Pink and Blue — A Colorful Chemical Balancing Act

A Demonstration of LeChâtelier's Principle

Introduction

Demonstrate the power of balance in a reversible chemical reaction. Students observe the effects of concentration and temperature on the position of equilibrium for an endothermic chemical reaction, visualize how a chemical system responds to changes in order to restore equilibrium, and thus formulate their own version of LeChâtelier's Principle.

Concepts

- Reversible chemical reaction
- Chemical equilibrium
- LeChâtelier's Principle

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

Beral-type pipets, 3

Hot plate

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. 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. 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 as follows, 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.
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.

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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.

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 5–8
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, properties and changes of properties in matter
Content Standards: Grades 9–12
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, structure of atoms, structure and properties of matter, chemical reactions, motions and forces, conservation of energy and increase in disorder, interactions of energy and matter

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.
- Point out to students when the solutions pass through intermediate "lavender" stages, corresponding to the presence of equal amounts of the pink and blue species. Ask students to speculate on the composition of the solutions at this intemediate stage.

Discussion

Chemical equilibrium is a dynamic condition. At equilibrium the concentrations of reactants and products remain unchanged. This definition is often 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 expression. The equilibrium "constant" itself is also temperature dependent. The effect of concentration, temperature, and pressure changes on the position of equilibrium for a reversible 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 classic pink-and-blue reaction in this demonstration involves the formation of complex ions between Co^{2+} and water molecules or chloride ions, respectively.

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

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, due 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.

Addition of Cl^- ion (excess reactant) shifts *Equation 1* to the right (toward $CoCl_4^{2-}$ formation) to consume some of the added reactant and thus restore equilibrium. 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 $Co(H_2O)_6^{2+}$.

Addition of $AgNO_3$ to the blue solution of $CoCl_4^{2-}$ results in the formation of a 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—heat is 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).

Pink and Blue—A Colorful Chemical Balancing Act is available as a Chemical Demonstration Kit from Flinn Scientific, Inc.

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Consult the Flinn Scientific website for current prices.