Trends in the Periodic Table

Atomic Radius

The **atomic radius** of an element is half the distance between the nuclei of two atoms of that element when bonded by a single covalent bond.

Trend 1- Atomic Radii decrease across a period in the periodic table due to

- Increase in effective nuclear charge - the electrons in the outer shell are pulled closer to the increasingly positive nucleus whilst the number of inner shells remains the same across the period
  (Effective nuclear charge = nuclear charge - no. of electrons in inner shells)
- No increase in the shielding effect - same number of shells across a given period

Trend 2 - Atomic radii increase down a group due to

- Increase in the number of shells (moving down a group a new shell is added)
- Increase in the shielding effect - additional inner shells of electrons shield the outer electrons from the positive nucleus.

***Note** noble gases don't bond so they have no atomic radius**

- Positive ions are smaller than their parent atoms (i.e. have a smaller atomic radius). The positively charged nucleus has less electrons to pull on so electrons are pulled closer (also, the number of shells may decrease as an electron is removed)
- Negative ions are larger than their parent atoms - the gain in electrons and no increase in nuclear charge means that the outer electrons are held more loosely.

First Ionisation Energy

First ionisation energy is the minimum energy required to remove the most loosely held electron from one mole of gaseous atoms in their ground state (measured in $kJmol^{-1}$)

Trend 1 - Ionisation energy values increase across a period because:

- Atomic radii decrease across a period. As a result the most loosely held electron lies increasingly close to the positive nucleus, making it more difficult to remove.
- Effective nuclear charge increases across a period - nucleus has an increasingly strong attraction for the outermost shell of electrons

Trend 2 - Ionisation energy values decrease down a group because:

- Increase in atomic radii - outermost electron is increasingly further away from the nucleus and so less energy is needed to remove it.
- Increase in the shielding effect - more inner shells down a group shield the outermost electron from the positive nucleus

A **half-filled sublevel** has extra stability associated with it and therefore more energy is required to remove the outermost electron

Groups that have higher than expected ionisation energy values

- group II elements - have a full outermost (s) sublevel
- group V elements - have a half-filled outermost (p) sublevel
- group 0 elements - have a full outermost (p) sublevel
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How to explain trends in ionisation energy values within a particular element/How jumps in ionisation energy values provide evidence for the existence of energy levels

1. There is a steady increase in ionisation energy values as electrons are removed from the same shell. This is due to an increase in effective nuclear charge.
2. A large increase in ionisation energy values occurs when an electron is removed from a new sublevel. This is due to the extra stability associated with having a full outermost sublevel.
3. A very large increase in ionisation energy occurs when an electron is removed from a new shell - elements with a fully filled outer shell have very high stability

These jumps in ionisation energy values are evidence for the existence of energy levels.

Electronegativity

Electronegativity is the relative attraction that an atom in an molecule has for the shared pair of electrons in a single covalent bond (electronegativity has no unit)

Trend 1 - Electronegativity values increase across a period because:

1. Decreasing atomic radius - nucleus gets increasingly close to the shared pair of electrons and so the force of attraction between them grows stronger
2. Increase in effective nuclear charge - increased attraction between the nucleus and the shared pair of electrons

Trend 2 - Electronegativity values decrease down a group because

1. Increasing atomic radius - shared pair of electrons are further from the nucleus
2. Decreasing effective nuclear charge - decreased attraction between the nucleus and the shared pair of electrons

***Note** noble gases do not have electronegativity values as they do not form bonds*

Group trends

Group I – alkali metals

- The atoms of the elements in group I all have one electron in their outer shell, thus their chemical properties are similar. In reaction with a non-metal, an alkali metal atom loses its outer electron and becomes an ion with a single positive charge, +1.
- The more easily the outer electron is lost, the more reactive the metal.
- Since the number of electron shells and atomic radii increase down the group, the outer electron becomes further away from the positively charged nucleus and so the force of attraction between them grows smaller. This results in the electron being more and more easily lost. ⇒ Reactivity increases down the group.

Group VII – the halogens

- The halogens have 7 electrons in their outermost shell, thus their chemical properties are similar. In reaction with a metal each halogen atom gains an electron and becomes an ion with a single negative charge, -1.
- The more easily these outer electrons are gained the more reactive the metal.
- Since atomic radii increase down the group, the outermost shell of electrons becomes increasingly further from the nucleus. As a result, the force of attraction between the positively charged nucleus and an electron from another atom becomes weaker. Hence, it becomes harder to attract and gain an electron moving down the group. ⇒ Reactivity decreases down the group.

Sample Exam Questions

TRENDS IN THE PERIODIC TABLE

2018 Q4(b)

(*b*) State and give the reason for the trend in atomic radii across the second period of the periodic table.

(b) *State:* decrease in atomic radius
Reason: increasing nuclear charge, which means the electrons are more attracted to the nucleus

2017 Q5

(c)(i) *describe:* atomic radius is increasing
Explain: an electron is entering a new shell
(ii)nuclear charge is increasing

(d) *first ionisation energy:* the minimum energy required to remove the most loosely bound electron from a gaseous atom in its ground state

-Ionisation energy increases due to a decreasing atomic radius and increasing nuclear charge, which make the electrons closer to the nucleus and more strongly attracted to the nucleus
(e)(i) There are six values of ionisation energy, showing that there are six electrons present in a carbon atom
(ii) -The first 4 ionisation energies are increasing steadily, showing that these electrons are all in the same outer energy level
-There is a large jump between the 4th and 5th ionisation energies, showing that the 5th electron is from the innermost energy level
-The 5th and 6th electrons have a similar ionisation energy, showing that they are in the same energy level

2014

b) *First ionisation energy is the minimum energy required to remove the most loosely held electron from one mole of gaseous atoms in their ground state.*

i) Silicon has a <u>smaller atomic radius</u> than aluminium. This means that the outermost electron in silicon is closer to its positively charged nucleus than the outermost electron in aluminium is.

Also, the <u>effective nuclear charge in silicon is greater</u> than that in aluminium because of a greater positive nuclear charge but no additional inner shells. Hence, the silicon nucleus has a greater attraction for the outermost electron than the aluminium nucleus does and so more energy is required to remove the outer electron in silicon.

ii) Silicon has a <u>greater atomic radius</u> than carbon. This means that the outermost electron is further from the nucleus in silicon than in carbon. As a result the force of attraction between the nucleus and the outermost electron is less in silicon than in carbon and so less energy is required to remove this outer electron.

c) i)

1. Ionisation energy values steadily increase as electrons are removed from the outermost shell of the silicon atom (1-4) due to an increasing effective nuclear charge.

2. A large increase in ionisation energy occurs as electrons are removed from a new energy level (4-5). The stability associated with having a full outer octet of electrons means that more energy is required to remove an electron.

3. Ionisation energy values steadily increase as the eight electrons are removed from silicon's 2nd energy level

4. A very large increase in ionisation energy occurs as an electron is removed from silicon's 1st energy level. These electrons require a huge amount of energy to remove as they are held very closely to the nucleus and are not shielded from the by inner shells of electrons.

These jumps in ionisation energies provide energy for the existence of energy levels.

ii) Line emission/absorption spectra from elements

2013

a) First ionisation energy is the minimum energy required to remove the most loosely held electron from one mole of gaseous atoms in their ground state.

i) B = helium

P = sulfur

x = 900

ii) R = argon, S = potassium

- *Argon is very stable as it has a full outer octet of electrons. As a result a large amount of energy is required to remove the outermost electron, i.e. its first ionisation energy is high.*
- *Potassium, however, has just one electron in its outermost shell. By losing this electron, potassium can achieve a full outer octet of electrons, i.e. the electron is readily lost. This results in potassium having a much lower first ionisation energy value than argon.*

iii) H=oxygen, G= nitrogen

- *Nitrogen has a half filled outer (p) sublevel and so has an extra stability associated with it. This stability means that more energy is required to remove the outermost electron*
- *Oxygen, however, does not have a full or half filled outer sublevel and so does not have this stability associated with it.*
  This results in the ionisation energy of nitrogen being slightly higher than that of oxygen.

2015

(*h*) In the periodic table, identify an element
(*i*) in the same period as magnesium but with larger atoms,
(*ii*) in the same group as magnesium but with smaller atoms.

i) Sodium (Na)

ii) Beryllium (Be)

2012

(*b*) Define *atomic radius* (*covalent radius*).
State and explain the trend in atomic radii (covalent radii) across the second period of the periodic table of the elements. (12)

(*c*) Give **one** reason why electronegativity values exhibit a general increase across the second period of the periodic table. (3)

b) i) The atomic radius of an element is half the distance between the nuclei of two atoms of that element when bonded by a single covalent bond.

ii) Atomic radii decrease across the (second period of the) periodic table.

- ⇨ *This is due to an increase in effective nuclear charge across the period - the electrons in the outer shell are pulled closer to the increasingly positive nucleus whilst the number of inner shells remains the same across the period*

c) Electronegativity increases due to the decreasing atomic radius across the period - the nucleus gets increasingly close to the shared pair of electrons and so the force of attraction between them gets stronger

2011

(*c*) One of the most useful features of the periodic table of the elements is that it allows trends in the properties of the elements to be compared.
Explain why (*i*) the alkali metals are all reactive, (*ii*) the reactivity of the alkali metals increases down the group. (9)

i) All alkali metals have one electron in their outer shell. By losing this electron, the atom can achieve a full outer shell of electrons. As a result this electron is readily lost by the alkali metal. This is what is meant by reactivity.

ii) Since the number of electron shells and atomic radii increase down the group, the outer electron becomes further away from the positively charged nucleus and so the force of attraction between them grows smaller. This results in the electron being more and more easily lost.

- ⇨ *Reactivity increases down the group.*