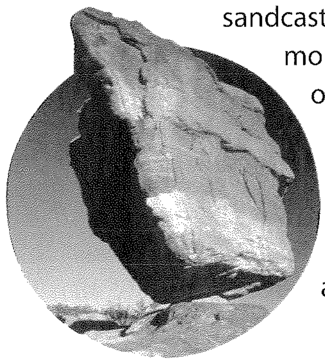
5.2 Electron Arrangement in Atoms

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.



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.

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. The 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.

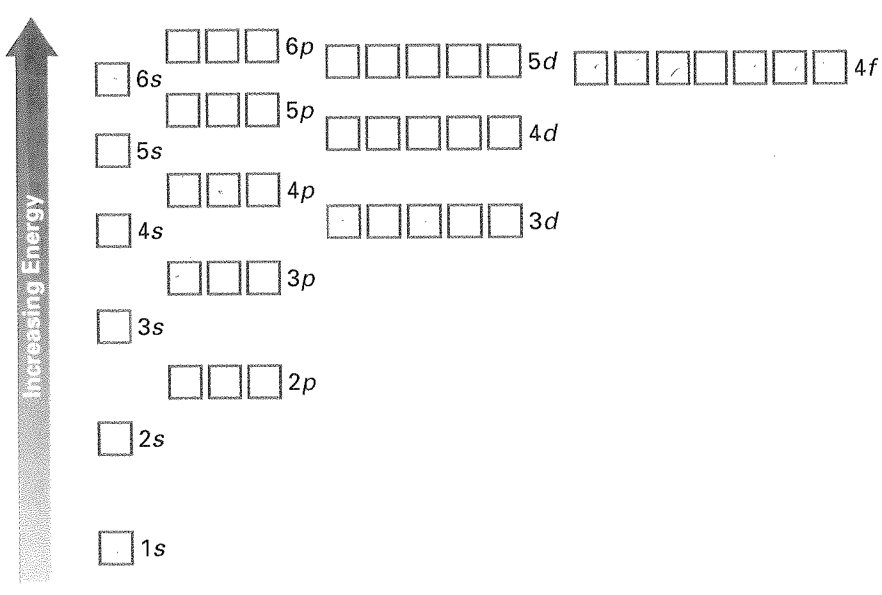


Figure 5.7 This aufbau diagram shows the energy levels of the various atomic orbitals. Orbitals of greater energy are higher on the diagram. **Using Tables** ***Which is of higher energy, a 4d or a 5s orbital?***

The 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 4*s* orbital is lower in energy than a 3*d* 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 $\boxed{\uparrow\downarrow}$.

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: $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$. 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.

Table 5.3

Electron Configurations for Some Selected Elements

Element	Orbital filling 1*s*	2*s*	$2p_x$	$2p_y$	$2p_z$	3*s*	Electron configuration
H	↑						$1s^1$
He	↑↓						$1s^2$
Li	↑↓	↑					$1s^22s^1$
C	↑↓	↑↓	↑	↑			$1s^22s^22p^2$
N	↑↓	↑↓	↑	↑	↑		$1s^22s^22p^3$
O	↑↓	↑↓	↑↓	↑	↑		$1s^22s^22p^4$
F	↑↓	↑↓	↑↓	↑↓	↑		$1s^22s^22p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^22s^22p^6$
Na	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^63s^1$

Interactive Textbook

Simulation 2 Fill atomic orbitals to build the ground state of several atoms.

with ChemASAP

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, 1*s*, has one electron, then a second electron of opposite spin. The next orbital to fill is 2*s*. It also has one electron, then a second electron of opposite spin. One electron then occupies each of the three 2*p* orbitals of equal energy. The remaining electron now pairs with an electron occupying one of the 2*p* orbitals. The other two 2*p* 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 1*s* orbital, the electron configuration is written $1s^1$. For helium, with two electrons in a 1*s* orbital, the configuration is $1s^2$. For oxygen, with two electrons in a 1*s* orbital, two electrons in a 2*s* orbital, and four electrons in 2*p* orbitals, it is $1s^2 2s^2 2p^4$. Note that the sum of the superscripts equals the number of electrons in the atom.

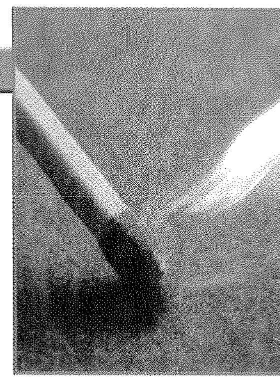
When the configurations are written, the sublevels within the same principal energy level are generally written together. This is not always the same order as given on the aufbau diagram. The 3*d* sublevel, for example, is written before the 4*s* sublevel, even though the aufbau diagram shows the 4*s*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

Go Online NSTA SCILINKS
For: Links on Electron Configuration
Visit: www.SciLinks.org
Web Code: cdn-1052

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 (1*s*) first, then continue placing electrons in each orbital with the next higher energy.

The electron configuration of phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$.
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.

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

Interactive Textbook
Problem Solving 5.9
Solve Problem 9 with the help of an interactive guided tutorial.
with ChemASAP

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

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:

$$\text{Cr } 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^4\, 4s^2$$
$$\text{Cu } 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^9\, 4s^2$$

The correct electron configurations are as follows:

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

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 3*d* and 4*s* 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 5*f* and 6*d*, 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.

Checkpoint ***How are energy differences and exceptions to the aufbau principle related?***

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 for some elements differ from those assigned using the aufbau principle.
12. Use Figure 5.7 to arrange the following sublevels in order of decreasing energy: 2*p*, 4*s*, 3*s*, 3*d*, and 3*p*.
13. Why does one electron in a potassium atom go into the fourth energy level instead of squeezing into the third energy level along with the eight already there?

Writing Activity

Modeling the Pauli Exclusion Principle 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.

Interactive Textbook

Assessment 5.2 Test yourself on the concepts in Section 5.2.

with ChemASAP