4.9 Atomic Mass: The Average Mass of an Element's Atoms

An important part of Dalton's atomic theory was that all atoms of a given element have the same mass. But as we just learned, the atoms of a given element may have different masses (because of isotopes). So Dalton was not completely correct. We can, however, calculate an average mass—called the **atomic mass**—for each

CHEMISTRY IN THE ENVIRONMENT

Radioactive Isotopes at Hanford, Washington

Nuclei of the isotopes of a given element are not all equally stable. For example, naturally occurring lead is composed primarily of Pb-206, Pb-207, and Pb-208. Other isotopes of lead also exist, but their nuclei are unstable. Scientists can make some of these other isotopes, such as Pb-185, in the laboratory. However, within seconds Pb-185 atoms emit a few energetic subatomic particles from their nuclei and change into different isotopes of different elements (which are themselves unstable). These emitted subatomic particles are called **nuclear radiation**, and the isotopes that emit them are termed **radioactive**. Nuclear radiation, always associated with unstable nuclei, can be harmful to humans and other living organisms because the energetic particles interact with and damage biological molecules. Some isotopes, such as Pb-185, emit significant amounts of radiation only for a very short time. Others, however, remain radioactive for a long time—in some cases millions or even billions of years.

The nuclear power and nuclear weapons industries produce by-products containing unstable isotopes of several different elements. Many of these isotopes emit nuclear radiation for a long time, and their disposal is an environmental problem. For example, in Hanford, Washington, which for 50 years produced fuel for nuclear weapons, 177 underground storage tanks contain 55 million gallons of highly radioactive nuclear waste. Certain radioactive isotopes within that waste will produce nuclear radiation for the foreseeable future. Unfortunately, some of the underground storage tanks are aging, and leaks have allowed some of the waste to seep into the environment. While the danger from short-term external exposure to this waste is minimal, ingestion of the waste through contamination of drinking water or food supplies would pose significant health risks. Consequently, Hanford is now the site of the largest environmental cleanup project in U.S. history. The U.S. government expects the project to take more than 20 years and cost about $10 billion.

Radioactive isotopes are not always harmful, however, and many have beneficial uses. For example, technetium-99 (Tc-99) is often given to patients to diagnose disease. The radiation emitted by Tc-99 helps doctors image internal organs or detect infection.

CAN YOU ANSWER THIS? *Give the number of neutrons in each of the following isotopes: Pb-206, Pb-207, Pb-208, Pb-185, Tc-99.*

◀ Storage tanks at Hanford, Washington, contain 55 million gallons of high-level nuclear waste. Each tank pictured here holds 1 million gallons.

element. The atomic mass of each element is listed in the periodic table directly beneath the element's symbol; it represents the average mass of the atoms that compose that element. For example, the periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77% chlorine-35 (mass 34.97 amu) and 24.23% chlorine-37 (mass 36.97 amu). Its atomic mass is:

Some books call this *average atomic mass* or atomic weight instead of simply *atomic mass*.

$$\text{Atomic mass} = (0.7577 \times 34.97\ \text{amu}) + (0.2423 \times 36.97\ \text{amu})$$
$$= 35.45\ \text{amu}$$

Notice that the atomic mass of chlorine is closer to 35 than 37 because naturally occurring chlorine contains more chlorine-35 atoms than chlorine-37 atoms. Notice also that when percentages are used in these calculations, they must always be converted to their decimal value. To convert a percentage to its decimal value, divide by 100. For example:

$$75.77\% = 75.77/100 = 0.7577$$
$$24.33\% = 24.23/100 = 0.2423$$

In general, atomic mass is calculated according to the following equation:

$$\begin{aligned}\text{Atomic mass} = {} & (\text{Fraction of isotope 1} \times \text{Mass of isotope 1}) + \\ & (\text{Fraction of isotope 2} \times \text{Mass of isotope 2}) + \\ & (\text{Fraction of isotope 3} \times \text{Mass of isotope 3}) + \ldots\end{aligned}$$

where the fractions of each isotope are the percent natural abundances converted to their decimal values. Atomic mass is useful because it allows us to assign a characteristic mass to each element and, as we will see in Chapter 6, it allows us to quantify the number of atoms in a sample of that element.

EXAMPLE 4.9 Calculating Atomic Mass

Gallium has two naturally occurring isotopes: Ga-69 with mass 68.9256 amu and a natural abundance of 60.11%, and Ga-71 with mass 70.9247 amu and a natural abundance of 39.89%. Calculate the atomic mass of gallium.

	SOLUTION
Convert the percent natural abundances into decimal form by dividing by 100.	$\text{Fraction Ga-69} = \dfrac{60.11}{100} = 0.6011$ $\text{Fraction Ga-71} = \dfrac{39.89}{100} = 0.3989$
Use the fractional abundances and the atomic masses of the isotopes to compute the atomic mass according to the atomic mass definition given earlier.	$\text{Atomic mass} = (0.6011 \times 68.9256\ \text{amu}) + (0.3989 \times 70.9247\ \text{amu})$ $= 41.4321\ \text{amu} + 28.2919\ \text{amu}$ $= 69.7231 = 69.72\ \text{amu}$

▶SKILLBUILDER 4.9 | Calculating Atomic Mass

Magnesium has three naturally occurring isotopes with masses of 23.99, 24.99, and 25.98 amu and natural abundances of 78.99%, 10.00%, and 11.01%. Calculate the atomic mass of magnesium.

▶FOR MORE PRACTICE Example 4.14; Problems 95, 96.

CONCEPTUAL CHECKPOINT 4.6

A fictitious element is composed of isotopes A and B with masses of 61.9887 and 64.9846 amu, respectively. The atomic mass of the element is 64.52. What can you conclude about the natural abundances of the two isotopes?

(a) The natural abundance of isotope A must be greater than the natural abundance of isotope B.

(b) The natural abundance of isotope B must be greater than the natural abundance of isotope A.

(c) The natural abundances of both isotopes must be about equal.

(d) Nothing can be concluded about the natural abundances of the two isotopes from the given information.