The University of Texas at Austin
Department of Chemistry
College of Natural Sciences

Cobalt Complexes

Performer's Version

Safety Hazards

- Personal Protective Equipment:
  - Safety glasses/goggles
  - Nitrile gloves
  - Chemical & flame retardant lab coat
  - Face shield (optional, but advised)
- Physical Hazards
  - Cobalt(II) chloride is very toxic to aquatic life with long lasting effects.
- Chemical Hazards
  - Cobalt(II) chloride is harmful if swallowed, may cause an allergic skin reaction and may cause allergy or asthma symptoms or breathing difficulties if inhaled, and causes serious eye damage. Cobalt(II) chloride is suspected of causing genetic defects and may damage fertility or the unborn child. Cobalt(II) chloride may cause cancer by inhalation.
  - Hydrochloric acid is corrosive, may cause eye and skin damage; irritant, may cause severe respiratory irritation.

Materials

- 200mL 0.1M Cobalt(II) chloride solution
- 200mL Hydrochloric acid, concentrated (12M)
- 500mL deionized water
- Ice
- 600mL beaker
- 2x 400mL beakers
- Insulated bucket
- Glass stir rod
- Stir/Hot plate
- Extension cord
- 3x glass test tubes (25mm dia. x 150mm length) (*Alternate Procedure only)*
- 3x Rubber stoppers (*Alternate Procedure only)*
- Parafilm (*Alternate Procedure only)*

Safety Data Sheet(s)

- Cobalt(II) chloride hexahydrate
- Hydrochloric acid, concentrated (12M)

Procedure

1. Stir 200mL of prepared 0.1M Cobalt(II) chloride solution with either a glass stir rod or on a stir plate, and add 200mL of concentrated hydrochloric acid to the 600mL beaker. This should turn the solution purple.
2. If the solution is too blue, dilute with deionized water until the desired purple color is obtained.
3. Place half of the solution in a 400mL beaker and heat it on a hot plate.
4. Place the remaining solution in the other 400mL beaker and place it in the insulated bucket of ice and allow it to cool.
5. The heated solution should turn blue, and the cooled solution should turn pink.

Alternate Procedure

1. Heat a beaker of deionized water on a hot plate to create a hot water bath. Do not boil the water.
2. Place one test tube in the hot water bath.
3. Place another test tube in the insulated bucket of ice and allow it to cool.
4. The heated solution should turn blue, and the cooled solution should turn pink.

The University of Texas at Austin
Department of Chemistry
College of Natural Sciences

Supplemental Information & Pedagogy

This demonstration showcases chemical equilibrium and thermochromism using cobalt(II) chloride in aqueous and chloride-rich environments. When cobalt(II) chloride ($CoCl_2$) dissolves in water, it forms the pink hexaaquacobalt(II) complex ion:

$$[Co(H_2O)_6]^{2+} + 4Cl^- \rightleftharpoons [CoCl_4]^{2-} + 6H_2O$$

This equilibrium is temperature-dependent and endothermic in the forward direction. The pink $[Co(H_2O)_6]^{2+}$ complex is favored at lower temperatures and higher water concentrations, while the blue $[CoCl_4]^{2-}$ complex is favored at higher temperatures and higher chloride ion concentrations.

In this demonstration, a concentrated hydrochloric acid solution is added to an aqueous $CoCl_2$ solution to supply excess chloride ions, shifting the equilibrium toward formation of the blue tetrachlorocobaltate(II) ion. When this solution is heated in a warm water bath, the increased temperature provides energy that drives the endothermic reaction forward, intensifying the blue color. Conversely, cooling the system (such as in an ice bath) shifts the equilibrium in the exothermic reverse direction, regenerating the pink $[Co(H_2O)_6]^{2+}$ complex. These color changes can be reversed repeatedly, making for a striking and pedagogically rich visual.

The demonstration powerfully illustrates Le Châtelier's Principle: a system at equilibrium will respond to external changes by shifting in the direction that counteracts the disturbance. In this case, heating the solution shifts equilibrium to the right, resulting in a higher concentration of the blue complex; meanwhile, cooling shifts equilibrium to the left, resulting in a higher concentration of the pink complex. Beyond the thermochromism, adding Cl^- via the addition of HCl favors the blue complex, whereas removing Cl^- (e.g., by adding $AgNO_3$, which forms a white AgCl precipitate) would shift the system back to favor the pink complex.

This equilibrium system also underscores the importance of complex ion formation in coordination chemistry. The formation of colored complexes with varying ligands – in this case, water vs. chloride – demonstrates how ligand identity affects both color and stability of metal complexes.

The reversible color change of cobalt chloride has practical use in humidity and temperature indicators. For instance, cobalt chloride is used in silica gel packets where it appears blue when dry and pink when moist, reflecting the equilibrium shift driven by water acting as a ligand. It also appears in thermochromic materials used for temperature-sensitive inks, novelty items, and laboratory indicators. These applications rely directly on the same principles demonstrated here—temperature- and ligand-driven equilibrium shifts resulting in observable color changes.