Cobalt Complex Equilibrium

Temperature and Concentration Effects:

Purpose

To demonstrate the shift in equilibrium caused by a change of temperature and a change in concentration.

Materials

0.4 M $CoCl_2$ (5.2g/100 mL H_2O)	hot plate
0.2 M $CoCl_2$ (2.6g/100 mL H_2O)	ice bath
0.1 M $AgNO_3$ (1.7g/100 mL H_2O)	2 Deep Petri dishes
concentrated. HCl	distilled water
250 mL beaker	dropper (2)
100 mL beakers (4)	

Procedure

Temperature effects

I. Place 100 mL 0.4 M $CoCl_2$ in 250 mL beaker

II. Heat the solution in a flask until it turns blue.

III. Fill a large test tube with the blue solution.

IV. Immerse lower half of tube in an ice-salt bath (-18 C).

V. Note the bottom half of solution turns pink.

Concentration Effects

1. Place 20 mL 0.2 M $CoCl_2$ in 100 mL beaker. Add 40 mL conc HCl. Note blue color.
2. Place 2 Deep Petri dishes on overhead projector. Add $CoCl_2$ solution from 1 to each.
    a. To Petri #1 add distilled water until color changes to clear.
    b. To Petri #2 add $AgNO_3$ drop wise until precipitate forms.
    Note pink color of solution.

Additional Information

1. The equilibrium is

$$\text{heat} + \underset{\text{pink}}{[Co(H_2O)_6]^{2+}}(aq) + 4\ Cl^-(aq) \rightleftharpoons \underset{\text{blue}}{[CoCl_4]^{2-}}(aq) + 6\ H_2O(l)$$

2. Temperature effects

   a. Heat causes shift in equilibrium to products

   b. Cooling shifts equilibrium to reactants

3. Concentration effects

   a. Excess Cl^- causes formation of a more blue complex (tetrachlorocobalt (II))

   b. excess H_2O shifts equilibrium to pink complex (hexaaquocobalt (II))

   c. Silver nitrate removes Cl^- (forming Ag Cl) causing equilibrium to shift toward the pink complex.

Disposal

Solutions should be placed in properly labeled waste container with UI# 96482.

Reference

Summerlin and Ealey, Chemical Demonstrations