2A – The electron

Evidence for shells

History of the atom

- The model of the atom has changed as our observations of its behaviour and properties have increased.
- A model is used to explain observations. The model changes to explain any new observations.
- The stages in the development of the atom:-

George Johnstone Stoney (1891) Electrolysis. The charge of an electron.

Joseph J Thompson (1897) The cathode ray tube and e/m deflection. The mass / charge of an electron.

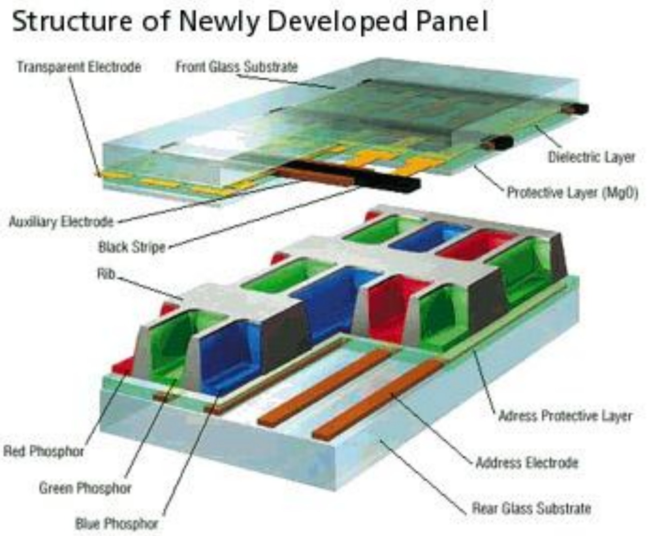
Robert Milikan (1909) Oil drop experiment. The mass / charge of an electron.

Joseph J Thompson 'Plum – Pudding' model of an atom.

Geiger, Rutherford and Marsden (1909) Alpha particle deflection. The nuclear model.

Henry Moseley (1913) Atomic number

Plasma displays



- A plasma unit has 100,000's of tiny cells (pixels) filled with a mixture of neon and xenon gasses.
- A single pixel is made up of three coloured sub-pixels, red, green and blue.
- Each sub pixel is driven by its own electrode, stimulates the gas to release ultraviolet light photons.
- The photons interact with a phosphor material coated on the inside wall of the cell.

- Phosphors are substances that give off light when they are exposed to other light eg. Ultraviolet light. The phosphors in a pixel give off coloured light when they are charged.
- The cells are situated in a grid like structure.
- The varying intensity of the current can create millions of different combinations of red, green and blue across the entire spectrum of colour.

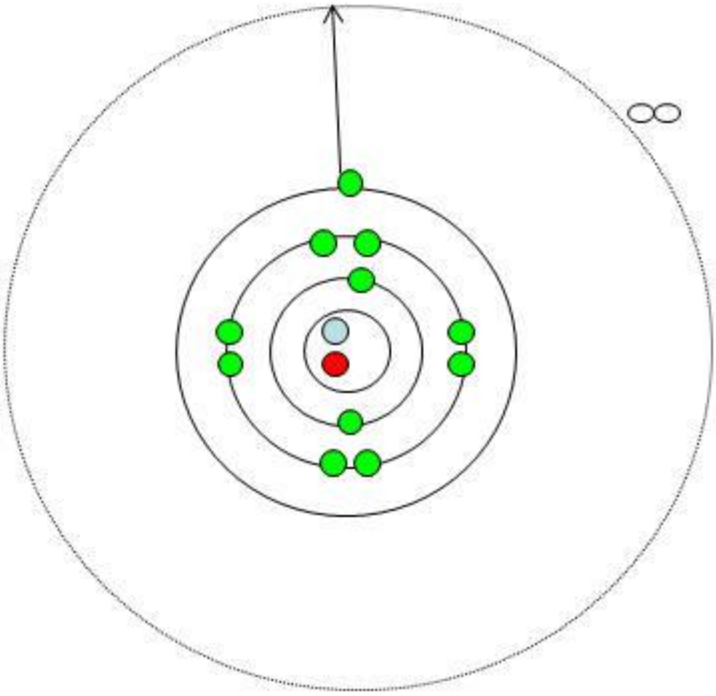
- How atoms are able to give off light comes down to understanding the movement of electrons in atoms and ions:

$$\mathbf{Ne_{(g)} \rightarrow Ne^{+}_{(g)} + e^{-}}$$

$$\mathbf{Xe_{(g)} \rightarrow Xe^{+}_{(g)} + e^{-}}$$

- When electrons are removed energy is required.

Ionisation energies:



- To form positive ions, electrons must be completely removed i.e. ionisation.
- To do this the electron must completely escape the attraction of the atom. i.e. reach n=∞.
- At n=∞ the electron has sufficient energy to escape the attraction from the nucleus.

Definition:

The 1st ionisation energy of an element is the energy required to remove 1 electrons from each atom in 1 mole of gaseous atoms to form 1 mole of gaseous 1+ ions

1st ionisation energy:

$Ca_{(g)} \rightarrow Ca^{+}_{(g)} + e^{-}$ **1st IE = +590 KjMol^{-1}**

2nd ionisation energy:

$Ca^{+}_{(g)} \rightarrow Ca^{2+}_{(g)} + e^{-}$ **2nd IE = +1150 KjMol^{-1}**

Factors affecting ionisation energy

1) The distance of the electron from the nucleus

- The further an electron is from the nucleus, the lower the force of attraction:

$$\mathbf{F \propto 1/d^2}$$

- This means that the electron will be easier to remove which means the ionisation energy will be lower.

2) Size of the positive nuclear charge

- The more protons in the nucleus, the higher the nuclear charge, the harder it is to remove an electron, the higher the ionisation energy.

3) The ‘shielding’ effect by full inner shells

- A full inner shell of electrons will repel electrons in outer shells.
- These ‘shields’ affect of the attraction from the nucleus on outer electrons.
- The more inner shells the greater the shielding.

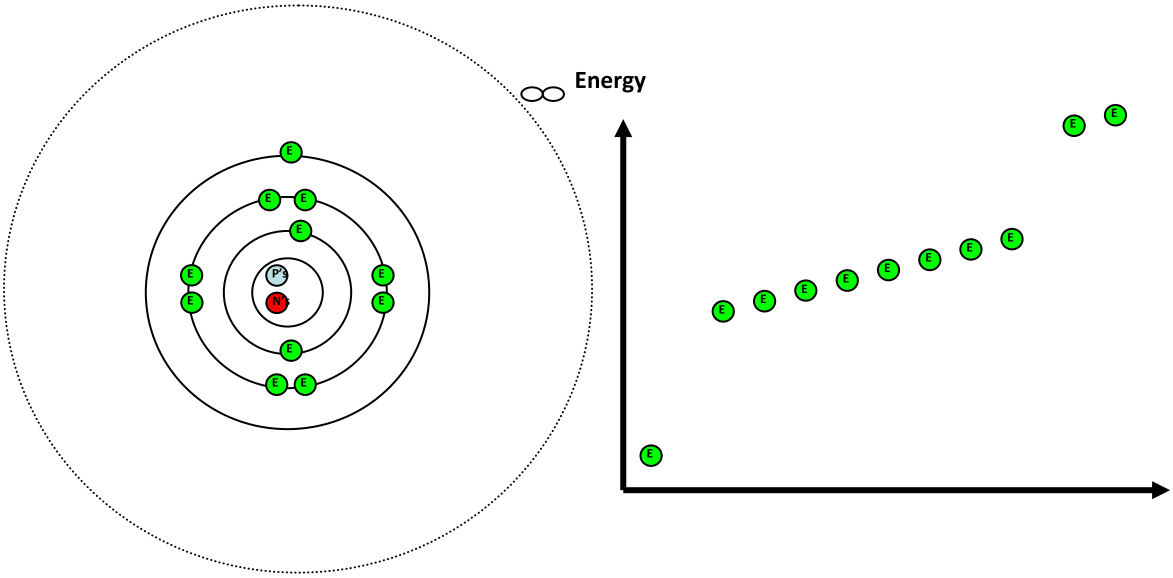
Ionisation energies down a Group:

- As you go down a Group, the outer electron shell is further from the nucleus - attraction decreases.
- The more inner shells the greater the shielding.
- As you go down Group 2, ionisation energy decreases.

Ionisation energies across a Period:

- As you go across a Period, electrons are removed from the same electron shell, shielding is the same.
- This means that the No of electrons : protons decreases.
- This increases the attraction pulling the electron shell in slightly.
- This increases attraction.
- This increases the successive ionisation energies

Successive ionisation energies



- The successive ionisation energies can tell us which group an element is in.
- For potassium the easiest electron(s) to remove before a large jump tells us the group.
- This can also be done by looking at data.

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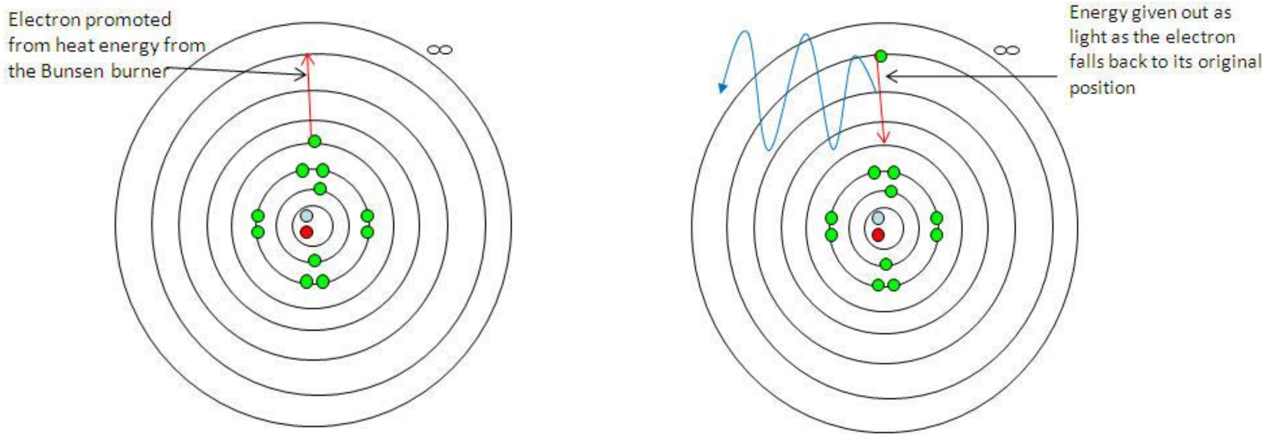
Shells and orbitals

Flame colours and emission spectra

Energy levels or shells -

- We have already seen what happens when an electron is removed from an atom - the atom becomes and ion.
- What if we supply enough energy to promote electrons to higher empty electron shells but not remove it?
- One way to look at the arrangement of electrons in an atom is to disturb them and see what happens when they go back to their original arrangement.
- This is done by heating compounds in a non-luminous Bunsen flame and studying the characteristic colours emitted when the electrons fall back to their original arrangement.
- The electrons gain energy from the Bunsen and lose energy (as light) as they go back to their original arrangement:



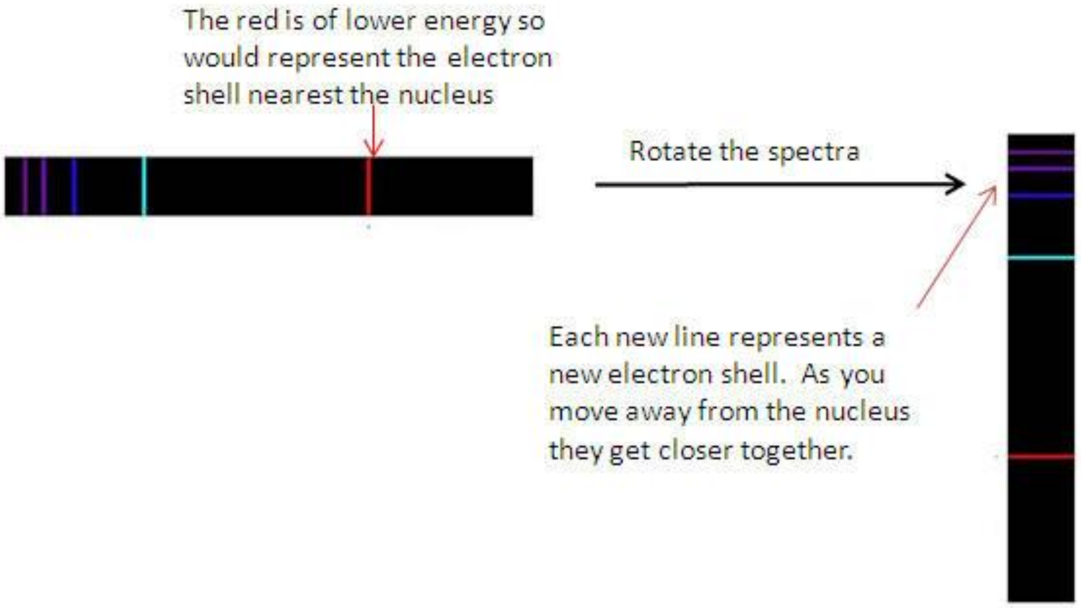
- This happens for any electron being promoted to any higher electron shell and falling back to any lower electron shell.
- This gives us the series of lines as electrons give out only specific energy when they move to a lower electron shell.
- A spectroscope shows us the specific energies characteristic to that element.

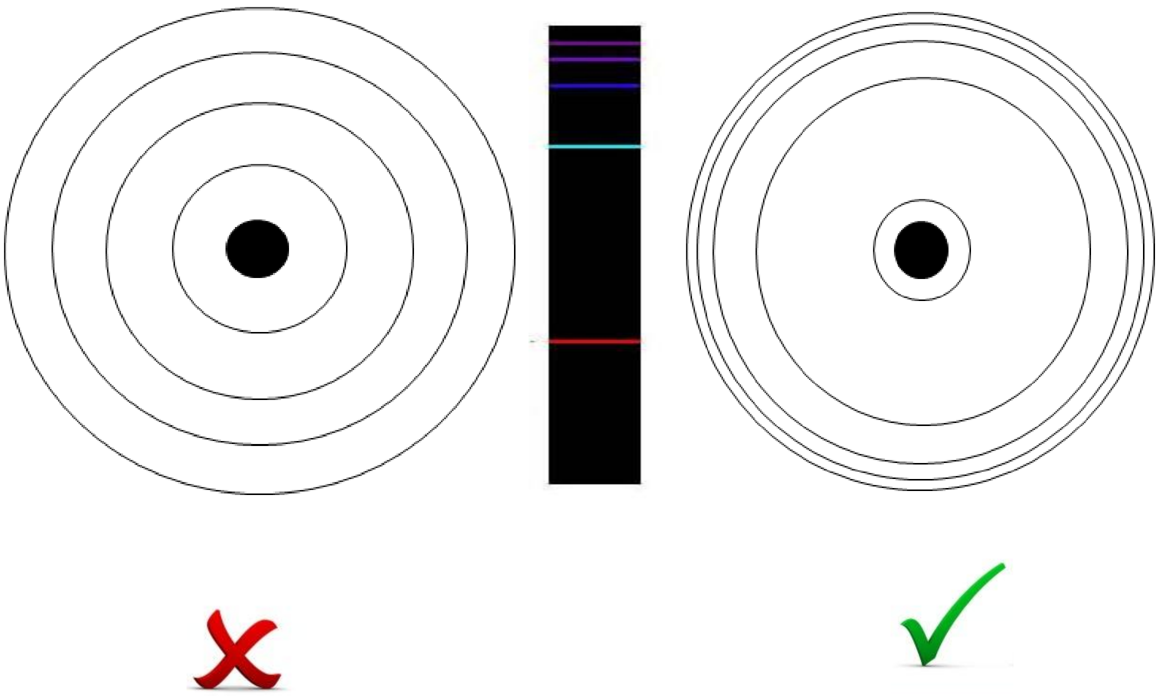
Looking at the line emission spectra 2 thing become apparent:

1) That the lines are specific colours representing specific energies / electron shells.

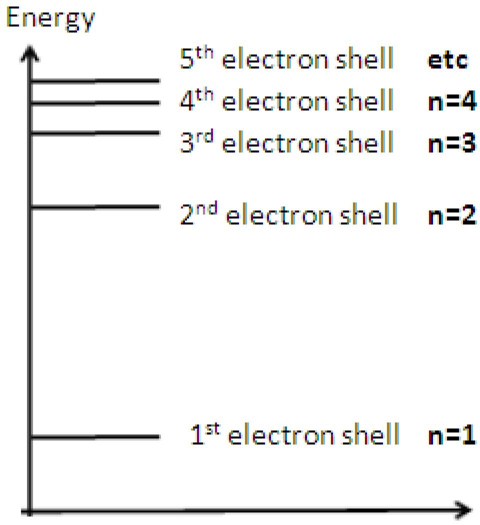
2) That as the move up in energy (towards the violet end) they get closer together (converge)

- This tells us that the electron shells are not nicely spread out but get closer together as they get further away from the nucleus:



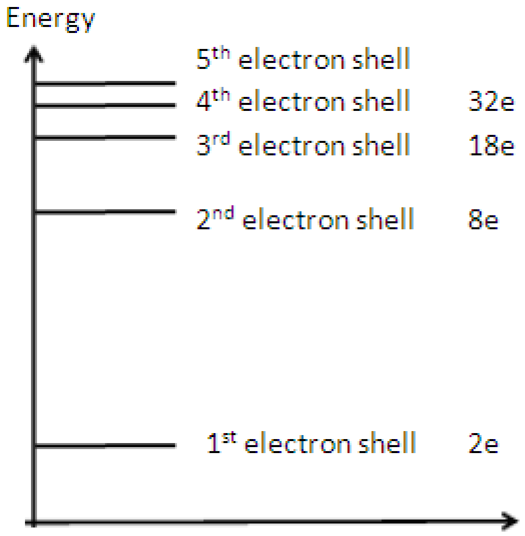


- That the electron shells can be converted into an energy level diagram showing the **Principle Quantum Number, n**



We call these diagrams energy level diagrams and each electron shell is now called '**The Principle Quantum Number, n**'

- The number of electrons in each **Principle Quantum Number, n can be calculated:**

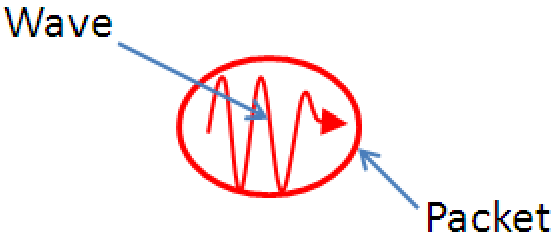


At GCSE you only filled the first 2 electron shells: 2.8. But if you keep filling they can hold more electrons (because they are bigger). We will look at this in more detail later.

To work out the number of electrons $= 2n^2$

Atomic orbitals

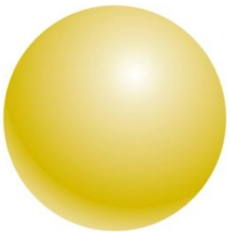
- At GCSE you will have been given a planetary model of the atom. This model assumes electrons are solid particles.
- Electrons are not solid particles, (more like a wave packet of energy or electron cloud):



- And so the planetary model has been replaced with the orbital model.
- These are regions around the nucleus in which the electrons are found.
- **Each electron shell is made up of orbitals.**
- **Each orbital can hold a maximum of 2 electrons.**
- Imagine each electron as an electron cloud with the shape of an orbital.
- Each electron in an orbital would be an electron cloud of the same shape.
- **2 electrons in an orbital would be the same shape but twice as dense.**

s - Orbitals:

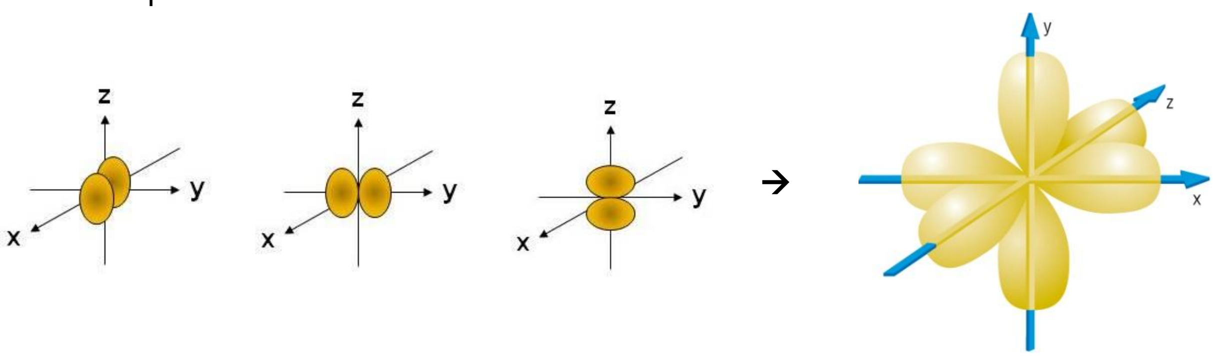
- These are spherical in shape.



- All electron shells contain an 's' orbital.
- As each orbital can hold a maximum of 2e the s - orbital has a total of 2e in every electron shell.

p - Orbitals:

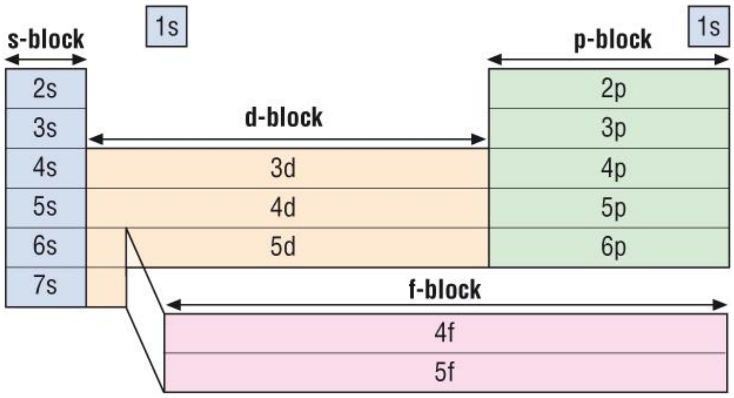
- Each p orbital is the shape of a dumb - bell.
- Each p orbital can hold a maximum of 2 electrons.



- As **there are 3 p orbitals**, this means that the p orbitals in a shell can hold a maximum of **6 electrons**.
- All electron shells from n = 2 upwards contain 3 p orbitals.

d and f orbitals:

- These orbitals are more complex. You do not need to know the shapes of these.
- Each shell from n = 3 upwards contains 5 d orbitals (5 x 2 = 10 electrons).
- Each shell from n = 4 upwards contains 7 f orbitals (7 x 2 = 14 electrons).



A summary of orbitals:

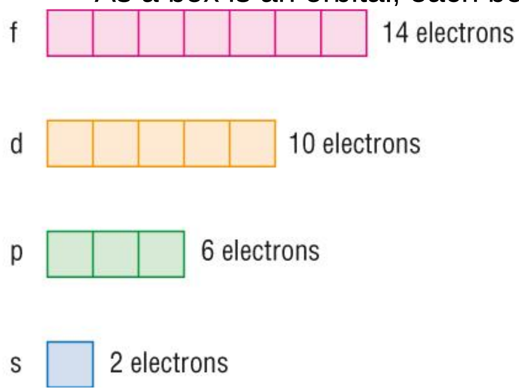
- **Each orbital can hold a maximum of 2e**

Complete the table below:

Electron shell	Principle Quantum Number	Types of orbital	No electrons in s shell	No electrons in p shell	No electrons in d shell	No electrons in f shell	Total
1							
2							
3							
4							

Representing electrons in orbitals

- With different types of orbitals and having different shapes, we represent an orbital with a box.
- As a box is an orbital, each box can hold 2e.



- If electrons are negative and repel each other how can they occupy the same orbital?
- This can only be explained by looking in more detail at the electrons.

- Electrons spin on their axis.

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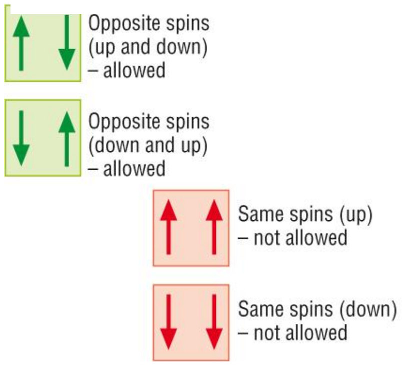
pin one way and the other spins in the opposite

hich we represent with an arrow.
re an orbital as there is now some attraction



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<u>Sub - shells and energy levels</u>

Sub - shells

- An electron shell is made up from orbitals.
- Orbitals of the same kind are called sub - shells.
- This means that each sub - shell contains the same type of orbitals, each of which holds a maximum of 2e:

<u>n = 1 shell: maximum 2 electrons</u>

Sub - shell	**1s**
Orbital	
Electrons	**2e**

<u>n = 2 shell: maximum 8 electrons</u>

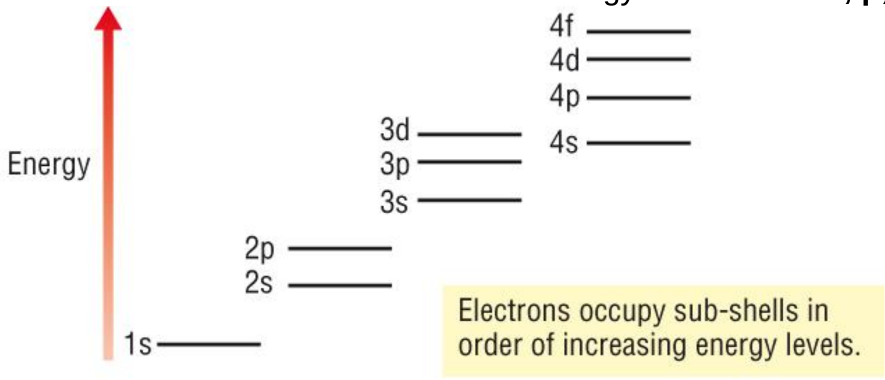
Sub - shell	**2s**		**2p**	
Orbital				
Electrons	**2e**	**2e**	**2e**	**2e**

<u>n = 3 shell: maximum 18 electrons</u>

Sub - shell	**3s**		**3p**				**3d**		
Orbital									
Electrons	**2e**	**2e**	**2e**	**2e**	**2e**	**2e**	**2e**	**2e**	**2e**

Electrons energy levels

- Within an electron shell, the sub - shells have slightly different energies.
- The sub - shell increases in energy in the order: **s, p, d and f**.



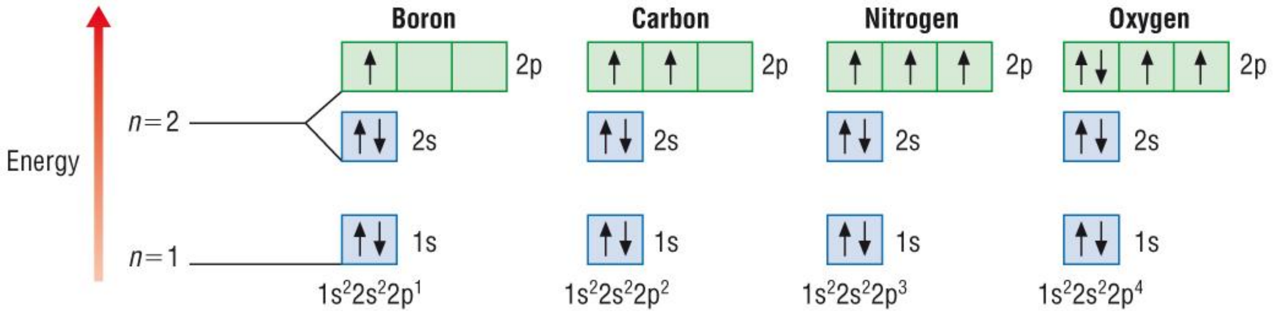
Filling shells and sub - shells:

- The **electron configuration** is the arrangement of electrons within an atom.
- This can be worked out by following a set of rules called **The Aufbau Principle**:

The Aufbau Principle:

1. **Electrons are added one at a time to 'build up' the atom.**
2. **The lowest available energy level fills first.**
3. **Each energy level must be full before the next, higher energy level can be filled.**
4. **Each orbital in a sub - shell is filled by single electrons before pairing up.**
5. **Each orbital can hold 2e of opposite spin**

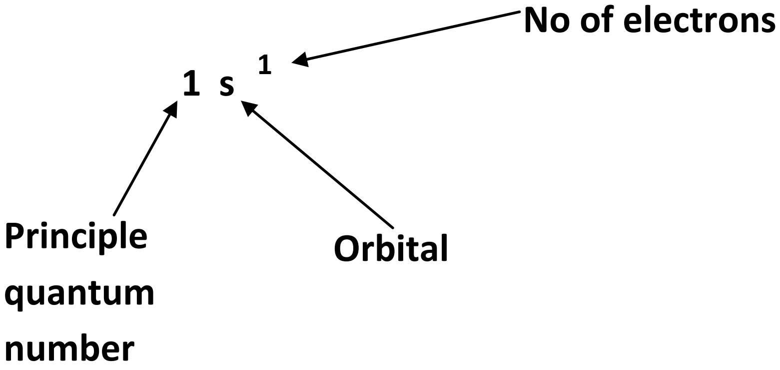
Filling the orbitals - Using the Aufbau Principle:



- The first 3 electrons in the p sub - shell spin in one direction and occupy p_x, p_y and p_z.
- After this the electrons have to pair up with those spinning the opposite way.
- In a sub - shell, electrons will remain unpaired in the orbitals until they have to pair up. This is the same in the d orbital

Electron configuration

- We use a shorthand to show how the electrons are arranged in an atom.
- Hydrogen is the simplest atom so we will use this to look at the shorthand:



Element	Orbitals occupied	Electron configuration
B	$1s^2 2s^2 2p_x^1$	$1s^2 2s^2 2p^1$
C	$1s^2 2s^2 2p_x^1 2p_y^1$	$1s^2 2s^2 2p^2$
N	$1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$	$1s^2 2s^2 2p^3$
O	$1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$	$1s^2 2s^2 2p^4$

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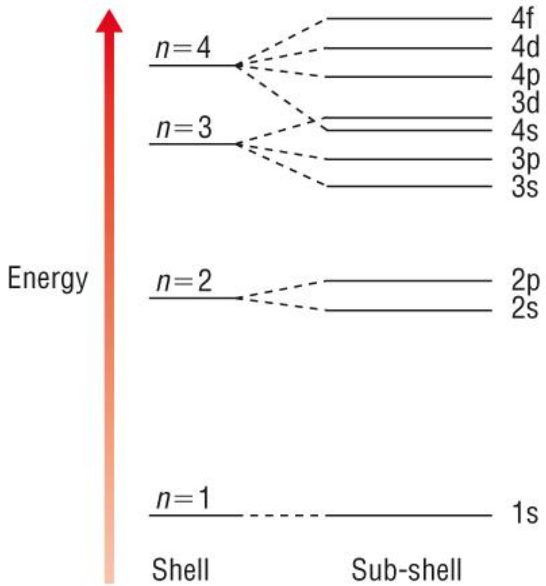
Electrons and the Periodic Table

Electron shells overlap

- From previous knowledge, we know that:

1. **The larger the value of n, the higher the energy level and the further from the nucleus.**
2. **As we move from n = 1 to n = 4, a new type of sub shell is added.**
3. **Within a shell the energies of the sub - shells increases in the order s, p, d and f.**

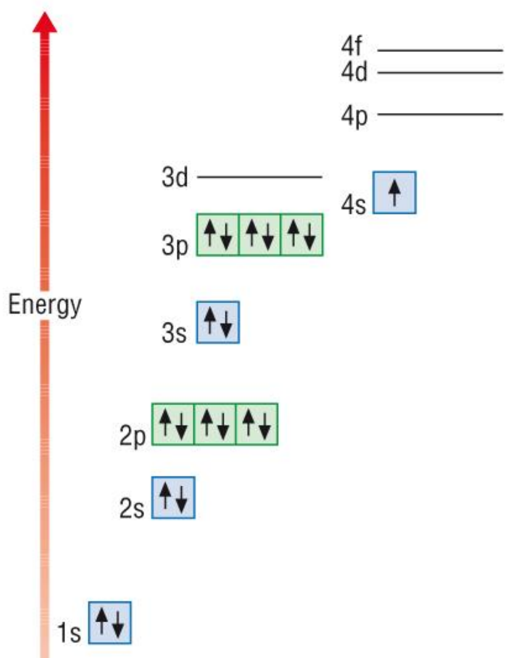
- After 3p it gets a bit complicated:



- **The 4s energy level is below the 3d energy level.**
- **This means that the 4s orbital fills before the 3d orbitals. (according to the Aufbau Principle)**

- Remember the orbitals fill up in energy level order from the bottom up:

Electronic configuration of potassium:
$1s^22s^22p^63s^23p^64s^1$



Sub shells and the Periodic Table:

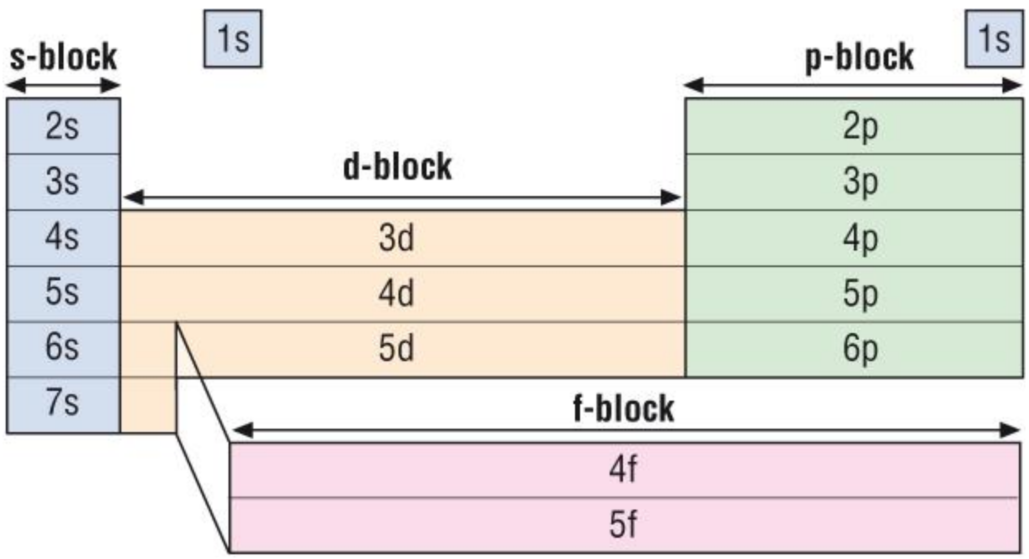
- **Fill in the following table:**

Element	No electrons	Electron configuration	What Period is this element in?	What is its highest Principle Quantum Number	What is the highest sub - shell	What Groups is the element in
H						
He						
Li						
Be						
B						
C						
N						
O						
F						
Ne						
Na						
Mg						
Al						

Things to notice:

- **The Period the element is in on the Periodic Table is equivalent to its highest Principle Quantum Number, n (and the electron shell of the outer electrons)**
- **All s orbital elements are in Groups 1 and 2 (except H and He)**
- **All p orbital elements are in groups 3,4,5,6,7 and 0**

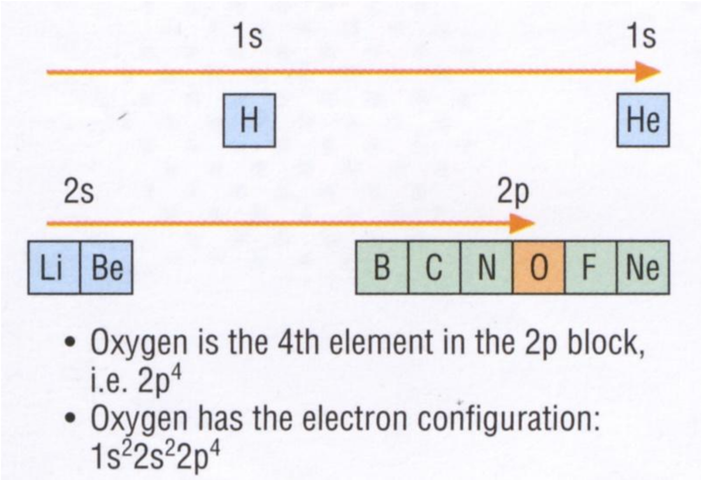
Extending further:



- **The Periodic Table is arranged in blocks of 2, 6, 10 and 14**
- **s block elements are all in Gp1 and 2 = 2**
- **p block elements are all in Gp3,4,5,6,7 and 0 = 6**
- **d block elements are all in the Transition Metals = 10**
- **f block elements are all in the Lanthanides and Actinides = 14**

Using the Periodic Table for electron configurations

- So, the electron configuration can be worked out from the Periodic Table, filling from left to right then top to bottom:



<u>Other examples:</u>

1) Na **$1s^2$ $2s^2$ $2p^6$ $3s^1$**

2) Sc **$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^1$ $4s^2$** (remember 4s fills before 3d)

Shortening an electron configuration

- For atoms with many electrons, the electron configuration can be long:

Sc $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^1$ $4s^2$

- We can write a shorthand version by using the closest noble gas configuration for the inner shells:

Sc [$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$]$3d^1$ $4s^2$

Inner shells

[Ar] has the same electron configuration as the inner shells

[Ar] $3d^1$ $4s^2$ where [Ar] represents the electronic configuration of argon.

Electronic configurations in ions

- Follow the same principle as for atoms but add / remove electrons:-

Eg Sodium ion Na^+

Na atom = $1s^2\ 2s^2\ 2p^6\ 3s^1$

To make a sodium ion 1 electron is removed

Na ion = $1s^2\ 2s^2\ 2p^6$

Eg Chlorine ion Cl^-

Cl atom = $1s^2\ 2s^2\ 2p^6\ 3s^2 2p^5$

To make a chlorine ion 1 electron is added

Cl^- ion = $1s^2\ 2s^2\ 2p^6\ 3s^2 2p^6$

Things to note:

- **All s and p block ions have a noble gas configuration.**
- **Electrons are taken from 4s before 3d (as it fills 3d before 4s)**

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