MEGA LECTURE

IB DP CHEMISTRY

Structure 2: Models of Bonding and Structure

Structure 2.3 The Metallic Model

Revision Notes · Standard and Higher Level

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Original notes prepared for the IB Diploma Programme Chemistry course (first assessment 2025)

What the syllabus requires

Structure 2.3 completes the trio of bonding models. Metals hold their atoms together in a distinctive way, and the model must account for the whole family of characteristic metallic properties. Use this checklist as a final revision sweep.

Understanding	You should be able to...
The metallic model	Describe a metal as a lattice of positive ions (cations) held together by a ‘sea’ of delocalised electrons.
Origin of the bond	Explain metallic bonding as the electrostatic attraction between the lattice of cations and the delocalised electrons.
Bond strength	Explain trends in metallic bond strength using the number of delocalised electrons and the charge density (charge/radius) of the cation.
Physical properties	Explain electrical and thermal conductivity, malleability, ductility, high melting and boiling points and lustre in terms of the model.
Alloys	Describe alloys as mixtures of a metal with other elements, and explain why they are usually harder and stronger than the pure metal.
Transition metals (HL)	Explain the involvement of d electrons in the bonding of transition elements and relate this to their high melting points and densities.

> **Exam note:** The core metallic model, bond-strength trends, properties and alloys are common to SL and HL. The role of d electrons in transition-metal bonding is **HL only** and is flagged (HL) throughout these notes.

1. The metallic bonding model

A pure metal is built from atoms of a single element, all with low ionisation energies. When these atoms pack together in the solid, each atom releases its outer (valence) electrons, which are no longer tied to any one atom but are free to move throughout the whole structure. What remains is a regular three-dimensional **lattice of positive ions** (cations), embedded in a **‘sea’ of delocalised electrons**.

Metallic bonding is defined as the **electrostatic attraction between the lattice of cations and the sea of delocalised electrons**. This attraction operates in all directions (it is non-directional) and binds the whole lattice into a single giant structure. There are no discrete molecules and no fixed bonds between named pairs of atoms — the electrons are shared communally by every cation at once.

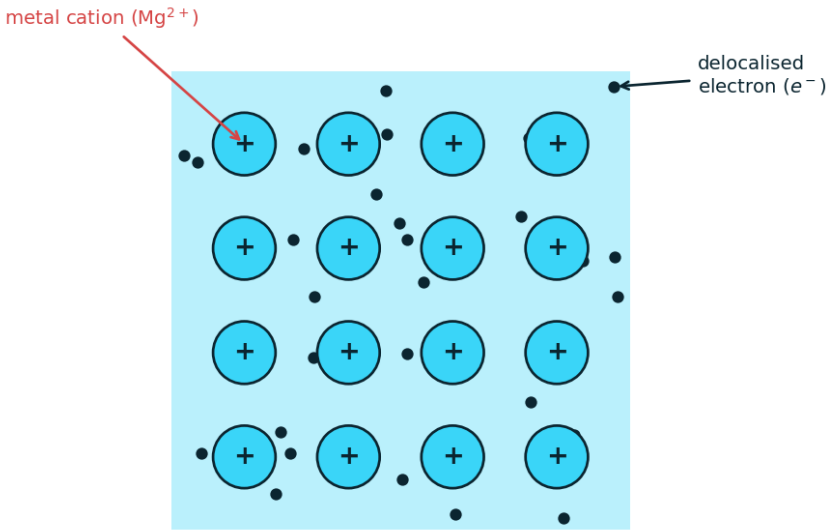

Figure 1. The metallic model: a regular lattice of positive cations held within a mobile 'sea' of delocalised electrons. Metallic bonding is the electrostatic attraction between the cations and this shared electron sea, acting in all directions.

Building the picture

- Each cation carries a charge equal to the number of valence electrons it has donated: sodium gives Na^+, magnesium gives Mg^{2+}, aluminium gives Al^{3+}.
- The delocalised electrons are not lost from the metal; they belong to the whole sample and keep it electrically neutral overall.
- The cations are fixed at lattice points but the electron sea is mobile — this mobility is the single feature that explains most metallic properties.
- Metallic bonding is generally strong; the giant lattice means large amounts of energy are needed to break it apart.

> **Key phrase:** Examiners look for the exact wording 'lattice of cations in a sea of delocalised electrons' and 'electrostatic attraction between the cations and the delocalised electrons'. Learn both.

2. What makes a metallic bond strong?

The strength of metallic bonding — measured by properties such as melting point, boiling point and hardness — depends on how strongly the electron sea is attracted to the cations. Two factors control this.

Factor	Effect on bond strength
Number of delocalised (valence) electrons per atom	More valence electrons released means a denser electron sea and more charge on each cation, so the electrostatic attraction is greater and the bond is stronger.
Charge density of the cation (charge ÷ ionic radius)	A higher charge and/or a smaller ionic radius concentrates the positive charge, pulling the electron sea in more tightly. High charge density means stronger bonding.

Trend across a period

Moving left to right across Period 3 (Na → Mg → Al), each atom donates one more valence electron (1, 2 then 3). The cation charge rises (Na^{+}, Mg^{2+}, Al^{3+}) while the ionic radius falls, so charge density increases. Both changes strengthen the bond, so melting point rises in the order Na < Mg < Al.

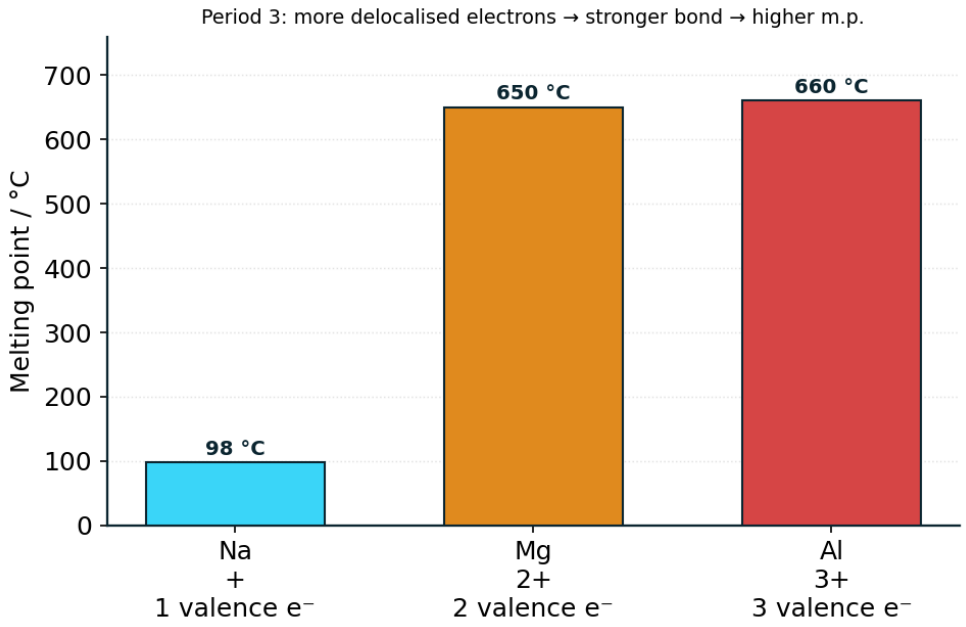

Figure 2. Across Period 3 each atom donates more valence electrons (1 → 3) and the cation charge density rises, so the metallic bond strengthens and the melting point increases Na < Mg < Al.

Trend down a group

Descending Group 1 (Li → Na → K → Rb → Cs), the number of delocalised electrons per atom stays at one, but the cations get steadily larger. The valence electrons are further from the nucleus and more shielded, so charge density falls and the electron sea is held less tightly. Metallic bonding weakens and melting point decreases down the group.

Metal	Valence e⁻ donated	Cation	Relative charge density	Melting point / °C
Sodium, Na	1	Na^{+}	low	98
Magnesium, Mg	2	Mg^{2+}	medium	650
Aluminium, Al	3	Al^{3+}	high	660
Potassium, K	1	K^{+}	very low	63

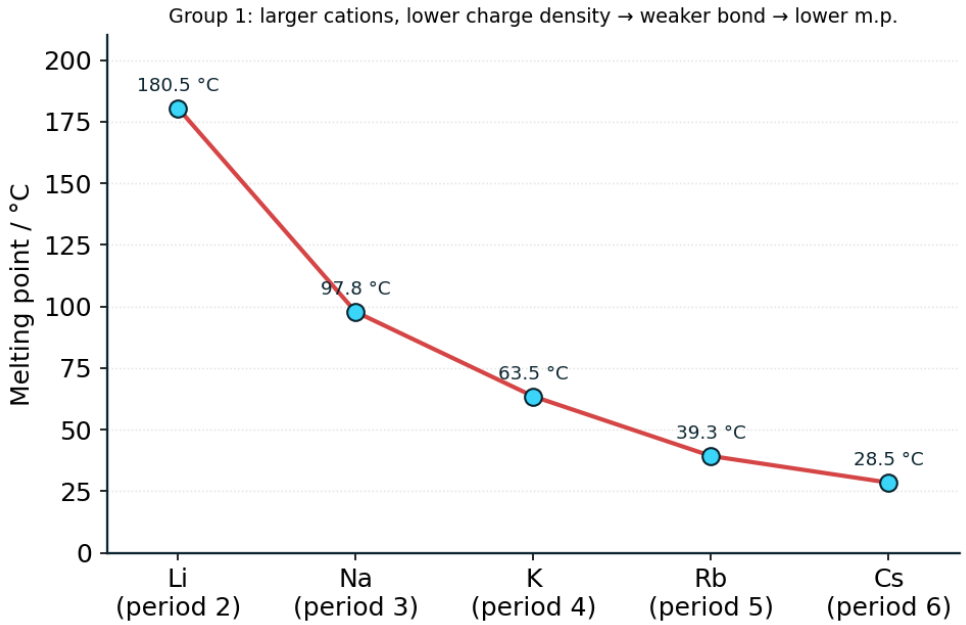

Figure 3. Down Group 1 the cations become larger and their charge density falls, so the electron sea is held less tightly: metallic bonding weakens and the melting point decreases steadily from lithium to caesium.

> **Watch out:** Down Group 1 the charge is constant at 1+, so you must argue from increasing ionic radius (lower charge density), not from charge. Across a period you can use both more electrons and higher charge density.

3. Physical properties explained by the model

The value of a good model is that one picture — mobile electrons around fixed cations — explains a whole set of properties. Learn to *explain* each one, not just state it.

Electrical conductivity

Metals conduct electricity in both the solid and the liquid state because the delocalised electrons are free to move. When a potential difference is applied, the electrons drift through the lattice, carrying charge. Crucially the ions themselves do not move, so conduction does not change the chemical composition of the metal.

Thermal conductivity

Metals are excellent conductors of heat. The mobile electrons gain kinetic energy at the hot end and transfer it rapidly as they move through the lattice; vibrations of the cations pass energy along as well. The free electrons make metals conduct heat far faster than non-metals, which rely on lattice vibrations alone.

Malleability and ductility

Metals can be hammered into sheets (malleable) and drawn into wires (ductile). When a force is applied, layers of cations slide over one another to new positions. Because the electron sea is non-directional and simply flows with the ions, the same attractive environment is restored after slipping — the metallic bond is not broken and the metal changes shape without shattering.

High melting and boiling points

Most metals are solids with high melting and boiling points. The giant lattice is held by strong electrostatic attraction between the cations and the electron sea, so a large amount of energy is needed to overcome these attractions and separate the particles. The stronger the metallic bonding (section 2), the higher the melting point.

Lustre (shininess)

Freshly cut or polished metals are shiny. The delocalised electrons absorb and re-emit light over a wide range of frequencies, reflecting it from the smooth surface — giving metals their characteristic metallic lustre.

Property	Explanation in one line
Electrical conductivity (solid & liquid)	Delocalised electrons are free to move and carry charge.
Thermal conductivity	Mobile electrons transfer kinetic energy quickly through the lattice.
Malleable and ductile	Layers of cations slip while the electron sea keeps them bonded.
High melting / boiling point	Strong electrostatic attraction in a giant lattice needs much energy to break.
Lustrous	Delocalised electrons reflect light of many frequencies.

4. Alloys

An **alloy** is a mixture of a metal with one or more other elements (metals or non-metals) blended together, usually by melting and mixing. Alloys are not compounds: the components are not chemically bonded in fixed ratios, so an alloy retains metallic bonding and metallic properties.

Why alloys are usually harder and stronger

In a pure metal all the atoms are the same size, so the layers of cations are regular and can slide over one another easily — this is why pure metals are relatively soft. In an alloy the added atoms are a different size. They disrupt the regular arrangement of the layers, so the layers can no longer slip past one another so easily. This makes the alloy harder, stronger and less malleable than the pure metal, while it still conducts electricity because the electron sea is preserved.

Two ways the atoms fit in (qualitative)

- **Substitutional alloy:** atoms of similar size replace some of the host metal atoms at lattice points (for example the copper and zinc atoms in brass).
- **Interstitial alloy:** much smaller atoms squeeze into the gaps (interstices) between the larger host atoms (for example small carbon atoms in the iron lattice of steel).

Both types distort the lattice and hinder the sliding of layers, which is why both kinds of alloy are harder than the parent metal.

Alloy	Main components	Type	Typical use
Steel	Iron + carbon	Interstitial	Construction, tools, car bodies
Stainless steel	Iron + chromium + nickel	Substitutional / interstitial	Cutlery, corrosion-resistant parts

Alloy	Main components	Type	Typical use
Brass	Copper + zinc	Substitutional	Instruments, fittings, decoration
Bronze	Copper + tin	Substitutional	Bearings, statues, medals

> **Exam tip:** A complete alloy answer names the disruption of the regular layers by different-sized atoms, and states that this stops the layers sliding — so the alloy is harder and stronger than the pure metal.

5. Transition metals and metallic bonding (HL)

The transition elements (the d block, such as Fe, Cu, Cr, Ni, W) show unusually strong metallic bonding. In the main-group metals only the outer s (and p) electrons are delocalised. In the transition metals the 3d electrons lie close in energy to the 4s electrons, so **d electrons can also be delocalised** into the electron sea.

Consequences of d-electron involvement

- More electrons are contributed per atom to the delocalised sea, giving a denser sea and greater charge on each cation.
- The electrostatic attraction between cations and electrons is therefore stronger, so metallic bonding is stronger than in the s-block metals.
- This gives transition metals their characteristically **high melting and boiling points** (for example iron melts at 1538 °C and tungsten at 3422 °C, the highest of any metal).
- Strong bonding pulls the cations close together, and the atoms have relatively high atomic masses packed into a small volume, so transition metals also have **high densities** (for example osmium and iridium are the densest elements).

Trend across the d block

Across a transition series the number of electrons available for delocalisation increases at first, so metallic bonding strengthens and melting points rise toward the middle of the series (around Cr, Mo, W). Toward the right of the block the d electrons increasingly pair up and are held more tightly by the nucleus, so fewer contribute to bonding; melting points fall again (zinc, with a full 3d subshell contributing little to the sea, melts at only 420 °C).

> **HL link:** Relate this back to Structure 1.3: it is the closeness in energy of the 3d and 4s subshells that allows d electrons to join the delocalised sea and makes transition-metal bonding so strong.

6. Comparing the three bonding models

Structure 2 asks you to compare ionic, covalent and metallic bonding. This synthesis table draws the three models together — a favourite source of exam questions.

Feature	Ionic	Covalent	Metallic
Particles present	Cations and anions	Atoms sharing electron pairs	Cations + delocalised electrons
Nature of attraction	Electrostatic, between oppositely charged ions	Shared pair of electrons between two nuclei	Electrostatic, between cations and electron sea
Formed between	Metal + non-metal	Non-metal + non-metal	Metal atoms (one element or an alloy)
Electrical conductivity	Only when molten or aqueous (ions free)	Usually none (no free charges)	Conducts as solid and liquid (free electrons)
Melting point	High (giant lattice)	Low (molecular) or very high (network)	Usually high (giant lattice)
Effect of force	Brittle — layers repel and shatter	Molecular solids soft; networks hard	Malleable and ductile — layers slip

> **Why the difference in force matters:** In ionic solids, forcing like charges next to like charges causes repulsion and the crystal shatters. In metals the electron sea flows with the moving cations, so no repulsion builds up and the metal bends instead of breaking — the reason metals are malleable but ionic solids are brittle.

7. Worked examples

Worked example 1 — explaining conductivity

Explain why solid magnesium conducts electricity, whereas solid magnesium chloride does not.

Magnesium is a metal: it has a lattice of Mg^{2+} cations in a sea of delocalised electrons. These electrons are free to move, so when a voltage is applied they drift through the solid and carry charge — magnesium conducts.

Magnesium chloride is ionic: its Mg^{2+} and Cl^{-} ions are locked in fixed positions in the solid lattice, so there are no mobile charge carriers and it does not conduct. (It would conduct only when molten or dissolved, when the ions become free to move.)

Worked example 2 — comparing melting points Na, Mg, Al

Explain the trend in melting points: Na (98 °C) < Mg (650 °C) < Al (660 °C).

Across Period 3 the number of valence electrons donated to the sea rises from 1 (Na) to 2 (Mg) to 3 (Al), so the electron sea becomes denser.

At the same time the cation charge increases (Na^+ → Mg^{2+} → Al^{3+}) and the ionic radius decreases, so charge density rises.

Both changes increase the electrostatic attraction between the cations and the delocalised electrons, so the metallic bonding gets stronger and more energy is needed to melt the lattice. Hence melting point increases Na → Mg → Al.

Worked example 3 — why alloys are harder

Pure copper is soft, but bronze (copper mixed with tin) is much harder. Explain, using the metallic model.

In pure copper all the atoms are the same size, so the layers of cations are regular and slide over one another easily when a force is applied — copper is soft and malleable.

In bronze the tin atoms are a different size from the copper atoms. They disrupt the regular layers so that the layers can no longer slip past each other easily. This makes bronze harder and stronger, while it still conducts electricity because the delocalised electron sea is unchanged.

Worked example 4 — melting points down Group 1

Explain why the melting points of the Group 1 metals decrease from lithium to caesium.

Every Group 1 atom donates just one valence electron, so the number of delocalised electrons per atom is the same throughout the group — charge cannot be the deciding factor.

Down the group the cations get larger (increasing ionic radius) and the valence electrons are more shielded from the nucleus, so charge density falls. The electron sea is held less tightly, metallic bonding weakens, and less energy is needed to melt the metal — so melting point decreases down the group.

Worked example 5 (HL) — transition vs s-block

Suggest why iron (1538 °C) has a much higher melting point than potassium (63 °C).

Potassium is an s-block metal that delocalises only its single 4s electron, giving a thin electron sea and a large, low-charge K^+ cation — weak bonding.

In iron the 3d and 4s subshells are close in energy, so d electrons as well as s electrons are delocalised. Iron therefore contributes several electrons per atom to a dense electron sea, and forms small, highly charged cations. The electrostatic attraction is much stronger, so iron has far stronger metallic bonding and a much higher melting point.

Common pitfalls

- **Electron ‘sea’ vs ions:** the metallic model has cations in a sea of *mobile electrons* — not free-moving positive and negative ions. Free-moving ions belong to the *molten ionic* model.
- Metals conduct as a *solid* because it is the electrons (not the ions) that move; ionic solids do not conduct because their ions are fixed, and conduct only when molten or aqueous.
- Do not call a metal cation an ‘atom’: it has lost its valence electrons to the sea, so it is positively charged.
- Down a group, argue from ionic radius / charge density, not from charge (the charge is constant).
- An alloy is a *mixture*, not a compound; it keeps metallic bonding and still conducts electricity.
- When explaining malleability, say the layers *slip* and the bond is *not broken* — do not say the metal ‘melts’ or the bonds break and reform.

Quick reference

Idea	One-line summary
Metallic bond	Electrostatic attraction between the lattice of cations and the sea of delocalised electrons.
Stronger when	More valence electrons delocalised, and higher cation charge density (higher charge, smaller radius).
Across a period	More electrons + higher charge density → stronger bond → higher melting point.
Down a group	Larger cations → lower charge density → weaker bond → lower melting point.
Conductivity	Delocalised electrons move → conducts as solid and liquid.
Malleable/ductile	Layers of cations slip; electron sea keeps them bonded.
Alloys	Different-sized atoms disrupt layers → harder and stronger than pure metal.
Transition metals (HL)	d electrons also delocalised → stronger bonding, higher m.p. and density.

8. Test yourself

Attempt all eight without notes, then check against the full worked answers below.

1. Describe, in terms of particles and forces, what is meant by metallic bonding.

2. Explain why metals are good conductors of electricity in the solid state.

3. Explain, in terms of the metallic model, why metals are malleable while ionic solids are brittle.

4. Aluminium has a higher melting point than sodium. Explain this using two factors.

5. State what an alloy is and explain why steel is harder than pure iron.

6. Explain why the melting points of the Group 1 metals fall from lithium to caesium.

7. (HL) Explain why transition metals generally have higher melting points and densities than Group 1 metals.

8. Distinguish between a substitutional and an interstitial alloy, giving one example of each.

Answers

1. Metallic bonding is the electrostatic attraction between a regular lattice of positive metal ions (cations) and a ‘sea’ of delocalised electrons that are free to move throughout the whole structure. The attraction acts in all directions and holds the giant lattice together.

2. A metal contains delocalised (free) electrons. When a potential difference is applied these electrons drift through the lattice and carry charge, so the solid conducts. The cations remain fixed, so the metal is not chemically changed by conducting.

3. In a metal, applying a force makes layers of cations slide over one another; the mobile electron sea flows with them, so the metallic bonding is maintained and the metal changes shape (malleable). In an ionic solid, sliding a layer brings like-charged ions next to each other; the repulsion splits the crystal, so it is brittle.

4. Aluminium donates 3 valence electrons per atom to the electron sea while sodium donates only 1, so aluminium has a denser electron sea. Aluminium also forms a smaller, more highly charged Al^{3+} cation (higher charge density) than Na^{+}. Both give a stronger electrostatic attraction, so aluminium has stronger metallic bonding and a higher melting point.

5. An alloy is a mixture of a metal with one or more other elements. In steel, small carbon atoms sit between the iron cations and disrupt the regular layers of the lattice, so the layers can no longer slide over each other easily. This makes steel harder and stronger than pure iron.

6. Each Group 1 atom delocalises one electron, so the number of delocalised electrons per atom is constant. Down the group the cations become larger, so charge density decreases and the electron sea is held less tightly. Metallic bonding weakens and less energy is needed to melt the metal, so melting point falls Li → Cs.

7. (HL) In transition metals the 3d and 4s electrons are close in energy, so d electrons are also delocalised. This gives a denser electron sea and more highly charged cations, so the electrostatic attraction — and hence the metallic bonding — is much stronger than in Group 1 metals (which delocalise only one electron). This stronger bonding, together with heavy atoms packed closely, gives higher melting points and higher densities.

8. In a *substitutional* alloy, atoms of similar size to the host replace some of its atoms at lattice points (e.g. zinc in copper, giving brass). In an *interstitial* alloy, much smaller atoms fit into the gaps between the host atoms (e.g. carbon in iron, giving steel). Both distort the lattice and stop layers sliding, making the alloy harder.