Chemistry Fundamentals

Lecture 17: Introduction to Chemical Reactions

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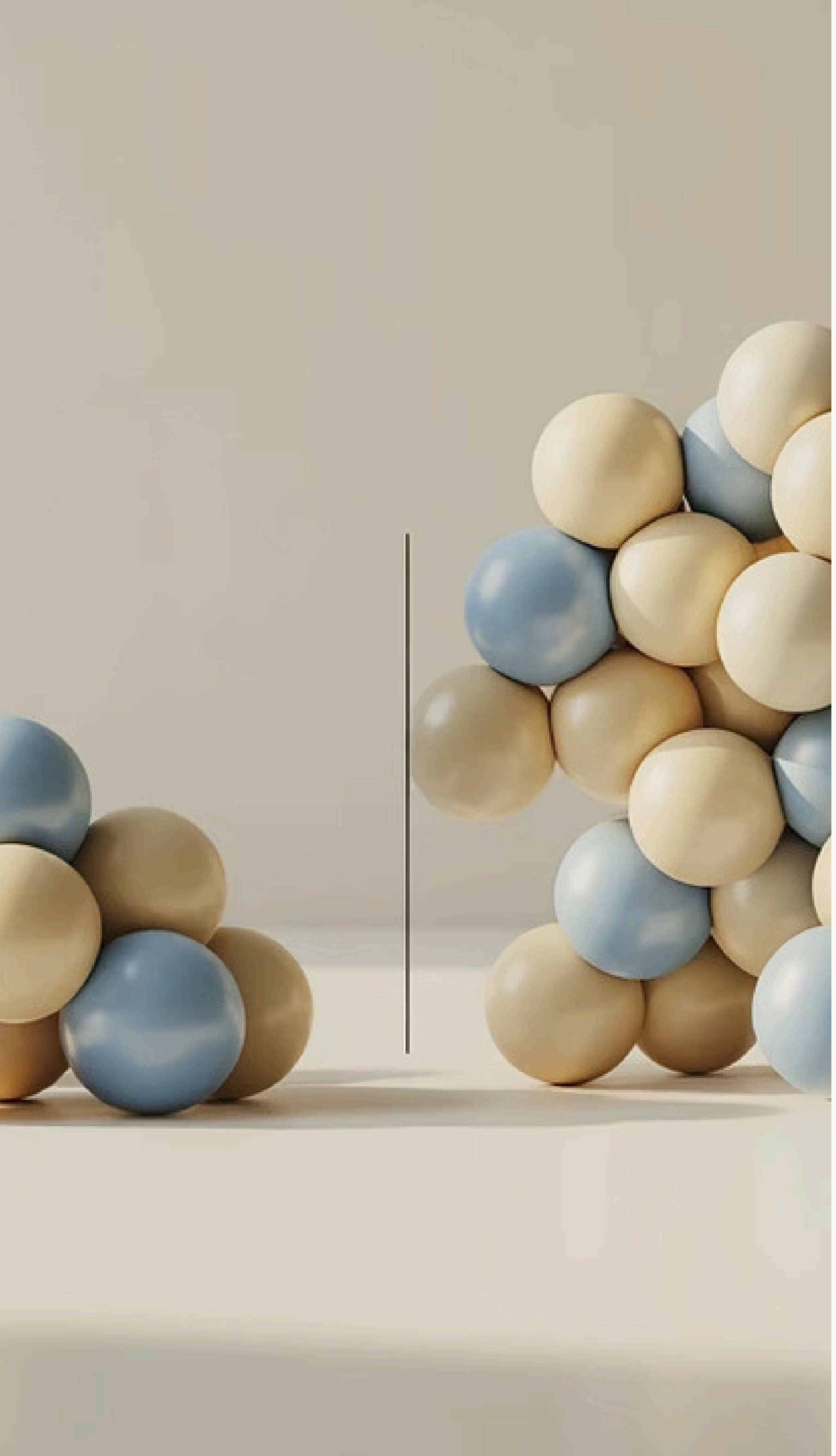


What Are Chemical Reactions?

Definition

Process where reactants are transformed into products with different properties

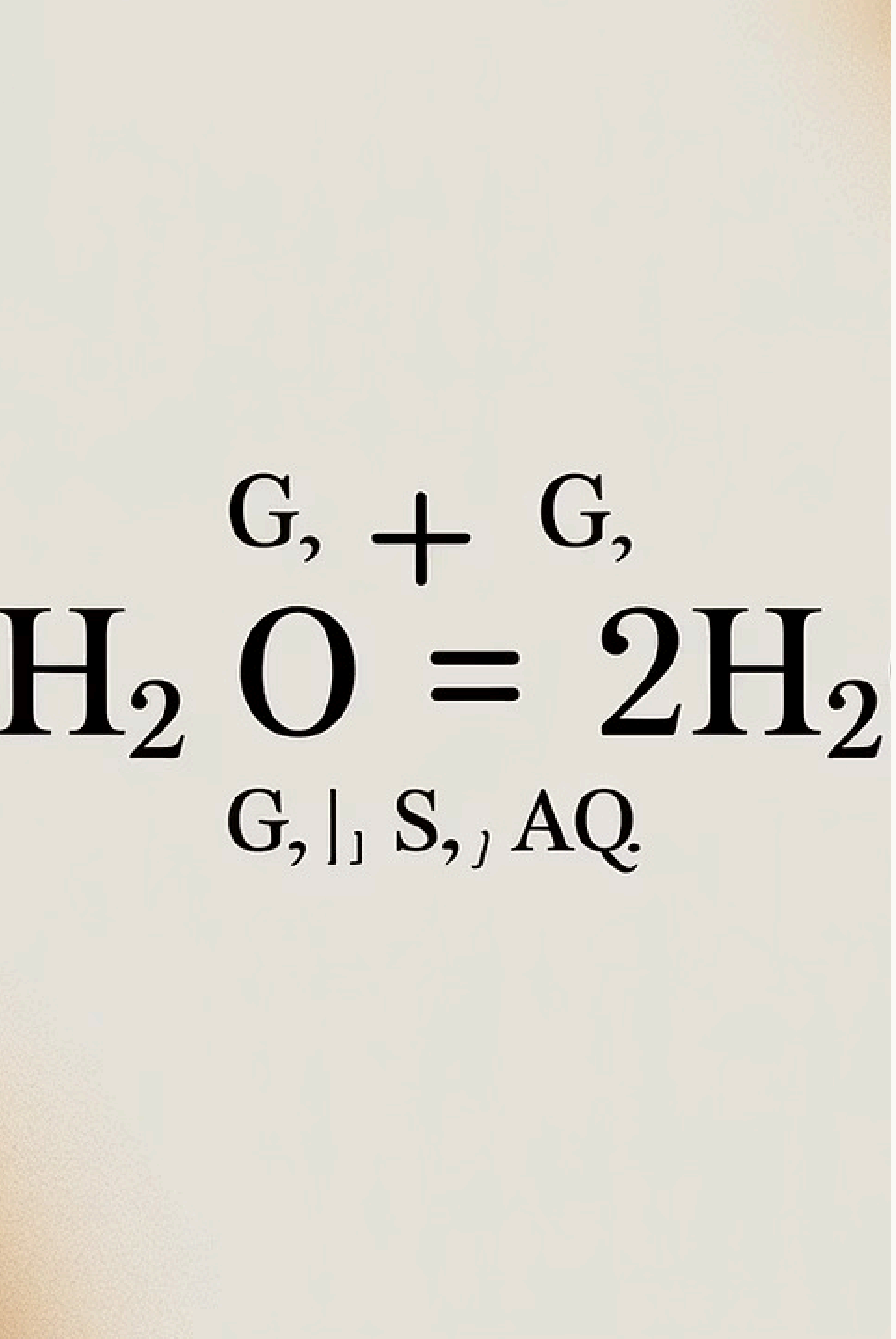
Key Characteristics

- Bonds are broken and formed
- Energy is absorbed or released
- New substances with different properties are created
- Atoms are rearranged but conserved

Signs of Chemical Reactions

- Color change
- Temperature change
- Gas production (bubbling)
- Precipitate formation
- Light emission

Everyday examples include cooking (Maillard reaction), digestion, photosynthesis ($CO_2 + H_2O \rightarrow$ glucose $+ O_2$), and rusting ($Fe + O_2 \rightarrow Fe_2O_3$).



Chemical Equations – The Language of Chemistry

Basic Format: Reactants → Products

Example: $H_2 + Cl_2 \rightarrow 2HCl$

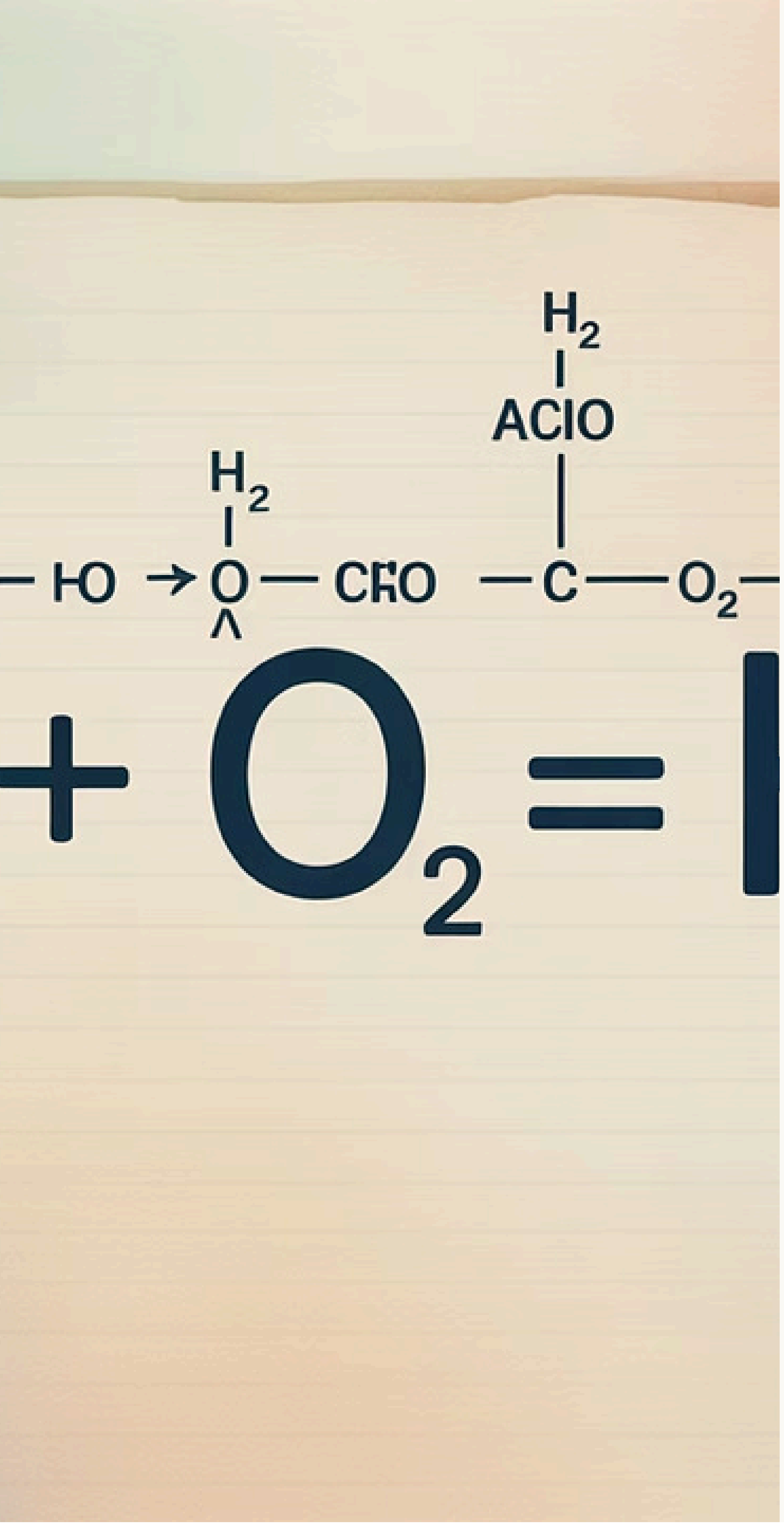
Essential Components

- Reactants: starting materials (left side)
- Products: substances formed (right side)
- Arrow: shows direction of reaction
- Coefficients: numbers showing relative amounts

State Symbols

- (s) = solid
- (l) = liquid
- (g) = gas
- (aq) = aqueous (dissolved in water)

Example with States: $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$



Balancing Chemical Equations – Conservation of Mass

Law of Conservation of Mass: Atoms cannot be created or destroyed in chemical reactions

Balancing Rules

1. Count atoms of each element on both sides
2. Adjust coefficients (never change subscripts)
3. Start with most complex molecule
4. Balance metals first, then nonmetals, then hydrogen and oxygen
5. Use whole number coefficients

Step-by-Step Example: $Fe + O_2 \rightarrow Fe_2O_3$

- Step 1: $Fe + O_2 \rightarrow Fe_2O_3$ (unbalanced)
- Step 2: Balance Fe: $2Fe + O_2 \rightarrow Fe_2O_3$
- Step 3: Balance O: $2Fe + 3O_2 \rightarrow Fe_2O_3$ (still unbalanced)
- Step 4: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ (balanced)

Common Mistakes: Changing subscripts instead of coefficients, using fractions

Types of Chemical Reactions - Synthesis

Synthesis (Combination) Reactions: $A + B \rightarrow AB$

General Pattern: Two or more reactants combine to form one product

Examples

- $2H_2 + O_2 \rightarrow 2H_2O$ (water formation)
- $N_2 + 3H_2 \rightarrow 2NH_3$ (ammonia synthesis - Haber process)
- $2Na + Cl_2 \rightarrow 2NaCl$ (salt formation)

Industrial Applications

- Haber process: ammonia for fertilizers
- Steel production: $Fe + C \rightarrow$ steel alloys
- Cement production: $CaO + SiO_2 \rightarrow$ calcium silicate

Energy Considerations: Usually exothermic (release energy)

Driving Forces: Formation of more stable products, decrease in system energy, entropy changes

Types of Chemical Reactions – Decomposition

Decomposition Reactions: AB → A + B

General Pattern: One reactant breaks down into two or more products

Examples

- $2H_2O \rightarrow 2H_2 + O_2$ (electrolysis of water)
- $CaCO_3 \rightarrow CaO + CO_2$ (limestone decomposition)
- $2HgO \rightarrow 2Hg + O_2$ (mercury oxide decomposition)

Energy Requirements

Usually endothermic (require energy input)

Common Energy Sources:

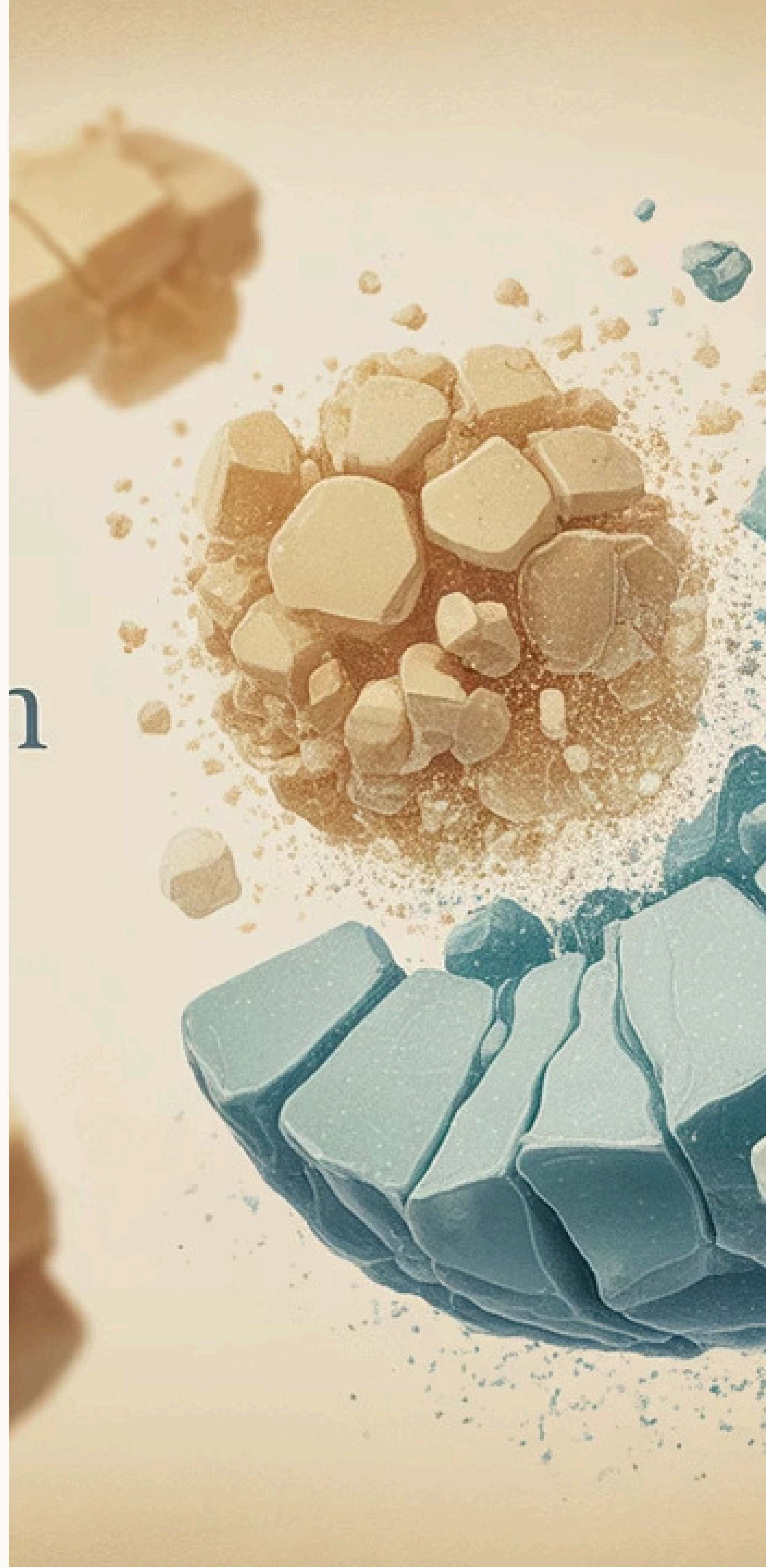
- Heat (thermal decomposition)
- Light (photochemical decomposition)
- Electricity (electrolysis)

Applications

Biological: Cellular respiration, digestion

Industrial: Metal extraction from ores, production of lime from limestone

Safety Note: Some decomposition reactions can be explosive



Types of Chemical Reactions – Single and Double Replacement

Single Replacement: A + BC → AC + B

- More reactive element replaces less reactive one
- Example: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
- Activity series determines if reaction occurs

Double Replacement: AB + CD → AD + CB

- Ions "switch partners"
- Example: $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
- Usually occurs in aqueous solution

Driving Forces: Precipitate formation, gas evolution, acid-base neutralization

Example: $Pb(NO_3)_2 + 2KI \rightarrow PbI_2(s) + 2KNO_3$ (yellow PbI_2 precipitate forms)

Acid-Base Example: $HCl + NaOH \rightarrow NaCl + H_2O$ (neutralization reaction)

Combustion Reactions – A Special Category

Definition: Reaction with oxygen producing heat and light

Hydrocarbon Combustion: $C_xHy + O_2 \rightarrow CO_2 + H_2O$

Complete vs. Incomplete Combustion

Complete (excess oxygen):

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Incomplete (limited oxygen):

$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$

Danger: CO is toxic

Balancing Combustion Equations

1. Balance carbon atoms first
2. Balance hydrogen atoms second
3. Balance oxygen atoms last
4. Adjust coefficients to whole numbers

Energy Release: Highly exothermic reactions

Real-World Applications: Fossil fuel burning for energy, cellular respiration, rocket propulsion

Environmental Impact: CO_2 production contributes to greenhouse effect

Predicting Reaction Products and Spontaneity

Factors Affecting Reactions

- Thermodynamics: energy changes
- Kinetics: reaction rate
- Activation energy barriers

Predicting Ionic Reactions

- Use solubility rules
- Consider gas formation
- Apply acid-base principles

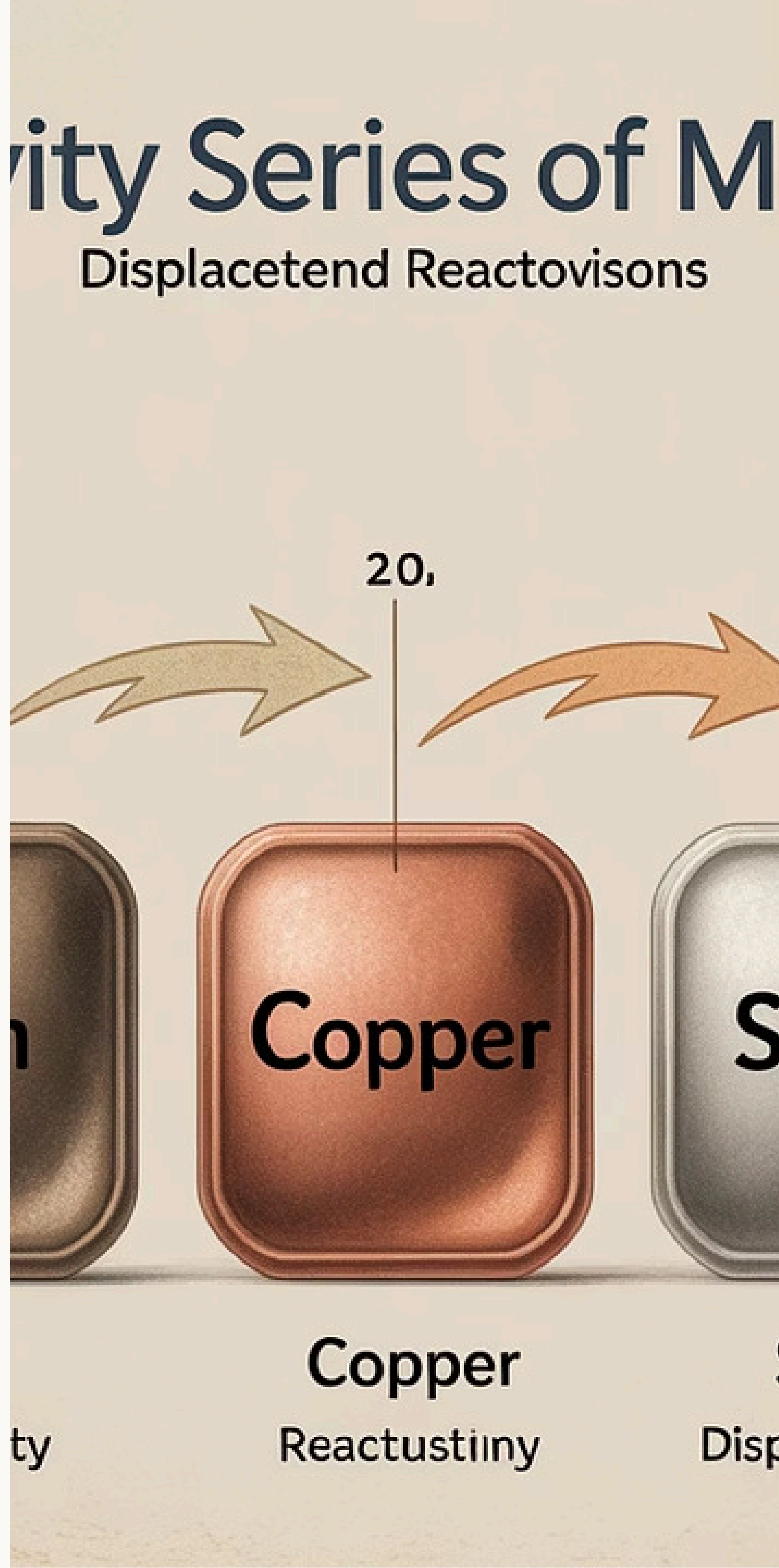
Solubility Rules (Key Points)

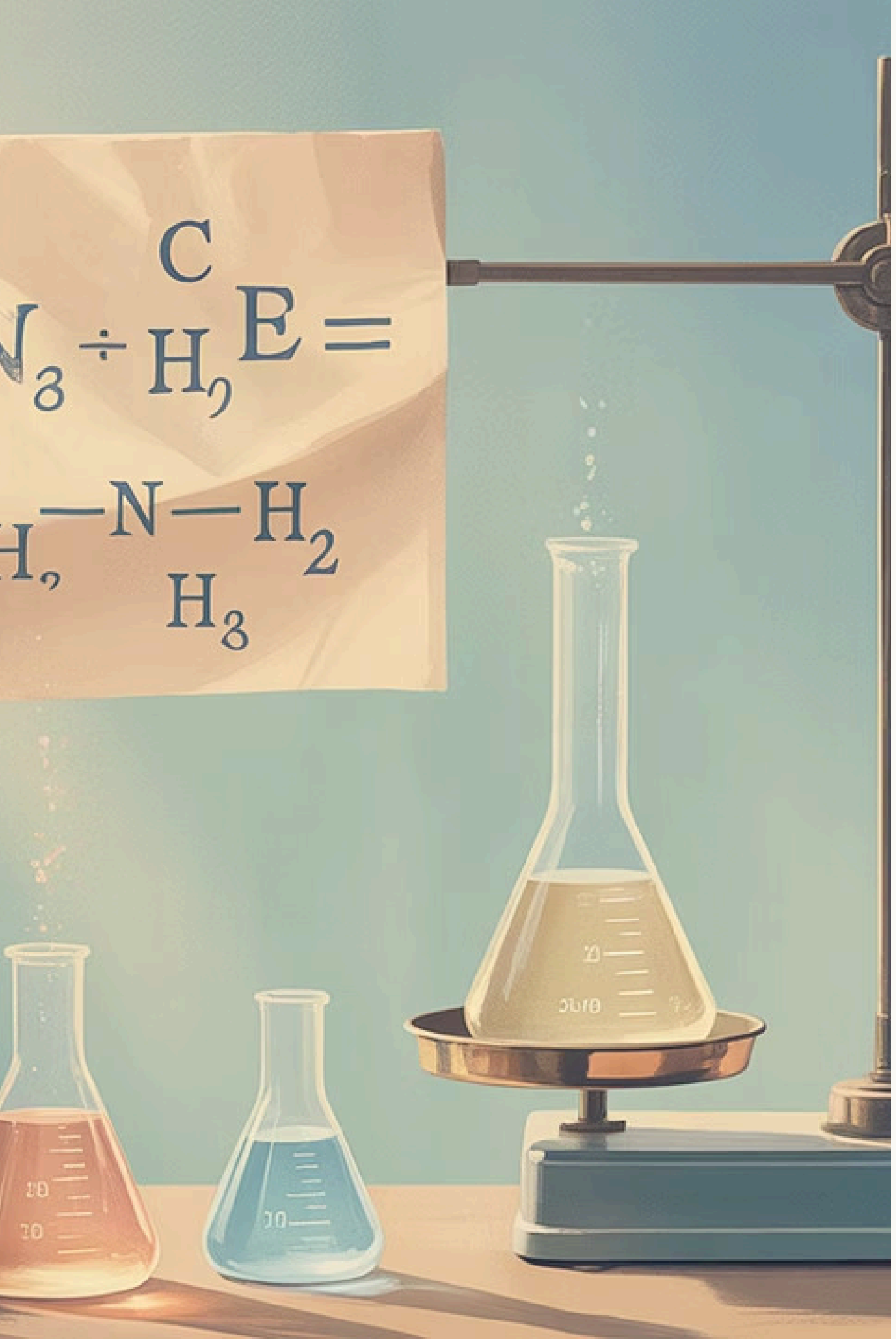
- Most nitrates (NO_3^-) are soluble
- Most chlorides are soluble except $AgCl$, $PbCl_2$, Hg_2Cl_2
- Most sulfates are soluble except $BaSO_4$, $PbSO_4$, $CaSO_4$

Activity Series Applications: Metals above hydrogen produce H_2 from acids; more reactive metals displace less reactive ones (e.g., Fe > Cu, so $Fe + CuSO_4 \rightarrow$ reaction occurs)

Spontaneity Indicators: Formation of precipitate, gas evolution, large energy release, increase in entropy

Practice Strategy: Learn patterns rather than memorizing every possible reaction





Coming Up Next:

Basic Stoichiometry

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