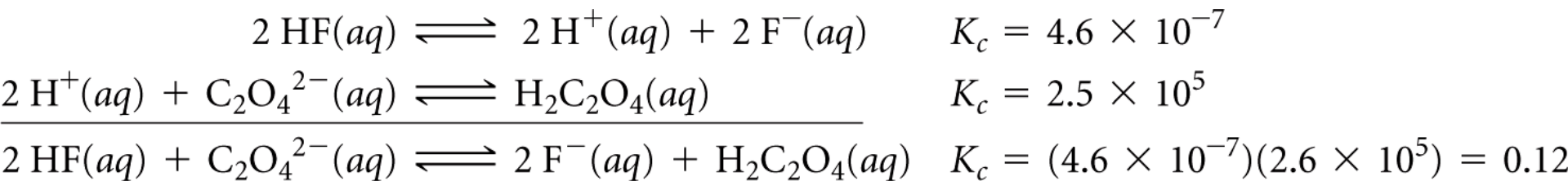
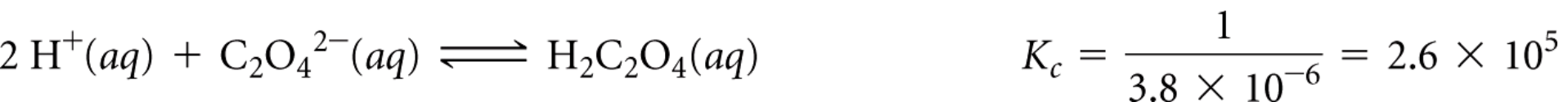
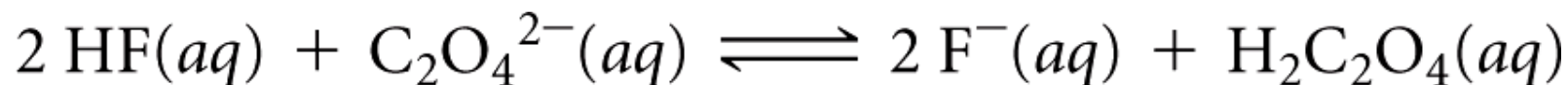
C. $K_p = [\text{H}_2\text{O}(g)]$

D. $K_p = 1/P_{\text{H}_2\text{O}(g)}$





On the basis of the above reactions, determine K_c for the below reaction



Equilibrium

Write the equilibrium-constant expression for the reaction $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$

A. $K_c = [\text{NH}_4^+][\text{OH}^-]$

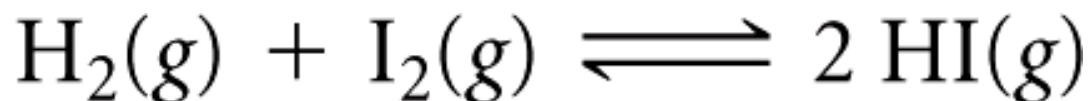
B. $K_c = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$

C. $K_c = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3][\text{H}_2\text{O}]}$

D. $K_c = \frac{[\text{NH}_3][\text{H}_2\text{O}]}{[\text{NH}_4^+][\text{OH}^-]}$



At 448 °C the equilibrium constant K_c for the below reaction is 50.5. Predict in which direction the reaction proceeds to reach equilibrium if we start with 2.0×10^{-2} mol of HI, 1.0×10^{-2} mol of H_2 , and 3.0×10^{-2} of I_2 in a 2.00-L container.



The initial concentrations are

$$[HI] = 2.0 \times 10^{-2} \text{ mol}/2.00 \text{ L} = 1.0 \times 10^{-2} M$$

$$[H_2] = 1.0 \times 10^{-2} \text{ mol}/2.00 \text{ L} = 5.0 \times 10^{-3} M$$

$$[I_2] = 3.0 \times 10^{-2} \text{ mol}/2.00 \text{ L} = 1.5 \times 10^{-2} M$$

The reaction quotient is

$$Q_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.0 \times 10^{-2})^2}{(5.0 \times 10^{-3})(1.5 \times 10^{-2})} = 1.3$$

Because $Q_c < K_c$, the concentration of HI must increase and the concentrations of H_2 and I_2 must decrease to reach equilibrium; the reaction as written proceeds left to right to attain equilibrium.

Use Le Châtelier's principle to explain why the equilibrium vapor pressure of a liquid increases with increasing temperature.

- A. Evaporation is an exothermic process and an increasing temperature shifts the equilibrium $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{O(g)}$ to the right and increases the equilibrium vapor pressure.
- B.

 Evaporation is an endothermic process and an increasing temperature shifts the equilibrium $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{O(g)}$ to the right and increases the equilibrium vapor pressure.
- C. Evaporation is an exothermic process and an increasing temperature shifts the equilibrium $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{O(g)}$ to the left and increases the equilibrium vapor pressure.
- D. Evaporation is an endothermic process and an increasing temperature shifts the equilibrium $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{O(g)}$ to the left and increases the equilibrium vapor pressure.



At equilibrium, the rate of the forward reaction is _____ the rate of the reverse reaction.

- ☒ a. equal to
- ☐ b. slower than
- ☐ c. faster than
- ☐ d. the reverse of

The reaction quotient Q is usually represented by

- ☐ a. $[\text{reactants}] / [\text{products}]$.
- ☒ b. $[\text{products}] / [\text{reactants}]$.
- ☐ c. $[\text{reactants}] \times [\text{products}]$.
- ☐ d. $[\text{reactants}] + [\text{products}]$.



If the value of the equilibrium constant is large, then _____ will mostly be present at equilibrium.

- a. reactants
- ☒ b. products
- c. catalysts
- d. shrapnel

If the value of the equilibrium constant is small, then _____ will mostly be present at equilibrium.

- ☒ a. reactants
- b. products
- c. catalysts
- d. shrapnel



Q = the reaction quotient; K = the equilibrium constant
At equilibrium, which is true?

- a. $Q > K$
- b. $Q < K$
- ☒ c. $Q = K$
- d. $Q^2 = K$

Equilibrium constants typically have units of

- a. M.
- b. M^2 .
- c. M^{1-} .
- ☒ d. None of the above



Reaction quotients for heterogeneous equilibria do not include concentrations of

- a. pure liquids.
- b. pure solids.
- ☒ c. Both of the above
- d. Neither of the above



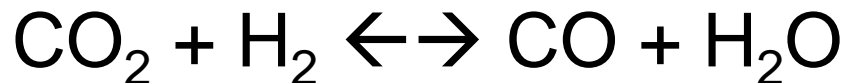
$[\text{HA}] = 1.65 \times 10^{-2} \text{ M}$ and $[\text{H}^+] = [\text{A}^-] = 5.44 \times 10^{-4} \text{ M}$ at equilibrium. $K_c = \underline{\hspace{2cm}}$.

- a. 1.79×10^{-2}
- b. 1.79×10^{-3}
- c. 1.79×10^{-4}
- ☒ d. 1.79×10^{-5}



$K_P = K_C$ when

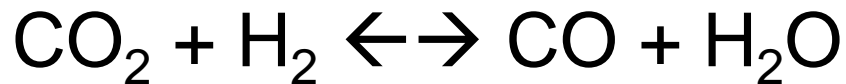
- a. the reaction is at equilibrium.
- b. the reaction is exothermic.
- c. all of the gases present are at the same pressure.
- ☒ d. the number of moles of gas on both sides of the balanced equation is the same.



If all species are gases and H_2 is added, the concentration of CO at equilibrium will

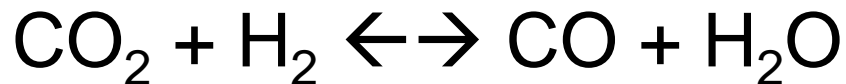
- ☒ a. increase.
- b. decrease.
- c. remain unchanged.
- d. disappear.





If all species are gases and H_2O is added, the equilibrium concentration of CO

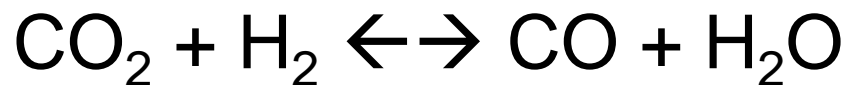
- a. will increase.
- ☒ b. will decrease.
- c. will remain unchanged.
- d. will disappear.



If all species are gases and CO_2 is removed, the $[\text{CO}]$ at equilibrium will

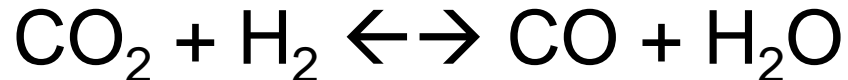
- a. increase.
- ☒ b. decrease.
- c. remain unchanged.
- d. disappear.





Increasing the temperature of this endothermic reaction will _____ [CO] at equilibrium.

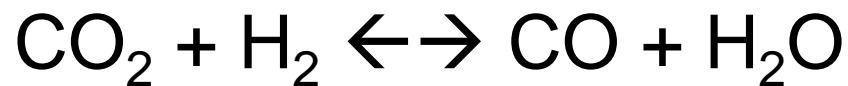
- ☒ a. increase
- ☐ b. decrease
- ☐ c. not change
- ☐ d. eradicate



If all species are gases and the container is compressed, the amount of CO will

- ☐ a. increase.
- ☐ b. decrease.
- ☒ c. remain unchanged.
- ☐ d. vanish.





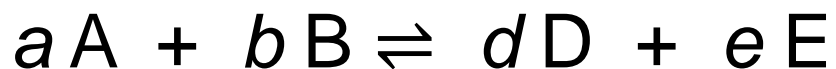
Adding a catalyst to this reaction will cause the [CO] at equilibrium to

- a. increase.
- b. decrease.
- ☒ c. remain unchanged.
- d. cease to exist.



Summary of Chapter 15

Equilibrium constant (K),
Homogeneous equilibria, Heterogeneous equilibria



$$K_c = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$

Reaction quotient

Le Châtelier's Principle



**Thank You for Your
Attention!
Any Questions?**

