AS 91390: Demonstrate understanding of thermochemical principles and the properties of particles and substances

Explaining why atomic radii decrease (get smaller) going across a period

Across a period, the effective nuclear charge increases.

Both have the
- same number of electron shells
- same energy levels
- same shielding of outer e's by inner e's
- valence e's in same energy level

but ___ has a greater
- nuclear charge
- no of protons

so there is a stronger attraction for the valence electrons, bringing them in closer, resulting in a smaller radius.

Example: Cl < Na

Explaining why cations are smaller than the parent atoms they were formed from

_____ is smaller than _____ because it has lost electron(s) from an entire valence shell so the electrons are
- only in ____ shells instead of ___ shells
- in less shells

Example: Na^+ < Na

Explaining why anions are larger than the parent atoms they were formed from

_____ is larger than _____ because it has gained electron(s) leading to increased electron-electron repulsion increasing the size of the electron cloud so __ > __

Example: Cl^- > Cl

Explaining why atomic radii increase (get bigger) going down a group

_____ is larger than _____ because the added electron(s) go into a whole new valence shell & shielding is greater (so less attraction of electrons to the nucleus).

Example: K > Na

Explaining why molecules and ions have a particular shape

There are | 2 3 4 5 6 | • electron pairs* • regions of -ve charge | around the central atom which **repel** to take a | • linear • trigonal planar • tetrahedral • trigonal bipyramidal • octahedral | arrangement | • to minimise repulsion • to get as far apart as possible

* double/triple bond = one region of negative charge

There are ___ bonding electron pairs **and** ___ lone pairs which results in a | • linear • v-shaped • trigonal planar • trigonal pyramidal • tetrahedral • trigonal bipyramidal • see-saw / distorted tetrahedron • t- • octahedral • square pyramidal • square planar | shape | • molecule • ion

Explaining why a covalent molecule is polar or non-polar

The | _ - _ bond(s) is/are | • polar • non polar | because of EN difference** | & molecule is | • symmetrical • a symmetrical | so | • the dipoles cancel out / don't cancel out • centre of +ve and -ve charge coincide / don't coincide | & so the molecule is | • polar • non-polar

**(state __ is more electronegative than __ OR draw the $\delta+$ and $\delta-$ above a bond OR use $\longrightarrow$ to indicate)

Explaining why first ionisation energy increases across a period but decreases down a group

__ and __ have the same number of shells but __ has

- greater nuclear charge
- greater number of protons

so greater electrostatic attraction between nucleus and valence electrons

__ has the greatest number of electron shells, so its electrons are further from the nucleus, so less energy is required to remove a valence electron. It also has greater shielding effect of an additional shell between the valence shell and the nucleus, so less energy is required to remove the outermost electron.

E.g. Ca < Mg < Cl

Explaining why electronegativity increases across a period but decreases down a group

Valence electrons are

- added to same shell
- the same distance from nucleus

but in ___ there is a greater

- number of protons
- nuclear attraction

so val. electron(s) is/are more strongly attracted

But going down a group nuclear charge increases but so does shielding too, so the atomic radii increase causing less attraction of electrons to the nucleus.

E.g. why F > O > N and why Cl < F

Extension:

Explaining why TM can form compounds with variable oxidation states / can form a range of oxides and a group 1 or 2 metal like Ca cannot

__ only has __ valence electron(s) which it loses to form a __ + ion but ___ has valence electrons in 3d & 4s which it can lose to form stable ions or share covalently

E.g. Ca and Mn