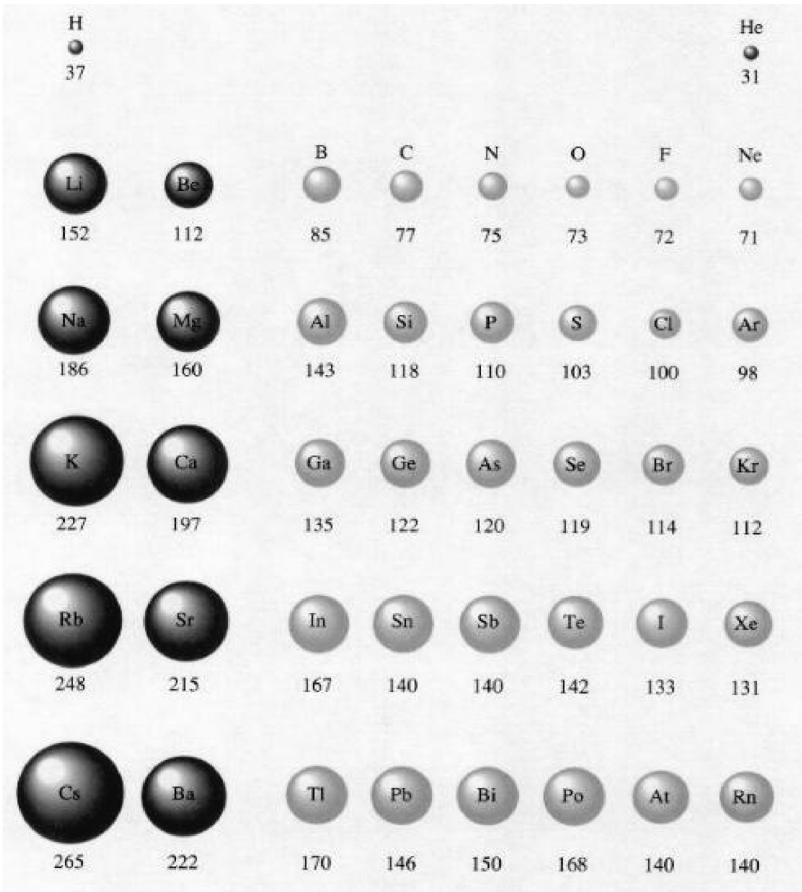
CHEMISTRY 90780

Describe properties of particles and thermochemical principles

- periodic trends in atomic radius, ionisation energy, and electronegativity, and comparison of atomic and ionic radii

Atomic radius decreases

Atomic radius increases



<u>Atomic radius</u> – **decreases** as you go **left** $\longrightarrow$ **right** across the periodic table.

As the <u>atomic number increases</u> so does the <u>positive charge</u> in the nucleus (more protons).

The electron shells are pulled in with <u>more force</u> and the <u>radius decreases</u>. As the valence shell is pulled closer to the nucleus the <u>atomic radius decreases</u>.

Atomic radius **increases** as you **go down a group** in the Periodic Table.

As the number of electron shells increases the valence electrons get further from the nucleus and are less strongly attracted by the increasing number of protons in the nucleus. The atomic radius gets **bigger**.

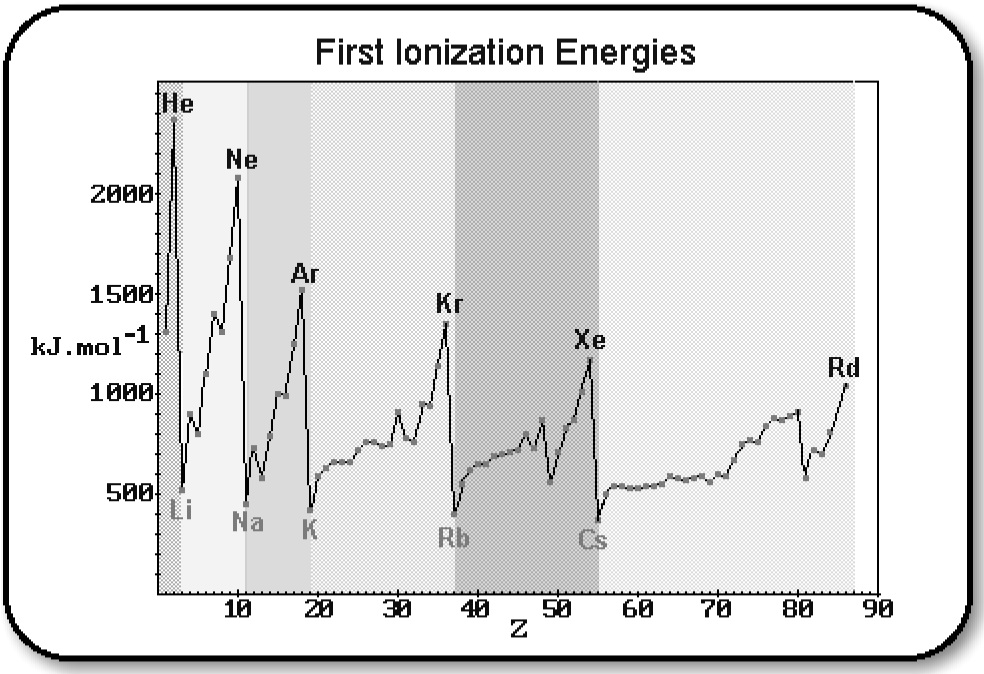
Eg Li $1s^2\ 2s^1$

Na $1s^2\ 2s^2\ 2p^6\ 3s^1$ atomic radius increases due to extra electron shell.

<u>Ionisation energy –</u> 1st ionisation energy is the energy needed to remove a mole of electrons from a mole of gaseous atoms from the outermost electron (valence) shell.

Example $Na_{(g)} \longrightarrow Na^{+}_{(g)} + e$

Ionisation energy **increases** from left ⟶ right across the periodic table.



The ionisation energy will increase from **left to right across** a period because the greater number of protons (higher nuclear charge) attract the orbiting electrons more strongly, thereby increasing the energy required to remove one of the electrons.

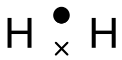
Going **down a group** on the periodic table, the ionisation energy will ***decrease*,** due to the **greater number of shells**, so the valence electrons **further from the protons, which attract them less strongly thereby requiring less energy to remove them**.

Also, outer electrons are **<u>shielded from the pull</u>** of the positive protons in the nucleus by the <u>electron shells that are closer to the nucleus</u>. This makes it easier to remove the outer electron (<u>less energy needed</u>).

Electronegativity - this is a measure of the attraction that an atom has for a pair of electrons in a chemical bond.

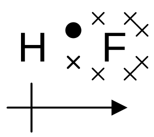
Example 2 hydrogen atoms have on average **the same attraction** for the bonding electron pair. The covalent bond is non-polar.

H - H



Example hydrogen and fluorine have **different electronegativity values**. Fluorine is more electronegative than hydrogen.

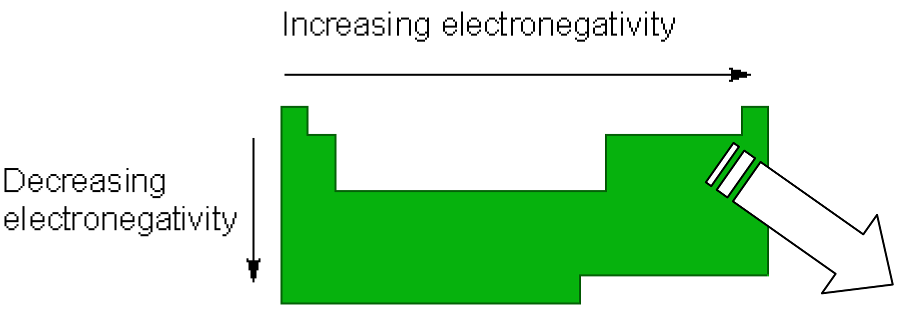
$\delta+$ $\delta-$
H - F

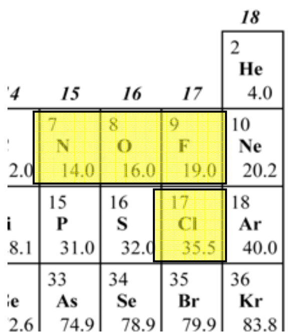


The **unequal pull** results in a polar covalent bond with the fluorine atom at the negative end of the dipole and the hydrogen at the positive end.

Electronegativity **increases passing from left to right along** a period, and decreases on descending a group. Hence, <u>fluorine</u> is the most electronegative of the elements. The nuclear charge increases, meaning electrons are pulled closer to the more electronegative atom.

Electronegativity **decreases as you go down a group**. Increasing the distance of the bonding electrons from the nucleus decreases the attraction. Also, the shielding effect reduces the pull of the nucleus on the bonding electron pair.





The four most electronegative atoms are F O N & Cl.