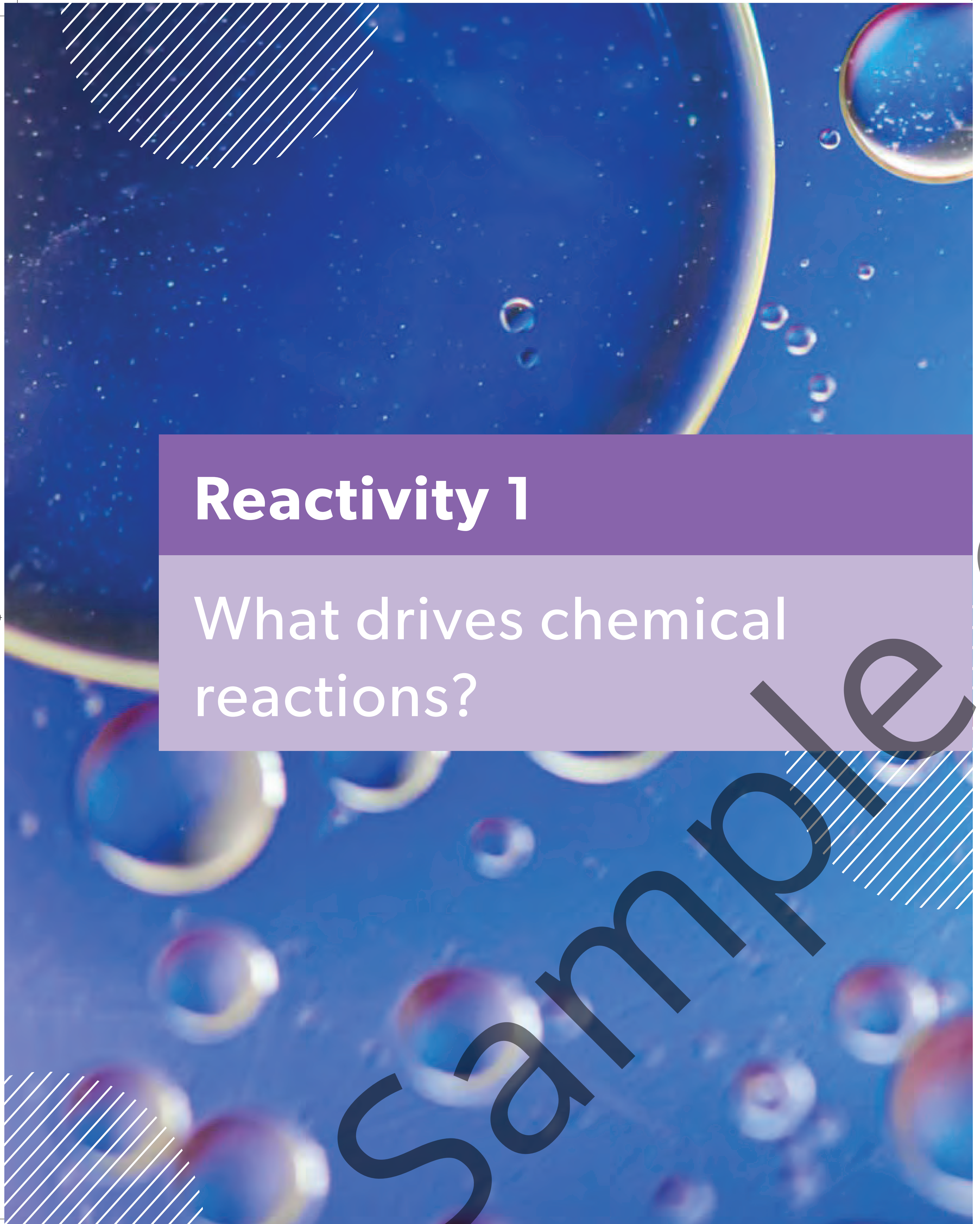


Reactivity 1

What drives chemical reactions?

Reactivity 1.1 Measuring enthalpy changes

What can be deduced from the temperature change that accompanies chemical or physical change?

Chemistry involves the study of chemical reactions and physical changes of state of the elements and their compounds. Conservation of energy is a fundamental principle of science, which is examined through observation and experimentation. The use of models, empirical or experimental data, the language of mathematics and scientific terminology, all contribute to our understanding of energy changes associated with chemical reactions.

An understanding of the relationships that exist between chemistry and energy involves understanding how energy is transferred between a chemical system and the surroundings. This information can in turn be used to develop an understanding of the relative stability of reactants and products, leading to better control over the progress of the reaction being studied.

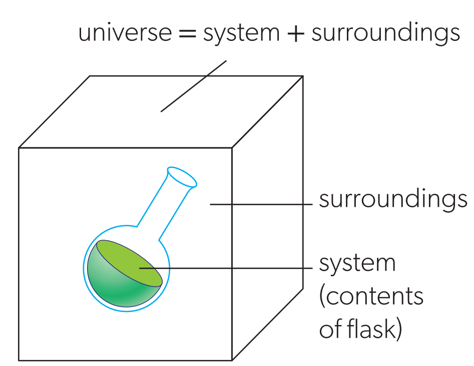
Understandings

Reactivity 1.1.1—Chemical reactions involve a transfer of energy between the system and the surroundings, while total energy is conserved.

Reactivity 1.1.2—Reactions are described as endothermic or exothermic, depending on the direction of energy transfer between the system and the surroundings.

Reactivity 1.1.3—The relative stability of reactants and products determines whether reactions are endothermic or exothermic.

Reactivity 1.1.4—The standard enthalpy change for a chemical reaction, $\Delta H^{\ominus}$, refers to the heat transferred at constant pressure under standard conditions and states. It can be determined from the change in temperature of a pure substance.



▲ Figure 1 The universe is the combination of the system and its surroundings

Energy transfer in chemical reactions (*Reactivity 1.1.1*)

In a chemical reaction, total energy is conserved. Chemical potential energy is stored in the chemical bonds of **reactants** and **products**, while the temperature of the reaction mixture is a function of the kinetic energy of the atoms, ions and molecules present.

All chemical reactions involve energy changes. Energy may be released into the **surroundings** from the **reaction system** or it may be absorbed by the reaction system from the surroundings. Most commonly, the energy is transferred in the form of heat, but it may also be in the form of sound or light.

In an **open system**, the transfer of matter and energy is possible across its boundary (for example, matter can be added to a beaker, and energy can be transferred through its sides). A **closed system** allows no transfer of matter, though energy may be transferred across the boundary. In an **isolated system**, matter and energy can neither enter nor exit, but can only move around inside.

▲ Figure 2 In each of the above scenarios, energy is transferred. In hot springs, energy is transferred as heat, and in fireworks, as heat, sound and light. Both scenarios are open systems

Models

Most industrial processes take place in open or closed systems. The loss of heat energy during an industrial process not only affects the efficiency of the chemical reaction, but also contributes to a loss of useful heat, an increase in thermal pollution and greenhouse gas emissions. Thermography can be used to model heat flow and loss from structures in chemical industries as a heat map, where red is hot and purple is cold.

▲ Figure 3 Thermograph of industrial engineering system

What are the advantages of modelling heat distribution and transfer? How can chemical engineers use the data collected to improve the efficiency of industrial processes? How does this help our environment?

What is the difference between heat and temperature?

Temperature, T, is an example of a **state function**. For a state function, any change in value is independent of the pathway between the initial and final measurements.

For example, if you take the temperature of the water in a swimming pool early in the morning (the initial value) and then again in the afternoon (final value), this does not tell you the complete story of any temperature fluctuations that may have occurred throughout the day. The calculation of the temperature change is simple:

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

▲ Figure 4 If you record the temperature of a pool at the beginning and the end of a day, this does not give you an indication of the heating and cooling that has occurred throughout the day, only the overall temperature change

Other examples of state functions include volume, enthalpy and pressure.

Heat, q, is a form of energy that is transferred from a warmer body to a cooler body, as a result of the **temperature gradient**. Heat is sometimes referred to as **thermal energy**. It can be transferred by the processes of conduction, convection and radiation.

Heat has the ability to do **work**. When heat is transferred to an object, the result is an increase in the average **kinetic energy** of its particles. This results in an increase in temperature and potentially a phase change, for example, a change of state from liquid to gas.

At **absolute zero**, 0 K (−273.15 °C), all motion of the particles theoretically stops and the **entropy**, S, of a system reaches its minimum possible value. The absolute temperature (in kelvin) is proportional to the average kinetic energy of the particles of matter. As the temperature increases, the kinetic energy or motion of the particles also increases.

Entropy is defined and explored in *Reactivity 1.4*.

ATL Communication skills

Your communication skills will develop incrementally throughout the entire chemistry programme. Communication skills include consistent and accurate application of scientific terminology. Your use of the terms "heat" and "temperature" in explanations will demonstrate your understanding of these concepts.

For example, consider the simple reaction between magnesium metal and hydrochloric acid.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

In this reaction, heat is released from the reaction system into the surroundings, and the temperature of the aqueous solution rises. When you think about heat, you are considering the transfer of thermal energy in the system. When you refer to the temperature of a system, you are describing the average kinetic energy of the particles within that system.

▲ Figure 5 Magnesium ribbon reacting with hydrochloric acid

Thermochemistry is the study of heat changes that occur during chemical reactions. Heat changes are often described in terms of **enthalpy**. At constant pressure, the **enthalpy change**, ΔH, is defined as the heat transferred by a closed system to the surroundings during a chemical reaction. The terms "enthalpy change" and "heat of reaction" are commonly used when describing the thermodynamics of a reaction. The most common unit of enthalpy change is kJ.

Activity

Imagine a glass of water containing ice cubes sitting in the summer sun. It will undergo a change in enthalpy.

- Is the glass of water an open, closed or isolated system?
- Identify the system and the surroundings.
- Explain the movement of heat in the form of energy, between the system and the surroundings.
- What would you observe on the outside of the glass? Explain this observation in terms of a change of state and movement of energy.

Exothermic and endothermic reactions (*Reactivity 1.1.2*)

A chemical reaction in which heat is transferred from the system to the surroundings is defined as an **exothermic** reaction. In contrast, chemical reactions in which heat is absorbed into the system from the surroundings are defined as **endothermic** reactions.

When a chemical reaction takes place, the atoms of the reactants are rearranged to create new products. Chemical bonds in the reactants are broken, and new chemical bonds are made to form products. Energy is absorbed by the reaction system to break the chemical bonds, and therefore bond breaking is an endothermic process. This energy is termed the **bond dissociation energy** and it can be quantified for each type of bond. Energy is released into the surroundings when new chemical bonds are made, and therefore bond making is an **exothermic** process. The **transfer of energy** between the surroundings and the system is an important part of your understanding of the energy changes in a reaction.

Exothermic reactions have a negative enthalpy change, and endothermic reactions have a positive enthalpy change. The sign of the enthalpy change is defined from the perspective of the system and not the surroundings. For example, in an exothermic reaction, heat is being lost by the system and so the enthalpy change is negative. For an endothermic reaction, heat is absorbed by the system, so the enthalpy change is positive.

To determine whether a reaction is endothermic or exothermic, we can use a calorimeter. A calorimeter is any apparatus used to measure the amount of heat being exchanged between the system and the surroundings. In the school laboratory, experiments focus on the change in temperature, ΔT, of the reaction solvent, which in most cases is water.

Observations

In the laboratory, observations can be made using human senses, or with the aid of instruments such as data-logging equipment. The application of digital technology to collect data is one of the essential skills in the study of chemistry.

Energy profiles (*Reactivity 1.1.3*)

The energy profiles for an endothermic or exothermic reaction enable you to examine the progress of a reaction as it proceeds from reactants to products. Energy profiles are a visual representation of the enthalpy change during a reaction. From an energy profile, you can determine the enthalpy of the reactants and the products, the **activation energy** (E_a), and the enthalpy change for the reaction.

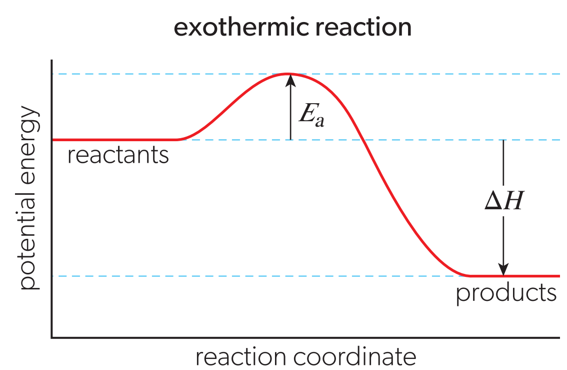
Activation energy, E_a, is the minimum energy required for the reaction to take place. You will study activation energy in *Reactivity 2.2*.

Many chemical reactions are exothermic. In these reactions, energy is released from the system to the surroundings. The reactants of this reaction are at a higher energy level and considered to be lower in stability. Products for exothermic reactions are at a lower energy level and considered to be more energetically favourable.

Consider the reaction between zinc and copper(II) sulfate solution. It is a single displacement reaction involving the displacement of the copper(II) ion by zinc:

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

Measured quantities of copper(II) sulfate solution and zinc are mixed in a calorimeter. The mixture is stirred, and the change in temperature of the solution is measured using a thermometer or data-logging equipment. In this reaction, heat is generated by the reaction system. This results in a heat transfer to the surroundings, so the temperature of the solution increases. The reaction is therefore exothermic.



▲ Figure 6 Using a thermometer or a temperature probe, you would observe an increase in the temperature of the reaction mixture in an exothermic reaction. The enthalpy of the products is lower than that of the reactants. You would describe the products as being energetically more stable than the reactants

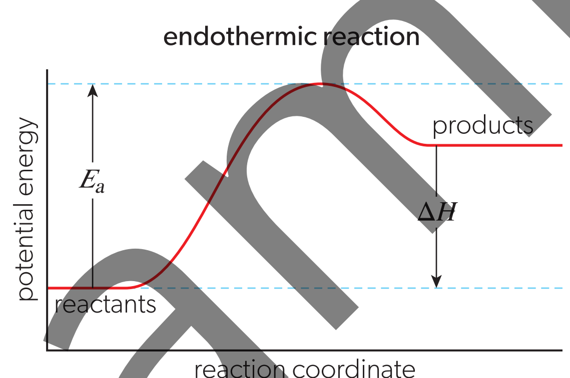
The graph in figure 6 is an example of an **energy profile**.

If you consider an endothermic reaction, the products of the reaction are at a higher energy level and therefore are less stable than the reactants.

Ammonium nitrate, NH_4NO_3, is an important component of fertilizers. When the solid dissolves in water to form aqueous ammonium and nitrate ions, the temperature of the solution decreases.

$$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

This heat is absorbed by the reaction system from the surroundings. The apparatus containing the reaction will feel cold to touch. This is an example of an endothermic reaction (figure 7).



▲ Figure 7 Using a thermometer or a temperature probe, you would observe a decrease in the temperature of the reaction mixture in an endothermic reaction. The enthalpy of the products is greater than that of the reactants. The products are described as being energetically less stable than the reactants

Global impact of science

Developments in science may have ethical, environmental, political, social, cultural and economic consequences, which must be considered during decision making. The pursuit of science may have unintended consequences. German chemist Fritz Haber was awarded the Nobel Prize in Chemistry in 1918 for developing a method to chemically extract nitrogen from the air by reacting it with hydrogen. Haber's discovery allowed for the large scale production of fertilizers that began during the green revolution and continues today. However, his process also provided Germany with a source of ammonia that was used for the production of explosives during the First World War. The global impact of science is evident in Haber's research.

ATL Communication skills

Communication skills cover a wide range of skills and forms of communication. Your ability to effectively communicate verbally and in written form most often comes to mind when you are thinking about improving communication skills. However, communication also involves your ability to read and write different forms of texts intended for different audiences. In science, you need to be able to write formal laboratory reports using specific terminology and accepted writing styles. Another form of communication you would utilize in writing reports and answering examination questions, is your ability to sketch graphs and extract data and meaningful information from graphs. Can you read and analyse the energy profiles that represent exothermic and endothermic reactions? Could you accurately sketch these diagrams, including all of the components?

Practice questions

1. Barium hydroxide, $Ba(OH)_2$, reacts with ammonium chloride, NH_4Cl:

$$Ba(OH)_2(s) + 2NH_4Cl(s) \rightarrow BaCl_2(aq) + 2NH_3(g) + 2H_2O(l) \quad \Delta H = +164\ \text{kJ mol}^{-1}$$

Which of the following is correct for this reaction?

	Temperature	Enthalpy	Stability
A	increases	products have lower enthalpy than the reactants	products are less stable than the reactants
B	decreases	products have lower enthalpy than the reactants	products are more stable than the reactants
C	decreases	products have higher enthalpy than the reactants	products are less stable than the reactants
D	increases	products have higher enthalpy than the reactants	products are more stable than the reactants

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Standard temperature and pressure (STP) conditions are denoted by the symbol ⊖. STP is a temperature of 273.15 K and a pressure of 100 kPa. Standard ambient temperature and pressure (SATP) refer to more practical reaction conditions of 298.15 K and 100 kPa. STP and SATP conditions are given in the section 2 of the data booklet.

Substance	Specific heat capacity / kJ kg^{-1} K^{-1}
water	4.18
ethanol	2.44
copper	0.385

▲ Table 1 The specific heat capacities of water, ethanol and copper

Standard enthalpy change, $\Delta H^{\ominus}$ *(Reactivity 1.1.4)*

The **standard enthalpy change for a reaction**, $\Delta H^{\ominus}$, refers to the heat transferred at constant pressure under standard conditions and states. It can be determined from the change in temperature of a pure substance. The units of $\Delta H^{\ominus}$ are kJ mol^{-1}.

To calculate $\Delta H^{\ominus}$ for a reaction, you therefore need to find the change in heat. When calculating the amount of heat lost or gained by a pure substance such as water, you need to know the **specific heat capacity**, c, of that substance.

The specific heat capacity of a pure substance is defined as the amount of heat needed to raise the temperature of 1 kg of that substance by 1 °C or 1 K. For example, the specific heat capacity of ethanol is 2.44 kJ kg^{-1} K^{-1}, so it takes 2.44 kJ to raise the temperature of 1 kg of ethanol by 1 K. The lower the specific heat capacity of a given substance, the higher the rise in temperature when the same amount of heat is transferred to the sample.

Specific heat capacity is an **intensive property** that does not vary in magnitude with the size of the system being described. For example, a 10 cm^3 sample of copper has the same specific heat capacity as a 1 ton block of copper.

When you heat up a pure substance, the rise in temperature is dependent on:

- its identity
- its mass
- the amount of heat supplied.

Practice questions

2. Using table 1, calculate how much energy is required to raise the temperature of the following by 1 K.
   a. 1 kg of water
   b. 1000 kg of copper
3. When equal masses of two different substances, X and Y, absorb the same amount of energy, their temperatures rise by 5 °C and 10 °C, respectively. Which of the following is correct?
   a. The specific heat capacity of X is twice that of Y.
   b. The specific heat capacity of X is half that of Y.
   c. The specific heat capacity of X is one fifth that of Y.
   d. The specific heat capacity of X is the same as Y.
4. Using table 1, state which of the following statements is correct.
   a. More heat is needed to increase the temperature of 50 g of water by 50 °C than 50 g of ethanol by 50 °C.
   b. If the same heat is supplied to equal masses of ethanol and water, the temperature of the water increases more.
   c. If equal masses of water at 20 °C and ethanol at 50 °C are mixed together, the final temperature is 35 °C.
   d. If equal masses of water and ethanol at 50 °C cool down to room temperature, ethanol liberates more heat.

Specific heat capacity is used to calculate the heat, Q, of a system using the relationship:

$$Q = mc\Delta T$$

where m is mass of the reaction mixture in kg and ΔT is the change in temperature of the surroundings in K.

Heat, Q, is related to enthalpy change, ΔH, by the following equation:

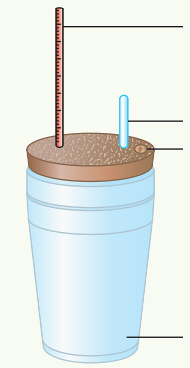
$$\Delta H = -\frac{Q}{n}$$

where n is the number of moles of the **limiting reactant**. In a reaction, the limiting reactant is the reacting substance with the least stoichiometric amount present, which therefore limits how much product can be formed. In contrast, the other reacting substances are said to be in **excess**.

Measurement

Performing reactions in a polystyrene coffee cup to measure the enthalpy change is a convenient experimental procedure. This method introduces systematic errors that can be analysed and the effect of their directionality assessed.

Systematic errors are a consequence of the experimental procedure. Their effect on empirical data is constant and always in the same direction. With the coffee-cup calorimeter, the measured change in enthalpy for a reaction will always be lower in magnitude than the actual value, as some heat will be transferred between the contents and the surroundings in every experiment.



▲ Figure 8 A coffee-cup calorimeter

Worked example 1

1. When a 1.15 g sample of anhydrous lithium chloride, LiCl, was added to 25.0 g of water in a coffee-cup calorimeter, a temperature rise of 3.80 K was recorded. Calculate the enthalpy change of dissolution for 1 mol of lithium chloride. Assume that the heat capacity of lithium chloride itself is negligible.
2. 180.0 J of heat is transferred to a 100.0 g sample of iron, resulting in a temperature rise from 22.0 °C to 26.0 °C. Calculate the specific heat capacity of iron.

Solution

1. $Q = mc\Delta T$

   $= 0.025 \text{ kg} \times 4.18 \text{ kJ kg}^{-1} \text{ K}^{-1} \times 3.80 \text{ K}$

   $= 0.397 \text{ kJ}$

   Now you need to convert to energy gained for 1 mol of LiCl.

   $n(\text{LiCl}) = \dfrac{1.15 \text{ g}}{42.39 \text{ g mol}^{-1}} = 0.0271 \text{ mol}$

   $\Delta H = -\dfrac{Q}{n} = \dfrac{-0.397 \text{ kJ}}{0.0271 \text{ mol}} = -14.6 \text{ kJ mol}^{-1}$

2. First, determine the change in temperature, ΔT:

   $\Delta T = (299 - 295) \text{ K} = 4 \text{ K}.$

   Substitute the values into $Q = mc\Delta T$:

   $0.180 \text{ kJ} = 0.100 \text{ kg} \times c \times 4 \text{ K}$

   Make c the subject of the equation and solve:

   $c = \dfrac{0.180 \text{ kJ}}{0.100 \text{ kg} \times 4 \text{ K}} = 0.450 \text{ kJ kg}^{-1} \text{ K}^{-1}$

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Practice questions

5. Calculate the energy absorbed by water when the temperature of 30 g of water is raised by 30 °C. The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.
6. 0.675 kJ of heat is transferred to 125 g of copper metal. Copper metal has a specific heat capacity of 385 J kg^{-1} K^{-1}. Calculate the change in temperature of the copper metal.

Investigation to find the enthalpy change for a reaction

In this skills task, we will look at the method used to calculate the enthalpy change for the exothermic metal displacement reaction between zinc and copper(II) sulfate:

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

Relevant skills

- Tool 1: Measuring variables
- Tool 1: Applying techniques
- Tool 2: Applying technology to process data
- Tool 3: Processing uncertainties
- Tool 3: Graphing
- Inquiry 1: Controlling variables
- Inquiry 2: Processing data
- Inquiry 3: Evaluating

Materials

- electronic balance
- coffee-cup calorimeter
- measuring cylinder
- thermometer or temperature probe
- 1.0 mol dm^{-3} copper(II) sulfate solution
- zinc powder

Method

1. Using an electronic balance, accurately measure the mass of 25 cm^3 of 1.0 mol dm^{-3} $CuSO_4$ solution. Transfer the solution to the coffee-cup calorimeter.
2. Using a thermometer or a temperature probe, record the temperature of the solution every 30 seconds for up to three minutes, or until a constant temperature is achieved.
3. At three minutes, introduce between 1.3 g and 1.4 g of zinc powder, record the exact mass of zinc and commence stirring.
4. Continue to take temperature readings for up to five minutes after the maximum temperature has been reached.
5. Produce a temperature versus time graph to determine the change in temperature.
6. Use your value of ΔT to calculate the heat released, Q, and the enthalpy change for the reaction, ΔH.

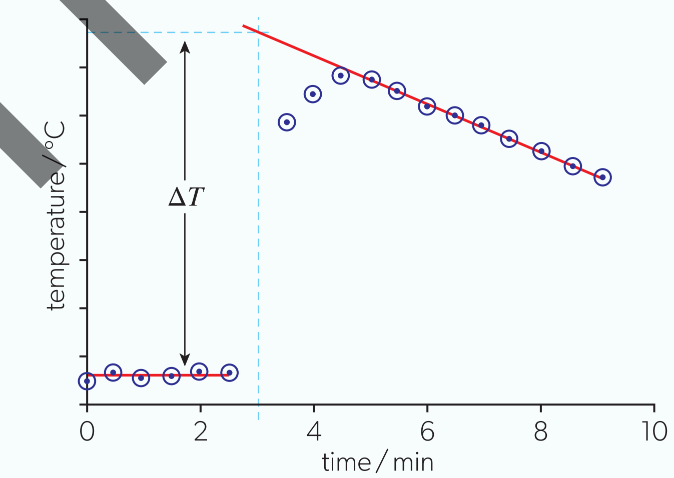
Assumptions and errors

A number of assumptions are made when using this method:

- The heat released from the reaction is completely transferred to the water.
- The coffee cup acts as an insulator against heat loss to the surroundings. However, the coffee cup also has a heat capacity and heat is transferred to it from the water. It would be difficult to quantify the heat capacity of a polystyrene cup, so it is assumed to be zero.
- The maximum temperature reached is an accurate representation of the heat evolved during the reaction.
- The specific heat capacity of an aqueous solution is the same as that of water.

Loss of heat from the system to the surroundings is the main source of error in this experiment and one that is difficult to quantify. The change in temperature, ΔT, calculated from a graph will include a systematic or directional error. This loss of heat means that the maximum temperature recorded will be lower than the theoretical value, making the calculated value of Q lower than the actual value. The effect of errors on the result of subsequent calculations is important in considering improvements in experimental procedures.

An accepted method of calculating the maximum temperature to compensate for systematic errors in data is to look at the cooling section of the curve after the reaction is complete, and extrapolate this back to the moment when zinc is introduced at 3 minutes, as shown in figure 9. A more accurate value for ΔT can then be determined.



▲ Figure 9 Example of a temperature vs time graph for a calorimetry experiment

Worked example 2

A coffee-cup calorimeter was used to measure the temperature change for the reaction between zinc powder and a 1.0 mol dm^{-3} solution of copper(II) sulfate. The following results were recorded:

Mass of copper(II) sulfate solution / g	28.8
Mass of zinc / g	1.37
ΔT / °C	39.0

Determine the amount of heat released and the enthalpy change for this reaction.

Solution

First, use $Q = mc\Delta T$ to determine the amount of heat released:

$$Q = 0.0288 \text{ kg} \times 4.18 \text{ kJ kg}^{-1} \text{ K}^{-1} \times 39.0 \text{ K}$$

$$= 4.69 \text{ kJ}$$

Then, determine the limiting reactant for the reaction.

$$\text{Number of moles of zinc, } n(\text{Zn}) = \frac{m}{M_r}$$

$$= \frac{1.37 \text{ g}}{65.38 \text{ g mol}^{-1}}$$

$$= 0.0210 \text{ mol}$$

$$\text{Number of moles of copper(II) sulfate, } n(CuSO_4) = c \times v$$

$$= 1.00 \text{ mol dm}^{-3} \times 0.0288 \text{ dm}^3$$

$$= 0.0288 \text{ mol}$$

Zinc is present in a smaller amount, so it is the limiting reactant. You can calculate the enthalpy change of reaction from $\Delta H = -\frac{Q}{n}$:

$$\Delta H = -\frac{4.69 \text{ kJ}}{0.0210 \text{ mol}} = -223 \text{ kJ mol}^{-1}$$

TOK

In theory of knowledge, there are 12 concepts in focus. These are: evidence, certainty, truth, interpretation, power, justification, explanation, objectivity, perspective, culture, values and responsibility. Scientists perform experiments and process the raw data to enable us to draw conclusions. We compare experimental and theoretical values. What concepts do we utilize when justifying our conclusions? How do we use evidence? Are our judgments subjective or objective? When analysing and appraising experimental limitations, how do assumptions have an impact on our perceptions?

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Combustion of primary alcohols

You can determine the enthalpy change of combustion of common alcohols in a school laboratory. After repeating the experiment several times with a homologous series of alcohols, you can subsequently analyse this data and identify patterns.

Relevant skills

- Tool 1: Recognise and address the relevant safety, ethical or environmental issues in an investigation
- Tool 1: Measuring temperature and mass
- Tool 1: Calorimetry
- Inquiry 1: Appreciate when and how to insulate against heat loss or gain
- Inquiry 2: Identify and record relevant qualitative observations and sufficient relevant quantitative data

Materials

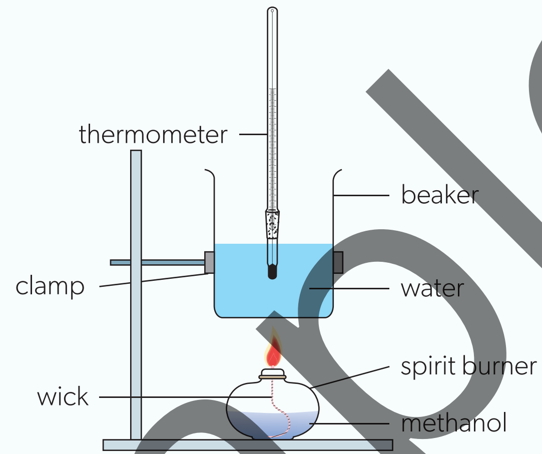
- five spirit burners, each containing one of the following alcohols: methanol, ethanol, propan-1-ol, butan-1-ol and pentan-1-ol
- electronic balance
- beaker or metal calorimeter
- tripod
- temperature probe or thermometer

Safety

Alcohols should be handled and disposed of with care because they are generally flammable, hazardous, and volatile.

Instructions

1. Using suitable sources, identify the hazards and complete a risk assessment for this experiment. In your risk assessment, you should:
    - identify the **hazards**
    - assess the level of **risk**
    - determine relevant **control measures**
    - identify suitable **disposal methods** aligned with your school's health and safety policies.
2. Determine the initial mass of the spirit burners using an electronic balance.
3. Accurately determine the mass of 30 cm^3 of water contained in a 250 cm^3 beaker or metal calorimeter.
4. Using either a temperature probe or a thermometer, determine and record the initial temperature of the water.
5. Ignite a spirit burner under the beaker or calorimeter and allow the alcohol to burn to heat the water. The period over which it burns can be set in one of two different ways:
    a. allow each alcohol to burn until a temperature change of 30 °C is reached
    b. allow each alcohol to burn for a period of two minutes.
6. Determine the final mass of each spirit burner immediately after the flame is extinguished. Take extra care because the burner will be hot.
7. Use your values of ΔT of the water and Δm of the burner to calculate the heat released, Q, and the enthalpy change of combustion, ΔH, for each alcohol.



▲ Figure 10 A typical arrangement of experimental apparatus for an enthalpy of combustion experiment

Assumptions and errors

- Heat loss to the environment is negligible.
- All the alcohols are pure, and they undergo complete combustion.

ATL Research skills

Cite your sources fully, according to your school's citing and referencing system.

Worked example 3

A metal calorimeter was used to measure the temperature change for the combustion of methanol. The following results were recorded:

Mass of water / g	31.2
Change in mass of methanol / g	0.348
ΔT / °C	30.0

Determine the heat released and the enthalpy change of combustion for methanol.

Solution

First, use $Q = mc\Delta T$ to determine the amount of heat released:

$$Q = 0.0312 \text{ kg} \times 4.18 \text{ kJ kg}^{-1} \text{ K}^{-1} \times 30.0 \text{ K}$$
$$= 3.91 \text{ kJ}$$

Methanol reacts with oxygen in a combustion reaction. Methanol is the limiting reactant for this reaction because oxygen is present in excess in the air.

$$\text{Number of moles of methanol, } n(CH_3OH) = \frac{m}{M_r}$$
$$= \frac{0.348 \text{ g}}{32.05 \text{ g mol}^{-1}}$$
$$= 0.0109 \text{ mol}$$

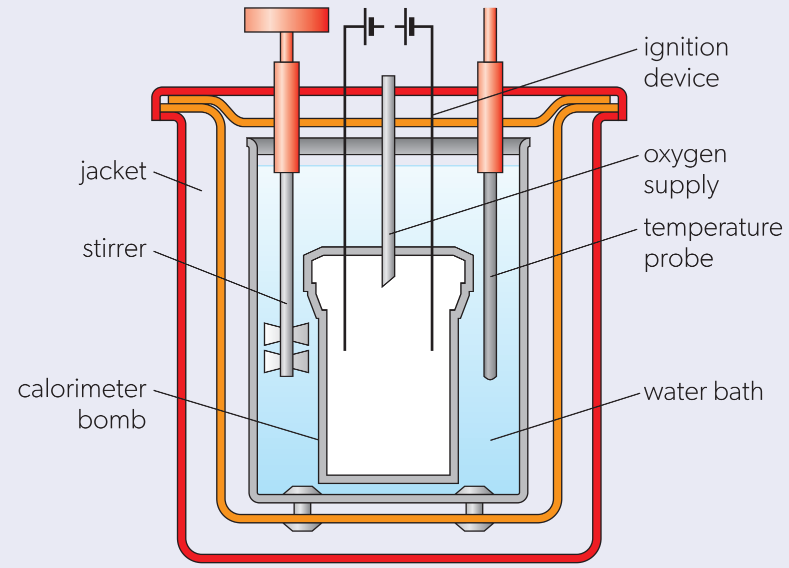
You can calculate the enthalpy change of reaction from $\Delta H = -\frac{Q}{n}$:

$$\Delta H = -\frac{3.91 \text{ kJ}}{0.0109 \text{ mol}} = -359 \text{ kJ mol}^{-1}$$

Thermochemistry experiments provide a useful set of raw data, and involve experimental procedures that can be evaluated for random and systematic errors. The identification of the systematic errors and examination of their directionality are essential aspects of the analysis of experimental results. Calorimetry experiments typically give a smaller change in temperature than is predicted from theoretical values. This is the result of heat loss from the system, which is difficult to measure. Scientists usually make the assumption that the heat lost to the environment is negligible. TOK helps us to understand our judgments of discrepancies between experimental and theoretical values.

ATL Thinking skills

Calorimetry experiments conducted in research laboratories utilize the same principles as the calorimetry experiments described in this chapter. The instrument used is called a **bomb calorimeter** (figure 11). A sample is burned inside a chamber (called a "bomb"), and the resulting temperature change of the surrounding water is measured.



▲ Figure 11 Diagram of a bomb calorimeter used in research laboratories to determine the energy content in food

1. Study the diagram carefully and list all the features that are labelled.
2. Deduce the purpose of each feature.
3. Consider why the measurements obtained with a bomb calorimeter are highly accurate and precise.
4. What properties of water make it suitable for calorimetry experiments?

Thermometric titration

The neutralization reaction between an acid and a base is exothermic. In this skills task, you will determine the unknown concentration of hydrochloric acid by measuring the change in temperature while sodium hydroxide is added to the acid. The temperature will reach a maximum when the acid and base are mixed together in stoichiometric amounts.

Relevant skills

- Tool 1: Calorimetry and acid–base titration
- Tool 2: Use sensors
- Tool 3: Calculate and interpret percentage error
- Tool 3: Understand the significance of uncertainties in raw and processed data
- Tool 3: Propagate uncertainties and state them to an appropriate level of precision
- Tool 3: Extrapolate graphs
- Tool 3: Systematic and random error
- Inquiry 1: Appreciate when and how to insulate against heat loss or gain
- Inquiry 3: Identify and discuss sources of systematic and random error

Safety

- Wear eye protection.
- Sodium hydroxide solution is corrosive.
- Hydrochloric acid is corrosive.

Materials

- two 250 cm^3 polystyrene cups
- thermometer or temperature probe
- graduated pipette and filler
- burette
- ~50.0 cm^3 sodium hydroxide solution of known concentration
- 30.0 cm^3 hydrochloric acid of unknown concentration.

Method

1. Read through the safety, materials and method. Use this information, and relevant safety data, to complete a risk assessment for this practical work and show it to your teacher.
2. Review the titration, percentage error and uncertainties sections in the *Skills* chapter.
3. Rinse and fill the burette with sodium hydroxide solution. Record its concentration.
4. Add 25 cm^3 of acid solution to the cup and place it under the burette. Nest it inside a second cup, for additional thermal insulation. For safety, these cups should be placed inside a beaker to avoid tipping over.
5. Position the temperature probe in the acid and record the initial temperature of the acid in the cup.
6. Add a small volume (~5 cm^3) of sodium hydroxide solution to the acid, stirring gently. Record the highest temperature reached with this addition.
7. Continue adding small volumes of sodium hydroxide solution and recording the temperature until the temperature decreases over several consecutive readings.
8. Clear up as instructed by your teacher.

Questions

1. Plot a graph showing temperature vs volume of sodium hydroxide solution added.
2. Extrapolate the two sections of the graph to find the maximum temperature reached during the titration.
3. Determine the concentration of the acid, along with absolute and percentage uncertainties. Make sure you state all values to an appropriate level of precision.
4. Calculate the percentage error of your experimental acid concentration.
5. Determine the enthalpy of neutralization, along with absolute and percentage uncertainties. State all values to an appropriate level of precision.
6. Calculate the percentage error of your experimental enthalpy of neutralization.
7. Comment on the relative impacts of systematic and random error on the values obtained for the acid concentration of the acid and enthalpy of neutralization.
8. Suggest and explain two improvements that could be made to this methodology.

Extension

Discuss how the identity of the acid affects the enthalpy of neutralization. Consider other strong acids such as nitric or sulfuric acid, or weak acids such as ethanoic acid. If you have time, test your ideas after discussing them with your teacher.

Measurement

Experimental enthalpy values can be assessed in terms of their accuracy and precision. Random errors in measurement lead to imprecision, whereas systematic errors cause inaccuracy. What are some of the sources of random and systematic errors in an enthalpy of neutralization experiment? To what extent are these errors quantifiable?

End of topic questions

1. Using your knowledge from the *Reactivity 1.1* topic, answer the guiding question as fully as possible:

   What can be deduced from the temperature change that accompanies chemical or physical change?

Multiple-choice questions

2. Which is correct for the following reaction?

   $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
   $\Delta H = -1049$ kJ mol^{-1}

   A Reactants are less stable than products and the reaction is endothermic.
   B Reactants are more stable than products and the reaction is endothermic.
   C Reactants are more stable than products and the reaction is exothermic.
   D Reactants are less stable than products and the reaction is exothermic.

3. Which statement is correct?

   A In an exothermic reaction, the products have more energy than the reactants.
   B In an exothermic reversible reaction, the activation energy of the forward reaction is greater than that of the reverse reaction.
   C In an endothermic reaction, the products are more stable than the reactants.
   D In an endothermic reversible reaction, the activation energy of the forward reaction is greater than that of the reverse reaction.

4. Which statement is correct for this reaction?

   $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
   $\Delta H = -26.6$ kJ mol^{-1}

   A 13.3 kJ are released for every mole of Fe produced.
   B 26.6 kJ are absorbed for every mole of Fe produced.
   C 53.2 kJ are released for every mole of Fe produced.
   D 26.6 kJ are released for every mole of Fe produced.

5. In which reaction do the reactants have a lower energy than the products?

   A $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
   B $HBr(g) \rightarrow H(g) + Br(g)$
   C $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$
   D $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

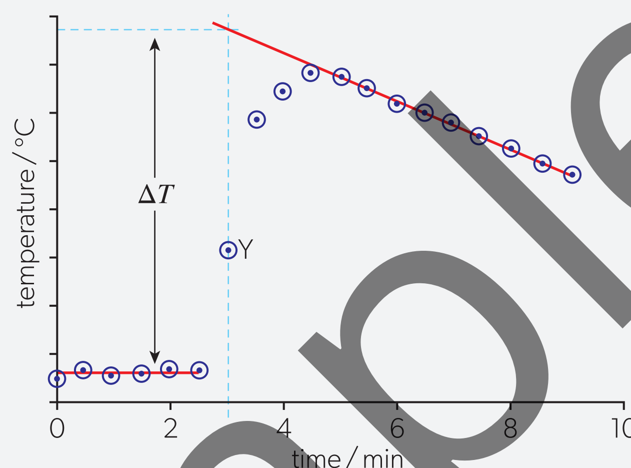
Extended-response questions

6. Nitrogen dioxide and carbon monoxide react according to the following equation:

   $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$
   $\Delta H = -226$ kJ mol^{-1}

   a. Calculate the enthalpy change for the reverse reaction.
   b. State the equation for the reaction of NO_2 in the atmosphere to produce acid deposition.

7. Powdered zinc was reacted with 25.00 cm^3 of 1.000 mol dm^{-3} copper(II) sulfate solution in an insulated beaker. Temperature was plotted against time.



   a. Estimate the time at which the powdered zinc was placed in the beaker.
   b. State what point Y on the graph represents.

   The maximum temperature used to calculate the enthalpy of reaction was chosen at a point on the extrapolated (red) line.

   c. State the maximum temperature that should be used, and outline **one** assumption made in choosing this temperature on the extrapolated line.
   d. To determine the enthalpy of reaction, the experiment was carried out five times. The same volume and concentration of copper(II) sulfate was used but the mass of zinc was different each time. Suggest, with a reason, if zinc or copper(II) sulfate should be in excess for each trial.

   The formula $q = mc\Delta T$ was used to calculate the amount of energy released. The values used in the calculation were $m = 25.00$ g and $c = 4.18$ J g^{-1} K^{-1}.

   e. State an assumption made when using these values for m and c.
   f. Predict, giving a reason, how the final enthalpy of reaction calculated from this experiment would compare with the theoretical value.

8. A potato chip (crisp) was ignited, and the flame was used to heat a test tube containing water.

Mass of water / g	7.8
Mass of chip / g	1.2
Initial temperature / °C	21.3
Final temperature / °C	22.6

   a. Calculate the heat required, in kJ, to raise the temperature of the water, using data in the table above and from section 2 of the data booklet.
   b. Determine the enthalpy of combustion of the potato chip, in kJ g^{-1}.