

Chapter 17 Thermochemistry Study Guide

17.1 – 17.2 Thermochemical Equations

1. Make the following conversions:
 - a. 444 cal to joules = $1.86 \times 10^3 \text{ J}$
 - b. 1.8 kJ to joules = $1.8 \times 10^3 \text{ J}$
 - c. 0.45 kJ to calories = $1.1 \times 10^2 \text{ cal}$
2. Classify each of these processes as endothermic or exothermic:
 - a. condensing steam – **exo**
 - b. burning alcohol – **exo**
 - c. evaporating alcohol – **endo**
 - d. baking a potato – **endo**
3. The specific heat capacity for silver is $0.24 \text{ J/g}^\circ\text{C}$. Calculate the energy required to raise the temperature of 150.0 g Ag from 273 K to 298 K. Calculate the molar heat capacity of silver.

$$\text{Energy: } q = (150.0)(.24)(298-273) = 9.0 \times 10^2 \text{ J}$$

$$\text{Molar heat capacity: } (0.24 \text{ J/g}^\circ\text{C})(107.87 \text{ g/mol Ag}) = 26 \text{ J/mol}$$

4. It takes 585 J of energy to raise the temperature of 125.6 g Hg from 20.0°C to 53.5°C . Calculate the specific heat capacity and the molar heat capacity of Hg.

$$\text{Specific heat capacity: } C = q/m\Delta t = 585/(125.6)(53.5-20) = 0.139 \text{ J/g}^\circ\text{C}$$

$$\text{Molar heat capacity: } (0.139 \text{ J/g}^\circ\text{C})(200.59 \text{ g/mol Hg}) = 27.9 \text{ J/mol}$$

5. A 46.2-g sample of copper is heated to 95.4°C and then placed in a calorimeter containing 75.0 g water at 19.6°C . The equilibrium temperature in the calorimeter is 21.8°C . Calculate the specific heat capacity of copper; assuming that all the heat lost by the copper is gained by the water.

$$C = 0.203 \text{ J/g}^\circ\text{C}$$

6. A 15.0-g sample of nickel metal is heated to 100.0°C and dropped into 55.0 g of water, initially at 23.0°C . Assuming that all the heat lost by the nickel is absorbed by the water; calculate the final temperature of the nickel and the water. The specific heat of nickel is $0.444 \text{ J/g}^\circ\text{C}$.

$$\begin{array}{rcl} 15.0(.444)(100.0 - T_f) & = & 55.0(4.184)(T_f - 23.0) \\ 6.66 - 6.66T_f & = & 230.12T_f - 5292.76 \\ 5958.76 & = & 236.78 T_f \\ & \mathbf{25.2^\circ\text{C}} & = T_f \end{array}$$

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Gary S. Thorpe



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Table of Contents Thermochemistry Test Study Guide Answers

1. Understanding the eBook Thermochemistry Test Study Guide Answers
 - The Rise of Digital Reading Thermochemistry Test Study Guide Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Thermochemistry Test Study Guide Answers
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Thermochemistry Test Study Guide Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Thermochemistry Test Study Guide Answers
 - Personalized Recommendations
 - Thermochemistry Test Study Guide Answers User Reviews and Ratings
 - Thermochemistry Test Study Guide Answers and Bestseller Lists
5. Accessing Thermochemistry Test Study Guide Answers Free and Paid eBooks
 - Thermochemistry Test Study Guide Answers Public Domain eBooks
 - Thermochemistry Test Study Guide Answers eBook Subscription Services
 - Thermochemistry Test Study Guide Answers Budget-Friendly Options
6. Navigating Thermochemistry Test Study Guide Answers eBook Formats
 - ePub, PDF, MOBI, and More
 - Thermochemistry Test Study Guide Answers Compatibility with Devices
 - Thermochemistry Test Study Guide Answers Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Thermochemistry Test Study Guide Answers
 - Highlighting and Note-Taking Thermochemistry Test Study Guide Answers
 - Interactive Elements Thermochemistry Test Study Guide Answers
8. Staying Engaged with Thermochemistry Test Study Guide Answers

- Joining Online Reading Communities
- Participating in Virtual Book Clubs
- Following Authors and Publishers Thermochemistry Test Study Guide Answers
- 9. Balancing eBooks and Physical Books Thermochemistry Test Study Guide Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Thermochemistry Test Study Guide Answers
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Thermochemistry Test Study Guide Answers
 - Setting Reading Goals Thermochemistry Test Study Guide Answers
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Thermochemistry Test Study Guide Answers
 - Fact-Checking eBook Content of Thermochemistry Test Study Guide Answers
 - Distinguishing Credible Sources
- 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
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