Decomposition Of Ni3 Produces Nitrogen Gas And Iodine. How Many Grams Of The Reactant Would Be Required

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Understanding chemical reactions and stoichiometry is fundamental in chemistry, especially when it involves predicting the amount of reactants needed to produce desired products. One such reaction involves the decomposition of Ni3, which yields nitrogen gas (N_2) and iodine (I_2). In this comprehensive guide, we will explore the chemical equation, analyze the molar relationships, and demonstrate how to determine the precise amount of reactant required in grams to produce a specific quantity of nitrogen gas and iodine. Whether you're a student preparing for exams, a professional chemist, or a science enthusiast, this detailed explanation aims to clarify the process step-by-step.

Understanding the Decomposition of Ni3: The Reaction Equation

Before calculating the amount of reactant needed, it is essential to understand the chemical nature of the decomposition process.

The Chemical Identity of Ni3

- Ni3 is not a common or standard chemical notation for a nickel compound.
- It is likely a shorthand or a typo, but for this exercise, let's assume Ni3 refers to a trimeric nickel compound, possibly analogous to a compound like nickel(III) nitride or a similar complex.
- Alternatively, it may refer to a generic nickel species involved in decomposition leading to nitrogen and iodine.

Given the context, and for the purpose of this problem, we will assume the reaction involves a hypothetical compound, Ni3, which decomposes into nitrogen gas and iodine.

The Balanced Chemical Equation

Suppose the decomposition reaction is represented as:

\[
\text{Ni}_3 \rightarrow \text{N}_2 + \text{I}_2
\]

However, this is an oversimplification and unlikely to be chemically accurate. For an instructive purpose, let's hypothesize a more chemically plausible reaction:

\[
\text{NiI}_3 \rightarrow \text{N}_2 + \text{I}_2 + \text{Ni}
\]

But since the problem states that Ni3 decomposes to produce nitrogen gas and iodine, perhaps a generic reaction is:

\[
\text{Ni}_3 \text{X}_x \rightarrow \text{N}_2 + \text{I}_2 + \text{other products}
\]

Given the ambiguity, and for clarity, let's define a hypothetical balanced reaction:

Hypothetical Reaction:

\[
\text{Ni}_3 \text{(some compound)} \rightarrow \text{N}_2 + \text{I}_2
\]

Alternatively, if the decomposition produces N_2 and I_2 directly from Ni3, the reaction could be:

\[
\text{Ni}_3 \rightarrow \text{N}_2 + \text{I}_2
\]

But this is not chemically consistent unless Ni3 is a compound containing nitrogen and iodine.

Assumption and Approach for Calculation

Given the ambiguity, we will proceed with a hypothetical balanced chemical equation explicitly involving Ni3 decomposing into nitrogen gas and iodine:

Assumed Balanced Equation:

\[

$$
\text{Ni}_3 \rightarrow 1\, \text{N}_2 + 3\, \text{I}_2
$$

Note: This is an illustrative example designed to show the process of calculating reactant amounts based on stoichiometry.

Step-by-Step Calculation of Reactant Mass

To determine how many grams of Ni3 are required to produce a specific amount of nitrogen gas and iodine, follow these steps:

1. Write the Balanced Chemical Equation

$$
\text{Ni}_3 \rightarrow \text{N}_2 + 3\, \text{I}_2
$$

- 1 mole of Ni3 decomposes to produce 1 mole of N_2 and 3 moles of I_2.

2. Identify the Molar Mass of Ni3

- Atomic mass of Nickel (Ni) ≈ 58.69 g/mol
- Molar mass of Ni3 = 3 × 58.69 g/mol ≈ 176.07 g/mol

3. Determine the Moles of Products Desired

Suppose the goal is to produce a certain amount of nitrogen gas, say x grams.

- Convert grams of N_2 to moles:

$$
\text{Moles of N}_2 = \frac{\text{grams of N}_2}{\text{molar mass of N}_2}
$$

- Molar mass of N_2 ≈ 28.02 g/mol

Similarly, for iodine, if you want to produce y grams:

$$
\text{Moles of I}_2 = \frac{\text{grams of I}_2}{253.81\, \text{g/mol}}
$$

Calculating the Reactant Required Based on Product Desired

Scenario 1: Producing a Specific Mass of Nitrogen Gas

Suppose you want to produce 50 grams of nitrogen gas.

Step 1: Convert grams of N_2 to moles:

$$
\text{Moles of N}_2 = \frac{50\, \text{g}}{28.02\, \text{g/mol}} \approx 1.78\, \text{mol}
$$

Step 2: Use the molar ratio from the balanced equation:

$$
\text{1 mol Ni}_3 \rightarrow \text{1 mol N}_2
$$

Thus, moles of Ni3 needed = moles of N_2 produced = 1.78 mol

Step 3: Calculate the mass of Ni3 required:

$$
\text{Mass of Ni}_3 = 1.78\, \text{mol} \times 176.07\, \text{g/mol} \approx 313.4\, \text{g}
$$

Result: To produce 50 grams of nitrogen gas, approximately 313.4 grams of Ni3 are required.

Scenario 2: Producing a Specific Mass of Iodine Gas

Suppose you want to produce 100 grams of iodine (I_2).

Step 1: Convert grams of I_2 to moles:

$$
\text{Moles of I}_2 = \frac{100\, \text{g}}{253.81\, \text{g/mol}} \approx 0.394\, \text{mol}
$$

Step 2: Use the molar ratio:

$$
\text{1 mol Ni}_3 \rightarrow 3\, \text{I}_2
$$

Thus, moles of Ni3 needed = $\frac{0.394}{3} \approx 0.131\, \text{mol}$

Step 3: Calculate the mass of Ni3:

$$
\text{Mass of Ni}_3 = 0.131\, \text{mol} \times 176.07\, \text{g/mol} \approx 23.07\, \text{g}
$$

Result: To produce 100 grams of iodine gas, approximately 23.07 grams of Ni3 are required.

Additional Considerations in Real-World Calculations

While the above calculations provide a clear method for determining reactant amounts based on product quantities, several practical factors should be considered:

Reaction Efficiency

- Actual reactions rarely proceed with 100% efficiency.
- Adjust calculations to account for reaction yield:

$$
\text{Adjusted grams of Ni}_3 = \frac{\text{theoretical grams}}{\text{reaction efficiency}}
$$

- For example, if the yield is 80%, multiply the theoretical amount by 1/0.8 = 1.25.

Purity of Reactants

- Commercial reactants may contain impurities.
- Adjust calculations based on the purity percentage.

Safety and Handling

- Decomposition reactions involving nitrogen and iodine can produce hazardous conditions.
- Ensure proper safety protocols are followed.

Environmental Considerations

- Manage waste materials responsibly.
- Use environmentally friendly methods where possible.

Summary and Final Thoughts

The decomposition of Ni3 to produce nitrogen gas and iodine involves understanding the molar relationships dictated by the balanced chemical equation. By identifying the molar masses of the reactant and products, and applying stoichiometric ratios, you can accurately determine how many grams of Ni3 are necessary to generate a desired amount of N_2 or I_2. This process underscores the importance of precise measurements and calculations in chemical synthesis and laboratory practices.

In summary:
- Write and balance the chemical equation.
- Convert desired product mass to moles.
- Use molar ratios to find the required moles of reactant.
- Convert moles of reactant to grams based on molar mass.
- Adjust for reaction efficiency and purity as needed.

Understanding these principles not only aids in laboratory preparation but also enhances the comprehension of chemical reactions' quantitative aspects, essential for research, industrial applications, and academic pursuits.

Note: For specific reactions involving real compounds, always refer to verified chemical equations and data. The hypothetical scenario presented here serves educational purposes to illustrate the process of stoichiometric calculations related to decomposition reactions.

Frequently Asked Questions

What is the balanced chemical equation for the decomposition of Ni_3 that produces nitrogen gas and iodine?

The balanced equation is $2\ Ni_3 \rightarrow 3\ Ni + N_2 + 3\ I_2$.

How do I calculate the mass of Ni_3 needed to produce a specific amount of nitrogen gas?

Use stoichiometry: convert the desired moles of N_2 to moles of Ni_3 using the balanced equation, then multiply by the molar mass of Ni_3 to find the required mass.

What is the molar mass of Ni_3, and how does it relate to the amount needed for the reaction?

The molar mass of Ni_3 is approximately 222.63 g/mol (3 × 58.69 g/mol). It helps convert moles of Ni_3 to grams needed for the reaction.

If I want to produce 10 grams of nitrogen gas, how much Ni_3 do I need to react?

First, calculate moles of N_2: 10 g / 28.02 g/mol ≈ 0.357 mol. From the balanced equation, 2 mol Ni_3 produce 1 mol N_2, so moles of Ni_3 needed: 2 × 0.357 = 0.714 mol. Then, grams of Ni_3: 0.714 mol × 222.63 g/mol ≈ 159 g.

What are the practical applications of understanding the decomposition of Ni_3 to produce nitrogen and iodine?

This knowledge is useful in industrial processes involving nitrogen and iodine production, materials synthesis, and understanding reaction mechanisms involving nickel compounds.

Are there any safety considerations when handling the decomposition of Ni_3?

Yes, nitrogen gas can displace oxygen, potentially causing suffocation, and iodine vapors are toxic and corrosive. Proper ventilation and protective equipment are essential.

How can I experimentally determine the amount of Ni_3 needed for a reaction?

By measuring the desired amount of nitrogen gas to be produced, applying stoichiometry based on the balanced equation, and calculating the corresponding mass of Ni_3 required.

What factors could affect the accuracy of the calculated amount of Ni_3 in the reaction?

Purity of reactants, incomplete reactions, measurement errors, and reaction conditions such as temperature and pressure can impact the accuracy of the calculations.

Additional Resources

Decomposition Of Ni_3 Produces Nitrogen Gas And Iodine. How Many Grams Of The Reactant Would Be Required

Understanding chemical reactions at a molecular level is fundamental in both academic research and industrial applications. One intriguing reaction involves the decomposition of a compound denoted as Ni_3, which results in the formation of nitrogen gas (N_2) and iodine (I_2). This process raises important questions about the stoichiometry involved—specifically, how much of the initial reactant, Ni_3, must be used to produce a desired quantity of nitrogen and iodine. This article explores the detailed mechanisms behind this decomposition, provides a step-by-step calculation to determine the required mass of Ni_3, and discusses the broader implications of such reactions in chemical manufacturing and laboratory settings.

Understanding the Decomposition of Ni_3: The Chemical Equation

Before delving into calculations, it is essential to clarify the nature of the reactant Ni_3 and its decomposition pathway. Based on typical chemical nomenclature and known reactions involving nickel compounds, Ni_3 likely refers to a nickel-rich compound, possibly a trimeric nickel complex or a compound containing nickel and iodine.

Assumed Reaction Pathway:

Given that the products are nitrogen gas (N_2) and iodine (I_2), the reaction suggests the presence of nitrogen and iodine in the original compound. A plausible decomposition reaction could be represented as:

$Ni_3X_n \rightarrow N_2 + I_2$ + other products?

However, since the original statement specifies that Ni_3 produces N_2 and I_2 upon decomposition, and considering nickel's chemistry, a more accurate representation might involve a compound such as nickel iodide or a nitrogen-containing nickel complex.

For the purpose of this calculation, let us assume the decomposition involves a compound with the formula $Ni_3N_3I_3$, which contains nickel, nitrogen, and iodine in a balanced proportion. The decomposition might be simplified as:

$Ni_3N_3I_3 \rightarrow 3\ N_2 + 3\ I_2 + 3\ Ni$ (possibly as a separate phase)

But since our focus is on the overall mass of Ni_3 needed to produce a certain amount of N_2 and I_2, the key is to establish a balanced chemical equation that relates the reactant to the products.

Establishing the Balanced Chemical Equation

To proceed accurately, we need to specify a balanced equation that reflects the decomposition process.

Hypothetical balanced reaction:

$$
\mathrm{Ni_3N_3I_3} \rightarrow 3\, \mathrm{N_2} + 3\, \mathrm{I_2}
$$

In this hypothetical reaction:

- 1 mole of $Ni_3N_3I_3$ decomposes into 3 moles of nitrogen gas and 3 moles of iodine.

Note: This is a simplified assumption for illustration; actual reactions may vary depending on the specific compound.

Calculating the Molar Masses

To determine how many grams of Ni_3 are required to produce a desired amount of N_2 and I_2, we need the molar masses of the relevant species.

Species	Molar Mass (g/mol)
Nickel (Ni)	58.69
Nitrogen (N_2)	28.02 (per N_2 molecule)
Iodine (I_2)	253.81

Molar mass of $Ni_3N_3I_3$:

$$
(3 \times 58.69) + (3 \times 14.01) + (3 \times 126.90) = 176.07 + 42.03 + 380.70 = 598.80\, \text{g/mol}
$$

(Note: The molar mass of nitrogen atom is 14.01 g/mol, and iodine atom is 126.90 g/mol.)

Step-by-Step Calculation: From Desired Product to Required Reactant

Suppose a chemist wants to produce 10 grams of nitrogen gas (N_2) and 10 grams of iodine (I_2). How much Ni_3 compound must be decomposed?

1. Moles of Products Needed

- Moles of N_2:

$$
\text{Moles of N}_2 = \frac{10\, \text{g}}{28.02\, \text{g/mol}} \approx 0.357\, \text{mol}
$$

- Moles of I_2:

$$
\text{Moles of I}_2 = \frac{10\, \text{g}}{253.81\, \text{g/mol}} \approx 0.0394\, \text{mol}
$$

2. Determine the Limiting Reactant

From the balanced equation:

$$
\mathrm{Ni_3N_3I_3} \rightarrow 3\, \mathrm{N_2} + 3\, \mathrm{I_2}
$$

- 1 mol of $Ni_3N_3I_3$ produces 3 mol of N_2 and 3 mol of I_2.

Calculate the amount of $Ni_3N_3I_3$ needed for the desired amount of each product:

- For N_2:

$$
\text{Moles of Ni}_3\text{ needed} = \frac{0.357\, \text{mol N}_2}{3} \approx 0.119\, \text{mol}
$$

- For I_2:

$$
\text{Moles of Ni}_3\text{ needed} = \frac{0.0394\, \text{mol I}_2}{3} \approx 0.0131\, \text{mol}
$$

The smaller value indicates I_2 production requires fewer moles of $Ni_3N_3I_3$, so the limiting factor is I_2.

Therefore, to produce 10 g of I_2, only approximately 0.0131 mol of $Ni_3N_3I_3$ is needed.

Calculating the Mass of $Ni_3N_3I_3$ Required

$$
\text{Mass} = \text{moles} \times \text{molar mass} = 0.0131\, \text{mol} \times 598.80\, \text{g/mol} \approx 7.84\, \text{g}
$$

Thus, approximately 7.84 grams of $Ni_3N_3I_3$ would be required to produce 10 grams of I_2 and, correspondingly, about 30 grams of N_2 (since 3 mol N_2 per mol $Ni_3N_3I_3$).

Broader Implications and Practical Considerations

Industrial and Laboratory Relevance

The ability to precisely calculate the amount of reactant needed for specific product yields is vital in chemical manufacturing, environmental chemistry, and research labs. Whether synthesizing reactive gases or managing resource efficiency, understanding stoichiometry ensures optimal use of materials, minimizes waste, and enhances safety.

Safety Precautions

Decomposition reactions involving gases like nitrogen and iodine can be hazardous. Nitrogen gas is inert but can displace oxygen in enclosed spaces, posing asphyxiation risks. Iodine vapors are toxic and corrosive. Proper ventilation, protective equipment, and adherence to safety protocols are essential when conducting such reactions.

Potential Challenges

- Purity of Reactants: Impurities can affect reaction pathways and yields.
- Reaction Conditions: Temperature, pressure, and catalysts influence the decomposition process.
- Reaction Completeness: Not all reactions proceed to completion; kinetic factors can lead to partial decomposition.

Conclusion

The decomposition of Ni_3 (or assumed $Ni_3N_3I_3$) into nitrogen gas and iodine exemplifies the importance of stoichiometric calculations in chemistry. By understanding the balanced chemical equation, molar relationships, and desired product quantities, chemists can accurately determine how much of the initial compound is required. In our illustrative scenario, approximately 7.84 grams of $Ni_3N_3I_3$ would be necessary to produce 10 grams each of nitrogen

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and iodine gases. Such calculations underpin efficient and safe chemical
processes, emphasizing the blend of theoretical knowledge and practical
application in the field of chemistry.

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Note: The specific compound and reaction pathway used here are hypothetical
for illustrative purposes. Real-world applications should reference precise
chemical identities and experimental data.
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