

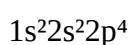
Electron configurations

1s 2p 3d and on and on. If these terms are confusing you, you're in the right place. If these terms don't even look familiar, you're either a chemistry student who needs a lot of help or are on the wrong website. Either way, if you keep reading, you'll be a pro at electron configurations by the end of this tutorial.

What's an electron configuration?

An electron configuration is a list that shows you where all of the electrons in an atom are located. That's it. It just tells you where the electrons are.

Let's see how this works out using the electron configuration of oxygen:



Believe it or not, this actually gives us some information about where electrons are. If you'll note, there are three terms in this configuration. Each has a number in front of it, a letter, and a superscript. For example, the first term is $1s^2$. The first number (a "1" in this case) tells you the energy level of the orbital the electron is in, the letter ("s") tells you the type of orbital the electron is in, and the superscript (a "2" in our example) tells us how many electrons are in that type of orbital. In other words, the entire electron configuration of oxygen literally says:

"There are two electrons in the 1s orbital, two electrons in the 2s orbital, and four electrons in the 2p orbitals"

Different electron configurations will have different numbers and letters and such, but they'll all follow the same general pattern.

Why do we care about electron configurations?

You probably *don't* care about electron configurations right now, and there's no particular reason you should. After all, you've just learned about orbitals and have no particular need for this information. Who *cares* if there are four electrons in the 2p orbitals of oxygen in the ground state? Not you!



The average student contemplating electrons.

This is one of those topics in chemistry that isn't immediately useful to know. It's like most other things you learn: If you don't use it for something, the information doesn't have any value.

Fortunately, you *will* be using electron configurations as you move on in chemistry. Unfortunately, you've got to learn electron configurations before you can use them. That's a bummer, and everybody who has studied chemistry knows how you feel.

To answer the question above in more detail, here's why we care about electron configurations:

- **It helps us to understand the reactivity of various elements.** It turns out that electron configurations give us information such as how many bonds something will want to make, as well as how many electrons something wants. That's handy, given that chemical reactions are really just quests for electrons.
- **It helps us to understand the weirdnesses of the periodic table.** Why do sodium and potassium have so many properties in common? Their electron configurations! Why do the transition metals all get kind of weird when it comes to their reactivity? Electron configurations!



Though it explains many weird things, it won't explain why this keeps turning up in your sock drawer.

- **It allows us to track where the electrons in a chemical reaction go.** Are electrons going from ground states to excited states? It's good to know what electrons we're talking about as well as what states they're moving from and to.

So, to answer your question, there's no reason you should care about electron configurations. Yet.

How to write electron configurations

Rather than having me talk about this at great length, it's probably better that I just do a bunch of examples and explain what I'm doing as I follow along. This will seem frustrating at first, but at some point you'll just understand it and be done with the whole thing. Trust me.

Before we get started, let's get a periodic table to make this easier. Oh, look! Here's one!

The periodic table is color-coded by groups:

- s-block (light blue):** Groups 1 and 2.
- p-block (light green):** Groups 13 through 18.
- d-block (light yellow):** Groups 3 through 10.
- f-block (light red):** Groups 1 through 14.

Large letters 's', 'p', 'd', and 'f' are placed near their respective blocks to indicate the type of orbital being filled.

Now, you may recognize this as being different than your normal periodic table. The main idea is the same, but you'll notice that various sections have been colored and labeled. These refer to the types of electrons that are present in the outer electrons of each element. For example, beryllium has s-electrons at the outside, while chlorine has p-electrons on the outside. I've also put numbers at the tops of each of these sections, for reasons we'll see in a moment.

Let's find hydrogen in this periodic table. I'll give you a hint: It's the first element.



Another example: Helium

Another example: Lithium

However, before we do that, we're going to have to consider something called the *aufbau principle*, which states the following:¹

What this means is the following:

The first two electrons in lithium are in the same place as the ones in helium. As a result, we start off with the term $1s^2$ to describe their location. You can also see that if you count from the top of the periodic table to lithium, you have to count two spaces over in the 1s section to get to lithium.

The third electron in lithium is in a new location – the 2s orbital, to be exact. As a result, the second term for this electron configuration is $2s^1$. Overall, this gives us an electron configuration of:



Another example: Beryllium

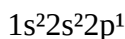
To get to beryllium, you first need to count across the 1s orbitals, and then down through the 2s orbitals. If you count where you're going, you'll find that you count two electrons in the 1s orbitals ($1s^2$) and two electrons in the 2s orbitals ($2s^2$). Put 'em together and you get:



Another example: Boron

Let's go through the periodic table again and start counting. To get to boron, you need to first go through the 1s and 2s orbitals in the same way as you did for beryllium, which means that your first two terms are $1s^2 2s^2$.

The next electron is in the p-section of the periodic table, which means that its term will have a p in it somewhere. The energy level of this electron is "2" (see the left side of the table) and there's only one electron of its kind. As a result, the term that describes it completely is $2p^1$. Put them together and get a full electron configuration of:



Another example: Carbon

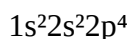
If you go through and do the same thing as above, you'll get the electron configuration $1s^2 2s^2 2p^2$.

Another example: Nitrogen

If you go through the same steps as above, you'll get the electron configuration $1s^2 2s^2 2p^3$.

By now, you've probably noticed that there's a lot of repetition between electron configurations. This is basically the same thing that the aufbau principle mentions above, except it's phrased all sciency by calling it "aufbau" instead of "adding up."

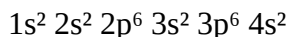
Do me a quick favor and, without even looking at a periodic table, guess what the electron configuration of oxygen (the next element) will be. If you've been paying attention to the electron configurations so far, you'll probably guess that it's:



Now, check your work. And pat yourself on the back. Because you've correctly realized that electron configurations are both easy to predict and kind of boring. Which means that you understand them pretty well!

Moving on to d- and f-orbitals

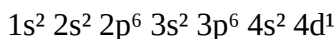
Let's write the electron configuration for calcium. Utilizing the ideas you've already learned, you can probably guess that it's:



Which is all I have to say on the matter.

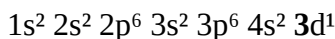
Another example: Scandium

Let's write out the electron configuration of scandium based on what we know:



That was easy enough, huh? Same thing as before. Easy, but also wrong.

You didn't think it would be completely simple, would you? It turns out that unlike the s- and p-orbitals, the number in front of the d-orbitals is actually one less than the number on the left side of the periodic table. As a result, even though there's a "4" written to the left of this row, the number of the d-orbital is actually 3. This gives us:



And that's the only difference between the d-orbitals and the other ones.

You may be asking yourself what the deal with the f-orbitals is. Just like the d-orbitals, the number in front of the f-orbitals is less than you'd expect, but this time it's less by 2 rather than 1. As a result, if we were to write the electron configuration of lanthanum, our last term would be $4f^1$ instead of $6f^1$. To help you remember this, I've written a quick song:

The s- and p's count normally

The d's are back by one

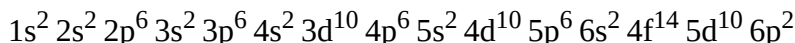
Them f- sons-o-bitches are back by two

And with that, my friends, yer done

I recommend you sing it loudly in class as much as possible.²

The final example I feel like discussing: Lead

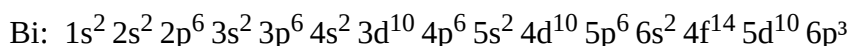
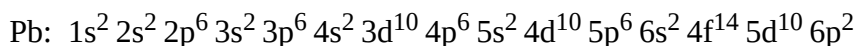
If you work out all of the terms for this, you'll find it's:



Not that much fun to write, but fortunately, we've got a way to make it easier...

Electron configurations: The short version

We just wrote out the electron configuration for lead. Now, let's put it next to the one for the next element, bismuth:



If you look at these two electron configurations, you can probably tell the difference between them after a few minutes of looking. In lead, the last term is $6p^2$, whereas in bismuth it's $6p^3$.

This really isn't surprising, given the aufbau principle idea that every element will have the electrons in the same place as the element before it, plus one. However, it should make you kind of wonder why anybody would go to the trouble of writing a bunch of terms that will all be the same, anyway.

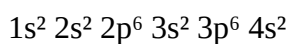
To get around this incredibly annoying situation, there's an abbreviated solution to writing electron configurations. Instead of just writing all of the terms in an electron configuration, just write the configuration only for the electrons past the most recent noble gas. Making this change, we find the electron configurations for Pb and Bi to be:

Pb: $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^2$

Bi: $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^3$

In these cases, the electron configuration is understood to be the same as before, except that the [Xe] part stands for "the electron configuration of Xe". Or, in other words, it's written instead of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$ when writing electron configurations.

This abbreviation method for electron configurations works for any element, though it's most useful for elements with lots of electrons. For example, if the electron configuration for calcium is:



we can just abbreviate it as $[\text{Ar}] 4s^2$. Which is pretty nice.

An important question: Doesn't this imply that only the outer electrons matter?

Yes it does. When we write that the electron configuration of calcium is $[\text{Ar}] 4s^2$, we're explicitly saying that we care about the 4s- electrons while we don't care at all about the ones that come before. This is because it's the outermost electrons in an element that are involved with chemistry, so we won't really ever need to refer to the ones that are earlier than the last noble gas. Sure, the electrons are still there doing their electron thing – it's just that their electron thing involves sitting in one place doing nothing. As a result, we just need to worry about the outer ones, which is what these electron configurations give us.

Footnotes:

1. Aufbau means "building up" in German, so don't capitalize it unless it's at the beginning of a sentence. It's not like there was a "Lou Aufbau" or anything like that.
2. If you liked this song, you might like the song I was listening to when I wrote it. Click [here](#) to listen to "My Name Is Mud" by Primus.