

IONISATION ENERGY

DEFINITIONS

The **first ionisation energy (I.E.)** is defined as the energy required to remove one electron from each atom in one mole of gaseous atoms.

e.g. for sodium atoms: $\text{Na(g)} \rightarrow \text{Na}^+(\text{g}) + \text{e}^-$

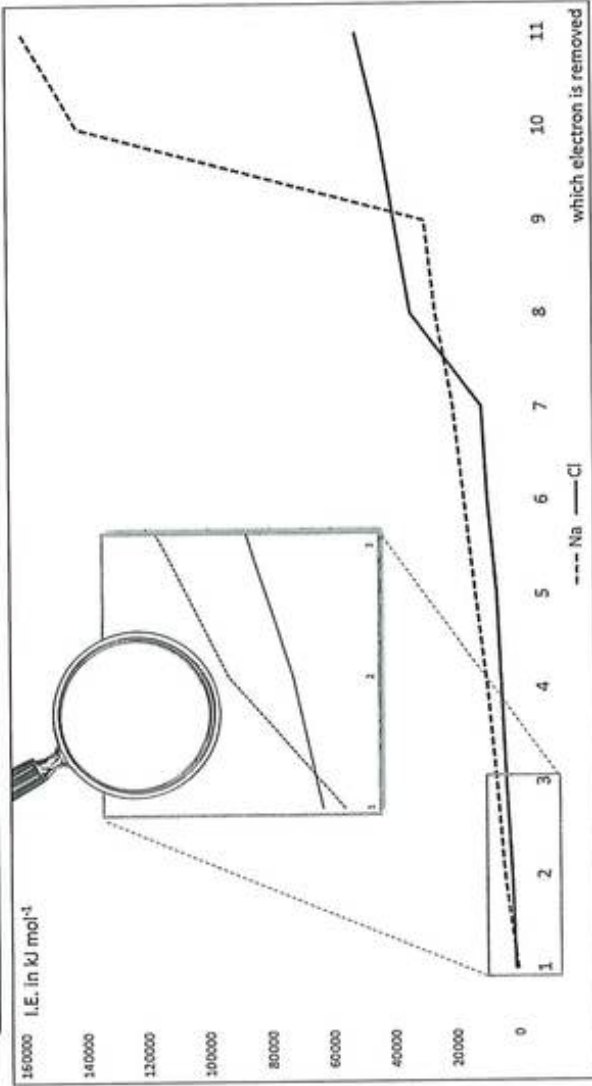
The **second I.E.** for sodium would involve this: $\text{Na}^+(\text{g}) \rightarrow \text{Na}^{2+}(\text{g}) + \text{e}^-$

The **third I.E.** for sodium would involve this: $\text{Na}^{2+}(\text{g}) \rightarrow \text{Na}^{3+}(\text{g}) + \text{e}^-$

EXAM TIP

You only need to define the **1st I.E.** but you may need to write equations for the 2nd, 3rd, 4th, etc.

EVIDENCE FOR SHELLS AND SUBSHELLS



The ionisation energies provide evidence for the electronic structure. For example, if you plot successive ionisation energies from 1^{st} to 11^{th} for Na and Cl, you see the energy rising smoothly but with sudden, dramatic increases at certain points.

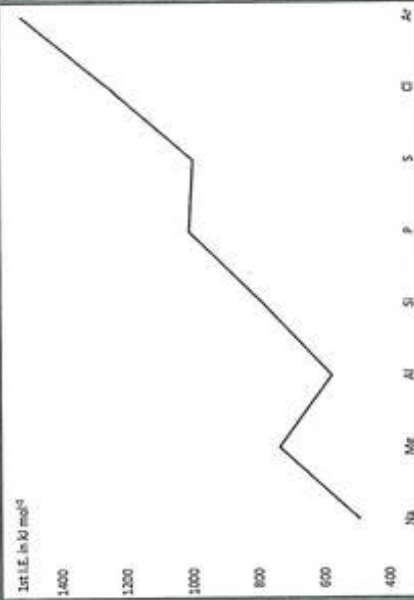
For Na, which is in group 1 (and has the electronic configuration 2,8,1), the first sharp increase is between the 1^{st} and 2^{nd} ionisation energies (look at the magnified portion of the graph) and then there is another from the 9^{th} to the 10^{th} . For Cl, in group 7 (with electronic configuration 2,8,7), the sharp rise is between the 7^{th} and 8^{th} ionisation energies.

Each time the ionisation energy jumps, it is because all electrons from a shell have been removed and a lot more energy is needed to remove the next electron as it is in a shell nearer to the nucleus.

TREND IN 1ST IONISATION ENERGY IN PERIOD 3

The graph on the left shows the 1^{st} ionisation energy of each of the elements in period 3 in order from Na to Ar. Ionisation energy generally increases because each element has one more electron in its nucleus than the last, and the extra positive charge attracts the electrons more strongly. However, there are two places in which the ionisation energy falls.

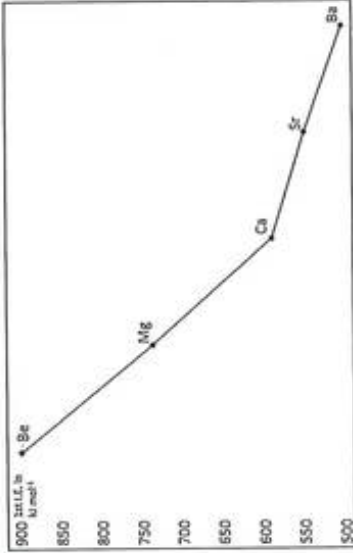
- Aluminium (electronic structure $1s^2 2s^2 2p^6 3s^2 3p^1$) has a lower 1^{st} ionisation energy than magnesium (electronic structure $1s^2 2s^2 2p^6 3s^2$). This is because the electron is being removed from a $3p$ orbital in aluminium, which takes less energy than removing an electron from the $3s$ orbital in magnesium.
- Sulfur (electronic structure $1s^2 2s^2 2p^6 3s^2 3p^4$) has a slightly lower 1^{st} ionisation energy than phosphorus (electronic structure $1s^2 2s^2 2p^6 3s^2 3p^3$). This is because the $3p^4$ configuration of sulfur involves one of three $3p$ orbitals having a pair of electrons in it. It takes less energy to remove a paired electron than an unpaired one due to repulsion between the two electrons.



TREND IN 1ST IONISATION ENERGY IN GROUP 2

The graph on the right shows the 1^{st} ionisation energy of each of the elements in group 2 in order of increasing atomic number from Be to Ba. Even though the atomic number, and so the positive charge of the nucleus, is increasing, the ionisation energy decreases. The reason for this is that in each successive element, the electron that is being removed (an outer-shell electron) is in a higher shell than in the previous element. There are two consequences to this that you need to describe when explaining the trend in ionisation energy:

- Each element has more electrons in the inner shells than the previous element and these **shield** the outer-shell electrons from the nucleus, reducing the force of attraction.
- In each element, the outer shell is **further from the nucleus** than in the previous one. The force of attraction decreases as the distance between the electron and the nucleus increases.



EXAM TIP

You may need to refer to trends in 1^{st} I.E. in group 2 and period 3 in order to explain how we know about the existence of shells and subshells.