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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Structure 2.3 The Metallic Model

Understandings 2.3.1 – 2.3.3

The metallic bond, properties of metals, bond strength and transition elements

Complete revision notes & memory points · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Core material. Understandings 2.3.1 and 2.3.2 are common to both levels. Every student must know this.

HL Additional higher level material. Understanding **2.3.3 (transition elements)** is HL only. SL students can skip section 4 and the questions marked HL.

Inside this booklet

1. What the syllabus actually asks of you
2. 2.3.1 — The metallic bond and the properties of metals
3. 2.3.2 — The strength of the metallic bond
4. HL 2.3.3 — Transition elements
5. Comparing the bonding models
6. Key points to memorise
7. Check yourself
8. Answers

1 What the syllabus actually asks of you

Structure 2.3 is built around one guiding question: *what determines the metallic nature and properties of an element?* Metals behave so differently from ionic and covalent substances — they conduct as solids, bend instead of shattering, and shine — that they need a bonding model of their own. This booklet sets out that model and uses it to explain the physical properties of metals, the trends in their melting points, and (at HL) the special behaviour of the transition elements.

Understanding	What it says	What you must be able to do	Level
2.3.1 The metallic bond	A metallic bond is the electrostatic attraction between a lattice of cations and delocalized electrons.	Explain the electrical conductivity, thermal conductivity and malleability of metals. Relate characteristic properties of metals to their uses.	SL+HL
2.3.2 Strength of the metallic bond	The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion.	Explain trends in the melting points of s- and p-block metals. A simple treatment in terms of the charge of the cations and electron density is required.	SL+HL
2.3.3 Transition elements	Transition elements have delocalized d-electrons.	Explain the high melting point and electrical conductivity of transition elements. (Their chemical properties are covered in Structure 3.1 and Reactivity 3.4.)	HL

Linking questions and Nature of Science attached to Structure 2.3

- **Tool 1, Inquiry 2, Structure 3.1:** what experimental data demonstrate the physical properties of metals, and trends in these properties in the periodic table? → sections 2.4 and 3
- **Reactivity 3.2:** what trends in reactivity of metals can be predicted from the periodic table? → section 3.4
- **Structure 2.4:** what are the features of metallic bonding that make it possible for metals to form alloys? → section 2.6
- HL **Structure 3.1:** why is the trend in melting points of metals across a period less evident across the d-block? → section 4.3
- **Nature of Science:** atomic theory explains the distinct properties of metals; and the use of metals raises ethical questions about extraction and processing → sections 2.6 and 2.7

What the data booklet gives you — and what it does not

- **Given:** the periodic table, from which you can read each metal's group (and so the charge on its ion) and electron configuration; ionic radii, which you need to compare metallic bond strength.
- **Not given:** the definition of the metallic bond, the explanations of metallic properties, or the factors that control bond strength. Melting points will be supplied in a question when you need them; you are expected to explain trends, not recall values.

2 Structure 2.3.1 — The metallic bond SL + HL

2.1 Where the delocalized electrons come from

Metals lie on the left of the periodic table and in the d-block. Their atoms have only a few electrons in their outer shell and **low ionization energies**, so these valence electrons are held only loosely by the nucleus. In a compound, a metal loses its valence electrons to a non-metal and forms positive ions (Structure 2.1). In the pure element there is no non-metal to accept them, so the

valence electrons simply leave their own atoms and spread through the whole structure. Electrons that are not fixed to one atom, but are shared across many, are **delocalized** — the same idea you met for graphite and benzene in Structure 2.2.

What remains of each atom is a positive ion, a **cation**, and the cations pack together in a regular, three-dimensional **lattice**. The delocalized electrons move freely through this lattice. The tendency of an element to lose its valence electrons in this way is what chemists mean by its **metallic character**.

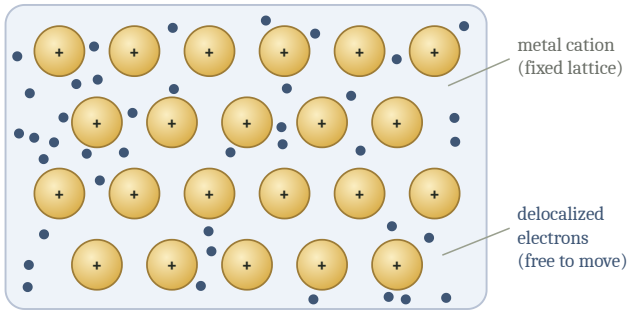


Figure 1 — The metallic model: a lattice of cations in a "sea" of delocalized electrons.

Definition to learn word-for-word

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

The attraction acts between each cation and the electrons all around it, so it is **non-directional**: it does not depend on which ion sits next to which. This one feature explains most metallic properties.

Wording that loses marks

- "Attraction between positive ions and negative ions." There are no anions in a metal.
- "Attraction between nuclei and electrons." The lattice is made of **cations** — the nuclei together with their inner electrons.
- "Free electrons" on their own, without saying they are **delocalized** and that it is the **attraction** between them and the cations that is the bond.

2.2 Explaining the properties of metals

Property	Explanation from the metallic model	Typical uses
Good electrical conductor	The delocalized electrons are mobile charged particles. When a potential difference is applied, they move through the lattice towards the positive terminal, carrying charge. Metals conduct both as solids and when molten.	copper wiring and cables; aluminium overhead power lines; gold-plated electrical contacts
Good thermal conductor	The mobile delocalized electrons carry kinetic energy quickly through the structure, and the closely packed cations pass vibrations efficiently to their neighbours.	saucepans and cooking pans; car radiators; heat sinks in computers
Malleable (can be hammered into shape) and **ductile** (can be drawn into wires)	When a force is applied, layers of cations can slide over one another. Because the metallic bond is non-directional, the delocalized electrons still surround and hold the cations in their new positions, so the bond is not broken (see Figure 2).	car body panels, cans and foil; wires and cables
High melting and boiling points (most metals)	Strong electrostatic attraction between the cations and the delocalized electrons throughout the lattice; a lot of energy is needed to overcome it.	tungsten lamp filaments; steel and titanium in engines and turbines
Shiny, lustrous	The delocalized electrons absorb and re-emit (reflect) visible light.	jewellery; mirrors; decorative fittings
Insoluble in water and organic solvents	No solvent interaction can replace the strong metallic bonding. (Reactive metals such as sodium do not dissolve in water; they *react* with it.)	water pipes (copper); containers

High density is common but not universal: most transition metals are dense, but the group 1 metals lithium, sodium and potassium float on water.

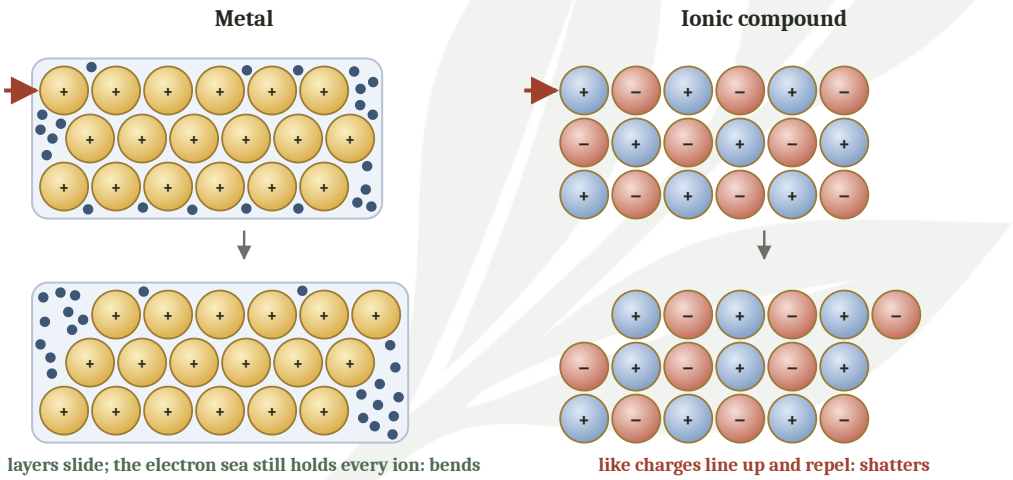


Figure 2 — Why metals bend but ionic solids shatter. In a metal the bond is non-directional. In an ionic lattice, shifting one layer brings ions of the same charge side by side, and the repulsion splits the crystal (Structure 2.1).

Model answer: "Explain why metals are malleable." (2 marks)

When a force is applied, layers of cations slide past one another **[1]**. The metallic bonding is non-directional, so the delocalized electrons continue to attract the cations in their new positions and the metallic bond is not broken **[1]**.

2.3 Relating properties to uses

Choosing a metal for a job means matching its properties to the need, and usually balancing several properties against cost. Aluminium is a good example: it is a good conductor of heat and electricity, highly malleable and ductile, and has a **low density**. Its surface also forms a thin, protective oxide layer that resists corrosion. That combination explains its use in cooking foil (malleable, good thermal conductor, non-toxic), aircraft bodywork (low density, strong as an alloy, corrosion resistant), cooking pans (thermal conductivity) and tent frames (low density and strength). Copper is the usual choice for electrical wiring because it combines very high conductivity with ductility at a reasonable cost.

2.4 What experimental data show

This is the Tool 1 / Inquiry 2 linking question. In the laboratory, the metallic properties can be demonstrated directly:

- **electrical conductivity**, by placing a strip or wire of the metal in a simple circuit with a bulb or ammeter (it conducts as a solid, unlike an ionic compound);
- **thermal conductivity**, by heating one end of rods of equal size made of different metals and timing how long wax at the other end takes to melt;
- **malleability and ductility**, by hammering or bending samples;
- **melting points**, which are compared from data tables and show clear trends down groups and across periods (section 3).

2.5 Metals compared with other conductors

Metals are the only elements that conduct well as solids *and* as liquids, because the delocalized electrons are present in both states. Ionic compounds conduct only when molten or dissolved, when their ions are free to move. Of the non-metals, only graphite and graphene conduct well (Structure 2.2), again because of delocalized electrons, but only within their layers.

2.6 Alloys

Because the metallic bond is non-directional, other elements can be mixed into a molten metal and stay held in the lattice as it solidifies. The product is an **alloy**, a mixture of a metal with other metals or with non-metals. Steel is iron with a small amount of carbon; bronze, which gave the Bronze Age its name, is copper with tin. The different-sized atoms disrupt the regular layers, so the layers can no longer slide easily: alloys are usually **harder and stronger** than the pure metal. Alloys are studied in Structure 2.4.

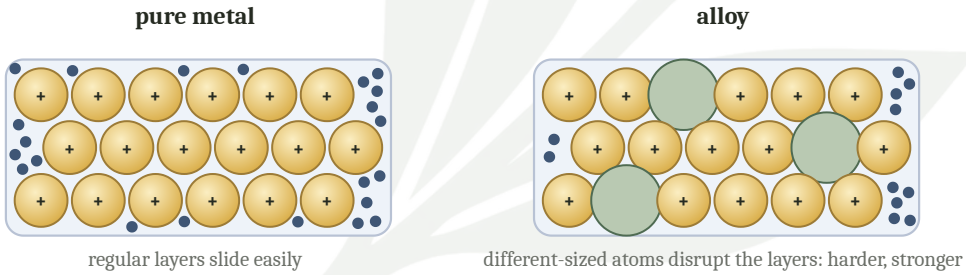


Figure — A pure metal compared with an alloy.

> **Nature of Science — theory that explains and predicts**
>
> The metallic model is a good example of atomic theory explaining observed properties: one simple picture accounts for conductivity, malleability, lustre and melting points. The same model guides the design of new materials, such as alloys that combine the useful features of several metals for aircraft, medical implants and electronics.

2.7 Metals, society and the environment

Nature of Science and the global impact of science

As technology develops, the uses of metals keep growing — from mobile phones and solar panels to the gold-coated beryllium mirrors of the James Webb Space Telescope. Extracting metals from their ores has real costs: open-cast mining destroys landscapes, waste rock and "tailings" accumulate in spoil heaps, and toxic materials can leach out and pollute water. Scientists have a responsibility to consider these social, political, economic and environmental consequences. Reusing and recycling metals, and avoiding waste, reduce the impact. Despite their apparent abundance, metals are a precious resource.

3 Structure 2.3.2 — The strength of the metallic bond SL + HL

3.1 What makes a metallic bond stronger?

The metallic bond is an electrostatic attraction, so its strength depends on how much charge is involved and how close together the charges are. Three linked factors decide it:

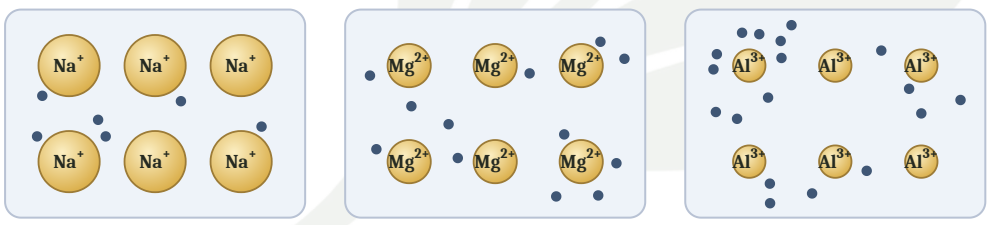
- **the number of delocalized electrons per atom** — more electrons give a greater density of negative charge;
- **the charge on the cation** — a higher charge (equal to the number of electrons each atom has given up) attracts the electrons more strongly;
- **the radius of the cation** — a smaller ion lets the delocalized electrons come closer to its nucleus, so the attraction is stronger.

The rule in one sentence

The **greater the charge** on the cations (and so the greater the density of delocalized electrons) and the **smaller the radius** of the cations, the **stronger** the metallic bond and the **higher** the melting point.

Melting point is the usual experimental measure of metallic bond strength, because more energy is needed to disrupt a more strongly bonded lattice.

3.2 Across a period: Na, Mg and Al



Na^+: radius 102 pm
1 delocalized e^- per atom

Mg^{2+}: radius 72 pm
2 delocalized e^- per atom

Al^{3+}: radius 54 pm
3 delocalized e^- per atom

Figure 3 — From sodium to aluminium the cation charge rises, the number of delocalized electrons per atom rises, and the cation radius falls.

Metal	Electron configuration	Cation	Delocalized e⁻ per atom	Ionic radius / pm	Melting point / °C	Boiling point / °C
sodium, Na	[Ne] $3s^1$	Na^+	1	102	98	≈ 880
magnesium, Mg	[Ne] $3s^2$	Mg^{2+}	2	72	650	≈ 1090
aluminium, Al	[Ne] $3s^2$ $3p^1$	Al^{3+}	3	54	660	≈ 2520

From Na to Mg to Al, each atom gives up one more electron, so the cation charge and the density of delocalized electrons increase, while the cation radius decreases. The attraction between the cations and the delocalized electrons gets stronger, so the metallic bond is stronger. Magnesium's melting point (650 °C) is far above sodium's (98 °C).

A detail worth knowing. Aluminium melts at only 660 °C, barely higher than magnesium, because melting disrupts the lattice without separating the atoms completely. Boiling points, where the metallic bonding is fully overcome, show the trend much more clearly: Al ≈ 2520 °C, far above Mg. Metals also become denser across the period (Na 0.97, Mg 1.74, Al 2.70 g cm^{-3}), because the smaller, more highly charged ions are pulled closer together.

3.3 Down a group

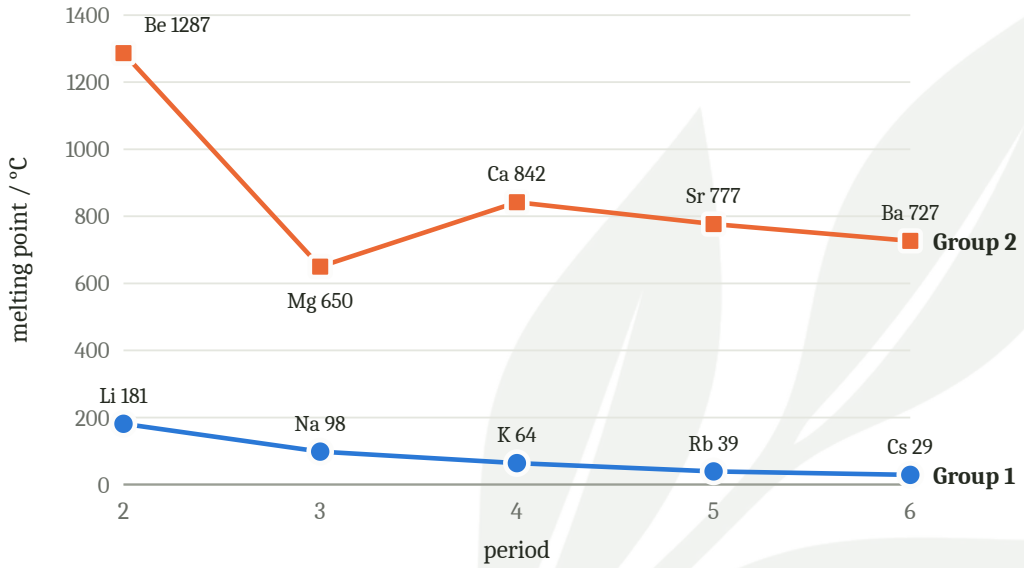


Figure 4 — Melting points of the group 1 and group 2 metals.

Group 1. Every group 1 metal forms a 1+ ion and contributes one delocalized electron per atom, so the charge does not change down the group. The ionic radius does increase (Li^+ 76, Na^+ 102, K^+ 138, Rb^+ 152, Cs^+ 167 pm), because each element has one more occupied shell. The delocalized electrons are further from the nucleus, the attraction is weaker, and the melting point falls steadily from lithium (181 °C) to caesium (29 °C).

Group 2. Every group 2 metal forms a 2+ ion with two delocalized electrons per atom, so its metallic bonding is stronger than that of the group 1 metal in the same period: compare lithium (181 °C) with beryllium (1287 °C). Down group 2 the cations get larger and the general trend is again a decrease in melting point. Magnesium is lower than the trend suggests, because it has a different crystal structure from calcium and strontium; this detail is beyond the course.

3.4 Melting point and reactivity

This is the Reactivity 3.2 linking question. When a metal reacts it loses its valence electrons, so the same factors that weaken the metallic bond — a lower charge and a larger radius, so a weaker hold on the outer electrons — also make the metal lose electrons more easily and react more vigorously. The two trends therefore run in parallel:

	Melting point	Attraction between cations and delocalized electrons	Reactivity
Left to right across a period	increases	stronger	decreases
Down a group	decreases	weaker	increases

Caesium, with the weakest metallic bonding in group 1, is the most reactive; lithium the least. Reactivity is explained more fully in terms of ionization energy in Structure 3.1 and Reactivity 3.2.

Worked example — predict and explain the order of melting points of potassium, calcium and sodium

Charge: K and Na form 1+ ions with one delocalized electron per atom; Ca forms a 2+ ion with two. **Radius:** K^+ (138 pm) is larger than Na^+ (102 pm). Ca^{2+} (100 pm) is similar in size to Na^+ but has twice the charge.

Order: K < Na < Ca. Calcium has the strongest metallic bonding, because of its higher charge and greater density of delocalized electrons; potassium has the weakest, because its large cation holds the delocalized electrons least strongly. (Data: K 64 °C, Na 98 °C, Ca 842 °C.)

Model answer: "Explain why magnesium has a higher melting point than sodium." (3 marks)

Magnesium forms Mg^{2+} ions and has two delocalized electrons per atom, whereas sodium forms Na^+ ions with one **[1]**. Mg^{2+} is also smaller than Na^+ **[1]**. So the electrostatic attraction between the cations and the delocalized electrons is stronger in magnesium, and more energy is needed to overcome it **[1]**.

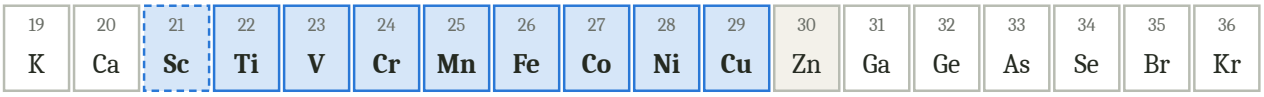
Examiner traps in 2.3.1 – 2.3.2

- Explaining conductivity with "free ions" or "mobile ions". In a solid metal it is the **delocalized electrons** that move.
- Saying malleability happens because "the bonds are weak". It happens because the bonding is **non-directional** and is not broken when layers slide.
- Explaining a melting point difference with "more electrons" alone. Mention the cation **charge** or electron density **and** the cation **radius**, and link them to the strength of the electrostatic attraction.
- Using **atomic** radius when the guide asks about the **cation** (ionic) radius. The trends are the same, but use the right term.
- Saying melting "breaks the metallic bonds completely". Boiling does that; melting only disrupts the lattice.

4 Structure 2.3.3 — Transition elements HL

4.1 What counts as a transition element?

> **Definition to learn**
>
> A **transition element** is an element whose atoms have an **incomplete d sublevel**, or which can give rise to **cations with an incomplete d sublevel**.



transition elements of period 4 (shaded): Sc–Cu

In period 4 the d-block runs from scandium to zinc, but only scandium to copper are transition elements:

- **Zinc is not** a transition element. Its atom is [Ar] $3d^{10}$ $4s^2$ and its only ion, Zn^{2+}, is [Ar] $3d^{10}$. Both have a complete d sublevel.
- **Copper is** a transition element, even though its atom is [Ar] $3d^{10}$ $4s^1$, because Cu^{2+} is [Ar] $3d^9$, with an incomplete d sublevel.
- **Scandium** is debated. Its atom has an incomplete d sublevel ([Ar] $3d^1$ $4s^2$), but its only common ion, Sc^{3+}, has no d electrons at all. Structure 3.1 discusses this.

4.2 Delocalized d electrons

Transition elements bond by the same metallic model as other metals, with one important difference. In a transition element, the outer 4s and the 3d sublevels are very close in energy, so electrons from **both** the 4s and the 3d sublevel can become delocalized. Transition elements therefore contribute **more delocalized electrons** per atom than the s-block metals, and their cations are also **relatively small**, because the d electrons are added to an inner shell and shield the nucleus poorly.

Property	Explanation	Examples
High melting points	The large number of delocalized electrons (from 4s and 3d) and the small, highly charged cations give a very strong electrostatic attraction, so a great deal of energy is needed to disrupt the lattice.	Iron 1538 °C, titanium 1668 °C, compared with potassium 64 °C and calcium 842 °C in the same period. Tungsten (3422 °C) has the highest melting point of any metal and is used for lamp filaments and furnace parts.
High electrical conductivity	The large number of delocalized electrons gives many mobile charge carriers.	Silver is the best electrical conductor of all, followed by copper and gold. Copper is the most common metal in electrical cables; gold is used for contacts and connectors on circuit boards.

4.3 Why the melting point trend across the d-block is less clear

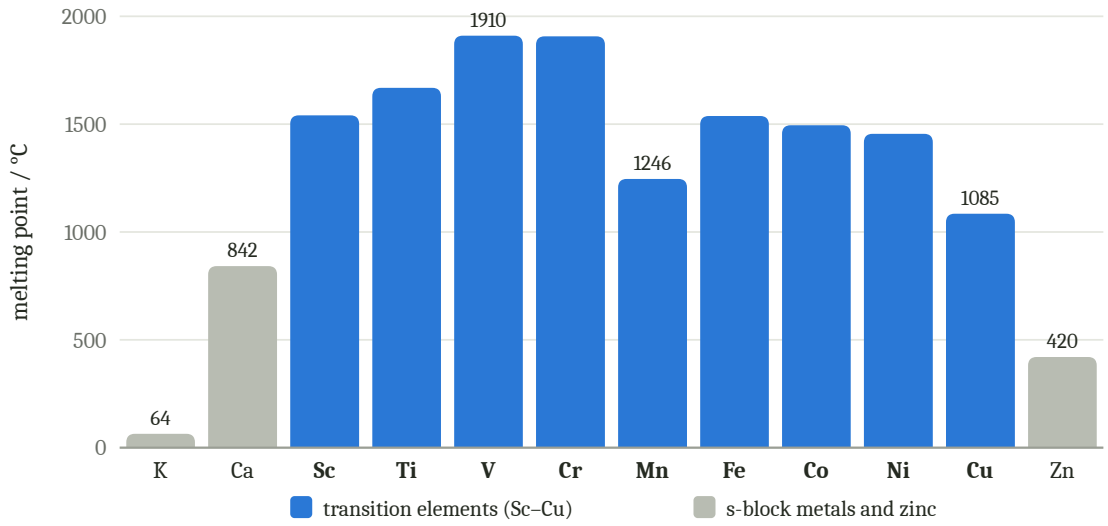


Figure 5 — Melting points of the period 4 metals. Values in °C.

This is the Structure 3.1 linking question. Across the s- and p-block metals of a period the melting point rises steadily, because each step adds one more electron to the outer shell and increases the cation charge (section 3.2). Across the transition elements the pattern is much flatter and less regular: melting points rise to a maximum around vanadium and chromium, dip sharply at manganese, then fall gradually towards copper and zinc. Two reasons:

- The extra electrons go into the **inner 3d sublevel**, so the radius of the atoms and ions changes only slightly across the series. One of the two factors that controls bond strength hardly varies.
- All the transition elements can delocalize a **large** but varying number of electrons from both the 4s and 3d sublevels, so the number of delocalized electrons does not simply rise by one at each step as it does for Na, Mg and Al.

The result is that the trend is harder to predict than for the s-block elements. Zinc, with a full 3d sublevel that does not take part in bonding, has a much lower melting point (420 °C) than the transition elements.

Model answer: "Explain why transition elements generally have higher melting points than s-block metals." (2 marks)

Transition elements can delocalize electrons from both the 4s and the 3d sublevels, because these are close in energy, so they have more delocalized electrons per atom **[1]**. Together with their relatively small cations, this gives a stronger electrostatic attraction between the cations and the delocalized electrons, which needs more energy to overcome **[1]**.

Challenge — silver is the best conductor, so why is copper used for wiring?

Silver is far more expensive and much less abundant than copper, and the improvement in conductivity is small. It is used only where small amounts make a real difference, such as in some contacts and components inside electronic devices. This is a typical example of choosing a material by balancing properties, cost and availability.

Examiner traps in 2.3.3

- Calling zinc a transition element because it is in the d-block. Neither Zn nor Zn^{2+} has an incomplete d sublevel.
- Explaining the high melting points with "more electrons" without saying that the **d electrons** are **delocalized**.
- Expecting a smooth trend in melting points across the transition series, as across Na → Mg → Al.

5 Comparing the bonding models SL + HL

Structure 2.1 to 2.3 have now introduced the three bonding models. This table summarises how each one explains the behaviour of a typical substance. Structure 2.4 then shows that most real substances lie somewhere between these idealised types.

	Metallic	Ionic	Covalent network	Simple molecular
Particles	cations and delocalized electrons	cations and anions	atoms	molecules
Bonding	attraction between a lattice of cations and delocalized electrons	attraction between oppositely charged ions	covalent bonds throughout	covalent bonds within molecules; intermolecular forces between them
Directional?	no	no	yes	yes (within molecules)
Melting point	usually high; varies widely	high	very high	low
Conducts when solid?	yes	no	no (except graphite, graphene)	no
Conducts when molten?	yes	yes	no	no
Mechanical behaviour	malleable, ductile	hard, brittle	very hard (diamond, SiO_2)	soft
Examples	Na, Cu, Fe, Al	NaCl, MgO	diamond, SiO_2	I_2, H_2O, CO_2

6 Key points to memorise

Everything below is either wording that examiners look for, or knowledge you must carry into the exam yourself.

A Definitions — near-verbatim

- **Metallic bond:** the electrostatic attraction between a lattice of cations and delocalized electrons.
- **Delocalized electrons:** electrons not fixed to one atom but free to move through the whole structure.
- **Malleable:** can be hammered or pressed into shape. **Ductile:** can be drawn into wires.
- **Alloy:** a mixture of a metal with other metals or non-metals.
- HL **Transition element:** an element whose atoms have an incomplete d sublevel, or which can give rise to cations with an incomplete d sublevel.

B Properties in one line each

- **Electrical conductivity:** delocalized electrons are mobile and carry charge.
- **Thermal conductivity:** mobile delocalized electrons and closely packed cations transfer energy efficiently.
- **Malleability and ductility:** layers of cations slide; the non-directional bond is not broken.
- **High melting point:** strong electrostatic attraction between cations and delocalized electrons.
- **Lustre:** delocalized electrons reflect light.

C Bond strength and trends

- Stronger metallic bond: **higher cation charge** (more delocalized electrons) and **smaller cation radius**.
- Across a period (Na → Mg → Al): stronger bonding, higher melting point, lower reactivity.
- Down a group: weaker bonding, lower melting point, higher reactivity.
- Group 2 metals have stronger metallic bonding than group 1 metals in the same period.

D HL Transition elements

- Period 4 transition elements: Sc to Cu. Zinc is not one (Zn and Zn^{2+} both have $3d^{10}$).
- 4s and 3d close in energy, so d electrons are delocalized too: more delocalized electrons, high melting points, high conductivity.
- Melting points across the d-block show no simple trend: small changes in radius, varying numbers of delocalized electrons.

E Command terms — what each one is actually asking for

- **Describe** (metallic bonding) — name the particles (lattice of cations, delocalized electrons) and the force (electrostatic attraction).
- **Explain** (a property) — link the property to a specific feature of the model: mobile electrons, non-directional bonding, strength of attraction.
- **Explain** (a melting point difference) — compare cation charge or electron density **and** cation radius for both metals, then link to the strength of attraction and the energy needed.
- **Suggest** (a use) — match one or more properties to the need, and consider cost and other properties.

7 Check yourself

Cover the answers on the next pages. Marks are shown in brackets where a question is written in exam style. Questions marked HL are for Higher Level students.

1. Which is the best description of metallic bonding?

   A the attraction between cations and anions B the attraction between a lattice of cations and delocalized electrons

   C the attraction between nuclei and shared pairs of electrons D the attraction between molecules

2. Describe the metallic model of bonding and explain how it accounts for the electrical conductivity of metals. **[3]**

3. Explain, with reference to the metallic model, why metals are good thermal conductors and why they are ductile. **[4]**

4. Explain why a sheet of copper bends when it is hammered, whereas a crystal of sodium chloride shatters. **[3]**

5. Aluminium is a widely used metal. Suggest which properties make it suitable for: (a) cooking foil (b) aircraft bodywork (c) overhead power cables (d) tent frames

6. Which combination gives the strongest metallic bond?

   A small cation charge, small radius B small cation charge, large radius C large cation charge, small radius D large cation charge, large radius

7. Which metal has the strongest metallic bonding?

   A Na B K C Mg D Al

8. Which element is most likely to be a metal?

	Electrical conductivity when molten	Electrical conductivity when solid
A	good	good
B	good	poor
C	poor	good
D	poor	poor

9. Which series shows the metals in order of decreasing strength of metallic bonding?

   A Cs > Rb > K > Na B Na > K > Rb > Cs C Rb > Na > K > Cs D Na > Cs > K > Rb

10. Lithium melts at 181 °C and beryllium at 1287 °C. Explain the difference. **[3]**

11. Explain why calcium has a higher melting point than strontium. **[2]**

12. Explain why the reactivity of the group 1 metals increases down the group, using the same ideas that explain their melting points. **[2]**

13. Explain why alloys are generally harder than the pure metals they are made from. **[2]**

14. HL Which are transition elements? I. zirconium, Zr II. palladium, Pd III. lead, Pb

    A I and II only B I and III only C II and III only D I, II and III

15. HL Explain why zinc is not classified as a transition element, while copper is. **[2]**

16. HL Explain, with reference to the metallic model, why transition elements have high electrical conductivity. **[2]**

17. HL Suggest why the melting points of the period 4 transition elements do not increase steadily from scandium to copper, as the melting points do from sodium to aluminium. **[2]**

8 Answers

1. **B.**
2. A lattice of cations surrounded by delocalized electrons **[1]**, held together by the electrostatic attraction between them **[1]**. The delocalized electrons are mobile and move through the lattice when a potential difference is applied, carrying charge **[1]**.
3. Thermal: the mobile delocalized electrons carry kinetic energy through the metal **[1]**, and the closely packed cations pass on vibrations efficiently **[1]**. Ductile: when the metal is pulled, layers of cations slide over each other **[1]**; the bonding is non-directional, so the delocalized electrons still hold the cations in their new positions and the metal does not break **[1]**.
4. In copper, layers of cations slide over one another and the non-directional metallic bonding holds them in their new positions **[1]**. In sodium chloride, moving a layer brings ions of the same charge next to each other **[1]**; they repel, and the crystal splits **[1]**.
5. (a) very malleable; good thermal conductor; non-toxic (b) low density; strong (as an alloy); resists corrosion because of its oxide layer (c) good electrical conductor; low density; ductile (d) low density; strong; resists corrosion.
6. **C.**
7. **D.** Al^{3+} has the highest charge (three delocalized electrons per atom) and the smallest radius.
8. **A.** Metals conduct as solids and as liquids. (B describes an ionic compound.)
9. **B.** All form 1+ ions; the radius increases from Na to Cs, so the bonding gets weaker.
10. Beryllium forms Be^{2+} ions with two delocalized electrons per atom; lithium forms Li^{+} with one **[1]**. Be^{2+} is also smaller than Li^{+} **[1]**. So the attraction between the cations and the delocalized electrons is much stronger in beryllium, and more energy is needed to overcome it **[1]**.
11. Both form 2+ ions with two delocalized electrons per atom, but Ca^{2+} is smaller than Sr^{2+} because it has one fewer occupied shell **[1]**. The delocalized electrons are closer to the nucleus, so the electrostatic attraction is stronger in calcium **[1]**.
12. Down the group the atoms and ions get larger, so the valence electron is further from the nucleus and more shielded **[1]**. It is held less strongly, which both weakens the metallic bond (lower melting point) and makes it easier to lose the electron in a reaction (higher reactivity) **[1]**.
13. The atoms of the added element are a different size from those of the metal and disrupt the regular layers of cations **[1]**, so the layers cannot slide over each other as easily **[1]**.
14. **A.** Zirconium and palladium are d-block transition elements; lead is a p-block metal.
15. Zinc atoms ([Ar] $3d^{10}$ $4s^{2}$) and Zn^{2+} ions ([Ar] $3d^{10}$) both have a complete d sublevel **[1]**. Copper forms Cu^{2+} ([Ar] $3d^{9}$), which has an incomplete d sublevel **[1]**.
16. The 3d and 4s sublevels are close in energy, so electrons from both can be delocalized **[1]**. This gives a large number of mobile delocalized electrons to carry charge **[1]**.
17. Across the transition series the electrons are added to the inner 3d sublevel, so the radius changes only slightly **[1]**; and the number of delocalized electrons is large but does not increase by one at each step, so bond strength does not rise steadily **[1]**.

Where this booklet leads

- **Structure 2.4** — from models to materials: the bonding triangle and alloys.
- **Structure 3.1** — periodic trends, ionization energy and (HL) the properties of transition elements.
- **Reactivity 3.2** — trends in the reactivity of metals; HL **Reactivity 3.4** — transition element complexes as Lewis acid–base products.

These notes are a free resource, shared for any IB student who finds them useful.

If a particular idea here still isn't clicking, specialist one-to-one support in Physics, Chemistry and Maths from a past A Level examiner is available at iBLaurel — call or message **+44 7404 710097**, or use the contact form at **www.iBLaurel.com**.

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Syllabus references are to the IB Chemistry guide, first assessment 2025. Melting points and ionic radii are standard literature values.