Answer on Question #49761 – Chemistry - Inorganic Chemistry

How do you write electron configurations for an element and its ion give examples explain step by step?

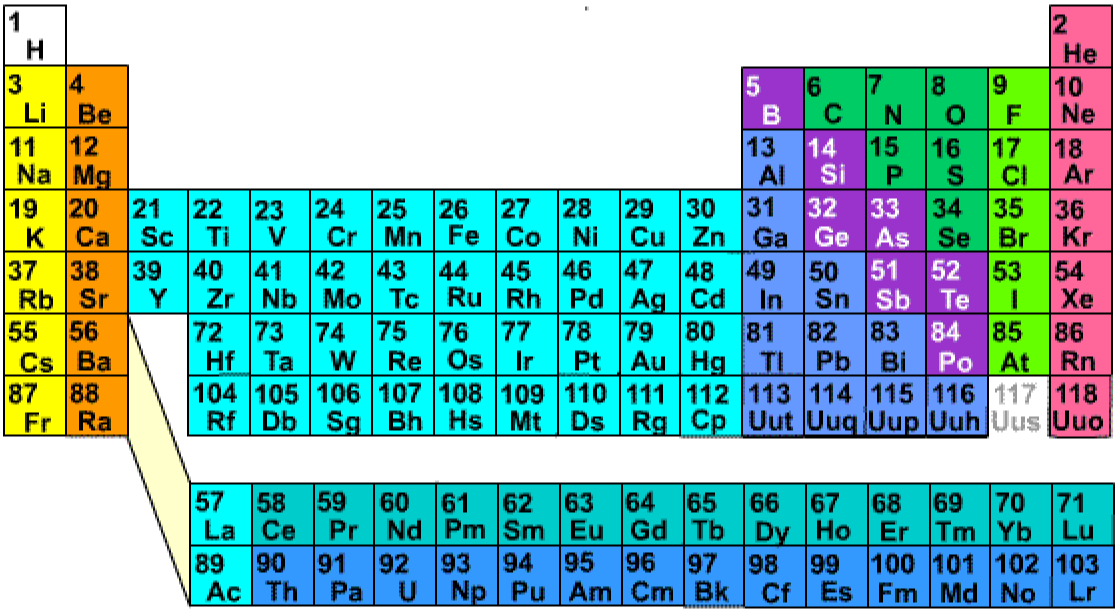
Solution:

Since all electrons do NOT orbit the nucleus per se as first believed, scientists now use the word "orbital" to describe the shape of the space an electron occupies in an atom, based on the combination of its energy, orientation in space and angular momentum. You simply need to know two things: (1) the standard electron configuration pattern and (2) how many electrons your particular atom has. [A little later, there will be one more piece of information you will need to know, but first let's start with the basics.]

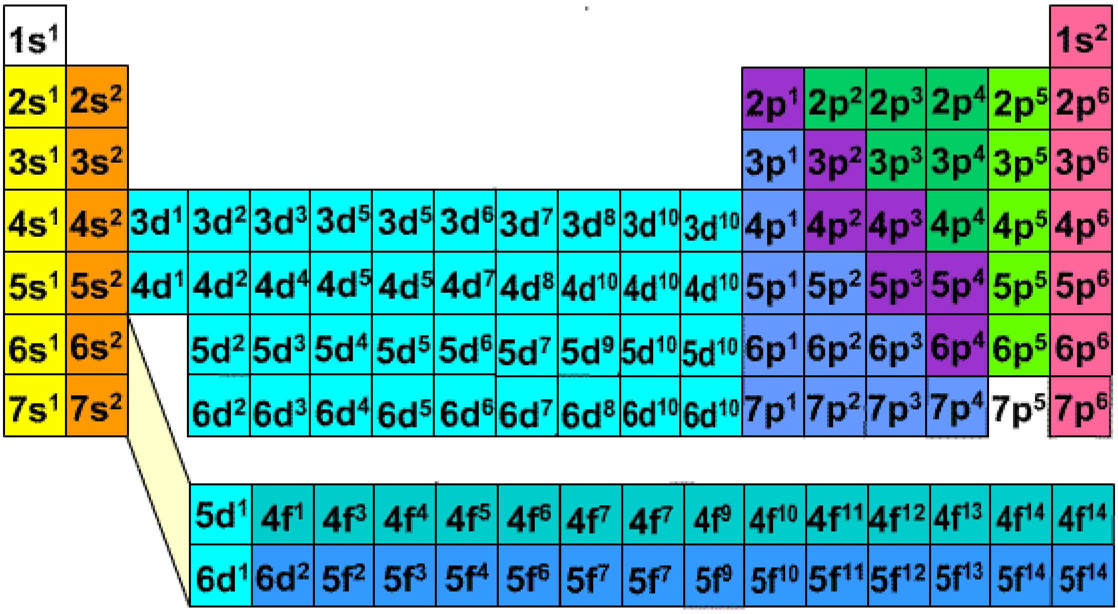
The word "configuration" means "the arrangement of the parts of something." An electron configuration describes "the arrangement of electrons in space around a nucleus." Electrons generally do not orbit the nucleus as scientists used to think. Rather, they hover, wiggle and jiggle, vibrate back and forth, and move in certain patterns based on three factors: (1) their energies, (2) their orientations in space (in other words, exactly where they fit in space relative to the other electrons in the atom) and (3) their angular momentum. (Angular momentum may be thought of as describing how wildly electrons swing about in a circular path around some center point. Some electrons may not necessarily be orbiting completely around the atom but might possibly just be making little circles out to the side, somewhat like if you were making arm circles out to the sides of your body.)

You probably already know that the elements of the periodic table are organized by number from 1 - 117. The number of the element, called the atomic number, is the number of protons found in the nucleus of exactly one atom of that element. The atomic number is also the number of electrons an atom has WHEN THE ATOM IS NEUTRAL. It is very important to know at this point that some of the outer electrons in an atom are able to jump onto or off of other atoms according to certain physical laws, so the number of electrons an atom has can change from moment to moment. The protons, however, are found deep inside the atom, in the nucleus, and CANNOT come and go. Therefore, the number of protons in an atom fixes the identity of the element, regardless of the number of electrons an atom has from one moment to the next.

Let's look now at the Periodic Table.



As we study electron configuration (the arrangement of electrons in atoms), we discover a pattern. Study the periodic table below and see what you can discover.



Notice the s, p, d and f blocks. Notice the sequence of main energy levels. Notice the superscripts which tell how many electrons are in each orbital sublevel. The pattern becomes more complex with each succeeding layer, and we have not included all the details here, but you should still be able to recognize a general trend. We have chosen to place in each element box the one term in the electron configuration pattern which shows movement in that section. In some cases, this is the last and highest term in the configuration. In other cases it is not.

The **standard electron configuration pattern** is as follows:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10}$$

The letters, i.e., "*s*, *spsps*, *dps*, *dps*, *fdps*, *fd*," refer to the orbital sublevels in which the electrons are found. Each MAIN energy level has a *number* of sublevels equal to its value. If you look more closely at the sequence again,

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10}$$

you will notice the following pattern:

MAIN Energy Level	**Has this number of Orbital Sublevels**	**Identity of Orbital Sublevels *see text below**
1	**1**	**s**
2	**2**	**s p**
3	**3**	**s p d**
4	**4**	**s p d f**
5	**5**	**s p d f (g)***
6	**6**	**s p d (f) (g) (h)***
7	**7**	**s (p) (d) (f) (g) (h) (i)***

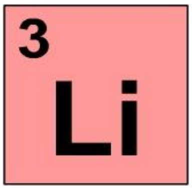
MAIN energy level "1" has only 1 sublevel, the "s" orbital sublevel. MAIN energy level "2" has two sublevels, both the "s" and "p" sublevels. MAIN energy level "3" has three sublevels, "s", "p" and "d." The pattern continues with MAIN energy levels "4, 5, 6 and 7," but the higher sublevels of "5, 6 and 7" (those shown in parentheses above) do not appear in the standard electron configuration pattern, because there are only 117 known elements, so after 117, we don't have any more electrons to put into the higher orbitals.
Each succeeding MAIN energy level has more orbital sublevels because as we move out from the nucleus of the atom, there is more room for more electrons which can have the same MAIN amount of energy but have different orientations in space and different angular momenta.

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^2\,5f^{14}\,6d^{10}$$

Now look at the small black raised numbers.
The small raised numbers shown above in black are called "superscripts." Superscripts represent the maximum numbers of electrons each orbital sublevel can hold. An orbital sublevel may hold LESS electrons than its maximum, but it can *never* hold MORE.

How to Use the Standard Electron Configuration Pattern to write Electron Configurations for Different Elements and Ions

Once you can write the standard electron configuration pattern, you can then write the electron configuration for any atom or ion. (An ion is an atom with an *unequal* number of protons and electrons.) To write the electron configuration, you just count the number of electrons you have and use as many orbital sublevels as you need to hold all your electrons. Be sure to always fill the lowest energy levels first.



For example, let's say you have an atom of lithium. Lithium's atomic number is 3. So a neutral atom of lithium has 3 protons and 3 electrons. We would need space for 3 electrons. Write:

$$1s^2\,2s^1$$

There are 2 electrons in the 1s sublevel and 1 electron in the 2s sublevel. 2 + 1 = 3. That's it!

Now let's try fluorine, which has an atomic number of 9. A neutral atom of fluorine has 9 protons and 9 electrons. We need enough space for 9 electrons. The 1s orbital can hold 2 electrons and the 2s orbital can hold 2 more electrons. The five remaining electrons must go into the next orbital, the 2p orbital. The 2p orbital can hold up to 6, but we only have 5. So the following would be the correct electron configuration for a neutral atom of fluorine. 2 + 2 + 5 = 9.

$$1s^2\,2s^2\,2p^5$$

33
As

What if we had an atom of arsenic, with atomic number 33? We would fill up all the orbital sublevels in the electron configuration pattern until we got to the 4p orbital sublevel. Then we would have only 3 electrons left. 2 + 2 + 6 + 2 + 6 + 2 + 10 + 3 = 33. The electron configuration for arsenic, then, is as follows:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$$

Li^{+1}

Now, what if lithium were an ion, rather than an atom? (Remember, ions have *unequal* numbers of protons and electrons.) The lithium ion has a +1 charge. The lithium ATOM has 3 electrons. How many electrons does the lithium ION have? Remember, electrons come and go, because they are on the "outside" of the atom, but the protons, which are held tighly together in the nucleus of the atom just stay there. So with 3 protons, how many electrons would we need to have in order to get a +1 charge? We would need a number of electrons which is 1 less than the number of protons. **Answer: 2.**
3 positive charges (from the protons) and 2 negative charges (from the electrons) add up to a total of +1. So we have only 2 electrons, and the electron configuration for the lithium ION is:

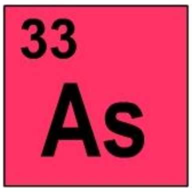
$$1s^2$$

F^{-1}

How about fluorine? As an ion, does it gain electrons or lose electrons? Its closest noble gas is neon, with 10 protons and 10 electrons. The elements always tend to be most stable in the "noble gas electron configuration." So fluorine would take on one EXTRA electron to have a total of 10 electrons like neon. 2 + 2 + 6 = 10.
9 positive charges (from the protons) and 10 negative charges (from the electrons) add up to a total of -1. So we have 10 electrons and the electron configuration for the fluorine ION is:

$$1s^2 2s^2 2p^6$$

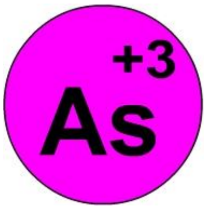
Because of arsenic's position on the periodic table, it can make four different ions: As+3, As-3, As+5 and As+1. Electron configurations for the first two ions of arsenic may be written using the basic rules we have already learned.



Here is the electron configuration again for neutral arsenic so you can easily compare it with the As+3 and As-3 ions.

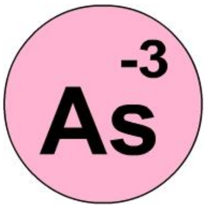
$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$$

33 electrons. (An equal number of protons and electrons.)



$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$$

30 electrons. (A positive charge means that electrons have been LOST. Notice that the 3 electrons in 4p are missing.)



$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$$

36 electrons. (A negative charge means that electrons have been GAINED. Notice that 4p has 3 EXTRA electrons, making a total of 6 and completely filling the 4p sublevel.)