Average or Apparent Mass of an Element

Isotopes and Atomic Mass

Introduction

At the beginning of the 19th century, John Dalton proposed his new atomic theory—all atoms of the same element are identical and equal in mass. It was a simple yet revolutionary theory. It was also not quite right. The discovery of radioactivity in the 20th century made it possible to study the actual structure and mass of atoms. Gradually, evidence was obtained that atoms of the same element could have different masses. These atoms were called isotopes. How are isotopes different from one another? What is the relationship between the atomic mass of an element and the mass of each isotope?

Concepts

- Isotope
- Mass number
- Atomic and electron structure
- Atomic mass

Materials

Bags, zipper-lock, 15
Balance, centigram (0.01-g precision)
Kidney beans, 12 oz
Labeling pen or marker
Lima beans, 20 oz
Navy beans, 4 oz
Weighing dishes or small cups, 4

Safety Precautions

Although the materials used in this activity are considered nonhazardous, please observe all normal laboratory safety guidelines. The food-grade items that have been brought into the lab are considered laboratory chemicals and are for lab use only. Do not taste or ingest any materials in the chemistry laboratory. Wash hands thoroughly with soap and water before leaving the laboratory.

Preparation

"Bean bag" isotopes may be mixed in any proportion to prepare samples for analysis. The mixtures analyzed in the *Sample Data* section were prepared by mixing navy beans, kidney beans, and lima beans in the following proportion: 100 g navy beans, 280 g kidney beans, and 500 g lima beans (*Note:* 1 oz = 28.35 g). The mixture (880 g total mass) was shaken in a large zipper-lock bag to mix the "isotopes" and divided into fifteen 50-g samples for student use. The samples are obviously not homogeneous—do not expect different student groups to obtain identical results for the percent abundance of each isotope. The percent abundance for the samples analyzed ranged from 22–28% for navy beans, 36–41% for kidney beans, and 33–38% for lima beans.

Sample Data *(Student data will vary.)*

Data Table*

	Trial 1		Trial 2	
"Bean Bag" Isotope (Bg)	Numbers of Atoms	Total Mass of Atoms	Numbers of Atoms	Total Mass of Atoms
1	19	5.62 g	16	4.72 g
2	28	15.60 g	29	16.31 g
3	26	28.32 g	27	29.45 g

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Results Table

	Trial 1		Trial 2	
"Bean Bag" Isotope (Bg)	**Average Mass**	**Percent Abundance**	**Average Mass**	**Percent Abundance**
1	0.296 g	26.0%	0.295 g	22.2%
2	0.557 g	38.4%	0.562 g	40.3%
3	1.089 g	35.6%	1.091 g	37.5%

*The data for two trials are shown in order to demonstrate the range of results that might be obtained in a typical classroom setting.

Procedure

1. Sort the atoms in the "bean bag" element sample (Bg) into three groups (1, 2, and 3) according to the type of bean. (Assume that each type of bean represents a different isotope and that each bean represents a separate atom.) Place each group into a separate weighing dish or small cup.
2. Count and record in the Data Table the number of Bg atoms in each group.
3. Measure the total mass of Bg atoms belonging to each group. Record the mass to the nearest 0.01 g in the Data Table. *Note:* Zero (tare) the balance with an empty weighing dish, then add all of the Bg atoms in one group to the weighing dish and record the mass. Do this for each group.

Disposal

None required. Save the "bean bag" samples for repeat use. The beans may begin to chip or break after repeated use. Discard any beans that are broken or otherwise very different from others.

Tips

- A wide variety of items may be used to simulate atoms in this exercise. The advantages of using beans are they are relatively non-perishable, many varieties are readily available, and students will not be tempted to eat them in the lab.
- The official Nobel Prize web site (www.nobel.se) contains biographies of the Nobel Prize–winning scientists and their presentation speeches. F. W. Aston's speech, for example, describes the history leading to the discovery of isotopes.
- Most of the isotopes used in medicine are manufactured radioisotopes. These artificial isotopes do not occur in nature. They are produced in nuclear reactors or cyclotrons by bombarding a non-radioactive target element with high energy particles—neutrons, protons, alpha particles, etc.
- Isotopes have nearly identical chemical properties. They differ, however, in their physical and nuclear properties. Differences in the physical properties of isotopes are used in various processes to separate isotopes. Two examples of physical processes used to separate isotopes include the distillation of water to obtain "heavy water" (deuterium oxide, D_2O) and the diffusion of gaseous UF_6, uranium hexafluoride, to obtain U-235 enriched uranium.
- Mass spectrometry is used in toxicology studies to detect and analyze drugs that might have been used in a crime. The technique is also used in arson investigations to detect the presence of accelerants that might have been used to set a fire. *Note to teachers:* The following Web sites are good starting points: "What is Mass Spectrometry?" tutorial is available at the American Society for Mass Spectrometry, www.asms.org/whatisms/ "A History of Mass Spectrometry" tutorial is available at the Scripps Center for Mass Spectrometry, http://masspec.scripps.edu. Finally, http://i-mass.com is a com mercial, international mass spectrometry web resource that contains many interesting feature articles on applications of mass spectrometry.

Discussion

Two experiments in the early 20th century suggested the possible existence of isotopes. The first was work by J. J. Thomson with positively charged atoms in gas discharge tubes. When the positively charged atoms were bent by electric and magnetic fields and then allowed to strike a photographic film, they left curved "spots" on the film at an angle that depended on the mass and charge of the atoms. In 1912, Thomson found that when the gas in the tube was neon, he obtained two curves or spots. The major spot corresponded to neon atoms with a mass of about 20 atomic mass units (amu). There was also a much fainter spot, however, corresponding to atoms with a mass of about 22 amu.

The second experiment suggesting the existence of isotopes came from studies of radioactivity. One of the products obtained from the radioactive decay of uranium is lead. When the atomic mass of lead deposits in radioactive uranium minerals was analyzed, it was found to be different from the atomic mass of lead in lead ore.

In 1913, Frederick Soddy, professor of chemistry at the University of Glasgow in Scotland, used the term isotope to define atoms of the same element that have the same properties but different atomic masses. The word isotope was derived from Greek words meaning "same place"—isotopes occupy the same place in the periodic table (they are the same element) even though they have different masses. Soddy received the Nobel Prize in Chemistry in 1921 for his work on isotopes.

Final proof for the existence of isotopes came from the work of Francis W. Aston at Cambridge University. Aston built a more accurate version of the apparatus that Thomson had earlier used to study charged atoms. In 1919, Aston measured the precise masses of the two isotopes of neon. Aston received the Nobel Prize in Chemistry in 1922 for his discovery of isotopes.

The modern definition of isotopes depends on the number of subatomic particles in atoms. Isotopes have the same number of protons but different numbers of neutrons. Since the identity of an element depends on the number of protons (the atomic number), isotopes have the same chemical properties. Isotopes are thus chemically indistinguishable from one another—they undergo the same reactions, form the same compounds, etc. Isotopes have different mass numbers. The mass number is the total number of protons and neutrons in an atom.

Chlorine, for example, occurs in the form of two isotopes, chlorine-35 and chlorine-37, where 35 and 37 are the mass numbers. Chlorine consists of 75.8% chlorine-35 atoms and 24.2% chlorine-37 atoms. The atomic mass of an element represents the *weighted average* of the masses of the isotopes. Equation 1 shows the atomic mass calculation for the element chlorine. The mass of each isotope is equal to its mass number.

$$\text{Atomic mass (chlorine)} = (0.758)(35.0 \text{ amu}) + (0.242)(37.0 \text{ amu}) = 35.5 \text{ amu} \qquad \textit{Equation 1}$$

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Evidence, models, and explanation
Constancy, change, and measurement

Content Standards: Grades 9–12
Content Standard B: Physical Science, structure of atoms
Content Standard G: History and Nature of Science, historical perspectives

Answers to Pre-Lab Questions *(Student answers will vary.)*

1. Silicon occurs in nature in the form of three isotopes, Si-28, Si-29, and Si-30. Determine the number of protons, neutrons, and electrons in each isotope of silicon.

Isotope	**Protons**	**Neutrons**	**Electrons**
Silicon-28	*14*	*14*	*14*
Silicon-29	*14*	*15*	*14*
Silicon-30	*14*	*16*	*14*

2. "The atomic mass of chlorine represents the mass of the most common naturally occurring isotope of chlorine." Decide whether this statement is true or false and explain why.

This statement is false—none of the atoms in a sample of chlorine has the same mass as the atomic mass. There are only two masses of individual chlorine atoms, either 35 or 37 amu. The atomic mass of chlorine (35.5 amu) represents the ***weighted average*** *of these two masses based on the number of atoms of each type as they are found in nature.*

Answers to Post-Lab Questions *(Student answers will vary.)*

1. Determine the average mass of each Bg isotope. Enter the results in the Results Table.

   Sample calculation for isotope 1, Trial 1: Average mass = 5.62 g/19 = 0.296 g.
   See the Results Table for the results of other calculations.

2. What is the total number of "bean bag" (Bg) atoms in the original sample? Calculate the *percent abundance* of each isotope: Divide the number of atoms of each isotope by the total number of atoms and multiply the result by 100. Enter the results to one decimal place in the Results Table.

   In Trial 1, the total number of atoms was (19 + 28 + 26) = 73.

   Sample calculation for isotope 1, Trial 1: Percent abundance = (19/73) × 100 = 26.0%

   See the Results Table for the results of other calculations.

3. The atomic mass of the "bean bag" element (Bg) represents a *weighted average* of the mass of each isotope and its relative abundance. Use the following equation to calculate the atomic mass of Bg. *Note:* Divide the percent abundance of each isotope by 100 to obtain its relative abundance.

   Relative abundance = Percent abundance/100

   $$\text{Atomic mass} = (\text{rel. abundance}_{\text{isotope 1}} \times \text{mass}_{\text{isotope 1}}) + (\text{rel. abundance}_{\text{isotope 2}} \times \text{mass}_{\text{isotope2}}) + (\text{rel. abundance}_{\text{isotope 3}} \times \text{mass}_{\text{isotope 3}})$$

   Trial 1: *Atomic mass = (0.260 × 0.296 g) + (0.384 × 0.557 g) + (0.356 × 1.089 g) = 0.679 g*

   Trial 2: *Atomic mass = (0.222 × 0.295 g) + (0.403 × 0.562 g) + (0.375 × 1.091 g) = 0.701 g*

4. How many Bg atoms in the original sample would you expect to have the same mass as the calculated atomic mass of the element? Explain.

   None! The atomic mass is a weighted average and does not represent the actual mass of any of the atoms in the sample.
   Note to teachers: *This answer applies to the actual samples analyzed in this exercise. Depending on the composition of "bean bag" samples prepared for student use, it is possible for the weighted average to be similar in value to the average mass of one of the types of beans. To avoid confusion, care should be taken to ensure that the average atomic mass does not coincide with any of the individual isotope masses.*

5. The isotopes of magnesium (and their percent abundance) are Mg-24 (79.0%), Mg-25 (10.0%), and Mg-26 (11.0%). Calculate the atomic mass of magnesium. *Note:* The mass of each isotope is equal to the mass number. Thus, the mass of an atom of Mg-24 is 24.0 amu.

   Atomic mass (Mg) = (0.790 × 24.0 amu) + (0.100 × 25.0 amu) + (0.110 × 26.0 amu) = 24.3 amu

6. Copper (atomic mass 63.5) occurs in nature in the form of two isotopes, Cu-63 and Cu-65. The percent aboundance is 75% Cu-63 and 25% Cu-65. Explain why the atomic mass of copper is not equal to 64, halfway between the mass of copper-63 and copper-65.

   The atomic mass is "weighted" toward the mass of the more abundant isotope, Cu-63.

Reference

This activity is from *Flinn ChemTopic™ Labs*, Volume 3, Atomic and Electron Structure; Cesa, I., Ed., Flinn Scientific: Batavia, IL.

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Average or Apparent Mass of an Element* activity, presented by Jesse Bernstein, is available in *Isotopes and Atomic Mass* and *Inquiry Lab Activities*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Average or Apparent Mass of an Element* are available from Flinn Scientific, Inc.

Materials required to perform this activity are available in the *"Bean Bag" Isotopes—Student Activity Kit* available from Flinn Scientific. Materials may also be purchased separately.

Catalog No.	Description
AP6633	"Bean Bag" Isotopes—Student Activity Kit
OB2141	Balance, Centigram, 0.01g-Precision

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

Data Table

Average of Apparent Mass of an Element

Data Table

"Bean Bag" Isotope (Bg)	Number of Atoms	Total Mass of Atoms
1		
2		
3		

Results Table

"Bean Bag" Isotope (Bg)	Average Mass of Isotopes	Percent Abundance
1		
2		
3		

Post-Lab Questions *(Use a separate sheet of paper to answer the following questions.)*

1. Determine the average mass of each Bg isotope. Enter the results in the Results Table.
2. What is the total number of "bean bag" (Bg) atoms in the original sample? Calculate the *percent abundance* of each isotope: Divide the number of atoms of each isotope by the total number of atoms and multiply the result by 100. Enter the results to one decimal place in the Results Table.
3. The atomic mass of the "bean bag" element (Bg) represents a *weighted average* of the mass of each isotope and its relative abundance. Use the following equation to calculate the atomic mass of Bg. *Note:* Divide the percent abundance of each isotope by 100 to obtain its relative abundance.

   Relative abundance = Percent abundance/100

   $$\text{Atomic mass} = (\text{rel. abundance}_{\text{isotope 1}} \times \text{mass}_{\text{isotope 1}}) + (\text{rel. abundance}_{\text{isotope 2}} \times \text{mass}_{\text{isotope2}}) + (\text{rel. abundance}_{\text{isotope 3}} \times \text{mass}_{\text{isotope 3}})$$

4. How many Bg atoms in the original sample would you expect to have the same mass as the calculated atomic mass of the element? Explain.
5. The isotopes of magnesium (and their percent abundance) are Mg-24 (79.0%), Mg-25 (10.0%), and Mg-26 (11.0%). Calculate the atomic mass of magnesium. *Note:* The mass of each isotope is equal to the mass number. Thus, the mass of an atom of Mg-24 is 24.0 amu.
6. Copper (atomic mass 63.5) occurs in nature in the form of two isotopes, Cu-63 and Cu-65. The percent aboundance is 75% Cu-63 and 25% Cu-65. Explain why the atomic mass of copper is not equal to 64, halfway between the mass of copper-63 and copper-65.