4.3

1 FOCUS

Objectives

4.3.1 **Explain** what makes elements and isotopes different from each other.

4.3.2 **Calculate** the number of neutrons in an atom.

4.3.3 **Calculate** the atomic mass of an element.

4.3.4 **Explain** why chemists use the periodic table.

Guide for Reading

Build Vocabulary L2

Graphic Organizer Have students use the vocabulary for this section to build a concept map that links and relates the terms.

Reading Strategy L2

Visualize Have students revisit Figure 4.10 when they are tackling Sample Problem 4.2. Ask how the visual relates to the problem.

2 INSTRUCT

Connecting to Your World

Have students study the photograph and read the text that opens the section. Ask, **What characteristics can you use to classify different apples?** *(Sample answers: color, size, taste, texture, cooking qualities)* **Which of these involve the apple's chemistry?** *(taste, cooking qualities)*

Atomic Number

Use Visuals L1

Table 4.2 Point out that the atomic number is equal to the number of protons for each element. Ask, **Why must the number of electrons equal the number of protons for each element?** *(Atoms are electrically neutral.)* **What is the relationship between the number of protons and the number of neutrons?** *(The number of neutrons tends to rise with the number of protons.)*

4.3 Distinguishing Among Atoms

Guide for Reading

Key Concepts

- What makes one element different from another?
- How do you find the number of neutrons in an atom?
- How do isotopes of an element differ?
- How do you calculate the atomic mass of an element?
- Why is a periodic table useful?

Vocabulary

atomic number
mass number
isotopes
atomic mass unit (amu)
atomic mass
periodic table
period
group

Reading Strategy

Building Vocabulary As you read the section, write a definition of each vocabulary term in your own words.

Connecting to Your World Fruits and vegetables come in different varieties. For example, a grocery store might sell three varieties of apples: Granny Smith, Red Delicious, and Golden Delicious. Apple varieties can differ in color, size, texture, aroma, and taste. Just as apples come in different varieties, a chemical element can come in different "varieties" called isotopes. In this section, you will learn how one isotope of an element differs from another.

Atomic Number

Atoms are composed of protons, neutrons, and electrons. Protons and neutrons make up the nucleus. Electrons surround the nucleus. How, then, are atoms of hydrogen, for example, different from atoms of oxygen? Look at Table 4.2. Notice that a hydrogen atom has one proton, but an oxygen atom has eight protons. **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. For each element listed in Table 4.2, the number of protons equals the number of electrons. Remember that atoms are electrically neutral. Thus, the number of electrons (negatively charged particles) must equal the number of protons (positively charged particles).

Table 4.2

Atoms of the First Ten Elements

Name	Symbol	Atomic number	Protons	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 4.3.

Section Resources

Print

- **Guided Reading and Study Workbook,** Section 4.3
- **Core Teaching Resources,** Section 4.3 Review
- **Transparencies,** T48–T56
- **Small-Scale Chemistry Laboratory Manual,** Lab 6

Technology

- **Interactive Textbook with ChemASAP,** Problem-Solving 4.15, 4.17, 4.20, 4.21, 4.24, Assessment 4.3
- **Go Online,** Section 4.3

CONCEPTUAL PROBLEM 4.1

Understanding Atomic Number

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

1 Analyze ***Identify the relevant concepts.***

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

2 Solve ***Apply the concepts to this problem.***

The atomic number of nitrogen is 7, which means that a neutral nitrogen atom has 7 protons and 7 electrons.

Practice Problems

15. Complete the table.

Element	Atomic number	Protons	Electrons
K	19	(a)	19
(b)	(c)	(d)	5
S	16	(e)	(f)
V	(g)	23	(h)

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

a. fluorine (atomic number = 9)
b. calcium (atomic number = 20)
c. aluminum (atomic number = 13)

Interactive Textbook

Problem-Solving 4.15 Solve Problem 15 with the help of an interactive guided tutorial.

with ChemASAP

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. The total number of protons and neutrons in an atom is called the **mass number.** Look again at Table 4.2 and note the mass numbers of helium and carbon. 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 4.2 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 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 ($16 - 8 = 8$). **The number of neutrons in an atom is the difference between the mass number and atomic number.**

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

The composition of any atom can be represented in shorthand notation using atomic number and mass number. Figure 4.8 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. The mass number is the superscript.

You can also refer to atoms by using the mass number and the name of the element. For example, $^{197}_{79}\text{Au}$ may be written as gold-197.

Checkpoint ***How do you calculate mass number?***

Figure 4.8 Au is the chemical symbol for gold. **Applying Concepts** ***How many electrons does a gold atom have?***

CONCEPTUAL PROBLEM 4.1

Answers

15. a. 19 **b.** B **c.** 5 **d.** 5 **e.** 16 **f.** 16 **g.** 23 **h.** 23

16. a. 9 protons and 9 electrons
b. 20 protons and 20 electrons
c. 13 protons and 13 electrons

Practice Problems Plus L2

Chapter 4 Assessment problem 47 is related to Conceptual Problem 4.1.

Mass Number

Discuss L2

Discuss why the mass number of an element is defined as the number of protons and neutrons. Explain that chemists have arbitrarily assigned a value of one atomic mass unit to represent the mass of one twelfth of a carbon-12 atom. Ask, **How do you find the number of neutrons in an atom?** *(Subtract the atomic number from the mass number.)*

Differentiated Instruction

Less Proficient Readers L1

Have students make a list of familiar elements and describe at least one use for each element. Have students combine their lists into a master list on the chalkboard.

Answers to...

Figure 4.8 79 electrons

Checkpoint Mass number = number of neutrons + atomic number

Section 4.3 (continued)

Sample Problem 4.1

Answers

17. a. 8 **b.** 16 **c.** 61 **d.** 45 **e.** 125

18. a. $^{12}_{6}C$ **b.** $^{19}_{9}F$ **c.** $^{9}_{4}Be$

Practice Problems Plus L2

The element strontium has four isotopes, strontium-84, strontium-86, strontium-87, and strontium-88. **Given that strontium has an atomic number of 38, how many neutrons are in each of these isotopes?** *(46, 48, 49, 50)*

Isotopes

CLASS Activity

Applications of Isotopes L2

Purpose To learn about practical applications of isotopes

Materials Library or Internet access

Procedure Have students use the library or Internet to find an isotope of an element that has a practical or everyday use.

Expected Outcome Students' research will most likely focus on applications of radioisotopes such as carbon-14 (used in archaeological dating), americium-241 (used in smoke alarms), iodine-131 (used in the treatment of thyroid disorders), and cobalt-60 (used in the treatment of some cancers). Point out that the instability of these isotopes is what makes them useful. Radioisotopes and radioactivity are discussed in Chapter 25.

SAMPLE PROBLEM 4.1

Determining the Composition of an Atom

How many protons, electrons, and neutrons are in each atom?

	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	Unknowns
• atomic number	• number of protons = ?
• mass number	• 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?***

For each atom, the mass number equals the number of protons plus the number of neutrons. The results make sense.

Practice Problems

17. How many neutrons are in each atom?

a. $^{16}_{8}O$ **b.** $^{32}_{16}S$ **c.** $^{108}_{47}Ag$

d. $^{80}_{35}Br$ **e.** $^{207}_{82}Pb$

18. Use Table 4.2 to express the composition of each atom in shorthand form.

a. carbon-12 **b.** fluorine-19 **c.** beryllium-9

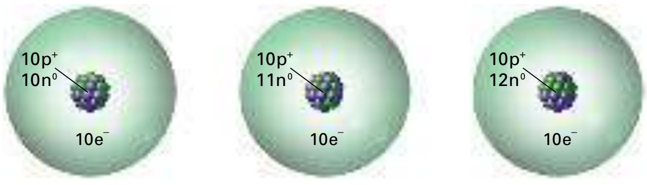
Interactive Textbook

Problem-Solving 4.17
Solve Problem 17 with the help of an interactive guided tutorial.

with ChemASAP

Isotopes

Figure 4.9 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. **Isotopes** are atoms that have the same number of protons but different numbers of neutrons. **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 chemical behavior.



Neon-20
10 protons
10 neutrons
10 electrons

Neon-21
10 protons
11 neutrons
10 electrons

Neon-22
10 protons
12 neutrons
10 electrons

Figure 4.9 Neon-20, neon-21, and neon-22 are three isotopes of neon, a gaseous element used in lighted signs. **Comparing and Contrasting** ***How are these isotopes different? How are they similar?***

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 ($^{1}_{1}H$) or simply hydrogen. The second isotope has one neutron and a mass number of 2. It is called either hydrogen-2 ($^{2}_{1}H$) or deuterium. The third isotope has two neutrons and a mass number of 3. This isotope is called hydrogen-3 ($^{3}_{1}H$) or tritium.

Checkpoint ***What are three known isotopes of hydrogen?***

CONCEPTUAL PROBLEM 4.2

Writing Chemical Symbols of Isotopes

Diamonds are a naturally occurring form of elemental carbon. Two stable 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 ***Identify the relevant concepts.***

Isotopes are atoms that have the same number of protons but different numbers of neutrons. The composition of an atom can be expressed by writing the chemical symbol, with the atomic number as a subscript and the mass number as a superscript.

2 Solve ***Apply the concepts to this problem.***

Based on Table 4.2, 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 $^{13}_{6}C$.

Practice Problems

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

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

Interactive Textbook

Problem-Solving 4.20
Solve Problem 20 with the help of an interactive guided tutorial.
with **ChemASAP**

CONCEPTUAL PROBLEM 4.2

Answers

19. $^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$

20. Chromium-50 has 26 neutrons; chromium-52 has 28 neutrons; chromium-53 has 29 neutrons.

Practice Problems Plus L2

Calculate the number of neutrons in each of the following radioactive isotopes.

a. $^{14}_{6}C$ *(8)*

b. $^{40}_{19}K$ *(21)*

c. $^{238}_{92}U$ *(146)*

d. $^{99}_{42}Mo$ *(57)*

Discuss L2

Point out that isotopes of an element are chemically the same. Ask students why that is the case. *(because the electrons, not the neutrons, determine the atom's chemical properties)*

Answers to...

Figure 4.9 They have different numbers of neutrons but the same number of protons.

Checkpoint hydrogen-1, hydrogen-2, hydrogen-3

Section 4.3 (continued)

Atomic Mass

Use Visuals

L1

Table 4.3 Point out that the average atomic masses listed in Table 4.3 are based on the masses of stable isotopes and their percent abundance in Earth's crust. Have students study the average atomic masses in the table. Ask, **Which elements exist predominantly as one natural isotope?** *(those with atomic masses closest to a whole number)* **Which element has substantial amounts of each of its natural isotopes?** *(chlorine)*

FYI

Students may ask why the masses (in amu) of most of the isotopes in Table 4.3 are not whole-number values like mass numbers. Although the mass number of carbon-12 exactly equals its mass in amu, this is generally not the case for other isotopes due to the mass defect. The mass defect is the difference between the mass of a nucleus and the sum of the masses of its component protons and neutrons. (This difference can also be expressed in terms of binding energy, based on the relationship $E = mc^2$, in which E = energy, m = mass, and c = the speed of light.) Because mass defect varies with different elements, the masses of isotopes other than carbon-12 in amu (a unit based on the mass of a carbon-12 atom) will generally not be whole numbers. For example, the mass of hydrogen-2 is 2.0141 amu; the mass of oxygen-16 is 15.995 amu. Chapter Assessment problem 82 asks students to calculate the mass defect of an atom in grams by comparing the actual mass of the atom to the sum of the masses of the atom's component protons, neutrons, and electrons.

Atomic Mass

A glance back at Table 4.1 on page 106 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 of 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. On the other hand, a nickel-60 atom has about five times the mass of a carbon-12 atom.

A carbon-12 atom has six protons and six neutrons in its nucleus, and its mass is set as 12 amu. The six protons and six neutrons account for nearly all of this mass. Therefore the mass of a single proton or a single

Table 4.3

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	1.0079
	$^{2}_{1}H$	0.015	2.0141	
	$^{3}_{1}H$	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	15.999
	$^{17}_{8}O$	0.037	16.995	
	$^{18}_{8}O$	0.204	17.999	
Sulfur	$^{32}_{16}S$	95.002	31.972	32.06
	$^{33}_{16}S$	0.76	32.971	
	$^{34}_{16}S$	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	

Crystalline sulfur

114 Chapter 4

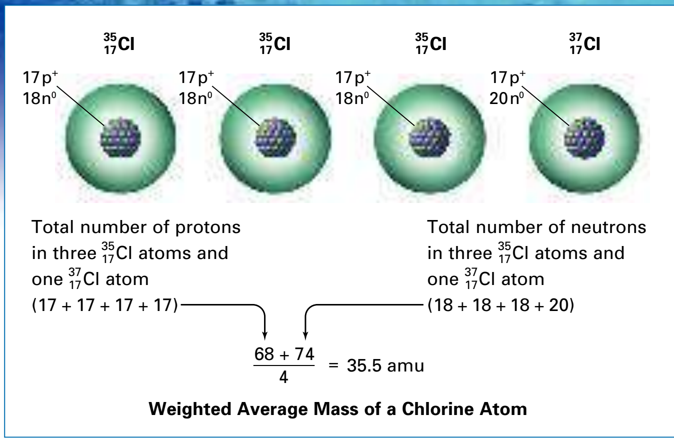


Figure 4.10 Chlorine is a reactive element used to disinfect swimming pools. Chlorine occurs as two isotopes: chlorine-35 and chlorine-37. Because there is more chlorine-35 than chlorine-37, the atomic mass of chlorine, 35.453 amu, is closer to 35 than to 37. **Evaluating** ***How does a weighted average differ from an arithmetic mean?***

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.

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 4.3, almost all naturally occurring hydrogen (99.985%) is hydrogen-1. The other two isotopes are present in trace amounts. Notice that the atomic mass of hydrogen listed in Table 4.3 (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 4.3: chlorine-35 and chlorine-37. If you calculate the arithmetic mean of these two masses ((34.969 amu + 36.966 amu)/2), you get an average atomic mass of 35.968 amu. However, this value is higher than the actual value of 35.453. To explain this difference, 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 4.10. 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.

Checkpoint ***What is the atomic mass of an element?***

Go Online NSTA SCILINKS

For: Links on Isotopes
Visit: www.SciLinks.org
Web Code: cdn-1043

Go Online NSTA SCILINKS

Download a worksheet on **Isotopes** for students to complete, and find additional teacher support from NSTA SciLinks.

Facts and Figures

Determining Relative Abundance of an Element's Isotopes

How do you determine relative abundance for an element's isotopes? Look at a sample of the element in a mass spectrometer. The display shows a peak at the mass of each isotope. The height of the peak shows the relative abundance of each isotope.

Answers to...

Figure 4.10 In an arithmetic mean, all the numbers in the calculation are weighted equally. A weighted average takes into account the varying weights of the numbers in the data set.

Checkpoint Atomic mass is the weighted average mass of the atoms in a naturally occurring sample of the element.

CONCEPTUAL PROBLEM 4.3

Answers

21. boron-11

22. Silicon-28 must be by far the most abundant. The other two isotopes must be present in very small amounts.

Practice Problems Plus L2

Argon has three isotopes with mass numbers 36, 38, and 40, respectively. **Which of these isotopes is the most abundant?** *(argon-40)*

Relate L2

Teachers sometimes evaluate a student's performance based on the weighted average for different work, for example, a term paper worth 20%, a midterm worth 30%, and a final exam worth 50%. If an A = 4.0, a B = 3.0, and a C = 2.0, have students calculate a grade for a student who receives a B on the term paper, a C on the midterm, and a B on the final exam. *(The student would receive a score of 2.7, a C+.)*

CONCEPTUAL PROBLEM 4.3

Using Atomic Mass to Determine the Relative Abundance of Isotopes

The atomic mass of copper is 63.546 amu. Which of copper's two isotopes is more abundant: copper-63 or copper-65?

1 Analyze ***Identify the relevant concepts.***

The atomic mass of an element is the 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.

2 Solve ***Apply the concepts to this problem.***

The atomic mass of 63.546 amu is closer to 63 than to 65. Because the atomic mass is a weighted average of the isotopes, copper-63 must be more abundant than copper-65.

Practice Problems

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

22. 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.

Interactive Textbook

Problem-Solving 4.21 Solve Problem 21 with the help of an interactive guided tutorial.

with ChemASAP

Now that you know that the atomic mass of an element is a weighted average of the masses of its isotopes, you can determine 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. **To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abundance, expressed as a decimal, and then add the products.** The resulting sum is the weighted average mass of the atoms of the element as they occur in nature.

You can calculate the atomic masses listed in Table 4.3 based on the given masses and natural abundances of the isotopes for each element. For example, carbon has two stable isotopes: carbon-12, which has a natural abundance of 98.89%, and carbon-13, which has natural abundance of 1.11%. The mass of carbon-12 is 12.000 amu; the mass of carbon-13 is 13.003 amu. The atomic mass is calculated as follows.

$$\begin{aligned}\text{Atomic mass of carbon} &= (12.000\ \text{amu} \times 0.9889) + (13.003\ \text{amu} \times 0.0111)\\ &= 12.011\ \text{amu}\end{aligned}$$

Checkpoint ***What three values must be known in order to calculate the atomic mass of an element?***

Facts and Figures

Carbon-14 Dating

All living organisms contain carbon-12 and carbon-14 in a fixed ratio. After an organism dies, this ratio changes as the carbon-14 decays. Paleontologists and archaeologists use this fact to establish the age of fossils and ancient artifacts.

SAMPLE PROBLEM 4.2

Calculating Atomic Mass

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.

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 products.

2 Calculate ***Solve for the unknown.***

for ^{10}X: $10.012 \text{ amu} \times 0.1991 = 1.993 \text{ amu}$
for ^{11}X: $11.009 \text{ amu} \times 0.8009 = \underline{8.817 \text{ amu}}$
for element X: atomic mass = 10.810 amu

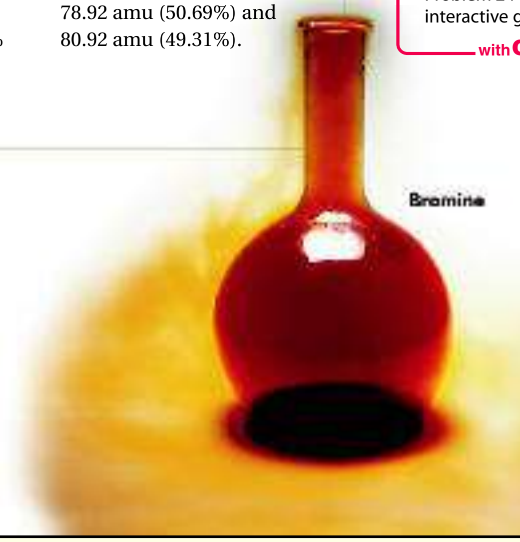
3 Evaluate ***Does the result make sense?***

The calculated value is closer to the mass of the more abundant isotope, as would be expected.

Practice Problems

23. 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.

24. 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%).



Bromine

CHEMath

Percents

A percent is a shorthand way of expressing a fraction whose denominator is 100. For example, 85% is equivalent to 85/100 or 0.85.

When working with percents, it is usually necessary to convert percents to fractions or decimals before using them in a calculation. For instance, if the natural percent abundance of an isotope is 35.6%, then there are 35.6 g of that isotope in 100 g of the element.

Math Handbook

For help with percents, go to page R72.

Interactive Textbook

Problem-Solving 4.24 Solve Problem 24 with the help of an interactive guided tutorial.

with ChemASAP

Sample Problem 4.2

Answers

23. 63.6 amu
24. 79.91 amu

Practice Problems Plus L2

Chlorine has two isotopes, chlorine-35 (atomic mass 5 34.97 amu, relative abundance 5 75.77%) and chlorine-37 (atomic mass 5 36.97 amu, relative abundance 5 24.23%). **Calculate the atomic mass of chlorine.** *(35.45 amu)*

CHEMath

Percents

Students will need to have a basic understanding of percents in order to calculate atomic mass. Atomic mass is a weighted average that takes into account natural percent abundance. Show students that the relative percent abundances of the isotopes add up to 100%.

Math Handbook

For a math refresher and practice, refer students to percents, page R72.

Answers to...

Checkpoint The number of stable isotopes, the mass of each isotope, and the natural abundance (expressed as a percent) of each isotope.

The Periodic Table—A Preview

Use Visuals L1

Figure 4.11 Have students study the periodic table. Explain that just as items in a supermarket are arranged by food type, such as breads, dairy products, and produce, the chemical elements in the periodic table are arranged in groups that have similar chemical properties. Ask, **Where do you find elements with similar chemical behaviors?** *(in a column)* **What elements are similar chemically to oxygen, O?** *(S, Se, Te, Po)*

FYI

The modern periodic table is discussed in more detail in Chapter 6 and also in Appendix A: Elements Handbook

Careers *in* Chemistry

Archaeologist

Encourage students to find more information on archaeological careers on the Internet. An interesting possibility is nautical archaeology, which sends archaeologists into lakes and oceans. Here's a site for nautical archaeology with links to other web pages. http://www.schoolship.org/careers/archaeologist.html.

Go Online PHSchool.com

Have students research chemistry-related careers. They can then construct a table that describes the nature of the job, requirements, employment outlook, working conditions, and other necessary information.

	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub		114 Uuq				

57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No

Figure 4.11 Elements are arranged in the modern periodic table in order of atomic number. **Interpreting Diagrams** ***How many elements are in Period 2? In Group 2A?***

The Periodic Table—A Preview

Now that you can differentiate between atoms of different elements and also between isotopes of the same element, you need to understand how the elements are organized with respect to each other. A **periodic table** is an arrangement of elements in which the elements are separated into groups based on a set of repeating properties. **A periodic table allows you to easily compare the properties of one element (or a group of elements) to another element (or group of elements).**

Figure 4.11 shows the most commonly used form of the modern periodic table, sometimes called the long form. Each element is identified by its symbol placed in a square. The atomic number of the element is shown centered above the symbol. Notice that the elements are listed in order of increasing atomic number, from left to right and 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.

Each horizontal row of the periodic table is called a **period.** There are seven periods in the modern periodic table. The number of elements per period ranges from 2 (hydrogen and helium) in Period 1, to 32 in Period 6. Within a given period, the properties of the elements vary as you move across it from element to element. This pattern of properties then repeats as you move to the next period.

Each vertical column of the periodic table is called a **group,** or family. Elements within a group have similar chemical and physical properties. Note that each group is identified by a number and the letter A or B. For example, Group 2A contains the elements beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). You will learn more about specific trends in the periodic table in Chapter 6.

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Careers in Chemistry

Archaeologist

Archaeologists are detectives of the past, sifting for clues that uncover secrets of past civilizations. Archaeologists excavate ancient cities and dwellings looking for artifacts that indicate what kinds of foods ancient people ate, what types of tools they used, and how they interacted with one another as a society. Often, archaeologists must draw conclusions based on indirect evidence.

Knowing when an event occurred or when an artifact was made can provide important information. Archaeologists use techniques such as radiometric-dating, in which a sample is dated by measuring the concentration of certain isotopes, such as carbon-14. This method tells them the age of a sample within a certain range, and is used with the greatest accuracy for samples no more than 10,000 years old.

Archaeologists also perform chemical tests on artifacts to determine their composition. For example, archaeologists might analyze the glazes used on pottery, or the dyes used in clothing. Archaeologists may also use chemicals to preserve artifacts that have been unearthed, so that the artifacts can be examined without being damaged.

Archaeology requires a background in both history and science. Archaeologists often spend as much time in the laboratory studying their finds as they do out in the field excavating sites. Archaeologists take courses in archaeological techniques, biology, anatomy, chemistry, math, and history.

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4.3 Section Assessment

25. **Key Concept** What distinguishes the atoms of one element from the atoms of another?

26. **Key Concept** What equation tells you how to calculate the number of neutrons in an atom?

27. **Key Concept** How do the isotopes of a given element differ from one another?

28. **Key Concept** How is atomic mass calculated?

29. **Key Concept** What makes the periodic table such a useful tool?

30. What does the number represent in the isotope platinum-194? Write the symbol for this atom using superscripts and subscripts.

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

32. 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.** $^{78}_{34}Se$, $^{80}_{34}Se$

33. Name two elements that have properties similar to those of the element calcium (Ca).

Elements Handbook

Elements Within You Read page R5 of the Elements Handbook. Identify the five most abundant elements in the human body, and locate them on the periodic table.

Interactive Textbook

Assessment 4.3 Test yourself on the concepts in Section 4.3.

with **ChemASAP**

Section 4.3 Assessment

25. Atoms of different elements contain different numbers of protons.

26. mass number − atomic number = number of neutrons

27. They have different mass numbers and different numbers of neutrons.

28. For each isotope, multiply its atomic mass by its % abundance, then add the products.

29. It allows you to compare the properties of the elements.

30. Mass number, $^{194}_{78}Pt$

31. The atomic mass is the weighted average of the masses of its isotopes.

32. **a.** lithium-6: 3 p^+, 3 e^-, 3 n^0 ; lithium-7: 3 p^+, 3 e^-, 4 n^0 **b.** calcium-42: 20 p^+, 20 e^-, 22 n^0; calcium-44: 20 p^+, 20 e^-, 24 n^0 **c.** selenium-78: 34 p^+, 34 e^-, 44 n^0; selenium-80: 34 p^+, 34 e^-, 46 n^0

33. any two: beryllium (Be), magnesium (Mg), strontium (Sr), barium (Ba), radium (Ra)

FYI

Radiometric dating and its applications are discussed in more detail in Chapter 25.

3 ASSESS

Evaluate Understanding L2

Write the symbols for isotopes of an element not described in the chapter. Ask students what the superscripts and subscripts refer to and the differences between the atoms shown. Ask, **How would changing the value of the subscript change the chemical properties of the atom?** *(The subscript designates the number of protons in the atoms of that isotope. Changing the number of protons would change the chemical identity of the isotope to that of another element.)*

Reteach L1

Review the concept of weighted averages. Work through the following calculations. If 75% of chlorine atoms are ^{35}Cl species and 25% are ^{37}Cl species, this implies that for a sample of 100 atoms, 75 atoms are ^{35}Cl and 25 atoms are ^{37}Cl species. The combined masses of these atoms would be (75 × 35 amu) + (25 × 37 amu) = 3550 amu for 100 atoms, or 35.5 amu for one atom.

Elements Handbook

Oxygen (Period 2, Group 6A), carbon (Period 2, Group 4A), hydrogen (Period 1, Group 1A), nitrogen (Period 2, Group 5A), and calcium (Period 4, Group 2A).

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 4.3.

with **ChemASAP**

Answers to...

Figure 4.11 There are eight elements in Period 2. There are 6 elements in Group 2A.

Small-Scale LAB

The Atomic Mass of Candium L2

Objective After completing this activity, students will be able to:

- make measurements to calculate the relative abundances of three types of candy in a mixture.
- use their data to calculate the atomic mass of a candium particle.

Skills Focus Measuring, calculating

Prep Time 15 minutes

Materials mass balance, coated candies (3 different brands), small plastic cups or containers

Advance Prep Prepare in advance a large mixture of the three candies and half fill a clean 3.5-ounce plastic cup for each student. Each sample will contain about 50 total pieces.

Class Time 30 minutes

Safety Discourage students from eating the candies after the experiment. Contamination can easily occur in a lab even if you have taken every precaution to keep the candy free of contamination.

Teaching Tips This lab is similar to the longer Small-Scale Lab "Isotopes and Atomic Mass" found in the *Small-Scale Chemistry Laboratory Manual*.

Expected Outcome Sample data are listed below.

Total Mass: 13.16 g; 13.83 g; 15.40 g; 42.39 g (total).
Number: 15; 13; 20; 48 (total).
Average mass (grams): 0.8773 g; 1.064 g; 0.7700 g; 0.8831 g (total).
Relative abundance: 0.3125, 0.2708, 0.4167, 1.000 (total).
Percent abundance: 31.25%; 27.08%; 41.67%; 100% (total).
Relative mass: 0.2742 g; 0.2883 g; 0.3208 g; 0.8833 g (total).

Analyze

1–5. See Expected Outcome

6. Percent abundance is parts per hundred. Relative abundance is parts per one, or the decimal form of percent. The individual percent abundances add up to 100. The individual relative abundances add up to 1.

7. Relative abundance tells you the decimal fraction of particles.

Small-Scale LAB

The Atomic Mass of Candium

Purpose

To analyze the isotopes of "candium" and to calculate its atomic mass.

Materials

- sample of candium
- balance
- pencil
- paper

Procedure

Obtain a sample of "candium" that contains three different brands of round, coated candy. Treat each brand of candy as an isotope of candium. Separate the three isotopes into groups labeled A, B, and C, and measure the mass of each isotope. Count the number of atoms in each sample. Make a table similar to the one below to record your measured and calculated data.

	A	B	C	Totals
Total mass (grams)				
Number				
Average mass (grams)				
Relative abundance				
Percent abundance				
Relative mass				

Analyze

Using the experimental data, record the answers to the following questions below your data table.

1. Calculate the average mass of each isotope by dividing its total mass by the number of particles of that isotope.
2. Calculate the relative abundance of each isotope by dividing its number of particles by the total number of particles.
3. Calculate the percent abundance of each isotope by multiplying the relative abundance from Step 2 by 100.
4. Calculate the relative mass of each isotope by multiplying its relative abundance from Step 2 by its average mass.
5. Calculate the weighted average mass of all candium particles by adding the relative masses. This weighted average mass is the atomic mass of candium.
6. Explain the difference between percent abundance and relative abundance. What is the result when you total the individual relative abundances? The individual percent abundances?
7. The percent abundance of each kind of candy tells you how many of each kind of candy there are in every 100 particles. What does relative abundance tell you?
8. Compare the total values for rows 3 and 6 in the table. Explain why the totals differ and why the value in row 6 best represents atomic mass.
9. Explain any differences between the atomic mass of your candium sample and that of your neighbor. Explain why the difference would be smaller if larger samples were used.

You're the Chemist

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Analyze It!** Determine the atomic mass of a second sample of candium. How does it compare with the first? Suggest reasons for any differences between the samples.
2. **Design It!** Design and test methods to produce identical samples of candium. Try measuring mass or volume as a means of counting. Test these methods by counting each kind of candy in each sample you produce. Which method of sampling gives the most consistent results?

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8. The total in row 3 is an average that ignores the relative abundances of particles. The total in row 6 is a weighted average that best represents atomic mass because it considers differences in mass and abundance among the particles.

9. Another student might not have had the same relative abundance of each candy.

You're the Chemist

1. Any differences are probably due to small variations in the numbers of each kind of candy in the samples, which affects the relative abundances.

2. The larger the samples, the better the results with any of the methods. Mass is likely to provide better results than volume.