

> Chapter 12

Measuring enthalpy change

CHAPTER OUTLINE

In this chapter you will:

- understand the difference between heat and temperature
- understand the difference between exothermic and endothermic reactions
- understand what is meant by the term *stability*
- sketch and interpret potential energy profiles
- understand the term *standard enthalpy change* for a reaction
- calculate standard enthalpy changes from experimental data.

KEY TERMS

Make sure you understand the following key terms before you do the exercises.

heat: the energy that flows from something at a higher temperature to something at a lower temperature because of the temperature difference between them

exothermic reaction: a chemical reaction that results in the release of heat to the surroundings – the reaction vessel gets hotter; ΔH for an exothermic reaction is negative

endothermic reaction: a chemical reaction in which heat is taken in from the surroundings – the reaction vessel gets colder; ΔH for an endothermic reaction is positive

enthalpy change (ΔH): the heat energy exchanged with the surroundings at constant pressure

system/surroundings: system refers to the chemicals themselves, whereas the surroundings refers to the solvent, the air and the apparatus – all that surrounds the chemicals

internal energy (sometimes called chemical energy): the name given to the total amount of energy (kinetic and potential) in a sample of a substance

potential energy profile diagram: a diagram showing the change in the potential energy (y-axis) of a system as a reaction proceeds (x-axis is the reaction coordinate)

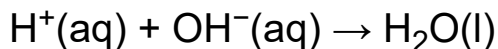
calorimetry: experimental determination of the heat given out/taken in during chemical reactions/physical processes

specific heat capacity: the energy required to raise the temperature of 1 g of substance by 1 K (1 °C). It can also be defined as the energy to raise the temperature of 1 kg of substance by 1 K. Specific heat

capacity has units of $\text{J g}^{-1} \text{K}^{-1}$ or $\text{J g}^{-1} \text{°C}^{-1}$. Units that are also encountered are $\text{kJ kg}^{-1} \text{K}^{-1}$ or $\text{J kg}^{-1} \text{K}^{-1}$

standard enthalpy change of neutralisation (ΔH_n):

the enthalpy change when one mole of H_2O molecules is formed when an acid (H^+) reacts with an alkali (OH^-) under standard conditions, i.e



the enthalpy change of neutralisation is always exothermic

KEY EQUATIONS

$$Q = mc\Delta T$$

where

Q = heat energy, J

m = mass, g

c = specific heat capacity, $\text{J g}^{-1}\text{K}^{-1}$

ΔT = temperature change, K or $^{\circ}\text{C}$

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

where

ΔH = enthalpy change, J mol^{-1}

Q = heat energy, J

n = number of moles, mol

Exercise 12.1 Heat and temperature

In this exercise, you will explore the difference between **heat** and temperature. Temperature was first introduced in [Chapter 1](#).

1 What is temperature a measure of?

Average Kinetic energy

2 What is the SI unit of temperature?

Kelvin (K)

3 What is the SI unit of heat energy?

Joules (J)

4 What is the usual unit of heat energy used in chemistry?

Kilojoules (kJ)

5 Which has a higher amount of chemical energy:

a a block of silver with a mass of 50 g at a temperature of 300 K or a block of copper with a mass of 50 g at a temperature of 300 K?

Same

b 150 g of water at a temperature of 350 K or a block of copper with a mass of 150 g at a temperature of 300 K?

Water, higher temp.

c a block of nickel with a mass of 100 g at a temperature of 300 K or a block of aluminium with a mass of 50 g at a temperature of 300 K?

Aluminium, energy shared between less matter.

Exercise 12.2 Exothermic and endothermic reactions

Some reactions, like combustion and respiration, are easily recognised as being **exothermic**. Some reactions are easily recognised as being **endothermic**, such as thermal decomposition and photosynthesis. In this exercise, these terms are explored and ideas about the exchange of heat energy (the **enthalpy change**) between the **system** and the surroundings, are introduced. In this exercise, you will also practise drawing and labelling **potential energy profile diagrams**, which are a useful way of representing energy changes during a chemical reaction.

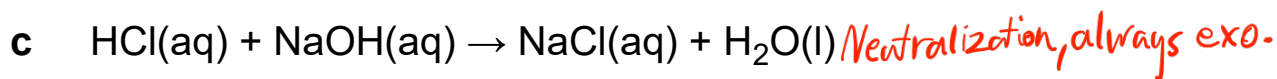
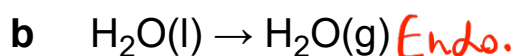
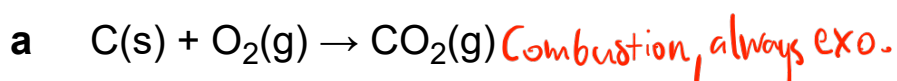
- 1 During a chemical reaction, there is an interchange between the **internal energy** of the substances and heat energy. Identify which statement describes an endothermic reaction and which describes an exothermic reaction.
 - a During the reaction, chemical energy is converted into heat energy.
Exo
 - b During the reaction, heat energy is converted into chemical energy.
Endo
- 2 H is the symbol used for enthalpy. What symbol is used to represent enthalpy *change*?
 ΔH
- 3 Explain why the sign for the enthalpy change of an exothermic reaction is negative, even though the temperature increases.

Energy is released from the substance, so change is negative, and it goes into the surroundings, where the temperature is being measured, so temperature increases.

4 Which of the following statements are **true** about exothermic reactions?

- ☒ a The products are more stable than the reactants.
- ☒ b Total energy is conserved.
- ☒ c The temperature of the surroundings will increase.
- ☒ d There will be a flow of energy from the system to the surroundings.
- ☒ e The sign of the enthalpy change will be positive.

5 Suggest whether the following reactions are exothermic or endothermic:



6 Sort the following statements into those that are true for exothermic reactions and those that are true for endothermic reactions.

Exo a Temperature increases during the reaction.

Endo b Temperature decreases during the reaction.

Endo c $\Delta H > 0$

Exo d $\Delta H < 0$

Exo e The products are more stable than the reactants.

Endo f The reactants are more stable than the products.

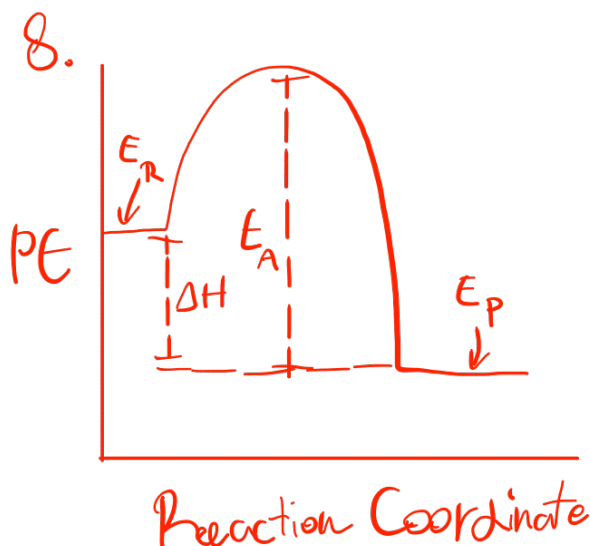
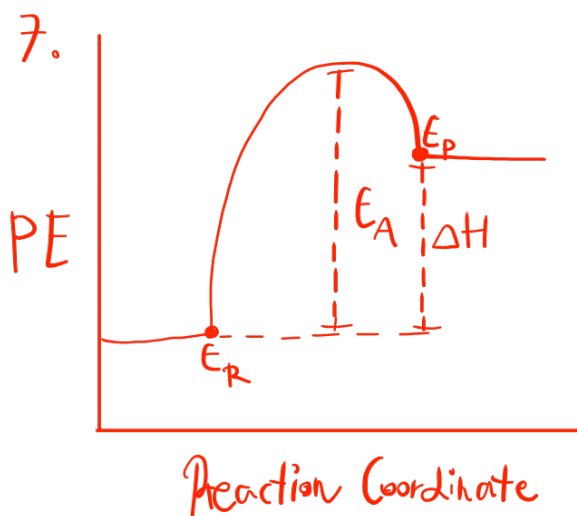
Both **g** Total energy is conserved.

7 Sketch the potential energy profile (energy level diagram) for an endothermic reaction. Include the following in your diagram:

- a** labels on both axes
- b** the energy of the reactants and the products
- c** the overall enthalpy change for the reaction
- d** the activation energy of the reaction.

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Exercise 12.3 Enthalpy changes and standard conditions

Enthalpy changes have different values, depending on the conditions under which they are measured. Standard conditions are used to make these energy changes more easily compared. In this syllabus, you do not need to calculate the difference in enthalpy changes under different conditions; you only need to be aware that there are differences. In this exercise, you will look at what is meant by *standard conditions*.

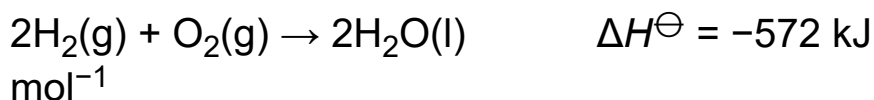
1 What are *standard conditions*?

Set conditions for ease of calculations in reactions.

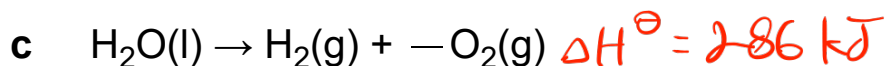
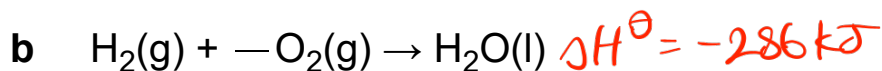
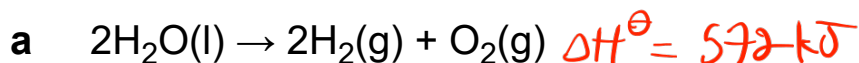
2 What symbol is used to denote that a value has been measured or calculated under standard conditions?

⊖ symbol, called the Plimsoll symbol.

3 Given the standard enthalpy change for the reaction



calculate the standard enthalpy changes for the following reactions:



Exercise 12.4 Measuring enthalpy changes

In this exercise, you will practise calculating enthalpy changes from experimental **calorimetry** data. The exercise begins with practise using the expression $Q = mc\Delta T$. This expression is used to calculate enthalpy of combustion, and then it is applied to enthalpy changes in solution, such as neutralisation reactions. In all of the questions, you can assume that the density of any solution is 1.00 g cm^{-3} and the **specific heat capacity** of a solution is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

- 1 A block of copper of mass 20.0 g was heated using an electric heating element. 384 J were required to raise the temperature of the block by 50.0°C . Calculate the specific heat capacity of copper. $384 = 20.0 \cdot C \cdot 50; C = 0.384 \text{ J kg}^{-1} \text{ K}^{-1}$
- 2 Calculate the temperature change for the reactions below.

specific heat capacity of water = $4.18 \text{ J g}^{-1} \text{ K}^{-1}$, of Ag = $0.240 \text{ J g}^{-1} \text{ K}^{-1}$, of Al = $0.900 \text{ J g}^{-1} \text{ K}^{-1}$

TIP

The formula $Q = mc\Delta T$ should be used here.

TIP

Pay attention to the units.

$$\Delta T = \frac{Q}{mc}$$

- a 2000 J of heat energy is transferred to 100 g of water $\Delta T = \frac{2000}{100 \cdot 4.18} = 4.78$
- b 2000 J of heat energy is transferred to 100 g of silver $\Delta T = \frac{2000}{100 \cdot 0.240} = 83.3$
- c 7.5 kJ of heat energy is transferred to 100 g aluminium. $\Delta T = \frac{7500}{100 \cdot 0.9} = 83.3$
- 3 The apparatus in Figure 12.1 was used to determine the enthalpy change of combustion of a liquid fuel with the formula $C_5H_{11}OH$.

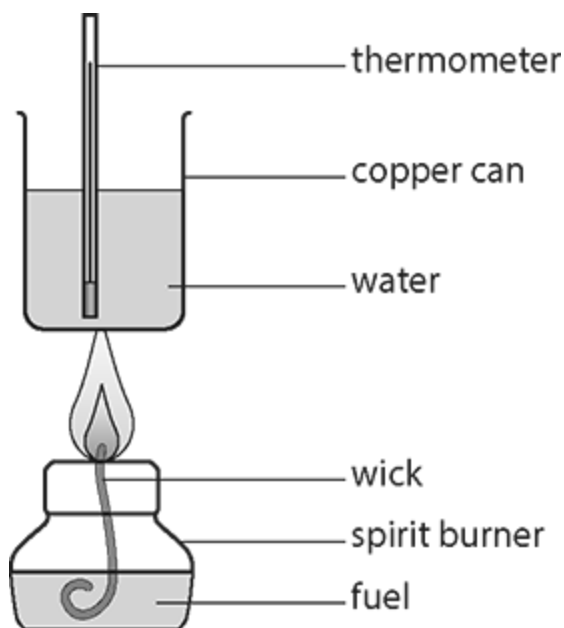


Figure 12.1: Apparatus to determine the enthalpy of combustion of a liquid fuel.

The following results were obtained:

mass of empty copper can = 212.3 g

mass of can and water before heating = 326.5 g

temperature of water before heating = 21.5 °C

mass of spirit burner before lighting = 104.0 g

mass of spirit burner at the end of experiment = 99.3 g

temperature of water after heating = 77.5 °C

specific heat capacity of the water = 4.18 J g⁻¹ K⁻¹

- a** Calculate the enthalpy change of combustion of the fuel in kJ mol⁻¹.
 - b** Explain why copper is a good choice of material for the can.
 - c** Suggest two possible sources of experimental error in the experiment.
 - d** Suggest ways to reduce the experimental errors suggested in part **c**.
 - e** Other than experimental errors, suggest two reasons why the value for the enthalpy change of combustion of this fuel differs from the values quoted in data books.
- 4** 20.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid was added to 20.0 cm³ of 1.00 mol dm⁻³ sodium hydroxide solution. The temperature rose by 6.8 °C. Calculate the amount of energy released by this reaction.

TIP

In question 5, note that one substance is in excess, so you will need to calculate the number of moles of water formed.

- 5 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid was added to 40.0 cm³ of 1.00 mol dm⁻³ sodium hydroxide solution. The temperature rose by 6.1 °C. Calculate the **enthalpy change of neutralisation** for this reaction.

TIP

Always include the equation for the reaction, as you need to check that the ratio of the number of moles of each substance matches their ratio in the equation to find out if one is in excess.

- 6 30.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid was added to 20.0 cm³ of 1.00 mol dm⁻³ calcium hydroxide solution. The temperature rose by 8.2 °C. Calculate the enthalpy change of neutralisation.

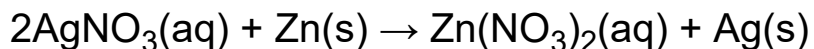
TIP

The mass of the solution formed in the reaction in question 7 could be debated; is it 25.0 g or 27.0 g?

Most past IB papers have allowed either answer, although the answer that excludes the mass of the solid is usually the primary answer.

- 7 Calculate the temperature change when 2.00 g of sodium hydrogencarbonate is dissolved in 25.0 cm³ of water if the standard enthalpy change of solution for sodium hydrogencarbonate is +18 kJ mol⁻¹.

- 8** 50.0 cm³ of 0.200 mol dm⁻³ silver nitrate solution was put in a calorimeter and 0.400 g of zinc powder added. The equation for the reaction is as follows:



The temperature of the solution rose by 8.6 °C. Calculate the enthalpy change for the reaction per mole of zinc that reacts. Ignore the heat capacity of the metals.

- 9** In question 8, you are told to ignore the heat capacity of the metals. What are the exact contents of the calorimeter at the end of this experiment, and hence, why are the specific heat capacities of the metals ignored?
- 10** A significant source of error in calorimetry experiments is due to heat losses, particularly if the reaction occurs slowly.
- a** Classify heat loss as a systematic or random error.
 - b** How can calorimetry experiments be modified to estimate a more accurate and quantifiable temperature change?

EXAM-STYLE QUESTIONS

- 1 Deduce the heat energy supplied if 50 g of water with a specific heat capacity of $4.2 \text{ J g}^{-1} \text{ K}^{-1}$ increases in temperature by 10°C .

A 210 J
B 420 J
C 2.1 kJ
D 4200 J

- 2 What will be the temperature change if 100 g of copper with a specific heat capacity of $0.385 \text{ J g}^{-1} \text{ K}^{-1}$ is supplied with 770 J of heat energy?

A 50 K
B 20 K
C 10 K
D 5 K

- 3 Which statement is correct?

A Temperature is a measure of the amount of heat energy something has.
B Heat energy is a measure of the average kinetic energy of the particles.
C Temperature is a measure of the average kinetic energy of the particles.

D Heat and temperature have the same meaning.

- 4** In an experiment, 0.50 g of ethanol was used to raise the temperature of 100 g of water by 35 °C. Which calculation shows how to find the correct enthalpy of combustion of ethanol?

specific heat capacity of water = 4.18 J g⁻¹ K⁻¹

specific heat capacity of ethanol = 2.46 J g⁻¹ K⁻¹

molar mass of ethanol = 46.08 g mol⁻¹

A $\Delta H = \frac{100.50 \times 4.18 \times 35 \times 0.50}{46.08}$

B $\Delta H = - \frac{46.08 \times 4.18 \times 35}{100}$

C $\Delta H = \frac{100 \times (4.18 + 2.46) \times 35}{46.08}$

D $\Delta H = - \frac{100 \times 4.18 \times 35 \times 46.08}{0.50}$

- 5** In an experiment, 50 cm³ of a solution of HCl with a concentration of 1 mol dm⁻³ was reacted with 50 cm³ of a solution of NaOH with a concentration of 1 mol dm⁻³. The temperature of the solution increased by 12 °C. Calculate the temperature rise if 50 cm³ of the same HCl solution was reacted with

50 cm³ of a solution of NaOH with a concentration of 0.5 mol dm⁻³.

- A 6 °C
- B 12 °C
- C 24 °C
- D 3 °C

6 Which of the following statements is **not** correct about endothermic reactions?

- A The surroundings lose energy.
- B The enthalpy change will be positive.
- C The enthalpy change will be negative.
- D The temperature of the surroundings decreases.

7 Which statement is correct about the energy profile diagram in Figure **12.2**?

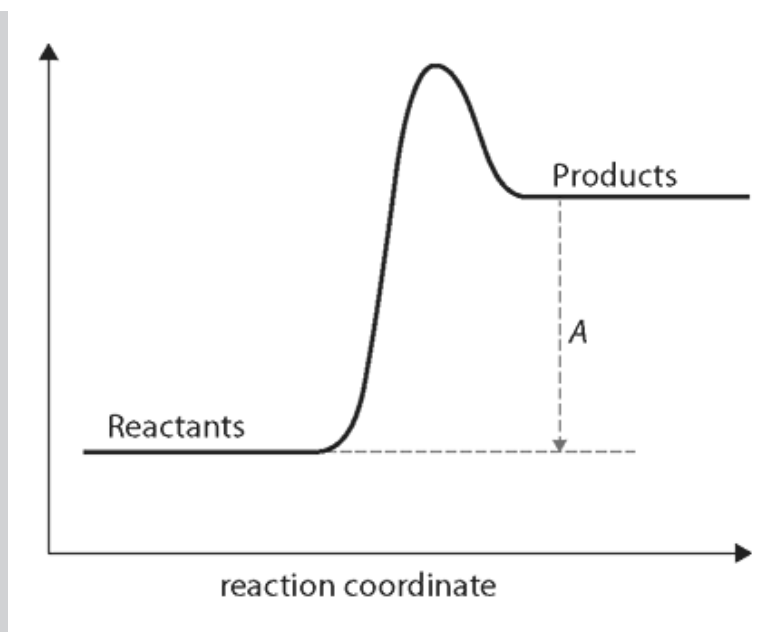
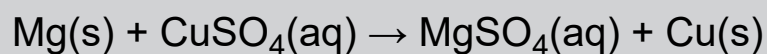


Figure 12.2: Energy profile diagram.

- A** The reaction is exothermic.
 - B** The products are more stable than the reactants.
 - C** The y-axis shows the potential energy of the system.
 - D** Arrow A is the enthalpy change for the reaction.
- 8** In an experiment to determine the enthalpy change of a reaction, 50.0 cm^3 of a solution of copper(II) sulfate with a concentration of $0.100 \text{ mol dm}^{-3}$ was placed in a polystyrene cup and the temperature recorded every 30 s. After 3 min, 2.00 g of magnesium powder was added to the cup, and the temperature

was measured every 30 s for a further 7 min.



The data shown in Figure 12.3 were obtained.

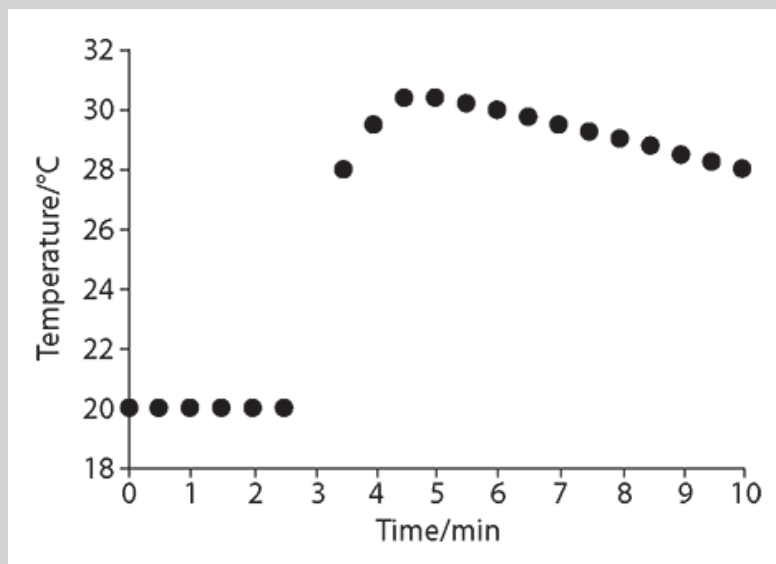


Figure 12.3: Temperature change for the reaction between CuSO_4 and Mg.

a Use the graph to find the temperature change for the reaction. [1]

b Calculate the heat energy produced by the reaction. [1]

Assume that the specific heat capacity of the mixture is the same as water:
 $4.18 \text{ J g}^{-1} \text{ K}^{-1}$

c Calculate the number of moles of magnesium and the number of moles

of copper(II) sulfate present and state which reactant was the limiting reactant.

[2]

- d** Calculate the enthalpy change for the reaction.

[2]

- e** What is the advantage of obtaining the temperature change for the reaction graphically, rather than simply measuring the initial temperature and the maximum temperature?

[1]

- f** Estimate the temperature change for the reaction if the reaction were repeated using 100 cm³ of copper(II) sulfate of concentration 0.100 mol and 2.0 g of magnesium powder. Justify your answer.

[3]