Calculate The Average Atomic Mass For Gallium If 60.00% Of Its Atoms Have A Mass Of 68.926 Amu And 40.00%

Calculate The Average Atomic Mass For Gallium If 60.00% Of Its Atoms Have A Mass Of 68.926 Amu And 40.00% of its atoms have a different mass involves understanding the concept of weighted averages in atomic physics. Gallium, a metal element with the atomic number 31, naturally occurs as a mixture of isotopes—atoms with the same number of protons but different numbers of neutrons, resulting in different atomic masses. When analyzing such mixtures, scientists often calculate the average atomic mass to better understand the element's properties, especially for applications in chemistry, physics, and materials science. This article provides a comprehensive guide on how to perform this calculation, focusing on the specific data provided: 60.00% of gallium atoms have an atomic mass of 68.926 amu, and 40.00% have another mass.

Understanding Atomic Mass and Isotopic Composition

What Is Atomic Mass?

Atomic mass, measured in atomic mass units (amu), represents the average mass of atoms of an element, considering all their isotopes and their relative abundances. It reflects the weighted average that accounts for the natural distribution of isotopes. For example, if an element has isotopes with different masses, the overall atomic mass is calculated by multiplying each isotope's mass by its relative abundance and summing these values.

Isotopic Distribution and Its Significance

Elements often consist of isotopes, which differ in neutron count. Gallium, for instance, primarily exists as two isotopes:

- Gallium-69 (mass ≈ 68.9256 amu)
- Gallium-71 (mass ≈ 70.9247 amu)

However, in the context of your question, we're given specific percentages and masses, which simplifies the calculation process for the average atomic mass.

Calculating the Average Atomic Mass of Gallium

Step 1: Gather the Data

The data provided states:

- 60.00% of gallium atoms have a mass of 68.926 amu
- 40.00% of gallium atoms have another known mass (the question implies this, but it should be specified; for this example, assume it's 71.000 amu)

Note: Ensure the mass for the second isotope is known or provided to perform an accurate calculation.

Step 2: Convert Percentages to Decimal Form

To perform calculations, convert the percentages to decimal form:

- 60.00% = 0.6000
- 40.00% = 0.4000

Step 3: Apply the Weighted Average Formula

The formula for average atomic mass (A_{avg}) is:
$$A_{avg} = (f_1 \times m_1) + (f_2 \times m_2)$$
where:

- f_1 and f_2 are the fractional abundances of each isotope
- m_1 and m_2 are the atomic masses of each isotope

Step 4: Perform the Calculation

Using the assumed data:

- $f_1 = 0.6000$, $m_1 = 68.926$ amu
- $f_2 = 0.4000$, $m_2 = 71.000$ amu

Calculate:

$$ A_{avg} = (0.6000 \times 68.926) + (0.4000 \times 71.000) $$

$$
A_{avg} = (41.3556) + (28.4000) = 69.7556 \, \text{amu}
$$

Therefore, the average atomic mass of gallium, based on these isotope distributions, is approximately 69.756 amu.

Understanding the Significance of the Calculation

Why Is Calculating the Average Atomic Mass Important?

Knowing the average atomic mass of gallium helps chemists:

- Determine molar masses for stoichiometric calculations
- Predict chemical behavior and reactivity
- Identify isotopic compositions in natural samples
- Support research in nuclear physics and materials science

Real-World Applications of Gallium's Atomic Mass

Gallium's properties make it valuable in:

- Semiconductor technology (used in LEDs and solar cells)
- Medical imaging and diagnostics
- Research in superconductivity and quantum computing

Accurate knowledge of its atomic mass ensures precise scientific computations and material design.

Additional Considerations in Atomic Mass

Calculations

Isotope Abundance Variability

The relative abundance of isotopes can vary across sources and natural samples. Therefore, the average atomic mass might differ slightly based on the sample analyzed.

Isotopic Mass Precision

Atomic masses are measured with high precision, but small measurement errors can influence the calculated average. When performing calculations, use the most accurate data available.

Using Atomic Mass Data from Reliable Sources

Always refer to reputable sources like:

- International Union of Pure and Applied Chemistry (IUPAC)
- National Institute of Standards and Technology (NIST)
- Scientific journals and publications

Summary of the Calculation Process

To summarize, calculating the average atomic mass of gallium involves:

1. Gathering isotope masses and their relative abundances
2. Converting percentages to decimal fractions
3. Applying the weighted average formula
4. Performing the multiplication and addition operations

This process provides an accurate estimate of the atomic mass, essential in various scientific and industrial applications.

Conclusion

Calculating the average atomic mass of gallium when given the percentages and masses of its isotopes is a fundamental skill in chemistry. By understanding

the principles of isotopic distribution and weighted averages, scientists can accurately determine an element's atomic mass, which influences everything from chemical reactions to material properties. Remember to use precise data and reliable sources to ensure accuracy in your calculations. Whether you're studying gallium for academic purposes or applying it in technological innovations, mastering this calculation is crucial for interpreting and utilizing atomic data effectively.

Frequently Asked Questions

How do you calculate the average atomic mass of gallium given the percentages and atomic masses?

Multiply each isotope's atomic mass by its percentage (expressed as a decimal), then sum these values to find the average atomic mass.

What is the atomic mass contribution of the 60.00% of gallium atoms with a mass of 68.926 amu?

It contributes 0.60 × 68.926 = 41.3556 amu to the average atomic mass.

Given 40.00% of gallium atoms have an unknown mass, how do you find the average atomic mass?

First, determine the atomic mass of the remaining isotope by considering the total average and then calculate the weighted contribution.

What is the formula to find the average atomic mass when given percentages and isotope masses?

Average atomic mass = (percentage of isotope 1 × mass of isotope 1) + (percentage of isotope 2 × mass of isotope 2).

If 60.00% of gallium atoms have a mass of 68.926 amu, and the average atomic mass is to be calculated, what is the value for the remaining 40.00%?

First, calculate the contribution of the known percentage, then subtract from the total average to find the contribution of the remaining percentage, which can be used to find its atomic mass.

What is the calculated average atomic mass of gallium based on the given data?

The average atomic mass is approximately 69.105 amu, calculated as (0.60 × 68.926) + (0.40 × 69.753) = 69.105 amu.

How does the percentage composition affect the calculation of the average atomic mass?

The percentage determines the weight of each isotope's atomic mass in the overall average, so higher percentages have a greater impact.

Why is it important to convert percentages to decimals when calculating average atomic mass?

Because the calculation involves multiplying the atomic mass by the fractional representation of the percentage, which is the decimal form.

Can you explain the step-by-step process to compute the average atomic mass for gallium with the given data?

Yes. First, convert percentages to decimals, multiply each isotope's atomic mass by its decimal percentage, then sum the results to obtain the average atomic mass.

What are common errors to avoid when calculating average atomic mass from isotope data?

Common errors include mixing up percentages and decimals, forgetting to weight each isotope properly, or miscalculating the contribution of each isotope.

Additional Resources

Atomic Mass Calculation

Calculating the average atomic mass of an element is an essential process in chemistry that provides insight into the element's isotopic composition and its properties. In this article, we delve into the detailed methodology behind determining the average atomic mass of gallium (Ga) based on specific isotopic data. Our focus is on a scenario where 60.00% of gallium atoms have an atomic mass of 68.926 atomic mass units (amu), and the remaining 40.00% have an unknown mass, which we will calculate. This comprehensive exploration will not only guide you through the step-by-step calculation but also elucidate the underlying principles, making it a valuable resource for

students, educators, and science enthusiasts alike.

Understanding Atomic Mass and Isotopic Distribution

Before diving into the calculations, it’s crucial to understand what atomic mass signifies and how isotopic distribution influences the overall atomic mass of an element.

What Is Atomic Mass?

Atomic mass, measured in atomic mass units (amu), reflects the average mass of an atom of an element, considering all its isotopes and their respective abundances. Since atoms of an element can exist in different isotopic forms—variants with different neutron counts—the atomic mass is not a fixed value but an average weighted by the natural abundance of each isotope.

Isotopes and Their Significance

Isotopes are atoms of the same element with identical proton counts but differing neutron counts. For gallium:

- The most common isotopes are Ga-69 and Ga-71.
- However, in this scenario, we are provided specific isotopic masses and their respective abundances, focusing on one isotope with a known mass (68.926 amu) and the other with an unknown mass.

The average atomic mass formula accounts for this distribution:

$$
\text{Average Atomic Mass} = \sum_{i} (\text{abundance}_i \times \text{mass}_i)
$$

where each isotope's contribution is weighted by its relative abundance.

Given Data and Problem Setup

In our specific problem:

- 60.00% of gallium atoms have an atomic mass of 68.926 amu.
- The remaining 40.00% have an unknown atomic mass, which we will denote as M.

Our goal is to calculate the average atomic mass of gallium considering this isotopic distribution.

Key Data Points:

Isotope	Abundance	Atomic Mass (amu)
Isotope 1	60.00%	68.926
Isotope 2	40.00%	M

Expressed numerically:

- $\text{Abundance}_1 = 0.6000$
- $\text{Mass}_1 = 68.926\, \text{amu}$
- $\text{Abundance}_2 = 0.4000$
- $\text{Mass}_2 = M$ (unknown, to be calculated)

Step-by-Step Calculation Process

Calculating the average atomic mass involves applying the weighted average formula, which directly incorporates the isotopic abundances and their respective masses.

Step 1: Write the Average Atomic Mass Equation

$$
\text{Average Atomic Mass} = (0.6000 \times 68.926) + (0.4000 \times M)
$$

This equation states that 60% of the atoms contribute to the average based on their mass, and 40% contribute based on the unknown mass M.

Step 2: Determine the Known Components

Calculate the contribution of the known isotope:

$$
0.6000 \times 68.926 = 41.3556\, \text{amu}
$$

This is the partial contribution to the average from the isotope with known mass.

Step 3: Use Additional Data or Assumptions to Find M

In typical problems, if the overall average atomic mass of gallium is known—say, from standard atomic weight data—this can help determine the unknown isotope’s mass.

Common atomic weight of gallium: approximately 69.723 amu.

Assuming this is the intended average atomic mass for the purpose of this calculation:

$$
\text{Average Atomic Mass} = 69.723\, \text{amu}
$$

Set up the equation:

$$
69.723 = 41.3556 + 0.4000 \times M
$$

Calculating the Unknown Isotope's Mass

Step 4: Solve for M

Rearranged formula:

$$
0.4000 \times M = 69.723 - 41.3556
$$

$$
0.4000 \times M = 28.3674
$$

$$

Divide both sides by 0.4000:

$$
M = \frac{28.3674}{0.4000} = 70.9185\, \text{amu}
$$

Result:

$$
\boxed{
M \approx 70.919\, \text{amu}
}
$$

This value represents the atomic mass of the second isotope of gallium contributing to the overall average.

Interpretation and Significance of the Results

The calculated atomic mass M, approximately 70.919 amu, aligns closely with the atomic mass of gallium's second isotope, Ga-71 (which has an atomic mass around 70.924 amu). This consistency reinforces the validity of the calculation and highlights the importance of isotopic contributions in defining an element's atomic weight.

Implications:

- The calculated isotopic mass indicates that the second isotope is likely Ga-71, matching real-world data.
- The average atomic mass of gallium, based on these proportions, is approximately 69.723 amu, consistent with standard atomic weight values.

Additional Insights and Applications

Understanding how to compute the average atomic mass from isotopic data is foundational in several scientific fields:

- Chemistry and Material Science: Helps in characterizing materials and understanding isotopic compositions.
- Nuclear Physics: Assists in identifying isotopic abundances in natural

samples.
- Environmental Science: Used in tracing sources of elements through isotopic signatures.

Practical Tips:

- Always verify the known average atomic mass for the element to guide calculations.
- Convert percentage abundances to decimal form for calculations.
- Remember that isotopic masses are often close to whole numbers but may include decimal points due to nuclear binding energy differences.

Conclusion

Calculating the average atomic mass of gallium given specific isotopic abundances demonstrates the interplay between isotopic composition and atomic weight. By applying the weighted average formula and utilizing known data, we found that if 60% of atoms have a mass of 68.926 amu, the remaining 40% must have a mass of approximately 70.919 amu to yield an average atomic mass consistent with real-world values.

This process exemplifies core principles of atomic structure and isotopic distribution, emphasizing the importance of precise measurements and calculations in understanding the fundamental properties of elements. Whether for academic purposes, research, or practical applications, mastering this calculation enhances your comprehension of the atomic world and the nuanced nature of elemental masses.

Note: Always cross-reference calculated values with experimental data for accuracy, especially when dealing with isotopic compositions and atomic weights.

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