

**SCH4U-C**



Structure and Properties  
of Matter



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# Introduction

Chemistry is the study of the properties and applications of substances that are all around us. A good place to begin this study is to look at the structure of the fundamental building block of matter—the atom—as you will do in Lesson 1. You can then consider different ways in which atoms link together to form the substances around us. In the lessons that follow, you will learn that the chemical and physical properties of these substances are a direct result of how substances are put together. In other words, the properties of any substance depend on the chemical bonds within that substance. In this unit, you will consider two types of bonding. Understanding the nature of bonding in a substance is the first key step in being able to explain the substance's properties. This understanding is critical because properties ultimately determine how the substance is used, as well as its potential impact on human health and the environment.

Our knowledge of the atom has led to many amazing technologies. One of the most useful applications of these technologies is medical imaging. You will begin your study of the atom by first considering some medical imaging techniques such as the X-ray, a type of ray that can safely penetrate the body to reveal its inner secrets.

## Overall Expectations

After completing this unit, you will be able to

- evaluate the benefits and costs to society and the environment of advances in the study of new materials
- investigate the physical properties of a variety of substances and predict the type of chemical bonds that occur within each substance
- describe how the properties of a substance can be predicted and explained through an understanding of the atomic and molecular structure of the substance

You can access this online [Periodic Table](#) and this [Scientific Calculator](#) any time you need them. You will find links to them on the Course Materials page.

**SCH4U-C**



Atomic Structure



## Introduction

In your local hockey championship, there are only 30 seconds left in the game and your team is down by one goal. You are about to take the most difficult slapshot of your recreational hockey career when suddenly you feel a “pop” in your left knee. The trip to the hospital is long and painful. Once you arrive, the doctor immediately orders an X-ray of the knee. You receive good news from the X-ray—no broken bones. However, you are in a lot of pain, so the doctor needs to view the soft tissue inside the knee as well. So she orders magnetic resonance imaging (MRI). The MRI isn’t promising. It indicates that you have torn the anterior cruciate ligament (ACL), which is perhaps the most crucial of the four ligaments connecting the bones of the knee joint. That amateur hockey career you’ve been dreaming about may be on hold for a while.

MRIs and X-ray imaging are technologies that would not have been possible without some of the basic research that was done by early chemists. You will learn about some of their early discoveries concerning atomic structure in this lesson.

## Planning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Introduction to Atomic Structure and Development of the Atomic Model	$\frac{3}{4}$
Spectroscopy and the Atom	1
Quantum Mechanical Model of the Atom	1
Electronic Configurations	$\frac{3}{4}$
Key Questions	1

## What You Will Learn

After completing this lesson, you will be able to

- assess the benefits to society of technologies that are based on the principles of atomic structures
- write electronic configurations for any element of the periodic table
- explain how experimental observations and inferences led to the development of the modern model of the atom
- identify the characteristic properties of elements in each of the *s*-, *p*-, and *d*-blocks of the periodic table, and explain the relationship between the position of an element in the periodic table, its properties, and its electronic configuration

# Introduction to Atomic Structure

X-rays and MRIs are both examples of medical imaging technologies. However, they differ in their applications. X-rays are ideal for examining dense tissue, like bone, while MRIs are best suited for examining soft tissue, like ligaments. X-rays and MRIs both rely on portions of the electromagnetic spectrum in order to operate. In earlier courses, you learned that the electromagnetic spectrum represents all possible forms of light energy (Figure 1.1). The energy of a particle increases inverse to its wavelength, or in other words, particles with shorter wavelengths have higher energy.

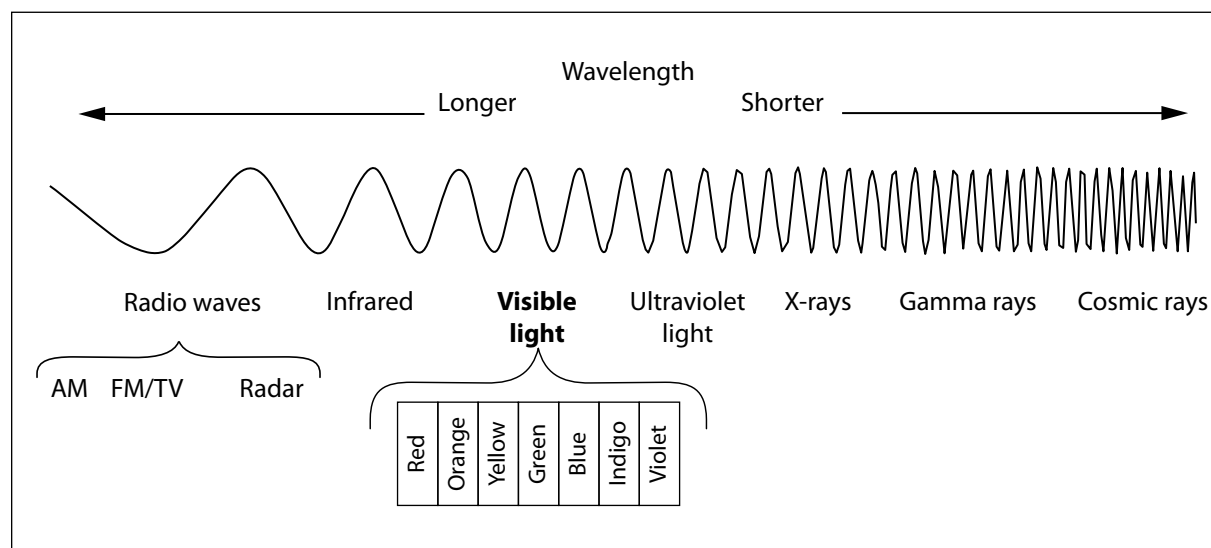


Figure 1.1: The electromagnetic spectrum summarizes all forms of electromagnetic radiation (including light) in order of decreasing wavelength.

X-rays are a high-energy form of electromagnetic radiation (an invisible type of light energy). Because of their energy, X-rays easily pass through soft tissue like skin, lungs, and muscles. However, X-rays cannot penetrate dense material like bone. Consequently, bones appear light in colour, in an X-ray image, while soft tissue appears dark. MRIs use low-energy radio waves, powerful magnets, and an unusual property of the hydrogen nucleus, to generate images of the body's soft tissue.

How can an MRI reveal the structure of soft tissue? MRI images are based on hydrogen atoms, which are plentiful in soft tissue. Experimental evidence suggests that the nucleus of a hydrogen atom spins and wobbles in a cyclical fashion. In other words, it “precesses” (or spins), much like a tiny rotating top (Figure 1.2). The angle of this spin (or precession) is random, from one atom to the next.



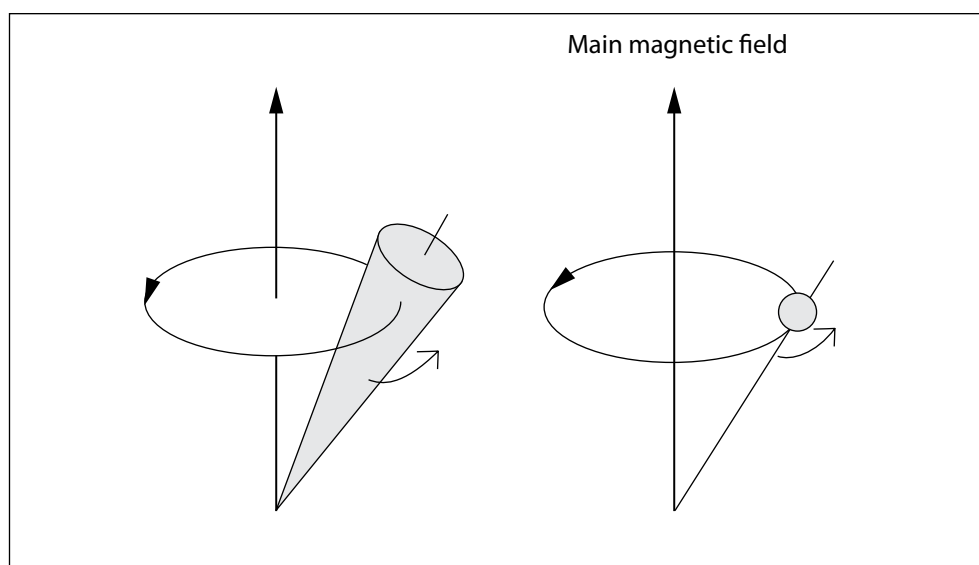


Figure 1.2: The spinning nucleus of the hydrogen atom wobbles in a cyclical fashion (or precesses), like a spinning top.

During an MRI, the patient is exposed to radio waves specifically “tuned” to the precession frequency of the hydrogen atom’s nucleus. As hydrogen nuclei in the body absorb this energy, they align and precess in the same direction. When the radio waves are turned off, the hydrogen nuclei slowly return to their natural random alignment and release the energy they had absorbed, again in the form of radio waves. Detectors surrounding the patient collect this energy and use it to generate an image of soft tissue in the body, such as the brain.

X-rays and MRIs are just two of the many technologies that have resulted from research into the atom’s structure. Our society has clearly benefited from this research. However, fundamental research is expensive, and sometimes it seems that not all people have equal access to the benefits resulting from this research. For example, a recent study found that well-off Ontarians were 38% more likely to have MRIs done than their poorer counterparts.

Now you will examine the atom in greater detail in order to see what other beneficial applications are possible.

## Development of the Atomic Model

By the turn of the twentieth century (1900), the atom was thought to consist of a sphere of positive charge with negative electrons dispersed inside it, much like chocolate chips spread throughout a chocolate-chip muffin (Figure 1.3). This model was created as a result of numerous experiments conducted by scientists during the nineteenth century.

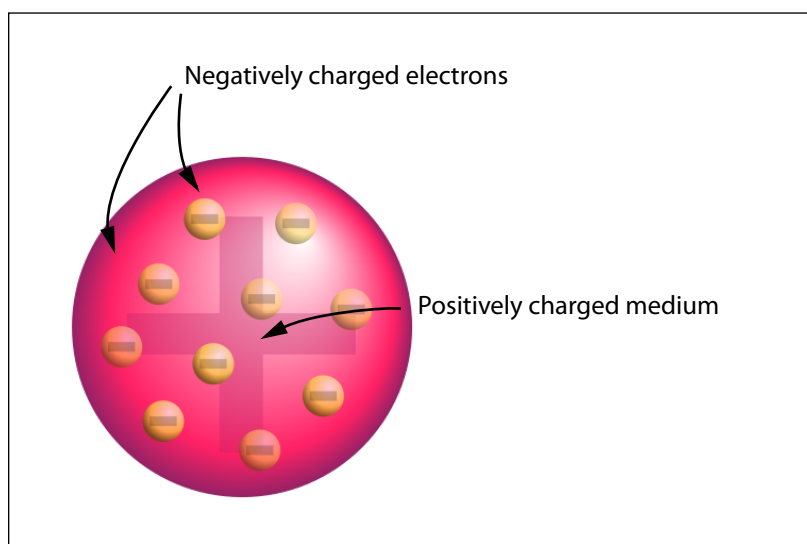


Figure 1.3: The model of the atom in 1900

Source: [http://commons.wikimedia.org/wiki/File:Plum\\_pudding\\_atom.svg](http://commons.wikimedia.org/wiki/File:Plum_pudding_atom.svg)

However, this model of the atom was about to undergo a radical transformation. The first major change came about as a result of one of the most famous experiments in the history of chemistry—the Rutherford gold foil experiment. In this experiment, Ernest Rutherford bombarded a very thin sheet of foil, made of gold, with alpha particles. (Alpha particles are positively charged particles, each containing two protons and two neutrons, and are symbolized by the Greek letter,  $\alpha$ . See Figures 1.4 (a) and (b).) Rutherford's source of alpha particles was a sample of the radioactive element, polonium. Elements that are radioactive have unstable nuclei. These atoms spontaneously decay (or fall apart), emitting tiny fragments of themselves. One of the decayed products of polonium is an alpha particle.

Figure 1.4 shows the set-up for the Rutherford gold foil experiment. Given the understanding of the atom at the time, Rutherford expected that most of the positive alpha particles would pass straight through the atoms in the gold foil. In fact, most did. However, much to his surprise, a small fraction of alpha particles deflected at wide angles and fewer still rebounded back toward the source. Rutherford explained these observations by inferring that these few alpha particles must have hit a small object of dense positive charge.

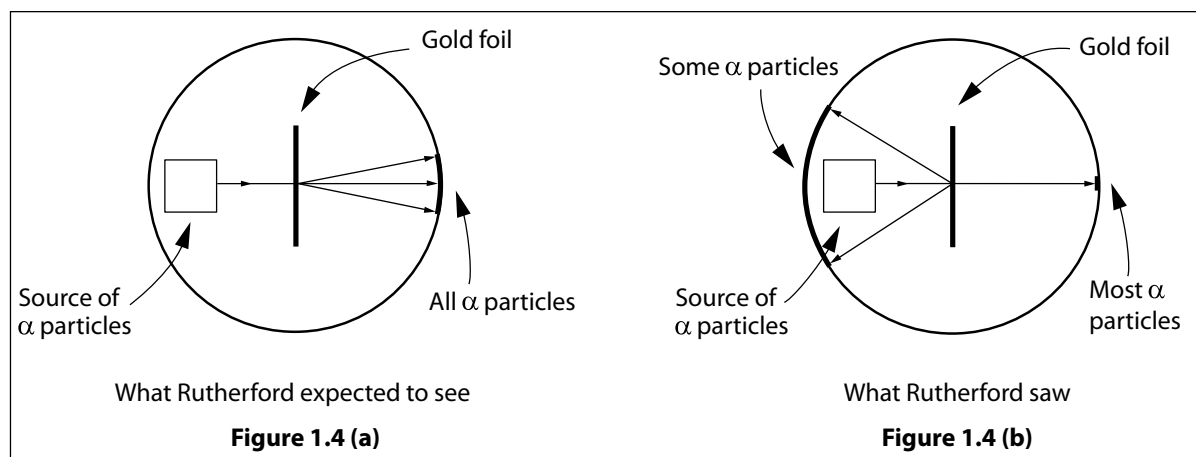


Figure 1.4: (a) Rutherford predicted that most of the alpha particles would pass directly through the gold foil, without deflecting. (b) Most did. However, a few were widely deflected.

Rutherford had discovered the nucleus! Rutherford also inferred that since most alpha particles went through the foil without deflecting, most of the atom must consist of empty space. As a result of Rutherford's work, the model of the atom changed to a small, dense, positive core called the nucleus, surrounded by orbiting electrons (Figure 1.5). Interestingly, if a typical atom could be enlarged to a diameter of one kilometre, then the atom's nucleus would be the size of a marble measuring one centimetre.

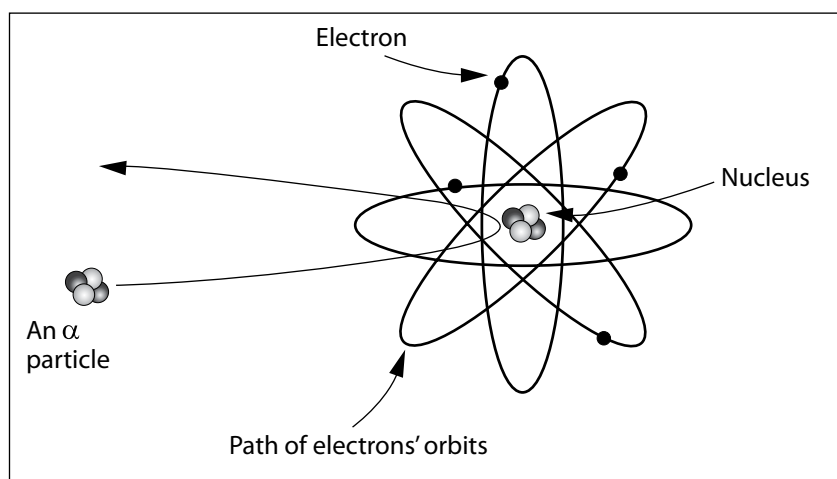


Figure 1.5: The Rutherford model of the atom

## Support Questions

Be sure to try the Support Questions on your own before looking at the suggested answers provided. Click on each “Suggested answer” button to check your work.

1. In December 2007, the nuclear reactor at Chalk River, Ontario (near Ottawa), belonging to Atomic Energy of Canada Limited (AECL), was shut down for maintenance. This reactor produces much of North America’s supply of technetium-99m, an important radioactive chemical used in the diagnosis and treatment of many major diseases.
  - a) Research the properties of technetium-99m and its medical applications, and then write a brief paragraph about these properties.
  - b) Why can’t technetium-99m be stockpiled, like other medications?
  - c) Critics argue that the Chalk River reactor is old and potentially dangerous. However, building a new nuclear reactor is very expensive. In your opinion, should Canadian tax dollars be used to replace the reactor? Explain your reasoning.

## Spectroscopy and the Atom

The next major development in our understanding of the atom came from a branch of chemistry called spectroscopy. Spectroscopy is the study of energy that is absorbed or emitted by matter. By the mid-nineteenth century, chemists already knew that each element of the periodic table (which is reproduced at the end of Unit 1) emits its own characteristic spectrum of light when it absorbs energy from some external source.

For example, a sample of hydrogen gas emits purplish light when electricity is passed through it. When this purplish light is passed through a prism, it is found to be made up of two violet lines of colour, a blue-green line, and a red line (Figure 1.6). Regardless of the source of the hydrogen, these lines always appear in the same position, with the same light wavelengths. The red wavelength is always 656.3 nanometres (abbreviated as nm), the blue-green wavelength is 486.1 nm, the blue-violet one is 434.1 nm, and the violet one is emitted at 410.2 nm. This pattern of lines emitted by elements became known as a line spectrum.

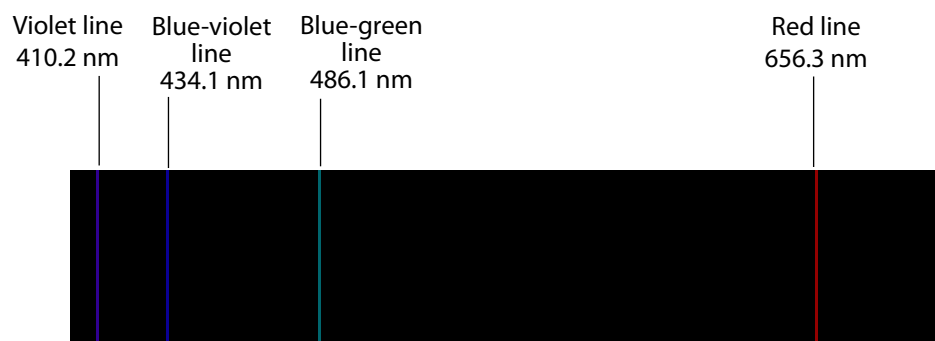


Figure 1.6: The line spectrum of hydrogen

Source: [http://en.wikipedia.org/wiki/File:Hydrogen\\_Emission\\_Spectrum.svg](http://en.wikipedia.org/wiki/File:Hydrogen_Emission_Spectrum.svg)

The origin of these spectral lines baffled scientists until the early twentieth century. In 1913, the Danish physicist, Niels Bohr, proposed an explanation for the origin of line spectra that required the current model of the atom to be changed. Bohr suggested that the distinct colours in line spectra were emitted by electrons. He proposed that electrons orbit the nucleus in very specific orbits, and that each orbit has a distinct energy. When the electron absorbs energy from some external source, it jumps to a higher orbit or energy level. A “jump” from one level to another is called an electron transition. The attraction of the nucleus eventually pulls the electron back to a lower energy level. When it does, the energy that the electron absorbed is emitted. The energy emitted by the atom, according to Bohr, is the difference in energy between the two energy levels that the electron occupied. In the case of hydrogen, the four lines of the visible spectrum correspond to the electron transitions shown in Figure 1.7.

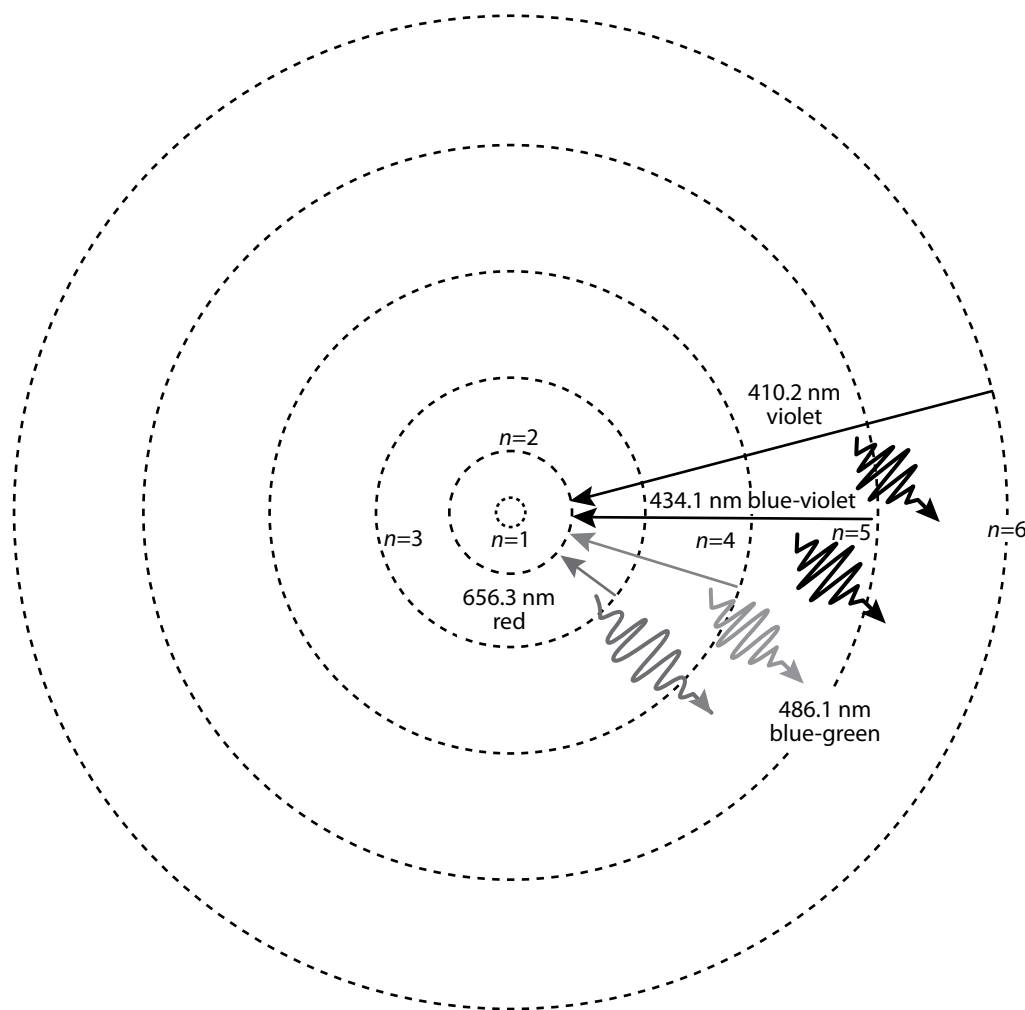


Figure 1.7: The four colours in the hydrogen spectrum are the result of four specific electron transitions that produce light in the visible spectrum.

Many other electron transitions are also possible—but the light they emit, such as infrared or ultraviolet light, isn't in the visible spectrum.

Now, go to the website of the University of Colorado in Boulder and search for “Models of the Hydrogen Atom.” Click on this link to open the applet. (If you have trouble locating this applet, do a web search using the terms “Colorado” and “phet” and “models of the hydrogen atom.”) You will start by selecting “Prediction,” and then “Bohr,” followed by “Show spectrometer,” and “Show electron energy level diagram.” Finally, you should select “White,” and then turn on the white light source by clicking on the red button, to begin the experiment. Observe the spectrometer bar graph (to see the light that hydrogen emits) and the electron energy-level diagram (to see the electron transitions that produce this light). If you are not getting results fast enough, try adjusting the slider at the very bottom of the page to “fast.” This applet will help you to understand how scientists first conceived of the electron levels or shells, and it will help you to answer the Support Questions later on in the lesson.

## X-rays



Figure 1.8: X-rays are used to create images like this one.

Source: [http://commons.wikimedia.org/wiki/File:Rad\\_160601\\_Skull\\_X-ray\\_False\\_color.nevit.jpg](http://commons.wikimedia.org/wiki/File:Rad_160601_Skull_X-ray_False_color.nevit.jpg)

One method of producing X-rays also involves electron transitions. Some elements will produce X-rays when electron beams pass through them. Inside an X-ray machine, atoms of a suitable element are first bombarded with a beam of electrons. This causes some of the inner orbit electrons to be “bumped” out of the atom, forming a temporary “hole.” The resulting electron hole can be filled by any electron in the atom’s intermediate or outer orbits. If the electron comes from an outer orbit, flashes of high-energy X-rays are emitted. The reason that these emissions have high energy is because the energy difference between the two electron orbits is large.

### Support Questions

- Ernest Rutherford used the element polonium in his gold foil experiment. What property of the element polonium makes it extremely dangerous to work with?
- Rank the following forms of electromagnetic radiation in order of increasing energy: visible light, gamma rays, radio waves, infrared waves
- Compare the energy of red, blue-green, blue-violet, and violet light illustrated in Figure 1.7.
- Explain why atomic emission spectra are sometimes called “fingerprints” of the elements.

6. There are only four lines in the visible spectrum of hydrogen. Explain why an electron transition from  $n = 6$  to  $n = 1$  would not appear in the visible spectrum shown in Figure 1.6.

## Limitations of the Bohr Model of the Atom

The Bohr model of the atom successfully explains the origin of the hydrogen line spectrum. In fact, it also explains the spectra of other single electron systems, such as  $\text{He}^+$  and  $\text{Li}^{2+}$ . However, the Bohr model is somewhat limited in explaining the properties of multi-electron atoms. For example, look at Figure 1.9, which shows emission spectra of hydrogen, mercury, and neon.

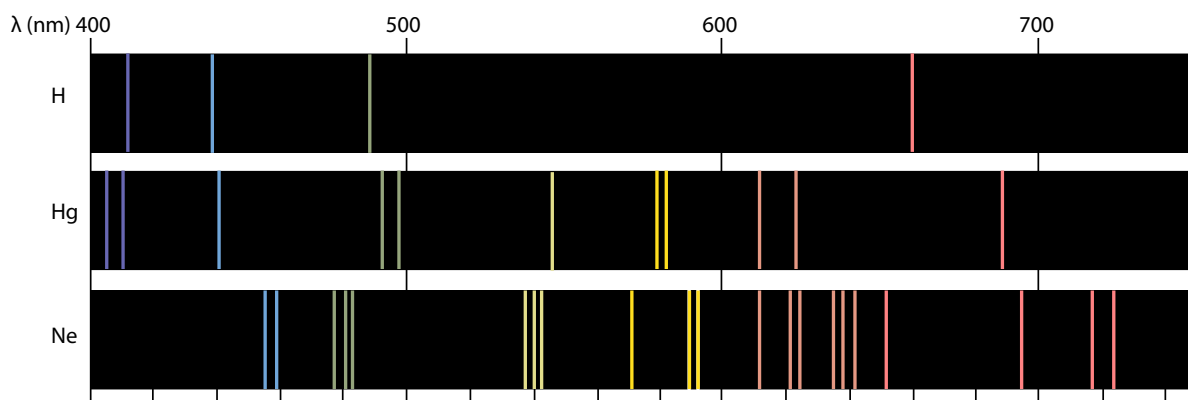


Figure 1.9: The emission spectrum of hydrogen, mercury, and neon

Note that the spectra of the larger atoms are far more complex than the spectrum of hydrogen. The large spaces between spectral lines can be successfully explained by Bohr's idea of energy levels. However, the small spaces between some of the lines can't be explained using the Bohr model. These spectral lines are results of electron transitions between energy levels. Therefore, small gaps between spectral lines suggest that there are sub-levels between the atom's main energy levels. Attempts to include sub-levels in the Bohr model had only limited success. It soon became clear that this model of the atom had to be revised.



## Quantum Mechanical Model of the Atom

The development of the next atomic model began with the discoveries of some rather peculiar properties of the electron. In other science courses, you may have learned that electrons travel as waves. Yet, you have also read that the electron is a negatively charged particle. Understandably, this is very confusing. Are they waves or particles?

### Are Electrons Waves or Particles?

Many electrical phenomena, such as static electricity, can easily be explained by negatively charged particles (electrons) moving from one place to another. However, by the 1920s, there was considerable evidence to support the idea that electrons could behave as a wave of energy as well. For example, in 1927, C. J. Davisson and L. H. Germer conducted an experiment in which electrons were fired at a nickel crystal. A detector was used to trace the path of the electrons after their collision with the crystal.

The results of the experiment produced a ring pattern identical to the one produced when X-rays were fired at the crystal (Figure 1.10). The production of rings like this is a characteristic property of waves, known as diffraction. X-rays are waves that are part of the electromagnetic spectrum.

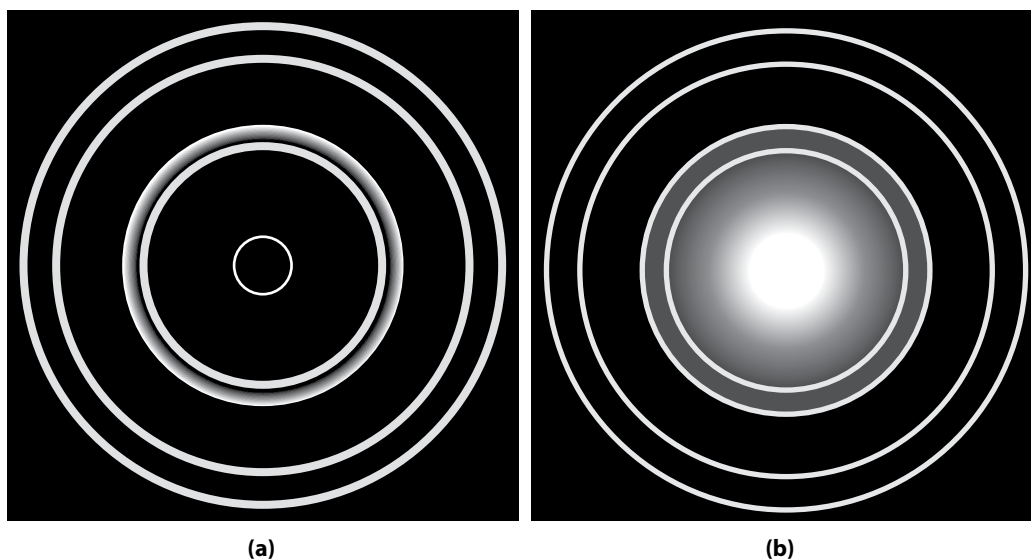


Figure 1.10: Diagram (a) shows the typical pattern of rings produced when X-rays are scattered by a nickel crystal. Diagram (b) shows a similar pattern, produced when electrons are scattered by the crystal.

The dual nature of the electron (behaving both as a particle and a wave) was one of the key ideas that led to the current model of the atom—the quantum mechanical model. A detailed description of this model requires an understanding of complex mathematics and therefore is far beyond the scope of this course. One of the key ideas in this model is that the exact location and speed of the electron can never be simultaneously known. Consequently, it is inaccurate

to describe the electron as a particle orbiting a fixed path around the nucleus, as in Bohr's idea of an orbit. Instead, quantum mechanics uses the idea of probability to describe the location of the electron. According to quantum mechanics, there is a region of space around the nucleus in which the probability of finding the electron is high. These high probability regions or electron clouds are called orbitals (Figure 1.11).

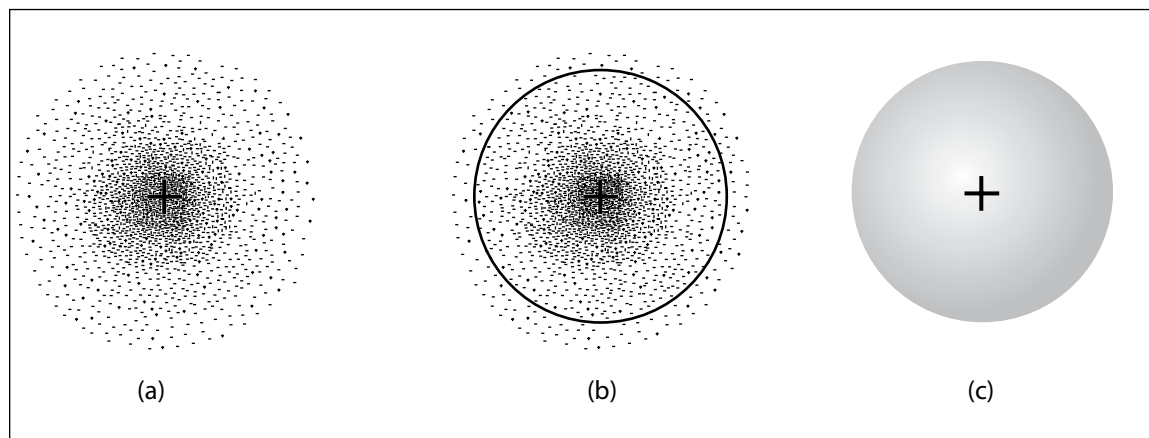


Figure 1.11: Orbitals are regions of space around the nucleus where the probability of finding the electron is high. Note that this is a spherically symmetrical orbital or “s orbital.”

Here's an analogy to help you visualize what an orbital is. Suppose that the plus sign in Figure 1.11 (a) is a beehive. Each dot in the picture represents flowers visited by the bees of the hive. Note that the density of visits is greatest near the hive and then decreases as the distance from the hive increases. Similarly, in an atomic orbital, each dot represents the location of the electron at any given moment. The electron density is greatest at the centre of the orbital—the nucleus. Electron density decreases as the distance from the nucleus increases. Theoretically, there is no outer limit to how far the electron can be away from the nucleus. However, for practical purposes, each orbital has an outer boundary, as shown in Figure 1.11 (b) and (c). This area within the orbital represents the region around the nucleus in which the probability of finding the electron is about 95%.

The shape and energy of orbitals is determined by solutions to a complex mathematical equation known as the Schrödinger wave equation. The results of this equation show that there are four different orbitals, which are identified by the letters *s*, *p*, *d*, and *f*.

## Orbitals

Figure 1.12 represents the four different possible orbitals. Each orbital represents a region of space around the nucleus where the electron can be found. The *s* orbitals are arranged in a spherically symmetrical manner around the nucleus. There is only one type of *s* orbital. Each *s* orbital can accommodate two electrons.

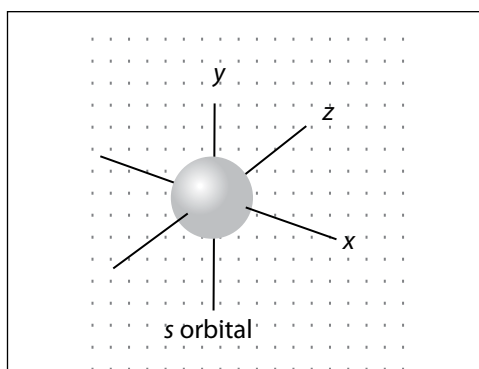


Figure 1.12: The  $s$  orbital is spherical. The centre of this three-dimensional graph is the nucleus.

There are three different  $p$  orbitals. Each looks like a dumbbell that is arranged around one of the three coordinate axes:  $x$ ,  $y$ , and  $z$ . The orbitals can each hold up to two electrons. This means that a complete set of  $p$  orbitals can hold up to six electrons.

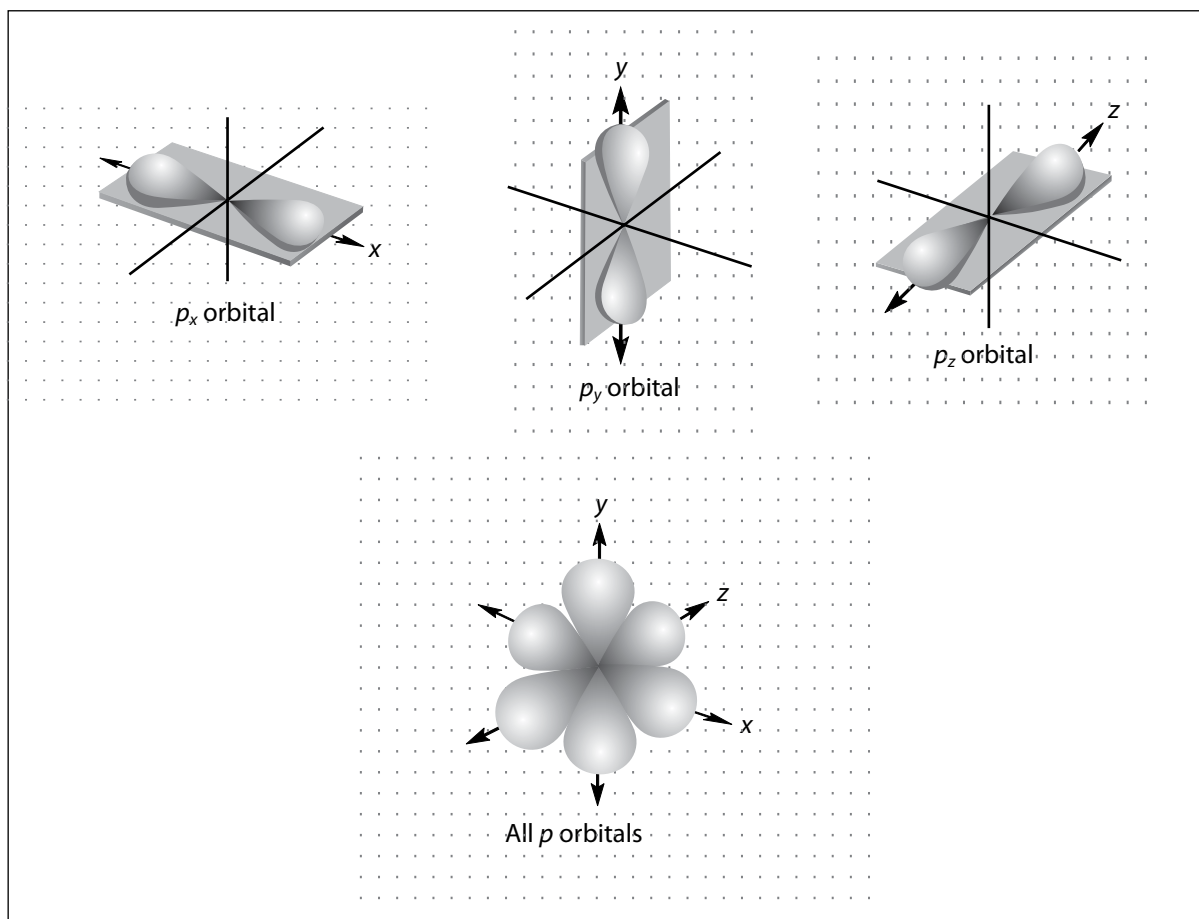


Figure 1.13: There are three different  $p$  orbitals— $p_x$ ,  $p_y$ , and  $p_z$ .

There are also five *d* orbitals and seven *f* orbitals. The complex, three-dimensional shape of these orbitals, which will not be illustrated, is beyond the scope of this course. Like *s* and *p* orbitals, each *d* and *f* orbital can hold up to two electrons. Therefore, *d* orbitals can accommodate up to 10 electrons, while *f* orbitals can accommodate 14. Table 1.1 summarizes the four types of orbitals and the number of electrons they can accommodate.

**Table 1.1: Four orbitals—*s*, *p*, *d*, and *f***

Orbital	<i>s</i>	<i>p</i>	<i>d</i>	<i>f</i>
Quantity of orbitals	1	3	5	7
Maximum number of electrons they can hold (2 electrons per orbital)	2	6	10	14

## What Do Atoms Look Like?

Hydrogen is the simplest of all atoms. Since it has an atomic number of 1, it has only one proton and one electron. The single electron of hydrogen occupies an *s* orbital. Therefore, the hydrogen atom is always shown as being spherical. Most atoms (other than hydrogen and helium) consist of a nucleus surrounded by a number of different orbitals blended together. Since orbitals are three-dimensional waves, it's difficult to imagine what the “blended” orbitals may look like. To get a sense of how this “blending” works, you can review a two-dimensional model of waves “blending” together. When two approaching wave pulses meet, they momentarily blend together to form one large pulse (Figure 1.14). But as two pulses continue, they emerge and continue in the direction they were travelling. Unlike water waves, however, the blended orbitals in atoms stay together.

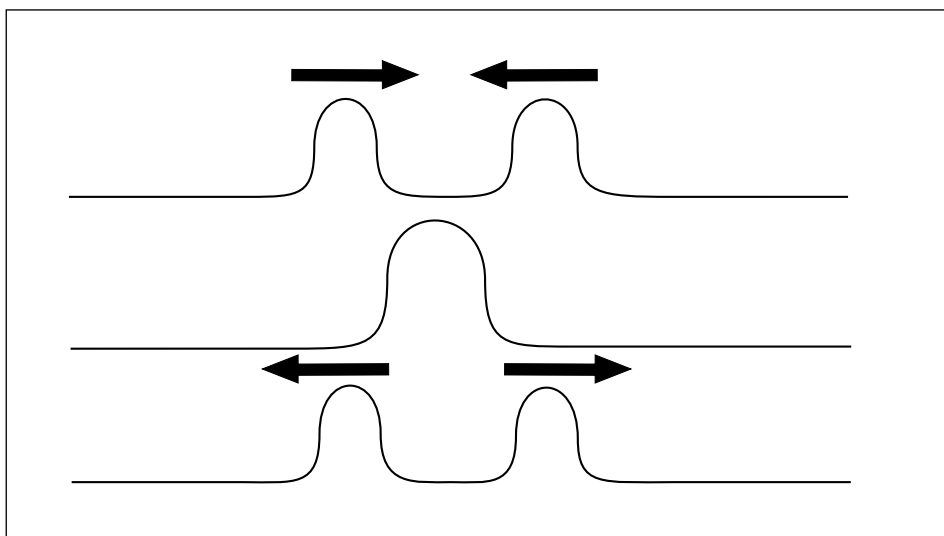


Figure 1.14: Two-dimensional wave pulses meet and blend to form one large pulse (middle). Since atomic orbitals are three-dimensional waves, different orbitals blend together to form the shape of an atom.

An atomic orbital is like a three-dimensional wave. Therefore, the shape of an atom is the result of the blending of its orbitals. If an atom contained  $s$  and  $p$  orbitals, these orbitals would combine and blend together like the combining wave pulses in Figure 1.15, to produce a new wave that was a blend of the two shapes. The result would be an atom that was neither spherical nor shaped like a dumbbell, but rather something in between.

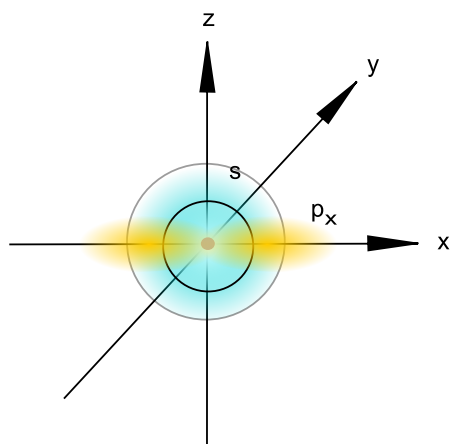


Figure 1.15: An  $s$  and a  $p$  orbital combine to produce a new orbital that is a “blend” of both sphere and dumbbell.

Source: [http://commons.wikimedia.org/wiki/File:Orbitale\\_s\\_px.svg](http://commons.wikimedia.org/wiki/File:Orbitale_s_px.svg)

In summary, the quantum mechanical model suggests that atoms consist of a nucleus surrounded by a combination or “blend” of orbitals containing electrons. In order to explain the properties of the elements, it is necessary to know the energy of these orbitals and their electrons.

## Energy-level Diagrams

The energy-level diagram (Figure 1.16) shows the relative energies of the orbitals in an atom under normal conditions. Comparing the energies of the orbitals and the number of electrons they occupy is useful to explain and predict the properties of the elements of the periodic table.

Each circle represents one orbital. Each orbital can accommodate up to two electrons and is labelled with a number followed by a letter, such as  $2s$ . The number in the orbital label gives the energy level of the orbital. For example, the first energy level ( $n = 1$ ) contains only one orbital— $1s$ . This orbital has the lowest energy of all. You can think of it as the orbital closest to the nucleus. The  $s$  orbital is analogous to the Bohr idea of the first orbit around the atom. The second energy level ( $n = 2$ ) contains two types of orbitals— $s$  and  $p$ . Since they are slightly different in energy, they are often referred to as sub-levels of the second energy level.

We can use a simple mathematical formula to calculate the number of orbitals (sub-levels) associated with a given level,  $n$ . This formula is  $n^2$ . For example, if we are looking at  $n = 2$ , the formula,  $n^2$ , tells us that there will be  $2^2$  sub-levels or orbitals associated with the second level in the atom. So for  $n = 2$ , we will have four orbitals associated with  $n = 2$ . The first of these four would be the  $s$  orbital and the other three would be the group of

three orbitals of the  $p$  type (for the second level,  $n = 2$ , we would call that group of  $p$  orbitals, the  $2p$  group of orbitals).

If we moved to the fourth level,  $n = 4$ , the formula  $n^2$  shows us that there will be  $4^2$  orbitals associated with the fourth level. Those 16 orbitals will contain all four types of orbitals,  $s$ ,  $p$ ,  $d$ , and  $f$ . The first one will be the  $s$  orbital at the fourth level (called  $4s$ ), the next three orbitals will be the group of  $p$  orbitals (three of them called the  $4p$  group). The next type, the  $d$  type, would have a group of five orbitals of the  $d$  type (five of them called the  $4d$  group), and the remaining seven orbitals of the 16 would be the group of  $f$  type orbitals (seven of them called the  $4f$  group).

Table 1.2 shows the orbitals that would be associated with a given level. We will use this table later.

**Table 1.2: Listing of orbitals associated with energy levels,  $n$**

Level, $n$	Number of orbitals, $n^2$	List of orbitals associated with level $n$
1	$1^2 = 1$	$1s$
2	$2^2 = 4$	$2s$ $2p$
3	$3^2 = 9$	$3s$ $3p$ $3d$
4	$4^2 = 16$	$4s$ $4p$ $4d$ $4f$
5	16 (maximum)	$5s$ $5p$ $5d$ $5f$
6	16	$6s$ $6p$ $6d$ $6f$
7	16	$7s$ $7p$ $7d$ $7f$

If you look at Figure 1.16, you can see those 16 orbitals associated with the fourth level ( $4s$ ,  $4p$ ,  $4d$ , and  $4f$ ). Locate them and count them. Notice how the  $s$ ,  $p$ ,  $d$ , and  $f$  orbitals have progressively higher energy levels. This means that electrons in a  $4p$  orbital will have higher energy than electrons found in a  $4s$  orbital. Figure 1.16 illustrates the relative energy levels of all orbitals that could exist in a given atom.

Note that as you proceed to higher energy levels, the space between energy levels decreases. This results in some rather unusual orbital arrangements. For example, Figure 1.16 indicates that the  $3d$  orbitals are closer in energy to the  $4s$  and  $4p$  orbitals than they are to the  $3s$  and  $3p$  orbitals.

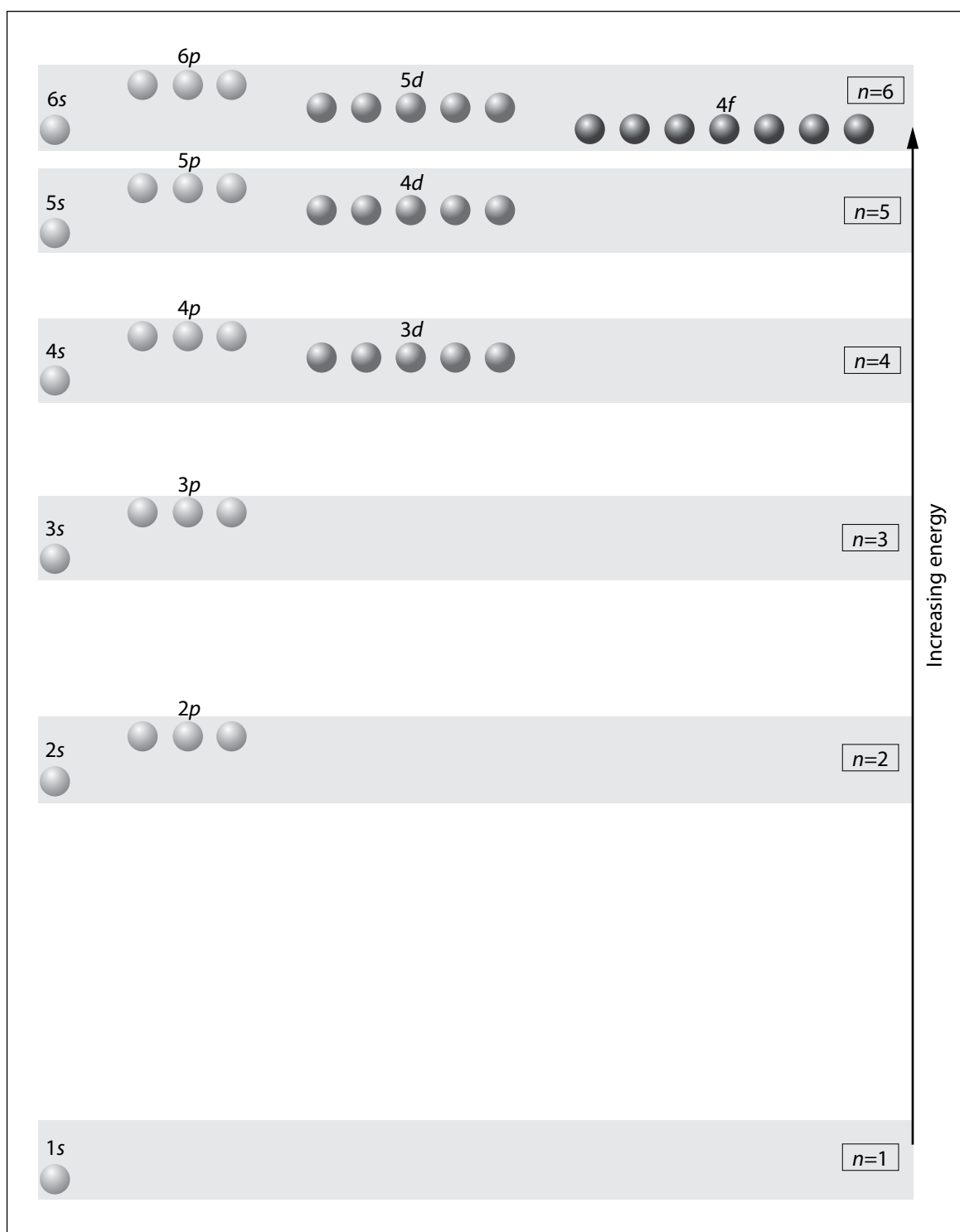


Figure 1.16: Master energy-level diagram

## Rules for Filling Energy-level Diagrams

As you'll soon see, these diagrams are a powerful tool to help explain trends in the properties of the elements. But first, some basic rules need to be established for filling these diagrams with electrons.

### 1. Pauli Exclusion Principle ("Opposite Spin")

When filling these diagrams, electrons are often represented as arrows (Figure 1.17). The direction of the arrow is used to represent a property of the electron called its spin. Experimental evidence suggests that the electron spins, much like a toy top. Further evidence suggests that there are two possible spin states for the electron. These are represented by either an up arrow or a down arrow. If there is only one electron in the orbital, as is the case for hydrogen, then the spin state of the electron can be either up or down. However, if the orbital contains a second electron, the two electrons must have opposite spins. Chemists have found that electron repulsion is minimized if the two electrons spin in opposite directions. A second electron will be excluded from entry if it does not have the opposite spin. This is known as the Pauli exclusion principle.



Figure 1.17: The energy-level diagrams for hydrogen and helium

**Note:** In this lesson, you will need to draw diagrams such as the one above. The best option is to draw them with your word processor's drawing tools, if possible. You may also draw them on paper.

If you are drawing with a pen or pencil, you should insert your drawings into a word-processing document for online submission by scanning (or taking a picture of) your drawing. For instance, the camera on a cellphone can be used to photograph each drawing, and you could then email or transfer your drawings to your computer for insertion in a document. If possible, reduce the size (especially the file size) of your scan or photograph so that your online submission doesn't become too large (with too many megabytes).

### 2. Aufbau Principle ("Lowest Fill First")

When orbitals are filled with electrons, orbitals with the lowest energy are filled first.

Chemists refer to this as the Aufbau principle. Therefore, the two electrons of helium are normally found in the 1s orbital. The third and fourth electrons in beryllium are found in the 2s orbital, because it is the lowest level after the 1s orbital has been filled by two electrons.



The order that orbitals are filled must be from lowest energy level (1s) and upward. Figure 1.16 could be used to determine the order by listing the orbitals from the bottom of the diagram to the top in the order that they are found on that figure. However, this is tedious, and a simpler tool can be used to produce the correct order of filling the orbitals. This tool is a slight modification of Table 1.2. By modifying this tool, we can create what is termed, “the diagonal rule” for filling orbitals. A series of parallel, diagonal lines is drawn on the table as shown in Figure 1.18.

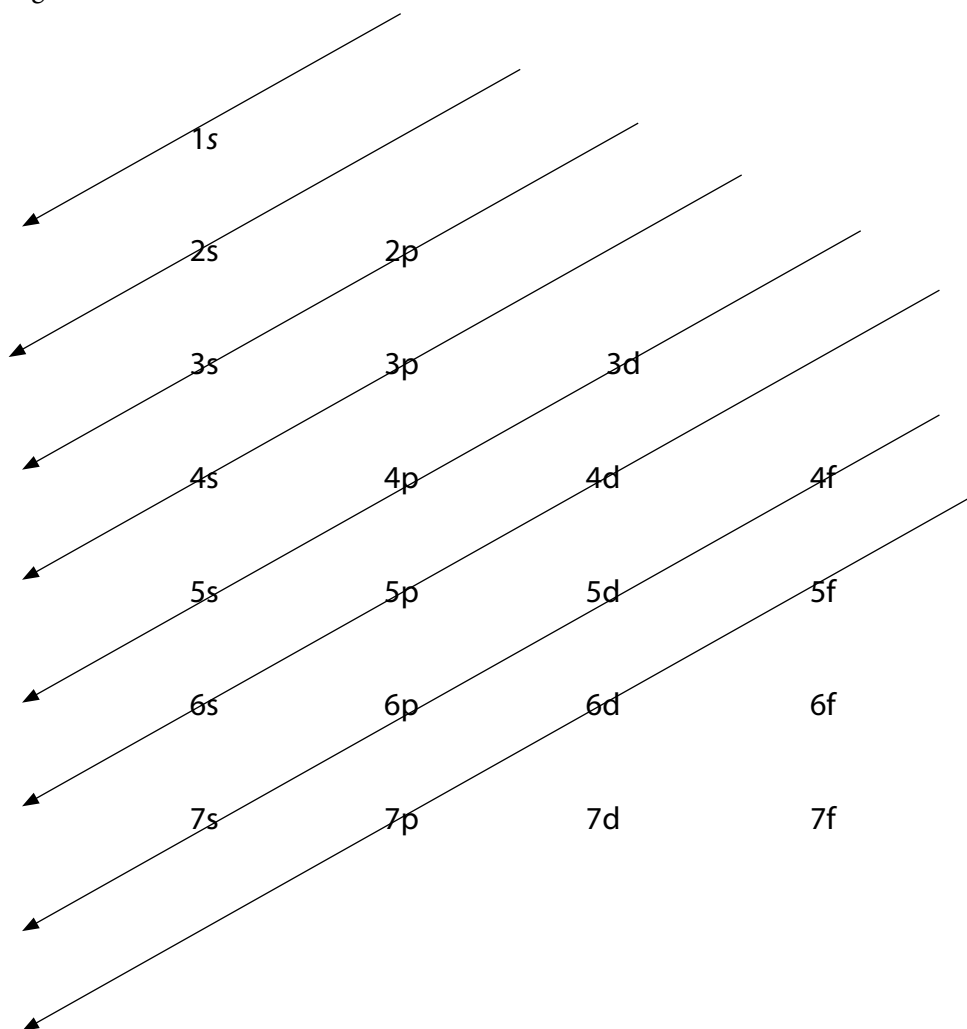


Figure 1.18: The diagonal rule for orbital order

The order to fill orbitals can now be listed by tracing along each diagonal line starting with the top diagonal and reading downward to the left along that diagonal. That first diagonal contains just the 1s orbital, so our list begins with 1s. We now move down to the second diagonal and read along it, reading again downward to the left again. The next orbital to fill is the 2s. Reading along the third diagonal now we have 2p, then 3s. The next diagonal gives us 3p and 4s. The fifth diagonal produces 3d, 4p, then 5s. Hopefully the pattern becomes clear. The order of filling produced by this diagonal rule will appear as follows:

[1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d...]

This tool will allow us to describe the location of electrons, in general, for any atom on the periodic table. As with many chemistry concepts, this tool is not foolproof. You may encounter exceptions to this rule in your studies.

## Example

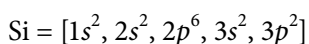
Let's now make use of this tool to describe the electron arrangement for a few elements on the periodic table. We will begin with a simple atom: silicon, Si, with atomic number 14.

## Solution

As silicon has 14 electrons, we will want to describe in which orbitals these 14 electrons would be found. From Figure 1.18, we have established, the diagonal rule, the order that orbitals will be filled with electrons, so we simply follow that order of filling in the list above.

Recall that each type of orbital can accommodate two electrons. Consequently any *s* orbital could hold a maximum of two electrons. Because *p* orbitals exist in groups of three in a set, a group of *p* orbitals—such as the *2p* group or the *3p* group—could accommodate a total of six electrons. A group of *d* orbitals—such as the *3d* group or the *4d* group—containing five orbitals in the group can accommodate 10 electrons, while any group of *f* orbitals—such as the *4f* or *5f* group—could accommodate 14 electrons, since *f* orbitals exist in groups of seven in a set. Review this from the earlier discussion of orbital types and commit this idea to memory. Orbital types (*s*, *p*, *d*, and *f*) being in sets of (1, 3, 5, and 7 orbitals) in each set respectively, accommodate (2, 6, 10 and 14) electrons respectively.

Let's now use these ideas to describe the electron arrangement (configuration) for the 14 electrons found in any silicon atom. The electron arrangement would be described this way:



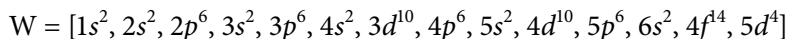
In that listing, one of the entries reads "*2p*<sup>6</sup>". The *2p* part is the set of *p* orbitals associated with the second level, *n*=2. Remember that *p* orbitals are always in a set of three. The 6, shown as a superscript, tells us that those *p* orbitals contain six electrons, of course implying that those three orbitals must be full. If you add all of the superscripts, the result is 14 (electrons).

## Example

A second example is the element tungsten, W, with atomic number 74.

## Solution

If we follow the order of orbital filling from the diagonal rule, the electron arrangement (configuration) of the 74 electrons would be written this way:



When the elements are arranged according to increasing atomic number, as they are in the modern periodic table, they align into four distinct sets (Figure 1.25). The elements of groups 1 and 2 (including helium) have electronic configurations ending in an *s* orbital. These elements are sometimes referred to as the “*s*-block” elements.

Groups 13 through 18 are sometimes called the “*p*-block” elements because their configurations end in *p*. The transition metals are *d*-block elements. The lanthanides (elements 57 to 71) and actinides (elements 89 to 103) make up the *f*-block elements. Each row (also called a period) on the periodic table corresponds to an energy level. The outermost electrons of the second row of elements, for example, are in the second energy level.

### 3. Hund's Rule (“Align Spins”)

When electrons are placed within a sub-level in which the orbitals have identical energy (*p*, *d*, or *f*), electron repulsion is minimized if each orbital receives one electron before pairing occurs. This is known as Hund's rule. For example, Figure 1.19 shows the preferred arrangement for nitrogen's *2p* electrons.

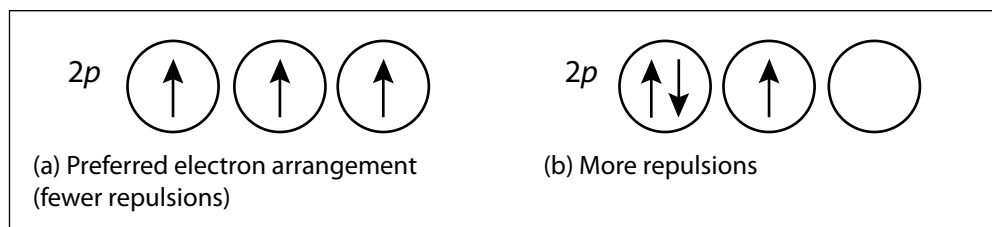


Figure 1.19: For orbitals that are identical in energy, such as *p* orbitals, having the electrons spread throughout all of the orbitals and spinning in the same direction (a) results in fewer repulsions than in the (b) arrangement.

Now you can apply the principles you've learned so far to a few small atoms: lithium, nitrogen, and oxygen.

Since lithium has an atomic number of 3, it contains three electrons and three protons. Therefore, the *1s* orbital fills first because it is lower in energy than the *2s* (due to the Aufbau principle), and the two electrons in the *1s* orbital have an opposite spin (due to the Pauli exclusion principle).

In nitrogen, the *1s* and *2s* orbitals fill first because they are the lowest in energy. The remaining electrons go into the *2p* orbitals because they are next lowest in energy. Note that each of the *2p* orbitals gets one electron before pairing occurs (due to Hund's rule).

The eighth electron in oxygen can go into any of the *2p* orbitals because they are equivalent in energy. However, its spin must be opposite (due to the Pauli exclusion principle).

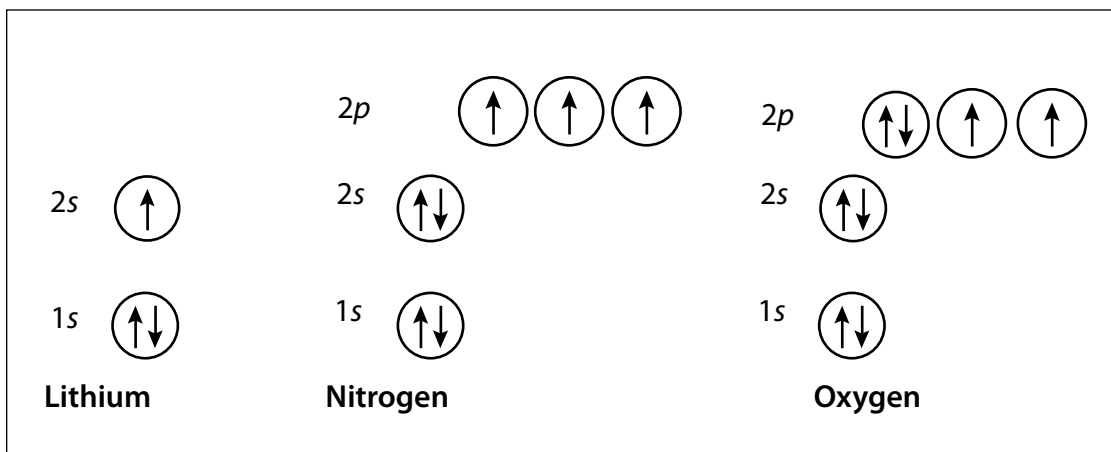


Figure 1.20: Energy-level diagrams for lithium, nitrogen, and oxygen

## Example

Draw the energy-level diagram for fluorine.

## Solution

Since fluorine has atomic number 9, it has nine electrons in its orbitals.

1. The first two electrons occupy the lowest orbital—1s.

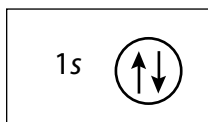


Figure 1.21

2. The next orbital to be filled is the 2s orbital.

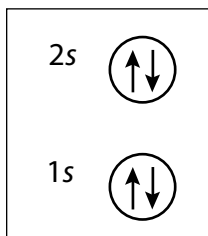


Figure 1.22

3. The remaining five electrons are added to the  $2p$  orbital.

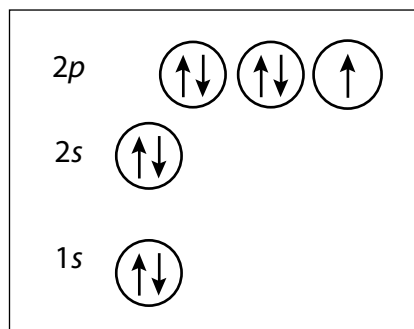


Figure 1.23

## Support Questions

7. How do the terms “orbit” and “orbital” differ?
8. Write the electron arrangement for each element using the diagonal rule, and draw the energy-level diagram for each. Indicate whether each element belongs in the  $s$ -block,  $p$ -block,  $d$ -block, or  $f$ -block of elements.
  - a) aluminum, Al
  - b) chlorine, Cl
  - c) scandium, Sc

## Electronic Configurations

Energy-level diagrams for large atoms are complex and tedious to draw. An abbreviated form of these diagrams, called electronic configurations, can be used to convey the electron arrangements of the elements. Like energy-level diagrams, electronic configurations give the correct order of orbitals from lowest to highest energy, as well as the number of electrons that occupy each orbital. This information is listed in order from left to right. For example, the energy-level diagram of fluorine is shown in Figure 1.24.

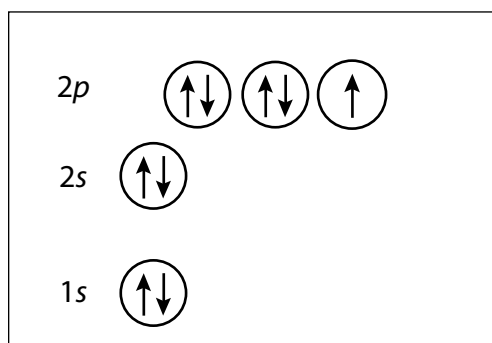


Figure 1.24: Energy-level diagram of fluorine

The corresponding electronic configuration for fluorine is  $1s^2 2s^2 2p^5$  and is read as: 1s two, 2s two, 2p five.” Note that the orbitals are given in order of increasing energy, as in the energy-level diagram. The superscript above each orbital letter indicates the number of electrons occupying the orbital.

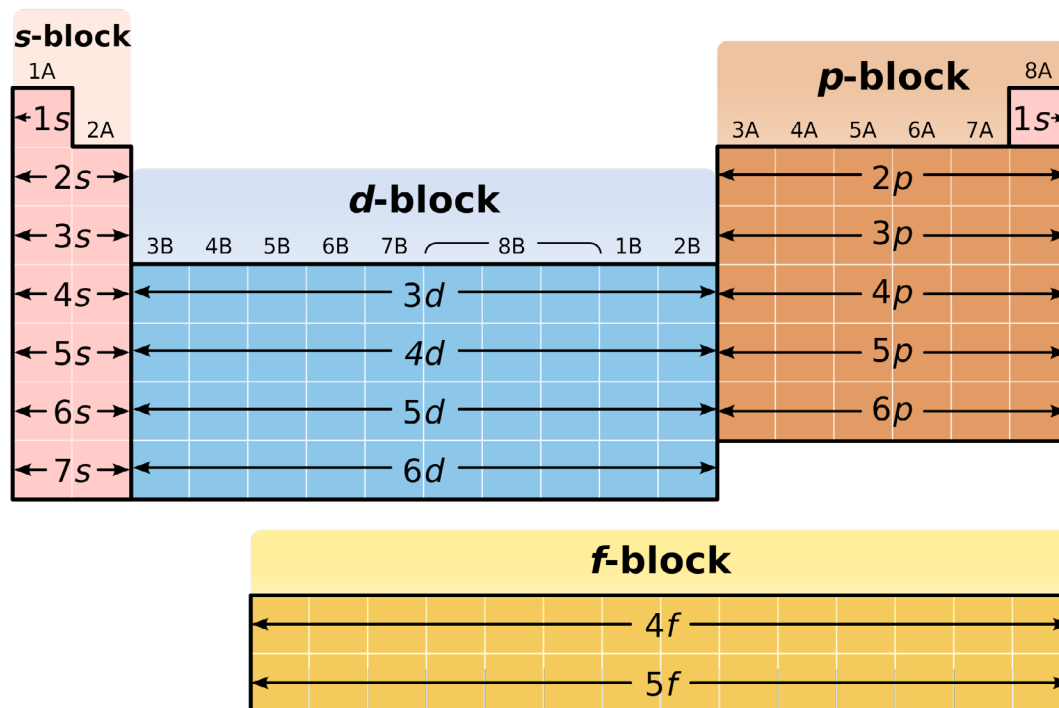


Figure 1.25: The periodic table is divided into the s sub-level block, p sub-level block, d sub-level block, and f sub-level block.

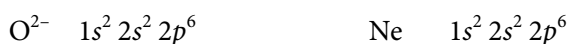
Source: [http://commons.wikimedia.org/wiki/File:Periodic\\_Table\\_structure.svg](http://commons.wikimedia.org/wiki/File:Periodic_Table_structure.svg)

## Electron Arrangements of Ions

The procedures for drawing energy-level diagrams and electronic configurations also apply to ions. As you learned in earlier courses, ions are atoms or groups of atoms that have an excess of either positive or negative charge. Atoms form ions by either gaining or losing electrons. The number of electrons gained or lost can often be predicted by the element's location on the periodic table. For example, the element oxygen tends to gain two electrons in its chemical reactions to produce the oxide ion,  $\text{O}^{2-}$ . The atom is named oxygen, while the ion is named oxide. The configurations of oxygen and oxide are:



The oxide ion has the same electronic configuration as the nearest noble gas element neon:



Oxide and neon are said to be isoelectronic, meaning that they have the same number of electrons. Experimental evidence shows that atoms or ions with filled outer orbitals are chemically stable. For example, neon is so chemically stable that it doesn't react at all under normal conditions. Similarly, lone oxide and fluoride ions are also non-reactive. However, individual oxygen and fluorine atoms are among the most reactive elements in the periodic table. Consequently, the atoms of many elements tend to gain or lose electrons so that they have the same electron arrangements as their nearest noble gas.

### Example

Draw the energy-level diagram and write the electronic configuration for magnesium and the ion that magnesium forms.

### Solution

The nearest noble gas to magnesium is the element neon. Both the energy-level diagram and the electronic configuration indicate that magnesium has two more electrons than neon.

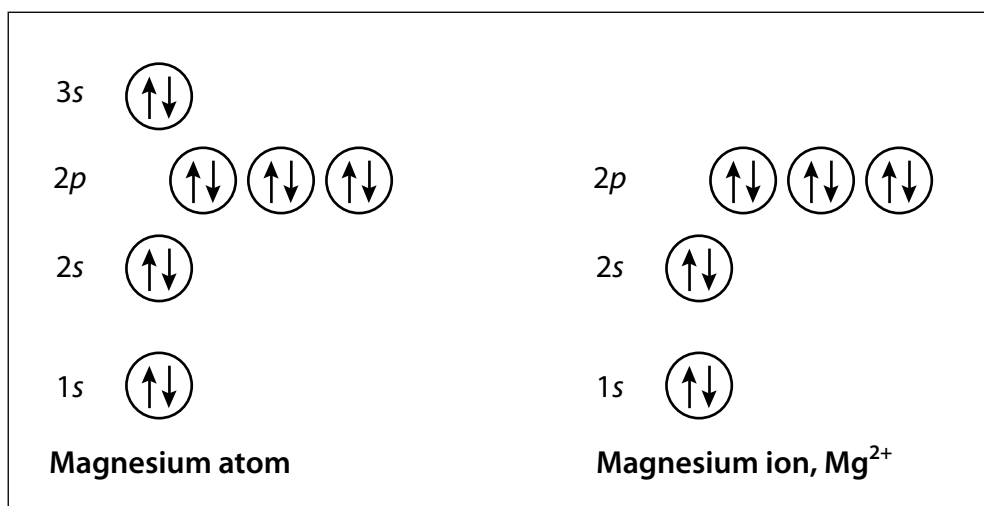


Figure 1.26: Magnesium loses its 3s electrons when it forms  $\text{Mg}^{2+}$ .

Electronic configurations:



Therefore, to achieve an electron arrangement like that of neon, magnesium must lose its two electrons in the 3s orbital. As a result, the magnesium ion has the chemical symbol  $\text{Mg}^{2+}$ .

## Support Questions

9. Write the electronic configuration for:
  - a) selenium, Se
  - b) strontium, Sr
10. Based on their electronic configurations, write the chemical symbol of the ion that the elements in Support Question 9 most commonly form. Justify your prediction.
11. Sketch the energy-level diagram for phosphorus.
12. Use your diagram in Support Question 11 to
  - a) predict why  $\text{P}^{4-}$  is not common.
  - b) explain why  $\text{P}^{3-}$  is more likely to exist than  $\text{P}^{5+}$ .



The “Careers in Chemistry” box below is the first in a series of profiles about career opportunities in chemistry. As you continue through the course, you will read about the career paths of three inspirational Canadian chemists. Some of these scientists have dedicated most of their lives to their chosen subject. As you read about their career paths and what motivated them to pursue a career in chemistry, you may get some ideas that will help you with your career choices.

## Careers in Chemistry

In the introduction to this course, you had the opportunity to view the video "Chemistry: All About You." This video was made to celebrate 2011 as the International Year of Chemistry. As you saw from the video, chemical research solves many complex global problems involving food, water, health, energy, and transportation.

The goal of the International Year of Chemistry (IYC) is to increase public awareness and appreciation of chemistry. One hundred years ago, Madame Marie Curie was awarded the Nobel Prize for her pioneering work in chemistry and 2011 marks the 100th anniversary of the founding of the International Association of Chemical Societies. If you want to learn more about IYC and the International Association of Chemical Societies, key these titles into the search engine of your browser and you will quickly be directed to the appropriate websites.

The following table highlights three careers in chemistry and the education and/or training needed to achieve these goals in Ontario. It's quite easy to do some basic research into a career in chemistry that interests you. Simply key the keywords in your browser (for example, “waste management technician educational requirements”) and insert the data you find into a table like the one that follows.

**Table 1.3: Careers in Chemistry**

<b>Profession</b>	<b>Educational requirements</b>	<b>Professional designation</b>
Chemical engineering technologist	<ul style="list-style-type: none"> <li>• Ontario Secondary School Diploma with a majority of senior credits</li> <li>• Grade 12 English</li> <li>• Grade 11 Biology (C) or (U)</li> <li>• Grade 12 Chemistry (C)</li> <li>• Grade 12 Mathematics</li> </ul>	Ontario College Advanced Diploma in Chemical Engineering technology (three years of full-time study)
Nurse	<ul style="list-style-type: none"> <li>• Ontario Secondary School Diploma with a majority of senior credits</li> <li>• Grade 12 English or French</li> <li>• Grade 12 Biology</li> <li>• Grade 12 Chemistry</li> <li>• Grade 12 Physics</li> <li>• Grade 12 Mathematics</li> </ul>	<ul style="list-style-type: none"> <li>• Practical Nurse Diploma (three years of full-time study)</li> <li>• Bachelor of Nursing (BN) or Bachelor of Nursing Science (BScN) (four years of full-time study)</li> </ul>
Pharmacist	<ul style="list-style-type: none"> <li>• Successful completion of specific required subjects at the appropriate level (Grade 12U or university)</li> <li>• English, Physics, Biology, Chemistry (Organic and Physical, Biochemistry, Calculus, Statistics, Humanities, or Social Science)</li> </ul> <p>(This normally requires a minimum of two years of study at the university level.)</p> <ul style="list-style-type: none"> <li>• PCAT (Pharmacy College Admission Test)</li> </ul>	Bachelor's Degree in Pharmacy

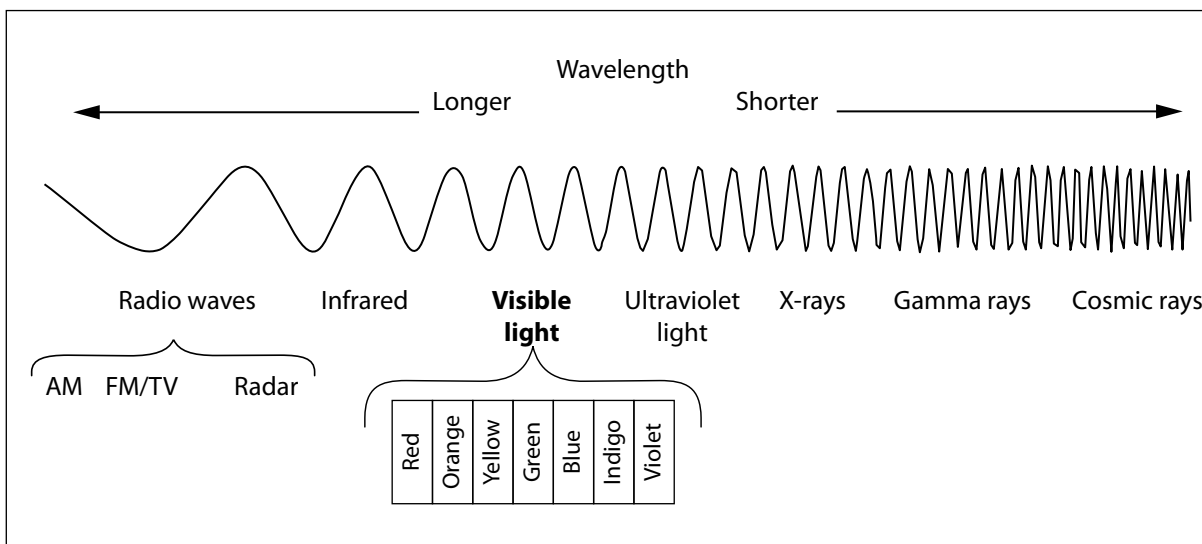
**Note:** Admission requirements change from time to time. The information above was correct when this course was written (2011–2012).

**Heads Up:** Remember, atomic mass as shown in the periodic table has been averaged based on the occurrence of isotopes.

## Key Questions

Now work on your Key Questions in the [online submission tool](#). You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

1. The development of MRI imaging technology is one useful spinoff of basic research into the structure of the atom. Research, however, is expensive. Many people argue that the high cost of research outweighs its potential benefits. Provide one argument for and one argument against increasing current funding for atomic-structure research. Use specific examples from this lesson in your answer to support each position. (6 marks)
2. The emission spectrum of an unknown element contains two lines—one in the visible portion of the spectrum, and the other, ultraviolet. Based on the following figure and on what you have learned about Niels Bohr's model of the atom, account for the difference in energy between these lines. (3 marks)



3. a) Draw the energy-level diagram and write the electronic configuration for zirconium (element 40). (3 marks)
- b) Use the diagram to explain why the  $\text{Zr}^{4+}$  ion exists. (2 marks)
- c) Compare the stability of the Zr atom and the Zr ion. Account for any differences in their stability. (2 marks)

Now go on to Lesson 2. Send your answers to the Key Questions to the ILC when you have completed Unit 1 (Lessons 1 to 4).