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

Period	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

Period														
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.**

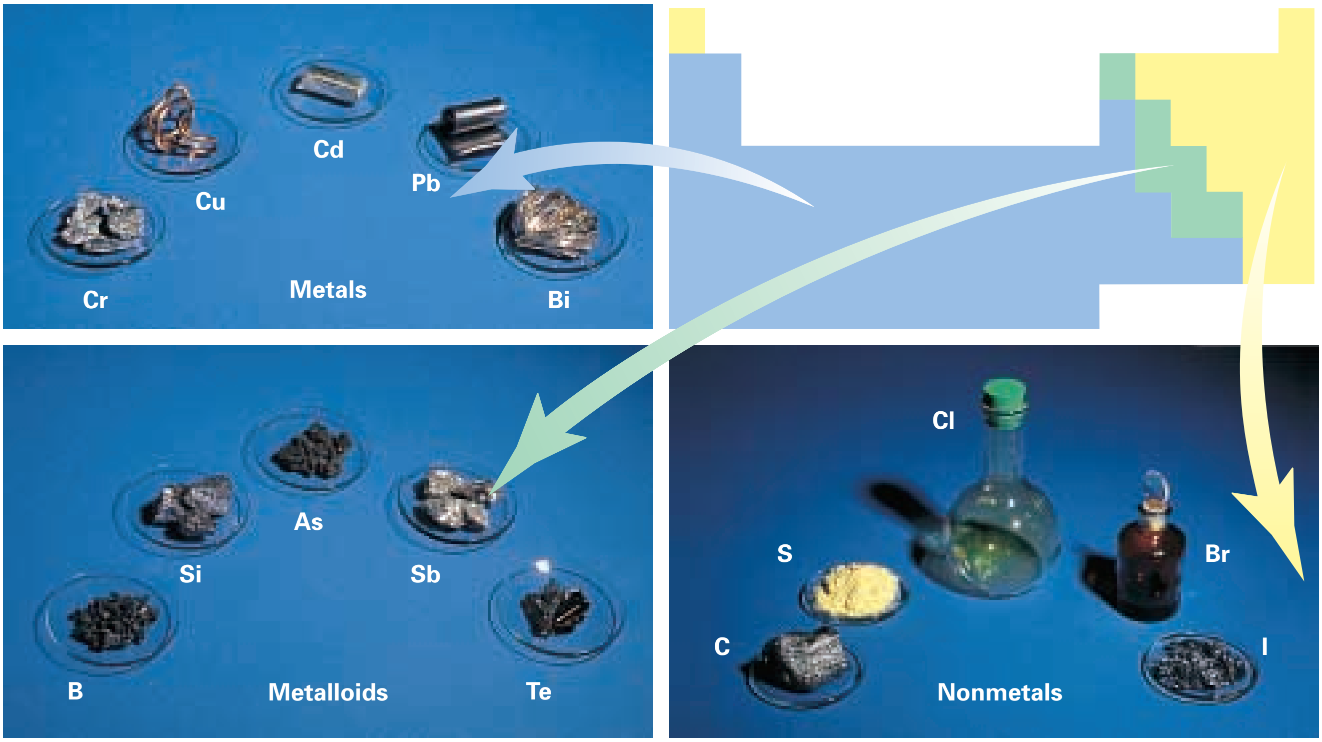

Figure 2.7 Several examples from each of the three main classes of elements are shown here. Find where they appear in the periodic table in Figure 2.6.

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

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.

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?

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.

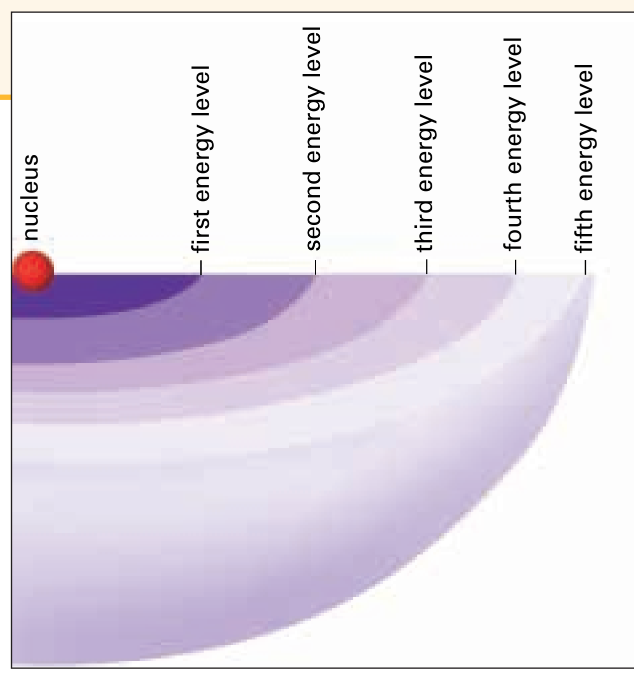

Figure 2.8 Energy levels of an atom from the fifth period

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.

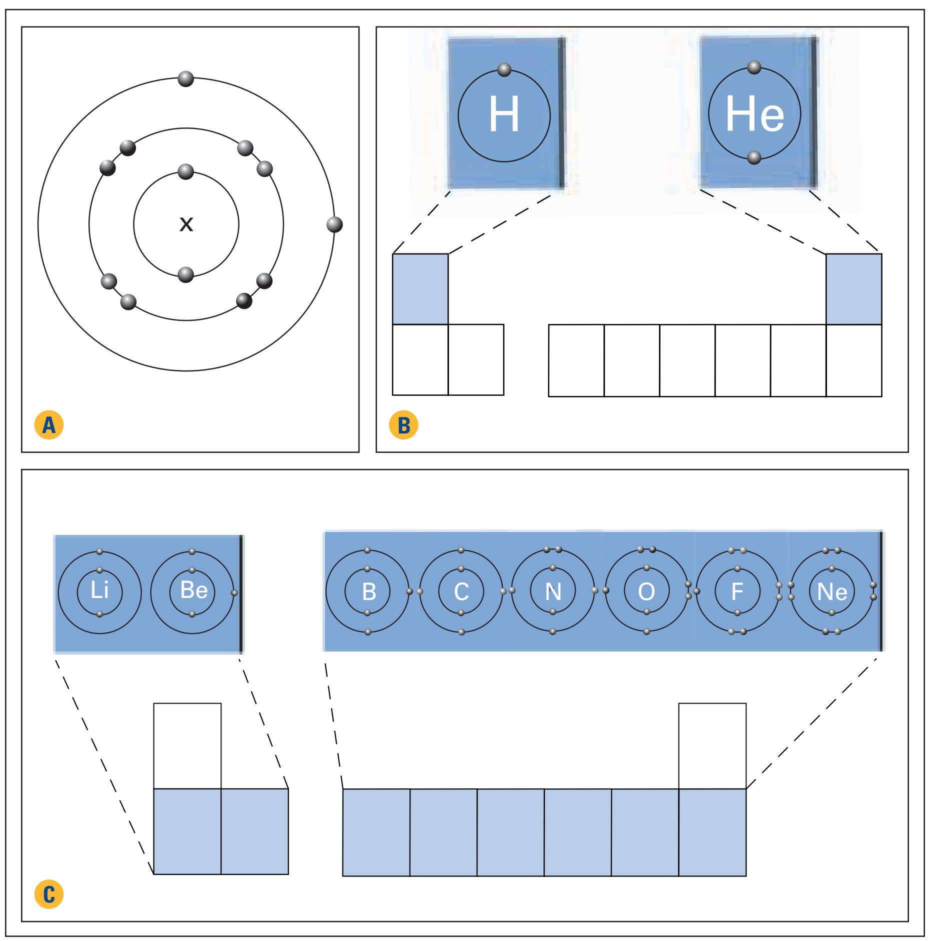

Figure 2.9 **(A) A Bohr-Rutherford diagram (B) Hydrogen and helium have a single energy level. (C) The eight Period 2 elements have two energy levels.**

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	2															3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA	IIA															IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	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?