Thermochemistry Review - Answers

- Using the data below, calculate the amount of energy required to melt a 16.2 g aluminum pop can. Assume that the aluminum is at an initial temperature of 20°C.
 - Molar heat of fusion of Al(s) = 10.7 kJ/mol

 - Melting point of Al(s) = 660°C
 Heat capacity of Al(s) = 0.900 J/g°C

mol = 16.2 q - 26.98154 q/mol = 0.600 mol ΔH = (20°C → 660°C) + (melt) $\Delta H = (mc\Delta T) + (n\Delta H fus)$ $\Delta H = (16.2 \text{ g})(0.900 \text{ J/g}^{\circ}\text{C})(640^{\circ}\text{C}) + (0.600 \text{ mol})(10.7 \text{ kJ/mol})$ $\Delta H = 9331.2 \text{ J} \cdot 6.42 \text{ kJ}$ $\Delta H = 9.33 \text{ kJ} + 6.42 \text{ kJ}$

Given the thermochemical equations:

AH = 15.8 kJ

 $C(s) \cdot O_t(q) \rightarrow CO_t(q)$ $\Delta H = -394 \text{ kJ}$ $\Delta H = -286 \text{ kJ}$ $H_r(a) \cdot \{O_r(a) \rightarrow H_rO(l)$ C(s) • 2 H₂(q) • ↑ O₂(q) → CH₂OH(I) ΔH = -239 kJ

Calculate the ΔH for the reaction: $CH_2OH(1) - {}^3/{}_2O_2(q) \rightarrow CO_2(q) + 2H_2O(1)$

 $CH_2OH(1) \rightarrow C(s) + 2 H_2(q) + \frac{1}{2} O_2(q)$ $\Delta H = (-1)(-239 kJ)$ $C(s) \cdot O_2(q) \rightarrow CO_2(q)$ $\Delta H = (1)(-394 \text{ kJ})$ 2 H₂(a) + O₂(a) → 2 H₂O(l) $\Delta H = (2)(-286 \text{ kJ})$

 $CH_3OH(1) + {}^3/_2 O_2(q) \rightarrow CO_2(q) + 2 H_2O(1)$ $\Delta H = -727 \text{ kJ}$

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- 1. Understanding the eBook Energy And Thermochemistry Review Answers
 - The Rise of Digital Reading Energy And Thermochemistry Review Answers
 - Advantages of eBooks Over Traditional Books
- 2. Identifying Energy And Thermochemistry Review 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 Energy And Thermochemistry Review Answers
 - User-Friendly Interface
- 4. Exploring eBook Recommendations from Energy And Thermochemistry Review Answers
 - Personalized Recommendations
 - Energy And Thermochemistry Review Answers User Reviews and Ratings
 - Energy And Thermochemistry Review Answers and Bestseller Lists
- 5. Accessing Energy And Thermochemistry Review Answers Free and Paid eBooks
 - Energy And Thermochemistry Review Answers Public Domain eBooks
 - Energy And Thermochemistry Review Answers eBook Subscription Services
 - Energy And Thermochemistry Review Answers Budget-Friendly Options
- 6. Navigating Energy And Thermochemistry Review Answers eBook Formats

- o ePub, PDF, MOBI, and More
- Energy And Thermochemistry Review Answers Compatibility with Devices
- Energy And Thermochemistry Review Answers Enhanced eBook Features
- 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Energy And Thermochemistry Review Answers
 - Highlighting and Note-Taking Energy And Thermochemistry Review Answers
 - Interactive Elements Energy And Thermochemistry Review Answers
- 8. Staying Engaged with Energy And Thermochemistry Review Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Energy And Thermochemistry Review Answers
- 9. Balancing eBooks and Physical Books Energy And Thermochemistry Review Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Energy And Thermochemistry Review Answers
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Energy And Thermochemistry Review Answers
 - Setting Reading Goals Energy And Thermochemistry Review Answers
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Energy And Thermochemistry Review Answers
 - Fact-Checking eBook Content of Energy And Thermochemistry Review Answers
 - Distinguishing Credible Sources
- 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
- 14. Embracing eBook Trends
 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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