Electron Configurations

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to…)

- Write the ground-state electron configuration for any element on the periodic table.

Success Criteria:

- Levels and sublevels are in the correct order.
- Each sublevel has the correct number of electrons.

Tier 2 Vocabulary: level, spin

Language Objectives:

- Explain the parts of an electron configuration.

Notes:

The electron configuration for an element is a list of all of the energy sub-levels that have electrons in them (in the ground state) in order from lowest to highest energy, and the number of electrons they contain.

For example, consider a neutral nitrogen atom with its seven electrons.

- The first two electrons occupy the 1s sublevel—the one with the lowest energy. We denote these two electrons as $1s^2$.
- The next two electrons occupy the 2s sublevel. We denote these two electrons as $2s^2$.
- The last three electrons are in the 2p sublevel. We denote these three electrons as $2p^3$.
- The complete electron configuration for nitrogen is therefore $1s^2\ 2s^2\ 2p^3$.

If this already makes sense, great! The next few pages explain where these numbers come from in more detail.

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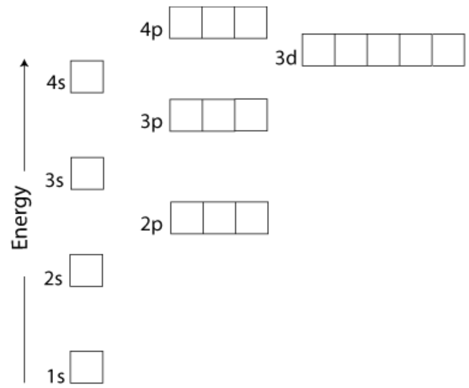
The following chart shows each orbital from 1s through 4p. Orbitals with the lowest energy are shown at the bottom, and those with the highest energy are at the top.



This can be a little confusing, because the levels overlap. For example, notice that the 3*d* sub-level is higher in energy than the 4*s* sub-level.

This means that the electrons will fill sub-levels in the following order:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, ...

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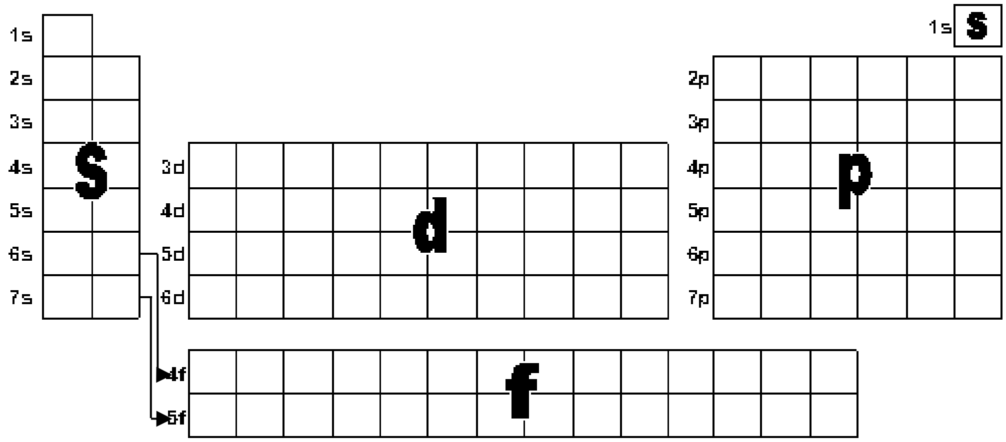
You can use the periodic table as the "road map":



To determine the electron configuration, imagine that you are placing one electron in each element's box until you "use up" all of the electrons. The last box is the element that you are writing the electron configuration for.

The two columns on the left* correspond with the "*s*" sublevels. The six columns on the right correspond with the "*p*" sublevels. The ten columns of the transition metals correspond with the "*d*" sublevels. The fourteen columns below the rest of the table correspond with the "*f*" sublevels.

As you move through the positions in order, you are moving through the sub-levels from lowest to highest energy.

Remember that the *"s"* sub-levels start with 1s, the *"p"* sub-levels start with 2p, the *"d"* sub-levels start with 3d, and the *"f"* sub-levels start with 4f. The "gotchas" are:

- The 3d sub-level is in row 4, right after 4s.
- The 4f sub-level is in row 6, right after 6s.

* For the purpose of electron configurations, helium should be in the "s" block, next to hydrogen.

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Writing Electron Configurations

An element has electrons that correspond with <u>each</u> of the available slots, from the beginning of the periodic table (where hydrogen is located) up to where that element is located.

If we were to represent an electron as an arrow, we could represent two electrons in a 1s sub-level like this: $\underset{1s}{\underline{\uparrow\downarrow}}$. The 1s sub-level has one orbital, which is represented by the one blank. The two electrons are represented as arrows. Because two electrons sharing an orbital have opposite spins, we represent them with one arrow pointing up and the other arrow pointing down.

We could represent five electrons in a 2p orbital like this: $\underline{\uparrow\downarrow}\ \underset{2p}{\underline{\uparrow\downarrow}}\ \underline{\uparrow\ }$. The 2p sub-level has 3 orbitals, represented by the 3 blanks. Two of those orbitals have two electrons in them, and the third one has only one electron.

We could represent all 13 of the electrons in aluminum like this:

$$\underset{1s}{\underline{\uparrow\downarrow}}\quad \underset{2s}{\underline{\uparrow\downarrow}}\quad \underline{\uparrow\downarrow}\ \underset{2p}{\underline{\uparrow\downarrow}}\ \underline{\uparrow\downarrow}\quad \underset{3s}{\underline{\uparrow\downarrow}}\quad \underline{\uparrow\ }\ \underset{3p}{\underline{\quad}}\ \underline{\quad}$$

This diagram shows the <u>electron configuration</u> of aluminum.

<u>electron configuration</u>: a description of which levels and sub-levels the electrons in an element are occupying.

Notice that we have to show all three of the orbitals (blanks) in the 3p sub-level, even if some of those orbitals don't have any electrons in them.

<u>ground state</u>: when all of the electrons in an atom are in the lowest-energy sublevel that has an available "slot".

<u>Pauli Exclusion Principle</u>: every electron in an atom has a different quantum state from every other electron. In plain English, this means that something has to be different about each electron, whether it's the level, sub-level, which orbital it's in, or its spin.

<u>aufbau principle</u>: in the ground state, each electron in an atom will occupy the lowest available energy state. This means that you start with the lowest sub-level (1s) and work your way up until you've placed all the electrons.

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<u>Hund's Rule</u>: electrons don't pair up in orbital until they have to. (This is kind of like siblings not wanting to share a room if there's an empty room available.) For example, the electron configuration for nitrogen would be:

	1s	2s	2p		
Wrong:	↑↓	↑↓	↑↓	↑	__
Right:	↑↓	↑↓	↑	↑	↑

If you don't need to draw every electron, you can use a shorter form, in which you just write the level and sub-level, and use a superscript for the number of electrons in the sub-level.

For example, $\underset{1s}{\underline{\uparrow\downarrow}}$ would become 1s^2, and $\underset{2p}{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}$ would become 2p^6.

The electron configuration for aluminum would go from the orbital notation version:

$$\underset{1s}{\underline{\uparrow\downarrow}}\quad \underset{2s}{\underline{\uparrow\downarrow}}\quad \underset{2p}{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}\quad \underset{3s}{\underline{\uparrow\downarrow}}\quad \underset{3p}{\underline{\uparrow}\ \underline{}\ \underline{}}$$

to the "standard" version:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$$

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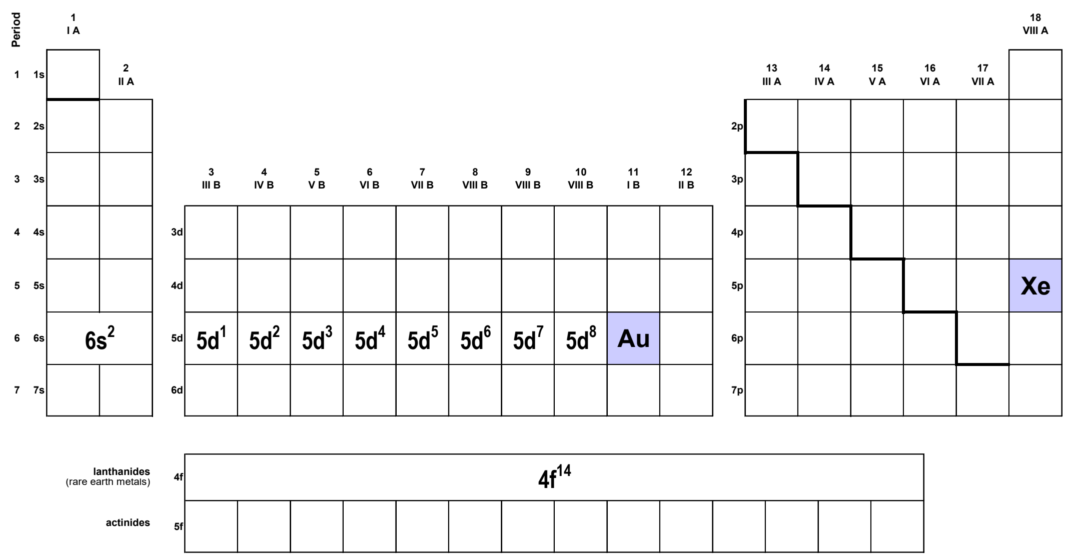
The shorter version can still get tediously long for elements with a lot of electrons. For example, the electron configuration for gold (Au) is:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^9$$

To shorten this even more, you're allowed to use the element in the *last* column of a row as an abbreviation for all of the electrons through the end of that row.

In our example, gold (Au) is in the 6th row of the periodic table:



This means we're allowed to start from xenon (Xe) at the end of the previous (5th) row, and add on only the parts that come after Xe. This gives us:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5\ d^9$$

Noble gas configuration: $[Xe]\ 6s^2\ 4f^{14}\ 5d^9$

This notation is called the <u>noble gas configuration</u>, because the elements in the last column (the ones you start from) are called the noble gases.

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Homework Problems

Give the electron configuration (orbital notation—with the arrows) for each of the following elements:

1. carbon

2. potassium

3. silicon

4. silver

For each of the following electron configurations, name the element.

5.

1s	2s	2p		
↑↓	↑↓	↑↓	↑↓	↑

6.

1s	2s	2p			3s	3p			4s	3d				
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑

Each of the following electron configurations has something wrong with it. For each one:

- State what the mistake is.
- Re-write the electron configuration correctly, keeping the total number of electrons the same.

7.

1s	2p		
↑↓	↑↓	↑	↑

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8.

1s	2s	2p	
↑↓	↑↓	↑↓	↑↓

9.

1s	2s	2p			3s	3p		
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	

For each of the following elements, give the "standard" electron configuration (*e.g.,* $1s^2\ 2s^2\ 2p^6\ 3s^1$).

10. boron (B)

11. phosphorus (P)

12. vanadium (V)

13. strontium (Sr)

For each of the following electron configurations, give the element.

14. $1s^2\ 2s^2\ 2p^6\ 3s^2$

15. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^6$

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Electron Configurations

For each of the following elements, give the "noble gas" electron configuration (*e.g.,* [Ar] $4s^2$ $3d^5$).

16. zirconium (Zr)

17. platinum (Pt)

18. dysprosium (Dy)

19. gallium (Ga)

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