

Principles of Inorganic Solids and Material Chemistry

Introduction

- Solids are not just static blocks of matter; they have intricate internal structures that dictate their strength, conductivity, and appearance.
- Inorganic solids differ from organic ones because they are not primarily based on carbon-hydrogen bonds.
- This module explores how atoms arrange themselves in crystals and how different bonding types create materials with vastly different properties (like why diamond is hard but graphite is soft).
- We will also look at the "Band Theory" to understand how some solids conduct electricity while others block it.

Learning Objectives

By the end of this module, you will be able to:

- Characterize inorganic solids based on their crystal structures and unit cells.
- Classify solids into Ionic, Metallic, Covalent Network, and Molecular categories based on their bonding.
- Apply **Band Theory** to distinguish between conductors, semiconductors, and insulators.
- Discuss the concept of crystal defects and how they influence material properties.

Key Concepts and Definitions

Term	Definition
Crystal Lattice	A highly ordered, repeating 3D arrangement of atoms, ions, or molecules in a solid.
Unit Cell	The smallest repeating building block of a crystal lattice (like a single brick in a wall).
Allotrope	Different physical forms of the same element in the same state (e.g., Carbon exists as Diamond, Graphite, and Buckyballs).
Coordination Number	The number of nearest neighbor atoms or ions surrounding a central particle in a crystal.
Band Gap	The energy difference between the valence band (where electrons live) and the conduction band (where electrons move freely).

Detailed Discussion

Crystal Structures and Unit Cells

Solids can be **Crystalline** (ordered structure) or **Amorphous** (disordered structure, like glass). In crystalline solids, we study the "Unit Cell" to understand the whole.

Common Cubic Unit Cells:

1. Simple Cubic (SC):

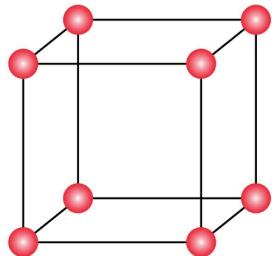
- Atoms are only at the 8 corners.
- Very much empty space; rare in nature (e.g., Polonium).
- Atoms per cell: 1

2. Body-Centered Cubic (BCC):

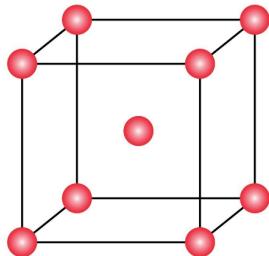
- Atoms at 8 corners + 1 atom in the dead center.
- More packed than simple cubic.
- Examples: Iron (Fe), Sodium (Na).

- Atoms per cell: 2
3. **Face-Centered Cubic (FCC):**
- Atoms at 8 corners + 1 atom on each of the 6 faces.
 - Most efficient packing (closest packed).
 - Examples: Copper (Cu), Aluminum (Al), Gold (Au).
 - Atoms per cell: 4

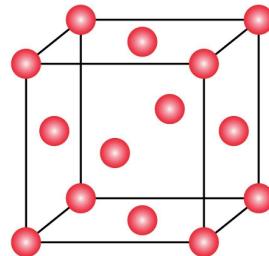
Three Types of Cubic Unit Cells



Simple cubic



Body-centered cubic



Face-centered cubic

Types of Inorganic Solids

The properties of a solid depend entirely on what holds it together.

1. Ionic Solids

- **Bonding:** Electrostatic attraction between Cations (+) and Anions (-).
- **Properties:** Hard, brittle, high melting points, non-conductive as solids (but conduct when melted/dissolved).
- **Example:** Sodium Chloride (NaCl), Magnesium Oxide (MgO).

2. Metallic Solids

- **Bonding:** "Sea of electrons." Metal cations float in a soup of delocalized electrons.
- **Properties:** Malleable (bendable), ductile (can make wires), good conductors of heat/electricity.
- **Example:** Copper (Cu), Iron (Fe).

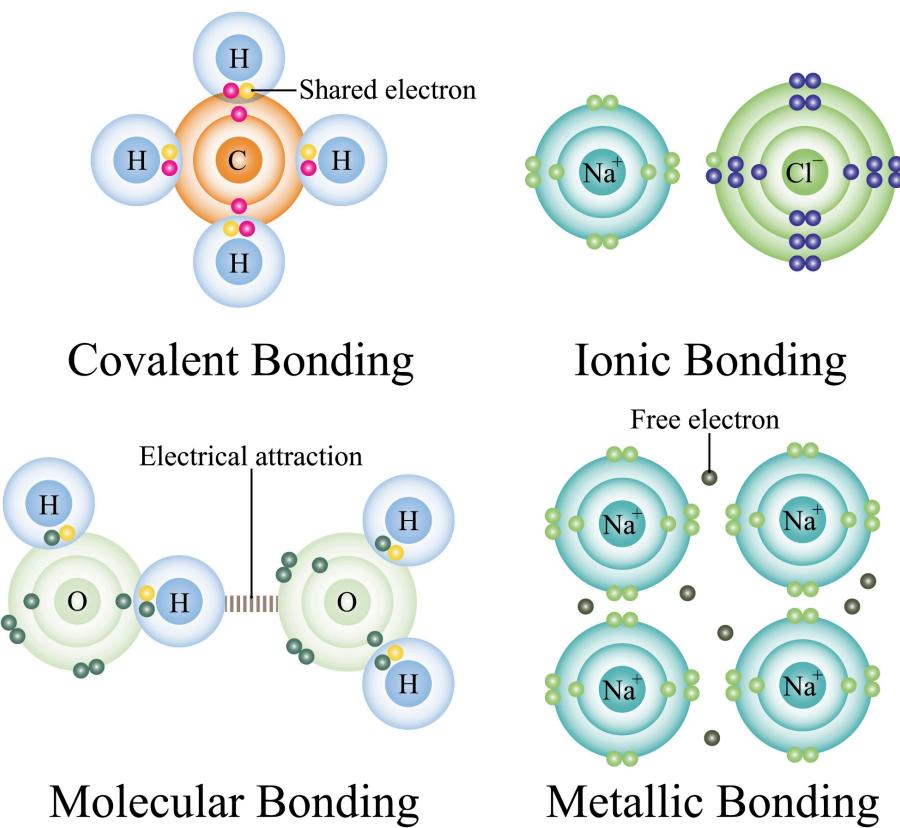
3. Covalent Network Solids

- **Bonding:** Atoms are bonded by a massive, continuous network of covalent bonds. Essentially one giant molecule.
- **Properties:** Extremely hard, very high melting points, usually insulators.
- **Example:** Diamond (C), Quartz (SiO₂).

4. Molecular Solids

- **Bonding:** Held together by weak intermolecular forces (Van der Waals, H-bonds).
- **Properties:** Soft, low melting points, poor conductors.
- **Example:** Ice (H₂O), Dry Ice (CO₂), Iodine (I₂).

Types of Chemical Bonding



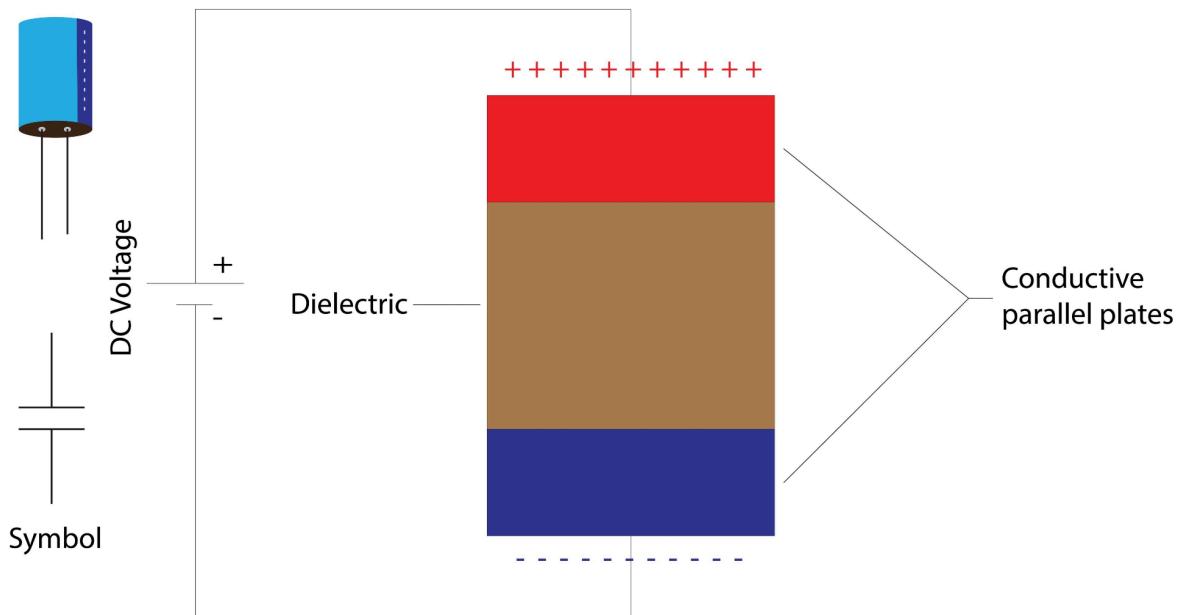
Band Theory of Solids

Why do metals conduct electricity while rubber does not? We use Band Theory to explain this.

- **Valence Band:** The energy band occupied by valence electrons (lower energy).
- **Conduction Band:** The energy band where electrons can move freely to conduct electricity (higher energy).
- **The Gap:** The space between these bands determines conductivity.

Classifications:

1. **Conductor (Metals):** The Valence and Conduction bands **overlap**. Electrons flow freely with no energy cost.
2. **Semiconductor (Silicon):** There is a **small gap**. Electrons can jump across if given energy (heat or light).
3. **Insulator (Glass/Wood):** There is a **large gap**. Electrons are trapped in the valence band and cannot cross.



References

1. Chang, R., & Goldsby, K. A. (2016). *Chemistry* (12th ed.). New York, NY: McGraw-Hill Education.
2. Zumdahl, S. S., & Zumdahl, S. A. (2014). *Chemistry* (9th ed.). Belmont, CA: Brooks/Cole, Cengage Learning.
3. Brown, T. L., LeMay, H. E., Bursten, B. E., Murphy, C. J., & Woodward, P. M. (2017). *Chemistry: The Central Science* (14th ed.). Pearson.