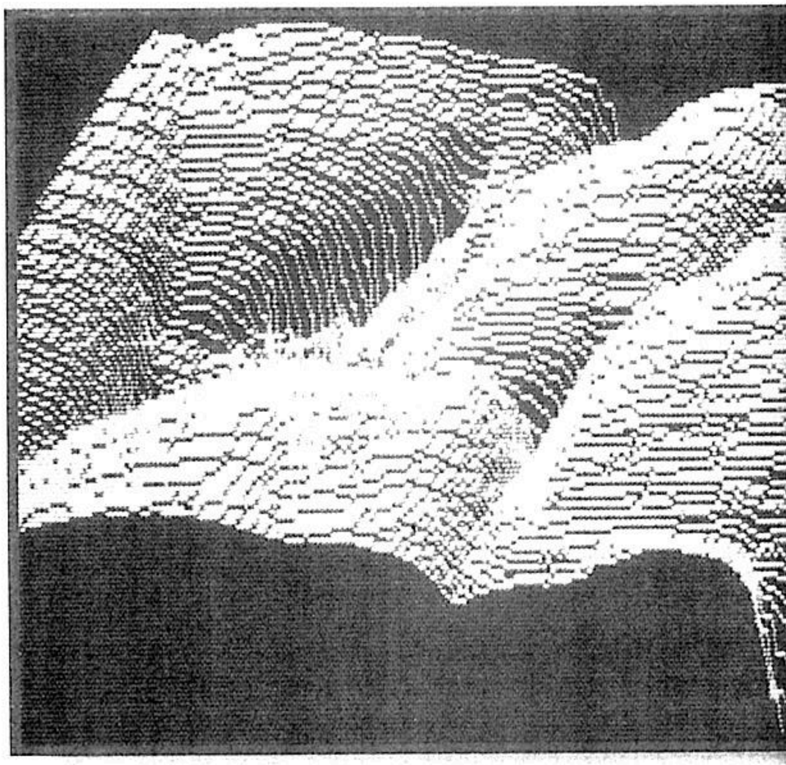
Chapter 5

ATOMIC STRUCTURE AND THE PERIODIC TABLE



Each bright spot on this image represents one gold atom.

FEATURES

DISCOVER IT! ELECTRIC CHARGE

You need four 25-cm lengths of clear plastic tape and a metric ruler.

1. Firmly stick two of the 25-cm pieces of tape side-by-side, about 10 cm apart, on your desk top. Leave 2 to 3 cm of tape sticking over the edge of the desk. Grasp the free ends of the tapes and pull sharply upward to peel the tape pieces off of the desk. Slowly bring the pieces, which have similar charges, toward one another. What do you observe?
2. Pull the third and fourth pieces of tape between your thumb and forefinger several times, as if you were trying to clean each one. Slowly bring these two pieces of tape, which now have similar charges, toward one another. What do you observe?
3. Predict what might happen if you brought a piece of tape pulled from your desk top close to a piece of tape pulled between your fingers. Try it! What happens?

Do you think the pieces of tape used in Step 1 have the same charge as the pieces used in Step 2? Explain. After reading about charged particles in this chapter, return to this activity and re-evaluate your answer.

ATOMS

In 1981, Swiss scientists Gerd Binnig and Heinrich Rohrer completed the construction of the scanning tunneling microscope. Their device, which they first tested using a specimen of gold, produces an image of individual atoms, often seen as rows of bright spots on a monitor as seen here. The scanning tunneling microscope is used today to study how atoms are arranged on the surface of many different materials. Binnig and Rohrer won the Nobel Prize for Physics in 1986 for their invention. **Why was an image of individual atoms considered such an important breakthrough?**

objectives

- Summarize Dalton's atomic theory
- Describe the size of an atom

key terms

- Dalton's atomic theory
- atom

Early Models of the Atom

Have you ever been asked to believe in something you could not see? Using your unaided eyes, you cannot see the tiny fundamental particles that make up matter. Yet all matter is composed of such particles. Democritus of Abdera, a teacher who lived in Greece during the fourth century B.C., first suggested the existence of these particles, which he called atoms. He believed that these atoms were indivisible and indestructible. Although Democritus's ideas agreed with later scientific theory, they were not useful in explaining chemical behavior. They also lacked experimental support because scientific testing was unknown at that time.

The real nature of atoms and the connection between observable changes and events at the atomic level were not established for more than 2000 years after Democritus. The modern process of discovery regarding atoms began with John Dalton (1766–1844), an English schoolteacher. Unlike Democritus, Dalton performed experiments to test and correct his atomic theory. Dalton studied the ratios in which elements combine in chemical reactions. Based on the results of his experiments, Dalton formulated hypotheses and theories to explain his observations. The result was **Dalton's atomic theory**, which includes the ideas illustrated in **Figure 5.1** and listed below.

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of the same element are identical. The atoms of any one element are different from those of any other element.

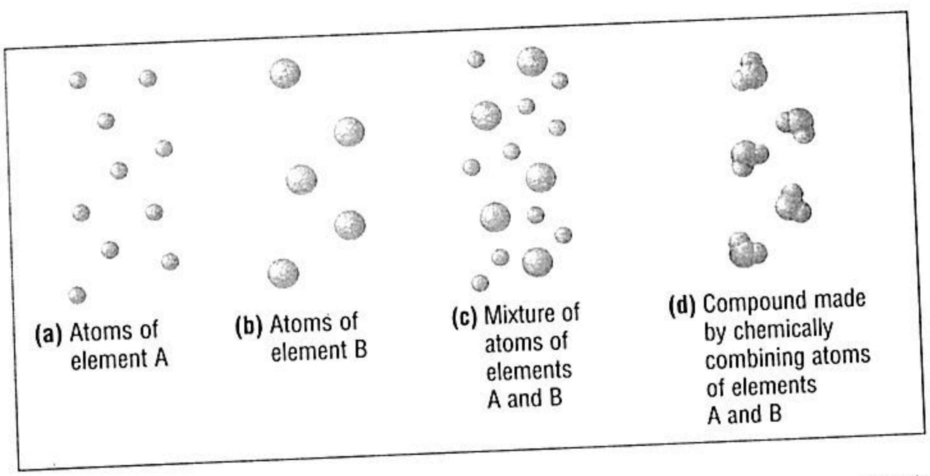


Figure 5.1
According to Dalton's atomic theory, an element is composed of only one kind of atom, and a compound is composed of particles that are chemical combinations of different kinds of atoms.

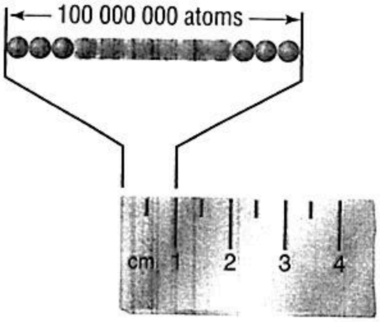


Figure 5.2
If 100 000 000 copper atoms were placed side by side, they would form a line 1 cm long. What is this number of atoms written in scientific notation?

3. Atoms of different elements can physically mix together or can chemically combine with one another in simple whole-number ratios to form compounds.
4. Chemical reactions occur when atoms are separated, joined, or rearranged. Atoms of one element, however, are never changed into atoms of another element as a result of a chemical reaction.

Just How Small Is an Atom?

A coin the size of a penny and composed of pure copper (Cu) illustrates Dalton's concept of the atom. Imagine grinding the copper coin into a fine dust. Each speck in the small pile of shiny red dust would still have the properties of copper. If by some means you could continue to make the copper dust particles smaller, you would eventually come upon a particle of copper that could no longer be divided and still have the properties of copper. This final particle is called an **atom,** defined in its modern sense as the smallest particle of an element that retains the properties of that element.

Copper atoms are very small. A pure copper coin the size of a penny contains about 2.4×10^{22} atoms. By comparison, Earth's population is only about 6×10^{9} people. There are about 4×10^{12} as many atoms in the coin as there are people on Earth. If you could line up 100 000 000 copper atoms side by side, they would produce a line only 1 cm long, as shown in **Figure 5.2.**

Does seeing individual atoms seem impossible? Despite their small size, individual atoms are observable with the proper instrument. As you read in the introduction to this section, and as **Figure 5.3** shows, a scanning tunneling microscope provides a visual image of individual atoms. Individual atoms can even be moved around and arranged in patterns. The ability to move individual atoms holds future promise for the creation of atomic-sized electronic devices, such as circuits and computer chips. This atomic-scale technology could someday be applied to communications and space exploration.

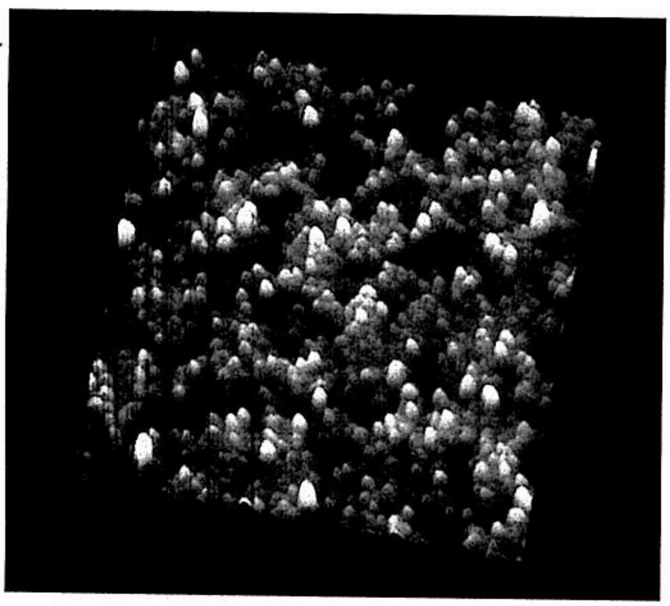


Figure 5.3
A scanning tunneling microscope can be used to view the surface of individual atoms, such as the gold atoms shown here.

section review 5.1

1. In your own words, state the main ideas of Dalton's atomic theory.
2. Characterize the size of an atom.
3. Democritus and Dalton both proposed that matter consists of atoms. How did their approaches to reaching that conclusion differ?

Chem ASAP! **Assessment 5.1** Check your understanding of the important ideas and concepts in Section 5.1.

STRUCTURE OF THE NUCLEAR ATOM

Atoms are so small they can only be visualized with a scanning tunneling microscope. How do scientists study the even smaller particles that make up atoms? Scientists study the makeup of atoms by breaking them apart! The atoms are accelerated to tremendous speeds—nearly the speed of light—in giant devices called particle accelerators. Then the atoms are smashed into one another, causing the particles within them to be released. Although scientists cannot actually see the particles, they can use the device shown here, called a bubble chamber, to see the tracks these particles make. **What are the component particles that scientists have discovered within atoms?**

objectives

- Distinguish among protons, electrons, and neutrons in terms of relative mass and charge
- Describe the structure of an atom, including the location of the protons, electrons, and neutrons with respect to the nucleus

key terms

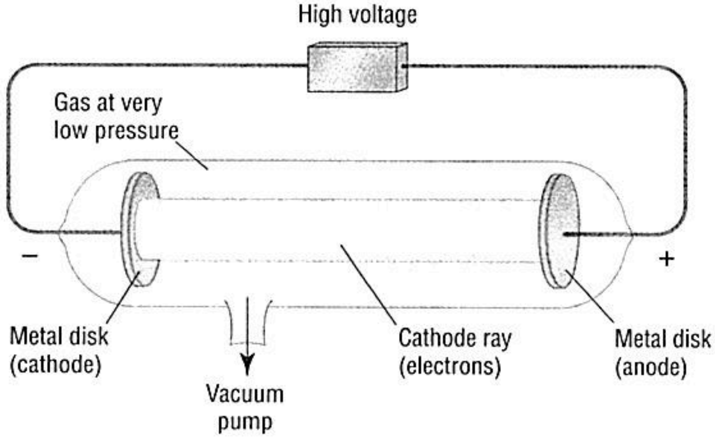
- electrons
- cathode ray
- protons
- neutrons
- nucleus

Electrons

Much of Dalton's atomic theory is accepted today. One important change, however, is that atoms are now known to be divisible. They can be broken down into even smaller, more fundamental particles. Dozens of kinds of subatomic particles are unleashed when powerful devices known as atom smashers are used to fracture atoms. You will now learn about three kinds of subatomic particles.

Electrons are negatively charged subatomic particles. The English physicist J. J. Thomson (1856–1940) discovered electrons in 1897. Thomson performed experiments that involved passing electric current through gases at low pressure. He sealed the gases in glass tubes fitted at both ends with metal disks called electrodes. **Figure 5.4** shows the kind of apparatus he used. The electrodes were connected to a source of high-voltage electricity. One electrode, the anode, became positively charged. The other electrode, the cathode, became negatively charged. A glowing beam formed between the electrodes. This beam, which traveled from the cathode to the anode, is called a **cathode ray.**

Thomson found that cathode rays are attracted to metal plates that have a positive electrical charge. Plates that carry a negative electrical charge repel the rays. **Figure 5.5** on page 110 shows the deflection of cathode rays. Thomson knew that opposite charges attract and like charges repel, so he proposed that a cathode ray is a stream of tiny negatively charged



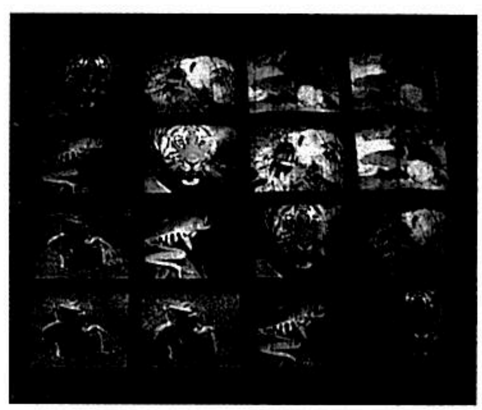
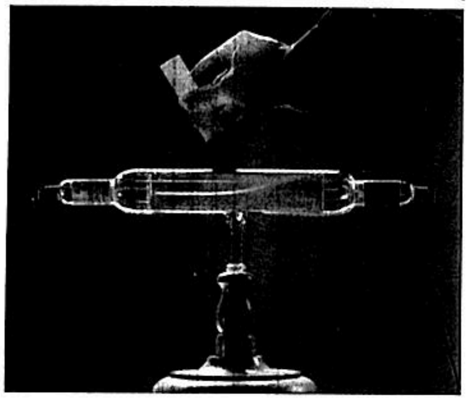


Figure 5.4
In a cathode-ray tube, electrons travel as a ray from the cathode (−) to the anode (+). A television tube is a specialized type of cathode-ray tube.



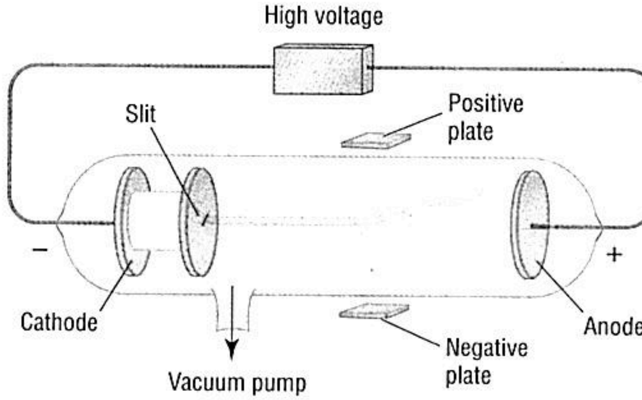


Figure 5.5
Cathode rays are deflected by a magnet and attracted by a positively charged plate. This shows that the particles that make up the rays are negatively charged.

particles moving at high speed. Thomson named these particles electrons. Thomson also showed that the production of cathode rays did not depend on the kind of gas in the cathode-ray tube or the type of metal used for the electrodes. He concluded that electrons must be parts of the atoms of all elements. By 1900, Thomson and others had determined that an electron's mass is about 1/2000 the mass of a hydrogen atom.

The American scientist Robert A. Millikan (1868–1953) carried out experiments that allowed him to find the quantity of charge carried by an electron. He also determined the ratio of the charge to the mass of an electron. Millikan used these two values to calculate an accurate value for the mass of the electron. Millikan's values for electron charge and mass, reported in 1916, are very similar to those accepted today. An electron carries exactly one unit of negative charge, and its mass is 1/1840 the mass of a hydrogen atom.

Protons and Neutrons

If cathode rays are electrons given off by atoms, what remains of the atoms that have lost the electrons? For example, after a hydrogen atom (the lightest kind of atom) loses an electron, what is left? You can think through this problem using four simple ideas about matter and electric charges. First, atoms have no net electric charge; they are electrically neutral. (One important piece of evidence for electrical neutrality is that you do not receive an electric shock every time you touch something!) Second, electric charges are carried by particles of matter. Third, electric charges always exist in whole-number multiples of a single basic unit; that is, there are no fractions of charges. Fourth, when a given number of negatively charged particles combines with an equal number of positively charged particles, an electrically neutral particle is formed.

Considering all of this information, it follows that a particle with one unit of positive charge should remain when a typical hydrogen atom loses an electron. Evidence for such a positively charged particle was found in 1886, when E. Goldstein observed a cathode-ray tube and found rays traveling in the direction opposite to that of the cathode rays. He called these rays canal rays and concluded that they were composed of positive particles. Such positively charged subatomic particles are called **protons.** Each proton has a mass about 1840 times that of an electron.

In 1932, the English physicist James Chadwick (1891–1974) confirmed the existence of yet another subatomic particle: the neutron. **Neutrons** are

Table 5.1

Properties of Subatomic Particles				
Particle	Symbol	Relative electrical charge	Relative mass (mass of proton = 1)	Actual mass (g)
Electron	e^-	1−	1/1840	9.11×10^{-28}
Proton	p^+	1+	1	1.67×10^{-24}
Neutron	n^0	0	1	1.67×10^{-24}

subatomic particles with no charge but with a mass nearly equal to that of a proton. Thus the fundamental building blocks of atoms are the electron, the proton, and the neutron. **Table 5.1** summarizes the properties of these subatomic particles.

The Atomic Nucleus

When subatomic particles were discovered, scientists wondered how these particles were put together in an atom. This was a difficult question to answer, given how tiny atoms are. Most scientists thought it likely that the electrons were evenly distributed throughout an atom filled uniformly with positively charged material. In 1911, Ernest Rutherford (1871–1937) and his coworkers at the University of Manchester, England, decided to test this theory of atomic structure. Their test used relatively massive alpha particles, which are helium atoms that have lost their two electrons and have a double positive charge because of the two remaining protons. In the experiment, illustrated in **Figure 5.6,** Rutherford directed a narrow beam of alpha particles at a very thin sheet of gold foil. According to the prevailing theory, the alpha particles should have passed easily through the gold, with only a slight deflection due to the positive charge thought to be spread out in the gold atoms.

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Animation 4
Take a look at Rutherford's gold-foil experiment, its results, and its conclusions.

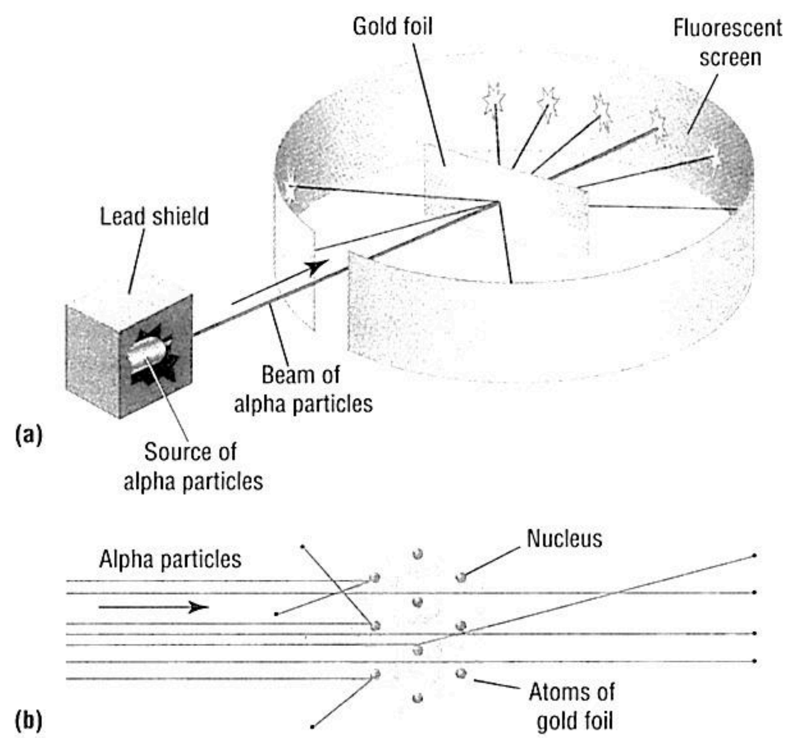


Figure 5.6
(a) *To learn more about the nature of the atom, Rutherford and his coworkers aimed a beam of alpha particles at a sheet of gold foil surrounded by a fluorescent screen. They found that most of the particles passed through the foil with no deflection at all. A few particles were greatly deflected.*
(b) *Rutherford concluded that most of the alpha particles pass through the gold foil because the atom is mostly empty space. The mass and positive charge are concentrated in a small region of the atom. Rutherford called this region the nucleus. Particles that approach the nucleus closely are greatly deflected.*

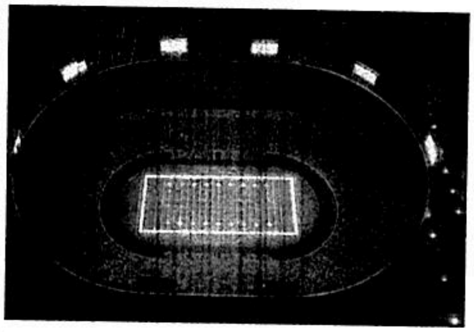


Figure 5.7
If an atom were the size of this stadium, then its nucleus would be about the size of a marble!

To everyone's surprise, the great majority of alpha particles passed straight through the gold atoms, without deflection. Even more surprisingly, a small fraction of the alpha particles bounced off the gold foil at very large angles. Some even bounced straight back toward the source. Rutherford later recollected, "It was about as credible as if you had fired a 15-inch shell at a piece of tissue paper, and it came back and hit you."

Based on the experimental results, Rutherford suggested a new theory of the atom. He proposed that the atom is mostly empty space, thus explaining the lack of deflection of most of the alpha particles. He concluded that all the positive charge and almost all the mass are concentrated in a small region that has enough positive charge to account for the great deflection of some of the alpha particles. He called this region the nucleus. The **nucleus** is the central core of an atom and is composed of protons and neutrons. It is tiny compared with the atom as a whole. See **Figure 5.7.**

MINI LAB

Using Inference: The Black Box

PURPOSE

To determine the shape of a fixed object inside a sealed box without opening the box.

MATERIALS

- box containing a regularly shaped object fixed in place and a loose marble

PROCEDURE

1. Do not open the box.
2. Carefully manipulate the box so that the marble moves around the fixed object.
3. Gather data (clues) that describe the movement of the marble.
4. Sketch a picture of the object in the box, showing its shape, size, and location within the box.
5. Repeat this activity with a different box containing a different object.

ANALYSIS AND CONCLUSIONS

1. Find a classmate who had the same lettered box that you had. Compare your findings and try to come to agreement about the shape and location of the fixed object.
2. What experiment that contributed to a better understanding of the atom does this activity remind you of?

section review 5.2

4. What are the charges and relative masses of the three main subatomic particles?
5. Describe the basic structure of an atom.
6. Describe Thomson's, Millikan's, and Rutherford's contributions to atomic theory. Include their experiments if appropriate.

Chem ASAP! Assessment 5.2 Check your understanding of the important ideas and concepts in Section 5.2.

DISTINGUISHING BETWEEN ATOMS

section 5.3

"Rose is a rose is a rose is a rose," wrote the famous author Gertrude Stein. Of course, this is not true when it comes to the color of roses, which can range from the familiar red or white to yellow or even lavender. In all, there are more than 13 000 colorful varieties of roses. Just as roses come in different varieties,a given chemical element can come in different "varieties" called isotopes.
What is an isotope, and how does one isotope of an element differ from another?

objectives

- Explain how the atomic number identifies an element
- Use the atomic number and mass number of an element to find the numbers of protons, electrons, and neutrons
- Explain how isotopes differ and why the atomic masses of elements are not whole numbers
- Calculate the average atomic mass of an element from isotope data

key terms

- atomic number
- mass number
- isotopes
- atomic mass unit (amu)
- atomic mass

Atomic Number

Atoms are composed of electrons, protons, and neutrons. Protons and neutrons make up the small, dense nucleus. Electrons surround the nucleus and occupy most of the volume of the atom. How, then, are atoms of hydrogen, for example, different from atoms of oxygen? Examine **Table 5.2.** Compare the entries for hydrogen and oxygen. You should notice that a hydrogen atom has one proton in its nucleus, but an oxygen atom has eight protons in its nucleus. Elements are different because they contain different numbers of protons.

The **atomic number** of an element is the number of protons in the nucleus of an atom of that element. Because all hydrogen atoms have one proton, the atomic number of hydrogen is 1. Similarly, because all oxygen atoms have eight protons, the atomic number of oxygen is 8. The atomic number identifies an element.

Look again at **Table 5.2.** For each element listed, the number of protons equals the number of electrons. Remember that atoms are electrically neutral. Thus the number of electrons (negatively charged particles) in an atom must equal the number of protons (positively charged particles) in the nucleus. A hydrogen atom has one electron, and an oxygen atom has eight electrons. How is the number of electrons for a neutral atom of a given element related to the atomic number of that element?

Table 5.2

Atoms of the First Ten Elements

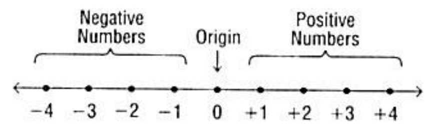
Name	Symbol	Atomic number	Composition of the nucleus: Protons	Composition of the nucleus: Neutrons*	Mass number	Number of electrons
Hydrogen	H	1	1	0	1	1
Helium	He	2	2	2	4	2
Lithium	Li	3	3	4	7	3
Beryllium	Be	4	4	5	9	4
Boron	B	5	5	6	11	5
Carbon	C	6	6	6	12	6
Nitrogen	N	7	7	7	14	7
Oxygen	O	8	8	8	16	8
Fluorine	F	9	9	10	19	9
Neon	Ne	10	10	10	20	10

* Number of neutrons in the most abundant isotope. Isotopes are introduced later in Section 5.3.

CHEMath

USING POSITIVE AND NEGATIVE NUMBERS

In this chapter you are introduced to protons and electrons, particles with charges of 1+ and 1− respectively. In addition to these charges, you will encounter a variety of positive and negative numbers throughout this course. The table below offers a review of basic mathematical operations related to positive and negative numbers.

BASIC CONCEPTS	**EXAMPLE**
Positive and negative numbers are used in many applications, including temperature scales. On a number line, the positive numbers are to the right of the *origin* (zero), and the negative numbers are to the left of the origin.	
The *absolute value* of any number is its distance from the origin. Absolute values are always positive (or zero).	$\lvert 3 \rvert$ = (absolute value of 3) = 3 $\lvert -2 \rvert$ = (absolute value of −2) = 2
The *opposite* of any number is a number having the same absolute value but the opposite sign.	The opposite of 8 is −8. The opposite of −5 is −(−5), or 5.
ADDITION	
If the two numbers are both positive, add them as usual.	$5 + 16 = 21$
If the two numbers are both negative, add their absolute values and take the opposite.	To find $-6 + (-5)$, note that $6 + 5 = 11$. Therefore, $-6 + (-5) = -11$.
It the two numbers have opposite signs, subtract the smaller absolute value from the larger absolute value. Use the sign of the number with the larger absolute value.	To find $6 + (-3)$: $6 - 3 = 3$, so $6 + (-3) = 3$. To find $6 + (-8)$: $8 - 6 = 2$, so $6 + (-8) = -2$. To find $(-6) + 3$: $6 - 3 = 3$, so $(-6) + 3 = -3$. To find $(-6) + 8$: $8 - 6 = 2$, so $(-6) + 8 = 2$.
SUBTRACTION	
Subtracting any number is the same as adding its opposite.	$12 - 23 = 12 + (-23) = -11$ $6 - (-4) = 6 + 4 = 10$
MULTIPLICATION AND DIVISION	
If the two numbers have the same sign (both positive or both negative), the product or quotient is positive.	$3 \times 5 = 15$ $-24 \div (-8) = 3$
If the two numbers have opposite signs, the product or quotient is negative.	$4 \times (-3) = -12$ $-20 \div 5 = -4$
CHEMISTRY CONNECTION	
An *ion* is a charged atom. The net charge of an ion is (number of protons) − (number of electrons).	If an ion has 3 protons and 5 electrons, its net charge is $3 - 5 = -2$.

Practice Problems

Practice using positive and negative numbers by performing each calculation.

A. $12 + (-4)$

B. $-5 - (-10)$

C. $4 \times (-10)$

D. $-21 \div (-7)$

E. $-36 + 17$

F. $8.7 - 11.3$

G. -3.1×0.20

H. $22.5 \div (-3.60)$

I. $-24.8 - (-4.7)$

J. $-5.5 \times (-1.8)$

K. Find the net charge of an ion with 15 protons and 18 electrons.

L. Find the net charge of an ion with 9 protons and 8 electrons.

M. An ion with a charge of 3− loses 2 electrons. What is its charge?

N. An ion with a charge of 1+ gains 3 electrons. What is its charge?

Sample Problem 5-1

The element nitrogen (N) has an atomic number of 7. How many protons and how many electrons are in a neutral nitrogen atom?

1. *ANALYZE* ***List the known and the unknowns.***

Known:
- atomic number = 7

Unknowns:
- number of protons = ?
- number of electrons = ?

The atomic number gives the number of protons, which in a neutral atom equals the number of electrons.

atomic number = number of protons = number of electrons

2. *CALCULATE* ***Solve for the unknowns.***

atomic number = 7 = number of protons
= number of electrons

3. *EVALUATE* ***Do the results make sense?***

The relationships have been applied correctly. The seven electrons are needed to balance the seven protons.

Practice Problems

7. How many protons and electrons are in each atom?

a. fluorine
b. aluminum
c. calcium

8. Complete the table.

Element	Atomic number	Protons	Electrons
K	19	[illegible]	19
____	____	____	5
____	16	____	____
____	____	23	____

Mass Number.

You know that most of the mass of an atom is concentrated in its nucleus and depends on the number of protons and neutrons. Look again at **Table 5.2** and note the number of protons and neutrons in helium and in carbon. The total number of protons and neutrons in an atom is called the **mass number.** A helium atom has two protons and two neutrons, so its mass number is 4. A carbon atom, which has six protons and six neutrons, has a mass number of 12.

If you know the atomic number and mass number of an atom of any element, you can determine the atom's composition. **Table 5.2** shows that an oxygen atom has an atomic number of 8 and a mass number of 16. Because the atomic number equals the number of protons, which equals the number of electrons, an oxygen atom has eight protons and eight electrons. The mass number of oxygen is 16 and is equal to the number of protons plus the number of neutrons. The oxygen atom, then, has eight neutrons, which is the difference between the mass number and the atomic number (16 − 8 = 8). For any atom, the number of neutrons can be determined from the following equation.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

The composition of any atom can be represented in shorthand notation using atomic number and mass number. **Figure 5.8** on page 116 shows how an atom of gold is represented using this notation. The chemical symbol Au appears with two numbers written to its left. The atomic number is the subscript (a number positioned lower). The mass number is the superscript (a number positioned higher). How many neutrons does this gold (Au) atom have?

Figure 5.8
Au is the chemical symbol for gold. How many electrons does a gold atom have?

You can also use the mass number and the name of the element to designate atoms. For example, atoms of hydrogen with a mass number of 1 may be designated hydrogen-1. Atoms of gold with a mass number of 197 are designated gold-197.

Sample Problem 5-2

How many protons, electrons, and neutrons are in the following atoms?

	Atomic number	Mass number
a. Beryllium (Be)	4	9
b. Neon (Ne)	10	20
c. Sodium (Na)	11	23

1. *ANALYZE* ***List the knowns and the unknowns.***

Knowns:

- For each atom, the atomic number and mass number are known.

Unknowns:

- number of protons = ?
- number of electrons = ?
- number of neutrons = ?

Use the definitions of atomic number and mass number to calculate the numbers of protons, electrons, and neutrons.

2. *CALCULATE* ***Solve for the unknowns.***

number of electrons = atomic number

a. 4 **b.** 10 **c.** 11

number of protons = atomic number

a. 4 **b.** 10 **c.** 11

number of neutrons = mass number − atomic number

a. $9 - 4 = 5$ **b.** $20 - 10 = 10$ **c.** $23 - 11 = 12$

3. *EVALUATE* ***Do the results make sense?***

The relationships among atomic number, number of protons, number of neutrons, number of electrons, and mass number have been applied correctly.

Practice Problems

9. How many neutrons are in each atom?

a. $^{16}_{8}O$ **c.** $^{108}_{47}Ag$ **e.** $^{207}_{82}Pb$

b. $^{32}_{16}S$ **d.** $^{80}_{35}Br$

10. Use **Table 5.2** and **Figure 5.8** to express the composition of each atom in shorthand form.

a. carbon-12

b. fluorine-19

c. beryllium-9

11. For each atom in Problem 9, identify the number of electrons.

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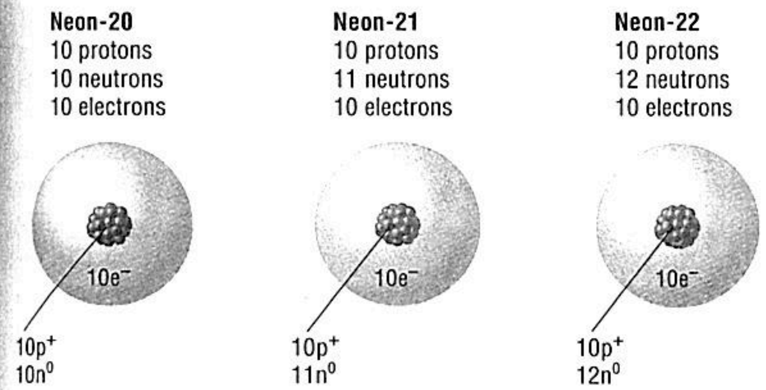
Problem-Solving 9

Solve Problem 9 with the help of an interactive guided tutorial.

Isotopes

Figure 5.9 shows that there are three different kinds of neon atoms. How do these atoms differ? All have the same number of protons (10) and electrons (10), but they each have different numbers of neutrons. Atoms that have the same number of protons but different numbers of neutrons are called **isotopes.** Because isotopes of an element have different numbers of neutrons, they also have different mass numbers. Despite these differences, isotopes are chemically alike because they have identical numbers of protons and electrons, which are the subatomic particles responsible for



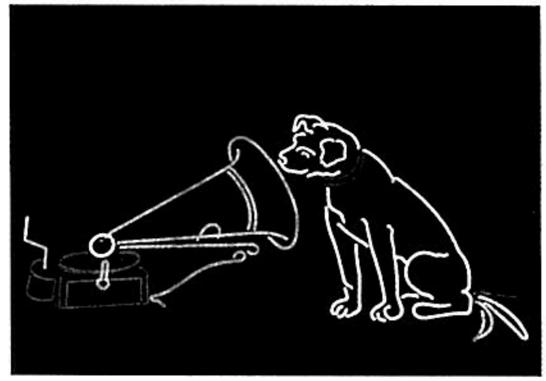


Figure 5.9
Neon-20, neon-21, and neon-22 are three isotopes of neon, a gaseous element used in lighted signs. How are these isotopes different? How are they the same?

chemical behavior. How does the discovery of isotopes contradict Dalton's atomic theory?

There are three known isotopes of hydrogen. Each isotope of hydrogen has one proton in its nucleus. The most common hydrogen isotope has no neutrons. It has a mass number of 1 and is called hydrogen-1 (^1_1H) or simply hydrogen. The second isotope has one neutron and a mass number of 2. It is called either hydrogen-2 (^2_1H) or deuterium. The third isotope has two neutrons and a mass number of 3. This isotope is called hydrogen-3 (^3_1H) or tritium.

Sample Problem 5-3

Two isotopes of carbon are carbon-12 and carbon-13. Write the symbol for each isotope using superscripts and subscripts to represent the mass number and the atomic number.

1. *ANALYZE* ***Plan a problem-solving strategy.***
   Knowns:
   - two isotopes of carbon: carbon-12 and carbon-13

   Unknowns:
   - each isotope's symbol

   Write the symbol for carbon and place the mass number to the left of the symbol as a superscript. Place the atomic number to the left of the symbol as a subscript.

2. *SOLVE* ***Apply the problem-solving strategy.***
   Based on **Table 5.2,** the symbol for carbon is C and the atomic number is 6. The mass number for each isotope is given by its name. For carbon-12, the symbol is $^{12}_{6}C$. For carbon-13, the symbol is $^{13}_{6}C$.

3. *EVALUATE* ***Do the results make sense?***
   The concepts of atomic number and mass number have been applied correctly, and the information is positioned properly next to the chemical symbols.

Practice Problems

12. Three isotopes of oxygen are oxygen-16, oxygen-17, and oxygen-18. Write the complete symbol for each, including the atomic number and mass number.

13. The three isotopes of chromium are chromium-50, chromium-52, and chromium-53. How many neutrons are in each isotope, given that chromium always has an atomic number of 24?

Chem ASAP!

Problem-Solving 13
Solve Problem 13 with the help of an interactive guided tutorial.

Atomic Mass

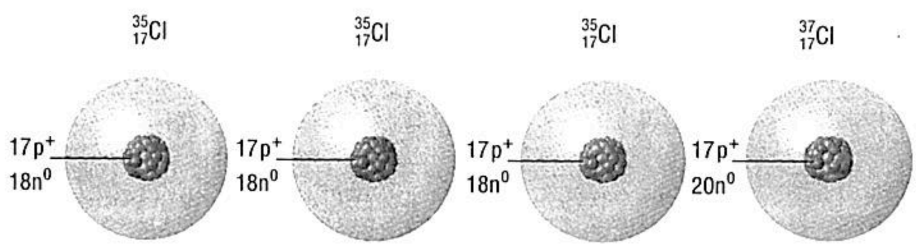
A glance back at **Table 5.1** on page 111 shows that the actual mass of a proton or a neutron is very small (1.67×10^{-24} g). The mass of an electron is 9.11×10^{-28} g, which is negligible in comparison. Given these values, the mass of even the largest atom is incredibly small. Since the 1920s, it has been possible to determine these tiny masses by using a mass spectrometer. With this instrument, the mass of a fluorine atom was found to be 3.155×10^{-23} g, and the mass of an arsenic atom was found to be 1.244×10^{-22} g. Such data about the actual masses of individual atoms can provide useful information, but, in general, these values are inconveniently small and impractical to work with. Instead, it is more useful to compare the relative masses of atoms using a reference isotope as a standard. The isotope chosen is carbon-12. This isotope of carbon was assigned a mass of exactly 12 atomic mass units. An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Using these units, a helium-4 atom, with a mass of 4.0026 amu, has about one-third the mass of a carbon-12 atom. How many carbon-12 atoms would have about the same mass as a nickel-60 atom?

A carbon-12 atom has six protons and six neutrons in its nucleus, and its mass is set as 12 amu. The twelve protons and neutrons account for nearly all of this mass. Therefore the mass of a single proton or a single neutron is about one-twelfth of 12 amu, or about 1 amu. Because the mass of any single atom depends mainly on the number of protons and neutrons in the nucleus of the atom, you might predict that the atomic mass of an element should be a whole number. However, that is not usually the case. For example, the atomic mass of chlorine (Cl) is 35.453 amu. How can such an atomic mass be explained? The explanation involves the relative abundance of the naturally occurring isotopes of the element.

In nature, most elements occur as a mixture of two or more isotopes. Each isotope of an element has a fixed mass and a natural percent abundance. Consider the three isotopes of hydrogen discussed earlier in this section. According to **Table 5.3**, almost all naturally occurring hydrogen (99.985%) is hydrogen-1. The other two isotopes are present in trace amounts. Notice that

Figure 5.10
Chlorine is a reactive element used to disinfect swimming pools. It is made up of two isotopes: chlorine-35 and chlorine-37. Because there is more chlorine-35 than chlorine-37, the atomic mass of chlorine, 35.453 amu, is closer to 35 than to 37; it is a weighted average.

Ratio of chlorine atoms in natural abundance: three $^{35}_{17}Cl$ to one $^{37}_{17}Cl$



Total number of protons in three $^{35}_{17}Cl$ atoms and one $^{37}_{17}Cl$ atom
(17 + 17 + 17 + 17)

Total number of neutrons in three $^{35}_{17}Cl$ atoms and one $^{37}_{17}Cl$ atom
(18 + 18 + 18 + 20)

$$\frac{68 + 74}{4} = 35.5 \text{ amu}$$

Weighted Average Mass of a Chlorine Atom

Table 5.3

Natural Percent Abundance of Stable Isotopes of Some Elements				
Name	**Symbol**	**Natural percent abundance**	**Mass (amu)**	**"Average" atomic mass**
Hydrogen	$^{1}_{1}H$	99.985	1.0078	
	$^{2}_{1}H$	0.015	2.0141	1.0079
	$^{3}_{1}H$	negligible	3.0160	
Helium	$^{3}_{2}He$	0.0001	3.0160	4.0026
	$^{4}_{2}He$	99.9999	4.0026	
Carbon	$^{12}_{6}C$	98.89	12.000	12.011
	$^{13}_{6}C$	1.11	13.003	
Nitrogen	$^{14}_{7}N$	99.63	14.003	14.007
	$^{15}_{7}N$	0.37	15.000	
Oxygen	$^{16}_{8}O$	99.759	15.995	
	$^{17}_{8}O$	0.037	16.995	15.999
	$^{18}_{8}O$	0.204	17.999	
Sulfur	$^{32}_{16}S$	95.002	31.972	
	$^{33}_{16}S$	0.76	32.971	32.06
	$^{34}_{16}S$	4.22	33.967	
	$^{36}_{16}S$	0.014	35.967	
Chlorine	$^{35}_{17}Cl$	75.77	34.969	35.453
	$^{37}_{17}Cl$	24.23	36.966	
Zinc	$^{64}_{30}Zn$	48.89	63.929	
	$^{66}_{30}Zn$	27.81	65.926	
	$^{67}_{30}Zn$	4.11	66.927	65.38
	$^{68}_{30}Zn$	18.57	67.925	
	$^{70}_{30}Zn$	0.62	69.925	

the atomic mass of hydrogen in **Table 5.3** (1.0079 amu) is very close to the mass of hydrogen-1 (1.0078 amu). The slight difference takes into account the larger masses, but smaller amounts, of the other two isotopes of hydrogen.

Now consider the two stable isotopes of chlorine listed in **Table 5.3:** chlorine-35 and chlorine-37. If you calculate the arithmetic mean of these two masses, you would get an average atomic mass of 35.968 amu ((34.969 amu + 36.966 amu)/2). A quick comparison with the value in **Table 5.3** indicates that 35.968 amu is higher than the actual value. To explain this, you need to know the natural percent abundance of the isotopes of chlorine. Chlorine-35 accounts for 75% of the naturally occurring chlorine atoms; chlorine-37 accounts for only 25%. See **Figure 5.10.**

LINK TO HUMANITIES

Philosophy of Science

Modern philosophers search for wisdom in many areas of human life, such as medicine, the arts, and the sciences. Today's philosophers of science are primarily concerned with the critical analysis of scientific concepts and the ways in which these concepts are expressed. These philosophers analyze such concepts as number, space, force, and organism. The search for wisdom leads philosophers of science to debate questions that textbooks may lead you to believe are settled. For example, is the modern scientific method the only correct way to examine and explain the natural world? Or is there a better way, as yet unknown? In spite of the successes of the scientific method in explaining nature, philosophers of science are still studying such questions. These studies could greatly affect science if an improved scientific method were suggested. For this reason, philosophical questions and answers about the methods and concepts of science could be an important aid to scientific progress.

The **atomic mass** of an element is a weighted average mass of the atoms in a naturally occurring sample of the element. A weighted average mass reflects both the mass and the relative abundance of the isotopes as they occur in nature.

Sample Problem 5-4

Which isotope of copper is more abundant: copper-63 or copper-65? (The atomic mass of copper is 63.546 amu.)

1. ***ANALYZE*** ***Plan a problem-solving strategy.***

   Knowns:
   - isotopes of copper: copper-63 and copper-65
   - atomic mass of copper = 63.546 amu

   Unknown:
   - isotope that is more abundant

   This problem can be solved without any calculations by analyzing the atomic mass of copper relative to the masses of the two isotopes.

2. ***SOLVE*** ***Apply the problem-solving strategy.***
   The atomic mass of 63.546 amu is closer to 63 than to 65. Thus, because the atomic mass is a weighted average of the isotopes, copper-63 must be more abundant than copper-65.

3. ***EVALUATE*** ***Does the result make sense?***
   The relative abundance is reflected in the atomic mass, so it is reasonable that copper-63 must be more abundant.

Practice Problems

14. Boron has two isotopes: boron-10 and boron-11. Which is more abundant, given that the atomic mass of boron is 10.81?

15. There are three isotopes of silicon; they have mass numbers of 28, 29, and 30. The atomic mass of silicon is 28.086 amu. Comment on the relative abundance of these three isotopes.

Now that you know that the atomic mass of an element is a weighted average of the masses of its isotopes, you can calculate atomic mass based on relative abundance. To do this, you must know three values:

- the number of stable isotopes of the element,
- the mass of each isotope, and
- the natural percent abundance of each isotope.

Table 5.3 on the previous page shows these values for a few elements. For other elements, you can use standard chemistry reference books. Once you have these values for an element, multiply the atomic mass of each isotope by its abundance, expressed as a decimal, then add the results. Sample Problem 5-5 illustrates this method.

Sample Problem 5-5

Element X has two natural isotopes. The isotope with a mass of 10.012 amu (^{10}X) has a relative abundance of 19.91%. The isotope with a mass of 11.009 amu (^{11}X) has a relative abundance of 80.09%. Calculate the atomic mass of this element.

Sample Problem 5-5 (cont.)

1. ***ANALYZE*** ***List the knowns and the unknown.***

Knowns:
- isotope ^{10}X:
  mass = 10.012 amu
  relative abundance = 19.91% = 0.1991
- isotope ^{11}X:
  mass = 11.009 amu
  relative abundance = 80.09% = 0.8009

Unknown:
- atomic mass of element X = ?

The mass each isotope contributes to the element's atomic mass can be calculated by multiplying the isotope's mass by its relative abundance. The atomic mass of the element is the sum of these contributions.

2. ***CALCULATE*** ***Solve for the unknown.***

for ^{10}X:	10.012 amu × 0.1991 =	1.993 amu
for ^{11}X:	11.009 amu × 0.8009 =	8.817 amu
for element X:	atomic mass =	10.810 amu

3. ***EVALUATE*** ***Does the result make sense?***
The calculated value is closer to the mass of the more abundant isotope, as would be expected.

Practice Problems

16. The element copper has naturally occurring isotopes with mass numbers of 63 and 65. The relative abundance and atomic masses are 69.2% for mass = 62.93 amu, and 30.8% for mass = 64.93 amu. Calculate the average atomic mass of copper.

17. Calculate the atomic mass of bromine. The two isotopes of bromine have atomic masses and relative abundance of 78.92 amu (50.69%) and 80.92 amu (49.31%).

Chem ASAP!
Problem-Solving 17
Solve Problem 17 with the help of an interactive guided tutorial.

section 5.3 review

18. Explain how the atomic number of an element identifies the element.

19. How can atomic number and mass number be used to find the numbers of protons, electrons, and neutrons?

20. An atom is identified as platinum-195.

a. What does the number represent?

b. Symbolize this atom using superscripts and subscripts.

21. How are isotopes of the same element alike? How are they different?

22. Determine the number of protons, electrons, and neutrons in each of the five isotopes of zinc.

23. List the number of protons, neutrons, and electrons in each pair of isotopes.

a. $^{6}_{3}Li$, $^{7}_{3}Li$ **b.** $^{42}_{20}Ca$, $^{44}_{20}Ca$ **c.** $^{78}_{34}Se$, $^{80}_{34}Se$

24. The atomic masses of elements are generally not whole numbers. Explain why.

25. How is the atomic mass of an element calculated from isotope data?

26. Using the data for nitrogen listed in **Table 5.3**, calculate the average atomic mass of nitrogen. Show your work.

Chem ASAP! **Assessment 5.3** Check your understanding of the important ideas and concepts in Section 5.3.

portfolio project

An instrument called a mass spectrometer can be used to determine the masses and relative abundances of isotopes. Find out how this instrument works and make a poster to summarize your findings.

THE PERIODIC TABLE: ORGANIZING THE ELEMENTS

How do you know where to find products in the supermarket? From your experience, you probably know that different types of products are arranged according to similar characteristics in aisles or sections of aisles. Such a classification structure makes finding and comparing products easy. **Is there a way of arranging more than 100 known elements?**

objectives

- Describe the origin of the periodic table
- Identify the position of groups, periods, and the transition metals in the periodic table

key terms

- periodic table
- periods
- periodic law
- group
- representative elements
- metals
- alkali metals
- alkaline earth metals
- transition metals
- inner transition metals
- nonmetals
- halogens
- noble gases
- metalloids

Development of the Periodic Table

About 70 elements had been discovered by the mid-1800s, but until the work of the Russian chemist Dmitri Mendeleev (1834–1907), no one had found a way to relate the elements in a systematic, logical way. Mendeleev listed the elements in columns in order of increasing atomic mass. He then arranged the columns so that the elements with the most similar properties were side by side. He thus constructed the first **periodic table,** an arrangement of the elements according to similarities in their properties. As you can see in **Figure 5.11,** Mendeleev left blank spaces in the table because there were no known elements with the appropriate properties and masses.

Mendeleev and others were able to predict the physical and chemical properties of the missing elements. Eventually these elements were discovered and were found to have properties similar to those predicted.

In 1913, Henry Moseley (1887–1915), a British physicist, determined the atomic number of the atoms of the elements. Moseley arranged the elements in a table by order of atomic number instead of atomic mass. That is the way the periodic table is arranged today.

			Ti = 50	Zr = 90	? = 180.
			V = 51	Nb = 94	Ta = 182.
			Cr = 52	Mo = 96	W = 186.
			Mn = 55	Rh = 104,4	Pt = 197,4
			Fe = 56	Ru = 104,4	Ir = 198.
			Ni = Co = 59	Pl = 106,6	Os = 199
H = 1			Cu = 63,4	Ag = 108	Hg = 200
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112	
	B = 11	Al = 27,4	? = 68	Ur = 116	Au = 197?
	C = 12	Si = 28	? = 70	Sn = 118	
	N = 14	P = 31	As = 75	Sb = 122	Bi = 210
	O = 16	S = 32	Se = 79,4	Te = 128?	
	F = 19	Cl = 35,5	Br = 80	I = 127	
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133	Tl = 204
		Ca = 40	Sr = 87,6	Ba = 137	Pb = 207.
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75,6	Th = 118?		

Figure 5.11
A version of Dmitri Mendeleev's periodic table is shown here.

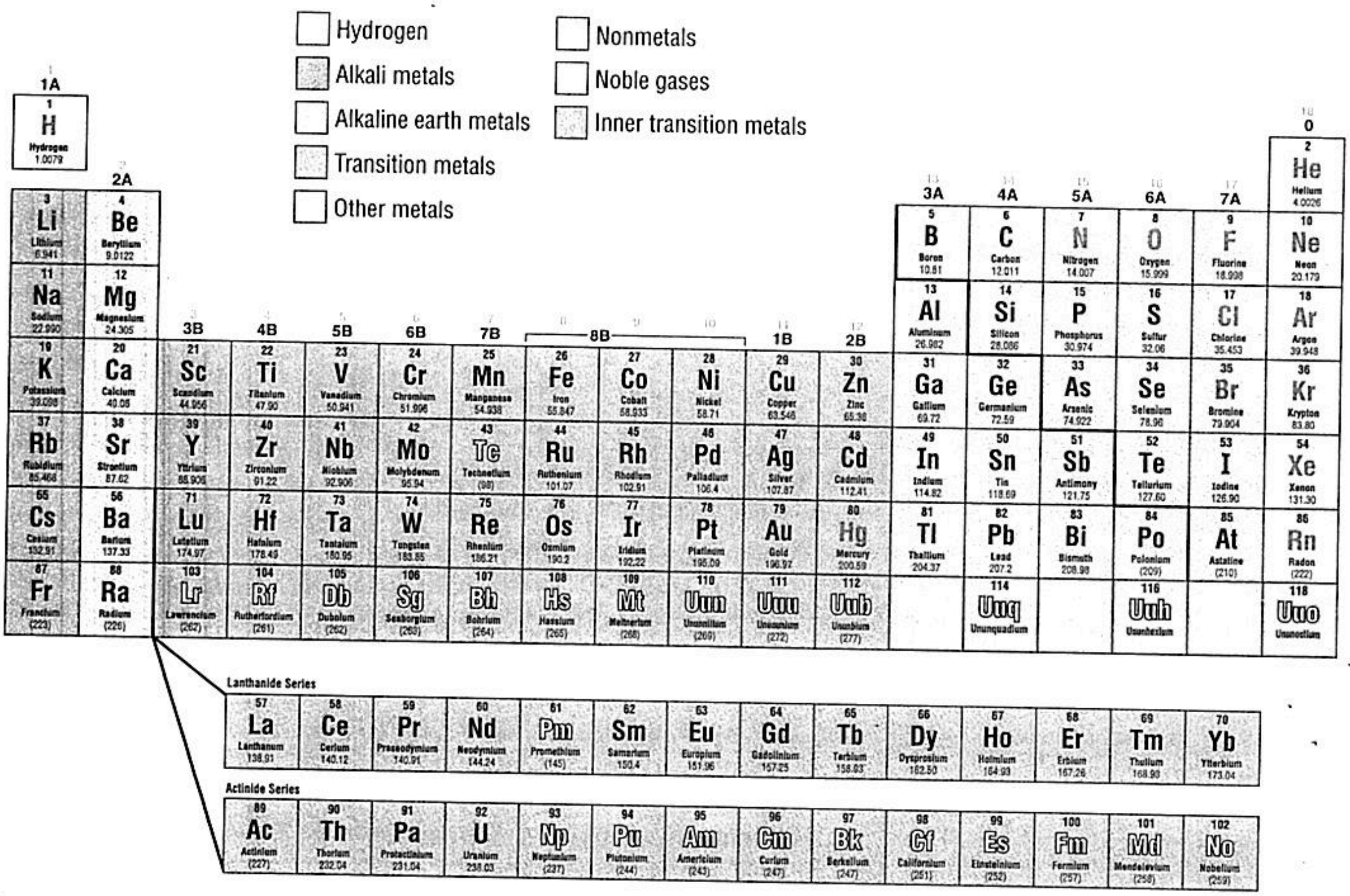


Figure 5.12

Elements are arranged in the modern periodic table in order of atomic number. The symbols are color coded according to the natural state of the elements: red for gases, black for solids, green for liquids (mercury and bromine), and white for elements that do not occur naturally.

The Modern Periodic Table

The most commonly used form of the modern periodic table, sometimes called the long form, is shown in **Figure 5.12.** Each element is identified by its symbol placed in a square. The atomic number of the element is shown centered above the symbol. The atomic mass and the name of the element are shown below the symbol. Notice that the elements are listed in order of increasing atomic number, from left to right and from top to bottom. Hydrogen (H), the lightest element, is in the top left corner. Helium (He), atomic number 2, is at the top right. Lithium (Li), atomic number 3, is at the left end of the second row.

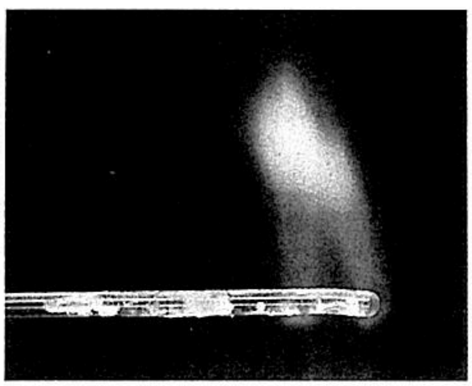
The horizontal rows of the periodic table are called **periods.** There are seven periods. The number of elements per period ranges from 2 (hydrogen and helium) in period 1, to 32 in period 6. The properties of the elements within a period change as you move across it from element to element. The pattern of properties within a period repeats, however, when you move from one period to the next. This gives rise to the **periodic law:** When the elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties. The arrangement of the elements into periods has an important consequence. Elements that have similar chemical and physical properties end up in the same column in the periodic table.

Each vertical column of elements in the periodic table is called a **group,** or family. The elements in any group of the periodic table have similar physical and chemical properties. Each group is identified by a number and the letter A or B. Look at the first column on the left. It includes the elements H, Li, Na, K, Rb, Cs, and Fr. This first column is designated Group 1A. Except for hydrogen, all of the Group 1A elements react vigorously, even explosively, with water. The next column to the right, Group 2A, starts with Be. Next comes Group 3A, toward the right of the table. The Group A elements are made up of Group 1A through Group 7A and Group 0 (the group at the far right). Group A elements are called the **representative elements** because they exhibit a wide range of both physical and chemical properties.

The representative elements can be divided into three broad classes. The first are **metals,** which have a high electrical conductivity and a high luster when clean. They are ductile (able to be drawn into wires) and malleable (able to be beaten into thin sheets). Except for hydrogen, the representative elements on the left side of the periodic table are metals. The Group 1A elements are called the **alkali metals,** and the Group 2A elements are called the **alkaline earth metals.** Most of the remaining elements that are not Group A elements are also metals. These include the **transition metals** and the **inner transition metals,** which together make up the Group B elements. Copper, silver, gold, and iron are familiar transition metals. The inner transition metals, which appear below the main body of the periodic table, are also called the rare-earth elements. Approximately 80% of all of the elements are metals. With one exception, all metals are solids at room temperature. **Figure 5.14** on page 126 shows the exception to this rule. What is the name, symbol, and physical state of this element?

Figure 5.13

The elements in the periodic table vary greatly in their properties.



Sodium emits bright yellow light during a flame test.

Potassium reacts violently with water.

Chromium, the principal element in chrome plating, resists corrosion.

Iodine exists as a solid and a vapor at 25°C.

Because of its malleability, silver is easily stamped into coins.

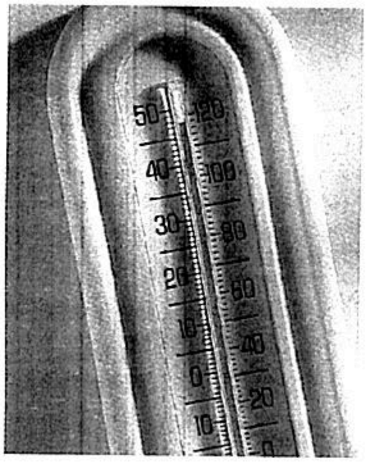


Figure 5.14
Mercury, a transition metal, is the only metallic element that is a liquid at room temperature. It is used in thermometers and barometers and as the electrical contact in a thermostat.

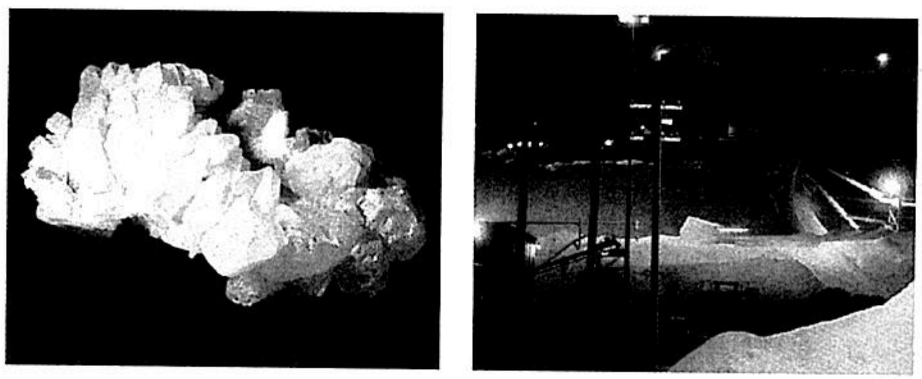


Figure 5.15
Sulfur is a low-melting point nonmetallic element that occurs as a crystalline solid or in the amorphous (formless) state. It is often mined through a process involving the pumping of hot water, which melts the sulfur. Sulfur is used primarily in the manufacture of sulfuric acid.

The nonmetals occupy the upper-right corner of the periodic table. **Nonmetals** are elements that are generally nonlustrous and that are generally poor conductors of electricity. Some of these elements, such as oxygen and chlorine, are gases at room temperature. Others, such as sulfur, shown in **Figure 5.15,** are brittle solids. One element, bromine, is a fuming dark-red liquid at room temperature. Two groups of nonmetals are given special names. The nonmetals of Group 7A are called the **halogens,** which include chlorine and bromine. The nonmetals of Group 0 are known as the **noble gases,** which are sometimes called the inert gases because they undergo few chemical reactions. The noble gas neon is used to fill the glass tubes of neon lights.

Notice the heavy stair-step line in **Figure 5.12.** This line divides the metals from the nonmetals. Most of the elements that border this line are **metalloids,** elements with properties that are intermediate between those of metals and nonmetals. Silicon and germanium are two important metalloids that are used in the manufacture of computer chips and solar cells.

Without the help of the periodic table, it would be quite difficult to learn and remember the chemical and physical properties of the more than 100 elements. Instead of memorizing their properties separately, you need only learn the general behavior and trends within the major groups. This gives you a useful working knowledge of the properties of most elements.

section 5.4 review

27. Describe how the periodic table was developed.

28. What criteria did Mendeleev use to construct his periodic table of the elements?

29. Relate group, period, and transition metals to the periodic table.

30. Identify each element as a metal, metalloid, or nonmetal.

a. gold **b.** silicon **c.** manganese **d.** sulfur **e.** barium

31. Which of the elements listed in the preceding question are representative elements?

32. Name two elements that have properties similar to those of the element calcium.

Chem ASAP! **Assessment 5.4** Check your understanding of the important ideas and concepts in Section 5.4.