Edexcel International AS Chemistry

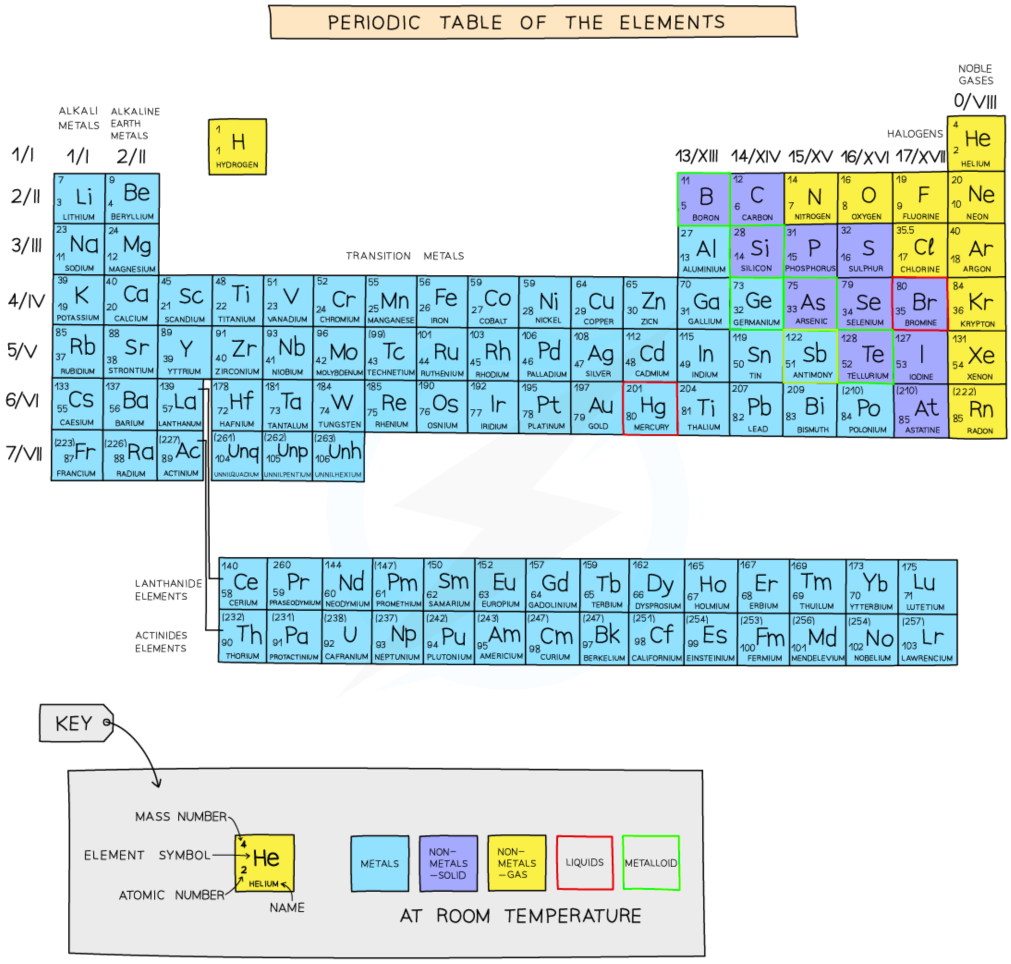
Your notes

The Periodic Table

Contents

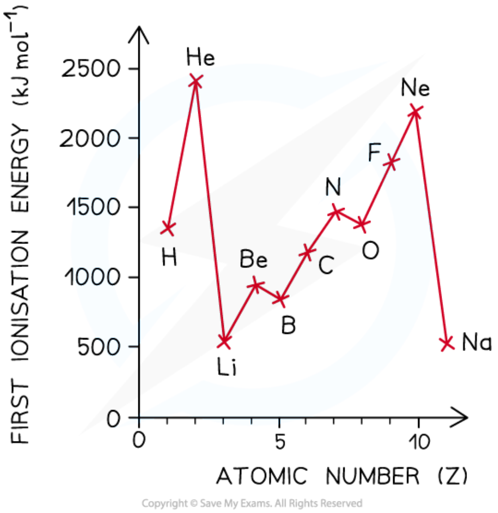
Ionisation Energy Across a Period

- Elements in the periodic table are arranged in order of increasing atomic number and placed in vertical columns (**groups**) and horizontal rows (**periods**)
- The elements across the periods show **repeating patterns** in chemical and physical properties
- This is called **periodicity**



All elements are arranged in the order of increasing atomic number from left to right

- Ionisation energies show a trend across a period of the Periodic Table
- As could be expected from their electron configuration, the group 1 metals have a relatively low ionisation energy, whereas the noble gases have very high ionisation energies



Graph showing first ionisation energies of elements 1 to 11

Ionisation energy across period 2 and 3

- The ionisation energy across a period generally **increases** due to the following factors:
  - Across a period the **nuclear charge increases**
  - This causes the **atomic radius** of the atoms to **decrease**, as the outer shell is pulled closer to the nucleus, so the distance between the nucleus and the outer electrons **decreases**
  - The **shielding** by inner shell electrons remain reasonably constant as electrons are being added to the same shell
  - It becomes **harder to remove an electron** as you move across a period; **more energy** is needed
  - So, the ionisation energy increases

Dips in the trend for period 2

- There is a slight **decrease** in IE_1 between **beryllium** and **boron** as the fifth electron in boron is in the 2p subshell, which is further away from the nucleus than the 2s subshell of beryllium
  - Beryllium has a first ionisation energy of 900 kJ mol^{-1} as its electron configuration is $1s^2\,2s^2$
  - Boron has a first ionisation energy of 800 kJ mol^{-1} as its electron configuration is $1s^2\,2s^2\,2p_x^1$
- There is a slight **decrease** in IE_1 between **nitrogen** and **oxygen** due to **spin-pair repulsion** in the $3p_x$ orbital of oxygen
  - Nitrogen has a first ionisation energy of 1400 kJ mol^{-1} as its electron configuration is $1s^2\,2s^2\,2p_x^1\,2p_y^1\,2p_z^1$

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- Oxygen has a first ionisation energy of 1310 kJ mol^{-1} as its electron configuration is $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$

- In oxygen, there are 2 electrons in the $2p_x$ orbital, so the repulsion between those electrons makes it slightly easier for one of those electrons to be removed

Dips in the trend for period 3

- There is again a slight decrease between **magnesium** and **aluminium** as the thirteenth electron in aluminium is in the 3p subshell, which is further away from the nucleus than the 3s subshell of magnesium
  - Magnesium has a first ionisation energy of 738 kJ mol^{-1} as its electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2$
  - Aluminium has a first ionisation energy of 578 kJ mol^{-1} as its electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1$
- There is a slight decrease in IE_1 between **phosphorus** and **sulfur** due to **spin-pair repulsion**in the $3p_x$ orbital of sulfur
  - Phosphorus has a first ionisation energy of 1012 kJ mol^{-1} as its electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$
  - Sulfur has a first ionisation energy of 1000 kJ mol^{-1} as its electron configuration is $1s^2 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$
- In sulfur, there are 2 electrons in the $3p_x$ orbital, so the repulsion between those electrons makes it slightly easier for one of those electrons to be removed

Ionisation Energy Down a Group

Ionisation energy down a group

- The ionisation energy down a group **decreases** due to the following factors:
  - The number of protons in the atom is increased, so the **nuclear charge** increases
  - But, the atomic radius of the atoms increases as you are adding more shells of electrons, making the atoms bigger
  - So, the **distance** between the nucleus and outer electron **increases** as you descend the group
  - The **shielding** by inner shell electrons **increases** as there are more shells of electrons
  - These factors outweigh the increased nuclear charge, meaning it becomes **easier to remove the outer electron** as you descend a group
  - So, the ionisation energy decreases

Summary Explanation for Trend in Ionisation Energy Across a Period and Down a Group

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Across a Period: Ionisation Energy Increases	Down a Group: Ionisation Energy Decreases
Increase in nuclear charge	Increase in nuclear charge
Shell number is the same Distance of outer electron to nucleus decreases	Increase in shells Distance of outer electron to nucleus increases **Shielding** effect increases, therefore, the attraction of valence electrons to the nucleus decreases
Shielding remains reasonably constant	Increased shielding
Decreased atomic/ionic radius	Increased atomic/ionic radius
The outer electron is held more tightly to the nucleus so it gets **harder** to remove it	The outer electron is held more loosely to the nucleus so it gets **easier** to remove it

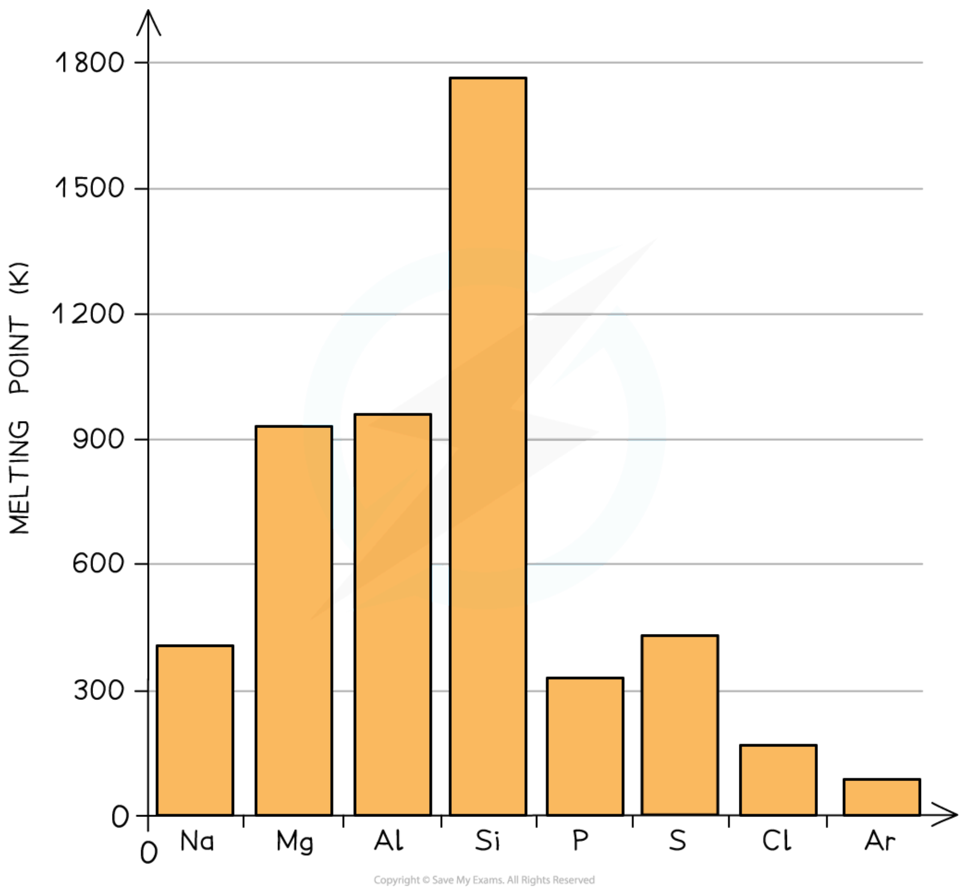
Trends in Melting & Boiling Point

Melting point

- Period 2 and 3 elements follow the **same pattern** in relation to their melting points

Melting Points of the Elements Across Period 3 Table

Period 3 Element	Na	Mg	Al	Si	P	S	Cl	Ar
Melting Point (K)	371	923	932	1683	317	392	172	84



Melting points of the period 3 elements

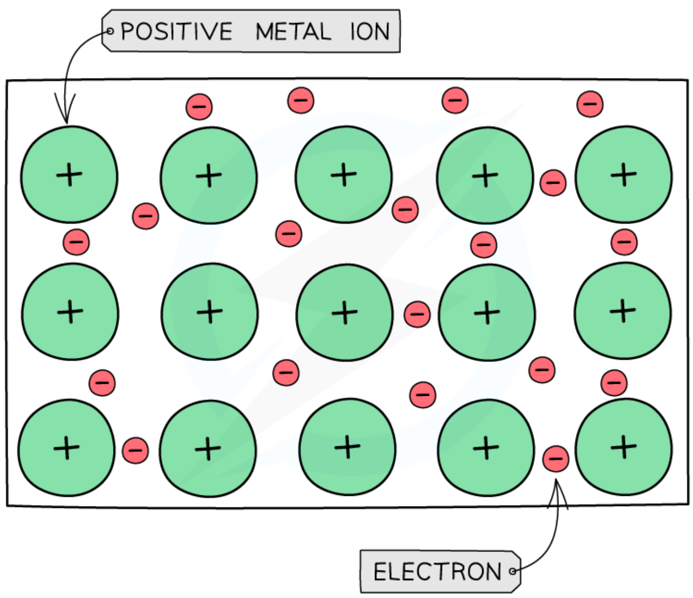
- A general increase in **melting point** for the Period 3 elements up to silicon is observed

- **Silicon** has the highest **melting point**
- After the Si element the melting points of the elements **decreases** significantly
- The above trends can be explained by looking at the bonding and structure of the elements

Bonding & Structure of the Elements Table

Period 3 Element	Na	Mg	Al	Si	P	S	Cl	Ar
Bonding	Metallic	Metallic	Metallic	Covalent	Covalent	Covalent	Covalent	–
Structure	Giant metallic	Giant metallic	Giant metallic	Giant molecular	Simple molecular	Simple molecular	Simple molecular	Simple molecular

- The table shows that **Na**, **Mg** and **Al** are metallic elements which form positive ions arranged in a **giant lattice** in which the ions are held together by a 'sea' of delocalised electrons



Metal cations form a giant lattice held together by electrons that can freely move around

Examiner Tips and Tricks

- **Remember:** At room temperature and pressure, metals (except for mercury) are **solid**
- This means that the lattice structure should:

- Have a regular arrangement of positive ions (rows and columns)
  - Have the ions tightly packed / close together
- The lattice structure should **not**:
  - Have large gaps between the positive ions
    - This could lose marks in an exam as examiners may not be satisfied that a solid is being shown
  - Have a random arrangement of particles
    - Examiners would consider a randomly arranged, close packed structure to be a liquid and penalise answers accordingly
- The delocalised electrons do not have to be specifically shown

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- The electrons in the 'sea' of delocalised electrons are those from the **valence shell** of the atoms
- **Na** will donate one electron into the 'sea' of delocalised electrons, **Mg** will donate two and **Al** three electrons
- As a result of this, the metallic bonding in **Al** is stronger than in **Na**
- This is because the electrostatic forces between a **3+ ion** and the larger number of negatively charged delocalised electrons is much larger compared to a **1+ ion** and the smaller number of delocalised electrons in Na
- Because of this, the **melting points increase** going from **Na** to **Al**
- **Si** has the highest melting point due to its giant molecular structure in which each Si atom is held to its neighbouring Si atoms by **strong covalent bonds**
- **P, S**, **Cl** and **Ar** are non-metallic elements and exist as **simple molecules** (P_4, S_8, Cl_2 and Ar as a single atom)
- The **covalent bonds within** the molecules are strong, however, **between** the molecules, there are only weak **instantaneous dipole-induced dipole forces**
- It doesn't take much energy to break these **intermolecular** forces
- Therefore, the melting points decrease going from **P** to **Ar** (note that the melting point of S is higher than that of P as sulphur exists as larger S_8 molecules compared to the smaller P_4 molecule)