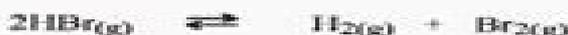


7. At a high temperature, 0.50 mol of HBr was placed in a 1.0 L container and allowed to decompose according to the reaction:



At equilibrium the $[\text{Br}_2]$ was measured to be 0.13 M. What is K_{eq} for this reaction at this temperature?

$$2\text{HBr} \rightleftharpoons \text{H}_2 + \text{Br}_2$$

[I]	0.50	0	0
[C]	-0.26	+0.13	+0.13
[E]	0.24	0.13	0.13

$$K_{eq} = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} = \frac{(0.13)^2}{(0.24)^2} = 0.29$$

Answer $K_{eq} = 0.29$

8. When 1.0 mol of $\text{NH}_3(g)$ and 0.40 mol of $\text{N}_2(g)$ are placed in a 5.0 L vessel and allowed to reach equilibrium at a certain temperature, it is found that 0.78 mol of NH_3 is present. The reaction is:



initial $[\text{NH}_3] = \frac{1.0 \text{ mol}}{5.0 \text{ L}} = 0.20 \text{ M}$ initial $[\text{N}_2] = \frac{0.40 \text{ mol}}{5.0 \text{ L}} = 0.080 \text{ M}$

Equil^m $[\text{NH}_3] = \frac{0.78 \text{ mol}}{5.0 \text{ L}} = 0.156 \text{ M}$

$$2\text{NH}_3 \rightleftharpoons 3\text{H}_2 + \text{N}_2$$

[I]	0.20	0	0.080
[C]	-0.044	+0.066	+0.022
[E]	0.156	0.066	0.102

- a) Calculate the equilibrium concentrations of all three species.

$[\text{NH}_3] = 0.16 \text{ M}$ $[\text{H}_2] = 0.066 \text{ M}$ $[\text{N}_2] = 0.10 \text{ M}$

- b) Calculate the value of the equilibrium constant at this temperature.

(use unrounded concs, then round to 2 sig)

$$K_{eq} = \frac{[\text{H}_2]^3 [\text{N}_2]}{[\text{NH}_3]^2} = \frac{(0.066)^3 (0.102)}{(0.156)^2} = 0.0012$$

Answer 1.2×10^{-3}

- c) How many moles of H_2 are present at equilibrium?

$$0.066 \text{ M} \times 5.0 \text{ L} = 0.33 \text{ mol}$$

Answer 0.33 mol

- d) How many moles of N_2 are present at equilibrium?

$$0.102 \text{ M} \times 5.0 \text{ L} = 0.51 \text{ mol}$$

Answer 0.51 mol

Modeling Chemistry Unit 3 1 Answers

ML Morrison



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Table of Contents Modeling Chemistry Unit 3 1 Answers

1. Understanding the eBook Modeling Chemistry Unit 3 1 Answers
 - The Rise of Digital Reading Modeling Chemistry Unit 3 1 Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Modeling Chemistry Unit 3 1 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 Modeling Chemistry Unit 3 1 Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Modeling Chemistry Unit 3 1 Answers
 - Personalized Recommendations
 - Modeling Chemistry Unit 3 1 Answers User Reviews and Ratings
 - Modeling Chemistry Unit 3 1 Answers and Bestseller Lists

5. Accessing Modeling Chemistry Unit 3 1 Answers Free and Paid eBooks
 - Modeling Chemistry Unit 3 1 Answers Public Domain eBooks
 - Modeling Chemistry Unit 3 1 Answers eBook Subscription Services
 - Modeling Chemistry Unit 3 1 Answers Budget-Friendly Options
6. Navigating Modeling Chemistry Unit 3 1 Answers eBook Formats
 - ePub, PDF, MOBI, and More
 - Modeling Chemistry Unit 3 1 Answers Compatibility with Devices
 - Modeling Chemistry Unit 3 1 Answers Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Modeling Chemistry Unit 3 1 Answers
 - Highlighting and Note-Taking Modeling Chemistry Unit 3 1 Answers
 - Interactive Elements Modeling Chemistry Unit 3 1 Answers
8. Staying Engaged with Modeling Chemistry Unit 3 1 Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Modeling Chemistry Unit 3 1 Answers
9. Balancing eBooks and Physical Books Modeling Chemistry Unit 3 1 Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Modeling Chemistry Unit 3 1 Answers
10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Modeling Chemistry Unit 3 1 Answers
 - Setting Reading Goals Modeling Chemistry Unit 3 1 Answers
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Modeling Chemistry Unit 3 1 Answers
 - Fact-Checking eBook Content of Modeling Chemistry Unit 3 1 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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