

**Principles of Chemical Reactivity: Equilibria**  
Chapter 13

Get the CH 223 Companion before lab!

Chemistry 223  
Professor Michael Russell  
<http://mhchem.org/223>

Last update: 4/29/24

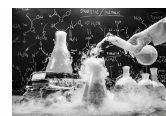
## CH 223: Lectures and Labs

Lectures: **MWF from 9 - 9:50 AM in AC 1303 (this room)**

- Lectures recorded, available soon afterwards
- Lecture notes to print available (under "Problem Sets and Handouts", [mhchem.org/223](http://mhchem.org/223)), get **CH 223 Companion** as soon as possible

Labs (Section 01): **Mondays from 1:10 - 5 PM**

- Start in room **AC 2501** ("the recitation")
- Move to **AC 2507** ("the lab") around **3 PM**
- For first day, bring a printed copy of the "Determination of an Equilibrium Constant (in class)" Lab ([mhchem.org/223](http://mhchem.org/223)), a pair of safety glasses (Dollar store ok) and your calculator



MAR

...more on Monday afternoon



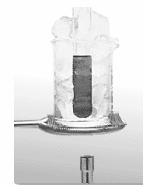
## CHEMICAL EQUILIBRIUM

$\rightleftharpoons$  = Bi-directional arrow

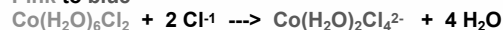
MAR

## Properties of an Equilibrium

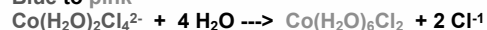
- Equilibrium systems are
- DYNAMIC** (in constant motion)
  - REVERSIBLE**
  - can be approached from either direction



Pink to blue

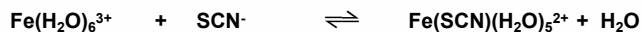
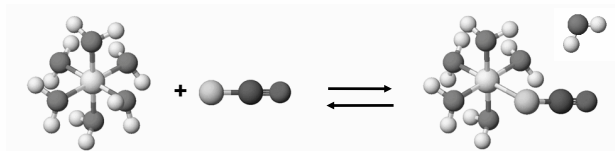


Blue to pink



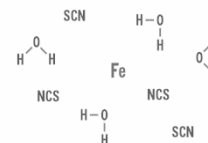
MAR

## Chemical Equilibrium

$$\text{Fe}^{3+} + \text{SCN}^{-} \rightleftharpoons \text{FeSCN}^{2+}$$


MAR

## Chemical Equilibrium

$$\text{Fe}^{3+} + \text{SCN}^{-} \rightleftharpoons \text{FeSCN}^{2+}$$


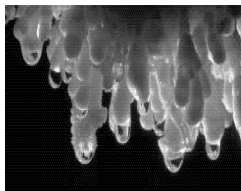
After a period of time, the concentrations of reactants and products are constant. The forward and reverse reactions continue after equilibrium is attained.

MAR

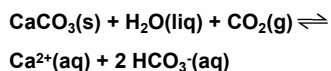
**Phase changes:**  
 $\text{H}_2\text{O(s)} \rightleftharpoons \text{H}_2\text{O(liq)}$



**Examples of Chemical Equilibria**



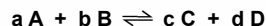
**Formation of stalactites and stalagmites:**



MAR

**The Equilibrium Constant**

For any type of chemical equilibrium of the type



the following is a **CONSTANT** (at a given T)

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

↑ conc. of products  
 ↓ conc. of reactants  
 ↑ equilibrium constant

**If K is known, we can predict concentrations of products or reactants.**

**The Equilibrium Constant: Example**

Equilibrium constants, K, come from *kinetic rate constants, k*

Example: For  $2 \text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$ ,

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

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

When rates equal,  $\text{rate}_f = \text{rate}_r$ , and

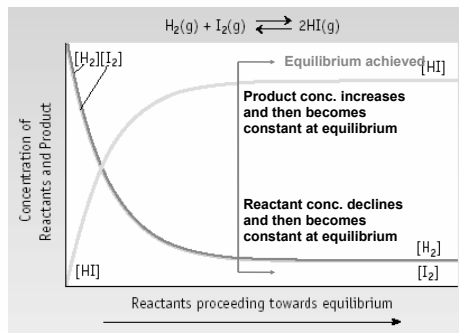
$$k_f[\text{NO}_2]^2 = k_r[\text{N}_2\text{O}_4]$$

$$\frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} = \frac{k_f}{k_r} = K \text{ (equilibrium)}$$

Product of two constants ( $k_f$  &  $k_r$ ) is itself a constant (K)!

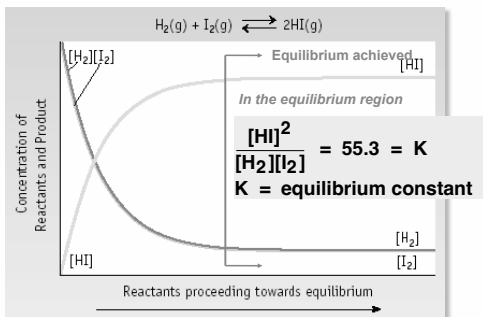
MAR

**The Equilibrium Constant: Example**



MAR

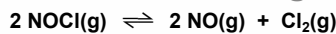
**The Equilibrium Constant: Example**



MAR

MAR

**Determining K**



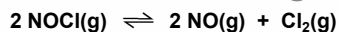
Place 2.00 mol of NOCl in a 1.00 L flask. At equilibrium you find 0.66 mol/L of NO. Calculate K.

Solution

Set of a table of concentrations (ICE)

	[NOCl]	[NO]	[Cl <sub>2</sub> ]
Initial	2.00	0	0
Change			
Equilibrium		0.66	

### Determining K



Place 2.00 mol of NOCl in a 1.00 L flask. At equilibrium you find 0.66 mol/L of NO. Calculate K.

Solution

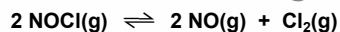
Set of a table of concentrations (ICE)

	[NOCl]	[NO]	[Cl <sub>2</sub> ]
Initial	2.00	0	0
Change	-0.66	+0.66	+0.33
Equilibrium	1.34	0.66	0.33

How to remember ICE:

MAR

### Determining K



	[NOCl]	[NO]	[Cl <sub>2</sub> ]
Initial	2.00	0	0
Change	-0.66	+0.66	+0.33
Equilibrium	1.34	0.66	0.33

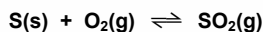
$$K = \frac{[\text{NO}]^2[\text{Cl}_2]}{[\text{NOCl}]^2}$$

$$K = \frac{[\text{NO}]^2[\text{Cl}_2]}{[\text{NOCl}]^2} = \frac{(0.66)^2(0.33)}{(1.34)^2} = 0.080$$

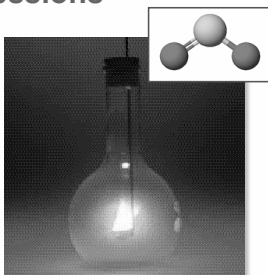
MAR

### Writing and Manipulating K Expressions

Solids and liquids *NEVER* appear in equilibrium expressions.



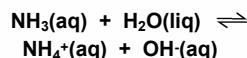
$$K = \frac{[\text{SO}_2]}{[\text{O}_2]}$$



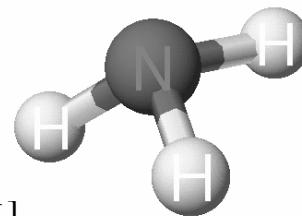
MAR

### Writing and Manipulating K Expressions

Solids and liquids *NEVER* appear in equilibrium expressions.



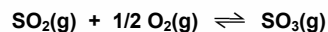
$$K = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$



MAR

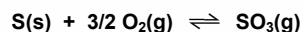
### Writing and Manipulating K Expressions

Adding equations for reactions



$$K_2 = \frac{[\text{SO}_3]}{[\text{SO}_2][\text{O}_2]^{1/2}}$$

NET EQUATION

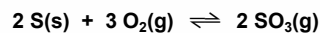
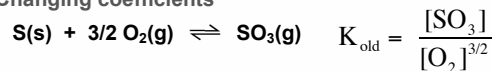


$$K_{\text{net}} = \frac{[\text{SO}_3]}{[\text{O}_2]^{3/2}} = K_1 \cdot K_2$$

MAR

### Writing and Manipulating K Expressions

Changing coefficients



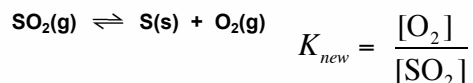
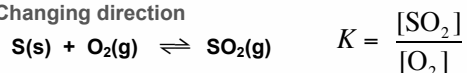
$$K_{\text{new}} = \frac{[\text{SO}_3]^2}{[\text{O}_2]^3}$$

$$K_{\text{new}} = \frac{[\text{SO}_3]^2}{[\text{O}_2]^3} = (K_{\text{old}})^2$$

MAR

## Writing and Manipulating K Expressions

Changing direction



$$\text{and: } K_{\text{new}} = \frac{[\text{O}_2]}{[\text{SO}_2]} = \frac{1}{K_{\text{old}}}$$

See Manipulating Equilibrium Constant Expressions

MAR

## K<sub>p</sub> Expressions

We have been writing K in terms of M (mol/L), designated by **K<sub>c</sub>**

Equilibrium constants expressed in terms of **gases** designated **K<sub>p</sub>** where all pressures in atm (760 mm Hg = 1 atm)

For:  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$

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

MAR

## Converting K<sub>c</sub> Into K<sub>p</sub>

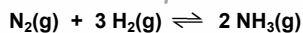
$$K_p = K_c(RT)^{\Delta n}$$

T = Temperature (K)

R = 0.082057 L atm mol<sup>-1</sup> K<sup>-1</sup>

Δn = change in moles of gas

Example:

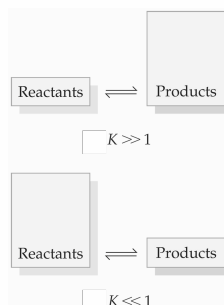


$$\Delta n = 2 - 4 = -2$$

see: Types of Equilibrium Constants

MAR

## The Meaning of K



If  $K \gg 1$ , the reaction is **product-favored**; product predominates at equilibrium.

If  $K \ll 1$ , the reaction is **reactant-favored**; reactant predominates at equilibrium.

MAR

## The Meaning of K

Can tell if a reaction is product-favored or reactant-favored.

For:  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$

$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = 3.5 \times 10^8$$

Conc. of products is **much greater** than that of reactants at equilibrium.

The reaction is strongly **product-favored**.

MAR

## The Meaning of K

For:  $\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

$$K_c = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-5}$$

Conc. of products is **much less** than that of reactants at equilibrium.

The reaction is strongly **reactant-favored**.

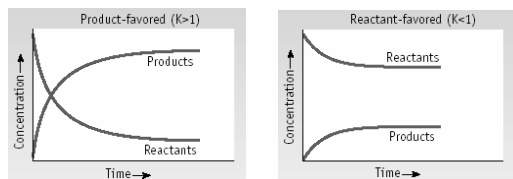


The reverse reaction  $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightleftharpoons \text{AgCl(s)}$  is **product-favored**,

$$K_{\text{rev}} = 1/1.8 \times 10^{-5} = 5.6 \times 10^4$$

MAR

### Product- or Reactant Favored



Product-favored  
 $K > 1$

Reactant-favored  
 $K < 1$

MAR

### The Reaction Quotient, Q

All reacting chemical systems can be characterized by their **REACTION QUOTIENT, Q**.



Under Any Reaction Conditions

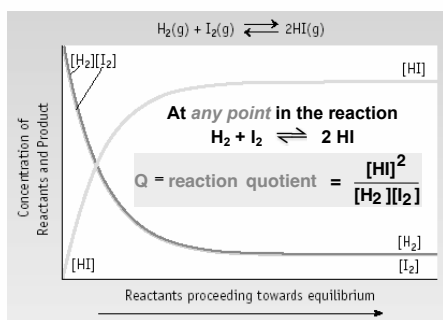
$$\text{Reaction quotient} = Q = \frac{\text{Product concentrations}}{\text{Reactant concentrations}} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

If  $Q = K$ , then system is at equilibrium.

If  $Q \neq K$ , then system is not at equilibrium

MAR

### The Reaction Quotient, Q



MAR

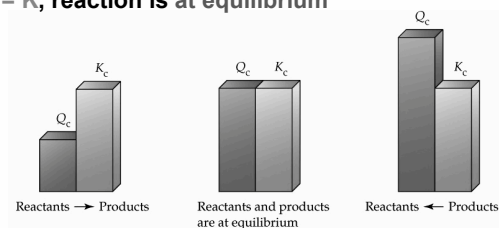
### The Reaction Quotient, Q

Comparing Q and K:

$Q < K$ , reaction shifts to the products (right)

$Q > K$ , reaction shifts to the reactants (left)

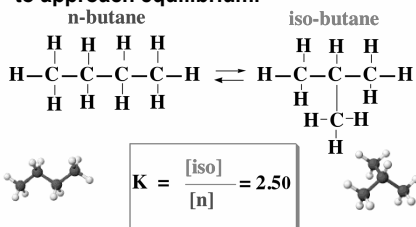
$Q = K$ , reaction is at equilibrium



MAR

### The Reaction Quotient, Q

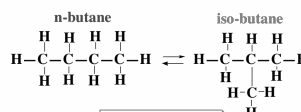
We can use Q to tell if a reaction is at equilibrium. If not at equilibrium, we can predict which way the reaction will move to approach equilibrium.



see [Understanding K and Q Handout](#)

MAR

### The Reaction Quotient, Q



If  $[\text{iso}] = 0.35 \text{ M}$  and  $[\text{n}] = 0.15 \text{ M}$ , are you at equilibrium?

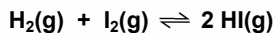
Which way does the reaction "shift" to approach equilibrium?

$$Q = \frac{\text{conc. of iso}}{\text{conc. of n}} = \frac{0.35}{0.15} = 2.3$$

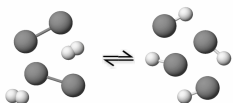
To **[iso]**!  $Q = 2.3$ ,  $Q < K$ , reaction will shift to "right" (product side, or **[iso]**)

MAR

## Typical Calculations

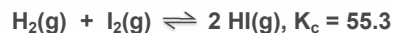


$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = 55.3$$



Place 1.00 mol each of  $\text{H}_2$  and  $\text{I}_2$  in a 1.00 L flask.  
Calc. equilibrium concentrations.

MAR

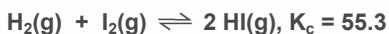


Step 1. Set up ICE table to define EQUILIBRIUM concentrations.

	$[\text{H}_2]$	$[\text{I}_2]$	$[\text{HI}]$
Initial			
Change			
Equilib			

where  $x$  is defined as am't of  $\text{H}_2$  and  $\text{I}_2$  consumed on approaching equilibrium.

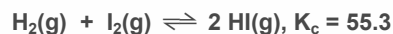
MAR



Step 2. Put equilibrium concentrations into  $K_c$  expression.

$$K_c = \frac{[2x]^2}{[1.00-x][1.00-x]} = 55.3$$

MAR



Step 3. Solve  $K_c$  expression - take square root of both sides.

$$7.44 = \frac{2x}{1.00-x}$$

$$7.44(1.00-x) = 2x$$

$$7.44 - 7.44x = 2x$$

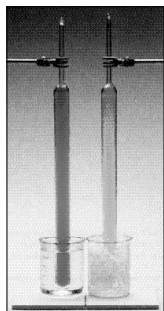
$$x = 7.44 / 9.44 = 0.788$$

Therefore, at equilibrium

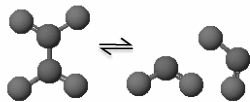
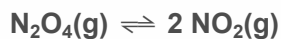
$$[\text{H}_2] = [\text{I}_2] = 1.00 - x = 0.21 \text{ M}$$

$$[\text{HI}] = 2x = 1.58 \text{ M}$$

MAR

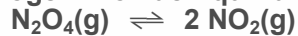


## Nitrogen Dioxide Equilibrium



MAR

## Nitrogen Dioxide Equilibrium



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.0059 \text{ at } 298 \text{ K}$$

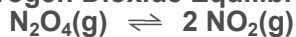
If initial concentration of  $\text{N}_2\text{O}_4$  is 0.50 M, what are the equilibrium concentrations?

Step 1. Set up an equilibrium table

	$[\text{N}_2\text{O}_4]$	$[\text{NO}_2]$
Initial	0.50	0
Change		
Equilib		

MAR

## Nitrogen Dioxide Equilibrium



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.0059 \text{ at } 298 \text{ K}$$

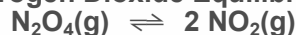
If initial concentration of  $\text{N}_2\text{O}_4$  is 0.50 M, what are the equilibrium concentrations?

Step 1. Set up an equilibrium table

	$[\text{N}_2\text{O}_4]$	$[\text{NO}_2]$
Initial	0.50	0
Change	-x	+2x
Equilib	0.50 - x	2x

MAR

## Nitrogen Dioxide Equilibrium



Step 2. Substitute into  $K_c$  expression and solve.

$$K_c = 0.0059 = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{(2x)^2}{(0.50 - x)}$$

$$\begin{aligned} \text{Rearrange: } 0.0059(0.50 - x) &= 4x^2 \\ 0.0029 - 0.0059x &= 4x^2 \\ 4x^2 + 0.0059x - 0.0029 &= 0 \end{aligned}$$

This is a **QUADRATIC EQUATION**

$$ax^2 + bx + c = 0$$

$$a = 4 \quad b = 0.0059 \quad c = -0.0029$$

CH 223: report **both** values of  $x$  when performing quadratic calculations

MAR

## Nitrogen Dioxide Equilibrium



Solve the quadratic equation for  $x$ .

$$ax^2 + bx + c = 0$$

$$a = 4 \quad b = 0.0059 \quad c = -0.0029$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-0.0059 \pm \sqrt{(0.0059)^2 - 4(4)(-0.0029)}}{2(4)}$$

$$x = -0.00074 \pm 1/8(0.046)^{1/2} = -0.00074 \pm 0.027$$

MAR

## Nitrogen Dioxide Equilibrium



$$x = \frac{-0.0059 \pm \sqrt{(0.0059)^2 - 4(4)(-0.0029)}}{2(4)}$$

$$x = -0.00074 \pm 1/8(0.046)^{1/2} = -0.00074 \pm 0.027$$

$$x = 0.026 \text{ or } -0.028$$

But a **negative** value is not reasonable.

**Conclusion:  $x = 0.026$**

$$[\text{N}_2\text{O}_4] = 0.50 - x = 0.47 \text{ M}$$

$$[\text{NO}_2] = 2x = 0.052 \text{ M}$$

CH 223: report **both** values of  $x$  when performing quadratic calculations

MAR

## Le Chatelier's Principle

Temperature, changes in volume, changes in pressure and changes in concentration affect equilibria.

The outcome is governed by **LE CHATELIER'S PRINCIPLE**

"...if a system at equilibrium is disturbed, the system tends to shift its equilibrium position to counter the effect of the disturbance."

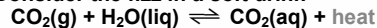


Henri Le Chatelier,  
1850 - 1936

MAR

Temperature effects change the numeric value of  $K$ 

Consider the fizz in a soft drink



Decrease  $T$ . What happens to equilibrium position? To value of  $K$ ?

$$K = [\text{CO}_2] / P(\text{CO}_2)$$

$K$  increases as  $T$  goes down because  $[\text{CO}_2]$  increases and  $P(\text{CO}_2)$  decreases.

Increase  $T$ . Now what?

Equilibrium shifts left and  $K$  decreases.

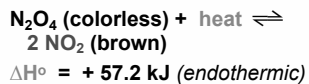
see: [Le Chatelier's Guide](#)



MAR



### Temperature Effects on Equilibrium

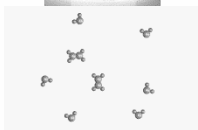
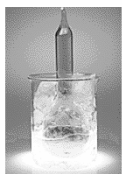


$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

$K_c = 0.00077$  at 273 K

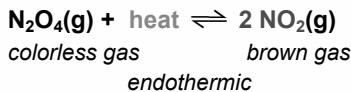
$K_c = 0.0059$  at 298 K

**K changes with temperature**



MAR

### Volume & Pressure Effects on Equilibrium



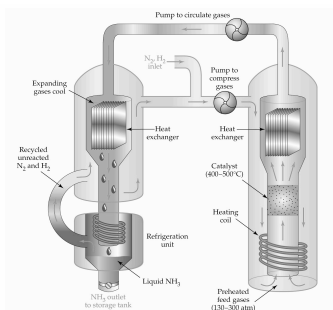
**Volume and Pressure changes affect equilibrium**

**Larger volume / Less Pressure = more molecules**

**Smaller volume / Higher Pressure = fewer molecules**

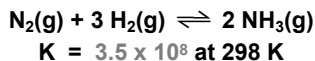


MAR



### NH<sub>3</sub> Production

**Catalyst for the Haber-Frisch reaction is mostly Fe with some KOH**

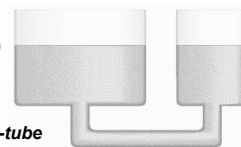
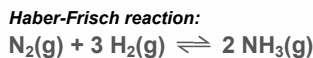


MAR

### EQUILIBRIUM AND EXTERNAL EFFECTS

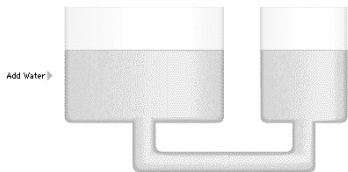
**Concentration changes ---> no change in K - only the position of equilibrium changes.**

**We can use the "water U-tube" to demonstrate this phenomena (and connect to Haber-Frisch)**

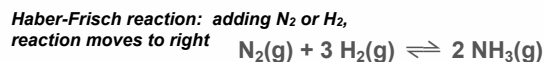


MAR

### Le Chatelier's Principle

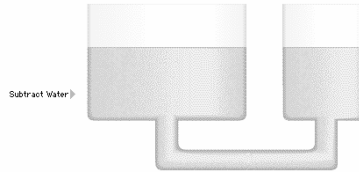


**Adding a "reactant" to a chemical system.**

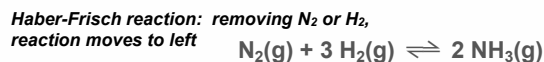


MAR

### Le Chatelier's Principle



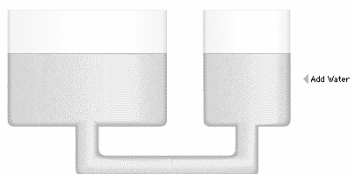
**Removing a "reactant" from a chemical system.**



MAR



### Le Chatelier's Principle



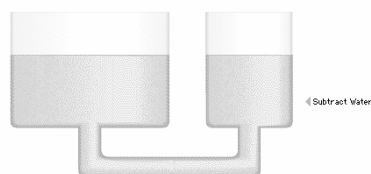
Adding a "product" to a chemical system.

Haber-Frisch reaction: *adding NH<sub>3</sub>, reaction moves to left*  

$$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$$

MAR

### Le Chatelier's Principle

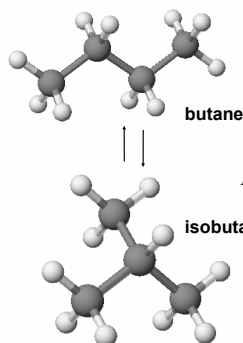


Removing a "product" from a chemical system.

Haber-Frisch reaction: *removing NH<sub>3</sub>, reaction moves to right*  

$$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$$

MAR



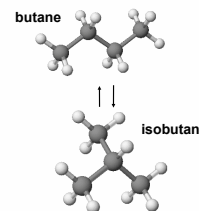
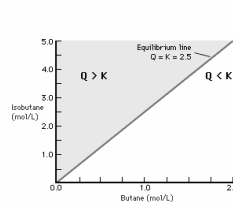
### Butane-Isobutane Equilibrium

$$K = \frac{[\text{isobutane}]}{[\text{butane}]} = 2.50$$

MAR

### Butane $\rightleftharpoons$ Isobutane

Assume you are at equilibrium with [iso] = 1.25 M and [butane] = 0.50 M. Now add 1.50 M butane. When the system comes to equilibrium again, what are [iso] and [butane]?  $K = 2.50$



MAR

### Butane $\rightleftharpoons$ Isobutane

Assume you are at equilibrium with [iso] = 1.25 M and [butane] = 0.50 M. Now add 1.50 M butane. When the system comes to equilibrium again, what are [iso] and [butane]?  $K = 2.50$

#### Solution

Calculate Q immediately after adding more butane and compare with K.

$$Q = \frac{[\text{isobutane}]}{[\text{butane}]} = \frac{1.25}{0.50 + 1.50} = 0.625$$

Q is **LESS THAN** K. Therefore, the reaction will shift to the \_\_\_\_\_.

MAR

### Butane $\rightleftharpoons$ Isobutane

You are at equilibrium with [iso] = 1.25 M and [butane] = 0.50 M. Now add 1.50 M butane.

#### Solution

Q is less than K, so equilibrium shifts right - away from butane and toward isobutane.

Set up concentration (ICE) table

	[butane]	[isobutane]
Initial		
Change		
Equilibrium		

MAR

Butane  $\rightleftharpoons$  Isobutane

You are at equilibrium with [iso] = 1.25 M and [butane] = 0.50 M. Now add 1.50 M butane.

Solution

$$K = 2.50 = \frac{[\text{isobutane}]}{[\text{butane}]} = \frac{1.25 + x}{2.00 - x}$$

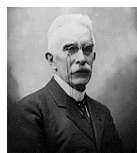
$$x = 1.07 \text{ M}$$

At the new equilibrium position,

[butane] = 0.93 M and [isobutane] = 2.32 M

Equilibrium has shifted toward isobutane.

MAR



## Le Chatelier's Principle - Overview

## Change T

- changes the value of K
- changes observed in P, V or concentrations at equilibrium

## Add or take away reactant or product

- K does not change
- Reaction adjusts to new equilibrium "position"

## Use a catalyst

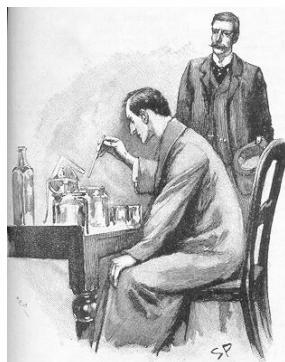
- reaction comes more quickly to equilibrium. K not changed.

MAR

## End of Chapter 13

See:

- [Chapter Thirteen Study Guide](#)
- [Chapter Thirteen Concept Guide](#)
- [Types of Equilibrium Constants](#)
- [Important Equations \(following this slide\)](#)
- [End of Chapter Problems \(following this slide\)](#)

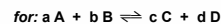


MAR

## Important Equations, Constants, and Handouts from this Chapter:

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

$$R = 0.082057 \text{ L atm mol}^{-1} \text{ K}^{-1}$$



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

↑  
equilibrium constant

Under Any Reaction Conditions

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

Product concentrations  
Reactant concentrations

Handouts:

- [Manipulating Equilibrium Constant Expressions](#)
- [Types of Equilibrium Constants](#)

MAR

Le Chatelier's Principle

## End of Chapter Problems: Test Yourself

1. Write an equilibrium constant expression for the following reaction:  $\text{C(s)} + \text{CO}_2(\text{g}) \rightleftharpoons 2 \text{CO}(\text{g})$
2.  $K = 5.6 \times 10^{-12}$  at 500 K for the dissociation of iodine molecules to iodine atoms:  $\text{I}_2(\text{g}) \rightleftharpoons 2 \text{I}(\text{g})$ . A mixture has  $[\text{I}_2] = 0.020 \text{ mol/L}$  and  $[\text{I}] = 2.0 \times 10^{-8} \text{ mol/L}$ . Is the reaction at equilibrium (at 500 K)? If not, which way must the reaction proceed to reach equilibrium?
3. The reaction:  $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$  was examined at 250 °C. At equilibrium,  $[\text{PCl}_5] = 4.2 \times 10^{-5} \text{ mol/L}$ ,  $[\text{PCl}_3] = 1.3 \times 10^{-2} \text{ mol/L}$ , and  $[\text{Cl}_2] = 3.9 \times 10^{-3} \text{ mol/L}$ . Calculate K for the reaction.
4. The equilibrium constant for the dissociation of iodine molecules to iodine atoms:  $\text{I}_2(\text{g}) \rightleftharpoons 2 \text{I}(\text{g})$  is  $3.76 \times 10^{-3}$  at 1000 K. Suppose 0.105 mol of  $\text{I}_2$  is placed in a 12.3 L flask at 1000 K. What are the concentrations of  $\text{I}_2$  and I when the system comes to equilibrium?
5. Dinitrogen trioxide decomposes to NO and  $\text{NO}_2$  in an endothermic process ( $\Delta H = 40.5 \text{ kJ/mol}$ ):  $\text{N}_2\text{O}_3(\text{g}) \rightleftharpoons \text{NO}(\text{g}) + \text{NO}_2(\text{g})$ . Predict the effect of the following changes on the position of the equilibrium (left, right, or no change):
  - a. adding more  $\text{N}_2\text{O}_3(\text{g})$
  - b. adding more  $\text{NO}_2(\text{g})$
  - c. increasing the volume of the reaction flask
  - d. lowering the temperature

MAR

## End of Chapter Problems: Answers

1.  $K_c = [\text{CO}]^2 / [\text{CO}_2]$  or  $K_p = P_{\text{CO}}^2 / P_{\text{CO}_2}$
2. No,  $Q < K$ , reaction will proceed to the right (more products)
3.  $K = 1.2$
4.  $[\text{I}_2] = 0.00614 \text{ M}$ ,  $[\text{I}] = 0.00480 \text{ M}$
5. a. right b. left c. right d. left

MAR