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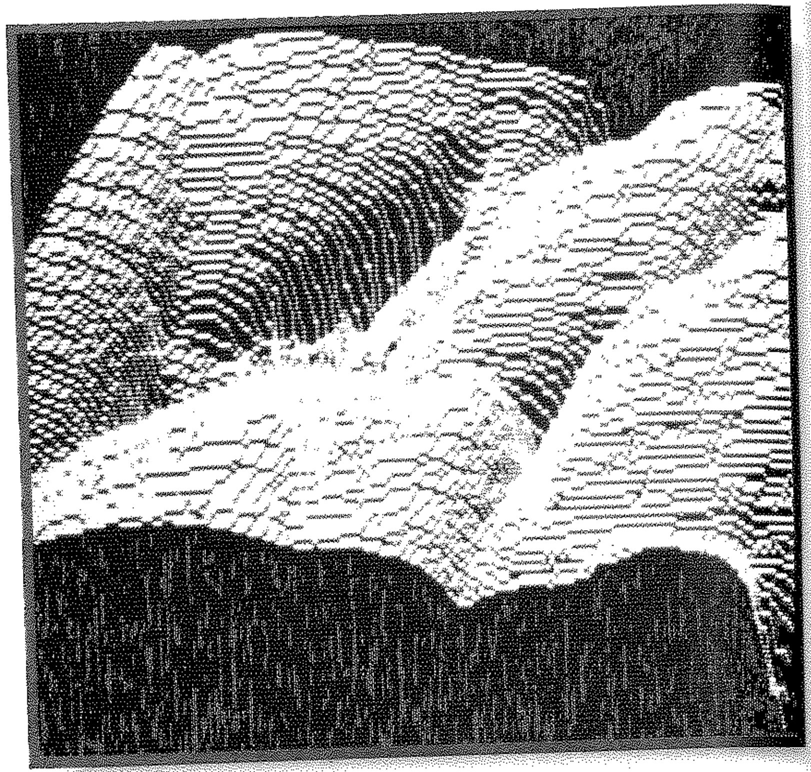


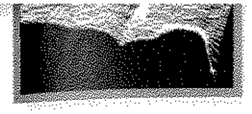
Each bright spot on this image represents one gold atom.

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?

used in Step 1 have the same



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?**

theory
- Describe the size of an atom

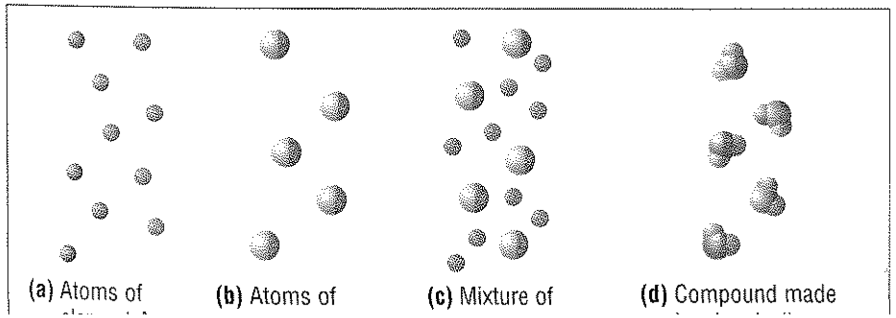
key terms
- Dalton's atomic theory
- atom

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.



(a) Atoms of
(b) Atoms of
(c) Mixture of
(d) Compound made

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.

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.

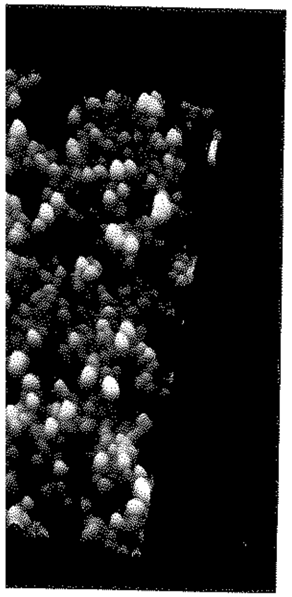


Figure 5.3

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?

ning 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?**

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

electrons, and neutrons in terms of relative mass and charge

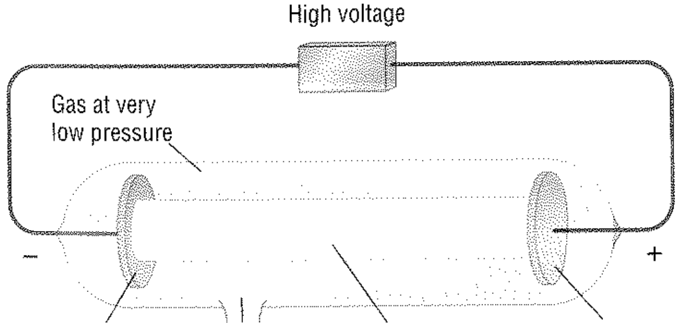
- Describe the structure of an atom, including the location of the protons, electrons, and neutrons with respect to the nucleus

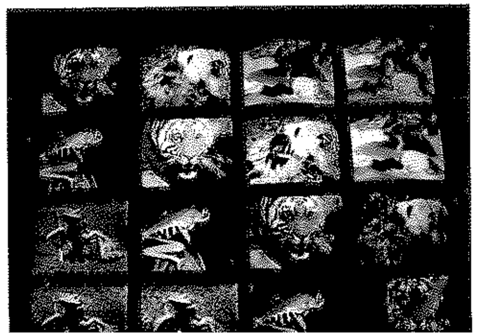
key terms

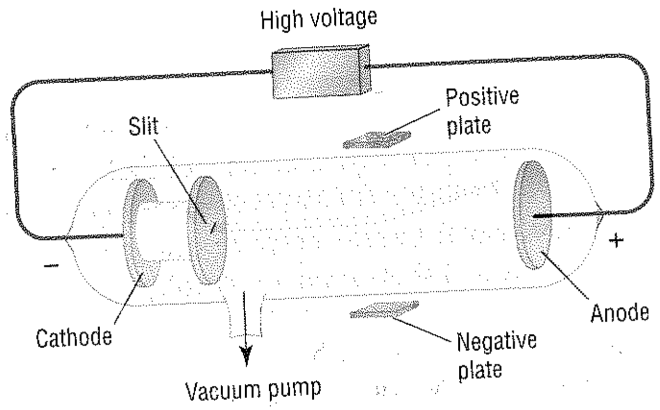
- electrons
- cathode ray
- protons
- neutrons
- nucleus

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.







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,

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.

Chem ASAP!

Animation 4

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

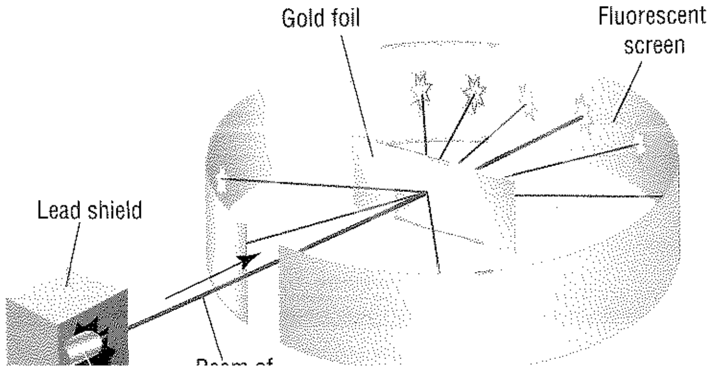


Figure 5.6

(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.*

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.**

: The Black Box

OCEDURE

Do not open the box.

Carefully manipulate the box so that the marble moves around the fixed object.

Gather data (clues) that describe the movement of the marble.

Sketch a picture of the object in the box, showing its shape, size, and location within the box.

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

relative masses of the three main subatomic

DISTINGUISHING BETWEEN ATOMS

"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

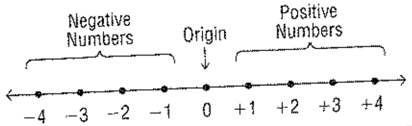
Atoms of the First Ten Elements

			Composition of the nucleus			
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

HEMath

AND NEGATIVE NUMBERS

ıced to protons and electrons, particles with charges of 1+ and 1− respectively. In
u will encounter a variety of positive and negative numbers throughout this course.
w of basic mathematical operations related to positive and negative numbers.

	EXAMPLE
ers are used in many appli- ıre scales. On a number line, the right of the *origin* (zero), ıre to the left of the origin.	
ıumber is its distance from are always positive (or zero).	|3| = (absolute value of 3) = 3 |−2| = (absolute value of −2) = 2
ər is a number having the ıe opposite sign.	The opposite of 8 is −8. The opposite of −5 is −(−5), or 5.
th positive, add them as usual.	5 + 16 = 21
th negative, add their he opposite.	To find −6 + (−5), note that 6 + 5 = 11. Therefore, −6 + (−5) = −11.
ɔpposite signs, subtract the m the larger absolute value. ər with the larger	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.
is the same as	12 − 23 = 12 + (−23) = −11 6 − (−4) = 6 + 4 = 10
DIVISION	
the same sign (both positive oduct or quotient is positive.	3 × 5 = 15 −24 ÷ (−8) = 3
opposite signs, the product	4 × (−3) = −12 −20 ÷ 5 = −4
TION	
n. The net charge of an ion is	If an ion has 3 protons and 5 electrons, its net charge is 3 − 5 = −2.

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:	*Unknowns:*
• atomic number = 7	• 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	____	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

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.

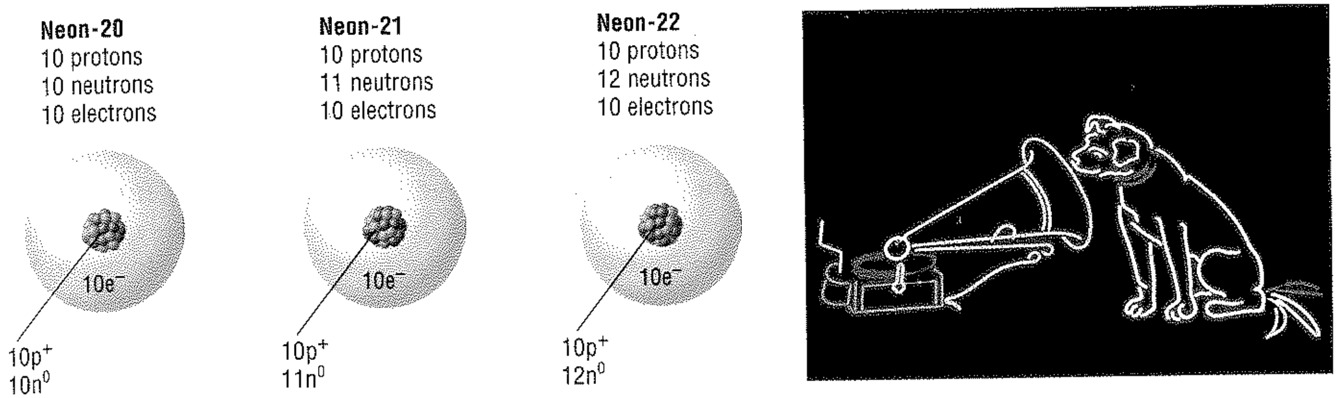


Figure 5.9
Neon-20, neon-21, and neon-22 are three isotopes of neon, a gaseous element used in lighted signs. How are these isotopes different? How are they the same?

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 ($^{1}_{1}H$) 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.

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

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 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

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

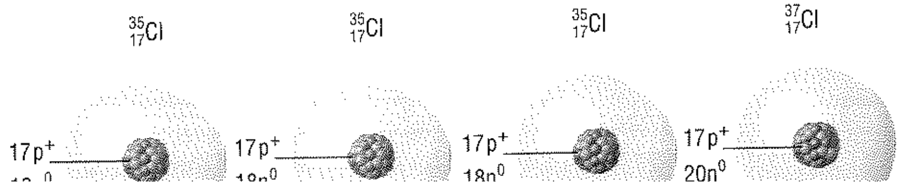


Table 5.3

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	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	
Zinc	${}^{64}_{30}Zn$	48.89	63.929	65.38
	${}^{66}_{30}Zn$	27.81	65.926	
	${}^{67}_{30}Zn$	4.11	66.927	
	${}^{68}_{30}Zn$	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

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.

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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

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: $10.012 \text{ amu} \times 0.1991 = 1.993 \text{ amu}$
   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

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.

portfolio project

An instrument called a mass spectrometer can be used to determine the