

## Thermochemical Equations and Calorimetry ANSWERS

A **Thermochemical equation** is simply a *balanced chemical equation* that *also* includes the change in energy that occurs in the equation. For example, the thermochemical equation for the synthesis of ammonia is:



The coefficients in the equation represent moles, so what this means is that if *1* mole of  $\text{N}_2$  reacts with *3* moles of  $\text{H}_2$  and *2* moles of  $\text{NH}_3$  are produced, then 46.2 kJ will be given off.

But how can you tell that heat was released?

The  $\Delta\text{H}$  of a reaction refers to the energy change that occurred in the reaction. The “ $\Delta$ ” means “change in” and the H refers to the “heat content” (also called “enthalpy”) of the chemicals.

$\Delta\text{H}$  is always the H of the products – H of the reactants.

If a reaction is **endothermic**, this means that the chemicals in the reaction absorbed heat from their surroundings, so that the products have more energy than the reactants. Since  $\Delta\text{H} = \text{H}_{\text{products}} - \text{H}_{\text{reactants}}$ , and in an endothermic reaction the products have more stored energy than the reactants (after all, they absorbed some heat), the  $\Delta\text{H}$  will always be **positive** for an endothermic reaction. This makes sense because heat was added, so you would expect the change in heat content to be positive.

In the reaction at the top of the page, the  $\Delta\text{H}$  is negative. This means that the reaction is **exothermic** and that as the reaction occurred heat was released to the surroundings. The reactants gave off heat energy, so the products must have less energy than the reactants. And since  $\Delta\text{H} = \text{H}_{\text{products}} - \text{H}_{\text{reactants}}$  and since the products have less heat content than the reactants, the  $\Delta\text{H}$  must always be **negative** for an exothermic reaction. This makes sense because the chemicals in the reaction lost heat.

Again referring to the thermochemical equation above, it tells us that when 2 moles of  $\text{NH}_3$  are synthesized, 46.2 kJ are given off to the surroundings. What would be the  $\Delta\text{H}$  if 2 moles of  $\text{NH}_3$  were *decomposed*? Well, since 46.2 kJ are released when 2 moles of  $\text{NH}_3$  are *formed*, then 46.2 kJ of heat would have to be *absorbed* in order to do the opposite. Thus, the  $\Delta\text{H}$  for the reverse reaction would be +46.2 kJ.

How thermochemical equations are used:

### Determining the amount of heat absorbed or released in a reaction.



Example: Given the thermochemical equation above, determine how much heat would be released if only *1* mole of  $\text{NH}_3$  was synthesized.

Well, according to the thermochemical equation 46.2 kJ are released when *2 moles* are synthesized, then half this—or 23.2 kJ—would be released when *1 mole* of  $\text{NH}_3$  is synthesized.

$$1 \text{ mole } \text{NH}_3 \times \frac{46.2 \text{ kJ}}{2 \text{ moles } \text{NH}_3} = 23.1 \text{ kJ}$$

You'll notice that the negative sign is not in front of the 46.2 kJ. *The sign + or – only needs to appear in the thermochemical equation, nowhere else.*

# Thermochemistry Packet Answers

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### **Table of Contents Thermochemistry Packet Answers**

1. Understanding the eBook Thermochemistry Packet Answers
  - The Rise of Digital Reading Thermochemistry Packet Answers
  - Advantages of eBooks Over Traditional Books
2. Identifying Thermochemistry Packet 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 Packet Answers
  - User-Friendly Interface
4. Exploring eBook Recommendations from Thermochemistry Packet Answers
  - Personalized Recommendations

- Thermochemistry Packet Answers User Reviews and Ratings
- Thermochemistry Packet Answers and Bestseller Lists
- 5. Accessing Thermochemistry Packet Answers Free and Paid eBooks
  - Thermochemistry Packet Answers Public Domain eBooks
  - Thermochemistry Packet Answers eBook Subscription Services
  - Thermochemistry Packet Answers Budget-Friendly Options
- 6. Navigating Thermochemistry Packet Answers eBook Formats
  - ePub, PDF, MOBI, and More
  - Thermochemistry Packet Answers Compatibility with Devices
  - Thermochemistry Packet Answers Enhanced eBook Features
- 7. Enhancing Your Reading Experience
  - Adjustable Fonts and Text Sizes of Thermochemistry Packet Answers
  - Highlighting and Note-Taking Thermochemistry Packet Answers
  - Interactive Elements Thermochemistry Packet Answers
- 8. Staying Engaged with Thermochemistry Packet Answers
  - Joining Online Reading Communities
  - Participating in Virtual Book Clubs
  - Following Authors and Publishers Thermochemistry Packet Answers
- 9. Balancing eBooks and Physical Books Thermochemistry Packet Answers
  - Benefits of a Digital Library
  - Creating a Diverse Reading Collection Thermochemistry Packet Answers
- 10. Overcoming Reading Challenges
  - Dealing with Digital Eye Strain
  - Minimizing Distractions
  - Managing Screen Time
- 11. Cultivating a Reading Routine Thermochemistry Packet Answers
  - Setting Reading Goals Thermochemistry Packet Answers
  - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Thermochemistry Packet Answers
  - Fact-Checking eBook Content of Thermochemistry Packet 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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