HL IB Chemistry

Your notes

The Metallic Model

Contents

Properties of Metals & Their Uses

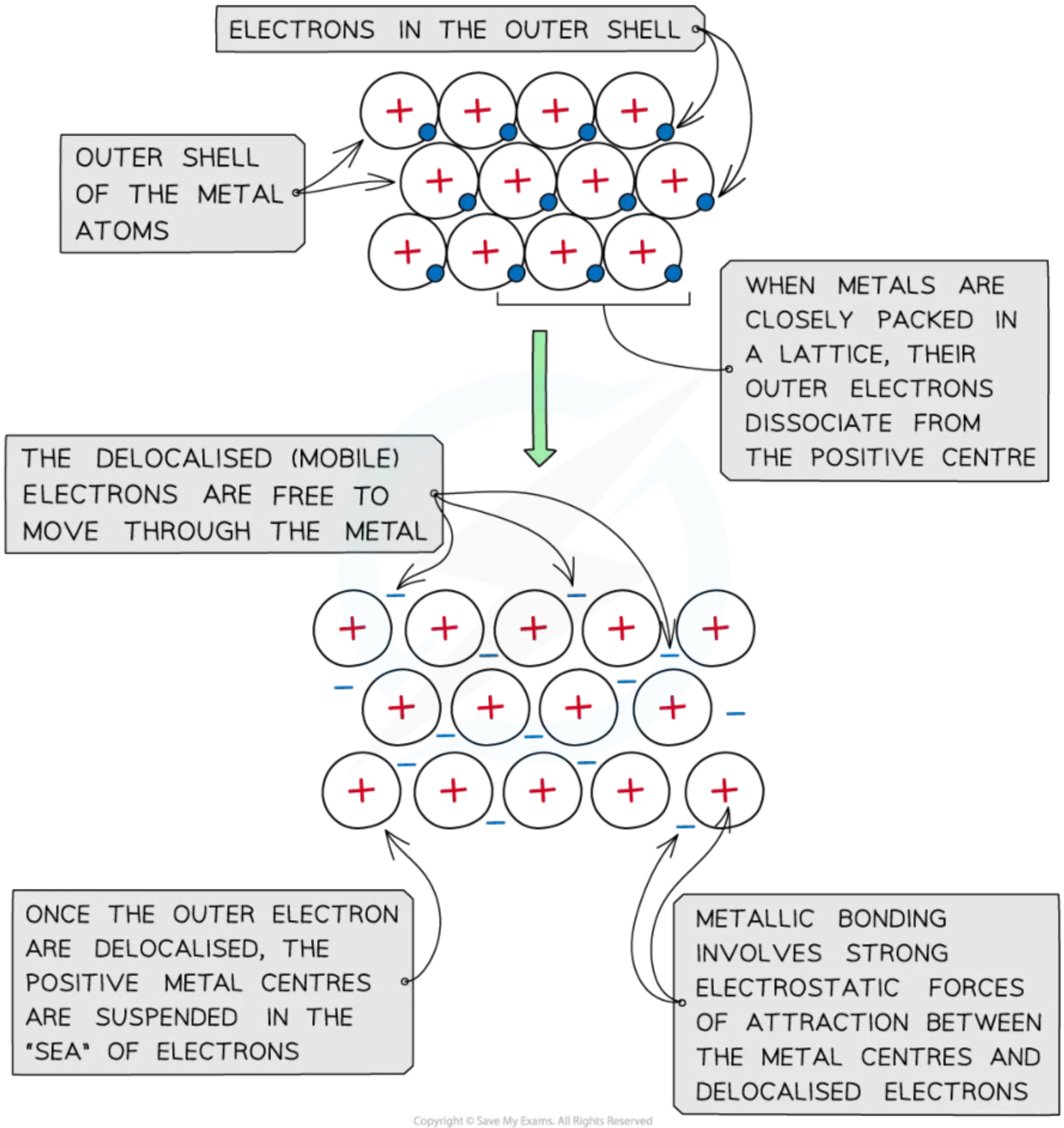
Properties of Metals & Their Uses

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

Your notes



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

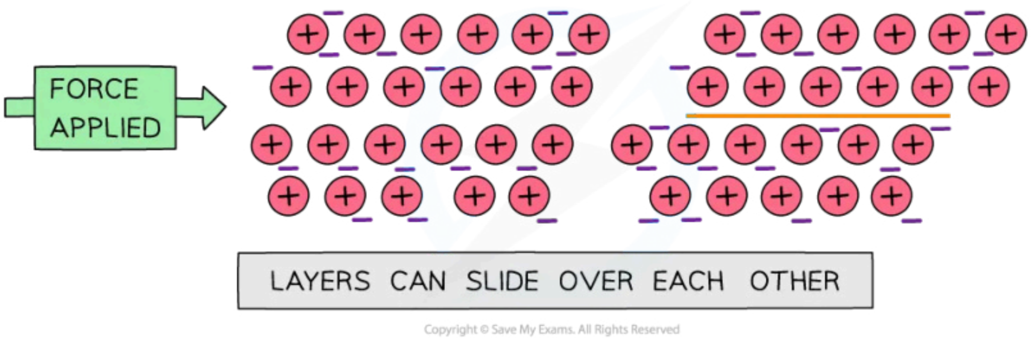
What are the properties of metals?

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

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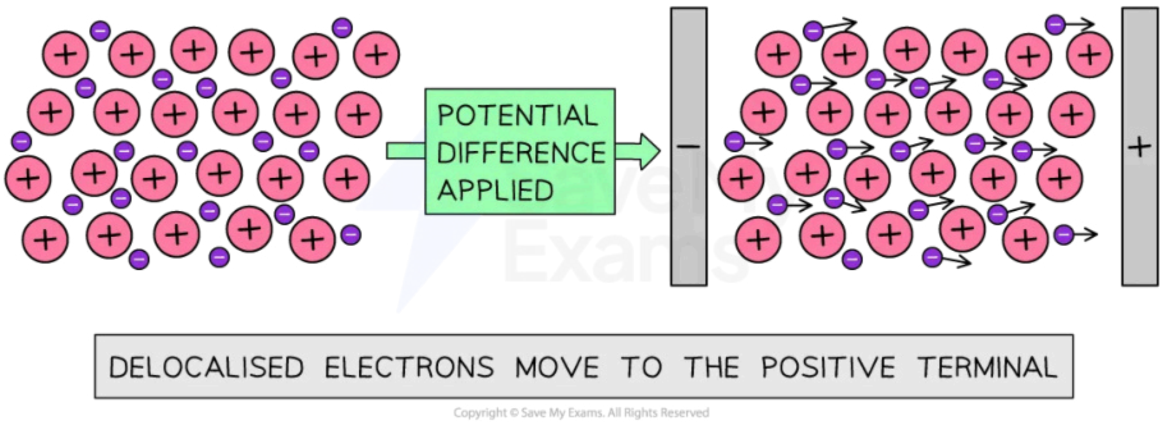

The delocalised electrons move towards the positive terminal when a potential difference is applied

- 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
- Reactivity
- Toxicity
- Lustre
- Corrosion resistance
- 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

Your notes

s & p Block Elements

Your notes

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

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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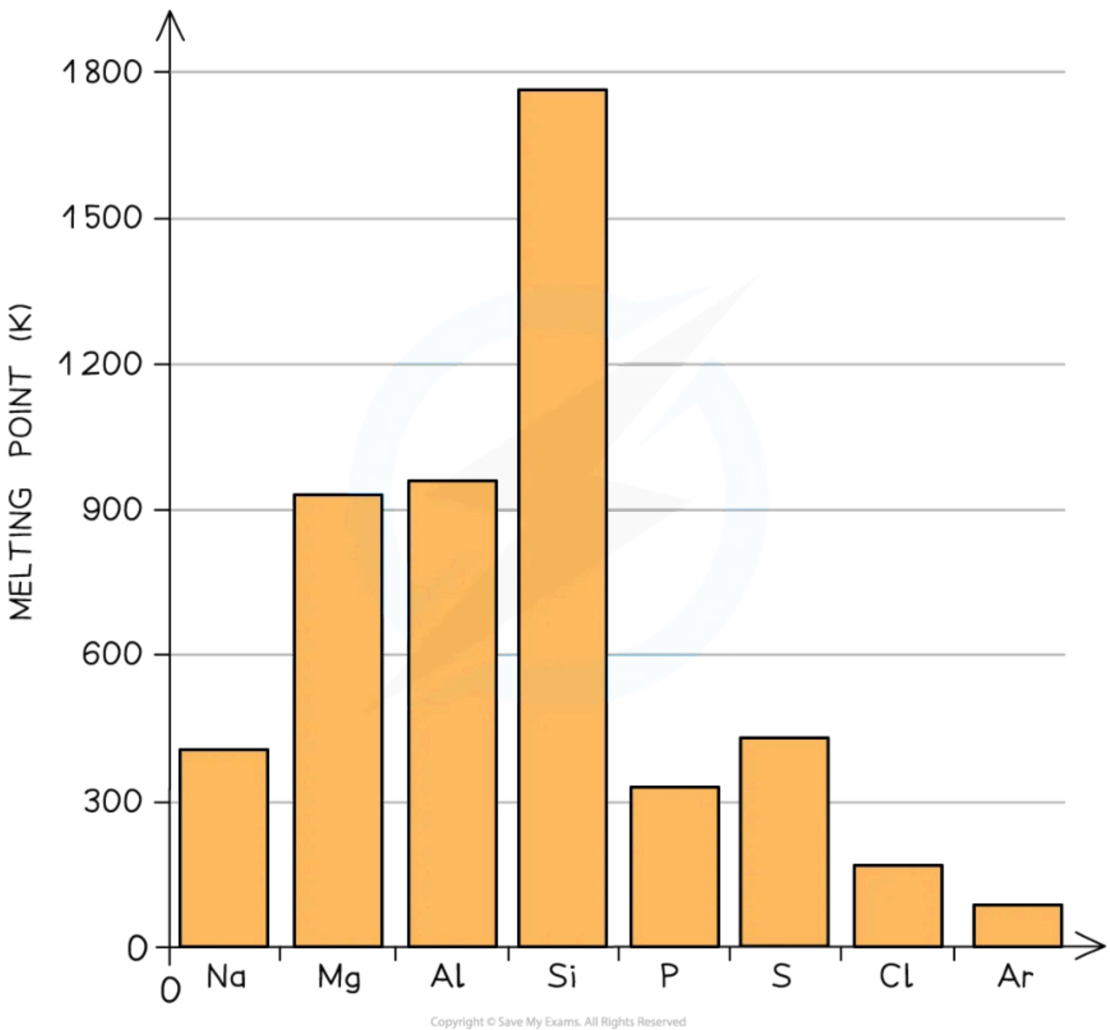
Your notes

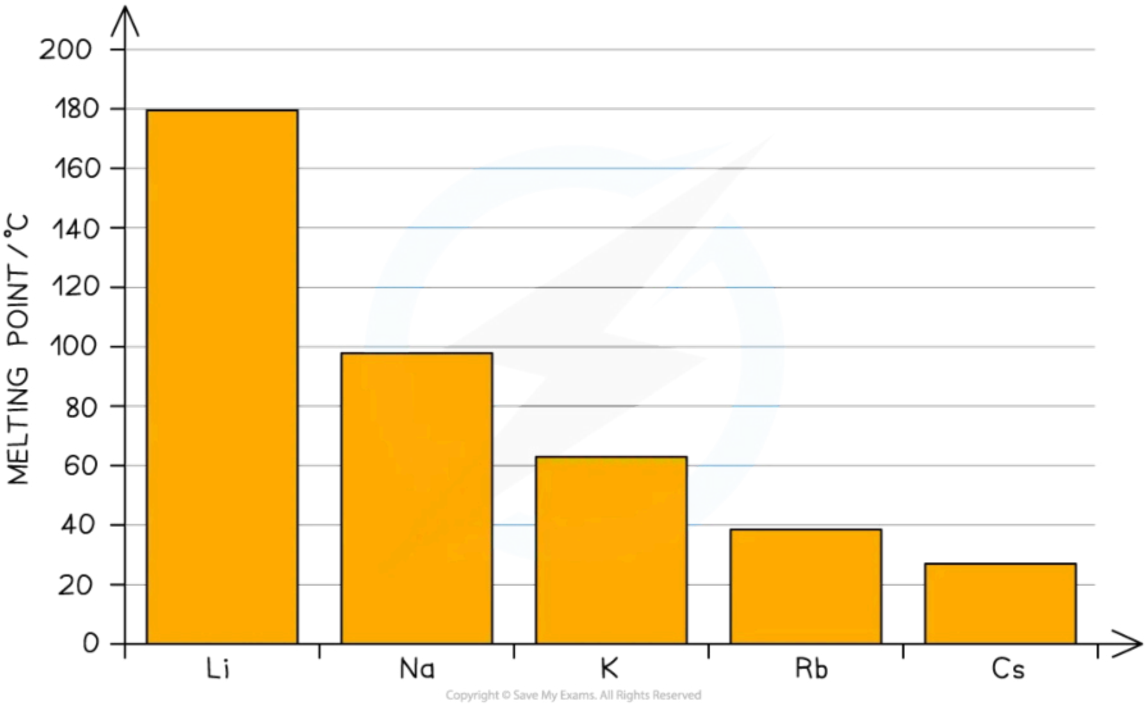


Melting points as you go across a period. The metallic bonding gets stronger from Na to Al

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

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Physical Properties of Transition Elements (HL)

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Physical Properties of Transition Elements

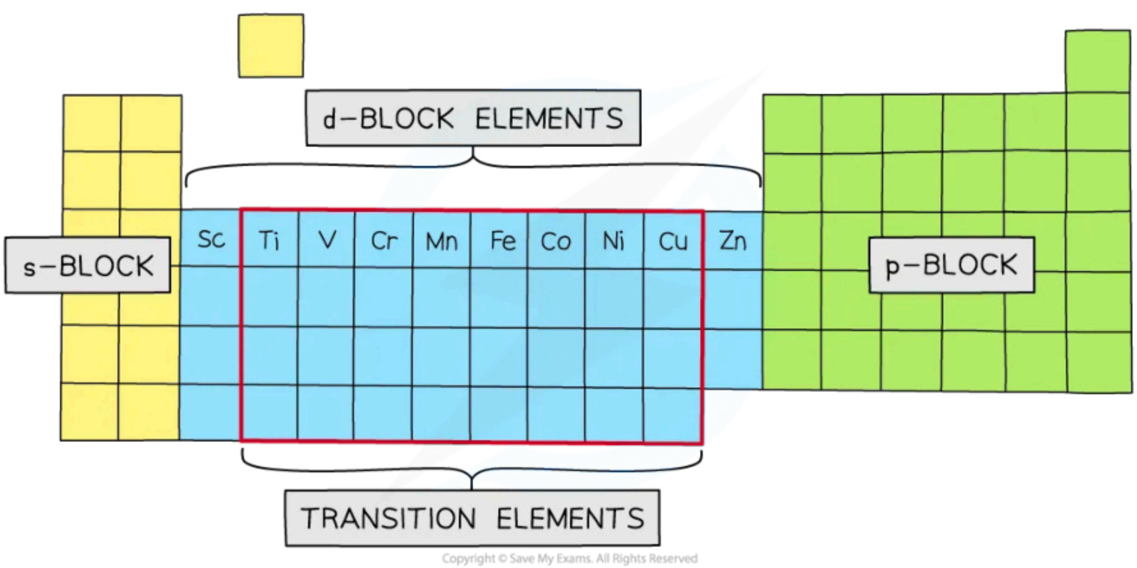
What are transition metals?

- 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 [Ar] $3d^0$
  - Zinc only forms the ion Zn^{2+}, configuration [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



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

- 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

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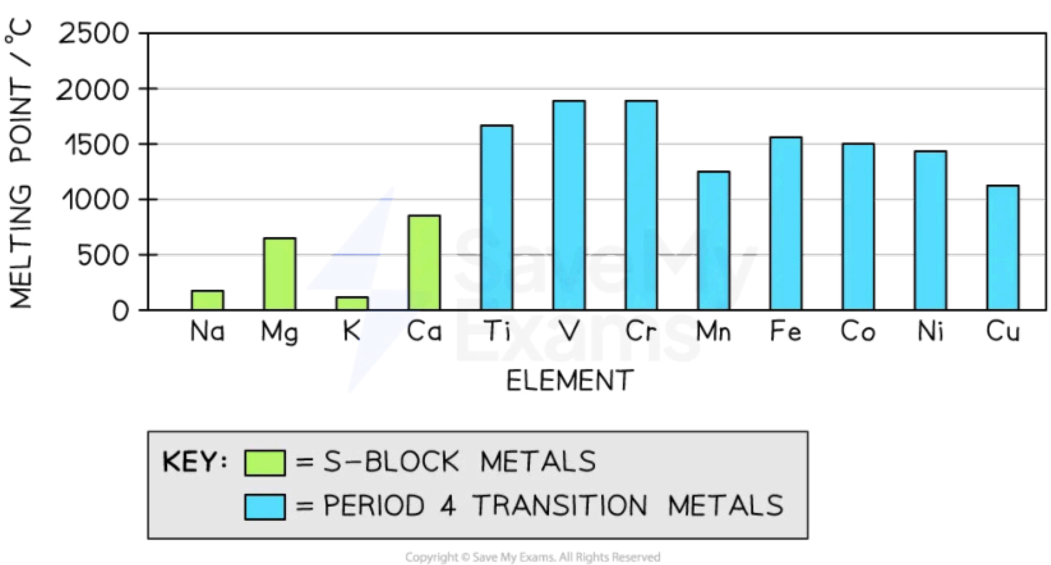
Why do transition metals have high melting points?

- 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



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:
  1. Silver

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

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