Electrons in atoms

Q-1) What is electronic configuration?

> The arrangement of e^- in an atom is called electron configuration.

Electrons are found in energy levels / quantum shells (symbol n)

The lowest energy level n = 1 is closest to the nucleus. n = 2 is further away & so on...

Each quantum shell can hold a maximum no. of e^-.

shell 1 = max 2

shell 2 = max 8

shell 3 = max 18

shell 4 = max 32

Q-2) What is Ionisation energy? (ΔH_i / IE)

> Ionisation energy is the energy required to remove one electron from each atom (from the outer shell) in a gaseous state.

it's measured in $kJ\,mol^{-1}$

We can continue to remove e^-, until only the nucleus is left.

This is called successive ionisation energies.

eg: 1st IE

$$\left.\begin{array}{l} Ca(g) \rightarrow Ca^+(g) + e^- \\ \text{2nd IE} \\ Ca^+(g) \rightarrow Ca^{2+}(g) + e^- \end{array}\right\}$$

only ONE e^- can be removed at a time

Q-3) Factors that influence IE?

> The size of nuclear charge.

: as the no. of p^+ increases, the nuclear charge increases, so there is greater force of attraction between p^+ and e^- ∴ more energy is required to break these forces.

∴ as nuclear charge increases, IE increases.

> Distance of outer electrons from nucleus

: as distance of e^- away from nucleus increases, forces of attraction between p^+ and e^- decrease.

∴ the further away the e^- is from the nucleus, the lower the IE.

> Shielding effect of inner electrons

: full inner shell e^- prevent the full p^+ nuclear charge being felt by the outer shell e^-, so there are lower forces of attraction.

∴ the greater the shielding effect, the lower the IE.

(more no. of e^- between outer shell + nucleus)

Q-4) Ionisation energy graphs / results.

> Look for a big jump in IE; the element is in that group? (group just before the jump)

A big jump in IE shows that the next e^- is in the next energy level (closer to the nucleus).

Q-5) What are sub-shells?

> The quantum shells are split into sub-shells.

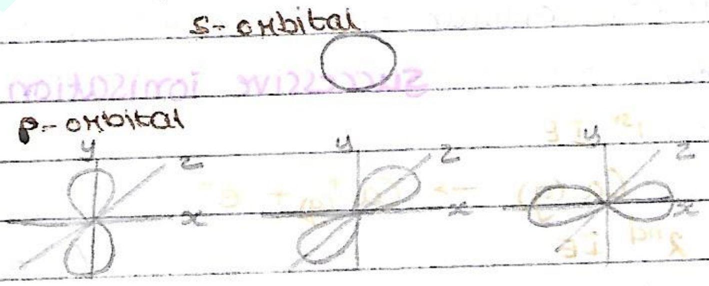
s < p < d < f.

s = $2e^-$

p = $6e^-$

d = $10e^-$

f = $14e^-$



Q-6) What are atomic orbitals?

> Atomic orbitals are region of space around the nucleus of an atom which can be occupied by a maximum of $2e^-$.

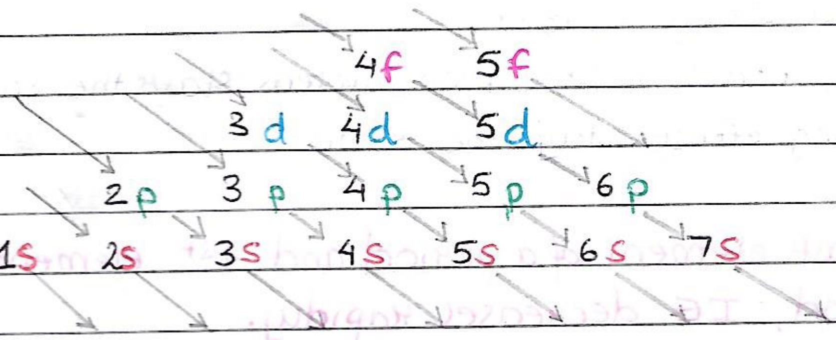
S = 1 orbital

p = 3 orbitals

d = 5 orbitals

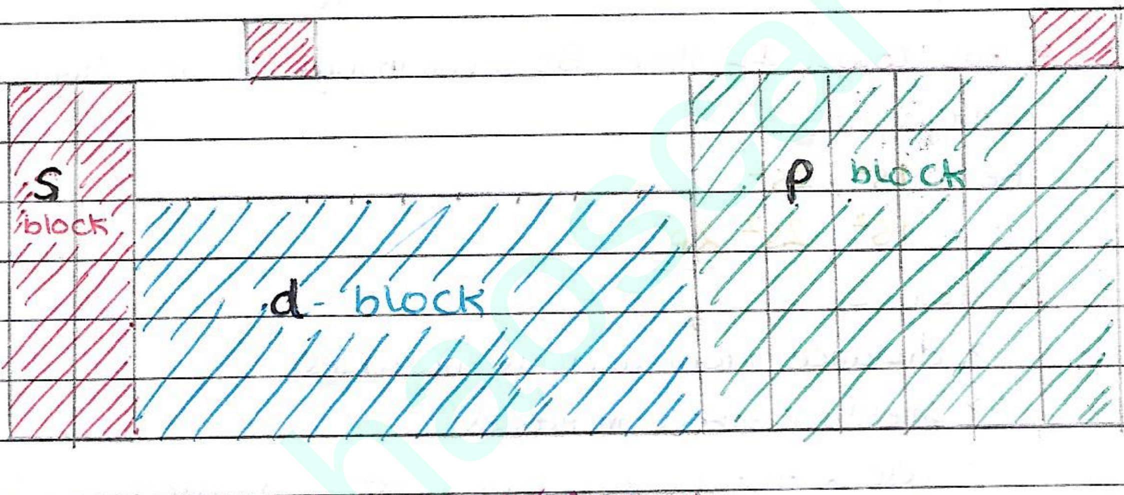
f = 7 orbitals.

Q-7) Filling of orbitals.



eg: $_{18}Ar = 1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6$

↳ ∴ Ar is in p-block, 3^{rd} period & 6^{th} grp in p (group 8).



* $_{24}Cr : 1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^{1}} \underline{3d^{5}}$
* $_{29}Cu : 1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^{1}} \underline{3d^{10}}$

Another way of representing e⁻ in orbitals is using boxes where each box can have upto 2e⁻.

eg: oxygen. $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$

$1s^2, 2s^2\ 2p^4$

Start filling each box with 1 e⁻ at a time in each orbital.

Q-8) Patterns in IE in the periodic table.

> Patterns accross the period.

Accross the period, IE increases.

> nuclear charge increases.
> same period ∴ distance from nucleus stays the same.
> shielding effect stays the same.

Between last element of a period and 1st element of the next period, IE decreases rapidly.

> distance from nucleus increases
> shielding effect increases.
> these 2 ~~effects~~ factors outweigh increase in nuclear charge.

B has lower IE than Be, even though it has higher nuclear charge.

Be = $1s^2, 2s^2$

B = $1s^2, 2s^2\ \underline{2p^1}$

∴ for B.

> distance from nucleus increases
> shielding effect increases
> these 2 factors outweigh increase in nuclear charge

O has lower IE than N, even though both e⁻ are removed from a 2p sub shell.

O = [↑↓] [↑↓] [↑↓ | ↑ | ↑]

N = [↑↓] [↑↓] [↑ | ↑ | ↑]

> The e⁻ removed from O has a pair of e⁻. This extra repultion between the e⁻ pair causes less energy needed to remove e⁻. It's called spin pair repultion.

Patterns down the group

Down the group, IE decreases

> distance away from nucleus increases
> shielding effect increases.
> these 2 factors outweigh the increase in nuclear charge.