

The metallic model

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout teaches Structure 2.3, The metallic model, for the course first assessed in 2025. Sections marked **AHL** are for Higher Level only.

Prerequisites: electron configurations and ionization energy from Structure 1.3, and the ionic and covalent models from Structures 2.1 and 2.2. You will compare melting points from the data booklet.

The metallic bond

■ Why can a copper wire bend and conduct?

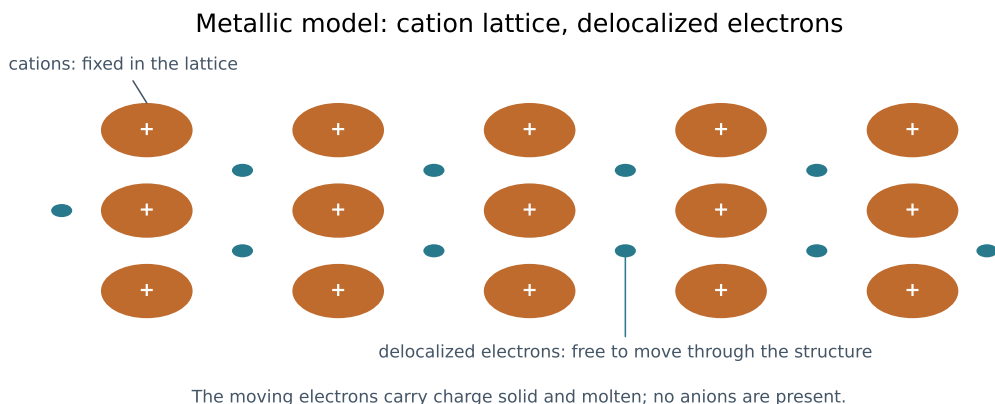
A copper wire carries current when it is cold and when it is red hot. You can bend it, hammer it flat or pull it into a thinner wire, and it does not shatter. An ionic crystal does neither: it conducts only when molten and it breaks under a hammer. The **metallic model** explains both properties of metals with one picture.

■ Cations in a sea of electrons

Metal atoms have few outer electrons and low ionization energies, so they lose their outer electrons easily. In a solid metal:

- the atoms lose their outer electrons and become positive ions, **cations**, packed in a regular **lattice**;
- the lost electrons are not attached to any one ion. They are **delocalized electrons**, free to move through the whole structure, often called a “sea” of electrons.

The **metallic bond** is the electrostatic attraction between a lattice of cations and the delocalized electrons.



Fixed cations in a sea of delocalized electrons; the moving electrons conduct

Three points the examiners look for:

- There are **no anions** in a metal. The negative charge is carried by electrons.
- The attraction acts in every direction: the bonding is **non-directional**
- Each ion is attracted to many electrons, and each electron to many ions, so there are no separate bonds between pairs of atoms.

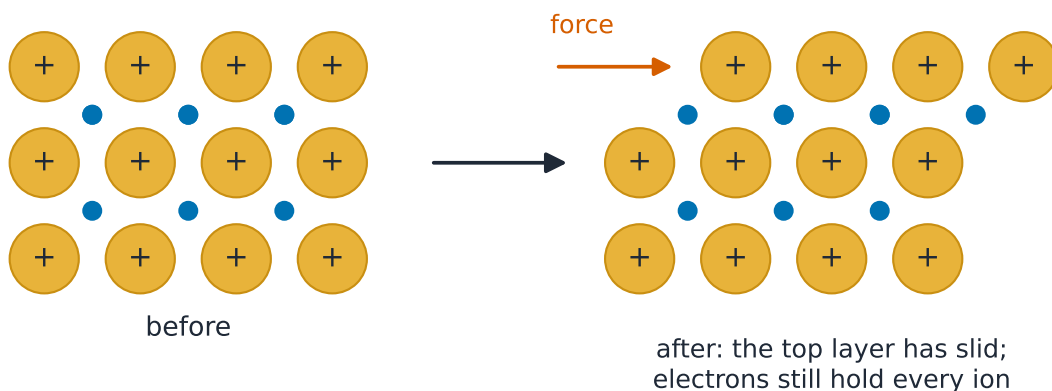
Properties of metals

■ Explaining each property

To explain any property, name the structure, the particles, the force and the mobile charge carriers.

Property	Explanation
electrical conductivity in solid and liquid	delocalized electrons move through the lattice when a voltage is applied
thermal conductivity	delocalized electrons (and vibrating ions) pass kinetic energy quickly through the metal
malleability and ductility	layers of cations can slide over each other; the electrons move with them, so the bonding is unchanged
high melting and boiling points (most metals)	strong attraction between many cations and the electron sea in a giant lattice
shiny (lustre)	delocalized electrons absorb and re-emit light

Malleable means it can be hammered into shape; ductile means it can be drawn into a wire.



The bonding is non-directional, so a shifted layer is held just as strongly as before

■ Metals compared with ionic compounds

	Metal	Ionic compound
particles	cations and delocalized electrons	cations and anions
conducts when solid	yes, by electrons	no: ions are fixed
conducts when molten	yes, by electrons	yes, by ions
effect of a blow	layers slide, metal bends	like charges meet, crystal shatters

Notice the different charge carriers: always say **electrons** for a metal and **ions** for a molten ionic compound.

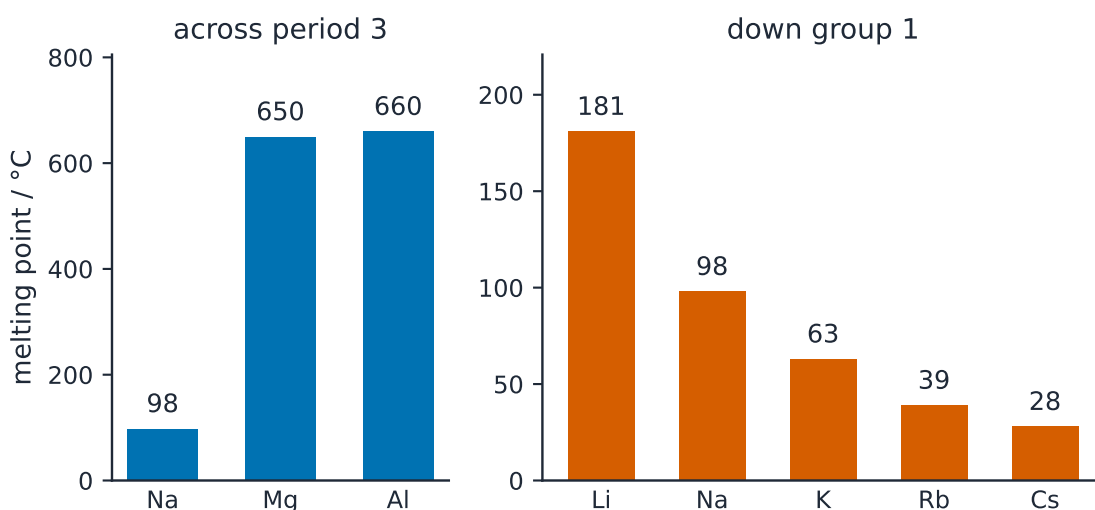
Strength of the metallic bond

■ What makes the bond stronger

The metallic bond is stronger when the attraction between the cations and the electron sea is stronger. Two factors decide this:

1. **Charge of the cation and number of delocalized electrons.** A larger charge gives more delocalized electrons per ion and a stronger attraction.
2. **Radius of the cation.** A smaller ion lets the electrons come closer to the nucleus, so the attraction is stronger.

A stronger metallic bond means a higher melting point and boiling point.



The two charts use different vertical scales. Charge raises the melting point across period 3; size lowers it down group 1.

■ Across period 3: Na, Mg, Al

- Na^+ gives one delocalized electron per ion, Mg^{2+} two and Al^{3+} three.
- The ions also get smaller: Na^+ 102 pm, Mg^{2+} 72 pm, Al^{3+} 54 pm.
- Both changes strengthen the attraction, so the melting point rises from 98 °C to 650 °C and 660 °C.

Boiling points show the trend more clearly than melting points: Na 883 °C, Mg 1090 °C, Al 2519 °C. Melting only loosens the lattice, but boiling separates the ions from each other completely.

■ Down group 1: Li to Cs

Every group 1 metal gives one delocalized electron per ion. Down the group the cations get larger, so the delocalized electrons are further from each nucleus and the attraction is weaker. The melting point falls from 181 °C (Li) to 28 °C (Cs): caesium melts in a warm hand.

■ Worked example

Explain why magnesium has a higher melting point than sodium.

- **Structure:** both are giant metallic lattices; melting overcomes the metallic bonding.
- **Ions:** Mg^{2+} has a higher charge than Na^+ and is smaller.
- **Electrons:** magnesium has two delocalized electrons per atom, sodium only one.
- **Conclusion:** the attraction between cations and delocalized electrons is stronger in magnesium, so more energy is needed to melt it.

AHL: Transition elements

■ Delocalized d electrons

The **transition elements** have partly filled 3d sublevels. The 4s and 3d sublevels are very close in energy, so both the 4s and some of the 3d electrons can become delocalized. More delocalized electrons per ion and a high ionic charge give a much stronger metallic bond than in the s-block metals.

Metal	Block	Delocalized electrons per atom (approx.)	Melting point / °C
potassium	s	1	63
calcium	s	2	842
iron	d	more than 2 (4s and 3d)	1538
chromium	d	more than 2 (4s and 3d)	1907

So transition metals have:

- **high melting points** and high boiling points;

- high **density** , because the ions are small and heavy and pack closely;
- high **electrical conductivity**, because there are many delocalized electrons to carry charge;
- great strength and hardness, which is why iron, titanium and their alloys are used for structures.

Zinc, with a full 3d sublevel, does not use d electrons in this way: it melts at only 420 °C, much lower than its neighbours.

Practice

1. Define the metallic bond.
2. Explain why solid copper conducts electricity but solid copper(II) chloride does not.
3. Explain why a metal bends when hammered, but a crystal of sodium chloride shatters.
4. Explain why aluminium has a higher boiling point than sodium.
5. Explain why the melting point of the group 1 metals decreases from lithium to caesium.
6. Predict, with a reason, which has the higher melting point: potassium or calcium.
7. **AHL**. Explain why iron has a much higher melting point than calcium.

Answers.

1. The electrostatic attraction between a lattice of positive ions (cations) and delocalized electrons.
2. Copper contains delocalized electrons that can move through the lattice and carry charge. In solid copper(II) chloride the ions are held in fixed positions and there are no mobile electrons, so no charge can move.
3. In a metal the bonding is non-directional: layers of cations slide and the delocalized electrons move with them, so the attraction holds the new arrangement. In NaCl a shift brings ions of the same charge next to each other, they repel and the crystal splits.
4. Al^{3+} has a higher charge and a smaller radius than Na^+ , and aluminium has three delocalized electrons per atom against one. The attraction between cations and delocalized electrons is stronger, so more energy is needed to separate the ions.
5. Each ion gives one delocalized electron, but the ions get larger down the group. The delocalized electrons are further from the nucleus, the attraction is weaker, and less energy is needed to melt the metal.
6. Calcium: Ca^{2+} has a higher charge and is smaller than K^+ , and gives two delocalized electrons per atom, so its metallic bond is stronger (842 °C against 63 °C).

7. **AHL.** Iron is a transition element. Its 4s and 3d electrons are close in energy, so 3d electrons are delocalized as well as 4s ones. More delocalized electrons and small, highly charged ions make the metallic bond much stronger.

Exam tips

- Define the metallic bond as an attraction between **cations** and **delocalized electrons**; never mention anions or “molecules”.
- Metals conduct by **delocalized electrons** in both solid and liquid states; molten ionic compounds conduct by **ions**.
- For malleability, say that layers **slide** and the **non-directional** bonding is kept.
- Compare metallic bond strength with **charge** (number of delocalized electrons) and **ionic radius**, and link it to the energy needed.
- **AHL.** For transition metals, say that **d electrons** are delocalized too.