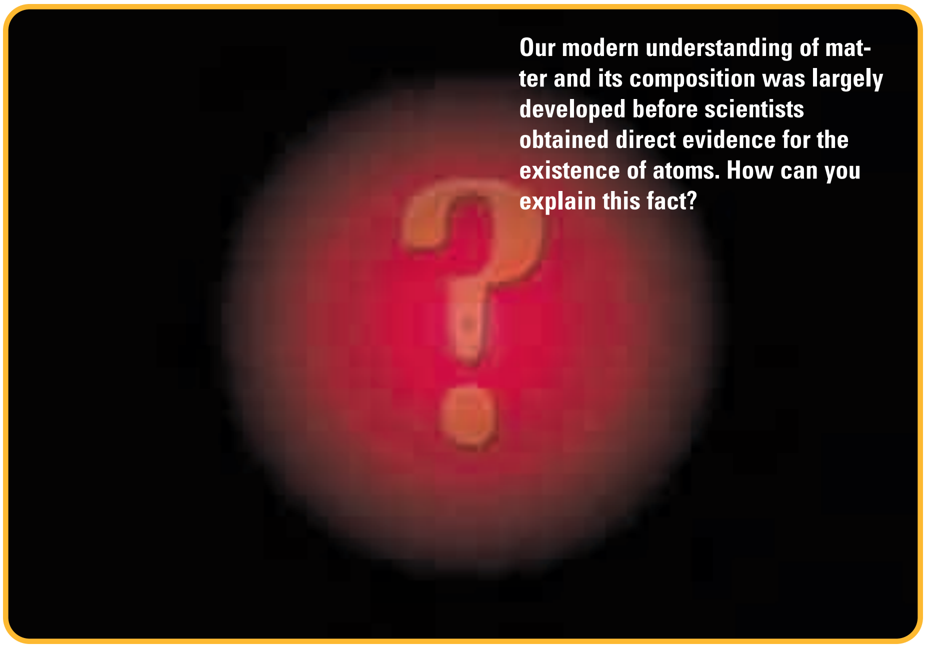
CHAPTER 2

Elements and the Periodic Table

Today, if you want to travel by air across the country or overseas, you take an airplane. During the first three decades of the twentieth century, you would have boarded a hydrogen-filled balloon such as the one shown in the black-and-white photograph. Large and small airships such as these, called dirigibles, were common sights in the skies above many North American and European cities. Unfortunately, during a landing in Lakehurst, New Jersey in 1937, the hydrogen in one of these airships, the *Hindenburg,* ignited. The resulting explosion killed 36 people, and marked the end of the use of hydrogen for dirigibles.

Gas-filled airships and balloons like the one shown in the colour photograph now use helium gas instead of hydrogen. Helium, unlike hydrogen, does not burn. In fact, helium is a highly unreactive gas. What is it about hydrogen that makes it so reactive? Why is helium so unreactive? The answer lies in the structure of the atoms of these elements. In previous science courses, you traced the history of our understanding of atoms and their structure. You also learned how chemists use properties to arrange elements, and the atoms of which they are made, into a remarkable tool called the periodic table. This chapter highlights and expands on key ideas from your earlier studies. By the end of the chapter, you will have a greater understanding of the properties of elements at an atomic level. This understanding is a crucial foundation for concepts that you will explore in your chemistry course this year.



Our modern understanding of matter and its composition was largely developed before scientists obtained direct evidence for the existence of atoms. How can you explain this fact?

Chapter Preview

Concepts and Skills You Will Need

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

- expressing the results of calculations to the appropriate number of decimal places and significant digits (Chapter 1, section 1.2)

2.1 Atoms and Their Composition

Section Preview/ Specific Expectations

In this section, you will

- **define** and **describe** the relationships among atomic number, mass number, atomic mass, isotope, and radioisotope
- **communicate** your understanding of the following terms: *atom, atomic mass unit (u), atomic number (Z), mass number (A), atomic symbol, isotopes, radioactivity, radioisotopes*

Technology LINK

How scientists visualize the atom has changed greatly since Dalton proposed his atomic theory in the early nineteenth century. Technology has played an essential role in these changes. At a library or on the Internet, research the key modifications to the model of the atom. Create a summary chart to show your findings. Include the scientists involved, the technologies they used, the discoveries they made, and the impact of their discoveries on the model of the atom. If you wish, use a suitable graphics program to set up your chart.

Elements are the basic substances that make up all matter. About 90 elements exist naturally in the universe. The two smallest and least dense of these elements are hydrogen and helium. Yet hydrogen and helium account for nearly 98% of the mass of the entire universe!

Here on Earth, there is very little hydrogen in its pure elemental form. There is even less helium. In fact, there is such a small amount of helium on Earth that it escaped scientists' notice until 1895.

Regardless of abundance, any two samples of hydrogen — from anywhere on Earth or far beyond in outer space — are identical to each other. For example, a sample of hydrogen from Earth's atmosphere is identical to a sample of hydrogen from the Sun. The same is true for helium. This is because *each element is made up of only a single kind of atom*. For example, the element hydrogen contains only hydrogen atoms. The element helium contains only helium atoms. What, however, is an atom?

The Atomic Theory of Matter

John Dalton was a British teacher and self-taught scientist. In 1809, he described atoms as solid, indestructible particles that make up all matter. (See Figure 2.1.) Dalton's concept of the atom is one of several ideas in his atomic theory of matter, which is outlined on the next page. Keep in mind that scientists have modified several of Dalton's ideas, based on later discoveries. You will learn about these modifications at the end of this section. See if you can infer what some of them are as you study the structure of the atom on the next few pages.

Figure 2.1 This illustration shows an atom as John Dalton (1766–1844) imagined it. Many reference materials refer to Dalton's concept of the atom as the "billiard ball model." Dalton, however, was an avid lawn bowler. His concept of the atom was almost certainly influenced by the smooth, solid bowling balls used in the game.

Dalton's Atomic Theory (1809)

- All matter is made up of tiny particles called atoms. An atom cannot be created, destroyed, or divided into smaller particles.
- The atoms of one element cannot be converted into the atoms of any another element.
- All the atoms of one element have the same properties, such as mass and size. These properties are different from the properties of the atoms of any other element.
- Atoms of different elements combine in specific proportions to form compounds.

CHECKPOINT

The atomic theory was a convincing explanation of the behaviour of matter. It explained two established scientific laws: the law of conservation of mass and the law of definite composition.

- *Law of conservation of mass*: During a chemical reaction, the total mass of the substances involved does not change.
- *Law of definite proportion*: Elements always combine to form compounds in fixed proportions by mass. (For example, pure water always contains the elements hydrogen and oxygen, combined in the following proportions: 11% hydrogen and 89% oxygen.)

How does the atomic theory explain these two laws?

The Modern View of the Atom

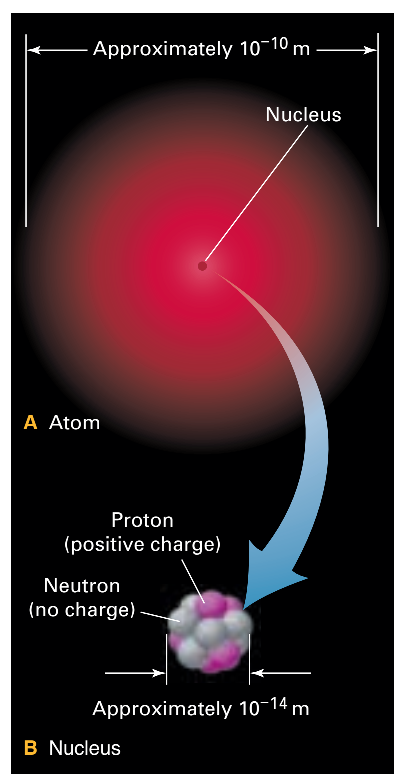
An **atom** is the smallest particle of an element that still retains the identity and properties of the element. For example, the smallest particle of the writing material in your pencil is a carbon atom. (Pencil "lead" is actually a substance called graphite. Graphite is a form of the element carbon.)

An average atom is about 10^{-10} m in diameter. Such a tiny size is difficult to visualize. If an average atom were the size of a grain of sand, a strand of your hair would be about 60 m in diameter!

Atoms themselves are made up of even smaller particles. These *subatomic particles* are protons, neutrons, and electrons. Protons and neutrons cluster together to form the central core, or *nucleus*, of an atom. Fast-moving electrons occupy the space that surrounds the nucleus of the atom. As their names imply, subatomic particles are associated with electrical charges. Table 2.1 and Figure 2.2 summarize the general features and properties of an atom and its three subatomic particles.

Table 2.1 Properties of Protons, Neutrons, and Electrons

Subatomic particle	Charge	Symbol	Mass (in g)	Radius (in m)
electron	1–	e^-	9.02×10^{-28}	smaller than 10^{-18}
proton	1+	p^+	1.67×10^{-24}	10^{-15}
neutron	0	n^0	1.67×10^{-24}	10^{-15}



Expressing the Mass of Subatomic Particles

As you can see in Table 2.1, subatomic particles are incredibly small. Suppose that you could count out protons or neutrons equal to

602 000 000 000 000 000 000 000 (or 6.02×10^{23})

and put them on a scale. They would have a mass of about 1 g. This means that one proton or neutron has a mass of

$$\frac{1\text{ g}}{6.02 \times 10^{23}} = 0.000\,000\,000\,000\,000\,000\,000\,001\,66\text{ g}$$
$$= 1.66 \times 10^{-24}\text{ g}$$

It is inconvenient to measure the mass of subatomic particles using units such as grams. Instead, chemists use a unit called an **atomic mass unit** (symbol **u**). A proton has a mass of about 1 u, which is equal to 1.66×10^{-24} g.

Figure 2.2 **This illustration shows the modern view of an atom. Notice that a fuzzy, cloud-like region surrounds the atomic nucleus. Electrons move rapidly throughout this region, which represents most of the atom's volume.**

CHEM FACT

The number 6.02×10^{23} is called the Avogadro constant. In Chapter 5, you will learn more about the Avogadro constant.

CHEM FACT

A proton is about 1837 times more massive than an electron. According to Table 2.1, the mass of an electron is 9.02×10^{-28} g. This value is so small that scientists consider the mass of an electron to be approximately equal to zero. Thus, electrons are not taken into account when calculating the mass of an atom.

The Nucleus of an Atom

All the atoms of a particular element have the same number of protons in their nucleus. For example, all hydrogen atoms—anywhere in the universe—have one proton. All helium atoms have two protons. All oxygen atoms have eight protons. Chemists use the term **atomic number** (symbol ***Z***) to refer to the number of protons in the nucleus of each atom of an element.

As you know, the nucleus of an atom also contains neutrons. In fact, the mass of an atom is due to the combined masses of its protons and neutrons. Therefore, an element's **mass number** (symbol ***A***) is the total number of protons and neutrons in the nucleus of one of its atoms. Each proton or neutron is counted as one unit of the mass number. For example, an oxygen atom, which has 8 protons and 8 neutrons in its nucleus, has a mass number of 16. A uranium atom, which has 92 protons and 146 neutrons, has a mass number of 238.

Information about an element's protons and neutrons is often summarized using the chemical notation shown in Figure 2.3. The letter X represents the **atomic symbol** for an element. (The atomic symbol is also called the *element symbol*.) Each element has a different atomic symbol. All chemists, throughout the world, use the same atomic symbols. Over the coming months, you will probably learn to recognize many of these symbols instantly. Appendix G, at the back of this book, lists the elements in alphabetical order, along with their symbols. You can also find the elements and their symbols in the periodic table on the inside back cover of this textbook, and in Appendix C. (You will review and extend your understanding of the periodic table, in section 2.2.)

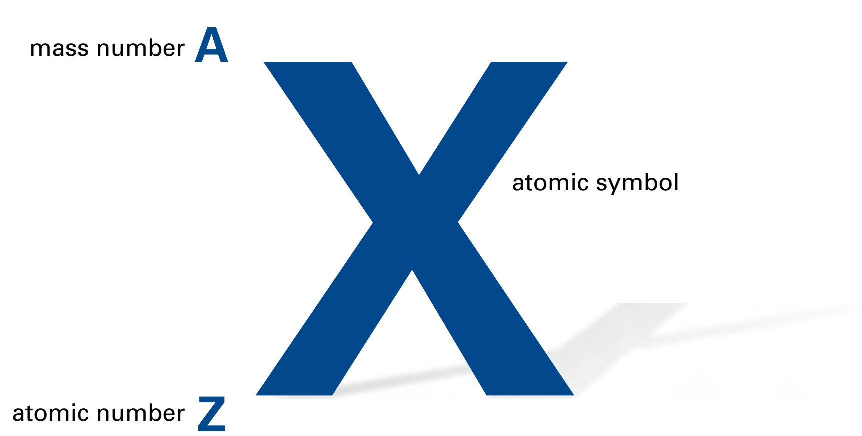


Figure 2.3

Notice what the chemical notation in Figure 2.3 does, and does not, tell you about the structure of an element's atoms. For example, consider the element fluorine: $^{19}_{9}F$. The mass number (the superscript 19) indicates that fluorine has a total of 19 protons and neutrons. The atomic number (subscript 9) indicates that fluorine has 9 protons. Neither the mass number nor the atomic number tells you how many neutrons fluorine has. You can calculate this value, however, by subtracting the atomic number from the mass number.

$$\begin{aligned}\text{Number of neutrons} &= \text{Mass number} - \text{Atomic number}\\ &= A - Z\end{aligned}$$

Thus, for fluorine,

$$\begin{aligned}\text{Number of neutrons} &= A - Z \\ &= 19 - 9 \\ &= 10\end{aligned}$$

Now try a few similar calculations in the Practice Problem below.

Practice Problems

1. Copy the table below into your notebook. Fill in the missing information. Use a periodic table, if you need help identifying the atomic symbol.

Chemical notation	Element	Number of protons	Number of neutrons
$^{11}_{5}B$	(a)	(b)	(c)
$^{208}_{82}Pb$	(d)	(e)	(f)
(g)	tungsten	(h)	110
(i)	helium	(j)	2
$^{239}_{94}Pu$	(k)	(l)	(m)
$^{56}_{26}$(n)	(o)	26	(p)
(q)	bismuth	(r)	126
(s)	(t)	47	60
$^{20}_{10}$(u)	(v)	(w)	(x)

Math LINK

Expressing numerical data about atoms in units such as metres is like using a bulldozer to move a grain of sand. Atoms and subatomic particles are so small that they are not measured using familiar units. Instead, chemists often measure atoms in nanometres ($1\ \text{nm} = 1 \times 10^{-9}\ \text{m}$) and picometres ($1\ \text{pm} = 1 \times 10^{-12}\ \text{m}$).

- Convert the diameter of a proton and a neutron into nanometres and then picometres.
- Atomic and subatomic sizes are hard to imagine. Create an analogy to help people visualize the size of an atom and its subatomic particles. (The first sentence of this feature is an example of an analogy.)

Using the Atomic Number to Infer the Number of Electrons

As just mentioned, the atomic number and mass number do not give you direct information about the number of neutrons in an element. They do not give you the number of electrons, either. You can infer the number of electrons, however, from the atomic number. The atoms of each element are electrically neutral. This means that their positive charges (protons) and negative charges (electrons) must balance one another. In other words, *in the neutral atom of any element, the number of protons is equal to the number of electrons*. For example, a neutral hydrogen atom contains one proton, so it must also contain one electron. A neutral oxygen atom contains eight protons, so it must contain eight electrons.

Isotopes and Atomic Mass

All neutral atoms of the same element contain the same number of protons and, therefore, the same number of electrons. The number of neutrons can vary, however. For example, most of the oxygen atoms in nature have eight neutrons in their atomic nuclei. In other words, most oxygen atoms have a mass number of 16 (8 protons + 8 neutrons). As you can see in Figure 2.4 on the next page, there are also two other naturally occurring forms of oxygen. One of these has nine neutrons, so $A = 17$. The other has ten neutrons, so $A = 18$. These three forms of oxygen are called isotopes. **Isotopes** are atoms of an element that have the same number of protons but different numbers of neutrons.

Web LINK

www.school.mcgrawhill.ca/resources/

The atomic symbols are linked to the names of the elements. The links are not always obvious, however. Many atomic symbols are derived from the names of the elements in a language other than English, such as Latin, Greek, German, or Arabic. With your classmates, research the origin and significance of the name of each element. Go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. See if you can infer the rules that are used to create the atomic symbols from the names of the elements.

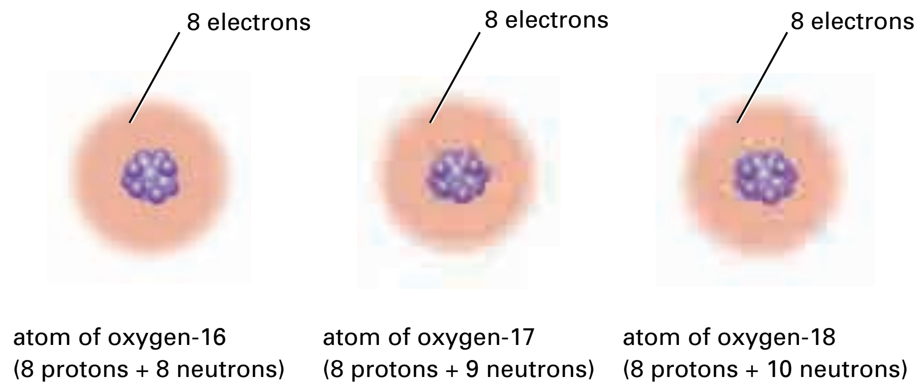


Figure 2.4 **Oxygen has three naturally occurring isotopes. Notice that oxygen-16 has the same meaning as $^{16}_{8}O$. Similarly, oxygen-17 has the same meaning as $^{17}_{8}O$ and oxygen-18 has the same meaning as $^{18}_{8}O$.**

CHEM FACT

Radioisotopes decay because their nuclei are unstable. The time it takes for nuclei to decay varies greatly. For example, it takes billions of years for only half of the nucleus of naturally occurring uranium-238 to decay. The nuclei of other radioisotopes — mainly those that scientists have synthesized — decay much more rapidly. The nuclei of some isotopes, such as sodium-22, take about 20 years to decay. For calcium-47, this decay occurs in a matter of days. The nuclei of most synthetic radioisotopes decay so quickly, however, that the radioisotopes exist for mere fractions of a second.

CHECKPOINT

When chemists refer symbolically to oxgyen-16 atoms, they often leave out the atomic number. They write ^{16}O. You can write other isotopes of oxygen, and all other elements, the same way. Why is it acceptable to leave out the atomic number?

The isotopes of an element have very similar chemical properties because they have the same number of protons and electrons. They differ in mass, however, because they have different numbers of neutrons.

Some isotopes are more unstable than others. Their *nuclei* (plural of *nucleus*) are more likely to decay, releasing energy and subatomic particles. This process, called **radioactivity**, happens spontaneously. All uranium isotopes, for example, have unstable nuclei. They are called radioactive isotopes, or **radioisotopes** for short. Many isotopes are not radioisotopes. Oxygen's three naturally occurring isotopes, for example, are stable. In contrast, chemists have successfully synthesized ten other isotopes of oxygen, all of which are unstable radioisotopes. (What products result when radioisotopes decay? You will find out in Chapter 4.)

Electrons in Atoms

So far, much of the discussion about the atom has concentrated on the nucleus and its protons and neutrons. What about electrons? What is their importance to the atom? Recall that electrons occupy the space surrounding the nucleus. Therefore, they are the first subatomic particles that are likely to interact when atoms come near one another. In a way, electrons are on the "front lines" of atomic interactions. The number and arrangement of the electrons in an atom determine how the atom will react, if at all, with other atoms. As you will learn in section 2.2, and throughout the rest of this unit, electrons are responsible for the chemical properties of the elements.

Revisiting the Atomic Theory

John Dalton did not know about subatomic particles when he developed his atomic theory. Even so, the modern atomic theory (shown on the next page) retains many of Dalton's ideas, with only a few modifications. Examine the comments to the right of each point. They explain how the modern theory differs from Dalton's.

The atomic theory is a landmark achievement in the history of chemistry. It has shaped the way that all scientists, especially chemists, think about matter. In the next section, you will investigate another landmark achievement in chemistry: the periodic table.

The Modern Atomic Theory

- All matter is made up of tiny particles called atoms. Each atom is made up of smaller subatomic particles: protons, neutrons, and electrons.

  Although an atom is divisible, it is still the smallest particle of an element that has the properties and identity of the element.

- The atoms of one element cannot be converted into the atoms of any another element by a chemical reaction.

  Nuclear reactions (changes that alter the composition of the atomic nucleus) may, in fact, convert atoms of one element into atoms of another.

- Atoms of one element have the same properties, such as average mass and size. These properties are different from the properties of the atoms of any other element.

  Different isotopes of an element have different numbers of neutrons and thus different masses. As you will learn in Chapter 5, scientists treat elements as if their atoms have an average mass.

- Atoms of different elements combine in specific proportions to form compounds.

  This idea has remained basically unchanged.

Section Review

1 C Copy the table below into your notebook. Use a graphic organizer to show the relationship among the titles of each column. Then fill in the blanks with the appropriate information. (Assume that the atoms of each element are neutral.)

Element	Atomic number	Mass number	Number of protons	Number of electrons	Number of neutrons
(a)	(b)	108	(c)	47	(d)
(e)	(f)	(g)	33	(h)	42
(i)	35	(j)	(k)	(l)	45
(m)	79	179	(n)	(o)	(p)
(q)	(r)	(s)	(t)	50	69

2 K/U Explain the difference between a stable isotope and a radioisotope. Provide an example other than oxygen to support your answer.

3 K/U Examine the information represented by the following pairs: $^{3}_{1}H$ and $^{3}_{2}He$; $^{14}_{6}C$ and $^{16}_{7}N$; $^{19}_{9}F$ and $^{18}_{9}F$.

(a) For each pair, do both members have the same number of protons? electrons? neutrons?

(b) Which pair or pairs consist of atoms that have the same value for Z? Which consists of atoms that have the same value for A?

4 C Compare Dalton's atomic theory with the modern atomic theory. Explain why scientists modified Dalton's theory.

5 C In your opinion, should chemistry students learn about Dalton's theory if scientists no longer agree with it completely? Justify your answer.

CHEM FACT

Not all chemists believed that Dalton's atoms existed. In 1877, one skeptical scientist called Dalton's atoms "stupid hallucinations." Other scientists considered atoms to be a valuable *idea* for understanding matter and its behaviour. They did not, however, believe that atoms had any physical reality. The discovery of electrons (and, later, the other subatomic particles) finally convinced scientists that atoms are more than simply an idea. Atoms, they realized, must be matter.

2.2 Atoms, Elements, and the Periodic Table

Section Preview/ Specific Expectations

In this section, you will

- **state**, in your own words, the periodic law
- **describe** elements in the periodic table in terms of energy levels and the electron arrangements
- **use** Lewis structures to represent valence electrons
- **communicate** your understanding of the following terms: *energy levels, periodic trends, valence electrons, Lewis structures, stable octet, octet*

Language LINK

The term *periodic* means "repeating in an identifiable pattern." For example, a calendar is periodic. It organizes the days of the months into a repeating series of weeks. What other examples of periodicity can you think of?

By the mid 1800's, there were 65 known elements. Chemists studied these elements intensively and recorded detailed information about their reactivity and the masses of their atoms. Some chemists began to recognize patterns in the properties and behaviour of many of these elements. (See Figure 2.5.)

Other sets of elements display similar trends in their properties and behaviour. For example, oxygen (O), sulfur (S), selenium (Se), and tellurium (Te) share similar properties. The same is true of fluorine (F), chlorine (Cl), bromine (Br), and iodine (I). These similarities prompted chemists to search for a fundamental property that could be used to organize all the elements. One chemist, Dmitri Mendeleev (1834–1907), sequenced the known elements in order of increasing atomic mass. The result was a table of the elements, organized so that elements with similar properties were arranged in the same column. Because Mendeleev's arrangement highlighted periodic (repeating) patterns of properties, it was called a *periodic table*.

The modern periodic table is a modification of the arrangement first proposed by Mendeleev. Instead of organizing elements according to atomic mass, the modern periodic table organizes elements according to atomic number. According to the **periodic law**, *the chemical and physical properties of the elements repeat in a regular, periodic pattern when they are arranged according to their atomic number.*

Figures 2.6 and 2.7 outline the key features of the modern periodic table. Take some time to review these features. Another version of the periodic table, containing additional data, appears on the inside back cover of this textbook, as well as in Appendix C.

Figure 2.5 These five elements share many physical and chemical properties. However, they have widely differing atomic masses.

lithium, Li — Sodium, Na — Potassium, K — Rubidium, Rb — Cesium, Cs

Shared Physical Properties

- soft
- metallic (therefore malleable, ductile, and good conductors of electricity)

Shared Chemical Properties

- are very reactive
- react vigorously (and explosively) with water
- combine with chlorine to form a white solid that dissolves easily in water

MAIN GROUP ELEMENTS

TRANSITION ELEMENTS

MAIN-GROUP ELEMENTS

- metals (main group)
- metals (transition)
- metals (inner transition)
- metalloids
- nonmetals

	1 (IA)	2 (IIA)	3 (IIIB)	4 (IVB)	5 (VB)	6 (VIB)	7 (VIIB)	8	9 (VIIIB)	10	11 (IB)	12 (IIB)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
1	1 H 1.01																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.13											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Uun (269)	111 Uuu (272)	112 Uub (277)		114 Uuq (285)		116 Uuh (289)		118 Uuo (293)

INNER TRANSITION ELEMENTS

6	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
7	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

- Each element is in a separate box, with its atomic number, atomic symbol, and atomic mass. (Different versions of the periodic table provide additional data and details.)
- Elements are arranged in seven numbered periods (horizontal rows) and 18 numbered groups (vertical columns).
- Groups are numbered according to two different systems. The current system numbers the groups from 1 to 18. An older system numbers the groups from I to VIII, and separates them into two categories labelled A and B. Both of these systems are included in this textbook.
- The elements in the eight A groups are the main-group elements. They are also called the representative elements.

The elements in the ten B groups are known as the transition elements. (In older periodic tables, Roman numerals are used to number the A and B groups.)

- Within the B group transition elements are two horizontal series of elements called inner transition elements. They usually appear below the main periodic table. Notice, however, that they fit between the elements in Group 3 (IIIB) and Group 4 (IVB).
- A bold "staircase" line runs from the top of Group 13 (IIIA) to the bottom of Group 16 (VIA). This line separates the elements into three broad classes: metals, metalloids (or semi-metals), and non-metals. (See Figure 2.7 on the next page for more information.)
- Group 1 (IA) elements are known as *alkali metals*. They react with water to form alkaline, or basic, solutions.
- Group 2 (IIA) elements are known as *alkaline earth metals*. They react with oxygen to form compounds called oxides, which react with water to form alkaline solutions. Early chemists called all metal oxides "earths."
- Group 17 (VIIA) elements are known as *halogens*, from the Greek word *hals*, meaning "salt." Elements in this group combine with other elements to form compounds called salts.
- Group 18 (VIIIA) elements are known as *noble gases*. Noble gases do not combine in nature with any other elements.

Figure 2.6 The basic features of the periodic table are summarized here. Most of your work in this course will focus on the representative elements.

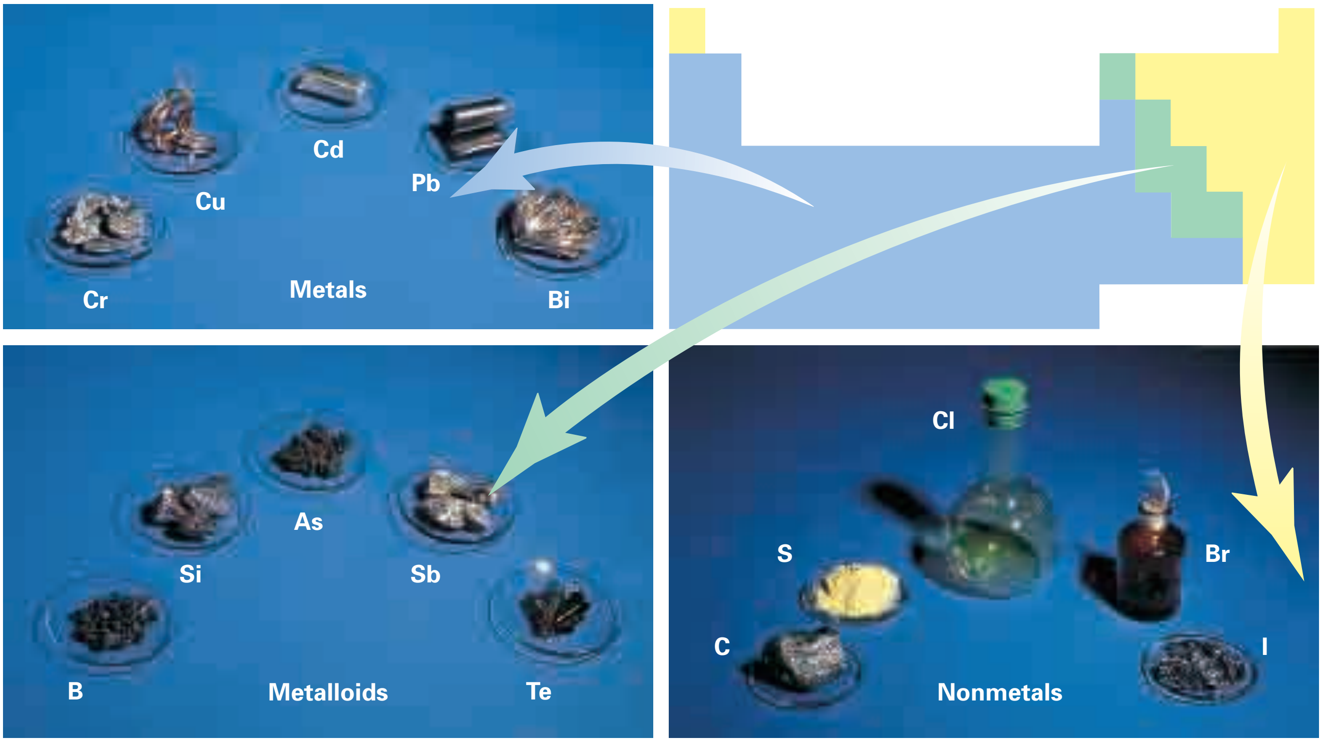


Figure 2.7 Several examples from each of the three main classes of elements are shown here. Find where they appear in the periodic table in Figure 2.6.

Practice Problems

2. Identify the name and symbol of the elements in the following locations of the periodic table:

(a) Group 14 (IVA), Period 2

(b) Group 11 (IB), Period 4

(c) Group 18 (VIIIA), Period 6

(d) Group 1 (IA), Period 1

(e) Group 12 (IIB), Period 5

(f) Group 2 (IIA), Period 4

(g) Group 17 (VIIA), Period 5

(h) Group 13 (IIIA), Period 3

History LINK

Mendeleev did not develop his periodic table in isolation. He built upon work that had been done by other chemists, in other parts of the world, over several decades. Research other ideas that were proposed for organizing the elements. Include Mendeleev's work in your research. What was it about his arrangement that convinced chemists to adopt it?

Electrons and the Periodic Table

You have seen how the periodic table organizes elements so that those with similar properties are in the same group. You have also seen how the periodic table distinguishes among metals, non-metals, and metalloids. Other details of the organization of the periodic table may seem baffling, however. Why, for example, are there different numbers of elements in the periods?

The reason for this, and other details of the periodic table's organization, involves the number and arrangement of electrons in the atoms of each element. To appreciate the importance of electrons to the periodic table, it is necessary to revisit the structure of the atom.

In the following ExpressLab, you will observe elements in much the same way that scientists did in the early twentieth century. In doing so, these scientists set the stage for a new understanding of matter and the electrical structure of its atoms.

ExpressLab Observing the Spectra of Elements

In this activity, you will use a device called a *diffraction grating*. It separates light into banded patterns of colour (a spectrum). Different colours of light have different frequencies and wavelengths, so they have different amounts of energy. Red light is less energetic, for example, than blue light.

Safety Precautions

- Gas discharge tubes operate at a voltage that is high enough to cause serious injury. Observe them only from a safe distance, as determined by your teacher.

Materials

diffraction grating
incandescent light source
gas discharge tubes containing different elements

Procedure

1. Use the diffraction grating to observe the light that is emitted from an ordinary incandescent light bulb. Make a quick sketch to record your observations.
2. Observe the light that is emitted from the hydrogen gas discharge tube. **CAUTION** You should be about 1 m from the discharge tube. Come no farther than your teacher directs. Sketch your observations.
3. Observe the light that is emitted from the discharge tubes of other elements. Sketch your observations for each element.

Analysis

1. If the electrons in a discharge tube are moving everywhere in the space around the nucleus, their spectrum should look like the spectrum of an ordinary light bulb. What does hydrogen's spectrum look like? How do the spectra of the other elements compare with the spectrum of a light bulb and the spectrum of hydrogen?
2. Hydrogen has only one electron. Why, then, does its spectrum have four coloured lines?
3. Why is the light that is emitted by hydrogen different from the light that is emitted by the other elements? Explain the difference in terms of electrons.

Application

4. What do gas discharge tubes have in common with street lights? Do research to find out which gases are used in street lamps, and why certain gases are chosen for certain locations.

Electrons and Energy Levels

Electrons cannot move haphazardly. Their movement around an atomic nucleus is restricted to fixed regions of space. These regions are three-dimensional, similar to the layers of an onion.

Figure 2.8 shows a representation of these regions. Keep in mind that they are *not* solid. They are volumes of space in which electrons may be found. You may have heard these regions called *energy shells* or *shells*. In this textbook, they are called **energy levels**. An electron that is moving in a lower energy level is close to the nucleus. It has less energy than it would if it were moving in a higher energy level.

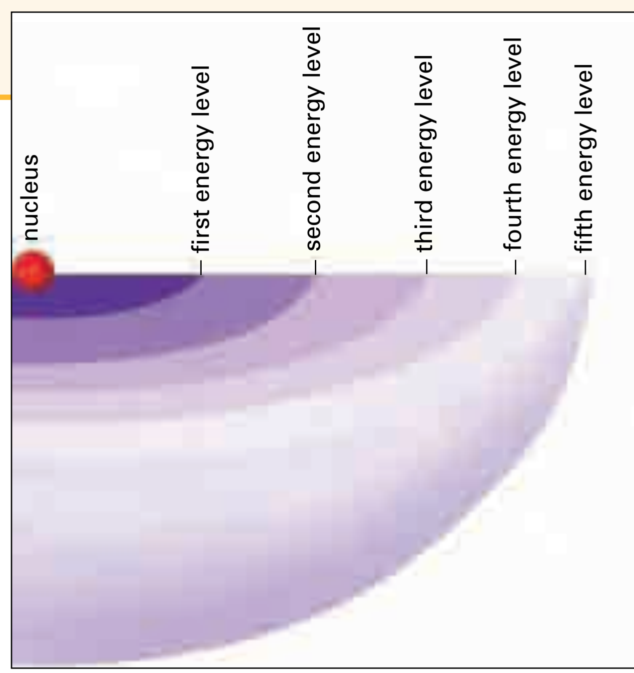


Figure 2.8 Energy levels of an atom from the fifth period

mind STRETCH

Examine the following illustration. Then answer these questions.

- Which book possesses more potential energy? Why?
- Can a book sit between shelves instead of on a shelf as shown?
- How does the potential energy of a book on a higher shelf change if it is moved to a lower shelf?
- How do you think this situation is related to electrons and the potential energy they possess when they move in different energy levels?

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has an interactive activity to help you assess your understanding of the relationship among elements, their atomic number, and their position in the periodic table.

There is a limit to the number of electrons that can occupy each energy level. For example, a maximum of two electrons can occupy the first energy level. A maximum of eight electrons can occupy the second energy level. The **periodic trends** (repeating patterns) that result from organizing the elements by their atomic number are linked to the way in which electrons occupy and fill energy levels. (See Figure 2.9.)

As shown in Figure 2.9A, a common way to show the arrangement of electrons in an atom is to draw circles around the atomic symbol. Each circle represents an energy level. Dots represent electrons that occupy each energy level. This kind of diagram is called a Bohr-Rutherford diagram. It is named after two scientists who contributed their insights to the atomic theory.

Figure 2.9B shows that the first energy level is full when two electrons occupy it. Only two elements have two or fewer electrons: hydrogen and helium. Hydrogen has one electron, and helium has two. These elements, with their electrons in the first energy level, make up Period 1 of the periodic table.

As you can see in Figure 2.9C, Period 2 elements have two occupied energy levels. The second energy level is full when eight electrons occupy it. Neon, with a total of ten electrons, has its first and second energy levels filled. Notice how the second energy level fills with electrons as you move across the period from lithium to fluorine.

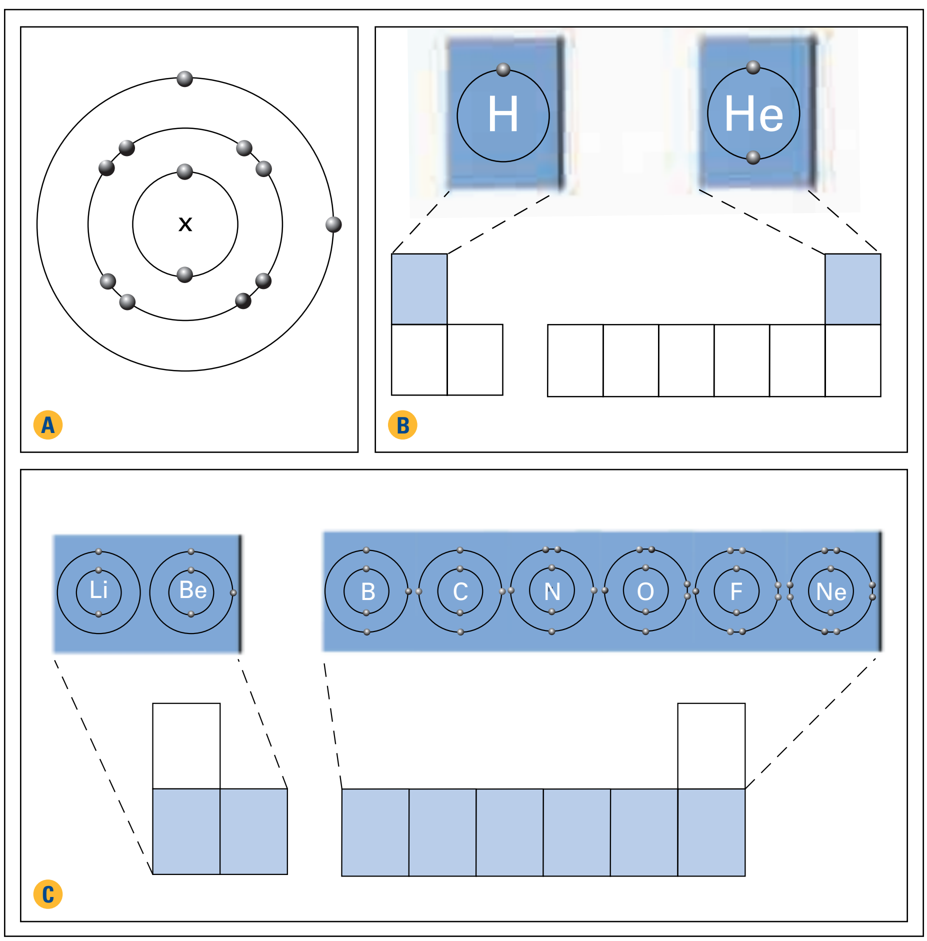


Figure 2.9 (A) A Bohr-Rutherford diagram (B) Hydrogen and helium have a single energy level. (C) The eight Period 2 elements have two energy levels.

Patterns Based on Energy Levels and Electron Arrangements

The structure of the periodic table is closely related to energy levels and the arrangement of electrons. Two important patterns result from this relationship. One involves periods, and the other involves groups.

The Period-Related Pattern

As you can see in Figure 2.9, elements in Period 1 have electrons in one energy level. Elements in Period 2 have electrons in two energy levels. This pattern applies to all seven periods. *An element's period number is the same as the number of energy levels that the electrons of its atoms occupy.* Thus, you could predict that Period 5 elements have electrons that occupy five energy levels. This is, in fact, true.

What about the inner transition elements—the elements that are below the periodic table? Figure 2.10 shows how this pattern applies to them. Elements 58 through 71 belong in Period 6, so their electrons occupy six energy levels. Elements 90 through 103 belong in Period 7, so their electrons occupy seven energy levels. Chemists and chemical technologists tend to use only a few of the inner transition elements (notably uranium and plutonium) on a regular basis. Thus, it is more convenient to place all the inner transition elements below the periodic table.

CHEM FACT

Energy levels and the arrangement of electrons involve ideas from theoretical physics. These ideas are beyond the scope of this course. Appendix D at the back of this book provides a brief introduction to these ideas. If you pursue your studies in chemistry next year and beyond, you will learn a more complete theory of electron arrangement.

	1 IA	2 IIA															3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H																															2 He
2	3 Li	4 Be																									5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg																									13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca															21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr															39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 BK	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub		114 Uuq		116 Unh		118 Uuo

Figure 2.10 The "long form" of the periodic table includes the inner transition metals in their proper place.

The Group-Related Pattern

The second pattern emerges when you consider the electron arrangements in the main-group elements: the elements in Groups 1 (IA), 2 (IIA), and 13 (IIIA) to 18 (VIIIA). All the elements in each main group have the same number of electrons in their highest (outer) energy level. The electrons that occupy the outer energy level are called **valence electrons**. The term "valence" comes from a Latin word that means "to be strong." "Valence electrons" is a suitable name because the outer energy level electrons are the electrons involved when atoms form compounds. In other words, valence electrons are responsible for the chemical behaviour of elements.

You can infer the number of valence electrons in any main-group element from its group number. For example, Group 1 (IA) elements have one valence electron. Group 2 (IIA) elements have two valence electrons. For elements in Groups 13 (IIIA) to 18 (VIIIA), the number of valence electrons is the same as the second digit in the current numbering system. It is the same as the only digit in the older numbering system. For example, elements in Group 15 (VA) have 5 valence electrons. The elements in Group 17 (VIIA) have 7 valence electrons.

Using Lewis Structures to Represent Valence Electrons

It is time-consuming to draw electron arrangements using Bohr-Rutherford diagrams. It is much simpler to use **Lewis structures** to represent elements and the valence electrons of their atoms. To draw a Lewis structure, you replace the nucleus and inner energy levels of an atom with its atomic symbol. Then you place dots around the atomic symbol to represent the valence electrons. The order in which you place the first four dots is up to you. You may find it simplest to start at the top and proceed clockwise: right, then bottom, then left.

Examine Figure 2.11, and then complete the Practice Problems that follow. In Chapter 3, you will use Lewis structures to help you visualize what happens when atoms combine to form compounds.

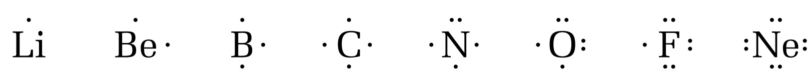


Figure 2.11 **Examine these Lewis structures for the Period 2 elements. Place a dot on each side of the element—one dot for each valence electron. Then start pairing dots when you reach five or more valence electrons.**

Practice Problems

3. Draw boxes to represent the first 20 elements in the periodic table. Using Figure 2.9 as a guide, sketch the electron arrangements for these elements.

4. Redraw the 20 elements from Practice Problem 2 using Lewis structures.

5. Identify the number of valence electrons in the outer energy levels of the following elements:

(a) chlorine
(b) helium
(c) indium
(d) strontium
(e) rubidium
(f) lead
(g) antimony
(h) selenium
(i) arsenic
(j) xenon

6. Use the periodic table to draw Lewis structures for the following elements: barium (Ba), gallium (Ga), tin (Sn), bismuth (Bi), iodine (I), cesium (Cs), krypton (Kr), xenon (Xe).

The Significance of a Full Outer Energy Level

The noble gases in Group 18 (VIIIA) are the only elements that exist as individual atoms in nature. They are extremely *un*reactive. They do not naturally form compounds with other atoms. (Scientists *have* manipulated several of these elements in the laboratory to make them react, however.) What is it about the noble gases that explains this behaviour?

Recall that chemical reactivity is determined by valence electrons. Thus, there must be something about the arrangement of the electrons in the noble gases that explains their *un*reactivity. All the noble gases have outer energy levels that are completely filled with the maximum number of electrons. Helium has a full outer energy level of two valence electrons. The other noble gases have eight valence electrons in the outer energy level. Chemists reason that having a full outer energy level must be a very stable electron arrangement.

What does this stability mean? It means that a full outer energy level is unlikely to change. Scientists have observed that, in nature, situations or systems of lower energy are favoured over situations or systems of higher energy. For example, a book on a high shelf has more potential energy (is less stable) than a book on a lower shelf. If you move a book from a high shelf to a lower shelf, it has less potential energy (is more stable). If you move a book to the floor, it has low potential energy (is much more stable).

When atoms have eight electrons in the outer energy level (or two electrons for hydrogen and helium), chemists say that they have a **stable octet**. Often this term is shortened to just **octet**. An octet is a very stable electron arrangement. As you will see in Chapter 3, an octet is often the result of changes in which atoms combine to form compounds.

Section Wrap-up

You have seen that the structure of the periodic table is directly related to energy levels and arrangements of electrons. The patterns that emerge from this relationship enable you to predict the number of valence electrons for any main group element. They also enable you to predict the number of energy levels that an element's electrons occupy. The relationship between electrons and the position of elements in the periodic table leads to other patterns, as well. You will examine several of these patterns in the next section.

Section Review

1. K/U State the periodic law, and provide at least two examples to illustrate its meaning.
2. K/U Identify the group number for each of these sets of elements. Then choose two of these groups and write the symbols for the elements within it.
   - alkali metals
   - noble gases
   - halogens
   - alkaline earth metals

3 (a) K/U Identify the element that is described by the following information. Refer to a periodic table as necessary.
- It is a Group 14 (IVA) metalloid in the third period.
- It is a Group 15 (VA) metalloid in the fifth period.
- It is the other metalloid in Group 15 (VA).
- It is a halogen that exists in the liquid state at room temperature.

(b) C Develop four more element descriptions like those in part (a). Exchange them with a classmate and identify each other's elements.

4 K/U What is the relationship between electron arrangement and the organization of elements in the periodic table?

5 C In writing, sketches, or both, explain to someone who has never seen the periodic table how it can be used to tell at a glance the number of valence electrons in the atoms of an element.

6 (a) K/U How many valence electrons are there in an atom of each of these elements?

neon	sodium	magnesium
bromine	chlorine	silicon
sulfur	helium	
strontium	tin	

(b) Present your answers from part (a) in the form of Lewis structures.
(c) Without consulting a periodic table, classify each element from part (a) as a metal, non-metal, or metalloid.

7 K/U How many elements are liquids at room temperature? Name them.

8 K/U Compare and contrast the noble gases with the other elements.

9 I An early attempt to organize the elements placed them in groups of three called triads. Examine the three triads shown below.

Triad 1	Triad 2	Triad 3
Mn	Li	S
Cr	Na	Se
Fe	K	Te

(a) Infer the reasoning for grouping the elements in this way.
(b) Which of the elements in these three triads still appear together in the same group of the modern periodic table?

10 MC Using print or electronic resources, or both, find at least one common technological application for each of the following elements:
(a) europium
(b) neodymium
(c) carbon
(d) nitrogen
(e) silicon
(f) mercury
(g) ytterbium
(h) bromine
(i) chromium
(j) krypton

11 (a) C Draw Lewis structures for each of these elements: lithium, sodium, potassium, magnesium, aluminum, carbon.
(b) Which of these elements have the same number of occupied energy levels?
(c) Which have the same number of valence electrons?

Periodic Trends Involving the Sizes and Energy Levels of Atoms

2.3

In section 2.1, you learned that the size of a typical atom is about 10^{-10} m. You know, however, that the atoms of each element are distinctly different. For example, the atoms of different elements have different numbers of protons. This means, of course, that they also have different numbers of electrons. You might predict that the size of an atom is related to the number of protons and electrons it has. Is there evidence to support this prediction? If so, is there a pattern that can help you predict the relative size of an atom for any element in the periodic table?

In Investigation 2-A, you will look for a pattern involving the size of atoms. Chemists define, and measure, an atom's size in terms of its radius. The radius of an atom is the distance from its nucleus to the *approximate* outer boundary of the cloud-like region of its electrons. This boundary is approximate because atoms are not solid spheres. They do not have a fixed outer boundary.

Figure 2.12 represents how the radius of an atom extends from its nucleus to the approximate outer boundary of its electron cloud. Notice that the radius line in this diagram is just inside the outer boundary of the electron cloud. An electron may also spend time beyond the end of the radius line.

Section Preview/ Specific Expectations

In this section, you will

- **use** your understanding of electron arrangement and forces in atoms to explain the following periodic trends: atomic radius, ionization energy, electron affinity
- **analyze** data involving atomic radius, ionization energy, and electron affinity to identify and describe general periodic trends
- **communicate** your understanding of the following terms: *ion, anion, cation ionization energy, electron affinity*

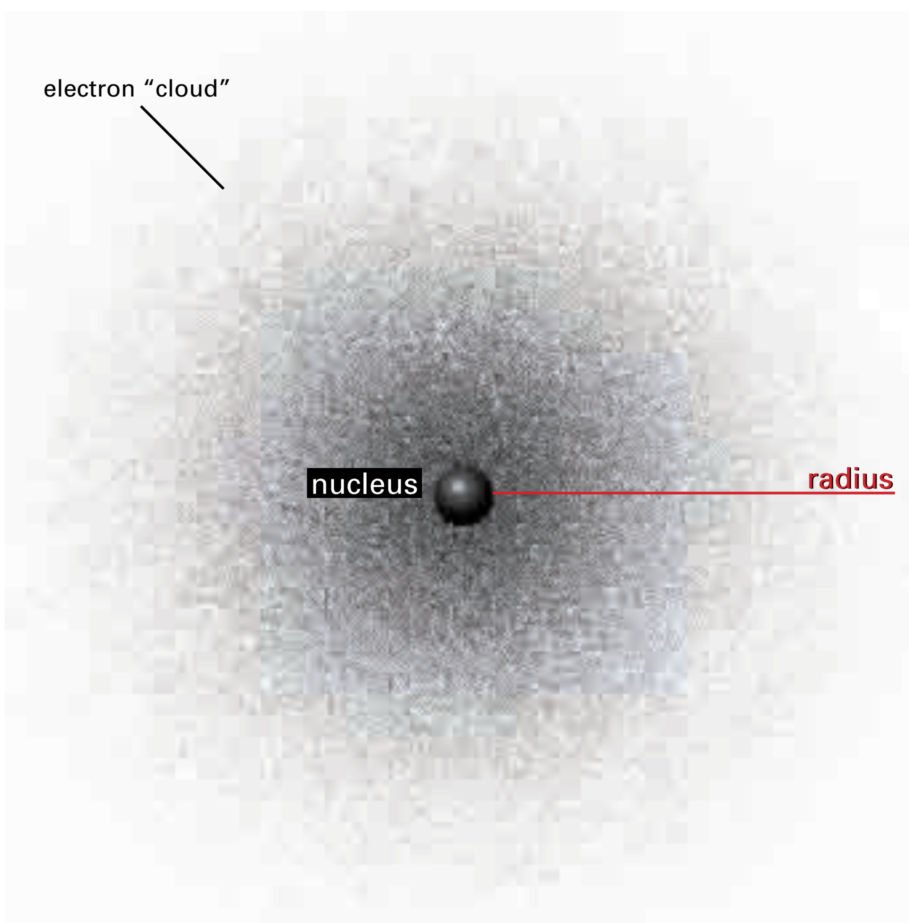


Figure 2.12 A representation of the radius of an atom

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Investigation 2-A

Analyzing Atomic Radius Data

Examine the main-group elements in the periodic table. Imagine how their size might change as you move down a group or across a period. What knowledge and reasoning can you use to infer the sizes of the atoms?

Question

How do the sizes of main-group atoms compare within a group and across a period?

Prediction

Predict a trend (pattern) that describes how the sizes of main-group atoms change down a group and across a period. Include a brief explanation to justify your prediction.

Safety Precautions

Be careful when handling any sharp instruments or materials that you choose to use.

Materials

to be decided in class

Procedure

1. The table below lists the atomic *radii* (plural of *radius*) for the main-group elements. Design different scale models that could help you visualize and compare the sizes of the atoms. Your models can be two-dimensional or three-dimensional, large or small.
2. Discuss your designs as a group. Choose a design that you think will best show the information you require.
3. Build your models. Arrange them according to their positions in the periodic table.

Atomic Radii of Main-Group Elements

Name of element	Atomic radius in picometres (pm)	Name of element	Atomic radius in picometres (pm)	Name of element	Atomic radius in picometres (pm)
aluminum	143	gallium	141	polonium	167
antimony	159	germanium	137	potassium	235
argon	88	helium	49	radon	134
astatine	145	hydrogen*	79	rubidium	248
barium	222	indium	166	selenium	140
beryllium	112	iodine	132	silicon	132
bismuth	170	krypton	103	sodium	190
boron	98	lead	175	strontium	215
bromine	112	lithium	155	sulfur	127
calcium	197	magnesium	160	tellurium	142
carbon	91	neon	51	thallium	171
cesium	267	nitrogen	92	tin	162
chlorine	97	oxygen	65	xenon	124
fluorine	57	phosphorus	128		

*Quantum mechanical value for a free hydrogen atom

Analysis

1. How do atomic radii change as you look from top to bottom within a group?
2. How do atomic radii change as you look from left to right across a period?
3. Compare your observations with your prediction. Explain why your results did, or did not, agree with your prediction.

Conclusion

4. State whether or not atomic radius is a periodic property of atoms. Give evidence to support your answer.

Application

5. Would you expect atoms of the transition elements to follow the same trend you observed for the main-group elements? Locate atomic radius data for the transition elements (not including the inner transition elements). Make additional models, or draw line or bar graphs, to verify your expectations.

Trends for Atomic Size (Radius)

There are two general trends for atomic size:

- *As you go down each group in the periodic table, the size of an atom increases*. This makes sense if you consider energy levels. As you go down a group, the valence electrons occupy an energy level that is farther and farther from the nucleus. Thus, the valence electrons experience less attraction for the nucleus. In addition, electrons in the inner energy levels block, or *shield*, the valence electrons from the attraction of the nucleus. As a result, the total volume of the atom, and thus the size, increases with each additional energy level.
- *As you go across a period, the size of an atom decreases*. This trend might surprise you at first, since the number of electrons increases as you go across a period. You might think that more electrons would occupy more space, making the atom larger. You might also think that repulsion from their like charges would force the electrons farther apart. The size of an atom decreases, however, because the positive charge on the nucleus also increases across a period. As well, without additional energy the electrons are restricted to their outer energy level. For example, the outer energy level for Period 2 elements is the second energy level. Electrons cannot move beyond this energy level. As a result, the positive force exerted by the nucleus pulls the outer electrons closer, reducing the atom's total size.

Figure 2.13 summarizes the trends for atomic size. The Practice Problems that follow give you a chance to apply your understanding of these trends.

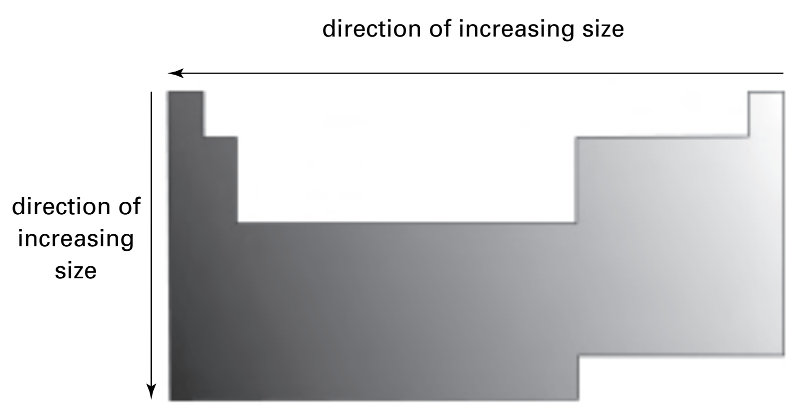


Figure 2.13 **Atomic size increases down a group and decreases across a period in the periodic table.**

Practice Problems

7. Using only their location in the periodic table, rank the atoms in each set by decreasing atomic size. Explain your answers.

(a) Mg, Be, Ba
(b) Ca, Se, Ga
(c) Br, Rb, Kr
(d) Se, Br, Ca
(e) Ba, Sr, Cs
(f) Se, Br, Cl
(g) Mg, Ca, Li
(h) Sr, Te, Se
(i) In, Br, I
(j) S, Se, O

Trends for Ionization Energy

A neutral atom contains equal numbers of positive charges (protons) and negative charges (electrons). The particle that results when a neutral atom gains electrons or gives up electrons is called an **ion**. Thus, an ion is a charged particle. *An atom that gains electrons becomes a negatively charged* **anion**. *An atom that gives up electrons becomes a positively charged* **cation**. Figure 2.14 shows the formation of ions for several elements. As you examine the diagrams, pay special attention to

- the energy level from which electrons are gained or given up
- the charge on the ion that is formed when an atom gains or gives up electrons
- the arrangement of the electrons that remain after electrons are gained or given up

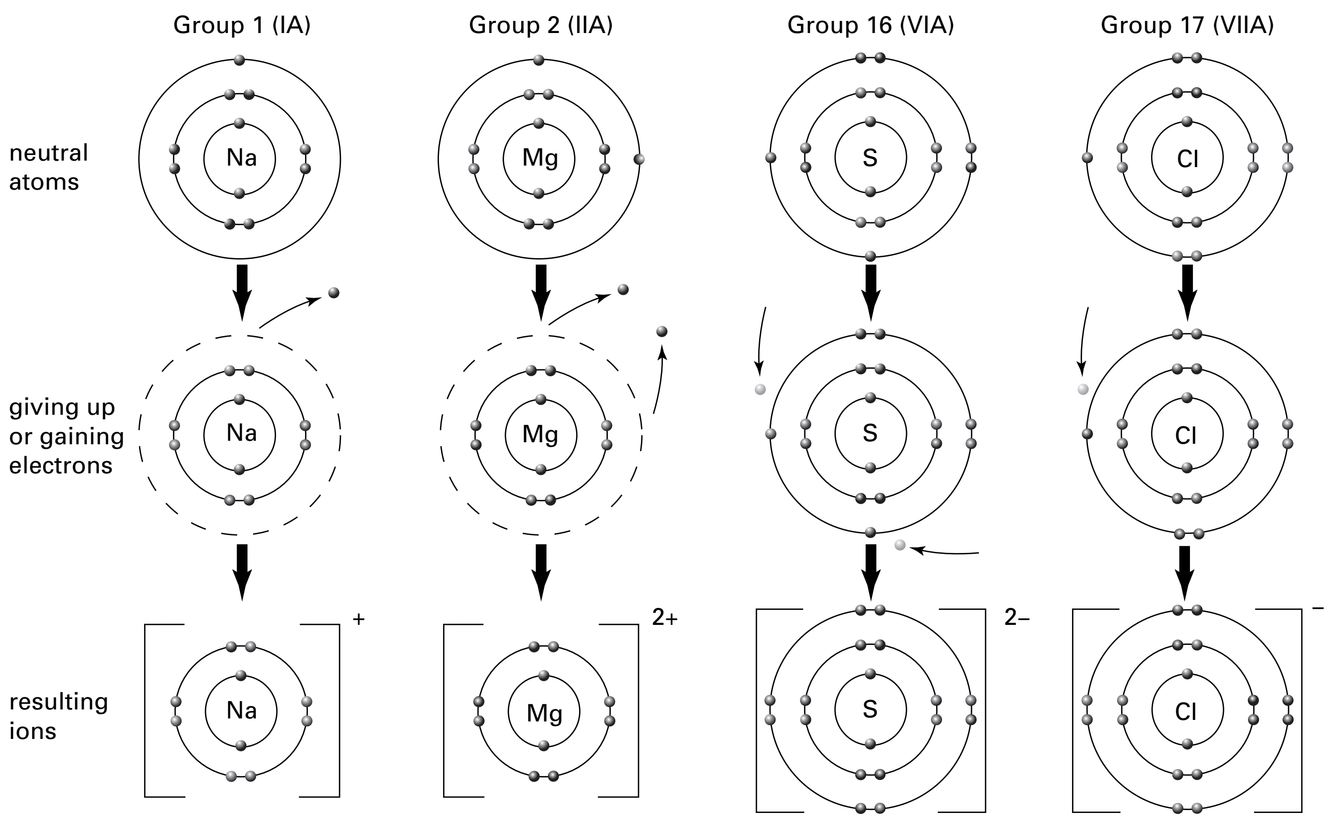


Figure 2.14 **These diagrams show the ions that are formed from neutral atoms of sodium, magnesium, sulfur, and chlorine. What other element has the same electron arrangement that sodium, magnesium, sulfide, and chloride ions have?**

Try to visualize the periodic table as a cylinder, rather than a flat plane. Can you see a relationship between ion formation and the electron arrangement of noble gases? Examine Figure 2.14 as well as 2.15 on the next page. The metals that are main-group elements tend to *give up* electrons and form ions that have the same number of electrons as the nearest noble gases. Non-metals tend to *gain* electrons and form ions that have the same number of electrons as the nearest noble gases. For example, when a sodium atom gives up its single valence electron, it becomes a positively charged sodium ion. Its outer electron arrangement is like neon's outer electron arrangement. When a fluorine atom gains an electron, it becomes a negatively charged ion with an outer electron arrangement like that of neon.

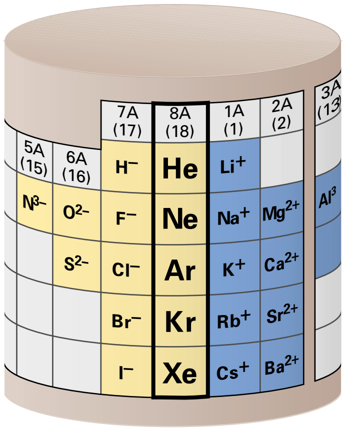


Figure 2.15 **Examine the relationship between ion charge and noble gas electron arrangement.**

Figure 2.15 can help you determine the charge on an ion. Count the number of groups an ion is from the nearest noble gas. That number is the charge on the ion. For example, aluminum is three groups away from neon. Thus, an aluminum ion has a charge of 3+. Sulfur is two groups away from argon. Thus, a sulfide ion has a charge of 2−. **Remember:** Metals form positive ions (cations) and non-metals form negative ions (anions).

It takes energy to overcome the attractive force of a nucleus and pull an electron away from a neutral atom. The energy that is required to remove an electron from an atom is called **ionization energy**. The bar graph in Figure 2.16 shows the ionization energy that is needed to remove one electron from the outer energy level of the atoms of the main-group elements. This energy is called the *first ionization energy.* It is measured in units of kJ/mol. A kilojoule (kJ) is a unit of energy. A mole (mol) is an amount of a substance. (You will learn about the mole in Unit 2.)

As you can see, atoms that give up electrons easily have low ionization energies. You would probably predict that the alkali metals of Group 1 (IA) would have low ionization energies. These elements are, in fact, extremely reactive because it takes so little energy to remove their single valence electron.

CHEM FACT

All elements, except hydrogen, have more than one electron that can be removed. Therefore, they have more than one ionization energy. The energy that is needed to remove a second electron is called the *second ionization energy.* The energy that is needed to remove a third electron is the *third ionization energy*, and so on. What trend would you expect to see in the values of the first, second, and third ionization energies for main-group elements? What is your reasoning?

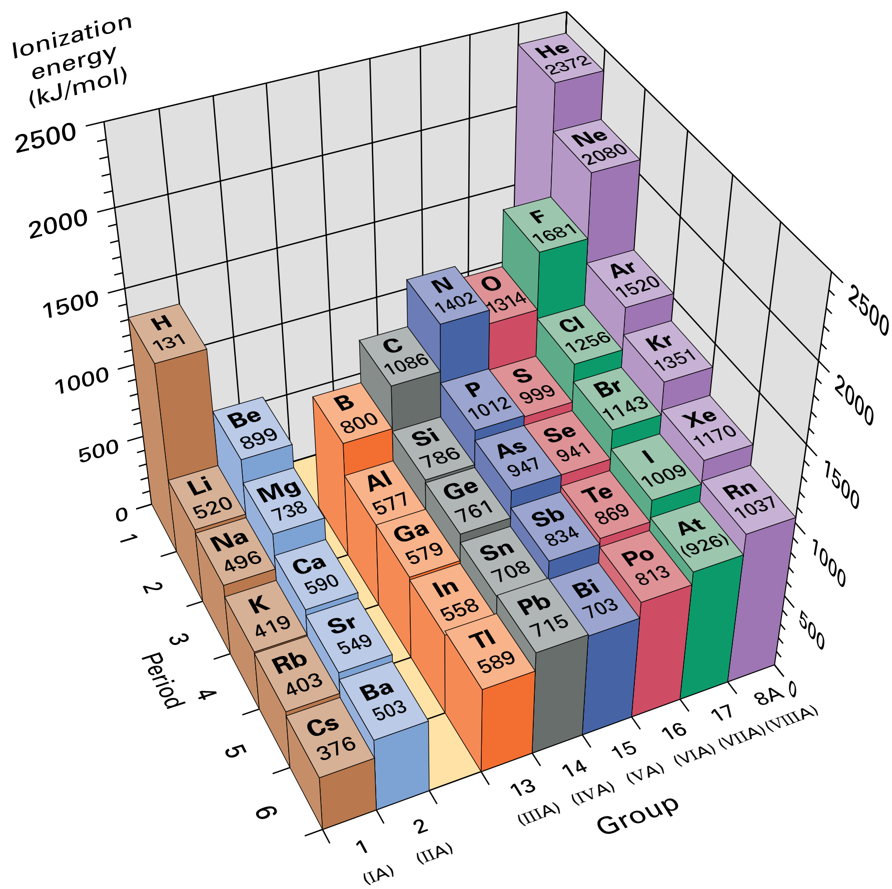


Figure 2.16 **This graph represents the first ionization energy for the main-group elements.**

Summarizing Trends for Ionization Energy

Although you can see a few exceptions in Figure 2.16, there are two general trends for ionization energy:

- *Ionization energy tends to decrease down a group.* This makes sense in terms of the energy level that the valence electrons occupy. Electrons in the outer energy level are farther from the positive force of the nucleus. Thus, they are easier to remove than electrons in lower energy levels.
- *Ionization energy tends to increase across a period.* As you go across a period, the attraction between the nucleus and the electrons in the outer energy level increases. Thus, more energy is needed to pull an electron away from its atom. For this trend to be true, you would expect a noble gas to have the highest ionization energy of all the elements in the same period. As you can see in Figure 2.16, they do.

Figure 2.17 summarizes these general trends for ionization energy. The Practice Problems below give you a chance to apply your understanding of these trends.

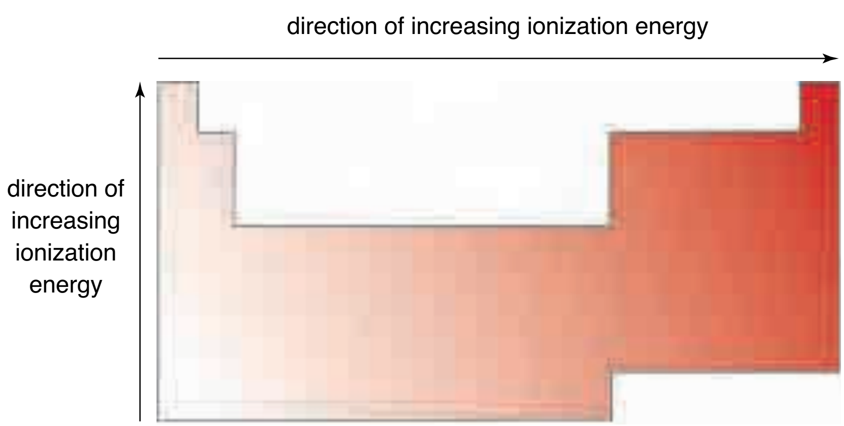


Figure 2.17 **Ionization energy tends to decrease down a group and increase across a period.**

Practice Problems

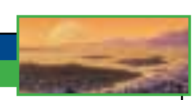
8. Using only a periodic table, rank the elements in each set by increasing ionization energy. Explain your answers.

(a) Xe, He, Ar	**(d)** Kr, Br, K
(b) Sn, In, Sb	**(e)** K, Ca, Rb
(c) Sr, Ca, Ba	**(f)** Kr, Br, Rb

9. Using only a periodic table, identify the atom in each of the following pairs with the *lower* first ionization energy.

(a) B, O	**(d)** F, N
(b) B, In	**(e)** Ca, K
(c) I, F	**(f)** B, Tl

COURSE CHALLENGE



Your understanding of periodic trends such as atomic radius and ionization energy will help you identify some unknown elements in the Chemistry Course Challenge at the end of this book.

Chemistry Bulletin

Science Technology Society Environment

Manitoba Mine Specializes in Rare Metals

At TANCO in Bernic Lake, Manitoba, miners are busy finding and processing two rare and very different metallic elements: tantalum and cesium. Both of these metals are important parts of "high-tech" applications around the world. They are used in nuclear reactors and as parts of aircraft, missiles, camera lenses, and surgical instruments like the one shown above.

Tantalum is found only in Canada, Australia, Brazil, Zaire, and China. TANCO (the Tantalum Mining Corporation of Canada) is the only mine in North America that produces tantalum. TANCO is also the world's main producer of cesium. Other than the fact that tantalum and cesium are both found at TANCO and both are used in high-tech applications, they share little in common.

Tantalum is a heavy, hard, and brittle grey metal. In its pure form, it is extremely ductile and can be made into a fine wire. This has proved useful for making surgical sutures. Another property that makes tantalum useful is its resistance to corrosion by most acids, due to its very limited reactivity. At normal temperatures, tantalum is virtually non-reactive. In fact, tantalum has about the same resistance to corrosion as glass. Tantalum can withstand higher temperatures than glass, however. It has a melting point of 3290 K—higher than the melting points of all other elements, except tungsten and rhenium. Tantalum's resistance to corrosion and high melting point make it suitable for use in surgical equipment and implants. For example, some of the pins that are used by surgeons to hold a patient's broken bones together are made of tantalum.

Tantalum is resistant to corrosion because a thin film of tantalum oxide forms when tantalum is exposed to oxygen. The metal oxide acts as a protective layer. The oxide also has special refractive properties that make it ideal for use in camera lenses.

Cesium is quite different from tantalum, but it, too, has many high-tech applications. Cesium is a silvery-white metal. It is found in a mineral called pollucite. Cesium is the softest of all the metals and is a liquid at just above room temperature. It is also the most reactive metal on Earth.

Cesium has a low ionization energy. It readily gives up its single valence electron to form crystalline compounds with all the halogen non-metals. Cesium is also very photoelectric. This means that it easily gives up its lone outer electron when it is exposed to light. Thus, cesium is used in television cameras and traffic signals. As well, it has the potential to be used in ion propulsion engines for travel into deep space.

Making Connections

1. Make a table to show the differences and similarities between tantalum and cesium. For each metal, add a column to describe how its different properties make it useful for specific applications.

2. Bernic Lake is one of the few locations where tantulum can be found. As well, it is the most important cesium source in the world. Research and describe what geographical conditions led to the presence of two such rare metals in one location.

Trends for Electron Affinity

In everyday conversation, if you like something, you may say that you have an affinity for it. For example, what if you enjoy pizza and detest asparagus? You may say that you have a high affinity for pizza and a low affinity for asparagus. If you prefer asparagus to pizza, your affinities are reversed.

Atoms are not living things, so they do not like or dislike anything. You know, however, that some atoms have a low attraction for electrons. Other atoms have a greater attraction for electrons. **Electron affinity** is a measure of the change in energy that occurs when an electron is added to the outer energy level of an atom to form a negative ion.

Figure 2.18 identifies the electron affinities of the main-group elements. If energy is released when an atom of an element gains an electron, the electron affinity is expressed as a negative integer. When energy is absorbed when an electron is added, electron affinity is low, and is expressed as a positive integer. Notice, for example, that fluorine has the highest electron affinity (indicated by a large, negative integer). This indicates that fluorine is very likely to be involved in chemical reactions. In fact, fluorine is the most reactive of all the elements.

Metals have very low electron affinities. This is especially true for the Group 1 (IA) and 2 (IIA) elements. Atoms of these elements form stable positive ions. A negative ion that is formed by the elements of these groups is unstable. It breaks apart into a neutral atom and a free electron.

Examine Figure 2.18. What trends can you observe? How regular are these trends?

1 (IA)	2 (IIA)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
H −72.8							**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 2.18 **The units for electron affinity are the same as the units for ionization energy: kJ/mol. High negative numbers mean a high electron affinity. Low negative numbers and any positive numbers mean a low electron affinity.**

Tools & Techniques

Analyzing the Ice Man's Axe

In September 1991, hikers in the Alps Mountains near the Austrian-Italian border discovered the body of a man who had been trapped in a glacier. He was almost perfectly preserved. With him was an assortment of tools, including an axe with a metal blade.

Scientists were particularly interested in the axe. At first, they believed that it was bronze, which is an alloy of copper and tin. There was a complication, however. Dating techniques that were used for the clothing and body suggested that the "Ice Man" was about 5300 years old. Bronze implements do not appear in Europe's fossil record until about 4000 years ago. Either Europeans were using bronze earlier than originally thought, or the axe was made of a different material. Copper was consistent with the Ice Man's age, since it has been used for at least the past 6000 years.

One technique to determine a metal's identity is to dissolve it in acid. The resulting solution is examined for evidence of ions. Scientists did not want to damage the precious artifact in any way, though.

The solution was an analytical technique called *X-ray fluorescence*. The object is irradiated with high-energy X-ray radiation. Its atoms absorb the radiation, causing electrons from a lower energy level to be ejected from the atom. This causes electrons from an outer energy level to "move in," to occupy the vacated space. As the electrons fall to a less energetic state, they emit X-rays. The electrons of each atom emit X-rays of a particular wavelength. Scientists use this energy "signature" to identify the atom.

Analysis by X-ray fluorescence revealed that the metal in the blade of the axe was almost pure copper.

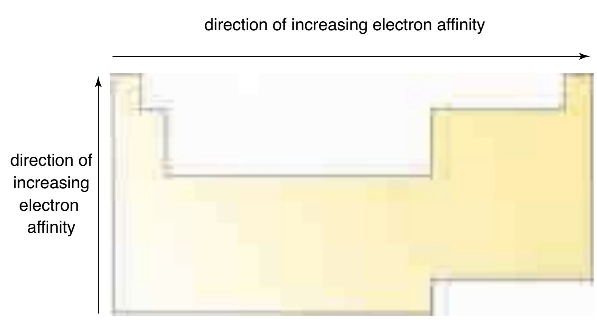


Figure 2.19 **Electron affinity tends to decrease down a group and increase across a period.**

The trends for electron affinity, shown in Figure 2.19, are more irregular than the trends for ionization energy and atomic radius. Nevertheless, the following *general* trends can be observed:

- *Electron affinity tends to decrease down a group*. For example, fluorine has a higher electron affinity than iodine.
- *Electron affinity tends to increase across a period*. For example, calcium has a lower electron affinity than sulfur.

ThoughtLab Design an Annotated Periodic Table

You have learned a great deal about the properties of the elements. In the following chapters, you will learn more. With your classmates, develop your own large-scale periodic table to record the properties and common uses of the elements.

Procedure

1. Use print and electronic resources (including this textbook) to find information about one element. Consult with your classmates to make sure that everyone chooses a different element.
2. Find the following information about your element:
   - atomic number
   - atomic mass
   - atomic symbol
   - melting point
   - boiling point
   - density
   - atomic radius
   - ionization energy
   - electron affinity
   - place and date discovered, and the name of the scientist who discovered it
   - uses, both common and unusual
   - hazards and methods for safe handling

   If possible, find a photograph of the element in its natural form. If this is not possible, find a photograph that shows one or more compounds in which the element is commonly found.
3. Record your findings on a sheet of notepaper or blank paper. Arrange all the sheets of paper, for all the elements, in the form of a periodic table on a wall in the classroom. Make sure that you leave space to insert additional properties and uses of your element as you learn about them during this course.

Analysis

1. What uses of your element did you know about? Which uses surprised you? Why?
2. Examine the dates on which the elements were discovered. What pattern do you notice? How can you explain this pattern?
3. Do you think that scientists have discovered all the naturally occurring elements? Do you think they have discovered all the synthetically produced elements? Give reasons to justify your opinions.

Section Wrap-up

Despite some irregularities and exceptions, the following periodic trends summarize the relationships among atomic size, ionization energy, and electron affinity:

- Trends for atomic size are the reverse of trends for ionization energy and electron affinity. Larger atoms tend to have lower ionization energies and lower electron affinities.
- Group 16 (VIA) and 17 (VIIA) elements attract electrons strongly. They do not give up electrons readily. In other words, they have a strong tendency to form negative ions. Thus, they have high ionization energies and high electron affinities.
- Group 1 (IA) and 2 (IIA) elements give up electrons readily. They have low or no attraction for electrons. In other words, they have a strong tendency to form positive ions. Thus, they have low ionization energies and low electron affinities.
- Group 18 (VIIIA) elements do not attract electrons and do not give up electrons. In other words, they do not naturally form ions. (They are very stable.) Thus, they have very high ionization energies and very low electron affinities.

The trends you have examined in this chapter have an enormous influence on the ability of atoms to combine and form compounds. In the next chapter, you will use these trends to help you understand and predict the kinds of compounds that atoms form. As well, you will learn about another periodic trend. This trend called electronegativity, is related to the formation of some of the most common compounds in your life, such as water, carbon dioxide, and sugar.

Section Review

1 K/U How does your understanding of electron arrangement and forces in atoms help you explain the following periodic trends?

(a) atomic radius
(b) ionization energy
(c) electron affinity

2 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of increasing atomic size. Explain your answer in each case.

(a) Mg, S, Cl
(b) Al, B, In
(c) Ne, Ar, Xe
(d) Rb, Xe, Te
(e) P, Na, F
(f) O, S, N

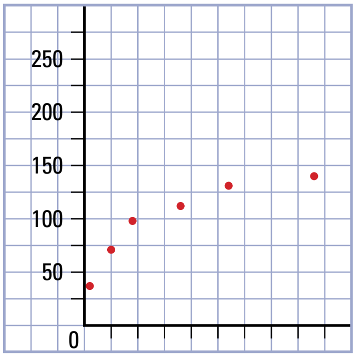
3 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of decreasing ionization energy. Explain your answer in each case.

(a) Cl, Br, I
(b) Ga, Ge, Se
(c) K, Ca, Kr
(d) Na, Li, Cs
(e) S, Cl, Br
(f) Cl, Ar, K

4 K/U Which element in each of the following pairs will have the lower electron affinity? Explain your answer in each case.

(a) K or Ca
(b) O or Li
(c) S or Se
(d) Cs or F

5 C The graph shows a periodic trend, but is only partially complete. Copy it into your notebook and fill in all the data and labels that will make it complete. Title the graph with the trend it shows.



What data does this graph need to be complete?

6 I Use your understanding of periodic trends to sketch the shape of a graph that shows a trend that is opposite to that shown in question 5. Label the x- and y-axes, and add any other labels that you think are necessary to represent the trend you are showing.

Unit Project Prep

Look ahead to the project at the end of this unit. Many common chemical products contain elements (as components of compounds) from some groups of the periodic table more than others. Examine ingredient labels from different chemical products. Which element groups are represented most frequently? Why might that be?

CHAPTER 2 Review

Reflecting on Chapter 2

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

- Identify the subatomic particles that make up atoms, as well as the theory that chemists use to explain the composition and behaviour of atoms.
- Use the periodic law to examine the structure and organization of the periodic table of the elements.
- Draw Lewis structures to model the arrangements of electrons in the outer energy levels of atoms.
- Identify periodic trends involving atomic size, ionization energy, and electron affinity.

Reviewing Key Terms

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

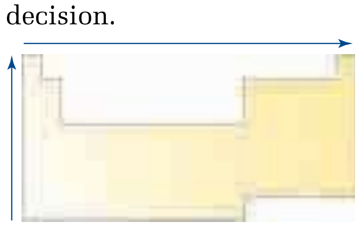
element
atom
atomic number (Z)
isotope
periodic law
periodic trend
Lewis structure
octet
ionization energy
atomic mass unit (u)
mass number (A)
radioisotope
energy level
valence electrons
stable octet
atomic radius
electron affinity

Knowledge/Understanding

1. Explain the difference between an atom and an element.
2. Compare protons, neutrons, and electrons in terms of their charge, their mass, and their size.
3. What information does the following notation express: $^{16}_{8}O$?
4. Write an equation that shows how to calculate the number of neutrons in a neutral atom if you know its mass number and its atomic number.
5. Use an example and the appropriate terminology to explain the difference between an isotope and a radioisotope.
6. In your notebook, copy the table below and fill in the missing information.

Symbol	Protons	Neutrons	Electrons	Charge
$^{14}_{7}N^{-3}$	(a)	(b)	(c)	(d)
(e)	34	45	36	(f)
$^{52}_{24}(g)^{3+}$	(h)	(i)	(j)	(k)
$^{(l)}_{(m)}(F)$	(n)	10	(o)	(p)

7. A cobalt atom has an atomic mass of 59 and an atomic number of 27. How many neutrons does it have? How many electrons does it have?
8. **(a)** Hydrogen atoms are lying side-by-side along a line that is 1 mm long. How many hydrogen atoms are there?
   (b) How many potassium atoms would lie side-by-side along the same 1 mm line?
9. Use a Venn diagram or a graphic organizer of your choice to compare Dalton's atomic theory with the more modern atomic theory that you learned about in this chapter.
10. Invent an entirely different name for the periodic table. Give reasons to support your choice.
11. Consider the following elements: H, Li, N, F, Co, Ag, Kr, I, Hg.
    (a) Sketch an outline of the periodic table, with these elements properly placed.
    (b) State the group number and period number each element belongs to.
    (c) Identify each element as a metal, metalloid, or non-metal.
    (d) Identify the state of each element at room temperature.
    (e) Draw the Lewis structure for each of these elements.
12. **(a)** Which of the following trends is best represented by this diagram: atomic size, ionization energy, or electron affinity? Justify your decision.



    (b) Sketch outlines of the periodic table to show the trends for the remaining two choices from part (a). Explain how these trends are related to the one in part (a).

13. Arrange the following elements into groups that share similar properties: Ca, K, Ga, P, Si, Rb, B, Sr, Sn, Cl, Bi, Br. How much confidence do you have in your groupings, and why?

14. Use a drawing of your choice to show clearly the relationship among the following terms: valence, stable octet, electron, energy level.

15. In what ways are periodic trends related to the arrangement of electrons in atoms?

Inquiry

16. Imagine hearing on the news that somebody has discovered a new element. The scientist who discovered this element claims that it fits between tin and antimony on the periodic table.
    (a) How likely is it for this claim to be true? Justify your answer.
    (b) Write at least three questions that you could ask this scientist. What is your reasoning for asking these questions? (In other words, what do you expect to hear that could help convince you that the scientist is right or that you are?)

17. Technetium, with an atomic number of 43, was discovered after Mendeleev's death. Nevertheless, he used the properties of manganese, rhenium, molybdenum, and ruthenium to predict technetium's properties.
    (a) Use a chemical database to find the following properties for the above-mentioned elemental "neighbours" of technetium:
    - atomic mass
    - appearance
    - melting point
    - density

    If you would like to truly follow in Mendeleev's footsteps, you could also look for the chemical formulas of the compounds that these elements form with oxygen and chlorine. (Such compounds are called oxides and chlorides.)
    (b) Use this data to predict the properties for technetium.
    (c) Consult a chemical data base to assess your predictions against the observed properties for technetium.

Communication

18. Explain how you would design a data base to display information about the atomic numbers, atomic masses, the number of subatomic particles, and the number of electrons in the outer energy levels of the main-group elements. If you have access to spreadsheet software, construct this table.

19. (a) Decide on a way to compare, in as much detail as you can, the elements sodium and helium. The following terms should appear in your answer. Use any other terms that you think are necessary to complete your answer fully.

atom	element
nucleus	proton
neutron	electron
energy level	valence
periodic table	periodic trend
group	period
atomic radius	electron affinity
ionization energy	

    (b) Modify your answer to part (a) so that a class of grade 4 students can understand it.

20. Element A, with three electrons in its outer energy level, is in Period 4 of the periodic table. How does the number of its valence electrons compare with that of Element B, which is in Group 13 (IIIA) and Period 6? Use Lewis structures to help you express your answer.

21. Which elements would be affected if the elements in Periods 1, 2, 3, and 4 were arranged based on their atomic mass, rather than their atomic number? Based on what you have learned in this chapter, how can you be reasonably sure that arranging elements by their atomic number is accurate?

Making Connections

22. "When she blew her nose, her handkerchief glowed in the dark." The woman who made this statement in the early 1900s was one of several factory workers who were hired to paint clock and watch dials with luminous paint. This paint glowed in the dark, because it contained radium (atomic number 88), which is

highly radioactive and toxic. Marie and Pierre Curie discovered radium in 1898. Chemists knew as early as 1906 that the element was dangerous. Nevertheless, it was used not only for its "glowing effects," but also as a medicine. In fact, several companies produced drinks, skin applications, and foods containing radium.

Choose either one of the topics below for research.

- the uses and health-related claims made for radium during the early 1900s
- the story of the so-called "radium girls"—the factory workers who painted clock and watch faces with radium paint

How does the early history of radium and its uses illustrate the need for people to understand the connections among science, technology, society, and the environment?

23. Have you ever heard someone refer to aluminum foil as "tin foil"? At one time, the foil was, in fact, made from elemental tin. Find out why manufacturers phased out tin in favour of aluminum. Compare their chemical and physical properties. Identify and classify the products made from or with aluminum. What are the technological costs and benefits of using aluminum? What health-related and environment-related issues have surfaced as a result of its widespread use in society? Write a brief report to assess the economic, social, and environment impact of our use of aluminum.

Answers to Practice Problems and Short Answers to Section Review Questions:

Practice Problems: 1. (a) boron **(b)** 5 **(c)** 6 **(d)** lead **(e)** 82 **(f)** 126 **(g)** $^{184}_{74}W$ **(h)** 74 **(i)** $^{4}_{2}He$ **(j)** 2 **(k)** plutonium **(l)** 94 **(m)** 145 **(n)** Fe **(o)** iron **(p)** 30 **(q)** $^{209}_{83}Bi$ **(r)** 83 **(s)** $^{154}_{47}Ag$ **(t)** silver **(u)** Ne **(v)** neon **(w)** 10 **(x)** 10 **2. (a)** carbon **(b)** copper **(c)** radon **(d)** hydrogen **(e)** cadmium **(f)** calcium **(g)** iodine **(h)** aluminum **3.** H and He have one occupied energy level; H has 1 electron, He has 2. Li, Be, B, C, N, O, F, and Ne have two energy levels. First energy level is filled with two electrons. Second energy of Li has 1 electron, and electrons increase by one, totaling 8 in outer energy level for Ne. Na, Mg, Al, Si, P, S, Cl, Ar have three occupied energy levels. First two energy levels are full. Third energy level of Na has 1 electron, and electrons increase by one, totaling 8 in outer energy level for Ar. K and Ca have four occupied energy levels. First three energy levels are full. K has 1 electron and Ca has two in outer energy level. **4.** The pattern of outer energy level electrons from Practice Problem 3 is repeated by placing dots around the atomic symbol for each element. **5. (a)** 7 **(b)** 2 **(c)** 3 **(d)** 2 **(e)** 1 **(f)** 4 **(g)** 5 **(h)** 6 **(i)** 5 **(j)** 8 **6.** Ba has two dots; Ga has 3 dots; Sn has 4 dots; Bi has 5 dots; I has 7 dots; Cs has 1 dot; Kr has 8 dots; Xe has 8 dots. **7. (a)** Ba, Mg, Be **(b)** Ca, Ga, Se **(c)** Rb, Br, Kr **(d)** Ca, Se, Br **(e)** Cs, Ba, Sr **(f)** Se, Br, Cl **(g)** Ca, Mg, Li **(h)** Sr, Te, Se **(i)** In, I, Br **(j)** Se, S, O **8. (a)** Xe, Ar, He **(b)** In, Sn, Sb **(c)** Ba, Sr, Ca **(d)** K, Br, Kr **(e)** Rb, K, Ca **(f)** Rb, Br, Kr **9. (a)** B **(b)** In **(c)** I **(d)** N **(e)** K **(f)** Tl

Section Review: 2.1: 1. (a) silver **(b)** 47 **(c)** 47 **(d)** 61 **(e)** arsenic **(f)** 33 **(g)** 75 **(h)** 33 **(i)** bromine **(j)** 80 **(k)** 35 **(l)** 35 **(m)** gold **(n)** 79 **(o)** 79 **(p)** 100 **(q)** tin **(r)** 50 **(s)** 119 **(t)** 50 **3. (a)** last pair has same protons and electrons, but different neutrons **(b)** last pair has same value for Z; first pair has same value for A **2.2: 2.** 1 (1A), 18 (8A), 17 (7A), 2 (2A) **3. (a)** Si; Sb; Te; Br **6. (a)** 8; 7; 6; 2; 1; 7; 2; 5; 2; 4. **(c)** non-metal; non-metal; non-metal; metal; metal; non-metal; non-metal; metal, metal; metalloid **7.** two: mercury and bromine **9. (b)** Triads 2 and 3 **11. (b)** Na, Mg, and Al have same number of energy levels, as do Li and C **(c)** Li, Na, and K have same number of valence electrons **2.3: 2. (a)** Cl, S, Mg **(b)** B, Al, In **(c)** Ne, Ar, Xe **(d)** Xe, Te, Rb **(e)** F, P, Na **(f)** O, N, S **3. (a)** Cl, Br, I **(b)** Se, Ge, Ga **(c)** Kr, Ca, K **(d)** Li, Na, Cs **(e)** Cl, Br, S **(f)** Ar, Cl, K **4. (a)** Ca, **(b)** Li **(c)** Se **(d)** Cs