Related concepts
Integral enthalpy of solution, Hess’s law, lattice energy, ion solvation, calorimetry.

Principle
When a solid electrolyte dissolves in water, a positive or negative heat effect occurs as a result of the destruction of the crystal lattice and the formation of hydrated ions. The enthalpy of hydration of copper sulphate can be calculated from the different heats of reaction measured when anhydrous and hydrated copper sulphate are separately dissolved in water.

Tasks
1. Record temperature-time curves for the dissolution of anhydrous copper sulphate and hydrated copper sulphate in water.
2. Calculate the hydration enthalpy of anhydrous copper(II)sulphate.

Equipment
Cobra3 Basic-Unit 12150.00 1
Power supply 12 V/2 A 12151.99 1
Data cable, RS232 14602.00 1
Temperature measuring module Pt 100 12102.00 1
Software Cobra 3 Temperature 14503.61 1
Temperature probe Pt 100 11759.01 1
Calorimeter, transparent 04402.00 1
Heating coil with sockets 04450.00 1
Work and power meter 13715.93 1

Figure 1. Experimental set-up.
Set-up and procedure

Set up the experiment as shown in Fig. 1 but for the time being do not connect the heating coil to the work and power meter. Prepare the two copper salts by grinding each of them separately to a fine powder in a mortar. Make sure that the anhydrous copper sulphate really is anhydrous by heating it in a porcelain dish over a Butane burner until it is completely white and allowing it to cool in a desiccator. Weigh 24.97 g (0.1 mol) of copper(II) sulphate and 15.96 g (0.1 mol) of anhydrous copper(II) sulphate in two separate beakers (weighing accuracy 0.01 g). Fill the calorimeter with 900 g of distilled water (weighing accuracy 0.1 g).

Connect the temperature probe to T1 of the measuring module. Call up the ‘Measure’ programme in Windows and enter <Temperature> as measuring instrument. Set the measuring parameters as shown in Fig. 2. Under <Diagram 1> select Temperature T0a, the appropriate range for the temperature and the X bounds and ‘auto range’. Now calibrate your sensor under <Calibrate> by entering a temperature value measured with a thermometer and pressing <Calibrate>. After having made these settings, press <Continue> to reach the field for the recording of measured values. Arrange the displays as you want them. Place the filled calorimeter on the magnetic stirrer, insert the oval magnetic stirrer bar and switch on the stirrer (Caution: Do not switch on the heating unit by mistake!). Insert the heating coil and the temperature probe into the lid of the calorimeter and fix them in position.

When temperature equilibrium has been reached (after approximately 10 min) start the measurement by pushing <Start measurement>. Wait 3 to 4 minutes, then add the first copper salt to the water by pouring it through the powder funnel which has been inserted in the opening of the lid. Make sure that the entire quantity of salt is added to the water without any loss. Continue to measure until a new equilibrium has been reached. Subsequently perform electrical calibration to determine the total heat capacity of the calorimeter. Supply 10 V AC to the work and power meter for the electric heating. Push the <Reset> button and then put the free ends of the heating coil connection cables into the output jacks. The system is now continuously heated and the quantity of energy supplied is measured. When the work and power meter shows approximately 4000 Ws, switch off the heating and read the exact quantity of electrical energy supplied. After a further three minutes, stop recording the temperature. Figs. 3 and 4 show the graphs as they are presented by the programme when the measurements are stopped. If you use <survey> from the toolbar you can read the temperature difference data.

Repeat the experiment to determine the enthalpy of solution of the second copper salt. At least two measurements for each salt should be performed to avoid errors and to calculate a mean value.

Theory and evaluation

The dissolving of a solid electrolyte in water is primarily determined by two simultaneously occurring processes: the destruction of the crystal lattice and the hydration of the ions. The destruction of the crystal lattice is an endothermic process because energy is required to break down the chemical bonds, whereas the hydration of the ions is exothermic. Depending on the type of lattice, and on both the radius and the charge of the ions (charge density), the resulting enthalpy of solution can be either endothermic or exothermic.
When a salt exists in both hydrated and dehydrated forms, and on assuming that when the hydrated salt dissolves only the degradation of the crystal lattice occurs, the enthalpy of hydration can be calculated using Hess's theorem.

\[ \Delta H_{\text{CuSO}_4} = \Delta L_{\text{H}_{\text{CuSO}_4}} - \Delta L_{\text{H}_{\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}}} \]

The integral enthalpy of solution can be calculated according to equation (3).

\[ Q_{\text{exp}} = Q_{\text{cal}} \cdot \frac{\Delta T_{\text{exp}}}{\Delta T_{\text{cal}}} \]

Data and results

- \( M_{\text{CuSO}_4} = 159.6 \text{ g} \cdot \text{mol}^{-1} \)
- \( M_{\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}} = 249.68 \text{ g} \cdot \text{mol}^{-1} \)
- \( \Delta L_{\text{H}_{\text{CuSO}_4}} = -66.2 \text{ kJ} \cdot \text{mol}^{-1} \)
- \( \Delta L_{\text{H}_{\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}}} = +11.5 \text{ kJ} \cdot \text{mol}^{-1} \)
- \( \Delta H_{\text{CuSO}_4} = -77.7 \text{ kJ} \cdot \text{mol}^{-1} \)
Determination of the hydration enthalpy of an electrolyte