Experiment 5

Heat and Temperature

"This coffee isn't hot enough!"

The Task

In this experiment you will study the heat flow associated with a range of processes and examine the relationship between heat and temperature.

Skills

By the end of the experiment you should be able to:
- use a temperature probe,
- understand how a calorimeter works.

Other Outcomes

- You will determine the heat capacities of different substances.
- You will appreciate the relationship between heat and temperature.

The Assessment

You will be assessed on your ability to calibrate the temperature probe correctly.
See Skill 14.

Introduction

Transformations are the heart of chemistry – reactants are converted into products and one phase into another. Conservation of mass requires that any new material formed must have arisen by transforming material already present. Understanding which conditions are required to make such transformations occur is very important. Many transformations involve the flow of heat. In this experiment you will study the heat flow associated with a range of processes and examine the relationship between heat and temperature.

When energy in the form of heat flows into a material, its temperature goes up. The amount of heat required to make a sample's temperature increase by 1 °C (or 1 K) is called the <u>heat capacity</u> of that sample. The heat capacity depends on the type and amount of material present. If you double the mass of the sample you need twice as much heat to raise its temperature by 1 °C. Heat capacities of pure compounds are therefore often given in units of J g^{-1} or kJ mol^{-1}. In the case of a specific object (*e.g.* a piece of lab equipment), its total heat capacity is often calculated.

The SI unit of energy is the joule (symbol J). It takes 4.184 J to raise the temperature of 1 g of water by 1 °C.

Safety

Chemical Hazard Identification

potassium chloride – non-hazardous

sodium chloride – non-hazardous

magnesium chloride – non-hazardous

calcium chloride – irritant, avoid contact with the eyes and skin. Do not swallow.

ammonium chloride – irritant, avoid contact with the eyes and skin.

Risk Assessment and Control

Moderate risk.

There is a constant risk of burns and scalds due to Bunsen burners, boiling water baths and hot blocks of aluminium.

There is a risk of cuts from broken glass if the calorimeter or a stirring rod should break.

Waste Disposal

All of the salts can be washed down the sink with plenty of water. When you have finished the experiment, wash the solid aluminium metal with water and return it to the front bench.
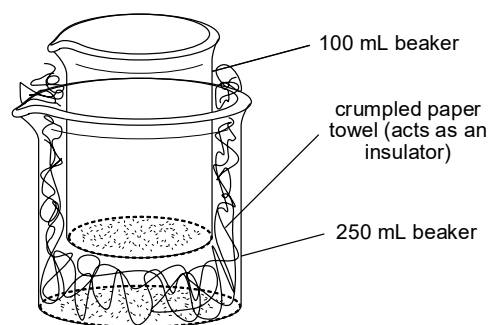
Experimental

> ### *This experiment is to be carried out in pairs.*

We'll start with a simple problem. We place something hot, a few degrees below 100 °C, into a fixed volume of water at room temperature. What is the final temperature of the water? Does it depend on what the hot material is?

To investigate changes of temperature in laboratory experiments, it's important to design the experiment to minimise unwanted heat exchange with the surroundings. This is done by constructing a simple calorimeter - a reaction vessel that is insulated from the surroundings.

Part A   Constructing a simple calorimeter

(A1)   Take a 250 mL beaker and half fill it with small loosely crumpled wads of dry paper towel.

(A2)   Insert a 100 mL beaker into the larger beaker, ensuring that the inner beaker is well separated from the large beaker everywhere, including the bottom, by the paper wads. This is to reduce heat loss. The small beaker is your *calorimeter*.

100 mL beaker

crumpled paper towel (acts as an insulator)

250 mL beaker

Part B   Calibration of the temperature probe

See Skill 14.

Part C   What changes as heat moves from one material to another?

In the following experiment we will measure the heat exchange between a hot aluminium sample and water at room temperature.

(C1)   Place 100 mL of tap water in a 400 mL beaker and heat it on a Bunsen burner until it reaches a steady boil. This is your hot water bath.

(C2)   Select an aluminium sample (approximately 30 g) from the front bench and weigh it using a top-loading balance (see Skill 3.1A). Record the mass in your logbook. Using metal tongs, carefully place the aluminium sample in your hot water bath and leave it until the water has been boiling for 5 minutes.

(C3)   Use the temperature probe to measure the temperature of the boiling water.

(C4)   Use a measuring cylinder to transfer 50.0 mL of tap water to your calorimeter. Use the temperature probe to measure the temperature to within 0.1 °C. **Record this value in your logbook**. Leave the temperature probe in the calorimeter.

(C5)   Quickly and very CAREFULLY, use the metal tongs to transfer the hot aluminium sample to your calorimeter, making sure that you do not add or lose any water. Start the stopwatch.

(C6)   In your logbook, record the time and temperature of the water in the calorimeter when the aluminium was added. Stir continuously with the temperature probe.

(C7)   Record in your logbook the temperature at 1 second, 15 seconds and 30 seconds and then at 30 second intervals until you reach 4 minutes.

(C8)   Remove the aluminium sample and refill the calorimeter with 50 mL of fresh water. Weigh it (*i.e.* the 100 mL beaker + water) using a top-loading balance and record the mass in your logbook.

(C9)   Place 30 mL of coloured water in a clean 100 mL beaker with a large wooden peg attached. Place this beaker in your hot water bath, making sure the level of the sample is lower than the level of water in the hot water bath. Leave for 10 minutes; you may need to add more water to the boiling water bath to keep the water level constant.

(C10)   Measure the temperature of the coloured water and record it in your logbook. Cool the temperature probe under some running water and then measure the temperature of the water in the calorimeter. Record the value in your logbook.

(C11)   Quickly and very CAREFULLY, use the metal tongs to pour the hot coloured water into your calorimeter and start stirring continuously with the temperature probe. Start the stopwatch.

(C12)   Record in your logbook the temperature at 1 second, 15 seconds and 30 seconds and then at 30 second intervals until you reach 4 minutes.

(C13)   Reweigh the calorimeter and record its mass in your logbook and hence calculate the mass of coloured water that was added to the calorimeter.

The key observations are the final (*i.e.* maximum) temperatures of the water in the calorimeter after addition of each of the two samples – the hot aluminium and the hot water.

For your logbook:

Plot the time dependence of the temperature for the two sets of measurements you made.

Any time dependent increase in the temperature of the water in the calorimeter reflects the finite rate at which heat flows from the hot sample you added. Any time dependent decrease in temperature corresponds to a failure of your calorimeter to prevent heat flowing to the surroundings.

Briefly comment on what information your plots of temperature vs *time provide with respect to these two aspects of heat flow.*

Based on the final temperatures that you measured and, without actually calculating a heat capacity, determine which of your two samples (aluminium or coloured water) has the higher heat capacity.

According to the SI Data Book, the specific heat capacity of aluminium is 0.895 J K^{-1} g^{-1}. Using the initial and final temperatures of the water in the calorimeter and the aluminium sample, calculate the heat capacity of the water+calorimeter (in J K^{-1}). Assume that your calorimeter provides perfect insulation so that all the heat that flowed out of the aluminium, flowed into the water+calorimeter. Clearly lay out all your calculations and reasoning in your logbook.

Determine the heat capacity of water, using the data collected from the hot water sample. How does your experimental value compare with the value given in the Introduction to this experiment? What do you think were the main sources of error?

Knowing the heat capacity of the water+calorimeter, you can determine the heat flow associated with any process taking place in the calorimeter by measuring its temperature change.

Dissolving Solids

The simple process of dissolving a salt in water is dependent on two important processes. Firstly, the break up of the crystal lattice (*i.e.* NaCl(s) → Na$^+$ + Cl$^-$) requires the input of energy to overcome the strong ionic attractions between the ions. Secondly, energy is released when the ions are hydrated and bonds are formed between the charged ions and the polar water molecules. Let's see what the overall energy change is for the following salts: NaCl, KCl, MgCl$_2$, CaCl$_2$, NH$_4$Cl

If the temperature of the water in your calorimeter increases, it means that the dissolution of the salt crystals is exothermic, *i.e.* it released heat. If the temperature of the water decreases, the dissolution process is endothermic, *i.e.* the salt absorbs heat when it dissolves and hence the temperature of the surroundings (the water) goes down.

Part D   Measuring the heat flow for dissolution

Your demonstrator will assign you two salts to study.

(D1)   Use a measuring cylinder to transfer 50.0 mL of tap water to your calorimeter. Weigh it (*i.e.* the 100 mL beaker + water) using a top-loading balance and record the mass in your logbook. Use the temperature probe to measure the temperature to within 0.1 °C. **Record this value in your logbook**. Leave the temperature probe in the calorimeter.

(D2)   Take a clean small dry beaker and weigh into it 2 g of any one of the salts listed above. Break-up any lumps of solid with a spatula.

(D3)   Add all the salt into the calorimeter, start the stopwatch and stir.

(D4)   Record in your logbook the temperature at 1 second, 15 seconds and 30 seconds and then at 30 second intervals until you reach 5 minutes.

(D5)   Reweigh the calorimeter and record its mass in your logbook and hence calculate the exact mass of salt that was added to the calorimeter.

(D6)   Discard the solution down the sink, wash and carefully dry the beaker and repeat steps (D1) - (D5) for the other salt.

For your logbook:

Calculate the heat of dissolution of each salt in J g^{-1} of added salt. You will need to use the heat capacity for the water+calorimeter from Part C. Lay out your complete calculations clearly in your logbook. Do not use scrap paper.

Which salts dissolved exothermically and which salts dissolved endothermically?

Did you find any correlations between the type of ions – cation or anion – and whether dissolution was exothermic or endothermic?

Group Discussion

Are you surprised that a spontaneous process can be endothermic, *i.e.* involving a spontaneous increase in energy (heat transfer from the water)? If yes, why? Where does your impression that spontaneous processes should decrease their energy come from?

What chemical processes occur when a salt dissolves in water?

From your experiment today, can you suggest which salts have stronger interactions in their crystal lattices relative to the interactions of the separated ions with water?

Conversely, can you suggest which salts have stronger interactions of their separated ions with water relative to the strength of their crystal lattices?