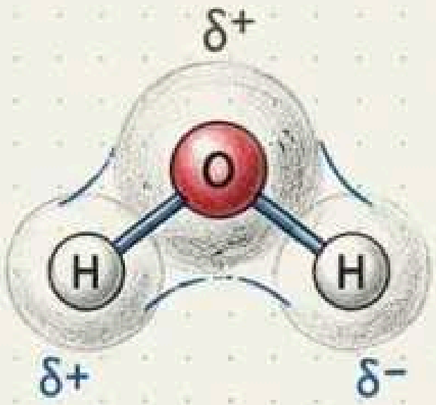
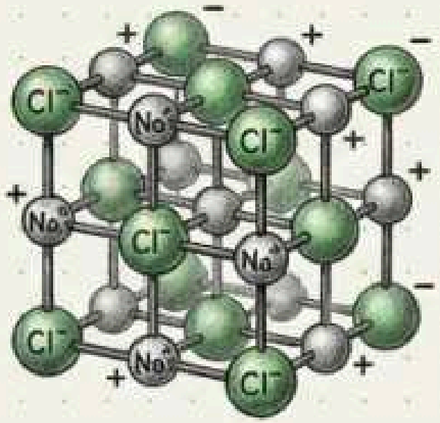


CHEM 101 B: Lecture 3

Chemical Bonding and Molecular Shape

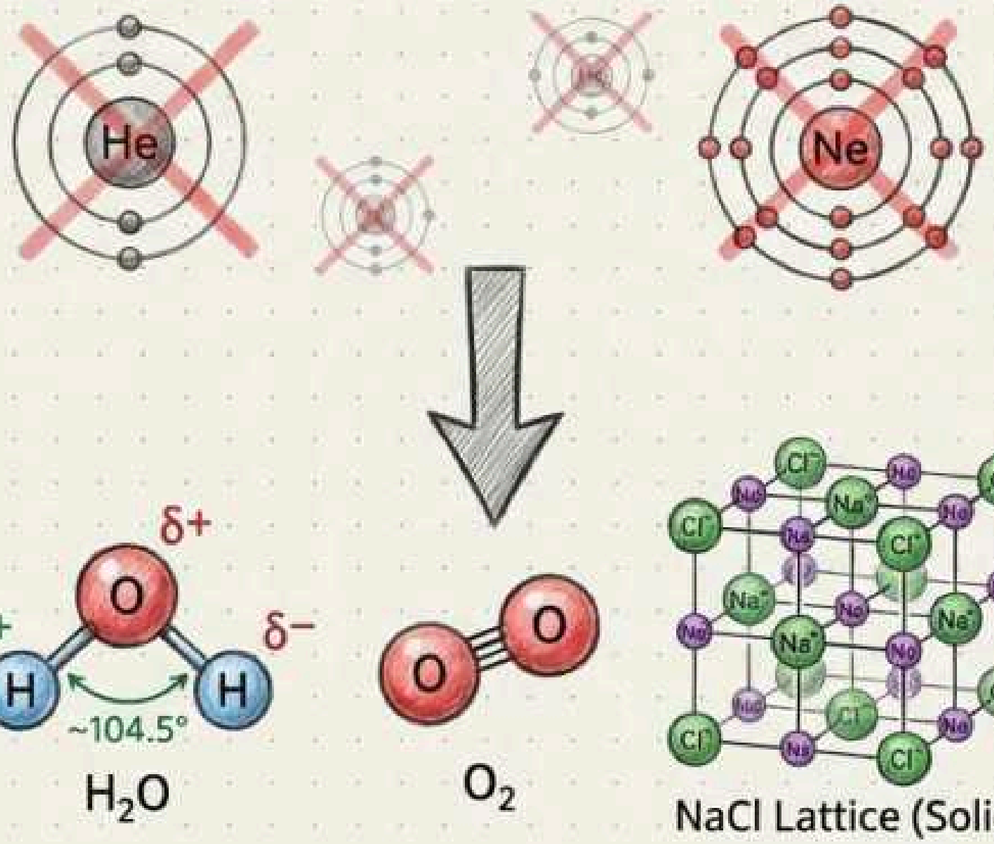


- At a Glance
  - Why atoms bond
  - Ionic, covalent & metallic bonds
  - Lewis structures
  - VSEPR shapes
  - Intermolecular forces & hydrogen bonding

Atoms Rarely Travel Alone

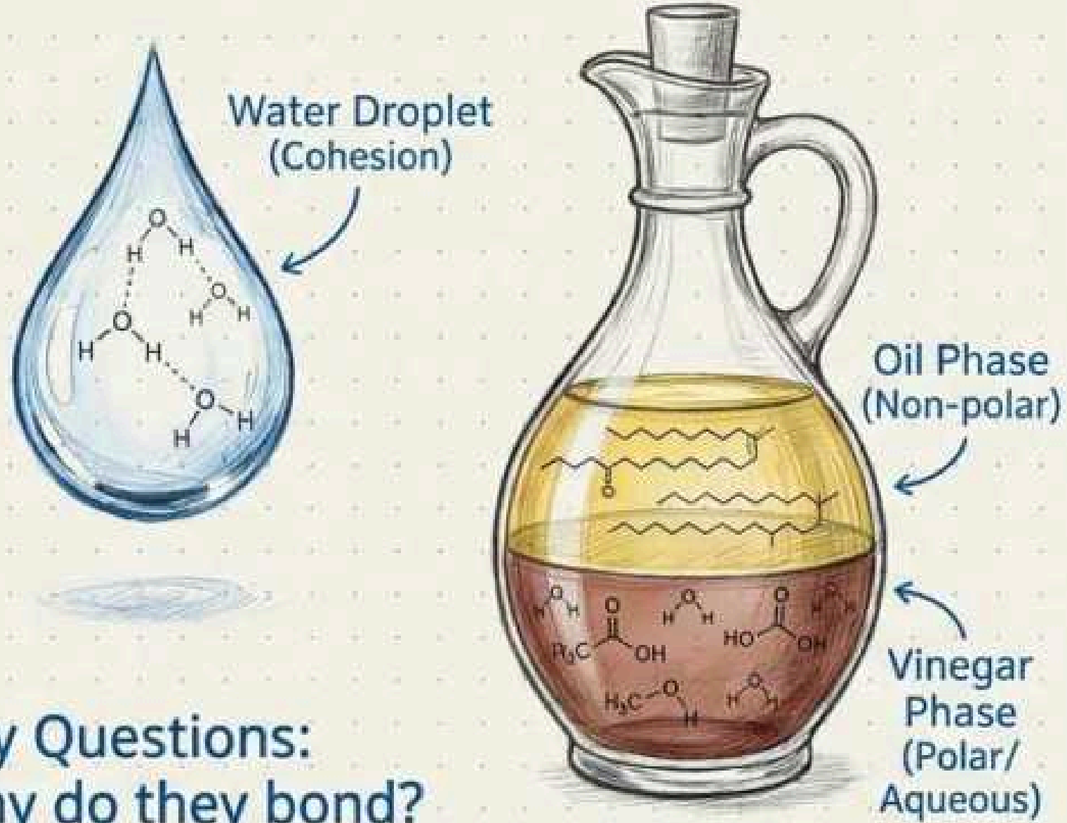
¶ 3.2

(Microscopic)



Fact: Almost everything around you is made of atoms joined together.

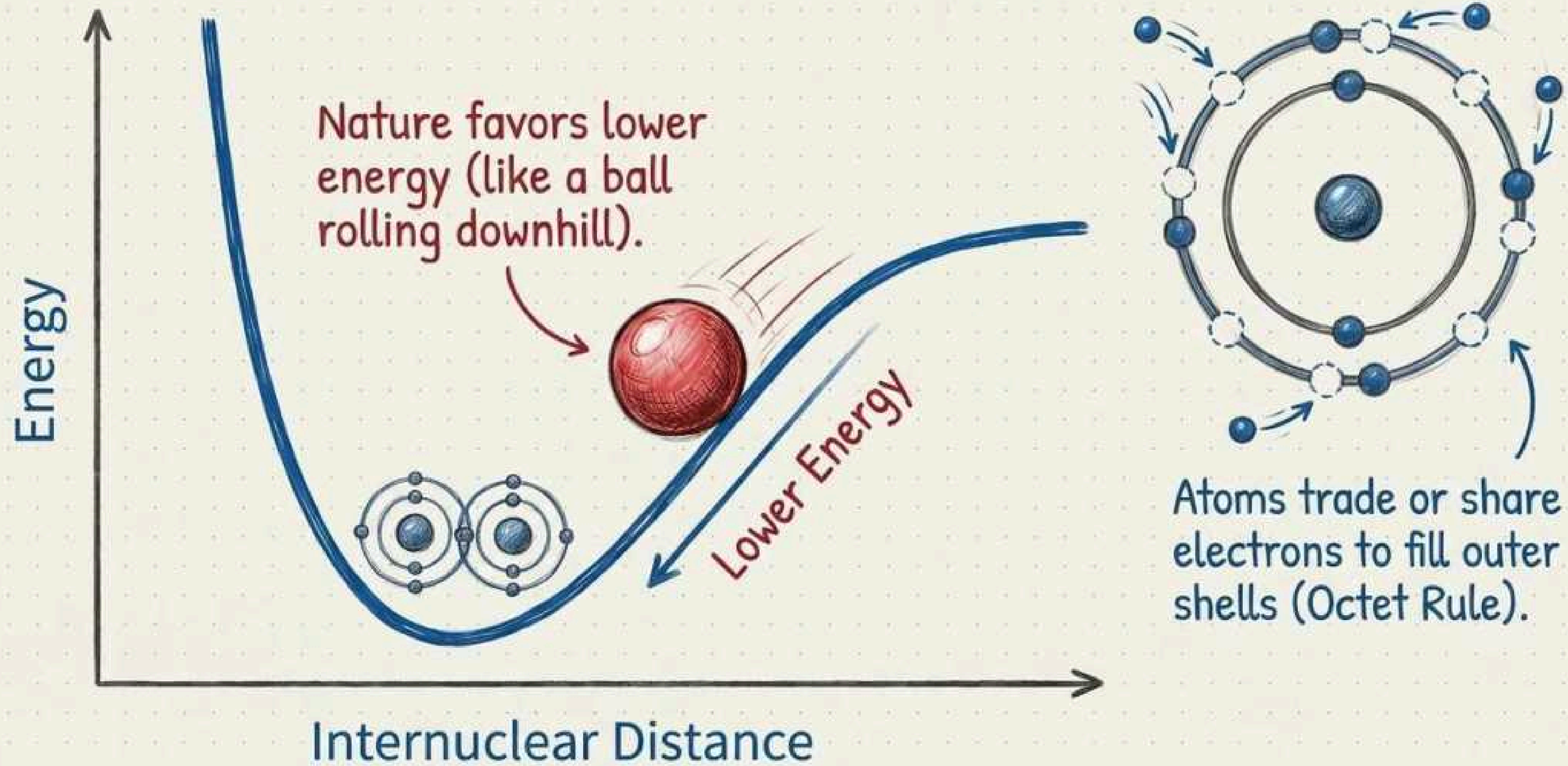
(Macroscopic)



Key Questions:
Why do they bond?
What shapes do they form?
Why are water and oil so different at room temperature?

The Drive for Lower Energy

¶ 3.3



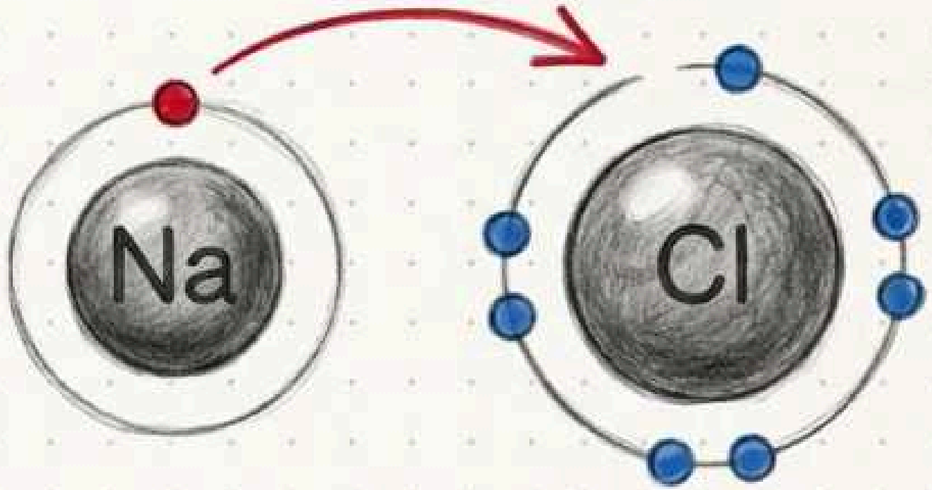
Atoms trade or share electrons to fill outer shells (Octet Rule).

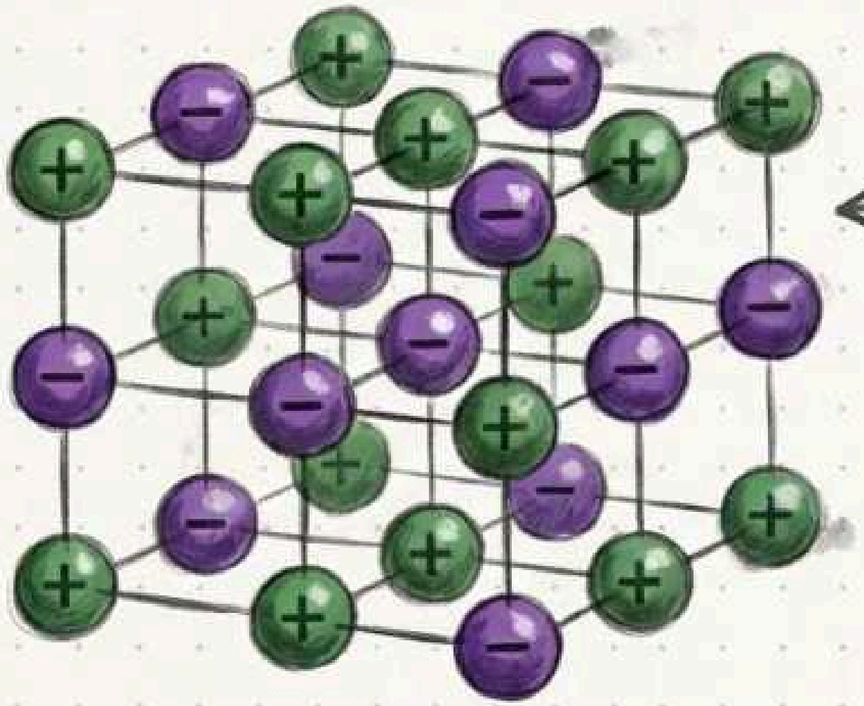
Ionic Bonding: Trading Electrons

¶ 3.4

Metal + Nonmetal —> Cations (+) + Anions (-)

$Na + Cl \rightarrow Na^{+} + Cl^{-}$

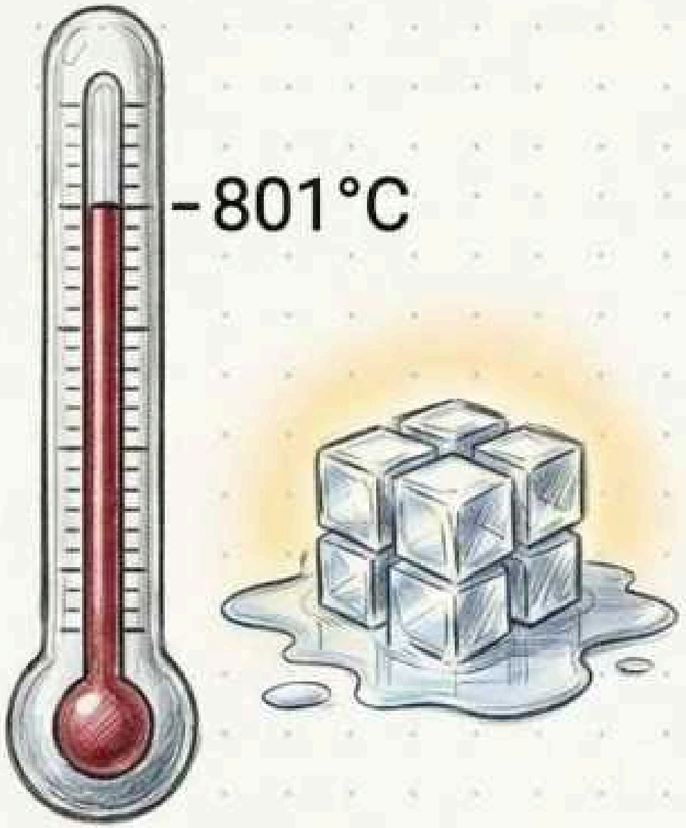




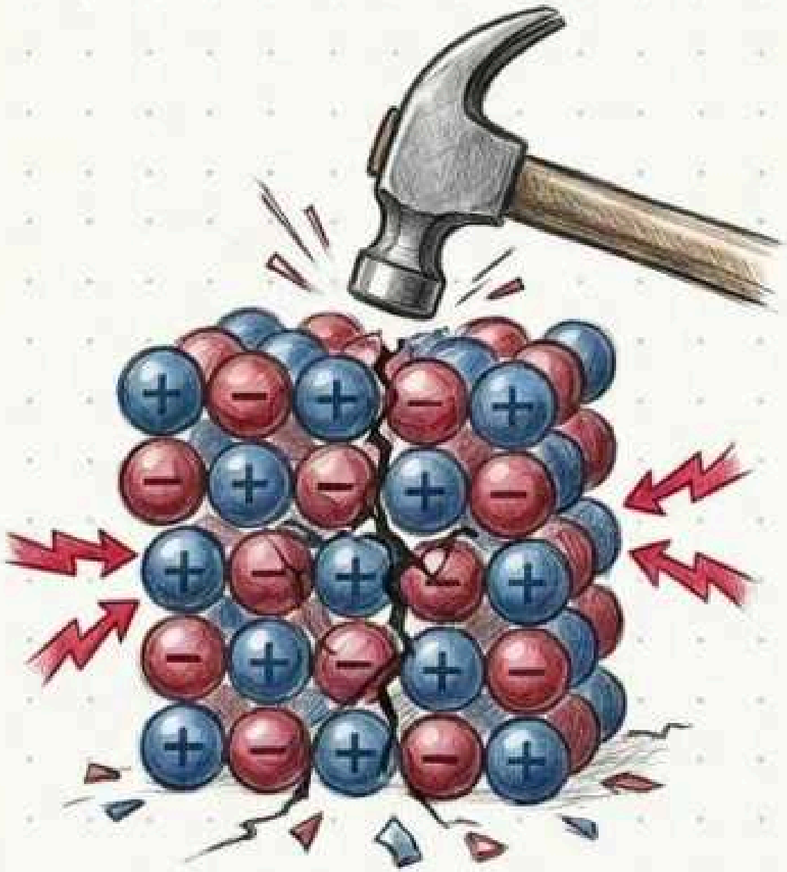
Solid salt is not a single molecule, but a vast 3D crystal lattice held by electrostatic attraction.

Properties of Ionic Compounds

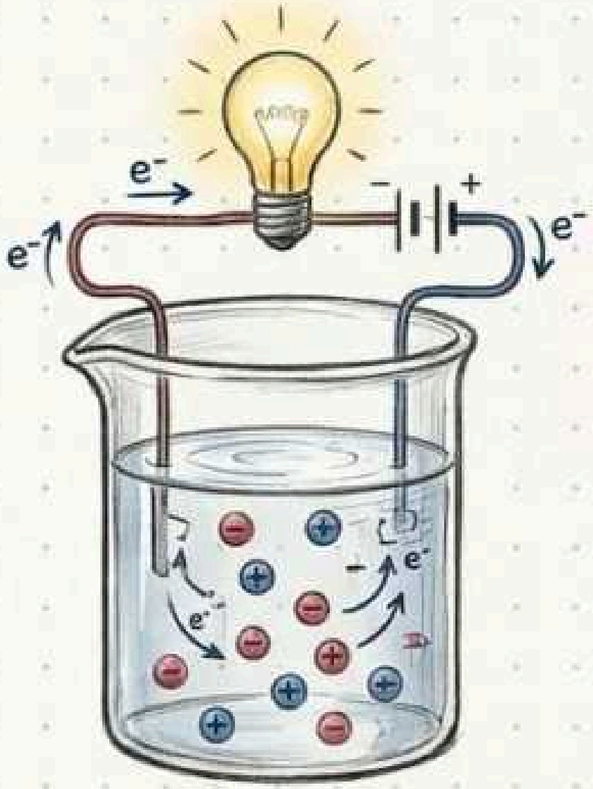
¶ 3.5



High Melting Point (strong attractions).



Hard but Brittle.

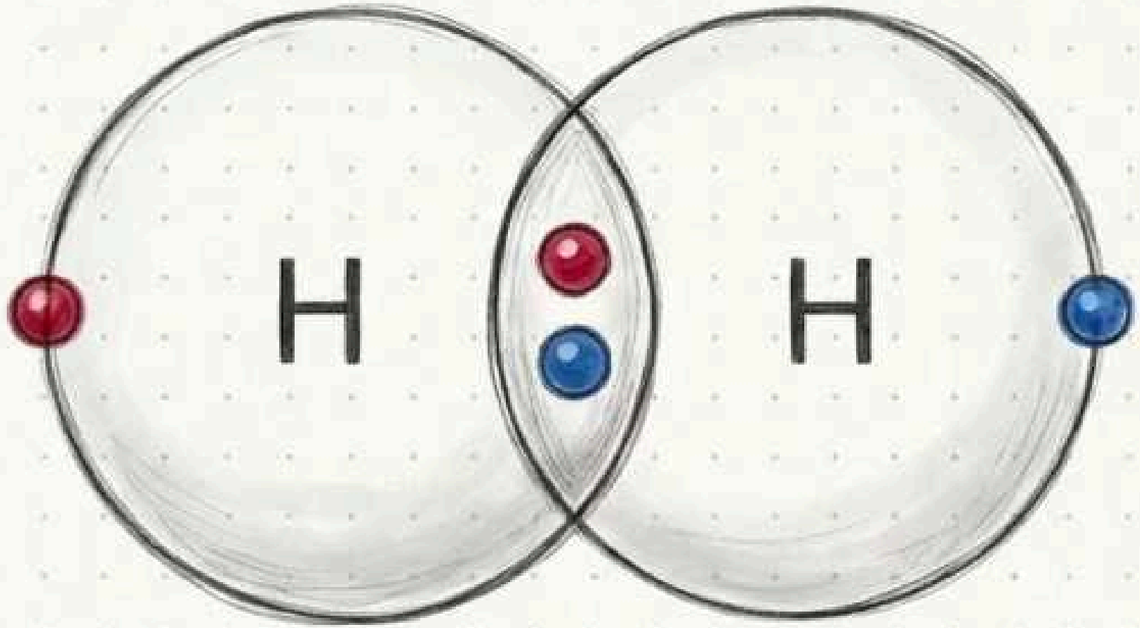


Non-conductive as solids; conductive when melted/dissolved (mobile ions).

Covalent Bonding: Sharing Electrons

¶ 3.6

H + H —> H2
(Molecule)



- Mechanism: Two atoms (nonmetals) share a pair of electrons.
- Key Insight: Sharing allows each atom to count both electrons, completing its shell.

Examples:
- Water
- Carbon Dioxide
- Sugar

Drawing Lewis Structures

¶ 3.7

The Rules

1. Count total valence electrons.
2. Join atoms with single bonds.
3. Place remaining electrons to reach 8 (or 2 for H).

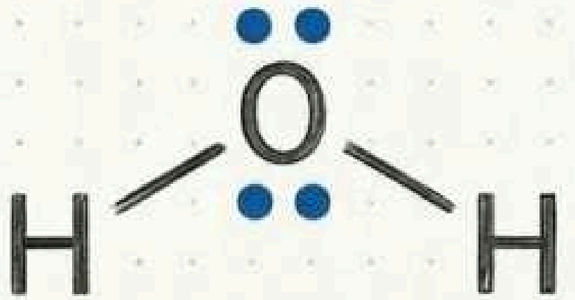
The Application

Step 1: Total valence = 8 (6 from O, 1x2 from H).

Step 2: Skeleton H – O - H

H—O—H

Step 3: Add lone pairs to Oxygen.



Result: 2 shared pairs (bonds), 2 lone pairs on Oxygen.

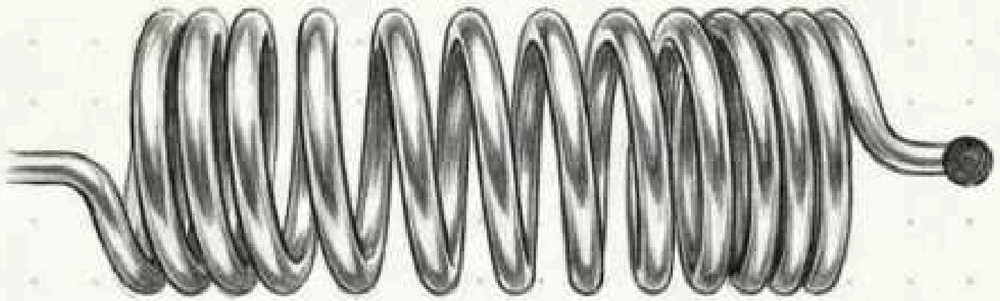
Multiple Bonds: Sharing More Pairs

¶ 3.8

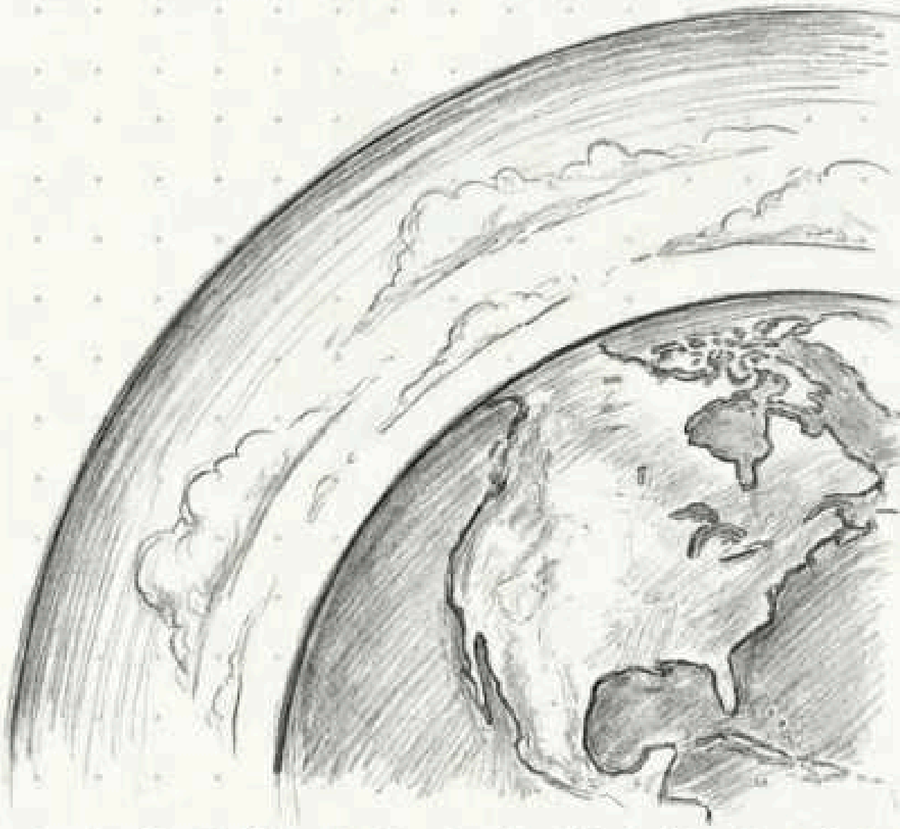
O=O

2 pairs shared (e.g., O_2)

N≡N



3 pairs shared (e.g., N_2)



Rule: More shared pairs = Shorter, stronger bond. The N_2 triple bond is so strong that atmospheric nitrogen (80% of air) is nearly inert.

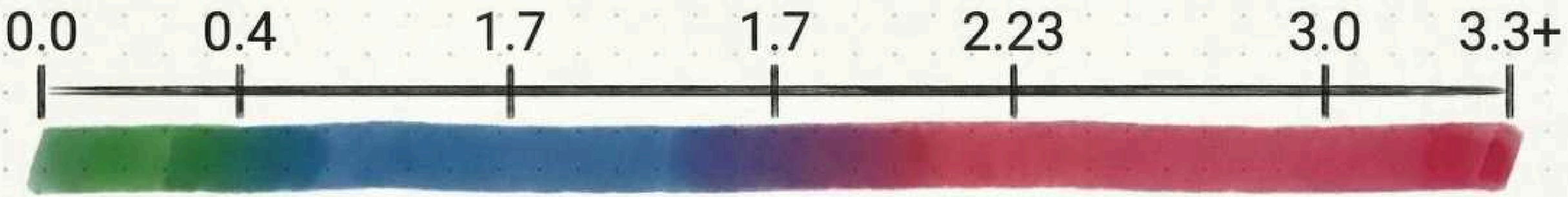
Electronegativity & Bond Types

¶ 3.9

$$\Delta EN = EN_1 - EN_2$$

Not all sharing is equal. More electronegative atoms pull harder on the shared pair.

NaCl → (3.16 - 0.93 = 2.23 → Ionic)



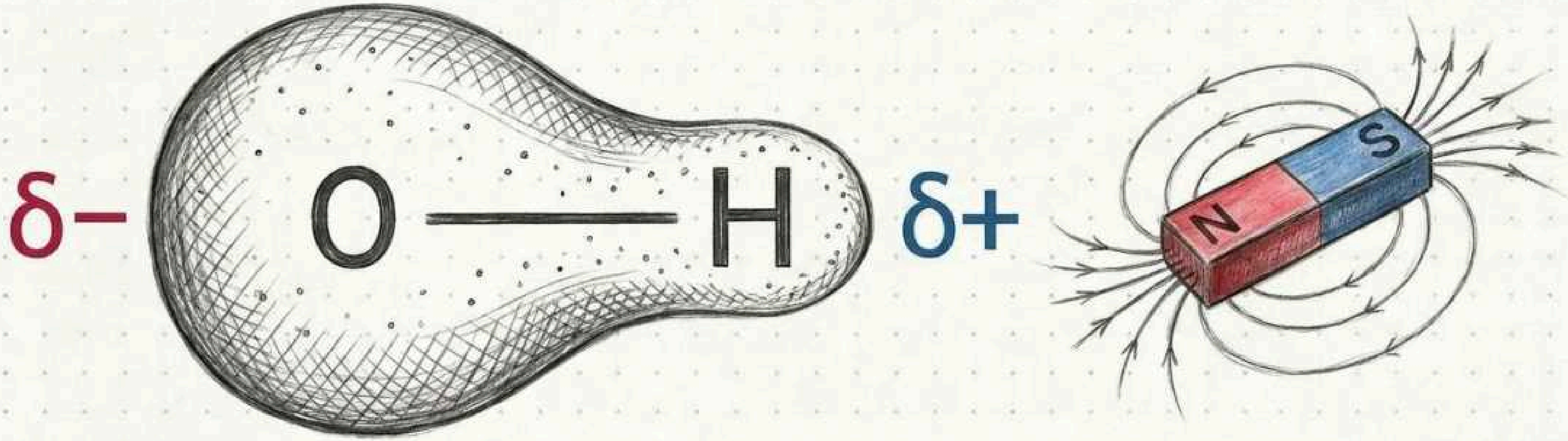
Nonpolar (< 0.4)

Polar Covalent (0.4 - 1.7)

Ionic (> 1.7)

The Polar Bond (Dipole)

¶ 3.10



Example: Water (O-H bond). ΔEN = 3.44 (O) - 2.20 (H) = 1.24

Mechanism: Shared electrons spend more time near Oxygen.

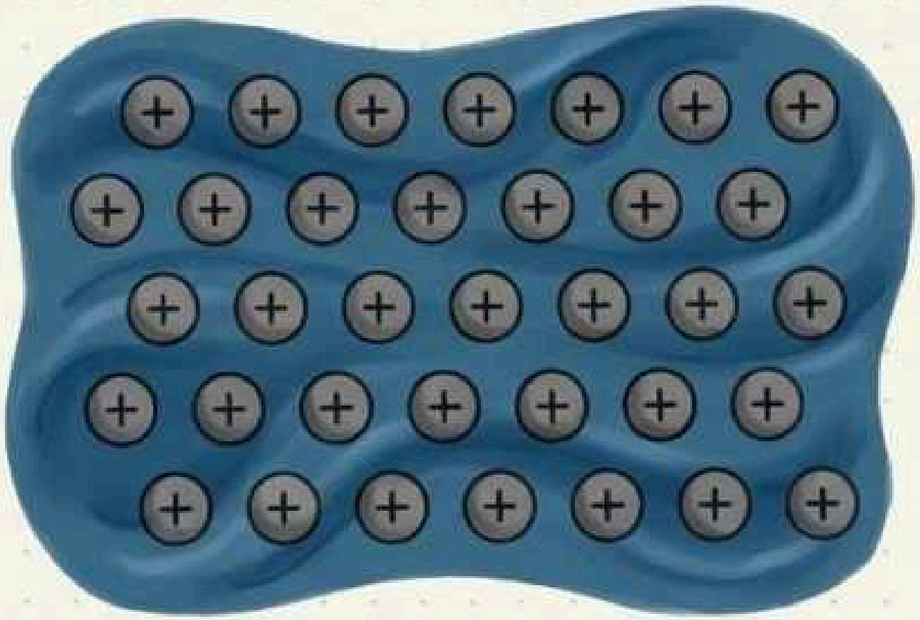
Result: Partial negative charge (δ-) on Oxygen, partial positive (δ+) on Hydrogen.

Definition: Dipole (a bond with two ends, like a tiny magnet).

Metallic Bonding: The Electron Sea

¶ 3.11

Microscopic



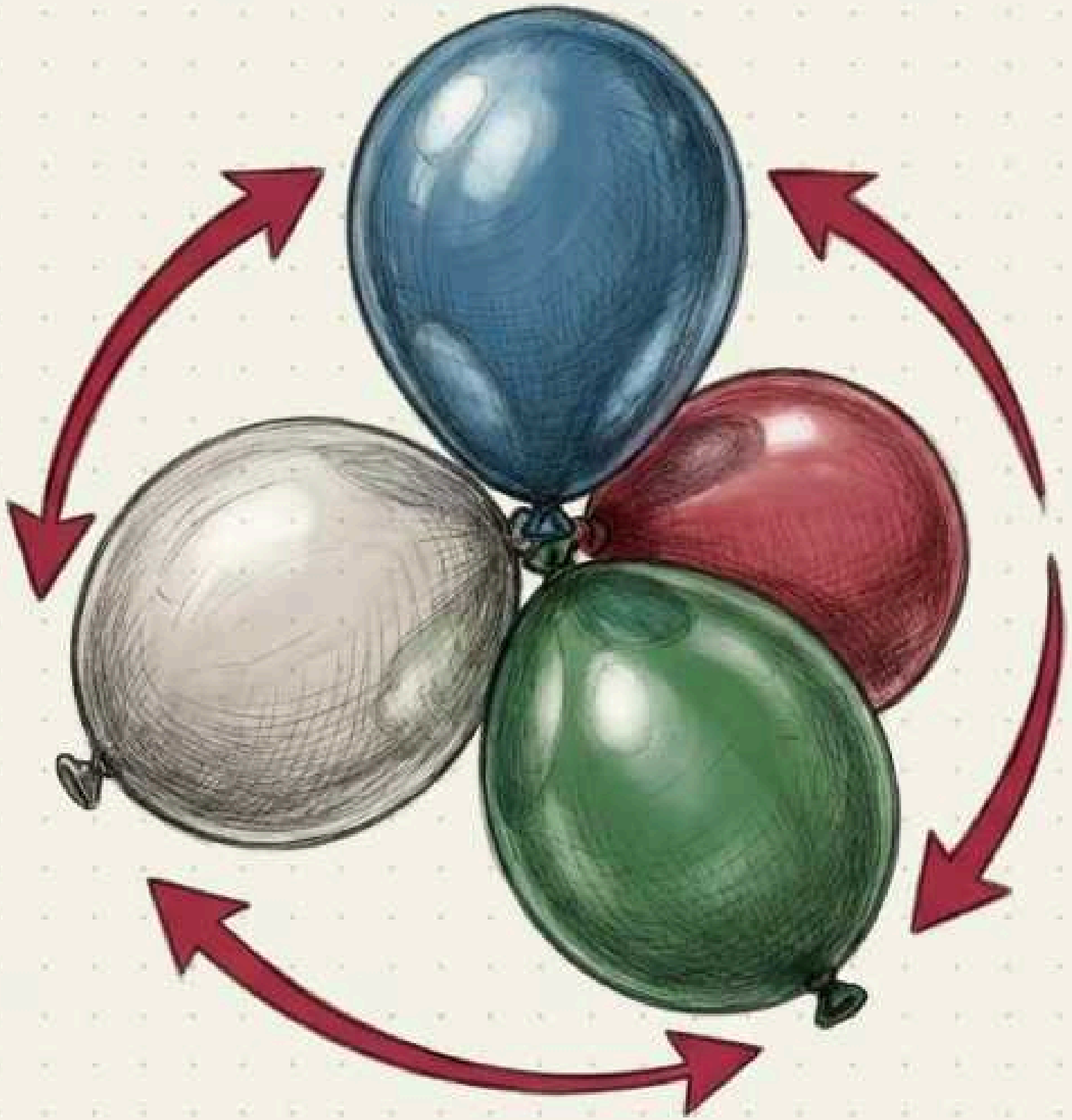
Macroscopic



- Mechanism: Atoms release valence electrons into a shared, mobile pool.
- Conductivity: Free-flowing electrons allow electricity to move effortlessly.
- Malleability: Can be hammered or drawn into wire without breaking (ions slide past each other smoothly).

Molecular Shape: VSEPR Theory

¶ 3.12



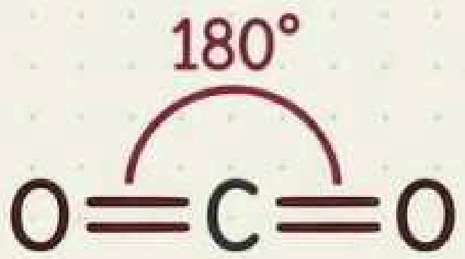
VSEPR: Valence Shell Electron Pair Repulsion.

Concept: Molecules are 3D. Electron pairs repel each other and spread as far apart as possible.

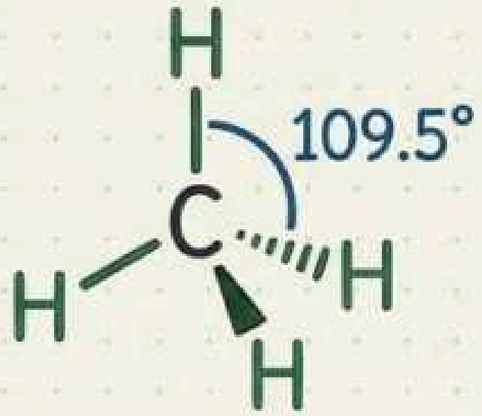
Rule: To predict shape, count groups of electrons (bonds + lone pairs) around the central atom to find the widest spatial arrangement.

The Four Primary Shapes

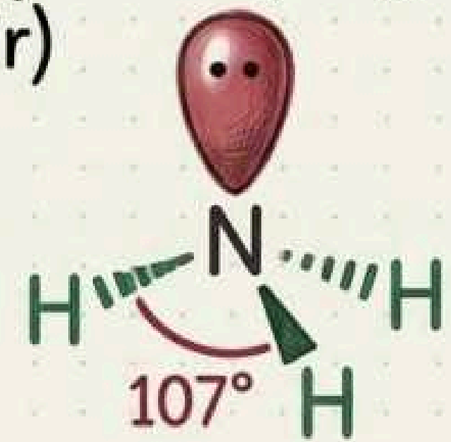
Linear (CO_2, 2 groups)



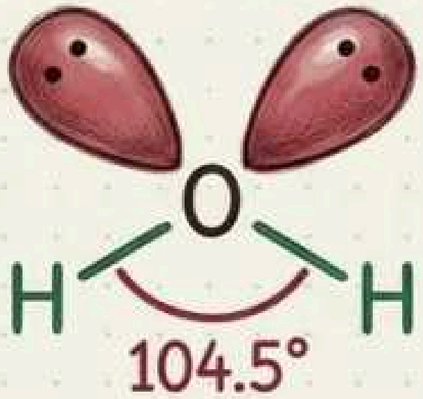
Tetrahedral (CH_4, 4 bonding pairs)



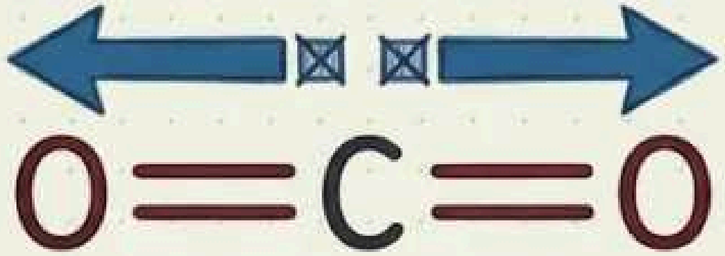
Trigonal Pyramid (NH_3, 3 bonds, 1 lone pair)



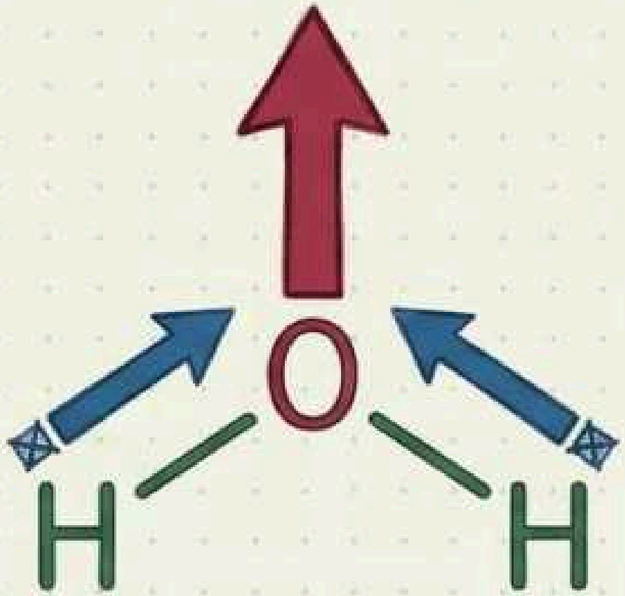
Bent (H_2O, 2 bonds, 2 lone pairs)



Shape Dictates Molecular Polarity



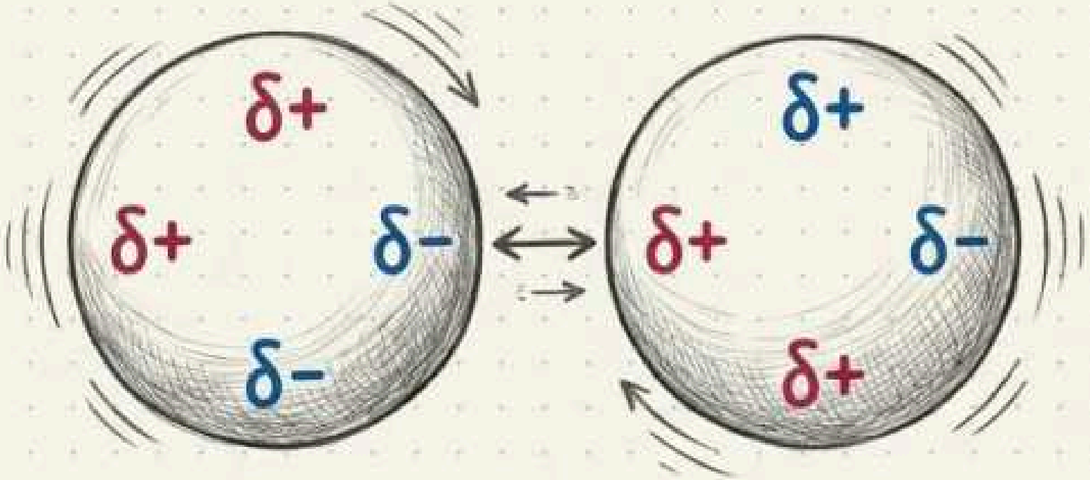
Carbon Dioxide (CO_2): Polar bonds point in opposite directions → Cancel out to zero → Nonpolar Molecule.



Water (H_2O): Bent shape → Bonds do not cancel → Polar Molecule overall.

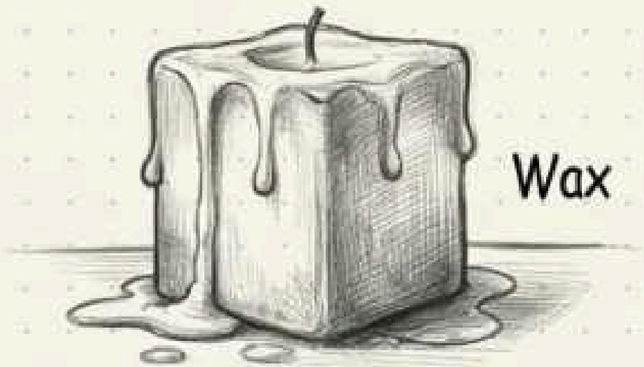
Key Insight: This single fact (a bent shape) explains why water is an extraordinary liquid.

Intermolecular Forces: London Dispersion

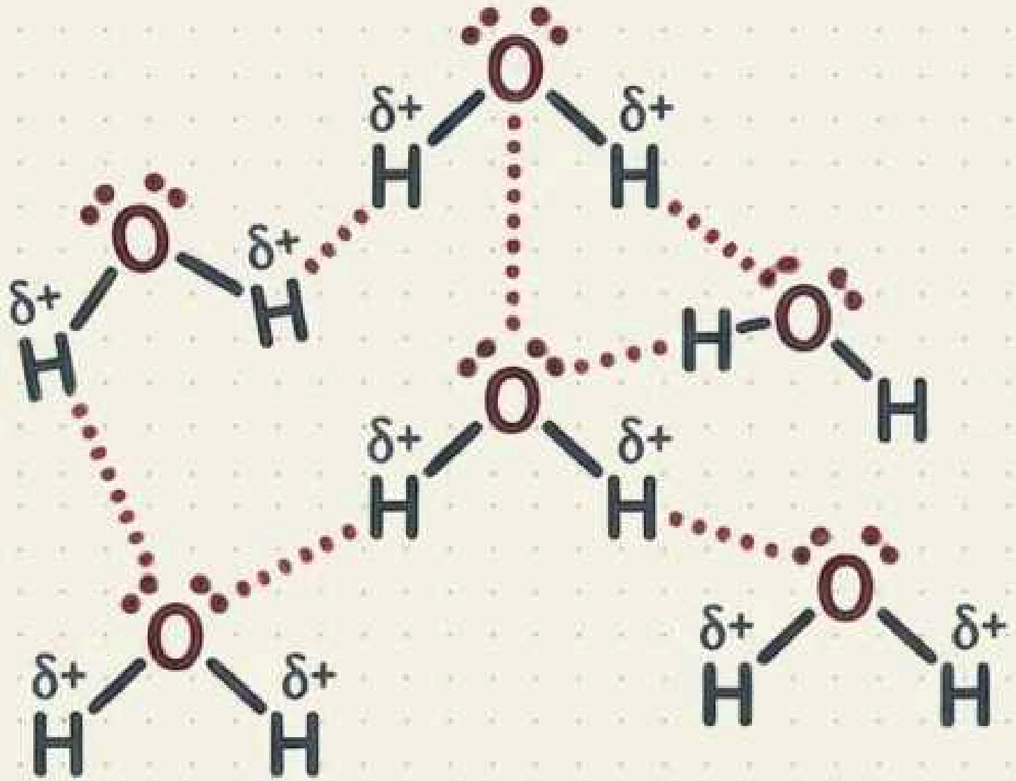


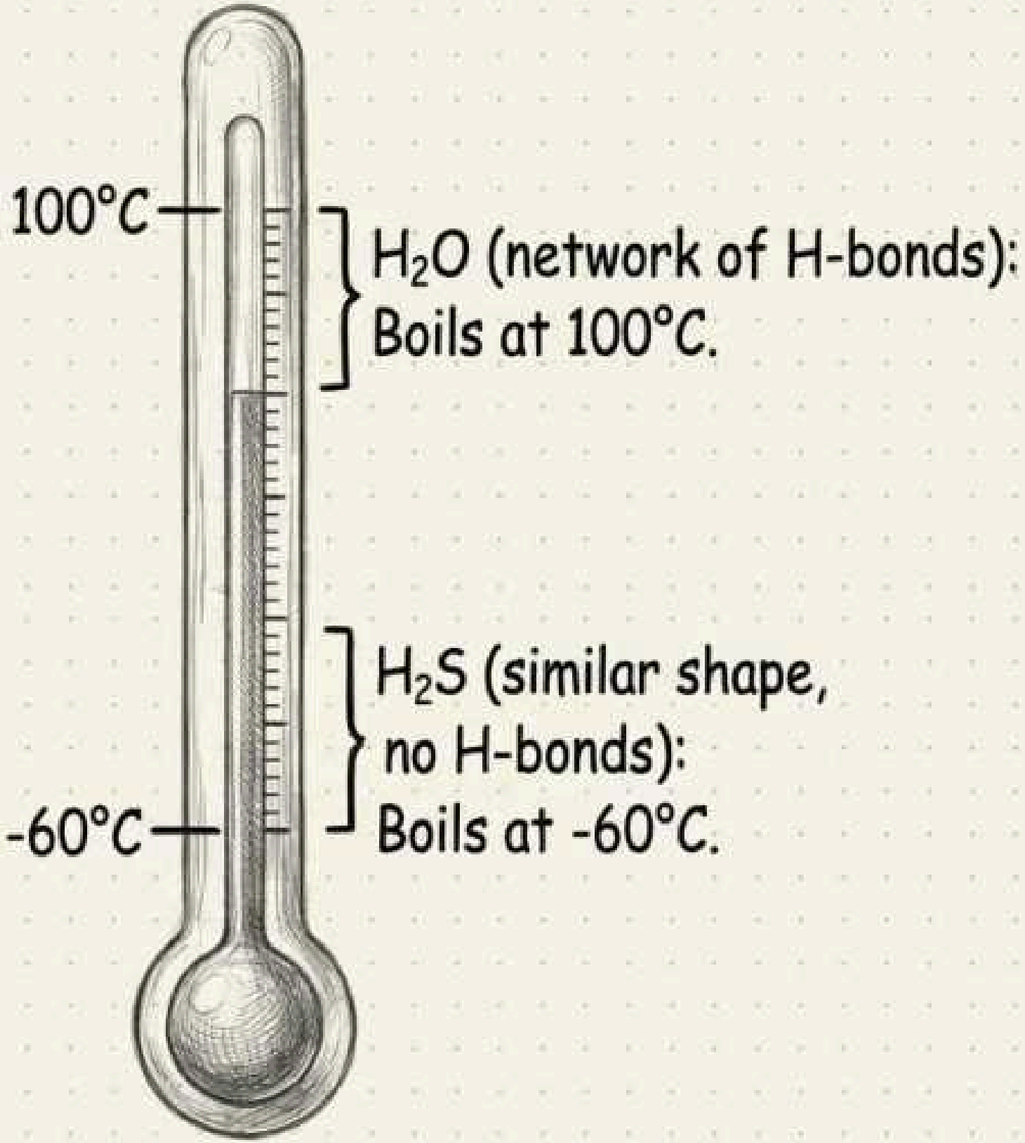
- **Concept:** Forces ***between*** molecules (weaker than bonds).
- **Mechanism:** Electrons shift about, creating brief, flickering charges (present in ALL molecules).
- **Rule:** Larger molecules = stronger dispersion forces (e.g., small Methane = gas; large Wax = solid).





Hydrogen Bonding



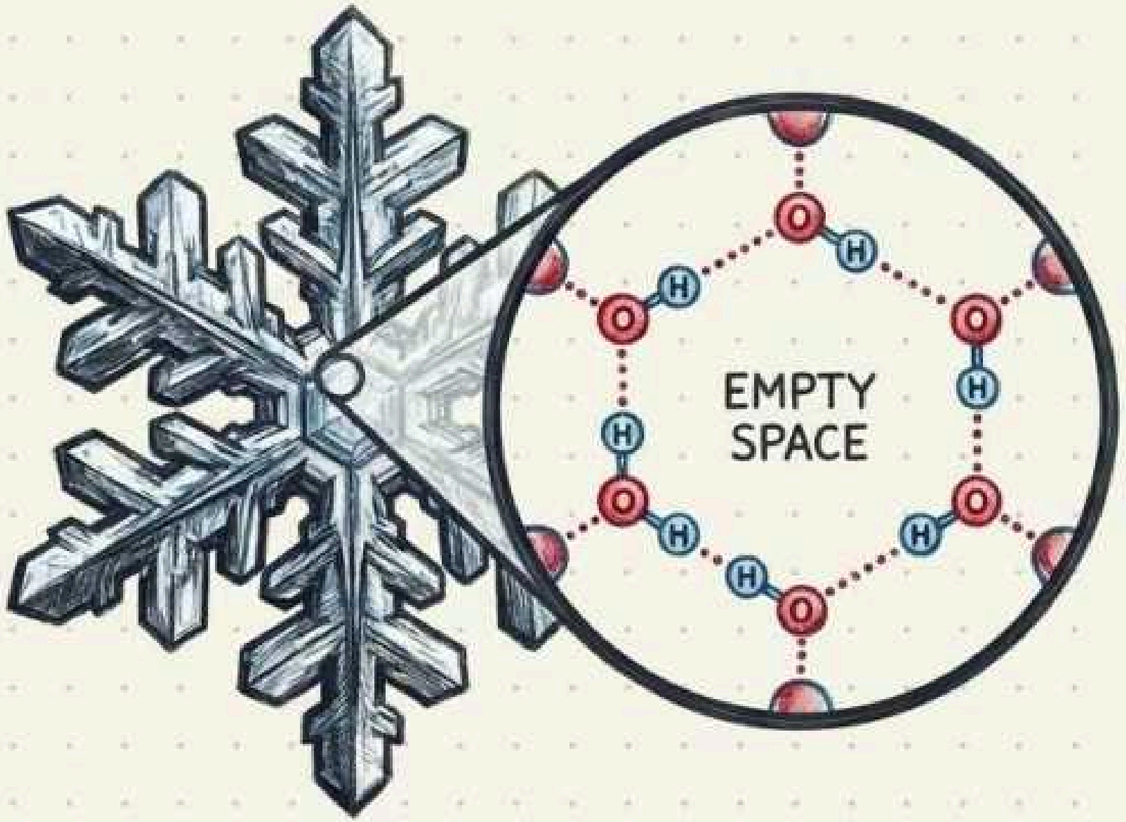


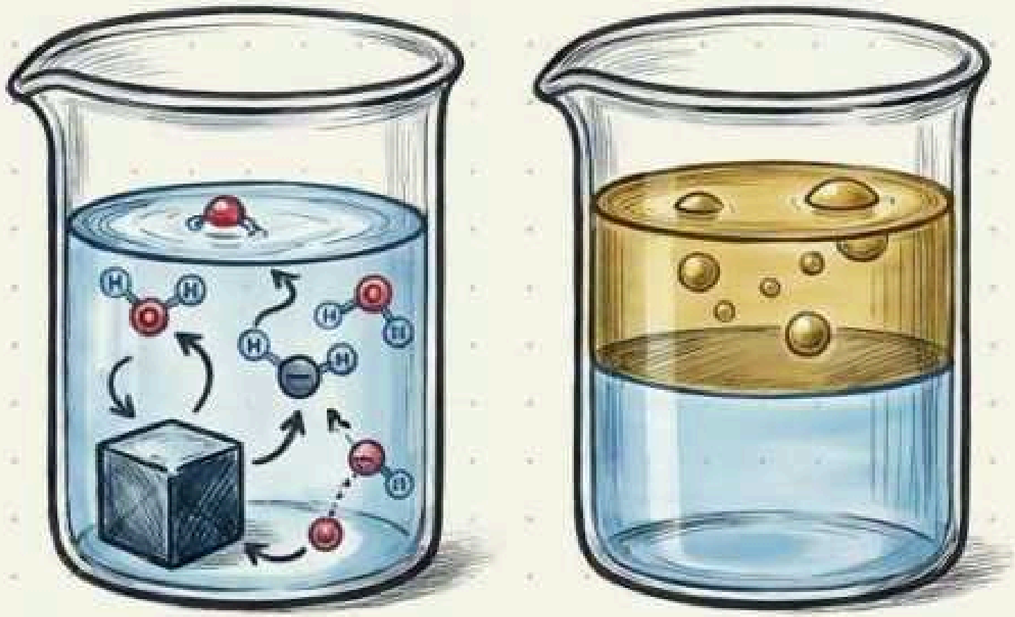
Concept: The strongest intermolecular force.

Rule: Occurs when H is bonded to N, O, or F, and attracts to a lone pair on a neighbor.

Macroscopic Effects of H-Bonds

¶ 3.17





Salt + Water
(Polar + Polar)

Oil + Water
(Nonpolar + Polar)

Ice Floats: H-bonds lock water into an open hexagonal pattern → takes up more space → 8% less dense than liquid.

Solubility Rule: "Like dissolves like." Polar water dissolves polar salt. Nonpolar oil only dissolves in other oils.

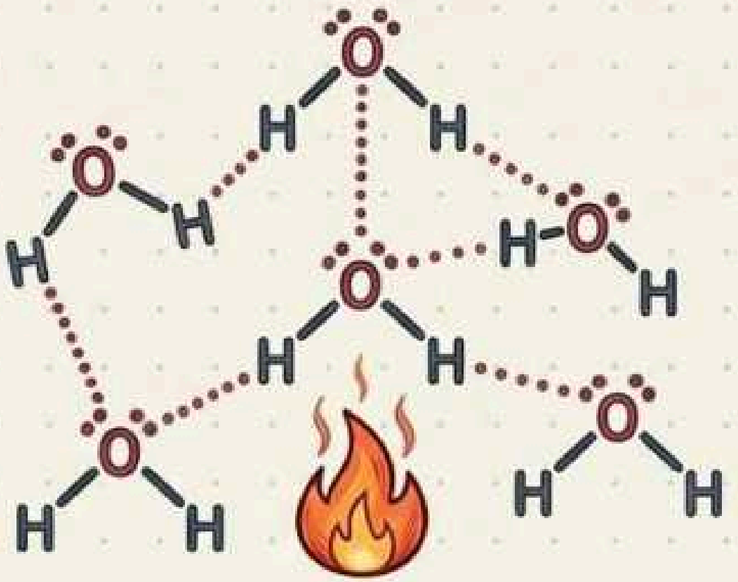
Bonds vs. Intermolecular Forces

Chemical Change



Bonds (Intramolecular): Hold atoms together *inside* a molecule.
Breaking them = Chemical Change.

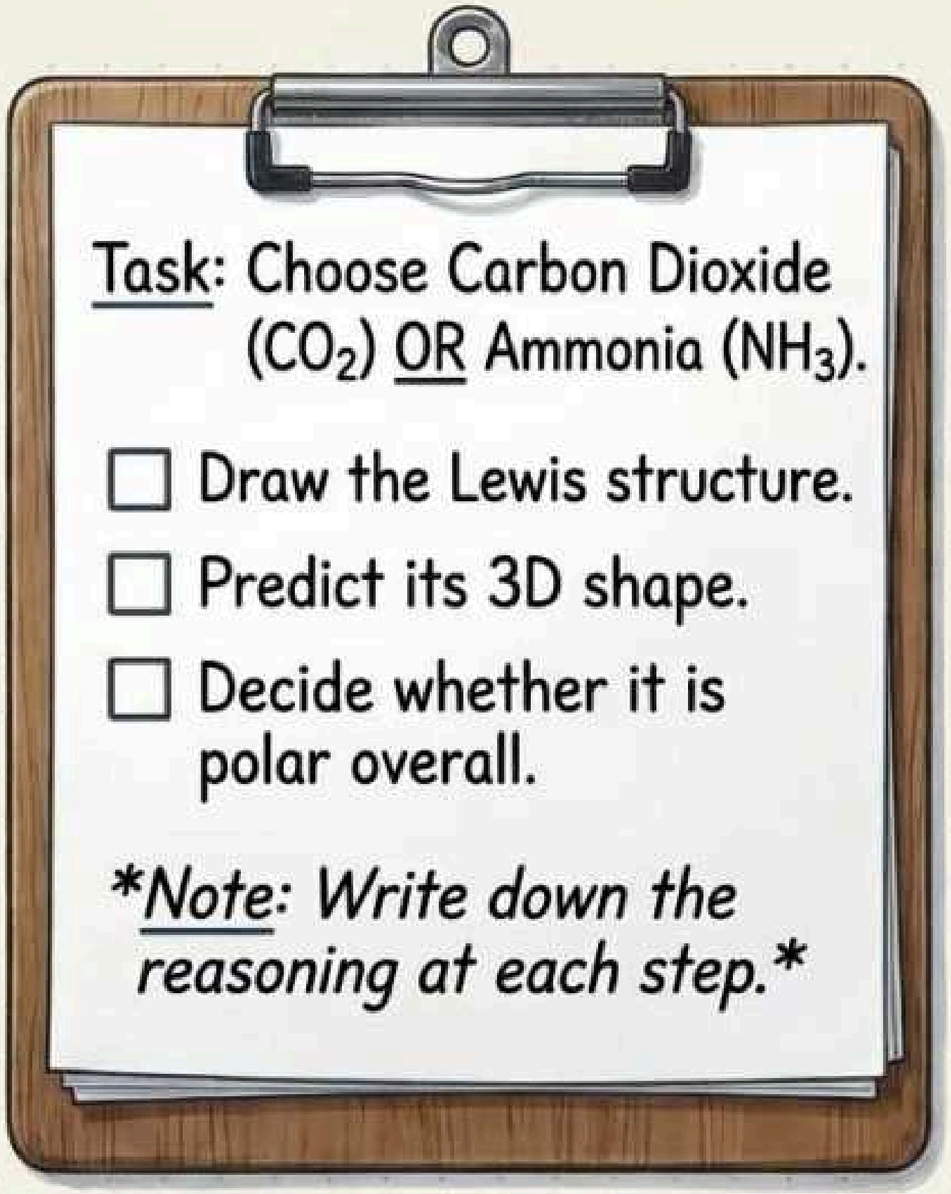
Physical Change



IMFs (Intermolecular): Act *between* molecules. Overcoming them (e.g., boiling) = Physical Change.

Crucial Distinction: Keep these two ideas apart!

Lecture 3 Assignment

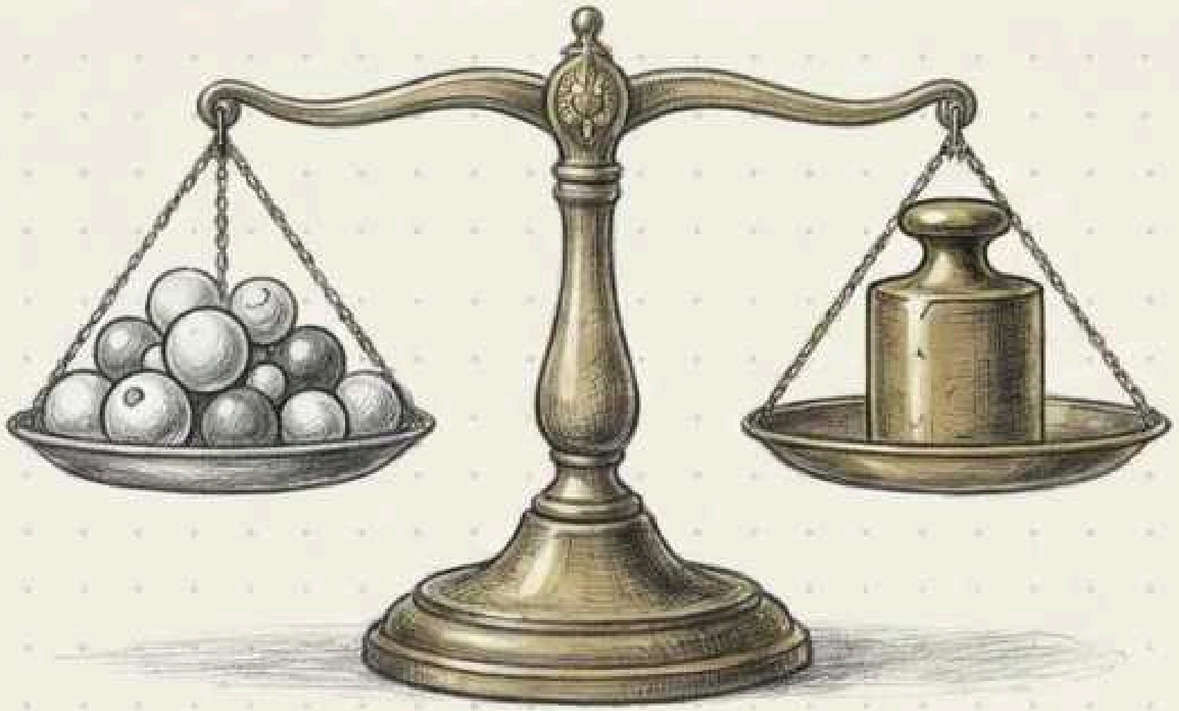


Task: Choose Carbon Dioxide (CO_2) OR Ammonia (NH_3).

- ☐ Draw the Lewis structure.
- ☐ Predict its 3D shape.
- ☐ Decide whether it is polar overall.

Note: Write down the reasoning at each step.

Coming Up in Lecture 4



$$2A + 3B \rightarrow C$$

Topic: Stoichiometry & Counting Atoms

- Concept: How chemists count atoms by weighing them.
- Concept: How to calculate exactly how much of each substance a reaction needs.