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Conclusion:

Net-ionic equation for the equilibrium:

(pink complex) $Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) \leftrightarrow CoCl_4^{2-}(aq) + 6H_2)(l)$ (blue complex)

1. The pink cobalt complex was favored when adding water because the increase in water shifted the equilibrium to the left, favoring the reverse reaction.
2. The blue cobalt complex, $CoCl_2$, was favored when the HCl or Calcium chloride was added because it dissociates in the solution, causing an increase in Cl^- ions, which is a common ion in both of the reagent, shifting the equilibrium to the right, favoring the forward.
3. When acetone was added, it effectively removed some of the water from the system, shifting the equilibrium to the right to replace the water, favoring the forward reaction and the blue cobalt complex.
4. When silver nitrate was added, it decreased the Cl^- ions because when the silver nitrate dissociate, it releases Ag^+ ions, which react with the free Cl^- ions forming AgCl. This causes the equilibrium to shift to the left, favoring the reverse reaction and the pink cobalt complex to form.
5.
6. Heating the solution favored the blue cobalt complex and cooling it favored the pink cobalt complex.
   $Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) + heat \leftrightarrow CoCl_4^{2-}(aq) + 6H_2)(l)$ $\Delta H=+50$ KJ/mol
   Since it turned blue when heating and pink when cooling, that means that heat is a reactant and the solution is endothermic. Adding heat causes the equilibrium to shift right, favoring the forward reaction and the blue cobalt complex. Removing heat shifts the equilibrium to the left, favoring the reverse reaction and the pink cobalt complex.

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