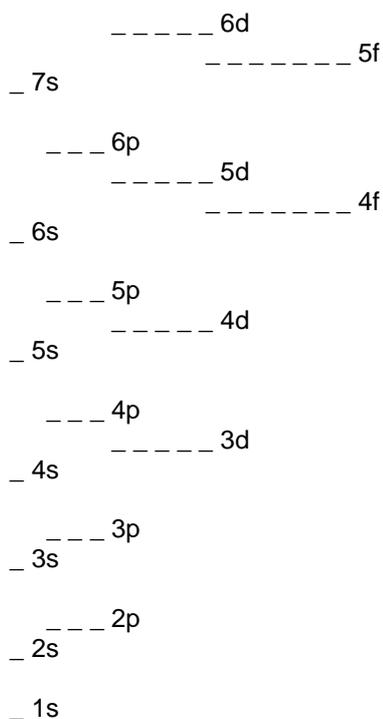


## Electron Configuration Notes

### Background Information:

The arrangement of electrons within the orbitals of an atom is known as the electron configuration. The most stable arrangement is called the ground-state electron configuration. This is the configuration where all of the electrons in an atom reside in the lowest energy orbitals possible. Keeping in mind that each orbital can accommodate a maximum of two electrons, we are able to predict the electron configurations of elements using the periodic table.

Basically, the distributions of orbitals can be laid out in the following fashion (read from the bottom up):



The bottom energy level is level 1 - it has the lowest energy. Each "\_" represents an orbital. You can see that there is 1 orbital for an s subshell. There are 3 orbitals for a p subshell, 5 for a d, and 7 for an f subshell. Each orbital can hold 2 electrons. Therefore, the s subshell can hold 2 electrons; the p can hold 6; the d can hold 10; and the f can hold 14. And thus, the first energy level can hold 2 electrons ( $1s = 2$ ), the second energy level can hold 8 electrons ( $2s2p = 2 + 6$ ), the third energy level can hold 18 electrons ( $3s3p3d = 2 + 6 + 10$ ), and the fourth energy level can hold 32 ( $4s4p4d4f = 2 + 6 + 10 + 14$ ).

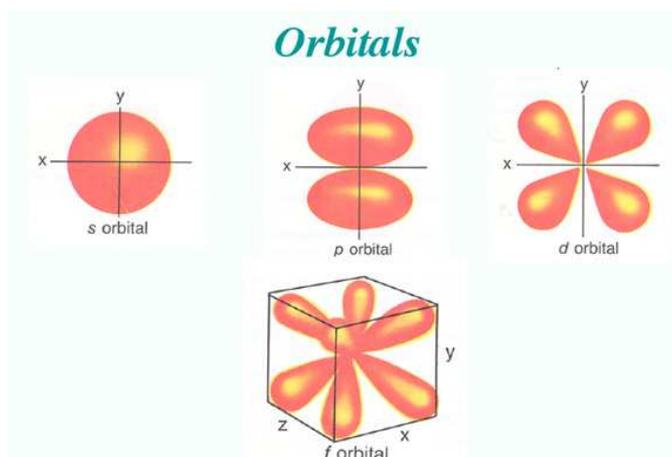
### Parts of an electron configuration:

- **Energy level** - a number (1, 2, 3 and so on)
- **Sublevel (orbital)** - a letter, either s, p, d, or f
- **Number of electrons** - a superscript number

**Analogy:** The energy level is like a driveway with cars in it, the sublevels are the type of cars in parking lot, and the orbitals are how many seats are in the car.

### What do the s, p, d and f orbitals look like?

The electron orbitals shown below represent a volume of space within which an electron would have a certain probability of being based on particular energy states and atoms. For example, in a simple lowest-energy state hydrogen atom, the electrons are most likely to be found within a sphere around the nucleus of an atom. In a higher energy state, the shapes become lobes and rings, due to the interaction of the quantum effects between the different atomic particles.



### How to write an electron configuration:

In a neutral atom, the number of electrons equals the number of protons of the atom. When the electrons fill the orbitals, they occupy the lowest energy orbitals that are available.

For example, hydrogen is atomic number 1 (has 1 proton). The one electron that it has occupies the lowest orbital, which is 1s. To write its electron configuration, it would be  $1s^1$ . In an orbital diagram, it would simply be a line with one up arrow in it, which represents the 1s orbital:

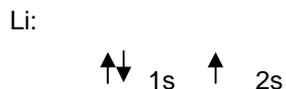
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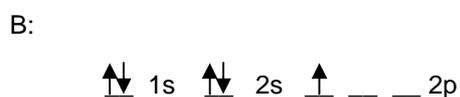
Likewise, helium has 2 protons and its electron configuration would be  $1s^2$ . Its orbital diagram would be a line with one up arrow and one down arrow.



Lithium has 3 protons and an electron configuration of  $1s^2 2s^1$ , so the orbital diagram will be:



Now things get trickier with higher orbitals. For example, Boron has an electron configuration of  $1s^2 2s^2 2p^1$  and the orbital diagram looks like this:



Now one might think that carbon, with an electron configuration of  $1s^2 2s^2 2p^2$  would have an orbital diagram of this:



But that is incorrect. It actually has one of this:



The rule for doing this is that when electrons are placed in a set of orbitals of equal energy, they are spread out to give as few paired electrons as possible. For carbon, that means that the two p electrons are in separate orbitals.

Basically, the above information can be summed up into 3 rules that define how electrons can be arranged in an atom's orbital.

### 1. The Aufbau Principle

- Each electron occupies the lowest energy orbital
  - i.e. Electrons are Lazy!
- All orbitals related to an energy level are of equal energy.
  - i.e. The three 2p orbitals are the same energy level.

## 2. Hund's Rule

- Single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals.
  - i.e. Electrons are unfriendly!

**Why?** Electrons, being unfriendly, fill up the empty orbitals before sharing orbitals. This is similar to seats on a bus – on a bus, you sit alone, rather than with a stranger, if there is an option.

## 3. Pauli Exclusion Principle

- **A maximum of two electrons may occupy a single orbital, but *only if the electrons have opposite spins.***
  - **Spin -- Electrons have an associated “spin,” either one way or the other, like a top.**
  - **These spins are called “spin up” and “spin down.”**

Exceptions to filling order are copper, chromium and some others- because half-filled shells are more stable, so an atom may borrow an electron from next s shell to half fill a lower energy level.

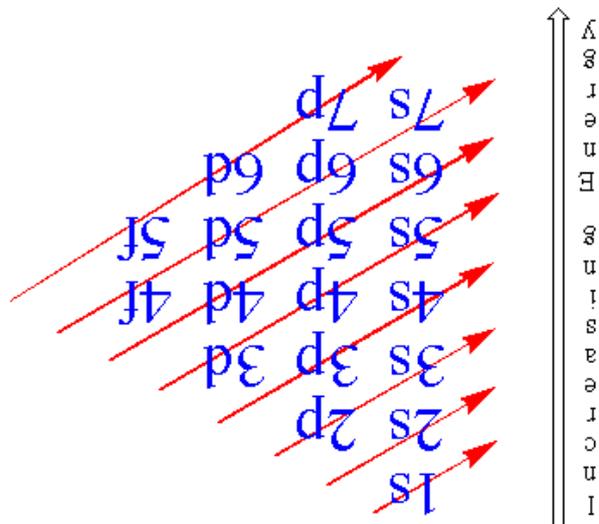
## Apartment Analogy

Imagine you are the landlord of a very strange apartment building. Your job is to fill the apartments in the building in the most efficient way possible. You are required by the owner of the building to fill the rooms in a certain way. The rules you have to follow are as strange as the building because quantum mechanics is not like anything you might have expected. The rules are summarized in the table below.

In the building the different floors are like the different energy levels (or shells) in an atom. The energy levels are numbered starting from one, just like the floors in an apartment. Each room corresponds to one orbital. The rooms have a capacity of two electrons (two people) each. In each room only a man and a woman may be paired together. In the strange world of quantum mechanics there are no same-gender room mates.

Apartment Rules	Electron Rules
<b>From the Bottom Up:</b> Rooms must be filled from the ground floor up. Fill the one room on the first floor before starting to put new tenants on the second floor. Then fill the s room before the p rooms. At higher floors the order might change a bit.	<b>Aufbau Principle:</b> the electrons fill the available orbitals from lowest energy to highest energy. In the ground state all the electrons are in the lowest possible energy level.
<b>Singles First:</b> the owner of the building wants to have the tenants spread out as much as possible. For that reason singles are placed in rooms before couples. If couples must be placed into a room then all of the other rooms on that floor must already have a single in them.	<b>Hund's Rule:</b> The electrons must be placed into the orbitals in such a way that no pairs are put together unless absolutely necessary. That is, single electrons must be placed into boxes first and then paired up if necessary.
<b>Opposite Gender Only:</b> When two people are placed in a room they must be of opposite genders. No men may room together and no women may room together. This is an arbitrary rule on the part of the owners: in a just world we wouldn't have to follow it. But quantum mechanics has nothing to do with justice.	<b>Pauli Exclusion Principle:</b> Electrons come in two varieties based on the direction they are 'spinning'. There is an Up spin and a Down spin. Up and Down spins are always paired together and Up-Up or Down-Down combinations are not allowed. No two electrons can ever be in the same place at the same time.

Graphically, the orbital diagram for electron filling can be depicted as follows, showing increasing energy from the 1s orbital down to the 7p orbital.



The rule is to start at the beginning of each arrow, and then follow it all of the way to the end, filling in the sublevels that it passes through. In other words, the order for filling in the sublevels becomes; 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p.

**Noble Gas Notation:**

When writing electron configurations, you may be required to write the simplified “noble gas notation” for a particular element. To do this, find the element on the periodic table, locate the previous noble gas and then add the rest of the electron configuration beyond the noble gas.

**For example:**

Calcium has an electron configuration of  $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$ .

The noble gas prior to calcium is argon, so:

Its noble gas notation would be:  $[Ar], 4s^2$ .

## How does electron configuration relate to the periodic table?

The periodic table is organized into Periods (rows), Groups 1-18 (columns) and Blocks (s, p, d and f). The periods and rows will likely be familiar to students but s, p, d and f blocks may be new.

The periodic table below shows the s, p, d and f view:

IA 1	"representative elements"										IIIA 13	IVA 14	V 15	VIA 16	VIIA 17	VIIIA 18																															
H	Li	Be	B	C	N	O	F	Ne																																							
Na	Mg	Al	Si	P	S	Cl	Ar																																								
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr																														
Cs	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe																														
Rb	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn																														
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	"main group elements"																																					
<table border="1"> <tr> <td>Ce</td><td>Pr</td><td>Nd</td><td>Pm</td><td>Sm</td><td>Eu</td><td>Gd</td><td>Tb</td><td>Dy</td><td>Ho</td><td>Er</td><td>Tm</td><td>Yb</td><td>Lu</td> <td>Lanthanides</td> </tr> <tr> <td>Th</td><td>Pa</td><td>U</td><td>Np</td><td>Pu</td><td>Am</td><td>Cm</td><td>Bk</td><td>Cf</td><td>Es</td><td>Fm</td><td>Md</td><td>No</td><td>Lr</td> <td>Actinides</td> </tr> </table>																		Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Lanthanides	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Actinides
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Lanthanides																																	
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Actinides																																	

## How does all of this relate to the formation of ions?

Why do ions form? Atoms (except the noble gases) can lower their energy, thus becoming more stable, by gaining or losing electrons in order to have full s and p subshells. The goal for all atoms is to have a full s and p subshell (NOT a full shell - the d and f don't need to be full to be stable). Noble gases don't have full d or f subshells. It is all about getting to an  $s^2p^6$  configuration! This configuration will lower energy and provide stability. We often say the motivation or goal for the atoms is to "be like a noble gas."

Take a look at sodium:

- Na -  $1s^22s^22p^63s^1$  not stable 😞
- Ne -  $1s^22s^22p^6$  stable 😊
- So sodium's goal is to lose that one outer electron to be like Ne.
- $Na^+$  -  $1s^22s^22p^6$  stable 😊

Take a look at chlorine:

- Cl -  $1s^22s^22p^63s^23p^5$  not stable 😞
- Ar -  $1s^22s^22p^63s^13p^6$  stable 😊
- So chlorine's goal is to gain one electron so it can be like Ar.
- $Cl^-$  -  $1s^22s^22p^63s^23p^6$  stable 😊