A Certain Sample Of Element Z Contains 60% Of 69Z And 40% Of 71Z. What Is The Relative Atomic Mass Of

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Understanding the concept of atomic mass and how it relates to isotopic composition is fundamental in chemistry. When analyzing a sample of an element, it's common to encounter different isotopes—atoms of the same element with varying numbers of neutrons. The average atomic mass of an element, known as the relative atomic mass, takes into account the abundance of its isotopes. In this article, we will explore how to calculate the relative atomic mass of an element given the isotopic composition in a sample, specifically focusing on a sample of element Z that contains 60% of isotope 69Z and 40% of isotope 71Z.

Understanding Isotopes and Atomic Mass

What Are Isotopes?

Isotopes are variants of a chemical element that have the same number of protons but different numbers of neutrons. Because neutrons contribute to atomic mass but not to chemical properties, isotopes differ primarily in their mass.

For example:
- 69Z: Isotope with atomic mass approximately 69 atomic mass units (amu).
- 71Z: Isotope with atomic mass approximately 71 amu.

Relative Atomic Mass

The relative atomic mass (also called atomic weight) of an element is the weighted average of the masses of its isotopes, based on their natural abundance.

Mathematically, it is expressed as:

$$
\text{Relative Atomic Mass} = \sum (\text{mass of isotope} \times \text{fractional abundance of isotope})
$$

Calculating the Relative Atomic Mass for the Sample

Given the sample:
- 60% of isotope 69Z
- 40% of isotope 71Z

Step 1: Convert percentage abundances to fractions

Isotope	Percentage	Fractional Abundance
69Z	60%	0.60
71Z	40%	0.40

Step 2: Use the isotope masses

Isotope	Mass (amu)
69Z	69
71Z	71

Step 3: Apply the formula for the average atomic mass

$$
\text{Atomic mass of sample} = (69 \times 0.60) + (71 \times 0.40)
$$

Calculating each term:

$$
69 \times 0.60 = 41.4
$$

$$
71 \times 0.40 = 28.4
$$

Adding these:

$$
41.4 + 28.4 = 69.8 \text{ amu}
$$

Result: The relative atomic mass of the element Z in this sample is approximately 69.8 amu.

Significance of the Calculation

Understanding how to calculate the relative atomic mass from isotopic abundances is vital in various fields, including:

- Chemical analysis: Determining the isotopic composition of elements in samples.
- Mass spectrometry: Interpreting spectral data based on isotopic distribution.
- Nuclear chemistry: Calculating atomic weights for nuclear reactions.
- Environmental science: Tracing sources of elements based on isotopic signatures.

This calculation also illustrates the importance of isotopic abundances in defining the properties of elements as they appear in nature.

Additional Considerations

Isotopic Abundance Variability

In nature, isotopic abundances can vary depending on the source of the sample. For example, isotopic compositions of elements like carbon, oxygen, and nitrogen differ in biological and geological samples. Such variations can influence the calculated atomic mass.

Refining Atomic Mass Calculations

For high-precision work, the atomic masses used are often the standard atomic weights provided by organizations like IUPAC, which account for the natural isotopic distribution. These standard atomic weights are weighted averages based on global isotopic abundances.

Examples of Similar Calculations

Example 1:
A sample contains 70% of isotope 50X and 30% of isotope 52X.
Isotope masses: 50 amu and 52 amu.

Calculate the relative atomic mass:

$$
(50 \times 0.70) + (52 \times 0.30) = 35 + 15.6 = 50.6\, \text{amu}
$$

Example 2:
An element has isotopic abundances: 80% of 100Y and 20% of 102Y.

Calculate the average atomic mass:

$$
(100 \times 0.80) + (102 \times 0.20) = 80 + 20.4 = 100.4\, \text{amu}
$$

Practical Implications and Applications

- Chemical Formulations: Accurate atomic masses are essential for stoichiometry calculations in chemical reactions.
- Isotope Labeling: Used in research and medicine, understanding isotopic composition helps in designing experiments and treatments.
- Environmental Monitoring: Isotopic ratios help track pollution sources and natural processes.
- Nuclear Energy: Precise atomic mass calculations are crucial for nuclear reaction calculations, including fission and fusion processes.

Conclusion

Calculating the relative atomic mass of an element based on its isotopic composition is a fundamental skill in chemistry. By understanding the percentages of isotopes present and their respective atomic masses, we can determine an accurate average atomic weight. In the specific case of element Z with 60% of isotope 69Z and 40% of isotope 71Z, the relative atomic mass is approximately 69.8 amu. This value provides insight into the typical atomic weight of the element in this sample, which can be used in further chemical calculations and analyses.

Summary of Key Points

- Isotopes are atoms of the same element with different neutron counts.
- The relative atomic mass is a weighted average based on isotopic abundance.
- Convert percentage abundances to fractions before calculations.

- Apply the formula: \(\sum (\text{mass} \times \text{fraction})\).
- Sample calculation: 69.8 amu for the given composition.

Understanding these concepts ensures accurate chemical analysis and helps in various scientific applications, from research to industry.

Frequently Asked Questions

What is the method to calculate the relative atomic mass of element Z given the sample composition?

The relative atomic mass is calculated using the weighted average based on the isotopic percentages: (percentage of isotope 69Z × 69) + (percentage of isotope 71Z × 71), divided by 100.

How do you interpret the percentages 60% and 40% in the context of isotopic composition?

They indicate that 60% of the sample consists of isotope 69Z and 40% consists of isotope 71Z, which are used as weights in calculating the average atomic mass.

What is the formula to find the relative atomic mass given isotopic abundances?

Relative atomic mass = (abundance of isotope 1 × mass of isotope 1 + abundance of isotope 2 × mass of isotope 2) / 100.

Calculate the relative atomic mass of element Z based on the given sample composition.

Relative atomic mass = $(60 \times 69 + 40 \times 71) / 100 = (4140 + 2840) / 100 = 6980 / 100 = 69.8$.

Why is it important to consider isotopic percentages when calculating atomic mass?

Because elements can have multiple isotopes with different masses, the average atomic mass reflects the weighted contribution of each isotope based on their natural abundance.

Can the relative atomic mass of element Z be

approximated as 70 based on this data?

No, the calculated value is approximately 69.8, which is close to but not exactly 70; thus, the precise value is 69.8.

What assumptions are made in calculating the relative atomic mass from isotopic percentages?

It assumes that the sample's isotopic distribution accurately represents the natural abundance and that the isotopic masses are known and constant.

How does this method of calculating atomic mass apply to real-world element analysis?

It allows scientists to determine the average atomic mass of elements with multiple isotopes, which is essential for accurate chemical calculations and understanding atomic structure.

Additional Resources

Atomic Mass Calculation: Deciphering the Composition of a Sample of Element Z

In the realm of chemistry, understanding the atomic mass of an element is fundamental to numerous scientific and industrial applications. Whether you're a student, a researcher, or a professional in the field, grasping how to determine an element's relative atomic mass based on sample composition is a critical skill. Today, we delve into a fascinating problem: a sample of a mysterious Element Z contains a mixture of two isotopes, 69Z and 71Z, in specific proportions. Our goal is to accurately compute the relative atomic mass of Element Z based on this data.

Understanding the Basics: Isotopes and Atomic Mass

Before diving into calculations, it's essential to grasp the foundational concepts underpinning this problem.

What Are Isotopes?

- Definition: Isotopes are variants of a particular chemical element that have the same number of protons but different numbers of neutrons.
- Implication: Since isotopes of an element differ in neutron count, they have different

atomic masses but exhibit similar chemical properties.
- Relevance: The natural occurrence of isotopes influences the atomic mass of elements. The weighted average of all isotopic masses, considering their relative abundances, yields the element's atomic weight.

The Concept of Relative Atomic Mass

- Definition: The relative atomic mass (or atomic weight) of an element is the weighted average of the masses of its isotopes, based on their natural abundances.
- Formula:

$$
\text{Atomic Mass of Element} = \sum (\text{Mass of Isotope} \times \text{Fractional Abundance})
$$

- Application: When isotopic abundances are known, calculating the atomic mass becomes a straightforward weighted average.

Analyzing the Sample Composition

The problem states:

> "A certain sample of element Z contains 60% of 69Z and 40% of 71Z."

Let's interpret what this entails.

Isotope Identification

- The two isotopes are:

Isotope	Mass Number	Atomic Mass (approximate)	Abundance in the sample
69Z	69	69 u (atomic mass units)	60%
71Z	71	71 u	40%

- The percentages indicate the relative abundance of each isotope within the sample, which are critical for calculating the average atomic mass.

Assumptions and Simplifications

- Since the isotopic masses are close to their mass numbers, the approximate atomic masses are taken as 69 u and 71 u.
- The percentages are converted into fractional form for calculation purposes:

$$
\text{Fraction of } 69Z = 0.60 \\
\text{Fraction of } 71Z = 0.40
$$

Calculating the Relative Atomic Mass of Element Z

With the given data, we can now perform the core calculation.

Step 1: Establish the Weighted Average Formula

The relative atomic mass (Ar) of Element Z is:

$$
\text{Ar} = (\text{Mass of } 69Z \times \text{Fractional abundance}) + (\text{Mass of } 71Z \times \text{Fractional abundance})
$$

Substituting the known values:

$$
\text{Ar} = (69 \times 0.60) + (71 \times 0.40)
$$

Step 2: Perform the Calculations

Calculating each term:

- $69 \times 0.60 = 41.4$
- $71 \times 0.40 = 28.4$

Adding the two:

$$
\text{Ar} = 41.4 + 28.4 = 69.8
$$

Step 3: Final Result and Interpretation

The relative atomic mass of Element Z is approximately:

$$
\boxed{69.8 \text{ u}}
$$

This value indicates that, based on the sample's isotopic composition, the average atomic weight of Element Z is 69.8 atomic mass units.

Implications and Broader Context

Understanding how isotopic composition influences atomic mass is not merely academic; it has practical and scientific significance.

Natural Isotopic Abundance vs. Sample Data

- The sample's isotopic distribution (60% and 40%) may differ from the natural abundance of isotopes in Element Z.
- In nature, isotope ratios vary depending on geological, biological, or synthesis factors.
- The calculated atomic mass (69.8 u) reflects the specific sample composition, which might differ from the standard atomic weight.

Applications in Scientific and Industrial Fields

- Isotope Analysis: Used in geochemistry, archaeology, and environmental science to trace processes.
- Nuclear Chemistry: Understanding isotopic composition is critical in nuclear reactors and radiometric dating.
- Material Science: Tailoring isotopic ratios to improve material properties or for specialized applications.

Additional Considerations and Complexities

While this calculation appears straightforward, real-world scenarios can introduce complexities:

- Multiple Isotopes: Elements often have more than two isotopes, requiring a weighted average across all.
- Measurement Precision: Accurate determination of isotopic abundances often involves mass spectrometry, which can detect minute differences.
- Isotopic Enrichment: Certain processes can enrich specific isotopes, thus shifting the average atomic mass.

Conclusion: Mastering Atomic Mass Calculations

In summary, determining the relative atomic mass of an element from a known isotopic composition involves:

- Recognizing the isotopic masses and their proportions.
- Applying the weighted average formula.
- Calculating to find the element's atomic weight.

For the sample of Element Z under discussion, containing 60% of 69Z and 40% of 71Z, the computed atomic mass of approximately 69.8 u provides a clear insight into its isotopic makeup. This process underscores the importance of understanding isotopic distributions in various scientific applications, from fundamental chemistry to advanced research.

In essence, mastering the calculation of atomic masses based on isotopic composition is an invaluable skill that bridges theoretical knowledge and practical application, offering insights into the very fabric of matter.

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