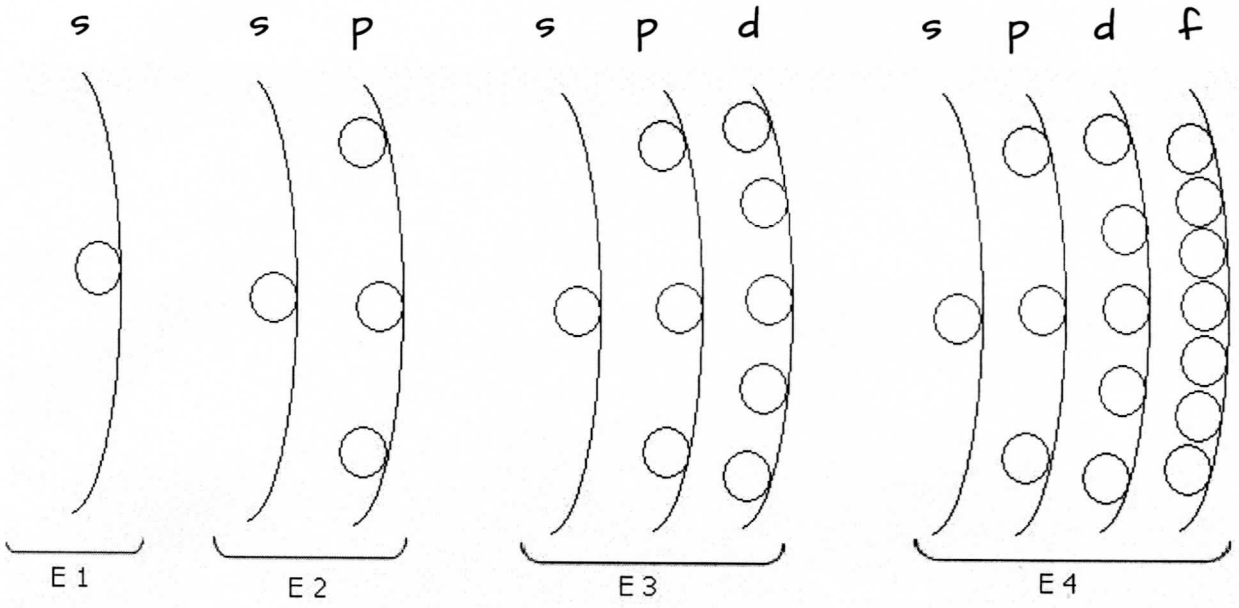
Chemistry Lecture #22: Electron Configuration

In the previous lecture we learned that electrons occupy energy levels, and that energy levels have sublevels. We learned that sublevels have orbitals.

Suppose an atom was stripped of its electrons. Then suppose we add the electrons back to the atom one at a time. Where would the electrons go? Would they go to the 1st or 3rd energy level? If it went to the 3rd energy level, would it occupy the s, p, or d sublevel? And if it occupied the d sublevel, which one of the five orbitals would it occupy?



Electrons fill the orbitals according to a pattern. This pattern can be determined if we first draw the following diagram:

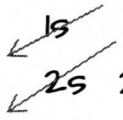
1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f

Notice that all we've done is write the energy levels and the sublevels that are in each energy level.

Next, we draw diagonal arrows through the diagram that slant from top right to bottom left. The first arrow goes through 1s.

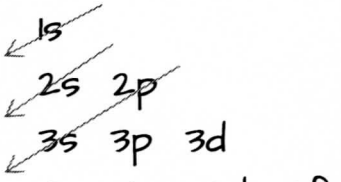
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f

The second arrow goes through 2s.

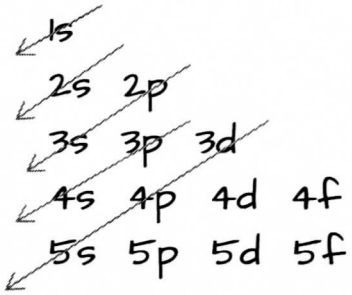


2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f

The third arrow goes through 2p and 3s.



Each time we draw a new arrow, it must go through the first term of the next horizontal row. Following this pattern, we end up with

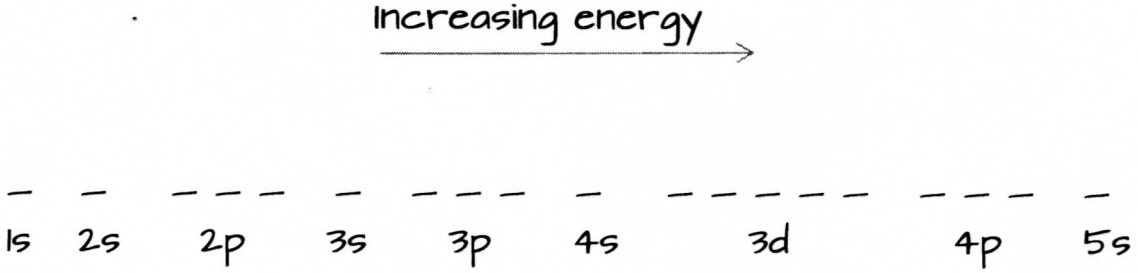


We could draw more lines, but this is all we need to do.

Next, we'll write out the terms in the order that the arrows went through them. The first arrow went through 1s. The 2nd went through 2s. The third went through 2p, then 3s. If we write out all of the terms that were hit by arrows, we get

1s 2s 2p 3s 3p 4s 3d 4p 5s

Finally, we will draw horizontal lines above the terms. The number of lines we draw depends on the number of orbitals in the sublevel. S sublevels have one orbital or one line, p sublevels have 3, and d sublevels have 5.



The above diagram is what we will use to determine the location of the electrons that orbit the nucleus. You need to either memorize the diagram or memorize the procedure for creating the diagram.

We will draw up and down arrows on the horizontal lines to indicate that an electron is in an orbital. In general, we will fill in the lines from left to right. This is because the left side of the diagram has the energy levels that are closest to the nucleus. Electrons want to get close to the nucleus, so the left side gets filled in first. As you go from left to right across the diagram, energy increases.

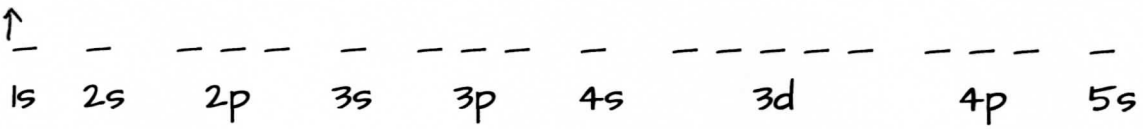
Filling in the diagram from left to right is also known as the Aufbau principle. It just means that electrons fill the lower energy levels first before filling the higher levels.

Here are the rules for filling the orbitals:

1. Choose the lowest energy level "n."
2. Choose the lowest sublevel, s, p, d, or f.
3. Fill the orbital with a maximum of 2 electrons, and have them spin in opposite directions.
4. Before a second electron can be placed in any orbital, all the orbitals of that sublevel must contain at least one electron, and spin in the same direction (Hund's rule).

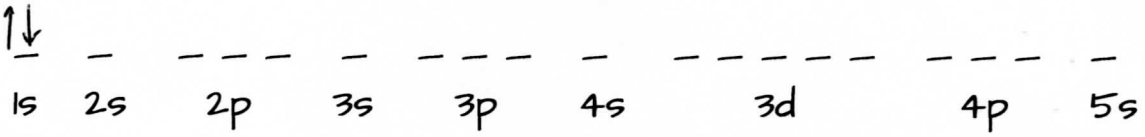
To understand how to use these rules, let's draw the electron configurations for the first ten elements. You'll need to use a periodic chart.

Draw the electron configuration for hydrogen (atomic # = 1).



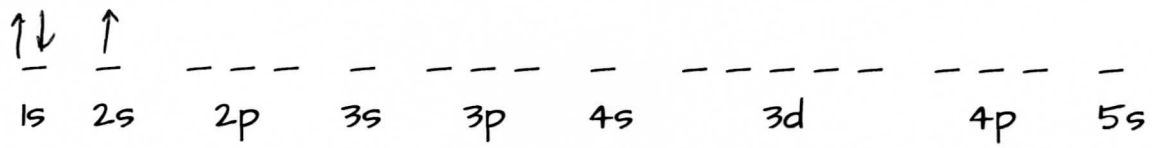
Hydrogen has 1 electron that goes into the 1s orbital.

Draw the electron configuration for helium (atomic # = 2).



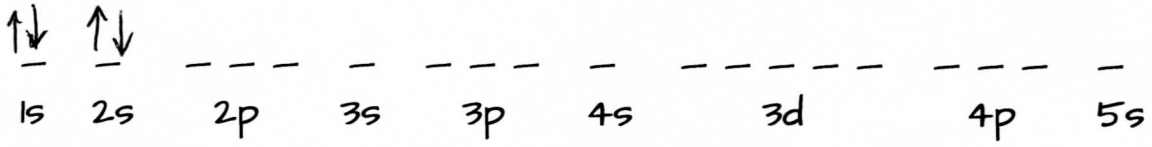
The 1s orbital has 2 arrows since each orbital can hold a maximum of 2 electrons. Notice that the arrows need to be pointed in opposite directions (Pauli Exclusion Principle).

Draw the electron configuration for lithium (atomic # = 3).

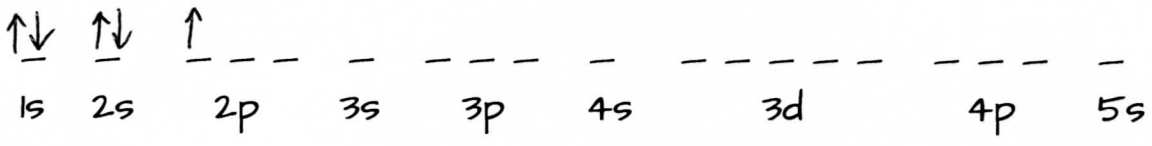


We have to put the third electron into the 2s orbital since the 1s orbital can only hold 2 electrons.

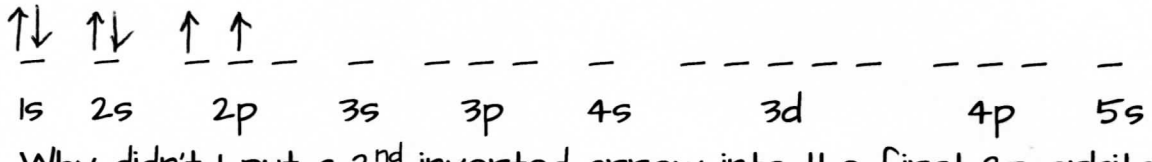
Draw the electron configuration of beryllium (atomic # = 4).



Draw the electron configuration of boron (atomic # = 5).

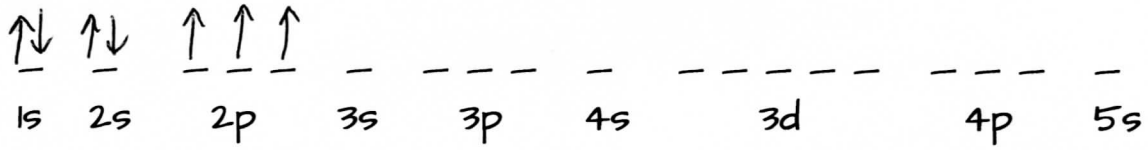


Draw the electron configuration of carbon (atomic # = 6)
Hint: remember rule #4.

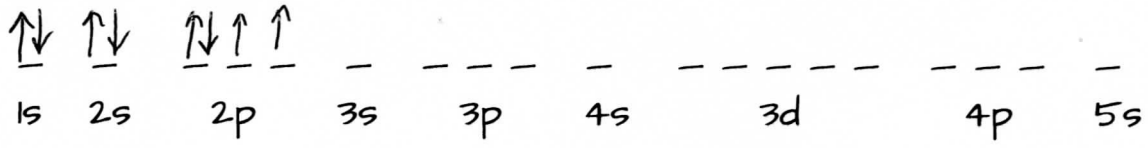


Why didn't I put a 2nd inverted arrow into the first 2p orbital? Why did I place the last arrow right side up into the second 2p orbital? Because rule #4 says that before I can put a 2nd arrow into the first 2p orbital, all the other 2p orbitals need at least 1 electron (and need to spin in same direction).

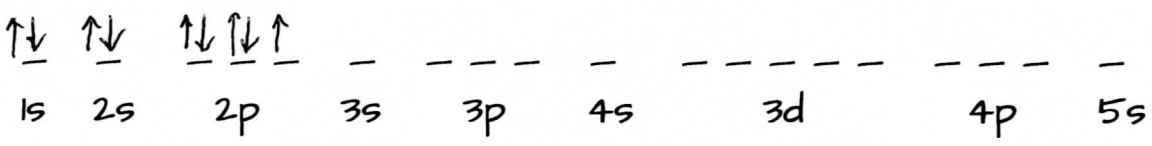
Draw the electron configuration of nitrogen (atomic # = 7).



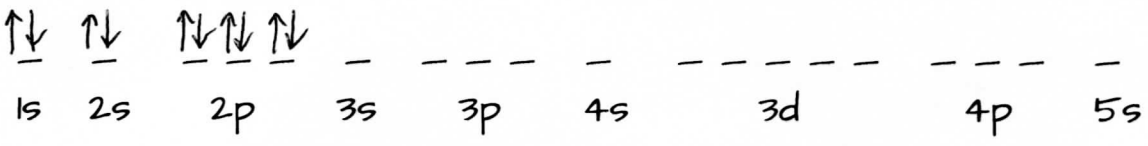
Draw the electron configuration of oxygen (atomic # = 8).



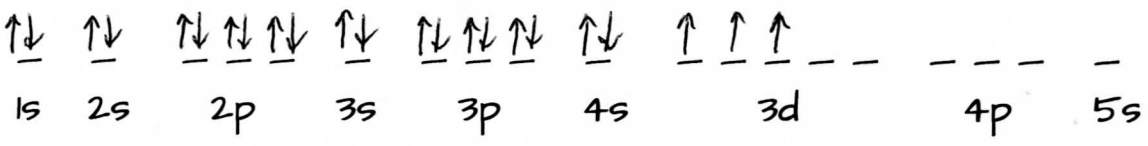
Draw the electron configuration of fluorine (atomic # = 9)



Draw the electron configuration of neon (atomic # = 10).

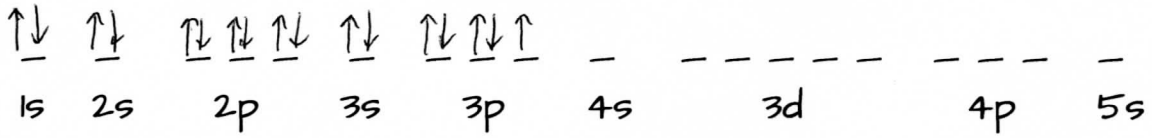


Draw the electron configuration of vanadium (atomic # = 23)



Instead of drawing arrows, an abbreviated form of the electron configuration uses superscripts. The number of the superscript is the number of arrows. If the 3p orbitals hold 5 arrows, you would write $3p^5$.

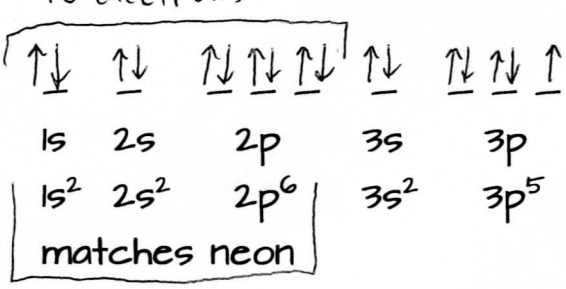
Thus, the electron configuration of chlorine (atomic # = 17) which looks like this:



Would be abbreviated as $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$.

We can abbreviate the electron configuration further if we use the configurations of the elements on the far right vertical column of the periodic chart (He, Ne, Ar, Kr, Xe, and Rn).

For example, the first part of the configuration of chlorine matches the configuration of neon. Atomic # of neon = 10.



So instead of writing the complete configuration as $1s^2 2s^2 2p^6 3s^2 3p^5$, we write [Ne] in place of $1s^2 2s^2 2p^6$ and write [Ne] $3s^2 3p^5$ as the configuration for chlorine.