



HL IB Chemistry


Your notes

The Metallic Model

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Properties of Metals & Their Uses



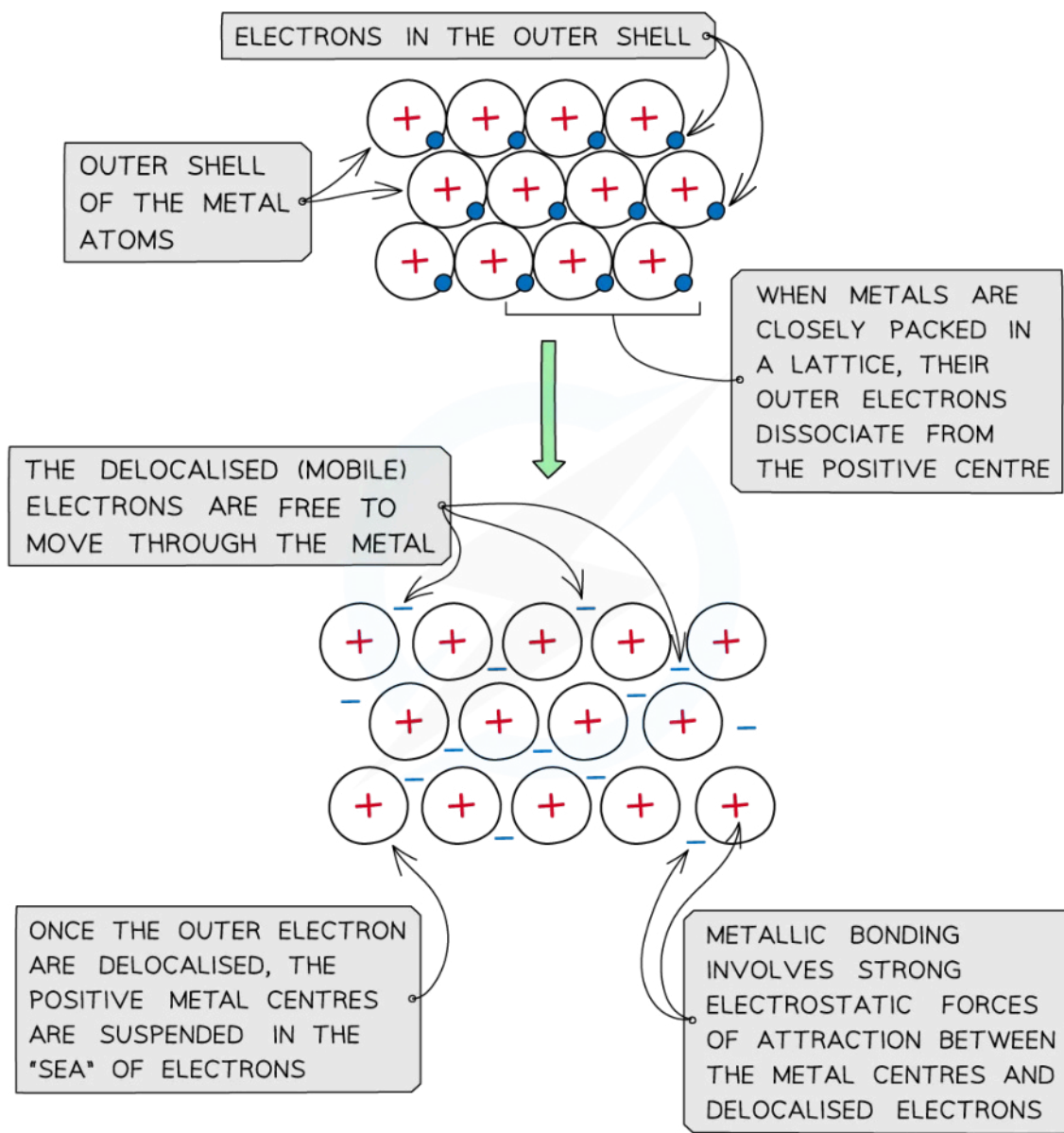
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Properties of Metals & Their Uses

What is metallic bonding?

- **Metal** atoms are tightly packed together in **lattice** structures
- When the metal atoms are in **lattice** structures, the electrons in their outer shells are free to move throughout the structure
- The free-moving **electrons** are called '**delocalised**' electrons and they are not bound to their atom
- When the electrons are **delocalised**, the metal atoms become **positively** charged
- The positive charges **repel** each other and keep the neatly arranged lattice in place
- There are very **strong electrostatic forces** between the positive metal centres and the 'sea' of delocalised electrons

Metallic bonding diagram



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The structure of metallic bonding has positive metal ions suspended in a 'sea' of delocalised electrons

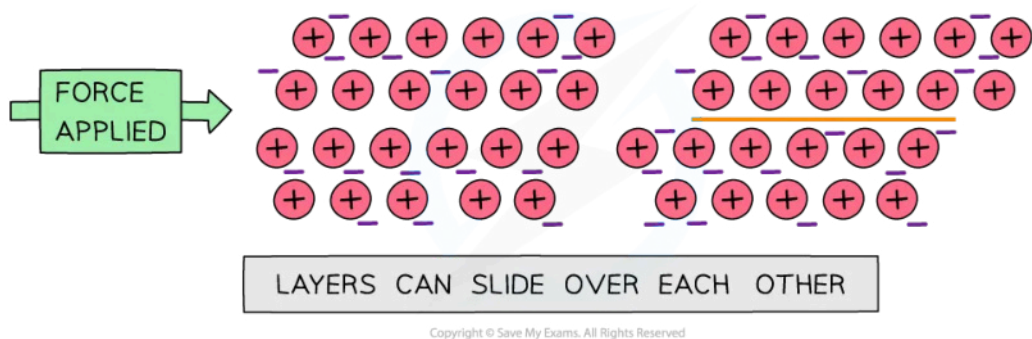
What are the properties of metals?

Malleability

- Metallic compounds are **malleable**
- When a force is applied, the metal layers can slide
- The **attractive forces** between the metal ions and electrons act in all directions
- So when the layers slide, the metallic bonds are re-formed

- The lattice is not broken and has changed shape

How metals are malleable diagram



Atoms are arranged in layers so the layers can slide when force is applied

Strength

- Metallic compounds are **strong** and **hard**
 - Due to the strong attractive forces between the metal ions and delocalised electrons

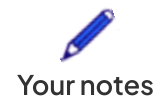
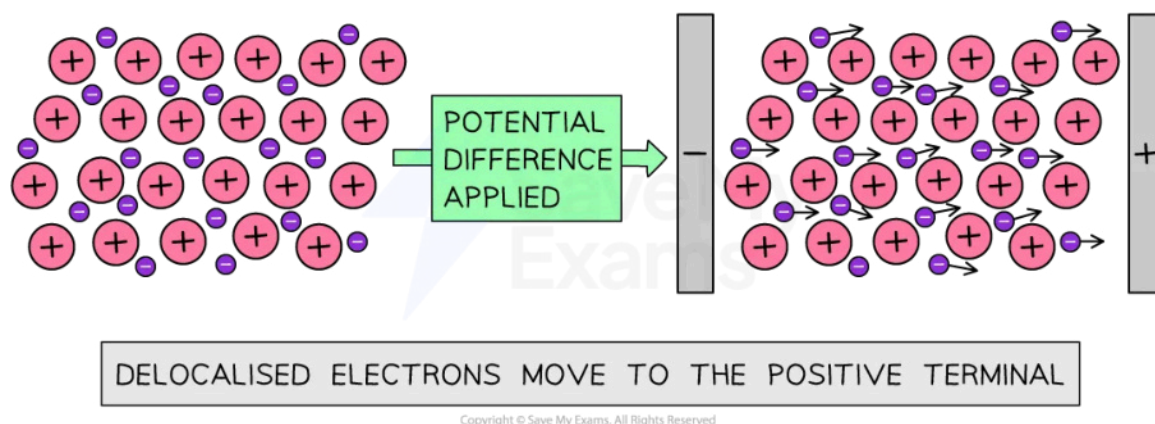
Electrical conductivity

- Metals can **conduct electricity** when in the **solid** or **liquid** state
 - In the solid and liquid states, there are **mobile electrons** which can freely move around and conduct electricity
- When a **potential difference** is applied to a metallic lattice, the delocalised electrons **repel** away from the negative terminal and move towards the positive terminal
 - As the number of outer electrons increases across a Period, the number of **delocalised charges** also increases:
 - Sodium = 1 outer electron
 - Magnesium = 2 outer electrons
 - Aluminium = 3 outer electrons
 - Therefore, the ability to conduct electricity also increases across a period

How metals conduct electricity diagram



Your notes



The delocalised electrons move towards the positive terminal when a potential difference is applied

- Since the bonding in metals is **non-directional**, it does not really matter how the **cations** are oriented relative to each other

Thermal conductivity

- Metals are **good thermal conductors** due to the behaviour of their cations and their delocalised electrons
 - When metals are heated, the cations in the metal lattice **vibrate** more vigorously as their **thermal energy increases**
 - These vibrating cations transfer their **kinetic energy** as they collide with neighbouring cations, effectively conducting heat
 - The delocalised electrons are not bound to any specific atom within the metal lattice and are free to move throughout the material
 - When the cations vibrate, they transfer kinetic energy to the electrons
 - The delocalised electrons then carry this increased kinetic energy and **transfer** it rapidly throughout the metal, contributing to its high thermal conductivity.

Melting and boiling point

- Metals have **high melting** and **boiling points**
 - This is due to the **strong electrostatic forces of attraction** between the cations and delocalised electrons in the metallic lattice
 - These require large amounts of energy to **overcome**
 - As the number of **mobile charges** increases across a Period, the melting and boiling points increase due to stronger electrostatic forces

Uses of metals

- The metal chosen for a particular job can be based on considering the following list of metal properties:
 - Malleability / ductility
 - Electrical conductivity
 - Thermal conductivity
 - Melting / boiling point
 - Strength
 - Strength-to-weight ratio

- Density
 - Toxicity
 - Corrosion resistance
 - Reactivity
 - Lustre
 - Sonority
- For example:
 - Aluminium is used in food cans because it is non-toxic and resistant to corrosion and acidic food stuffs
 - Copper is used in electrical wiring because it is a good electrical conductor and malleable / ductile
 - Stainless steel is used for cutlery as it is strong and resistant to corrosion



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s & p Block Elements

Trends in s & p Block Metals

What determines the strength of metallic bonds?

- Not all metallic bonds are equal
- There are several factors that affect the **strength** of a metallic bond:

The charge on the metal ion

- The **greater the charge** on the metal ion, the greater the number of electrons in the sea of delocalised electrons and the greater the **charge difference** between the ions and the electrons
- A greater charge difference leads to a **stronger** electrostatic attraction, and therefore a stronger metallic bond
- This effect can be seen in melting point data across a period, as the charge on the metal ion **increases** without a significant change in ionic radius:

Melting point data of the Period 3 metals

Group	1	2	3 (13)
Metal	Sodium	Magnesium	Aluminium
Melting point / K	371	923	933

The melting point of the metal increases moving across a period, from left to right

The radius of the metal ion

- Metal ions with **smaller ionic radii** exert a greater attraction on the sea of delocalised electrons
- This greater attraction means a **stronger** metallic bond, requiring more energy to break
- This can be seen in data from metals, descending a group, where the charge on the ion remains constant but the ionic radius increases:

Melting point data of the Group 1 metals

Period	3 (13)	4	5
Metal	Sodium	Potassium	Rubidium
Melting point / K	371	336	312

The melting point of the metal decreases moving down a group



Your notes

Trends in Melting Points of Metals

- The **strength** of electrostatic attraction can be increased by:
 - **Increasing** the number of **delocalised electrons** per metal atom
 - **Increasing** the number of **positive charges** on the metal centres in the lattice
 - **Decreasing** the **size** of the metal ions
- These factors can be seen in the trends across a period and down a group

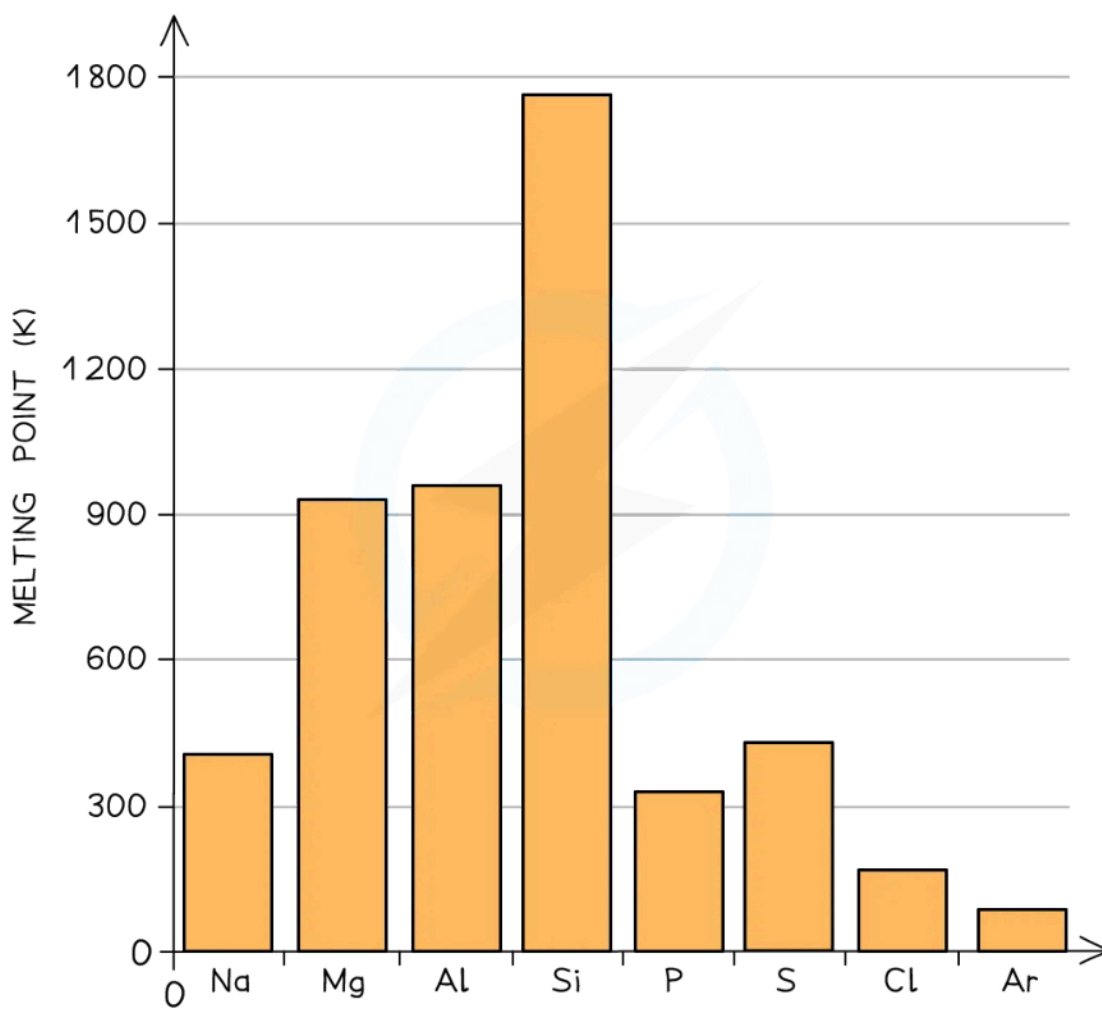
Melting points of metals across a period

- If you compare the electron configuration of sodium, magnesium and aluminium you can see the number of valence electrons increases
 - **Na** = $1s^2 2s^2 2p^6 3s^1$
 - **Mg** = $1s^2 2s^2 2p^6 3s^2$
 - **Al** = $1s^2 2s^2 2p^6 3s^2 3p^1$
- Aluminium ions are also a smaller size than magnesium ions or sodium ions and these two factors lead to **stronger** metallic bonding which can be seen in the melting points
- The **stronger** the metallic bonding, the **more energy** is needed to break the metallic lattice and so the **higher** the melting point
- As we go across Period 3, we can see the effect of stronger metallic bonding on the metals
 - **Remember:** Only the first three elements have metallic bonding in this graph

Melting point of elements across a period chart



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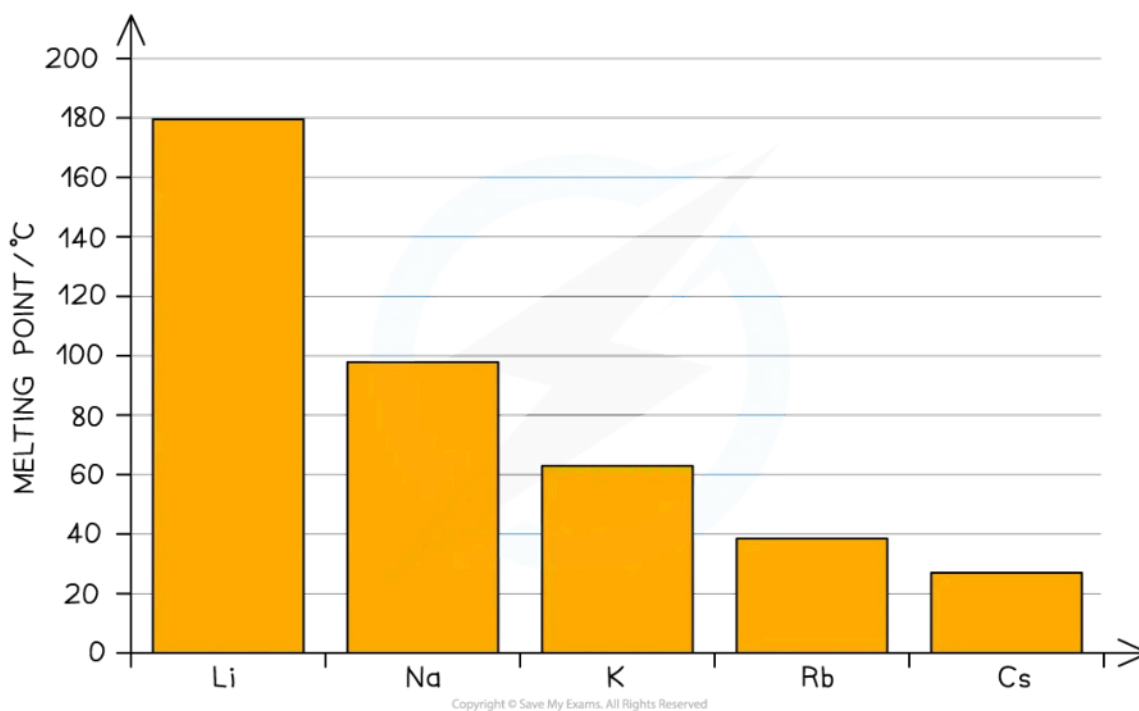


Melting points as you go across a period. The metallic bonding gets stronger from Na to Al

Melting points of metals down a group

- As you go **down the group**, the size of the cation increases
 - This **decreases the attraction** between the outer electrons and the metallic lattice
 - Therefore, there is a reduction in the melting point

Melting point of metals down a group chart



Melting points as you go down a group of metals. The metallic bonding gets weaker from Li to Cs

Examiner Tip

- You see from the chart that the melting point of aluminium is not that much higher than magnesium
- It is a reminder to us that these are trends and not rules about melting points and sometimes there are other factors which can result in subtle differences from what was expected
- One factor here is the metal packing structure, which can also influence the melting point
 - This is beyond what is required in the IB Chemistry syllabus, you just need to learn and explain the broad trends



Your notes

Physical Properties of Transition Elements (HL)

Physical Properties of Transition Elements

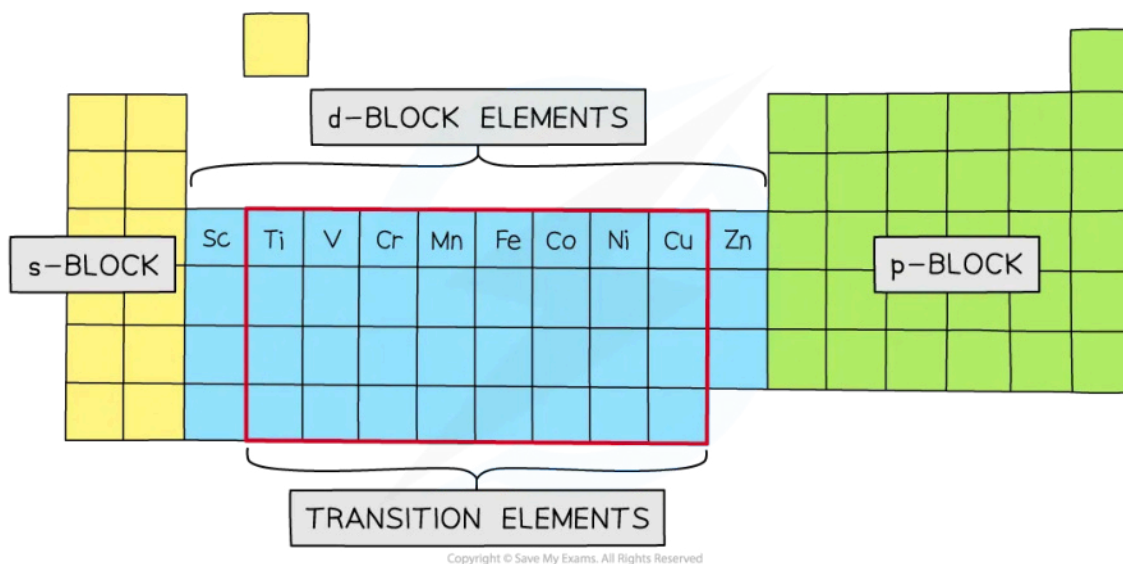
What are transition metals?

- The definition of a transition metal is an element with an incomplete d-subshell or an element that can form at least one stable cation with an incomplete d-subshell
- This definition distinguishes them from d-block elements because scandium and zinc do not fit the definition
 - Scandium only forms the ion Sc^{3+} , configuration $[\text{Ar}] 3d^0$
 - Zinc only forms the ion Zn^{2+} , configuration $[\text{Ar}] 3d^{10}$
- The elements of the first transition series are therefore titanium to copper

Where are transition metals on the Periodic Table?

- The transition metals are located in the d-block
 - Period 4: From titanium to copper
 - Period 5: From zirconium to silver
 - Periods 6 and 7 are complicated by the presence of the f-block lanthanides and actinides

Location of transition metals in the Periodic Table



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The transition elements and the d-block elements

Structure and properties of transition metals

- Like other metals, transition metals have a metallic lattice structure
 - Layers of positive ions within a sea of delocalised electrons



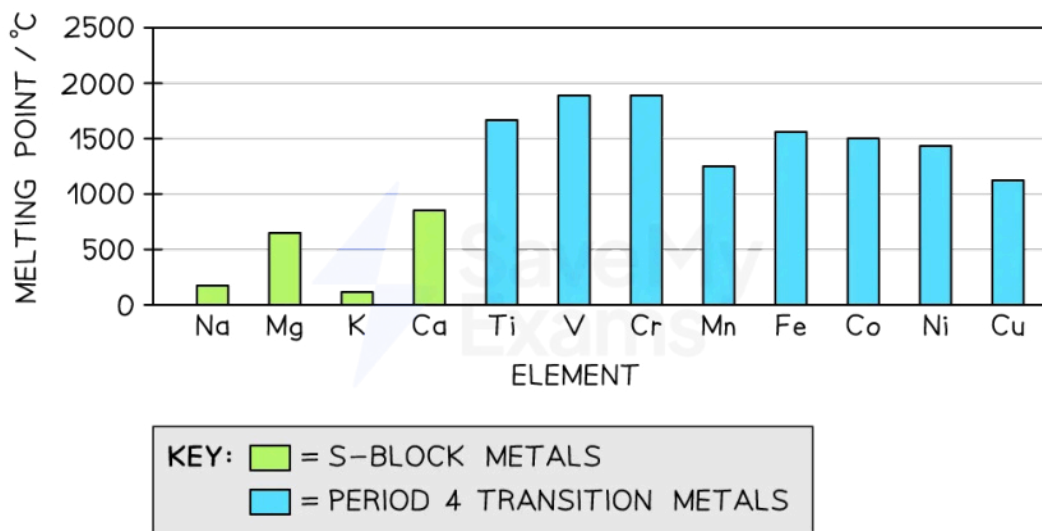
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- Since the 3d and 4s subshells are so close in energy, the transition metals are able to delocalise their d-electrons to form metallic bonds
- This causes transition metals to have particularly good electrical conductivity and high melting points

Why do transition metals have high melting points?

- The ability to delocalise the d-electrons means that transition metals have a greater electron density
 - This means that the electrostatic forces of attraction between the large positive charge of the cations and the sea of delocalised electrons are strengthened
 - The stronger forces of attraction result in a higher melting point as more energy is required to overcome them
- The melting points of s-block metals range from 27 °C for francium to 839 °C for calcium
 - As the following graph shows, all of the Period 4 transition metals have higher melting points than Group 1 and Group 2 metals
 - There is an exception to the lower melting points of s-block metals with a melting point of 1,287 °C for beryllium, due to the small size of a beryllium atom resulting in strong metallic bonding

Melting point graph



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The Period 4 transition metals have higher melting points than s-block metals

Why do transition metals have high electrical conductivity?

- Transition metals have a large number of delocalised electrons
- Therefore, more electrons are able to move when a potential difference is applied
- This causes transition metals to have high electrical conductivity
- The three most conductive metals are:
 - Silver

2. Copper - the most used metal in electrical cables due to a combination of cost and conductivity
3. Gold

- For more information about other characteristic properties of transition metals, see our revision note on the [Characteristic Properties of Transition Elements](#)



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