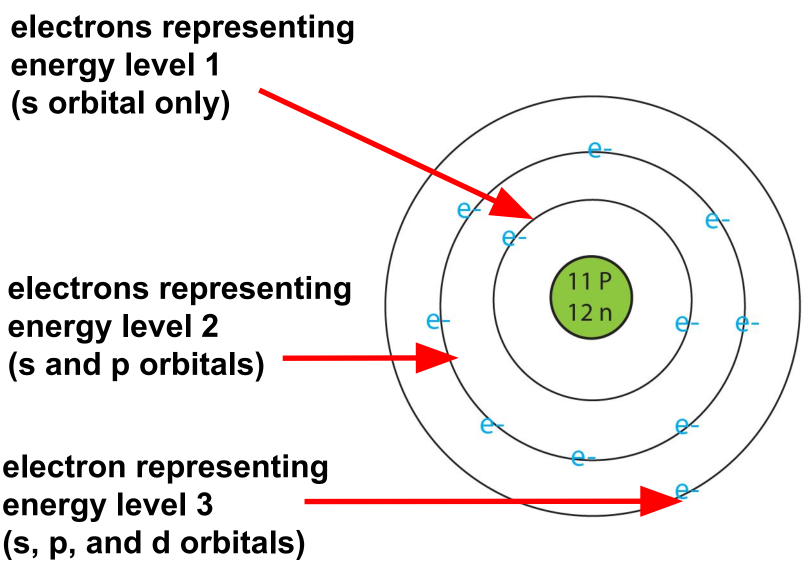
Explore Student Guide

Name: ______________________________ Date: __________

Part I: Principal Energy Levels and Sublevels

As you already know, all atoms are made of subatomic particles, including protons, neutrons, and electrons. Positive protons and neutral neutrons are found in the nucleus of the atom. The nucleus is where the positive charge of the atom is found. The negatively charged electrons orbit around the nucleus, but they do not do this randomly. In fact, the arrangement of the electrons surrounding the nucleus is fairly ordered, with each electron occupying a specific atomic orbital. Orbitals are the specific regions around the nucleus of an atom in which electrons travel. In general, this region is called the electron cloud. Within this cloud are specific subshells identified as the s, p, d, or f subshell. These subshells are further broken down into orbitals. The subshells and orbitals depend on what period and group the element is located in. The electrons can be identified by electron configurations, or the arrangement of an element's electrons in various orbitals around the nucleus. Electrons are arranged to achieve the lowest energy state.

The diagram on this page shows the electron orbitals for the element sodium (Na). A neutral atom of this element would have 11 electrons: two in the 1s orbital, two in the 2s orbital, six in the 2p orbital, and one in the 3s orbital. Each "layer" represents an electron energy level. Period 1 on the periodic table represents the first energy level, or n = 1. Period 2 represents the second energy level, or n = 2, and so on.



These specific electron subshells and orbitals may be easily identified on the periodic table. To help you start to distinguish the correct electron orbitals on the periodic table, create a periodic table with your class.

Explore Student Guide

Part I: Principal Energy Levels and Sublevels, continued

Create a Periodic Table

1. Get into your four class groups. Each group will represent a different electron subshell and orbitals. Group 1 will represent the s orbitals. Group 2 will represent the p orbitals. Group 3 will represent the d orbitals. Group 4 will represent the f orbitals.
2. Proceed in order. Group 1 will start by writing $1s^1$ on their colored self-adhesive note card. Place this note card in the hydrogen space on your periodic table (PT). This represents that there is one electron in the s orbital in the first principal energy level.
3. Then write $1s^2$ on another self-adhesive and place this in the helium space on the PT. This represents that there are two electrons in the s orbital in the first principal energy level. This energy level is now full.
4. Then write a $2s^1$ on a self-adhesive and place this on the lithium space, followed by a $2s^2$ in the beryllium space.
5. Group 2 will now write a $2p^1$ on a self-adhesive and place this in the boron space.
6. Fill the remainder of the second period until reaching neon, which will have a $2p^6$ self-adhesive. You have now completed the electron orbital sublevels for the first two principal energy levels.
7. Continue following this pattern until reaching argon at the end of the third period.
8. Group 1 will continue the pattern for the fourth principal energy level by filling in potassium and calcium.
9. Group 3 should get ready. Scandium (Sc) is next, but you have not filled the third principal energy level yet. This element is represented by a $3d^1$. You must fill the 3d orbitals before you can move on to the remainder of energy level 4.

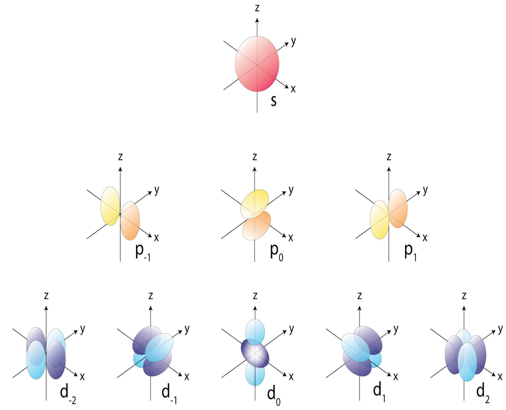
   NOTE: There are several d block elements that will not follow the pattern in periods 4 and 5, but you will find out why and adjust these later. For now, continue the pattern.

10. Once you get to the sixth period, group 4 needs to get ready. Starting with the first lanthanide series element lanthanum (La), you will place a 4f. The fourth principal energy level is the first to have f orbitals, but electrons do not start to fill them until after the 6s orbital has been filled.
11. For the elements in periods 6 and 7, know that the electron configurations do not follow a similar systematic pattern as before. Just fill in each element with the correct period number and orbital letter. For example, all of the d-block elements in period 6 will have a 5d. All f-block elements starting in period 6 will contain a 4f. They will not contain superscripted electron numbers. However, fill them in to see the pattern.

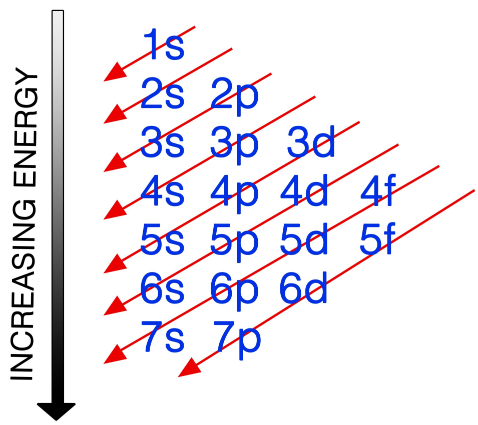
Explore Student Guide

Part II: Orbital Filling Diagrams

Electron subshells have specific shapes. These shapes correspond to the most probable areas where electrons can be found. The images to the right show the shapes of the s, p, and d orbitals (f orbitals are not shown). As you can see, there are three different axes (x, y, and z) to help one visualize the spatial orientations of these orbitals. The s orbitals have a spherical shape, the p orbitals have a dumbbell shape, and most of the d orbitals have a cloverleaf shape. The f orbitals are much more complicated.



The electrons found in the electron orbitals will fill the lowest energy levels first. There is a specific order to which principal energy levels and subshells are filled when creating electron configurations. Look at the graphic to the right. You can follow the direction of the arrows to help you in this process. Level 1s is filled before 2s. Level 2p is filled before 3s. However, notice where the pattern changes. Level 3p will be filled, then 4s, then 3d, then 4p, and so on. As you can see, the orbitals are not filled by following a number sequence, but instead are filled by following an increase in energy sequence. Study this image and become familiar with how to use it, as you will refer to it extensively when writing the electron configurations of elements.



To be able to correctly configure the electrons of elements, you need to understand three main rules:

1. **The Aufbau Principle.** This principle states that the electrons in any electron orbital will occupy the lowest energy level first. In this case, remember that each period on the periodic table represents a principal energy level. Period 1 is principal energy level 1, represented by n = 1. Period 2 represents principal energy level 2, represented by n = 2.

   Refer to the graphic above to follow the correct pattern of increasing energy levels.

2. **The Pauli Exclusion Principle.** This principle states that no two electrons in a particular atom can be in exactly the same energy state. An electron's energy state is defined by its principal energy level, its electron subshell (s, p, d, or f), the spatial orientation of the orbital it is in, and a property called spin. Electrons can spin in only two directions, clockwise or counterclockwise, which are referred to as up and down. Each electron will be represented by an arrow pointing up (↑) or by an arrow pointing down (↓). As a result of this principle, each electron orbital can hold no more than two electrons.

Explore Student Guide

Part II: Orbital Filling Diagrams, continued

3. **Hund's Rule.** When the electrons occupy the electron orbitals of the same energy subshell, they will occupy each orbital with the same spin direction. This especially comes into play with the p, d, and f orbitals. Let's begin to solve this puzzle by looking at an orbital filling diagram for the element oxygen.

The 1s orbital is filled first, followed by the 2s orbital. See how the 2p orbital is filled with electrons with the same spin before moving back to the $2p_x$ box?

Element	1s	2s	$2p_x$	$2p_y$	$2p_z$
Oxygen	↑↓	↑↓	↑↓	↑	↑

Part III: Writing Electron Configurations

You are now ready to write the electron configurations of any element. The written electron configurations are simply numerical representations of all that you have done. It may be written in one of two ways. First, you can write it as the complete electron configuration, which represents each electron that may be found in that element. Second, you can write it with a noble gas core. Instead of writing the entire configuration, we can use a symbol for the noble gas in the principal energy level (period) just before the element for part of the configuration. This is represented by the chemical symbol of the noble gas shown in brackets. Look at the example for chlorine below.

Chlorine is in the third period, principal energy level 3, or n = 3. The complete electron configuration for chlorine is:

$$\mathbf{1s^2 2s^2 2p^6 3s^2 3p^5}$$

The noble gas that occurs in the principal energy level above chlorine is neon. Neon's electron configuration is $1s^2 2s^2 2p^6$. This may be represented by the symbol [Ne]. Therefore, the electron configuration for chlorine using the noble gas core is:

$$\mathbf{[Ne]3s^2 3p^5}$$

Explore Student Guide

Part III: Writing Electron Configurations, continued

Remember, elements like to have the most stable electron configuration possible. This means that the electrons will fill a sublevel to maximize stability. There are many exceptions to the electron configuration rules that you have learned. We encounter the first two exceptions in the third principal energy level. See the complete electron configurations for chromium (Cr) and copper (Cu) below. For both of these exceptions, the 4s subshell and the 3d subshell are nearly identical in energy (i.e., almost the same distance from the nucleus, with 3d being just a little higher in energy), but the 3d orbitals are larger in size than 4s. Moving an electron from 4s to 3d requires energy, but the decrease in the electron to electron repulsion (since the electrons are farther away from each other in the 3d orbitals) more than compensates for the energy needed to move the electron.

$$\mathbf{Cr = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5}$$
$$\mathbf{Cu = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}}$$

Part IV (Extension): Lewis Valence Electron Dot Structures and Electron Configurations

Valence electrons are the electrons in the outermost (highest energy level) s and p subshells. The s and p subshells must be of the same principal energy level. Since the valence electrons are the highest occupied s and p subshell, the maximum number of valence electrons for any element is eight.

Valence electrons are important in chemistry as they are the electrons involved in bond formation (the connections between two or more elements that make up a substance) and are in large part responsible for periodic trends such as ionization energy and atomic radius.

In 1916, Gilbert Lewis introduced the concept of dots drawn around the symbol of an element to represent the valence electrons. Just as in the noble gas core electron configuration where the noble gas's symbol represents the electron configuration up to that element, the symbol of the element in a Lewis valence electron dot structure represents all of the electrons in that element except for the valence electrons. Lewis valence electron dot structures are commonly known as electron dot structures (or diagrams) or Lewis dot diagrams (or structures).

Let's see how the three ways to represent chlorine relate to each other.

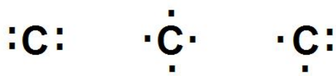
$\mathbf{1s^2 2s^2 2p^6 3s^2 3p^5}$	$\mathbf{[Ne]3s^2 3p^5}$	
electron configuration	noble gas core	Lewis dot diagram

Explore Student Guide

Part IV (Extension): Lewis Valence Electron Dot Structures and Electron Configurations, continued

There will be a maximum of two dots on each side of the element's symbol. It does not matter how they are arranged. For carbon, each of the following would be the correct Lewis dot diagram and each one is used in chemistry, depending on the problem.

Lewis dot diagrams will be used extensively when you study bonding and the valence shell electron pair repulsion theory, also known as the VSEPR theory.



Part V: Effects of Chemical Properties From Configurations

Lithium has an electron configuration of $1s^2 2s^1$. For any element to be stable, it must have a full outermost shell. Lithium has one electron in its second energy level, and therefore has two options.

- Gain seven more electrons to fill energy level 2.
- Lose one electron to empty energy level 2, and have a full level 1.

Lithium will choose the most energetically favorable option. This will be to lose the one electron it has in level 2. This works because level 1 already has a full shell of two electrons.