Some additional examples of the atomic masses of the naturally occurring isotopes of several elements are given in **Table 4** on the next page. Isotopes of an element may occur naturally, or they may be made in the laboratory (*artificial isotopes*). *Although isotopes have different masses, they do not differ significantly in their chemical behavior*.

The masses of subatomic particles can also be expressed on the atomic mass scale (see **Table 1**). The mass of the electron is 0.000 5486 amu, that of the proton is 1.007 276 amu, and that of the neutron is 1.008 665 amu. Note that the proton and neutron masses are close to but not equal to 1 amu. You have learned that the mass number is the total number of protons and neutrons that make up the nucleus of an atom. You can now see that the mass number and relative atomic mass of a given nuclide are quite close to each other. They are not identical because the proton and neutron masses deviate slightly from 1 amu and the atomic masses include electrons. Also, as you will read in Chapter 21, a small amount of mass is changed to energy in the creation of a nucleus from its protons and neutrons.

HISTORICAL CHEMISTRY

Discovery of Element 43

The discovery of element 43, technetium, is credited to Carlo Perrier and Emilio Segrè, who artificially produced it in 1937. However, in 1925, a German chemist named Ida Tacke reported the discovery of element 43, which she called masurium, in niobium ores. At the time, her discovery was not accepted because it was thought technetium could not occur naturally. Recent studies confirm that Tacke and coworkers probably did discover element 43.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as indicated in **Table 4.** The percentage of each isotope in the naturally occurring element on Earth is nearly always the same, no matter where the element is found. The percentage at which each of an element's isotopes occurs in nature is taken into account when calculating the element's average atomic mass. **Average atomic mass** *is the weighted average of the atomic masses of the naturally occurring isotopes of an element.*

The following is a simple example of how to calculate a *weighted average.* Suppose you have a box containing two sizes of marbles. If 25% of the marbles have masses of 2.00 g each and 75% have masses of 3.00 g each, how is the weighted average calculated? You could count the number of each type of marble, calculate the total mass of the mixture, and divide by the total number of marbles. If you had 100 marbles, the calculations would be as follows.

$$25 \text{ marbles} \times 2.00 \text{ g} = 50 \text{ g}$$
$$75 \text{ marbles} \times 3.00 \text{ g} = 225 \text{ g}$$

Adding these masses gives the total mass of the marbles.

$$50 \text{ g} + 225 \text{ g} = 275 \text{ g}$$

Dividing the total mass by 100 gives an average marble mass of 2.75 g.

A simpler method is to multiply the mass of each marble by the decimal fraction representing its percentage in the mixture. Then add the products.

$$25\% = 0.25 \qquad 75\% = 0.75$$
$$(2.00 \text{ g} \times 0.25) + (3.00 \text{ g} \times 0.75) = 2.75 \text{ g}$$