$^{\circ}C = (^{\circ}F - 32) *5/9$ 

 $^{\circ}F = ^{\circ}C * 9/5 + 32$ 

 $K = ^{\circ}C + 273$ 

Calculate the Gibbs free energy change (ΔG) for the following chemical reaction:

ATP ADP + Pi

The reaction occurs at 68 °F, the change in heat ( $\Delta H$ ) = 19,070 cal, and the change in entropy ( $\Delta S$ ) = 90 cal/K.

68 °F = 293 K, therefore the equation is set up in the following way:

 $\Delta G = \Delta H - T\Delta S$ 

 $\Delta G = 19,070 \text{ cal} - (293 \text{ K}) (90 \text{ cal/K})$ 

 $\Delta G = -7300 \text{ cal} = -7.3 \text{ kcal}$ 

Calculate the Gibbs free energy change (ΔG) for the following chemical reaction:

glutamate + NH<sub>3</sub> glutamine + H<sub>2</sub>O

The reaction occurs at 68 °F, the change in heat  $(\Delta H)$  = 4103 cal, and the change in entropy  $(\Delta S)$  = 2.4 cal/K.

66 °F = 293 K, therefore the equation is set up in the following way:  $\Delta G = \Delta H - T\Delta S$   $\Delta G = 4103$  cal -293 K (2.4 cal/ K)  $\Delta G = 3399.8$  cal =3.4 kcal

3. Would either of the reactions above occur spontaneously? If so, which one(s) and why?

The reaction with ATP in #1. The change in Gibbs free energy ( $\Delta G$ ) is negative for exothermic reactions and can only be negative for spontaneous reactions. However, the activation energy required would still cause this reaction to occur slowly.

4. Are either of the above reactions endergonic? If so, which one(s) and why?

The reaction with glutamate and NH<sub>2</sub> in #2. The change in  $\Delta G$  (free energy) is positive, meaning free energy is absorbed, making this reaction endergonic.

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#### **Table of Contents Energy Practice Problems And Answers**

- 1. Understanding the eBook Energy Practice Problems And Answers
  - The Rise of Digital Reading Energy Practice Problems And Answers
  - Advantages of eBooks Over Traditional Books
- 2. Identifying Energy Practice Problems And Answers
  - Exploring Different Genres
  - o Considering Fiction vs. Non-Fiction
  - $\circ \ \ Determining \ Your \ Reading \ Goals$
- 3. Choosing the Right eBook Platform
  - $\circ \ \ Popular \ eBook \ Platforms$
  - Features to Look for in an Energy Practice Problems And Answers
  - User-Friendly Interface
- 4. Exploring eBook Recommendations from Energy Practice Problems And Answers
  - Personalized Recommendations

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