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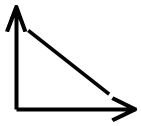
Quick Summary Series - HL/SL Chemistry - Exclusive for Chemyst Tuition Centre students only

Atomic Radius

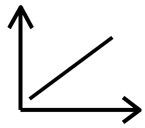
Across Period

Atomic Radius **decreases** as the **nuclear charge increases across the period.** The inner electron **shielding effect remains constant.** Distance from nucleus to valence electron decreases..



Down the group

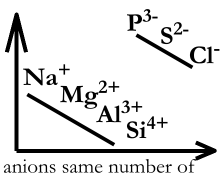
Atomic Radius **increases** as the **number of shells increases** leading to a **greater inner electron shielding effect.** Distance from nucleus to valence electron increases.



Ionic Radius

Across Period

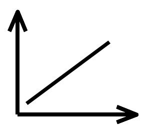
Cationic Radius **smaller** than atomic as cations have one less shell. The inner electron **shielding effect is lesser with same nuclear charge.**



Anionic Radius **bigger** than atomic as anions same number of shell but higher no. of electrons. **Repulsion between electrons** increases the distance from nucleus to valence electrons.

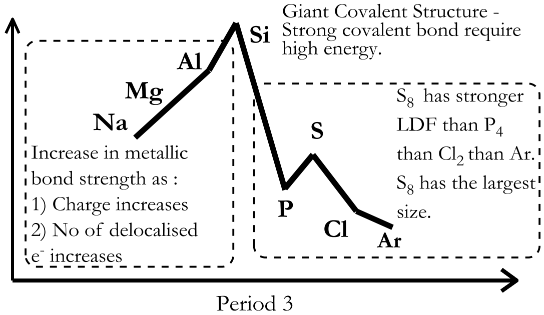
Down the group

Ionic Radius **increases** as the **number of shells increases** leading to a **greater inner electron shielding effect.** Distance from nucleus to valence electron increases.



Melting Point

Across Period

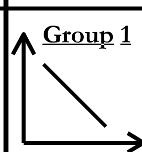


Giant Covalent Structure - Strong covalent bond require high energy.

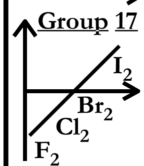
Increase in metallic bond strength as :
1) Charge increases
2) No of delocalised e^- increases

S_8 has stronger LDF than P_4 than Cl_2 than Ar. S_8 has the largest size.

Down the group



Metal ions get larger down the group, the nucleus becomes further from the delocalised e^- and the attraction becomes weaker. Less energy is required to break apart the lattice going **down group 1, melting point decreases.**

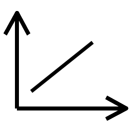


As the relative molecular masses of the X_2 halogen molecules increase,the London dispersion forces between molecules get stronger. More energy must be supplied to separate the molecules from each other. Going **down group 17, melting point increases.**

Electro-negativity

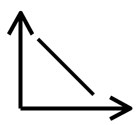
measure of the attraction of an atom in a molecule for the electron pair in the covalent bond of which it is a part.

Across Period



Electronegativity increases across a period. The increase in nuclear charge across the period with no significant change in inner electron shielding effect increases the attraction between nucleus and shared electron.

Down the group

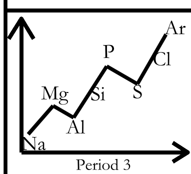


Electronegativity decreases down a group. The size of the atoms increases down a group thus nucleus of an atom is further away from the shared electron, hence attraction decreases.

First Ionisation Energy

$(X_{(g)} \rightarrow X_{(g)}^+ + e^-)$

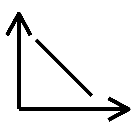
Across Period



Al : $1s^22s^22p^63s^23p^1$, an electron in a 3p orbital is further away from the nucleus than an electron in a 3s orbital.
S: $1s^22s^22p^63s^23p^4$, in the 3p orbital there is inter-electronic repulsion between the electrons which makes it easier to remove an electron.

IE(1st) generally **increases** as the nuclear charge increases leading to a greater attraction from the nucleus towards the valence electron. The inner electron shielding effect remains constant.

Down the group

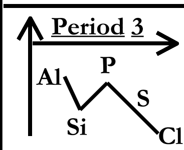


IE (1st) decreases as the number of shells increases leading to a greater inner electron shielding effect. Distance from nucleus to valence electron increases and it is easier to remove electron.

First Electron Affinity

$(X_{(g)} + e^- \rightarrow X_{(g)}^-)$

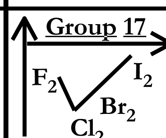
Across Period



The general trend is that the **electron affinity becomes more exothermic across period.** This is because of an increase in nuclear charge and a decrease in atomic radius from left to right across the period.

P less exothermic electron affinity than Si as P has three unpaired electrons in three separate p orbitals and when one electron is added this electron must be paired up in the same p orbital as another electron – this introduces an extra repulsion.

Down the group



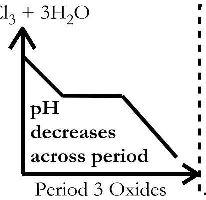
The general trend is that **electron affinity decreases down a group**, but it can be seen that chlorine has the most exothermic value for electron affinity. The electron affinity becomes less exothermic from Cl to I as the size of the atom increases

Electron–electron repulsion also affects the electron affinity and as the atom gets smaller the electrons are, on average, closer together and there is more electron–electron repulsion. (F_2 is an anomaly)

pH (period 3 oxides)

$Al_2O_3 + 2\,NaOH + 3H_2O \rightarrow 2\,NaAl(OH)_4$
$Al_2O_3 + 6\,HCl \rightarrow 2AlCl_3 + 3H_2O$

$SiO_2 + CaO \rightarrow CaSiO_3$

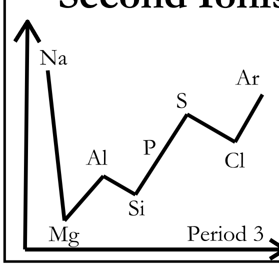


Basic
$Na_2O + H_2O \rightarrow 2\,NaOH$ pH = 13-14
$MgO + H_2O \rightarrow Mg(OH)_2$ pH = 12-13

Amphoteric $Al_2O_3 + H_2O \rightarrow$ No reaction, pH = 7

Acidic
$SiO_2 + H_2O \rightarrow$ No reaction, pH = 7
$P_4O_6 + 6H_2O \rightarrow 4H_3PO_3$, pH = 4-5
$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$, pH = 4-5
$SO_2 + H_2O \rightarrow H_2SO_3$, pH = 3-4
$SO_3 + H_2O \rightarrow H_2SO_4$, pH = 2-3
$Cl_2O_7 + H_2O \rightarrow 2HClO_4$, pH = 1-2

Second Ionisation Energy



Si^+ : $1s^22s^22p^63s^23p^1$, an electron in a 3p orbital is further away from the nucleus than an electron in a 3s orbital.
Cl^+: $1s^22s^22p^63s^23p^4$, in the 3p orbital there is inter-electronic repulsion between the electrons which makes it easier to remove an electron.