learning objectives

1. Define atomic weight
2. Calculate atomic weight from percent abundance
3. Manipulate the atomic weight equation to calculate various unknown variables
4. Distinguish between atomic weight, atomic number, and mass number

Atomic Weight

The atomic weight is the mass of an atom, typically expressed in atomic mass units (amu). For an isotope, it is the mass of the nucleus, that is the mass of the protons and neutrons, as the mass of the electrons are considered negligible. In their natural state only 21 elements exist as single isotopes, that is a sample has nuclei of only one isotope, and these are called the mononuclidic elements. Most elements exist as a mixture of nuclei from multiple isotopes, and these are labeled as the polynuclidic elements. The atomic weight of a mononuclidic element is that mass of that nuclide.

For a polynuclidic element the atomic weight is the average weight based on the fractional abundance of each isotope, and this is the value given on the periodic table. Copper has two isotopes, $^{63}\text{Cu}$ (69.15%, mass=62.9300 amu) and $^{65}\text{Cu}$ (30.85%, mass = 64.928 amu), and so the respective mole fractions are 0.6915 and 0.3085, resulting in an average atomic weight of 63.55 amu, even though there is not a single atom that weighs 63.55 amu.

\[
\underbrace{0.6915}_{\text{fraction}} \underbrace{(62.9300 \text{ amu})}_{\text{mass}} + \underbrace{0.3085}_{\text{fraction}} \underbrace{(64.928 \text{ amu})}_{\text{mass}} = \underbrace{63.55 \text{ amu}}_{\text{average mass}} \quad \text{note: } 0.6915 + 0.3085 = 1
\]

Figure \(\PageIndex{1}\): Natural samples of copper contain two isotopes, and its atomic weight is to four significant digits is 63.55 amu, even though there is not a single atom of copper that weighs 63.55 amu.

Exercise \(\PageIndex{1}\)

The atomic weight of chorine is ____________ and the atomic number of chlorine-35 is______________.

a. 35, 17  
b. 17, 35  
c. 35.4527; 17  
d. 35.4527; 35  

Answer

C) the atomic weight is the average of mass of all isotopes of chlorine atoms and found below the symbol on the
periodic table. The atomic number is the number of protons in all chlorine atoms and is found on the top of the symbol in the periodic table.

Mononuclidic Elements

There are 21 elements with natural occurring samples that have only have one nucleus and are thus classified as mononuclidic. Nineteen of these are stable and two are radioactive (Bi & Th). The term monoisotopic refers to having only 1 stable nucleus, of which there are 26, nineteen of which are mononuclidic, but seven of the monoisotopic elements also have radioactive isotopes and so are not mononuclidic, but polynuclidic.

For a mononuclidic element, the atomic weight is the weight of its isotope. Sodium is a stable mononuclidic element that consists of just one isotope, $^{23}_{11}$Na.

![Figure](PagelIndex1): Sodium only has one isotope, and note, the uncertainty of this value is expressed by the (2) in 22.989 769 28(2), indicating that the last value has an uncertainty of +/- .000 000 02. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community. (cc. 4.0), p. 1864.

The following lists the 21 stable mononuclidic elements:

<table>
<thead>
<tr>
<th>Element Symbol</th>
<th>Stable Isotope</th>
<th>Relative Atomic Mass</th>
<th>Mole Fraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{23}_{11}$Na</td>
<td>$^{23}_{11}$Na</td>
<td>22.989 769 29</td>
<td>1</td>
</tr>
</tbody>
</table>

Table (PagelIndex1): List of the Mononuclidic elements, note, $^{232}_{90}$Th has a half life of $(1.4x10^{10}$years$)$, while $^{209}_{83}$Bi has one of $(2x10^{19}$years$)$, and so even though they are radioatcive, natural samples only have one isotope (Th is 99.98% $^{232}_{90}$Th) and 0.02% $^{230}_{90}$Th).

Polynuclidic Elements

The vast majority of elements are polynuclidic, that is, a sample contains a mixture of nuclei from different isotopes. For some samples the fraction of each isotope is constant, while for others it varies between samples. So there are two types of polynuclidic elements, those with invariant isotopic distributions, and those with variable isotopic distributions.
Constant Isotopic Distributions

For example, copper has two stable isotopes, and is approximately 69% $^{63}\text{Cu}$ and 31% $^{65}\text{Cu}$ by number (not mass). This ratio is constant from sample to sample and so it is invariant. The atomic weight of copper is based on this isotropic distribution and comes out to 63.546 amu, even though there is not a single atom of copper that weighs 63.546 amu (69% of them weigh 63 amu and 31% weigh 65).

<table>
<thead>
<tr>
<th>Stable Isotope</th>
<th>Relative Atomic Mass</th>
<th>Mole Fraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{63}\text{Cu}$</td>
<td>62.929 597</td>
<td>0.6915</td>
</tr>
<tr>
<td>$^{65}\text{Cu}$</td>
<td>64.927 790</td>
<td>0.3085</td>
</tr>
</tbody>
</table>

Figure \(\PageIndex{3}\): Copper has two stable isotopes as indicated in the pie diagram. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p. 1901.

Lead has 4 stable isotopes as shown in figure . Lead is a fairly large element and has many radioisotopes, in fact it has 29 isotopes with a half life of less than one hour, seven between an hour and a year, and two with half lives over a year ($^{202}\text{Pb} = 5.25\times 10^4$ years & $^{210}\text{Pb} = 22.3$ years), but these are insignificant to the thousandth position (note the four mole fractions add to one).

<table>
<thead>
<tr>
<th>Stable Isotope</th>
<th>Relative Atomic Mass</th>
<th>Mole Fraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{204}\text{Pb}$</td>
<td>203.973 043</td>
<td>0.014</td>
</tr>
<tr>
<td>$^{206}\text{Pb}$</td>
<td>205.974 465</td>
<td>0.241</td>
</tr>
<tr>
<td>$^{207}\text{Pb}$</td>
<td>206.975 897</td>
<td>0.221</td>
</tr>
<tr>
<td>$^{208}\text{Pb}$</td>
<td>207.976 652</td>
<td>0.524</td>
</tr>
</tbody>
</table>

Figure \(\PageIndex{4}\): Lead has 4 stable isotopes. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p 2007.
Variable Isotopic Distributions

Until recently, it was thought that the isotopic composition of an element was constant, but now it is realized that these can vary across samples. To date, the following 13 elements have been identified to have a range of isotopic distributions: hydrogen, lithium, boron, carbon, nitrogen, oxygen, magnesium, silicon, sulfur, chlorine, argon, bromine, and thallium. In 2009 the IUPAC Commission on Isotopic Abundances and Atomic Weights introduced an interval notation on atomic weights that expressed the natural occurring variability in the isotopic composition of the elements.

Hydrogen consists of two isotopes, protium ($^1\text{H}$, a single proton) and deuterium ($^2\text{H}$, a proton and a neutron). The isotopic composition of different samples vary within the defined ranges, so the percent protium ranges from 99.999% to 99.972%. There is a third naturally occurring radioactive isotope of hydrogen, $^3\text{H}$ (tritium), which is very rare and has a mole fraction that is so small that it is insignificant compared to protium and deuterium.

\[
\begin{array}{|c|c|c|}
\hline
\text{Stable Isotope} & \text{Relative Atomic Mass} & \text{Mole Fraction} \\
\hline
^1\text{H} & 1.007 825 0322 & [0.999 72, 0.999 99] \\
^2\text{H} & 2.014 101 7781 & [0.000 01, 0.000 28] \\
\hline
\end{array}
\]

Figure \(\PageIndex{5}\): Hydrogen’s Isotopic distribution can vary from sample to sample and so IUPAC introduced an interval notation. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p 1839.

These differences can actually show up in the hydrogen atomic weight in river water, where for example, water in the mountains will have a higher proportion of the heavier isotope, and this makes sense, as the heavier isotope would condense into rain easier than the lighter, and water at lower elevations would tend to have more of the lighter isotope. This can be seen in figure 2.3.6.

Figure \(\PageIndex{6}\): This map shows the variance of hydrogen’s isotopic distribution from various water samples across the U.S. Adopted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p. 1840.

Note: The sum of the percent natural abundances of all the isotopes of any given element must total 100%.
Some naturally occurring and artificially produced isotopes are radioactive. All atoms heavier than Bismuth ($^{209}_{83}Bi$) are radioactive. However, there are many lighter nuclides that are radioactive. For example, hydrogen has two naturally occurring stable isotopes, $^1H$ and $^2H$ (deuterium), and a third naturally occurring radioactive isotope, $^3H$ (tritium).

### Measuring Isotopic Abundances

The Mass Spectrometer is an instrument that allows us to measure the mass of atoms, or more accurately, the mass to charge ratio. In figure 2.3.8 you can see chlorine gas entering an mass spectrometer. The chlorine has multiple isotopes and is hit with a stream of ionizing electrons which break the bond of Cl₂ and strip electrons off the chlorine atoms causing monatomic ions. These are then accelerated down the chamber until they reach a magnetic field that deflects the particles. The angle of deflection depends on both the mass of the particle and the magnetic field strength, with the lighter particles being deflected more for a constant magnetic field strength (the lighter $^{35}Cl^+$ ions are deflected more than the heavier $^{37}Cl^+$ ions.) At the end of the chamber is an exit hole with a detector, and as the magnetic field intensity is increased the deflection angle changes, which separates the particles. Note, the mass spectrum in figure 2.3.8 (b) gives the relative abundance of each isotope, with the peak normalized to the isotope with the highest abundance. So if this ratio was 3:1 that means there are 3 particles of $^{35}Cl$ for every particle of $^{37}Cl$, and the fractional abundance would be 0.75 for $^{35}Cl$ and 0.25 for $^{37}Cl$.

![Mass spectrometer diagram](image)

*Figure 2.3.8: Determining Relative Atomic Masses Using a Mass Spectrometer. Image used with permission (CC BY-SA-NC; Anonymous by request).*

Below is a video from YouTube describing the mass spectrometer
Video \(\PageIndex{1}\): 2:43 YouTube uploaded by Evagating on how a mass spectrometer works.

Here is a bar chart showing the relative abundance of 4 isotopes of strontium

![The mass spectrum for strontium](image.png)

*Figure \(\PageIndex{9}\): Bar chart showing the relative isotopic percent abundance for the four isotopes of strontium.*

The mass spectrum of strontium has four different peaks, varying in intensity. The four peaks indicate that there are four
isotopes of strontium. The four isotopes of strontium have isotopic mass numbers of 84, 86, 87, and 88, and relative abundances of 0.56%, 9.86%, 7.00%, and 82.58%, respectively. The intensity of the peak corresponds to the abundance. \(^{\text{84}}\text{Sr}\) has the smallest peak, which corresponds to its relative abundance of 0.56%, whereas \(^{\text{88}}\text{Sr}\) has the largest peak, which corresponds to its relative abundance of 82.58%. This indicates that \(^{\text{88}}\text{Sr}\) is the isotope that occurs in highest amounts.

Exercise \(\PageIndex{2}\)
If the first isotope (Isotope 1) has a mass of 129.588amu and the second isotope (Isotope 2) has a mass of 131.912 amu, which isotope has the greatest natural abundance?

a. Isotope 1
b. Isotope 2
c. There are equal amounts.
d. Not enough information provided.
e. none of the above

Answer

B) Isotope 2. Although it is algebraically possible to calculate the particular percent abundances for both isotopes, there is not need to spend that much time on this problem if you know the principle behind it. The average is 131.244 amu. It looks like the mass of Isotope 2 (131.912amu) is closer to the average than the mass of isotope 1 (129.588 amu). This indicates that isotope 2 impacted the average much more than isotope 1 and has a greater percent abundance.

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Isotopic Abundance Calculations

For elements with more than one isotope there are a variety of problems you need to be able to solve. You not only need to be able to calculate the average mass from the isotopic abundance and masses, but go backwards, using the atomic weight on the periodic table as the average atomic mass. Rewriting eq. 2.3.1 in terms of algebraic variables gives us a feel for the types of problems. Consider an element with two isotopes A & B, each having a mass of \(m_A\) or \(m_B\) and a mole fraction of \(X_A\) or \(X_B\).

\[
\underbrace{X_A}_\text{fraction A} \underbrace{(m_A)}_\text{ mass A} + \underbrace{X_B}_\text{fraction B} \underbrace{(m_B)}_\text{ mass B} = \underbrace{m_{ave}}_\text{average mass}
\]

In general the equation for an element with \([n]\) isotopes is:

\[
\sum_{i=1}^{n}X_im_i=m_{ave}
\]

Note, the equation for two isotopes has 5 unknowns, and you can get \(m_{ave}\) from the periodic table, so you need to be
given three of the other variables to solve the equation (one equation can only be solved for one unknown). But there is another equation we can use to reduce the number of unknowns. That is the sum of the fractions equals 1

\[X_A + X_B = 1\]

or in general

\[\sum_{i=1}^{n}X_i = 1\]

In the following problem we look

you not only need to be able to calculate the average atomic mass, but also need to be use the average mass on the periodic table to work backwards and calculate percent

Example: Calculating Average Atomic Mass

What is the average atomic mass of Neon, given that it has 3 isotopes with the follow percent abundances;

\[{}^{20}\text{Ne} = 19.992 \text{ amu (90.51%)}, {}^{21}\text{Ne} = 20.993 \text{ amu (0.27%)}, {}^{22}\text{Ne} = 21.991 \text{ amu}.\]

What we know: since you know what the element is, you can solve this without doing any math by using the periodic table, but you need to be able to do the math because it might be an unknown, and that is the only way you can figure out the correct significant figures.

Since Ne-20 has the greatest percent abundance, it should have the most impact on your average. Therefore, we expect the average atomic mass to be closer to the mass of Ne-20 (about 19.992 amu). Click the following video tutor to see if we estimated correctly.
Video \(\PageIndex{2}\): 4:53 minute YouTube solving above problem.

**Answer:** According to the correct number of significant figures, we came up with 20.18 amu as the average atomic weight even though the average atomic weight from the periodic table is 20.179 amu. However, it is still a good check to make sure that you are on the right path.

**Check Yourself:** We predicted earlier that our answer should be closer to the mass of Ne-20 (19.992 amu) instead of Ne-21 or Ne-22 because it has the greatest natural abundance, and thus, impacts the average more. We can see that the math does align with our logic!

Example \(\PageIndex{1}\): Isotopic Mass Calculation

Chlorine has two isotopes, with 75.53% being \(^{35}\text{Cl}\) with an isotopic mass of 34.969 amu, what is the mass of the other isotope?

**What we know:** In this case, you have the average atomic mass (from the periodic table). You are trying to find the mass of the individual isotope. You also know that the individual isotopes have to add up to 100%.
Answer: The correct answer is 36.9 amu.

After watching the following two YouTubes, you should do the following worksheets, which were designed as in class activities for the prep course, and so give more step-by-step instructions than we are using. You will want to be able to bypass many of the "steps" in the handouts as these were designed for the preparatory chem 1300 class, which spent more time on this material. Also, note the handouts are posted as PDF files, and you can use the hypothes.is annotation system to comment on them, and if you tag them with the tag f19c1402c2, they will appear on the table of contents for this chapter.

Isotope Abundance Worksheets:

Isotope Abundance Worksheet

Key to Isotope Abundance Worksheet

Vocabulary

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- Anonymous
- Ronia Kattoum & November Palmer (UA Little Rock)