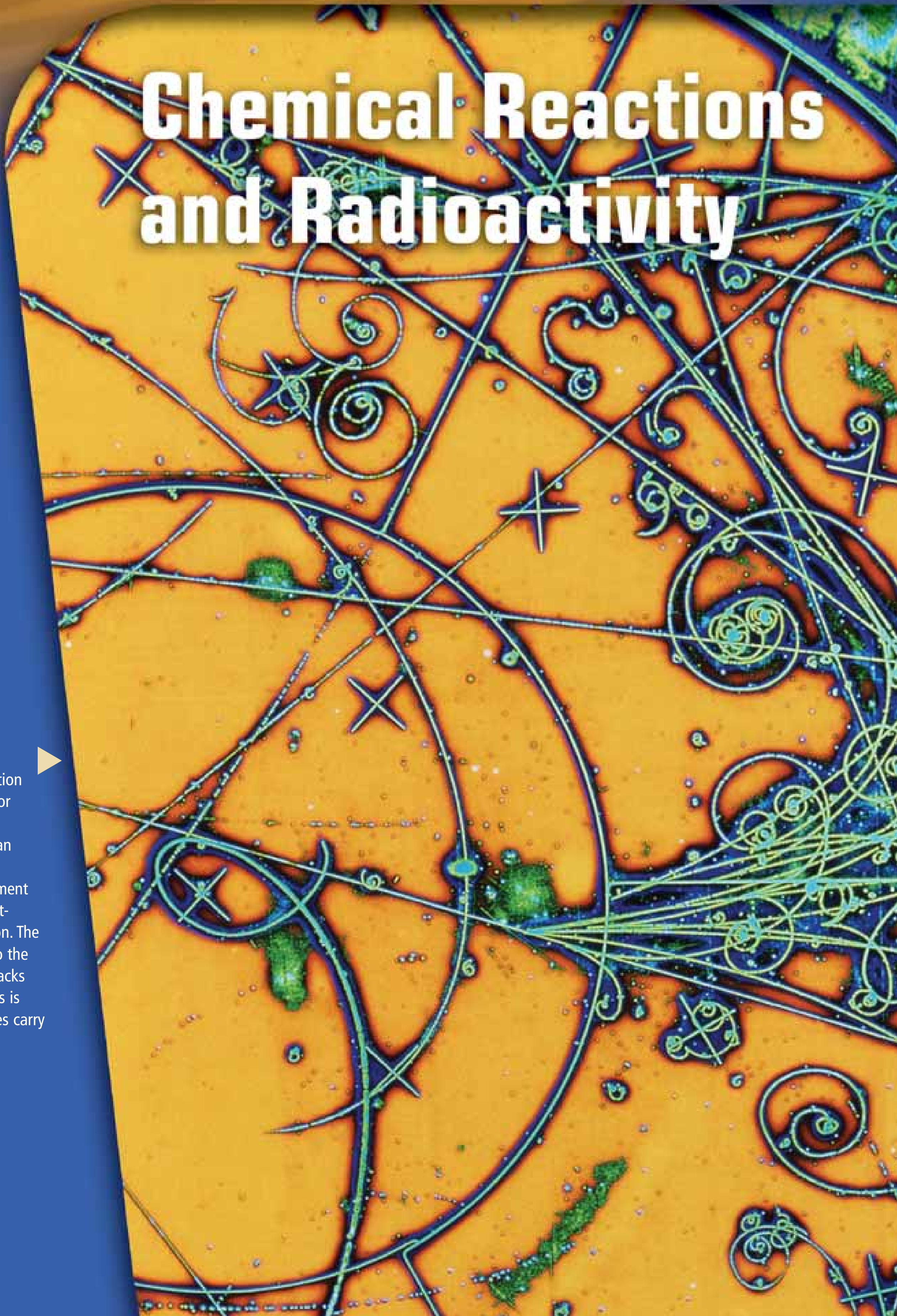
UNIT

2 Chemical Reactions and Radioactivity



What does this photograph look like to you? Is it a portion of a stained glass window or some kind of elaborate metalwork? In fact, this is an image produced by a cloud chamber, a scientific instrument that reveals the path of fast-moving particles of radiation. The colours have been added to the image. Notice that many tracks take the form of spirals. This is evidence that those particles carry an electric charge.

Key Ideas

4 **Atomic theory explains the formation of compounds.**

4.1 Atomic Theory and Bonding
4.2 Names and Formulas of Compounds
4.3 Chemical Equations

5 **Compounds are classified in different ways.**

5.1 Acids and Bases
5.2 Salts
5.3 Organic Compounds

6 **Chemical reactions occur in predictable ways.**

6.1 Types of Chemical Reactions
6.2 Factors Affecting the Rate of Chemical Reactions

7 **The atomic theory explains radioactivity.**

7.1 Atomic Theory, Isotopes, and Radioactive Decay
7.2 Half-Life
7.3 Nuclear Reactions

Getting Started

Dr. Ross Chapman maps methane hydrates on the sea floor using ROPOS, operated out of Sidney, British Columbia.

The methane trapped in the ice burns.

Could it ever be possible to light chunks of ice on fire? The answer is yes, if the chunks contain methane hydrates. Hydrates are chemicals that easily join with water molecules. Methane is the principal component of the natural gas used to heat many homes in Canada.

Methane hydrates can be found in ocean trenches, which are the deepest parts of the ocean where the sea floor is slowly moving down into Earth's crust. Along the west coast of British Columbia, the floor of the Pacific Ocean is sliding beneath North America. Deep inside Earth, heat and pressure cause the formation of methane gas. When the methane gas bubbles seep to the surface of the sea floor, they encounter ice and soak into it, becoming trapped in "water cages." This is why it is possible to literally light this special form of ice on fire—the methane trapped in the water burns.

At the University of Victoria, Dr. Ross Chapman is mapping Canadian methane hydrate reserves with the help of the submersible shown behind him in the photograph. The submersible is called ROPOS—Remotely Operated Platform for Ocean Science. Working 80 km off the west coast of Vancouver Island at a depth of nearly 1000 m, ROPOS cuts off chunks of methane hydrates and stores them under pressure. This prevents the chunks from decomposing while being brought to the surface. As long as the pressure remains high, the methane will remain trapped in the ice.

internet connect

Find out more about methane hydrates. Visit **www.bcscience10.ca**.

Calcium Metal in Water

Find Out ACTIVITY

In this activity, you will add calcium metal to water. Work safely and cooperatively. Use this opportunity to increase your laboratory skills.

Safety

- Avoid touching the calcium. Calcium reacts with moisture, including the moisture on your hands.
- Follow your teacher's directions regarding using open flames.
- Tie back long hair.
- Be sure to wear eye protection.
- Avoid touching all reactants and products.
- Wash your hands and equipment thoroughly after completing this activity.
- Do not remove any materials from the science room.

Materials

- medium diameter test tube
- test tube rack
- large diameter test tube (to fit over medium test tube)
- water
- matches or flame striker
- candle or Bunsen burner
- sandpaper or triangular file
- calcium metal
- paper towel
- test tube tongs
- wooden splints

Step 4

What to Do

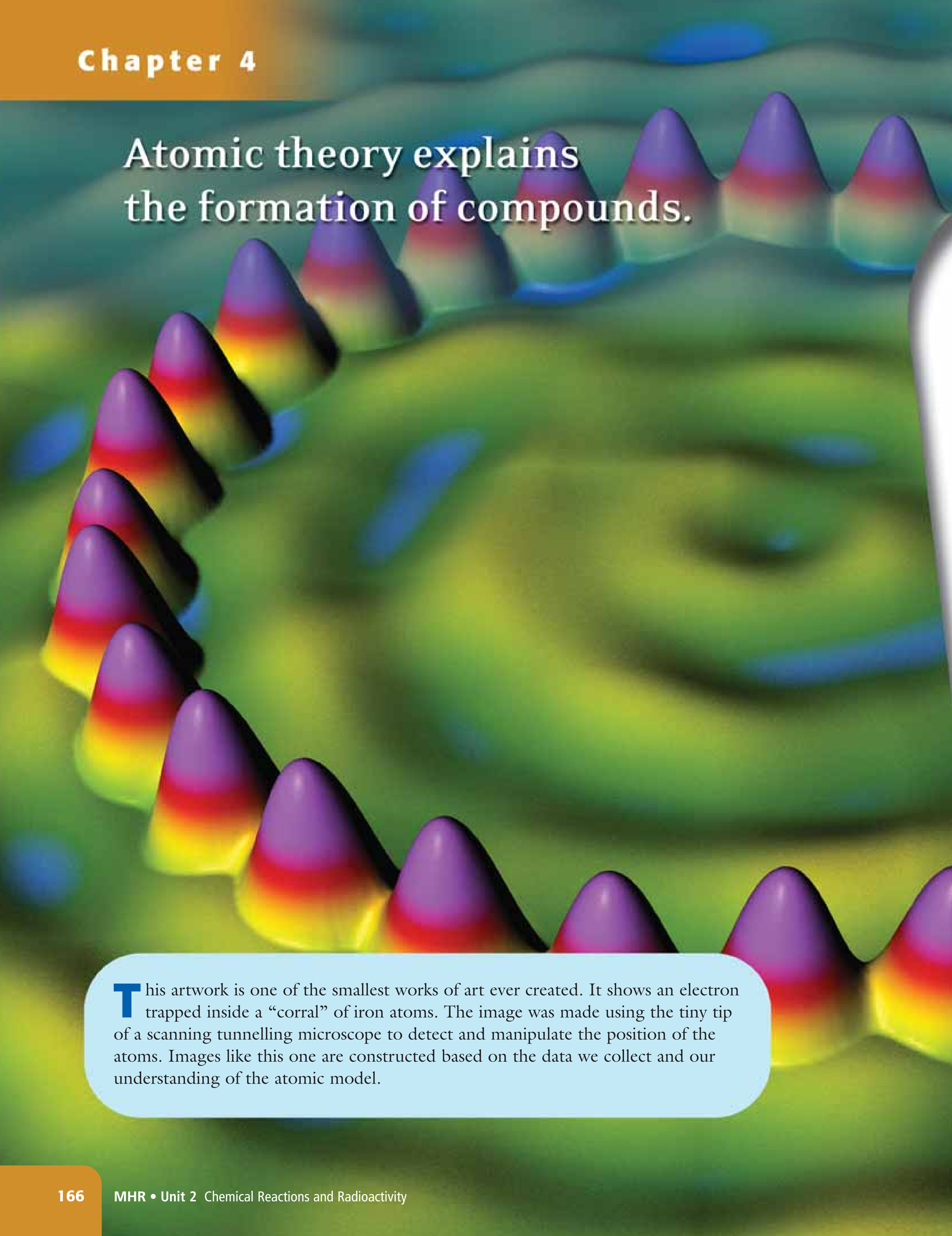
1. Your teacher will review the detailed safety information on pages xxii to xxv before you begin this activity.
2. Create a data table to record observations you make during the activity. Give your table a title.
3. Place the medium test tube in the test tube rack. Make sure the large test tube will fit over the medium test tube. Set the large test tube aside, such as in the spine of your opened textbook. Place water in the medium test tube to a depth of about 3 cm.
4. Set up a lit candle or Bunsen burner according to your teacher's instructions.
5. Use sandpaper or a triangular file to expose a fresh piece of the calcium metal surface. Do this over a piece of wet paper towel. Be sure to wear gloves. Do not touch the calcium with your bare hands. Observe.
6. Place the calcium metal into the water in the medium test tube. Slide the large test tube over the mouth of the medium one.
7. Observe for about 30 s, then use test tube tongs to lift the large test tube off the medium one. Keeping the large test tube turned upside down, bring a lit wooden splint near the mouth of the large test tube. Firmly hold on to the large test tube. Be prepared for a reaction, and do not drop the test tube!
8. Repeat the experiment if time permits.
9. Record your observations.
10. Clean up and put away the equipment you have used. Follow your teacher's instructions for disposal of wastes.

What Did You Find Out?

1. Reflect on this activity in terms of laboratory safety. What safety issues are important to this activity?
2. (a) What physical changes did you observe?
   (b) What chemical changes do you think happened? Explain.

Chapter 4

Atomic theory explains the formation of compounds.



This artwork is one of the smallest works of art ever created. It shows an electron trapped inside a "corral" of iron atoms. The image was made using the tiny tip of a scanning tunnelling microscope to detect and manipulate the position of the atoms. Images like this one are constructed based on the data we collect and our understanding of the atomic model.

What You Will Learn

In this chapter, you will
- **distinguish** between atoms, ions, and molecules
- **describe** the arrangement of electrons in atoms, ions, and molecules
- **write** names and formulas for ionic and covalent compounds
- **balance** chemical equations
- **explain** the law of conservation of mass as it applies to chemical reactions

Why It Is Important

The atomic theory gives us a way to visualize matter that is too small to be viewed directly so we can understand and predict the changes we see in our everyday world. By learning more about how molecules and compounds form and break apart, we can use chemicals more effectively in industry and develop better medicines.

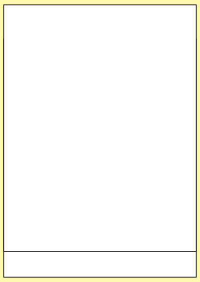
Skills You Will Use

In this chapter, you will
- **work co-operatively and safely** in a laboratory setting
- **observe** changes in properties of matter undergoing chemical change
- **classify** compounds as ionic or covalent
- **use models** to understand the structure of matter and the process of chemical changes

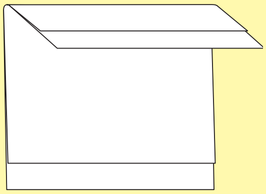
FOLDABLES™
Reading & Study Skills

Make the following Foldable and use it to take notes on what you learn in Chapter 4.

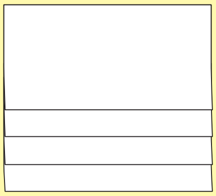
STEP 1 **Stack** two sheets of paper (22 cm by 28 cm) so that the back sheet is 2.5 cm lower than the front sheet.



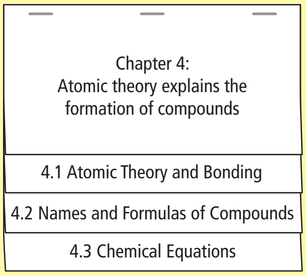
STEP 2 **Bring** the top of both sheets downward. **Align** the edges so that all of the layers or tabs are the same distance apart.



STEP 3 When all tabs are an equal distance apart, **fold** the papers and **crease** well.



STEP 4 **Open** the papers. **Glue** them together along the inner centre fold, or **staple** them along the mountain. **Label** as shown.



Organize As you progress through the chapter, summarize key points, record information, and define terms beneath the appropriate tabs. Use the back of the Foldable to practise drawing models of compounds, writing formulas, and writing and balancing equations.

4.1 Atomic Theory and Bonding

Atoms are composed of protons and neutrons, which make up the nucleus, and electrons, which surround the nucleus in patterns. Bohr diagrams show the arrangement of protons, neutrons, and electrons in atoms and also in ions. Ions are formed from atoms that have lost or gained electrons. Compounds can be ionic or covalent. Lewis diagrams show the arrangement of bonds within compounds.

Words to Know

atomic number
Bohr diagram
compound
covalent bonding
ionic bonding
ions
Lewis diagram
molecule
valence electrons

Did You Know?

Imagine dumping sand onto an area the size of a hockey arena until the ice rink is buried under 30 m of sand. Suppose that volume of sand represents 1 s. Half of the volume represents 0.5 s. How much sand would represent one femtosecond (10^{-15} s)? A femtosecond would be one single grain of sand.

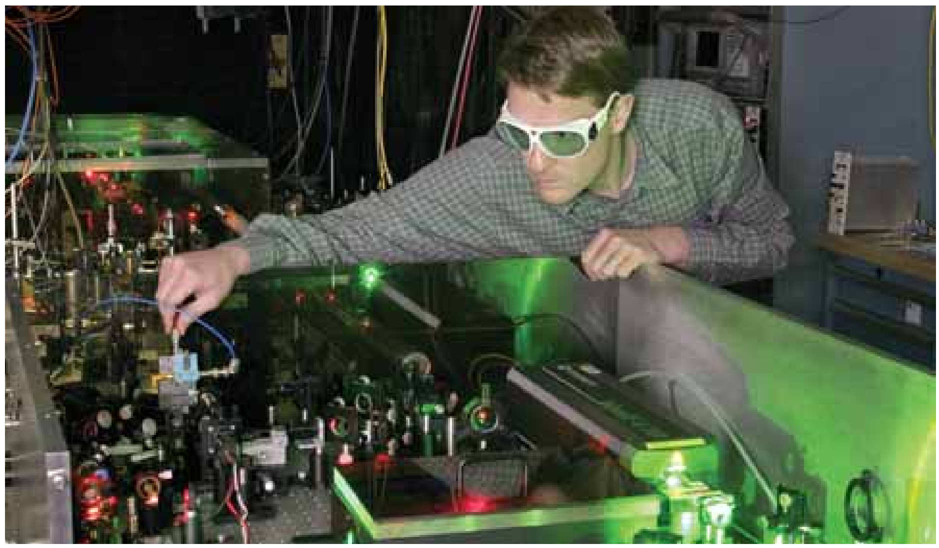


Figure 4.1A A researcher adjusts a component on a femtosecond laser at the National Institute of Standards and Technology. The apparatus is used for imaging chemical changes.

How could you capture the result of the collisions of atoms and molecules as they slam into each other during a chemical change? You could use the world's fastest camera and most powerful microscope—a femtosecond laser (Figure 4.1A).

In earlier science studies, you learned that a pure substance is made up of only one kind of matter. There are two categories of pure substances. An element is a pure substance that cannot be chemically broken down into simpler substances. A **compound** is a pure substance that is composed of two or more atoms combined in a specific way.

An **atom** is the smallest particle of any element that retains the properties of the element. How small is an atom? If you could line up 50 million atoms, the line would be about 1 cm long.

Because atoms are incredibly small, you cannot see an atom with regular light. With a femtosecond laser, the energy from an extremely fast pulse of laser light is used to position atoms and detect chemical changes. **Chemical changes** are changes in the ways the atoms and molecules in a pure substance are arranged and interconnected. To help get the clearest image, the chemicals are usually cooled as much as possible to slow down the particles.

People believed that atoms existed long before we were able to capture images of them. Two hundred years ago, John Dalton imagined that an atom could exist. With the invention of femtosecond laser technology, it is now possible to detect the movement of a single hydrogen atom (Figure 4.1B).

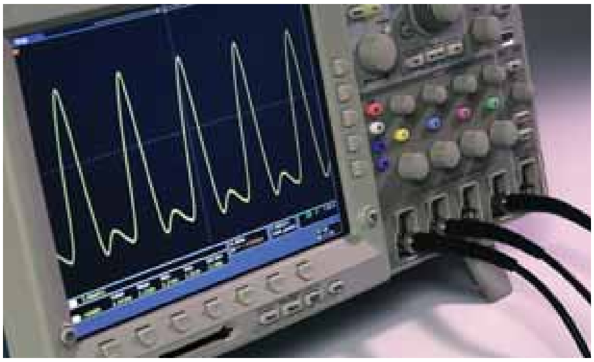


Figure 4.1B A digital oscilloscope measures the output and response time of the femtosecond laser.

4-1A Observing Chemical Changes

Find Out ACTIVITY

Teacher Demonstration

In this demonstration, you will observe chemical changes and reflect on some of the chemistry you learned in earlier science studies.

Safety

- Use a fume hood.
- Follow your teacher's directions regarding using open flames.
- The sparks from the burning steel wool are hot enough to cause burns.
- The light from burning magnesium is very bright and releases UV rays. Observe the light only from a distance and through glass, which absorbs UV.
- Persons with medical conditions that make them prone to seizure should not look directly at the light.
- Avoid touching all reactants and products.
- Wash your hands and equipment thoroughly after completing this activity.
- Do not remove any materials from the science room.

Materials

- fume hood
- steel wool
- 9.0 V dry cell
- magnesium ribbon
- large Pyrex® beaker
- tongs
- propane torch or Bunsen burner
- matches or flame striker
- hot pad

What to Do

1. Observe as your teacher uses the steel wool to make a short circuit between the terminals of the 9.0 V dry cell.
2. Record your observations.
3. Observe as your teacher ignites a small piece of magnesium ribbon inside a large Pyrex® beaker.
4. Record your observations.

What Did You Find Out?

1. Reflect on this activity in terms of laboratory safety. What safety issues are important to this activity?
2. Suggest which gas or gases in the air may have been responsible for the chemical changes that you saw.
3. New pure substances were formed in these chemical changes. Suggest what pure substance or substances formed in:
   (a) the first demonstration
   (b) the second demonstration

Atomic Theory

Subatomic particles are the particles that make up an atom. Through many experiments, scientists have found that individual atoms are composed of three subatomic particles, which are called protons, neutrons, and electrons (Figure 4.2). You can compare the size and mass of subatomic particles in Table 4.1.

- **Protons** are subatomic particles that have a 1+ (positive) electric charge.
- **Neutrons** are subatomic particles that do not have an electric charge.
- **Electrons** are subatomic particles that have a 1− (negative) electric charge.
- Protons and neutrons are held tightly together at the centre of the atom in a tiny region called the nucleus.
- Electrons exist in the region around the nucleus in regular patterns called shells or energy levels.
- Most of the volume of an atom is the region occupied by its electrons.
- Every neutral atom of the same element contains an equal number of protons and electrons. Since a proton counts as a 1+ electric charge and an electron counts as a 1− electric charge, the charges add up to zero, making an atom uncharged or neutral.

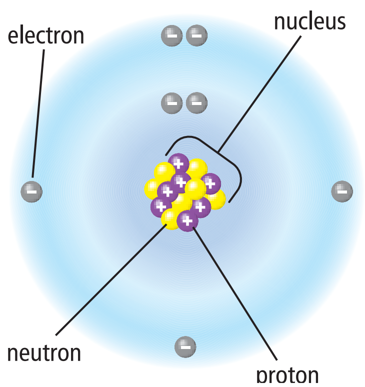


Figure 4.2 A model of the atom

Table 4.1 Subatomic Particles

Name	Symbol	Electric Charge	Location in the Atom	Relative Mass
Proton	p	1+	Nucleus	1836
Neutron	n	0	Nucleus	1837
Electron	e	1−	Surrounding the nucleus	1

The Nucleus

At the centre of each atom is a nucleus (Figure 4.2). The nucleus is tiny compared to the rest of the atom. Depending on the element, it would take between 10 000 and 100 000 nuclei lined up in a row to stretch across the diameter of one atom.

The simplest possible nucleus is found in the element hydrogen, which contains only one proton. A hydrogen atom may also have neutrons present in its nucleus. The nuclei of atoms of all other elements contain both protons and neutrons. For atoms with fewer than about 30 protons, the number of neutrons and the number of protons present in the nucleus are roughly equal. In heavier atoms, there are always more neutrons than protons. A stable nucleus is less likely to break apart from protons repulsing each other due to their positive charge. Extra neutrons help to make the nucleus stable by keeping the protons as far apart as possible. Very heavy atoms are unstable because the repulsion between the protons is so high.

Word Connect

The plural of nucleus is nuclei.

The electric charge on any nucleus is always positive, since the protons have a positive charge and the neutrons are not electrically charged. The **nuclear charge** is the term given to the electric charge on the nucleus, and it is simply found by counting the number of protons. For example, all carbon atoms possess six protons, so a carbon nucleus has a nuclear charge of 6+. The nuclear charge is the same as the **atomic number**. The number of protons and electrons in a neutral atom is equal.

$$\begin{array}{l}\text{atomic number}\\ \text{of a neutral atom}\end{array} = \text{number of protons} = \text{number of electrons}$$

The atomic number always identifies the element to which the atom belongs. For example, all atoms with an atomic number of 8 are oxygen atoms, and all oxygen atoms always have eight protons in the nucleus.

Did You Know?

If the nucleus of a carbon atom is represented by a loonie placed at centre ice, then the arena would represent the volume occupied by the electrons in their shells.

Reading Check

1. What is the electric charge on each of the three subatomic particles?
2. Which two subatomic particles have nearly equal mass?
3. Which particle determines the nuclear charge?
4. Which two subatomic particles are present in equal number in every atom of the same element?

Organization of the Periodic Table

Many facts about the elements are recorded in the periodic table of the elements (Figure 4.3). You may remember these points from earlier studies.

- Each element is listed according to its atomic number, left to right across each row and then row by row from top to bottom.
- Each row is also called a **period**. Each column (top to bottom) is called a **group** or **family**. For example, magnesium (Mg) is in group 2 of period 3.
- Metals are on the left side and in the middle of the table. Non-metals are in the upper right corner. The metalloids form a staircase toward the right side.
- Elements in the same chemical group or family have similar chemical properties. Four families you may have studied are:
  - the alkali metals (group 1 excluding hydrogen)—very reactive metals (e.g., sodium)
  - the alkaline earth metals (group 2)—somewhat reactive metals (e.g., calcium)
  - the halogens (group 17)—very reactive non-metals (e.g., chlorine)
  - the noble gases (group 18)—very unreactive gaseous non-metals (e.g., neon)
- The block of elements from group 3 through group 12 are collectively called the **transition metals**. They include familiar elements such as iron, nickel, copper, silver, and gold.

Periodic Table of the Elements

METALS ← → NON-METALS

Atomic Number	Symbol	Name	Atomic Mass	Ion charge(s)
22	Ti	Titanium	47.9	4+ 3+

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H Hydrogen 1.0 +																1 H Hydrogen 1.0 –	2 He Helium 4.0 0
3 Li Lithium 6.9 +	4 Be Beryllium 9.0 2+											5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0 3–	8 O Oxygen 16.0 2–	9 F Fluorine 19.0 –	10 Ne Neon 20.2 0
11 Na Sodium 23.0 +	12 Mg Magnesium 24.3 2+											13 Al Aluminum 27.0 3+	14 Si Silicon 28.1	15 P Phosphorus 31.0 3–	16 S Sulfur 32.1 2–	17 Cl Chlorine 35.5 –	18 Ar Argon 39.9 0
19 K Potassium 39.1 +	20 Ca Calcium 40.1 2+	21 Sc Scandium 45.0 3+	22 Ti Titanium 47.9 4+ 3+	23 V Vanadium 50.9 5+ 4+	24 Cr Chromium 52.0 3+ 2+	25 Mn Manganese 54.9 2+ 3+ 4+	26 Fe Iron 55.8 3+ 2+	27 Co Cobalt 58.9 2+ 3+	28 Ni Nickel 58.7 2+ 3+	29 Cu Copper 63.5 2+ 1+	30 Zn Zinc 65.4 2+	31 Ga Gallium 69.7 3+	32 Ge Germanium 72.6 4+	33 As Arsenic 74.9 3–	34 Se Selenium 79.0 2–	35 Br Bromine 79.9 –	36 Kr Krypton 83.8 0
37 Rb Rubidium 85.5 +	38 Sr Strontium 87.6 2+	39 Y Yttrium 88.9 3+	40 Zr Zirconium 91.2 4+	41 Nb Niobium 92.9 3+ 5+	42 Mo Molybdenum 95.9 2+ 3+	43 Tc Technetium (98) 7+	44 Ru Ruthenium 101.1 3+ 4+	45 Rh Rhodium 102.9 3+ 4+	46 Pd Palladium 106.4 2+ 4+	47 Ag Silver 107.9 1+	48 Cd Cadmium 112.4 2+	49 In Indium 114.8 3+	50 Sn Tin 118.7 4+ 2+	51 Sb Antimony 121.8 3+ 5+	52 Te Tellurium 127.6 2–	53 I Iodine 126.9 –	54 Xe Xenon 131.3 0
55 Cs Cesium 132.9 +	56 Ba Barium 137.3 2+	57 La Lanthanum 138.9 3+	72 Hf Hafnium 178.5 4+	73 Ta Tantalum 180.9 5+	74 W Tungsten 183.8 6+	75 Re Rhenium 186.2 4+ 7+	76 Os Osmium 190.2 3+ 4+	77 Ir Iridium 192.2 3+ 4+	78 Pt Platinum 195.1 4+ 2+	79 Au Gold 197.0 3+ 1+	80 Hg Mercury 200.6 2+ 1+	81 Tl Thallium 204.4 1+ 3+	82 Pb Lead 207.2 2+ 4+	83 Bi Bismuth 209.0 3+ 5+	84 Po Polonium (209) 2+ 4+	85 At Astatine (210) –	86 Rn Radon (222) 0
87 Fr Francium (223) +	88 Ra Radium (226) 2+	89 Ac Actinium (227) 3+	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (263)	107 Bh Bohrium (262)	108 Hs Hassium (265)	109 Mt Meitnerium (266)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Uub Ununbium (285)	113 Uut Ununtrium (284)	114 Uuq Ununquadium (289)	115 Uup Ununpentium (288)	116 Uuh Ununhexium (292)	117 Uus Ununseptium (?)	118 Uuo Ununoctium (294)
Alkali Metals	Alkaline Earth Metals															Halogens	Noble Gases

58 Ce Cerium 140.1 3+ 4+	59 Pr Praseodymium 140.9 3+ 4+	60 Nd Neodymium 144.2 3+	61 Pm Promethium (145) 3+	62 Sm Samarium 150.4 3+ 4+	63 Eu Europium 152.0 3+ 2+	64 Gd Gadolinium 157.3 3+	65 Tb Terbium 158.9 3+ 4+	66 Dy Dysprosium 162.5 3+	67 Ho Holmium 164.9 3+	68 Er Erbium 167.3 3+	69 Tm Thulium 168.9 3+ 2+	70 Yb Ytterbium 173.0 3+ 2+	71 Lu Lutetium 175.0 3+
90 Th Thorium 232.0 4+	91 Pa Protactinium 231.0 5+ 4+	92 U Uranium 238.0 6+ 4+ 5+	93 Np Neptunium (237) 5+ 3+ 4+ 6+	94 Pu Plutonium (244) 4+ 6+ 3+ 5+	95 Am Americium (243) 3+ 4+ 5+ 6+	96 Cm Curium (247) 3+	97 Bk Berkelium (247) 3+ 4+	98 Cf Californium (251) 3+	99 Es Einsteinium (252) 3+	100 Fm Fermium (257) 3+	101 Md Mendelevium (258) 2+ 3+	102 No Nobelium (259) 2+ 3+	103 Lr Lawrencium (262) 3+

Based on mass of C-12 at 12.00.

Any value in parentheses is the mass of the most stable or best known isotope for elements that do not occur naturally.

Figure 4.3 The periodic table of the elements

The Periodic Table and Ion Formation

When atoms gain or lose electrons, they become electrically charged particles called **ions**. Metal atoms, for example, lose electrons to form positively charged ions called *cations*. Many metals can form a cation only in one way. For example, aluminum forms a cation by losing three electrons to become Al^{3+}. The common ions formed by the elements are sometimes shown in the upper right-hand corner of the element's box in the periodic table (Figure 4.4A).

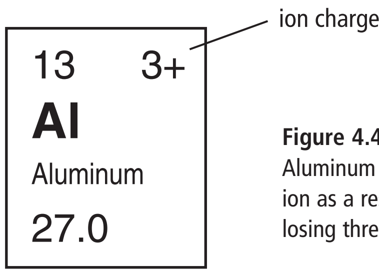


Figure 4.4A Aluminum forms a 3+ ion as a result of losing three electrons.

Some metals are **multivalent**, which means they can form ions in more than one way, depending on the chemical reaction they undergo. For example, iron is a multivalent element because it can lose two or three electrons to become Fe^{2+} ions and Fe^{3+} ions. Look at the periodic table to see which metals are multivalent. The most common charge is listed first at the top of the element's box in the periodic table (Figure 4.4B).

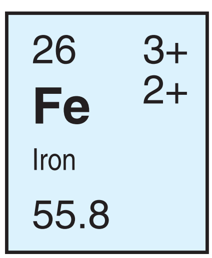


Figure 4.4B Iron is a multivalent element.

Many non-metals also form ions. However, since non-metal atoms, with very few exceptions, *gain* electrons, they form negative ions called *anions*. For example, the periodic table shows that chlorine will form a Cl^- ion (Figure 4.4C). This happens when a chlorine atom gains one electron.

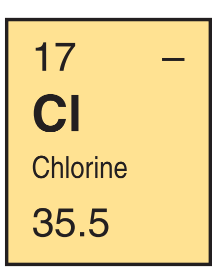


Figure 4.4C Chlorine forms a negative ion.

Reading Check

1. (a) What is another name for each row in the periodic table?
   (b) What is another name for each column or group in the periodic table?
2. Which group of metals is more reactive, the alkali metals or the alkaline earth metals?
3. List all five halogen elements.
4. Name the first and the last transition metal in period 4.

Bohr Diagrams

The periodic table represents patterns related to the arrangement of electrons in atoms. These patterns help explain how elements behave during a chemical change. For example, the periodic table can help you answer questions such as the following.

- What properties do elements in the same group (family) share?
- How can you predict the kinds of compounds that are likely to form?

The answer to both these questions comes from the electrons. Each shell can hold up to a certain number of electrons but no more. For example, the first shell can hold a maximum of two electrons. The second shell can hold a maximum of eight electrons. A **Bohr diagram** is a diagram that shows how many electrons are in each shell surrounding the nucleus. The Bohr diagram is named in honour of Niels Bohr (Figure 4.5), a Danish physicist who developed several models for showing the arrangement of electrons in atoms.

Figure 4.5 Niels Bohr (1885–1962) discovered that electrons were arranged in energy levels or shells in specific patterns. Bohr received a Nobel Prize in 1922 for his work on shells.

Figure 4.6 shows several types of Bohr diagrams, each representing an atom of potassium, which has 19 protons, 20 neutrons, and 19 electrons. Table 4.2 shows some examples of ion formations for lithium, aluminum, and chlorine.

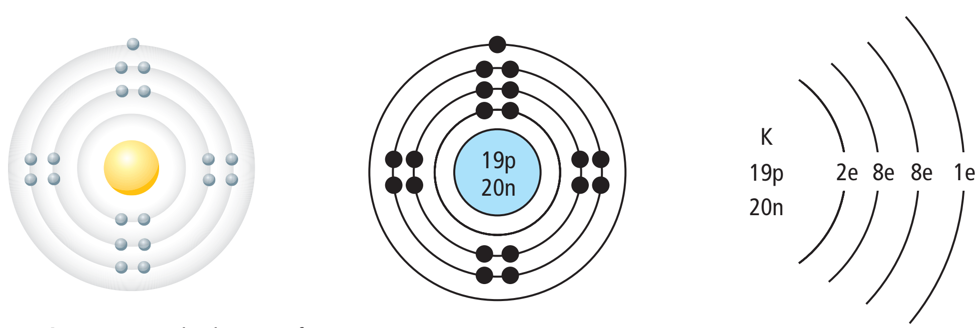


Figure 4.6 Bohr diagrams for potassium

Table 4.2 Ion Formations of Lithium, Aluminum, and Chlorine

	Lithium			Aluminum			Chlorine		
Atom	Li	3 p	2, 1	Al	13 p	2, 8, 3	Cl	17 p	2, 8, 7
Ion	Li^{+}	3 p	2	Al^{3+}	13 p	2, 8	Cl^{-}	17 p	2, 8, 8

Patterns of Electron Arrangement in Periods

The period number of an element equals the number of occupied shells of its atoms. Notice in Figure 4.7 that the two elements in period 1, hydrogen and helium, have a single occupied shell. The first shell of an atom can have a maximum of two electrons. Hydrogen has only one electron in its shell. Helium has a full set of two electrons in its shell.

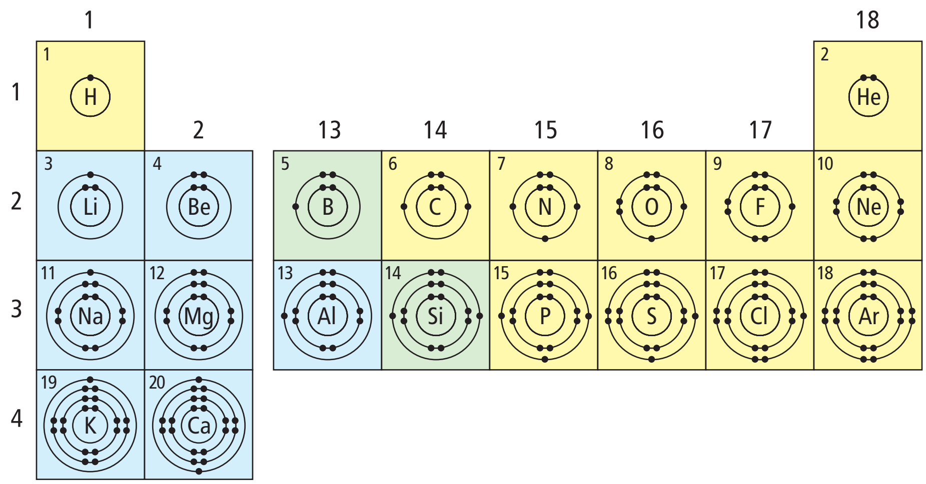


Figure 4.7 Occupied shells for individual atoms of the first 20 elements in the periodic table. Note that the diagrams do not represent the position or path of electrons.

The elements in period 2 all have two occupied shells. For each element in period 2, the first shell, which is closest to the nucleus, has a full set of two electrons. As you move from left to right across period 2, one more electron is added to the second shell of each atom. Notice that neon, the last element in period 2, has a full set of eight electrons in its second shell. The arrangement of eight electrons in the outermost shell is called a **stable octet**. "Octet" refers to a complete set of eight electrons.

Elements in period 3 have three occupied shells. The first two shells for each element in period 3 are full. As you move from left to right across period 3, one more electron is added to the third shell of each atom. Notice that argon has a stable octet in its outermost shell.

Patterns of Electron Arrangement in Groups

The outermost shell that contains electrons is called the **valence shell**. The electrons in the valence shell are called the **valence electrons**. Valence electrons are involved in chemical bonding. Examine the Group 1 elements in Figure 4.7 (lithium, sodium, and potassium). The atoms of each element in group 1, as well as hydrogen, have only one electron in their valence shell. Group 2 elements have two electrons in their valence shell. Group 13 elements have three electrons in their valence shell. Group 14 has four electrons, and so on through group 18. All group 18 elements have filled valence shells. Helium has two electrons filling its valence shell. Neon and argon each have eight electrons, or a stable octet, filling their valence shell. Figure 4.7 above shows that electrons can exist singly or in pairs. Electrons in completed shells (except for hydrogen) appear in pairs.

Did You Know?

Scientists were surprised to find that electrons could pair up inside atoms because all electrons have a negative charge and therefore should repel each other. The reason they can pair up is because of certain differences in the magnetic properties of each electron.

Practice Problems

Use the periodic table in Figure 4.3 on page 172 to help you answer the following questions.

1. Based on the patterns of the periodic table, identify the number of occupied shells for each of the following elements.
   (a) calcium, Ca
   (b) krypton, Kr
   (c) sulfur, S
   (d) iodine, I

2. Based on the patterns in the periodic table, identify the number of valence electrons for each of the following elements.
   (a) chlorine, Cl
   (b) magnesium, Mg
   (c) strontium, Sr
   (d) bromine, Br

Answers provided on page 591

Forming Compounds

When two atoms move close together, their valence electrons interact. A chemical bond forms between the atoms if the new arrangement of atoms and electrons is stable. The stability of an atom, ion, or compound is related to its energy; that is, lower energy states are more stable. The lowest energy is achieved when the atoms in the compound have the same arrangement of valence electrons as the arrangement for the noble gas to which they are closest in the periodic table.

When an atom forms a compound, it may acquire a valence shell like that of its closest noble gas in one of three ways.

- Atoms of metals may lose electrons to other atoms, forming **cations**.
- Atoms of non-metals may gain electrons from other atoms, forming **anions**.
- Atoms may share electrons.

Ionic bonding

Compounds are of two basic types, ionic and covalent. An ionic compound contains a positive ion (usually a metal) and a negative ion (usually a non-metal). In **ionic bonding**, one or more electrons transfers from each atom of the metal to each atom of the non-metal.

You can use Bohr diagrams to show ionic bonding for simple compounds. Calcium fluoride (CaF_2) has a ratio of ions of 1:2, as shown in Figure 4.8. In other words, each calcium fluoride has one calcium ion for every two fluoride ions.

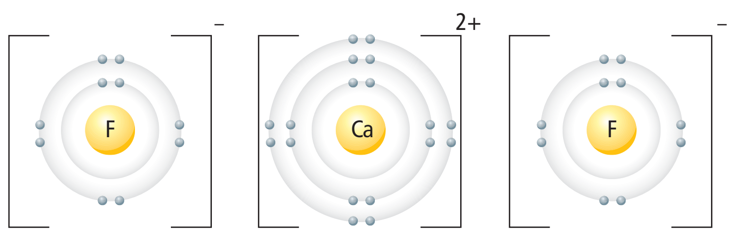


Figure 4.8 CaF_2 is represented using a Bohr diagram. Notice that each valence shell is full, making the element resemble the nearest noble gas.

The formation of the ionic compound sodium chloride is shown in Figure 4.9. Sodium chloride (NaCl) has a ratio of ions of 1:1. In other words, each sodium chloride is made up of one sodium ion and one chlorine ion. Notice that for sodium chloride, the valence shells are filled in both ions. Large square brackets are placed around the diagram with the ion charge shown just outside the end bracket.

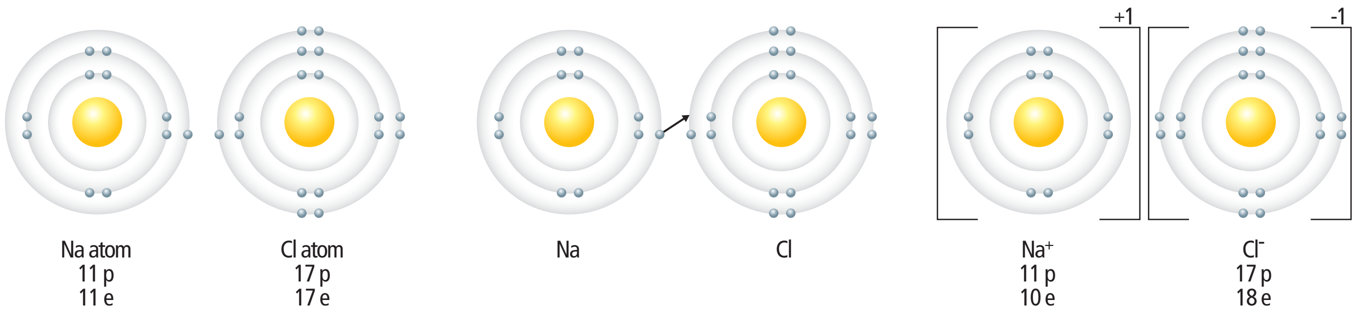


Figure 4.9 An ionic compound forms when an electron from a metal atom transfers to a non-metal atom, creating oppositely charged ions.

Covalent bonding

The atoms of many non-metals share electrons with other non-metal atoms. In **covalent bonding**, atoms overlap slightly, and one unpaired electron from each atom will pair together. Both atoms are attracted to the same pair of electrons, forming a covalent bond. A **covalent compound** is formed when non-metallic atoms share electrons to form covalent bonds. A covalent **molecule** is a group of atoms in which the atoms are bound together by sharing one or more pairs of electrons. The pair of electrons involved in a covalent bond are sometimes called the **bonding pair**. A pair of electrons in the valence shell that is not used in bonding is sometimes called a **lone pair**.

Bohr diagrams can be used to describe simple covalent compounds. Figure 4.10 shows the covalent compounds hydrogen fluoride (HF), water (H_2O), ammonia (NH_3), and methane (CH_4).

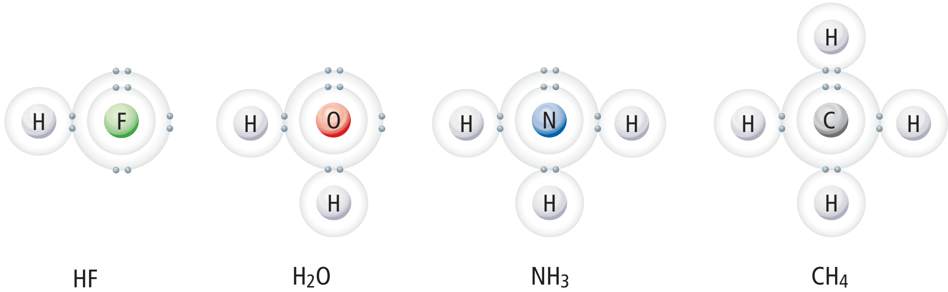


Figure 4.10 Bohr diagrams of HF, H_2O, NH_3, and CH_4

Word Connect

The prefix "co-" means sharing or together, and "–valent" refers to valence electrons. "Valence" is from a Latin word meaning strength or power.

Reading Check

1. How many valence electrons are there in an atom of carbon?
2. (a) How many electrons in phosphorus are paired?
   (b) How many are single?
3. How many electrons are there in an ion of S^{2-}?
4. What ion has 36 electrons and a charge of $1-$?

Lewis Diagrams

Gilbert Lewis (1875–1946) was a brilliant and influential American chemist who invented a concise method of showing bonding, particularly covalent bonding. A **Lewis diagram** is a diagram that illustrates chemical bonding by showing only an atom's valence electrons and the chemical symbol. Lewis diagrams are sometimes called Lewis structures or electron dot diagrams. Figure 4.11 compares Bohr diagrams and Lewis diagrams for the elements O, F, and Na.

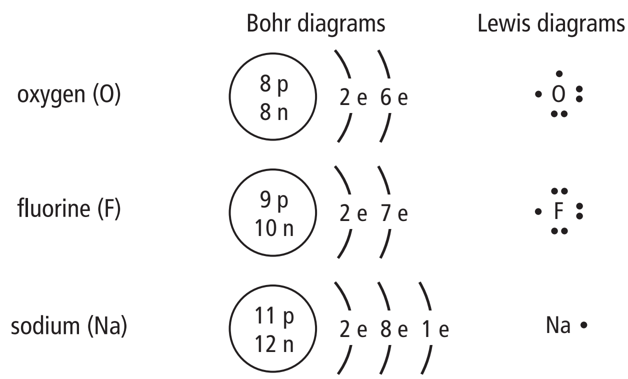


Figure 4.11 Comparison between Bohr diagrams and Lewis diagrams for atoms of oxygen, fluorine, and sodium

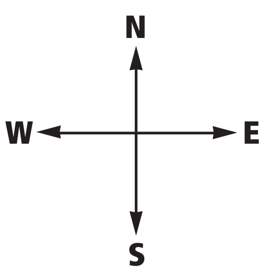


Figure 4.12A North, east, south and west are the four points of the compass.

You can follow these rules to draw a Lewis diagram:

- Dots representing electrons are placed around the element symbols at the points of the compass (Figure 4.12A).
- Electron dots are placed singly until the fifth electron is reached, then they are paired (Figure 4.12B).

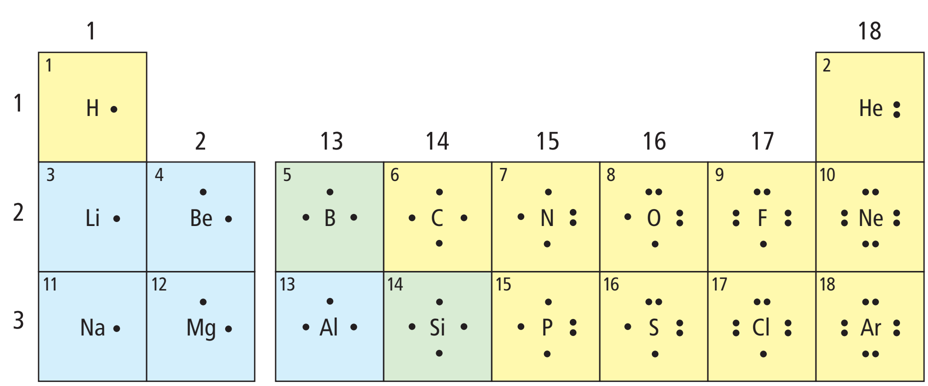


Figure 4.12B Lewis diagrams of the first 18 elements

Lewis diagrams of ions

You can follow these rules to draw ions in a Lewis diagram:

- For positive ions, one electron dot is removed from the valence shell for each positive charge of the ion. This usually means all the electron dots are removed. Only the element symbol remains encased in square brackets with a positive charge shown at the top right.
- For negative ions, one electron dot is added to each valence shell for each negative charge of the ion. This usually means the element's symbol is surrounded by eight electron dots (two electron dots for hydrogen). Square brackets are placed around this diagram with a negative charge shown at the top right (Figure 4.13).

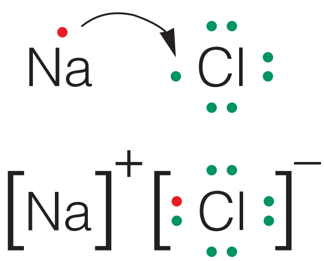
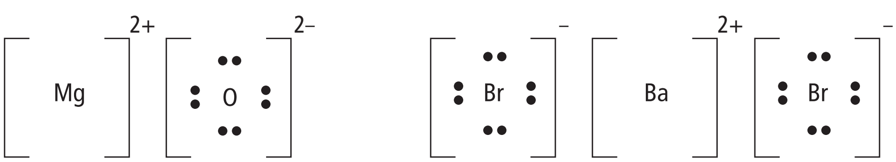


Figure 4.13 Lewis diagram for the formation of NaCl

Lewis diagrams of compounds

Lewis diagrams can be used to show ionic compounds, such as magnesium oxide and barium bromide (Figure 4.14).



MgO　　　　BaBr$_2$

Figure 4.14 Lewis diagram for MgO and BaBr$_2$. The lack of electron dots around Mg and Ba means that the previous shell is full. Another way to write it would be with eight dots around Mg and Ba, but that is not done because those eight electrons are not in the valence shell.

Lewis diagrams are also useful for showing covalent bonds. Figure 4.15 shows the covalent compound HF.

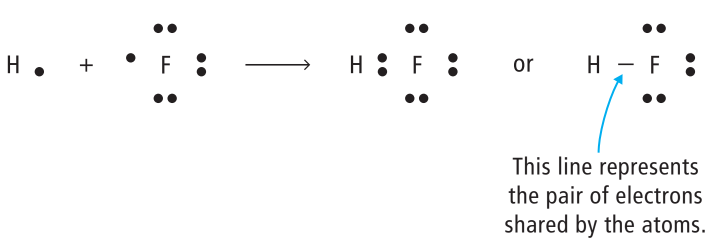


Figure 4.15 The electron from the hydrogen atom and the unpaired electron from the fluorine atom pair up in the HF molecule. Hydrogen has two paired electrons (in a full shell resembling the noble gas helium) and fluorine has four pairs of electrons (in a full shell resembling the noble gas neon).

Lewis diagrams of covalent molecules

The covalent molecules H_2O, NH_3, and CH_4 are shown below in Figure 4.16. Compare this illustration with the more complicated Bohr diagrams shown in Figure 4.10 on page 177.

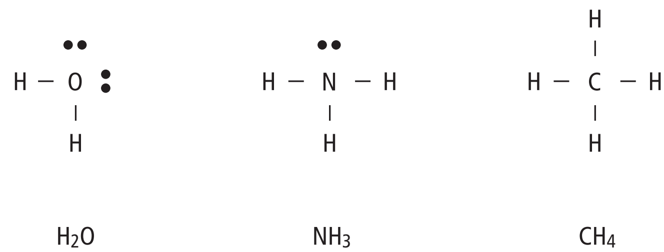


Figure 4.16 Lewis diagrams of H_2O, NH_3, and CH_4

Lewis diagrams of diatomic molecules

You can use Lewis diagrams to help explain why some of the non-metal elements exist as diatomic molecules. A **diatomic molecule** is a pair of atoms that are joined by covalent bonds. Diatomic elements form this way because the two-atom molecules are more stable than the individual atoms. For example, fluorine gas is diatomic. By joining together to form F_2, each fluorine atom can achieve a full valence shell of eight electrons (Figure 4.17). Other diatomic elements are hydrogen (H_2), nitrogen (N_2), oxygen (O_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

Suggested Activity

Think About It 4-1B on page 181

Explore More

Ozone, O_3, is a triatomic molecule that protects life on Earth from deadly ultraviolet radiation from the Sun. Try drawing a Lewis diagram for ozone. Then find out how it is produced in our atmosphere. Begin your search at **www.bcscience10.ca**.

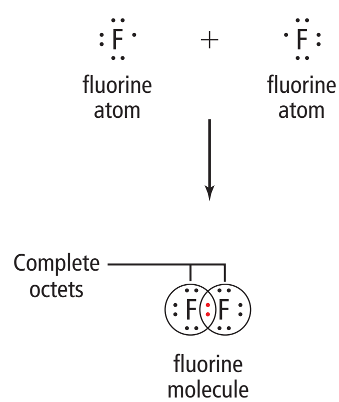


Figure 4.17 Two fluorine atoms share a pair of electrons to form a covalent bond. The shared electron pair gives each atom a complete octet.

4-1B Modelling Compounds

Bohr diagrams show all electrons present in an atom or ion, whereas Lewis diagrams show only valence electrons. In this activity, you will use Bohr diagrams and Lewis diagrams to describe bonding in compounds.

What to Do

Part 1 Bohr Diagrams

1. Copy and complete the following chart in your notebook. Refer to the periodic table in Figure 4.7 on page 172 to help you draw the Bohr diagrams.

Bohr Diagrams

Hydrogen	Lithium	Magnesium	Oxygen	Chlorine	Fluorine

2. Draw a Bohr model of Li_2O following these instructions.
   - Use a pencil to draw a Bohr diagram for an oxygen atom.
   - Draw two diagrams of lithium atoms, one to the left and one to the right of the oxygen atom. Notice that each lithium atom has one valence electron, and that the outer shell of the oxygen atom has room for two more valence electrons.
   - Erase the valence electrons from each of the lithium atoms and redraw them in the oxygen atom's valence shell.
   - Place a square box around each of the element symbols, and write 1^+ beside each lithium atom and 2^- beside the oxygen atom.
3. Follow the steps above to draw Bohr diagrams for these ionic compounds.
   (a) LiCl
   (b) MgO
   (c) $MgCl_2$
4. Follow the steps above to draw Bohr diagrams for these covalent compounds.
   (a) HF
   (b) H_2O
   (c) OF_2

Part 2 Lewis Diagrams

5. Copy and complete the following chart in your notebook. Refer to the periodic table in Figure 4.12B on page 178 to help you draw the Lewis diagrams. Find the number of valence electrons in each atom and arrange them as follows.
   - Dots representing electrons are placed around the element symbols at the points of the compass (north, east, south, and west)
   - Electron dots are placed singly until the fifth electron is reached, then they are paired. For example, oxygen has six electrons in its valence shell, so the fifth and sixth dots are each paired with one of the other four dots.

Lewis Diagrams

Hydrogen	Lithium	Magnesium	Oxygen	Chlorine	Fluorine

6. Draw Lewis diagrams for these ionic compounds.
   (a) Li_2O
   (b) LiCl
   (c) MgO
   (d) $MgCl_2$
7. Draw Lewis diagrams for these covalent compounds.
   (a) HF
   (b) H_2O
   (c) OF_2

What Did You Find Out?

1. Describe the information contained in a Bohr diagram compared with the information contained in a Lewis diagram.
2. (a) Which diagram do you find easier to use, a Bohr diagram or a Lewis diagram?
   (b) Why?

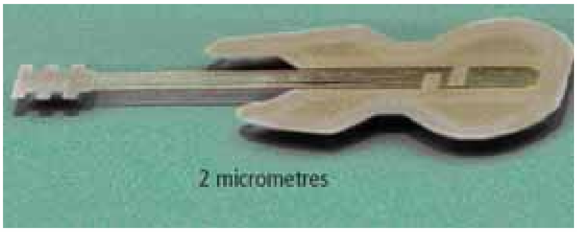
Wild, Weird, Wonderful

www Science

The Art of Science

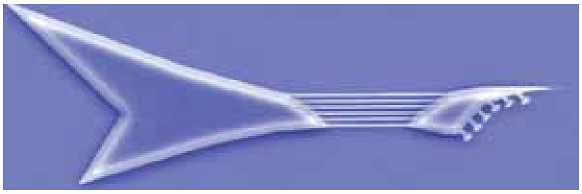
When you imagine a scientist at work, do you also picture an artist? Creativity is an essential part of both art and science. New technologies often open the door for new kinds of artistic experimentation and expression. For example, advances in electronics have changed the way we make and listen to music. Now, developments in nanotechnology are blending science, technology, and art at the microscopic scale. Nanotechnology takes its name from the nanometre, which is an extremely small unit of measurement equalling one billionth of a metre.

The tiny guitar shown below is a replica of a Fender Stratocaster and was made in a university laboratory in 1997. It was cut from a single crystal of silicon dioxide, a major component of glass and concrete. With a length of 10 micrometres, it is the world's smallest guitar. To get an idea of how small it is, picture this: five of these guitars could fit across the width of one human hair. Although it cannot be used to make music, it pushes the boundaries in the production of tiny devices.



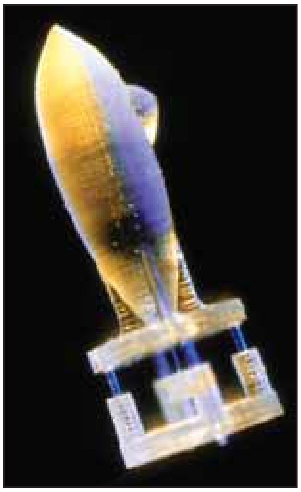
Nano guitar

In 2003, a slightly larger version of the guitar was produced, this time with working "strings." Although it is too small to be played by any mechanical device, the strings can be vibrated using lasers. The frequency of the sound is too high to be heard by humans. However, the same technology may be used to help construct timing devices used in cellphones.



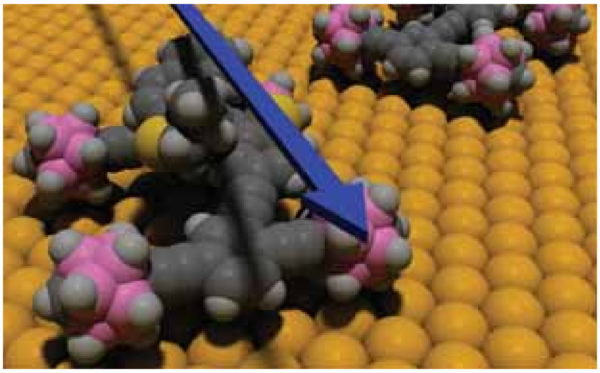
Nano guitar with strings

A tiny functioning submarine was recently produced by shining lasers through a liquid material containing acrylic, a kind of plastic. Computer-controlled lasers caused certain parts of the material to become solid while the surrounding parts remained liquid. At 4 mm in length, this submarine is large enough to be seen by the human eye. This kind of technology might one day lead to developing devices small enough to be used inside the body to deliver medicines or repair damaged tissues.



Nano submarine

There are plans to build machines made from individual atoms, including motors, wheels, and axles. Recently a "nano car" was designed and built from carbon atoms and driven on a "road" made of gold atoms, but it had no engine. It was pushed around using a tiny needle, which was only a few atoms wide at its tip.



Nano car

What will the future bring? Possible applications of nanotechnology in medicine and research seem endless. What issues do we need to consider in applying nano-technology?

Check Your Understanding

Checking Concepts

1. (a) What is one property that protons and neutrons have in common?
(b) What is one property that is different for protons and neutrons?

2. Which two subatomic particles are nearly equal in mass?

3. Which subatomic particle is nearly equal to the masses of the other two subatomic particles added together?

4. A bucket full of water has both mass and volume. Referring to the subatomic particles, explain what accounts for most of the:
(a) mass of the water
(b) volume of the water

5. Explain how an atom is composed of charged particles yet can have an overall charge of zero.

6. (a) What is the value of the nuclear charge on a neon atom?
(b) How is the nuclear charge determined?

7. Copy and complete the following chart in your notebook.

	Element	Atomic Number	Number of Protons	Number of Electrons
(a)	Pb	82		
(b)			8	
(c)				30
(d)	Fe			
(e)		47		
(f)				17

8. For each of Cs, S, Kr, C, Fe, and Hg, name its:
(a) period
(b) group

9. List four chemical family names, working from left to right across the periodic table.

Understanding Key Ideas

10. Name the subatomic particle(s) that best fit each of the following descriptions.
(a) has a negative charge
(b) has an electric charge
(c) surrounds the nucleus in a regular pattern
(d) has an electric charge of zero
(e) is present in the nucleus
(f) The number of this particle is always the same as the atomic number.

11. How is a covalent compound different from an ionic compound?

12. Compare a Bohr diagram and a Lewis diagram. Explain how they are:
(a) similar
(b) different

13. Draw Bohr diagrams for:
(a) diatomic molecules H_2 and F_2
(b) covalent compounds H_2O and HCl
(c) ionic compounds KF and Li_2O

14. Draw Lewis diagrams for:
(a) diatomic molecules H_2 and F_2
(b) covalent compounds H_2O and HCl
(c) ionic compounds NaF, $BeCl_2$ and Li_2O

Pause and Reflect

Think back over the information you have learned about atoms in this section. Illustrate and explain your understanding of the current model for the atom. How have your ideas changed from your earlier understanding of the atom?