

## 7.12: Electron Configurations

The several electrons that surround an atomic nucleus will have various orbital angular momenta - that is, each electron will have a certain value of the orbital angular momentum quantum number  $l$ .

An electron with  $l = 0$  is called an  $s$  electron.

An electron with  $l = 1$  is called a  $p$  electron.

An electron with  $l = 2$  is called a  $d$  electron.

An electron with  $l = 3$  is called an  $f$  electron.

This curious and admittedly illogical notation derives from the early study of the spectra of the alkali (e.g. Na, K) and alkali earth (e.g. Mg, Ca) elements in which four series of lines were noted, which, at the time, were called the "Sharp", "Principal", "Diffuse" and "Fundamental" series. Only later, when atomic structure was better understood, were these series associated with what we now know to be electrons with  $l = 0, 1, 2, 3$ . (San Francisco Police Department, SFPD, is NOT a good mnemonic to use in trying to remember them.) After  $l = 2$ , the letters go  $g, h, i, k, \dots$  etc.,  $j$  being omitted.

An electron with, for example,  $l = 1$ , is often described as being "in a  $p$ -orbital". Bear in mind, however, the meaning of the shapes described by the wavefunctions as discussed in section 9.

Electrons with  $n = 1, 2, 3, 4$ , etc., are said to be in the " $K, L, M, N$ , etc." shells. This almost as curious and not very logical notation derives from the early study of x-ray spectra, in which various observed absorption edges or groups of emission lines were labelled  $K, L, M, N$ , etc., and, with subsequent knowledge, these have since been associated with electrons with the principal quantum number being 1, 2, 3, 4. Presumably, since it wasn't initially known which x-ray absorption edge would ultimately prove to be the "first" one, it made good sense to start the notation somewhere near the middle of the alphabet.

The restrictions on the values of the quantum numbers together with the Pauli exclusion principle enable us to understand the *electron configurations* of the atoms. For example the electron configuration of copper, Cu, in its ground state is

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s \quad (7.12.1)$$

This is usually pronounced, including by myself, "one-s-squared, two-s-squared, two-p-to-the-sixth..." etc, but we shall soon see that this is certainly not how it *ought* to be pronounced, and I shall not discourage the reader who wants to do it properly, while I continue with my slovenly ways.

What it means is as follows:

The 1 refers to electrons with  $n = 1$ ; that is to  $K$ -shell electrons. The notation  $s^2$  that follows indicates that there are two  $s$ -electrons; that is, two electrons with zero orbital angular momentum. (In a Bohr-Sommerfeld model presumably they'd either have to be motionless, or else move two and fro in a straight line through the nucleus! We don't have that difficulty in a wave-mechanical model.) Now recall that  $l$  can have integral values only up to  $n - 1$ , so that the only electrons possible in the  $K$ -shell are  $s$ -electrons with  $l = 0$ . Consequently the only possible value of  $m_l$  is zero. There are the two possible values for  $m_s$ , however, namely  $+1/2$  and  $-1/2$  so two  $s$ -electrons (but no more than two) are possible in the  $K$ -shell. Thus the  $K$ -shell is *full*, and the quantum numbers for the two  $K$ -shell electrons are:

$n$	$l$	$m_l$	$m_s$
1	0	0	$+1/2$
1	0	0	$-1/2$

(7.12.2)

Next we come to  $2s^2 2p^6$ . This indicates that there are two  $s$ -electrons in the  $L$ -shell and six  $p$ -electrons in the  $L$ -shell. Let's look at their quantum numbers:

$$\begin{array}{cccc}
 n & l & m_l & m_s \\
 \\ 
 2 & 0 & 0 & +1/2 \\
 2 & 0 & 0 & -1/2 \\
 2 & 1 & -1 & +1/2 \\
 2 & 1 & -1 & -1/2 \\
 2 & 1 & 0 & +1/2 \\
 2 & 1 & 0 & -1/2 \\
 2 & 1 & 1 & +1/2 \\
 2 & 1 & 1 & -1/2
 \end{array} \tag{7.12.3}$$

And so on, and so on and so on... Many physics and chemistry books give the configurations of the ground states of all the atoms in the periodic table, and such tables are well worth careful study. Recall that, for a given  $n$ ,  $l$  can take values only up to  $n - 1$ , so that, for example, the  $N$ -shell ( $n = 3$ ) can have only  $s$ ,  $p$ , and  $d$  electrons. Within a given shell, there can be only two  $s$ -electrons, six  $p$ -electrons, ten  $d$ -electrons,  $2(2l + 1)$  electrons with orbital angular momentum quantum number  $l$ . A given shell can hold only  $2n^2$  electrons each having its unique set of four quantum numbers. You will observe that copper, in its ground configuration, has full  $K$ ,  $L$ , and  $M$  shells, plus one outer electron in its  $N$  shell.

As said at the beginning of this chapter, we expect only to introduce some of the words and ideas encountered in spectroscopy. Careful study of more detailed textbooks will be necessary, and I strongly recommend trying yourself to build up the ground configurations of at least the first 30 elements, up to zinc. Compare your efforts with tables in the books, and you will find that the ground configurations of a very few of the first thirty elements may not be exactly what you predicted.

---

This page titled [7.12: Electron Configurations](#) is shared under a [CC BY-NC 4.0](#) license and was authored, remixed, and/or curated by [Jeremy Tatum](#) via [source content](#) that was edited to the style and standards of the LibreTexts platform.