

Name:

Chemical Bonding	Objectives
7. Trends in The Periodic Table	-define and explain atomic radius -explain the general trends in values of atomic radii (covalent radii only) <ul style="list-style-type: none"> • down a group • across a period (main group elements only) -define and explain first ionisation energy -explain the general trends in first ionisation energy values: <ul style="list-style-type: none"> • down a group • across a period (main group elements) and -explain the exceptions to the general trends across a period -define and explain second and successive ionisation energies -describe how second and successive ionisation energies provide evidence for energy levels -recognise the relationship and trends in successive ionisation energies of an individual element -explain how chemical properties of elements depend on their electronic structure -explain how atomic radius, screening effect and nuclear charge account for general trends in properties of elements in groups I and VII -recognise the trends in electronegativity values down a group and across a period -explain the general trends in electronegativity values <ul style="list-style-type: none"> • down a group • across a period

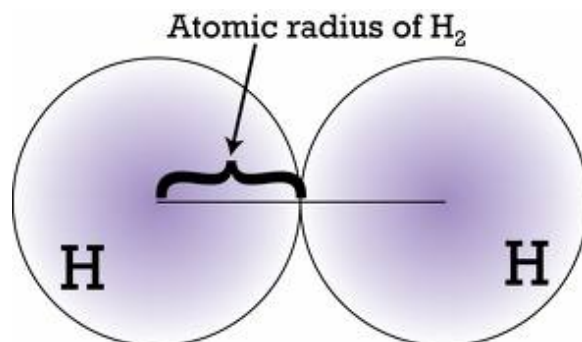
Trends in Atomic Radii:

Defⁿ: The **atomic radius** of an atom is defined as half the distance between the nuclei of two atoms of the same element that are joined by a single covalent bond.

The values of the atomic radius increases going down groups in the periodic table.

Reasons:

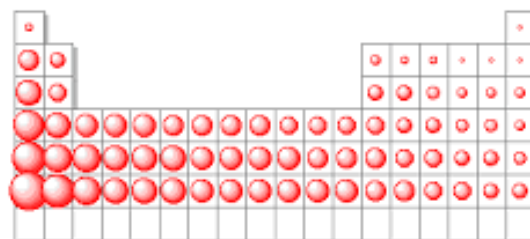
1. Each time we move down the periodic table we add an extra energy level further away from the nucleus, making it bigger.
2. With these extra energy levels further from the nucleus as we move down the periodic table, the **screening effect** of the inner electrons reduces (cancels out) the pull the positive nuclear charge has on the outer electrons.



The values of the atomic radius decreases going across periods in the periodic table.

Reasons:

1. The effective nuclear charge of the nucleus increases going across a period (more positive charge as we have more protons in the nucleus), which pulls the outer electrons closer to the nucleus.
2. No increase in screening effect as all elements in the same period have the same outer energy level.



Trends in Ionisation Energy:

Defⁿ: The **First Ionisation Energy** of an atom is the minimum energy required to completely remove the most loosely bound electron from a neutral gaseous atom in its ground state.

This means the energy that's needed to pull the loosest (furthest from the nucleus) electron off an atom.

The values of the first ionisation energy decrease going down groups in the periodic table.

Reasons:

1. Increasing atomic radius as we go down a group means that the outermost electron gets further and further from the nucleus, making it easier to remove.
2. The increasing screening effect of the inner electrons as we go down a group also makes the most loosely bound electron easier to remove.

The values of the first ionisation energy increase going across periods in the periodic table.

Reasons:

1. Increasing effective nuclear charge as we go across a period means that the most loosely bound electron gets pulled more strongly by the nucleus as we go across the group. More energy is then needed to remove this electron.
2. Decreasing atomic radius means that the most loosely bound electron gets closer to the nucleus of the atom as we go across the period. This means the nucleus has a stronger pull on the electron, so it becomes more difficult to remove the electron.

Exceptions to the Trend Across a Period:

Look at the graph of ionisation energies for the first 20 elements in the periodic table.

Look at the trend from Li to Ne, the second period of the periodic table.

Generally the ionisation energy increases, however there are exceptions.

From Be to B there is a dip.

To explain this, look at the electron configuration of Be and B:

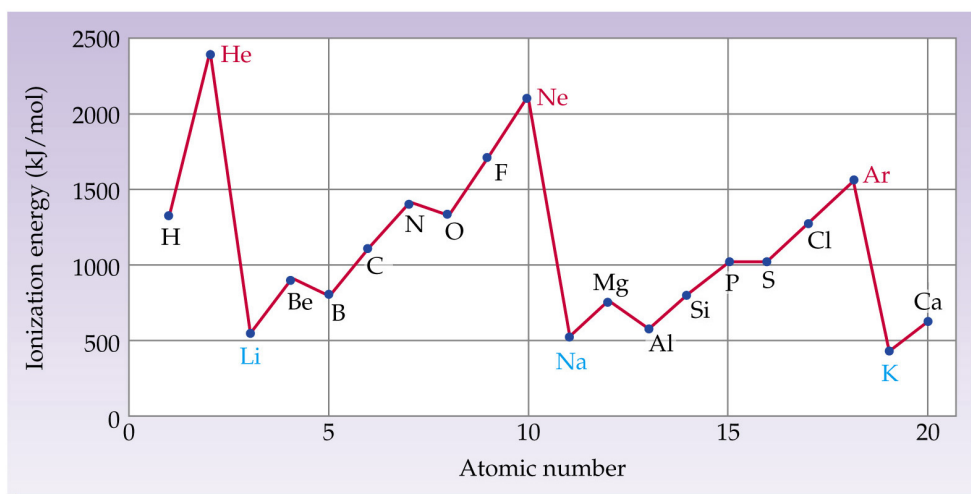


The outermost sublevel of Be (2s) is full, whereas the outermost sublevel of B (2p) has only 1 electron. Atoms whose outermost sublevel is **half full** or **completely full** have **extra stability**. As Be has a completely full outer sublevel, its ionisation energy is particularly high.

Likewise for the dip from N to O, look at their electron configurations:



N has a half filled outer sublevel (2p), O does not have a half filled or completely filled outer sublevel, meaning that N has extra stability. This explains why N has a particularly high ionisation energy.



Ionisation Energies are Used as Evidence for the Existence of Energy Levels:

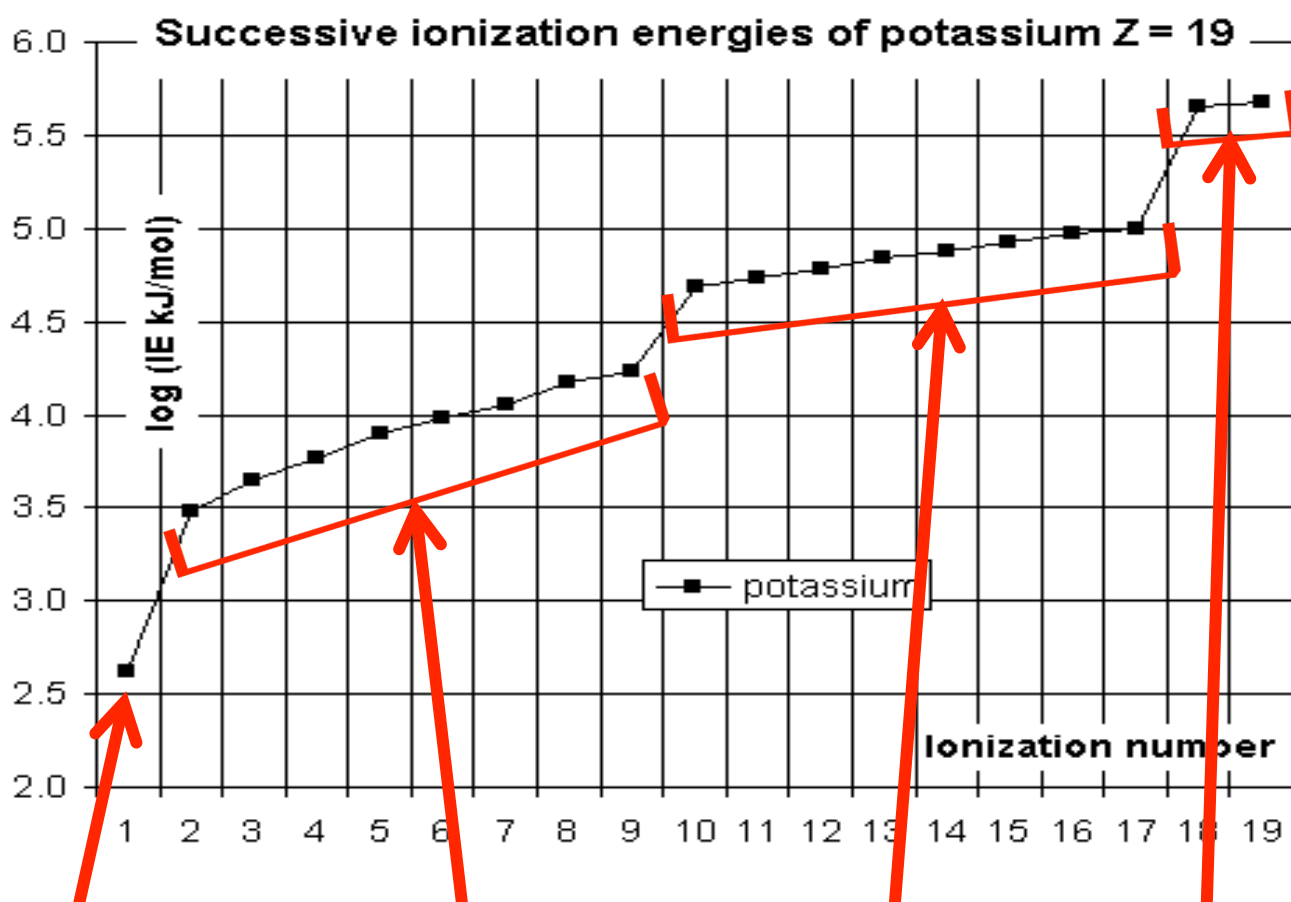
Along with the Defⁿ given for the *First Ionisation Energy*, there is another one needed:

Defⁿ: **Second Ionisation Energy** is the energy required to remove an electron from an ion with one positive charge in the gaseous state.

This means the energy needed to remove the second loosest electron, after the loosest has already been removed. This definition can keep going to define the third, fourth, etc., ionisation energies.

Look at the graph below which shows a plot of the log of ionisation energy vs. number of electrons removed from an atom of Potassium: (the log is used because the energy values get very, very big, very, very quickly! Don't worry about this)

Keep in mind the the electron configuration of Potassium, K, is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$



<p>The first ionisation is quite low as the first electron removed comes from the n=4 energy level, which is furthest away from the nucleus.</p>	<p>The second electron is much more difficult to remove than the first (large spike in energy required). This suggests that an electron is being removed from a filled energy level (closer to the nucleus). This is true as electrons are now being removed from the n=3 energy level. Notice there are eight electron removals with similar energy. This is because there are 8 electrons in the n=3 energy level.</p>	<p>Large increase because we are removing an electron from the filled n=2 energy level (closer to the nucleus). There are 8 electrons in the n=2 energy level, explaining why they have similar ionation energies.</p>	<p>Large increase because we are removing an electron from the filled n=1 energy level (closer to the nucleus). There are 2 electrons in the n=1 energy level, explaining why they have similar ionation energies.</p>
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Trends In Electronegativity:

The values of electronegativity decrease going down groups in the periodic table.

Reasons:

1. Increasing atomic radius. This means that as you go down a group, the atom gets bigger, meaning that the atom has a weaker pull on electrons in a bond as they are further away.
2. The Screening Effect of the inner electrons increases going down a group. This reduces the pull that the nucleus has on outer electrons involved in bonding.

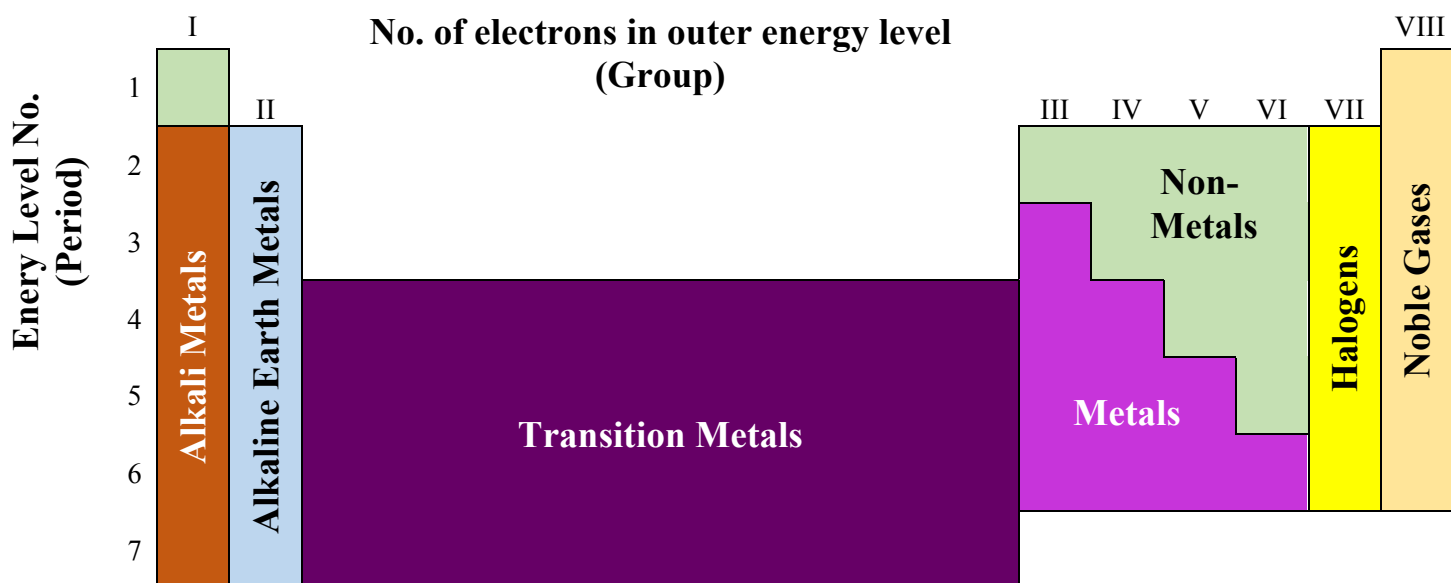
The values of electronegativity increase going across periods in the periodic table.

Reasons:

1. Increasing effective nuclear charge. This increases the pull of the nucleus on the outer electrons involved in bonding.
2. Decreasing atomic radius. As you go across a period, atomic radius gets smaller, meaning that the nucleus is closer to the outer electrons involved in bonding. This results in the nucleus having a stronger pull (attraction) on the outer bonding electrons.

Trends Within Groups:

The names of certain groups in the periodic table need to be known. They are shown below:

**Trends in the Alkali Metals (Group I):**

All Alkali Metals are very reactive because they have only 1 electron in their outer energy level, which is easy to remove as can be seen by their Ionisation Energies.

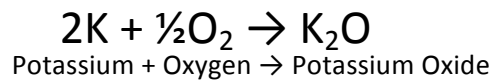
Trend: The chemical reactivity of the Alkali Metals increases going *down* the group.

Reasons:

1. As you go down the group, you add an extra energy level further from the nucleus with every step down. This increases the screening effect on the inner electrons, making the outer electron easier to remove.
2. By adding an extra energy level with each step down, we also increase the atomic radius, reducing the pull that the nuclear charge has on the outer electron, making it easier to remove.

Important Reactions with Alkali Metals:**1. Reaction with Oxygen (O₂):**

All Alkali Metals react with oxygen to form oxides:



Any other Alkali Metal can replace the Potassium in this reaction.

Notes:

- This reaction is the reason why Alkali Metals are stored under oil – so they won't react with oxygen in the air.
- This reaction causes Alkali Metals to lose their shiny colour when exposed to air – the oxides are dull.

2. Reaction with Water (H₂O):

All Alkali Metals react with water vigorously to form hydroxides and Hydrogen gas (H₂):



Any other Alkali Metal can replace the Sodium in this reaction.

Notes:

- This reaction demonstrates the increasing reactivity of the Alkali Metals going down a group. Potassium reacts much more vigorously with water than Lithium does.

Trends in the Halogens (Group VII):

All Halogens are very reactive because they only need 1 extra electron to fill their outer energy level.

Trend: The reactivity of the Halogens increases going *up* the group.

Reason:

1. The Halogens are the most electronegative of all elements. The values of their electronegativities increase going up the group, meaning that the higher up the group we go, the greater the pull they have on electrons, making it easier for them to react and fill their outer energy level.