

**Cambridge International**

**AS and A Level Chemistry (9701)**

Practical booklet 4

Enthalpy changes

**Introduction**

Practical work is an essential part of science. Scientists use evidence gained from prior observations and experiments to build models and theories. Their predictions are tested with practical work to check that they are consistent with the behaviour of the real world. Learners who are well trained and experienced in practical skills will be more confident in their own abilities. The skills developed through practical work provide a good foundation for those wishing to pursue science further, as well as for those entering employment or a non-science career.

The science syllabuses address practical skills that contribute to the overall understanding of scientific methodology. Learners should be able to:

1. plan experiments and investigations
2. collect, record and present observations, measurements and estimates
3. analyse and interpret data to reach conclusions
4. evaluate methods and quality of data, and suggest improvements.

The practical skills established at AS Level are extended further in the full A Level. Learners will need to have practised basic skills from the AS Level experiments before using these skills to tackle the more demanding A Level exercises. Although A Level practical skills are assessed by a timetabled written paper, the best preparation for this paper is through extensive hands-on experience in the laboratory.

The example experiments suggested here can form the basis of a well-structured scheme of practical work for the teaching of AS and A Level science. The experiments have been carefully selected to reinforce theory and to develop learners’ practical skills. The syllabus, scheme of work and past papers also provide a useful guide to the type of practical skills that learners might be expected to develop further. About 20% of teaching time should be allocated to practical work (not including the time spent observing teacher demonstrations), so this set of experiments provides only the starting point for a much more extensive scheme of practical work.

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**Practical 4 – Guidance for teachers**

**Enthalpy changes**

**Aim**

To determine the enthalpy change for a metal displacement reaction by adding a known mass of zinc powder to excess of aqueous copper (II) sulfate solution and recording the rise in temperature.

**Outcomes**

Syllabus section 5.1(a), 5.1(b)(i), 1.5(b)(i) and (iii), 5.1(c), 5.2(a) and 6.1(a)(ii) as well as experimental skills 2, 3 and 4

Further work: syllabus section 7.2(b) and experimental skill 1

**Skills included in the practical**

|  |  |
| --- | --- |
| **AS Level skills** | **How learners develop the skills** |
| MMO collection | set up and use the apparatus to the level of precision indicated  |
| MMO quality | obtain results that are close to those of an experienced chemist |
| PDO recording | record the times and temperatures with appropriate headings and units  |
| PDO display | show the level of precision of their thermometer readingsshow working in the calculation and use significant figures appropriate to the precision of measurements |
| PDO layout | results clearly tabulated |
| ACE analysis | calculate numbers of moles of reactants and the enthalpy change of the reaction |
| ACE conclusions | determine whether an exo- or endothermic reaction has taken place |
| ACE improvements | suggest ways to improve the accuracy of the procedure suggest ways in which to extend the investigation to answer a new question  |

**Method**

* Learners must wear eye protection for this investigation.
* This technique can be used to determine enthalpies of reaction for a variety of reactions involving solutions. The reaction must be fast enough to be complete within a few minutes (otherwise there is too much heat loss) and it is normal that one of the reactants is used in excess.
* Temperatures are always measured to half of the smallest graduation on the thermometer. In examinations, thermometers reading to 1 oC are specified, so learners should be trained to read them to the nearest 0.5 oC. Solutions must always be stirred thoroughly before any temperature is recorded and the bulb of the thermometer must be totally immersed in the solution.
* Learners should be taught how to carry out calculations of Δ*H* from the temperatures they measure. They should understand why it important to identify which reagent is in excess. They should remember that, for exothermic reactions, Δ*H* is negative and that Δ*H* is positive when the reaction proceeds with a decrease in temperature.
* Foamed plastic (polystyrene) cups provide a reasonable level of insulation against heat loss. Plotting a cooling curve provides a correction for heat that is inevitably lost while a reaction is taking place. Method B (see Worksheet) therefore provides more accurate values of Δ*H*.
* Learners should become familiar with using the apparatus provided so they can perform experiments accurately.
* **Extension involving Hess’ Law**: learners could react magnesium turnings **[F]** with excess CuSO4(aq), and determine the enthalpy change for this reaction. Using the Δ*H* value for the reaction of zinc with CuSO4(aq), Hess’ Law can be used to determine the Δ*H* value for the reaction of magnesium turnings with zinc sulfate solution.
* These techniques can be used for **further work** when studying strong and weak acids and alkalis. Learners can plan a suitable method and concentrations of reactants to use, in order to investigate whether the overall enthalpy change of reaction differs for strong and weak acids and alkalis when mixed in different combinations in neutralisation reactions.

**Results**

* Learners should tabulate results showing a consistent number of decimal places in the balance readings and all thermometer readings recorded to either #.0 or #.5 °C. Headings should be unambiguous with units shown as / g or (g) and / °C or (°C), as specified in the syllabus.

**Interpretation and evaluation**

* The equations *n* = *m/A*r and *n* = *cV* can be introduced or revised.
* The appropriate number of significant figures can be discussed. Answers to 2 – 4 sf are appropriate given the precision of the measuring instruments. The accuracy of the procedure is not as great as for a titration exercise, and in some experiments Δ*T* may be < 10.0 °C, so answers to 2 sf are acceptable, especially if a 1 dp balance is used.
* Errors in measurements made with balances, measuring cylinders and thermometers can be discussed. The difference between needing one reading (as in a measuring cylinder) and two readings (for the mass of zinc or the change in temperature) can be linked to maximum percentage errors. Heat energy losses can be discussed.
* Ways of reducing the inaccuracies can be discussed.

(i) A lid with a hole for the thermometer can be used to reduce heat energy losses by convection. (The plastic cup should be of foamed plastic so would be a very poor conductor of heat. The supporting beaker further reduces heat losses by conduction.)

(ii) A thermometer calibrated to 0.2 °C can be used so that the maximum error is ± 0.1 °C (max % error for Δ*T* = 37 ºC is 2.7% when using a thermometer accurate to ± 0.5 °C).

(iii) The volume of CuSO4(aq) can be measured with a burette (more precisely calibrated than a measuring cylinder) so that the volume of solution heated would be more accurate.

(iv) The amounts of reactants can be scaled up to decrease errors in volume and mass measurements. In any proposed scaling the CuSO4 must remain in excess. (The use of a balance accurate to more dp is not significant unless the balance used in the experiment was to 1 dp.)

**Typical results**

Initial mass of weighing boat / g = 2.54

Mass of weighing boat and zinc / g = 4.57

Final mass of weighing boat / g = 2.55

Mass of zinc added / g = 2.02

Volume of aqueous CuSO4 used = 40 cm3

Initial temperature reading / ºC = 22.0

Final temperature reading / ºC = 59.0

Temperature change / ºC = 37.0

**Calculation** (unrounded values may be carried in the calculator)

Heat energy given out =*mc*Δ*T* = 40 x 4.18 x 37 = 6186 = 6.19 x 103 J

Moles of Zn = 2.02 / 65.4 = 3.09 x 10–2 mol

Δ*H* = – 6186 / (3.09 x 10–2 x 103) = – 200 kJ mol

Graph of results for **Method B**

Temperature rise at 3 minutes = 61.0 – 22.0 = 39.0 °C; Δ*H*  = – 211 kJ mol–1

**Questions**

Discussion of which atoms / ions are taking part in the reactions leads to a suitable Hess’ Law diagram.



(Using the results and data given in the learner section,

Δ*H* = –320 – (–200) = –120 kJ mol‑1;

theoretical value = – (–153.89) + (–466.85) = –312.96 = –313 kJ mol–1.)

**Extension**

If the extension practical exercise is carried out then possible reasons for the larger difference in Δ*H* for the experiment with magnesium turnings with copper sulfate solution from the theoretical value using Δ*H*f~~o~~ compared with that for zinc powder can be discussed. Learners may note some bubbling when magnesium is added to copper sulfate solution showing a competing reaction (releasing hydrogen) is occurring.

**Further work**

The students can plan an experiment to determine the enthalpy change of neutralisation reactions. This can be carried out in groups with each using a different combination of strong and weak acids and alkalis. This supports learning objective 5.1(b)(ii) and links up with learning objective 7.2(b) in addition to those listed above. Discussion can help learners decide what volumes and concentrations would give them a sensible temperature rise.

**Practical 4 – Information for technicians**

**Enthalpy changes**

**Each learner will require:**

|  |  |  |
| --- | --- | --- |
|  | (a) | Eye protection |
|  | (b) | 250 cm3 beaker |
|  | (c) | 1 x foamed plastic (polystyrene) cup approximately 150 cm3 |
|  | (d) | 1 x thermometer (–10 ºC to +110 ºC at 1 ºC) |
|  | (e) | 1 x weighing boat or 100 cm3 beaker |
|  | (f) | 1 x 50 cm3 measuring cylinder |
|  | (g) | access to a balance reading to **at least** 1 dp |
| **[H] [N]** | (h) | 50.0 cm3 1.0 mol dm–3 copper(II) sulfate |
| **[F] [N]** | (i) | 2.0 ± 0.1 g zinc powder (supplied in a stoppered container) |

**Extension:**

Apparatus and 1.0 mol dm–3 copper(II) sulfate as above

|  |  |  |
| --- | --- | --- |
| **[F]** | (j) | 0.8 ± 0.1 g magnesium turnings (supplied in a stoppered container) |

**Hazard symbols**

|  |  |
| --- | --- |
| **C** = corrosive substance | **F** = highly flammable substance |
| **H** = harmful or irritating substance | **O** = oxidising substance |
| **N** = harmful to the environment | **T** = toxic substance |

**Practical 4 – Worksheet**

**Enthalpy changes**

**Aim**

To determine the enthalpy change for a metal displacement reaction by adding a known mass of zinc powder to excess aqueous copper(II) sulfate solution and recording the rise in temperature.

**Method A**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Safety:*** Wear eye protection.
* 1.0 mol dm–3 copper(II) sulfate **[H] [N]**
* zinc powder **[F] [N]**

**Hazard symbols**

|  |  |
| --- | --- |
| **H** = harmful or irritating substance | **F** = highly flammable substance |
| **N** = harmful to the environment |  |

 |

1. Weigh the container with zinc.

2. Support the plastic cup in a 250 cm3 beaker.

3. Using a 50 cm3 measuring cylinder, pour 40 cm3 of the copper(II) sulfate solution into the plastic cup.

4. Measure the temperature of the copper(II) sulfate solution. Record your thermometer reading.

5. Add all the zinc from the container to the copper(II) sulfate solution in the plastic cup.

6. Use the thermometer to stir the mixture gently while the reaction takes place.

 (It is usually poor practice to stir a mixture with a thermometer. However, it is acceptable here because it is less likely that the thermometer will break with a plastic cup being used.)

7. Measure and record the highest temperature reached.

8. Reweigh the empty container (with any traces of zinc remaining). Record this mass.

 Calculate and record the mass of zinc used.

9. Record the change in temperature and the mass of zinc powder added.

10. If there is time, repeat the whole experiment.

**Method B**

Carry out steps 1 – 3 as in Method A

4. Draw a table for your results. You will be taking thermometer readings almost every ½ minute for 8 minutes.

5. Measure the temperature of the copper(II) sulfate solution. Record your thermometer reading. This is the temperature at time = 0.

6. Start the stop clock and take the temperature of the copper(II) sulfate solution every ½ minute for 2½ minutes. Record all your thermometer readings.

7. At 3 minutes add all the zinc to the copper(II) sulfate solution in the plastic cup.

8. Use the thermometer to stir the mixture gently while the reaction takes place.

9. Measure and record the temperature of the mixture every ½ minute from 3½ minutes to 8 minutes.

10. Reweigh the empty container (with any traces of zinc remaining). Record this mass.

 Calculate and record the mass of zinc used.

11. Plot a graph of temperature / °C (*y*-axis) against time / minutes (*x*-axis).

12. Draw two lines of best fit, one before adding zinc and the other after the highest temperature is reached.

13. Extrapolate the two lines to 2½ minutes and find the theoretical change in temperature at this time.

**Results**

Record **all** your observations.

Tabulate your results showing a consistent number of decimal places in the balance readings and all thermometer readings recorded to either #.0 or #.5 °C. Headings should be unambiguous with units shown as / g or (g) and / °C or (°C), as specified in the syllabus.

**Interpretation and evaluation**

**Calculation – methods A and B**

**Data required:** specific heat capacity, c, of the solution may be taken as 4.18 J cm–3 °C–1;

*A*r: Zn, 65.4; equation for the reaction,

Zn(s) + CuSO4(aq) → ZnSO4(aq) + Cu(s)

1. Calculate the heat energy given out in J.

2. Calculate the number of moles of zinc reacting.

3. Show, by calculation, that copper(II) sulfate was in excess.

4. Use your answers to 1 and 2 to calculate the enthalpy change, in kJ mol–1, of the reaction between zinc and aqueous copper(II) sulfate given above. (Remember to include the sign to show whether the reaction was exothermic or endothermic.)

**Points to consider**

1. (i) What is the error in a single thermometer reading?

 (ii) What is the percentage error in your temperature rise?

 (iii) What changes could you make to reduce this error?

2. Other errors in this experiment may arise from

* imprecise volume of aqueous copper(II) sulfate;
* heat losses.

 What changes could you make to reduce these errors?

**Questions**

1. A second experiment was carried out using magnesium instead of zinc.

 It was found that 0.40 g of Mg raised the temperature of 40 cm3 of copper sulfate solution by 31.5 °C. Calculate the enthalpy change for this reaction.

2. Use your own results and the answer to question 1 to calculate the enthalpy change for the following reaction.

Mg(s) + ZnSO4(aq) → MgSO4(aq) + Zn(s)

 You will need to construct a Hess’ Law energy cycle.

3. Some enthalpies of formation, in kJ mol-1, of aqueous ions are given below.

|  |  |
| --- | --- |
| metal ion | ΔHf~~o~~ / kJ mol–1 |
| Cu2+(aq) | +64.8 |
| Mg2+(aq) | –466.9 |
| Zn2+(aq) | –153.9 |

 Use the data above to calculate the enthalpy changes for the following reactions.

* Zn(s) + CuSO4(aq) → ZnSO4(aq) + Cu(s)
* Mg(s) + CuSO4(aq) → MgSO4(aq) + Cu(s)
* Mg(s) + ZnSO4(aq) → MgSO4(aq) + Zn(s)

Compare these to the values that you have determined in your experiment and that you have calculated in Questions 1 and 2.