Chemistry Fundamentals

Lecture 11: Chemical Bonding Fundamentals

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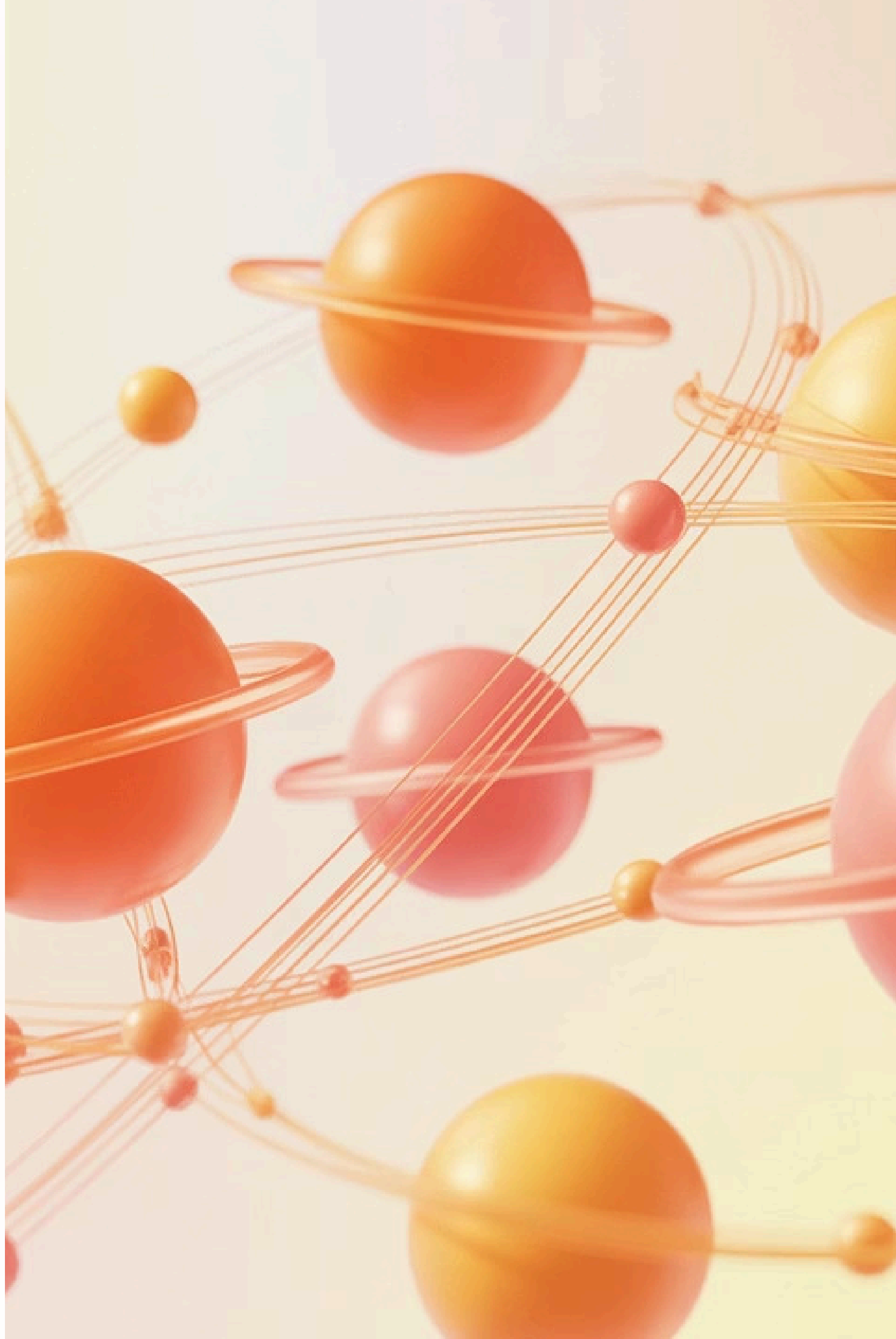
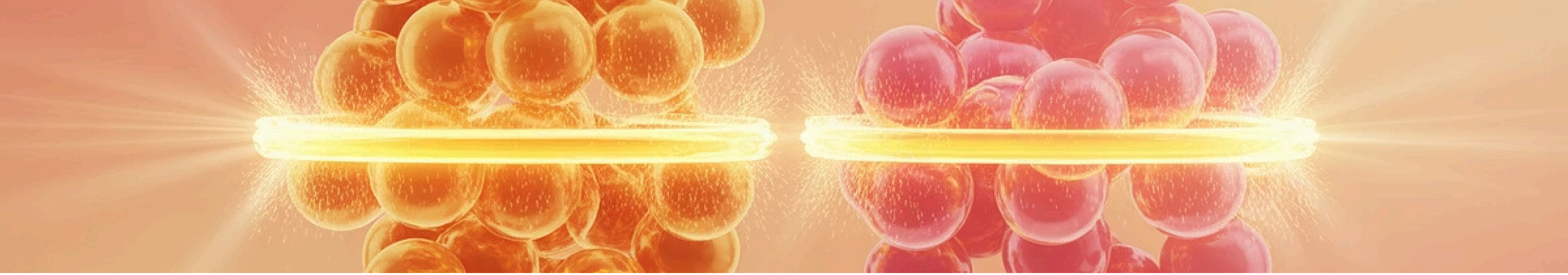




Why Atoms Bond - The Drive for Stability

Fundamental Principle

Atoms bond to achieve lower energy and greater stability

Noble Gas Rule

Atoms tend to gain, lose, or share electrons to achieve noble gas configuration

Electron Involvement

Valence electrons participate in bonding through various mechanisms

Bonding involves electrostatic attraction between oppositely charged particles, with atoms striving to achieve complete outer electron shells (octet rule). This results in three main types of bonds: ionic, covalent, and metallic, seen in everyday materials like salt (NaCl), water (H_2O), and metals (Fe).

Ionic Bonding - Electron Transfer

Ionic bonding occurs through the complete transfer of electrons from a metal to a nonmetal. Metal atoms lose electrons to become positively charged cations, while nonmetal atoms gain electrons to become negatively charged anions.

The resulting electrostatic attraction between these oppositely charged ions forms ionic compounds with distinctive properties:

- High melting points
- Conduct electricity when molten
- Brittle crystal structure

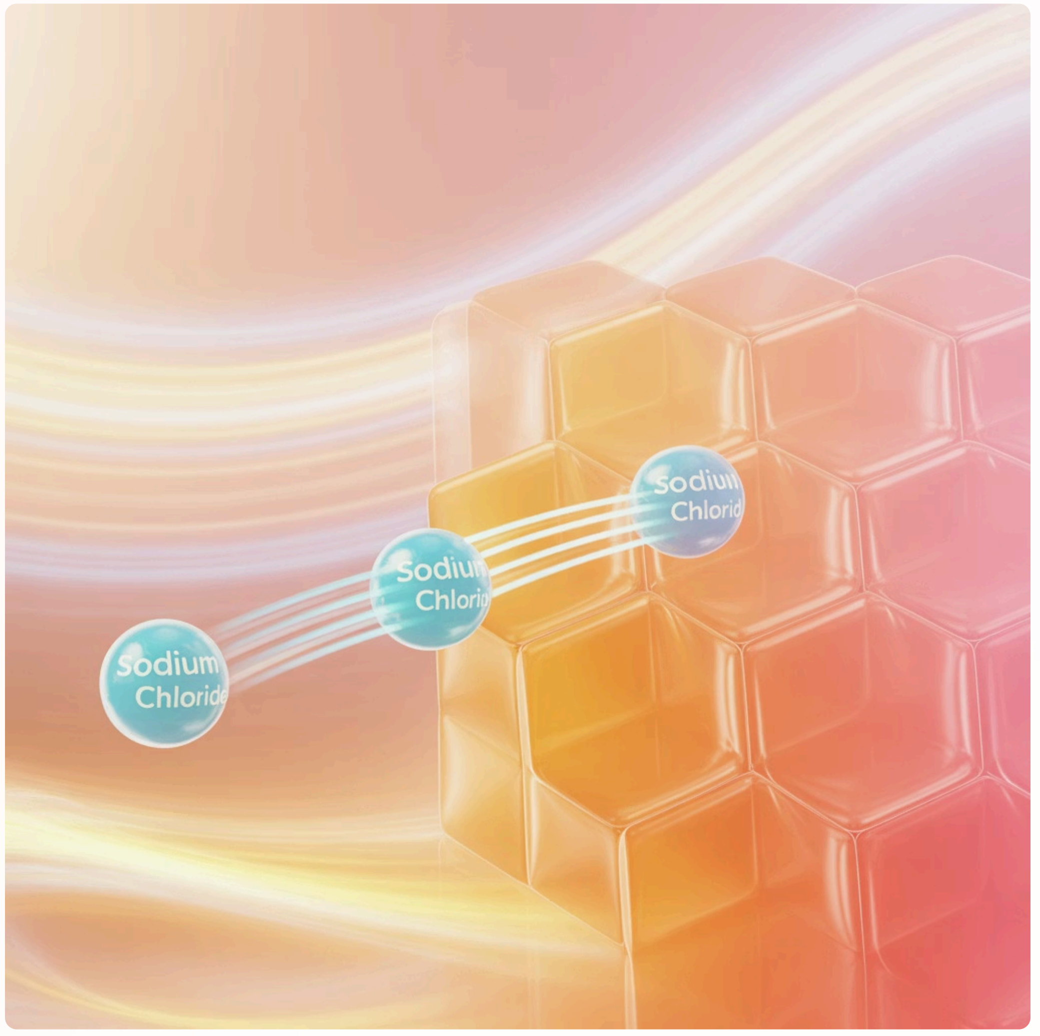
Example: Sodium Chloride (NaCl)

$Na \rightarrow Na^+ + e^-$ (loses 1 electron)

$Cl + e^- \rightarrow Cl^-$ (gains 1 electron)

$Na^+ + Cl^- \rightarrow NaCl$ (electrostatic attraction)

Both ions achieve noble gas configuration, with sodium resembling neon and chloride resembling argon.



Covalent Bonding - Electron Sharing

Covalent bonding involves the sharing of electron pairs between atoms, typically occurring between nonmetals with similar electronegativity.

1

Hydrogen Molecule (H_2)

Each H atom contributes 1 electron to create a shared pair

Both atoms achieve helium configuration (2 electrons)

Results in a single bond (H-H)

2

Water Molecule (H_2O)

Oxygen shares electrons with two hydrogen atoms

Oxygen achieves octet (8 electrons)

Each hydrogen achieves duet (2 electrons)

Covalent compounds typically have lower melting points than ionic compounds and don't conduct electricity. Bonds can be single, double, or triple depending on the number of shared electron pairs.

Electronegativity and Bond Polarity

Electronegativity measures an atom's ability to attract electrons in a bond, ranging from 0.7 (Cs) to 4.0 (F) on the Pauling scale. It increases across periods and decreases down groups in the periodic table.

The difference in electronegativity between bonded atoms determines bond polarity:

- 0.0-0.4: Nonpolar covalent (equal sharing)
- 0.4-1.7: Polar covalent (unequal sharing)
- >1.7: Ionic (electron transfer)

Example: H-Cl Bond

Hydrogen: 2.1

Chlorine: 3.0

Difference: 0.9

Result: Polar covalent bond (δ^{+}H-Clδ^{-})

Lewis Structures - Mapping Electron Arrangement

Lewis structures show valence electrons and bonding patterns in molecules using element symbols, dots for lone pairs, and lines for bonds.

Count total valence electrons

Add up all valence electrons from each atom

Arrange atoms

Place least electronegative atom in center

Connect atoms with single bonds

Each bond uses 2 electrons

Complete octets with remaining electrons

Add lone pairs to outer atoms first

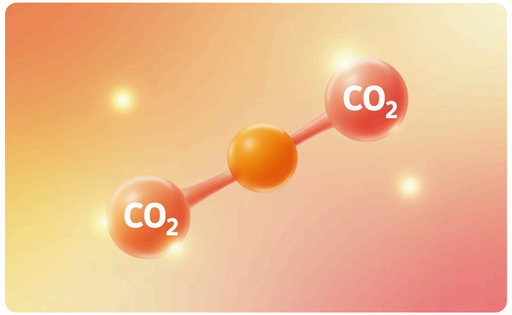
Form multiple bonds if needed

Convert lone pairs to bonds if octets aren't complete

Example: Water (H_2O) has 8 total valence electrons (6 from O + 1 from each H), resulting in H-O-H with 2 lone pairs on oxygen.

Molecular Shapes and VSEPR Theory

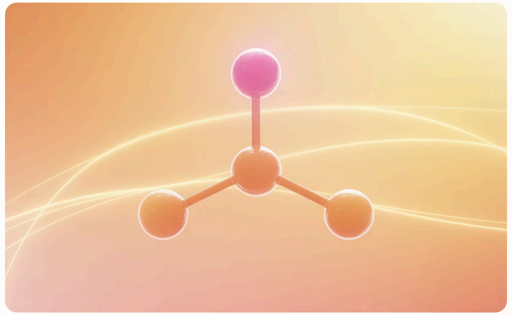
VSEPR (Valence Shell Electron Pair Repulsion) theory predicts molecular shapes based on the principle that electron pairs repel and arrange to minimize repulsion.



Linear

2 electron pairs (180°)

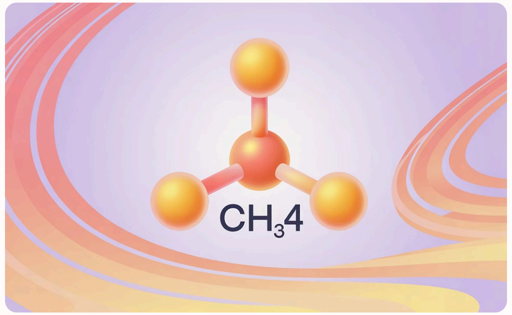
Examples: BeF_2, CO_2



Trigonal Planar

3 electron pairs (120°)

Example: BF_3



Tetrahedral

4 electron pairs (109.5°)

Example: CH_4

Bent

2 bonding + 2 lone pairs

Example: H_2O (104.5°)

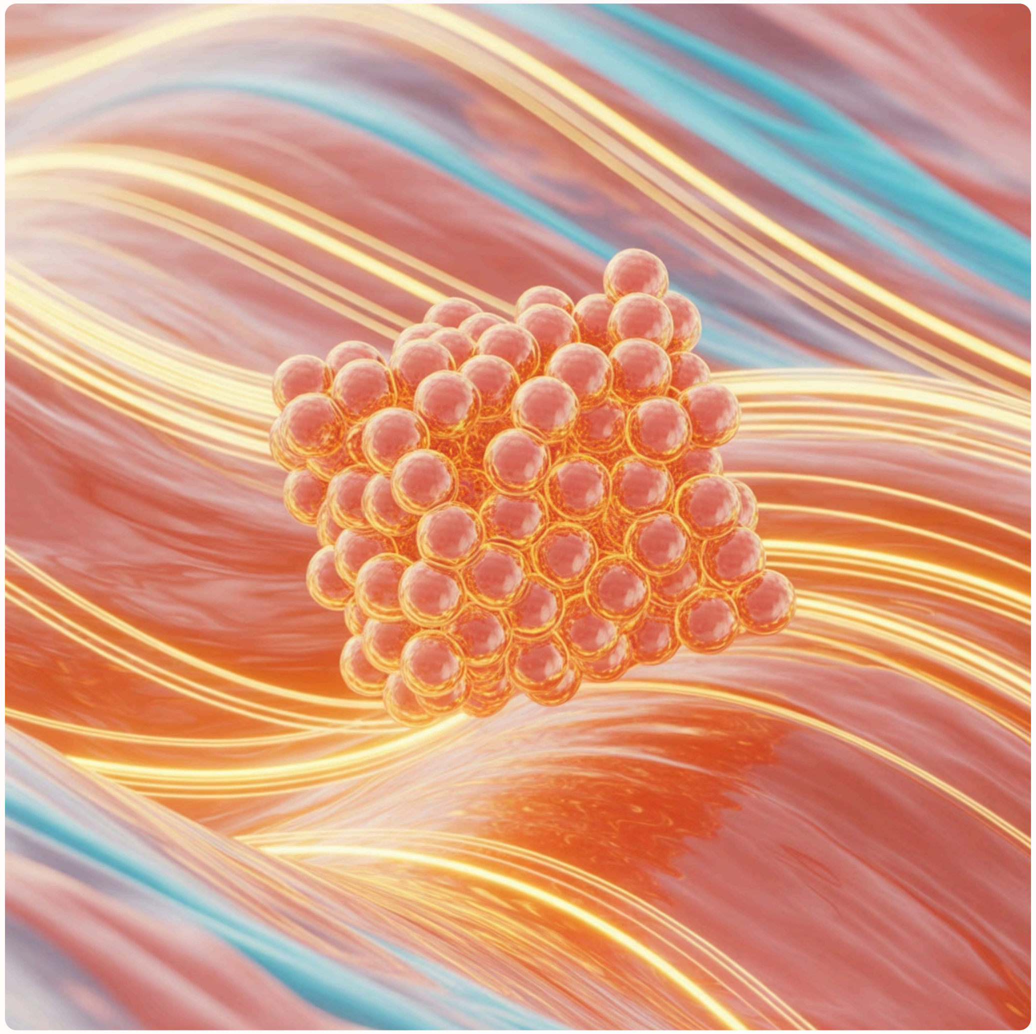
Lone pairs occupy more space than bonding pairs, affecting bond angles and molecular shape.

Metallic Bonding - Electron Sea Model

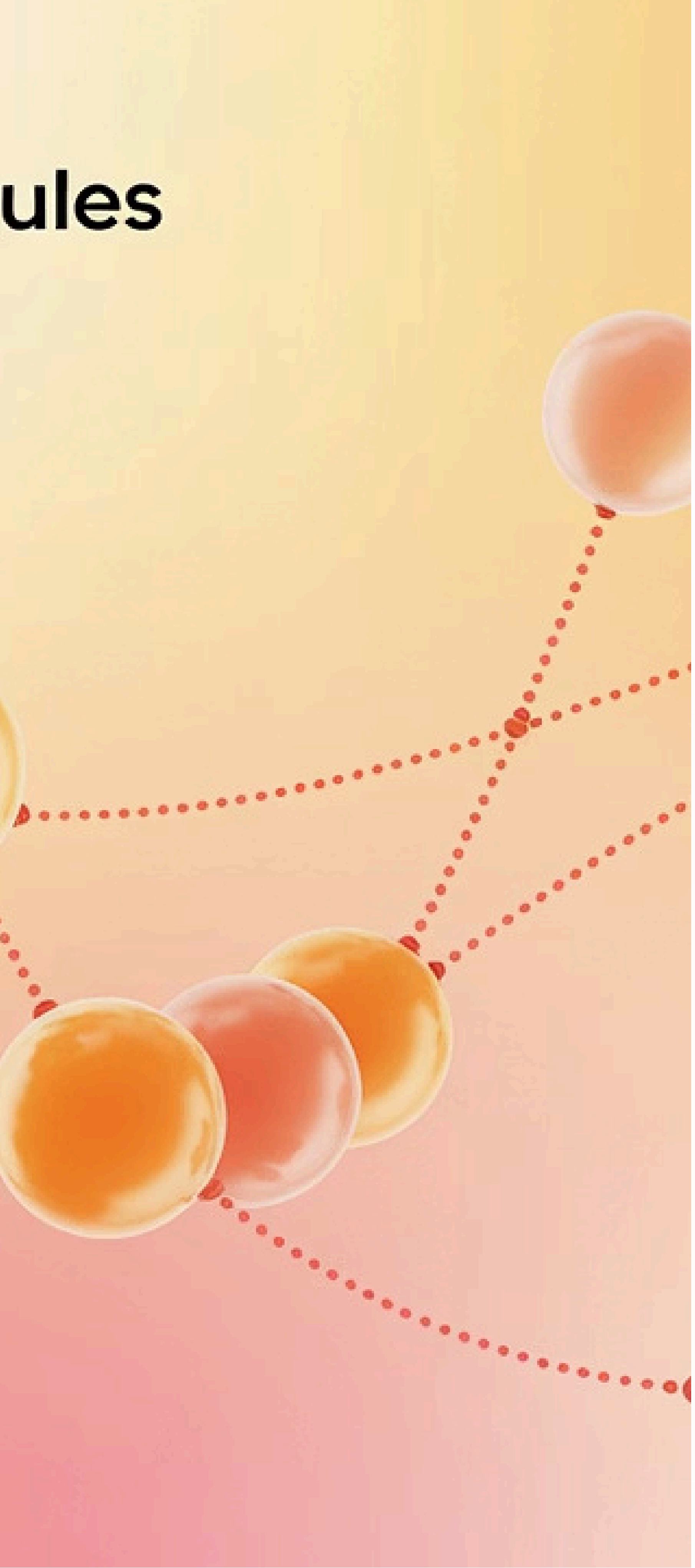
Metallic bonding occurs when metal atoms release their valence electrons into a "sea" of mobile electrons. The metal atoms become positively charged cations arranged in a regular pattern, while the delocalized electrons move freely throughout the structure.

This unique bonding explains key metallic properties:

- Electrical conductivity: Mobile electrons carry current
- Thermal conductivity: Electrons transfer kinetic energy
- Malleability: Layers of atoms slide past each other
- Metallic luster: Electrons absorb and re-emit light



Common examples include copper wiring, aluminum foil, and steel structures. Alloys are mixtures of metals with modified properties, such as bronze (copper + tin) or steel (iron + carbon).



Intermolecular Forces - Attractions Between Molecules

Intermolecular forces (IMFs) are weak attractions between separate molecules that influence physical properties like boiling points and solubility.

London Dispersion Forces

Temporary dipoles in all molecules

Weakest IMF, increases with molecular size

Example: Interactions between hexane molecules

Dipole-Dipole Forces

Permanent dipoles in polar molecules

Moderate strength

Example: Interactions between acetone molecules

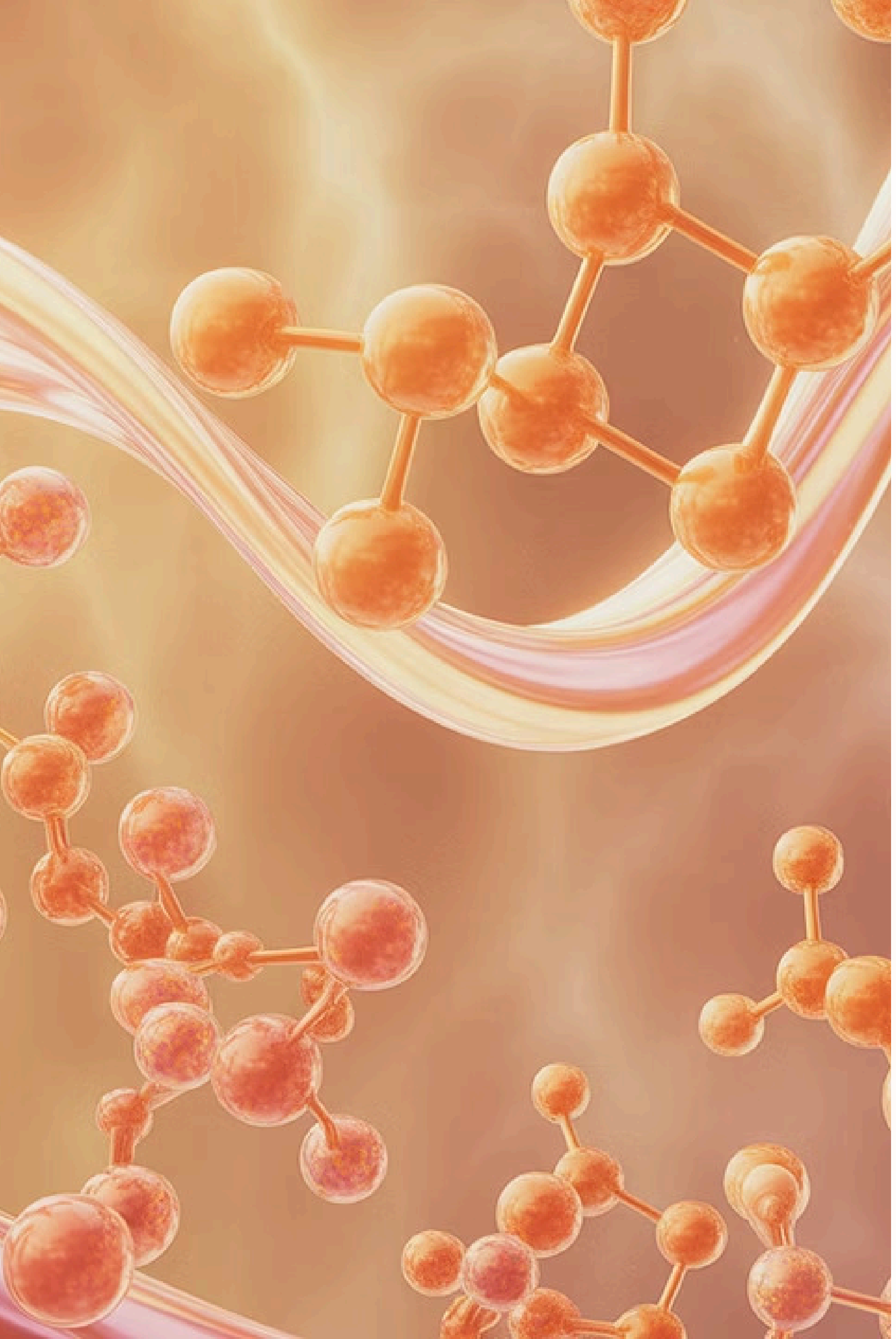
Hydrogen Bonding

H attached to N, O, or F

Strongest IMF

Example: Water's high boiling point

Strength comparison: Covalent bonds > hydrogen bonds > dipole-dipole > London forces. These forces are crucial in biological systems, influencing protein folding, DNA structure, and cell membrane formation.



Next Lecture: Molecular Compounds

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