Bean Bag Isotopes
Relative Abundance and Atomic Mass

Introduction

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

Concepts

- Isotope
- Percent abundance
- Mass number
- Atomic mass

Background

Two lines of evidence in the early 20th century suggested the possible existence of isotopes. The first came from work by J. J. Thomson with "positive rays," positively charged streams of atoms generated in gas discharge tubes. When these positive rays were bent or deflected in the presence of 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. Although these results were consistent with the existence of two types of neon atoms having different masses, they were not precise or accurate enough to be conclusive.

The second line of evidence 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 significantly different from the atomic mass of lead in lead ore. The actual composition of the lead atoms seemed to be different, depending on their origin.

In 1913, Frederick Soddy, professor of chemistry at the University of Glasgow, coined the term *isotope* to define atoms of the same element that have the same chemical properties but different atomic masses. The word isotope was derived from Greek words meaning "same place" to denote the fact that 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 investigations into the nature and origin of isotopes.

Conclusive proof for the existence of isotopes came from the work of Francis W. Aston at Cambridge University. Aston built a modified, more accurate version of the "positive ray" apparatus that Thomson had earlier used to study ions. In 1919, Aston obtained precise measurements of the major and minor isotopes of neon, corresponding to mass numbers of 20 and 22 respectively. Aston received the Nobel Prize in Chemistry in 1922 for his discovery of isotopes.

The modern definition of isotopes is based on knowledge of the subatomic particle structure of atoms. Isotopes have the same number of protons but different numbers of neutrons. Since the identity of an element depends only 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 are distinguished from one another based on their mass number, defined as the sum of the number of protons and neutrons in the nucleus of the atom.

Chlorine, for example, occurs naturally in the form of two isotopes, chlorine-35 and chlorine-37, where 35 and 37 represent the mass numbers of the isotopes. Each isotope of chlorine has a characteristic *percent abundance* in nature. Thus, whether it is analyzed from underground salt deposits or from seawater, the element chlorine always contains 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 in a naturally occurring sample of the element. Equation 1 shows the atomic mass calculation for the element chlorine. The mass of each isotope is equal to its mass number, *to one decimal place precision*.

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

Experiment Overview

The purpose of this experiment is to investigate the mass properties and relative abundance of isotopes for the "bean bag" element (symbol, Bg) and to calculate the atomic mass of this element.

e-Lab Questions

1. Neutrons were discovered in 1932, more than 10 years after the existence of isotopes was confirmed. What property of electrons and protons led to their discovery? Suggest a possible reason why neutrons were the last of the three classic subatomic particles to be discovered.

2. 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.

3. *"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.

terials

Balance, centigram (0.01-g precision)
"Bean bag" element, symbol Bg, approximately 50 g
Weighing dishes or small cups, 4
Labeling pen or marker

ety 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.

Procedure

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

Data Table

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

Results Table

"Bean Bag" Isotope (Bg)	Average Mass	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 to three significant figures. 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 $= \dfrac{\text{Percent abundance}}{100}$

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

4. How many Bg atoms in the original sample would be expected 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:* To one decimal place, 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. Use this information to calculate the percent abundance of each copper isotope.

7. Explain why the atomic mass of copper is not exactly equal to 64, midway between the mass numbers of copper-63 and copper-65.

8. Radioactive isotopes (radioisotopes) are widely used in medicine. Because isotopes have identical chemical properties, the reaction and distribution of radioisotopes in the body is similar to that of their natural isotopes. Iodine-131, for example, is an artificial radioisotope that is used to diagnose thyroid disorders. When administered to a patient, the radioisotope is taken up by the thyroid gland, where it is incorporated into the thyroid hormones, just as iodine in the diet would be. Based on where the following elements are likely to be found in the body, match each radioisotope with its medical use.

Sodium-24	_____	a. studies of bone formation
Phosphorus-32	_____	b. red blood cell studies
Calcium-47	_____	c. tracing blood circulation
Iron-55	_____	d. genetics (DNA) research

9. *(Optional)* Aston called the instrument he designed to measure the masses of atoms the mass spectrograph. Modern versions of Aston's mass spectrograph, called mass spectrometers, are workhorse instruments in chemical analysis, including forensics. Look up *mass spectrometry* on the Internet and briefly describe two applications of this technology in forensic analysis.