

Find The Molecular Formula For A Compound That Contains 4.90 % N And 11.2 % O. The Molecular Mass Of

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Understanding how to determine the molecular formula of a compound based on its percent composition is a fundamental skill in chemistry. This process involves converting percentage data into the empirical formula, and then using the molecular mass to find the molecular formula. In this article, we will walk through this process step-by-step for a compound that contains 4.90% nitrogen (N) and 11.2% oxygen (O). Although the molecular mass is not specified in the initial data, we will discuss how to proceed once that information becomes available or how to estimate it if needed.

Understanding Percent Composition and Its Significance

What is Percent Composition?

Percent composition indicates how much of each element is present in a compound by mass. It is expressed as a percentage of the total mass of the compound. For instance, a compound with 4.90% N contains 4.90 grams of nitrogen per 100 grams of compound.

Why Percent Composition Matters

The percent composition provides the first clue in deducing the empirical formula of a compound. By converting these percentages into mole ratios, chemists can determine the simplest whole-number ratio of atoms within the compound, known as the empirical formula.

Step-by-Step Process to Find the Empirical and Molecular Formulas

Step 1: Convert Percentages to Masses

- Assuming a 100-gram sample of the compound, the masses of elements are:
- Nitrogen: 4.90 g
- Oxygen: 11.2 g
- Remaining mass (if any) corresponds to other elements or the total mass; in this case, since only N and O are given, the rest of the mass is not specified, but for the purpose of calculation, focus on these percentages.

Step 2: Convert Masses to Moles

Use atomic masses of elements to convert the masses into molar quantities:

- **Atomic masses:** N $\approx$ 14.01 g/mol; O $\approx$ 16.00 g/mol
- Number of moles of nitrogen:

$$n_{\text{N}} = 4.90 \text{ g} / 14.01 \text{ g/mol} \approx 0.35 \text{ mol}$$

- Number of moles of oxygen:

$$n_{\text{O}} = 11.2 \text{ g} / 16.00 \text{ g/mol} \approx 0.70 \text{ mol}$$

Step 3: Determine the Mole Ratio

Find the simplest whole-number ratio by dividing each mole quantity by the smallest number of moles:

Smallest number of moles = 0.35 mol (N)

$$\text{N: } 0.35 / 0.35 = 1$$

$$\text{O: } 0.70 / 0.35 = 2$$

Thus, the empirical formula ratio is N_1O_2 , or simply NO_2 .

Step 4: Establish the Empirical Formula

Based on the mole ratio, the empirical formula of the compound is NO_2 .

Calculating the Molecular Formula

Step 1: Find the Empirical Formula Mass

- Empirical formula: NO_2
- Empirical formula mass:

$$1 \times 14.01 + 2 \times 16.00 = 14.01 + 32.00 = 46.01 \text{ g/mol}$$

Step 2: Use the Molecular Mass to Find the Multiplier

The molecular formula is a whole-number multiple of the empirical formula. To find this multiple (n), use:

$$n = (\text{Molecular mass}) / (\text{Empirical formula mass})$$

Since the molecular mass is not provided, you must either be given it or estimate it based on other data or context.

Step 3: Determine the Molecular Formula

- If the molecular mass is known, multiply each subscript in the empirical formula by n to get the molecular formula.
- For example, if the molecular mass is 92 g/mol:

$$n = 92 / 46.01 \approx 2$$

- Thus, the molecular formula would be $(\text{NO}_2)_2 = \text{N}_2\text{O}_4$.

Additional Considerations and Practical Applications

Estimating the Molecular Mass When Unknown

If the molecular mass is not specified, other techniques such as mass spectrometry can determine it. Alternatively, if you have data such as molar volume, density, or boiling point, these can sometimes help estimate molecular weight.

Common Mistakes to Avoid

- Using incorrect atomic masses.
- Forgetting to reduce the mole ratios to the simplest whole numbers.
- Assuming the entire compound is composed solely of N and O when other elements may be present.

Summary and Final Remarks

To find the molecular formula of a compound based on percent composition, you need to convert mass percentages into mole ratios, determine the empirical formula, and then use the molecular mass to find the molecular formula. In the case discussed here, with 4.90% N and 11.2% O, the empirical formula is N_1O_2 . Once the molecular mass is known, multiply the empirical formula subscripts by the appropriate factor to determine the molecular formula. This approach is fundamental in analytical chemistry and helps in understanding the composition and structure of unknown compounds.

Practical Example Recap

1. Assume a 100 g sample; convert percentages to masses.
2. Convert masses to moles using atomic weights.
3. Calculate the mole ratio and simplify to the smallest whole numbers.

4. Determine the empirical formula from the mole ratio.
5. Use the molecular mass to find the multiple of the empirical formula, thus obtaining the molecular formula.

Understanding these steps enhances your ability to analyze unknown compounds and deduce their chemical formulas accurately. Whether for academic purposes, research, or industrial applications, mastering this process is essential for any chemist.

Frequently Asked Questions

How do I determine the empirical formula from the given percentages of N and O?

Convert the percentages to moles by dividing each by their atomic masses (N: 14.01 g/mol, O: 16.00 g/mol). Then, divide each by the smallest number of moles to find the ratio, which gives the empirical formula.

What is the first step to find the molecular formula once I have the empirical formula?

Calculate the molar mass of the empirical formula and divide the given molecular mass by this value to find a factor. Multiply the empirical formula subscripts by this factor to get the molecular formula.

Given 4.90% N and 11.2% O, how do I find the empirical formula?

Assuming a 100 g sample, there are 4.90 g N and 11.2 g O. Convert to moles: $N = 4.90 / 14.01 \approx 0.35 \text{ mol}$, $O = 11.2 / 16.00 \approx 0.70 \text{ mol}$. Divide both by the smallest (0.35), resulting in N: 1, O: 2. Empirical formula: NO_2 .

How do I use the molecular mass to find the molecular formula from the empirical formula?

Calculate the empirical formula mass (e.g., NO_2 : $14.01 + 2 \times 16.00 = 46.01 \text{ g/mol}$). Divide the given molecular mass by this value to find the multiple (e.g., if molecular mass is 92 g/mol, $92 / 46.01 \approx 2$). Multiply the subscripts in the empirical formula by this multiple to get the molecular formula (N_2O_4).

What is the significance of the percentage composition in molecular formula determination?

It provides the ratio of elements in the compound, allowing you to derive the empirical formula. Knowing the molecular mass then helps to scale the empirical formula to the actual molecular formula.

Can I determine the molecular formula without knowing the molecular mass?

No, the molecular mass is essential to determine how many empirical units are in a molecule. Without it, you can only identify the empirical formula, not the exact molecular formula.

Additional Resources

Find The Molecular Formula For A Compound That Contains 4.90 % N And 11.2 % O. The Molecular Mass Of: Unlocking the Structural Secrets of a Chemical Compound

In the realm of chemical analysis, determining the molecular formula of a compound is a fundamental step toward understanding its structure, reactivity, and potential applications. When a compound's elemental composition is known—particularly the percentages of certain elements—scientists can deduce its molecular formula through systematic calculations. Today, we explore how to find the molecular formula of a compound based on its nitrogen (N) and oxygen (O) content, with additional information about its molecular mass. This process combines principles of percentage composition, mole calculations, and molecular weight determination, offering a window into the molecular architecture of the substance.

Understanding the Basics: Percentage Composition and Molecular Formula

Before delving into calculations, it's essential to grasp what percentage composition implies and how it relates to the molecular formula.

Percentage Composition: The percentage of an element in a compound indicates how much of the total mass is due to that element. For instance, a compound with 4.90% nitrogen means that in 100 grams of the compound, 4.90 grams are nitrogen.

Molecular Formula: The molecular formula specifies the number of each type of

atom in a molecule of the compound. It is a multiple of the empirical (simplest) formula, which represents the simplest whole-number ratio of atoms.

The process typically involves three key steps:

1. Determine the mass of each element in a hypothetical 100 g sample.
2. Convert these masses to moles to find the mole ratio.
3. Derive the empirical formula and then scale it to the molecular formula using molar mass.

Step 1: Calculating the Masses of Nitrogen and Oxygen

Given the percentage composition:

- Nitrogen (N): 4.90%
- Oxygen (O): 11.2%

Assuming a 100 g sample simplifies calculations:

- Mass of N = 4.90 g
- Mass of O = 11.2 g

Remaining mass in the sample is attributable to other elements, likely carbon or hydrogen, but since only nitrogen and oxygen are specified, the remaining percentage ($100\% - 4.90\% - 11.2\% = 83.9\%$) must be accounted for by other elements present in the compound. For this analysis, we will focus solely on N and O, assuming the rest of the composition is either not specified or negligible for our calculation. If more data were provided, we could incorporate additional elements.

Step 2: Convert Masses to Moles

The next step involves converting the masses into moles, using atomic weights:

- Atomic weight of Nitrogen (N): approximately 14.01 g/mol
- Atomic weight of Oxygen (O): approximately 16.00 g/mol

Calculate moles:

- Moles of N = $4.90 \text{ g} / 14.01 \text{ g/mol} \approx 0.350 \text{ mol}$
- Moles of O = $11.2 \text{ g} / 16.00 \text{ g/mol} = 0.700 \text{ mol}$

This ratio of moles indicates the relative number of atoms in the empirical formula.

Step 3: Determine the Empirical Formula

To find the simplest whole-number ratio, divide both mole quantities by the smaller of the two:

- For N: $0.350 \text{ mol} / 0.350 \text{ mol} = 1$
- For O: $0.700 \text{ mol} / 0.350 \text{ mol} = 2$

Thus, the empirical formula is approximately N_1O_2 , which we can write as NO_2 .

Step 4: Calculate the Empirical Formula Mass

Sum the atomic weights:

- N: $14.01 \text{ g/mol} \times 1 = 14.01 \text{ g/mol}$
- O_2 : $16.00 \text{ g/mol} \times 2 = 32.00 \text{ g/mol}$

Empirical formula mass = $14.01 + 32.00 = 46.01 \text{ g/mol}$

Step 5: Incorporate the Molecular Mass to Find the Molecular Formula

Now, if the molecular mass of the compound is known (which the beginning of the article suggests is a key part of the query), the molecular formula can be determined.

Suppose the molecular mass is given or measured—say, for example, it is 92.02 g/mol . To find the molecular formula:

- Divide the molecular mass by the empirical formula mass:

$92.02 \text{ g/mol} \div 46.01 \text{ g/mol} \approx 2$

- Multiply the empirical formula subscripts by this factor:

- N: $1 \times 2 = 2$

- O: $2 \times 2 = 4$

Therefore, the molecular formula is N_2O_4 .

If the exact molecular mass differs, the same approach applies: divide the molecular mass by the empirical formula mass, and multiply the subscripts accordingly.

Interpreting the Results and Practical Considerations

What does the molecular formula tell us?

The molecular formula N_2O_4 indicates the molecule contains two nitrogen atoms and four oxygen atoms. It is a known compound—dinitrogen tetroxide—a highly reactive, colorless gas used as an oxidizer in rocket propulsion.

Additional factors to consider:

- Purity of sample: Real-world samples may contain impurities affecting the percentage composition.

- Presence of other elements: If other elements are present, the initial assumptions need revising.

- Measurement accuracy: Precise elemental analysis and molecular mass measurement are crucial for accurate determination.

Implications for chemistry and industry:

Accurate molecular formulas allow chemists to understand the compound's reactivity and potential applications—a vital step in fields ranging from pharmaceuticals to materials science.

Summary: The Systematic Approach to Finding Molecular Formulas

The process exemplified above underscores a systematic approach:

1. Use percentage composition to find the mass of each element in a hypothetical sample.

2. Convert these masses to moles using atomic weights.

3. Derive the empirical formula by calculating mole ratios.
4. Calculate the empirical formula mass.
5. Use the known molecular mass to find the multiple of the empirical formula, thereby obtaining the molecular formula.

This methodology is foundational in analytical chemistry, enabling scientists to decode the molecular architecture of unknown compounds accurately.

Conclusion: Bridging Composition to Structure

Determining the molecular formula from elemental percentages and molecular mass is a cornerstone of chemical analysis. It transforms raw compositional data into a tangible structural insight, guiding further research and application development. Whether in academic laboratories or industrial settings, mastering this process empowers chemists to unlock the secrets held within molecular structures, advancing science and technology alike.

Note: The specific molecular mass for the compound was not provided in the initial query. If a particular molecular mass value is available, replace the hypothetical 92.02 g/mol in the calculations with that value to obtain the exact molecular formula.

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