Copper and Le Chatelier

There is a well-known experiment that is used to great effect to show the effect of heat on the position of an equilibrium. The solution used contains cobalt chloride and hydrochloric acid resulting in an equilibrium between cobalt ions that have either pink aqueous or blue chloride ligands.

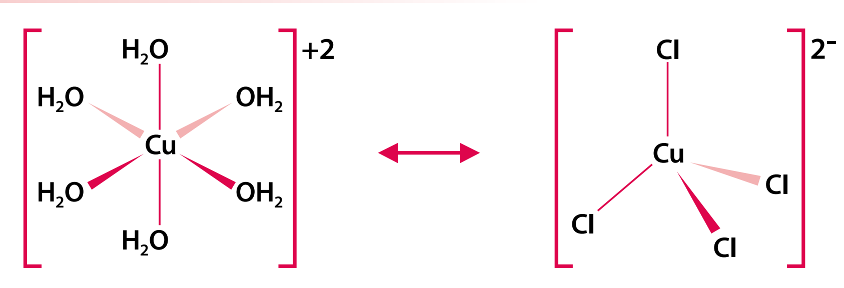

Figure 1 *- Structures of water and chloride complexes of copper.*

Cobalt, however, is significantly carcinogenic so here is a safer version using copper rather than cobalt: one that students can carry out themselves with negligible risk.

Background

The effect of temperature on the position of an equilibrium can easily be seen by observing the colour changes of the octahedral hexaaquacopper(II) cation and the tetrahedral tetrachlorocopper(II) anion as the sample is moved from a low temperature to a high temperature (see Figure 1).

The change in co–ordination number from 6 to 4 comes about as a result of the larger size of the chloride ion ligand resulting from the complex ion shape changing from octahedral to tetrahedral.

The colour of the complex changes from blue to green (In fact the chloride complex is actually a yellow-brown but is rendered green due to residual blue).

The complex changes from a cationic complex ion (2^+, unaffected by water) to an anionic complex ion 2^-, (2^+ countered by 4 x Cl^-). But there is no oxidation state change at all, copper is in the +2 state throughout the reaction.

The change is not as clear as using cobalt but it is certainly observable.

What you will need

- 1.0 M copper(II) chloride solution
- Distilled water
- Saturated sodium chloride solution
- 100 cm^3 measuring cylinder
- 250 cm^3 beaker
- 100 cm^3 beaker of hot water
- 100 cm^3 beaker of iced water

Optional
- Hair dryer
- Small square of polystyrene with hole cut in it (diameter of the test tube)

Preparation of the solution

1) Prepare a 1.0 M solution of the copper chloride in the large beaker by dissolving 17.0 g in 100 cm^3 of distilled water.
2) Place 20 cm^3 of the solution in each of three test tubes.
3) Add a few cm^3 of the sodium chloride solution to one tube until the solution turns bright green.
4) Taking this tube and one of your originals as your end points, add a small amount of the sodium chloride solution until its colour is a green/blue - half-way between the two others. (See Figure 2)

The activity

1) Take 2 beakers, one containing boiling (or very hot) water and the other with iced water.
2) Divide your 'balanced' solution between two test tubes and place one in each beaker.
3) Leave for a minute or two and then inspect the colours.
4) Now swap them over and see what happens.

OR

1) Place the polystyrene collar about a third of the way down your test tube.
2) Place the test tube in the beaker of iced water to a depth of about a third its length. At the same time use the hair dryer to heat the top third of the solution above the polystyrene.
3) (The polystyrene is to stop the middle third of the test tube heating up).
4) Observe the colour changes! (See Figure 3)

If desired, test tubes of the solution can be prepared for students to carry out the experiment for themselves.

The equilibrium equation can be expressed as follows:

$$\underbrace{Cu(H_2O)_6^{2+}(aq)}_{\text{Blue}} + 4Cl^-(aq) \longleftrightarrow \underbrace{CuCl_4^{2-}(aq)}_{\text{Green}} + 6H_2O(l) \qquad \Delta H+ve$$

Formula 1 *- Equilibrium equation.*

Figure 2 *- Left to right, Cl, mix, H_2O.*

At Higher level, if they are given the equilibrium equation and the colours of the ions in solution, their observations of the colours at low and higher temperatures can lead them to predict if the ΔH for the forward reaction is positive or negative.

It must be said that the colour changes are not quite as clear as with cobalt chloride, which could perhaps be done as a demonstration to begin or end the lesson, but it is significantly safer.

Figure 3 *- Result of using a hairdryer.*

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