

Periodic trends

Atomic radius: radius of an atom

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca	Ga	Ge	As	Se	Br

Atomic radius increases going down any group of the periodic table.

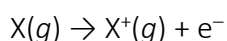
- ◆ Going down the group, electrons are added.
- ◆ Valence electrons are in energy levels further from the nucleus, resulting in the increased radius.
- ◆ **“Atoms have more shells of electrons.”**

Going across any row (period) of the periodic table, atoms get smaller in size from left to right.

- ◆ Increasing number of protons in the nucleus.
- ◆ Protons increase the positive charge of the nucleus, which draws the electrons that are being added *to the same shell* in more tightly.
- ◆ **“Nuclear charge increases (while shielding remains same) causing increasing attraction to the nucleus.”**

Ionisation energy:

First IE: Energy required to remove the outermost electron from one mole of gaseous atoms **OR**



Ionisation energies decrease down the group.

- ◆ Going down the group the element has more electron shells
- ◆ Valence electrons are further from the nucleus so less energy is required to remove a valence electron.
- ◆ Additional shells mean a greater shielding effect between the valence shell and the nucleus, so less energy is required to remove the outermost electron.

E.g. 1st IE Ca (2.8.8.2) < 1st IE Mg (2.8.2)

Ionisation energies increase across the table.

- ◆ Element has same number of electron shells (same shielding / screening) but nuclear charge (number of protons in nucleus) is increasing.
- ◆ Greater electrostatic attraction between nucleus and valence electrons.

E.g. 1st IE Cl (2.8.7) > 1st IE Mg (2.8.2)

2nd, 3rd, 4th IE (Only for Scholarship)

The *n*th ionization energy is the energy required to strip it of the *n*th electron after the first *n* - 1 electrons have been removed. It is considered a measure of the "reluctance" of an atom or ion to surrender an electron, or the "strength" by which the electron is bound; the greater the ionisation energy, the more difficult it is to remove an electron.

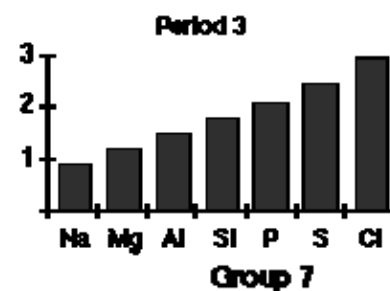
Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.

Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to caesium and francium which are the least electronegative at 0.7.

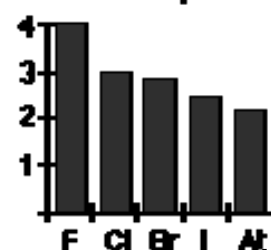
Electronegativity increases across a group

- ◆ nuclear charge increases (while shielding remains same) causing increasing attraction to the nucleus.



Electronegativity decreases down a group

- ◆ whilst nuclear charge and shielding both increase, electron shells are added and so radii increase causing decreasing attraction of electrons to the nucleus.



No electronegativity difference between two atoms leads to a pure non-polar covalent bond. A small electronegativity difference leads to a polar covalent bond.

A large electronegativity difference leads to an ionic bond.

Atomic and ionic radii

Metal ions are smaller than metal atoms.

- ◆ Metals form positive ions by the loss of electron(s).
- ◆ The ion has one less energy level.
- ◆ The same number of protons is also now attracting *less electrons* more strongly.

E.g. Radius of (Na^+) < radius (Na), because the ion Na^+ has: 1 fewer shell / subshell / energy level; And 11 protons are now also attracting only 10 electrons.

Non-metal ions are larger than non-metal atoms.

- ◆ Non-metals form negative ions by the gain of electron(s).
- ◆ There is now greater $e^- - e^-$ repulsion between the electrons. (“Repulsive forces between the valence electrons increase the size of the electron cloud”).
- ◆ The same number of protons is now attracting *more electrons* less strongly.

E.g. Radius (N^{3-}) > radius (N), because N^{3-} has: more valence e^- causing increased electrostatic repulsion between the valence electrons; And 7 protons are now attracting 10 electrons.

Comparing 2 iso-electronic ions:

Anions are bigger than their iso-electronic cation

Isoelectronic means same electron arrangement Eg Ca^{2+} and Cl^- , both with an electron arrangement of 2.8.8.

E.g. Cl^- is much larger than Ca^{2+}

Although both have same number of shells or electron arrangement, the Ca^{2+} has more protons (20) or nuclear charge is greater, so the electrostatic attraction between the valence electrons and the nucleus is stronger, making the Ca^{2+} ion smaller than the Cl^- ion. (Cl has 17 protons).

E.g. F^- and Na^+ have same electron configuration but Na^+ has more protons in the nucleus. This results in increased electrostatic attraction between nucleus and valence electrons so radius of Na^+ is smaller than F^- .