The Genesis spacecraft mission (2001-2004) primary objective was to determine the isotope ratios of oxygen in the sun. These ratios provide important evidence about the origin of the Earth and other inner "terrestrial" planets.

Oxygen is the third most common element in the universe, after hydrogen and helium. It is unusual among light elements because it has three stable isotopes:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Symbol</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Typical% on Earth</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen-16</td>
<td>8\textsuperscript{16}O</td>
<td>8</td>
<td>8</td>
<td>99.757%</td>
</tr>
<tr>
<td>Oxygen-17</td>
<td>8\textsuperscript{17}O</td>
<td>8</td>
<td>9</td>
<td>0.038%</td>
</tr>
<tr>
<td>Oxygen-18</td>
<td>8\textsuperscript{18}O</td>
<td>8</td>
<td>10</td>
<td>0.205%</td>
</tr>
</tbody>
</table>

The three stable isotopes may occur in different percentages, or "isotopic signatures"\textsuperscript{[1]} in different locations, so you can tell a rock from Mars from a rock on Earth by measuring isotopic abundances with a mass spectrometer. Rocks on the Moon, however, are similar to those on Earth, and this is taken as evidence that the Moon may actually be a chunk of the Earth that broke off as a result of a collision with a Mars-sized planet\textsuperscript{[2]}.
Isotope "delta" Values

The analysis is based on "δ^{18}O" and "δ^{17}O" values, which represent very small differences in isotopic ratios. Delta values are measured in parts per thousand (or "per mil", ‰). The more negative the value, the more of the lighter isotope is present, so positive values represent material with more of the heavy isotope. Different rocks on Earth have different "δ^{18}O" and "δ^{17}O" values, but "δ^{18}O" is always twice "δ^{17}O". The same is true on the Moon, which suggests that the Moon broke from the Earth eons ago. But Mars has a larger "δ^{18}O" than Earth, and an asteroid named Vesta has slightly smaller "δ^{18}O" value, as shown in the Figure. All of these are consistent with more or less common origins, with subsequent processes affecting the isotope ratios slightly.

![Oxygen Isotope Ratios](image)

Figure 2 Comparision of oxygen isotope ratios on Mars, Earth, and Vesta meteorites

But the Allende meteorite, a fireball which fell in Mexico in 1969, contains nuggets with entirely different oxygen isotope ratios. These nuggets are thought to be the first particles to form from the "solar nebula" about 4.567 billion years ago, before the planets were formed. They are enriched in the lighter isotopes, and "δ^{18}O" is not twice "δ^{17}O", but they're the same. To find out if the Sun, and presumably the solar nebula, actually has an isotopic composition much different than the planets, the Genesis spacecraft collected samples of the "solar wind".

Samples of the Solar Wind show the same enrichment in the 80 ratio as Allende, so the inclusions are close to the bulk composition of the solar system. This implies that Earth, the Moon, Mars, and asteroids all formed from -enriched material. The mechanism of their formation is a subject of intense debate.

But is there a "normal" isotopic abundance ratio for an element? If the abundance of oxygen isotopes can vary by ~20‰ (2%), how can we have a single "atomic weight" for the element?

The "Normal" Isotopic Ratio: Atomic Weights

All atoms of a given element do not necessarily have identical masses. But all elements combine in definite mass ratios,
so they behave as if they had just one kind of atom. In order to solve this dilemma, we define the **atomic weight** as the weighted average mass of all naturally occurring (occasionally radioactive) isotopes of the element.

A weighted average is defined as

\[
\text{Atomic Weight} = \sum \left( \frac{\% \text{ abundance isotope } i}{100} \right) \times \text{mass of isotope } i
\]

Similar terms would be added for all the isotopes. Since the abundances change from place to place, IUPAC has established "normal" abundances which are most likely to be encountered in the laboratory. This important document that reports these values can be found at the [IUPAC site](#). The abundances are also usually listed on the [Table of the Nuclides](#) which lists all isotopes for all elements. Surprisingly, a good number of elements have isotopic abundances that vary quite widely, so that atomic weights based on them have only 3 or 4 digit precision.

The atomic weight calculation is analogous to the method used to calculate grade point averages in most colleges:

\[
\text{GPA} = \sum \left( \frac{\text{Credit Hours Course } i}{\text{total credit hours}} \right) \times \text{Grade in Course } i
\]

Example \(\PageIndex{1}\): The Atomic Weight of Oxygen

The IUPAC Report gives the following data for oxygen:

- 99.757\% \(\text{^{16}O}\) whose isotopic weight is 15.9949146.
- 0.038\% \(\text{^{17}O}\) whose isotopic weight is 16.9991315.
- 0.205\% \(\text{^{18}O}\) whose isotopic weight is 17.9991604.

Calculate the atomic weight of an average naturally occurring sample of carbon.

**Solution**

\[
\dfrac{99.757}{100} \times 15.9949 + \dfrac{0.038}{100} \times 16.99913 + \dfrac{0.205}{100} \times 17.99916 = 15.9994
\]
Isotope Signatures of the Planets

The fact that the $\delta^{18}O$ and $\delta^{17}O$ values on Allende do not follow the expected 2:1 ratio is an example of a mass-independent fractionation, which can't be explained by the usual mass dependent kinetic isotope effect. One explanation would be that the oxygen in the meteorite, and the solar wind, arises by photochemical decomposition of CO, and the wavelength required depends on the oxygen isotope in the CO\textsuperscript{[7]}. This may account for the differences among bodies in the solar system. Another possibility is that matter from a supernova prior to the existence of our solar system impacted the oxygen isotope ratios. There is much that isn't known about the development of our solar system.

While we use the "normal" 99.757% abundance for $\text{^{16}}O$ to calculate the standard atomic weight, the abundance varies from 99.7384% - 99.7756% (other oxygen isotope abundances vary similarly). This allows the "isotopic fingerprinting" to be used for material on Earth.

Example $\PageIndex{2}$: Atomic Weight of Hydrogen

Naturally occurring hydrogen is found to consist of two isotopes:

99.9885% $\text{^{1}}H$ whose isotopic weight is 1.007825.

0.0115% $\text{^{2}}H$ whose isotopic weight is 2.01410178.

Calculate the atomic weight of an average naturally occurring sample of hydrogen.

Atomic Weight =

$$\dfrac{98.9885 \times 1.007825}{100.00} + \dfrac{0.0115 \times 2.01410178}{100.00} = 1.00794$$

Defining the Mole

The SI definition of the mole also depends on the isotope $\text{^{12}}C$ and can now be stated. One mole is defined as the amount of substance of a system which contains as many elementary entities as there are atoms in exactly 0.012 kg of $\text{^{12}}C$. The elementary entities may be atoms, molecules, ions, electrons, or other microscopic particles. This official definition of the mole makes possible a more accurate determination of the Avogadro constant than was reported earlier. The currently accepted value is $N_A = 6.02214179 \times 10^{23}$ mol$^{-1}$. This is accurate to 0.00000001 percent and contains five more significant figures than $6.022 \times 10^{23}$ mol$^{-1}$, the number used to define the mole previously. It is very seldom, however, that more than four significant digits are needed in the Avogadro constant. The value $6.022 \times 10^{23}$ mol$^{-1}$ will certainly suffice for most calculations needed.

From ChemPRIME: 4.13: Average Atomic Weights
References


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