

Complete Combustion Of A 0.350 G Sample Of A Compound In A Bomb Calorimeter Releases 14.0 KJ Of Heat.

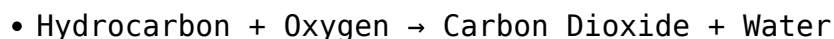
Complete Combustion Of A 0.350 G Sample Of A Compound In A Bomb Calorimeter Releases 14.0 KJ Of Heat

Complete combustion of a 0.350 g sample of a compound in a bomb calorimeter releases 14.0 kJ of heat. This statement summarizes a fundamental experiment in thermochemistry, providing valuable insights into the energy content of chemical substances. Understanding how to interpret this data allows chemists to determine the molar enthalpy of combustion, which is crucial for applications ranging from fuel efficiency to environmental impact assessments. In this article, we will explore the principles behind combustion reactions, how bomb calorimetry operates, calculations to determine molar enthalpy, and the significance of these findings in scientific and industrial contexts.

Understanding Combustion Reactions

What Is Complete Combustion?

Complete combustion occurs when a hydrocarbon or other organic compound reacts fully with oxygen, producing carbon dioxide and water as the primary products. The general reaction can be summarized as:



For example, the combustion of a generic hydrocarbon, C_xH_y , can be represented as:



This process releases energy, which is harnessed in various energy sources such as fuels.

Significance of Complete Combustion Data

Knowing the heat released during complete combustion helps in:

- Calculating the energy content of fuels
- Designing efficient combustion systems
- Assessing environmental impacts

The precise measurement of heat released, as in the case of the 0.350 g sample, provides a basis for these applications.

Bomb Calorimetry: The Method for Measuring Heat of Combustion

Principles of Bomb Calorimetry

A bomb calorimeter is a device used to measure the heat released or absorbed during a chemical reaction, especially combustion. The key features include:

- A sealed, oxygen-filled container (the "bomb") where the reaction occurs
- An insulated environment to prevent heat exchange with surroundings
- A thermometer and a calorimeter bucket containing water to absorb heat

When the sample combusts inside the bomb, the heat released is transferred to the surrounding water, causing a temperature change that can be measured precisely.

Procedure of a Typical Bomb Calorimetry Experiment

1. Weigh the sample accurately (here, 0.350 g).
2. Place the sample in the combustion chamber within the bomb.
3. Fill the bomb with pure oxygen at a known pressure.
4. Submerge the bomb in water within the calorimeter.
5. Initiate combustion via an electric ignition system.
6. Record the temperature change (ΔT) of the water.
7. Use the data to calculate the heat released.

Advantages of Bomb Calorimetry

- Accurate measurement of combustion heat
- Controlled environment minimizes heat loss
- Suitable for solid and liquid samples

Calculations: From Heat Released to Molar Enthalpy of Combustion

Given Data

- Mass of sample, $m = 0.350 \text{ g}$
- Heat released, $Q = 14.0 \text{ kJ}$

Step 1: Determine Moles of the Sample

To find the molar enthalpy of combustion, first calculate the number of moles:

$$\text{Number of moles } (n) = \frac{\text{mass}}{\text{molar mass}}$$

Since the molar mass (M) of the compound is not specified, we consider it as $M \text{ g/mol}$.

$$n = \frac{0.350 \text{ g}}{M \text{ g/mol}}$$

Step 2: Calculate Molar Enthalpy of Combustion

The molar enthalpy of combustion (ΔH_{comb}) is expressed as:

$$\Delta H_{\text{comb}} = \frac{\text{Heat released}}{\text{moles of sample}} = \frac{Q}{n}$$

Substituting the values:

$$\Delta H_{\text{comb}} = \frac{14.0 \text{ kJ}}{0.350 / M} = \frac{14.0 \text{ kJ}}{0.350} \times M$$

$$\Delta H_{\text{comb}} = \left(\frac{14.0}{0.350} \right) \times M = 40.0 \times M \text{ kJ/mol}$$

This expression shows that the molar enthalpy depends on the molar mass of the compound.

Note:

- To obtain a numerical value, the molar mass must be known.
- For typical hydrocarbons, molar masses range from about 16 g/mol (methane) to over 200 g/mol (large hydrocarbons).

Interpreting the Results and Significance

Understanding the Energy Content

The heat released per mole (ΔH_{comb}) provides a measure of the compound's energy density. Higher values indicate a more energy-rich compound, which is advantageous in fuel applications.

Applications of Combustion Data

- Fuel Industry: Optimizing fuel formulations for maximum energy output.
- Environmental Impact: Estimating CO₂ emissions based on combustion energy.
- Material Science: Developing new compounds with desired combustion properties.

Limitations and Considerations

- The experiment assumes complete combustion; incomplete combustion would yield less heat.
- Heat losses, impurities, and experimental errors can affect measurements.
- The calorimeter's calibration is crucial for accurate results.

Practical Example: Estimating the Molar Enthalpy of Combustion of a Hydrocarbon

Suppose the compound is methane (CH₄), with a molar mass of 16.04 g/mol.

1. Calculate moles of methane:

$$n = \frac{0.350 \text{ g}}{16.04 \text{ g/mol}} \approx 0.0218 \text{ mol}$$

2. Calculate molar enthalpy:

$$\Delta H_{\text{comb}} = \frac{14.0 \text{ kJ}}{0.0218} \approx 642.2 \text{ kJ/mol}$$

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This aligns with known values for methane's combustion enthalpy, demonstrating the validity of the experimental approach.

Conclusion

The complete combustion of a 0.350 g sample of a compound in a bomb calorimeter releasing 14.0 kJ of heat provides essential data for understanding the compound's energy content. By applying thermodynamic principles and precise calculations, scientists can determine the molar enthalpy of combustion, which is pivotal in various fields including energy production, environmental science, and material development. Accurate experimentation and analysis enable the optimization of fuels and materials, contributing to more sustainable and efficient technologies. Understanding the thermochemistry behind combustion reactions not only deepens scientific knowledge but also supports practical advancements in energy management and environmental stewardship.

Frequently Asked Questions

What is the significance of complete combustion in calorimetry experiments?

Complete combustion ensures that all the fuel reacts fully with oxygen, releasing the maximum amount of heat, which is essential for accurate calorimetric measurements.

How can the heat released from the combustion of 0.350 g of a compound be used to determine its calorific value?

The calorific value can be calculated by dividing the heat released (14.0 kJ) by the mass of the compound (0.350 g), giving energy per gram of the substance.

What does the heat release of 14.0 kJ indicate about the energy content of the compound?

It indicates that burning 0.350 g of the compound releases 14.0 kJ of energy, reflecting its combustion energy or calorific value.

Why is a bomb calorimeter suitable for measuring the heat of combustion?

A bomb calorimeter provides a sealed, oxygen-rich environment that ensures complete combustion and accurate measurement of heat released during the reaction.

How would you calculate the molar enthalpy of combustion for the compound from the given data?

First, determine the molar mass of the compound, then calculate the number of moles combusted, and finally divide the total heat released (14.0 kJ) by the number of moles to find the molar enthalpy of combustion.

What factors could affect the accuracy of the heat measurement in this calorimetry experiment?

Factors include heat losses to the surroundings, incomplete combustion, calibration errors of the calorimeter, and impurities in the sample.

Additional Resources

Complete Combustion of a 0.350 g Sample of a Compound in a Bomb Calorimeter Releases 14.0 kJ of Heat

Introduction

Complete combustion is a fundamental chemical process where a substance reacts fully with oxygen, producing carbon dioxide and water as the primary products. The energy released during this process is central to understanding the thermodynamic properties of compounds and their potential uses as fuels. In this context, analyzing the combustion of a small sample – specifically, 0.350 grams – of an unknown compound within a bomb calorimeter provides valuable insights into the compound's energy content, bonding characteristics, and overall thermodynamic profile.

This article explores the detailed methodology, calculations, and implications of the complete combustion of the given sample, which releases 14.0 kilojoules (kJ) of heat energy. It aims to elucidate the scientific principles, interpret the data, and discuss the broader significance of such calorimetric measurements.

Understanding the Experimental Setup

The Role of a Bomb Calorimeter

A bomb calorimeter is an insulated device designed to measure the heat released or absorbed during chemical reactions, especially combustion. Its key features include:

- Sealed chamber ("bomb"): Contains the reactants and oxygen, ensuring complete combustion under constant volume.
- Calorimetric vessel: Encloses the bomb, surrounded by a water bath to absorb heat.
- Thermometer and Stirring Mechanism: Ensures uniform temperature distribution.
- Electrical Ignition System: Initiates the combustion reaction.

The primary advantage of a bomb calorimeter is its ability to accurately measure the heat of combustion at constant volume, which is vital for calculating the energy content of fuels or organic compounds.

Procedure Summary

In a typical experiment:

1. A sample of known mass (here, 0.350 g) is placed inside the bomb.
2. The bomb is filled with pure oxygen at a specified pressure.
3. The system is sealed and submerged in water.
4. The sample is ignited electrically, and the combustion occurs.
5. The temperature change (ΔT) of the water bath is recorded.
6. The heat released (Q) is determined based on the temperature change and calibration constants.

Data Interpretation and Thermodynamic Principles

The Heat of Combustion

The key data point from the experiment is that 14.0 kJ of heat is released for the combustion of 0.350 g of the compound. This value, often called the enthalpy of combustion ($\Delta H_{\text{combustion}}$), provides information about the energy stored within the chemical bonds of the compound.

Significance of Complete Combustion

Complete combustion ensures all carbon in the sample is converted to CO_2 , and all hydrogen is converted to H_2O . This maximizes energy release and simplifies calculations since the products are well-defined, and the energy change can be directly related to the chemical structure.

Calculations and Analytical Insights

1. Molar Energy Content

To understand the energy density of the compound, it is essential to determine the molar heat of combustion, which requires calculating the number of moles of the compound combusted.

- Moles of compound (n):

$$n = \frac{\text{mass of sample}}{\text{molar mass (M)}} = \frac{0.350 \text{ g}}{M}$$

- Energy released per mole (ΔH_m):

$$\Delta H_m = \frac{\text{Heat released}}{\text{moles}} = \frac{14.0 \text{ kJ}}{n} = \frac{14.0 \text{ kJ}}{\frac{0.350 \text{ g}}{M}} = \frac{14.0 \text{ kJ}}{0.350 \text{ g}} \times M$$

Thus,

$$\boxed{\Delta H_m = 40.0 \text{ kJ/mol} \times M}$$

where M is the molar mass of the compound in g/mol.

2. Estimating the Molar Mass

Without the exact chemical formula, we can attempt to estimate the molar mass based on typical combustion energy values. For example:

- Hydrocarbons often release approximately 30–50 kJ/g.
- The measured energy release (14.0 kJ for 0.350 g) translates to:

$$\text{Energy per gram} = \frac{14.0 \text{ kJ}}{0.350 \text{ g}} \approx 40 \text{ kJ/g}$$

This value suggests the compound could be similar in energy content to certain hydrocarbons or organic molecules.

3. Determining the Molar Mass Range

Assuming the molar combustion energy (ΔH_m) is typical for hydrocarbons (~800–1100 kJ/mol), then:

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$M \approx \frac{\Delta H_m}{40} \approx \frac{800}{40} = 20$,
g/mol
or
 $M \approx \frac{1100}{40} = 27.5$,
g/mol

This suggests the compound could be a light hydrocarbon or similar organic molecule with a molar mass in the 20–28 g/mol range.

Broader Implications and Applications

Energy Content and Fuel Potential

The energy released per unit mass (about 40 kJ/g) indicates the compound's potential as a fuel source. For context:

- Gasoline has an energy density of approximately 44–46 kJ/g.
- Methane (CH₄) releases about 55.5 kJ/g.

Therefore, the compound in question exhibits a significant energy density, making it a candidate for fuel applications, pending further structural analysis.

Thermodynamic Insights

- Enthalpy of combustion reflects the strength of chemical bonds within the compound.
- The magnitude of the heat released correlates with the degree of oxidation and the stability of the bonds broken during combustion.
- In practical applications, understanding these thermodynamic parameters aids in designing efficient energy sources and optimizing combustion processes.

Environmental and Safety Considerations

- Complete combustion minimizes the production of harmful pollutants like carbon monoxide (CO) and unburned hydrocarbons.
- The calorimetric data also inform safety protocols in handling and storage, especially if the compound is volatile or reactive.

Limitations and Further Investigations

While the calorimetric measurement provides valuable data, several factors warrant further exploration:

- Chemical Identity: Without the exact chemical structure, assumptions about molar mass and energy content remain estimative.
- Reaction Completeness: Ensuring full combustion is crucial; incomplete combustion would underestimate the energy content.
- Heat Losses: Minor heat losses to surroundings can affect measurements; calibration and correction are necessary.
- State of the Compound: The physical form, purity, and initial conditions influence the results.

Further experiments, such as spectroscopic analysis or molecular weight determination, would complement calorimetric data and facilitate precise characterization.

Conclusion

The complete combustion of a 0.350 g sample of an unknown compound in a bomb calorimeter releasing 14.0 kJ of heat offers a window into the compound's thermodynamic and energetic properties. Through careful calculations, estimates of molar mass and energy density suggest that the compound is a light hydrocarbon or a similarly energetic organic molecule with potential application as a fuel source.

This experiment underscores the importance of calorimetry in chemical analysis, bridging fundamental thermodynamics with practical energy considerations. As energy demands grow and environmental concerns intensify, such insights contribute to the development of cleaner, more efficient fuels and deepen our understanding of chemical energy storage in organic compounds.

References

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Note: The precise chemical identity of the compound would enable more accurate calculations and interpretations. Nonetheless, the principles and methodology outlined here serve as a foundation for analyzing calorimetric data in chemical thermodynamics.

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