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. In this section, you will learn that energy and stability play an important role in determining how electrons are configured in an atom.

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 the atom, electrons and the nucleus interact to make the most stable arrangement possible. The ways in which electrons are arranged into 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. The three rules are as follows.

Aufbau Principle  According to the aufbau principle, electrons occupy the orbitals of lowest energy first. Suppose you compare how electrons fill the orbitals in hydrogen with how they fill the orbitals in helium. Continue to examine how each atom's orbitals are filled compared with the atom that has one more proton and one more electron.

Using Tables Which is of higher energy, a 4d or a 5s orbital?

Vocabulary

aufbau principle
Pauli exclusion principle
Hund's rule

Reading Strategy

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

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?

Electron Configurations

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 \( 4p \).) Which has higher energy, \( 4s \) or \( 3d \)? (\( 3d \))
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

<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 2p</td>
<td>1s^2 2s^1</td>
</tr>
<tr>
<td>C</td>
<td>1s 2s 2p</td>
<td>1s^2 2s^2 2p^2</td>
</tr>
<tr>
<td>N</td>
<td>1s 2s 2p 3s</td>
<td>1s^2 2s^2 2p^3</td>
</tr>
<tr>
<td>O</td>
<td>1s 2s 2p</td>
<td>1s^2 2s^2 2p^4</td>
</tr>
<tr>
<td>F</td>
<td>1s 2s 2p</td>
<td>1s^2 2s^2 2p^5</td>
</tr>
<tr>
<td>Ne</td>
<td>1s 2s 2p</td>
<td>1s^2 2s^2 2p^6</td>
</tr>
<tr>
<td>Na</td>
<td>1s 2s 2p 3s</td>
<td>1s^2 2s^2 2p^6 3s^1</td>
</tr>
</tbody>
</table>

Look at Figure 5.7. Each box \(\square\) represents an atomic orbital. The various orbitals for any sublevel of a principal energy level are always of equal energy. Further, within a principal energy level the s sublevel is always the lowest-energy sublevel. Yet the range of energy levels within a principal energy level can overlap the energy levels of another principal level. Notice again in Figure 5.7 that the filling of atomic orbitals does not follow a simple pattern beyond the second energy level. For example, the 4s orbital is lower in energy than a 3d orbital.

**Pauli Exclusion Principle** According to the Pauli exclusion principle, an atomic orbital may describe at most two electrons. For example, either one or two electrons can occupy an s orbital or a p orbital. To occupy the same orbital, two electrons must have opposite spins; that is, the electron spins must be paired. Spin is a quantum mechanical property of electrons and may be thought of as clockwise or counterclockwise. A vertical arrow indicates an electron and its direction of spin (↑ or ↓). An orbital containing paired electrons is written as \(\uparrow\downarrow\).

**Hund’s Rule** When you use the aufbau diagram to decide how electrons occupy orbitals of equal energy, one electron enters each orbital until all the orbitals contain one electron with the same spin direction. Hund’s rule states that electrons occupy orbitals of the same energy in a way that makes the number of electrons with the same spin direction as large as possible. For example, three electrons would occupy three orbitals of equal energy as follows: \(\uparrow\downarrow\uparrow\). Second electrons then occupy each orbital so that their spins are paired with the first electron in the orbital. Thus each orbital can eventually have two electrons with paired spins.

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**Differentiated Instruction**

**Gifted and Talented** Students may be interested in learning that an electron in an orbital is described precisely by four quantum numbers, the first being the principle quantum number, \(n\). Direct them to a college text and ask them to name and explain the three other quantum numbers. \(l\), \(m\), and \(m_s\) denote the shape of the orbital; the orientation of the orbital in space; and 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\(^1\). 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 orbitals with the lowest energy (1s) first, then continue placing electrons in each orbital with the next higher energy.

   P 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^3\)

   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. In this case, the 3d sublevel is written before the lower-energy 4s sublevel.

**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\(^6\)
    b. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)
    c. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)

9.  a. 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^3\); 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.
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.

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.

Writing 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...

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

Section 5.2 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. 3s and 3p orbitals are already filled, so the last electron must go to the next higher energy orbital, which is 4s.

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:

- Cr: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s²
- Cu: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s²

The correct electron configurations are as follows:

- Cr: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹ 4p¹
- Cu: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹ 4p¹

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 5f and 6d, 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.
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
- pencil
- black construction paper
- white notebook paper

Procedure
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 slit about 2 mm wide. Remove the top of a cereal box and tape the construction paper so that they are parallel and form a narrow slit. Very narrow slits do not let in enough light to see a bright spectrum. Remind students to tape up stray light leaks.

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 mm apart are these lines on the paper? By how many mm 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 mm 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 given that $v = c/\lambda$, where $c = 2.998 \times 10^8$ m/s.
5. The energy (E) of a quantum of light an atom emits is related to its frequency (v) by $E = h \times v$. Use the frequency value for each line and $h = 6.63 \times 10^{-34}$ J·s to calculate its corresponding energy.

You’re The Chemist
1. Design It! 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 daylight 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 It! Design and carry out 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 It! 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.

v = 6.63 × 10^{-34} \text{ J·s} × 5.08 × 10^{14} \text{ s}^{-1} = 3.37 \times 10^{-19} \text{ J}

Sunlight has more red and more violet and lacks the distinct mercury emission lines of fluorescent light. Typical measurements are: Violet line to edge of red = 59 mm; 436 nm + (59 mm × 3.67 nm/mm) = 653 nm; Violet line to edge of violet = 4 mm. 436 nm – (4 × 3.67 nm/mm) = 421 nm

2. Translucent colored file tabs from an office supply store work well as filters. A yellow filter transmits green, yellow, orange and red and absorbs violet and blue. A green filter transmits blue, green and yellow and absorbs violet, orange and red. A blue filter transmits violet, blue and green and absorbs yellow, orange and red. A red filter transmits yellow, orange and red and absorbs violet, blue and green.

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

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

Analyse
1. Most fluorescent lights display 5 lines: violet, blue, green, yellow, and red. All fluorescent lights display violet, green and yellow. Older fluorescent lights will display only two distinct lines, one violet and one blue. A diffuse yellow line is also evident.
2. A typical cereal box will yield a measurement of 30 mm between the violet and green line: 546 – 436 = 110 nm; 110 nm/30 mm = 3.67 nm/mm
3. Blue: 436 nm + (16 mm × 3.67 nm/mm) = 436 + 95 = 531 nm; Green: 436 nm + (30 mm × 3.67 nm/mm) = 436 + 110 = 546 nm; Yellow: 436 nm + (42 × 3.67 nm/mm) = 436 + 154 = 590 nm; Red: 436 nm + (50 × 3.67 nm/mm) = 436 + 184 = 620 nm
4. Violet: v = 3.00 × 10^{17} \text{ nm/s} /436 nm = 6.88 × 10^{14} \text{ s}^{-1}; Blue: v = 3.00 × 10^{17} \text{ nm/s} /495 nm = 6.06 × 10^{14} \text{ s}^{-1}; Green: v = 3.00 × 10^{17} \text{ nm/s} /546 nm = 5.49 × 10^{14} \text{ s}^{-1}; Yellow: v = 3.00 × 10^{17} \text{ nm/s} /590 nm = 5.08 × 10^{14} \text{ s}^{-1}; Red: v = 3.00 × 10^{17} \text{ nm/s} /620 nm = 4.84 × 10^{14} \text{ s}^{-1}
5. Violet: v = 6.63 × 10^{-34} \text{ J·s} × 6.88 × 10^{14} \text{ s}^{-1} = 4.56 × 10^{-19} \text{ J}; Blue: v = 6.63 × 10^{-34} \text{ J·s} × 6.06 × 10^{14} \text{ s}^{-1} = 4.02 × 10^{-19} \text{ J}; Green: v = 6.63 × 10^{-34} \text{ J·s} × 5.49 × 10^{14} \text{ s}^{-1} = 3.64 × 10^{-19} \text{ J}; Yellow: