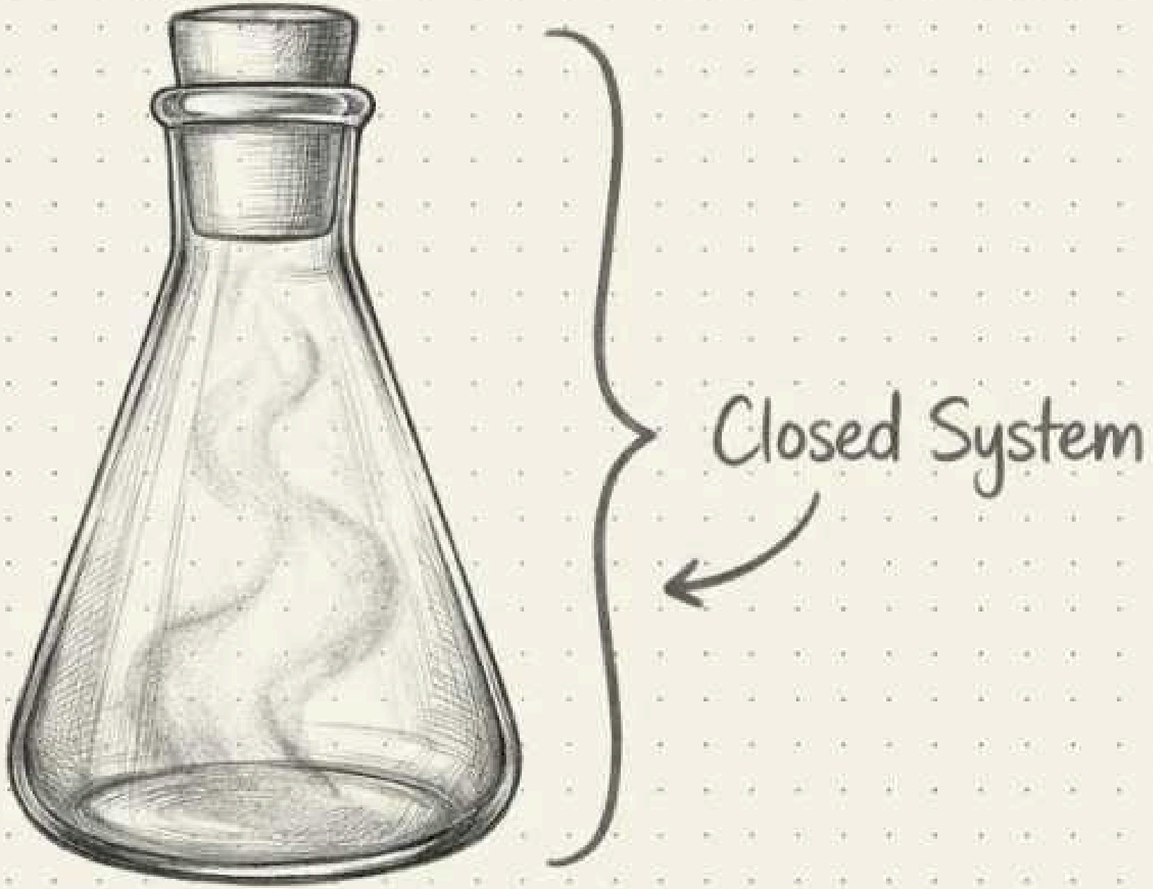
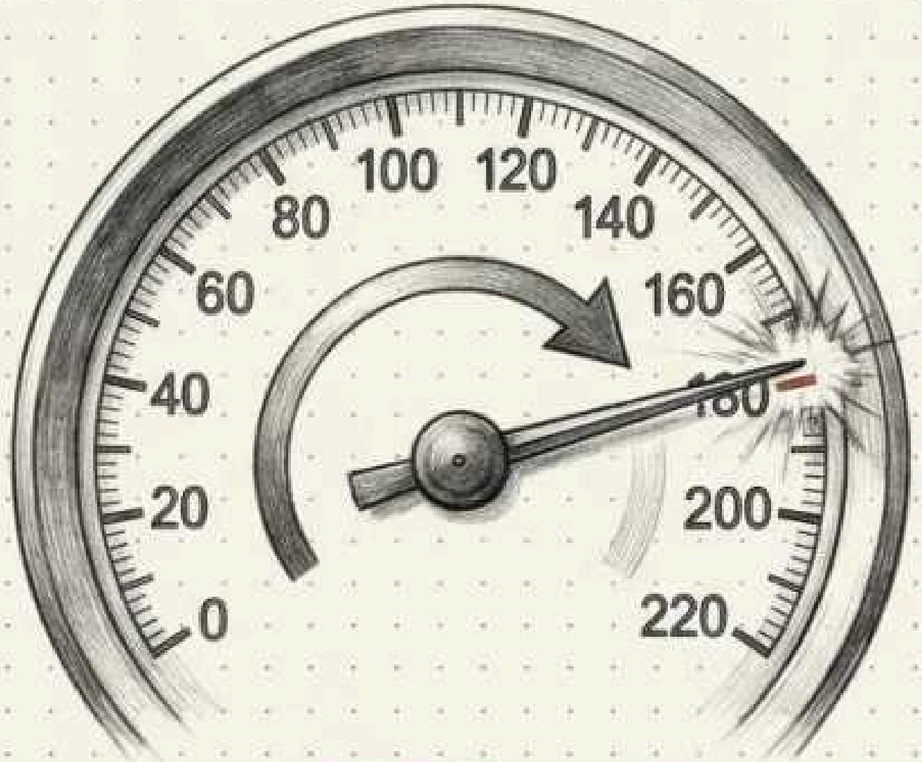
CHEM 102 B Lecture 2: Chemical Equilibrium

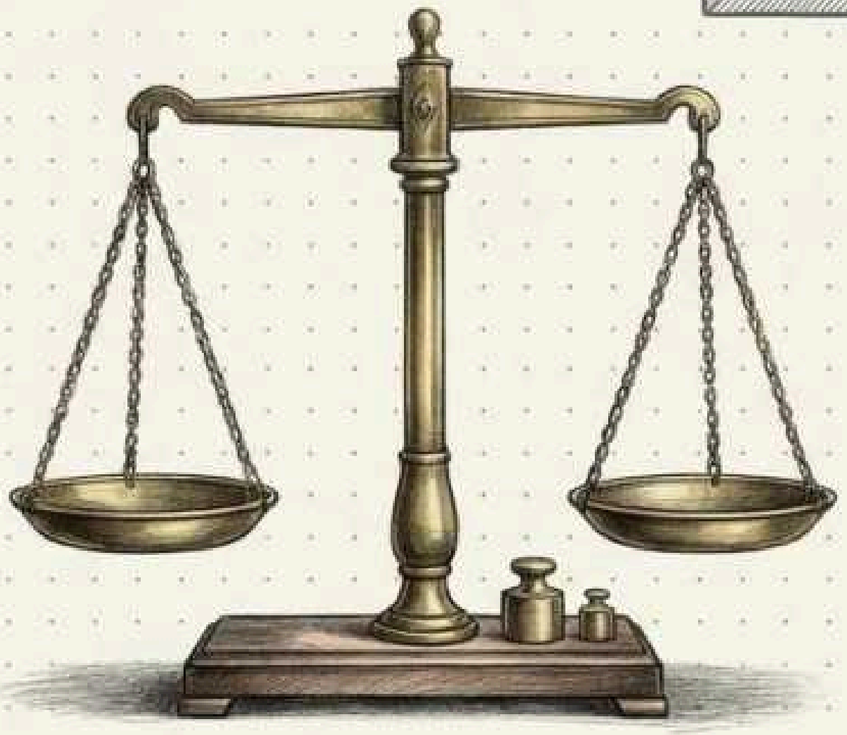
At a glance: The equilibrium constant, K and Q, ICE tables, Le Chatelier's principle, and the Haber process.



¶ 2.2



How fast does it go?



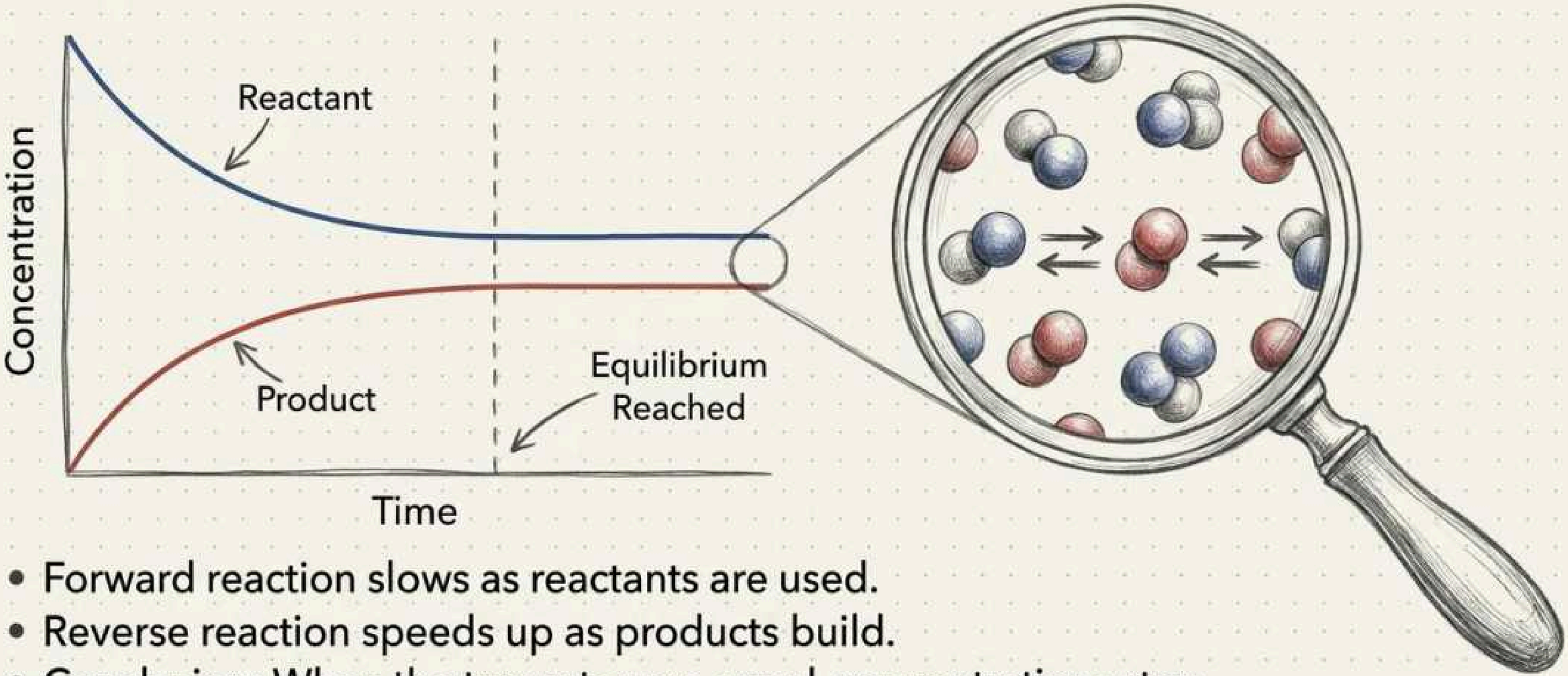
How FAR does it go?

Many reactions do not run to completion.

Reactants and products settle into a balance. We will describe this balance with one number, and calculate exactly what is in the flask.

Dynamic Equilibrium

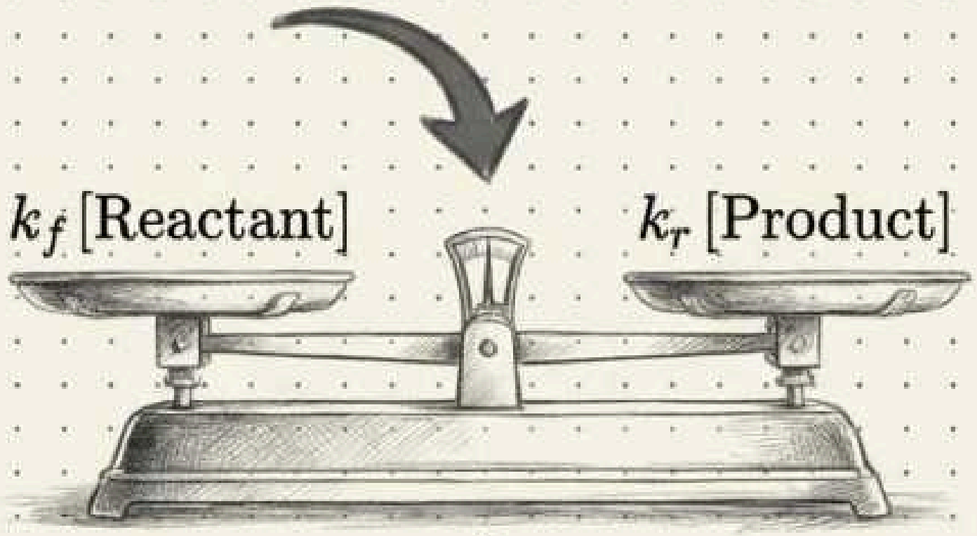
¶ 2.3



- Forward reaction slows as reactants are used.
- Reverse reaction speeds up as products build.
- Conclusion: When the two rates are equal, concentrations stop changing. Both reactions go on, but there is no net change.

¶ 2.4

$$Rate_{forward} = k_f\,[\text{Reactant}] \qquad Rate_{reverse} = k_r\,[\text{Product}]$$



$$k_f\,[\text{Reactant}] = k_r\,[\text{Product}]$$

Rearranging...

$$\frac{k_f}{k_r} = \frac{[\text{Product}]}{[\text{Reactant}]} = K$$

Setting the two rates equal and rearranging gives a ratio of concentrations that is constant at one temperature.

The Law of Mass Action

¶ 2.5

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

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

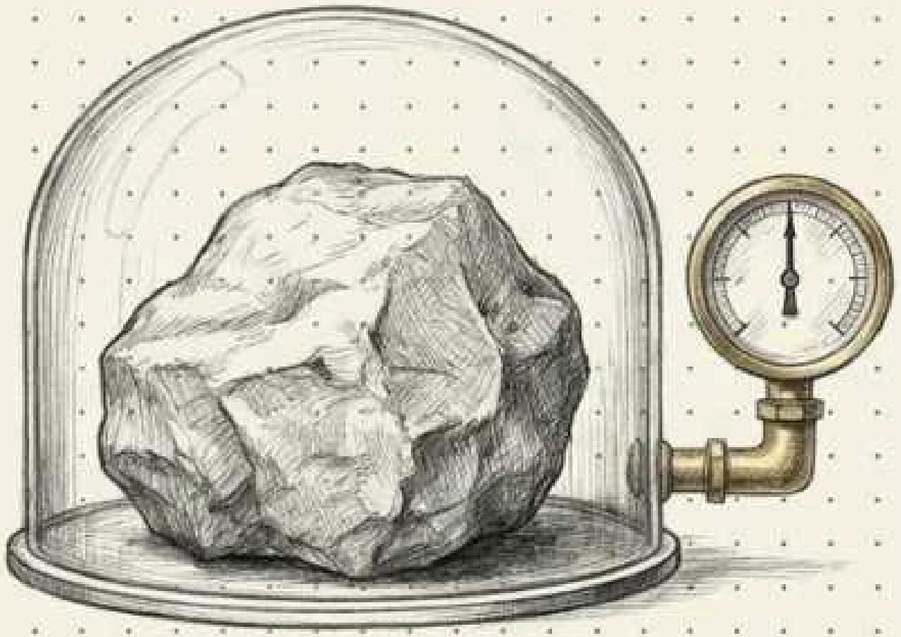
Products go on top!

Rule: Each concentration is raised to its coefficient.

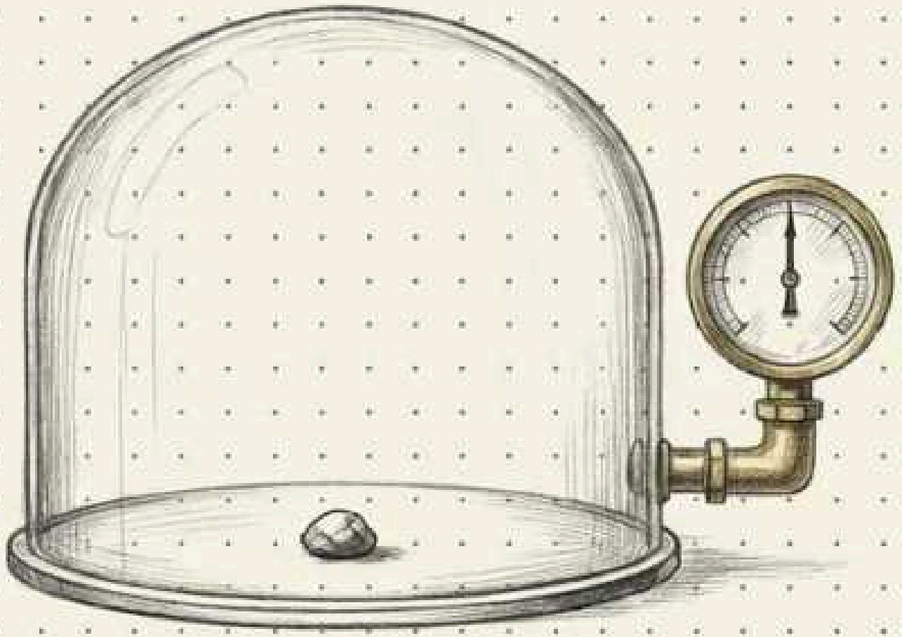
Excluding Solids and Liquids

¶ 2.6

Dome A



Dome B



$$\mathrm{CaCO_3}(s) \rightleftharpoons \mathrm{CaO}(s) + \mathrm{CO_2}(g)$$

$$K = [\mathrm{CO_2}]$$

Pure solids and liquids are left out of *K* because their concentrations do not change. A heap of limestone and a small pile give the same CO_2 pressure.

K_p vs K_c

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

Change in moles of gas.

Example Reaction:

$$N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$$

(2 moles product) - (1 mole reactant) = Δn of +1

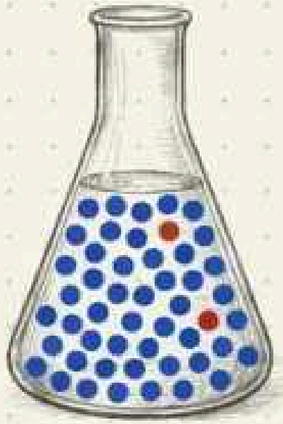
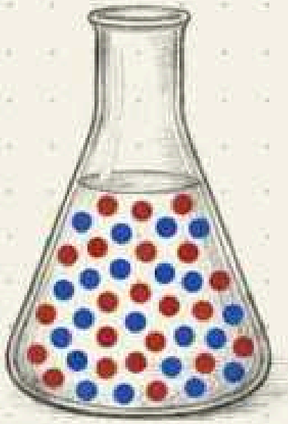
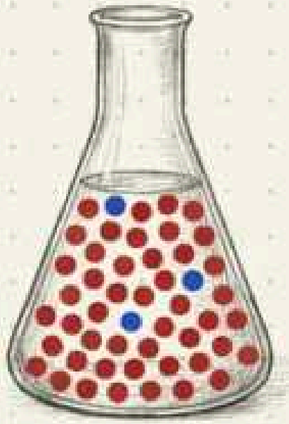
Given at 350K:
$K_c = 0.36$



Calculation:
$K_p \approx 10.3$

The Story Told by K

¶ 2.8

$K \ll 1$	$K \approx 1$	$K \gg 1$
		
Reactants favored. Example: $N_2 + O_2$	A real mixture.	Products favored. Example: $H_2 + O_2 \rightleftharpoons H_2O$

Tiny K at room temp, so air does not turn to poison

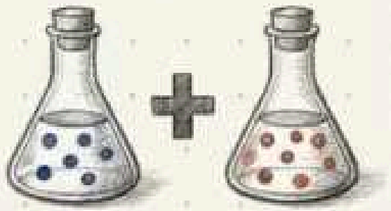


WARNING: K says NOTHING about speed!

Hydrogen and oxygen have an **enormous** K, but can **sit for years** until a **spark arrives**.

Manipulating K

¶ 2.9

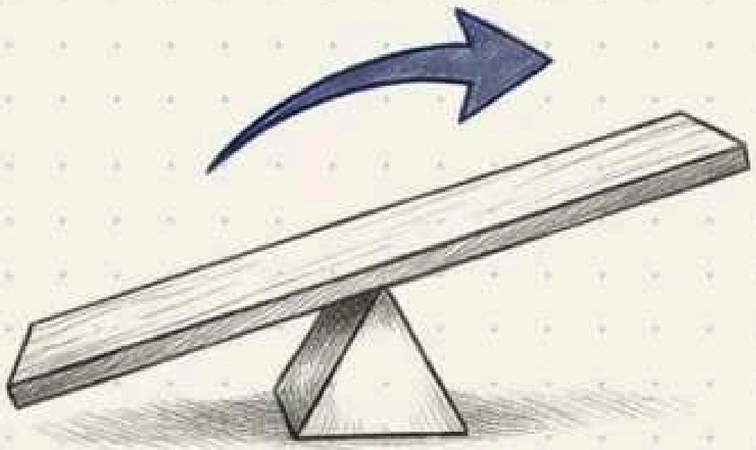
	Reverse a reaction	→	$\text{New } K = \dfrac{1}{K_{\text{old}}}$
	Multiply equation by 2	→	$\text{New } K = (K_{\text{old}})^2$
	Add two reactions	→	Multiply their K values $(K_1 \times K_2)$

These rules let you build one constant from others, just as you add enthalpies.

The Reaction Quotient (Q)

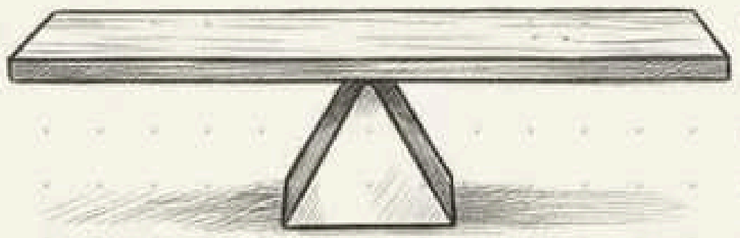
¶ 2.10

$Q < K$



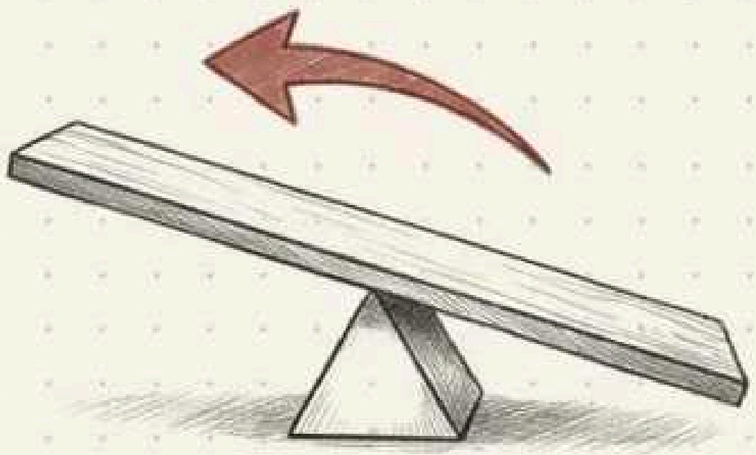
Too little product.
Moves forward.

$Q = K$



Equilibrium.

$Q > K$



Too much product.
Moves back.

Q has the same form as K, but uses current concentrations.
Comparing Q with K tells you which way the reaction will move.

Calculating Q: An Example

Given:

$N_2O_4 \rightleftharpoons 2\ NO_2$

$K = 0.36$

Flask holds:

0.30 M NO_2 and

0.10 M N_2O_4

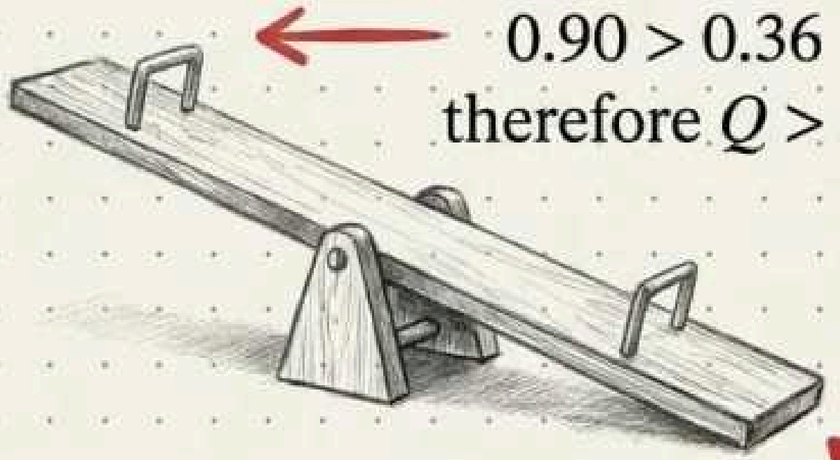
Equation:

$$Q = \frac{[NO_2]^2}{[N_2O_4]}$$

$$Q = \frac{(0.30)^2}{0.10}$$

$$Q = 0.90$$

Conclusion:



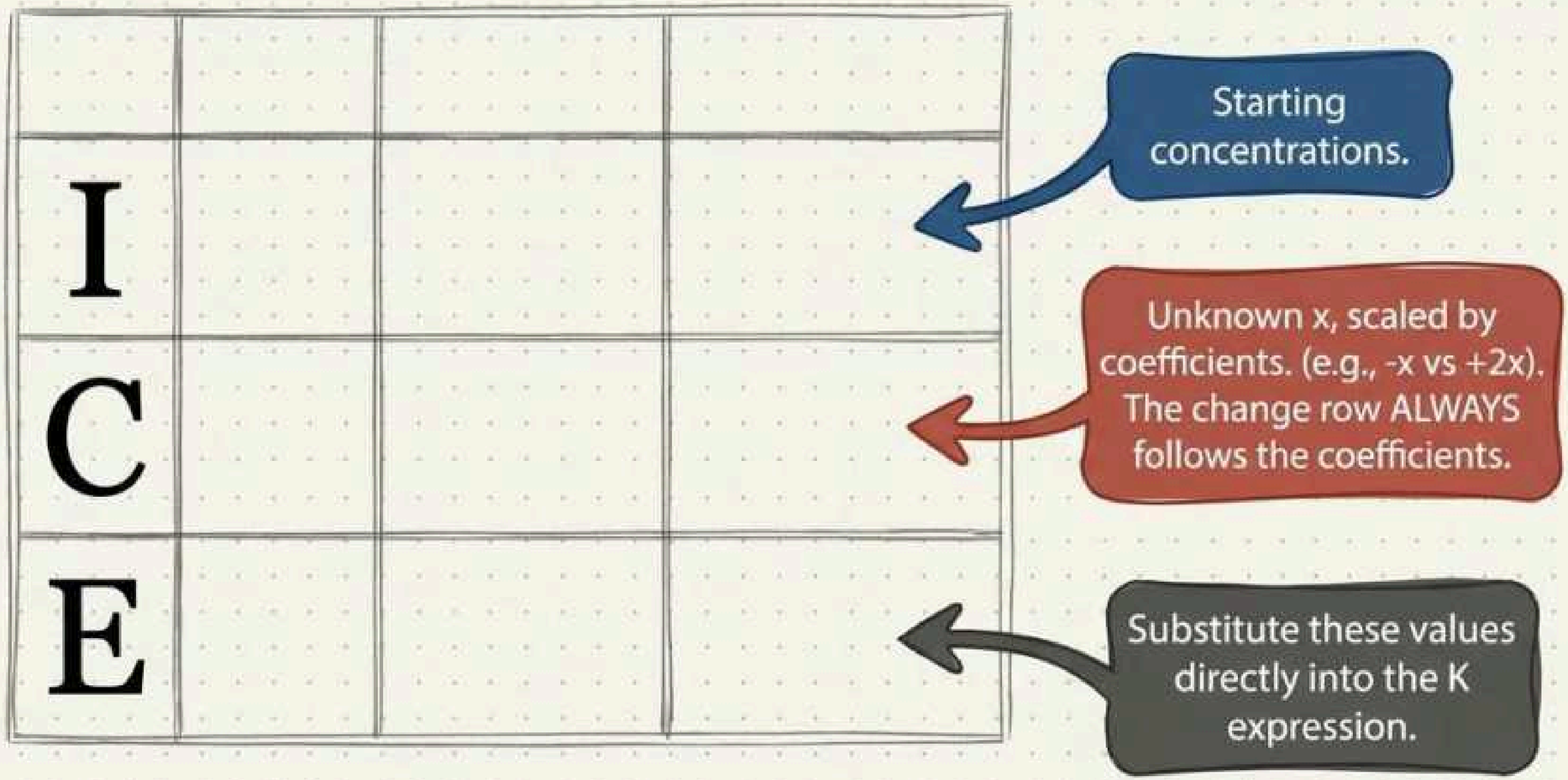
$0.90 > 0.36$

therefore $Q > K$

There is too much nitrogen dioxide. The reaction shifts BACK toward dinitrogen tetroxide.

Anatomy of an ICE Table

The main skill: It works for almost every equilibrium problem.



Setting Up the ICE Table

	$N_2O_4 \rightleftharpoons 2\ NO_2$	
I	0.100	0
C	-x	+2x
E	0.100 - x	2x

$$K = \frac{(2x)^2}{0.100 - x} = 0.36$$

Setting up the math to solve for x.

Solving the Quadratic

¶ 2.14

$$4x^2 + 0.36x - 0.036 = 0$$

Multiply out

↓

$$x = 0.060$$

← Apply quadratic formula

↓

$$[\mathrm{N_2O_4}] = 0.100 - 0.060 = 0.040\ M$$
$$[\mathrm{NO_2}] = 2(0.060) = 0.120\ M$$

Equilibrium Concentrations

$$\frac{(0.12)^2}{0.04} = 0.36$$ ✓

The math perfectly matches the known K of 0.36.

The Perfect Square Shortcut

¶ 2.15

Reaction: H2 + I2 ⇌ 2 HI (K = 50, Start with 1 M each)

$$50 = \sqrt{\frac{(2x)^2}{(1-x)^2}}$$

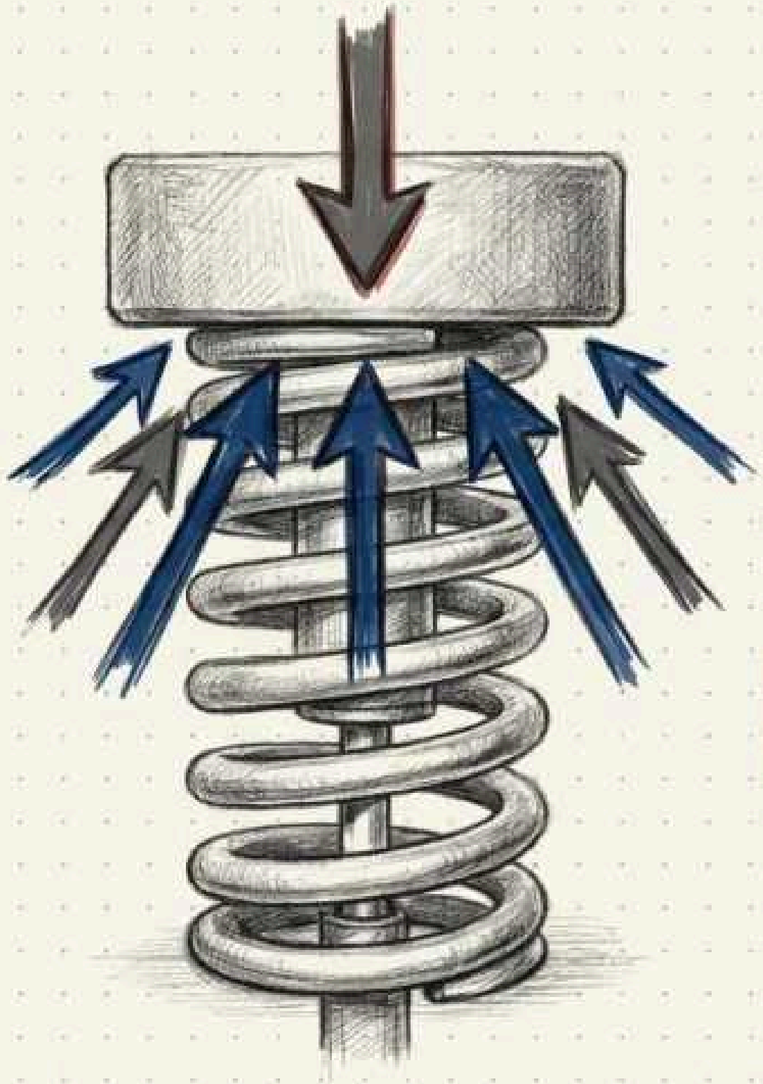
$$\sqrt{50} \approx 7.07 = \frac{2x}{1-x}$$

$$x = 0.78$$

Sometimes the equation is easier. When both sides are perfect squares, take the square root to skip the quadratic formula.

Le Chatelier's Principle (1884)

¶ 2.16

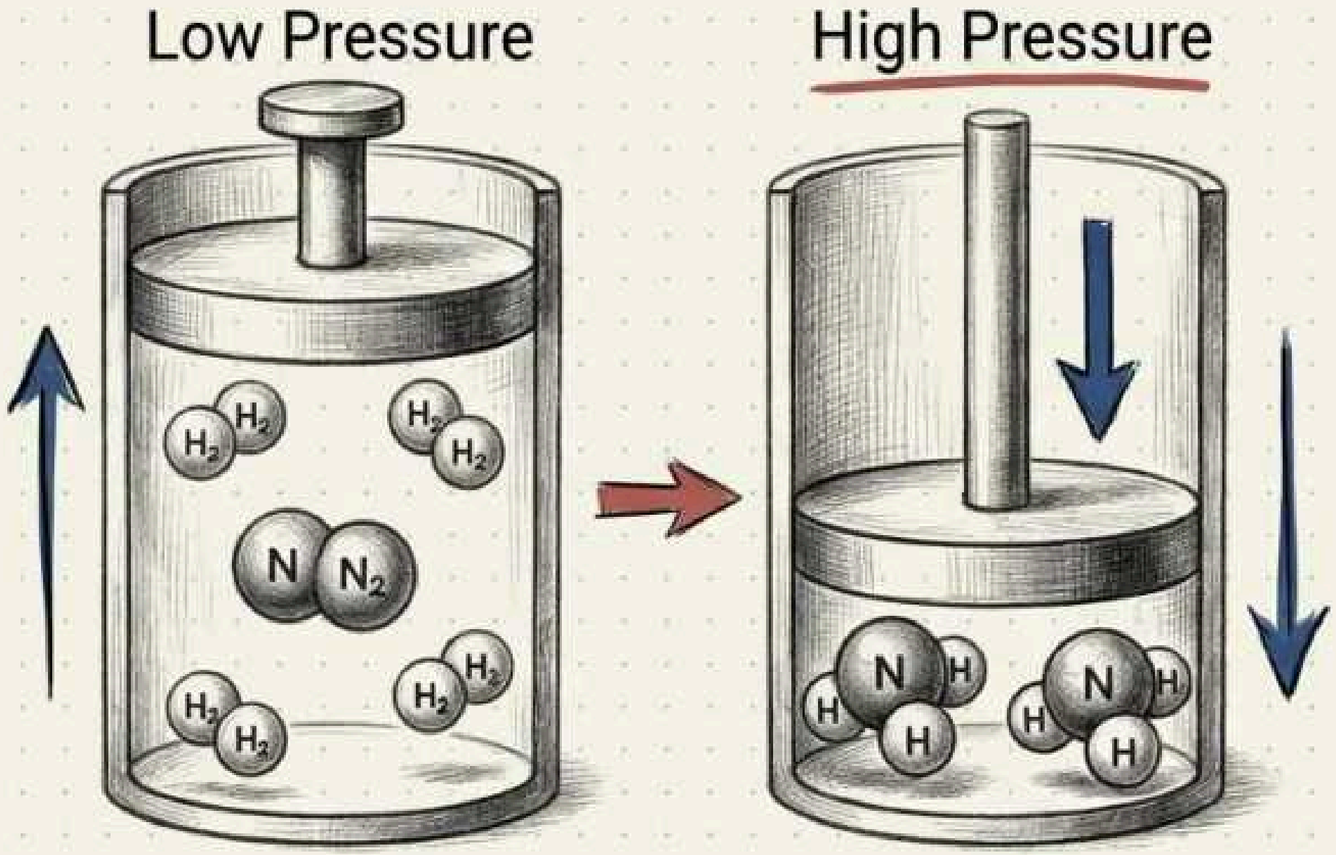


The system shifts to oppose the change.

Add Reactant → $Q < K$ → System shifts Forward (toward products)

The reaction moves until Q rises back to K.

Le Chatelier: Pressure & Volume



- **Concept:** Pressure matters only for gases. Compress the mixture, and it shifts toward the side with fewer gas molecules.
- **Application:** In the Haber process, high pressure favors ammonia.

Adding inert gas at constant volume changes nothing, because no concentration has changed.

Le Chatelier: Temperature

¶ 2.18



Endothermic reactions

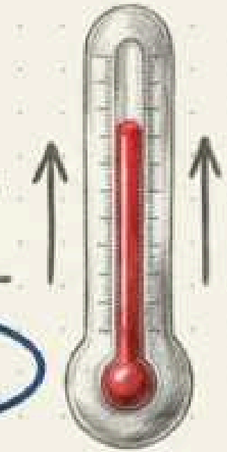
HEAT + Reactants ⇄ Products

Raises K

Exothermic

Reactants ⇄ Products + HEAT

Lowers K

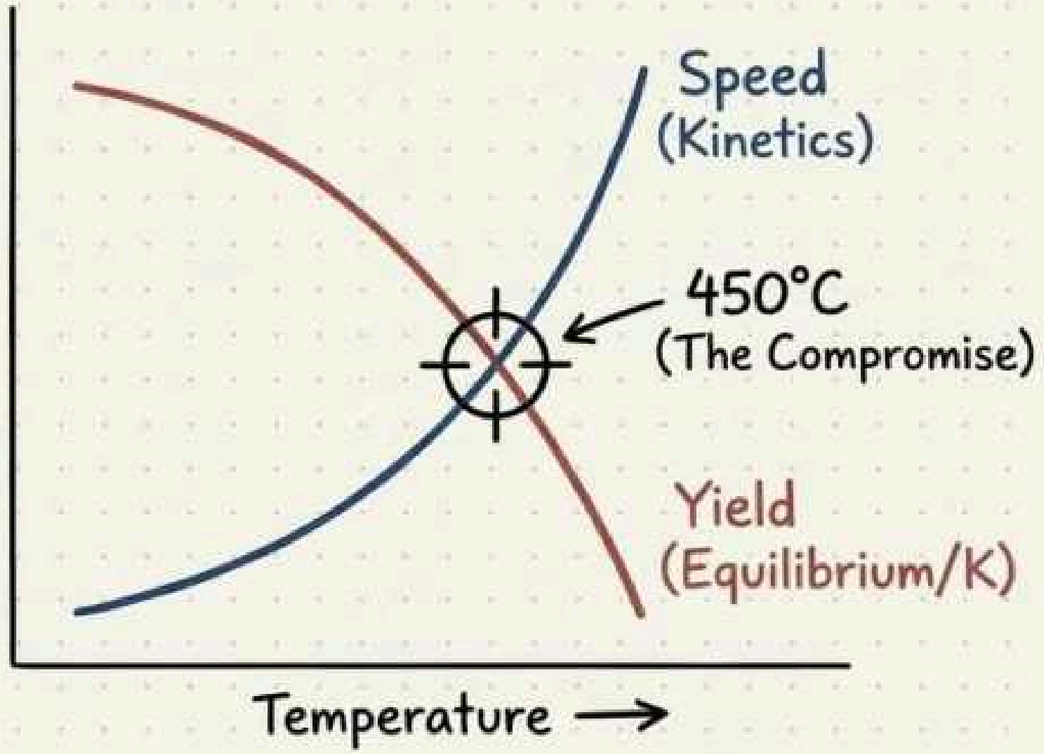


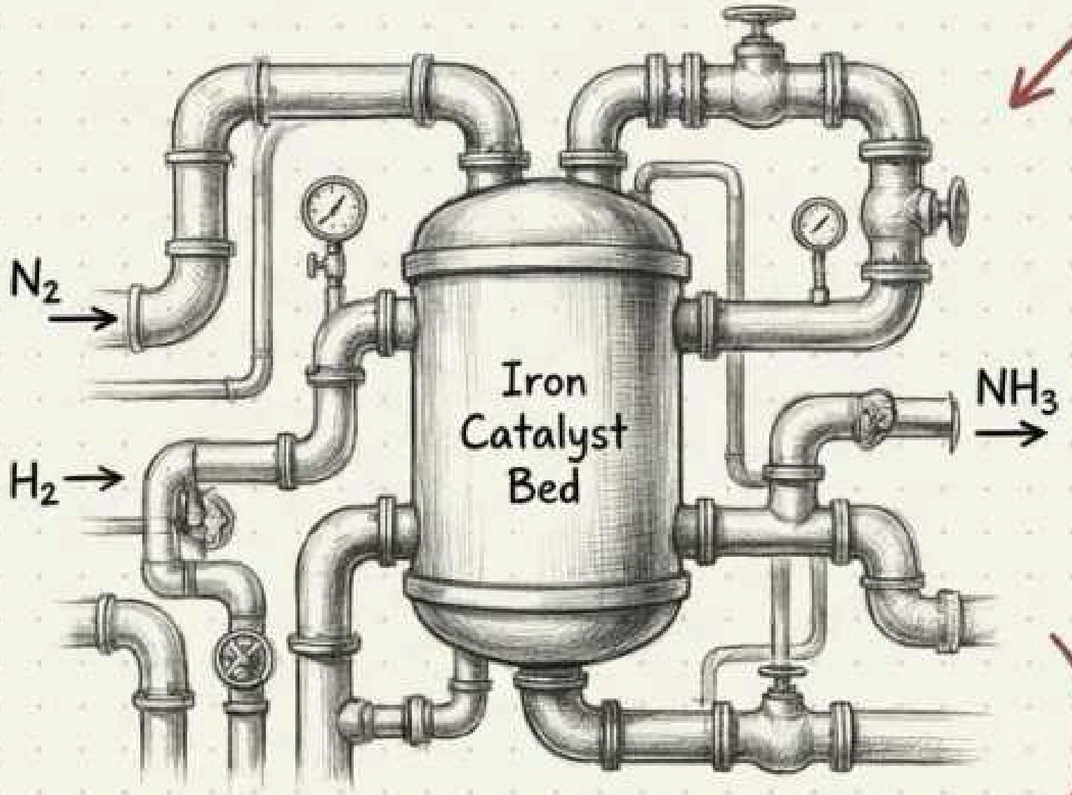
Main Takeaway: Temperature is different. Treat heat as a reactant or product.

Crucial Note: Concentration, pressure, and catalysts DO NOT change the value of K at all.

¶ 2.19

The Haber Process Compromise





The Problem: Making ammonia is exothermic (low temp gives bigger K). But low temp is too slow.

The Solution: Industry uses 450°C, high pressure, an iron catalyst, and removes NH3 as it forms.

Conclusion: Kinetics and equilibrium both shape the design.

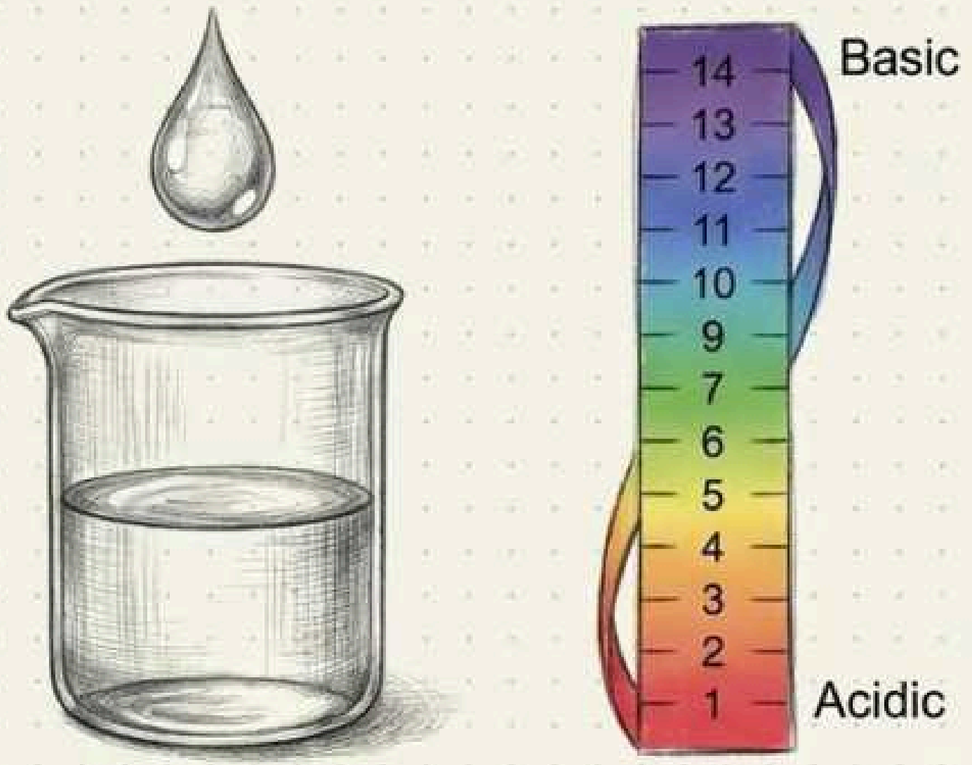
¶ 2.20

Conclusion & Looking Ahead

The Summary Checklist

- ✓ Describe an equilibrium.
- ✓ Predict an equilibrium.
- ✓ Calculate an equilibrium.

The Preview



Looking Ahead: The next two lectures apply these tools to water.

Teaser: Acids and bases turn out to be equilibria, and the ICE table will solve their pH.