Week 8

A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

Contents 目录

A-Level Chemistry Name: _____ Score: _____

1. In an exothermic reaction:
   - ☐ heat is given out, the products are lower in energy, and ΔH is negative
   - ☐ heat is taken in and ΔH is positive
   - ☐ the products are higher in energy
   - ☐ no energy changes
2. Which statement is correct?
   - ☐ breaking bonds takes energy in (endothermic); making bonds gives energy out (exothermic)
   - ☐ breaking bonds gives energy out
   - ☐ making bonds takes energy in
   - ☐ neither breaking nor making changes energy
3. Hess's law states that the total enthalpy change for a reaction:
   - ☐ is the same no matter which route is taken from reactants to products
   - ☐ depends on the route taken
   - ☐ is always zero
   - ☐ is always positive
4. The activation energy is:
   - ☐ the least energy the particles need before they can react
   - ☐ the energy given out by the reaction
   - ☐ the difference between products and reactants
   - ☐ always zero
5. Using bond energies, ΔH of reaction equals:
   - ☐ sum of bonds broken − sum of bonds made
   - ☐ sum of bonds made − sum of bonds broken
   - ☐ bonds broken × bonds made
   - ☐ bonds made ÷ bonds broken
6. Hess's law works because energy is conserved.

   ☐ True ☐ False

7. A reaction that releases heat to the surroundings has a negative value of _____.

_____

8. Breaking bonds is endothermic and making bonds is exothermic.

☐ True ☐ False

9. The law stating that enthalpy change is independent of the route taken is _____ law.

_____

10. In an energy cycle, the direct route and the indirect route:

☐ have the same total enthalpy change

☐ have different enthalpy changes

☐ cannot be compared

☐ are always endothermic

5. Chemical energetics

Answers · 答案

A-Level Chemistry

1. **heat is given out, the products are lower in energy, and ΔH is negative**

   Exothermic reactions release heat; products sit below reactants, so ΔH is negative.

2. **breaking bonds takes energy in (endothermic); making bonds gives energy out (exothermic)**

   Energy is needed to break bonds and released when bonds form; ΔH is the difference.

3. **is the same no matter which route is taken from reactants to products**

   Because energy is conserved, the overall ΔH depends only on the start and end states, not the path.

4. **the least energy the particles need before they can react**

   Activation energy is the height of the "hill" — the minimum energy needed to start the reaction.

5. **sum of bonds broken − sum of bonds made**

   $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds made})$; energy in to break minus energy out to make.

6. **True**

   Conservation of energy means the total enthalpy change is fixed by the start and end states alone.

7. **ΔH**

   Exothermic reactions have $\Delta H < 0$.

8. **True**

   ΔH = bonds broken − bonds made.

9. **Hess's**

   Enthalpy is a state function.

10. **have the same total enthalpy change**

    Both routes start and end at the same states, so their total ΔH must be equal (e.g. $\Delta Hr = \Delta H_1 + \Delta H_2$).

A-LEVEL CHEMISTRY • EXERCISE SHEET

5.1 Enthalpy change

Name: _____ Class: _____ Date: _____ **Total: 25 marks**

KEY WORDS enthalpy change 焓变 • exothermic 放热 • bond energy 键能 • reaction pathway diagram 反应路径图 • activation energy 活化能

Objective

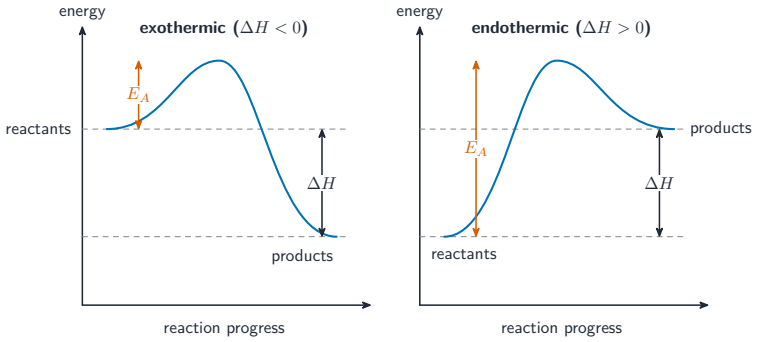
Build the skills to answer exam questions on **enthalpy change** 焓变 — **exothermic** 放热 and **endothermic** 吸热 reactions, the standard enthalpy changes, **bond energy** 键能 calculations, and $q = mc\Delta T$.

You must be able to:

- interpret reaction pathway diagrams and define the standard enthalpy changes
- calculate ΔH_r from bond energies
- calculate enthalpy changes from experiment using $q = mc\Delta T$ and $\Delta H = -mc\Delta T/n$

1 Worked examples

Exothermic and endothermic

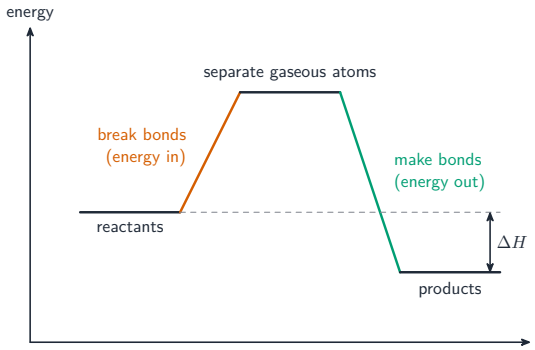


Exothermic: products lower, $\Delta H < 0$. Endothermic: products higher, $\Delta H > 0$. The hill height is the activation energy

Standard conditions: 298 K, 101 kPa ($^{\ominus}$). Know ΔH_f (formation), ΔH_c (combustion), ΔH_{neut} (neutralisation).

Bond energies

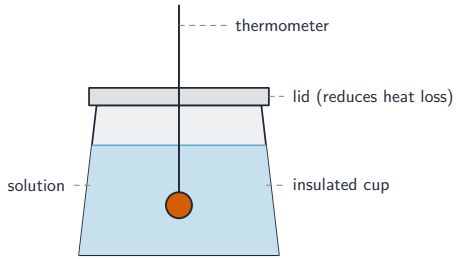
$$\Delta H_r = \sum(\text{bonds broken}) - \sum(\text{bonds made})$$



Breaking bonds takes energy in; making bonds gives energy out

Example. $H_2 + Cl_2 \rightarrow 2HCl$ (H–H 436, Cl–Cl 242, H–Cl 431): $\Delta H = (436 + 242) - (2 \times 431) = -184$ kJ mol^{-1}.

Measuring ΔH ($q = mc\Delta T$)



Measure the temperature change of a known mass of solution

$q = mc\Delta T$; $\Delta H = -q/n$. **Example.** 0.50 g methanol (M_r 32) heats 100 g water by 18 °C ($c = 4.18$): $q = 100 \times 4.18 \times 18 = 7520$ J; $n = 0.50/32 = 0.0156$; $\Delta H_c = -7520/0.0156 \approx -480$ kJ mol^{-1}.

2 Practice

2.1 State whether ΔH is positive or negative for an exothermic reaction, and what this means for the products' energy. [2]

2.2 Define the **enthalpy change of combustion**. [2]

3 Exam-style questions

3.1 In an endothermic reaction: [1]

- **A** ΔH is negative and heat is given out
- **B** ΔH is positive and heat is taken in
- **C** the products are lower in energy
- **D** no bonds are broken

3.2 Breaking a chemical bond is: [1]

- **A** exothermic (releases energy)
- **B** endothermic (takes in energy)
- **C** neither
- **D** always +431 kJ mol^{-1}

3.3 For $N_2 + 3H_2 \rightarrow 2NH_3$, use bond energies (kJ mol^{-1}: N≡N 945, H–H 436, N–H 391) to calculate ΔH_r. [3]

3.4 Burning 0.32 g of a fuel ($M_r = 32$) raises the temperature of 200 g of water by 9.0 °C ($c = 4.18$).

(a) Calculate the heat released, q, in J. [1]

(b) Calculate the enthalpy change of combustion per mole. [3]

3.5 A student burns 0.780 g of propan-1-ol, $CH_3CH_2CH_2OH$, and uses the heat released to warm 150 g of water from 21.0 °C to 48.5 °C. ($c = 4.18$ J g^{-1}K^{-1}; $M_r = 60.0$)

(a) Calculate the enthalpy change of combustion of propan-1-ol from these results. [4]

(b) The data book value is -2021 kJ mol^{-1}. Suggest **two** reasons for the difference, and state which is likely to be the larger. [3]

(c) Explain why the value is negative, and what that says about the bonds broken and the bonds made. [3]

(d) State **two** reasons why the value obtained this way is not a *standard* enthalpy change of combustion. [2]

Solutions

2.1 ΔH is negative; the products have less energy than the reactants.

2.2 the enthalpy change when one mole of a substance is burned completely in oxygen under standard conditions.

3.1 B. Endothermic = ΔH positive, heat taken in.

3.2 B. Breaking a bond requires energy (endothermic).

3.3 broken $= 945 + 3(436) = 2253$; made $= 6 \times 391 = 2346$; $\Delta H = 2253 - 2346 = -93$ kJ mol^{-1}.

3.4 (a) $q = 200 \times 4.18 \times 9.0 = 7524$ J.
(b) $n = 0.32/32 = 0.010$ mol; $\Delta H_c = -q/n = -7524/0.010 = -752$ kJ mol^{-1} (≈ -750).

3.5 (a) $q = mc\Delta T = 150 \times 4.18 \times 27.5 = 17\,240$ J $= 17.2$ kJ; $n = \dfrac{0.780}{60.0} = 0.0130$ mol;
$\Delta H_c = -\dfrac{17.2}{0.0130} = -1.33 \times 10^3$ kJ mol^{-1}.
(b) Heat is lost to the surroundings and to the apparatus instead of warming the water; combustion is incomplete, or some alcohol evaporates without burning; in an open calorimeter the **heat loss** is normally much the larger effect.
(c) Negative means energy is released to the surroundings, i.e. the reaction is exothermic; more energy is released when the bonds in CO_2 and H_2O form than is absorbed breaking the bonds in the alcohol and oxygen; so the products lie at a lower enthalpy than the reactants.
(d) The conditions are not standard — the temperature changes throughout and the pressure is not controlled at 100 kPa; and the alcohol is not necessarily burned completely to CO_2 and H_2O.

7 When 1.0 mol of ethanoic acid, CH_3COOH, in aqueous solution is neutralised by an excess of aqueous sodium hydroxide, 55 kJ mol^{-1} of energy is released.

Which statement about this reaction is correct?

A The reaction is exothermic because only bond breaking takes place.

B The reaction is exothermic because only bond forming takes place.

C The reaction is exothermic because more energy is given out in breaking bonds than is taken in to form bonds.

D The reaction is exothermic because more energy is given out in forming bonds than is taken in to break bonds.

8 In an experiment, 1.60 g of a fuel is burnt. 45.0% of the energy released is absorbed by 200 g of water. The temperature of the water rises from 18.0 °C to 66.0 °C.

What is the total energy released per gram of fuel burnt?

A 25 100 J **B** 55 700 J **C** 89 200 J **D** 143 000 J

9 When 0.47 g of a hydrocarbon is completely burnt in air, the energy released heats 200 g of water from 23.7 °C to 41.0 °C.

What is the amount of energy absorbed, in Joules, by the water?

A $0.47 \times 4.18 \times 17.3$

B $0.47 \times 4.18 \times (273 + 17.3)$

C $200 \times 4.18 \times 17.3$

D $200 \times 4.18 \times (273 + 17.3)$

10 One commercially available 'heat pad' contains iron, activated carbon and water. The 'heat pad' is activated by air. This causes the pad to get hotter.

Which statement describes the chemical reaction occurring in the 'heat pad' when it is exposed to air?

A The reaction is endothermic and iron gains electrons.

B The reaction is endothermic and iron loses electrons.

C The reaction is exothermic and iron gains electrons.

D The reaction is exothermic and iron loses electrons.

10 In an experiment to measure the enthalpy change of neutralisation of hydrochloric acid, 20 cm^3 of solution containing 0.04 mol of HC*l* is placed in a plastic cup of negligible heat capacity.

A 20 cm^3 sample of aqueous sodium hydroxide containing 0.04 mol of NaOH, at the same initial temperature, is added and the temperature rises by 15 K.

If the heat capacity per unit volume of the final solution is 4.2 J K^{-1} cm^{-3}, what is the enthalpy change of neutralisation of hydrochloric acid?

A $\frac{20 \times 4.2 \times 15}{0.04}$ J mol^{-1}

B $40 \times 4.2 \times 15 \times 0.08$ J mol^{-1}

C $\frac{40 \times 4.2 \times 15}{0.04}$ J mol^{-1}

D $\frac{20 \times 4.2 \times 15}{0.08}$ J mol^{-1}

2 When hydrated sodium thiosulfate is dissolved in water the temperature of the liquid changes. You will carry out an experiment to determine the enthalpy change, ΔH, when one mole of hydrated sodium thiosulfate dissolves in water.

FA 3 is hydrated sodium thiosulfate, $Na_2S_2O_3 \cdot 5H_2O$.

(a) Method

- Support the cup in the 250 cm^3 beaker.
- Use the 50 cm^3 measuring cylinder to transfer 30.0 cm^3 of distilled water into the cup.
- Measure the temperature of the water in the cup. Record this in the space for results.
- Weigh the container with **FA 3**. Record the mass.
- Tip all the **FA 3** into the cup.
- Stir the mixture until the minimum temperature is obtained. Record this temperature.
- Weigh the container with any residual **FA 3**. Record the mass.
- Calculate and record the mass of **FA 3** added.
- Calculate and record the temperature change.

Results

I	
II	
III	
IV	

[4]

(b) Calculations

(i) Calculate the amount, in mol, of hydrated sodium thiosulfate added.

amount of $Na_2S_2O_3 \cdot 5H_2O$ = ... mol [1]

(ii) Calculate the energy change, in J, in your experiment.

energy change = ... J [1]

(iii) Calculate the enthalpy change, ΔH, in $kJ\,mol^{-1}$, when 1.00 mol of hydrated sodium thiosulfate dissolves in water. Show your working.

ΔH = ... $kJ\,mol^{-1}$
sign *value* [1]

(c) The value calculated in **(b)(iii)** can be used to determine the enthalpy change, ΔH_r, for the following reaction.

$$Na_2S_2O_3(s) + 5H_2O(l) \rightarrow Na_2S_2O_3 \cdot 5H_2O(s)$$

Outline the method of one further experiment you would need to carry out to obtain the data necessary to calculate the value of ΔH_r.
Show how you would use your results from this experiment and **(b)(iii)** to calculate ΔH_r.

Do **not** carry out your experiment.

method .....

.....

.....

calculation .....

.....
[4]

[Total: 11]

Qualitative analysis

For each test you should record all your observations in the spaces provided.

Examples of observations include:
- colour changes seen
- the formation of any precipitate and its solubility (where appropriate) in an excess of the reagent added
- the formation of any gas and its identification (where appropriate) by a suitable test.

You should record clearly at what stage in a test an observation is made.

Where no change is observed, you should write 'no change'.

Where reagents are selected for use in a test, the name or correct formula of the element or compound must be given.

If any solution is warmed, a boiling tube must be used. If a solid is heated, a hard-glass test-tube must be used.

Rinse and reuse test-tubes and boiling tubes where possible.

No additional tests should be attempted.

1 Sodium carbonate can be manufactured using a two-step process. The first step involves making sodium hydrogencarbonate, $NaHCO_3$, which is then converted into sodium carbonate, Na_2CO_3, in the second step of the process by the following reaction.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

In this experiment you will determine the enthalpy change, ΔH_r, for this reaction. You will do this by calculating the enthalpy changes when separate samples of sodium hydrogencarbonate and sodium carbonate are added to excess hydrochloric acid, HC*l*(aq). You will then combine these values using Hess's law.

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_1$$

$$Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_2$$

FB 1 is 2.0 mol dm^{-3} hydrochloric acid, HC*l*.
FB 2 is sodium hydrogencarbonate, $NaHCO_3$.
FB 3 is sodium carbonate, Na_2CO_3.

(a) Method

Experiment 1
- Support a cup in the 250 cm^3 beaker.
- Use the 25 cm^3 measuring cylinder to transfer 20.0 cm^3 of **FB 1** into the cup.
- Weigh the container with **FB 2**. Record the mass.
- Measure the temperature of the acid in the cup. Record this temperature.
- Carefully add all of the **FB 2**, in small portions to avoid acid spray. Stir to dissolve. Record the lowest temperature.
- Reweigh the container with any residual **FB 2**. Record the mass.
- Calculate and record the mass of **FB 2** used.
- Calculate and record the decrease in temperature.

Experiment 2
- Support the second cup in the 250 cm^3 beaker.
- Use the 25 cm^3 measuring cylinder to transfer 20.0 cm^3 of **FB 1** into the second cup.
- Weigh the container with **FB 3**. Record the mass.
- Measure the temperature of the acid in the cup. Record this temperature.
- Carefully add all of the **FB 3**, in small portions to avoid acid spray. Stir to dissolve. Record the highest temperature.
- Reweigh the container with any residual **FB 3**. Record the mass.
- Calculate and record the mass of **FB 3** used.
- Calculate and record the increase in temperature.

Keep FB 1 for use in Question 3.

Results

I	
II	
III	
IV	

[4]

(b) Calculations

(i) Calculate the amount, in mol, of **FB 2** that reacts with **FB 1** and the amount, in mol, of **FB 3** that reacts with **FB 1**.

Experiment 1	**Experiment 2**
amount of **FB 2** = ... mol	amount of **FB 3** = ... mol

[2]

(ii) Calculate the energy changes, in J, for each reaction.

Experiment 1	**Experiment 2**
energy change = ... J	energy change = ... J

[1]

(iii) Calculate the enthalpy change, ΔH, in kJ mol^{-1} for each reaction.

Experiment 1	**Experiment 2**
ΔH_1 = ... kJ mol^{-1} *sign* *value*	ΔH_2 = ... kJ mol^{-1} *sign* *value*

[2]

(iv) Construct an enthalpy cycle and use Hess's law to determine the enthalpy change, ΔH_r, for the reaction shown.
Show your working.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

(If you were unable to calculate values in **(b)(iii)** then assume that ΔH_1 is +27.3 kJ mol^{-1} and that ΔH_2 is –24.9 kJ mol^{-1}. These may **not** be the correct values.)

ΔH_r = ... kJ mol^{-1}
sign *value*

[2]

[Total: 11]

A-LEVEL CHEMISTRY • EXERCISE SHEET

5.2 Hess's law

Name: _____ Class: _____ Date: _____ **Total: 20 marks**

KEY WORDS hess's law 盖斯定律 • energy cycle 能量循环 • enthalpy change 焓变 • enthalpy change of formation 生成焓变 • exothermic 放热

Objective

Build the skills to answer exam questions on **Hess's law** 盖斯定律 — constructing **energy cycles** 能量循环 and calculating enthalpy changes from formation and combustion data.

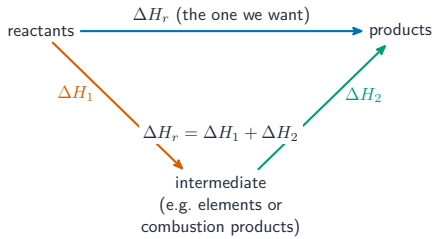
You must be able to:

- state Hess's law and construct simple energy cycles
- calculate an enthalpy change that cannot be measured directly
- use formation or combustion data (and bond energies) in a cycle

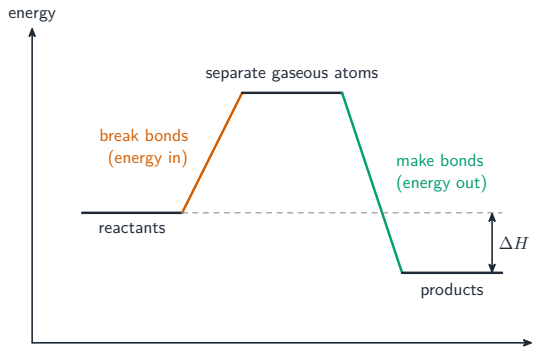
1 Worked examples

Hess's law

Hess's law: the total enthalpy change is the same whatever the route from reactants to products (energy is conserved).



The direct route equals the indirect route: $\Delta H_r = \Delta H_1 + \Delta H_2$



Bond energies feed an enthalpy (Hess) cycle

Using formation data

$$\Delta H_r = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$$

Example. $CaCO_3 \rightarrow CaO + CO_2$ (ΔH_f / kJ mol^{-1}: $CaCO_3$ −1207, CaO −635, CO_2 −394): $\Delta H_r = (-635 - 394) - (-1207) = +178$ kJ mol^{-1}.

Using combustion data

$$\Delta H_r = \sum \Delta H_c(\text{reactants}) - \sum \Delta H_c(\text{products})$$

2 Practice

2.1 State **Hess's law**. [1]

2.2 Give **one** reason Hess's law is useful (something it lets you find). [1]

3 Exam-style questions

3.1 Hess's law is a consequence of: [1]

- **A** the conservation of mass
- **B** the conservation of energy
- **C** the ideal gas equation
- **D** Avogadro's law

3.2 ΔH_r from formation data is calculated as: [1]

- **A** $\sum \Delta H_f(\text{reactants}) - \sum \Delta H_f(\text{products})$
- **B** $\sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$
- **C** $\sum \Delta H_c(\text{products}) - \sum \Delta H_c(\text{reactants})$
- **D** $mc\Delta T$

3.3 For $C_2H_4 + H_2 \rightarrow C_2H_6$, use ΔH_f (kJ mol^{-1}: C_2H_4 +52, C_2H_6 −85, H_2 0) to calculate ΔH_r. [3]

3.4 Explain why Hess's law allows the enthalpy change of formation of a compound to be found even when the reaction cannot be carried out directly. [2]

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus$ / kJ mol^{-1}
$CH_3OH(l)$	−239
$CO_2(g)$	−394
$H_2O(l)$	−286

(a) Write the equation for the complete combustion of methanol, then construct an energy cycle through the elements and use it to calculate $\Delta H_c^\ominus$ for methanol. [5]

(b) Explain why $\Delta H_f^\ominus$ of $O_2(g)$ is zero. [2]

(c) A school laboratory obtains −690 kJ mol^{-1}. State whether that is more or less exothermic than your answer, and give one reason for the difference. [2]

(d) State the principle that allows this calculation even though methanol is never actually burned along the route you drew. [2]

Solutions

2.1 the total enthalpy change of a reaction is the same regardless of the route taken.

2.2 e.g. it lets you find an enthalpy change that cannot be measured directly / calculate ΔH_r from formation or combustion data.

3.1 B. Hess's law follows from the conservation of energy.

3.2 B. $\sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$.

3.3 $\Delta H_r = \Delta H_f(C_2H_6) - [\Delta H_f(C_2H_4) + \Delta H_f(H_2)] = -85 - (52 + 0) = -137$ kJ mol^{-1}.

3.4 an indirect route (e.g. via combustion or formation of the elements) can be measured; by Hess's law the sum of the indirect steps equals the direct (unmeasurable) enthalpy change.

3.5 (a) $CH_3OH(l) + 1\frac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$. The cycle runs from the elements up to the reactants and up to the products, so $\Delta H_c = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$; $= [(-394) + 2(-286)] - [(-239) + 0] = -966 + 239 = -727$ kJ mol^{-1}.
(b) By definition the standard enthalpy of formation of an **element in its standard state** is zero — forming it from itself is no change at all.
(c) Less exothermic; heat escapes to the surroundings, so less energy appears to have been released per mole burned.
(d) Hess's law: the enthalpy change of a reaction is independent of the route taken, provided the initial and final states are the same.

9 The data shown are needed for this question.

$\Delta H_f^\ominus(P_4O_{10}(s)) = -3012\,\text{kJ mol}^{-1}$

$\Delta H_f^\ominus(H_2O(l)) = -286\,\text{kJ mol}^{-1}$

$\Delta H_f^\ominus(H_3PO_4(s)) = -1279\,\text{kJ mol}^{-1}$

What is $\Delta H^\ominus$ for the reaction shown?

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(s)$$

A $-9844\,\text{kJ mol}^{-1}$

B $-388\,\text{kJ mol}^{-1}$

C $-97\,\text{kJ mol}^{-1}$

D $+2019\,\text{kJ mol}^{-1}$

3 The alkanes are a homologous series of organic molecules. Alkanes are generally unreactive and are commonly used as fuels.

(a) Define homologous series.

.....

..... [2]

(b) Give two reasons to explain the general unreactivity of alkanes.

1 .....

2 .....

[2]

(c) Alkanes with low relative molecular mass, M_r, are more useful than those found in heavier crude oil fractions.

Name the process that is used to obtain alkanes with low M_r from heavier crude oil fractions.

..... [1]

(d) Hexane, C_6H_{14}, has four structural isomers.

Fig. 3.1 shows hexane and two of its structural isomers.

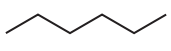
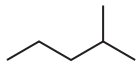
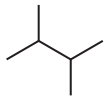
hexane	A	B	C	D
				

Fig. 3.1

(i) Complete Fig. 3.1 by drawing structures for **C** and **D**, the other two structural isomers of hexane. [2]

(ii) **A**, **B** and hexane have different boiling points.

Arrange **A**, **B** and hexane in order of increasing boiling point.

Explain your answer.

lowest ..... < ..... < ..... highest

.....

.....

....

[3]

(e) Hexane can be converted into compounds **E** and **F** at high temperature and pressure. Fig. 3.2 shows the reaction scheme involving hexane, **E** and **F**.

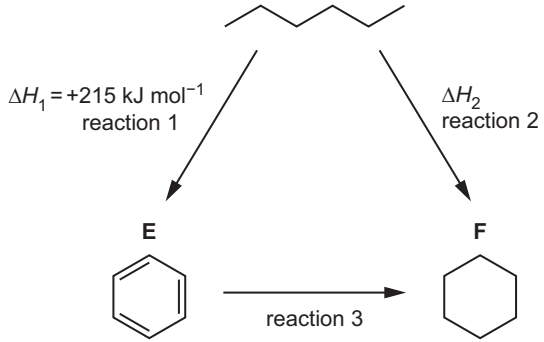


Fig. 3.2

(i) Identify a suitable reagent for reaction 3.

..... [1]

(ii) Use the data in Fig. 3.2 and in Table 3.1 to calculate the enthalpy change of reaction 2, ΔH_2.

Table 3.1

compound	enthalpy change of formation, ΔH_f / kJ mol^{-1}
	−156
	+48

ΔH_2 = ... kJ mol^{-1} [2]

(iii) Write an equation for the complete combustion of hexane.

..... [1]

[Total: 14]

1 Sodium carbonate can be manufactured using a two-step process. The first step involves making sodium hydrogencarbonate, $NaHCO_3$, which is then converted into sodium carbonate, Na_2CO_3, in the second step of the process by the following reaction.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

In this experiment you will determine the enthalpy change, ΔH_r, for this reaction. You will do this by calculating the enthalpy changes when separate samples of sodium hydrogencarbonate and sodium carbonate are added to excess hydrochloric acid, $HCl(aq)$. You will then combine these values using Hess's law.

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_1$$

$$Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_2$$

FB 1 is 2.0 mol dm^{-3} hydrochloric acid, HCl.
FB 2 is sodium hydrogencarbonate, $NaHCO_3$.
FB 3 is sodium carbonate, Na_2CO_3.

(a) Method

Experiment 1

- Support a cup in the 250 cm^3 beaker.
- Use the 25 cm^3 measuring cylinder to transfer 20.0 cm^3 of **FB 1** into the cup.
- Weigh the container with **FB 2**. Record the mass.
- Measure the temperature of the acid in the cup. Record this temperature.
- Carefully add all of the **FB 2**, in small portions to avoid acid spray. Stir to dissolve. Record the lowest temperature.
- Reweigh the container with any residual **FB 2**. Record the mass.
- Calculate and record the mass of **FB 2** used.
- Calculate and record the decrease in temperature.

Experiment 2

- Support the second cup in the 250 cm^3 beaker.
- Use the 25 cm^3 measuring cylinder to transfer 20.0 cm^3 of **FB 1** into the second cup.
- Weigh the container with **FB 3**. Record the mass.
- Measure the temperature of the acid in the cup. Record this temperature.
- Carefully add all of the **FB 3**, in small portions to avoid acid spray. Stir to dissolve. Record the highest temperature.
- Reweigh the container with any residual **FB 3**. Record the mass.
- Calculate and record the mass of **FB 3** used.
- Calculate and record the increase in temperature.

Keep FB 1 for use in Question 3.

Results

I	
II	
III	
IV	

[4]

(b) Calculations

(i) Calculate the amount, in mol, of **FB 2** that reacts with **FB 1** and the amount, in mol, of **FB 3** that reacts with **FB 1**.

Experiment 1	**Experiment 2**
amount of **FB 2** = ... mol	amount of **FB 3** = ... mol

[2]

(ii) Calculate the energy changes, in J, for each reaction.

Experiment 1	**Experiment 2**
energy change = ... J	energy change = ... J

[1]

(iii) Calculate the enthalpy change, ΔH, in kJ mol^{-1} for each reaction.

Experiment 1	**Experiment 2**
ΔH_1 = ... kJ mol^{-1} (*sign* *value*)	ΔH_2 = ... kJ mol^{-1} (*sign* *value*)

[2]

(iv) Construct an enthalpy cycle and use Hess's law to determine the enthalpy change, ΔH_r, for the reaction shown.
Show your working.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

(If you were unable to calculate values in **(b)(iii)** then assume that ΔH_1 is +27.3 kJ mol^{-1} and that ΔH_2 is –24.9 kJ mol^{-1}. These may **not** be the correct values.)

ΔH_r = ... kJ mol^{-1}
sign *value*

[2]

[Total: 11]