2.13 Ionisation Energies

Definition :**First ionisation energy**

The first ionisation energy is the energy required when one mole of gaseous atoms forms one mole of gaseous ions with a single positive charge

This is represented by the equation: $H_{(g)} \rightarrow H^{+}_{(g)} + e^{-}$

Always gaseous

The equation for 1st ionisation energy always follows the same pattern.
It does not matter if the atom does not normally form a +1 ion or is not gaseous

$Be_{(g)} \rightarrow Be^{+}_{(g)} + e^{-}$

$O_{(g)} \rightarrow O^{+}_{(g)} + e^{-}$

All values of ionisations are positive and therefore endothermic. This is because energy must be supplied to overcome the electrostatic attractive force between the nucleus and the electron.

Definition :**Second ionisation energy**

The second ionisation energy is the energy required when one mole of gaseous ions with a single positive charge forms one mole of gaseous ions with a double positive charge

This is represented by the equation: $Ti^{+}{}_{(g)} \rightarrow Ti^{2+}{}_{(g)} + e^{-}$

Factors that affect Ionisation energy

There are three main factors
1.The attraction of the nucleus
(The more protons in the nucleus the greater the attraction)
2. The distance of the electrons from the nucleus
(The bigger the atom the further the outer electrons are from the nucleus and the weaker the attraction to the nucleus)
3. Shielding of the attraction of the nucleus
(An electron in an outer shell is repelled by electrons in complete inner shells, weakening the attraction of the nucleus)

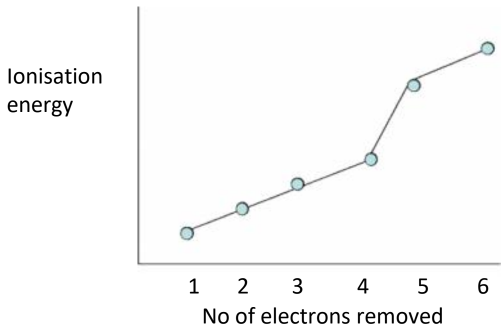
Many questions can be answered by application of these factors

Successive ionisation energies

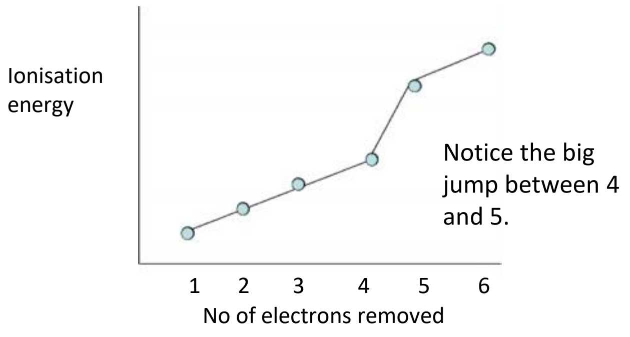
An element can have as many successive ionisation energies as it has electrons

The patterns in successive ionisation energies for an element give us important information about the electronic structure for that element.

Why are successive ionisation energies always larger?
The second ionisation energy of an element is always bigger than the first ionisation energy. When the first electron is removed a positive ion is formed. The ion increases the attraction on the remaining electrons and so the energy required to remove the next electron is larger.
Each successive ionisation energy is bigger than the previous one for the same reason.
Some of the increases are much bigger, however, and these big jumps gives us evidence for the main principle electron shells.



How are ionisation energies linked to the main electron energy levels ?

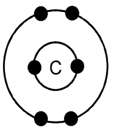


There is a big jump between the 4th and 5th ionisation energies.

Explanation

The fifth electron is in a inner main shell closer to the nucleus and therefore attracted much more strongly by the nucleus than the fourth electron.

It also does not have any shielding by inner complete shells of electron so is easier to remove.



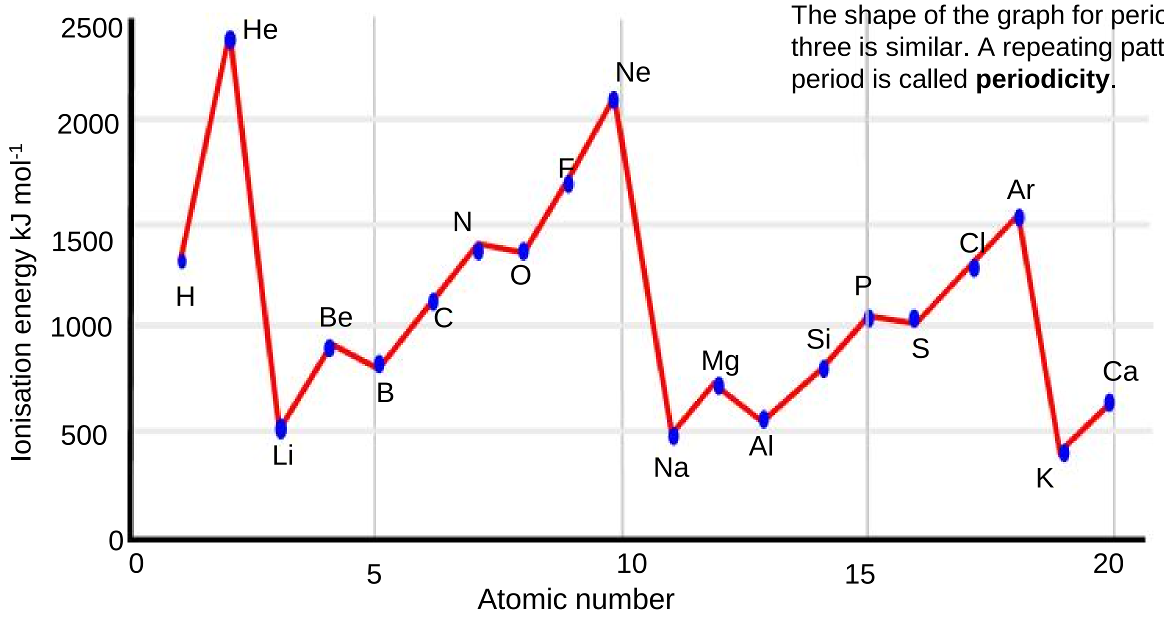
Example: What group must this element be in?

	1	2	3	4	5
Ionisation energy kJ mol^{-1}	590	1150	4940	6480	8120

Here there is a big jump between the 2nd and 3rd ionisations energies which means that this element must be in group 2 of the periodic table as the 3rd electron is removed from an electron shell closer to the nucleus with less shielding and so has a larger ionisation energy

The first Ionisation energy of the elements

The pattern in the first ionisation energy of each successive element in the periodic table also gives us useful information about electronic structure

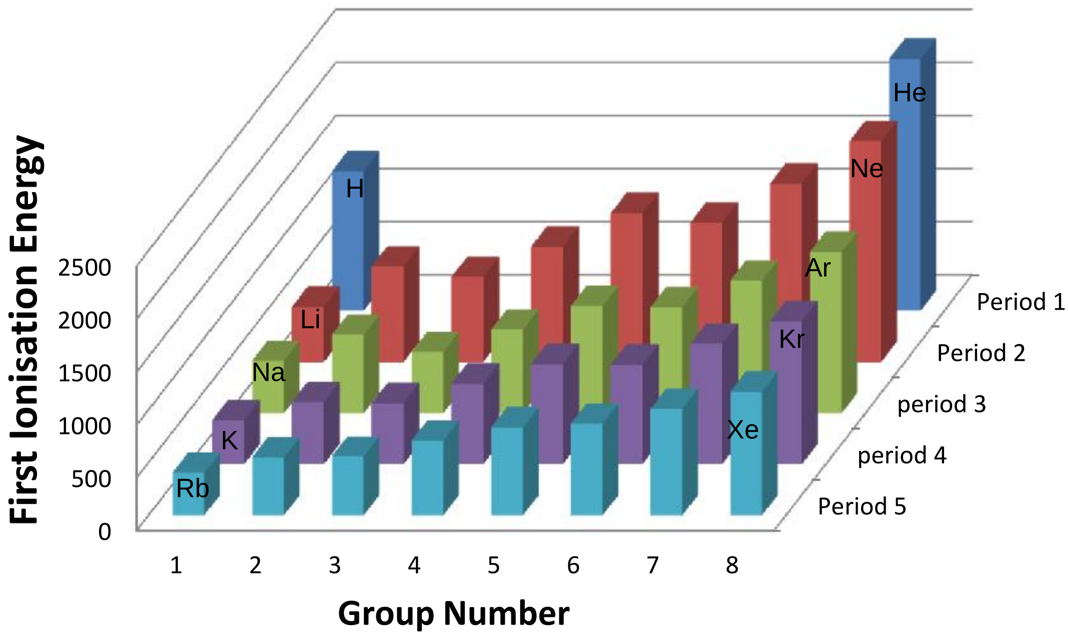


The shape of the graph for periods two and three is similar. A repeating pattern across a period is called **periodicity**.

Evidence for the main electron energy levels

The noble gases are always at the maximum peak for each period, but there is a decrease in ionisation energy down the group. (true of all groups). This is because as one goes down the group the outer electrons become further from the nucleus and become more shielded from the nuclear pull by complete inner shells

There is a large drop each time between the group 0 elements and the group1 elements.
This is because the element in group 1 will have its outer electron in a new shell further from the nucleus and is more shielded. So the group 1 element is easier to remove and has a lower ionisation energy.

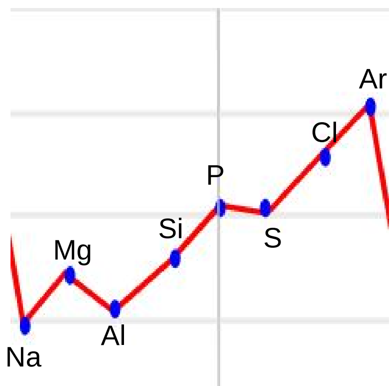


Notice the drop in ionisation energy down each group

As one goes down a group, the outer electrons are found in shells further from the nucleus and are more shielded so the attraction of the nucleus becomes smaller.

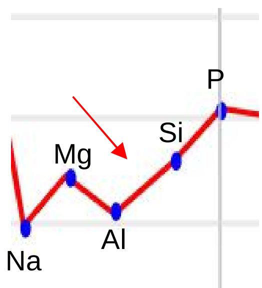
Helium has the biggest first ionisation energy because its first electron is in the first shell closest to the nucleus and has no shielding effects from inner shells. Helium has a bigger first ionisation energy than hydrogen as it has one more proton.

Evidence for the electron sub energy levels

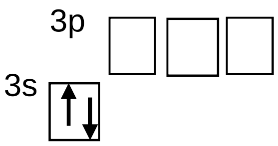


There is in general an increase in ionisation energy across a period. This is because as one goes across a period , the number of protons increases making the effective attraction of the nucleus greater. The electrons are being added to the same shell which has the same shielding effect and the electrons are pulled in closer to the nucleus.

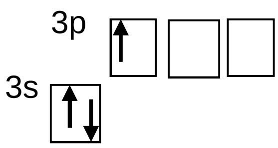
There are two small drops in the general trend, however, that provide further evidence for the existence of sub energy levels.



Notice the small drop between the group 2 elements and group 3 elements (Be + B , Mg + Al). Aluminium is starting to fill a 3p sub shell, whereas magnesium has its outer electrons in the 3s sub shell. The electrons in the 3p subshell are slightly easier to remove because the 3p electron sub shell is slightly higher in energy and they are also slightly shielded by the 3s electrons.

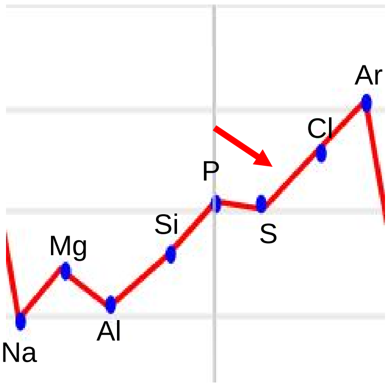


Magnesium
$1s^2 2s^2 2p^6 3s^2$

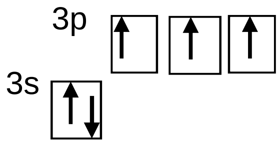


Aluminium
$1s^2 2s^2 2p^6 3s^2 3p^1$

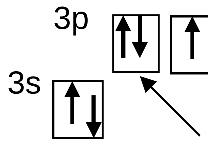
Why is there a small drop from P to S?



With sulfur there are 4 electrons in the 3p sub shell and the 4th is starting to doubly fill the first 3p orbital.
When the second electron is added to a **3p orbital** there is a slight **repulsion** between the two negatively charged electrons which makes the second electron easier to remove.



phosphorus $1s^2\,2s^2\,2p^6 3s^2 3p^3$

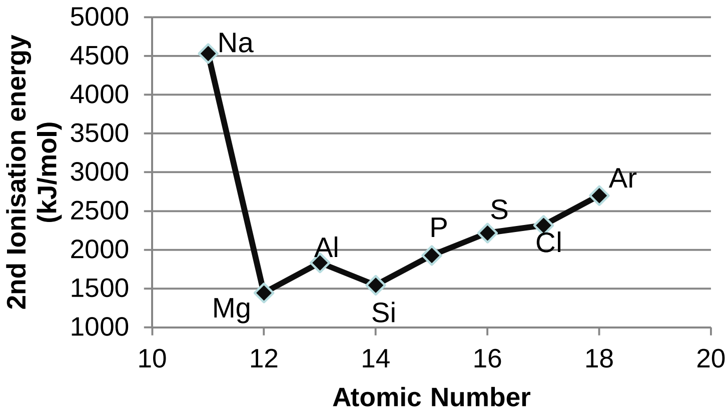


Two electrons of opposite spin in the same orbital

sulfur $1s^2\,2s^2\,2p^6 3s^2 3p^4$

Patterns in the second ionisation energy.

If the graph of second ionisation or each successive element is plotted then a similar pattern to the first ionisation energy is observed but all the elements will have shifted one to the left.



The group 1 elements are now at the peaks of the graph

Lithium would now have the second largest ionisation of all elements as its second electron would be removed from the first 1s shell closest to the nucleus and has no shielding effects from inner shells. Li has a bigger second ionisation energy than He as it has more protons.

Questions on Ionisation Energies

1) What is meant by the term *first ionisation energy*?
2) Write an equation to illustrate the process which occurs during the first ionisation of neon.
3) Explain why the value of the first ionisation energy of magnesium is higher than that of sodium.
4) Explain why the value of the first ionisation energy of neon is higher than that of sodium.
5) Write an equation to illustrate the process occurring when the **second** ionisation energy of magnesium is measured.
6) Explain why the **third** ionisation energy of magnesium is very much larger than the **second** ionisation energy of magnesium.
7) The Ne atom and the Mg^{2+} ion have the same number of electrons. Give two reasons why the first ionisation energy of neon is lower than the third ionisation energy of magnesium.
8) Sketch a graph to show the first 5 ionisation energies of aluminium,
9) State and explain the trend in the first ionisation energy of the elements Mg to Ba in Group II.
10) Explain why the **second** ionisation energy of **sodium** is greater than the **second** ionisation energy of **magnesium**.
11) Explain why the **second** ionisation energy of magnesium is greater than the **first** ionisation energy of magnesium.

12) Use your understanding of electron arrangement to complete the table by suggesting a value for the third ionisation energy of magnesium.

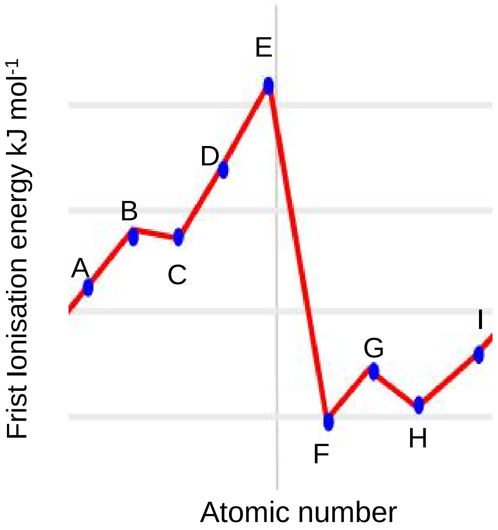
Ionisation energies of magnesium	1	2	3	4	5
kJ mol^{-1}	736	1450		10500	13629

13) State which of the first, second or third ionisations of aluminium would produce an ion with the electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^1$
14) Explain why the ionisation energy of every element is endothermic
15) Explain why the **second** ionisation energy of carbon is higher than the **first** ionisation energy of carbon.
16) Explain which element has the largest first ionisation energy
17) Identify the element in **Period 2** that has the highest **first** ionisation energy and give its electron configuration.
18) Identify the element in Period **1 or 2** that has the highest **second** ionisation energy
19) Why is the first ionisation energy of krypton is lower than the first ionisation energy of argon.
20) Explain why the first ionisation energy of rubidium is lower than the first ionisation energy of sodium.
21) Give one reason why the **second** ionisation energy of silicon is lower than the **second** ionisation energy of aluminium.
22) Predict the element in Period 3 that has the highest **second** ionisation energy. Give a reason for your answer.
23) State and explain the general trend in the values of the first ionisation energies of the elements Na to Ar.
24) State how, and explain why, the values of the first ionisation energies of the elements Al and S deviate from the general trend.
25) Explain why the first ionisation energy of boron is lower than that of beryllium
26) There is a similar general trend in first ionisation energies for the Period 4 elements gallium to krypton to that in period 3. State how selenium deviates from this general trend and explain your answer

27) The table below gives the successive ionisation energies of an element. Deduce the group in the Periodic Table that contains this element

Ionisation energies	1	2	3	4	5
kJ mol^{-1}	590	1150	4940	6480	8120

28 The sketch graph below shows the trend in first ionization energies for some elements in Periods two and three.



a) Select, from the elements A to I, the one that has atoms with five p electrons
b) Select, from the elements A to I, which one is a member of group 3
c) Select, from the elements A to I, which one form a +2 ion
d) Select, from the elements A to I, which one is likely to be very unreactive.
e) Select, from the elements A to I, which one(s) normally forms four covalent bonds per atom
f) Name Element D
g) What would the be the formula of the compound formed between element F and C

29) What is wrong/ incomplete with the answers to the following question 'Which element has the largest 1st Ionisation energy? '

a) "Helium's outer electron is close to the nucleus with little shielding so hard to remove"

b) "Helium is a noble gas and has a full shell of electrons. It would lose stability if an electron was removed."

c) "Helium has a really strong ionic bond between its electron and nucleus because it has no shielding."