

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  $K_{\text{eq}}$  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

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

Answer  $K_{\text{eq}} = 0.29$

8. When 1.0 mol of  $\text{NH}_3\text{(g)}$  and 0.40 mol of  $\text{N}_2\text{(g)}$  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  $\text{NH}_3$  is present. The reaction is:



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

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

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

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

[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.

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

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

(use unrounded concs, then round to 2 sigs)

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

Answer  $1.2 \times 10^{-3}$

- c) How many moles of  $\text{H}_2$  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  $\text{N}_2$  are present at equilibrium?

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

Answer 0.51 mol

# Modeling Chemistry U4 V2 0 Answer Sheet

**M Mosston**



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## **Table of Contents Modeling Chemistry U4 V2 0 Answer Sheet**

1. Understanding the eBook Modeling Chemistry U4 V2 0 Answer Sheet
  - The Rise of Digital Reading Modeling Chemistry U4 V2 0 Answer Sheet
  - Advantages of eBooks Over Traditional Books
2. Identifying Modeling Chemistry U4 V2 0 Answer Sheet
  - 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 U4 V2 0 Answer Sheet
  - User-Friendly Interface
4. Exploring eBook Recommendations from Modeling Chemistry U4 V2 0 Answer Sheet
  - Personalized Recommendations
  - Modeling Chemistry U4 V2 0 Answer Sheet User Reviews and Ratings
  - Modeling Chemistry U4 V2 0 Answer Sheet and Bestseller Lists
5. Accessing Modeling Chemistry U4 V2 0 Answer Sheet Free and Paid eBooks
  - Modeling Chemistry U4 V2 0 Answer Sheet Public Domain eBooks
  - Modeling Chemistry U4 V2 0 Answer Sheet eBook Subscription Services
  - Modeling Chemistry U4 V2 0 Answer Sheet Budget-Friendly Options
6. Navigating Modeling Chemistry U4 V2 0 Answer Sheet eBook Formats

- ePub, PDF, MOBI, and More
  - Modeling Chemistry U4 V2 0 Answer Sheet Compatibility with Devices
  - Modeling Chemistry U4 V2 0 Answer Sheet Enhanced eBook Features
7. Enhancing Your Reading Experience
- Adjustable Fonts and Text Sizes of Modeling Chemistry U4 V2 0 Answer Sheet
  - Highlighting and Note-Taking Modeling Chemistry U4 V2 0 Answer Sheet
  - Interactive Elements Modeling Chemistry U4 V2 0 Answer Sheet
8. Staying Engaged with Modeling Chemistry U4 V2 0 Answer Sheet
- Joining Online Reading Communities
  - Participating in Virtual Book Clubs
  - Following Authors and Publishers Modeling Chemistry U4 V2 0 Answer Sheet
9. Balancing eBooks and Physical Books Modeling Chemistry U4 V2 0 Answer Sheet
- Benefits of a Digital Library
  - Creating a Diverse Reading Collection Modeling Chemistry U4 V2 0 Answer Sheet
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
  - Minimizing Distractions
  - Managing Screen Time
11. Cultivating a Reading Routine Modeling Chemistry U4 V2 0 Answer Sheet
- Setting Reading Goals Modeling Chemistry U4 V2 0 Answer Sheet
  - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Modeling Chemistry U4 V2 0 Answer Sheet
- Fact-Checking eBook Content of Modeling Chemistry U4 V2 0 Answer Sheet
  - 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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