Appendix 2

8. Measuring the enthalpy changes of exothermic and endothermic reactions – Student Sheet

Chemical reactions are often accompanied by energy changes. These experiments show you how energy changes can be measured.

Intended lesson outcomes

By the end of this practical you should:
- have further developed skills in manipulating apparatus and accurate measurement of temperature;
- know that some chemical reactions are accompanied by energy changes in the form of heat energy;
- know that energy changes can be exothermic (ΔH, negative) or endothermic (ΔH, positive);
- be able to calculate enthalpy changes from your experimental results using enthalpy change = $mc\Delta T$.

Background information

To make a chemical bond, another bond must first be broken. It is the sum of the energy changes in making and breaking bonds that results in the overall energy change in a reaction. **Enthalpy** is the stored energy in a compound. It has a symbol **H**. If some of this stored energy is released during a reaction, the surroundings heat up. This is an exothermic reaction and we say that the change of enthalpy, ΔH, is negative. (Δ means *change of*.) However, if energy is absorbed during a reaction, then ΔH is positive and the surroundings cool down.

To measure enthalpy changes you are going to use a polystyrene cup as a calorimeter (Fig. 8.1). A calorimeter is any container that is used to measure energy changes from chemical reactions.
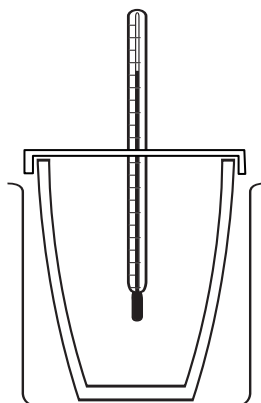
Figure 8.1

Question

1   The simple apparatus in Fig. 8.1 is very effective in minimising energy losses and this is important if we are going to measure accurately any energy changes arising from reactions.

Explain three ways in which heat energy loss is minimised using this apparatus.
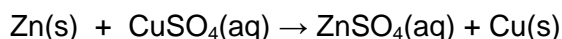
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Safety

(eye protection symbol)	**You must wear eye protection throughout this experiment**.
(harmful X symbol) (environment symbol)	Copper sulphate is **harmful** and **dangerous for the environment**
(flammable symbol)	Zinc powder is **highly flammable**.

Procedure

Experiment 1: The displacement reaction between zinc and copper(II) sulphate

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

Use a measuring cylinder to pour 50 cm^3 of 0.2 mol dm^{-3} copper(II) sulphate into the polystyrene cup. Weigh out about 1 g of powdered zinc onto a piece of paper. This is an excess (more than is needed to react with the copper(II) sulphate) so it does not need to be measured accurately. Record the initial temperature of the solution in the cup and then add the powdered zinc. Stir the solution gently and constantly with the thermometer until the temperature stops changing. Record the final temperature reached. Rinse the cup with water ready for use in Experiment 2.

Experiment 2: The reaction between citric acid and sodium hydrogen carbonate

$C_6H_8O_7(aq) + 3NaHCO_3(s) \rightarrow C_6H_8O_7Na_3(aq) + 3CO_2(g) + 3H_2O(l)$
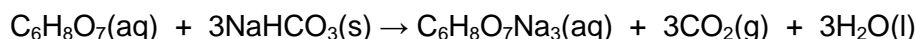
Use the same procedure as in Experiment 1 but this time use 25 cm^3 of 1 mol dm^{-3} citric acid and weigh out 8 g sodium hydrogen carbonate. This reaction froths up, so add the powdered solid slowly enough to prevent it coming over the top of the cup. Keep stirring and record the maximum temperature change.

Processing your results

We assume that all the energy in these reactions is exchanged with the surrounding water and no energy is transferred to the air, the glass of the thermometer, or the polystyrene cup. We know that the specific heat capacity of water, ***c***, is approximately 4.2 J g^{-1} K^{-1}. This means that 4.2 J of energy will raise the temperature of 1 g of water by 1 K. (This is the same as 1ºC.) As the solution is dilute, we also assume it has the same specific heat capacity of water and the same mass. 1 cm^3 of water weighs 1 g.

We are going to use the relationship:
enthalpy change = mass x specific heat capacity of the solution x temperature change
enthalpy change = $mc\Delta T$

For each experiment, calculate the enthalpy change that took place.

To complete the calculations, chemists express their results as enthalpy changes per mole of equation. In the first reaction, 1 mole of copper(II) sulphate appears in the balanced equation; in the second reaction there is 1 mole of citric acid.

Work out the enthalpy change per mole and indicate whether it is exothermic or endothermic by using '+' or '−' next to your enthalpy values.

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Questions

2 **(a)** In Experiment 1, zinc was in excess. Calculate the moles of copper(II) sulphate used and the moles of zinc used. Now use the reacting moles in the balanced equation to show that zinc is indeed the reactant that is in excess.

(b) Similarly, show that sodium hydrogen carbonate is the reagent in excess.

3 Write the ionic equations for the reactions in Experiments 1 and 2. Do not include any spectator ions.

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8. Measuring the enthalpy changes of exothermic and endothermic reactions – Teachers' Notes

In these two experiments students have the opportunity to accurately measure a temperature change and calculate the enthalpy changes of the reactions.

Learning outcomes

These are printed on the Student Sheet.

A suggested approach

Students will already have some familiarity with exothermic and endothermic reactions but it will probably be worth reminding them about this before they start. The concept of bond breaking and bond making and how this contributes to the overall enthalpy change is mentioned in the 'Background information' section on the Student Sheet, as is the requirement to minimise heat losses. Under the 'Procedure', the concept of having one reactant in excess is described but the reason is explored in Question 2.

The practical will take about 40 minutes, leaving time for the 'Processing your results' section to be tackled. Students could finish off calculations and answer the questions for homework.

Technical information

Requirements per student/group:

Polystyrene cup with lid
Glass beaker large enough to stand the polystyrene cup in but not much larger (see Fig. 8.1 on the Student Sheet)
Thermometer, $-5°C$ to $50°C$
50 cm^3 and 25 cm^3 measuring cylinders
1 mol dm^{-3} citric acid, approximately 30 cm^3
0.2 mol dm^{-3} copper(II) sulphate, approximately 60 cm^3 **(harmful)**
1 g zinc powder **(highly flammable)**
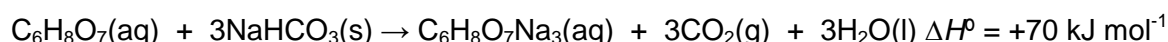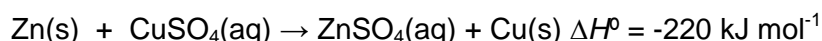8 g sodium hydrogen carbonate
Access to a balance weighing 0.01 g

Safety

The main points are included on the Student Sheet but it is the teacher's responsibility to ensure that a full risk assessment is carried out prior to the practical session. MSDS sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident.

Processing the results

Although this is described in the Student Sheet, students may need help in understanding the calculation per mole of reaction. The approximate values are given below.

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ ΔH^o = -220 kJ mol^{-1}

$C_6H_8O_7(aq) + 3NaHCO_3(s) \rightarrow C_6H_8O_7Na_3(aq) + 3CO_2(g) + 3H_2O(l)$ ΔH^o = +70 kJ mol^{-1}

Answers to questions on the Student Sheet

1   Polystyrene is a good insulator, so heat loss from the sides is minimised. The lid prevents energy loss from the top and the glass beaker provides another layer of insulation for the sides, as the air trapped in the beaker is also a poor conductor.

2   **(a)** If 1 g of powdered zinc is used, then moles of Zn = 1/65 = 0.015 mol.

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The moles of $CuSO_4$ in solution = 0.2 x 50/1000 = 0.01 mol.

Since the equation involves 1 mole of $CuSO_4$ reacting with 1 mole of Zn, the zinc is in excess.

(b) If 8 g of sodium hydrogen carbonate is used, moles $NaHCO_3$ = 8/84 = 0.095 mol.

Moles of citric acid = 1 x 25/1000 = 0.025 mol.

From the equation, 3 moles of $NaHCO_3$ exactly react with 1 mole citric acid. 3 x 0.025 = 0.075 mol. This is 0.02 mol less than the actual moles added, so $NaHCO_3$ is in excess.

(c) To ensure complete reaction:

$3Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

$C_6H_8O_7(aq) + 3HCO_3^-(s) \rightarrow C_6H_8O_7^{3-}(aq) + 3CO_2(g) + 3H_2O(l)$