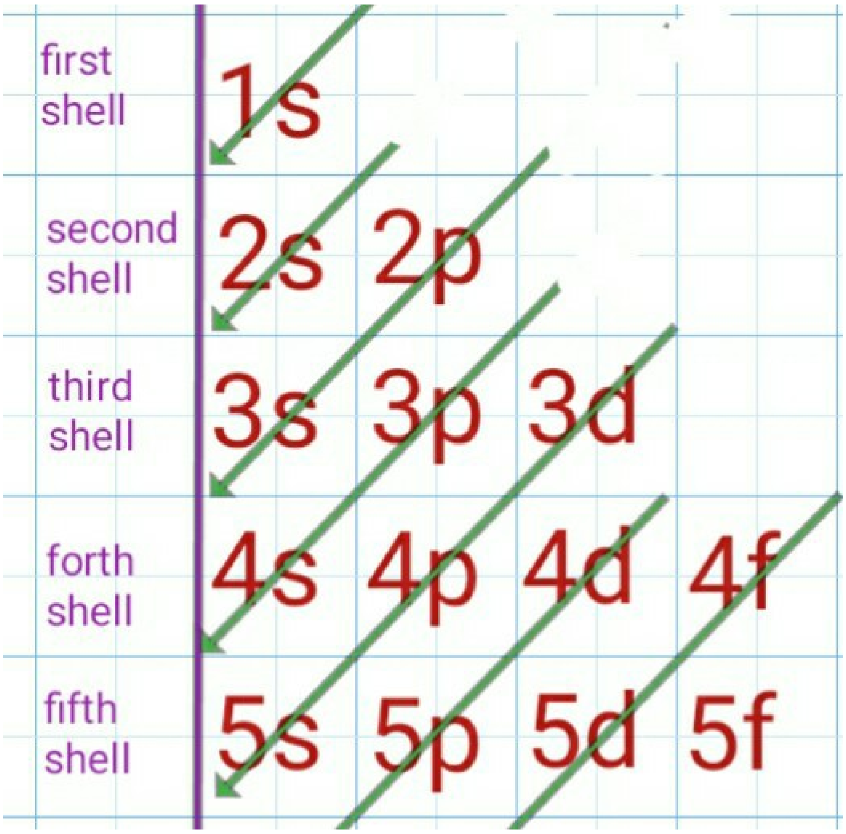

Electron Shells, Subshells and how they are filled (A level Chemistry)

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Electron Shells, Subshells and how they are filled (A level Chemistry)

Many chemistry students face problems on **electron shells, subshells and how they are filled.** However, the issue of filling shells and sub-shells is very straight forward except for some minor exceptions that we are going to highlight of course. For some students, the problem stem from the fact that they do not understand the properties of the shells themselves. That is where we are going to start.

What are electron shells

Electron shells are basically orbits followed by electrons as they circle around an atom's nucleus. They have different energy levels depending on how far away they are from the nucleus.

We can also define an electron shell as the set of electron states that share the same principal quantum number (n).

- The closest shell to the nucleus is the first shell" (also known as the "K shell") and has principal quantum number n = 1,
- followed by the "second shell" (also known as the "L shell"), n 2,
- then the "third shell" (also known as the "M shell"), n = 3, and so on.

Each shell can hold only a certain maximum number of electrons as follows:

- first shell, up to two electrons
- second shell, up to eight (2 + 6) electrons
- third shell, up to 18 (2 + 6 + 10) electrons and so on.

The maximum number of electrons a shell can hold is calculated by the following general formula:

`maximum number of electrons for an nth shell = 2n2`

This means that:

- first shell maximum = $2(1)^2 = 2$
- second shell maximum = $2(2)^2 = 8$
- third shell maximum = $2(3)^2 = 18$
- fourth shell maximum = $2(4)^2 = 32$

What are sub-shells

Sub-shells are groupings of atomic orbitals within an electron shell. Each electron shell consists of one or more subshells. The subshells themselves consists of atomic orbitals.

The table below shows some shells and the sub-shells they contain.

Electron Shell	Sub-shells
First shell (K shell)	1s
Second shell (L shell)	2s 2p
Third shell (M shell)	3s 3p 3d
Forth shell (N shell)	4s 4p 4d 4f

The first (K) shell has one subshell (1s); the second (L) shell has two subshells (2s and 2p) the third shell has three subshells (3s, 3p and 3d), and so on.

We can also define a subshell as the set of electron states with a
common azimuthal quantum number, ℓ, within a shell. Sub-shells are given names s, p, d and f
corresponding to ℓ values of 0, 1, 2, 3
respectively. For example, the 3d sub-shell has principal quantum number, n = 3 and azimuthal quantum number $\ell = 2$.

Each subshell subshell holds a maximum of $4\ell + 2$ electrons.

For example:

- s subshell maximum electrons = $4(0) + 2 = 2$

- p subshell maximum electrons = 4(1) + 2 = 6
- d subshell maximum electrons = 4(2) + 2 = 10
- f subshell maximum electrons = 4(3) + 2 = 14

And all that is summarised in the table below:

Subshell	azimuthal quantum number (ℓ)	Max electrons
s	0	2
p	1	6
d	2	10
f	3	14

What are orbitals

Orbitals are regions in space that can hold a certain number of electrons. Different sub-shells contain different numbers of orbitals, with each orbital containing a maximum of two electrons.

The table below shows the number of orbitals each subshell has:

Subshell	Number of orbitals
s	1
p	3
d	5
f	7

Since each orbital can hold a maximum of two electrons, the maximum number of electrons a sub-shell can hold is twice the number of the orbitals.

The sequence of filling sub-shells with electrons

The table below shows the (ℓ) value each subshell.

Subshell	azimuthal quantum number (ℓ)
s	0
p	1
d	2
f	3

The order in which the sub-shells are filled is deduced from the (n + ℓ) whereby n is the shell energy level.

According to the Madelung rule, sub-shells with a lower (n + ℓ) value are filled before those with higher (n + ℓ) values. In the case of equal (n + ℓ) values, the sub-shell with a lower n value is filled first.

Let us look at some of the (n + ℓ) of some of the sub-shells in the table below.

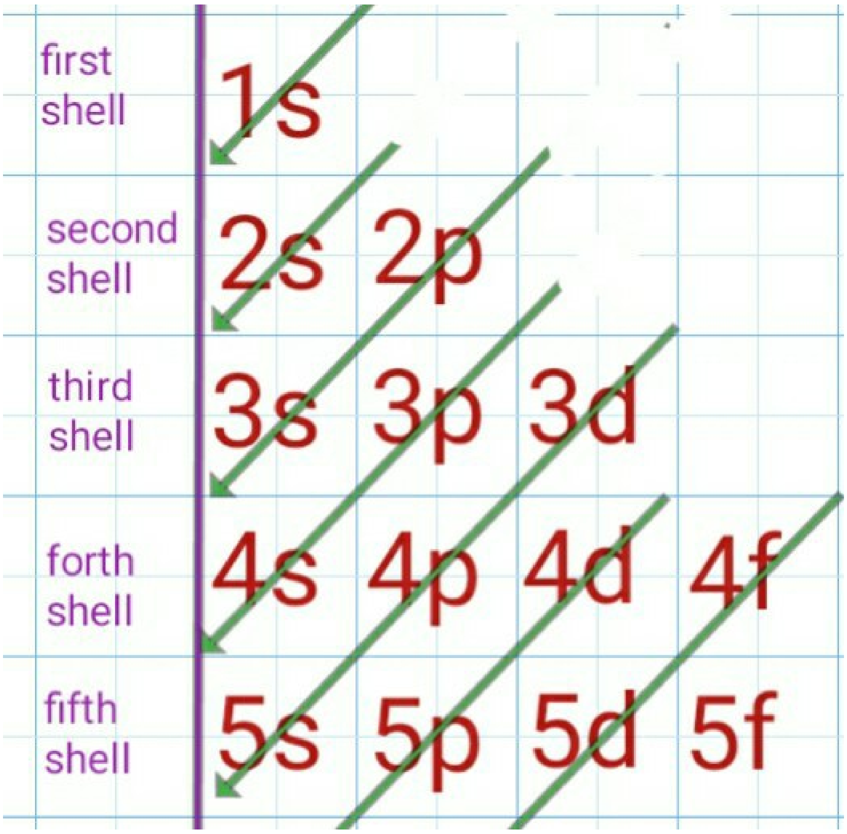
Sub-shell	n	ℓ	(n + ℓ) value
1s	1	0	1
2s	2	0	2
2p	2	1	3
3s	3	0	3
3p	3	1	4
3d	3	2	5
4s	4	0	4
4p	4	1	5
4d	4	2	6
4f	4	3	7
5s	5	0	5
5p	5	1	6

Sub-shell	n	ℓ	(n + ℓ) value
5d	5	2	7

Putting the sub-shells in order of increasing (n + ℓ) values we get the order of filling the sub-shells, ie:

- 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p…

A good way to visualise the order in which sub-shells are filled in is to create a digram as shown below and then cross out sub-shells with the same (n + ℓ) values. Basically that can be achieved by crossing out the sub-shells diagonally.



How to write electron configuration

The electron configuration is the arrangement of electrons in the orbitals of an atom, molecule or ion. Armed with knowledge of filling sub-shells from the previous section, we can use a standard notation to indicate the electron configurations of atoms and molecules.

Basically write electronic configurations following the sequence: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p…, moving to the next sub-shell when one sub-shell is full. The superscripts on each sub-shell shows the number of electrons in the respective sub-shell.

Element	Number of electrons	Electronic Configuration (showing sub-shells)
Hydrogen	1	$1s^1$
Helium	2	$1s^2$
Lithium	3	$1s^2\ 2s^1$
Beryllium	4	$1s^2\ 2s^2$
Boron	5	$1s^2\ 2s^2\ 2p^1$
Carbon	6	$1s^2\ 2s^2\ 2p^2$
Nitrogen	7	$1s^2\ 2s^2\ 2p^3$
Oxygen	8	$1s^2\ 2s^2\ 2p^4$
Fluorine	9	$1s^2\ 2s^2\ 2p^5$
Neon	10	$1s^2\ 2s^2\ 2p^6$
Sodium	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon	14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Phosphorus	15	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Sulphur	16	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Chlorine	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Argon	18	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Potassium	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Calcium	20	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

One thing you might have noticed with this type of notation is that as the number of electrons increase, the electronic configuration becomes quite lengthy. To solve that, we use shorthand notations based on the configuration of a noble gas immediately before the selected element.

For example, the electron configuration of calcium is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ and the noble gas immediately before it is argon which has a configuration $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$. So we replace the $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ in calcium with [Ar], the symbol of argon.

The electron configuration of calcium therefore becomes [Ar] $4s^2$.

After calcium comes the transition metals. The next 10 transition elements are called d-block elements because they have a partially filled 3d sub-shell. Almost all of them have a full 4s sub-shell except for copper and chromium. The following table shows the electron configurations of d-block elements using the shorthand notation.

Element	Number of electrons	Electronic Configuration
Sc	21	[Ar] $4s^2$ $3d^1$
Ti	22	[Ar] $4s^2$ $3d^2$
V	23	[Ar] $4s^2$ $3d^3$
Cr	24	[Ar] $4s^1$ $3d^5$
Mn	25	[Ar] $4s^2$ $3d^5$
Fe	26	[Ar] $4s^2$ $3d^6$
Co	27	[Ar] $4s^2$ $3d^7$
Ni	28	[Ar] $4s^2$ $3d^8$
Cu	29	[Ar] $4s^1$ $3d^{10}$
Zn	30	[Ar] $4s^2$ $3d^{10}$

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