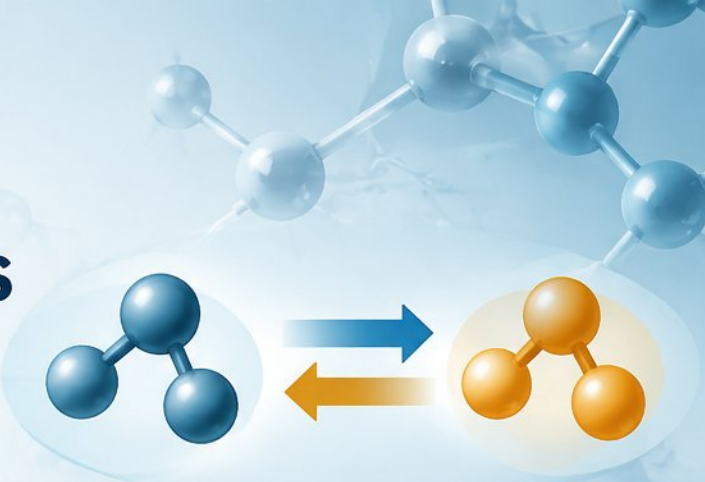
# Chemical Equilibrium and Le Chatelier's Principle: Introduction and Learning Goals



- **Course path:** reversible reactions → dynamic equilibrium → K and Q → disturbances → Le Chatelier → quantitative checks

- For chemistry learners, high school and early university students, and science teachers

- **Why it matters:** Haber–Bosch ammonia, blood pH and oxygen transport, environmental chemistry, everyday saturated solutions

- **Goal 1:** explain dynamic equilibrium as equal forward and reverse rates

- **Goal 2:** predict shift direction after a disturbance, justified by K/Q reasoning, not memorized rules









# Reversible Reactions and the Meaning of Dynamic Equilibrium

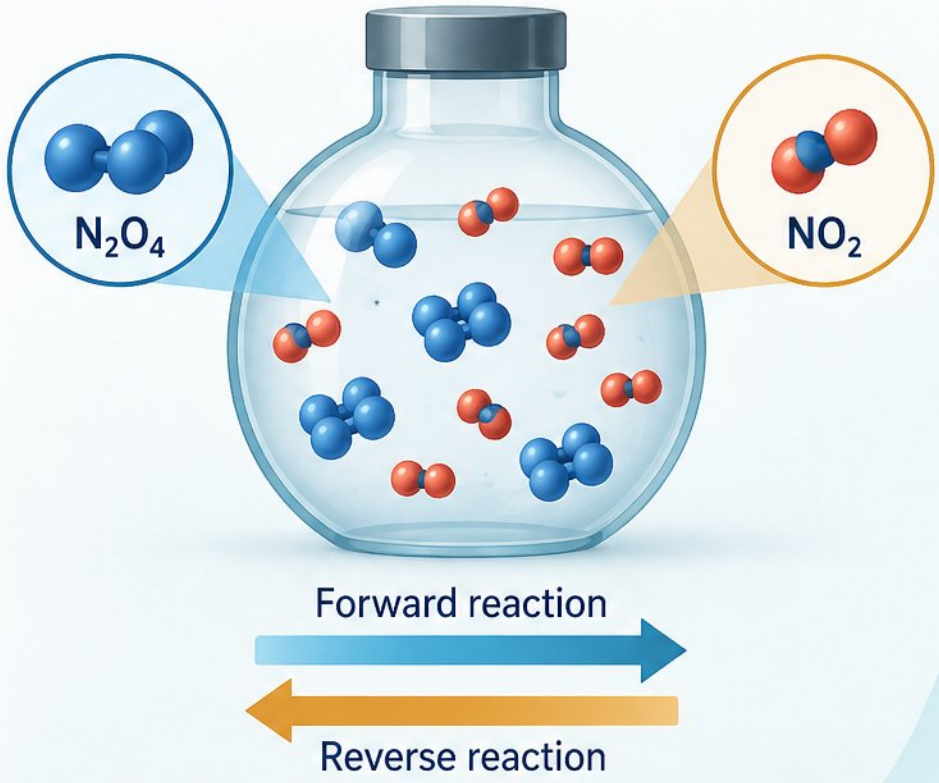
**Reversible reactions:**
forward and reverse occur together (⇌)

**Dynamic equilibrium:**
equal rates, not equal amounts

**Concentrations stay constant**
while both reactions continue

**Misconception:**
equilibrium does not mean the reaction has stopped

$$N_2O_4 \rightleftharpoons 2\,NO_2$$




**At dynamic equilibrium:**
macroscopic properties stay constant while microscopic reactions continue

# The Equilibrium Constant K: Expression, Meaning, and Manipulation

**1** Law of mass action: $K_c = [C]^c[D]^d / ([A]^a[B]^b)$; $K_P$ uses partial pressures

$$aA + bB \rightleftharpoons cC + dD$$

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

**2** $K_P = K_c(RT)^{\Delta n}$ links pressure-based and concentration-based constants

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

**3** Omit pure solids and liquids:
$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g),\ K = p_{CO_2}$



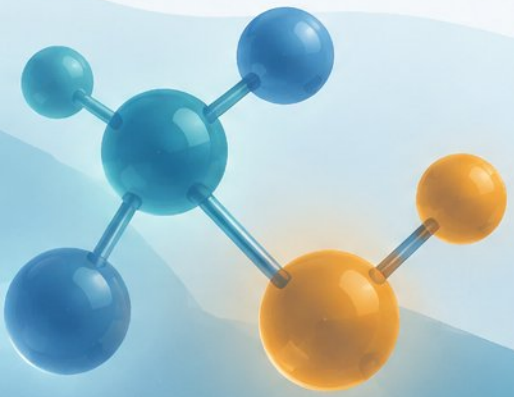
$CaCO_3(s)$ $CaO(s)$ $CO_2(g)$

**4** K >> 1 favors products; K << 1 favors reactants

**5** $K$ depends on temperature only — not concentration, pressure, or catalyst

**6** Reversing gives 1/K; multiplying coefficients by $n$ gives $K^n$

**7** $\Delta G° = -RT \ln K$ explains why $K$ is temperature-dependent



SAME REACTIONS.
A DEEPER
UNDERSTANDING.

# Reaction Quotient Q and the Three-Case Logic for Predicting Shift

- Q uses the K expression with current, non-equilibrium concentrations
- Q < K: reaction proceeds forward; Q > K: reaction proceeds in reverse
- Q = K: system is already at equilibrium
- Worked example: compute Q for $N_2O_4/NO_2$ or Haber–Bosch, then compare with K
- Q versus K is the rigorous check behind every Le Chatelier prediction



# ICE Tables and Quantitative Equilibrium Calculations

CHEMISTRY BUILDS BRIGHTER UNDERSTANDING

## Example: Set up an ICE table

$$A(g) + B(g) \rightleftharpoons C(g)$$

| | A(g) | B(g) | C(g) |
|---|---|---|---|
| Initial | $a$ | $b$ | $c$ |
| Change | $-x$ | $-x$ | $+x$ |
| Equilibrium | $a - x$ | $b - x$ | $c + x$ |

### Substitute into the equilibrium expression

$$K = \frac{c + x}{(a - x)(b - x)}$$

↓

Substitute equilibrium terms into $K$ and solve for $x$

### Small-x approximation

Valid only if $x < \sim 5\%$ of initial

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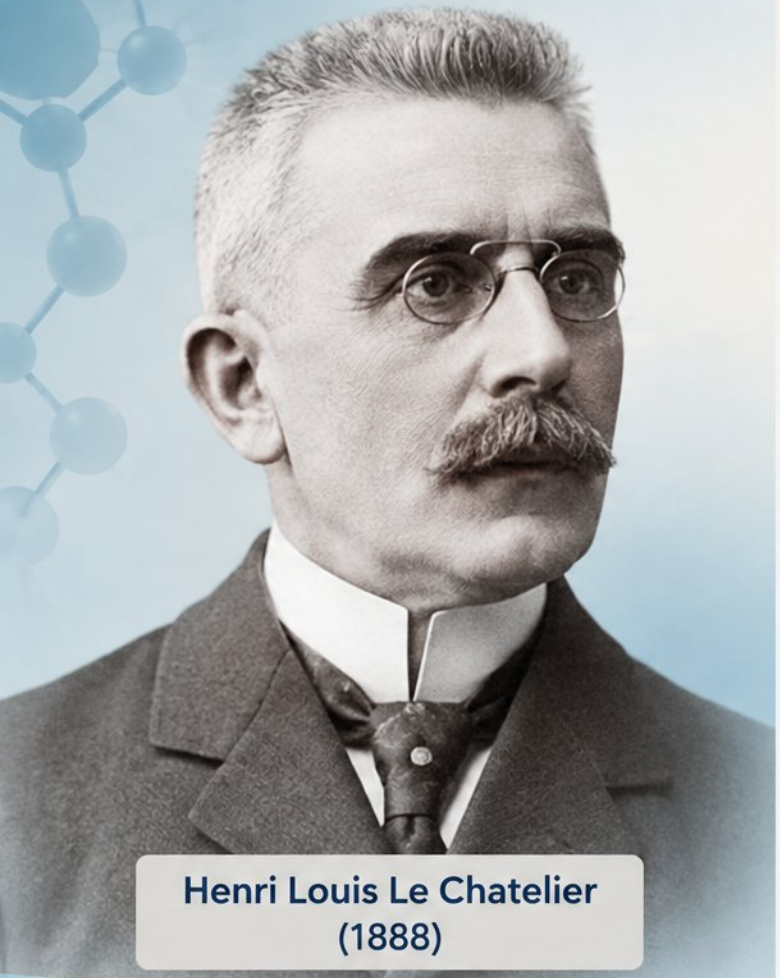
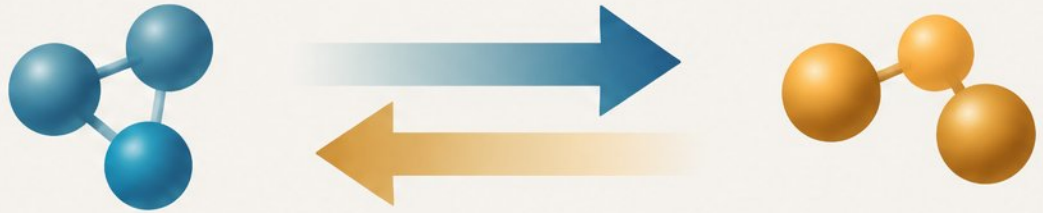
- ICE table: Initial → Change → Equilibrium, expressed through $x$
- Substitute equilibrium terms into $K$ and solve for $x$
- Small-x approximation valid only if $x < \sim 5\%$ of initial
- The Change step mirrors $Q$ vs. $K$ direction
- Always interpret what $x$ physically represents

# Le Chatelier's Principle: Statement, Logic, and Honest Caveats



Henri Louis Le Chatelier
(1888)

**Statement:** a disturbed equilibrium shifts to partially counteract the change



## Key Points and Caveats

- A qualitative heuristic, not a law; it restores Q = K, never fully
- K stays constant at fixed temperature; only equilibrium position moves
- Wording misleads for pressure and temperature changes
- Formulated by Le Chatelier in 1888; later authors documented ambiguity

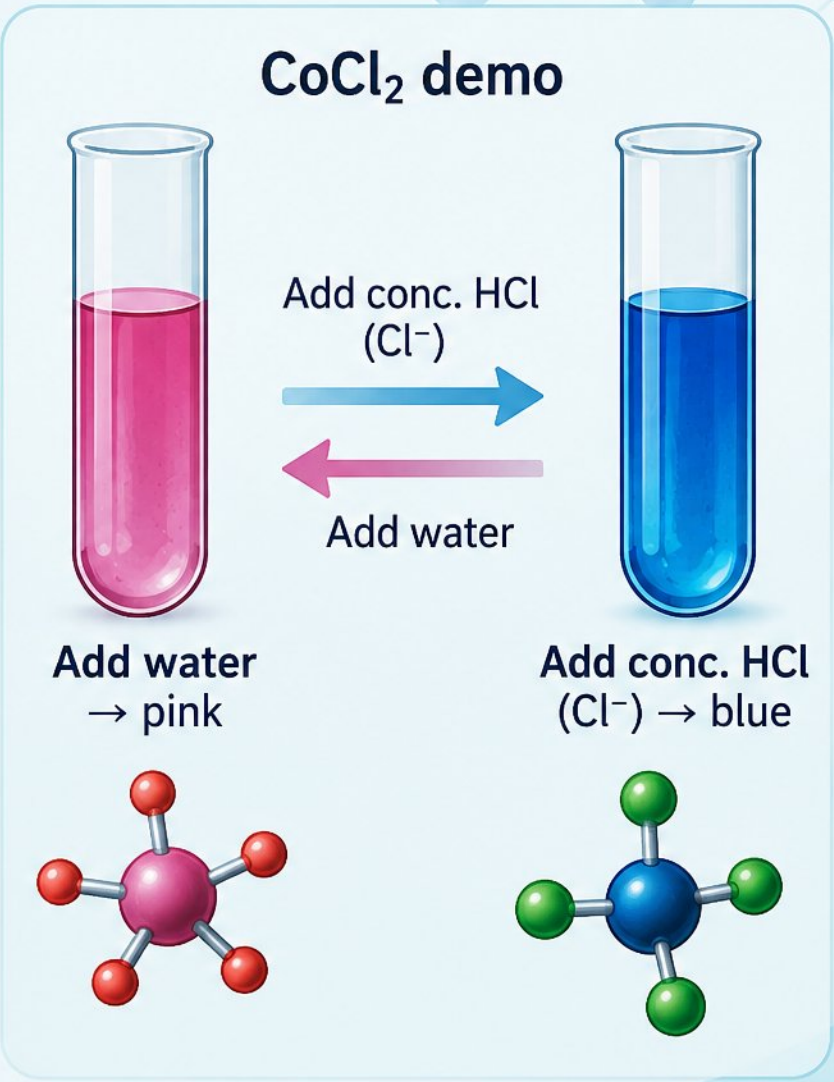
EQUILIBRIUM CONNECTS CHANGE

SAME PRINCIPLES.
BRIGHTER UNDERSTANDING.

# Concentration Effects, Explained Through Q versus *K*

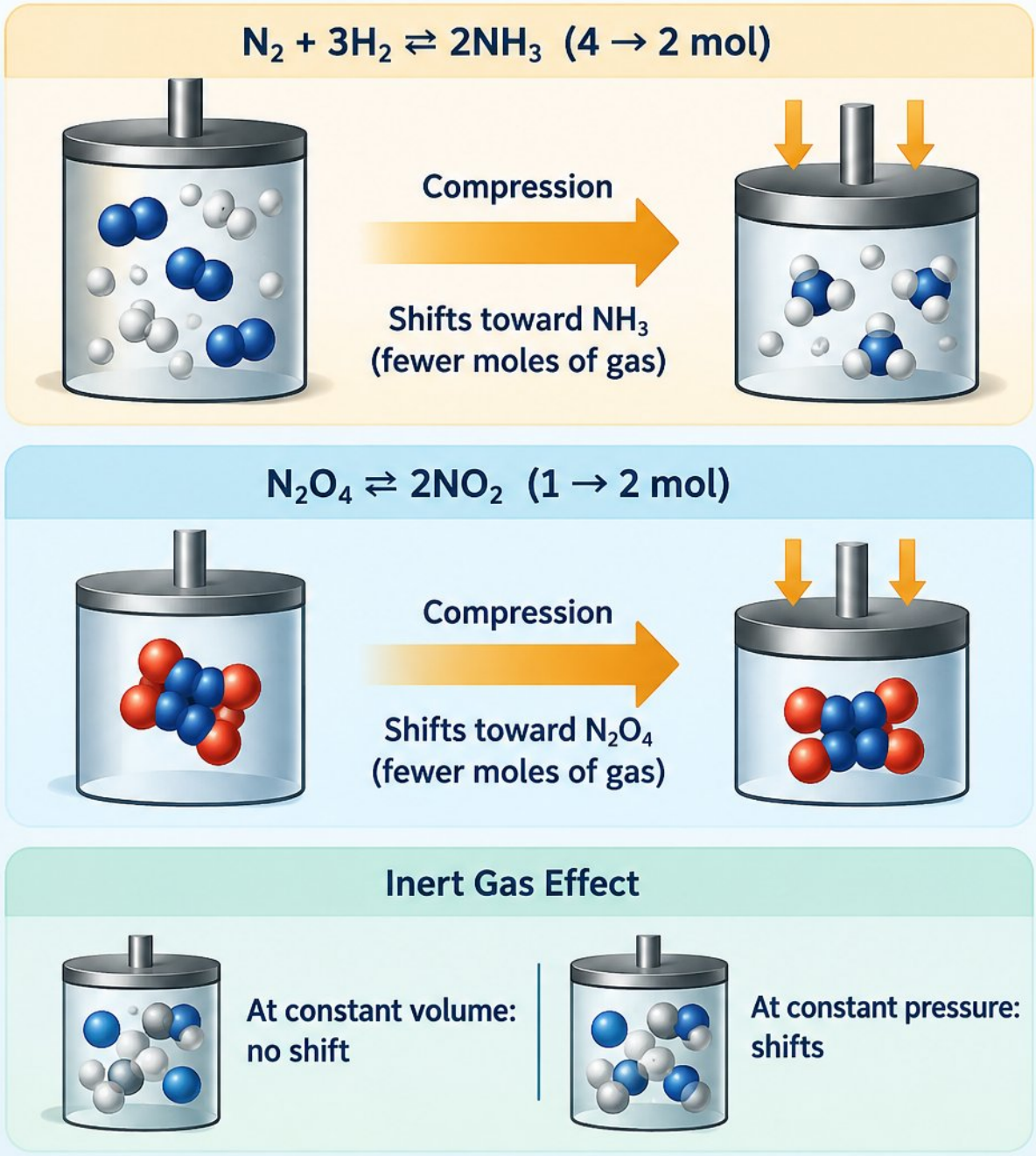
- Adding a reactant raises *Q*'s denominator → $Q < K$ → shifts forward
- Removing a product lowers *Q*'s numerator → $Q < K$ → also shifts forward
- Haber–Bosch: extra $N_2$ or $H_2$, or condensing out $NH_3$, favours more ammonia
- $CoCl_2$ demo: add conc. HCl ($Cl^-$) → blue; add water → pink
- Key point: the equilibrium position moves, but *K* stays constant

**The equilibrium position moves, but *K* stays constant.**



# Pressure, Volume, and Inert Gas Effects

- Compression shifts gas equilibria toward fewer moles of gas
- $\Delta n(\text{gas}) = 0$? Then pressure changes do almost nothing
- Haber–Bosch (4 → 2 mol): compression favors $NH_3$
- $N_2O_4 \rightleftarrows 2NO_2$ (1 → 2 mol): compression favors $N_2O_4$
- Inert gas at constant volume: no shift; at constant pressure: shifts
- Trust $Q$ versus $K$, not "counteract the volume change" wording



# Temperature Effects and the Distinct Role of Catalysts

- Temperature is the only common variable that changes K itself.
- Heating favors the endothermic direction; cooling favors the exothermic direction.
- $N_2O_4 \rightleftarrows 2NO_2$: hot bath deepens brown $NO_2$; ice bath fades it.
- Cobalt chloride: hot water → blue $[CoCl_4]^{2-}$; ice water → pink $[Co(H_2O)_6]^{2+}$.
- Catalysts speed up both directions equally — no shift in equilibrium position or K.
- Catalysts are industrially valuable because they shorten time to equilibrium.

## $N_2O_4 \rightleftarrows 2NO_2$



**Hot bath**
deepens brown $NO_2$

Ice bath
fades it

## Cobalt chloride

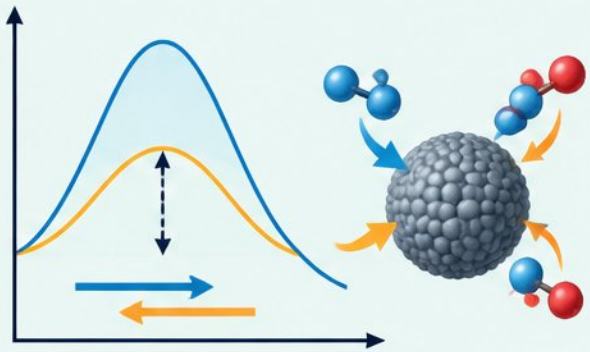


**Hot water →**
blue $[CoCl_4]^{2-}$

**Ice water →**
pink $[Co(H_2O)_6]^{2+}$

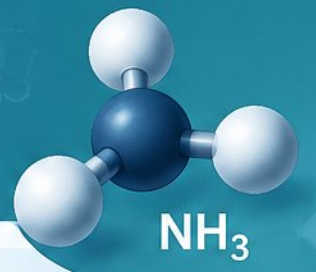
## Catalysts

Speed up both directions equally
— no shift in equilibrium position or K.



Catalysts are industrially valuable
because they shorten time to equilibrium.

# Industrial Case Study:
# Haber–Bosch Ammonia Synthesis



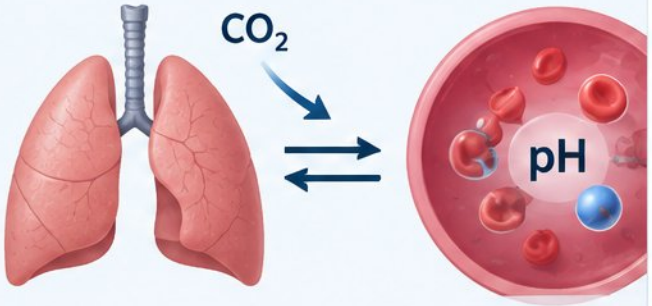

- $N_2 + 3H_2 \rightleftarrows 2NH_3$, $\Delta H = -92.4$ kJ/mol; exothermic, fewer gas moles
- Equilibrium favours high pressure and low temperature; kinetics favours high temperature
- Compromise: 400–500 °C (723–873 K), 100–400 bar, promoted iron ($K_2O$, $Al_2O_3$)
- Single-pass conversion ~15–25%; $NH_3$ condensed out, unreacted gases recycled → ~97% overall
- Discussion: which variable makes the largest practical difference, and why?



# Biological and Environmental Equilibrium Cases

## Bicarbonate Buffer

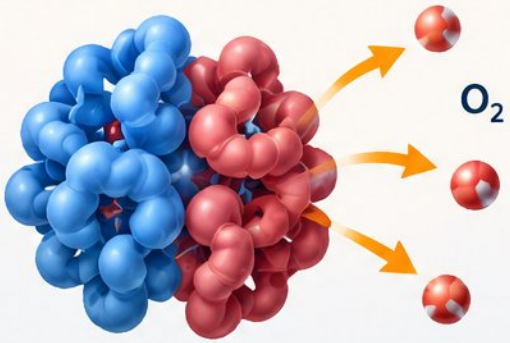
$CO_2 + H_2O \rightleftarrows H_2CO_3 \rightleftarrows H^+ + HCO_3^-$




Breathing adjusts blood pH

## Bohr Effect

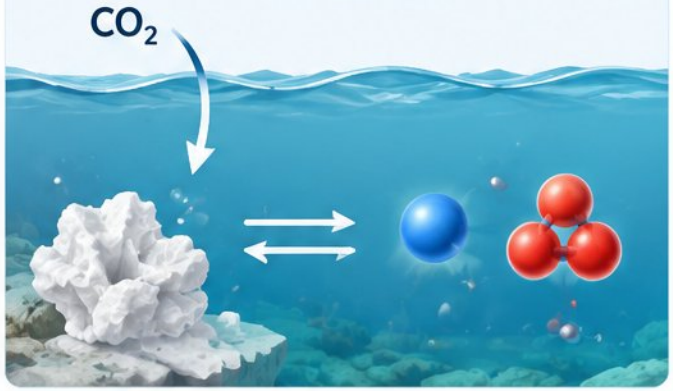
$CO_2$, low pH, heat release more $O_2$ to active tissue




## Ocean $CO_2$ and carbonate equilibrium

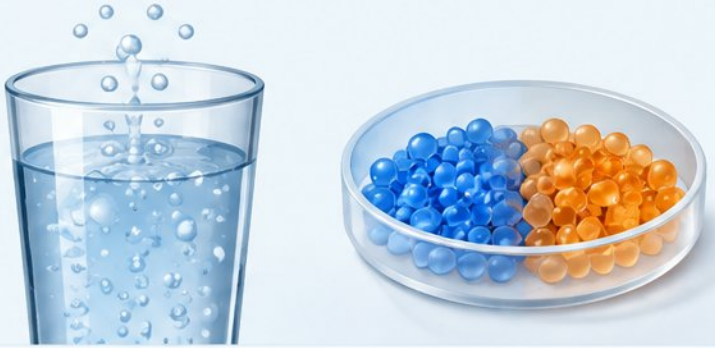
$CaCO_3 \rightleftarrows Ca^{2+} + CO_3^{2-}$

Acid deposition



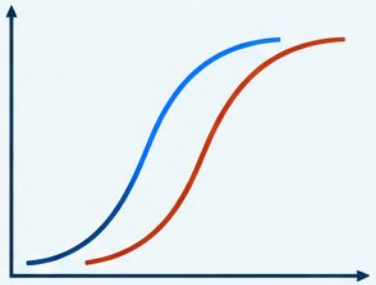

## Everyday

Carbonated drinks losing fizz; silica gel desiccant color change



## Data tasks

Hemoglobin saturation vs pH; $NH_3$ yield vs temperature and pressure



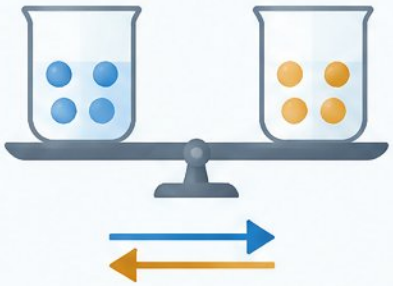


# Common Misconceptions and Diagnostic Teaching

Clearer ideas. Stronger reasoning. Deeper understanding.

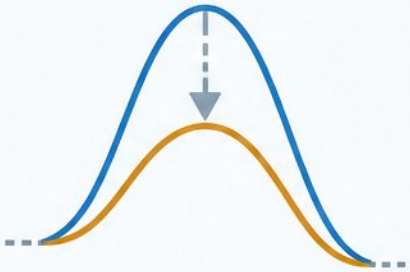
### “Equillibrium means equal concentrations”

— equal rates, not equal amounts



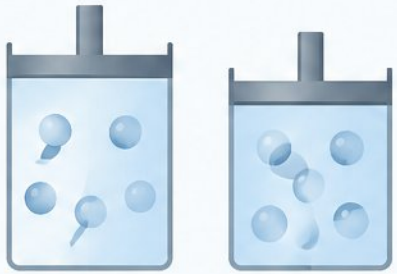
### “A catalyst increases yield”

— catalysts change rate, never *K* or position



### “Pressure always shifts toward fewer moles”

— fails when Δn(gas) = 0



### “K changes with concentration”

— *K* depends on temperature only




**Diagnostic routine:**
predict first, collect rationales, then check with Q versus K



# Classroom Demonstrations, Simulations, and Assessment Design

OBSERVE
REASON
APPLY

## $NO_2/N_2O_4$ Sealed Tubes



Hot water (70–80 °C) darkens gas

Ice water fades it

Use fume hood.

## Cobalt chloride demo



Pink $CoCl_2(aq)$ plus concentrated HCl gives violet control; test heat, ice, added HCl, added water.



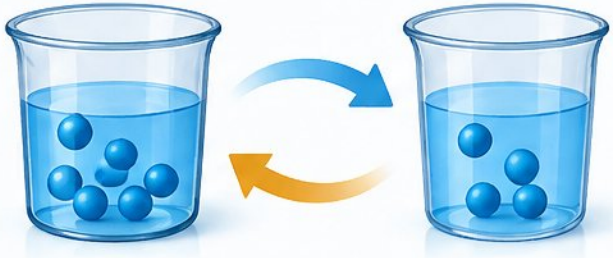
## Safety and Disposal




- Dispose of cobalt waste as hazardous waste; neutralize acidic solutions before disposal.



## Bucket-and-Beaker Role-Play



Bucket-and-beaker role-play shows dynamic equilibrium and approach before equations appear.

## Teaching Sequence

1. Qualitative observation
2. Q/K reasoning
3. ICE calculation
4. Le Chatelier's limits

## Assessment

- Conceptual multiple-choice items
- Predict-and-justify tasks
- Short ICE problems
- Every prediction checked with Q versus K.

# Wrap-Up: Key Takeaways and Next Steps




## Five Takeaways

1. dynamic equilibrium (equal rates, not equal amounts)
2. *K* depends on temperature only
3. *Q* versus *K* directs shift rigorously
4. Pressure/volume and temperature expose Le Chatelier's limits; catalysts change rate, never *K*
5.

## Q vs K: Direction of Shift




## Self-Check Quiz

Self-check quiz: prediction, *Q* calculation, temperature effect, catalyst item

## Next

Next: acid–base equilibria, $K_{sp}$, $\Delta G^\circ = -RT \ln K$

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