Electrons in Atoms

5.2 Electron Arrangement in Atoms

Connecting to Your World

Does this scene look natural to you? Surprisingly, it is. Arrangements like this are rare in nature because they are unstable. Unstable arrangements, whether the grains of sand in a sandcastle or the rock formation shown here, tend to become more stable by losing energy. If this rock were to tumble over, it would end up at a lower height. It would have less energy than before, but its position would be more stable.

Electron Configurations

Try to balance a pencil on its point. Each time you try, the pencil falls over. At the end of its fall, its energy has decreased. In most natural phenomena, change proceeds toward the lowest possible energy. In an atom, electrons and the nucleus interact to make the most stable arrangement possible. The ways in which electrons are arranged in various orbitals around the nuclei of atoms are called electron configurations.

Three rules—the aufbau principle, the Pauli exclusion principle, and Hund’s rule—tell you how to find the electron configurations of atoms. These three rules are as follows:

Aufbau Principle

According to the aufbau principle, electrons occupy the orbitals of lowest energy first. Look at the aufbau diagram in Figure 5.7. Each box represents an atomic orbital.

Figure 5.7 This aufbau diagram shows the energy levels of the various atomic orbitals. Orbitals of greater energy are higher on the diagram. Using Tables Which is of higher energy, a 4d or a 3s orbital?

Guide for Reading

Key Concepts

- What are the three rules for writing the electron configurations of elements?
- Why do actual electron configurations for some elements differ from those assigned using the aufbau principle?

Vocabulary

electron configurations
aufbau principle
Pauli exclusion principle
Hund’s rule

Reading Strategy

Building Vocabulary As you read the section, write a definition of each vocabulary term in your own words.

INSTRUCT

Because of gravity, rocks, such as the one shown, are less stable than rocks on level ground. Explain that electrons have a comparable ground level. Ask, What role do energy and stability play in the way that electrons are configured in an atom? (In the most stable atoms, electrons occupy the lowest possible energy orbitals.)

Electron Configurations

Use Visuals

Figure 5.7 Ask, Do all the orbitals at energy level n = 4 have the same energy? (No, 4p orbitals have higher energy than 4s, and 4d have higher energy than 4f.) Which has higher energy, 4s or 3d? (3d)

Answers to...

Figure 5.7 4d

Section Resources

Print

- Guided Reading and Study Workbook, Section 5.2
- Core Teaching Resources, Section 5.2
- Small-Scale Chemistry Laboratory Manual, Lab 7
- Transparencies, T58–T60

Technology

- Interactive Textbook with ChemASAP, Simulation 2, Problem-Solving 5.9, Assessment 5.2
- Go Online, Section 5.2
Section 5.2 (continued)

Go Online
NSTA SciLinks
Download a worksheet on Electron Configuration for students to complete, and find additional teacher support from NSTA SciLinks.

Discuss
Develop the electron configurations for several of the simpler elements. Introduce each rule governing the process as needed. Begin with hydrogen. Use the aufbau diagram to explain that electrons enter orbitals of lowest energy first. Show how the orbital notation \(1s^1\) describes the energy level, the orbital, and the number of electrons. Repeat the process for helium. Then continue with lithium, beryllium, and boron. Apply the Pauli exclusion principle to explain why additional orbitals must be used. When you reach carbon, explain and apply Hund’s rule. Complete the configurations for the second period elements.

Writing Electron Configurations

Purpose
Students gain practice in determining the correct order for filling orbitals.

Procedure
Have students work with a partner to develop the electron configurations for the third period elements. When the exercise is complete, ask students to compare the configurations of second and third period elements in preparation for future discussions of the periodic table.

Table 5.3
Electron Configurations for Some Selected Elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Orbital filling</th>
<th>Electron configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1s</td>
<td>1s^1</td>
</tr>
<tr>
<td>He</td>
<td>1s 2s</td>
<td>1s^2</td>
</tr>
<tr>
<td>Li</td>
<td>1s 2s 3s</td>
<td>1s^2 2s^1</td>
</tr>
<tr>
<td>C</td>
<td>1s 2s 2p</td>
<td>1s^2 2s^2p^2</td>
</tr>
<tr>
<td>N</td>
<td>1s 2s 2p 3s</td>
<td>1s^2 2s^2p^3</td>
</tr>
<tr>
<td>O</td>
<td>1s 2s 2p 3s</td>
<td>1s^2 2s^2p^4</td>
</tr>
<tr>
<td>F</td>
<td>1s 2s 2p 3s 4s</td>
<td>1s^2 2s^2p^5</td>
</tr>
<tr>
<td>Ne</td>
<td>1s 2s 2p 3s 4s</td>
<td>1s^2 2s^2p^6</td>
</tr>
<tr>
<td>Na</td>
<td>1s 2s 2p 3s 4s</td>
<td>1s^2 2s^2p^3s^1</td>
</tr>
</tbody>
</table>

Differentiated Instruction

Gifted and Talented
Students may be interested in learning that an electron in an orbital is described precisely by four quantum numbers, \(l\), \(m\), \(m_s\), and \(n\). (The azimuthal quantum number, \(l\), defines the shape of the orbital; the magnetic quantum number, \(m\), describes the orientation of the orbital in space; and the spin quantum number, \(m_s\), denotes the direction of electron spin, + or −.)
Look at the orbital filling diagrams of the atoms listed in Table 5.3. An oxygen atom contains eight electrons. The orbital of lowest energy, 1s, has one electron, then a second electron of opposite spin. The next orbital to fill is 2s. It also has one electron, then a second electron of opposite spin. One electron then occupies each of the three 2p orbitals of equal energy. The remaining electron now pairs with an electron occupying one of the 2p orbitals. The other two 2p orbitals remain only half filled, with one electron each.

A convenient shorthand method for showing the electron configuration of an atom involves writing the energy level and the symbol for every sublevel occupied by an electron. You indicate the number of electrons occupying that sublevel with a superscript. For hydrogen, with one electron in a 1s orbital, the electron configuration is written 1s. For helium, with two electrons in a 1s orbital, the configuration is 1s\(^2\). For oxygen, with two electrons in a 1s orbital, two electrons in a 2s orbital, and four electrons in 2p orbitals, it is 1s\(^2\)2s\(^2\)2p\(^4\). Note that the sum of the superscripts equals the number of electrons in the atom.

When the configurations are written, the sublevels within the same principal energy level are generally written together. This is not always the same order as given on the aufbau diagram. The 3d sublevel, for example, is written before the 4s sublevel, even though the aufbau diagram shows the 4s sublevel to have lower energy.

**CONCEPTUAL PROBLEM 5.1**

**Writing Electron Configurations**

Phosphorus, an element used in matches, has an atomic number of 15. Write the electron configuration of a phosphorus atom.

1. **Analyze** Identify the relevant concepts.
   Phosphorus has 15 electrons. There is a maximum of two electrons per orbital. Electrons do not pair up within an energy sublevel (orbitals of equal energy) until each orbital already has one electron.

2. **Solve** Apply concepts to this situation.
   Using Figure 5.7 on page 133, place electrons in the orbital with the lowest energy (1s) first, then continue placing electrons in each orbital with the next higher energy.

   The electron configuration of phosphorus is 1s\(^2\)2s\(^2\)2p\(^6\)2p\(^4\).
   The superscripts add up to the number of electrons. When the configurations are written, the sublevels within the same principal energy level are written together. This is not always the same order as given on the aufbau diagram.

**Practice Problems**

8. Write the electron configuration for each atom.
   a. carbon  b. argon  c. nickel

9. Write the electron configuration for each atom. How many unpaired electrons does each atom have?
   a. boron  b. silicon

**Answers**

8. a. 1s\(^2\)2s\(^2\)2p\(^2\)
   b. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)
   c. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)3d\(^8\)4s\(^2\)

9. a. 1s\(^2\)2s\(^2\)2p\(^1\); one unpaired electron
   b. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^2\); two unpaired electrons

**Practice Problems Plus**

What are the electron configurations for atoms of the following elements?

How many unpaired electrons does each atom have?

a. neon (1s\(^2\)2s\(^2\)2p\(^6\); no unpaired electrons)  b. sulfur (1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^4\); two unpaired electrons)

Students may find it helpful to use a diagonal diagram to write electron configurations until they become familiar with the order in which sublevels fill.

**Differentiated Instruction**

**Special Needs**

Encourage students to draw the aufbau diagram (Figure 5.7) on a 5 x 7 card and keep it handy when writing electron configurations. Also encourage them to make flash cards, each with the name and atomic number of an element on one side and its electron configuration on the other side. Pairs of students can use these to quiz each other.
Section 5.2 (continued)

Exceptional Electron Configurations

Discuss

Among the transition elements there are some exceptions to the filling rules. Exceptions can be explained by the atom’s tendency to keep its energy as low as possible. These exceptions help explain the unexpected chemical behavior of transition elements.

E ASSESS

Evaluate Understanding

Make a set of small cards, each with the symbol and atomic number of an element. Choose elements from throughout the periodic table. Have students choose cards and write the electron configurations for the elements.

Reteach

Have students work in small groups, using the set of element cards to practice writing more electron configurations. Have them refer to Figure 5.7, if necessary.

Check Activity

Students are likely to write that the two magnets would push each other apart. In the same way, electrons with the same spin would push apart and be unable to occupy the same orbital.

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 5.2.

Answers to...

Checkpoint

Exceptions to the aufbau principle are due to subtle electronic interactions in orbitals with similar energies.

5.2 Section Assessment

10. Key Concept: What are the three rules for writing the electron configuration of elements?

11. Key Concept: Explain why the actual electron configurations differ from those assigned using the aufbau principle.

12. Use Figure 5.7 to arrange the following sublevels in order of decreasing energy: 2p, 4s, 2s, 3p, and 4p.

13. Write a brief description of how trying to place two bar magnets pointing in the same direction alongside each other is like trying to place two electrons into the same orbital.

Exceptional Electron Configurations

Copper, shown in Figure 5.8, has an electron configuration that is an exception to the aufbau principle. You can obtain correct electron configurations for the elements up to vanadium (atomic number 23) by following the aufbau diagram for orbital filling. If you were to continue in that fashion, however, you would assign chromium and copper the following incorrect configurations:

\[
\begin{align*}
\text{Cr} & : 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 \\
\text{Cu} & : 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2
\end{align*}
\]

The correct electron configurations are as follows:

\[
\begin{align*}
\text{Cr} & : 1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1 \\
\text{Cu} & : 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1
\end{align*}
\]

These arrangements give chromium a half-filled d sublevel and copper a filled d sublevel. Filled energy sublevels are more stable than partially filled sublevels. Some actual electron configurations differ from those assigned using the aufbau principle because half-filled sublevels are not as stable as filled sublevels, but they are more stable than other configurations. This tendency overcomes the small difference between the energies of the 3d and 4s sublevels in copper and chromium.

Exceptions to the aufbau principle are due to subtle electron-electron interactions in orbitals with very similar energies. At higher principal quantum numbers, energy differences between some sublevels (such as 5s and 6s, for example) are even smaller than in the chromium and copper examples. As a result, there are other exceptions to the aufbau principle. Although it is worth knowing that exceptions to the aufbau principle occur, it is more important to understand the general rules for determining electron configurations in the many cases where the aufbau rule applies.

Figure 5.8 Copper is a good conductor of electricity and is commonly used in electrical wiring.

5.2 Section Assessment

10. aufbau principle, Pauli exclusion principle, Hund’s rule

11. Half-filled sublevels and filled sublevels are more stable than other configurations.

12. 3d, 4s, 3p, 3s, 2p

13. The 3s and 3p orbitals are already filled, so the last electron must go to the next higher energy orbital, which is 4s.
Atomic Emission Spectra

Purpose
To build a spectroscope and use it to measure the wavelengths, frequencies, and energies of atomic emission lines.

Materials
• cereal box
• diffraction grating
• tape
• ruler
• pencil
• scissors
• black construction paper
• white notebook paper
• spirit lamp (flashlight)
• fluorescent light

Procedure
1. Cut a square of black construction paper approximately 2 cm per side and tape a slit as shown. Cut a square of white notebook paper and tape the construction paper slit as shown. Cover the rest of the opening with white notebook paper. Place the construction paper slit as shown. Point the spectroscope toward a fluorescent light. Tape up any light leaks. Your lab hole (approximately 2 cm per side) and tape a diffraction grating across the opening with white notebook paper.
2. Use the cereal box to make a spectroscope for students to examine. Tape together two 2.0 cm × 10 cm strips of black construction paper so that they are parallel and form a narrow slit about 2 mm wide. Remove the top of a cereal box and tape the construction paper slit as shown. Cover the rest of the opening with white notebook paper. Cut a square hole (approximately 2 cm per side) and tape a diffraction grating over the hole as shown. Point the spectroscope toward a fluorescent light. Tape up any light leaks. Your lab partner should mark the exact positions of all the colored emission lines you see on the notebook paper. Measure the distance between the violet line and the other lines you have marked.

Analyze
1. List the number of distinct lines that you see as well as their colors.
2. Each line you see has a property called its wavelength. The prominent violet line has a wavelength of 436 nm and the prominent green line is 546 nm. How many nm apart are these lines on the paper? By how many nm do their wavelengths differ? How many nanometers of wavelength are represented by each millimeter you measured?
3. Using the nm/mm value you calculated in Step 2 and the nm distance you measured for each line from the violet reference line, calculate the wavelengths of all the other lines you see.
4. Each wavelength corresponds to another property of light called its frequency. Use the wavelength value of each line to calculate its frequency (v = c/λ) where c = 2.998 × 10^10 m/s = (2.998 × 10^10 m)/s.
5. The energy (E) of a quantum of light or atom emits is related to its frequency (ν) by E = h × ν. Use the frequency value for each line and h = 6.63 × 10^-34 J·s to calculate its corresponding energy.

You’re The Chemist
1. Design III Design and carry out an experiment to measure the longest and shortest wavelengths you can see in daylight. Use your spectroscope to observe light from a fluorescent light reflected off a white piece of paper. Caution: Do not look directly at the sun! Describe the differences in daylight and fluorescent light.
2. Design III Design and carry out an experiment to determine the effect of colored filters on the spectrum of fluorescent light or daylight. For each filter tell which colors are transmitted and which are absorbed.
3. Analyze III Use your spectroscope to observe various atomic emission discharge tubes provided by your teacher. Note and record the lines you see and measure their wavelengths.

Transmittance Table
Colors are transmitted and which are absorbed.

<table>
<thead>
<tr>
<th>Color</th>
<th>Transmitted</th>
<th>Absorbed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Violet</td>
<td>green, blue</td>
<td>yellow, orange, red</td>
</tr>
<tr>
<td>Yellow</td>
<td>green, red</td>
<td>blue, orange</td>
</tr>
<tr>
<td>Blue</td>
<td>red, green</td>
<td>yellow, orange</td>
</tr>
<tr>
<td>Red</td>
<td>green, yellow</td>
<td>blue, orange</td>
</tr>
</tbody>
</table>

Answers will vary. Provide students with gas discharge tubes containing hydrogen and noble gases.

Prep Time 10 minutes
Class Time 20 minutes
Teaching Tips
Have a completed spectroscope for students to examine. Have students bring in their own cereal boxes. Use holographic diffraction gratings of about 700 lines per millimeter. The diffraction grating needs to be positioned so that the lines on the grating are parallel to the slit. Very narrow slits do not let in enough light to see a bright spectrum.

Expected Outcome
Students calculate the wavelength, frequency, and energy of emissions from a fluorescent light.

Electrons in Atoms