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

An important part of Dalton's atomic theory is that all atoms of a given element have the same mass. In Section 1.8, we learned that because of isotopes, the atoms of a given element often have different masses, so Dalton was not completely correct. We can, however, calculate an average mass—called the **atomic mass**—for each element.

> Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*

The atomic mass of each element is listed directly beneath the element's symbol in the periodic table and represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*. For example, the periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). We can calculate its atomic mass:

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

Naturally occurring chlorine contains more chlorine-35 atoms than chlorine-37 atoms, so the weighted average mass of chlorine is closer to 35 amu than to 37 amu.

17
Cl
35.45
chlorine

> In this book, we use the atomic masses recommended by the International Union of Pure and Applied Chemistry (IUPAC) for users seeking an atomic mass value for an unspecified sample. Detailed studies of the atomic masses of many samples, however, have shown that atomic masses are not constants of nature because the exact isotopic abundances in any given sample depend on the history of the sample.

We generally calculate the atomic mass with the equation:

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

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

> When percentages are used in calculations, we convert them to their decimal value by dividing by 100.

Example 1.4 Atomic Mass

Copper has two naturally occurring isotopes: Cu-63 with mass 62.9396 amu and a natural abundance of

Copper has two naturally occurring isotopes: Cu-63 with mass 62.9396 amu and a natural abundance of 69.17%, and Cu-65 with mass 64.9278 amu and a natural abundance of 30.83%. Calculate the atomic mass of copper.

SOLUTION

Convert the percent natural abundances into decimal form by dividing by 100.

$$\text{Fraction Cu-63} = \frac{69.17}{100} = 0.6917$$
$$\text{Fraction Cu-65} = \frac{30.83}{100} = 0.3083$$

Calculate the atomic mass using the equation given in the text. (See Section E.4 for significant figure and rounding conventions.)

$$\begin{aligned}\text{Atomic mass} &= 0.6917(62.9396\text{ amu}) + 0.3083(64.9278\text{ amu})\\ &= 43.5\underline{3}53\text{ amu} + 20.0\underline{1}72\text{ amu} = 63.5\underline{5}25 = 63.55\text{ amu}\end{aligned}$$

The magnitude of the answer makes sense given that approximately two-thirds of the atoms have a mass of nearly 63 amu and one-third have a mass of nearly 65. The weighted average should be closer to 63 than 65.

FOR PRACTICE 1.4 Magnesium has three naturally occurring isotopes with masses of 23.99 amu, 24.99 amu, and 25.98 amu and natural abundances of 78.99%, 10.00%, and 11.01%, respectively. Calculate the atomic mass of magnesium.

FOR MORE PRACTICE 1.4 Gallium has two naturally occurring isotopes: Ga-69 with a mass of 68.9256 amu and a natural abundance of 60.11%, and Ga-71. Use the atomic mass of gallium from the periodic table to find the mass of Ga-71.

Interactive Worked Example 1.4 **Atomic Mass**

Conceptual Connection 1.8 **Atomic Mass**

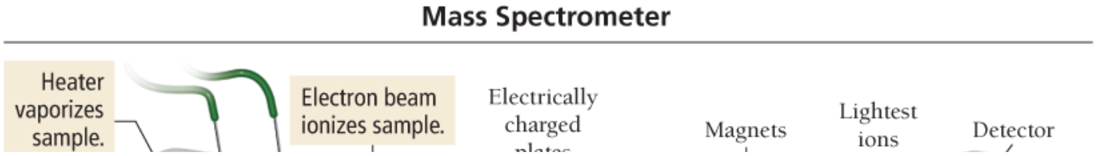
Interactive

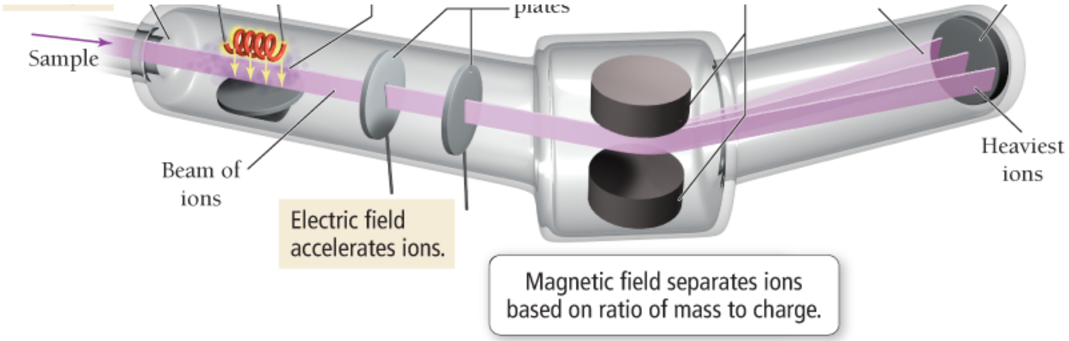
Mass Spectrometry: Measuring the Mass of Atoms and Molecules

The masses of atoms and the percent abundances of isotopes of elements are measured using **mass spectrometry**, a technique that separates particles according to their mass. In a mass spectrometer, such as the one in Figure 1.11, the sample (containing the atoms whose mass is to be measured) is injected into the instrument and vaporized. The vaporized atoms are ionized by an electron beam—the electrons in the beam collide with the atoms, removing electrons and creating positively charged ions. The ions are then accelerated into a magnetic field. When ions drift through a magnetic field, they experience a force that bends their trajectory. The amount of bending depends on the mass (and charge) of the ions—the trajectories of lighter ions are bent more than those of heavier ones (of the same charge).

Figure 1.11 **The Mass Spectrometer**

Atoms are converted to positively charged ions, accelerated, and passed through a magnetic field that deflects their path. The heaviest ions undergo the least deflection.

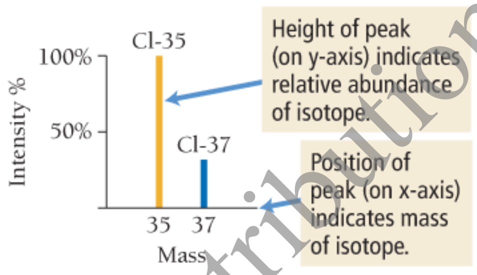



plates
Sample
Beam of ions
Electric field accelerates ions.
Magnetic field separates ions based on ratio of mass to charge.
Heaviest ions

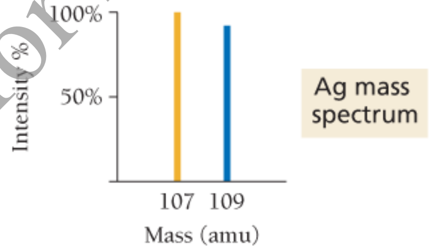
In the right side of the spectrometer shown in Figure 1.11, you can see three different paths, each corresponding to ions of different mass. Finally, the ions strike a detector and produce an electrical signal that is recorded. The result is the separation of the ions according to their mass, producing a mass spectrum such as the one in Figure 1.12. The *position* of each peak on the *x*-axis indicates the *mass of the isotope* that was ionized, and the *intensity* (indicated by the height of the peak) indicates the *relative abundance of that isotope*.

Figure 1.12 The Mass Spectrum of Chlorine

The position of each peak on the *x*-axis indicates the mass of the isotope. The intensity (or height) of the peak indicates the relative abundance of the isotope. The intensity of the highest peak is usually set to 100%, and the intensity of all other peaks is reported relative to the most intense one.



We can use the mass spectrum of an elemental sample to determine the atomic mass of that sample of the element. For example, consider the mass spectrum of a naturally occurring sample of silver:



The two peaks correspond to the two naturally occurring isotopes of silver. We determine the percent abundance of each isotope from the intensity of each line. However, the *total* intensity must be *normalized*—it must be made to equal 100%. We accomplish this by dividing the intensity of each peak by the total intensity:

$$\text{abundance of Ag-107} = \frac{100.0\%}{100.0\% + 92.90\%} \times 100\% = 51.84\%$$

$$\text{abundance of Ag-109} = \frac{92.90\%}{100.0\% + 92.90\%} \times 100\% = 48.16\%$$

Then we calculate the atomic mass of silver as usual.

$$\begin{aligned}\text{Ag atomic mass} &= 0.5184(106.905\ \text{amu}) + 0.4816\,(108.904\ \text{amu})\\ &= 55.4\underline{1}95\ \text{amu} + 52.4\underline{4}82\ \text{amu} = 107.8\underline{6}77 = 107.87\ \text{amu}\end{aligned}$$

We can also use mass spectrometry on molecules. Because molecules often fragment (break apart) during ionization, the mass spectrum of a molecule usually contains many peaks representing the masses of different parts of the molecule, as well as a peak representing the mass of the molecule as a whole. The fragments that form upon ionization, and therefore the corresponding peaks that appear in the mass spectrum, are specific to the molecule, so that a mass spectrum is like a molecular fingerprint. Mass spectroscopy can be used to identify an unknown molecule and to determine how much of it is present in a particular sample.

Since the early 1990s, researchers have successfully applied mass spectrometry to biological molecules, including proteins (the workhorse molecules in cells) and nucleic acids (the molecules that carry genetic information). For a long time, these molecules could not be analyzed by mass spectrometry because they were difficult to vaporize and ionize without destroying them, but modern techniques have overcome this problem. A tumor, for example, can now be instantly analyzed by mass spectrometry to determine whether it contains specific proteins associated with cancer.