Cobalt Complex Ions

LeChâtelier's Principle

Introduction

Demonstrate the power of balance in a reversible chemical reaction by showing how chemical equilibrium represents a true chemical balancing act and thus responds dramatically to anything that might upset that balance. This chemical demonstration allows students to—

- Observe the effect of concentration and temperature on the position of equilibrium for an endothermic chemical reaction.
- Infer how a chemical system responds to changes in order to restore chemical equilibrium, and thus to formulate their own version of LeChâtelier's Principle.
- Visualize the dynamic nature of chemical equilibrium.

Concepts

- Reversible chemical reaction
- LeChâtelier's Principle
- Equilibrium constant
- Reaction quotient

Materials

Cobalt(II) chloride solution, $CoCl_2$, 0.1 *M*, 20 mL

Hydrochloric acid, HCl, concentrated, 12 *M*, 10 mL

Silver nitrate solution, $AgNO_3$, 0.1 *M*, 3 mL

Water, distilled

Beakers, 400-mL, 2

Hot plate

Pipets, Beral-type, 3

Test tube rack

Test tubes, borosilicate glass, medium-size (approximately 19 × 150 mm), 5

Safety Precautions

Concentrated hydrochloric acid is highly toxic by ingestion or inhalation and is severely corrosive to skin and eyes; can cause severe body tissue burns. Avoid contact with skin or clothing. Cobalt(II) chloride solution is moderately toxic by ingestion; body tissue irritant. Silver nitrate solution is corrosive and will stain skin and clothing. Avoid contact with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Practice strict hygiene in the use of the chemicals involved in this demonstration. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Preparation

1. Obtain four medium-sized test tubes and label them: P1, P2, B1, and B2.
2. Measure approximately 5 mL of 0.1 M $CoCl_2$ solution into each of the four test tubes.
3. Slowly and carefully add 5 mL of concentrated HCl to test tubes B1 and B2. *Note:* The solution in test tubes B1 and B2 should turn blue.
4. Set aside test tubes P1 and B1 as control solutions.
5. Obtain a 400-mL beaker and fill it with tap water about half full. Use a hot plate to heat the water to 80–85 °C.
6. Obtain a second 400-mL beaker and prepare an ice water bath filled about half full.

Procedure

1. Add 5 mL of concentrated HCl to test tube P2 in approximately 0.5 mL increments until the solution turns blue in color.

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2. Once the solution has turned blue in test tube P2 add 5 mL of distilled water in approximately 0.5 mL increments until the solution reverts to its original pink color.
3. Add 2 mL of 0.1 M $AgNO_3$ solution to test tube B2. *Note:* A large amount of white precipitate will form while the supernatant will be pink.
4. Obtain test tube P2, which contains roughly 15 mL of pink solution, and place it in the hot water bath. The solution will gradually change in color from pink to lavender blue.
5. Using a test tube clamp, remove test tube P2 from the hot water bath and immediately immerse it in the ice water bath. The solution should revert to pink.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Solutions containing silver nitrate and silver chloride can be disposed of according to Flinn Suggested Disposal Method #11. Solutions containing cobalt(II) chloride can be disposed of according to Flinn Suggested Disposal Method #27f. Alternatively, the solutions can be combined and filtered to remove insoluble silver chloride, which can be dried and packaged for landfill disposal. The combined filtrate can then be neutralized to pH 3–10 according to Flinn Suggested Disposal Method #24b and saved in a disposal container reserved for heavy-metal waste.

Tips

- The demonstration is written on the 5-mL scale in order to avoid having to handle larger amounts of concentrated hydrochloric acid. The recommended scale should permit full student participation in an average-size classroom.
- The demonstration can also be done on a larger scale by adjusting the amounts of reagents proportionately. In this case it is recommended that the teacher "prime" the pink solution beforehand by adding half of the required acid amount during pre-lab preparation. This will reduce the amount of concentrated acid that needs to be handled and transferred during the demonstration itself.
- The solution colors can be distinguished more easily (especially in the lavender/blue transition) if the test tube is held against a white background; use a notecard or a piece of paper. The demonstration can also be projected to a classroom if the reactions are carried out in a small beaker or Petri dish placed on an overhead projector. Adjust the amounts of materials as necessary.
- Point out to students when the solutions pass through intermediate "lavender" stages, corresponding to the presence of roughly equal amounts of the pink and blue species in solution. Ask students to speculate on the composition of the solutions at this intermediate stage.
- The pink and blue equilibrium due to cobalt chloride in aqueous solution is the basis of moisture-sensitive paper such as Hydrion Humidicator Paper, which is used to measure the relative humidity in air. The paper strips are impregnated with deep blue anhydrous $CoCl_2$ powder that changes to bright pink when it is hydrated. The intermediate colors between blue and pink are clear and definite, and a color chart can be used to estimate relative humidity levels between 20 and 80 percent.
- An interesting example of LeChâtelier's Principle is found in the carbonic acid–bicarbonate buffering system in the blood which is able to control blood pH by responding in a vital manner to potentially dangerous changes in blood CO_2 levels. The ideas of respiratory acidosis (excess CO_2 due to restricted breathing) and respiratory alkalosis (reduced CO_2 due to hyperventilation) can be introduced.

Discussion

Chemical equilibrium is a dynamic condition—students often misinterpret it as being static. At equilibrium the concentrations of reactants and products remain unchanged. This standard definition is frequently misunderstood to mean that the concentrations of reactants and products have constant values. It is the ratio of product to reactant concentrations, governed by the stoichiometry of the balanced chemical equation, that is constant. The concentrations of individual reactants and products are affected by changes in the other terms in the equilibrium constant ratio or expression. And the equilibrium "constant" itself is temperature dependent. The effect of concentration, temperature, and pressure changes on the position of chemical equilibrium for a reversible chemical reaction is expressed intuitively in *LeChâtelier's Principle: "If the conditions of a system, initially at equilibrium, are changed, the equilibrium will shift in such a direction as to tend to restore the original conditions."*

The chemical reaction demonstrated herein involves the formation of complex ions between Co^{2+} and water molecules or chloride ions, respectively.

$$\underset{pink}{Co(H_2O)_6^{2+}} + 4Cl^- + heat \leftrightarrows \underset{blue}{CoCl_4^{2-}} + 6H_2O \qquad \textit{Equation 1}$$

A solution of cobalt(II) ion in water is pink, the color of the complex ion formed between Co^{2+} ions and water molecules. When chloride ion in the form of hydrochloric acid is added to the solution, the color changes to blue, corresponding to the formation of a charged coordination complex between Co^{2+} and chloride ions. This reaction is reversible and quickly reaches a position of chemical equilibrium, which is immediately evident by the color of the solution.

In terms of the position of equilibrium for this reaction, addition of Cl^- ion (excess reactant) shifts the equilibrium to the right (toward $CoCl_4^{2-}$ formation) to consume some of the added reactant and thus restore the equilibrium condition. If the blue solution corresponding to $CoCl_4^{2-}$ is diluted by the addition of water (a product of the above reaction), the effect is to shift the equilibrium back to the left, toward $Co(H_2O)_6^{2+}$. This observation requires a slightly different explanation, since technically the concentration of water (solvent) in an aqueous solution is constant. The effect can be explained in terms of the equilibrium constant expression (K_{eq}) for the reaction, which contains one term in the numerator but two terms in the denominator.

$$K_{eq} = \frac{[CoCl_4^{2-}]}{[Co(H_2O)_6^{2+}][Cl^-]^4} \qquad \textit{Equation 2}$$

Reducing each concentration term in the equilibrium constant by a factor of one-third, due to dilution with water in step 6, means that the concentration ratio in Equation 2 becomes greater than K_{eq}, and the reaction shifts back to reactants in order to make the ratio equal to the equilibrium constant.

Addition of $AgNO_3$ to the blue solution of $CoCl_4^{2-}$ results in the formation of a copious white precipitate of AgCl, via the reaction $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$, and a pink solution of $Co(H_2O)_6^{2+}$. Depletion of the chloride ion concentration due to the formation of insoluble AgCl shifts the equilibrium in Equation 1 back to the left, toward reactant formation, in order to offset the effect of this change. The effect of heat is explained by noting that reaction (1) is endothermic, so that heat may be thought of as a reactant in the reaction equation. Addition of excess reactant in the form of heat shifts the equilibrium in the direction in which heat is absorbed in order to "consume" the excess reactant. Adding heat shifts the reaction in Equation (1) to the right (blue), while removing heat shifts it back to the left (pink).

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
- Evidence, models, and explanation
- Evolution and equilibrium

Content Standards: Grades 9–12
- Content Standard A: Science as Inquiry
- Content Standard B: Physical Science, structure and properties of matter, chemical reactions, conservation of energy and increase in disorder, interactions of energy and matter

Answers to Worksheet Questions

1. Describe what you observed at the following stages in this demonstration:
   a. Adding HCl to the originally pink solution (P2)
   The solution turned blue.
   b. Adding distilled water to the originally pink solution (P2)
   The solution, which was now blue, turned back to pink.
   c. Placing the originally pink solution in a hot water bath (P2)
   The solution slowly turned blue.

d. Placing the originally pink solution in a cold water bath (P2)

The now-blue solution quickly turned back to pink.

e. Adding silver nitrate to the originally blue solution (B2)

A large amount of white precipitate formed and the solution turned pink.

2. Write the chemical equation for this reaction, in which complex ions form between CO^{2+} and either water molecules or chloride ions. *Hint:* Heat is a reactant.

$Co(H_2O)_6^{2+} + 4Cl^{-} + heat \rightarrow CoCl_4^{2-} + 6H_2O$

3. Using the equation you just wrote and what you observed during the demonstration, indicate whether the following steps shift the reaction to the left or the right.

a. Adding HCl

The reaction shifts to the right.

b. Adding distilled water

The reaction shifts to the left.

c. Placing the solution in a hot water bath

The reaction shifts to the right.

4. The addition of silver nitrate to the blue solution of $CoCl_4^{2-}$ results in the formation of a white precipitate and a pink solution. Identify both the precipitate and the solution.

The precipitate is silver chloride (AgCl), which is insoluble. The chloride ions are thus depleted, shifting the reaction back to the left and thereby forming the pink solution of $Co(H_2O)_6^{2+}$.

Acknowledgment

Special thanks to Jim and Julie Ealy, The Peddie School, Hightstown, NJ, who provided us with instructions for this activity.

Reference

Shakhashiri, B. Z., *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Vol. 1;* The University of Wisconsin Press: Madison, 1983; pp 280–285.

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Cobalt Complex Ions* activity, presented by Peg Convery, is available in *LeChâtelier's Principle*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Cobalt Complex Ions* are available from Flinn Scientific, Inc.

Materials required to perform this activity are available in the *Pink and Blue: A Colorful Chemical Balancing Act—Chemical Demonstration Kit* available from Flinn Scientific. Materials may also be purchased separately.

Catalog No.	Description
AP8471	Pink and Blue: A Colorful Chemical Balancing Act—Chemical Demonstration Kit
AP4656	Hydrion Humidicator Paper
C0242	Cobalt Chloride Solution, 0.1 M, 500 mL
H0031	Hydrochloric Acid Solution, 12 M, 100 mL
S0305	Silver Nitrate Solution, 0.1 M, 100 mL

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

Cobalt Complex Ions Worksheet

Discussion Questions

1. Describe what you observed at the following stages in this demonstration:

   a. Adding HCl to the originally pink solution (P2)

   b. Adding distilled water to the originally pink solution (P2)

   c. Placing the originally pink solution in a hot water bath (P2)

   d. Placing the originally pink solution in a cold water bath (P2)

   e. Adding silver nitrate to the originally blue solution (B2)

2. Write the chemical equation for this reaction, in which complex ions form between CO^{2+} and either water molecules or chloride ions. *Hint:* Heat is a reactant.

3. Using the equation you just wrote and what you observed during the demonstration, indicate whether the following steps shift the reaction to the left or the right.

   a. Adding HCl

   b. Adding distilled water

   c. Placing the solution in a hot water bath

4. The addition of silver nitrate to the blue solution of $CoCl_4^{2-}$ results in the formation of a white precipitate and a pink solution. Identify both the precipitate and the solution.