

Chapter 6

Equilibrium

Question 6.61:

The ionization constant of nitrous acid is 4.5×10^{-4} . Calculate the pH of 0.04 M sodiumnitrite solution and also its degree of hydrolysis.

Answer 6.61:

NaNO_2 is the salt of a strong base (NaOH) and a weak acid (HNO_2).



$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} \Rightarrow \frac{K_b}{K_a} = \frac{10^{-14}}{4.5 \times 10^{-4}} = 0.22 \times 10^{-10}$$

Now, If x moles of the salt undergo hydrolysis, then the concentration of various species present in the solution will be:

$$[\text{NO}_2^-] = .04 - x$$

$$[\text{HNO}_2] = x$$

$$[\text{OH}^-] = x$$

$$K_b = \frac{x^2}{0.04} = 0.22 \times 10^{-10} \Rightarrow x^2 = .0088 \times 10^{-10} \Rightarrow x = .093 \times 10^{-5}$$

$$\therefore [\text{OH}^-] = 0.093 \times 10^{-5} \text{ M} \Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{.093 \times 10^{-5}} = 10.75 \times 10^{-9} \text{ M}$$

$$\Rightarrow \text{pH} = -\log(10.75 \times 10^{-9}) = 7.96 \Rightarrow \frac{x}{0.04} = \frac{.093 \times 10^{-5}}{.04} = 2.325 \times 10^{-5}$$

Question 6.62:

A 0.02M solution of pyridinium hydrochloride has pH = 3.44. Calculate the ionization constant of pyridine.

Answer 6.62:

$$\text{pH} = 3.44$$

We know that,

$$\text{pH} = -\log [\text{H}^+]$$

$$\therefore [\text{H}^+] = 3.63 \times 10^{-4}$$

$$\text{Then } K_b = \frac{(3.63 \times 10^{-4})^2}{0.02} \quad (\because \text{Concentration} = 0.002 \text{ M})$$

$$\Rightarrow K_b = 6.6 \times 10^{-6}$$

$$\text{Now, } K_b = \frac{K_w}{K_a} \Rightarrow K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{6.6 \times 10^{-6}} = 1.51 \times 10^{-9}$$

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Table of Contents Chemistry Chapter 6 Answers

1. Understanding the eBook Chemistry Chapter 6 Answers
 - The Rise of Digital Reading Chemistry Chapter 6 Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Chemistry Chapter 6 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 Chemistry Chapter 6 Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Chemistry Chapter 6 Answers

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

- Fact-Checking eBook Content of Chemistry Chapter 6 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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