

UNIT 2

Structure and Properties

UNIT 2 CONTENTS

CHAPTER 3

Atoms, Electrons, and Periodic Trends

CHAPTER 4

Structure and Properties of Substances

UNIT 2 OVERALL EXPECTATIONS

- What is the quantum mechanical model of the atom, and how does a understanding of atomic structure enable chemists to explain the properties of substances and their chemical bonding?
- How do the properties of liquids and solids compare, and how can you use bonding theory to predict the shape of simple molecules?
- Which technologies have advanced our understanding of atomic and molecular theory, and how have those theories contributed to the development of other technologies and products?

Unit Project Prep

Look ahead to the project at the end of Unit 2. Plan in advance for the challenge by making a list of skills and information you will need as you progress through the unit.

In 1958, an astonished television audience watched as Dr. Harry Coover used a tiny drop of a substance called cyanoacrylate to lift a game show host off the ground. Cyanoacrylate is an amazingly powerful adhesive. This stunt helped launch the commercial career of superglue, now used for everything from building circuit boards to sealing human tissue in surgery.

Superglue is just one example of an astounding variety of compounds with useful properties. Whether synthesized by design or discovered by trial and error, each new compound does more than provide a potentially useful product. It also contributes to scientists' understanding of the relationship between structure at the molecular level and properties at the macroscopic level.

In this unit, you will investigate the structure and properties of atoms, ions, and molecules, and the natural and synthetic products that result from their interactions.





Atoms, Electrons, and Periodic Trends

Chapter Preview

- 3.1** The Nuclear Atomic Model
- 3.2** The Quantum Mechanical Model of the Atom
- 3.3** Electron Configurations and Periodic Trends

Prerequisite Concepts and Skills

Before you begin this chapter, review the following concepts and skills:

- identifying subatomic particles and their properties (from previous studies)
- describing the structure and organization of the periodic table (from previous studies)
- explaining periodic trends for properties such as atomic radius, first ionization energy, and electron affinity (from previous studies)

The northern lights, also called the *aurora borealis*, are a fleeting feature of the night sky in northern parts of the world. For centuries, people who saw the shimmering, dancing sheets of coloured lights were awed by the phenomenon, and tried to explain it. Inuit that lived near Hudson's Bay told that the lights were caused by torches held by spirits to lead souls of the dead to the afterlife. In Greenland and Labrador, people thought the lights could help predict the weather. In Europe in the Middle Ages, their appearance was an omen of impending war or plague.

Today, scientists explain the northern lights as resulting from streams of protons and electrons—plasma—emanating from the Sun. When the electrons from the plasma interact with gaseous atoms in Earth's upper atmosphere, the atoms emit coloured light.

Nineteenth-century atomic theory could not explain why the aurora exhibits a limited range of specific wavelengths, rather than the full visible spectrum. Early in the twentieth century, however, scientists developed a revolutionary new model of the atom. This model, and the theory that supports it, helped to account for many puzzling phenomena that the existing atomic theory had failed to explain. Among these phenomena are the characteristic colours of the northern lights.

In this chapter, you will learn about the developments that led to the modern model of the atom. You will also learn about the model itself and how it relates to periodic trends and the periodic table.

The green colours in the northern lights come from the interaction of accelerated electrons with oxygen atoms. The red colours usually come from nitrogen atoms. How do these interactions lead to atoms of different elements emitting light with characteristic wavelengths?



The Nuclear Atomic Model

3.1

Toward the close of the nineteenth century, chemists had two invaluable conceptual tools to aid them in their understanding of matter. The first was John Dalton's atomic theory, which you have studied intensively in previous chemistry courses. Dalton's atomic theory, first published in 1809, provided chemists with a framework for describing and explaining the behaviour of matter during chemical reactions. As you can see in Figure 3.1, the model of the atom that resulted from this theory was very simple.

The second conceptual tool was Dmitri Mendeleev's periodic table, which listed the known elements in order of increasing atomic mass. The resulting organizational chart arranged elements so that those with similar chemical properties were grouped together in the same column.

As you know, Dalton's atomic theory no longer applies in its original form, and Mendeleev's periodic table has undergone many changes. For example, scientists later discovered that atoms are not the most basic unit of matter because they *are* divisible. As well, the modern periodic table lists the elements in order of their atomic number, not their atomic mass. Of course, it also includes elements that had not been discovered in Mendeleev's time. Even so, in modified form, both of these inventions are still studied and used today in every chemistry course around the world.

When Mendeleev invented the periodic table, he was well-acquainted with Dalton's atomic theory. He knew nothing, however, about subatomic particles, and especially the electron, which is the foundation for the modern periodic table's distinctive shape. Because the original periodic table developed out of experimental observations, chemists did not need an understanding of atomic structure to develop it. (As you will see in section 3.3, however, the periodic table easily accommodates details about atomic structure. In fact, you will learn that the modern periodic table's distinctive design is a natural consequence of atomic structure.)

The First Step Toward the Modern Atomic Model

Chemists needed Dalton's atomic theory to advance their understanding of matter and its behaviour during chemical reactions. His atomic model, however, was inadequate for explaining the behaviour of substances. For example, Dalton designed a system of symbols to show how atoms combine to form other substances. Figure 3.2 on the next page shows several of these symbols. As you will no doubt notice, Dalton correctly predicted the formulas for carbon dioxide and sulfur trioxide, but ran into serious trouble with water, ammonia, and methane. Dalton's attempt at molecular modelling highlights a crucial limitation with his atomic model. Chemists could not use it to explain *why* atoms of elements combine in the ratios in which they do. This inability did not prevent chemists from pursuing their studies. It did, however, suggest the need for a more comprehensive atomic model.

Section Preview/ Specific Expectations

In this section, you will

- **describe** and **explain** the experimental observations that led to Rutherford's nuclear model of the atom
- **describe** and **explain** the modifications that Niels Bohr made to the nuclear model of the atom
- **distinguish** between a quantized model of the atom and the model proposed by Rutherford
- **communicate** your understanding of the following terms: *nuclear model, emission spectrum, absorption spectrum, quantum, photons*

CONCEPT CHECK

In your notebook, list the main ideas in Dalton's atomic theory. Explain how this theory enabled chemists to explain the three mass laws: the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

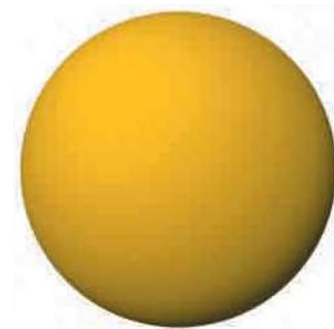


Figure 3.1 The model of the atom in 1809. The atom, as Dalton pictured it, was a tiny, solid, indestructible sphere.

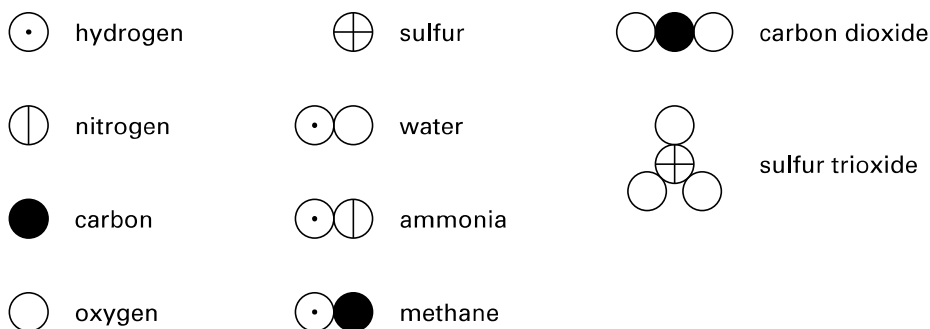


Figure 3.2 Examples of Dalton's system of symbols for atoms and molecules

The Discovery of the Electron Requires a New Atomic Model

In 1897, Dalton's idea of an indivisible atom was shattered with a startling announcement. A British scientist, Joseph John Thomson, had discovered the existence of a negatively charged particle with mass less than $\frac{1}{1000}$ that of a hydrogen atom. This particle was, of course, the electron. (Later calculations showed that the mass of an electron is $\frac{1}{1837}$ that of a hydrogen atom.) It took several years for chemists to consider the consequences of this discovery. They realized that if atoms contain electrons, atoms must also contain a positive charge of some kind to balance the negative charge. The atomic model that Thomson eventually proposed is shown in Figure 3.3. Keep in mind that scientists had not yet discovered the proton. Therefore, in this model, the *entire sphere* carries a uniform, positive charge.

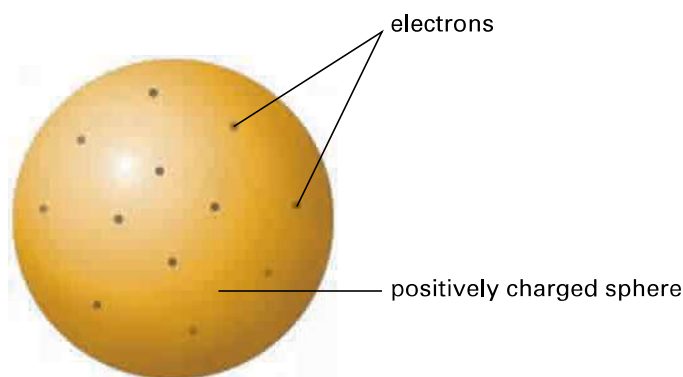


Figure 3.3 The atomic model in 1903. Thomson viewed the atom as a positively charged sphere embedded with sufficient numbers of electrons to balance (neutralize) the total charge.

Rutherford's Nuclear Model of the Atom

In the last four years of the nineteenth century, scientists in France—notably Henri Becquerel and Marie and Pierre Curie—discovered that certain elements are radioactive. That is, their atoms naturally emit positively charged particles (alpha particles), negatively charged particles (beta particles), and energy (gamma radiation).

Working in his laboratory at Montreal's McGill University, Ernest Rutherford studied the chemistry of radioactive elements intensively from 1898 to 1907. His efforts would lead to a Nobel Prize in Chemistry in 1908. Greater achievements lay ahead, however.



CHEM

FACT

Scientists initially described radioactivity solely in terms of radiation. The idea of radioactive *particles* first appeared around the turn of the twentieth century. In 1909, Ernest Rutherford reported confidently that the alpha particle was, in fact, a helium nucleus, ${}^4_2\text{He}$, with a 2+ charge. Scientists still had not discovered the proton by this time, so the nature of the helium nucleus (or any other atomic nucleus) was still unknown. In 1919, the existence of the proton was confirmed experimentally by, appropriately enough, Rutherford himself.

In 1909, two of Rutherford's students reported observations that cast doubts on Thomson's atomic model. As part of ongoing investigations into the nature and properties of radioactive emissions, they aimed alpha (α) particles at extremely thin metal foils. A small number of the alpha particles, about one in every 8000, were deflected significantly by the atoms that made up the metal foils. These observations were inconsistent with Thomson's atomic model. The researchers expected the alpha particles to pass through the metal atoms with, at most, deflections averaging $\frac{1}{200}$ of a degree. Therefore, deflections of 90° and more strained the credibility of Thomson's model. Rutherford encouraged further investigation. Either the observations and data were flawed, or Thomson's atomic model was.

As you may recall from earlier studies, the flaw lay in Thomson's model. In 1911, Rutherford published the results of the now-famous gold-foil experiment, shown in Figure 3.4. On the basis of this experiment, Rutherford suggested that the deflections he and his students observed were caused by an encounter between an alpha particle and an intense electric field at the centre of the atom.

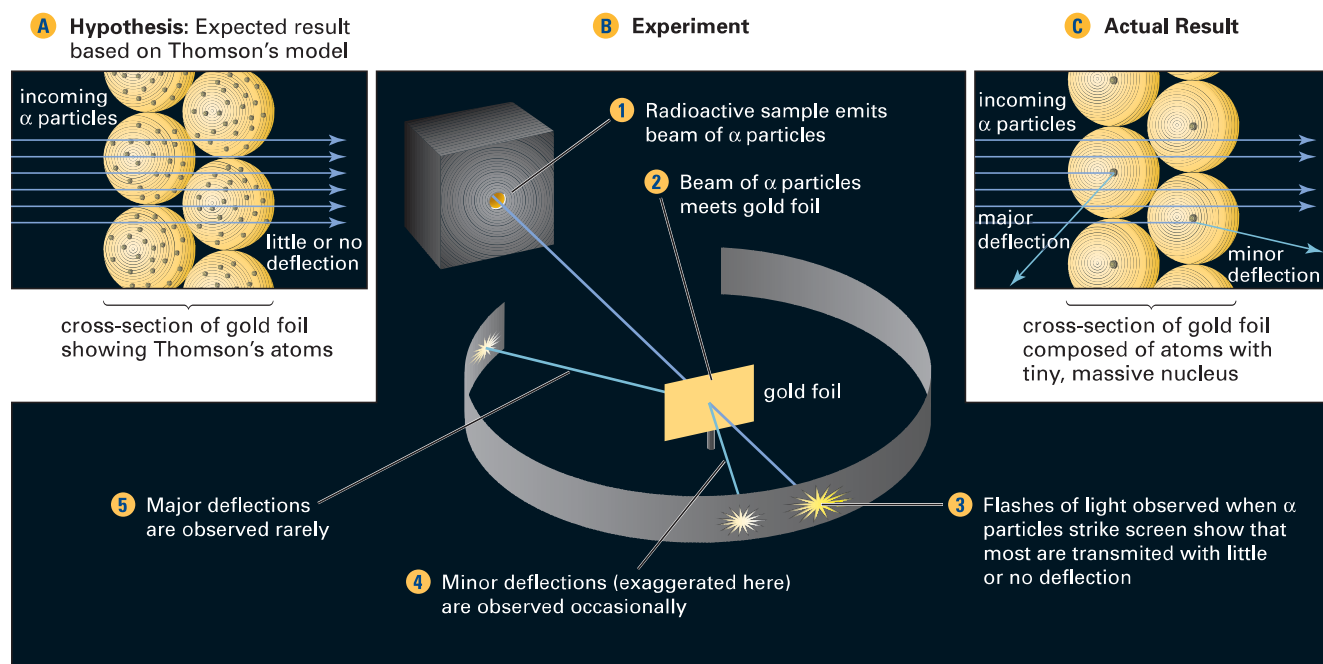


Figure 3.4 The hypothesis, experiment, and results of Rutherford's gold foil experiment. The experimental hypothesis and design owed much to the contributions of Rutherford's students, Hans Geiger (of Geiger-counter fame) and Ernest Marsden.

Rutherford performed several calculations that led him to an inescapable conclusion: the atom is made up mainly of empty space, with a small, massive region of concentrated charge at the centre. Soon afterward, the charge on this central region was determined to be positive, and was named the atomic nucleus. Because Rutherford's atomic model, shown in Figure 3.5 on the next page, pictures electrons in motion around an atomic nucleus, chemists often call this the **nuclear model** of the atom. You may also see it referred to as a planetary model because the electrons resemble the planets in motion around a central body.



CHEM

FACT

Contrary to common belief, Rutherford's 1911 paper on the discovery of the nucleus does *not* associate the nucleus with a positive charge. In fact, Rutherford acknowledged that the results of his experiment were independent of the sign of the charge. He suggested that the charge of the nucleus, as well as other questions raised by his paper, were subjects for later investigation. Sometime between 1911 and 1913, someone identified the charge on the nucleus as positive. Whether or not this person was Rutherford is not clear.

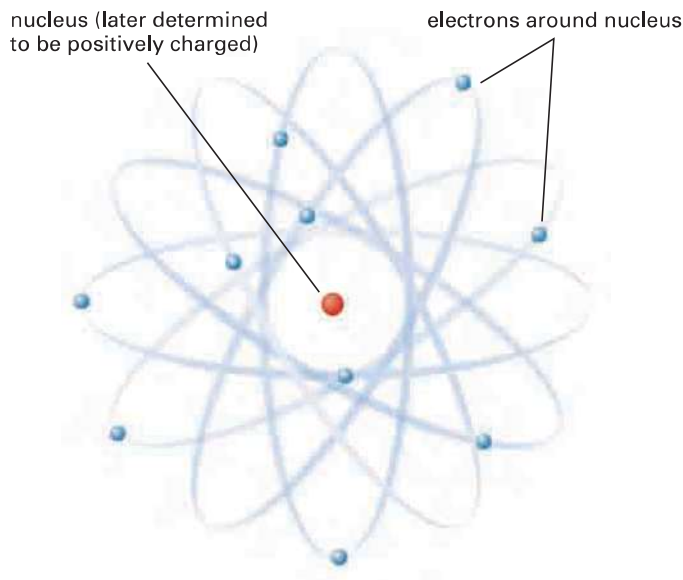


Figure 3.5 The atomic model in 1911. A Japanese scientist, Hantaro Nagaoka, proposed a similar, disk-shaped model with electrons orbiting a positively charged nucleus, in 1904. Rutherford notes in his 1911 paper that his results would be the same if Nagaoka's model were correct.

Prelude to a New Atomic Model

Rutherford's atomic model solved problems inherent in Thomson's atomic model, but it also raised others. For example, an atomic nucleus composed entirely of positive charges should fly apart due to electrostatic forces of repulsion. Furthermore, Rutherford's nuclear atom could not adequately explain the total mass of an atom. The discovery of the neutron, in 1932, eventually helped to settle these questions.

There was a more significant problem, however. Rutherford's atomic model seemed to contradict the laws of nineteenth-century physics. According to these assumptions, an electron in motion around a central body *must* continuously give off radiation. Consequently, one should be able to observe a continuous spectrum (a "rainbow") of light energy as the electron gives off its radiation.

Because the electron should also lose energy as a result of this radiation, the radius of its orbit should continuously decrease in size until it spirals into the nucleus. This predicted annihilation of the atom would occur in a fraction of a second. However, atoms were not seen to destabilize as predicted by this model.

Nineteenth-century physics could not explain why Rutherford's model corresponded to a stable atom. The solution to this problem marked a turning point in the history of chemistry and physics. To appreciate the scope of the problem, and the impact of its eventual solution, it is necessary to remain a short while longer in the nineteenth century. In the following pages, you will examine a phenomenon that scientists of the time could not explain.

The Problem of Atomic Spectra

You probably recall that visible light is part of a broader continuum of energy called the electromagnetic spectrum. Figure 3.6 reviews the key properties of the electromagnetic spectrum of energy.

Technology LINK

Chemists used spectral analysis during the nineteenth century to analyze substances and, sometimes, to discover new elements. Another common technique for analyzing substances, often used in conjunction with spectral analysis is a flame test. One of the foremost practitioners of this technique was a chemist named Robert Bunsen. Find out why he invented his famous burner. Carry on your research to investigate other ways that chemists use spectral analysis to examine the composition of substances. Select an appropriate medium to report your findings.

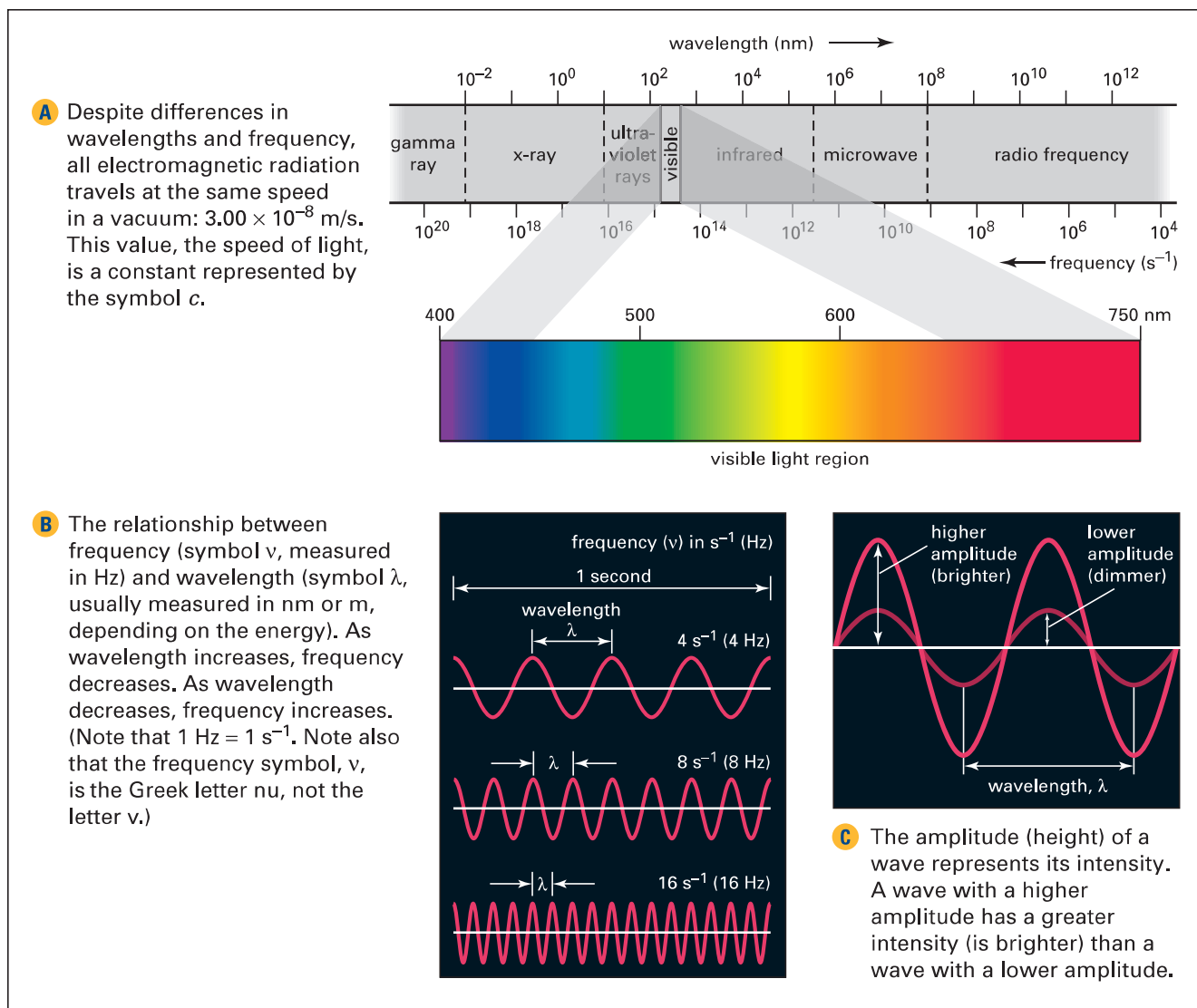


Figure 3.6 The electromagnetic spectrum and its properties

The visible portion of the electromagnetic spectrum is called a continuous spectrum, because the component colours are indistinct. They appear “smeared” together into a continuum of colour. According to nineteenth-century physics, part of the energy emitted by electrons *should* be observable as a continuous spectrum. This is not, however, the case. Instead, when atoms absorb energy (for example, when they are exposed to an electric current), you observe a pattern of discrete (distinct), coloured lines separated by spaces of varying length. See Figure 3.7. You can also observe this *line spectrum* for hydrogen, and for other atoms, in Investigation 3-A.



Figure 3.7 The discrete, coloured lines of this spectrum are characteristic of hydrogen atoms. No other atoms display this pattern of coloured lines.

Investigation **3-A**

Atomic Emission Spectra (Teacher Demonstration)

When a high voltage current is passed through a glass tube that contains hydrogen gas at low pressure, the gas glows with a pinkish-purple colour. If you observe this gas discharge tube through a spectroscope or a diffraction grating, you can see distinct, coloured lines. This type of spectrum is called a line spectrum or an emission spectrum. Chemists first observed spectra like this in the mid-1800s.

Question

Can the Rutherford model of the atom explain the emission spectra of elements?

Predictions

The Rutherford model of the atom described electrons as tiny masses in motion around a nucleus. According to this model, excited atoms should emit a continuous spectrum of light energy as their electrons give off radiation. In fact, according to this model, atoms should continually radiate energy whether they are excited or not. Read through the procedure and predict whether your observations will support or contradict the Rutherford model. Give reasons for your answer.

Safety Precautions



- A very high voltage is required to operate the gas discharge tubes. Do not come into contact with the source while viewing the tubes.
- Do not work with the gas discharge tubes yourself. Your teacher will demonstrate how it works.

Materials

gas discharge tube apparatus (teacher use only)
low pressure hydrogen gas tube and tubes for other available elements
spectroscope or diffraction grating

Procedure

1. Practise using the spectroscope or diffraction grating by observing an incandescent light bulb. Point the slit of the spectroscope toward the bulb and move the spectroscope until you can clearly see the spectrum.
2. Record the appearance of the spectrum from the incandescent bulb.
3. Your teacher will set up the gas discharge tube apparatus to demonstrate the emission spectra of hydrogen and other elements.
4. Observe the hydrogen discharge tube, and note its colour when high-voltage current is applied to it.
5. With the lights dimmed, examine the hydrogen emission spectrum with the aid of a spectroscope or diffraction grating. Make a careful sketch to record your observations. Note in particular the colours you see. Estimate their wavelengths. (Use the figure below to help you estimate the wavelengths.)
6. Follow steps 4 and 5 for the emission spectra produced by other gas discharge tubes that your teacher has available.



Analysis

1. According to Rutherford's model of the atom and nineteenth-century physics, electrons could emit or absorb any quantity of energy.
 - (a) Compare the spectrum you observed for the incandescent bulb to the spectrum that Rutherford's model predicts for hydrogen.
 - (b) Compare the spectrum you observed for hydrogen to the spectrum you observed for the incandescent bulb. Do your observations seem to support or contradict Rutherford's model? Explain your answer.
 - (c) How do you think Rutherford's model would have to be modified in order to account for your observations?
2. How do you think the lines you observed in the hydrogen emission spectrum relate to the energy of the electrons in a hydrogen atom? Include a diagram in your answer.
3. Compare the hydrogen emission spectrum to the other spectra you observed. Why do you think emission spectra are different for different elements?
4. Why do you think that scientists initially focused their attention on the hydrogen emission spectrum rather than the emission spectrum of another element?
5. Each element has a characteristic emission spectrum. Why is it only possible to observe the emission spectra of selected elements in a high school laboratory?

Conclusion

6. Reread your prediction and state whether it was accurate or not. Explain your answer in detail.

Applications

7. The emission spectrum of each element is characteristic of that element. In other words, it is a "fingerprint" that can be used to identify the element.
 - (a) Suggest reasons to explain why each element has a distinct and characteristic emission spectrum.
 - (b) Helium was discovered in the Sun before it was discovered on Earth. How do you think this was possible? Use print resources or the Internet to help you answer the question.
8. Examine the spectrum below. It is the *absorption spectrum* for sodium. Absorption spectra are produced when atoms absorb light of specific wavelengths.
 - (a) If you examined the emission spectrum for sodium in the laboratory, compare it to the absorption spectrum for sodium. (If you did not, look for a picture of the emission spectrum for sodium in a book or on the Internet.)
 - (b) In general, how do you think emission spectra compare to absorption spectra?
 - (c) Instruments based on absorption spectra are used extensively in analytical chemistry. One such instrument is an atomic absorption spectrophotometer. Use print and electronic resources to find out how the instrument works, and for what applications it is used.

Na



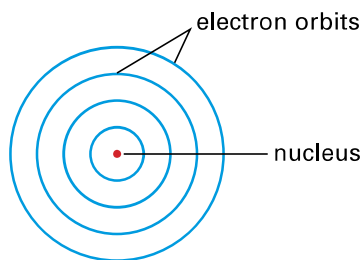


Figure 3.8 The model of the atom in 1913. Although Bohr pictured the atom in three dimensions, it is presented here, in two dimensions, to simplify some of Bohr's revolutionary concepts.

The Bohr Model of the Atom

Scientists of the nineteenth century lacked the concepts necessary to explain line spectra. Even in the first decade of the twentieth century, a suitable explanation proved elusive. This changed in 1913 when Niels Bohr, a Danish physicist and student of Rutherford, proposed a new model for the hydrogen atom. This model retained some of the features of Rutherford's model. More importantly, it was able to explain the line spectrum for hydrogen because it incorporated several new ideas about energy. As you can see in Figure 3.8, Bohr's atomic model pictures electrons in orbit around a central nucleus. Unlike Rutherford's model, however, in which electrons may move anywhere within the volume of space around the nucleus, Bohr's model imposes certain restrictions.

The Bohr Atomic Model

1. The atom has only specific, allowable energy levels, called stationary states. Each stationary state corresponds to the atom's electrons occupying fixed, circular orbits around the nucleus.
2. While in one of its stationary states, atoms do *not* emit energy.
3. An atom changes stationary states by emitting or absorbing a specific quantity of energy that is exactly equal to the difference in energy between the two stationary states.

On what basis could Bohr make these claims? In 1900, a physicist named Max Planck proposed an idea that was so revolutionary that he himself was unwilling to accept its implications. Planck suggested that *matter, at the atomic level, can absorb or emit only discrete quantities of energy*. Each of these specific quantities is called a **quantum** of energy. In other words, Planck said that the energy of an atom is *quantized*. Something that is quantized can exist only in certain discrete amounts. It is not continuous. In this sense, the rungs of a ladder are quantized, while the smooth slope of a ramp is not. Unlike ladders and ramps, however, a quantum is an extremely small "packet" of energy.

Although Planck said that the energy of matter is quantized, he continued to describe energy as travelling in the form of waves. He was, in fact, unwilling to consider that energy might have particle-like properties. In 1905, however, Albert Einstein was prepared to make just such an assertion. Light, according to Einstein, is also quantized. It occurs as quanta of electromagnetic energy that have particle-like properties. These particle-like "packets" of energy were later called **photons**. In Einstein's view, light (and, by extension, all electromagnetic energy) travels in the form of photons of energy. Light is emitted as photons of energy, and light is absorbed as photons of energy.

How Bohr's Atomic Model Explains the Spectrum for Atomic Hydrogen

If you re-read the key points of Bohr's atomic model above, you can see that he is applying the new ideas of quanta in his model. Bohr proposed that the energy that is emitted and absorbed by an atom must have specific values. The change in energy when an electron moves to higher or lower energy levels is not continuous. It is, rather, quantized.

Web LINK

www.mcgrawhill.ca/links/chemistry12

The concept of quantization enabled physicists to solve problems that nineteenth-century physics could not. One of these involved the thermal properties of solids when they are heated to incandescence. The other involved the induction of electrical current in metals when they are exposed to only specific frequencies of electromagnetic radiation. Use the Internet to investigate the phenomena known as blackbody radiation and the photoelectric effect. Write a brief report outlining the nature of these phenomena, why nineteenth-century physics could not explain them, and how the new idea of quantization could. Go to the web site above. Click on **Web Links** to find out where to go next.

Look again at the line spectrum for hydrogen shown in Figure 3.7. The energy that is associated with the coloured lines in this spectrum corresponds to the change in energy of an electron as it moves to higher or lower energy levels. For example, when a hydrogen atom is exposed to electrical current, or to another form of electromagnetic energy, its electron absorbs photons of energy. The atom is now said to be in an *excited state*. Another common way for atoms to become excited is through atomic collisions. For example, when two hydrogen atoms collide, some of the kinetic energy from one atom is transferred to the other atom. The electron of the second hydrogen atom absorbs this energy and is excited to a higher energy level.

When the electron of a hydrogen atom that has been excited to the third energy level falls to the second energy level, it emits light of certain energy. Specifically, an electron that makes a transition from the third energy level to the second energy level emits a photon of red light with a wavelength of 656 nm. Because line spectra result when atoms in an excited state *emit* photons as they fall to a lower energy level, these spectra are also called **emission spectra**.

Figure 3.9 shows the energy transitions that are responsible for the coloured lines in hydrogen's emission spectrum. Notice the use of the symbol n to designate the allowed energy levels for the hydrogen atom: $n = 1$, $n = 2$, and so on. This symbol, n , represents a positive integer (such as 1, 2, or 3), and is called a *quantum number*. You will learn more about the significance of quantum numbers in section 3.2.

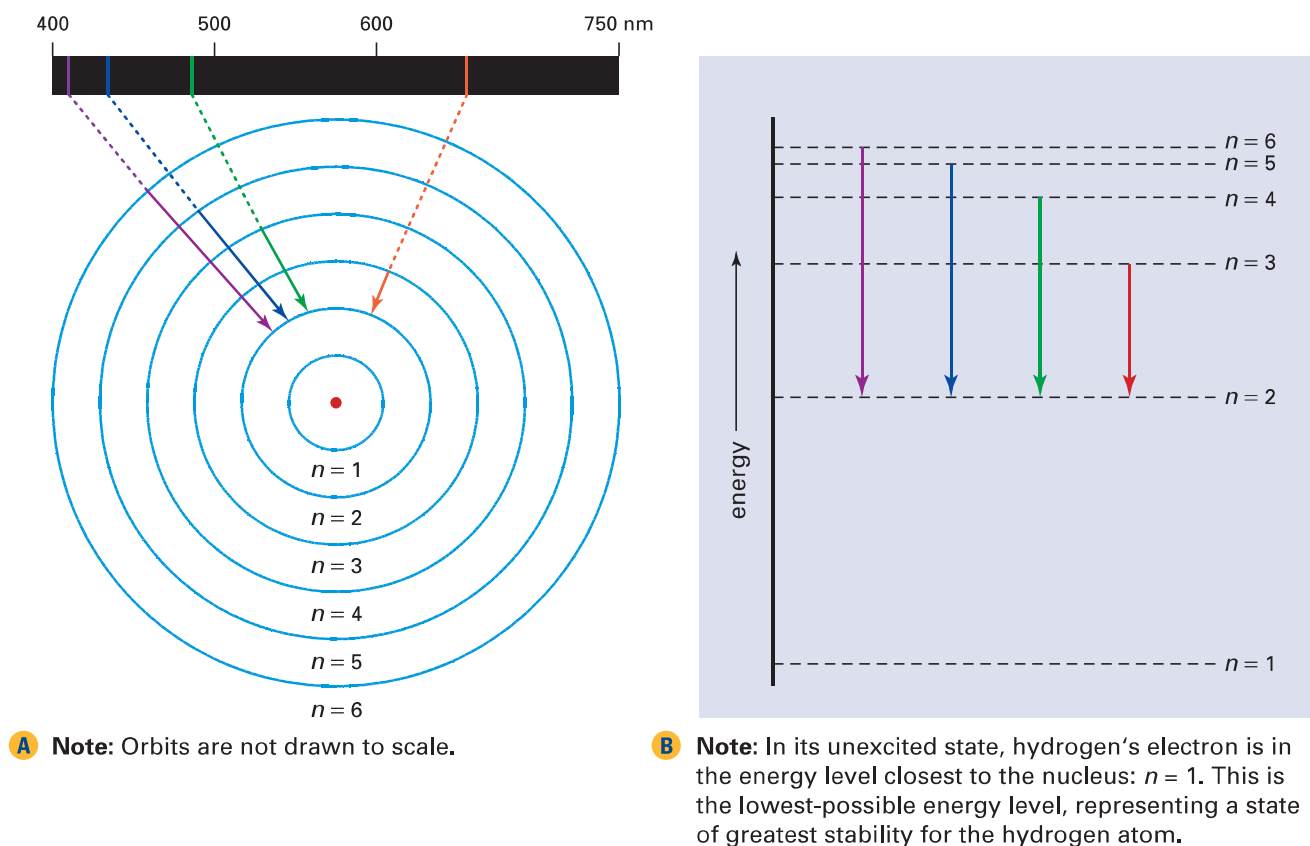


Figure 3.9 How the Bohr model explains the coloured lines in hydrogen's emission spectrum. When an excited electron falls from a higher energy level to a lower energy level (shown by the downward-pointing arrows), it emits a photon with a specific wavelength that corresponds to one of the coloured lines in the spectrum.

Another kind of spectrum that is related to an atom's emission spectrum is an **absorption spectrum**. This kind of spectrum results when electrons of atoms absorb photons of certain wavelengths, and so are excited from lower energy levels to higher energy levels. Figure 3.10 shows the absorption spectrum for hydrogen. As you can see, discrete, dark lines appear in the precise locations (wavelengths) as those of the coloured lines in hydrogen's emission spectrum.

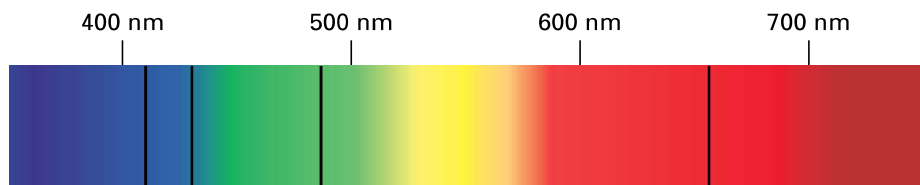


Figure 3.10 When light passes through gaseous hydrogen atoms, the electrons in the hydrogen atoms absorb photons of red, green, blue, and purple light at specific wavelengths. When the electrons return to lower energy levels, they re-emit photons with these same wavelengths. However, most of the photons are emitted in a different direction from the one in which the electrons absorbed the light. As a result, dark lines appear in the spectrum where photons of those particular wavelengths are absent.

The Successes and Limitations of the Bohr Atomic Model

Bohr's realization that the atom's energy is quantized—that electrons are restricted to specific energy levels (orbits)—was an astounding achievement. As you have seen, this model successfully predicted the coloured lines in the visible-light portion of hydrogen's emission spectrum. It also successfully predicted other lines, shown in Figure 3.11, that earlier chemists had discovered in the ultraviolet and infrared portions of hydrogen's emission spectrum.

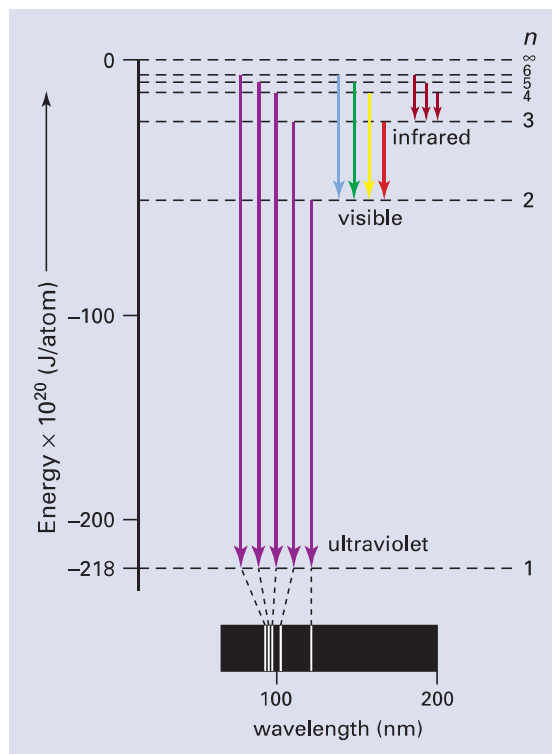


Figure 3.11 The ultraviolet series of lines appear as a result of electrons falling from higher energy levels to $n = 1$. The infrared lines appear as a result of electrons falling from higher energy levels to $n = 3$.

Math LINK

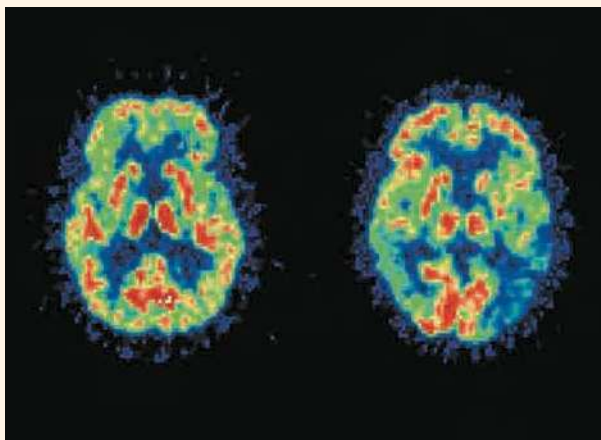
An atom has an infinite number of energy levels. In other words, n ranges from 1 to ∞ (infinity). What is the change in energy when a hydrogen atom is excited from $n = 1$ to $n = \infty$? Use the equation below to determine the value. Then consult the periodic table in Appendix C and examine hydrogen's atomic properties. To what property of hydrogen does this value correspond? Is this finding surprising to you? Explain why.

$$\begin{aligned}\Delta E &= E_{\text{final}} - E_{\text{initial}} \\ &= -2.18 \times 10^{-18} \text{ J} \\ &\quad \left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)\end{aligned}$$

There was a problem with Bohr's model, however. It successfully explained *only* one-electron systems. That is, it worked fine for hydrogen and for ions with only one electron, such as He^+ , Li^{2+} , and Be^{3+} . Bohr's model was unable, however, to explain the emission spectra produced by atoms with two or more electrons. Either Bohr's model was a coincidence, or it was an oversimplification in need of modification. Further investigation was in order.

Careers in Chemistry

Nuclear Medicine



This scan, a product of a nuclear medical technology called PET, shows changes in metabolic activity in different parts of the brain. The red colour indicates greater activity.

The discovery of the atomic nucleus affected all sciences, including biology and its applications in the field of medicine. During her graduate studies at the University of Western Ontario, Dr. Karen Goulenchyn excelled in theoretical mathematics, but switched to medicine because she was drawn to practical applications of science in people's lives. An interest in computers led her to nuclear medicine, which she now practices at the Civic Hospital in Ottawa.

Nuclear medicine is used chiefly in medical diagnosis. A radiopharmaceutical—a relatively harmless compound with a low dose of radiation—is swallowed or injected into the patient and tracked through the bloodstream by instruments such as a PET (positron emission tomography) camera. The nuclear physician can use the results to create a

three-dimensional computer image that evaluates the function as well as the structure of an organ. This procedure enables the physician to diagnose cancers and other tissue irregularities without the need for more invasive techniques such as exploratory surgery.

The dose of radiation that patients receive is similar to that of a diagnostic x-ray. However, because many patients are alarmed by the word "nuclear," Dr. Goulenchyn says that her team must explain the procedure tactfully "so that they don't run away." Some physicians ease fears by calling the discipline "molecular medicine."

Dr. Goulenchyn sees an increase of 6 to 12 percent of patients annually, primarily for suspected cancer or heart disease. Her greatest frustration is the time she must spend lobbying for new equipment, particularly for PET, a technique for measuring the concentrations of positron-emitting radioisotopes within the tissues. Positrons are a form of radiation identical to a beam of electrons, except that the charge of a positron is positive. PET scans can assess biochemical changes in the body, especially abnormal ones. These scans provide greater accuracy in determining whether a current or proposed therapy is effective.

A career in nuclear medicine, Dr. Goulenchyn notes, requires excellence in computers, science, and people skills, with a strong background in physiology. It also offers an unusual benefit for a physician: regular working hours. She appreciates the time this provides for her family, as well as for her work with the Canadian Association of Nuclear Medicine, for which she served as president from 1996 to 1998.

Section Summary

You have seen how scientists in the late nineteenth and early twentieth century developed and modified the atomic model. Changes in this model resulted from both experimental evidence and new ideas about the nature of matter and energy. By 1913, chemists and physicists had a working model that pointed tantalizingly in a promising direction. During the third decade of the twentieth century, the promise was fulfilled. In the next section, you will learn how physicists extended the ideas of Planck, Einstein, and Bohr to develop an entirely new branch of physics, and a new model of the atom.

Section Review

- 1 (a) **K/U** In what ways did Rutherford's model of the atom differ from Thomson's?
(b) On the basis of what evidence did Rutherford propose his model?
- 2 **K/U** Why did Rutherford's atomic model require modification? Be as specific as possible in your answer.
- 3 (a) **C** Imagine that you are Niels Bohr. Explain your ideas about the atom to scientists who are familiar with the work of Planck and Einstein.
(b) Imagine that you are a nineteenth-century scientist who has never heard of Planck and Einstein's ideas. You have just reviewed Bohr's proposed model for the atom. Write a paragraph describing your response to Bohr's model.
- 4 **C** List and arrange the types of electromagnetic radiation
(a) in order from longest wavelength to shortest wavelength
(b) in order from highest frequency to lowest frequency
- 5 **C** Compare visible light radiation to ultraviolet radiation. How are they different? How are they the same?
- 6 **I** Explain the significance, in terms of absorption or emission of photons, of the following statements.
(a) An electron moves from $n = 1$ to $n = 6$.
(b) An electron moves from $n = 5$ to $n = 2$.
(c) An electron moves from $n = 4$ to $n = 3$.
- 7 **C** Arrange the following in order from lowest energy to highest energy, and justify your sequence: $n = 7$, $n = 2$, $n = 5$, $n = 4$, $n = 1$.

The Quantum Mechanical Model of the Atom

3.2

You have seen how Bohr's model of the atom explains the emission spectrum of hydrogen. The emission spectra of other atoms, however, posed a problem. A mercury atom, for example, has many more electrons than a hydrogen atom. As you can see in Figure 3.12, mercury has more spectral lines than hydrogen does. The same is true for other many-electron atoms. Observations like these forced Bohr and other scientists to reconsider the nature of energy levels. The large spaces between the individual colours suggested that there are energy differences between individual energy levels, as stated in Bohr's model. The smaller spaces between coloured lines, however, suggested that there were smaller energy differences *within* energy levels. In other words, scientists hypothesized that there are *sublevels* within each energy level. Each of these sublevels has its own slightly different energy.

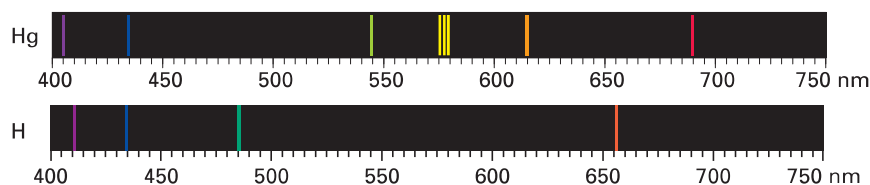


Figure 3.12 The emission spectrum for mercury shows that it has more spectral lines than the emission spectrum for hydrogen.

It was fairly straightforward to modify Bohr's model to include the idea of energy sublevels for the hydrogen spectrum and for atoms or ions with only one electron. There was a more fundamental problem, however. The model still could not explain the spectra produced by many-electron atoms. Therefore, a simple modification of Bohr's atomic model was not enough. The many-electron problem called for a new model to explain spectra of all types of atoms. However, this was not possible until another important property of matter was discovered.

The Discovery of Matter Waves

By the early 1920s, it was standard knowledge that energy had matter-like properties. In 1924, a young physics student named Louis de Broglie stated a hypothesis that followed from this idea. What if, de Broglie wondered, matter has wave-like properties?

He developed an equation that enabled him to calculate the wavelength associated with any object — large, small, or microscopic. For example, a baseball with a mass of 142 g and moving with a speed of 25.0 m/s has a wavelength of 2×10^{-34} m. Objects that you can see and interact with, such as a baseball, have wavelengths so small that they do not have any significant observable effect on the object's motion. However, for microscopic objects, such as electrons, the effect of wavelength on motion becomes very significant. For example, an electron moving at a speed of 5.9×10^6 m/s has a wavelength of 1×10^{-10} m. The size of this wavelength is greater than the size of the hydrogen atom to which it belongs. (The calculated atomic radius of hydrogen is 5.3×10^{-11} m.)

Section Preview/ Specific Expectations

In this section, you will

- **describe** the quantum mechanical model of the atom and its historical development
- **state** the meaning and significance of the first three quantum numbers
- **communicate** your understanding of the following terms: *quantum mechanical model of the atom, orbitals, ground state, principal quantum number (n), orbital-shape quantum number (l), magnetic quantum number (m_l)*

De Broglie's hypothesis of matter waves received experimental support in 1927. Researchers observed that streams of moving electrons produced diffraction patterns similar to those that are produced by waves of electromagnetic radiation. Since diffraction involves the transmission of waves through a material, the observation seemed to support the idea that electrons had wave-like properties.

The Quantum Mechanical Model of the Atom

In 1926, an Austrian physicist, Erwin Schrödinger, used mathematics and statistics to combine de Broglie's idea of matter waves and Einstein's idea of quantized energy particles (photons). Schrödinger's mathematical equations and their interpretations, together with another idea called Heisenberg's uncertainty principle (discussed below), resulted in the birth of the field of *quantum mechanics*. This is a branch of physics that uses mathematical equations to describe the wave properties of sub-microscopic particles such as electrons, atoms, and molecules. Schrödinger used concepts from quantum mechanics to propose a new atomic model: the **quantum mechanical model of the atom**. This model describes atoms as having certain allowed quantities of energy because of the wave-like properties of their electrons.

Figure 3.13 depicts the volume surrounding the nucleus of the atom as being indistinct or cloud-like because of a scientific principle called the *uncertainty principle*. The German physicist Werner Heisenberg proposed the uncertainty principle in 1927. Using mathematics, Heisenberg showed that it is impossible to know both the position and the momentum of an object beyond a certain measure of precision. (An object's momentum is a property given by its mass multiplied by its velocity.) According to this principle, if you can know an electron's precise position and path around the nucleus, as you would by defining its orbit, you cannot know with certainty its velocity. Similarly, if you know its precise velocity, you cannot know with certainty its position. Based on the uncertainty principle, Bohr's atomic model is flawed because you cannot assign fixed paths (orbits) to the motion of electrons.

Clearly, however, electrons exist. And they must exist *somewhere*. To describe where that "somewhere" is, scientists used an idea from a branch of mathematics called statistics. Although you cannot talk about electrons in terms of certainties, you *can* talk about them in terms of probabilities. Schrödinger used a type of equation called a wave equation to define the probability of finding an atom's electrons at a particular point within the atom. There are many solutions to this wave equation, and each solution represents a particular wave function. Each wave function gives information about an electron's energy and location within an atom. Chemists call these wave functions **orbitals**.

In further studies of chemistry and physics, you will learn that the wave functions that are solutions to the Schrödinger equation have no direct, physical meaning. They are mathematical ideas. However, the square of a wave function does have a physical meaning. It is a quantity that describes the probability that an electron is at a particular point within the atom at a particular time. The square of each wave function (orbital) can be used to plot three-dimensional probability distribution graphs for that orbital. These plots help chemists visualize the space in which electrons are most likely to be found around atoms. These plots are

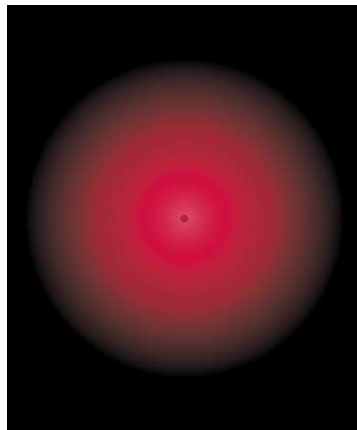


Figure 3.13 The model of the atom in 1927. The fuzzy, spherical region that surrounds the nucleus represents the volume in which electrons are most likely to be found.

also referred to as electron probability density graphs. Note: Although orbitals are wave functions without associated physical characteristics like shape and size, chemists often use the term orbitals when they *mean* three-dimensional probability distribution graphs. To simplify discussion, this textbook will discuss the size and shapes of orbitals. However, what is really meant is their associated probability distribution graph, which is calculated using the square of the wave function.

Each orbital has its own associated energy, and each represents information about where, inside the atom, the electrons would spend most of their time. Scientists cannot determine the actual paths of the moving electrons. However, orbitals indicate where there is a high probability of finding electrons.

Figure 3.14A represents the probability of finding an electron at any point in space when the electron is at the lowest energy level ($n = 1$) of a hydrogen atom. Where the density of the dots is greater, there is a higher probability of finding the electron. This graph is “fuzzy-looking” because the probability of finding the electron anywhere in the $n = 1$ energy level of a hydrogen atom is never zero. Farther from the nucleus, the probability becomes very small, but it will still never reach zero. Therefore, because the shape of the $n = 1$ orbital for hydrogen represents the level of probability of finding an electron, and since the probability never reaches zero, you have to select a “cut-off” level of probability. A level of probability is usually expressed as a percentage. Therefore, the contour line in Figure 3.14B defines an area that represents 95 percent of the probability graph. This two-dimensional shape is given three dimensions in Figure 3.14C. What this means is that, at any time, there is a 95 percent chance of finding the electron within the volume defined by the spherical contour.

CONCEPT CHECK

Distinguish clearly between an electron orbit, as depicted in Bohr's atomic model, and an electron orbital, as depicted in the quantum mechanical model of the atom.

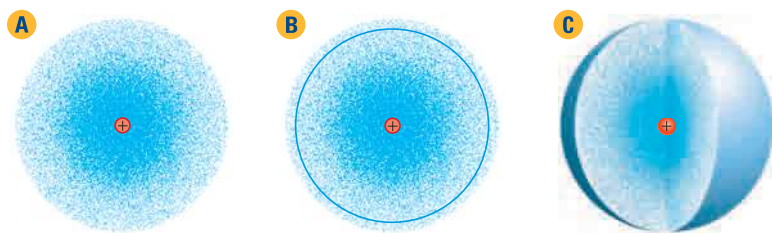


Figure 3.14 Electron density probability graphs for the lowest energy level in the hydrogen atom. These diagrams represent the probability of finding an electron at any point in this energy level.

Quantum Numbers and Orbitals

Figure 3.14 showed electron-density probabilities for the lowest energy level of the hydrogen atom. This is the most stable energy state for hydrogen, and is called the **ground state**. The quantum number, n , for a hydrogen atom in its ground state is 1. When $n = 1$ in the hydrogen atom, its electron is associated with an orbital that has a characteristic energy and shape. In an excited state, the electron is associated with a different orbital with its own characteristic energy and shape. This makes sense, because the electron has absorbed energy in its excited state, so its total energy increases and its motion changes. Figure 3.15 on the next page compares the sizes of hydrogen's atomic orbitals when the atom is in its ground state and when it is in an excited state.

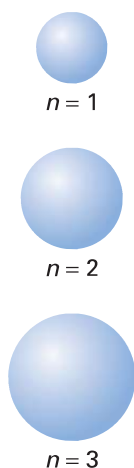


Figure 3.15 The relationship between orbital size and quantum number for the hydrogen atom. As n increases, the electron's energy increases and orbital size increases.

Orbitals have a variety of different possible shapes. Therefore, scientists use three quantum numbers to describe an atomic orbital. One quantum number, n , describes an orbital's energy level and size. A second quantum number, l , describes an orbital's shape. A third quantum number, m_l , describes an orbital's orientation in space. These three quantum numbers are described further below. The Concept Organizer that follows afterward summarizes this information. (In section 3.3, you will learn about a fourth quantum number, m_s , which is used to describe the electron inside an orbital.)

The First Quantum Number: Describing Orbital Energy Level and Size

The **principal quantum number (n)** is a positive whole number that specifies the energy level of an atomic orbital and its relative size. The value of n , therefore, may be 1, 2, 3, and so on. A higher value for n indicates a higher energy level. A higher n value also means that the size of the energy level is larger, with a higher probability of finding an electron farther from the nucleus. The greatest number of electrons that is possible in any energy level is $2n^2$.

The Second Quantum Number: Describing Orbital Shape

The second quantum number describes an orbital's shape, and is a positive integer that ranges in value from 0 to $(n - 1)$. Chemists use a variety of names for the second quantum number. For example, you may see it referred to as the angular momentum quantum number, the azimuthal quantum number, the secondary quantum number, or the orbital-shape quantum number.

Regardless of its name, the second quantum number refers to the energy sublevels within each principal energy level. The name that this book uses for the second quantum number is **orbital-shape quantum number (l)**, to help you remember that the value of l determines orbital shape. (You will see examples of orbital shapes near the end of this section.)

The value of n places precise limits on the value of l . Recall that l has a maximum value of $(n - 1)$. So, if $n = 1$, $l = 0$ (that is, $1 - 1$). If $n = 2$, l may be either 0 or 1. If $n = 3$, l may be either 0, 1, or 2. Notice that the number of possible values for l in a given energy level is the same as the value of n . In other words, if $n = 2$, then there are only two possible sublevels (two types of orbital shapes) at this energy level.

Each value for l is given a letter: s , p , d , or f .

- The $l = 0$ orbital has the letter s .
- The $l = 1$ orbital has the letter p .
- The $l = 2$ orbital has the letter d .
- The $l = 3$ orbital has the letter f .

To identify an energy sublevel (type of orbital), you combine the value of n with the letter of the orbital shape. For example, the sublevel with $n = 3$ and $l = 0$ is called the $3s$ sublevel. The sublevel with $n = 2$ and $l = 1$ is the $2p$ sublevel.

There are, in fact, additional sublevels beyond $l = 3$. However, for chemical systems known at this time, only the s , p , d , and f sublevels are required.

The Third Quantum Number: Describing Orbital Orientation

The **magnetic quantum number (m_l)** is an integer with values ranging from $-l$ to $+l$, including 0. This quantum number indicates the orientation of the orbital in the space around the nucleus. The value of m_l is limited by the value of l . If $l = 0$, m_l can be only 0. In other words, for a given value of n , there is only one orbital, of s type ($l = 0$). If $l = 1$, m_l may have one of three values: -1 , 0 , or $+1$. In other words, for a given value of n , there are three orbitals of p type ($l = 1$). Each of these p orbitals has the same shape and energy, but a different orientation around the nucleus. Notice that for any given value of l , there are $(2l + 1)$ values for m_l .

The total number of orbitals for any energy level n is given by n^2 . For example, if $n = 2$, it has a total of 4 (that is, 2^2) orbitals (an s orbital and three p orbitals). The Sample Problem below shows further use of this calculation.

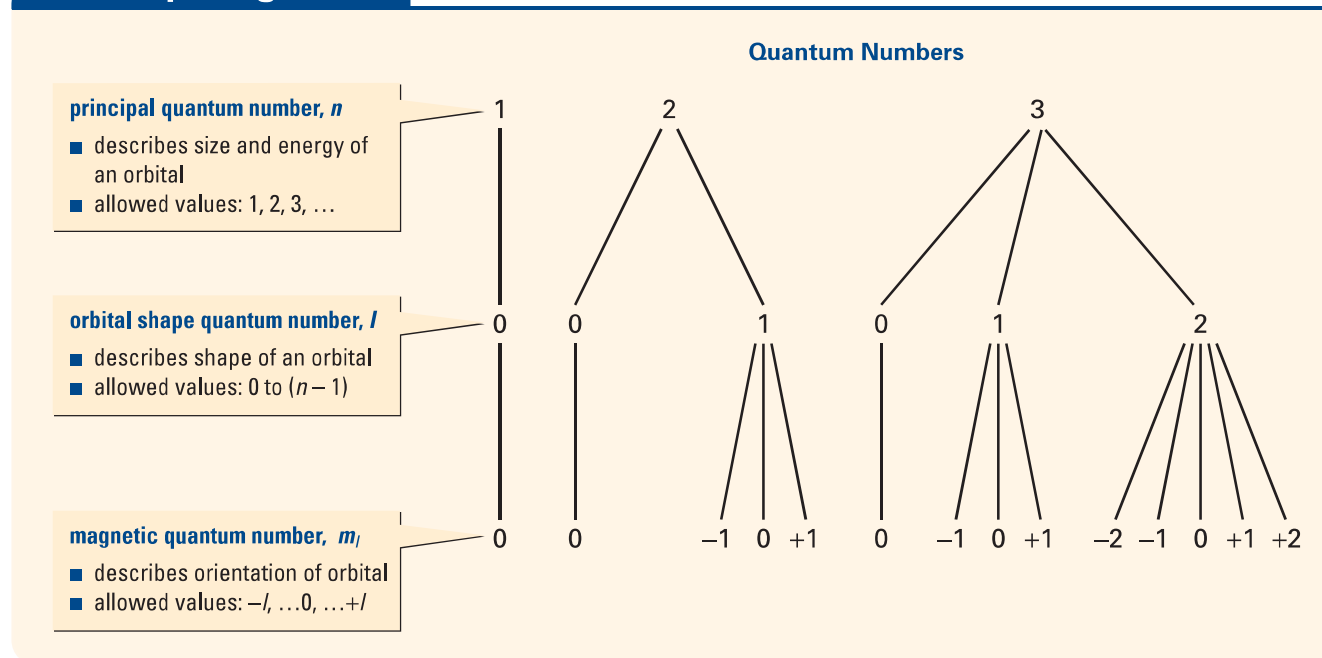


CHEM FACT

The letters used to represent energy sublevels are abbreviations of names that nineteenth-century chemists used to describe the coloured lines in emission spectra. These names are sharp, principal, diffuse, and fundamental.

Concept Organizer

The Relationship Among the First Three Quantum Numbers



Sample Problem

Determining Quantum Numbers

Problem

- If $n = 3$, what are the allowed values for l and m_l , and what is the total number of orbitals in this energy level?
- What are the possible values for m_l if $n = 5$ and $l = 1$? What kind of orbital is described by these quantum numbers? How many orbitals can be described by these quantum numbers?

Continued ...

Solution

- (a) The allowed values for l are integers ranging from 0 to $(n - 1)$. The allowed values for m_l are integers ranging from $-l$ to $+l$ including 0. Since each orbital has a single m_l value, the total number of values for m_l gives the number of orbitals.

To find l from n :

If $n = 3$, l may be either 0, 1, or 2.

To find m_l from l :

If $l = 0$, $m_l = 0$

If $l = 1$, m_l may be -1 , 0, $+1$

If $l = 2$, m_l may be -2 , -1 , 0, $+1$, $+2$

Since there are a total of 9 possible values for m_l , there are 9 orbitals when $n = 3$.

- (b) You determine the type of orbital by combining the value for n with the letter used to identify l . You can find possible values for m_l from l , and the total of the m_l values gives the number of orbitals.

To name the type of orbital:

$l = 1$, which describes a p orbital

Since $n = 5$, the quantum numbers represent a $5p$ orbital.

To find m_l from l :

If $l = 1$, m_l may be -1 , 0, $+1$

Therefore, there are 3 possible $5p$ orbitals.

Check Your Solution

- (a) Since the total number of orbitals for any given n is n^2 , when $n = 3$, the number of orbitals must be 9 (that is, 3^2).
- (b) The number m_l values is equivalent to $2l + 1$: $2(1) + 1 = 3$. Since the number of orbitals equals the number of m_l values, the answer of 3 must be correct.

Practice Problems

- What are the allowed values for l in each of the following cases?
 - $n = 5$
 - $n = 1$
- What are the allowed values for m_l , for an electron with the following quantum numbers:
 - $l = 4$
 - $l = 0$
- What are the names, m_l values, and total number of orbitals described by the following quantum numbers?
 - $n = 2$, $l = 0$
 - $n = 4$, $l = 3$
- Determine the n , l , and possible m_l values for an electron in the $2p$ orbital.
- Which of the following are allowable sets of quantum numbers for an atomic orbital? Explain your answer in each case.
 - $n = 4$, $l = 4$, $m_l = 0$
 - $n = 3$, $l = 2$, $m_l = 1$
 - $n = 2$, $l = 0$, $m_l = 0$
 - $n = 5$, $l = 3$, $m_l = -4$

Shapes of Orbitals

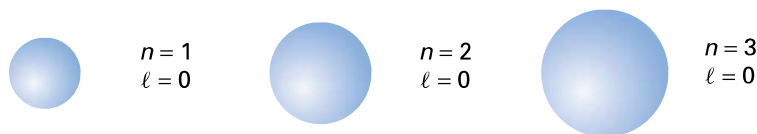
An orbital is associated with a size, a three-dimensional shape, and an orientation around the nucleus. Together, the size, shape, and position of an orbital represent the probability of finding a specific electron around the nucleus of an atom. Figure 3.16 shows the probability shapes associated with the s , p , and d orbitals. (The f orbitals have been omitted due to their complexity. You may study these orbitals in post-secondary school chemistry courses.)

Notice that the *overall shape* of an atom is a combination of all its orbitals. Thus, the overall shape of an atom is spherical. Be careful, however, to distinguish for yourself between the overall spherical shape of the atom, and the spherical shape that is characteristic of *only* the s orbitals.

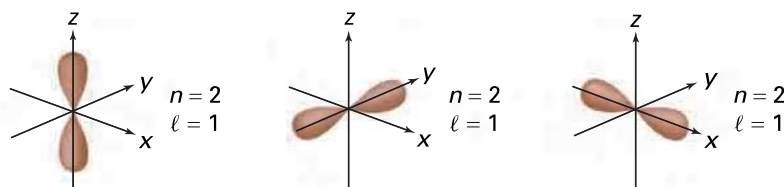
Finally, it is important to be clear about what orbitals are when you view diagrams such as those in Figure 3.16. Orbitals, remember, are solutions to mathematical equations. Those solutions, when manipulated, describe the motion and position of the electron in terms of probabilities. Contour diagrams, such as those shown here and in numerous print and electronic resources, *appear solid*. It therefore becomes easy to begin thinking about orbitals as physical “containers” that are “occupied” by electrons. In some ways, this is unavoidable. Try to remind yourself, now and then, of the following:

- Electrons have physical substance. They have a mass that can be measured, and trajectories that can be photographed. They exist in the physical universe.
- Orbitals are mathematical descriptions of electrons. They do not have measurable physical properties such as mass or temperature. They exist in the imagination.

s orbitals



p orbitals



d orbitals

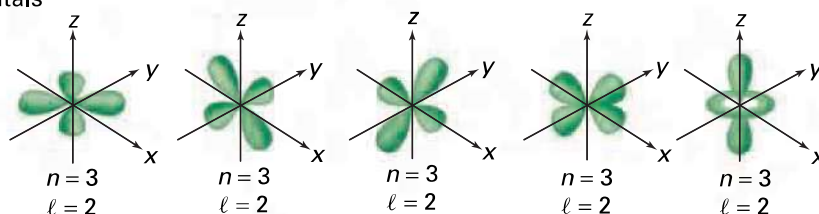


Figure 3.16 Shapes of the s , p , and d orbitals. Orbitals in the p and d sublevels are oriented along or between perpendicular x , y , and z axes.

Section Summary

In this section, you saw how the ideas of quantum mechanics led to a new, revolutionary atomic model—the quantum mechanical model of the atom. According to this model, electrons have both matter-like and wave-like properties. Their position and momentum cannot both be determined with certainty, so they must be described in terms of probabilities. An orbital represents a mathematical description of the volume of space in which an electron has a high probability of being found. You learned the first three quantum numbers that describe the size, energy, shape, and orientation of an orbital. In the next section, you will use quantum numbers to describe the total number of electrons in an atom and the energy levels in which they are most likely to be found in their ground state. You will also discover how the ideas of quantum mechanics explain the structure and organization of the periodic table.

Section Review

- 1 K/U** Explain how the quantum mechanical model of the atom differs from the atomic model that Bohr proposed.
- 2 K/U** List the first three quantum numbers, give their symbols, and identify the property of orbitals that each describes.
- 3 C** Design a chart that shows all the possible values of l and m_l for an electron with $n = 4$.
- 4 C** Agree or disagree with the following statement: The meaning of the quantum number n in Bohr's atomic model is identical to the meaning of the principal quantum number n in the quantum mechanical atomic model. Justify your opinion.
- 5 I** Identify any values that are incorrect in the following sets of quantum numbers.
 - $n = 1, l = 1, m_l = 0$; name: $1p$
 - $n = 4, l = 3, m_l = +1$; name: $4d$
 - $n = 3, l = 1, m_l = -2$; name: $3p$
- 6 I** Fill in the missing values in the following sets of quantum numbers.
 - $n = ?, l = ?, m_l = 0$; name: $4p$
 - $n = 2, l = 1, m_l = 0$; name: ?
 - $n = 3, l = 2, m_l = -2$; name: ?
 - $n = ?, l = ?, m_l = ?$; name: $2s$

Electron Configurations and Periodic Trends

3.3

For the hydrogen atom (and other single-electron systems), all orbitals that have the same value for n have the same energy. You can see this in the energy level diagram shown in Figure 3.17. Hydrogen's single electron is in its lowest (most stable) energy state when it is in the $1s$ orbital. In other words, the $1s$ orbital is the ground state for hydrogen. When a hydrogen atom is in an excited state, its electron may be found in any of its other orbitals, depending on the amount of energy absorbed.

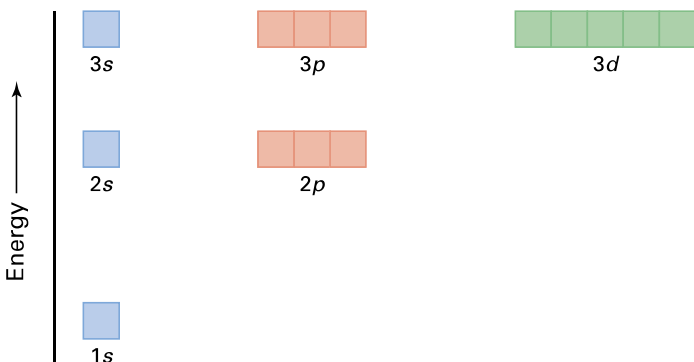


Figure 3.17 For the hydrogen atom, orbital energy depends only on the value of n . For example, all four $n = 2$ orbitals have the same energy. All nine $n = 3$ orbitals have the same energy. What must the value of l be for each of these orbitals?

Atoms with two or more electrons have orbitals with shapes that are similar to those of hydrogen. However, the interactions among the additional electrons result in orbitals with the same value of n , but different *energies*. For example, Figure 3.18 shows an energy level diagram for the lithium atom, which has three electrons. You can see that the $2s$ orbital has a lower energy than the $2p$ orbitals. Similarly, the $3p$ orbitals are lower in energy than the $3d$ orbitals. Notice, however, that all the orbitals within a sublevel have the same energy. For example, the three p orbitals in the $3p$ sublevel have the same energy.

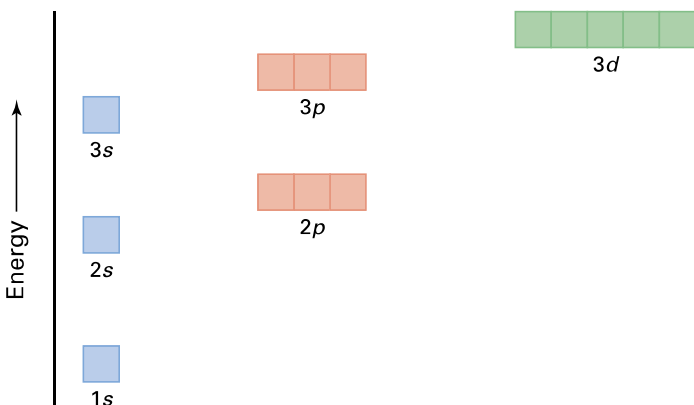


Figure 3.18 In the lithium atom, and for all other multi-electron atoms, orbitals in different energy sublevels differ in energy.

Section Preview/ Specific Expectations

In this section, you will

- **write** electron configurations and **draw** orbital diagrams for atoms of elements in the periodic table
- **explain** the significance of the exclusion principle and Hund's rule in writing electron configurations
- **list** characteristics of the s , p , d , and f blocks of elements in the periodic table
- **explain** the relationship between electron configurations and the position and properties of elements in the periodic table
- **communicate** your understanding of the following terms: *spin quantum number (m_s)*, *Pauli exclusion principle*, *electron configuration*, *aufbau principle*, *orbital diagram*, *Hund's rule*, *atomic radius*, *ionization energy*, *electron affinity*

Since low energy usually means stability, it is reasonable that the most stable energy state (ground state) for any atom is one in which its electrons are in the lowest possible energy level. As you know, this is the 1s orbital ($n = 1$). However, the electrons of most atoms are not “packed” into this orbital. In fact, experimental evidence shows that only two atoms have all their electrons in the 1s orbital: hydrogen and helium. To explain how and why this is the case, you must consider another property of the electron.

The Fourth Quantum Number: A Property of the Electron

As you learned from the previous section, three quantum numbers— n , l , and m_l —describe the energy, size, shape, and spatial orientation of an orbital. A fourth quantum number describes a property of the electron that results from its particle-like nature. Experimental evidence suggests that electrons spin about their axes as they move throughout the volume of their atoms. Like a tiny top, an electron can spin in one of two directions, each direction generating a magnetic field. The **spin quantum number** (m_s) specifies the direction in which the electron is spinning. This quantum number has only two possible values: $+\frac{1}{2}$ or $-\frac{1}{2}$.

In 1925, an Austrian physicist, Wolfgang Pauli, proposed that *only two electrons of opposite spin could occupy an orbital*. This proposal became known as the **Pauli exclusion principle**. What the exclusion principle does is place a limit on the total number of electrons that may occupy any orbital. That is, an orbital may have a maximum of two electrons only, each of which must have the opposite spin direction of the other. It may also have only one electron of either spin direction. An orbital may also have no electrons at all.

Summarizing the Four Quantum Numbers for Electrons in Atoms

Quantum Number Name	Symbol	Allowed Values	Property
principal	n	positive integers (1, 2, 3, etc.)	orbital size and energy
orbital-shape	l	integers from 0 to $(n - 1)$	orbital shape
magnetic	m_l	integers from $-l$ to $+l$	orbital orientation
spin	m_s	$+\frac{1}{2}$ or $-\frac{1}{2}$	electron spin direction

The Pauli Exclusion Principle and Quantum Numbers

Another way of stating the exclusion principle is that *no two electrons in an atom have the same four quantum numbers*. This important idea means that each electron in an atom has its own unique set of four quantum numbers. For example, compare the quantum numbers that distinguish a ground state hydrogen atom from a helium atom. (Recall that a helium atom has two electrons. Note also that m_s quantum number is given as $+\frac{1}{2}$. It could just as easily have a value of $-\frac{1}{2}$. By convention, chemists usually use the positive value first.)

Atom	Electron	Quantum numbers
Hydrogen	lone	$n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2}$
Helium	first	$n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2}$
	second	$n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$

Each of a lithium atom's three electrons also has its own unique set of quantum numbers, as you can see below. (Assume these quantum numbers represent a ground state lithium atom.)

Atom	Electron	Quantum numbers
Lithium	first	$n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2}$
	second	$n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$
	third	$n = 2, l = 0, m_l = 0, m_s = +\frac{1}{2}$

Consider the significance of lithium's quantum numbers. The first two electrons occupy the 1s orbital, and have the same sets of quantum numbers as helium's two electrons. According to the Pauli exclusion principle, the 1s orbital ($n = 1$) is now "full" because it contains the maximum number of electrons: two. For lithium's third electron, n cannot equal 1. Therefore, the next principal energy level for the electron is $n = 2$. If $n = 2$, l may have a value of 0 or 1. Remember that an atom's ground state represents its lowest energy state. Because an orbital with $l = 0$ (an s orbital) has a lower energy than an orbital with $l = 1$ (a p orbital), you would expect a high probability of finding lithium's third electron in the s orbital given by $n = 2, l = 0$. In fact, experimental evidence supports this expectation. You know that if $l = 0$, m_l has only one possible value: 0. Finally, by convention, m_s is $+\frac{1}{2}$.

It is possible to distinguish atoms by writing sets of quantum numbers for each of their electrons. However, writing quantum numbers for an atom such as uranium, which has 92 electrons, would be mind-bogglingly tedious. Fortunately, chemists have developed a "shortcut" to represent the number and orbital arrangements of electrons in each atom. As you will see shortly, these electron configurations, as they are called, are intimately connected to the structure and logic of the periodic table. In learning how to write electron configurations, you will discover this connection. You will also find out why orbital energies differ (for example, why a 2s orbital is lower in energy than a 2p orbital).

Note: Before proceeding, it is important to keep the following in mind. According to quantum mechanics:

- Electrons do not "occupy" orbitals. Nor do orbitals "contain" electrons.
- Electrons do not "fill" orbitals one at a time. Nor do electrons have properties that designate individual electrons as first, second, third, fourth, etc.
- Orbitals do not really have substance or shapes. Orbitals, recall, are mathematical equations that can be used to calculate probability densities for electrons. The shapes, and the impression of substance, result from the contour lines used to set a limit on the extent of those probabilities.

Nevertheless, it takes time and experience for most people to think about and talk about atoms in completely quantum mechanical terms. It is much simpler to think about orbitals as if they have substance. Therefore, in the text that follows, you will read about electrons "occupying" and "filling" orbitals. Strictly speaking, this terminology is not scientifically correct. However, it will help you understand some challenging concepts more easily.

CONCEPT CHECK

Write a set of quantum numbers for each electron of the following atoms: beryllium (Be), boron (B), and carbon (C). Look closely for evidence of a pattern. Write a short paragraph describing this pattern and explaining why you would have expected it.

An Introduction to Electron Configurations

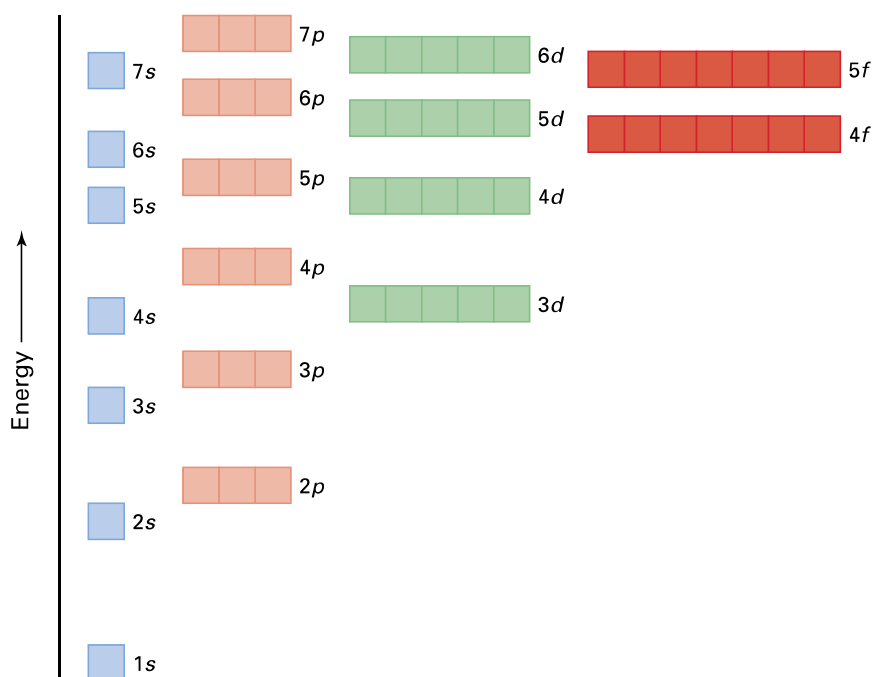


Figure 3.19 Atomic orbitals and their relative energies

Figure 3.19 shows the total number of orbitals in which electrons are likely to be found for all naturally occurring atoms and most synthetic atoms. The orbitals, represented by boxes, are listed in order of increasing energy. You will refer to this diagram often as you learn to write electron configurations.

An atom's **electron configuration** is a shorthand notation that shows the number and arrangement of electrons in its orbitals. Because the value of n ranges to infinity, there are an infinite number of electron configurations that are possible for each atom. For each atom, all but one of these represent the atom in an excited state. However, *an atom's chemical properties are mainly associated with its ground state electron configuration*. Therefore, unless stated otherwise, you can assume that any given electron configuration in this book represents an atom in its ground state. For example, the ground state electron configuration for a hydrogen atom is $1s^1$ (pronounced “one ess one”). The superscript 1 indicates that only one electron is in the s orbital. The ground state electron configuration for a helium atom is $1s^2$ (pronounced “one ess two”), where the superscript 2 indicates that there are two electrons in the s orbital. For a lithium atom, it is $1s^2 2s^1$. Again, the superscript numbers indicate the number of electrons in each of the orbitals.

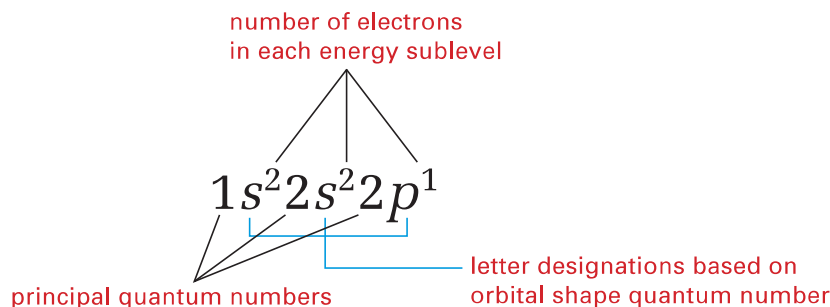
When you are learning to write electron configurations, it is helpful to start with the first element in the periodic table (hydrogen), and “build up” the electronic structure of its atom by adding an electron to its lowest available energy level. Then you move on to the next element (helium) by adding a proton (and several neutrons) to the nucleus, and by adding an electron to the appropriate orbital. This imaginary process of building up the ground state electronic structure for each atom, in order of atomic number, is called the **aufbau principle**. (*Aufbau* comes from a German word that means “to build up.”)

CONCEPT CHECK

Refer to the sets of quantum numbers for hydrogen and helium that you saw earlier. Then use the quantum numbers for lithium to infer why a lithium atom has the ground state electron configuration that it does.

Writing Electron Configurations

Electron configurations, as they are used in this book, provide information about the first two quantum numbers, n and l . (Electron configurations may also reflect the third quantum number, m_l , but this notation goes beyond the scope of this chemistry course.) The electron configuration below represents a boron atom in its ground state.



From the few examples of electron configurations that you have seen so far, you may have noticed that they do not include information about the spin direction of the electron. Once you have considered the electron configurations for a variety of atoms, you will see that you can safely infer where electrons with opposite spins have been paired together.

There is, however, a system that chemists use alongside electron configurations to help them plot and keep track of electrons in their orbitals. An **orbital diagram** uses a box for each orbital in any given principal energy level. (Some chemists use a circle or a line instead of a box.) An empty box \square represents an orbital in which there are no electrons (an “unoccupied” orbital). A box that has an upward-pointing arrow \uparrow represents an orbital with an electron that spins in one direction. A box with a downward-pointing arrow \downarrow represents an orbital with an electron that spins in the opposite direction. You can probably surmise that a box with oppositely pointing arrows $\uparrow\downarrow$ represents a filled orbital.

You are nearly ready to begin “building up” the electronic structures of the atoms, using both electron configurations and orbital diagrams. Below you will find some guidelines to help you. Refer to these guidelines and to Figure 3.19 as you work your way through the information that follows.

Guidelines for “Filling” Orbitals

1. Place electrons into the orbitals in order of increasing energy level.
2. Each set of orbitals of the same energy level must be completely filled before proceeding to the next orbital or series of orbitals.
3. Whenever electrons are added to orbitals of the same energy sublevel, each orbital receives one electron before any pairing occurs.
4. When electrons are added singly to separate orbitals of the same energy, the electrons must all have the same spin.

Guidelines 3 and 4 together comprise what is known as **Hund's rule**. If you “obey” Hund's rule, no two electrons can have the same set of quantum numbers. Thus, this rule follows from Pauli's exclusion principle.

Writing Electron Configurations for Periods 1 and 2

Below are the electron configurations for atoms of the first ten elements, along with their orbital diagrams.

Figure 3.20 presents the same basic information about the first ten elements, arranged in the form of the periodic table. Using these electron configurations, orbital diagrams, and Figure 3.20, you can probably follow the logic that leads to the electronic structure of each atom. As you examine and compare this information, note the following.

- The energy of each orbital (or each group of orbitals) increases as you move from left to right.
- Boron's fifth electron must go into the $2p$ energy sublevel. Since $l = 1$, m_l may be -1 , 0 , or $+1$. The fifth electron can go into any of these orbitals, because they all have the same energy. When you draw orbital diagrams, it is customary to place the electron in the first available box, from left to right.
- With carbon, you must apply Hund's rule. In other words, carbon's sixth electron must go into the next unoccupied $2p$ orbital. (Experimental evidence confirms that the spin of this sixth electron is, in fact, the same direction as the fifth electron.)
- With oxygen, as with helium's $1s$ orbital and beryllium's $2s$ orbital, the last-added (eighth) electron is paired with a $2p$ electron of opposite spin. In other words, the Pauli exclusion principle applies.

H	($Z = 1$)	$\boxed{\uparrow}$ 1s	$1s^1$
He	($Z = 2$)	$\boxed{\uparrow\downarrow}$ 1s	$1s^2$
Li	($Z = 3$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{}$ $\boxed{}$ $\boxed{}$ 1s 2s 2p	$1s^2 2s^1$
Be	($Z = 4$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{}$ $\boxed{}$ $\boxed{}$ 1s 2s 2p	$1s^2 2s^2$
B	($Z = 5$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{}$ $\boxed{}$ 1s 2s 2p	$1s^2 2s^2 2p^1$
C	($Z = 6$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ $\boxed{}$ 1s 2s 2p	$1s^2 2s^2 2p^2$
N	($Z = 7$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ 1s 2s 2p	$1s^2 2s^2 2p^3$
O	($Z = 8$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ 1s 2s 2p	$1s^2 2s^2 2p^4$
F	($Z = 9$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ 1s 2s 2p	$1s^2 2s^2 2p^5$
Ne	($Z = 10$)	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ 1s 2s 2p	$1s^2 2s^2 2p^6$











Period 1	1(IA) 1 H $1s^1$ 							18(VIII A) 2 He $1s^2$ 
	2(IIA) 3 Li $1s^2 2s^1$ 	4 Be $1s^2 2s^2$ 	13(IIIA) 5 B $1s^2 2s^2 2p^1$ 	14(IVA) 6 C $1s^2 2s^2 2p^2$ 	15(VA) 7 N $1s^2 2s^2 2p^3$ 	16(VIA) 8 O $1s^2 2s^2 2p^4$ 	17(VIIA) 9 F $1s^2 2s^2 2p^5$ 	10 Ne $1s^2 2s^2 2p^6$ 

Figure 3.20 How orbitals are filled for atoms of the first 10 elements of the periodic table

Electron Configurations and Orbital Diagrams for Period 3

To write the electron configurations and draw the orbital diagrams for atoms of Period 3 elements, you follow the same process as for Period 2. The practice problems below give you the chance to write and draw these representations. In doing so, you will observe a pattern that enables you to take a “shortcut” in writing electron configurations. As the atomic number increases, electron configurations become longer and longer. You know that chemical reactivity depends mainly on an atom’s valence electrons: the electrons in the outermost-occupied principle energy level. To highlight these electrons, you can use a simplified notation called a *condensed electron configuration*. This notation places the electron configuration of the noble gas of the previous period in square brackets, using its atomic symbol only. Then you continue with the configuration of the next energy level that is being filled.

The condensed electron configuration for a nitrogen atom, for example, is $[\text{He}]2s^2 2p^3$. The notation $[\text{He}]$ is used to represent $1s^2$. For a sodium atom ($Z = 11$), the condensed electron configuration is $[\text{Ne}]3s^1$. Here, $[\text{Ne}]$ represents, $1s^2 2s^2 2p^6$. Be aware that condensed electron configurations are simply convenient short forms. Thus, $[\text{Ne}]3s^1$ does *not* mean that a sodium atom is the same as a neon atom plus one electron. Sodium and neon are different elements because the nuclei of their atoms are completely different.

Practice Problems

- Use the aufbau principle to write complete electron configurations and complete orbital diagrams for atoms of the following elements: sodium, magnesium, aluminum, silicon, phosphorus, sulfur, chlorine, and argon (atomic numbers 11 through 18).
- Write condensed electron configurations for atoms of these same elements.

Continued ...

8. Make a rough sketch of the periodic table for elements 1 through 18, including the following information: group number, period number, atomic number, atomic symbol, and condensed electron configuration.
9. A general electron configuration for atoms belonging to any element of group 1 (IA) is ns^1 , where n is the quantum number for the outermost occupied energy level. Based on the patterns you can observe so far for elements 1 to 18, predict the general electron configuration for the outermost occupied energy levels of groups 2 (IIA), 13 (IIIA), 14 (IVA), 15 (VA), 16 (VIA), 17 (VIIA), and 18 (VIIIA).

Electron Configurations and Orbital Diagrams for Period 4

You may have noticed that period 3 ended with electrons filling the $3p$ energy sublevel. However, when $n = 3$, l may equal 0, 1, and 2. Perhaps you wondered what happened to the $3d$ orbitals ($l = 2$).

Electrons do not start occupying $3d$ orbitals until the $4s$ orbital is filled. Therefore, the $3d$ orbitals are filled starting in period 4. The reason for this change in the expected orbital filling order is that the $4s$ orbital has a lower energy than the $3d$ orbitals. Remember that usually you fill orbitals in order of increasing energy. (If necessary, refer to the Guidelines for “Filling” Orbitals and Figure 3.19 to refresh your memory.)

The chart below shows electron configurations and partial orbital diagrams for the 18 elements of period 4. You would expect the filling pattern shown for potassium ($Z = 19$) through vanadium ($Z = 23$). However, an unexpected deviation from the pattern occurs with chromium ($Z = 24$). The same thing happens with copper ($Z = 29$). All other configurations for period 4 conform to the aufbau principle.

Why do Cr and Cu have electron configurations that are different from what you would predict? The guidelines that you have been using state that you fill orbitals according to increasing energy. In most cases, this results in the most stable ground state configuration for the atom. Experimental evidence indicates, however, that some atoms achieve greater stability with electron configurations that do not conform to predicted patterns. For chromium, the greatest ground state stability results from a configuration in which its $4s$ and $3d$ orbitals are half-filled. For copper, a half-filled $4s$ orbital and a completely filled set of $3d$ orbitals gives the most stable configuration. Similar situations arise for a number of atoms in the remaining periods.

Electron Configurations and Partial Orbital Diagrams for Atoms of Period 4 Elements

Element	Atomic Number	Orbital Diagram (4s, 3d, and 4p Orbitals Only)	Complete Electron Configuration	Condensed Electron Configuration
K	$Z = 19$	$4s$ \uparrow $3d$ \square \square \square \square \square $4p$ \square \square \square	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1$	$[\text{Ar}] 4s^1$
Ca	$Z = 20$	$4s$ $\uparrow\downarrow$ $3d$ \square \square \square \square \square $4p$ \square \square \square	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2$	$[\text{Ar}] 4s^2$
Sc	$Z = 21$	$4s$ $\uparrow\downarrow$ $3d$ \uparrow \square \square \square \square \square $4p$ \square \square \square	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^1$	$[\text{Ar}] 4s^2 3d^1$
Ti	$Z = 22$	$4s$ $\uparrow\downarrow$ $3d$ \uparrow \uparrow \square \square \square \square $4p$ \square \square \square	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^2$	$[\text{Ar}] 4s^2 3d^2$

V	Z = 23	4s ↑↓	3d ↑ ↑ ↑	4p □ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^3$	$[\text{Ar}] 4s^2 3d^3$
Cr	Z = 24	↑	↑ ↑ ↑ ↑ ↑	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1 3d^5$	$[\text{Ar}] 4s^1 3d^5$
Mn	Z = 25	↑↓	↑ ↑ ↑ ↑ ↑	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^5$	$[\text{Ar}] 4s^2 3d^5$
Fe	Z = 26	↑↓	↑↓ ↑ ↑ ↑ ↑	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^6$	$[\text{Ar}] 4s^2 3d^6$
Co	Z = 27	↑↓	↑↓ ↑↓ ↑ ↑ ↑	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^7$	$[\text{Ar}] 4s^2 3d^7$
Ni	Z = 28	↑↓	↑↓ ↑↓ ↑↓ ↑ ↑	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^8$	$[\text{Ar}] 4s^2 3d^8$
Cu	Z = 29	↑	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1 3d^{10}$	$[\text{Ar}] 4s^1 3d^{10}$
Zn	Z = 30	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	□ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10}$	$[\text{Ar}] 4s^2 3d^{10}$
Ga	Z = 31	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ □ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^1$	$[\text{Ar}] 4s^2 3d^{10} 4p^1$
Ge	Z = 32	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑ □	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^2$	$[\text{Ar}] 4s^2 3d^{10} 4p^2$
As	Z = 33	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^3$	$[\text{Ar}] 4s^2 3d^{10} 4p^3$
Se	Z = 34	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^4$	$[\text{Ar}] 4s^2 3d^{10} 4p^4$
Br	Z = 35	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^5$	$[\text{Ar}] 4s^2 3d^{10} 4p^5$
Kr	Z = 36	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^6$	$[\text{Ar}] 4s^2 3d^{10} 4p^6$

Electron Configurations and the Periodic Table

So far, you have “built up” atoms of the first 36 elements of the periodic table. There are more than 80 elements still remaining. You do not, however, have to build up atoms of these elements to write their electron configurations. You can do it simply by consulting the periodic table! All you need is to recognize the significance of several patterns. You have already observed many of these.

In the process of examining the patterns outlined below, you will learn the filling order for atoms of elements in periods 5, 6, and 7. You will also see why the shape and organization of the periodic table is a direct consequence of the electronic structure of the atoms.

Arranging Elements by Electron Configurations

Figure 3.21 on the next page shows the entire periodic table, with certain segments colour-coded and labelled according to the type of orbital that is being filled. If you built up an atom of every element, filling orbitals in order of increasing energy, you would construct a chart with *precisely* this shape and organization. Figure 3.22 highlights the filling order, and therefore the energy order, of the orbitals when you read the periodic table from left to right.

Elements that appear in the s block and the p block are called either the *main group elements* or the *representative elements*. Chemists give them these names because, collectively, these elements are representative of a wide range of physical and chemical properties. Among the main group elements, for example, you will find metals, non-metals, metalloids,

highly reactive elements, moderately reactive elements, and unreactive elements. While most main group elements are solids at room temperature, roughly one quarter of them are gases, and one is a liquid.

Elements that appear in the *d* block are called the *transition elements*. They mark the transition from the *p* orbital filling order to the *d* orbital filling order. By the same reasoning, the *f* block elements are called the *inner transition elements*, because they mark a transition from the *d* orbital filling order to the *f* orbital filling order.

Figure 3.21 The long form of the periodic table, with the four energy sublevel blocks identified.

1 (IA)	2 (IIA)																18 (VIIIA)
1s	2s																
2s	2p																
3s	3p																
4s	4p																
5s	5p																
6s	6p																
7s	7p																
s block (main group elements)		f block (inner transition elements)						d block (transition elements)						p block (main elements)			

Patterns Involving Group Numbers and Period Numbers

Elements in a group have similar chemical properties because they have similar outer electron configurations. That is, they have the same number of valence electrons. This observation gives rise to three patterns that you can deduce from the periodic table.

1s	→	
2s	→	2p
3s	→	3p
4s	→	3d → 4p
5s	→	4d → 5p
6s	→	4f → 5d → 6p
7s	→	5f → 6d

Figure 3.22 Reading the periodic table from left to right, starting at the top and finishing at the bottom, displays the filling order and the energy order of the atomic orbitals.

- For main group elements, the last numeral of the group number is the same as the number of valence electrons. For example, phosphorus in group 15 has 5 valence electrons. Strontium in group 2 has 2 valence electrons. (If you use the roman-numeral system for numbering groups, the group number is the number of valence electrons.) The only exception to this pattern is helium (group 18), which has only two valence electrons. Nevertheless, helium's outermost occupied energy level is complete with two electrons.
- The n value of the highest occupied energy level is the period number. For example, atoms of elements that have electrons with $n = 3$ appear in period 3. Atoms of elements that have electrons with $n = 6$ appear in period 6.
- The square of the n value (n^2) equals the total number of orbitals in that energy level. Furthermore, since each orbital may have a maximum of two electrons, the maximum number of electrons in any principle energy level is $2n^2$. For example, with $n = 2$, there are four orbitals: one $2s$ orbital and three $2p$ orbitals. Squaring the n value gives $2^2 = 4$. The total number of electrons in this energy level is *eight*, given by $2n^2$. Notice that there are *eight* elements in period 2. This is no coincidence. Compare this result with the next period. Does the pattern hold? For period 3 elements, $n = 3$. The number of orbitals is $n^2 = 9$. (Verify that for yourself.) The number of electrons is $2n^2 = 18$. However, *there are only eight elements in period 3*. Why? The $4s$ orbital has a lower energy than the $3p$ orbitals, and thus fills first. The $3d$ orbitals fill after the $4s$ orbitals. Therefore, there are 18 elements in period 4. Recall that the $4f$ orbitals do not fill until after the $6s$ orbital fills.

Summarizing Characteristics of *s*, *p*, *d*, and *f* Block Elements

The *s* block includes hydrogen, helium, and the elements of groups 1 (IA) and 2 (IIA). Valence electrons occupy only the *ns* orbitals in this block. Hydrogen atoms and atoms of group 1 have partially filled *s* orbitals, represented by the general electron configuration ns^1 . Helium and atoms of group 2 elements have *s* orbitals filled with the maximum number of valence electrons, represented by the general notation ns^2 . Because *s* orbitals can hold a total of two electrons, the *s* block elements span two groups.

The *p* block includes elements of groups 13 (IIIA) through 18 (VIIIA). Electron configurations of the *p* block atoms take the general form ns^2np^a , where *a* represents a value ranging from 1 to 6. Because the three *p* orbitals can hold a maximum of six electrons, the *p* block elements span six groups.

Unique among the *p* block elements are the atoms of group 18 (VIIIA) elements: the noble gases, renowned for their chemical stability. Their electron configuration takes the general form ns^2np^6 , representing a ground state outer energy level with fully occupied *s* and *p* orbitals. The stability of this configuration is a key factor that drives the formation of chemical bonds.

The *d* block includes all the transition elements. In general, atoms of *d* block elements have filled *ns* orbitals, as well as filled or partially filled *d* orbitals. Generally, the *ns* orbitals fill before the $(n - 1)d$ orbitals. However, there are exceptions (such as chromium and copper) because these two sublevels are very close in energy, especially at higher values of *n*. Because the five *d* orbitals can hold a maximum of ten electrons, the *d* block spans ten groups.

The *f* block includes all the inner transition elements. Atoms of *f* block elements have filled *s* orbitals in the outer energy levels, as well as filled or partially filled *4f* and *5f* orbitals. In general, the notation for the orbital filling sequence is *ns*, followed by $(n - 2)f$, followed by $(n - 1)d$, followed by (for period 6 elements) *np*. However, there are many exceptions that make it difficult to predict electron configurations. Because there are seven *f* orbitals, with a maximum of fourteen electrons, the *f* block spans fourteen groups.

CONCEPT CHECK

A teacher asks the class to write electron configurations for magnesium, iron, and tin. Lois does not have a periodic table available. Edward's advice is to use a memory aid that is shown below. Write a short paragraph to explain how to use this memory aid.



Sample Problem

Determining Quantum Numbers

Problem

You do not have a periodic table. You are told that the condensed electron configuration for strontium is $[\text{Kr}]5s^2$. Identify the group number, period number, and orbital block in which strontium appears on the periodic table. Show your reasoning.

Solution

From the electron configuration, you can deduce the energy level of the valence electrons, which tells you the period number for strontium. The number of valence electrons and their sublevel enable you to deduce the group number and the block. Therefore:

Continued ...

The configuration for the valence electrons, s^2 , indicates that strontium is in group 2 (IIA). The value of 5 in $5s^2$ tells you that strontium is in period 5. The notation s^2 means that strontium's valence electrons fill the s orbital. Thus, strontium must be in the s block.

Check Your Solution

All elements in group 2 (IIA) have electron configurations that end with the notation s^2 . The link between n value and period number is correctly applied. Strontium has an electron configuration with the general notation ns^2 , which is characteristic of s block elements.



CHEM

FACT

The d block elements known to nineteenth century chemists posed a puzzle. Many of these metals were known to react with neutral molecules. In particular, chemists observed that nickel could bond with carbon monoxide. The resulting compound, nickel tetracarbonyl, $\text{Ni}(\text{CO})_4$, is a gas. At the time, the nature of the bonding in this compound was unexplainable. Even so, chemists could easily produce and decompose $\text{Ni}(\text{CO})_4$. The chemical reactions involved were used to purify nickel, and led to the establishment of a company that later became International Nickel Company, INCO.

Practice Problems

- Without looking at a periodic table, identify the group number, period number, and block of an atom that has the following electron configurations.
 (a) $[\text{Ne}]3s^1$ (b) $[\text{He}]2s^2$ (c) $[\text{Kr}]5s^24d^{10}5p^5$
- Use the aufbau principle to write complete electron configurations for the atom of the element that fits the following descriptions.
 (a) group 2 (IIA) element in period 4
 (b) noble gas in period 6
 (c) group 12 (IIB) element in period 4
 (d) group 16 (VIA) element in period 2
- Identify all the possible elements that have the following valence electron configurations.
 (a) s^2d^1 (b) s^2p^3 (c) s^2p^6
- For each of the elements below, use the aufbau principle to write the full and condensed electron configurations and draw partial orbital diagrams for the valence electrons of their atoms. You may consult the periodic table in Appendix C, or any other periodic table that omits electron configurations.
 (a) potassium (b) nickel (c) lead

Electron Configurations, Atomic Properties, and Periodic Trends

Electron configurations—representations of the electronic structure of atoms—help to determine the atomic and chemical properties of the elements. Properties such as atomic radius, ionization energy, metallic character, and electron affinity are periodic. That is, they follow recurring trends in the periodic table. In a previous chemistry course, you studied these properties and noted their general trends. With your newly developed understanding of the periodic table, you are in a position to re-examine these properties and consider the exceptions to the general trends. Before doing so, carry out the following ThoughtLab, which applies and synthesizes your understanding of the quantum mechanical model of the atom, quantum numbers, electron configurations, and the periodic arrangement of the elements.



The arrangement of elements in the periodic table is a direct consequence of the allowed values for the four quantum numbers. If the laws of physics allowed these numbers to have different values, the appearance of the periodic table would change. In this activity, you will examine the effect on the periodic arrangement of the elements when one of the quantum numbers has a different set of allowed values.

Part 1

Procedure

- Suppose, in a different universe, that the quantum number m_l has only one allowed value, $m_l = 1$. All other allowed values for the remaining quantum numbers are unchanged. Therefore, the set of allowed values for all quantum numbers is:

$$n = 1, 2, 3, \dots$$

$$l = 0, 1, 2, \dots, (n - 1)$$

$$m_l = 1$$

$$m_s = +\frac{1}{2}, -\frac{1}{2}$$

Write a memory aid for the relative order of orbital energies for the first six energy levels.

- Assume that the exclusion principle and Hund's rule are followed. Demonstrate that the periodic table for the first 30 elements would be as shown below:

		s block						
period	1	1	2					
	2	3	4					
	3	7	8					
	4	11	12	d block				
	5	17	18					
	6	23	24					
						p block		
						5	6	
						9	10	
				13	14	15	16	
				19	20	21	22	
				27	28	29	30	

Periodic Trends in Atomic Radius

As you know, atoms are not solid spheres. Their volumes (the extent of their orbitals) are described in terms of probabilities. Nevertheless, the size of an atom — its **atomic radius** — is a measurable property. Chemists can determine it by measuring the distance between the nuclei of bonded, neighbouring atoms. For example, for metals, atomic radius is half the distance between neighbouring nuclei in a crystal of the metal element. For elements that occur as molecules, which is the case for many non-metals, atomic radius is half the distance between nuclei of identical atoms that bonded together with a single covalent bond. In Figure 3.23, the radii of metallic elements represent the radius of an atom in a metallic crystal. The radii of all other elements represent the radius of an atom of the element participating in a single covalent bond with one additional, like atom.

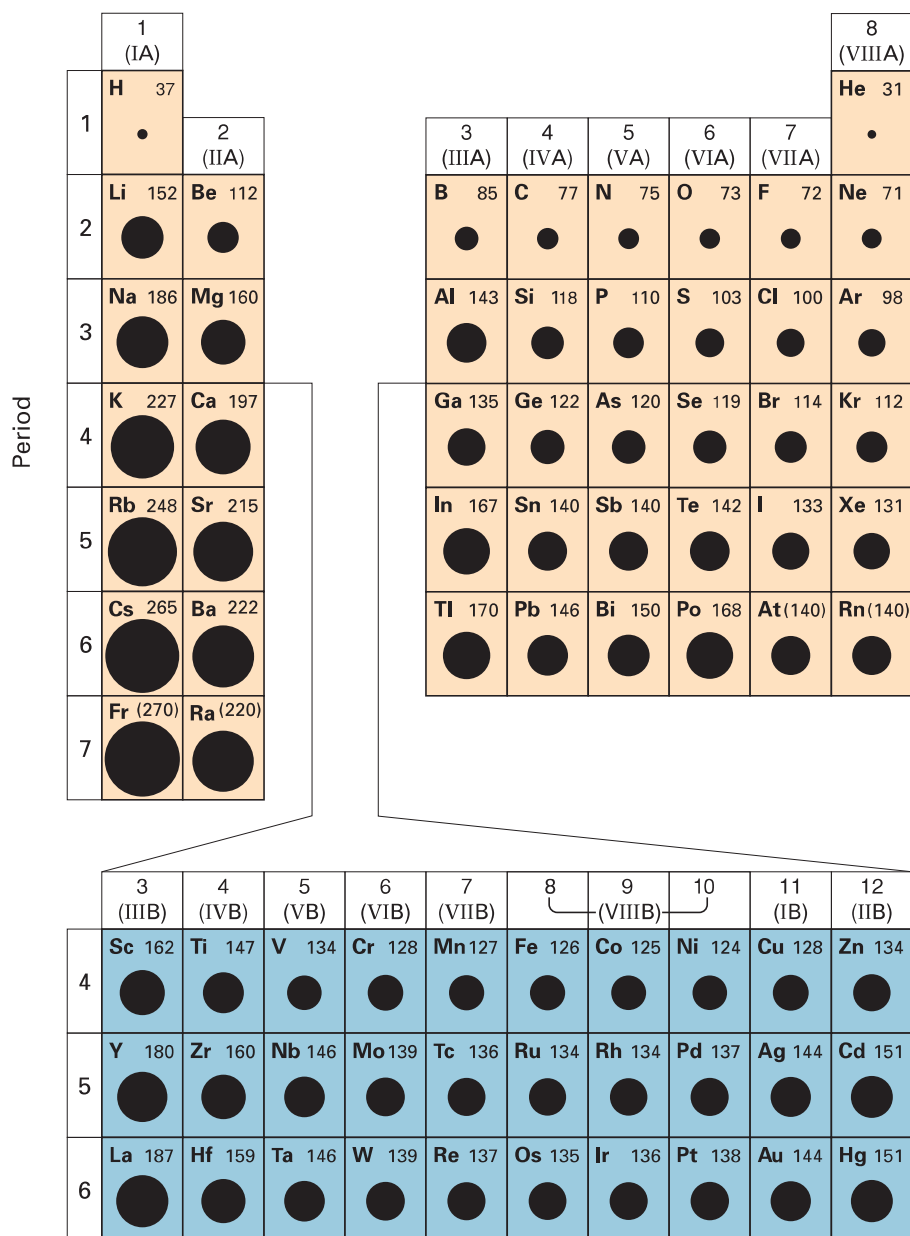


Figure 3.23 Representations of atomic radii for main group and transition elements. (Values for atomic radii are given in picometres. Those in parentheses have only two significant digits.)

Figure 3.23 shows the periodic trends associated with the atomic radius. You can see that atomic radii generally decrease across a period. Furthermore, atomic radii generally increase down a group. Two factors affect differences in atomic radii.

One factor affecting atomic radii is changing n . As n increases, there is a higher probability of finding electrons farther from their nucleus. Therefore, the atomic volume is larger. In other words, atomic radius tends to increase with increasing n , and decrease with decreasing n .

The other factor that affects atomic radii is changing nuclear charge — specifically, the *effective* nuclear charge. Effective nuclear charge is the *net* force of attraction between electrons and the nucleus they surround. Only hydrogen's single electron experiences the full positive charge of its nucleus. For hydrogen, the nuclear charge experienced by the electron is Z , its atomic number. For all other atoms, the nuclear charge that any given electron experiences is offset to some degree by other electrons. Depending on the electron density between the nucleus and any particular electron, the net force of attraction—the effective nuclear charge, Z_{eff} — may be large or small.

As Z_{eff} increases, electrons are attracted more strongly to the nucleus, and the size of the atom decreases. As Z_{eff} decreases, electrons experience a reduced force of attraction, and the size of the atom increases. Valence electrons, especially, experience a smaller Z_{eff} than inner electrons because the inner electrons shield or screen them from the attractive force of the nucleus.

For the main group elements, the combined influences of n , Z_{eff} , and this shielding effect have the following outcomes.

- n governs the trend of increasing atomic radius down a group. Down a group, atoms of each subsequent element have one more level of inner electrons, increasing the shielding effect. Because of the additional shielding, Z_{eff} for the valence electrons changes only slightly. The increase in atomic size, therefore, results from the increasing value of n .
- Z_{eff} governs the trend of decreasing atomic radius across a period. Across a period, atoms of each element have one more electron added to the same outer energy level — n does not change. The shielding effect changes only slightly. However, Z_{eff} changes considerably. The decrease in atomic size across a period, therefore, results from increasing Z_{eff} .

The atoms of transition elements do not display the same general trend as the main group elements. A key reason for this is that electrons are added to inner energy levels — the d orbitals — rather than to the outer energy levels. As a result, Z_{eff} changes relatively little, so atomic size remains fairly constant. In later chemistry courses, you will learn a more complete explanation for the atomic radii of transition-element atoms.

Periodic Trends in Ionization Energy

For the hydrogen atom, the difference in energy between $n = 1$ and $n = \infty$ represents the energy change associated with the complete removal of its electron. (By convention, when $n = \infty$, an electron is considered to be completely removed from its atom.) The energy needed to completely remove one electron from a ground state gaseous atom is called the **ionization energy**. Energy must be added to the atom to remove an electron in order to overcome the force of attraction exerted on the electron by the nucleus. Since multi-electron atoms have two or more electrons, they also have more than one ionization energy.

A gaseous atom's first ionization energy is the least energy required to remove an electron from the outermost occupied energy level. The second ionization energy is always greater than the first ionization energy, because the electron must be removed from a positively charged ion. The same reasoning applies for successive ionization energies.

First ionization energy (IE_1) is closely linked to an atom's chemical reactivity. Atoms with a low IE_1 tend to form cations during chemical reactions. Notice in Figure 3.24 that the atoms with the lowest IE_1 are those belonging to group 1 elements. These elements, also known as the alkali metals, are among the most reactive elements in the periodic table. Atoms of elements with high IE_1 tend to form anions: negatively charged ions. (A notable exception is the noble gases, which do not form ions naturally.)

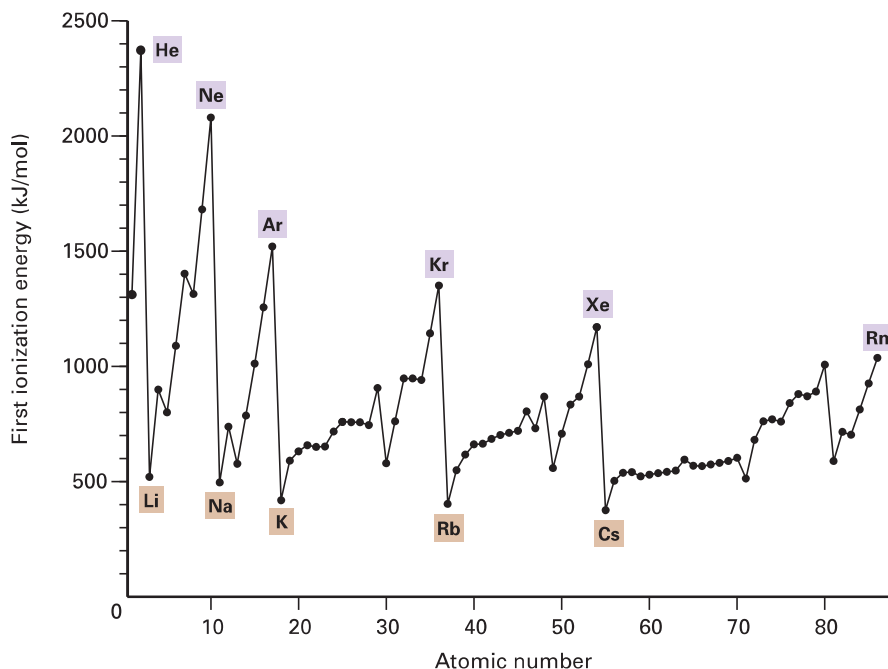


Figure 3.24 Periodic trends involving first ionization energy

Figure 3.24 illustrates several trends associated with ionization energy.

- Ionization energy generally decreases down a group. Notice that this trend is the inverse of the trend for atomic radius. The two trends are, in fact, linked. As atomic radius increases, the distance of valence electrons from the nucleus also increases. There is a decrease, therefore, in the force of attraction exerted by the nucleus on the valence electrons. Thus, less energy is needed to remove one such electron.
- Ionization energy generally increases across a period. Again, this trend is linked to the atomic radius. Across a period, the atomic radius decreases because Z_{eff} increases. The force of attraction between the nucleus and valence electrons is subsequently increased. Therefore, more energy is needed to remove one such electron.


Several minor variations in these trends occur. The variations involve boron and aluminum in group 13 (IIIA) and oxygen and sulfur in group 16 (VIA). The drop in IE_1 with group 13 occurs because electrons start to fill the np orbitals. These orbitals are higher in energy than the ns orbital, so the additional electron is more easily removed. Group 16 marks the first pairing of the p orbital electrons. The np^3 configuration of nitrogen is more stable than the np^4 configuration of oxygen. Electron repulsions

increase the orbital energy in oxygen. Therefore, less energy is required to remove the fourth *p* sublevel electron in oxygen.


Periodic trends in ionization energy are linked to trends involving the reactivity of metals. In general, the chemical reactivity of metals increases down a group and decreases across a period. These trends, as well as a further trend from metallic to non-metallic properties across a period, and increasing metallic properties down a group, are shown in Table 3.1.

Table 3.1 Periodic Trends in Reactivity of Metals


Reaction	With oxygen in the air	With water	Reaction with dilute acids
Group 1	React rapidly to form oxides with the general formula M_2O . These compounds are strong bases.	React vigorously with cold water, generating $H_{2(g)}$ and forming a strong base with the general formula MOH .	The reaction is dangerously violent, igniting $H_{2(g)}$ generated.
Group 2	React moderately to form basic oxides with the general formula MO .	React slowly in cold water, generating $H_{2(g)}$ and forming a strong base, $M(OH)_2$.	React quickly generating $H_{2(g)}$.
Group 13	React slowly. Aluminum forms an oxide coating that protects the metal; the compound Al_2O_3 can act as a base or an acid.	Al displaces $H_{2(g)}$ from steam.	Al reacts readily if the protective oxide coating is removed, to generate $H_{2(g)}$.
Group 14	C and Si are non-metals, and form acidic oxides. Ge is a metalloid, and forms an acidic oxide. SnO_2 can act as an acid or a base. PbO_2 is unreactive.	The most metallic elements, Sn and Pb, do not react with water.	Sn and Pb react slowly to form $H_{2(g)}$.
Transition Metals (for example, Fe, Co, Ni, Cu, Zn, Ag, Au)	Zn is the most reactive transition metal, and forms ZnO when the metal is burned in air. Ag and Au do not react.	Zn and Fe displace $H_{2(g)}$ from steam. Fe rusts slowly at room temperature.	Zn reacts readily to generate $H_{2(g)}$. Cu, Ag, and Au do not react.



Magnesium ribbon reacting with oxygen in air



Potassium reacting with water



Nickel reacting with dilute acid

Periodic Trends in Electron Affinity

Electron affinity is the change in energy that occurs when an electron is added to a gaseous atom. As was the case with ionization energy, there is more than one value for electron affinity. The first electron affinity results in the formation of an anion with a charge of 1[−]. For example, when a neutral fluorine atom acquires an electron, as in a bonding situation, the resulting electron configuration of the fluoride ion is the same as that of the noble gas, neon:



Fluorine is a highly reactive element. The relative ease with which it gains an electron when it forms bonds is reflected in its high electron affinity.

1 (IA)						18 (VIII A)	
H −72.8	2 (IIA)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	He (+21)
Li −59.6	Be (+241)	B −26.7	C −122	N 0	O −141	F −328	Ne (+29)
Na −52.9	Mg (+230)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+34)
K −48.4	Ca (+156)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+167)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+40)
Cs −45.5	Ba (+52)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

Figure 3.25 Electron affinities for the main group elements. Negative values mean that anions form with the release of energy. Positive values mean that energy is absorbed in order to form the anion. The positive values are estimated, because the anions formed for atoms of groups 2 and 18 are unstable.

Figure 3.25 shows electron affinities for atoms of the main group elements. Positive values mean that energy is required to add an electron. Negative values mean that energy is given off when an electron is added. High negative numbers mean a high electron affinity. Low negative numbers and positive numbers mean a low electron affinity.

Trends for electron affinity are more irregular than those for atomic radius and ionization energy, because factors other than atomic size and Z_{eff} are involved. In future chemistry courses, you will learn about these factors and how they explain the irregularities. However, the property of electron affinity is still significant when you consider it in combination with ionization energy. The trends that result from this combination are important for chemical bonding.

- Atoms of elements in group 17 (VIIA), and to a lesser degree group 16 (VIA), have high ionization energies and high electron affinities. As a result, it takes a great deal of energy to remove electrons from these atoms. However, these atoms attract electrons strongly, and form negative ions in ionic compounds.

- Atoms of elements in group 1 (IA) and group 2 (IIA) have low ionization energies and low electron affinities. Atoms of these elements give up electrons easily, but attract them poorly. Therefore, they form positive ions in ionic compounds.
- Atoms of elements in group 18 (VIIIA) have very high ionization energies and very low electron affinities. Therefore, in nature, they do not gain, give up or share electrons at all. (Under laboratory conditions, only the larger group 18 atoms can be made to form compounds.)

Section Summary

In this section, you have seen how a theoretical idea, the quantum mechanical model of the atom, explains the experimentally determined structure of the periodic table, and the properties of its elements. Your understanding of the four quantum numbers enabled you to write electron configurations and draw orbital diagrams for atoms of the elements. You also learned how to “read” the periodic table to deduce the electron configuration of any element.

Since the start of high school science courses, you have used the periodic table to help you investigate the composition and behaviour of the elements. Your early experiences with the periodic table were limited largely to the first 20 elements, because you could explain their electron structure without the concepts of orbitals and electron configurations. The modern, quantum mechanical model of the atom has broadened your understanding of the elements, the composition of their atoms, and their chemical and physical behaviour in the world around you. You will draw upon these concepts in the next chapter, as you expand your understanding of the forces that are responsible for the millions of kinds of matter in the universe: chemical bonds.

Section Review

- (a) **K/U** What is the fourth quantum number, and in what way is it different from the other three quantum numbers?

(b) Explain why the fourth quantum number is an example of a quantized value.
- C** Compare the orbitals of a hydrogen atom with the orbitals of all other atoms. In what ways are they similar? In what ways are they different?
- C** Identify the group number, period number, and orbital block of the periodic table for elements whose atoms have the following electron configurations.

(a) $[\text{Kr}]5s^24d^1$	(c) $[\text{He}]2s^22p^6$
(b) $[\text{Ar}]4s^23d^{10}4p^3$	(d) $[\text{Ne}]3s^23p^1$
- C** Which requires more energy: removing a valence electron from its atom or removing an electron from an inner energy level? Explain why.
- K/U** On which side of the periodic table would you find an element whose atom is likely to form a cation? Which atomic property is related to this question?

- 6 **K/U** In what ways are each of the following related to or affected by one another?
- (a) the aufbau principle, the exclusion principle, and Hund's rule
 - (b) electron configurations and orbital diagrams
 - (c) ionization energy and atomic radius
 - (d) ionization energy and electron affinity
 - (e) effective nuclear charge and valence electrons
 - (f) the periodic table and chemical properties of the elements
 - (g) electron configurations and the periodic table
 - (h) electron configurations and quantum numbers
 - (i) electron configurations, period number, and group number
 - (j) the quantum mechanical model of the atom and the periodic table
- 7 **C** Use the aufbau principle to write complete and condensed electron configurations for the most common ions for the elements listed below, and explain the significance of any patterns you observe in their electronic structures.
- (a) sodium
 - (b) calcium
 - (c) chlorine
 - (d) sulfur
- 8 **I** The following data lists the ionization energies for a given atom: $IE_1 = 738 \text{ kJ/mol}$; $IE_2 = 1451 \text{ kJ/mol}$; $IE_3 = 7733 \text{ kJ/mol}$. Predict the valence electron configuration for this atom, and explain your reasoning.

CHAPTER 3 Review

Reflecting on Chapter 4

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Compare the Rutherford, Bohr, and quantum mechanical models of the atom.
- Distinguish between the view of matter and energy in the macroscopic world (the world of everyday experience) and the view of matter and energy in the quantum mechanical world.
- Describe, using examples, the relationship between the four quantum numbers and the electron configurations of atoms.
- Explain the relationship between the electronic structure of atoms and the arrangement of elements in the periodic table.
- Illustrate the trends associated with several atomic and chemical properties of the elements.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

nuclear model	emission spectrum
absorption spectrum	quantum
photons	quantum mechanical model of the atom
orbitals	ground state
principal quantum number (n)	orbital-shape
magnetic quantum number (m_l)	quantum number (l)
Pauli exclusion principle	spin quantum number (m_s)
aufbau principle	electron configuration
Hund's rule	orbital diagram
ionization energy	atomic radius
	electron affinity

Knowledge/Understanding

1. Explain the experimental observations and inferences that led Rutherford to propose the nuclear model of the atom.
2. Explain how the Bohr atomic model differs from the Rutherford atomic model, and explain the observations and inferences that led Bohr to propose his model.
3. Both the Rutherford and Bohr atomic models have been described as planetary models. In what ways is this comparison appropriate? In what ways is this comparison misleading?

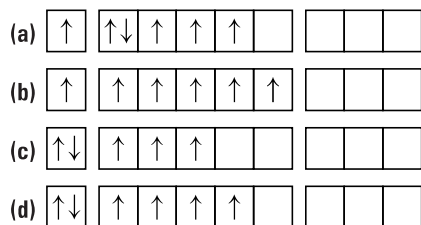
4. Briefly describe the contributions made by the following physicists to the development of the quantum mechanical model of the atom.

(a) Planck	(d) Heisenberg
(b) de Broglie	(e) Schrödinger
(c) Einstein	
5. List characteristics of the s , p , d , and f blocks to clearly distinguish among the atoms of their elements.
6. Explain how Pauli's exclusion principle and Hund's rule assist you in writing electron configurations.
7. Give the energy level and type of orbital occupied by the electron with the following set of quantum numbers: $n = 3$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$.
8. Differentiate between a ground state sulfur atom and an excited state sulfur atom. Use electron configurations to illustrate this difference.
9. Which of the following pairs of atoms would you expect to have a larger atomic radius? Explain your choice.

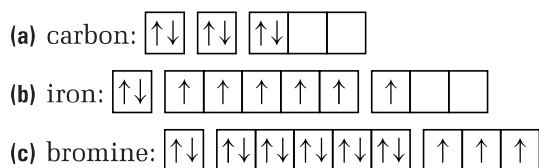
(a) Na, Mg	(d) Al, Ga
(b) Cl, K	(e) O, F
(c) Ca, Sr	(f) Cl, Br
10. Locate each of the following elements on a periodic table, and state the orbital block in which each element is found: U, Zr, Se, Rb, Re, Sr, Dy, Kr.
11. The chemical formulas for the oxides of lithium, beryllium, and boron are: Li_2O , BeO , and B_2O_3 . Write the formulas for the oxides of sodium, potassium, magnesium, calcium, aluminum, and gallium. Explain how you determined their formulas.
12. (a) How many electron-containing orbitals does an arsenic atom have in the ground state?
 (b) How many of the orbitals are completely filled?
 (c) How many of the orbitals are associated with the atom's fourth principal energy level?
 (d) How many possible orbitals could electrons occupy if the atom were excited?

Inquiry

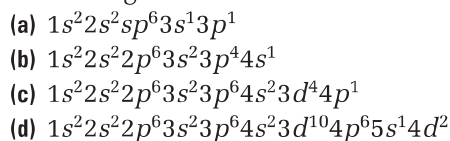
13. Which of the following is the correct orbital diagram for the third and fourth principal energy levels of a vanadium atom ($Z = 23$)? Justify your answer.



14. Each of the following orbital diagrams is incorrect. Identify the errors, explain how you recognized them, and use the aufbau principle to write electron configurations using the corrected orbital diagrams.



15. The electron configurations below represent atoms in excited states. Identify each atom, and write its ground state electron configuration.



16. Sketch a blank periodic table into your notebook to fill, roughly, half a page. Write the letters of the questions below in their appropriate place in the periodic table.
- (a) smallest atomic radius in period 3
- (b) lowest IE_1 in period 5
- (c) highest EA in period 17 (VIIA)
- (d) largest atomic radius in period 6
- (e) most metallic character in period 14 (IVA)
- (f) group 3 (IIIA) atom that forms the most basic oxide
- (g) forms a $2+$ ion with the electron configuration $[Ar]4s^2$
- (h) has a condensed ground state electron configuration of $[Ne]3s^2 3p^2$
- (i) period 4 element whose atom, in the ground state, has an outermost energy level that is completely full

- (j) a transition metal for which the aufbau principle predicts an electron configuration that would be incorrect (**Note:** There are several possible answers; you are required to provide only one.)

17. On your copy of the periodic table from question 16, colour-code and label the four orbital blocks.
18. Imagine that scientists have successfully synthesized element X, with atomic number 126. Predict the values of n and l for the outermost electron in an atom of this element. State the number of orbitals there would be in this energy sublevel.

Communication

19. In terms of the general periodic trends for groups and periods, explain why ranking the following pairs might prove tricky.
- (a) K and Sr, according to atomic size
- (b) Mn and Fe, according to first ionization energy
- (c) Na and Ca, according to metallic character
20. Explain why it is not possible to measure the size of an atom directly.
21. Identify elements whose atoms have the following valence electron configurations:
- (a) $5s^1$ (c) $3s^2$
- (b) $4s^2 3d^2$ (d) $4s^2 3d^{10} 4p^3$
22. Why are there no p block elements in period 1 of the periodic table?
23. At one time, scandium ($Z = 21$) was placed in the same group of the periodic table as aluminum ($Z = 13$).
- (a) Use the aufbau principle to write the ground state electron configurations for atoms of these two elements.
- (b) What ionic charge would each have in common? How might this have led chemists to place them originally in the same group?
- (c) Identify and explain all the evidence you can think of to support their current locations in the periodic table.

24. The values for ionization energy in the periodic table in Appendix C are first ionization energies. Construct a bar graph to show the relative sizes of IE_1 values for the main group elements. If available, use spreadsheet software to plot and render your graph.

Making Connections

25. Scientists have succeeded in synthesizing about 25 elements that do not exist in nature. All except one of these (technetium, $Z = 43$) are members of period 7. All are very dense and radioactive. Many, such as Rutherfordium ($Z = 104$) and Bohrium ($Z = 107$), are quite short-lived, with no known applications outside of the laboratory. An exception is the element plutonium ($Z = 94$). Research several applications of this element, and the properties that make it well-suited for these uses. What risks are associated with plutonium? Do the risks outweigh the benefits? Justify your opinions.
26. Synthesizing elements requires the use of equipment called particle accelerators. These machines, usually the size of buildings, are enormously expensive to build and operate. Do you think this money could be better spent elsewhere? Conduct research to help you examine this question thoroughly and thoughtfully. Present your findings in the form of an editorial that either supports or opposes theoretical scientific research of this kind.

(c) $Zn: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$ (d) $O: 1s^2 2s^2 2p^4$

12.(a) Sc, Y, La, Ac (b) N, P, As, Sb, Bi (c) Ne, Ar, Kr, Xe, Rn

13.(a) $K: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$; [Ar] $4s^1$;

(b) $Ni: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$; [Ar] $4s^2 3d^8$;

(c) $Pb: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$;

[Xe] $6s^2 4f^{14} 5d^{10} 6p^2$

Section Review 3.1: 1.(a) mostly empty space; contains a central, massive, concentration of charge. (b) scattering of alpha particles that pass through gold foil 4.(a) radio waves, microwaves, infrared, visible light, ultraviolet, x-rays, gamma rays (b) opposite of (a) 5. UV has higher energy and frequency, lower wavelength; both travel at the same speed and have wave and particle-like properties 6.(a) photons absorbed to excite electron to higher energy level (b) photons emitted as electron falls to lower energy level (c) photons emitted as electrons falls to lower energy level 7. $n = 1, n = 2, n = 4, n = 5, n = 7$ 3.2: 2. Principal quantum number, n , size and energy; orbital-shape quantum number, l , shape; magnetic quantum number, m_l , orientation 5.(a) $l = 0, 1s$ (b) $4f$ (c) $m_l \neq -2$, but could be 0, ± 1 6.(a) $n = 4, l = 1$ (b) $2p$ (c) $3d$ (d) $n = 2, l = 0$ 3.3: 1. magnetic spin quantum number, m_s ; not dependent on value of any other quantum number, always $\pm 1/2$ for any m_l value, not derived from quantum mechanics (describes electron, not orbital) 3.(a) group 3 (IIIB), period 5, d block (b) group 15 (VA), period 4, p block (c) group 18 (VIIIA), period 2, p block (d) group 13 (IIIA), period 3, p block 4. removal of inner electron, being closer to nucleus, Z_{eff} is greater 5. left, ionization energy 7.(a) $1s^2 2s^2 2p^6$, [Ne] (b) $1s^2 2s^2 2p^6 3s^2 3p^6$ [Ar] (c) $1s^2 2s^2 2p^6 3s^2 3p^6$ [Ar] (d) $1s^2 2s^2 2p^6 3s^2 3p^6$; stable ions have a noble gas electron configuration 8. ns^2 , large jump between IE_2 and IE_3 indicates that a noble gas electron configuration was attained after 2 electrons are removed

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: 1.(a) 0, 1, 2, 3, 4 (b) 0

2.(a) $\pm 4, \pm 3, \pm 2, \pm 1, 0$ (b) 0 3.(a) $2s$; 0; one (b) $4f$; $\pm 3, \pm 2, \pm 1, 0$; seven 4. $n = 2, l = 1, 1$ $m_l = 0, \pm 1$ 5. (b), (c), allowed

6. Na: $1s^2 2s^2 2p^6 3s^1$; Mg: $1s^2 2s^2 2p^6 3s^2$; Al: $1s^2 2s^2 2p^6 3s^2 3p^1$; Si: $1s^2 2s^2 2p^6 3s^2 3p^2$; P: $1s^2 2s^2 2p^6 3s^2 3p^3$; S: $1s^2 2s^2 2p^6 3s^2 3p^4$; Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$; Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$ 7. Na: [Ne] $3s^1$; Mg: [Ne] $3s^2$; Al: [Ne] $3s^2 3p^1$; Si: [Ne] $3s^2 3p^2$; P: [Ne] $3s^2 3p^3$; S: [Ne] $3s^2 3p^4$; Cl: [Ne] $3s^2 3p^5$; Ar: [Ne] $3s^2 3p^6$ 9. 2 (IIA): ns^2 ; 13 (IIIA): $ns^2 np^1$; 14 (IVA): $ns^2 np^2$; 15 (VA): $ns^2 np^3$; 16 (VIA): $ns^2 np^4$; 17 (VIIA): $ns^2 np^5$; 18 (VIIIA): $ns^2 np^6$ 10.(a) group 1 (IA), period 3, s block (b) group 2 (IIA), period 2, s block (c) group 17 (VIIA), period 5, p block

11.(a) Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

(b) Rn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$