Bonding and Structures

GCSE Chemistry

March 10, 2025

Ionic Bonding

Definition:

- Occurs when a metal atom reacts with a non-metal atom.
- Electrons are transferred to achieve a **full outer shell**.
- **Metal atoms lose electrons** to become positively charged ions.
- **Non-metal atoms gain electrons** to become negatively charged ions.

Key Details:

- Ions formed by Groups 1, 2, 6, and 7 have the electronic structure of noble gases.
- The charge on the ions relates to the group number of the element.

Ionic Compounds

Definition:

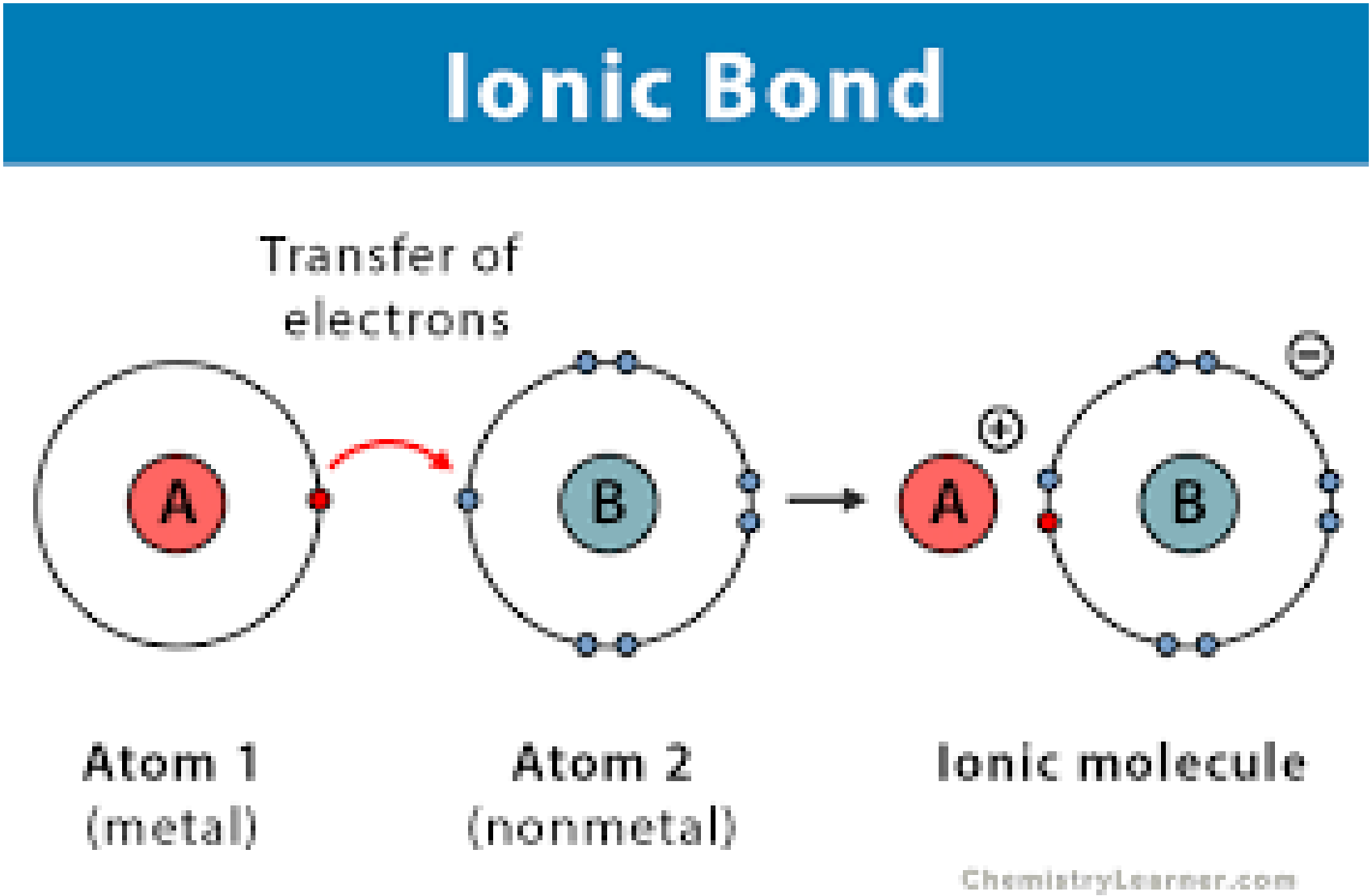
- Giant ionic structures (**giant ionic lattice**).
- Held together by **strong electrostatic attractions** between oppositely charged ions.
- These forces act in all directions.

Properties of Ionic Compounds

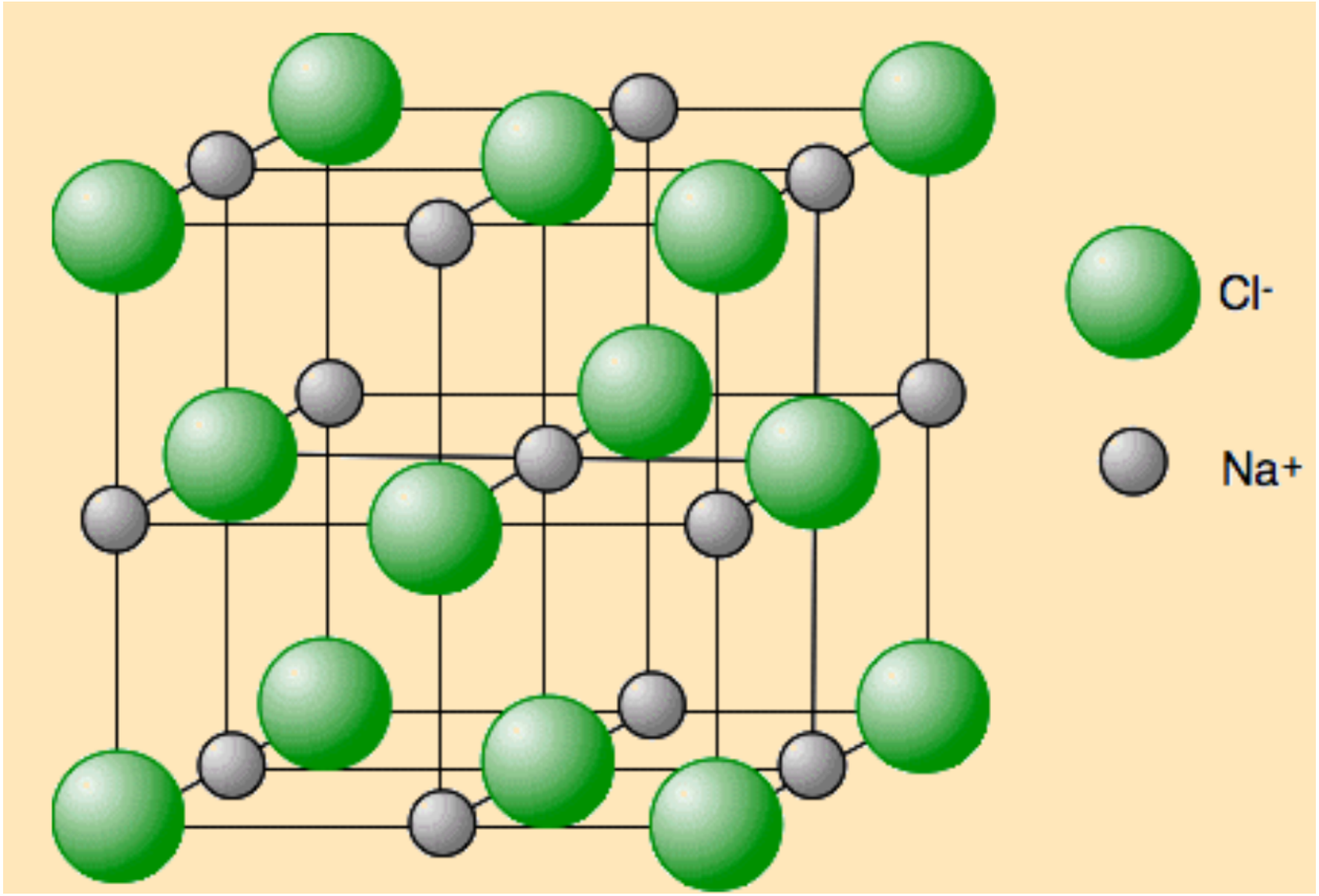
Key Properties:

- **High melting and boiling points:** Large amounts of energy are required to break strong electrostatic attractions.
- **Do not conduct electricity as solids:** Ions cannot move.
- **Conduct electricity when molten or dissolved:** Ions are free to move.

Ionic Bonding Diagram



Giant Ionic Lattice



Covalent Bonding

Definition:

- Atoms share pairs of electrons.
- Strong bonds form between atoms.

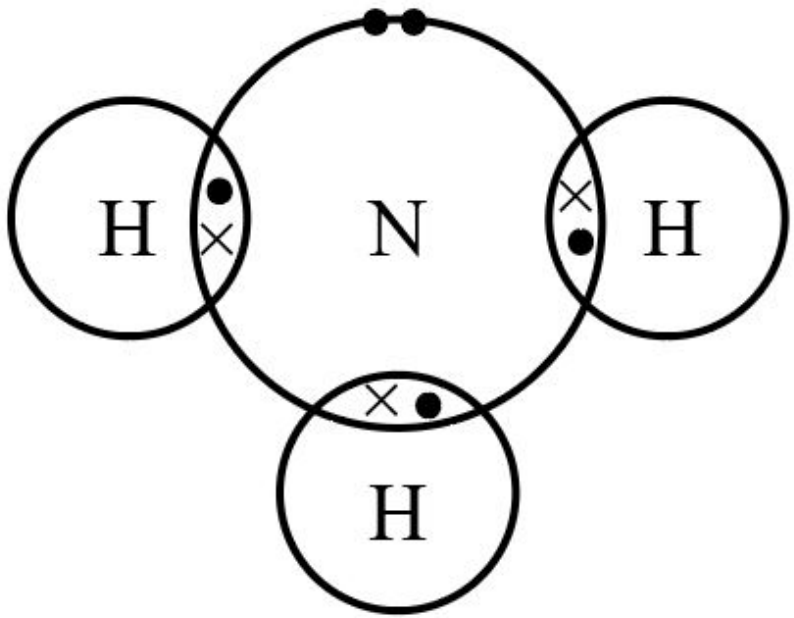
Examples:

- Covalently bonded substances include small molecules, large molecules (polymers), and giant covalent structures (e.g., diamond, graphite).

Representing Covalent Bonds

Example: Ammonia (NH_3)

- Diagram shows shared electron pairs between nitrogen and hydrogen atoms.



Properties of Small Covalent Molecules

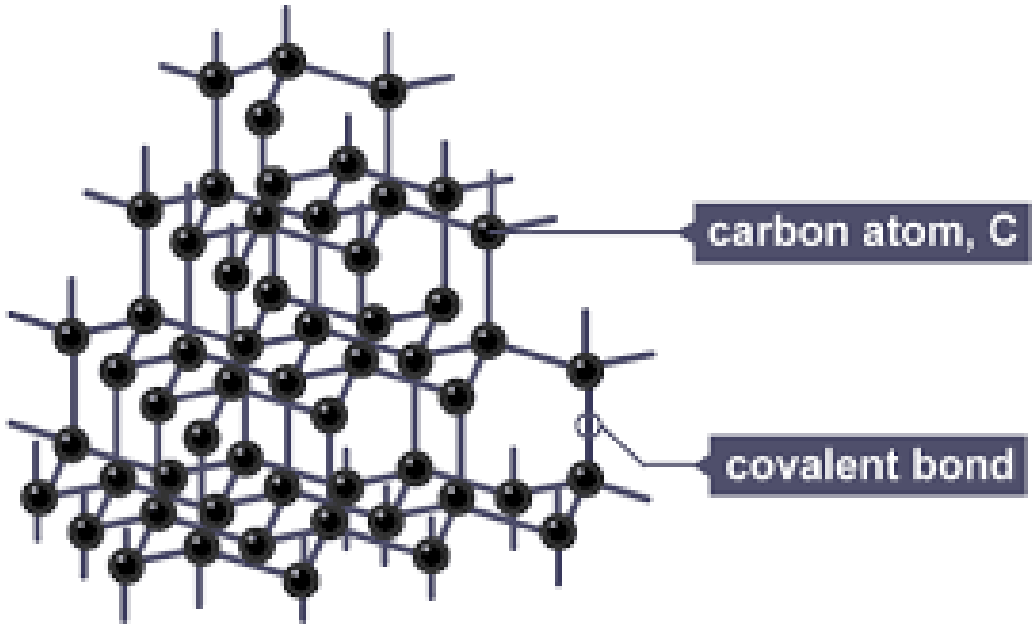
Key Features:

- Usually gases or liquids with low melting and boiling points.
- Weak intermolecular forces are easily broken, not the covalent bonds.
- Do not conduct electricity: No overall electric charge.
- Intermolecular forces increase with molecule size, leading to higher melting and boiling points.

Giant Covalent Structures

Key Features:

- All atoms are linked by strong covalent bonds.
- Very high melting and boiling points due to the strength of the bonds.



Diamond

Structure:

- Each carbon atom forms four covalent bonds with other carbon atoms.

Properties:

- Does not conduct electricity (no delocalised electrons).
- Extremely hard and has a very high melting point.

Graphite

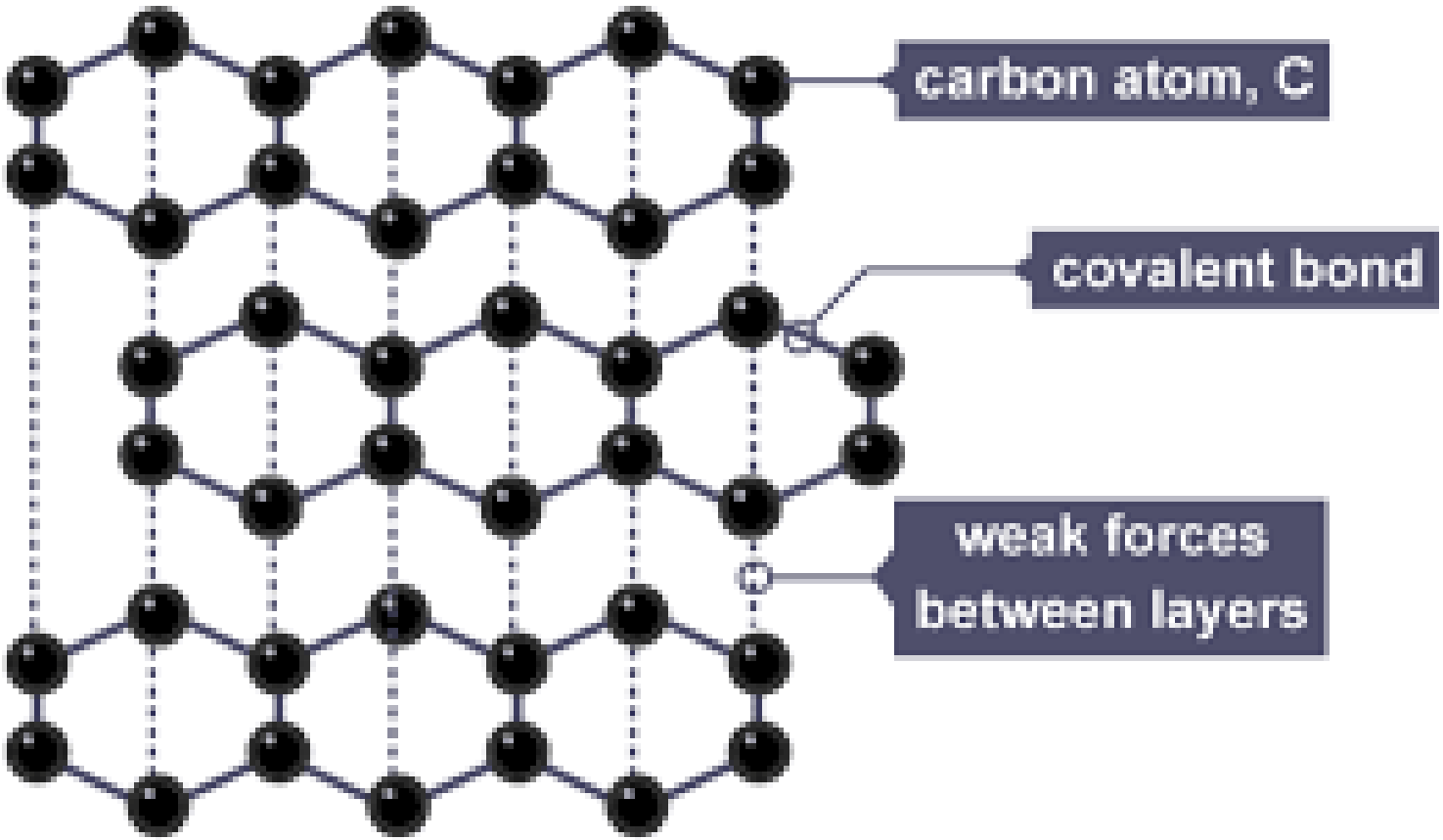
Structure:

- Each carbon atom forms three covalent bonds, creating layers of hexagonal rings.
- One electron is delocalised, allowing graphite to conduct electricity.

Properties:

- Good conductor of electricity.
- Layers slide over each other due to weak intermolecular forces, making graphite a good lubricant.
- Very high melting point.

Example



Graphene and Fullerenes

Graphene:

- Single layer of graphite.
- Used in electronics and composites.

Fullerenes:

- Molecules of carbon atoms with hollow shapes.
- Structure based on hexagonal, pentagonal, or heptagonal rings.
- Examples: Buckminsterfullerene (C_{60}), carbon nanotubes.

Carbon Nanotubes:

- Cylindrical fullerenes with high length-to-diameter ratios.
- Used in nanotechnology, electronics, and materials.

Metallic Bonding Overview

Key Features of Metallic Structures:

- Giant structures of atoms arranged in a regular pattern.
- Electrons in the outer shell are **delocalised**, allowing them to move freely.
- Delocalised electrons create strong **metallic bonds** through sharing.

Properties of Metallic Structures

Key Properties:

- **High Melting and Boiling Points:** Due to strong metallic bonds.
- **Good Conductors:** Delocalised electrons transfer heat and electricity.
- **Malleable and Ductile:** Atoms can slide over each other.

Why Conductivity Works:

- Electrons carry electrical charge through the structure.
- Heat energy is transferred efficiently by moving electrons.

Pure Metals

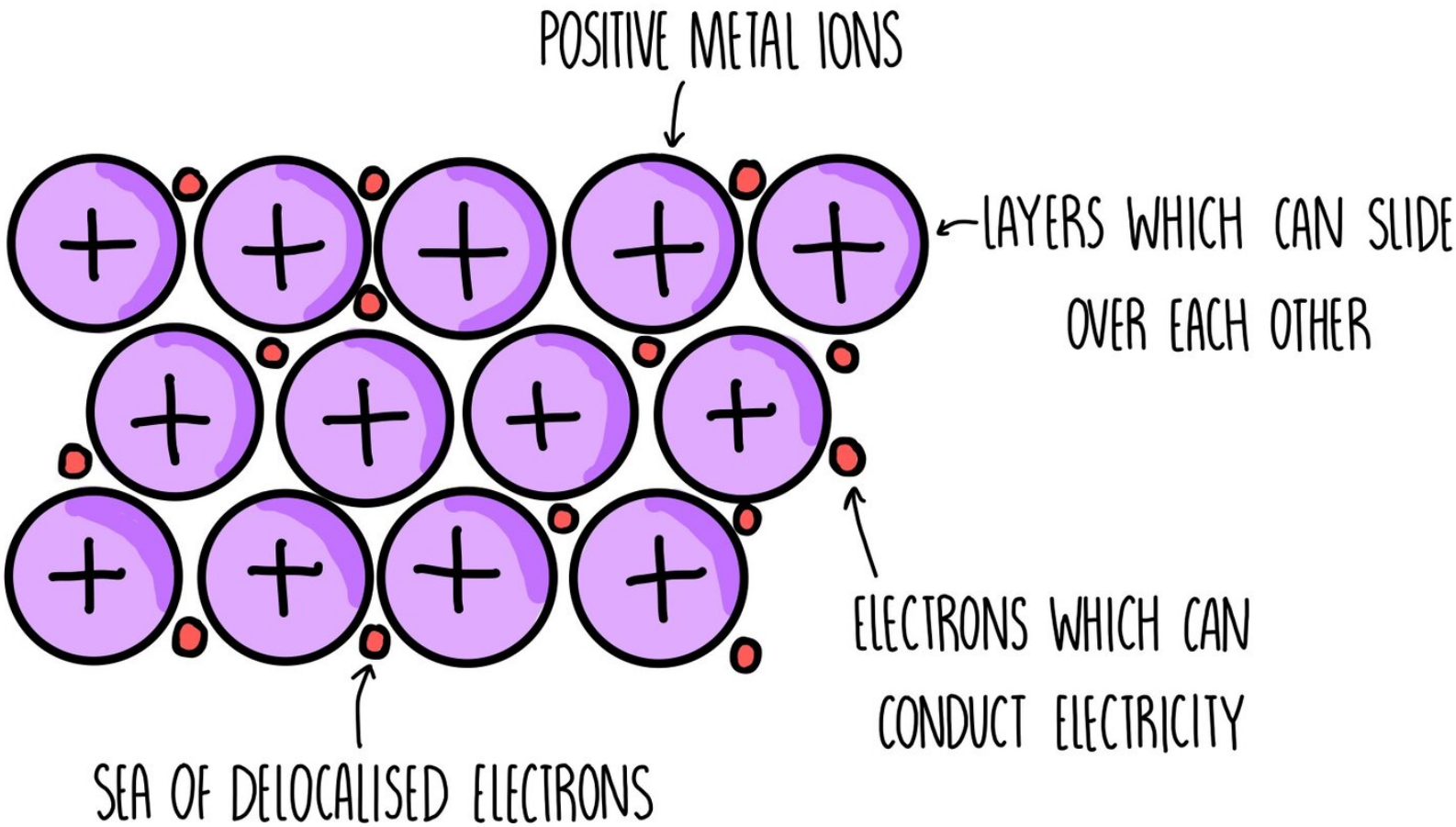
Key Features:

- Atoms are arranged in **layers**.
- Layers can **slide over each other**, making metals bend and shape easily.

Limitation of Pure Metals:

- Too **soft** for many uses.
- Requires strengthening through mixing with other elements.

Metallic Bond



Alloys

Definition:

- A mixture of **two or more elements**, at least one of which must be a metal.

Properties of Alloys:

- **Harder than pure metals:** Layers are distorted, preventing sliding.
- Customisable for specific properties, such as strength, corrosion resistance, or flexibility.

Why Are Alloys Harder?

Key Reasons:

- Different sized atoms distort the regular layers in the structure.
- Distorted layers cannot slide over each other easily.

Examples:

- **Steel:** An alloy of iron and carbon, used for construction.
- **Brass:** An alloy of copper and zinc, used for musical instruments.

Alloy Diagram