Teacher Guide: Electron Configuration

Learning Objectives

Students will …

- Know the relationship between the atomic number and the number of electrons in a neutral atom.
- Understand what is meant by *shell*, *subshell*, and *orbital*.
- Determine the electron configuration of various elements.
- Explain why atomic radii decrease across a period and increase down a group.

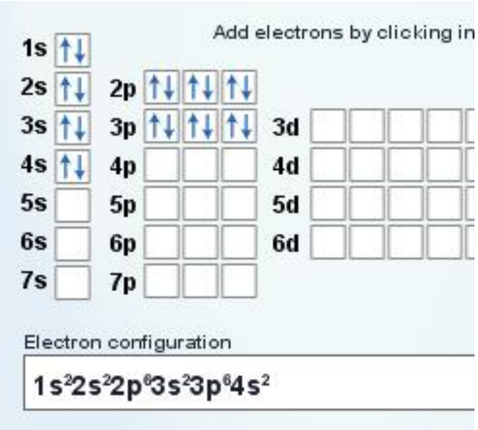
Vocabulary

atomic number, atomic radius, Aufbau principle, chemical family, diagonal rule, electron configuration, Hund's rule, orbital, Pauli exclusion principle, period, shell, spin, subshell

Lesson Overview

The way that electrons are arranged around the nuclei of atoms is one of the most important topics in chemistry. Electron configurations form the basis of chemical families, determine how elements react to form bonds, and explain some physical properties such as conductivity as well.

The *Electron Configuration* Gizmo™ allows students to discover how electrons are arranged in an atom of any element.



Electron configuration

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

The Student Exploration sheet contains three activities:

- Activity A – Students find the electron configurations of the first 10 elements in the periodic table.
- Activity B – Students explore patterns in atomic radii across a period and within a chemical family.
- Activity C – Students learn how to determine the electron configuration of any element.

Suggested Lesson Sequence

1. **Pre-Gizmo activity** (30 – 45 minutes)

   Before using the *Electron Configuration* Gizmo, have students go through the *Element Builder* Gizmo. This Gizmo allows students to construct the first 18 elements in the periodic table by adding protons, neutrons, and electrons to an atom.

   While doing this activity, students will see that electrons in these elements are arranged in three levels. The first level holds two electrons, and the second and third levels each hold eight. Seeing this pattern will help students understand the electron configurations for these elements when they do the first two activities of the *Electron Configuration* Student Exploration sheet.

2. **Prior to using the Gizmo** (10 – 15 minutes)
   Before students are at the computers, pass out the Student Exploration sheets and ask students to complete the Prior Knowledge Questions. Discuss student answers as a class, but do not provide correct answers at this point. Afterwards, if possible, use a projector to introduce the Gizmo and demonstrate its basic operations. Demonstrate how to take a screenshot and paste the image into a blank document.

3. **Gizmo activities** (15 – 20 minutes per activity)
   Assign students to computers. Students can work individually or in small groups. Ask students to work through the activities in the Student Exploration using the Gizmo. Alternatively, you can use a projector and do the Exploration as a teacher-led activity.

4. **Discussion questions** (15 – 30 minutes)
   As students are working or just after they are done, discuss the following questions:

   - Look at the first four shells. What is the relationship between the shell number and the number of subshells it contains? [The number of subshells is equal to the shell number. The fifth, sixth, and seventh shells have room for more subshells, but there are no elements with enough electrons to fill them.]
   - How many orbitals are there in the *s*, *p*, *d*, and *f* subshells? Is there a pattern?
   - How are the *s*, *p*, *d*, and *f* subshells shown in the periodic table?
   - Metals form bonds when they lose their outermost electrons. How could losing outer electrons lead to a stable electron configuration?

5. **Follow-up activities** (variable)
   There are several possible follow-up activities to the *Electron Configuration* Gizmo. First, hold a configuration contest with fun prizes. Check that all students are starting with a blank screen. Announce the name of an element; the first person to get the correct configuration stands up. Confirm the configuration and pass out the prize!

   Another natural follow-up to understanding electron configurations is to move on to chemical bonding. The *Ionic Bonds* and *Covalent Bonds* Gizmos address these types of chemical bonds.

   The key clue that led to the concept of shells and subshells was the spectrum of hydrogen. You can explore this connection with the *Bohr Model: Introduction* and *Bohr Model of Hydrogen* Gizmos. You can also demonstrate the spectra of various elements in class. (See the “Flame tests” link in the **Selected Web Resources** below.)

Scientific Background

While many images of atoms show electrons orbiting the nucleus in simple circles, the reality is much more complex. Electrons move in specific *shells*, or energy levels. Each shell is divided into *subshells*, and each subshell contains several orbitals. Orbitals hold up to two electrons.

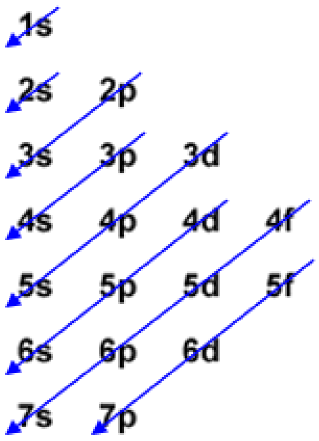
The subshells are designated by the letters *s*, *p*, *d*, and *f*. The *s* subshell contains one orbital and can hold two electrons. The *p* subshell contains three orbitals and can hold up to six electrons. The *d* subshell has five orbitals and can hold 10 electrons, and the *f* subshell has seven orbitals and can hold up to 14 electrons. Other subshells exist in theory but no atoms have enough electrons to fill them. The number of the shell determines how many subshells are present. The first shell has one subshell (1*s*), the second shell has two subshells (2*s*, 2*p*), and so on. The number of electrons in each shell is summarized below:

Shell	**Subshell**				**Total electrons**
	s	***p***	***d***	***f***	
1	2	—	—	—	2
2	2	6	—	—	8
3	2	6	10	—	18
4	2	6	10	14	32

The *Aufbau principle* states that electrons will occupy the lowest-energy orbital that is available. However, there is overlap between the energies of subshells in different shells, so the order of filling is complex. For example, 4*s* has lower energy than 3*d*, and 6*s* has lower energy than 5*p* and 4*d.* The order of filling subshells is described by the *diagonal rule*, shown at right. Follow each arrow from upper right to lower left, and then move to the top of the next arrow to the right. So the order is as follows:

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, etc.

When there are several available orbitals within a subshell, electrons will go to an unoccupied orbital if available. In this way, electrons resemble passengers getting on a bus: Usually passengers will sit in empty seats before they start to double up.



Diagonal rule

The position of every electron can be described by four *quantum numbers*. The first number gives the shell, the next the subshell, and the third number gives the orbital. The final quantum number describes the *spin* of the electron. Although electrons don't actually spin, they exhibit properties (such as magnetism) that are consistent with a clockwise or counterclockwise spin. The spin of the electron is represented by arrows pointing up (clockwise) or down (counter-clockwise). The *Pauli exclusion principle* states that no two electrons can have the same four quantum numbers. Therefore, if two electrons share an orbital, they must have opposite spins.

Selected Web Resources

Electron configurations: http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson36.htm
Spectral lines: http://www.colorado.edu/physics/2000/quantumzone/index.html
Elements as atoms: http://www.colorado.edu/physics/2000/elements_as_atoms/index.html
Introduction to quantum physics: http://www.faqs.org/docs/qp/chap01.html
Orbitals: http://www.chemguide.co.uk/atoms/properties/atomorbs.html
Flame tests: http://www.800mainstreet.com/spect/emission-flame-exp.html#Anchor-flame-head

Related Gizmos:
Element Builder: http://www.explorelearning.com/gizmo/id?424
Ionic Bonds: http://www.explorelearning.com/gizmo/id?514
Covalent Bonds: http://www.explorelearning.com/gizmo/id?512
Bohr Model: Introduction: http://www.explorelearning.com/gizmo/id?510
Bohr Model of Hydrogen: http://www.explorelearning.com/gizmo/id?506

Appendix: Development of the quantum atom

In the first half of the twentieth century, the standard atomic model underwent a rapid evolution from a “plum pudding” of positive material embedded with electrons to the complex quantum model that is accepted today.

The development of the modern atomic model began in 1909 when Ernest Rutherford conducted a key experiment. He bombarded a thin sheet of gold foil with positively charged particles called *alpha particles*. The vast majority of the alpha particles went straight through the foil, indicating that the gold atoms were made up mostly of empty space. But a few of the alpha particles bounced back from the gold foil, indicating the presence of a tiny, dense, positively charged *nucleus* within the atom. Rutherford hypothesized that the electrons orbited the nucleus like planets orbiting the Sun.

The next key piece of evidence came from *spectral lines*. When an element is heated, it emits colored light. When the light passes through a prism, its spectrum appears as a series of distinct spectral lines. The spectrum of hydrogen is shown below:

Photo by Jan Homann

In 1913, Niels Bohr explained spectral lines with a new atomic model. In this model, electrons orbit the nucleus at specific *energy levels*, or shells. When an electron moves from a higher shell to a lower shell, it emits a photon of light with a specific wavelength. The emitted wavelengths are seen as visible bands of light in the spectrum. Photons of the same wavelength could be absorbed by an electron and cause the electron to move from a lower to a higher shell.

During the 1920s, several experiments showed that tiny particles such as electrons could often exhibit wave-like behavior. For example, if you pass light through two small slits, you will see an *interference pattern* caused by the interaction of the light waves. Physicists found that the same interference patterns could be seen when electrons passed through two slits. Amazingly, the interference patterns still showed up even if the electrons passed through the slits one at a time!

This odd behavior, known as *wave-particle duality*, provided an explanation of Bohr’s energy levels. Louis de Broglie, a French physicist, determined that only orbits that were a multiple of the wavelength were allowed. Other orbit sizes would result in destructive interference and thus would not be stable.

In time, the full structure of the atom was determined with contributions from Werner Heisenberg, Erwin Schrödinger, Max Born, and other pioneers of quantum mechanics. One of their most surprising discoveries was that the position of an electron could not be described as a single point, but rather as a “probability cloud” that represented all of the possible locations of the electron. The probability cloud for each subshell has a unique shape, as shown below.