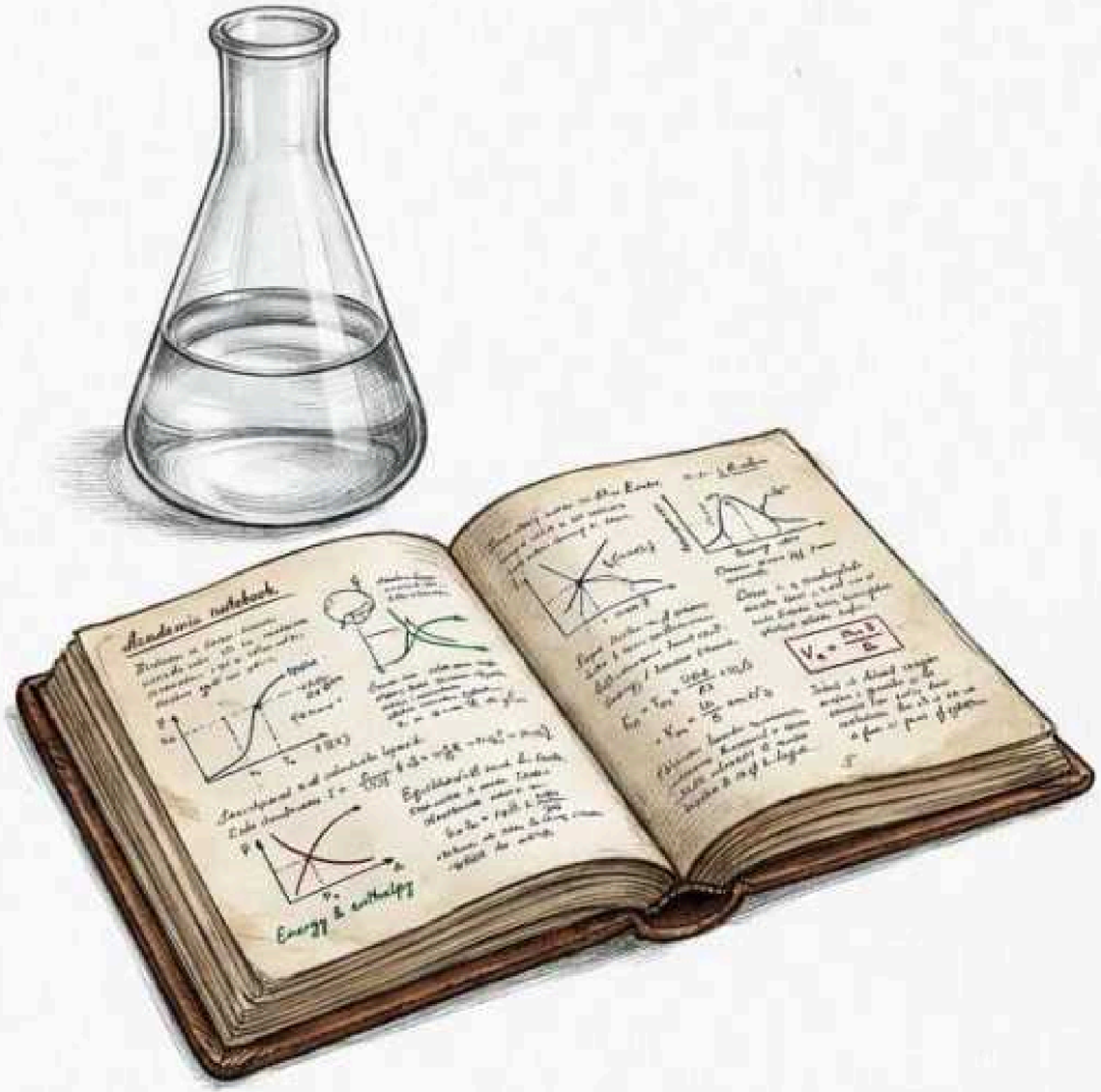
¶ 6.1



CHEM 101 B: Lecture 6

Energy, Equilibrium, Acids and Bases

Gary Anderson | Anderson Academy

At a Glance:

- Energy & enthalpy
- Exothermic & endothermic reactions
- Calorimetry & bond energies
- Collision theory, rates & catalysts
- Equilibrium, K & Le Chatelier
- Acids, bases, pH & buffers

¶ 6.2

What we know:
Composition,
Bonding,
Stoichiometry

→

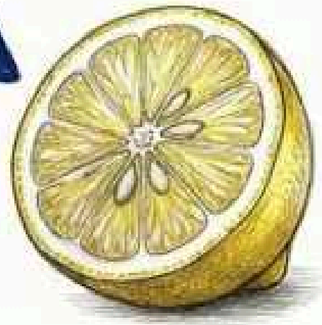
The Final Questions



How much energy does a reaction release or absorb?



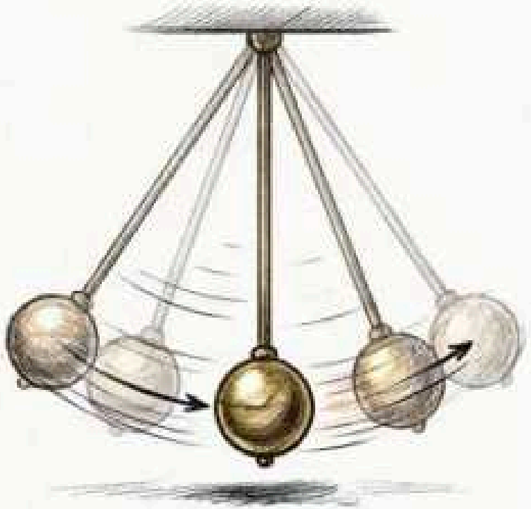
How fast does it go, and does it ever stop?



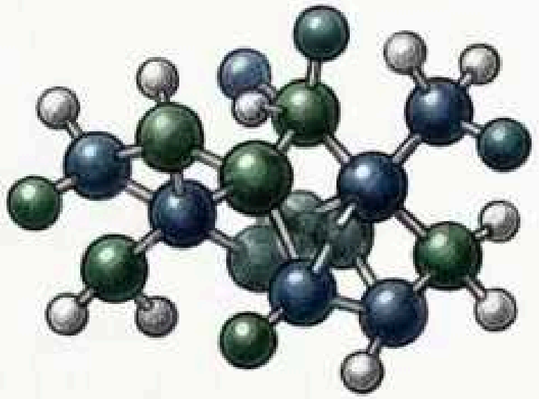
What makes a substance an acid or a base?

These ideas tie the whole course together.

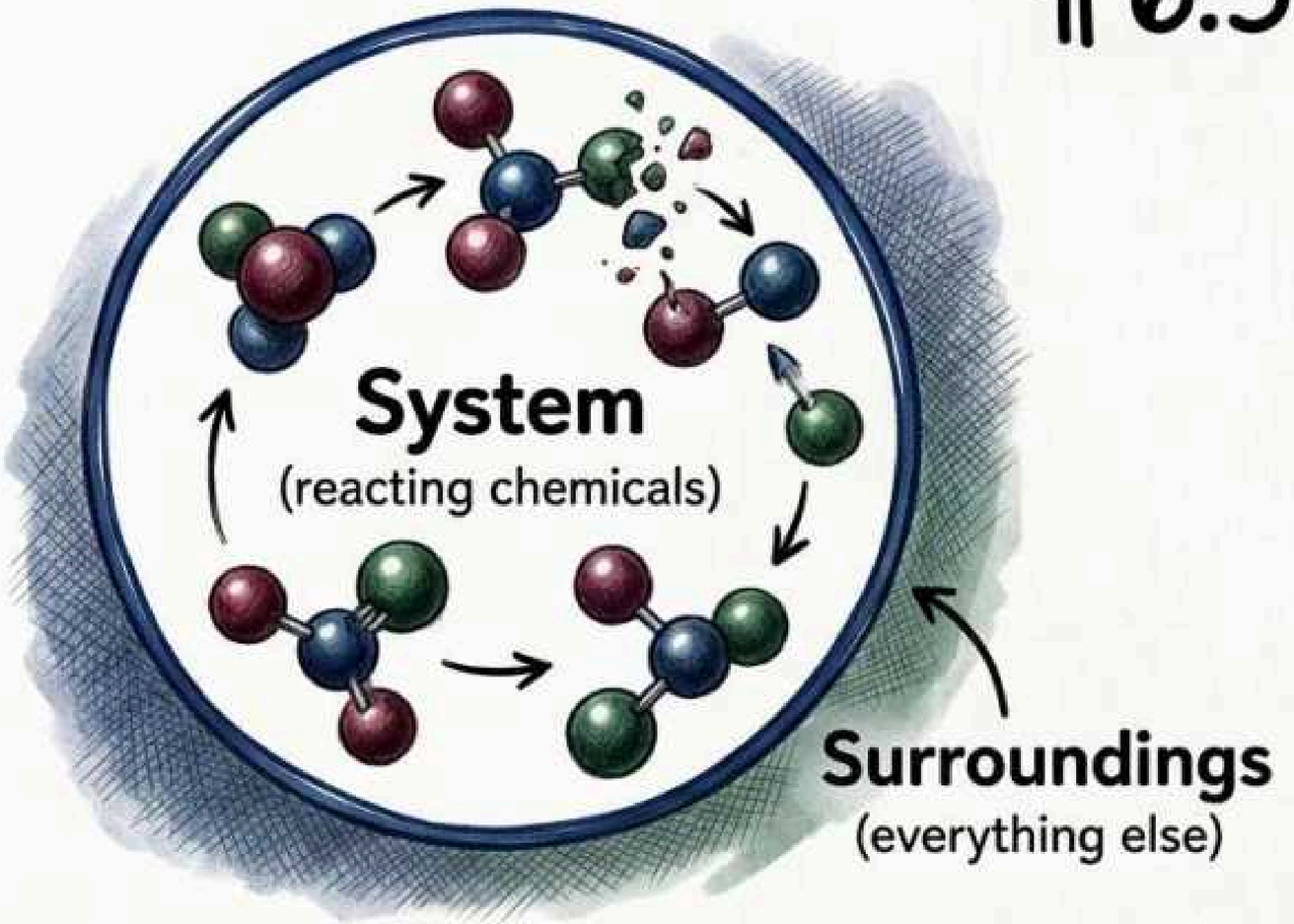
¶ 6.3



Kinetic Energy
(Energy of motion)



Potential Energy
(Stored energy due to position or arrangement)



1st Law of Thermodynamics:
Energy is never created or destroyed, only moved around.

¶ 6.4

Exothermic



Releases heat to surroundings.

Enthalpy (ΔH): Negative (-)

Example: Burning 1 mole of methane (CH_4) releases ≈ 890 kJ.

Endothermic



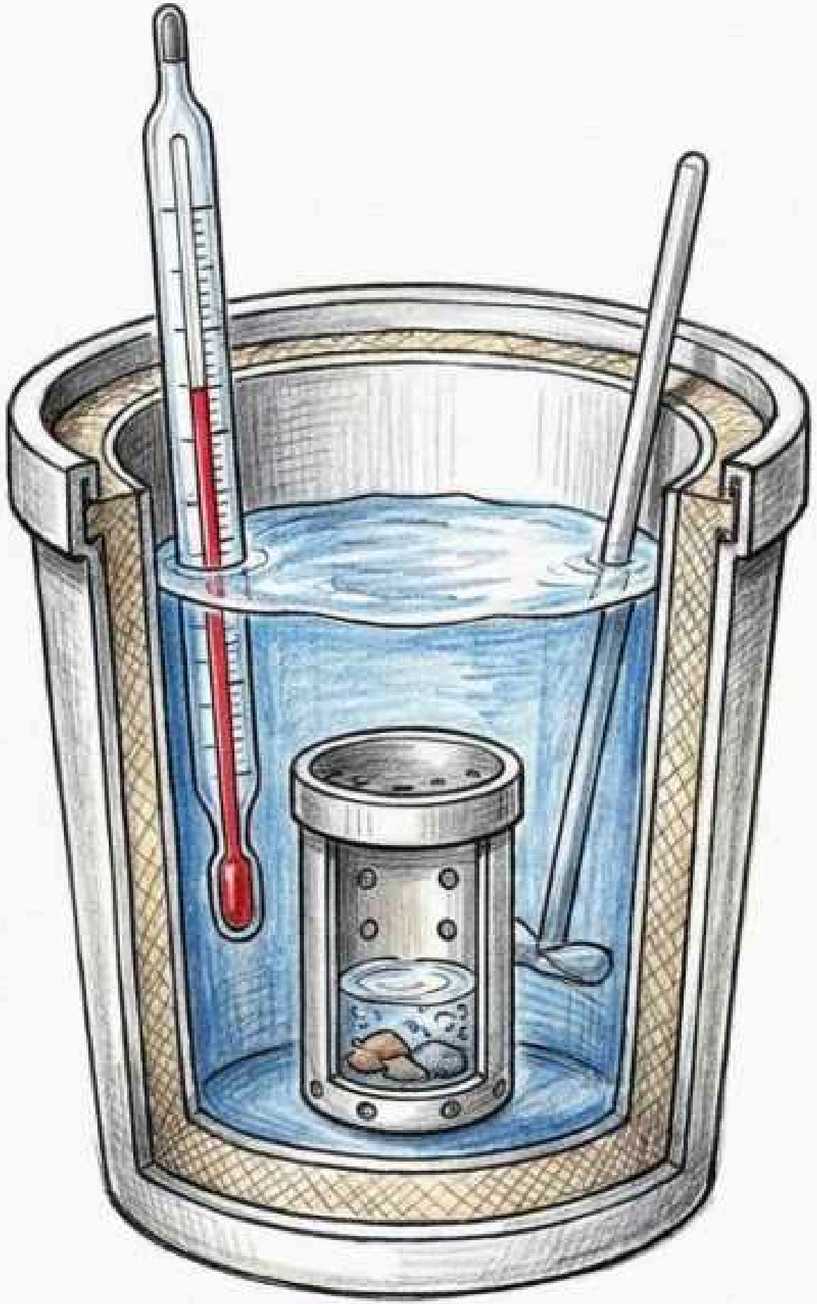
Absorbs heat from surroundings.

Enthalpy (ΔH): Positive (+)

Example: Dissolving 1 mole of ammonium nitrate (NH_4NO_3) absorbs ≈ 25 kJ.

Enthalpy change (ΔH) = Heat change at constant pressure.

¶ 6.5

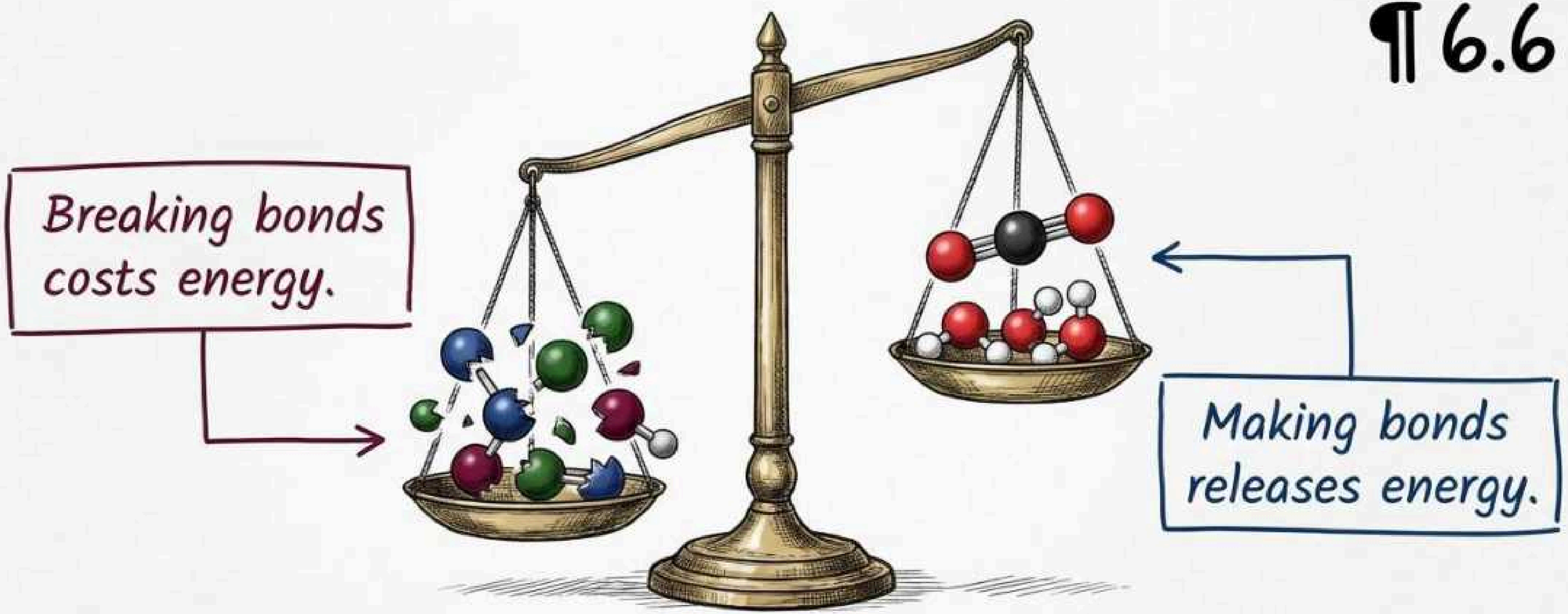


$$q = m \cdot c \cdot \Delta T$$

- q = Heat absorbed
- m = Mass (e.g., 100g)
- c = Specific heat capacity (Water = 4.184 J/g·°C)
- ΔT = Temperature change (e.g., 25°C)

Calculation: $100 \times 4.184 \times 25 = 10{,}460$ Joules (≈ 10.5 kJ)

¶ 6.6

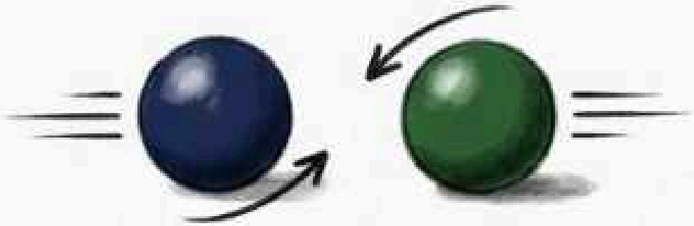


Core Principle: If the bonds formed in products are stronger than bonds broken in reactants, extra energy leaves as heat (Exothermic).

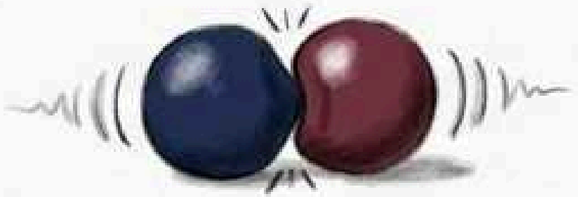
Real-World Application: Burning fuel. Carbon dioxide and water have exceptionally strong bonds, resulting in massive energy release.

Chemical Kinetics: Collision Theory

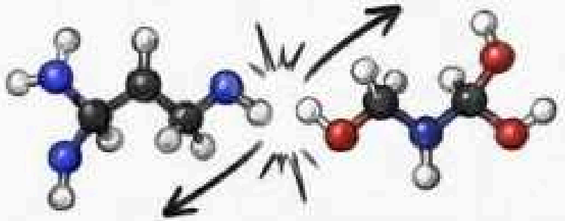
1. Must collide.



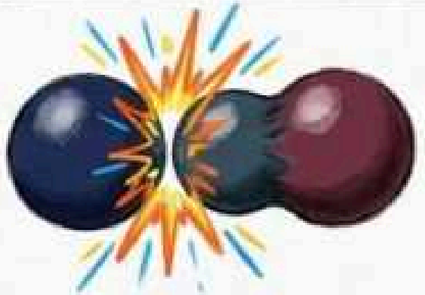
2. Must have enough energy.



3. Right orientation.



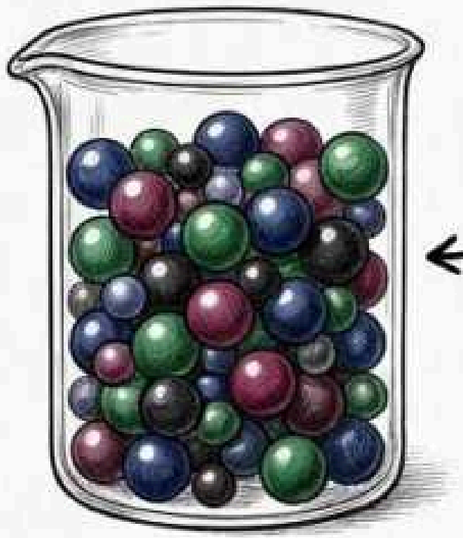
Success State:



Insight: Favorable reactions (like wood burning) need a spark because most initial collisions fail these three tests.

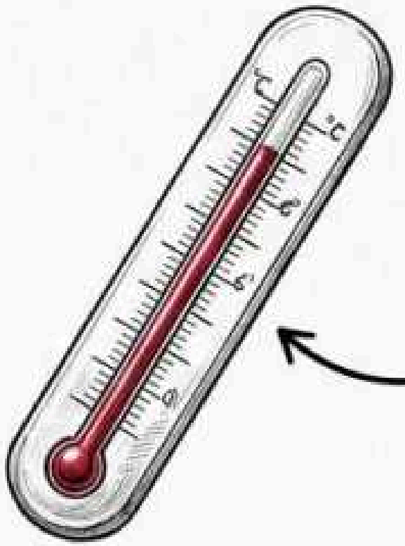
Four Factors That Speed Up Reactions

¶ 6.8



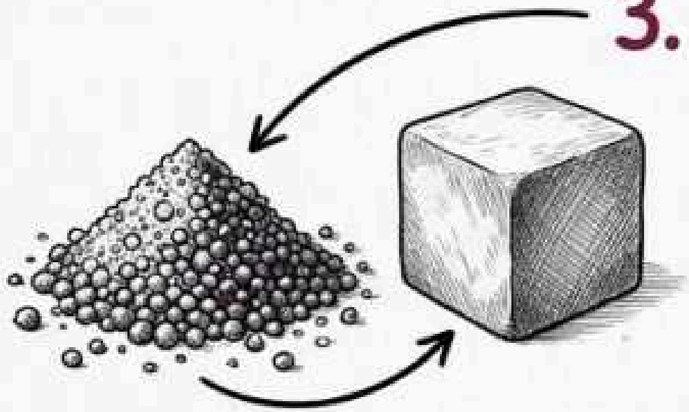
1. Concentration

More particles = more collisions.



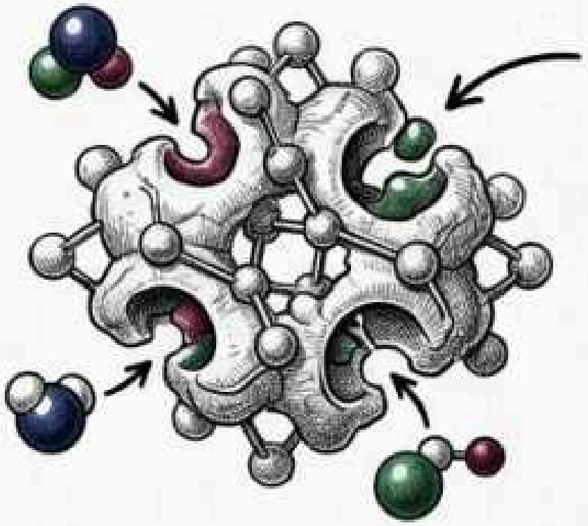
2. Temperature

Faster particles = more energetic collisions. (Rule: +10°C roughly doubles the rate).



3. Surface Area

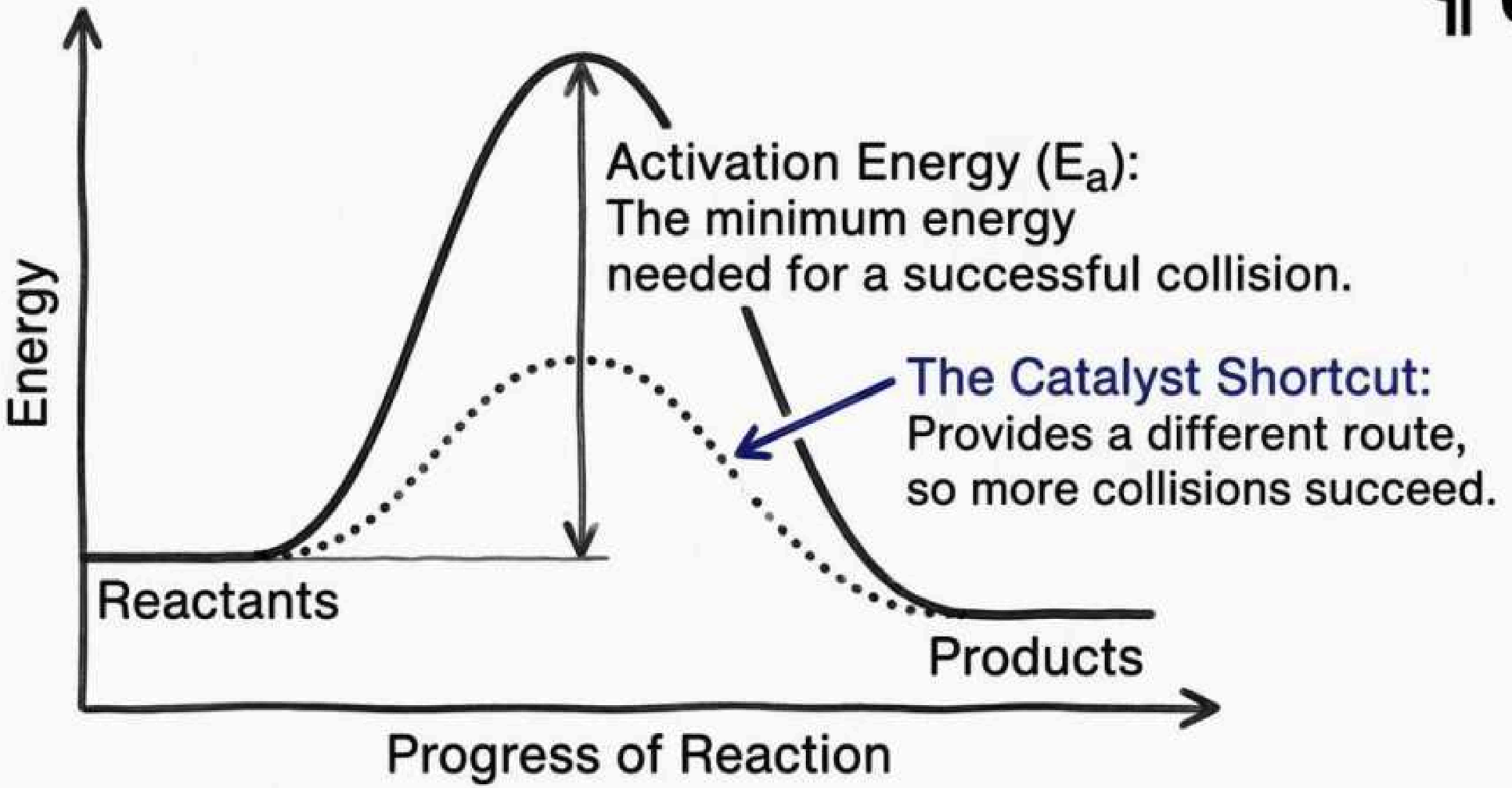
Powders expose far more reacting particles than a lump.



4. Catalyst

Speeds up the reaction without being used up itself.

¶ 6.9



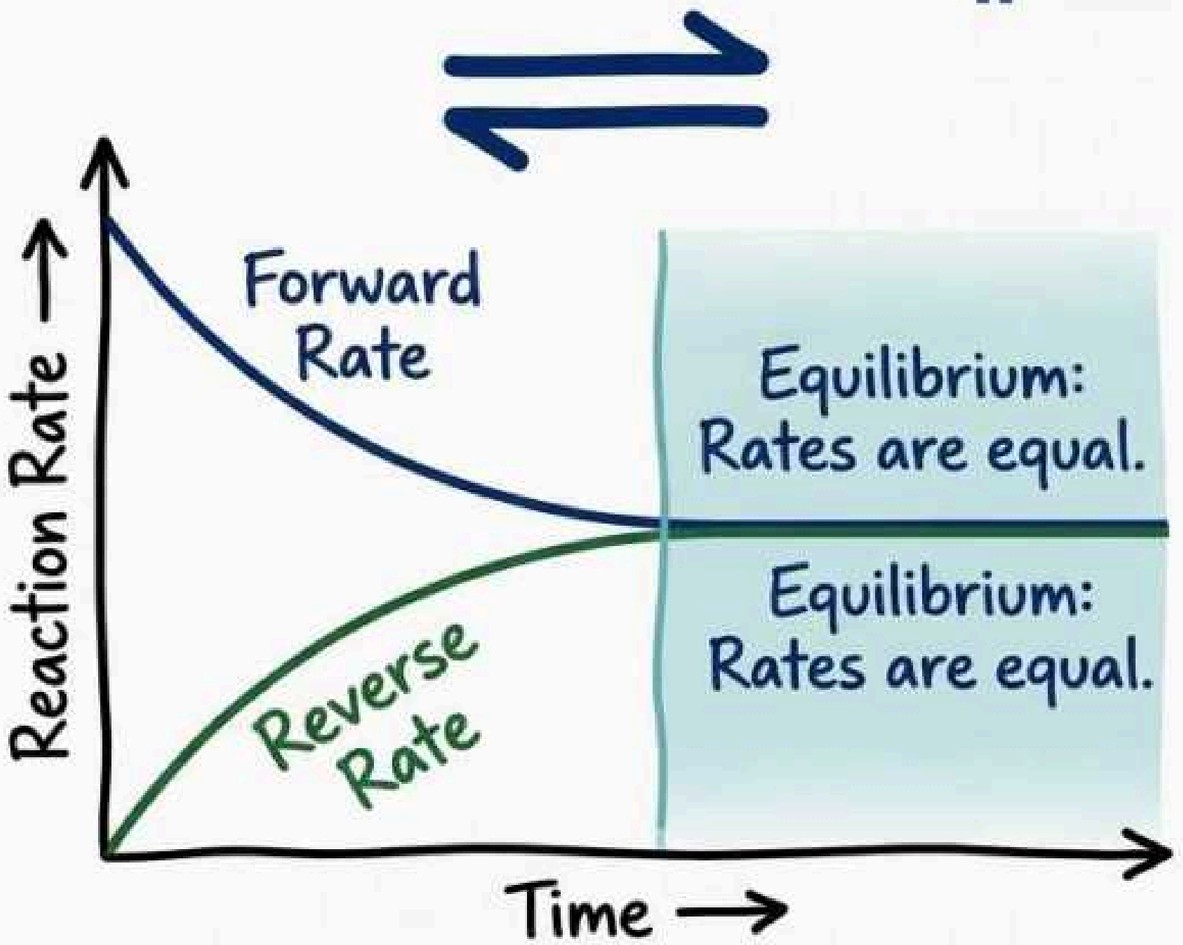
Real-World Catalysts:

- Biological: Enzymes make life's reactions possible at body temp.
- Mechanical: Catalytic converters use metals to clean car exhaust.

¶ 6.10

Key Principles:

- **Reversible:** Products turn back into reactants.
- **Convergence:** Forward slows as reactants are used; reverse speeds up as products build.
- **Equilibrium:** Reached when forward and reverse rates are equal.



Crucial Note: Equilibrium is dynamic, not static. Both reactions continue, but concentrations stop changing. Equal rates do not mean equal amounts!

The Equilibrium Constant (K)

¶ 6.11

$$N_2O_4 \rightleftharpoons 2NO_2$$

$$K = \frac{[\text{Products}]_{\text{coeff}}}{[\text{Reactants}]_{\text{coeff}}}$$

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

Calculation Example:

$[NO_2] = 0.12\ M$, $[N_2O_4] = 0.040\ M$

$$K = \frac{(0.12)^2}{0.040} = 0.36$$

¶ 6.12

Le Chatelier's Rule: The system shifts to oppose a disturbance.

Context: Large K = Products favoured. Small K = Reactants favoured.

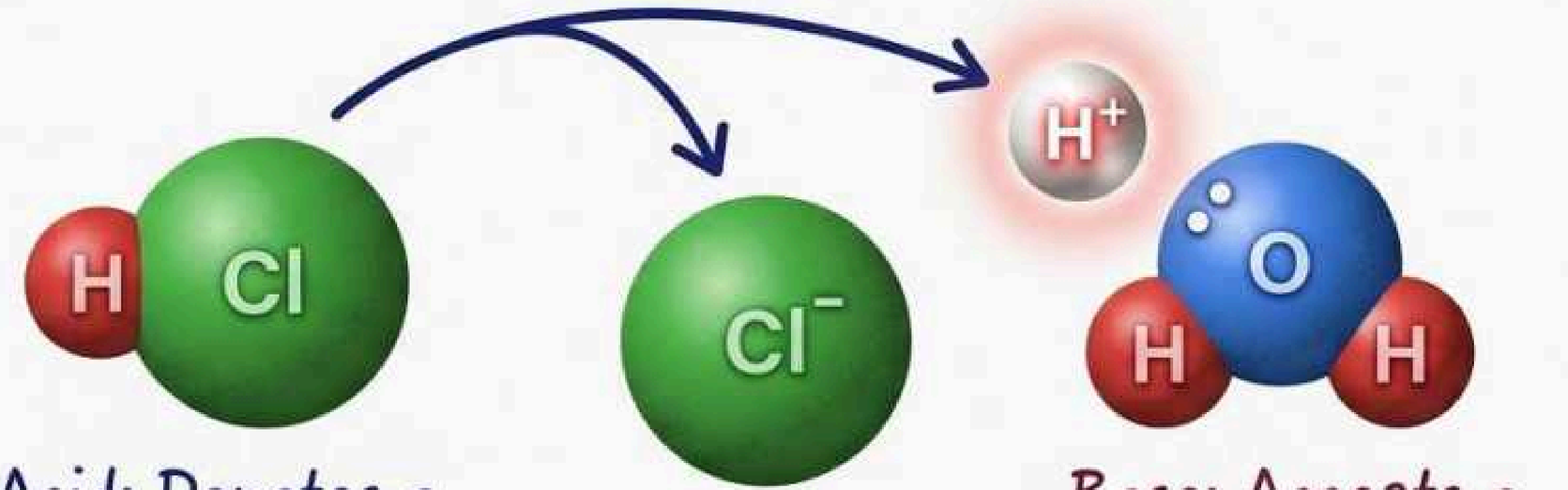


- Heat an endothermic reaction → Shifts toward products.
- Compress a gas mixture → Shifts toward fewer gas molecules.

Application: Industry constantly uses this rule to raise chemical yields.

§ 6.13

Brønsted-Lowry Definition: Acids & Bases



Acid: Donates a hydrogen ion (H^+)

Base: Accepts a hydrogen ion (H^+)

Old Clues vs. Modern Safety:

- Sour taste (acids) and slippery feel (bases) are historical clues.
- Never taste a chemical. Always use an indicator or a meter.

Water Ionization & pH Math

¶ 6.14

The Baseline: Even pure water ionizes slightly: $[H^+] = 1 \times 10^{-7}$ M (at room temp).

$$pH = -\log_{10}[H^+]$$

The Translation: The negative base-10 logarithm turns tiny, difficult numbers into easy-to-handle whole numbers.

Example 1

$[H^+] = 0.001\ M$

$pH = 3$

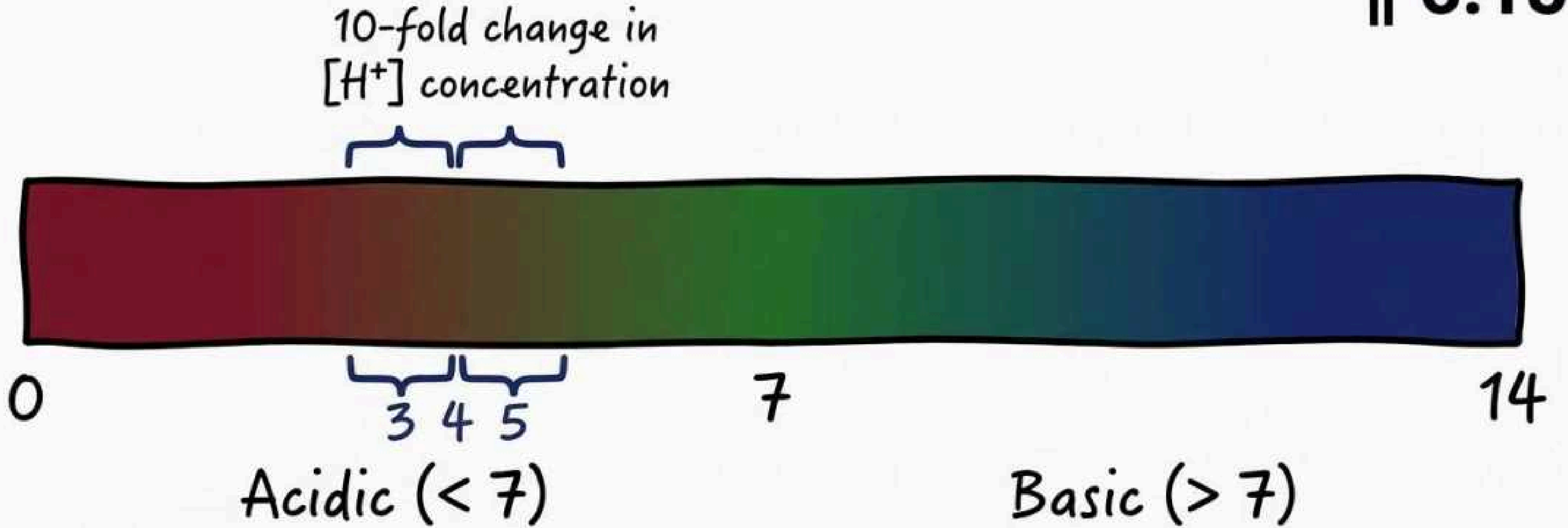
Example 2

Pure water:

$[H^+] = 1 \times 10^{-7}\ M$

$pH = 7$

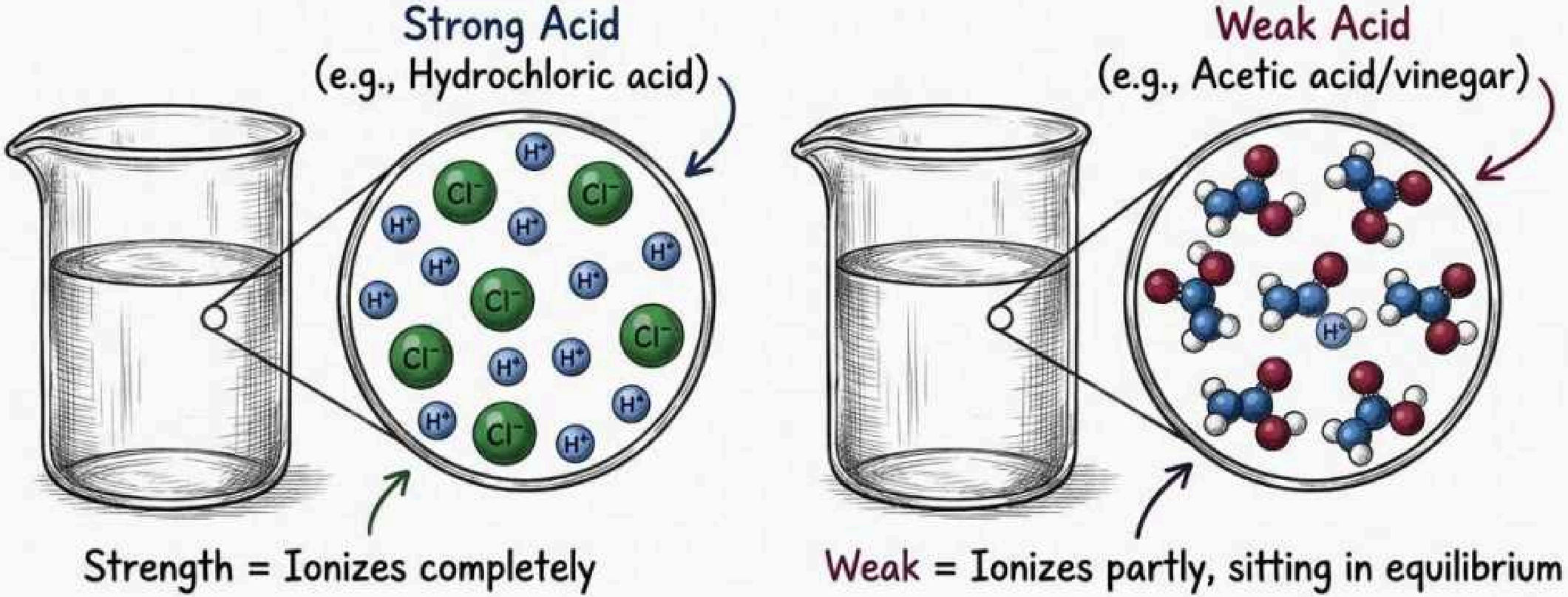
¶ 6.15



Logarithmic Steps:

- Every 1 pH unit = 10x jump in concentration.
- Example: pH 3 is 100x more acidic than pH 5.
- Example: A 0.01 M solution of hydrochloric acid has a pH of 2.

Strength vs. Concentration



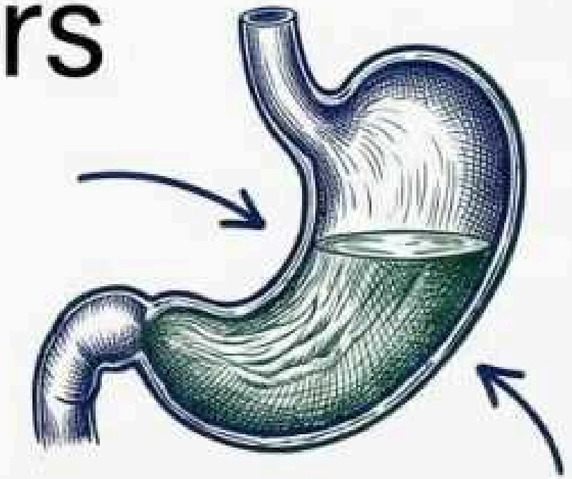
Crucial Distinction: Strength is how fully it ionizes. Concentration is how much total substance is dissolved.

The Partner Scale: pH + pOH = 14 (e.g., A pH of 11 means a pOH of 3)

¶ 6.17

Neutralization and Buffers

Acid + Base → Water + Salt

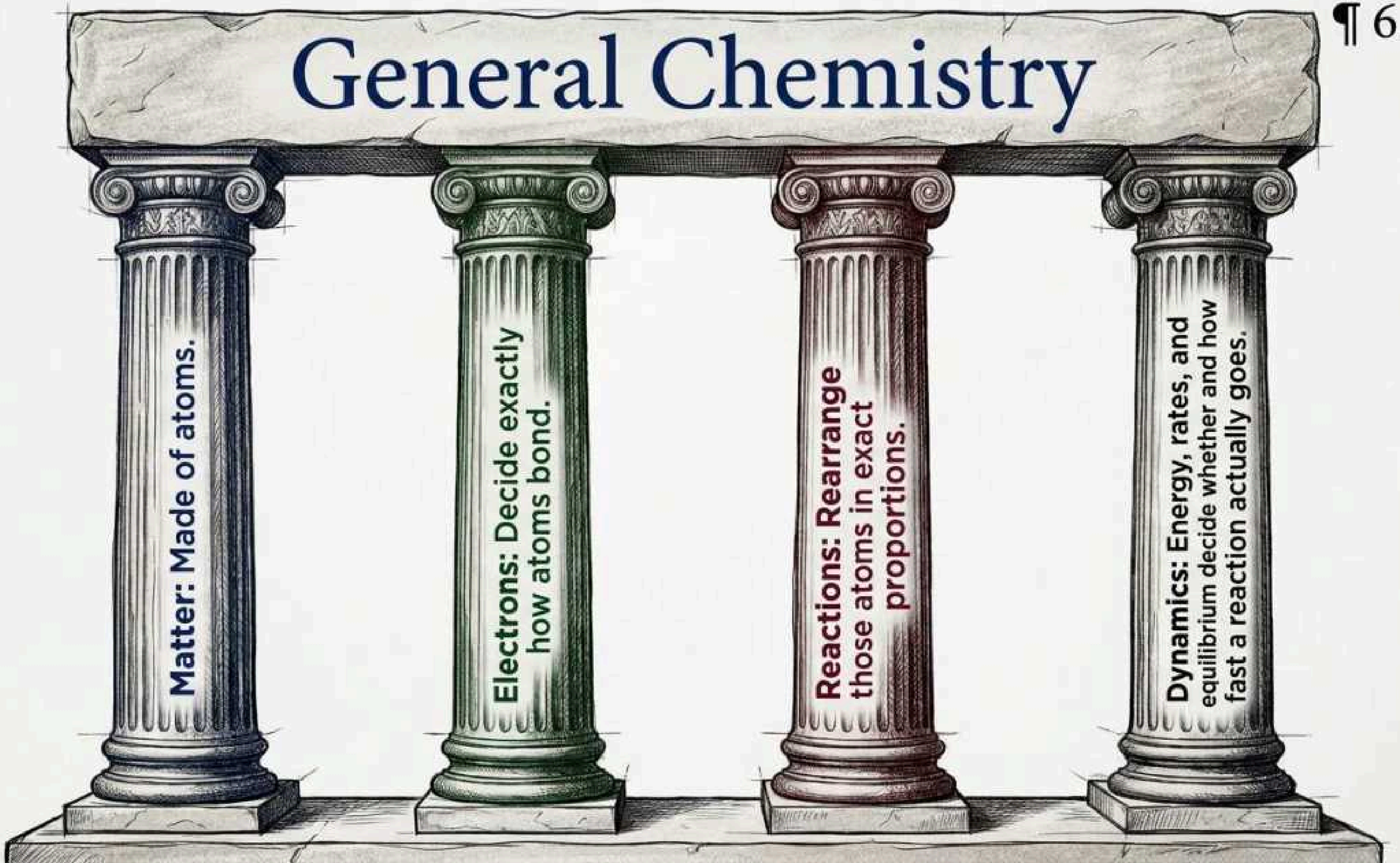


Application: Antacid tablets calming excess stomach acid.



Buffers: A mixture of a weak acid and its partner that actively resists changes in pH.

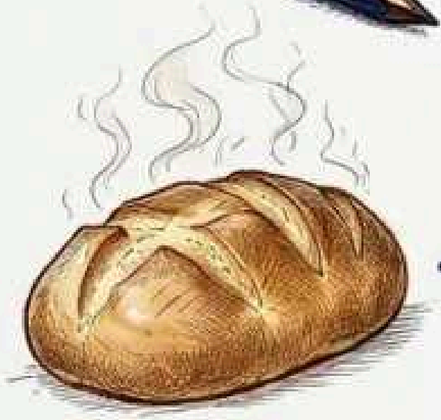
Biological Imperative: Human blood is buffered at ≈ 7.4. Even a small drift from this range can be dangerous.


General Chemistry
Matter: Made of atoms.
Electrons: Decide exactly how atoms bond.
Reactions: Rearrange those atoms in exact proportions.
Dynamics: Energy, rates, and equilibrium decide whether and how fast a reaction actually goes.

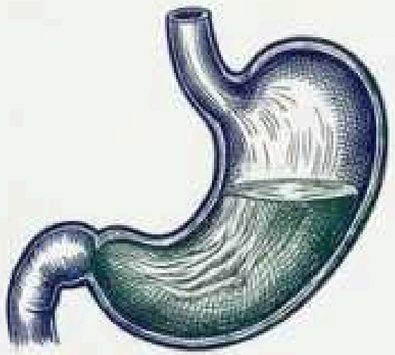
The Final Question



← Pick a reaction from daily life.



← Which of these course ideas does it use?



← Write it down, and keep it.

Contact

Thank you for taking Chemistry 101 B.

Questions or feedback:
garyanderson@mail.grokbot.com

X (Twitter): @garyanderson

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