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4 ATOMIC STRUCTURE

SECTION 4.1 DEFINING THE ATOM (pages 101–103)

This section describes early atomic theories of matter and provides ways to understand the tiny size of individual atoms.

▶ Early Models of the Atom (pages 101–102)

1. Democritus, who lived in Greece during the fourth century B.C., suggested that matter is made up of tiny particles that cannot be divided. He called these particles ____________________ .

2. List two reasons why the ideas of Democritus were not useful in a scientific sense. ____

____

3. The modern process of discovery about atoms began with the theories of an English schoolteacher named ____________________ .

4. Circle the letter of each sentence that is true about Dalton's atomic theory.

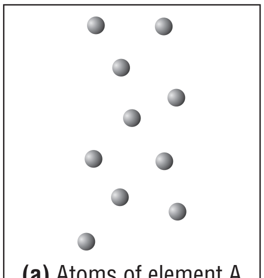
a. All elements are composed of tiny, indivisible particles called atoms.

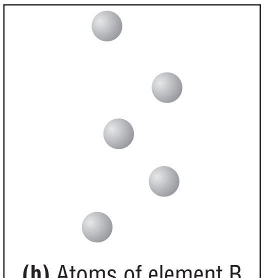
b. An element is composed of several types of atoms.

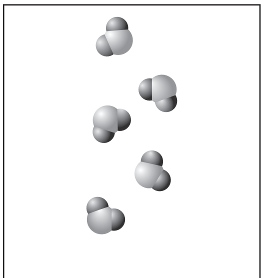
c. Atoms of different elements can physically mix together, or can chemically combine in simple, whole-number ratios to form compounds.

d. Chemical reactions occur when atoms are separated, joined, or rearranged; however, atoms of one element are never changed into atoms of another element by a chemical reaction.

5. In the diagram, use the labels *mixture* and *compound* to identify the mixture of elements A and B and the compound that forms when the atoms of elements A and B combine chemically.


(a) Atoms of element A


(b) Atoms of element B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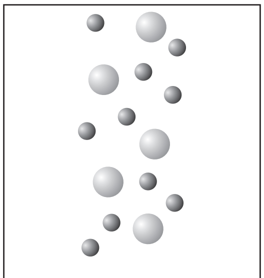
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▶ Sizing up the Atom (page 103)

6. Suppose you could grind a sample of the element copper into smaller and smaller particles. The smallest particle that could no longer be divided, yet still has the chemical properties of copper, is ________________________________ .

7. About how many atoms of copper when placed side by side would form a line 1 cm long? ______________________

SECTION 4.2 STRUCTURE OF THE NUCLEAR ATOM (pages 104–108)

This section describes the experiments that led to the discovery of subatomic particles and their properties.

▶ Subatomic Particles (pages 104–106)

1. How is the atomic theory that is accepted today different from Dalton's atomic theory? ____

____

2. Which subatomic particles carry a negative charge? ______________________

Match each term from the experiments of J. J. Thomson with the correct description.

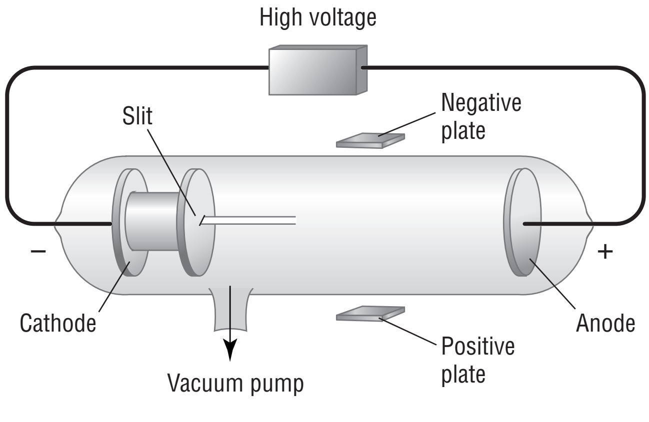
_______ **3.** anode **a.** an electrode with a negative charge

_______ **4.** cathode **b.** a glowing beam traveling between charged electrodes

_______ **5.** cathode ray **c.** an electrode with a positive charge

_______ **6.** electron **d.** a negatively charged particle

7. The diagram shows electrons moving from left to right in a cathode-ray tube. Draw an arrow showing how the path of the electrons will be affected by the placement of the negatively and positively charged plates.



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8. Thomson observed that the production of cathode rays did not depend on the kind of gas in the tube or the type of metal used for the electrodes. What conclusion did he draw from these observations?

____

9. What two properties of an electron did Robert Millikan determine from his experiments?

____

10. Circle the letter of each sentence that is true about atoms, matter, and electric charge.

   a. All atoms have an electric charge.

   b. Electric charges are carried by particles of matter.

   c. Electric charges always exist in whole-number multiples of a single basic unit.

   d. When a given number of positively charged particles combines with an equal number of negatively charged particles, an electrically neutral particle is formed.

11. Circle the letter next to the number of units of positive charge that remain if a hydrogen atom loses an electron.

   a. 0 **b.** 1 **c.** 2 **d.** 3

12. The positively charged subatomic particle that remains when a hydrogen atom loses an electron is called ______________ .

13. What charge does a neutron carry? ______________ .

14. Complete the table about the properties of subatomic particles.

Properties of Subatomic Particles

Particle	Symbol	Relative Electrical Charge	Relative Mass (mass of proton = 1)	Actual Mass (g)
Electron	e^-			9.11×10^{-28}
Proton	p^+			1.67×10^{-24}
Neutron	n^0			1.67×10^{-24}

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▶ The Atomic Nucleus (pages 106–108)

15. Is the following sentence true or false? An alpha particle has a double positive charge because it is a helium atom that has lost two electrons. ______________________

16. Explain why in 1911 Rutherford and his coworkers were surprised when they shot a narrow beam of alpha particles through a thin sheet of gold foil.

____

____

____

____

____

17. Circle the letter of each sentence that is true about the nuclear theory of atoms suggested by Rutherford's experimental results.

a. An atom is mostly empty space.

b. All the positive charge of an atom is concentrated in a small central region called the nucleus.

c. The nucleus is composed of protons.

d. The nucleus is large compared with the atom as a whole.

e. Nearly all the mass of an atom is in its nucleus.

SECTION 4.3 DISTINGUISHING AMONG ATOMS (pages 110–119)

This section explains how atomic number identifies an element; how to use atomic number and mass number to find the number of protons, neutrons, and electrons in an atom; how isotopes differ; and how to calculate average atomic mass.

▶ Atomic Number (page 110)

1. Circle the letter of the term that correctly completes the sentence. Elements are different because their atoms contain different numbers of _______ .

a. electrons

b. protons

c. neutrons

d. nuclei

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2. Complete the table showing the number of protons and electrons in atoms of six elements.

Atoms of Six Elements				
Name	**Symbol**	**Atomic Number**	**Number of Protons**	**Number of Electrons**
Hydrogen	H	1		
Helium	He		2	
Lithium	Li	3		
Boron	B	5		
Carbon	C	6		
Oxygen	O			8

▶ Mass Number (pages 111–112)

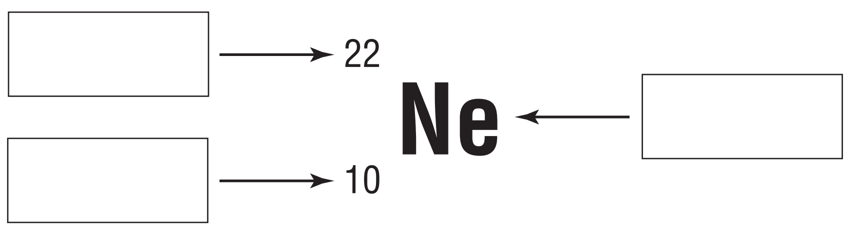
3. The total number of protons and neutrons in an atom is its

____________________ .

4. What is the mass number of a helium atom that has two protons and two neutrons? ____________________

5. How many neutrons does a beryllium atom with four protons and a mass number of nine have? ____________________

6. Place the labels *chemical symbol, atomic number,* and *mass number* in the shorthand notation below.



7. Designate the atom shown in Question 6 in the form "name of element"-"mass number." ____________________

8. How many protons, neutrons, and electrons are in the atom discussed in Questions 6 and 7? Protons: ______ Neutrons: ______ Electrons: ______

▶ Isotopes (pages 112–113)

9. How do atoms of neon-20 and neon-22 differ?

____

____

10. Neon-20 and neon-22 are called ______________ .

11. Is the following sentence true or false? Isotopes are chemically alike because they have identical numbers of protons and electrons. ______________

Match the designation of each hydrogen isotope with its commonly used name.

_____ **12.** hydrogen-1 **a.** tritium

_____ **13.** hydrogen-2 **b.** hydrogen

_____ **14.** hydrogen-3 **c.** deuterium

▶ Atomic Mass (pages 114–117)

15. Why is the atomic mass unit (amu), rather than the gram, usually used to express atomic mass?

____

____

16. What isotope of carbon has been chosen as the reference isotope for atomic mass units? What is the defined atomic mass in amu of this isotope?

____

17. Is the following sentence true or false? The atomic mass of an element is always a whole number of atomic mass units. ______________

18. Circle the letter of each statement that is true about the average atomic mass of an element and the relative abundance of its isotopes.

a. In nature, most elements occur as a mixture of two or more isotopes.

b. Isotopes of an element do not have a specific natural percent abundance.

c. The average atomic mass of an element is usually closest to that of the isotope with the highest natural abundance.

d. Because hydrogen has three isotopes with atomic masses of about 1 amu, 2 amu, and 3 amu, respectively, the average atomic mass of natural hydrogen is 2 amu.

19. Circle the letter of the correct answer. When chlorine occurs in nature, there are three atoms of chlorine-35 for every one atom of chlorine-37. Which atomic mass number is closer to the average atomic mass of chlorine?

a. 35 amu **b.** 37 amu

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20. In the periodic table, the elements are organized into groups based on

____ .

▶ The Periodic Table—A Preview (page 118)

21. What are the horizontal rows in the periodic table called?

____

Reading Skill Practice

Outlining can help you understand and remember what you have read. Prepare an outline of Section 4.3, *Distinguishing Among Atoms.* Begin with the headings in the textbook. Under each heading, write the main idea. Then list the details that support the main idea. Do your work on a separate sheet of paper.

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GUIDED PRACTICE PROBLEMS

Fill in the write-on lines and boxes provided as you work through the guided practice problems.

GUIDED PRACTICE PROBLEM 18 (page 112)

18. Use Table 4.2 to express the compositions of carbon-12, fluorine-19, and beryllium-9 in shorthand notation.

Analyze

Carbon-12

Step 1. The number of protons in an atom is called its ____________________ number.

The number of protons in an atom of carbon-12 is __________ .

Calculate

Step 2. The number of protons plus the number of neutrons in an atom is called its ____________________ number. For carbon-12, this number is __________ .

Step 3. The shorthand notation for carbon-12 is:

mass number → $\boxed{}$

C

atomic number → $\boxed{}$

Evaluate

Step 4. Except for hydrogen-1, the mass number of an isotope is always greater than its atomic number. Is the mass number reasonable? __________

Fluorine-19

Step 1. The atomic number of fluorine-19 is __________ .

Step 2. Its mass number is __________ .

Step 3. The shorthand notation for fluorine-19 is: $^{\boxed{}}_{\boxed{}}\mathrm{F}$

Step 4. Is your answer reasonable? Why?

____

____

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Beryllium-9

Step 1. The atomic number of beryllium-9 is __________ .

Step 2. Its mass number is __________ .

Step 3. The shorthand notation for beryllium-9 is: $^{\square}_{\square}\text{Be}$

Step 4. Is your answer reasonable? Why?

____

____

EXTRA PRACTICE (similar to Practice Problem 19, page 113)

19. Three isotopes of sulfur are sulfur-32, sulfur-33, and sulfur-34. Write the complete symbol for each isotope, including the atomic number and the mass number.

sulfur-32	sulfur-33	sulfur-34
$^{\square}_{\square}\text{S}$	$^{\square}_{\square}\text{S}$	$^{\square}_{\square}\text{S}$

GUIDED PRACTICE PROBLEM 23 (page 117)

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.

Analyze

Step 1. Will the average atomic mass be closer to 63 or to 65? Explain.

____

____

Solve

Step 2. For Cu-63: $69.2\% \times 62.93\text{ amu} = 0.692 \times 62.93\text{ amu} = \square$

Step 3. For Cu-65: $30.8\% \times 64.93\text{ amu} = \square \times \square = \square$

Step 4. Average mass: $43.6\text{ amu} + \square = \square$

Evaluate

Step 5. Explain why your answer is reasonable.

____

EXTRA PRACTICE (similar to Practice Problem 24, page 117)

24. Calculate the atomic mass of rubidium. The two isotopes of rubidium have atomic masses and relative abundancies of 84.91 amu (72.16%) and 86.91 amu (27.84%). ____________________