Pearson

1.5: Modern Atomic Theory and the Laws That Led to It

Key Concept Video **Atomic Theory**

Like most theories, Dalton's theory that all matter is composed of atoms grew out of observations and laws. The three most important laws that led to the development and acceptance of the atomic theory are the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

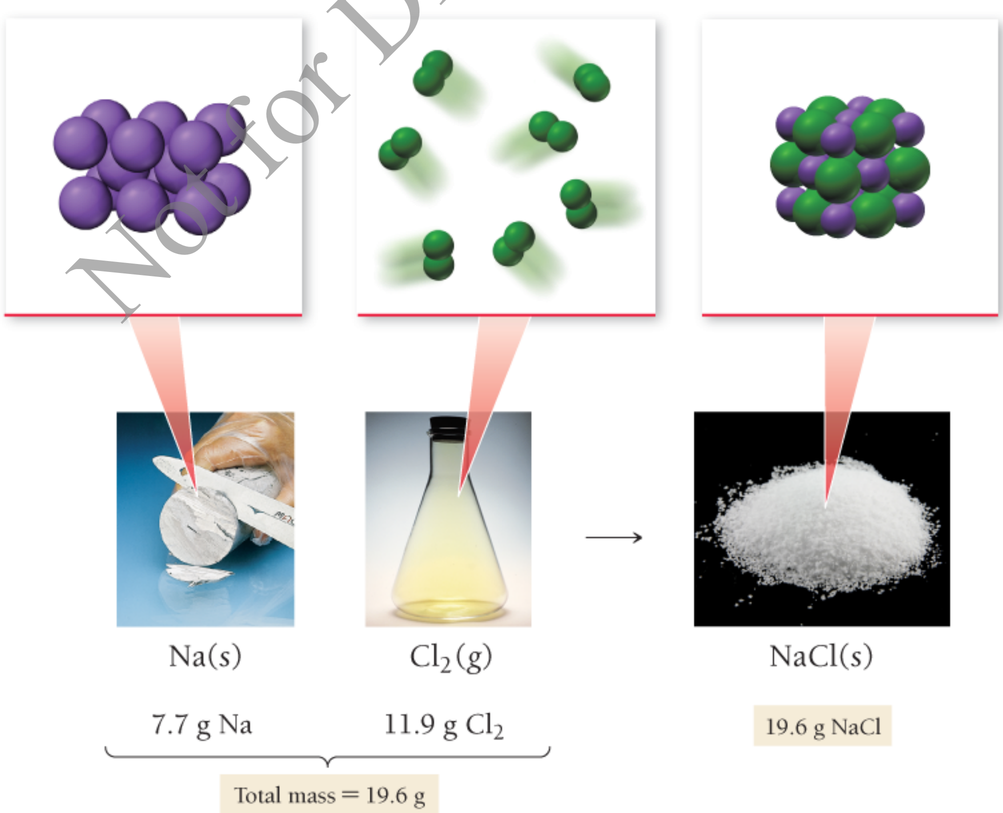
The Law of Conservation of Mass

Recall from Section 1.3 that in 1789 Antoine Lavoisier studied combustion and formulated the law of conservation of mass, which states:

> **In a chemical reaction, matter is neither created nor destroyed.**

> We will see in Chapter 20 that this law is a slight oversimplification. However, the changes in mass in ordinary chemical processes are so minute that we can ignore them for all practical purposes.

A **chemical reaction** (discussed more fully in Chapter 7) is a process in which one or more substances are converted into one or more different substances. The law of conservation of mass states that when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change. For example, consider the reaction between sodium and chlorine to form sodium chloride shown here:



Mass of reactants = Mass of product

The combined mass of the sodium and chlorine that react (the reactants) exactly equals the mass of the sodium chloride that forms (the product). This law is consistent with the idea that matter is composed of small, indestructible particles. The particles rearrange during a chemical reaction, but the number of particles is conserved because the particles themselves are indestructible (at least by chemical means).

Conceptual Connection 1.3 **The Law of Conservation of Mass**

The Law of Definite Proportions

Interactive

In 1797, a French chemist named Joseph Proust (1754–1826) made observations on the composition of compounds. He found that the elements composing a given compound always occur in fixed (or definite) proportions in all samples of the compound. In contrast, the components of a mixture can be present in any proportions whatsoever. He summarized his observations in the **law of definite proportions**:

> **All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.**

> The law of definite proportions is sometimes called the law of constant composition.

For example, the decomposition of 18.0 g of water results in 16.0 g of oxygen and 2.0 g of hydrogen, or an oxygen-to-hydrogen mass ratio of:

$$\text{mass ratio} = \frac{16.0 \text{ g O}}{2.0 \text{ g H}} = 8.0 \text{ or } 8:1$$

This ratio holds for any sample of pure water, regardless of its origin. The law of definite proportions applies to all compounds. Consider ammonia, a compound composed of nitrogen and hydrogen. Ammonia contains 14.0 g of nitrogen for every 3.0 g of hydrogen, resulting in a nitrogen-to-hydrogen mass ratio of:

$$\text{mass ratio} = \frac{14.0 \text{ g N}}{3.0 \text{ g H}} = 4.7 \text{ or } 4.7:1$$

Again, this ratio is the same for every sample of ammonia. The law of definite proportions hints at the idea that matter is composed of atoms. Compounds have definite proportions of their constituent elements because the atoms that compose them, each with its own specific mass, occur in a definite ratio. Because the ratio of atoms is the same for all samples of a particular compound, the ratio of masses is also the same.

Example 1.1 Law of Definite Proportions

Two samples of carbon dioxide decompose into their constituent elements. One sample produces 25.6 g of oxygen and 9.60 g of carbon, and the other produces 21.6 g of oxygen and 8.10 g of carbon. Show that these results are consistent with the law of definite proportions.

SOLUTION

To show this, calculate the mass ratio of one element to the other for both samples by dividing the mass of one element by the mass of the other. For convenience, divide the larger mass by the smaller one.

For the first sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{25.6}{9.60} = 2.67 \text{ or } 2.67:1$$

For the second sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{21.6}{8.10} = 2.67 \text{ or } 2.67 : 1$$

The ratios are the same for the two samples, so these results are consistent with the law of definite proportions.

FOR PRACTICE 1.1 Two samples of carbon monoxide decompose into their constituent elements. One sample produces 17.2 g of oxygen and 12.9 g of carbon, and the other sample produces 10.5 g of oxygen and 7.88 g of carbon. Show that these results are consistent with the law of definite proportions.

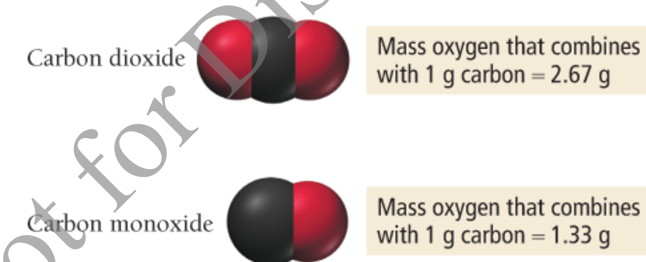
Answers to For Practice and For More Practice Problems are in Appendix IV.

The Law of Multiple Proportions

In 1804, John Dalton published his **law of multiple proportions**:

> **When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.**

Dalton already suspected that matter was composed of atoms, so that when two elements A and B combine to form more than one compound, an atom of A combines with either one, two, three, or more atoms of B $(AB_1, AB_2, AB_3, \text{etc.})$. Therefore, the masses of B that react with a fixed mass of A are always related to one another as small whole-number ratios. Consider the compounds carbon monoxide and carbon dioxide, two compounds composed of the same two elements: carbon and oxygen. We saw in Example 1.1 that the mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon. In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.



The ratio of these two masses is a small whole number.

$$\frac{\text{mass oxygen to 1 g carbon in carbon dioxide}}{\text{mass oxygen to 1 g carbon in carbon monoxide}} = \frac{2.67}{1.33} = 2$$

With the help of the molecular models, we can see why the ratio is 2:1—carbon dioxide contains two oxygen atoms to every carbon atom, while carbon monoxide contains only one. Of course, neither John Dalton nor Joseph Proust had access to any kind of modern instrumentation that could detect individual atoms—Dalton supported his atomic ideas primarily by using the *weights* of samples. *But the weights implied that matter was ultimately particulate; what else would explain why these ratios were always whole numbers?*

Example 1.2 Law of Multiple Proportions

Nitrogen forms several compounds with oxygen, including nitrogen dioxide and dinitrogen monoxide. Measurements of the masses of nitrogen and oxygen that form upon decomposing these compounds show that nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains that nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that these results are consistent with the law of multiple proportions.

SOLUTION

To show this, calculate the ratio of the mass of oxygen from one compound to the mass of oxygen in the other. Always divide the larger of the two masses by the smaller one.

$$\frac{\begin{array}{c}\text{mass oxygen to 1 g}\\ \text{nitrogen in nitrogen dioxide}\end{array}}{\begin{array}{c}\text{mass oxygen to 1 g}\\ \text{nitrogen in dinitrogen monoxide}\end{array}} = \frac{2.28}{0.570} = 4.00$$

The ratio is a small whole number (4); these results are consistent with the law of multiple proportions.

FOR PRACTICE 1.2 Hydrogen and oxygen form both water and hydrogen peroxide. The decomposition of a sample of water forms 0.125 g hydrogen to every 1.00 g oxygen. The decomposition of a sample of hydrogen peroxide forms 0.0625 g hydrogen to every 1.00 g oxygen. Show that these results are consistent with the law of multiple proportions.

Conceptual Connection 1.4 **The Laws of Definite and Multiple Proportions**

John Dalton and the Atomic Theory

Interactive

In 1808, John Dalton explained the laws discussed in this section with his **atomic theory**, which states that:

1. Each element is composed of tiny, indestructible particles called atoms.
2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
3. Atoms combine in simple, whole-number ratios to form compounds.
4. Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms only change the way that they are *bound together* with other atoms.

The most important idea presented in this section is that measurements of the relative weights of matter samples in three categories (of samples before and after a reaction; of different samples of the same compound; and of different compounds composed of the same elements) indicate that matter is particulate. These scientists—with the help of a balance—gathered and interpreted data that settled an age-old question: Is matter continuous or particulate? Thanks to the careful observations they made, the evidence confirmed the particulate view.