Note in each case that the ratio of the masses is always 3 to 1,
\[
\frac{1500}{500} = \frac{150}{50} = \frac{15}{5} = \frac{3}{1}
\]
which is the ratio of the masses of the individual components:
\[
\frac{\text{Mass of mint}}{\text{Mass of jelly bean}} = \frac{15}{5} = \frac{3}{1}
\]
Any two samples, one of mints and one of jelly beans, that have a mass ratio of 15/5 = 3/1 will contain the same number of components. And these same ideas apply also to atoms, as we will see in the next section.

### Atomic Masses: Counting Atoms by Weighing

**Objective:** To understand atomic mass and its experimental determination.

The balanced chemical equation for the reaction of solid carbon and gaseous oxygen to form gaseous carbon dioxide is as follows:

\[
C(s) + O_2(g) \rightarrow CO_2(g)
\]

Now suppose you have a small pile of solid carbon and want to know how many oxygen molecules are required to convert all of this carbon into carbon dioxide. The balanced equation tells us that one oxygen molecule is required for each carbon atom.

\[
C(s) + O_2(g) \rightarrow CO_2(g)
\]

1 atom reacts with 1 molecule to yield 1 molecule

To determine the number of oxygen molecules required, we must know how many carbon atoms are present in the pile of carbon. But individual atoms are far too small to see. We must learn to count atoms by weighing samples containing large numbers of them.

In the last section we saw that we can easily count things like jelly beans and mints by weighing. Exactly the same principles can be applied to counting atoms.

Because atoms are so tiny, the normal units of mass—the gram and the kilogram—are much too large to be convenient. For example, the mass of a single carbon atom is \(1.99 \times 10^{-23}\) g. To avoid using terms like \(10^{-23}\) when describing the mass of an atom, scientists have defined a much smaller unit of mass called the **atomic mass unit**, which is abbreviated **amu**. In terms of grams,

\[
1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}
\]

Now let’s return to our problem of counting carbon atoms. To count carbon atoms by weighing, we need to know the mass of individual atoms, just as we needed to know the mass of the individual jelly beans. Recall from Chapter 3 that the atoms of a given element exist as isotopes. The isotopes of carbon are \(^{12}\text{C}, \(^{13}\text{C}, \text{and }^{14}\text{C. Any sample of carbon contains a mixture of these isotopes, always in the same proportions. Each of these isotopes has a slightly different mass. Therefore, just as with the nonidentical jelly beans,} \)

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we need to use an average mass for the carbon atoms. The **average atomic mass** for carbon atoms is 12.01 amu. This means that any sample of carbon from nature can be treated as though it were composed of identical carbon atoms, each with a mass of 12.01 amu. Now that we know the average mass of the carbon atom, we can count carbon atoms by weighing samples of natural carbon. For example, what mass of natural carbon must we take to have 1000 carbon atoms present? Because 12.01 amu is the average mass,

\[
\text{Mass of 1000 natural carbon atoms} = (1000 \text{ atoms}) \left( \frac{12.01 \text{ amu}}{1 \text{ atom}} \right)
\]

\[
= 12,010 \text{ amu} = 1.201 \times 10^4 \text{ amu}
\]

Now let’s assume that when we weigh the pile of natural carbon mentioned earlier, the result is $3.00 \times 10^{20}$ amu. How many carbon atoms are present in this sample? We know that an average carbon atom has the mass 12.01 amu, so we can compute the number of carbon atoms by using the equivalence statement

\[
1 \text{ carbon atom} = 12.01 \text{ amu}
\]

to construct the appropriate conversion factor,

\[
\frac{1 \text{ carbon atom}}{12.01 \text{ amu}}
\]

The calculation is carried out as follows:

\[
3.00 \times 10^{20} \text{ amu} \times \frac{1 \text{ carbon atom}}{12.01 \text{ amu}} = 2.50 \times 10^{19} \text{ carbon atoms}
\]

The principles we have just discussed for carbon apply to all the other elements as well. All the elements as found in nature typically consist of a mixture of various isotopes. So to count the atoms in a sample of a given element by weighing, we must know the mass of the sample and the average mass for that element. Some average masses for common elements are listed in Table 6.1.

### Table 6.1

<table>
<thead>
<tr>
<th>Element</th>
<th>Average Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>1.008</td>
</tr>
<tr>
<td>Carbon</td>
<td>12.01</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>14.01</td>
</tr>
<tr>
<td>Oxygen</td>
<td>16.00</td>
</tr>
<tr>
<td>Sodium</td>
<td>22.99</td>
</tr>
<tr>
<td>Aluminum</td>
<td>26.98</td>
</tr>
</tbody>
</table>

### Calculating Mass Using Atomic Mass Units (amu)

Calculate the mass, in amu, of a sample of aluminum that contains 75 atoms.

**Solution**

To solve this problem we use the average mass for an aluminum atom: 26.98 amu. We set up the equivalence statement

\[
1 \text{ Al atom} = 26.98 \text{ amu}
\]

which gives the conversion factor we need:

\[
75 \text{ Al-atoms} \times \frac{26.98 \text{ amu}}{\text{Al-atoms}} = 2024 \text{ amu}
\]

**Self-Check Exercise 6.1**

Calculate the mass of a sample that contains 23 nitrogen atoms.
The opposite calculation can also be carried out. That is, if we know the mass of a sample, we can determine the number of atoms present. This procedure is illustrated in Example 6.2.

**Example 6.2**

**Calculating the Number of Atoms from the Mass**

Calculate the number of sodium atoms present in a sample that has a mass of 1172.49 amu.

**Solution**

We can solve this problem by using the average atomic mass for sodium (see Table 6.1) of 22.99 amu. The appropriate equivalence statement is

\[ \frac{1 \text{ Na atom}}{22.99 \text{ amu}} \]

which gives the conversion factor we need:

\[ \frac{1172.49 \text{ amu}}{22.99 \text{ amu}} \times \frac{1 \text{ Na atom}}{22.99 \text{ amu}} = 51.00 \text{ Na atoms} \]

**Self-Check Exercise 6.2**

Calculate the number of oxygen atoms in a sample that has a mass of 288 amu.

To summarize, we have seen that we can count atoms by weighing if we know the average atomic mass for that type of atom. This is one of the fundamental operations in chemistry, as we will see in the next section.

The average atomic mass for each element is listed on the inside front cover of this book.

**CHEMISTRY in ACTION**

**Counting Pennies Without Counting**

1. Obtain a sample of pennies in a sealed container. Do not count the pennies.
2. Obtain an empty container that is similar to the one containing the pennies and one additional penny. Devise a method using a balance to determine the number of pennies in the sealed container without opening it. Determine the number of pennies in the sealed container.
3. Repeat step 2 using ten additional pennies instead of one penny. Use your method to determine the number of pennies in the sealed container.
4. Open the sealed container and count the pennies.
5. Which method (using one penny or ten pennies) allowed you to more accurately determine the number of pennies in the sealed container? Why? How does this finding relate to counting atoms by weighing?