In order to determine the number of atoms in a sample of an element, we need to know the mass of one atom of the element. As a result of several different experimental techniques, we know that there are $6.022 \times 10^{23}$ atoms in an amount of an element equal to its atomic mass in grams. The average atomic mass of carbon is 12.011 amu (atomic mass units). Therefore, 12.011 grams of carbon contain $6.022 \times 10^{23}$ carbon atoms. The average mass of one carbon atom is $12.011 \text{ g}/6.022 \times 10^{23} \text{ atom}$ or $1.995 \times 10^{-23} \text{ g/atom}$.

In the laboratory, we work on a gram scale and, consequently, handle extremely large numbers of atoms at a time. Of course, industries work on a much larger scale, yet. For our convenience, we introduce a unit called the mole to represent a large quantity of items. We define the mole to be $6.022 \times 10^{23}$ units. This is similar to our use of the word "pair" to represent two items and "dozen" for twelve items. This number, $6.022 \times 10^{23}$, is called Avogadro's number in honor of Avogadro, who lived from 1776 to 1856 and was a pioneer in the physical sciences.

From the definition of the mole, we can state that one mole is the atomic mass of an atomic substance in grams. The metals are atomic substances. The unit that we associate with the atomic mass is either the amu or the grams/mole (g/mol). Most of the time the g/mol unit is more useful in solving problems. To summarize, we can say one mole of an atomic substance contains $6.022 \times 10^{23}$ atoms and has a mass of one gram atomic mass.

A gram atomic mass means an amount of the element equal to its atomic mass in grams. We usually shorten this to just "molar mass".

We frequently convert between grams and moles, and sometimes atoms. We perform this pattern of conversions again and again.

\[
\begin{array}{cccc}
\text{grams} & \text{= molar mass} & \times & 6.022 \times 10^{23} \\
\text{molecules} & \text{=} & \text{mole} & \text{=} \\
\text{atoms} & \text{=} & \text{molar mass} & + 6.022 \times 10^{23} \\
\end{array}
\]

The top line indicates the pattern from left to right, and the bottom line shows the pattern from right to left. If you are given the number of atoms of a substance and want to find the number of grams of the substance, you would divide the number of atoms by Avogadro's number and multiply by the molar mass.

Historically, atomic masses were obtained from the mass relationships between elements in forming compounds. For example, 1.000 g of carbon combines with 2.664 g of oxygen to form carbon dioxide, $\text{CO}_2$. The workers at the time did not know how many carbon atoms were in the 1.000 g, but they recognized that the 2.664 g of oxygen had twice as many atoms as the 1.000 g of carbon. Therefore each oxygen atom is $2.664/2$ or 1.332 times
beavier than a carbon atom. If we assign carbon atoms to have an average mass of 12.011 amu, then the mass of the oxygen atom is 1.032x12.011 or 16.00 amu. We can continue this process to obtain the atomic masses of the other elements by forming compounds with carbon or oxygen, etc. Currently, we measure atomic masses using an expensive instrument known as a double-focusing mass spectrometer.

The atomic masses to four significant figures are included in the periodic table which is listed under the program item "Tables". The atomic masses are the numbers below the elemental symbols. These masses are based on the carbon-12 isotope being assigned an atomic mass of exactly 12.000... amu.

EXAMPLE Find the number of lithium atoms in 5.00 grams of Li. We know that the atomic mass in grams contains $6.022 \times 10^{23}$ atoms. The atomic mass of lithium is 6.941 g/mol. We need to convert grams into atoms, so our conversion factor is $6.022 \times 10^{23}$ atoms/6.941 g.

\[
\frac{6.022 \times 10^{23} \text{ atoms}}{6.941 \text{ g Li}} = \frac{5.00 \text{ g Li}}{6.022 \times 10^{23} \text{ atoms}}
\]

? Li atoms = 5.00 g Li × \frac{6.022 \times 10^{23} \text{ atoms}}{6.941 \text{ g Li}} = 4.35 \times 10^{23} \text{ atoms}