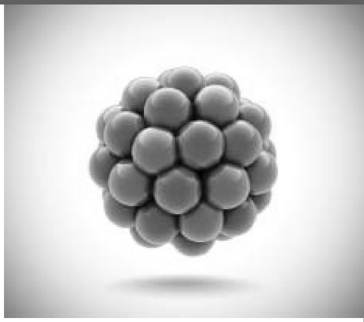


Inside Isotopes

Investigating Isotopic Distribution

NATIONAL MATH + SCIENCE INITIATIVE

MATERIALS

scoop

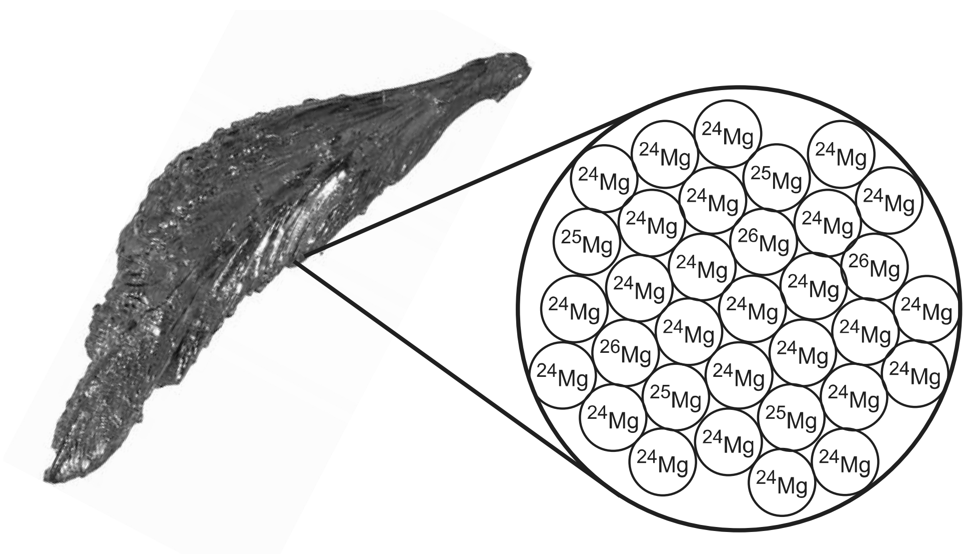
plate, paper

During a lecture in 1803, John Dalton proposed that all atoms of the same element are identical. Is this claim still part of our modern atomic theory? Not exactly. All atoms of a particular element have the same number of protons because this atomic number identifies the element. If they did not have the same number of protons, they would not be the same element. However, the individual atoms of an element may have a different number of neutrons.

Isotopes are atoms of the same element that differ in their number of neutrons. To distinguish the isotopes of an element, we include the mass number (protons + neutrons) in either the chemical symbol (^{24}Mg) or the name (magnesium-24).

Isotopes do not differ in their chemical behavior. Because an atom's chemical properties are a function of its electron arrangement, the number of neutrons does not affect the way the atom behaves in ordinary chemical reactions. The different isotopes of an element are indistinguishable in non-nuclear reactions.

Figure 1. Atoms of magnesium

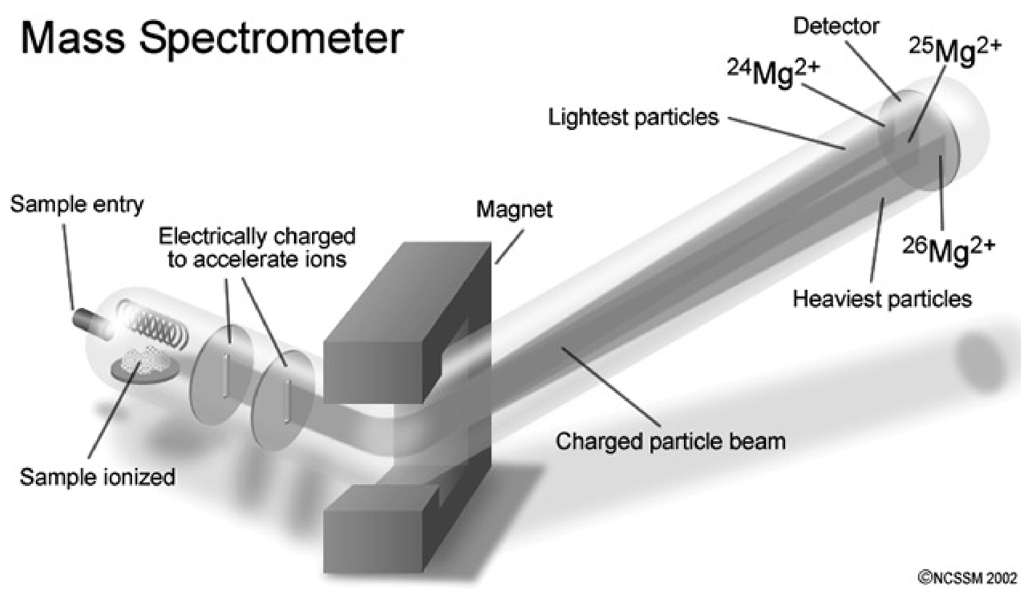


Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

All but 21 of the elements on the periodic table have multiple isotopes, and tin (Sn) wins the prize for the greatest number of stable isotopes (10). Because each isotope has a different mass, the overall atomic mass of an element must reflect an average of these isotope masses. But given the unequal distribution of the different isotopes, it would not be accurate to do a standard average calculation. Rather, we must use a *weighted average* that reflects the fact that some isotopes are very common and others are more obscure.

So how are the different isotopes identified and how do we know which ones are more or less abundant? Scientists use a tool called a mass spectrometer to make this sort of determination. At its most basic level, a **mass spectrometer** is like an atomic balance. It can provide us with the mass of a single atom as well as tell us how many of those atoms are present in a given sample. It operates on the idea that charged particles are deflected in a magnetic field to varying degrees based on their mass (Figure 2).

Figure 2. Mass spectrometer

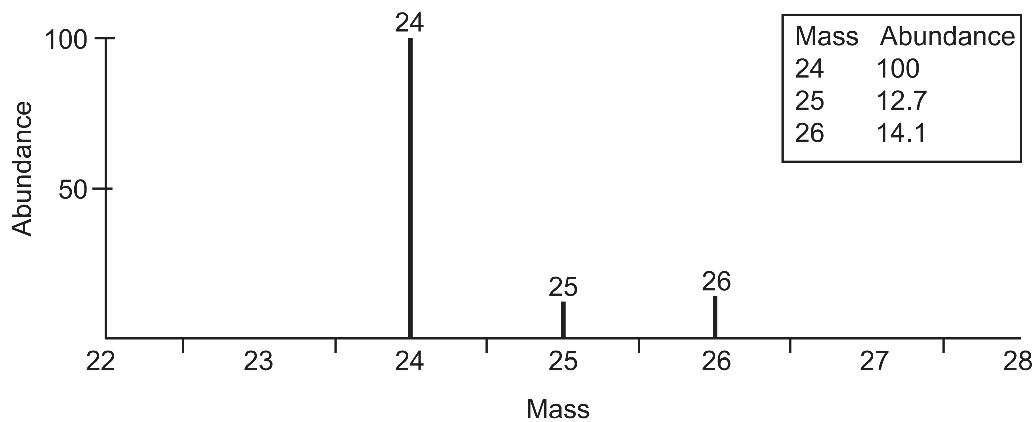


When an element sample is bombarded with electrons, positive ions are created and those cations travel in curved paths toward a negatively charged plate. As they travel toward the detector, they pass through a magnetic field and are deflected more or less based on their mass. The effect would be similar to throwing a ping-pong ball versus a baseball on a very windy day; the lighter one is deflected more.

Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

For the magnesium sample depicted in the examples here, three isotopes are present: ^{24}Mg, ^{25}Mg, and ^{26}Mg. The output of a mass spectrometer is a histogram-like graph (Figure 3) where, in this case, we see that ^{24}Mg is most abundant and that ^{25}Mg and ^{26}Mg are present in much smaller amounts. Without doing any calculations, we can predict that the average atomic mass of magnesium is somewhere between 24 and 26 amu but much closer to 24 amu based on the greater abundance of that isotope. What does your periodic table report as the mass of magnesium?

Figure 3. Mass spectrum of magnesium



If we want to quantitatively use the data from a mass spectrometer to calculate the average atomic mass, we must first adjust the abundance values by turning them into a percentage. This can easily be done using Equation 1:

$$\%\text{ abundance} = \frac{\text{abundance of one isotope}}{\text{total abundance of all isotopes}} \times 100 \qquad \text{(Eq. 1)}$$

Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

Table 1. Mass Spectrum of Magnesium

Isotope	Calculation	Abundance (%)
^{24}Mg	$\frac{100}{126.8} \times 100$	78.9
^{25}Mg	$\frac{12.7}{126.8} \times 100$	10.0
^{26}Mg	$\frac{14.1}{126.8} \times 100$	11.1

If needed, the percent abundance and masses can then be used to calculate the weighted average for the atomic mass of the element using Equation 2:

$$\text{weighted average atomic mass} = \left(\frac{\%\text{ abundance}}{100} \times \text{mass}\right) + \left(\frac{\%\text{ abundance}}{100} \times \text{mass}\right) + \ldots \quad \text{(Eq.2)}$$

For magnesium,

$$\left(\frac{78.9}{100} \times 24\right) + \left(\frac{10.0}{100} \times 25\right) + \left(\frac{11.1}{100} \times 26\right) = 24.3 \text{ amu}$$

PURPOSE

In this activity, you will model an element with three different isotopes using colored beads and analyze that data to calculate the average atomic mass of your sample. In addition, you will interpret a mass spectrum and relate the data to the percent abundance of isotopes of a given element.

Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

PROCEDURE

PART I

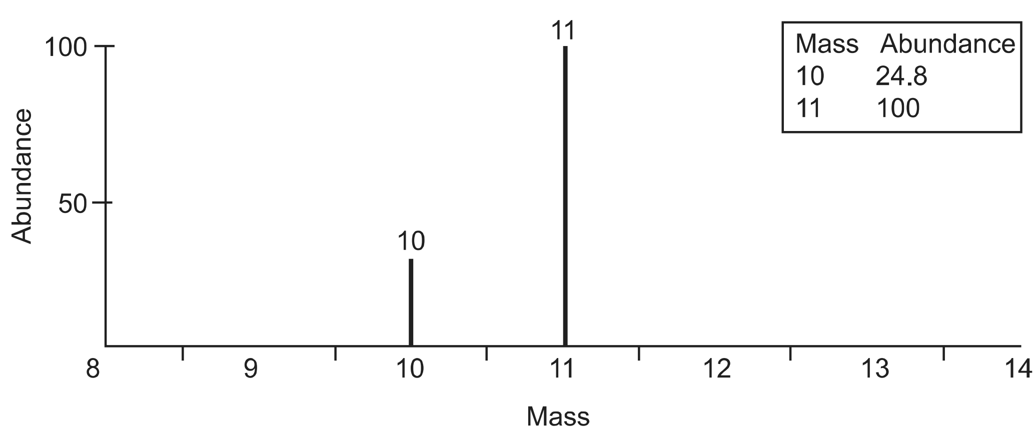
In this simulation, colored beads will represent the different isotopes of a fictitious element. Your task will be to determine the average atomic mass of this element based on the relative abundance and masses of the various isotopes.

1. Record the values for Table 2 that your teacher provides.
2. Obtain a representative sample of the element using the scoop your teacher has provided, and take the sample back to your lab station.
3. Sort the atoms from your sample by color and record the frequency of each color in Table 3 on your student answer page.
4. Create a visual representation of the different abundances by physically arranging the beads like a bar graph on your tabletop. Take a picture of your graph with a digital camera if your teacher permits.
5. Determine the percent abundance of each isotope in your sample and record the values in Table 3.
6. Return the beads to the class bowl when you are finished.

PART II

Given the mass spectrum shown in Figure 4, construct a representative sample using exactly 100 total beads and take a picture of your sample with a digital camera if your teacher permits. Record any necessary calculations in the Data and Observations section of your student answer pages.

Figure 4. Mass spectrum



Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

PRE-LAB EXERCISES

1. Define the following terms:
   a. Atomic number

   b. Mass number

   c. Isotope

   d. Average atomic mass

2. Write the chemical symbol and name for each of the isotopes listed:
   a. 30 protons and 34 neutrons

   b. 46 protons and 60 neutrons

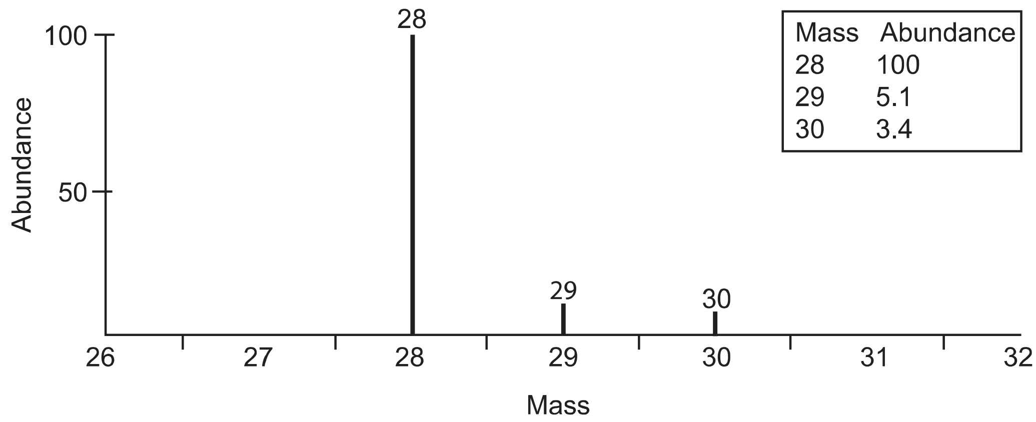
   c. 23 protons and 28 neutrons

Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

PRE-LAB EXERCISES (CONTINUED)

3. For the mass spectrum shown in Figure 5, calculate the percent abundances and the average atomic mass, and then identify the element.

Figure 5. Mass spectrum



Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.

PRE-LAB EXERCISES (CONTINUED)

4. Carbon has two naturally occurring stable isotopes, carbon-12 (^{12}C) and carbon-13 (^{13}C). Of all the carbon on Earth, 98.93% is carbon-12 and 1.07% is carbon-13.

   a. How many neutrons are in each isotope of carbon?

   b. Which isotope would be deflected more in a mass spectrometer, ^{12}C or ^{13}C? Explain.

   c. Based on the abundance percentages, should the average atomic mass of carbon atoms be closer to 12 or 13? Why?

   d. Calculate the average atomic mass of carbon based on these abundances.

Copyright © 2014 National Math + Science Initiative, Dallas, Texas. All rights reserved. Visit us online at www.nms.org.