Electronic Configuration part 2

In part 1 we learned about the subshell order up to and including the 3rd shell. You can write it out like this (I haven't put in the 3d subshell at this stage just to keep it simple): **1s 2s 2p 3s 3p**

Taking it one step further, we can also add the electrons to this notation in the form of a superscript e.g. 1s^2 or 2p^4 etc. 2p^4 just means there are 4 electrons in the 2p subshell.

This allows us to pick any element and fill up the subshells with electrons using the above subshell order. The end result is called the **electronic configuration** or electronic arrangement or electronic structure.

In my opinion, this is one of the most useful things that you will learn in AS chemistry and can help with many different types of exam question such as ionisation energies.

Sodium example

To get the electronic configuration of sodium:

1. look at the periodic table to get the total number electrons: sodium has 11 electrons.
2. put two electrons into the 1s subshell as it is the closest to the nucleus and lowest in energy → **1s^2**.
3. similarly add electrons to the next subshells from left to right until there are no electrons left. Two more go into the 2s, then 6 go into the 3 x 2p orbitals, which takes care of 10 electrons. The final electron goes into the 3s orbital on its own to give the electronic configuration:

$$1s^2\ 2s^2\ 2p^6\ 3s^1$$
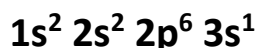
✓ When a shell, subshell or orbital can't hold any more electrons, we say that they are "full". I suppose that's common sense!

Larger elements

The subshell order above is ok for small elements that don't have many electrons. For larger elements, we need to introduce further orbitals.

It would be advisable to know the following subshell order up to 4p, which accommodates up to 36 electrons:

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p$$
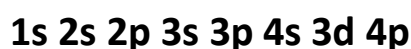
✓ The above order is a bit strange. The 4s orbital is at a **lower energy** than the 3d, which will make no sense. But you don't need to know why, so it is easier just accept it and make sure you get the order correct.

✓ You can write the order as 3d 4s if you prefer but it can cause problems. The 4s **must** be filled before the 3d. If you write 3d 4s you might forget to do this.

For example, bromine has 35 electrons. Using the same theory as above, we start with 1s and work our way up until we have used all the electrons:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$$
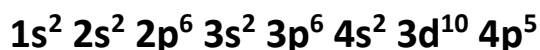
✓ One more subshell you might see is 'f'. It comes after d but you shouldn't really need to use it. But just incase, an f subshell has 7 f orbitals → 14 electrons.

Blocks

The periodic table is often referred to in "blocks". For example, groups 1 and 2 are known as the s-block. This refers to the subshell that the outer electrons are in e.g. sodium (yes, again!) **$1s^2\ 2s^2\ 2p^6\ 3s^1$.** The $3s^1$ is the outer electron and therefore it is in the "s-block".
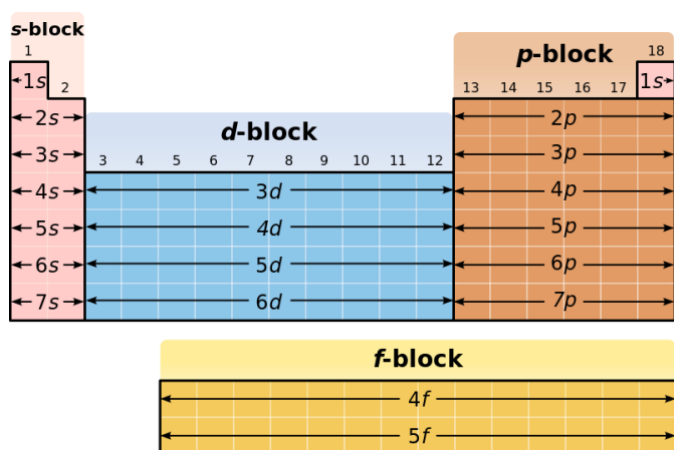
This pattern is seen throughout the periodic table. The only thing that changes is the shell number.

Filling Orbitals

As we mentioned way back at the start of part 1, the reason we introduced subshells and orbitals was to provide a bit more detail on where the electrons are.

Another thing we can do is to actually draw out the orbitals individually. This again provides us with more information, as we can see how the electrons are arranged in the **orbitals**. We can't see that from the electronic configuration.

One electron is normally represented by a ↑ or a ↑ symbol, you can use either.

When there are 2 electrons in an orbital they are written as: ↑↓ or ↑↓

When the electrons are drawn like this they are said to be **"paired"**. One points up and one points down, often called **"opposite spin"**. An electron on its own is called **"unpaired".**

Example: Nitrogen

Nitrogen has 7 electrons → an electronic configuration of **$1s^2\ 2s^2\ 2p^3$**.

✓ Always start with the electronic configuration as it tells you immediately which **subshells are full and which are not**. This will stop a lot of mistakes knowing this.

To draw the orbitals, you can use a box (as shown below) or simply a straight line. Remember that any s subshell has 1 orbital, any p subshell has 3 orbitals and any d subshell has 5 orbitals:
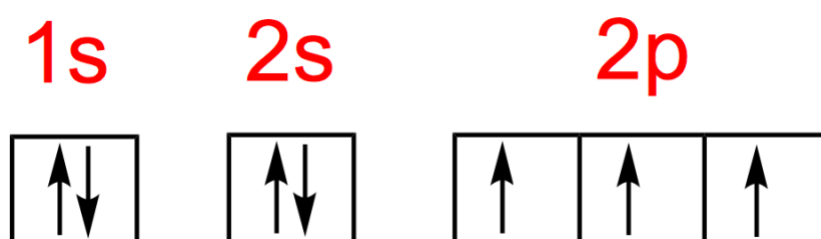
From the electronic configuration above, we know that the 1s and 2s subshells are full so we can put two "paired" electrons in each.

We also know that we have $2p^3$, which means that these orbitals are **not full** (we only have three electrons and the three p orbitals can hold six).

In this scenario we **don't** pair them up. Instead, we put **1 electron in each of the three p orbitals**. It is just more stable this way.

So the rule for orbitals in a subshell that is **not full**:

put one electron in each orbital and only then start to pair them up

Each orbital in a subshell must get one electron each before any orbital can have two electrons.

Example: Oxygen

Oxygen has 8 electrons → electronic configuration: **$1s^2\ 2s^2\ 2p^4$**

Oxygen has one more electron than nitrogen, which now goes into the first 2p orbital to give a pair of electrons. The diagram below shows that we have now got paired electrons in the first 2p orbital and 2 unpaired in the next two orbitals:
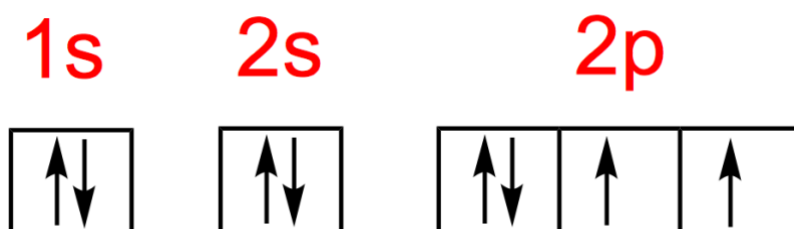
If we then went to fluorine we would have another electron in the second 2p orbital and then on to neon, we would fill all the 2p orbitals. Once we have filled the 2p orbitals we then need a new shell and we would add the 3s orbital.

✓ In the examples above I have drawn the orbitals from left to right (increasing in energy from left to right). You can also do it vertically i.e. the lowest energy orbitals at the bottom then the next one on top and so on.

Other Questions

I mentioned earlier that electronic configuration is very useful. This is because it can help you with questions on trends such as size, reactivity, ionisation energy and working out ionic formulas.

It's also useful when comparing elements. Questions often ask why one element does one thing in contrast to another. If you get stuck, write out the electronic configuration and look for differences between them. Any difference usually points you towards the answer.