

Equilibrium Constants and Le Châtelier's Principle: CoCl_2

INTRODUCTION

A well-known equilibrium system involves cobalt chloride in water, as the two cobalt-containing species have distinctly different colors (Equation 1). The equilibrium expression for this system is given in Equation 2.



Equation (2)

$$K = \frac{[\text{CoCl}_4^{2-}]}{[\text{Co}(\text{H}_2\text{O})_6^{2+}] \cdot [\text{Cl}^{-}]^4}$$

$$\frac{1}{1 \times 10^{-4}} = \frac{1}{1 \times 10^{-4}}$$

Changing the cobalt concentration (i.e., the ratio of cobalt ions to water) and the chloride concentration shifts the position in accord with Le Châtelier's principle—when disturbed, an equilibrium system shifts in a direction such that the disturbance is minimized.

The conversion of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ to $[\text{CoCl}_4]^{2-}$ involves converting an octahedral complex into a tetrahedral complex (Figure 1).



FIGURE 1 Structure of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{CoCl}_4]^{2-}$.

The bond dissociation energy for a cobalt(II) bond to a neutral oxygen atom is about the same as that for the bond between cobalt(II) and a chloride anion. The conversion from an octahedral complex to a tetrahedral complex involves breaking six bonds and forming four bonds. As breaking bonds is endothermic and forming bonds is exothermic, the net result is that the conversion from the pink octahedral complex to the blue tetrahedral complex is endothermic. Consequently, the equilibrium is temperature-sensitive, as shown in Equation 3. Thus, the application of heat or the removal of heat shifts the position of the equilibrium.



In this experiment, you will measure the room temperature effective equilibrium constant for this system, K of Equation 2. You will also examine the effects of the addition of water, the addition and removal of chloride, and changes in temperature on the position of the equilibrium to determine whether they are consistent with Le Châtelier's principle.

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Table of Contents Pkgacp Cer General Chemistry Lab Ch 122l

1. Understanding the eBook Pkgacp Cer General Chemistry Lab Ch 122l
 - The Rise of Digital Reading Pkgacp Cer General Chemistry Lab Ch 122l
 - Advantages of eBooks Over Traditional Books
2. Identifying Pkgacp Cer General Chemistry Lab Ch 122l
 - 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 Pkgacp Cer General Chemistry Lab Ch 122l
 - User-Friendly Interface
4. Exploring eBook Recommendations from Pkgacp Cer General Chemistry Lab Ch 122l
 - Personalized Recommendations
 - Pkgacp Cer General Chemistry Lab Ch 122l User Reviews and Ratings
 - Pkgacp Cer General Chemistry Lab Ch 122l and Bestseller Lists
5. Accessing Pkgacp Cer General Chemistry Lab Ch 122l Free and Paid eBooks
 - Pkgacp Cer General Chemistry Lab Ch 122l Public Domain eBooks
 - Pkgacp Cer General Chemistry Lab Ch 122l eBook Subscription Services
 - Pkgacp Cer General Chemistry Lab Ch 122l Budget-Friendly Options
6. Navigating Pkgacp Cer General Chemistry Lab Ch 122l eBook Formats

- ePub, PDF, MOBI, and More
 - Pkgacp Cer General Chemistry Lab Ch 122l Compatibility with Devices
 - Pkgacp Cer General Chemistry Lab Ch 122l Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Pkgacp Cer General Chemistry Lab Ch 122l
 - Highlighting and Note-Taking Pkgacp Cer General Chemistry Lab Ch 122l
 - Interactive Elements Pkgacp Cer General Chemistry Lab Ch 122l
 8. Staying Engaged with Pkgacp Cer General Chemistry Lab Ch 122l
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Pkgacp Cer General Chemistry Lab Ch 122l
 9. Balancing eBooks and Physical Books Pkgacp Cer General Chemistry Lab Ch 122l
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Pkgacp Cer General Chemistry Lab Ch 122l
 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
 11. Cultivating a Reading Routine Pkgacp Cer General Chemistry Lab Ch 122l
 - Setting Reading Goals Pkgacp Cer General Chemistry Lab Ch 122l
 - Carving Out Dedicated Reading Time
 12. Sourcing Reliable Information of Pkgacp Cer General Chemistry Lab Ch 122l
 - Fact-Checking eBook Content of Pkgacp Cer General Chemistry Lab Ch 122l
 - Distinguishing Credible Sources
 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
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