

Mastering Chemistry

- Book 3C
- Topic 8 Chemical Reactions and Energy



Content

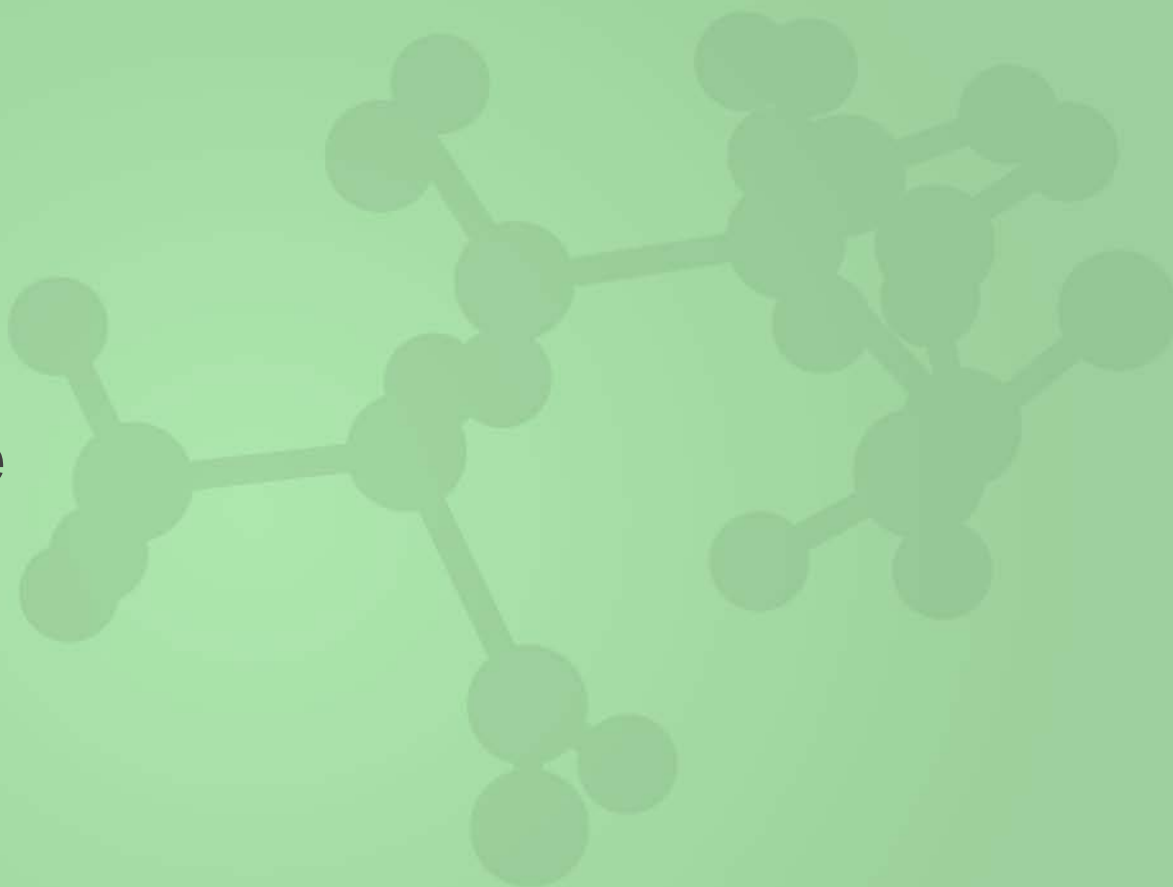
- ➡ 35.1 Hess's law
- ➡ 35.2 Determining the enthalpy change of formation of calcium carbonate
- ➡ 35.3 Enthalpy change of formation from enthalpy changes of combustion
- ➡ 35.4 Enthalpy change of reaction from enthalpy changes of formation

Continued on next page ➡



Content

- ➡ Key terms
- ➡ Summary
- ➡ Unit Exercise
- ➡ Topic Exercise





35.1 Hess's law (p.45)

- ◆ The techniques described in Unit 34 allow enthalpy changes to be determined directly in a single experiment. Unfortunately, the enthalpy changes of many reactions are very difficult to be determined directly.
- ◆ **Hess's law** (赫斯定律) comes to the rescue, allowing enthalpy changes to be determined indirectly.

Hess's law states that 'the total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same'.

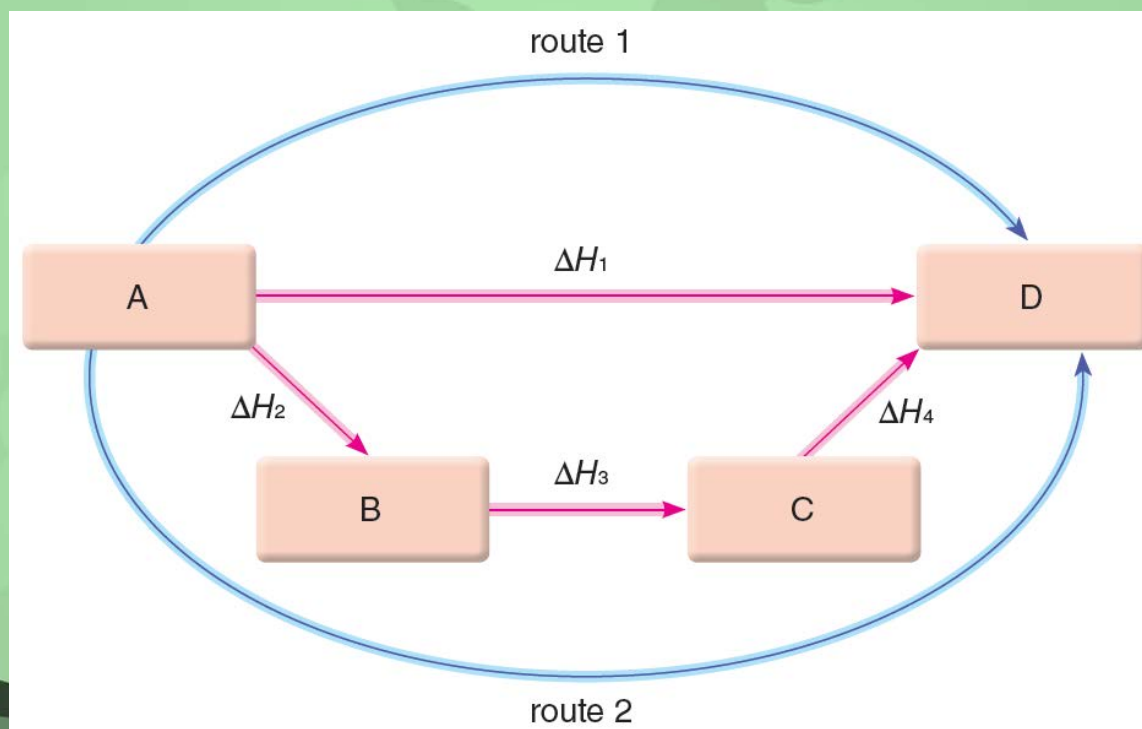


Hess's law [Ref.](#)



35.1 Hess's law (p.45)

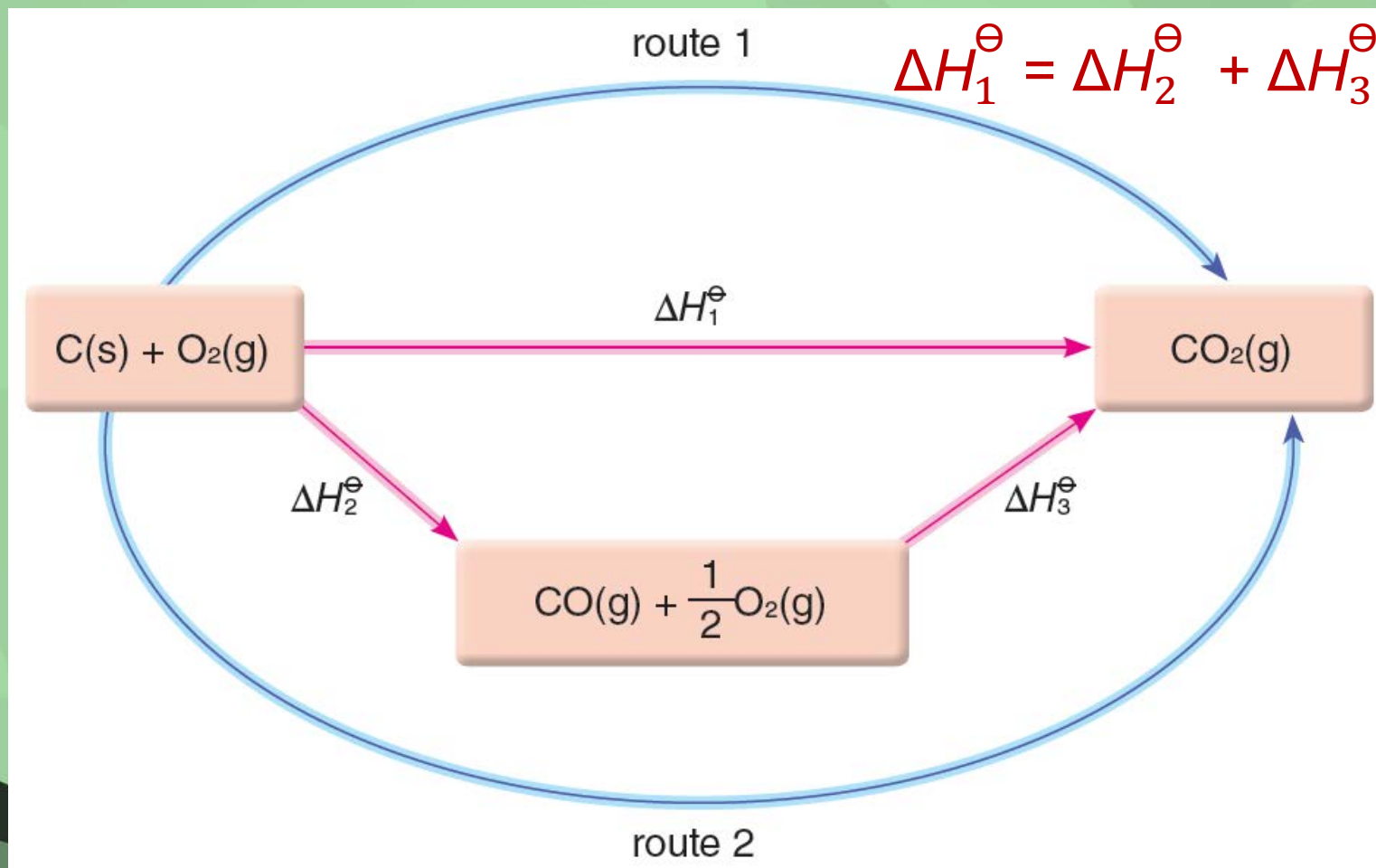
- ◆ The **enthalpy change cycle** (焓變循環) shown helps illustrate Hess's law.
- ◆ According to Hess's law, the enthalpy change in *route 1* will be equal to the sum of enthalpy changes of the reactions in *route 2*. In symbols: $\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$





35.1 Hess's law (p.45)

- Consider: carbon + oxygen \rightarrow carbon dioxide.
Two possible routes: direct *route 1* and indirect *route 2*

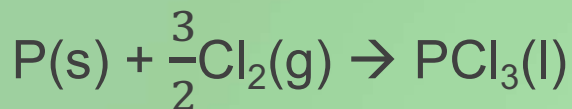




35.1 Hess's law (p.45)

Q (Example 35.1)

Use the following standard enthalpy change data:

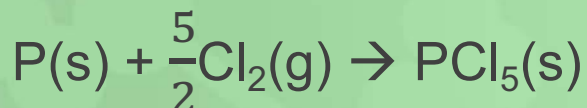


$$\Delta H_1^\ominus = -339 \text{ kJ mol}^{-1}$$



$$\Delta H_2^\ominus = -124 \text{ kJ mol}^{-1}$$

to work out the standard enthalpy change of the following reaction:



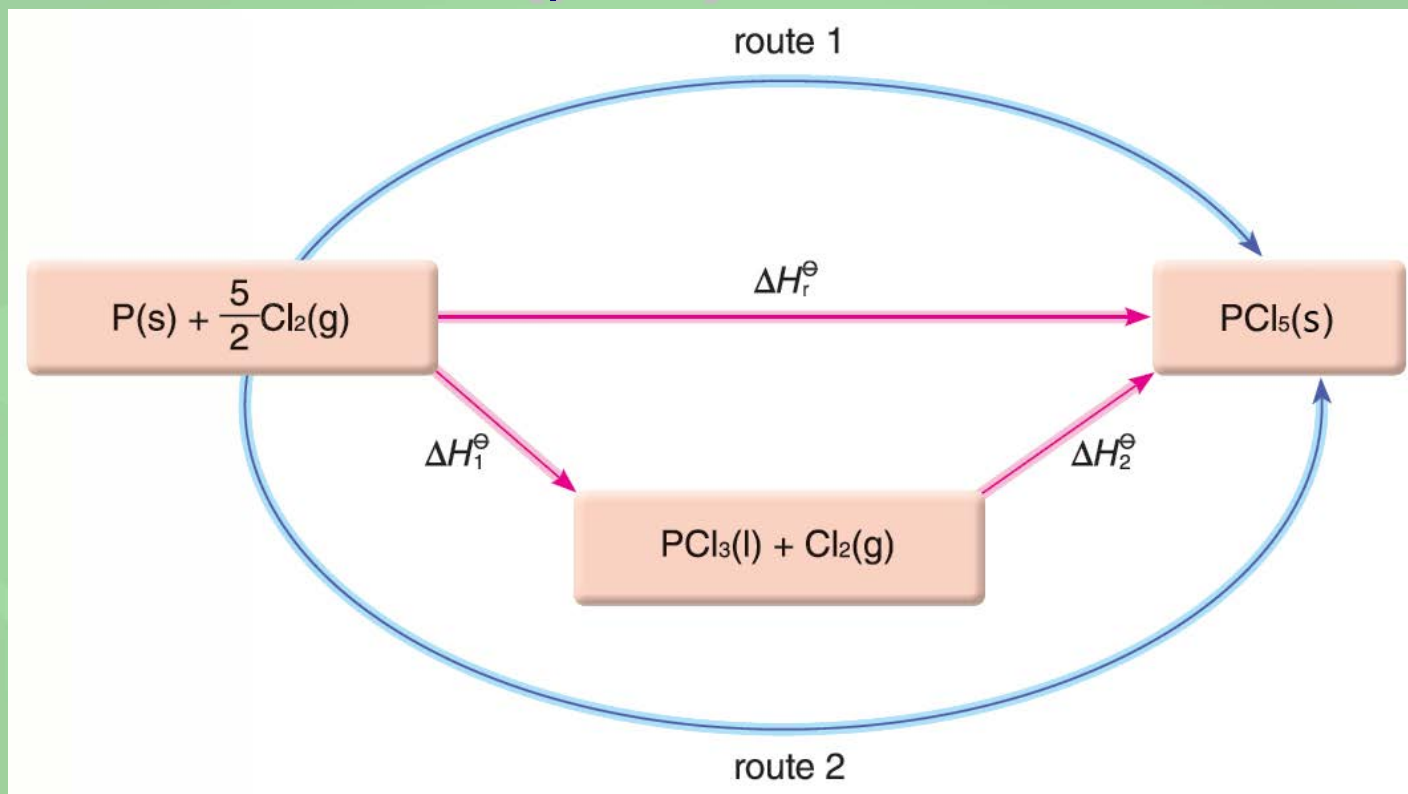
A

Method 1 Solving by an enthalpy change cycle

Based on the given data to construct an enthalpy change cycle in which the equation for the reaction with the enthalpy change you want to find is along the top.



35.1 Hess's law (p.45)



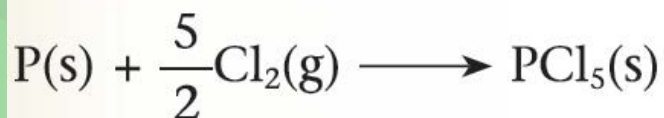
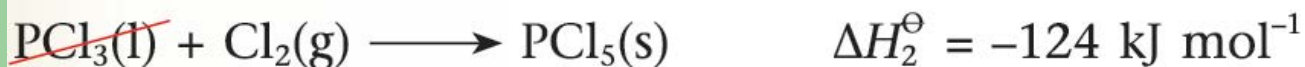
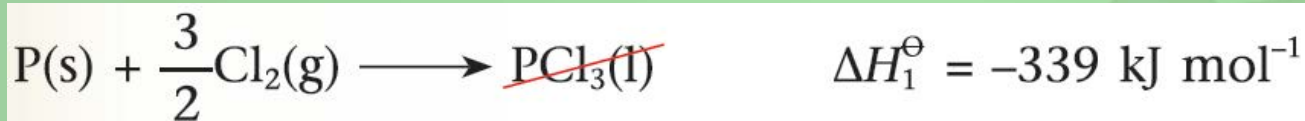
According to Hess's law: $\Delta H_r^\ominus = \Delta H_1^\ominus + \Delta H_2^\ominus$
 $= [(-339) + (-124)] \text{ kJ mol}^{-1} = -463 \text{ kJ mol}^{-1}$



35.1 Hess's law (p.45)

Method 2 Solving by manipulating equations

Add together the two equations:



$$\begin{aligned} \Delta H_r^\ominus &= [(-339) + (-124)] \text{ kJ mol}^{-1} \\ &= -463 \text{ kJ mol}^{-1} \end{aligned}$$

\therefore the enthalpy change of the reaction is -463 kJ mol^{-1} .



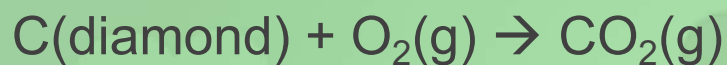
35.1 Hess's law (p.45)

Q (Example 35.2)

Use the following standard enthalpy change data:



$$\Delta H_1^\ominus = -394 \text{ kJ mol}^{-1}$$



$$\Delta H_2^\ominus = -395 \text{ kJ mol}^{-1}$$

to work out the enthalpy change for the conversion of graphite to diamond under standard conditions:



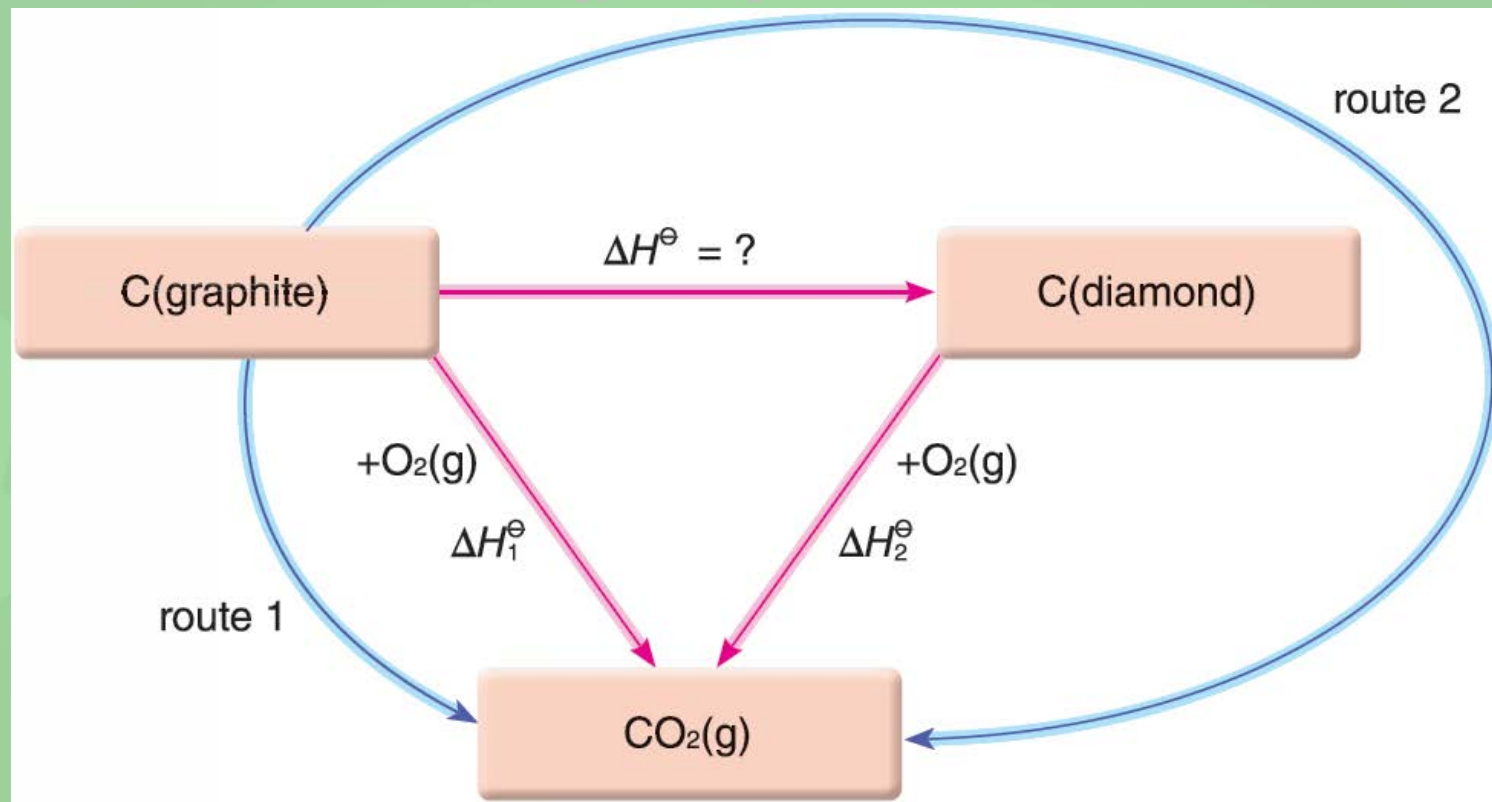
A

Method 1 Solving by an enthalpy change cycle

Based on the given data to construct an enthalpy change cycle in which the equation for the process with the enthalpy change you want to find is along the top.



35.1 Hess's law (p.45)



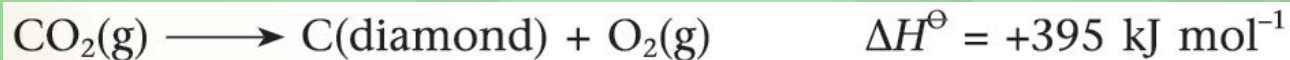
According to Hess's law: $\Delta H_1^\ominus = \Delta H^\ominus + \Delta H_2^\ominus$
 $\Delta H^\ominus = \Delta H_1^\ominus - \Delta H_2^\ominus$
 $= [(-394) - (-395)] \text{ kJ mol}^{-1} = +1 \text{ kJ mol}^{-1}$



35.1 Hess's law (p.45)

Method 2 Solving by manipulating equations

You need C(diamond) on the product side, so reverse the second equation and the sign of the enthalpy change as well.



Add together the first equation and the reversed equation:



Enthalpy change for the conversion = $[(-394) + (+395)] \text{ kJ mol}^{-1} = +1 \text{ kJ mol}^{-1}$
 \therefore the enthalpy change of the conversion of graphite to diamond under the standard conditions is $+1 \text{ kJ mol}^{-1}$.



35.1 Hess's law (p.45)

Q (Example 35.3)

The compound diborane (B_2H_6) is used as a rocket fuel. The equation for the combustion of diborane is shown below.



The enthalpy changes of three reactions under certain conditions are shown:

<u>Reaction</u>	<u>Enthalpy change</u>
(1) $2\text{B}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{B}_2\text{H}_6(\text{g})$	$+42 \text{ kJ mol}^{-1}$
(2) $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$	-286 kJ mol^{-1}
(3) $\text{B}(\text{s}) + \frac{3}{4} \text{O}_2(\text{g}) \rightarrow \frac{1}{2} \text{B}_2\text{O}_3(\text{s})$	-635 kJ mol^{-1}

Calculate the enthalpy change of combustion of diborane under the same conditions.



35.1 Hess's law (p.45)

A

Look at the equation for the combustion of $\text{B}_2\text{H}_6(\text{g})$. You need one $\text{B}_2\text{H}_6(\text{g})$ on the reactant side, so reverse equation (1), giving equation (1)'.



You need three $\text{H}_2\text{O}(\text{l})$ on the product side, so multiply equation (2) by 3, giving equation (2)'.



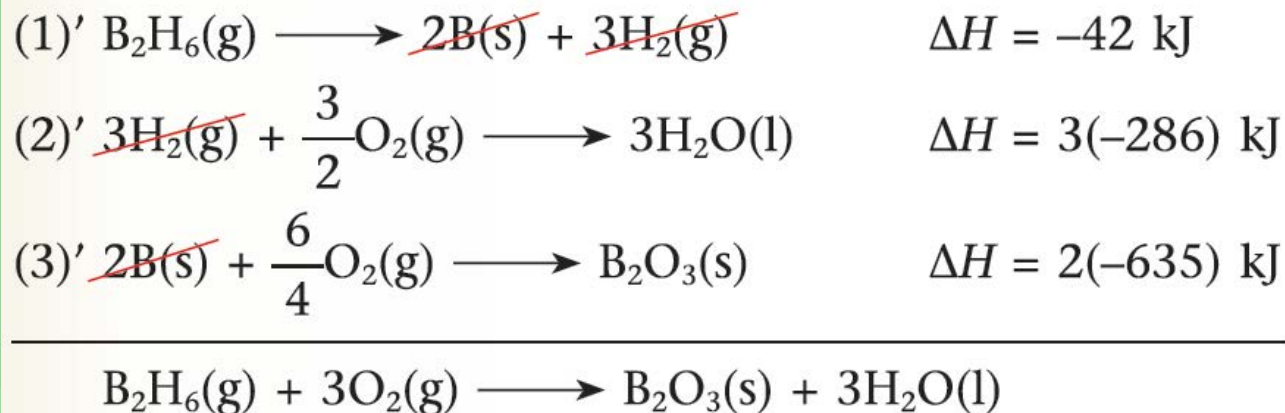
You need one $\text{B}_2\text{O}_3(\text{s})$ on the product side, so multiply equation (3) by 2, giving equation (3)'.





35.1 Hess's law (p.45)

Add together equations (1)', (2)' and (3)':



Enthalpy change of combustion of diborane = $[(-42) + 3(-286) + 2(-635)] \text{ kJ mol}^{-1}$
 $= -2\,170 \text{ kJ mol}^{-1}$

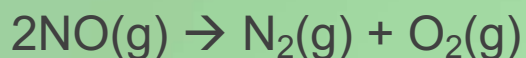
\therefore the enthalpy change of the combustion of diborane is $-2\,170 \text{ kJ mol}^{-1}$.



35.1 Hess's law (p.45)

Practice 35.1

1 Use the following standard enthalpy change data:



$$\Delta H^\ominus = -180 \text{ kJ mol}^{-1}$$

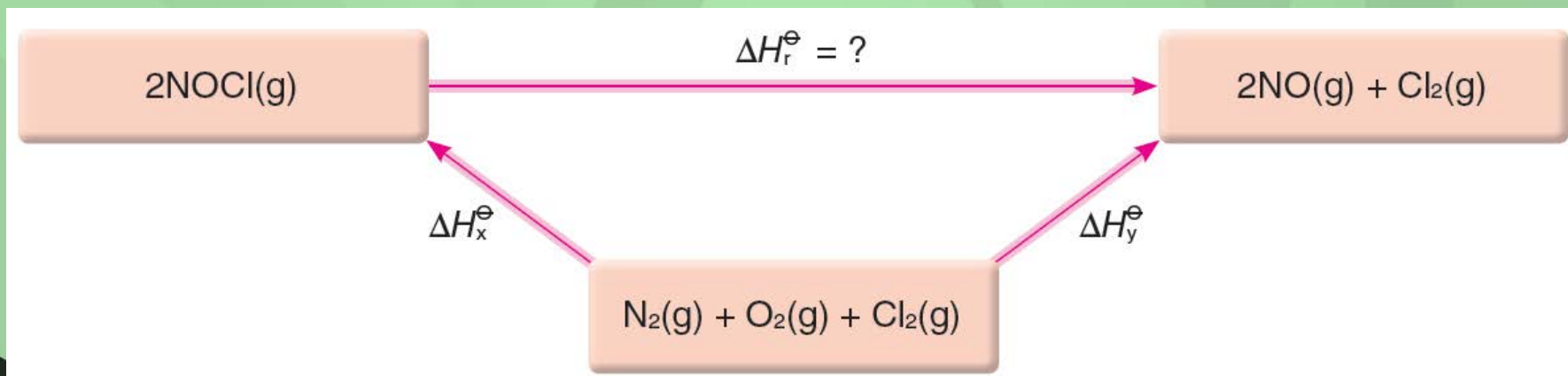


$$\Delta H^\ominus = +103 \text{ kJ mol}^{-1}$$

to work out the standard enthalpy change of the following reaction:



An enthalpy change cycle is constructed based on the given data:1





35.1 Hess's law (p.45)

a) Give the values of ΔH_x^\ominus and ΔH_y^\ominus respectively.

$$\Delta H_x^\ominus = +103 \text{ kJ}; \quad \Delta H_y^\ominus = +180 \text{ kJ}$$

b) Calculate the standard enthalpy change of the following reaction:



According to Hess's law:

$$\Delta H_x^\ominus + \Delta H_r^\ominus = \Delta H_y^\ominus$$

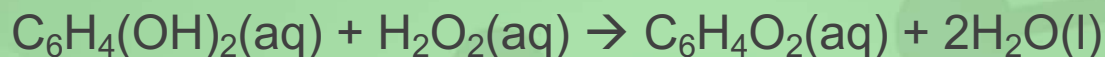
$$\Delta H_r^\ominus = \Delta H_y^\ominus - \Delta H_x^\ominus = [+180 - (+103)] \text{ kJ} = +77 \text{ kJ}$$



35.1 Hess's law (p.45)

2 When in danger, bombardier beetles can fire a hot and toxic mixture of chemicals at their attacker. This mixture contains quinone, $\text{C}_6\text{H}_4\text{O}_2$, a compound that is formed by the reaction of hydroquinone, $\text{C}_6\text{H}_4(\text{OH})_2$, with hydrogen peroxide, H_2O_2 .

The equation for the overall reaction is:



The enthalpy changes of three reactions under certain conditions are as follows:

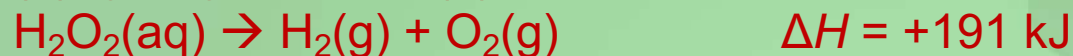
<u>Reaction</u>	<u>Enthalpy change</u>
(1) $\text{C}_6\text{H}_4(\text{OH})_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + \text{H}_2(\text{g})$	+177 kJ mol ⁻¹
(2) $\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\text{aq})$	-191 kJ mol ⁻¹
(3) $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$	-286 kJ mol ⁻¹

Calculate the enthalpy change of formation of $\text{C}_6\text{H}_4\text{O}_2(\text{aq})$ and $2\text{H}_2\text{O}(\text{l})$ from $\text{C}_6\text{H}_4(\text{OH})_2(\text{aq})$ and $\text{H}_2\text{O}_2(\text{aq})$ under the same conditions.



35.1 Hess's law (p.45)

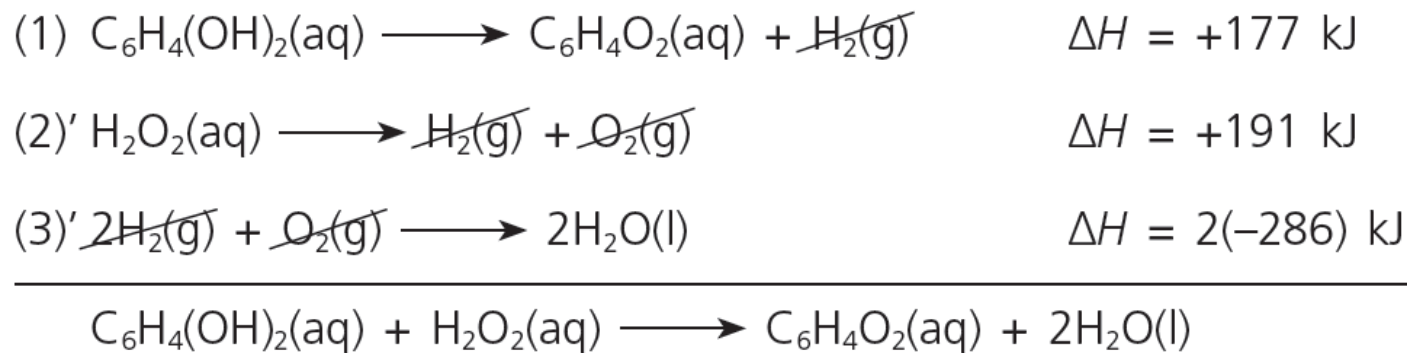
Look at equation (2). You need one $\text{H}_2\text{O}_2(\text{aq})$ on the reactant side, so reverse equation (2), giving equation (2)'.



Look at equation (3). You need two $\text{H}_2\text{O}(\text{l})$ on the product side, so multiply equation (3) by 2, giving equation (3)'.



Add together (1), (2)' and (3)':





35.1 Hess's law (p.45)

Enthalpy change of formation of $\text{C}_6\text{H}_4\text{O}_2(\text{aq})$ and $2\text{H}_2\text{O}(\text{l})$ from $\text{C}_6\text{H}_4(\text{OH})_2(\text{aq})$ and $\text{H}_2\text{O}_2(\text{aq})$

$$= [+177 + (+191) + 2(-286)] \text{ kJ (1)}$$

$$= -204 \text{ kJ (1)}$$

\therefore the enthalpy change of formation of $\text{C}_6\text{H}_4\text{O}_2(\text{aq})$ and $2\text{H}_2\text{O}(\text{l})$ from $\text{C}_6\text{H}_4(\text{OH})_2(\text{aq})$ and $\text{H}_2\text{O}_2(\text{aq})$ is -204 kJ .



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

- ♦ $\text{Ca(s)} + \text{C(graphite)} + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s})$
- ♦ Calcium carbonate cannot be formed directly from its constituent elements. Hence the enthalpy change of formation of calcium carbonate cannot be determined directly from experimental results. Instead, you can determine the enthalpy change indirectly based on Hess's law.



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

- ◆ You can calculate the enthalpy change of formation of calcium carbonate from the information below.
 - enthalpy change of reaction between calcium and hydrochloric acid;
 - enthalpy change of reaction between calcium carbonate and hydrochloric acid;
 - enthalpy change of formation of carbon dioxide; and
 - enthalpy change of formation of water.
- ◆ The first two items can be determined directly from experimental results while the last two items can be obtained from data books.



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

Step 1 ΔH for $\text{Ca(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$

◆ Mass of calcium used = 1.04 g

Initial temperature of the acid = 26.0 °C

Final temperature of the resulting solution = 54.5 °C

Heat capacity of the resulting solution = 425 J K⁻¹

Temperature change $\Delta T = (54.5 - 26.0) \text{ °C} = +28.5 \text{ °C} = +28.5 \text{ K}$

Enthalpy change for the quantities in this experiment = 425 J K⁻¹ x 28.5 K
= 12 100 J = 12.1 kJ

Number of moles of Ca used = $\frac{1.04 \text{ g}}{40.1 \text{ g mol}^{-1}} = 0.0259 \text{ mol}$

Enthalpy change of the reaction = $\frac{-12.1 \text{ kJ}}{0.0259 \text{ mol}} = -467 \text{ kJ mol}^{-1}$



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

Step 2 ΔH for $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

◆ Mass of calcium carbonate used = 2.52 g

Initial temperature of the acid = 26.0 °C

Final temperature of the resulting solution = 27.1 °C

Heat capacity of the resulting solution = 430 J K⁻¹

Temperature change $\Delta T = (27.1 - 26.0) \text{ °C} = +1.1 \text{ °C} = +1.1 \text{ K}$

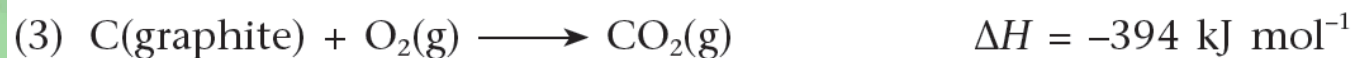
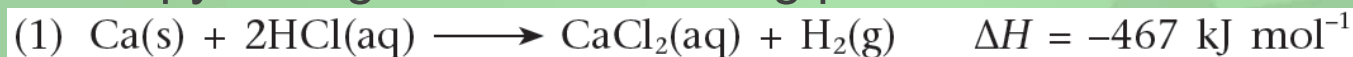
Enthalpy change for the quantities in this experiment
= 430 J K⁻¹ x 1.1 K = 473 J



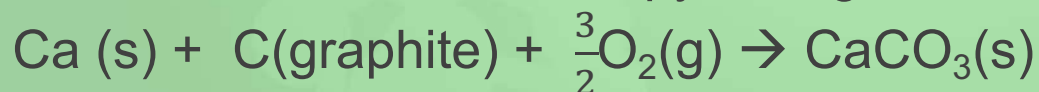
35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

Step 3 ΔH for $\text{Ca (s)} + \text{C(graphite)} + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s})$

Enthalpy changes of the following processes are known:



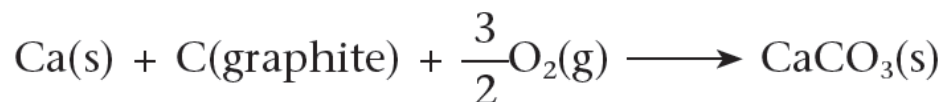
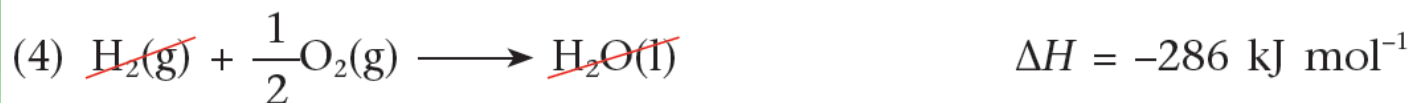
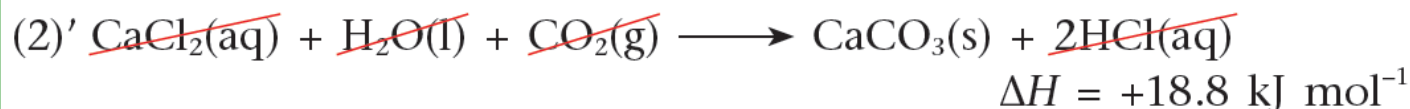
You need to find the enthalpy change of the following process:



Reverse equation (2) to give equation (2)'. Add together the equations as shown below:



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)



According to Hess's law:

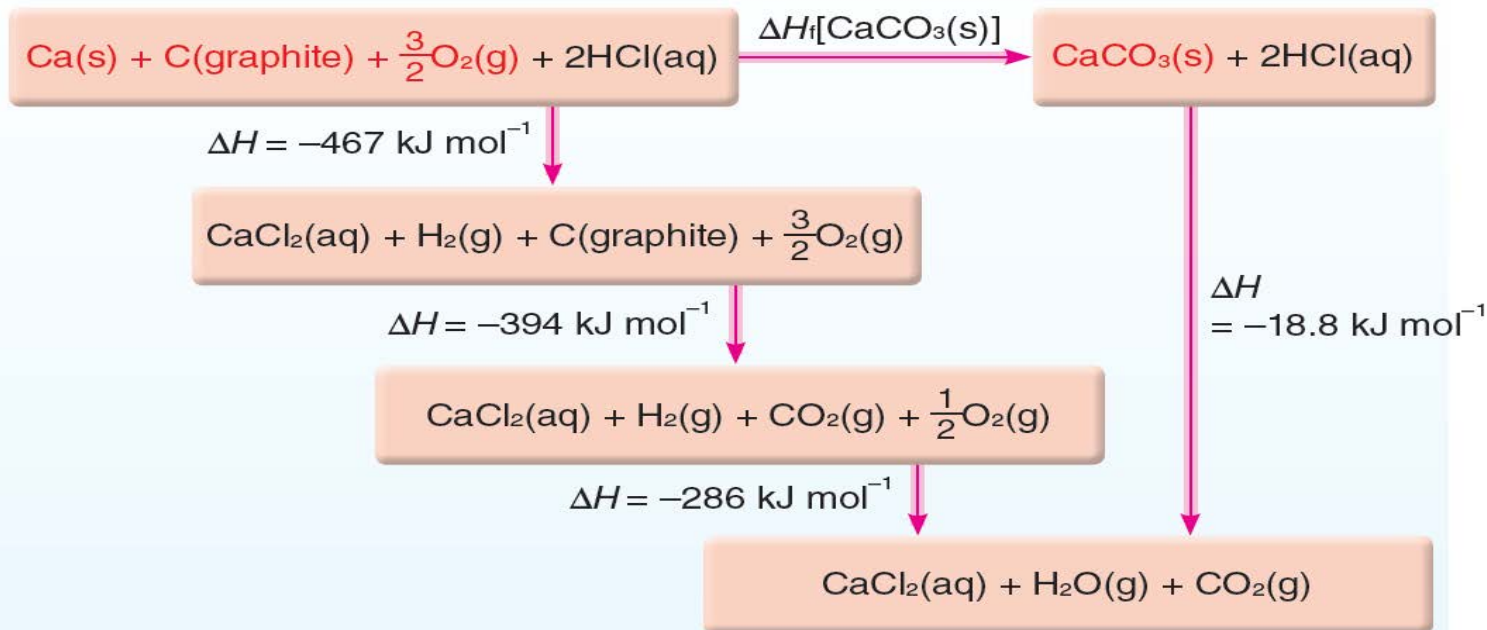
$$\Delta H_f^\ominus [\text{CaCO}_3\text{(s)}] = [(-467) + (+18.8) + (-394) + (-286)] \text{ kJ mol}^{-1} = -1\,128 \text{ kJ mol}^{-1}$$

\therefore the enthalpy change of formation of calcium carbonate is $-1\,128 \text{ kJ mol}^{-1}$.



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

You can construct an enthalpy change cycle based on the known enthalpy changes:



According to Hess's law:

$$\Delta H_f[\text{CaCO}_3\text{(s)}] + (-18.8 \text{ kJ mol})$$

$$= (-467 \text{ kJ mol}^{-1}) + (-394 \text{ kJ mol}^{-1}) + (-286 \text{ kJ mol}^{-1})$$

$$\Delta H_f[\text{CaCO}_3\text{(s)}] = [(-467) + (-394) + (-286) - (-18.8)] \text{ kJ mol}^{-1}$$

$$= -1128 \text{ kJ mol}^{-1}$$



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)



Determining the enthalpy change of formation of MgO

[Ref.](#)



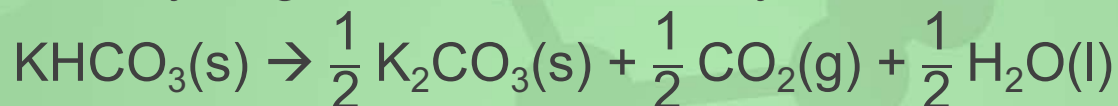
Determining the enthalpy change of thermal decomposition of KHCO_3



35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

Practice 35.2

Potassium hydrogencarbonate decomposes on heating to form potassium carbonate. It is difficult to measure the enthalpy change of the decomposition of potassium hydrogencarbonate directly.



One method of determining this enthalpy change is to react known amounts of potassium hydrogencarbonate and potassium carbonate, separately, with excess dilute hydrochloric acid.

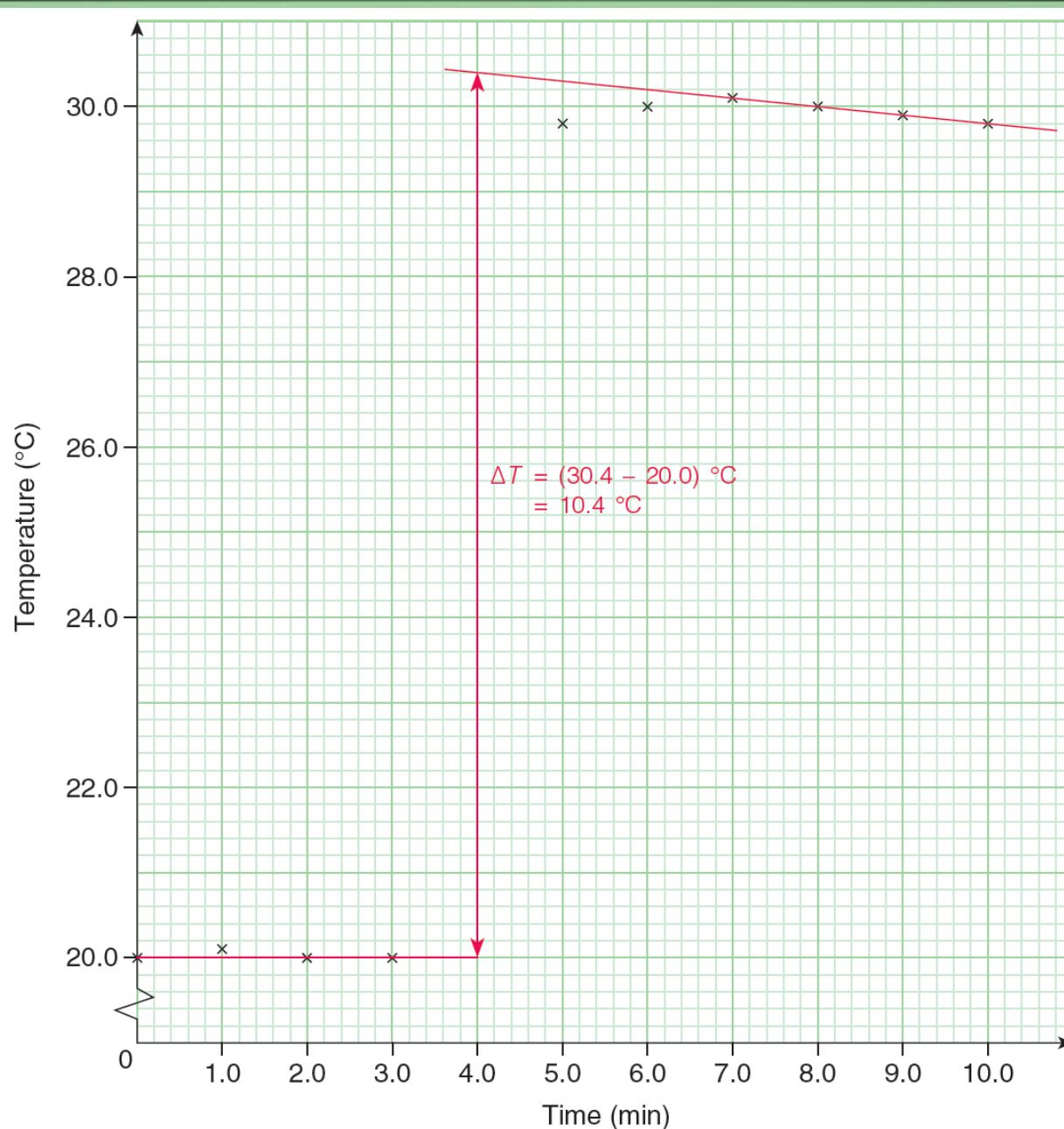


35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

- a) A student determined the temperature change when a known amount of potassium carbonate reacted with dilute hydrochloric acid.
- Dilute hydrochloric acid (excess) were placed in a polystyrene foam cup.
 - The temperature of acid was measured every minute for 3.0 minutes.
 - At 4.0 minutes, 0.0300 mole of potassium carbonate was added to the acid.
 - The temperature of the resulting solution was measured every minute for a further of 6.0 minutes.

The graph below shows the results obtained:

- i) What is the greatest temperature rise of the resulting solution as estimated from the graph?





35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

ii) Given that:

Mass of the resulting solution = 28.0 g

Specific heat capacity of the resulting solution = $4.3 \text{ J g}^{-1} \text{ K}^{-1}$

Calculate the enthalpy change for the above quantities in this experiment.

$$\begin{aligned}\text{Enthalpy change for the quantities in experiment} &= m \times c \times \Delta T \\ &= 28.0 \text{ g} \times 4.3 \text{ J g}^{-1} \text{ K}^{-1} \times 10.4 \text{ K} = 1\,250 \text{ J} = 1.25 \text{ kJ}\end{aligned}$$

iii) Calculate the enthalpy change of the reaction when 1 mole of potassium carbonate reacts with dilute hydrochloric acid.

$$\begin{aligned}\text{Enthalpy change of reaction between } \text{K}_2\text{CO}_3(\text{s}) \text{ and } \text{HCl}(\text{aq}) \\ &= \frac{-1.25 \text{ kJ}}{0.0300 \text{ mol}} = -41.7 \text{ kJ mol}^{-1}\end{aligned}$$



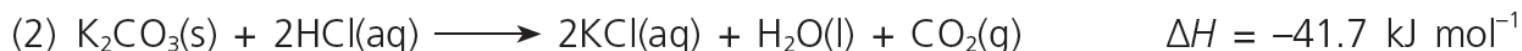
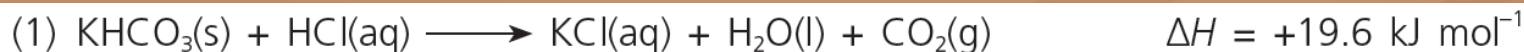
35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

- b) The enthalpy change of the reaction between potassium hydrogencarbonate and dilute hydrochloric acid was found by a similar method to be $+19.6 \text{ kJ mol}^{-1}$.

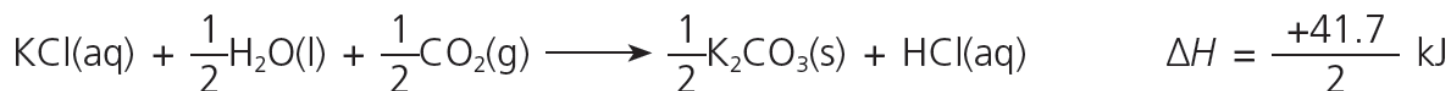
Calculate the enthalpy change of decomposition of potassium hydrogencarbonate under the experimental conditions.



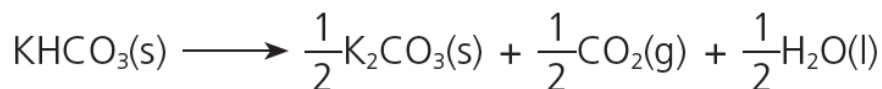
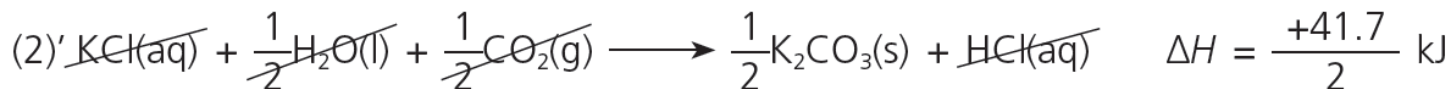
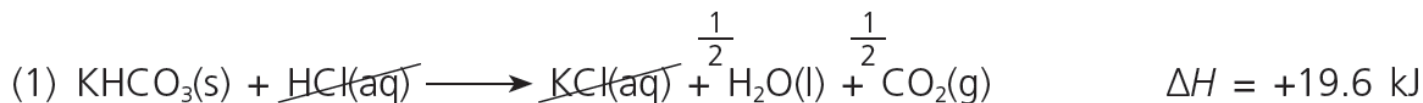
35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)



You need $\frac{1}{2}\text{K}_2\text{CO}_3(\text{s})$ on the product side, so multiply equation (2) by $\frac{1}{2}$ and reverse it, giving equation (2)'.



Add together (1) and (2)':





35.2 Determining the enthalpy change of formation of calcium carbonate (p.52)

According to Hess's law:

Enthalpy change of decomposition of $\text{KHCO}_3(\text{s})$

$$= \left(+19.6 + \frac{+41.7}{2} \right) \text{ kJ mol}^{-1}$$

$$= +40.5 \text{ kJ mol}^{-1}$$

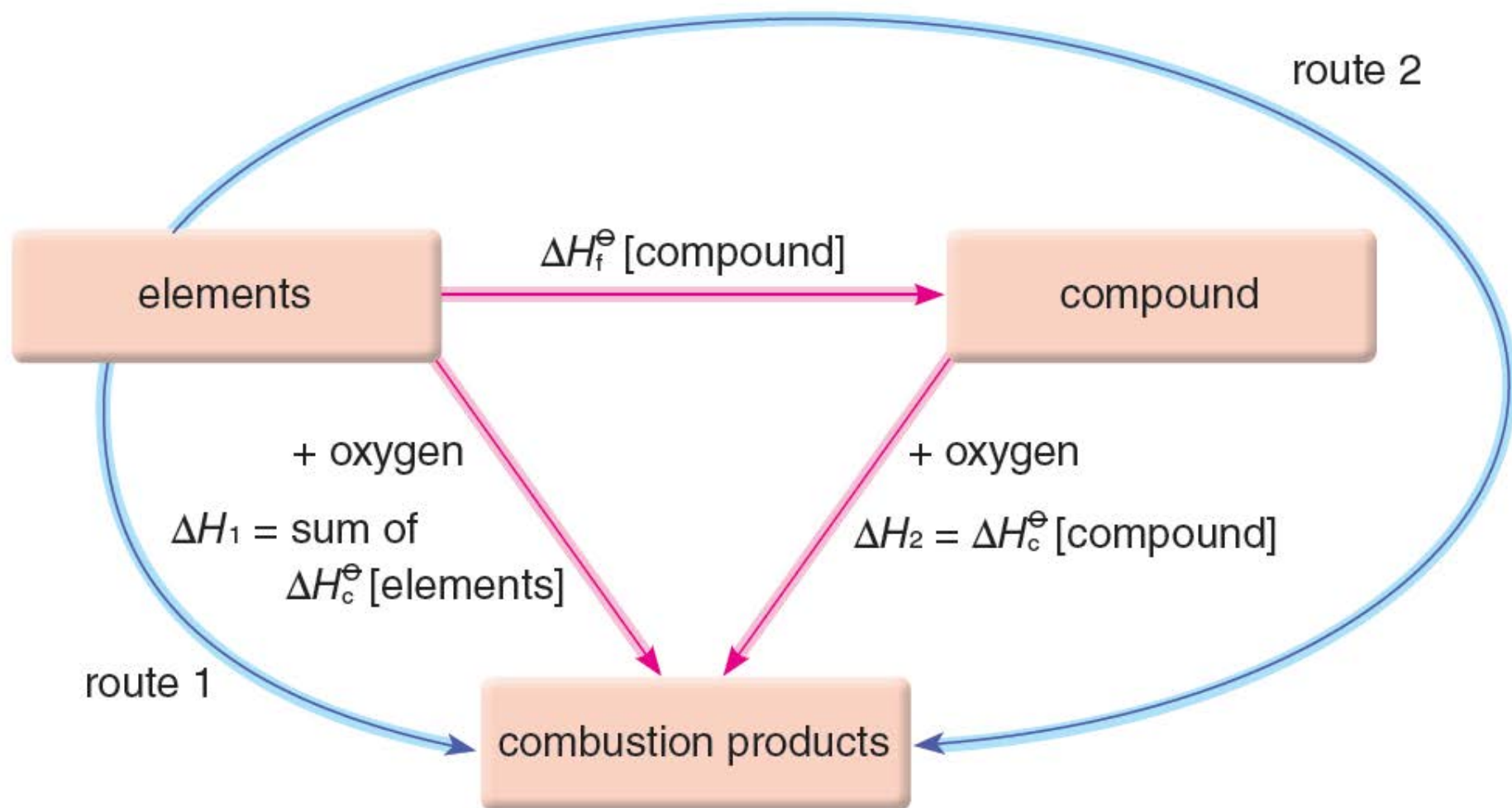


35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

- ◆ Enthalpy changes of formation of compounds are often difficult or impossible to measure directly. This is because the constituent elements often do not react directly to form the compound that you are interested in.
- ◆ For example, the following equation represents the formation of ethyne (C_2H_2) from its elements.
$$2\text{C}(\text{graphite}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$$
- ◆ Carbon and hydrogen do not react to give ethyne under normal conditions. However, all the chemical species involved burn readily in oxygen, so their enthalpy changes of combustion can be measured.



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)





35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

- According to Hess's law: $\Delta H_1 = \Delta H_f^\ominus + \Delta H_2$

Rearranging gives: $\Delta H_f^\ominus = \Delta H_1 - \Delta H_2$

So

$$\Delta H_f^\ominus [\text{compound}] = \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{compound}]$$



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

Q (Example 35.4)

Calculate the standard enthalpy change of formation of ethyne ($\text{C}_2\text{H}_2(\text{g})$) from the following standard enthalpy changes of combustion:

Substance	ΔH_c^\ominus (kJ mol ⁻¹)
$\text{C}_2\text{H}_2(\text{g})$	-1 300
$\text{C}(\text{graphite})$	-394
$\text{H}_2(\text{g})$	-286

A

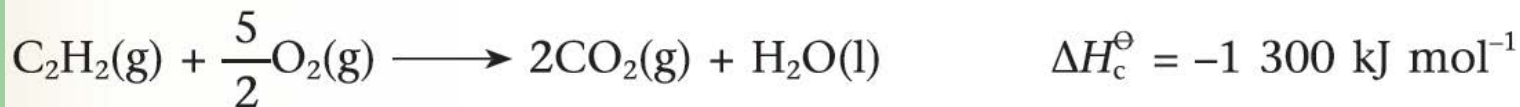
Method 1 Solving by an enthalpy change cycle

You are given the standard enthalpy change of combustion data and so, knowing the definition, you can write out the following thermochemical equations:



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

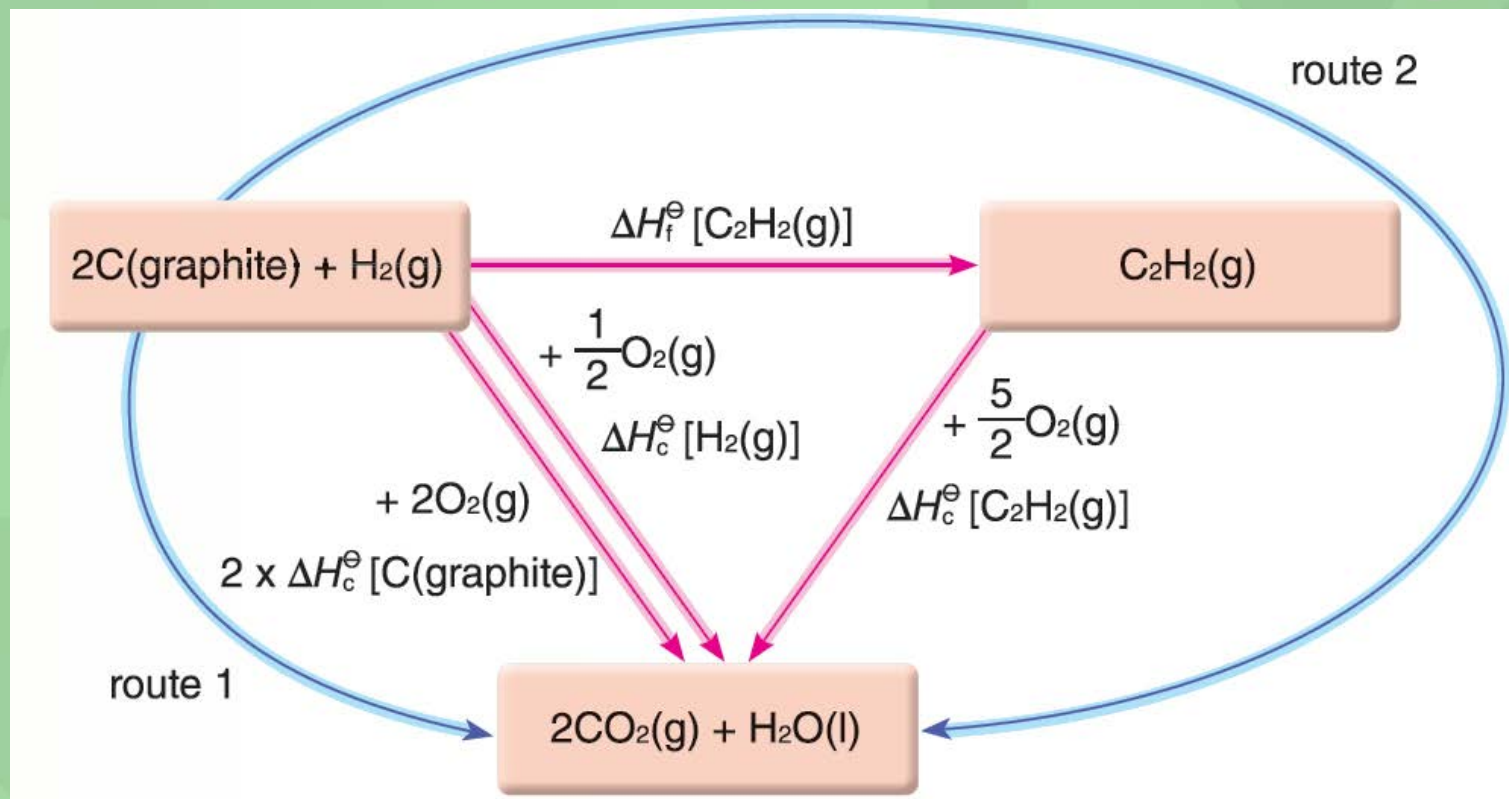
Calculate the standard enthalpy change of formation of ethyne ($\text{C}_2\text{H}_2(\text{g})$) from the following standard enthalpy changes of combustion:



These can be used to construct an enthalpy change cycle in which the equation for the reaction with the enthalpy change you want to find is along the top and arrows go down to the combustion products.



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)





35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

According to Hess's law:

$$\begin{aligned} 2 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + \Delta H_c^\ominus [\text{H}_2(\text{g})] &= \Delta H_f^\ominus [\text{C}_2\text{H}_2(\text{g})] + \Delta H_c^\ominus [\text{C}_2\text{H}_2(\text{g})] \\ \Delta H_f^\ominus [\text{C}_2\text{H}_2(\text{g})] &= 2 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + \Delta H_c^\ominus [\text{H}_2(\text{g})] - \Delta H_c^\ominus [\text{C}_2\text{H}_2(\text{g})] \\ &= [2(-394) + (-286) - (-1\,300)] \text{ kJ mol}^{-1} = +226 \text{ kJ mol}^{-1} \end{aligned}$$

Method 2 Solving by manipulating equations

$$\Delta H_f^\ominus [\text{compound}] = \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{compound}]$$

$$\begin{aligned} \Delta H_f^\ominus [\text{C}_2\text{H}_2(\text{g})] &= 2 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + \Delta H_c^\ominus [\text{H}_2(\text{g})] - \Delta H_c^\ominus [\text{C}_2\text{H}_2(\text{g})] \\ &= [2(-394) + (-286) - (-1\,300)] \text{ kJ mol}^{-1} = +226 \text{ kJ mol}^{-1} \end{aligned}$$

\therefore the standard enthalpy change of the formation of ethyne is $+226 \text{ kJ mol}^{-1}$.



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

Practice 35.3

1 The enthalpy change of formation of $\text{CH}_3\text{CHO}(\text{l})$ can be calculated by drawing an enthalpy change cycle based on the enthalpy changes of combustion ΔH_c^\ominus of $\text{CH}_3\text{CHO}(\text{l})$, $\text{C}(\text{graphite})$ and $\text{H}_2(\text{g})$, and applying Hess's law.

a) Draw, with labels, this enthalpy change cycle.

b) The standard enthalpy changes of combustion ΔH_c^\ominus of $\text{CH}_3\text{CHO}(\text{l})$, $\text{C}(\text{graphite})$ and $\text{H}_2(\text{g})$ are as follows:

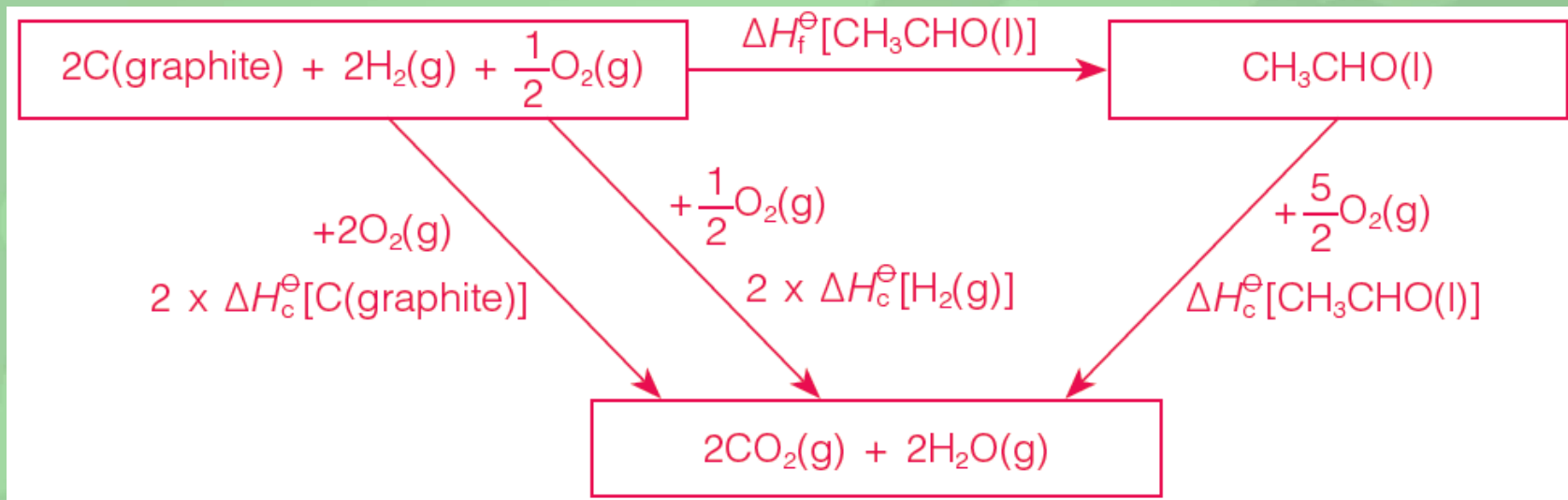
Substance	ΔH_c^\ominus (kJ mol^{-1})
$\text{CH}_3\text{CHO}(\text{l})$	-1 167
$\text{C}(\text{graphite})$	-394
$\text{H}_2(\text{g})$	-286

Calculate the standard enthalpy change of formation of $\text{CH}_3\text{CHO}(\text{l})$.



35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

a)



b)

According to Hess's law:

$$\begin{aligned}\Delta H_f^\ominus[\text{CH}_3\text{CHO}(\text{l})] + \Delta H_c^\ominus[\text{CH}_3\text{CHO}(\text{l})] &= 2 \times \Delta H_c^\ominus[\text{C}(\text{graphite})] + 2 \times \Delta H_c^\ominus[\text{H}_2(\text{g})] \\ \Delta H_f^\ominus[\text{CH}_3\text{CHO}(\text{l})] &= 2 \times \Delta H_c^\ominus[\text{C}(\text{graphite})] + 2 \times \Delta H_c^\ominus[\text{H}_2(\text{g})] - \Delta H_c^\ominus[\text{CH}_3\text{CHO}(\text{l})] \\ &= [2(-394) + 2(-286) - (-1\,167)] \text{ kJ mol}^{-1} \\ &= -193 \text{ kJ mol}^{-1}\end{aligned}$$

\therefore the enthalpy change of formation of $\text{CH}_3\text{CHO}(\text{l})$ is -193 kJ mol^{-1} .

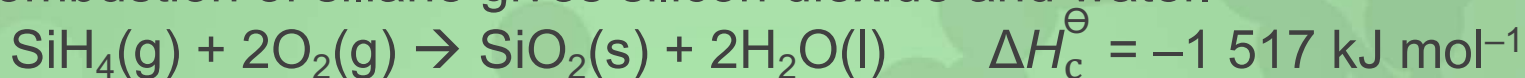


35.3 Enthalpy change of formation from enthalpy changes of combustion (p.57)

2 Silane (SiH_4) is formed in the reaction of silicon and hydrogen.



The combustion of silane gives silicon dioxide and water.



The standard enthalpy changes of combustion of silicon and hydrogen are -911 kJ mol^{-1} and -286 kJ mol^{-1} respectively.

Calculate the standard enthalpy change of formation of silane.

$$\begin{aligned} \Delta H_f^\ominus [\text{SiH}_4(\text{g})] &= \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{SiH}_4(\text{g})] \\ &= [(-911) + 2(-286) - (-1\,517)] \text{ kJ mol}^{-1} \\ &= +34 \text{ kJ mol}^{-1} \end{aligned}$$

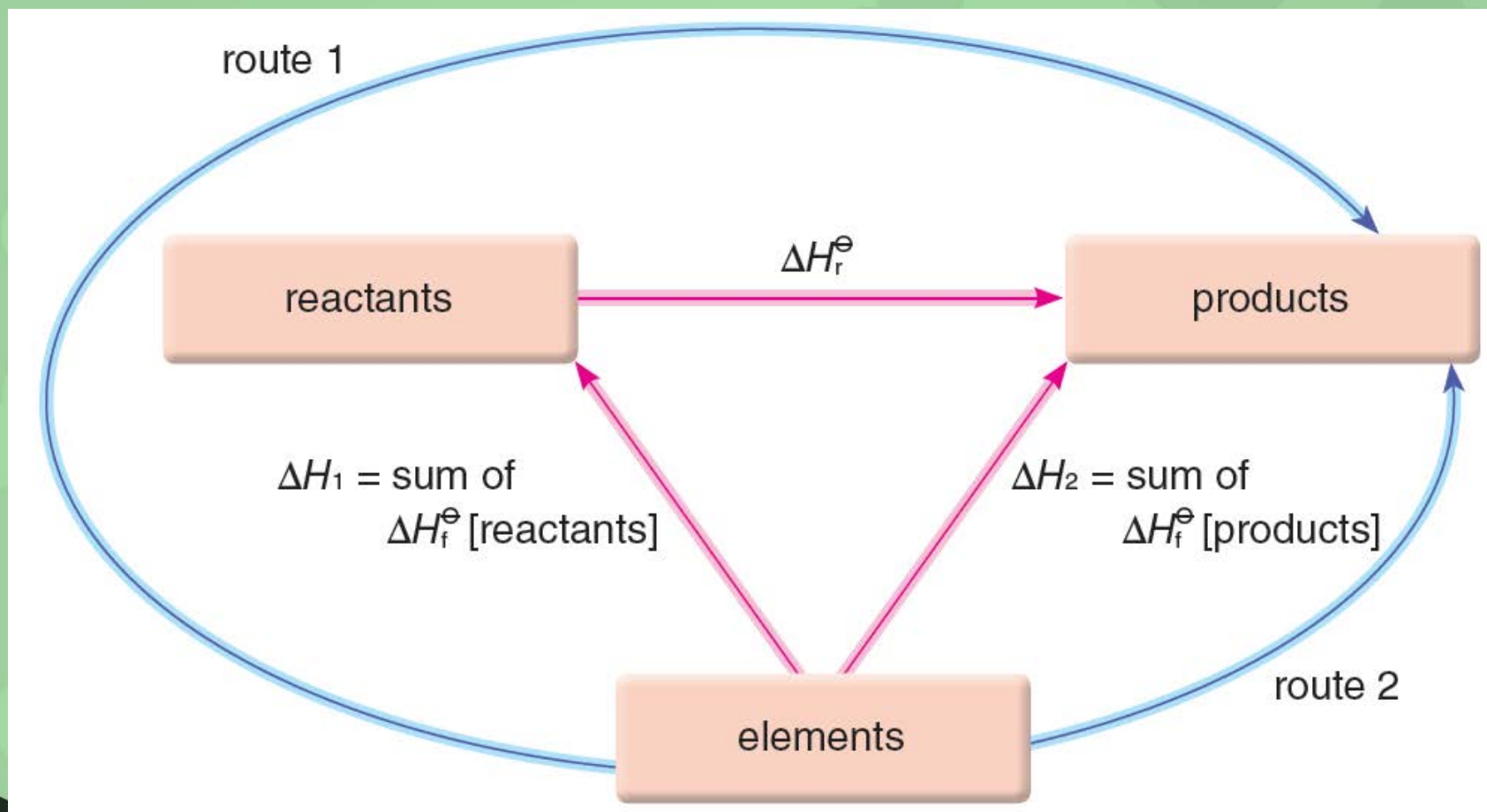


35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

- ◆ Data books contain data of standard enthalpy changes of formation for both inorganic and organic compounds. The great value of these data is that they allow chemists to calculate the standard enthalpy change for any reaction involving the substances listed in the tables.
- ◆ Thanks to Hess's law, it is easy to calculate the standard enthalpy change of a reaction from tabulated values of standard enthalpy changes of formation.



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)





35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

- According to Hess's law: $\Delta H_1 + \Delta H_r^\ominus = \Delta H_2$

Rearranging gives: $\Delta H_r^\ominus = \Delta H_2 - \Delta H_1$

So

$$\Delta H_r^\ominus = \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}]$$



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

Q (Example 35.5)

Iron is extracted from its ore according to the chemical equation below:



Calculate ΔH_r^\ominus from the ΔH_f^\ominus values below:

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
Fe ₂ O ₃ (s)	-826
CO(g)	-111
Fe(s)	0
CO ₂ (g)	-394

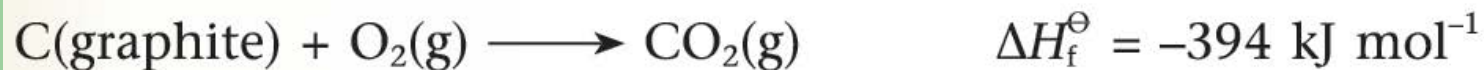
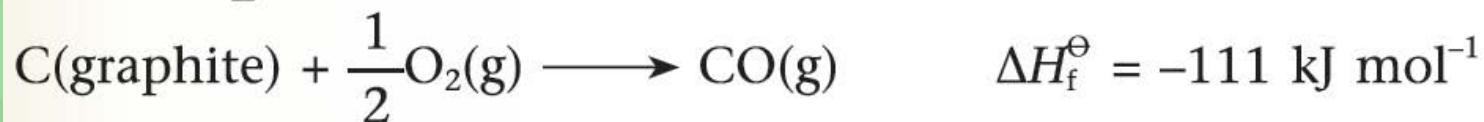
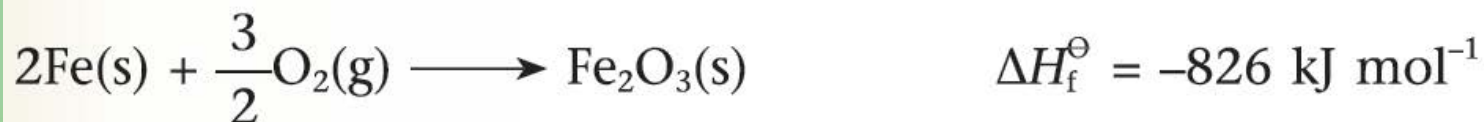


35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

A

Method 1 Solving by an enthalpy change cycle

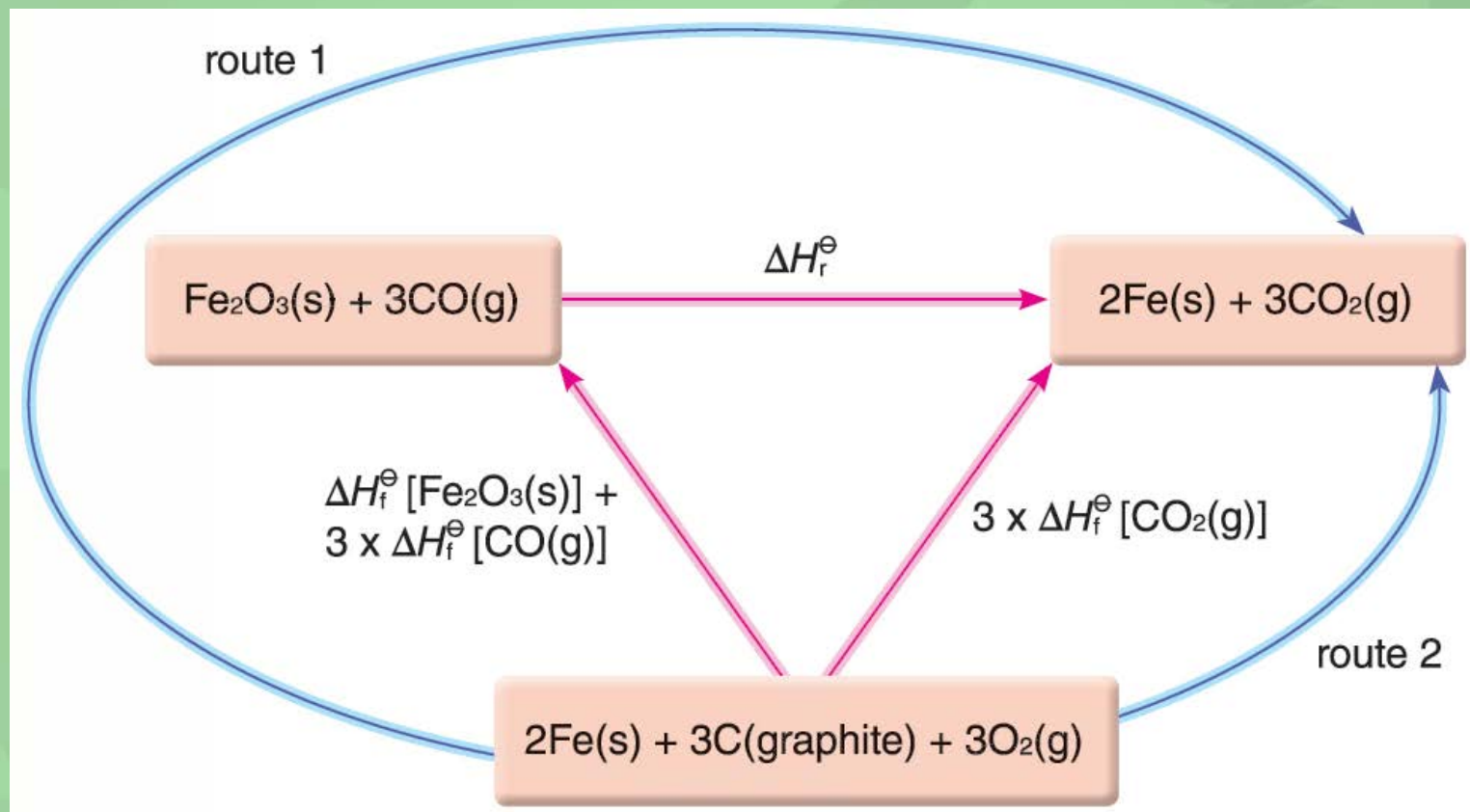
Given the ΔH_f^\ominus values, you can write out the thermochemical equations:



Use these to construct an enthalpy change cycle in which the equation for the reaction with the enthalpy change you want to find is along the top and arrows go up from the elements to the compounds.



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)





35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

According to Hess's law:

$$\Delta H_f^\ominus [\text{Fe}_2\text{O}_3(\text{s})] + 3 \times \Delta H_f^\ominus [\text{CO}(\text{g})] + \Delta H_r^\ominus = 3 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})]$$

$$\begin{aligned}\Delta H_r^\ominus &= 3 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] - \Delta H_f^\ominus [\text{Fe}_2\text{O}_3(\text{s})] + 3 \times \Delta H_f^\ominus [\text{CO}(\text{g})] \\ &= [(-394) - (-826) - 3(-111)] \text{ kJ mol}^{-1} = -23 \text{ kJ mol}^{-1}\end{aligned}$$

Method 2 Solving by manipulating equations

$$\Delta H_r^\ominus = \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}]$$

$$\begin{aligned}\Delta H_r^\ominus &= 3 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] - \Delta H_f^\ominus [\text{Fe}_2\text{O}_3(\text{s})] + 3 \times \Delta H_f^\ominus [\text{CO}(\text{g})] \\ &= [(-394) - (-826) - 3(-111)] \text{ kJ mol}^{-1} = -23 \text{ kJ mol}^{-1}\end{aligned}$$

\therefore the standard enthalpy change of this reaction is -23 kJ mol^{-1} .

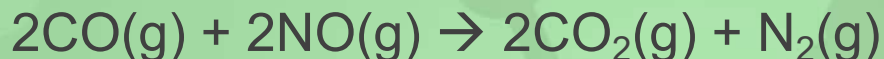


35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

Practice 35.4

1 Catalytic converters are installed in motor cars to convert nitrogen monoxide and carbon monoxide to less harmful substances.

The equation for the reaction involved in the conversion is shown below:



Constructing an enthalpy change cycle based on the enthalpy changes of formation ΔH_f^\ominus of CO(g) , NO(g) and $\text{CO}_2\text{(g)}$ and applying Hess's law can give the enthalpy change of the above reaction.

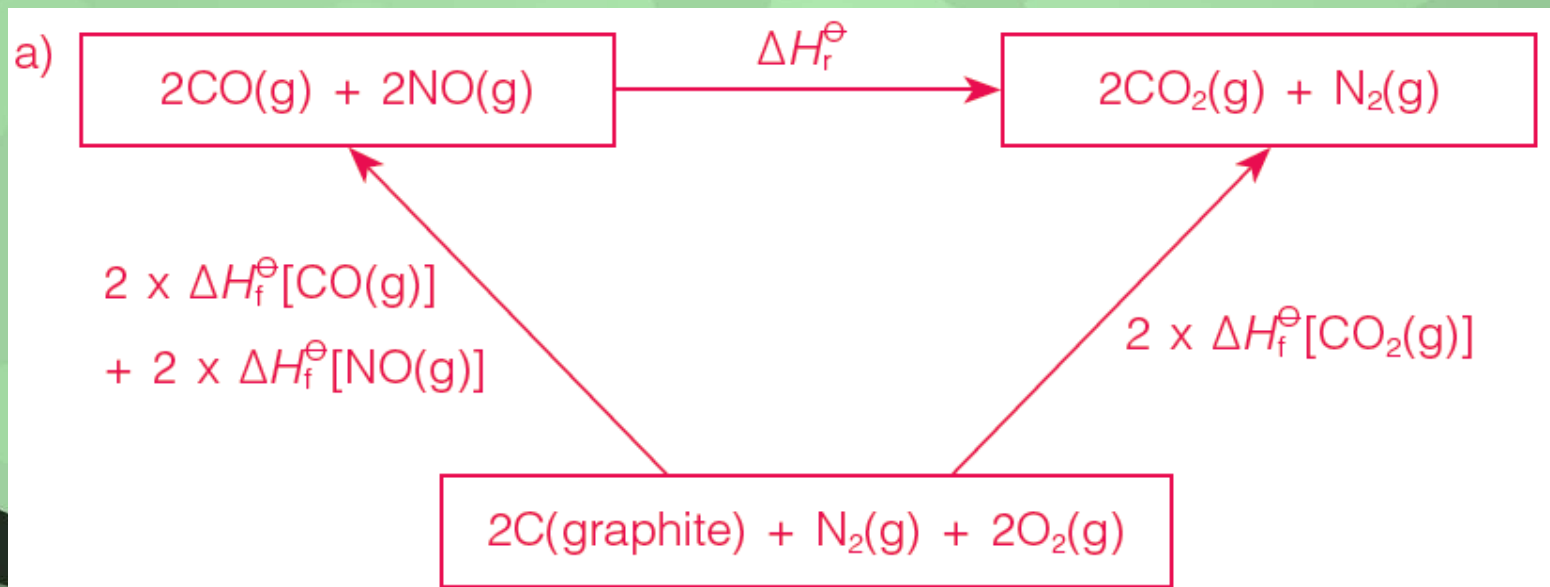
a) Draw, with labels, this enthalpy change cycle.

b) The standard enthalpy changes of formation of CO(g) , NO(g) and $\text{CO}_2\text{(g)}$ are as follows. Calculate the standard enthalpy change of the above reaction.



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
CO(g)	-111
NO(g)	+90
CO ₂ (g)	-394





35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

b) According to Hess's law:

$$2 \times \Delta H_f^\ominus[\text{CO(g)}] + 2 \times \Delta H_f^\ominus[\text{NO(g)}] + \Delta H_r^\ominus = 2 \times \Delta H_f^\ominus[\text{CO}_2\text{(g)}]$$

$$\Delta H_r^\ominus = 2 \times \Delta H_f^\ominus[\text{CO}_2\text{(g)}] - 2 \times \Delta H_f^\ominus[\text{CO(g)}] - 2 \times \Delta H_f^\ominus[\text{NO(g)}]$$

$$= [2(-394) - 2(-111) - 2(+90)] \text{ kJ mol}^{-1}$$

$$= -746 \text{ kJ mol}^{-1}$$

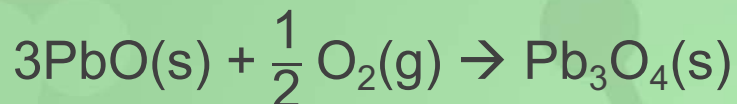
\therefore the standard enthalpy change of the reaction is -746 kJ mol^{-1} .



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

2 Lead forms several solid oxides, the most common of which are PbO, PbO₂ and Pb₃O₄.

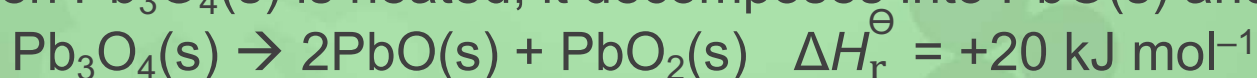
a) Pb₃O₄(s) can be made by heating PbO(s) in oxygen.



Calculate the standard enthalpy change of this reaction.

(Standard enthalpy change of formation of PbO(s) = -219 kJ mol^{-1} ;
standard enthalpy change of formation of Pb₃O₄(s) = -735 kJ mol^{-1})

b) When Pb₃O₄(s) is heated, it decomposes into PbO(s) and PbO₂(s).



Using this information, together with the data supplied in (a), calculate the standard enthalpy change of formation of PbO₂(s).



35.4 Enthalpy change of reaction from enthalpy changes of formation (p.61)

$$\begin{aligned}\text{a) } \Delta H_r^\ominus &= \Delta H_f^\ominus[\text{Pb}_3\text{O}_4(\text{s})] - 3 \times \Delta H_f^\ominus[\text{PbO}(\text{s})] \\ &= [(-735) - 3(-219)] \text{ kJ mol}^{-1} \\ &= -78 \text{ kJ mol}^{-1}\end{aligned}$$

\therefore the standard enthalpy change of the reaction is -78 kJ mol^{-1} .

$$\begin{aligned}\text{b) } \Delta H_r^\ominus &= 2 \times \Delta H_f^\ominus[\text{PbO}(\text{s})] + \Delta H_f^\ominus[\text{PbO}_2(\text{s})] - \Delta H_f^\ominus[\text{Pb}_3\text{O}_4(\text{s})] \\ \Delta H_f^\ominus[\text{PbO}_2(\text{s})] &= \Delta H_r^\ominus - 2 \times \Delta H_f^\ominus[\text{PbO}(\text{s})] + \Delta H_f^\ominus[\text{Pb}_3\text{O}_4(\text{s})] \\ &= [+20 - 2(-219) + (-735)] \text{ kJ mol}^{-1} \\ &= -277 \text{ kJ mol}^{-1}\end{aligned}$$

\therefore the standard enthalpy change of formation of $\text{PbO}_2(\text{s})$ is -277 kJ mol^{-1} .



Key terms (p.65)

Hess's law	赫斯定律	Enthalpy change cycle	焓變循環



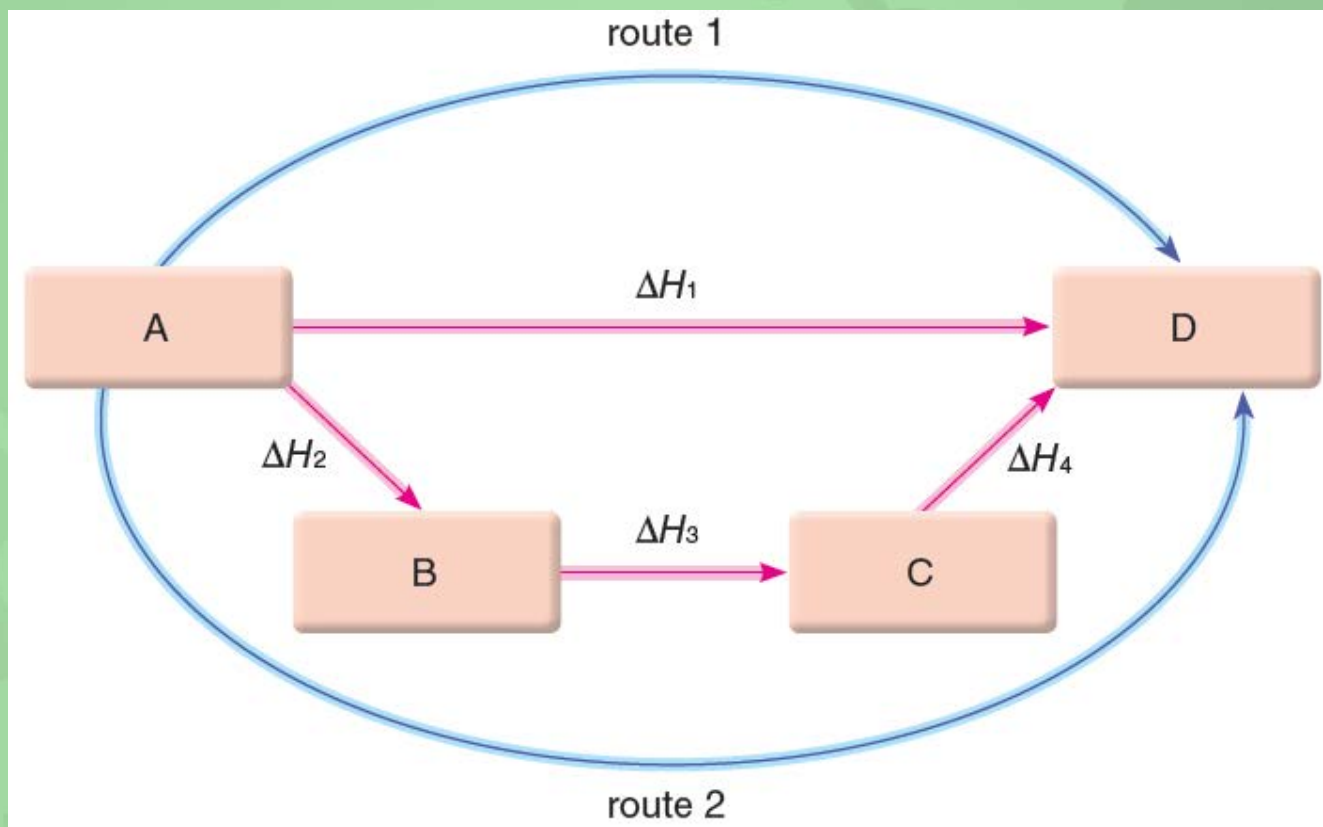
Summary (p.66)

- 1 Hess's law states that 'the total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same'.



Summary (p.66)

- 2 The enthalpy change cycle below helps illustrate the Hess's law. In symbols: $\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$





Summary (p.66)

- 3 Hess's law can be used to determine enthalpy changes of reactions that are difficult to be determined directly by experiments.
- 4 $\Delta H_f^\ominus [\text{compound}] = \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{compound}]$
- 5 $\Delta H_r^\ominus = \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}]$



Unit Exercise (p.67)

Note: Questions are rated according to ascending level of difficulty (from 1 to 5):



question targeted at level 3 and above;



question targeted at level 4 and above;



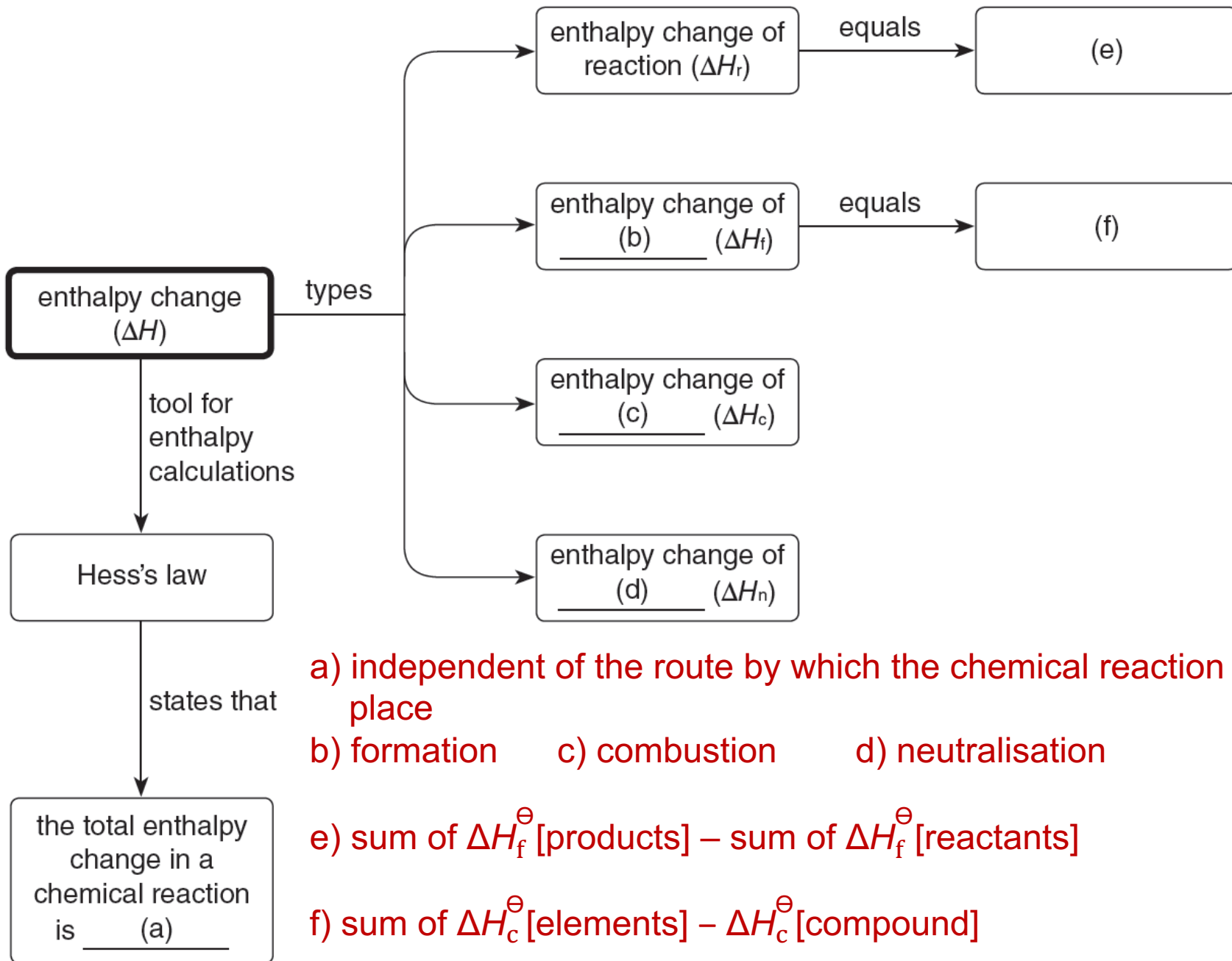
question targeted at level 5.

' * ' indicates 1 mark is given for effective communication.

Unit Exercise (p.67)

PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.

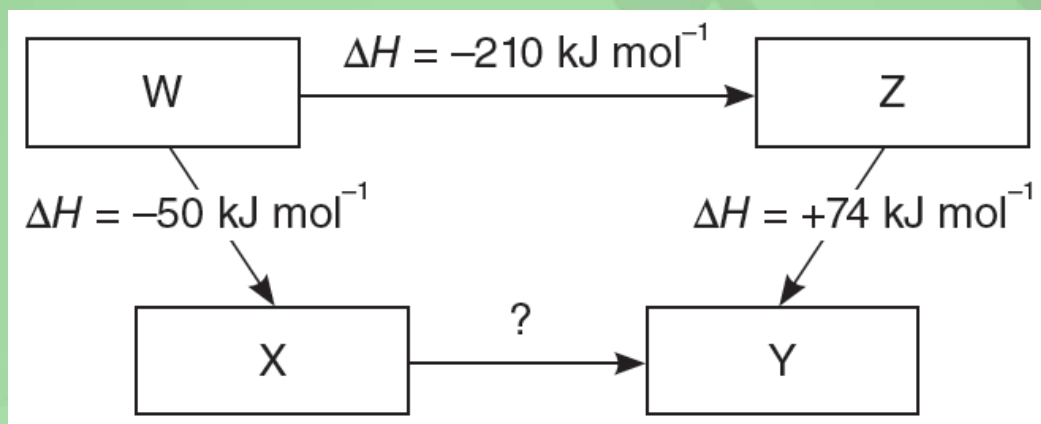




Unit Exercise (p.67)

PART II MULTIPLE CHOICE QUESTIONS

2 Consider the enthalpy change cycle shown below:



Answer: A

According to Hess's law, the ΔH value of the reaction X to Y is

- A -86 kJ mol^{-1}
- B -234 kJ mol^{-1}
- C $+86 \text{ kJ mol}^{-1}$
- D $+234 \text{ kJ mol}^{-1}$

Explanation:

According to Hess's law:

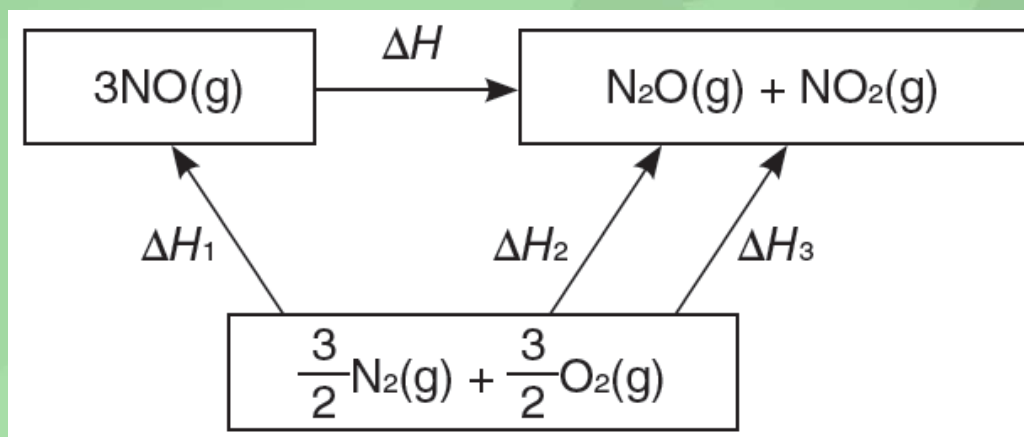
$$-50 \text{ kJ mol}^{-1} + \Delta H = -210 \text{ kJ mol}^{-1} + 74 \text{ kJ mol}^{-1}$$

$$\Delta H = -86 \text{ kJ mol}^{-1}$$



Unit Exercise (p.67)

3 The enthalpy change cycle for a decomposition of NO(g) is shown below:



Answer: A

Which of the following is correct?

A $\Delta H = \Delta H_2 + \Delta H_3 - \Delta H_1$

B $\Delta H = \Delta H_2 + \Delta H_3 + \Delta H_1$

C $\Delta H = \Delta H_2 - \Delta H_3 - \Delta H_1$

D $\Delta H = \Delta H_2 - \Delta H_3 + \Delta H_1$

Explanation:

According to Hess's law:

$$\Delta H_1 + \Delta H = \Delta H_2 + \Delta H_3$$

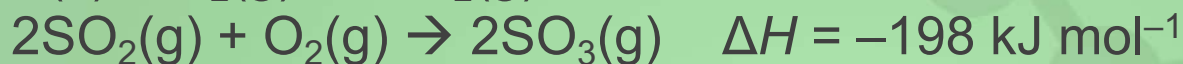
$$\Delta H = \Delta H_2 + \Delta H_3 - \Delta H_1$$



Unit Exercise (p.67)



4 The equations and enthalpy changes of two reactions used in the manufacture of sulphuric acid are shown below:



What is the enthalpy change of the reaction below?



A -99 kJ mol^{-1}

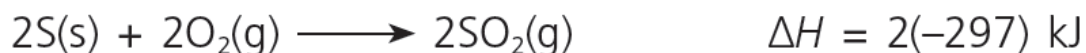
B -495 kJ mol^{-1}

C -693 kJ mol^{-1}

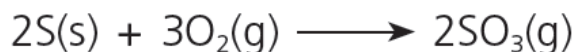
D -792 kJ mol^{-1}

Answer: D

Multiply equation (1) by 2 to give equation (1)'.



Add together (1)' and (2):



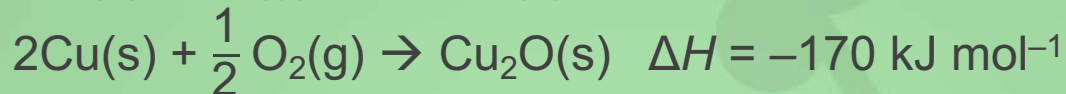
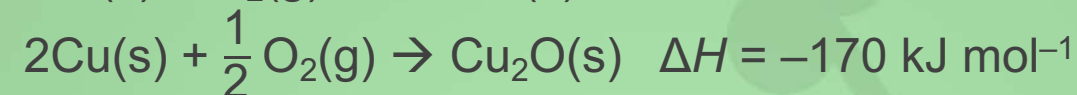
According to Hess's law:

$$\begin{aligned} \Delta H_r &= [2(-297) + (-198)] \text{ kJ mol}^{-1} \\ &= -792 \text{ kJ mol}^{-1} \end{aligned}$$



Unit Exercise (p.67)

5 Copper reacts with oxygen to form two oxides according to the equations below:



What is the enthalpy change of the reaction below?



A -284 kJ mol^{-1}

B -142 kJ mol^{-1}

C $+142 \text{ kJ mol}^{-1}$

D $+284 \text{ kJ mol}^{-1}$

Answer: D

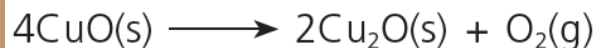
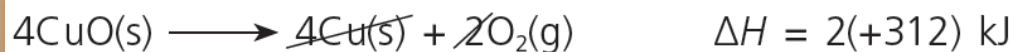
Multiply equation (1) by 2 and reverse it to give equation (1)'.



Multiply equation (2) by 2 to give equation (2)'.



Add together (1)' and (2)':



According to Hess's law:

$$\begin{aligned} \Delta H_r &= [2(+312) + 2(-170)] \text{ kJ mol}^{-1} \\ &= +284 \text{ kJ mol}^{-1} \end{aligned}$$

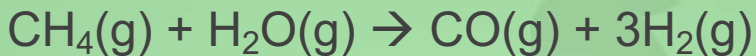


Unit Exercise (p.67)

6 Methanol can be produced from methane by a two-step process.



In the first step, methane is reacted with steam to produce carbon monoxide and hydrogen.



Consider the following enthalpy changes:

<u>Reaction</u>	<u>Enthalpy change</u>
$\text{CO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g})$	-283 kJ mol^{-1}
$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{H}_2\text{O}(\text{g})$	-242 kJ mol^{-1}
$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$	-803 kJ mol^{-1}

Answer: B

What is the enthalpy change of the reaction between methane and steam?

A $+360 \text{ kJ mol}^{-1}$

B $+206 \text{ kJ mol}^{-1}$

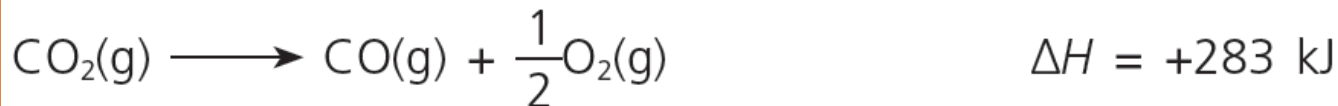
C -206 kJ mol^{-1}

D -360 kJ mol^{-1}

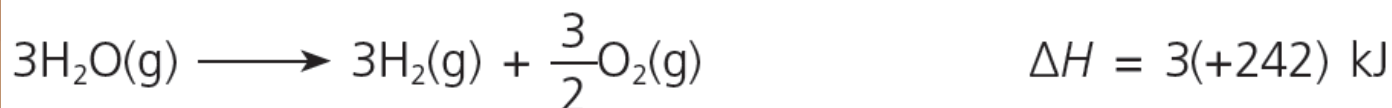


Unit Exercise (p.67)

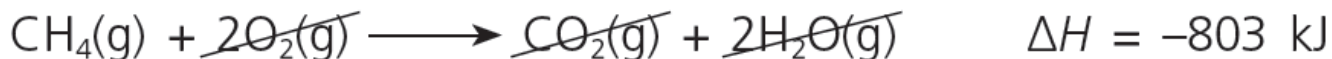
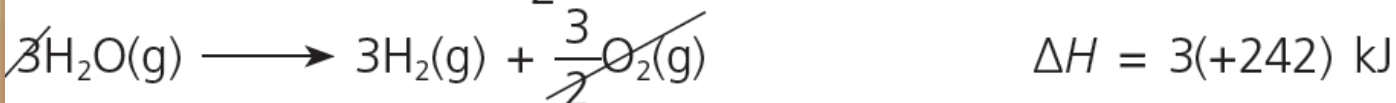
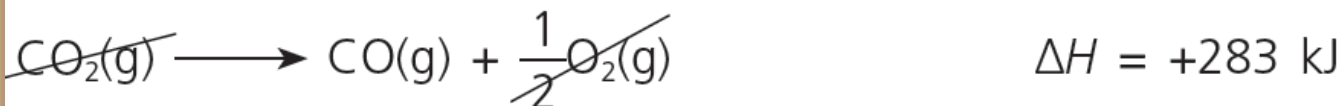
Reverse equation (1) to give equation (1)'.



Multiply equation (2) by 3 and reverse it to give equation (2)'.



Add together (1)', (2)' and (3):



According to Hess's law:

$$\begin{aligned}\Delta H_r &= [+283 + 3(+242) + (-803)] \text{ kJ mol}^{-1} \\ &= +206 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)

7 Given the following data:



Substance	ΔH_c^\ominus (kJ mol ⁻¹)
C(graphite)	-394
H ₂ (g)	-286
CH ₃ CH ₂ CH ₂ OH(l)	-2 021

What is the standard enthalpy change of formation of CH₃CH₂CH₂OH(l)?

A +1 341 kJ mol⁻¹

B -1 341 kJ mol⁻¹

C -305 kJ mol⁻¹

D +305 kJ mol⁻¹

Answer: C

$$\begin{aligned}\Delta H_f^\ominus[\text{CH}_3\text{CH}_2\text{CH}_2\text{OH(l)}] &= \text{sum of } \Delta H_c^\ominus[\text{elements}] - \Delta H_c^\ominus[\text{CH}_3\text{CH}_2\text{CH}_2\text{OH(l)}] \\ &= 3 \times \Delta H_c^\ominus[\text{C(graphite)}] + 4 \times \Delta H_c^\ominus[\text{H}_2(\text{g})] - \Delta H_c^\ominus[\text{CH}_3\text{CH}_2\text{CH}_2\text{OH(l)}] \\ &= [3(-394) + 4(-286) - (-2\,021)] \text{ kJ mol}^{-1} \\ &= -305 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)

8 Consider the reaction of $\text{MnO}_2(\text{s})$ with $\text{C}(\text{s})$.



The standard enthalpy changes of formation of $\text{CO}_2(\text{g})$ and $\text{MnO}_2(\text{s})$ are as follows:

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
$\text{CO}_2(\text{g})$	-394
$\text{MnO}_2(\text{s})$	-520

What is the standard enthalpy change of the above reaction?

Answer: B

- A +914 kJ mol⁻¹
- B +126 kJ mol⁻¹
- C -126 kJ mol⁻¹
- D -914 kJ mol⁻¹

$$\begin{aligned}\Delta H_r^\ominus &= \Delta H_f^\ominus[\text{CO}_2(\text{g})] - \Delta H_f^\ominus[\text{MnO}_2(\text{s})] \\ &= [(-394) - (-520)] \text{ kJ mol}^{-1} \\ &= +126 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)



9 Under standard conditions, complete combustion of 0.0180 mole of ethene ($\text{C}_2\text{H}_4(\text{g})$) gives 25.4 kJ of heat. What is the standard enthalpy change of formation of ethene?

(Standard enthalpy change of combustion of $\text{H}_2(\text{g}) = -286 \text{ kJ mol}^{-1}$;
standard enthalpy change of combustion of $\text{C}(\text{graphite}) = -394 \text{ kJ mol}^{-1}$)

A -705 kJ mol^{-1}

B -51 kJ mol^{-1}

C $+51 \text{ kJ mol}^{-1}$

D $+705 \text{ kJ mol}^{-1}$

Answer: C

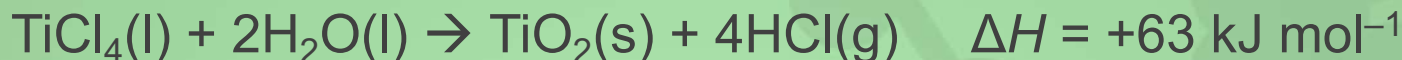
$$\begin{aligned}\Delta H_c^\ominus[\text{C}_2\text{H}_4(\text{g})] &= \frac{-25.4 \text{ kJ}}{0.0180 \text{ mol}} \\ &= -1\,411 \text{ kJ mol}^{-1}\end{aligned}$$

$$\begin{aligned}\Delta H_f^\ominus[\text{C}_2\text{H}_4(\text{g})] &= \text{sum of } \Delta H_c^\ominus[\text{elements}] - \Delta H_c^\ominus[\text{C}_2\text{H}_4(\text{g})] \\ &= 2 \times \Delta H_c^\ominus[\text{C}(\text{graphite})] + 2 \times \Delta H_c^\ominus[\text{H}_2(\text{g})] - \Delta H_c^\ominus[\text{C}_2\text{H}_4(\text{g})] \\ &= [2(-394) + 2(-286) - (-1\,411)] \text{ kJ mol}^{-1} \\ &= +51 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)

- 10 Consider the reaction between $\text{TiCl}_4(\text{l})$ and $\text{H}_2\text{O}(\text{l})$ under certain conditions.



The table below shows the enthalpy changes of formation under the same conditions.

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
$\text{H}_2\text{O}(\text{l})$	-286
$\text{TiO}_2(\text{s})$	-945
$\text{HCl}(\text{g})$	-92

Answer: D

What is the enthalpy change of formation of $\text{TiCl}_4(\text{l})$ under the same conditions?

- A +804 kJ mol⁻¹
- B +528 kJ mol⁻¹
- C -528 kJ mol⁻¹
- D -804 kJ mol⁻¹

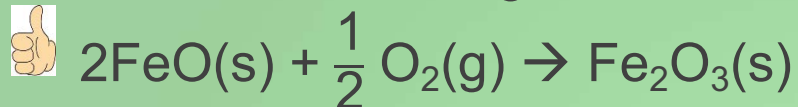
$$\Delta H_r = \Delta H_f[\text{TiO}_2(\text{s})] + 4 \times \Delta H_f[\text{HCl}(\text{g})] - \Delta H_f[\text{TiCl}_4(\text{l})] - 2 \times \Delta H_f[\text{H}_2\text{O}(\text{l})]$$

$$\begin{aligned}\Delta H_f[\text{TiCl}_4(\text{l})] &= \Delta H_f[\text{TiO}_2(\text{s})] + 4 \times \Delta H_f[\text{HCl}(\text{g})] - 2 \times \Delta H_f[\text{H}_2\text{O}(\text{l})] - \Delta H_r \\ &= [(-945) + 4(-92) - 2(-286) - (+63)] \text{ kJ mol}^{-1} \\ &= -804 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)

11 Given the following data:



$$\Delta H_r^\ominus = -280 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\ominus [\text{Fe}_2\text{O}_3\text{(s)}] = -824 \text{ kJ mol}^{-1}$$

What is the standard enthalpy change of formation of FeO(s)?

A $+544 \text{ kJ mol}^{-1}$

B $+272 \text{ kJ mol}^{-1}$

C -272 kJ mol^{-1}

D -544 kJ mol^{-1}

Answer: C

$$\Delta H_r^\ominus = \Delta H_f^\ominus [\text{Fe}_2\text{O}_3\text{(s)}] - 2 \times \Delta H_f^\ominus [\text{FeO(s)}]$$

$$2 \times \Delta H_f^\ominus [\text{FeO(s)}] = \Delta H_f^\ominus [\text{Fe}_2\text{O}_3\text{(s)}] - \Delta H_r^\ominus$$

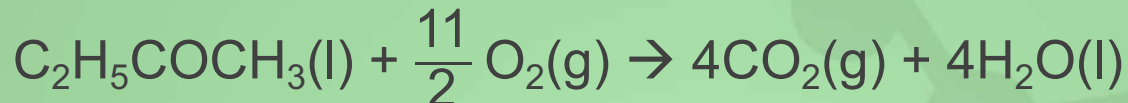
$$\begin{aligned}\Delta H_f^\ominus [\text{FeO(s)}] &= \frac{1}{2} [(-824) - (-280)] \text{ kJ mol}^{-1} \\ &= -272 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)



12 Under standard conditions, complete combustion of 0.0400 mole of butanone gives 97.6 kJ of heat.



Answer: A

What is the standard enthalpy change of formation of butanone?
(Standard enthalpy change of formation of $\text{H}_2\text{O}(\text{l}) = -286 \text{ kJ mol}^{-1}$;
standard enthalpy change of formation of $\text{CO}_2(\text{g}) = -394 \text{ kJ mol}^{-1}$)

- A -280 kJ mol^{-1}
- B $+280 \text{ kJ mol}^{-1}$
- C $-2\,010 \text{ kJ mol}^{-1}$
- D $+2\,010 \text{ kJ mol}^{-1}$

$$\begin{aligned}\Delta H_c^\ominus[\text{C}_2\text{H}_5\text{COCH}_3(\text{l})] &= \frac{-97.6 \text{ kJ}}{0.0400 \text{ mol}} \\ &= -2\,440 \text{ kJ mol}^{-1}\end{aligned}$$

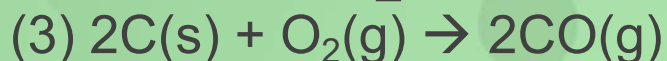
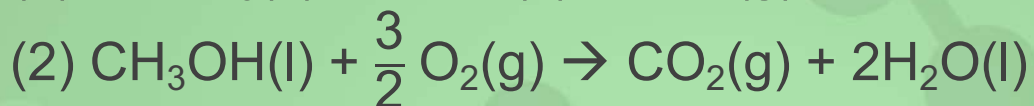
$$\Delta H_c^\ominus = 4 \times \Delta H_f^\ominus[\text{CO}_2(\text{g})] + 4 \times \Delta H_f^\ominus[\text{H}_2\text{O}(\text{l})] - \Delta H_f^\ominus[\text{C}_2\text{H}_5\text{COCH}_3(\text{l})]$$

$$\begin{aligned}\Delta H_f^\ominus[\text{C}_2\text{H}_5\text{COCH}_3(\text{l})] &= 4 \times \Delta H_f^\ominus[\text{CO}_2(\text{g})] + 4 \times \Delta H_f^\ominus[\text{H}_2\text{O}(\text{l})] - \Delta H_c^\ominus \\ &= [4(-394) + 4(-286) - (-2\,440)] \text{ kJ mol}^{-1} \\ &= -280 \text{ kJ mol}^{-1}\end{aligned}$$



Unit Exercise (p.67)

13 For which of the following reactions must the enthalpy change(s) be determined by INDIRECT methods?



A (1) only

B (2) only

C (1) and (3) only

D (2) and (3) only

(1) It is difficult to measure the heat taken in when heating a substance.

It is difficult to measure the temperature change of a solid.

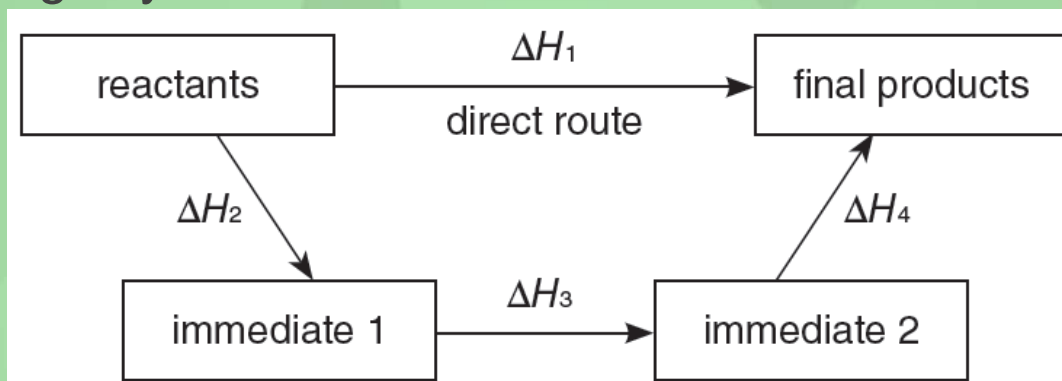
(3) In the reaction $2\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}(\text{g})$, further oxidation of $\text{CO}(\text{g})$ would occur as a side reaction (i.e. $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$).



Unit Exercise (p.67)

PART III STRUCTURED QUESTIONS

14 It is sometimes impossible to measure an enthalpy change directly. However, an enthalpy change can be measured indirectly using an enthalpy change cycle based on Hess's law.



Using the enthalpy change cycle,

a) write an equation relating ΔH_1 , ΔH_2 , ΔH_3 and ΔH_4 ;

$$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4 \quad (1)$$

b) explain your understanding of Hess's law.

Hess's law states that 'the total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same'. (1)



Unit Exercise (p.67)



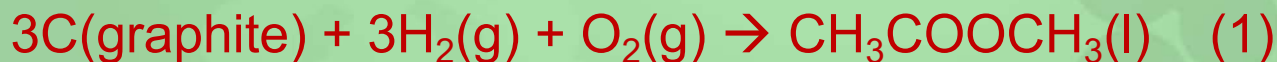
15 Standard enthalpy changes of combustion can be used to calculate the standard enthalpy change of formation of a compound by applying Hess's law.

a) State Hess's law.

Hess's law states that 'the total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same'. (1)

b) Methyl ethanoate is a liquid ester.

i) Write the chemical equation, including state symbols, for the reaction with an enthalpy change equals to the standard enthalpy change of formation of methyl ethanoate.





Unit Exercise (p.67)

- ii) Use the data below to calculate the standard enthalpy change of formation of methyl ethanoate.

Substance	ΔH_c^\ominus (kJ mol ⁻¹)
C(graphite)	-394
H ₂ (g)	-286
CH ₃ COOCH ₃ (l)	-1 592

$$\begin{aligned}\Delta H_f^\ominus [\text{CH}_3\text{COOCH}_3(\text{l})] &= \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{CH}_3\text{COOCH}_3(\text{l})] \\ &= 3 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + 3 \times \Delta H_c^\ominus [\text{H}_2(\text{g})] - \Delta H_c^\ominus [\text{CH}_3\text{COOCH}_3(\text{l})] \\ &= [3(-394) + 3(-286) - (-1\,592)] \text{ kJ mol}^{-1} \quad (1) \\ &= -448 \text{ kJ mol}^{-1}\end{aligned}$$

\therefore the standard enthalpy change of formation of methyl ethanoate is -448 kJ mol^{-1} . (1)



Unit Exercise (p.67)

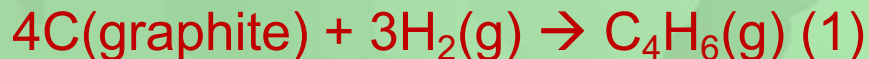
16 Buta-1,3-diene is a gaseous hydrocarbon with the molecular formula



C_4H_6 .

- a) Using buta-1,3-diene ($C_4H_6(g)$) as an example, state the meaning of the term 'standard enthalpy change of formation' with the aid of a chemical equation.

The standard enthalpy change of formation of a substance, ΔH_f^\ominus , is the enthalpy change when one mole of the substance is formed from its elements in their standard states under standard conditions. (1)



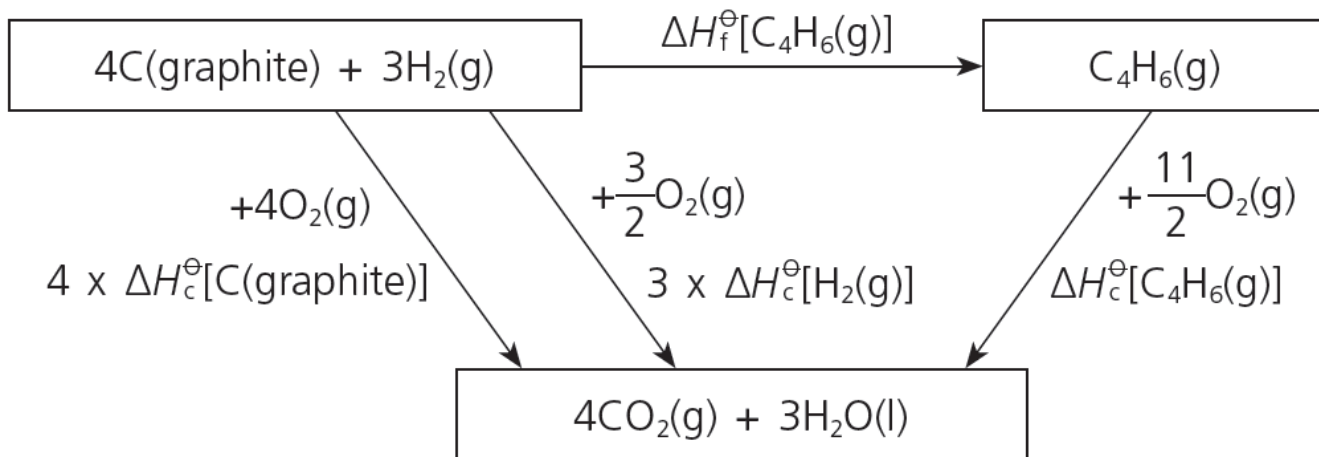
- b) Suggest why the enthalpy change of formation of $C_4H_6(g)$ CANNOT be determined directly by experiment.

$C_4H_6(g)$ CANNOT be formed directly from its constituent elements. / Carbon and hydrogen do not react to produce $C_4H_6(g)$ under normal conditions. (1)



Unit Exercise (p.67)

- c) Based on the enthalpy changes of combustion, ΔH_c , of $C_4H_6(g)$, $C(\text{graphite})$ and $H_2(g)$ to construct an enthalpy change cycle and applying Hess's law can give the enthalpy change of formation of $C_4H_6(g)$.
- i) Draw, with labels, this enthalpy change cycle.



(2)



Unit Exercise (p.67)

ii) Use the data below to calculate the standard enthalpy change of formation of buta-1,3-diene.

Substance	ΔH_c^\ominus (kJ mol ⁻¹)
C(graphite)	-394
H ₂ (g)	-286
C ₄ H ₆ (g)	-2 542

According to Hess's law:

$$\Delta H_f^\ominus [\text{C}_4\text{H}_6(\text{g})] + \Delta H_c^\ominus [\text{C}_4\text{H}_6(\text{g})] = 4 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + 3 \times \Delta H_c^\ominus [\text{H}_2(\text{g})]$$

$$\Delta H_f^\ominus [\text{C}_4\text{H}_6(\text{g})] = 4 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + 3 \times \Delta H_c^\ominus [\text{H}_2(\text{g})] - \Delta H_c^\ominus [\text{C}_4\text{H}_6(\text{g})]$$
$$= [4(-394) + 3(-286) - (-2\,542)] \text{ kJ mol}^{-1} \quad (1)$$

$$= +108 \text{ kJ mol}^{-1} \quad (1)$$

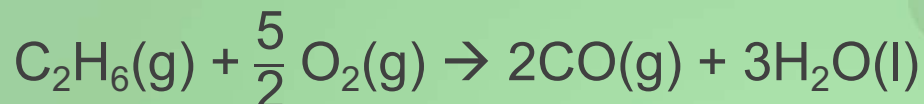
\therefore the standard enthalpy change of formation of C₄H₆(g) is +108 kJ mol⁻¹.



Unit Exercise (p.67)



17 Ethane burns in a limited supply of oxygen to give carbon monoxide and water.



The enthalpy change for this reaction cannot be determined directly, but can be found using the enthalpy changes of combustion ΔH_c of ethane and carbon monoxide, together with Hess's law.

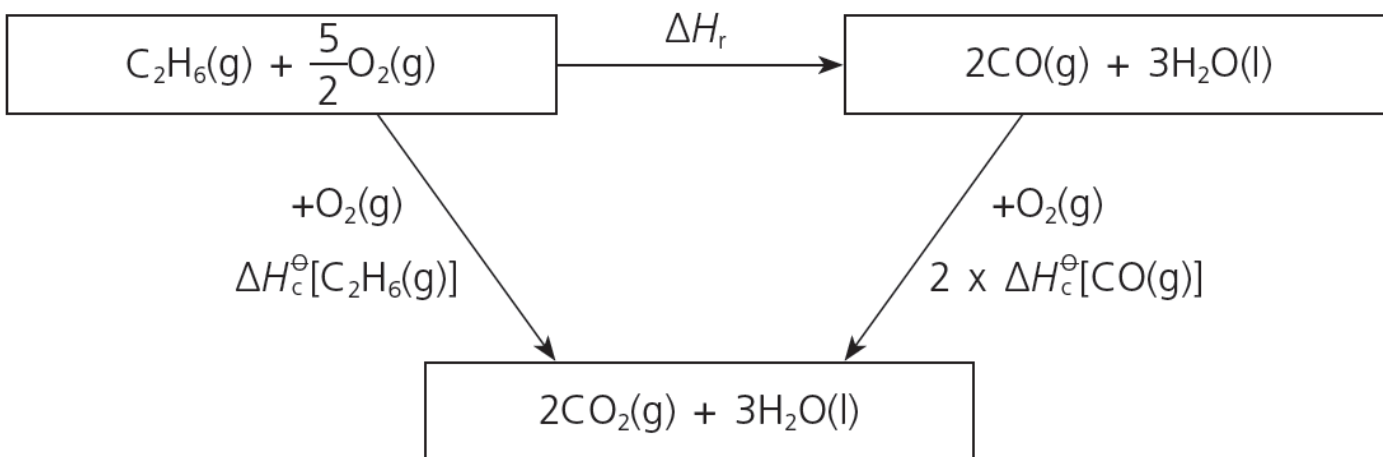
a) Explain why the enthalpy change of the above reaction CANNOT be determined directly.

It is impossible to stop the reaction at the formation of CO. The reaction produces some CO₂, and may produce some carbon / soot. (1)
The exact amounts of ethane to oxygen CANNOT be determined.



Unit Exercise (p.67)

- b) Draw an enthalpy change cycle, with labels, which would enable you to calculate the enthalpy change of combustion of ethane to carbon monoxide from ΔH_c^\ominus of ethane and carbon monoxide.



(2)



Unit Exercise (p.67)

c) Calculate the enthalpy change of the above reaction.

(Standard enthalpy change of combustion of ethane = $-1\,411\text{ kJ mol}^{-1}$;
standard enthalpy change of combustion of carbon monoxide = -283 kJ mol^{-1})

According to Hess's law:

$$\Delta H_r + 2 \times \Delta H_c^\ominus [\text{CO(g)}] = \Delta H_c [\text{C}_2\text{H}_6(\text{g})]$$

$$\Delta H_r = \Delta H_c^\ominus [\text{C}_2\text{H}_6(\text{g})] - 2 \times \Delta H_c^\ominus [\text{CO(g)}]$$

$$= [(-1\,411) - 2(-283)]\text{ kJ mol}^{-1} \text{ (1)}$$

$$= -845\text{ kJ mol}^{-1} \text{ (1)}$$

\therefore the enthalpy change of the reaction is -845 kJ mol^{-1} .



Unit Exercise (p.67)



18 A rocket uses the reaction between dinitrogen tetroxide (N_2O_4) and methylhydrazine (CH_3NHNH_2) to release a large amount of heat.



$$\Delta H_r = -4\,596 \text{ kJ mol}^{-1}$$

Some standard enthalpy changes of formation are shown in the table.

<u>Substance</u>	<u>ΔH_f^\ominus (kJ mol⁻¹)</u>
$\text{CH}_3\text{NHNH}_2(\text{l})$	+54
$\text{CO}_2(\text{g})$	-394
$\text{H}_2\text{O}(\text{g})$	-242



Unit Exercise (p.67)

- a) What is meant by the term 'standard enthalpy change of formation'?
Give the standard conditions.

The standard enthalpy change of formation of a substance, ΔH_f^\ominus , is the enthalpy change when one mole of the substance is formed from its elements in their standard states under standard conditions. (1)

Standard conditions: 298 K, 1 atmosphere (1)

- b) Write the equation, including state symbols, for the reaction which represents the standard enthalpy change of formation of $\text{N}_2\text{O}_4(\text{l})$.





Unit Exercise (p.67)

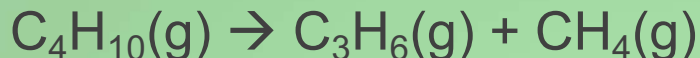
c) Calculate the standard enthalpy change of formation of $\text{N}_2\text{O}_4(\text{l})$ using the data above.

$$\begin{aligned}\Delta H_r^\ominus &= \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}] \\ &= 4 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] + 12 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{g})] - 4 \times \Delta H_f^\ominus [\text{CH}_3\text{NHNH}_2(\text{l})] - 5 \times \Delta H_f^\ominus [\text{N}_2\text{O}_4(\text{l})] \\ 5 \times \Delta H_f^\ominus [\text{N}_2\text{O}_4(\text{l})] &= 4 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] + 12 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{g})] - 4 \times \Delta H_f^\ominus [\text{CH}_3\text{NHNH}_2(\text{l})] - \Delta H_r^\ominus \\ \Delta H_f^\ominus [\text{N}_2\text{O}_4(\text{l})] &= \frac{1}{5}[4(-394) + 12(-242) - 4(+54) - (-4\,596)] \text{ kJ mol}^{-1} \quad (1) \\ &= -20 \text{ kJ mol}^{-1} \quad (1) \\ \therefore \text{the standard enthalpy change of formation of } \text{N}_2\text{O}_4(\text{l}) &\text{ is } -20 \text{ kJ mol}^{-1}.\end{aligned}$$



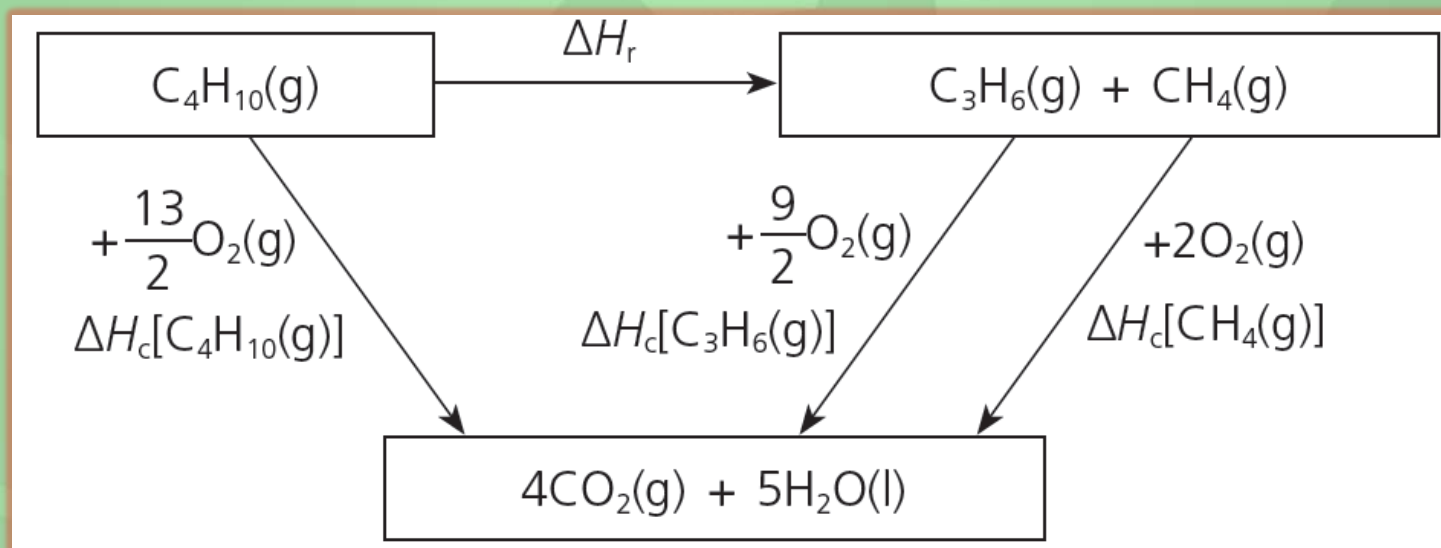
Unit Exercise (p.67)

19 The equation for a cracking reaction of butane is shown below.



Based on the enthalpy changes of combustion ΔH_c of $\text{C}_4\text{H}_{10}(\text{g})$, $\text{C}_3\text{H}_6(\text{g})$ and $\text{CH}_4(\text{g})$ to construct an enthalpy change cycle and applying Hess's law can give the enthalpy change of this cracking reaction.

a) Draw, with labels, this enthalpy change cycle.





Unit Exercise (p.67)

- b) The standard enthalpy changes of combustion ΔH_c^\ominus of $C_4H_{10}(g)$, $C_3H_6(g)$ and $CH_4(g)$ are given below.

	ΔH_c^\ominus (kJ mol ⁻¹)
$C_4H_{10}(g)$	-2 877
$C_3H_6(g)$	-2 058
$CH_4(g)$	-890

- i) State the standard conditions for 'standard enthalpy change'.
298 K; 1 atmosphere (1)
- ii) Calculate the standard enthalpy change of this cracking reaction.

According to Hess's law:

$$\Delta H_c^\ominus [C_4H_{10}(g)] = \Delta H_r^\ominus + \Delta H_c^\ominus [C_3H_6(g)] + \Delta H_c^\ominus [CH_4(g)]$$

$$\Delta H_r^\ominus = \Delta H_c^\ominus [C_4H_{10}(g)] - \Delta H_c^\ominus [C_3H_6(g)] - \Delta H_c^\ominus [CH_4(g)]$$

$$= [(-2\,877) - (-2\,058) - (-890)] \text{ kJ mol}^{-1} (1) = +71 \text{ kJ mol}^{-1} (1)$$

\therefore the standard enthalpy change of the cracking reaction is $+71 \text{ kJ mol}^{-1}$.



Unit Exercise (p.67)

20 The booster rocket of a spacecraft uses a mixture of Al(s) and NH₄ClO₄(s) as a fuel. Upon ignition, the fuel reacts to give Al₂O₃(s), AlCl₃(s), NO(g) and H₂O(g).



a) Complete and balance the chemical equation for the reaction between Al(s) and NH₄ClO₄(s).



b) Some standard enthalpy changes of formation are shown in the table below.

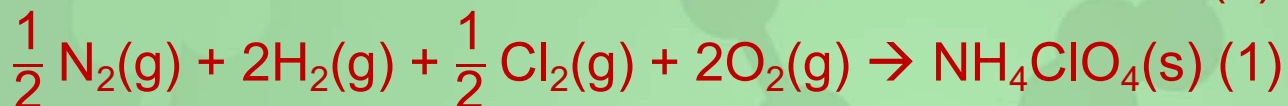
Substance	ΔH_f^\ominus (kJ mol ⁻¹)
NH ₄ ClO ₄ (s)	-295
Al ₂ O ₃ (s)	-1 676
AlCl ₃ (s)	-704
NO(g)	+90
H ₂ O(g)	-242



Unit Exercise (p.67)

- i) Using $\text{NH}_4\text{ClO}_4(\text{s})$ as an example, state the meaning of the term 'standard enthalpy change of formation', with the aid of a chemical equation.

The standard enthalpy change of formation of a substance, ΔH_f^\ominus , is the enthalpy change when one mole of the substance is formed from its elements in their standard states under standard conditions. (1)



- ii) Calculate the standard enthalpy change of the reaction between $\text{Al}(\text{s})$ and $\text{NH}_4\text{ClO}_4(\text{s})$.

$$\begin{aligned} \Delta H_r^\ominus &= \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}] \\ &= \Delta H_f^\ominus [\text{Al}_2\text{O}_3(\text{s})] + \Delta H_f^\ominus [\text{AlCl}_3(\text{s})] + 3 \times \Delta H_f^\ominus [\text{NO}(\text{g})] + 6 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{g})] - \\ &\quad 3 \times \Delta H_f^\ominus [\text{NH}_4\text{ClO}_4(\text{s})] \\ &= [(-1\,676) + (-704) + 3(+90) + 6(-242) - 3(-295)] \text{ kJ mol}^{-1} \quad (1) \\ &= -2\,677 \text{ kJ mol}^{-1} \quad (1) \end{aligned}$$

\therefore the standard enthalpy change of the reaction between $\text{Al}(\text{s})$ and $\text{NH}_4\text{ClO}_4(\text{s})$ is $-2\,677 \text{ kJ mol}^{-1}$.



Unit Exercise (p.67)

21 Energy exists in various forms.



Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is one important energy source for living things.

a) Write a chemical equation for the conversion of carbon dioxide gas and liquid water to solid glucose and oxygen gas.

b) The following standard enthalpy changes of formation are given:

$$\text{CO}_2(\text{g}) = -394 \text{ kJ mol}^{-1}, \text{H}_2\text{O}(\text{l}) = -286 \text{ kJ mol}^{-1},$$

$$\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) = -1\,274 \text{ kJ mol}^{-1}$$

Calculate the standard enthalpy change of the conversion in (a) above.

c) Green plants can convert carbon dioxide and water to glucose and oxygen. State the transformation of energy in this conversion.

(HKDSE, Paper 1B, 2018, 6(a))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



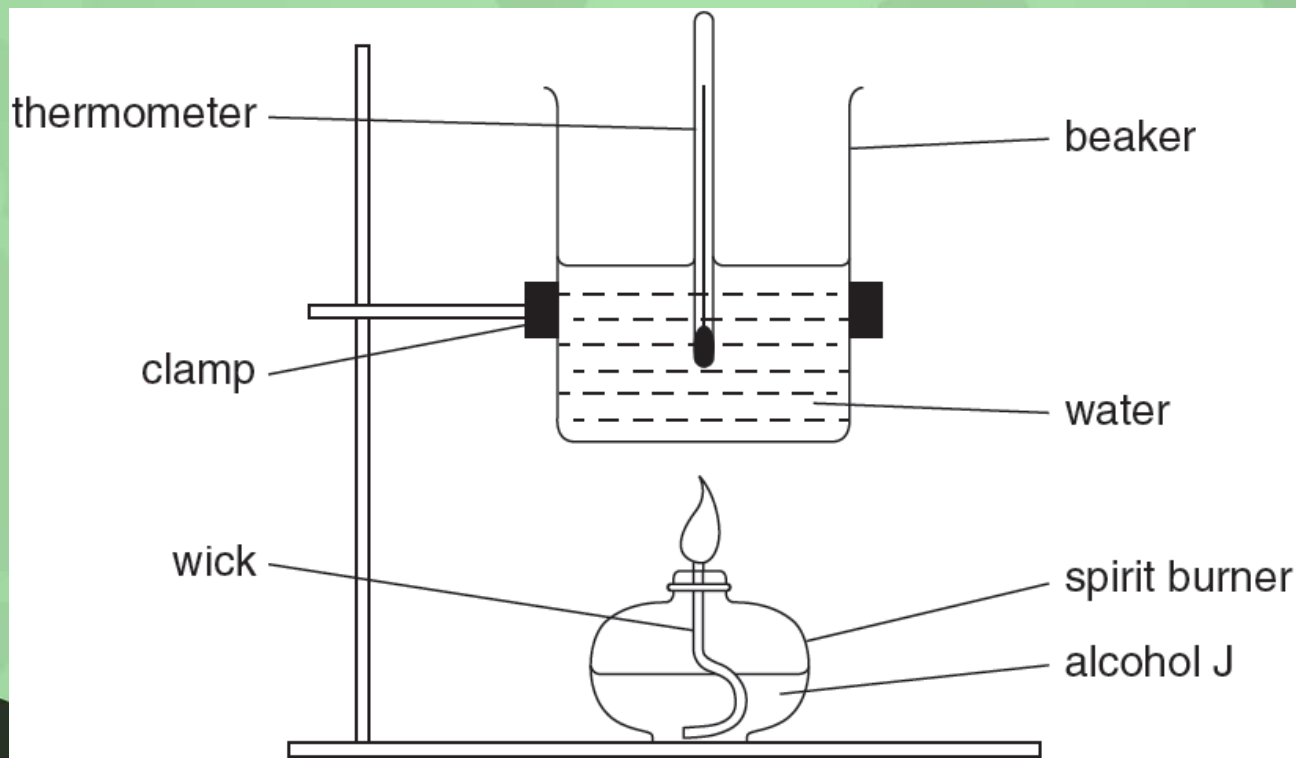
Unit Exercise (p.67)

22 A branched-chain alcohol J is a liquid and has the molecular formula $C_5H_{12}O$.



a) A student does an experiment to measure the enthalpy change of combustion, ΔH_c , of alcohol J.

i) The student burns alcohol J using the apparatus below.





Unit Exercise (p.67)

The student found that combustion of 1.54 g of alcohol J changes the temperature of 180 g of water from 22.8 °C to 75.3 °C.

The specific heat capacity of water is 4.18 J g⁻¹ K⁻¹.

- Calculate the amount, in mol, of alcohol J that burns.
- Calculate the enthalpy change of combustion, ΔH_c , of alcohol J, in kJ mol⁻¹.

Amount of heat

$$= m \times c \times \Delta T$$

$$= 180 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times (75.3 - 22.8) \text{ K} = 39\,500 \text{ J} = 39.5 \text{ kJ (1)}$$

Molar mass of alcohol J

$$= (5 \times 12.0 + 12 \times 1.0 + 16.0) \text{ g mol}^{-1} = 88.0 \text{ g mol}^{-1}$$

Number of moles of alcohol J burnt

$$= \frac{1.54 \text{ g}}{88.0 \text{ g mol}^{-1}} = 0.0175 \text{ mol}$$

Enthalpy change of combustion of alcohol J

$$= \frac{-39.5 \text{ kJ}}{0.0175 \text{ mol}} = -2\,260 \text{ kJ mol}^{-1} \text{ (1)}$$

\therefore the enthalpy change of combustion of alcohol J is $-2\,260 \text{ kJ mol}^{-1}$.



Unit Exercise (p.67)

ii) The calculated value of ΔH_c from this experiment is different from the value obtained from data books.

Apart from heat loss, suggest TWO reasons for the difference.
Assume that the calculation has been carried out correctly.

Any two of the following:

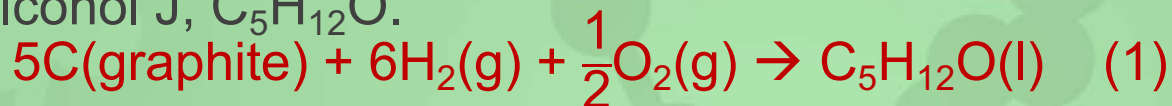
- Incomplete combustion (1)
- Non-standard conditions (1)
- Evaporation of alcohol / water (1)
- Heat capacity of beaker / apparatus neglected (1)



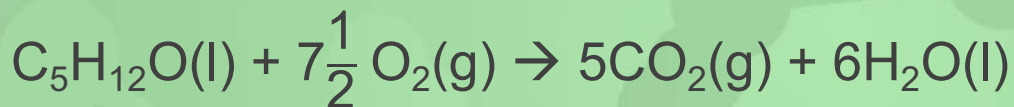
Unit Exercise (p.67)

b) The enthalpy change of combustion for alcohol J can also be determined indirectly from standard enthalpy changes of formation.

- i) Write an equation, including state symbols, for the chemical change that represents the standard enthalpy change of formation of the liquid alcohol J, C₅H₁₂O.



- ii) The equation for the complete combustion of alcohol J is shown below.



Enthalpy changes of formation, ΔH_f° , are shown in the table.

Substance	C ₅ H ₁₂ O(l)	CO ₂ (g)	H ₂ O(l)
ΔH_f° (kJ mol ⁻¹)	-366	-394	-286

Calculate the enthalpy change of combustion, ΔH_c° , of alcohol J from the information given above.

(OCR Advanced Subsidiary GCE, Chem. A, F322, Jun. 2015, 6(a)–(b))



Unit Exercise (p.67)

ii) $\Delta H_c [\text{C}_5\text{H}_{12}\text{O(l)}] = \text{sum of } \Delta H_f [\text{products}] - \text{sum of } \Delta H_f [\text{reactants}]$
 $= 5 \times \Delta H_f [\text{CO}_2(\text{g})] + 6 \times \Delta H_f [\text{H}_2\text{O(l)}] - \Delta H_f [\text{C}_5\text{H}_{12}\text{O(l)}]$
 $= [5(-394) + 6(-286) - (-366)] \text{ kJ mol}^{-1} \text{ (1)}$
 $= -3\,320 \text{ kJ mol}^{-1} \text{ (1)}$
 \therefore the enthalpy change of combustion of alcohol J is $-3\,320 \text{ kJ mol}^{-1}$.



Unit Exercise (p.67)

23 The first stage in the manufacture of nitric acid involves the oxidation of ammonia at 900 °C.



a) Calculate the standard enthalpy change of this reaction using the standard enthalpy changes shown below.

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
NH ₃ (g)	-46
NO(g)	+90
H ₂ O(g)	-242

$$\begin{aligned}\Delta H_r^\ominus &= \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}] \\ &= 4 \times \Delta H_f^\ominus [\text{NO}(\text{g})] + 6 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{g})] - 4 \times \Delta H_f^\ominus [\text{NH}_3(\text{g})] \\ &= [4(+90) + 6(-242) - 4(-46)] \text{ kJ mol}^{-1} \quad (1) \\ &= -908 \text{ kJ mol}^{-1} \quad (1)\end{aligned}$$

∴ the standard enthalpy change of the reaction is -908 kJ mol⁻¹.



Unit Exercise (p.67)

b) Explain why the enthalpy change calculated in (a) is NOT the standard enthalpy change of combustion of ammonia.

Any one of the following:

- H_2O is not in standard state (1)
- Combustion is not complete (1)
- Involves 4 moles of NH_3 (1)



Unit Exercise (p.67)

24 The enthalpy change of formation of $\text{MgCO}_3(\text{s})$ can be obtained using an indirect method. Firstly, the enthalpy change for the reaction of $\text{MgCO}_3(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq})$, and that of $\text{Mg}(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq})$ are respectively determined experimentally. After that, the enthalpy change of formation of $\text{MgCO}_3(\text{s})$ can be obtained through calculation with given enthalpy changes of formation of $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$.



- According to definition, under which condition could the 'heat change' of a reaction be regarded as the 'enthalpy change'?
- Explain why, instead of a direct method, an indirect method is used to obtain the enthalpy change of formation of $\text{MgCO}_3(\text{s})$.



Unit Exercise (p.67)

- c) In order to determine experimentally the enthalpy change for the reaction of $\text{MgCO}_3(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq})$, an accurate mass of $\text{MgCO}_3(\text{s})$ was firstly allowed to react with excess $\text{H}_2\text{SO}_4(\text{aq})$ in a polystyrene foam cup. The maximum rise in temperature of the mixture was then found. After calculation, the enthalpy change for the reaction can be obtained.
- i) Suggest one possible error for the above experimental procedure.
 - ii) Explain whether the enthalpy change for the reaction of $\text{CaCO}_3(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq})$ can be obtained using a similar experimental procedure.
- d) Using the information given below, calculate the standard enthalpy change of formation of $\text{MgCO}_3(\text{s})$.
- Standard enthalpy change for the reaction of $\text{MgCO}_3(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq}) = -50 \text{ kJ mol}^{-1}$
- Standard enthalpy change for the reaction of $\text{Mg}(\text{s})$ with $\text{H}_2\text{SO}_4(\text{aq}) = -467 \text{ kJ mol}^{-1}$
- Standard enthalpy change of formation of $\text{CO}_2(\text{g}) = -394 \text{ kJ mol}^{-1}$
- Standard enthalpy change of formation of $\text{H}_2\text{O}(\text{l}) = -286 \text{ kJ mol}^{-1}$

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

(HKDSE, Paper 1B, 2016, 7)



Unit Exercise (p.67)



25 It is very difficult to determine the enthalpy change of decomposition of copper(II) carbonate directly. A student carried out two experiments.

Experiment 1 Reaction of copper(II) carbonate with sulphuric acid



Experiment 2 Reaction of copper(II) oxide with sulphuric acid



In *Experiment 1*, excess $\text{H}_2\text{SO}_4(\text{aq})$ were placed in a polystyrene foam cup. The temperature of the acid was measured every 30 seconds for 2.5 minutes.

At 3.0 minutes, 3.58 g of $\text{CuCO}_3(\text{s})$ were added to the cup. The mixture was stirred and its temperature was measured for a further of 6.0 minutes.

a) Suggest why it is difficult to determine directly the enthalpy change of decomposition of copper(II) carbonate.

Anyone of the following:

- It is difficult to measure the heat taken in when heating a substance. (1)
- It is difficult to measure the temperature change of a solid. (1)

b) Suggest why using a polystyrene foam cup, rather than a glass beaker, was likely to give a more accurate answer for the enthalpy changes.

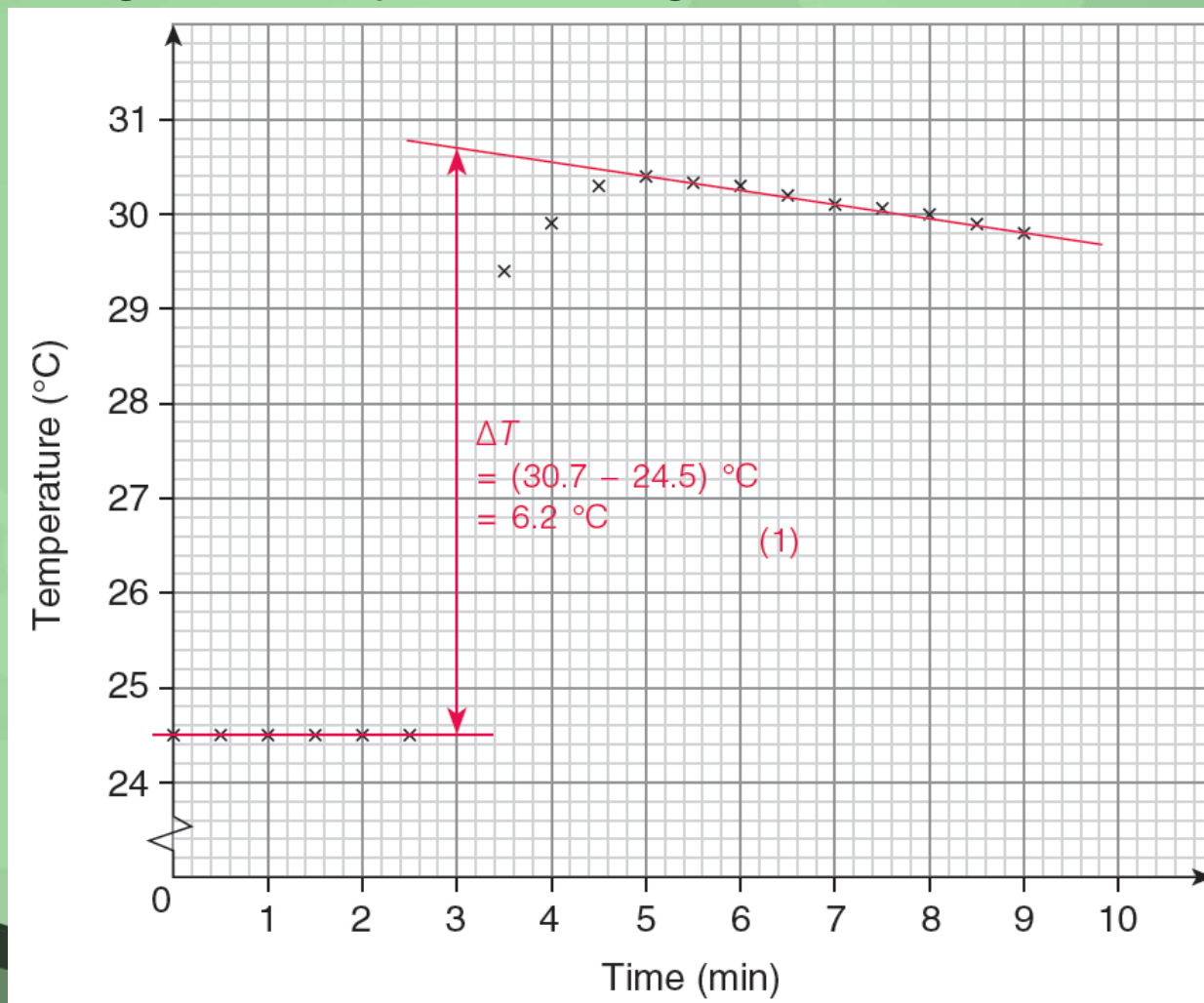
Polystyrene foam is a better thermal insulator / has a lower heat capacity.

Hence less heat transferred to the surroundings / cups. (1)



Unit Exercise (p.67)

- c) The recordings of temperature are shown in the graph below.
i) Deduce the greatest temperature change of the reaction mixture.





Unit Exercise (p.67)

- c) ii) Calculate the enthalpy change of reaction between copper(II) carbonate and sulphuric acid, in kJ mol^{-1} .

(Relative atomic masses: C = 12.0, O = 16.0, Cu = 63.5; mass of the resulting solution = 54.0 g; specific heat capacity of the resulting solution = $4.18 \text{ J g}^{-1} \text{ K}^{-1}$)

$$\text{Amount of heat released} = m \times c \times \Delta T = 54.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 6.2 \text{ K} \\ = 1\,400 \text{ J} = 1.40 \text{ kJ} \quad (1)$$

$$\text{Molar mass of CuCO}_3 = (63.5 + 12.0 + 3 \times 16.0) \text{ g mol}^{-1} = 123.5 \text{ g mol}^{-1}$$

$$\text{Number of moles of CuCO}_3 \text{ reacted} = \frac{3.58 \text{ g}}{123.5 \text{ g mol}^{-1}} = 0.0290 \text{ mol}$$

$$\text{Enthalpy change of reaction} = \frac{-1.40 \text{ kJ}}{0.0290 \text{ mol}} = -48.3 \text{ kJ mol}^{-1} \quad (1)$$

\therefore the enthalpy change of reaction between copper(II) carbonate and sulphuric acid is $-48.3 \text{ kJ mol}^{-1}$.

- iii) The enthalpy change obtained from this experiment was much less negative than the data book value. Suggest ONE reason for this difference.

Any one of the following:

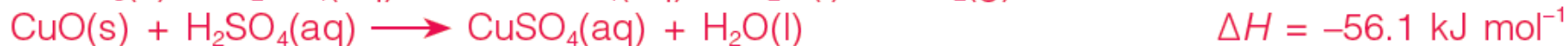
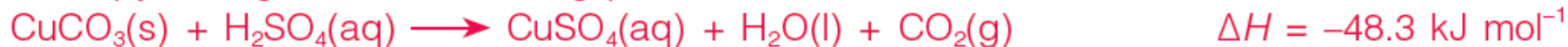
- Heat loss to the surroundings (1)
- The apparatus absorbs heat (1)



Unit Exercise (p.67)

- d) In another experiment performed under the same conditions, the enthalpy change of the reaction between CuO(s) and $\text{H}_2\text{SO}_4\text{(aq)}$ was found to be $-56.1 \text{ kJ mol}^{-1}$. Calculate the enthalpy change of decomposition of $\text{CuCO}_3\text{(s)}$ under the experimental conditions.

Enthalpy changes of the following processes are known:



Reverse equation (2) to give equation (2)'. Add together the equations as shown below:



According to Hess's law:

$$\text{Enthalpy change of decomposition of CuCO}_3\text{(s)} = (-48.3 + 56.1) \text{ kJ mol}^{-1} \quad (1)$$

$$= +7.8 \text{ kJ mol}^{-1} \quad (1)$$

\therefore the enthalpy change of decomposition of $\text{CuCO}_3\text{(s)}$ is $+7.8 \text{ kJ mol}^{-1}$.



Topic Exercise (p.77)

Note: Questions are rated according to ascending level of difficulty (from 1 to 5):



question targeted at level 3 and above;



question targeted at level 4 and above;



question targeted at level 5.

‘ * ’ indicates 1 mark is given for effective communication.



Topic Exercise (p.77)

PART I MULTIPLE CHOICE QUESTIONS

- 1 The specific heat capacity of gold is $0.13 \text{ J g}^{-1} \text{ K}^{-1}$.
10.0 g of gold at 25.0°C take in 33.8 J of heat. What is the final temperature of the gold?

- A 26.0°C
- B 43.9°C
- C 51.0°C
- D 60.5°C

Answer: C

Explanation:

Amount of heat = $m \times c \times \Delta T$

$$33.8 \text{ J} = 10.0 \text{ g} \times 0.13 \text{ J g}^{-1} \text{ K}^{-1} \times \Delta T$$

$$\Delta T = 26.0 \text{ K}$$

\therefore the final temperature of gold

$$= (25.0 + 26.0)^\circ \text{C} = 51.0^\circ \text{C}$$



Topic Exercise (p.77)

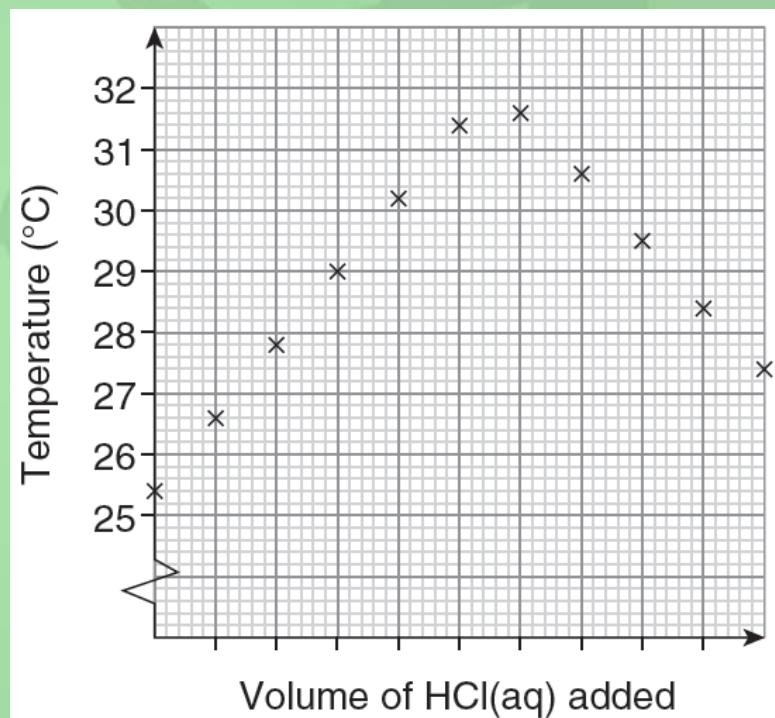
Answer: D



- 2 In an experiment, standard HCl(aq) was added from a burette to a known volume of NaOH(aq) placed in an expanded polystyrene cup. The graph below shows the temperatures of the mixture in the cup during the process:

What is the greatest temperature rise of the mixture in the cup as estimated from the graph above?

- A $2.0\text{ }^{\circ}\text{C}$
- B $4.6\text{ }^{\circ}\text{C}$
- C $6.2\text{ }^{\circ}\text{C}$
- D $6.6\text{ }^{\circ}\text{C}$



(HKDSE, Paper 1A, 2014, 12)



Topic Exercise (p.77)

3 Which of the following enthalpy changes CANNOT be measured directly by experiment?

Answer: A

- A Enthalpy change of formation of propane
- B Enthalpy change of combustion of hydrogen
- C Enthalpy change of formation of carbon dioxide
- D Enthalpy change of combustion of carbon monoxide

(HKDSE, Paper 1A, 2017, 26)



Topic Exercise (p.77)

4 Some camping stoves use butane (C_4H_{10}) as fuel. On complete combustion, 6.00 g of butane give 297 kJ of heat.

What is the enthalpy change of the reaction below?



(Relative atomic masses: H = 1.0, C = 12.0)

Answer: A

A $-5\,740\text{ kJ mol}^{-1}$

B $-4\,020\text{ kJ mol}^{-1}$

C $-2\,340\text{ kJ mol}^{-1}$

D $-1\,030\text{ kJ mol}^{-1}$

Explanation:

**Molar mass of C_4H_{10} = $(4 \times 12.0 + 10 \times 1.0)\text{ g mol}^{-1}$
= 58.0 g mol^{-1}**

Number of moles of C_4H_{10} combusted = $\frac{6.00\text{ g}}{58.0\text{ g mol}^{-1}}$

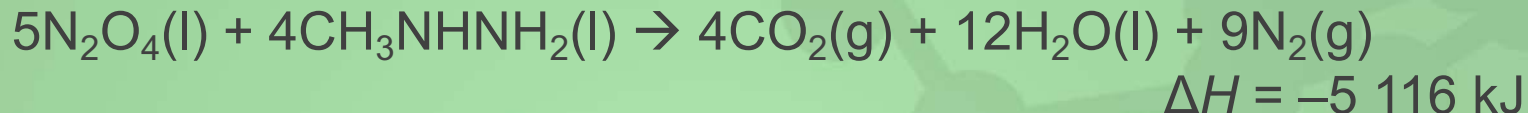
**Enthalpy change when 2 moles of C_4H_{10} combusted
= $-297\text{ kJ} \times \frac{2 \times 58.0\text{ g}}{6.00\text{ g}} = -5\,740\text{ kJ}$**



Topic Exercise (p.77)



5 $\text{N}_2\text{O}_4(\text{l})$ and $\text{CH}_3\text{NHNH}_2(\text{l})$ react according to the equation below when ignited.



How much heat would be released when 1 mole of each reactant are mixed and ignited?

Answer: A

- A 1 023 kJ
- B 1 955 kJ
- C 4 093 kJ
- D 5 116 kJ

Explanation:

When 1 mole of each reactant is mixed and ignited, 1 mole of N_2O_4 reacts with 0.8 mole of CH_3NHNH_2 .

$$\begin{aligned}\text{Amount of heat released} &= \frac{1}{5} \times 5\,116 \text{ kJ} \\ &= 1\,023 \text{ kJ}\end{aligned}$$



Topic Exercise (p.77)

Answer: A



6 Under standard conditions, complete combustion of 0.200 mol of carbon disulphide gives 216 kJ of heat.



Which of the following is the standard enthalpy change of formation of $\text{CS}_2(\text{l})$?

(Standard enthalpy change of formation of $\text{CO}_2(\text{g}) = -394 \text{ kJ mol}^{-1}$;
standard enthalpy change of formation of $\text{SO}_2(\text{g}) = -297 \text{ kJ mol}^{-1}$)

- A $+92 \text{ kJ mol}^{-1}$
- B $+181 \text{ kJ mol}^{-1}$
- C $+240 \text{ kJ mol}^{-1}$
- D $+475 \text{ kJ mol}^{-1}$

Standard enthalpy change of combustion of $\text{CS}_2(\text{l})$

$$\begin{aligned} &= \frac{-216}{0.200} \text{ kJ mol}^{-1} \\ &= -1\,080 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta H_c^\ominus[\text{CS}_2(\text{l})] &= \text{sum of } \Delta H_f^\ominus[\text{products}] - \text{sum of } \Delta H_f^\ominus[\text{reactants}] \\ &= \Delta H_f^\ominus[\text{CO}_2(\text{g})] + 2 \times \Delta H_f^\ominus[\text{SO}_2(\text{g})] - \Delta H_f^\ominus[\text{CS}_2(\text{l})] \end{aligned}$$

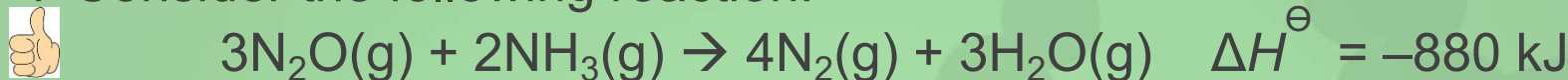
$$\begin{aligned} \Delta H_f^\ominus[\text{CS}_2(\text{l})] &= \Delta H_f^\ominus[\text{CO}_2(\text{g})] + 2 \times \Delta H_f^\ominus[\text{SO}_2(\text{g})] - \Delta H_c^\ominus[\text{CS}_2(\text{l})] \\ &= [(-394) + 2 \times (-297) - (-1\,080)] \text{ kJ mol}^{-1} \\ &= +92 \text{ kJ mol}^{-1} \end{aligned}$$



Topic Exercise (p.77)

Answer: B

7 Consider the following reaction:



Given the following data:

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
NH ₃ (g)	-45.9
H ₂ O(g)	-242

What is the standard enthalpy change of formation of N₂O(g)?A +246 kJ mol⁻¹B +81.9 kJ mol⁻¹C -81.9 kJ mol⁻¹D -246 kJ mol⁻¹

$$\begin{aligned}\Delta H_r^\ominus &= \text{sum of } \Delta H_f^\ominus[\text{products}] - \text{sum of } \Delta H_f^\ominus[\text{reactants}] \\ &= 3 \times \Delta H_f^\ominus[\text{H}_2\text{O}(\text{g})] - 3 \times \Delta H_f^\ominus[\text{N}_2\text{O}(\text{g})] - 2 \times \Delta H_f^\ominus[\text{NH}_3(\text{g})] \\ 3 \times \Delta H_f^\ominus[\text{N}_2\text{O}(\text{g})] &= 3 \times \Delta H_f^\ominus[\text{H}_2\text{O}(\text{g})] - 2 \times \Delta H_f^\ominus[\text{NH}_3(\text{g})] - \Delta H_r^\ominus \\ \Delta H_f^\ominus[\text{N}_2\text{O}(\text{g})] &= \frac{1}{3}[3(-242) - 2(-45.9) - (-880)] \text{ kJ mol}^{-1} \\ &= +81.9 \text{ kJ mol}^{-1}\end{aligned}$$



Topic Exercise (p.77)

8 Which of the following processes are endothermic?

- (1) Melting of wax
- (2) Cracking of heavy oil
- (3) Adding zinc powder to $\text{CuSO}_4(\text{aq})$

Answer: A

- A (1) and (2) only
- B (1) and (3) only
- C (2) and (3) only
- D (1), (2) and (3)

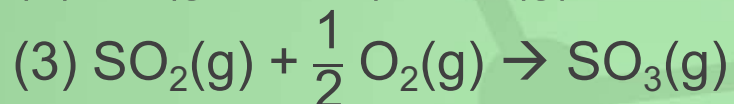
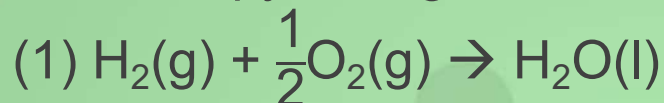
(HKDSE, Paper 1A, 2018, 22)



Topic Exercise (p.77)



9 For which of the following reactions does the value of ΔH^\ominus represent both the standard enthalpy change of combustion of a substance and the standard enthalpy change of formation of another substance?



A (1) only

B (2) only

C (1) and (3) only

D (2) and (3) only

Answer: A



Topic Exercise (p.77)

10 Which of the following combinations is / are correct?



<u>Chemical reaction</u>	<u>Enthalpy change of reaction</u>
(1) $2\text{H}_2\text{O}(\text{l}) \longrightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$	positive
(2) $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{CO}_2(\text{g})$	positive
(3) $2\text{Na}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \longrightarrow 2\text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$	negative

A (1) only

B (2) only

C (1) and (3) only

D (2) and (3) only

(HKDSE, Paper 1A, 2015, 18)



Topic Exercise (p.77)


Directions : Each question (Questions 11–12) consists of two separate statements. Decide whether each of the two statements is true or false; if both are true, then decide whether or not the second statement is a correct explanation of the first statement. Then select one option from A to D according to the following table :


- A Both statements are true and the 2nd statement is a correct explanation of the 1st statement.
- B Both statements are true but the 2nd statement is NOT a correct explanation of the 1st statement.
- C The 1st statement is false but the 2nd statement is true.
- D Both statements are false.



Topic Exercise (p.77)

1st statement

11  The standard enthalpy change of formation of a compound must be a negative value.

12  The standard enthalpy change of neutralization between HCl(aq) and NaOH(aq) is the same as that between CH₃COOH(aq) and NaOH(aq).

The standard enthalpy change of neutralisation between HCl(aq) and NaOH(aq) is -57 kJ mol^{-1} .

The standard enthalpy change of neutralisation between CH₃COOH(aq) (a weak acid) and NaOH(aq) is less than -57 kJ mol^{-1} . This is because some of the heat released in neutralisation is used to ionise the acid.

2nd statement

Under standard conditions, a compound must be energetically more stable than its constituent elements.

Answer: D

(HKDSE, Paper 1A, 2016, 24)

NaOH(aq) is a strong alkali.

Answer: C

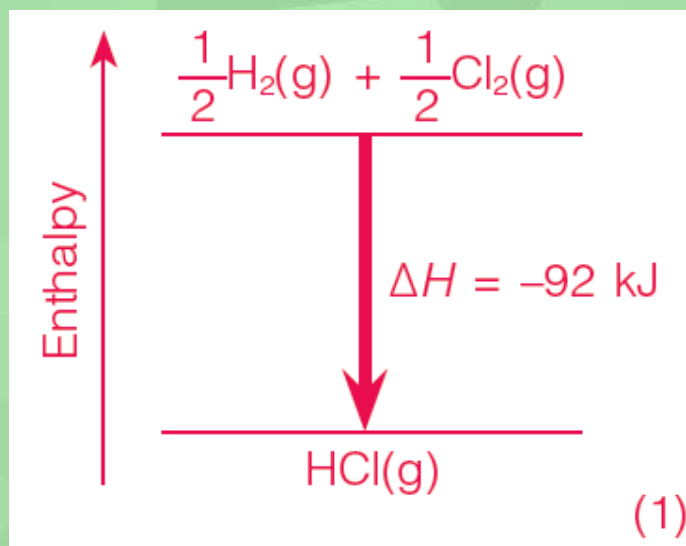


Topic Exercise (p.77)

PART II STRUCTURED QUESTIONS

13 The standard enthalpy change of formation of hydrogen chloride is -92 kJ mol^{-1} .

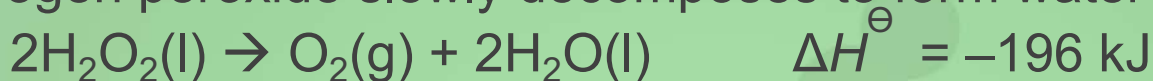
Draw a labelled enthalpy level diagram to represent the reaction which occurs when hydrogen chloride formed from its elements.





Topic Exercise (p.77)

14 Hydrogen peroxide slowly decomposes to form water and oxygen.



a) Calculate the standard enthalpy change of formation of $\text{H}_2\text{O}_2(\text{l})$.

(Standard enthalpy change of combustion of $\text{H}_2(\text{g}) = -286 \text{ kJ mol}^{-1}$)

$$\Delta H^\ominus = 2 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{l})] - 2 \times \Delta H_f^\ominus [\text{H}_2\text{O}_2(\text{l})]$$

$$2 \times \Delta H_f^\ominus [\text{H}_2\text{O}_2(\text{l})] = 2 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{l})] - \Delta H^\ominus$$

$$\Delta H_f^\ominus [\text{H}_2\text{O}_2(\text{l})] = \frac{1}{2}[2(-286) - (-196)] \text{ kJ mol}^{-1} \quad (1)$$

$$= -188 \text{ kJ mol}^{-1} \quad (1)$$

\therefore the standard enthalpy change of formation of $\text{H}_2\text{O}_2(\text{l})$ is -188 kJ mol^{-1} .

b) Explain, in terms of the breaking and forming of bonds, why the decomposition of hydrogen peroxide is exothermic.

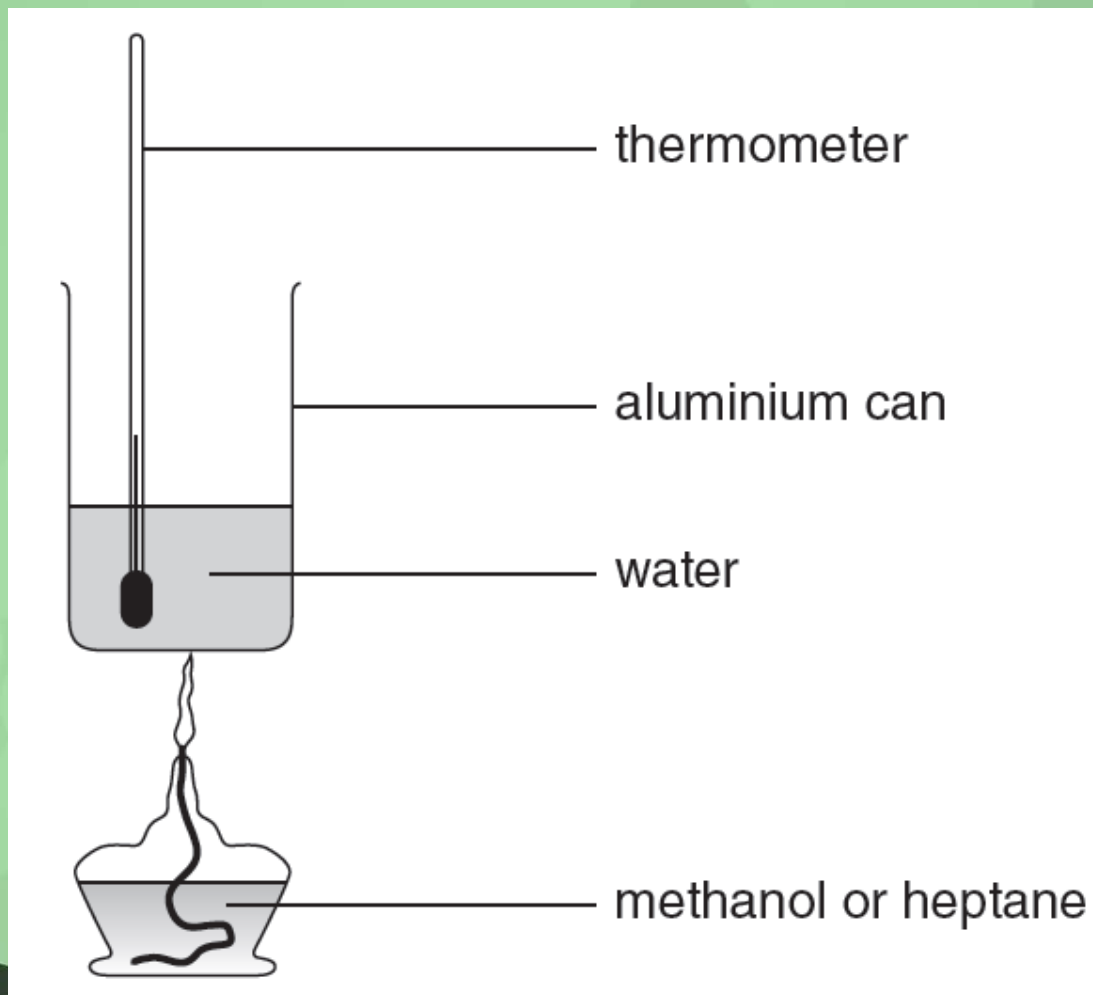
The total energy released in the bond forming process is greater than the total energy required in the bond breaking process. (1)



Topic Exercise (p.77)



15 Burning heptane (C_7H_{16}) releases energy. The enthalpy change of combustion of heptane was determined using the set-up shown below:





Topic Exercise (p.77)

Step (I): The aluminium can with a fixed mass of water was heated by burning methanol. The temperature of water increased by $18.5\text{ }^{\circ}\text{C}$ after 1.58 g of methanol was burnt.

Step (II): The aluminium can with the same mass of water in *Step (I)* was heated by burning heptane. The temperature of water increased by $25.8\text{ }^{\circ}\text{C}$ after 1.02 g of heptane was burnt.

- a) Given that, under the conditions of experiment, the enthalpy change of combustion of methanol is -715 kJ mol^{-1} , calculate the enthalpy change of combustion of heptane, in kJ mol^{-1} , under the same conditions.

(Relative molecular masses: methanol = 32.0, heptane = 100.0)

- b) Besides heat loss, suggest another source of error in the experiment.

(HKDSE, Paper 1B, 2018, 6(b))

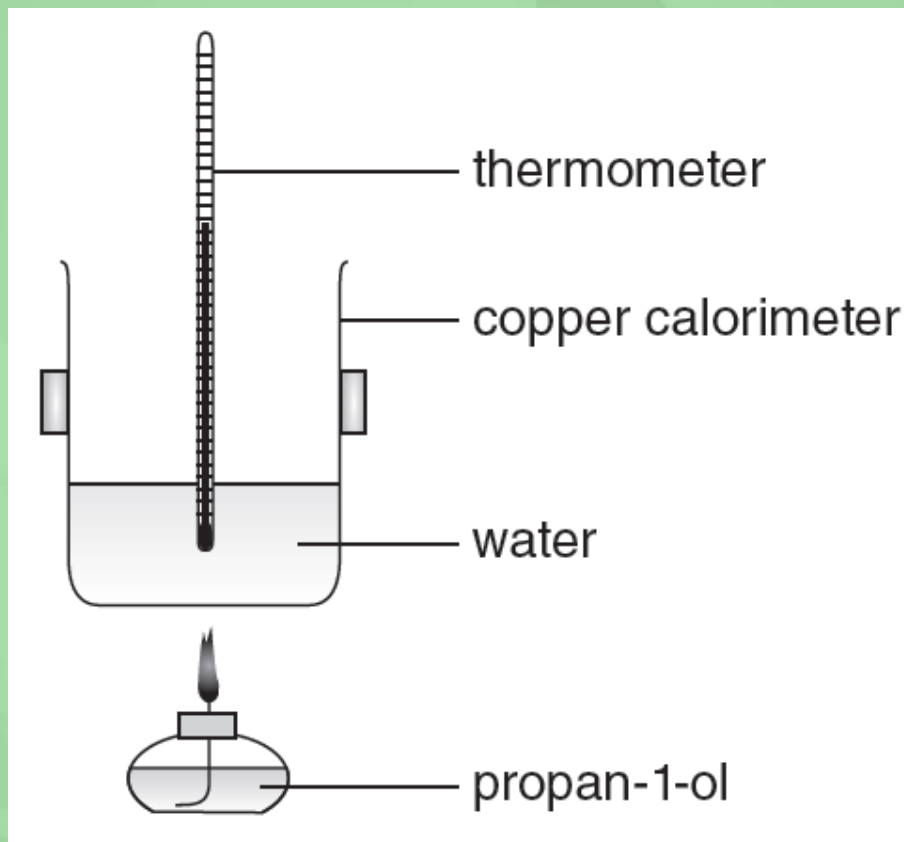
Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



Topic Exercise (p.77)



16 A student used the experimental set-up shown below to determine the enthalpy change of combustion of propan-1-ol.





Topic Exercise (p.77)

a) The student measured the mass of propan-1-ol burnt and the temperature rise of the water.

Give THREE other pieces of information which the student needed in order to calculate the enthalpy change of combustion of propan-1-ol.

Any three of the following:

- Relative molecular mass of propan-1-ol (1)
- Specific heat capacity of water (1)
- Mass / volume of water (1)
- Specific heat capacity of calorimeter (1)
- Mass of calorimeter (1)



Topic Exercise (p.77)

- b) The result obtained by this method was lower than the data book value. Suggest ONE way in which the experimental set-up could be modified in order to obtain a result closer to the data book value. Explain your answer.

Heat loss to the surroundings (1)

Use a lid / insulate the calorimeter / provide shielding around the burner (1)



Topic Exercise (p.77)

- c) Use the standard enthalpy changes of combustion given in the table to calculate the standard enthalpy change of formation of propan-1-ol.

<u>Substance</u>	<u>ΔH_c^\ominus (kJ mol⁻¹)</u>
C(graphite)	-394
H ₂ (g)	-286
C ₃ H ₇ OH(l)	-2 021

$$\begin{aligned}\Delta H_f^\ominus [\text{C}_3\text{H}_7\text{OH}(\text{l})] &= \text{sum of } \Delta H_c^\ominus [\text{elements}] - \Delta H_c^\ominus [\text{C}_3\text{H}_7\text{OH}(\text{l})] \\ &= 3 \times \Delta H_c^\ominus [\text{C}(\text{graphite})] + 4 \times \Delta H_c^\ominus [\text{H}_2(\text{g})] - \Delta H_c^\ominus [\text{C}_3\text{H}_7\text{OH}(\text{l})] \\ &= [3(-394) + 4(-286) - (-2\,021)] \text{ kJ mol}^{-1} \quad (1) \\ &= -305 \text{ kJ mol}^{-1} \quad (1)\end{aligned}$$

∴ the standard enthalpy change of formation of propan-1-ol is -305 kJ mol⁻¹.



Topic Exercise (p.77)



17 0.0200 mole of solid calcium hydroxide was allowed to react with 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid (excess) in a polystyrene foam cup. The experimental data obtained are shown below:

Initial temperature of the solution	24.6 °C
Final temperature of the resulting solution	35.5 °C
Mass of the resulting solution	52.0 g
Specific heat capacity of the resulting solution	4.3 J g ⁻¹ K ⁻¹
Molar mass of Ca(OH) ₂	74.1 g mol ⁻¹



Topic Exercise (p.77)

- a) Calculate the enthalpy change of the reaction, in kJ mol^{-1} .

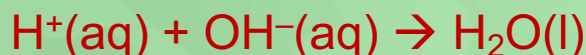
$$\text{Amount of heat released} = m \times c \times \Delta T = 52.0 \text{ g} \times 4.3 \text{ J g}^{-1} \text{ K}^{-1} \times (35.5 - 24.6) \text{ K} \\ = 2440 \text{ J} = 2.44 \text{ kJ} \text{ (1)}$$

$$\text{Enthalpy change of reaction} = \frac{-2.44 \text{ kJ}}{0.0200 \text{ mol}} = -122 \text{ kJ mol}^{-1} \text{ (1)}$$

- b) The experiment was repeated using 50.0 cm^3 of 1.00 mol dm^{-3} nitric acid instead of hydrochloric acid. Explain why the temperature change was the same in both experiments.

Nitric acid is a strong acid.

Thus, the same neutralisation occurs. (1)



- c) The experiment was repeated again using 25.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid. Predict the temperature change in this experiment.

The temperature change would double.

The same amount of heat was transferred to 25.0 cm^3 of solution, rather than 50 cm^3 solution. (1)



Topic Exercise (p.77)

18 The procedure below was used to determine the enthalpy change of the reaction of copper(II) sulphate solution with magnesium.



Procedure

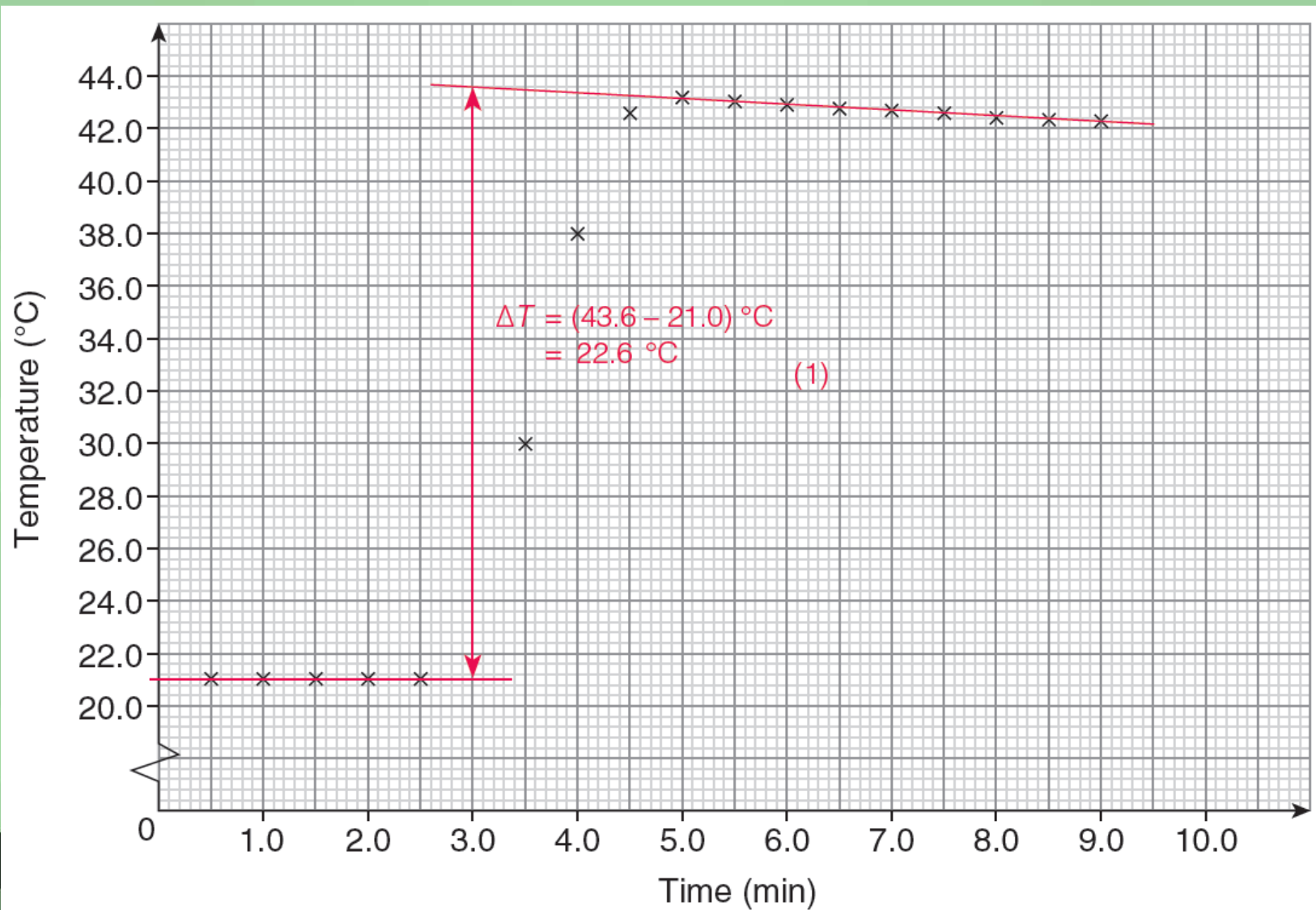
- 100.0 cm³ of 0.500 mol dm⁻³ CuSO₄(aq) were placed in a polystyrene foam cup and the temperature recorded every 30 seconds for 2.5 minutes.
- At 3.0 minutes, 0.520 g of magnesium powder was added and stirred.
- The temperature was recorded every 30 seconds for a further of 6.0 minutes.

The graph shows the results obtained.

- a) Estimate, from the graph, the maximum temperature change of the reaction mixture.



Topic Exercise (p.77)





Topic Exercise (p.77)

b) Show, by calculations, that Mg(s) was the limiting reactant.

(Relative atomic mass: Mg = 24.3)

$$\text{Number of moles of Mg} = \frac{0.520 \text{ g}}{24.3 \text{ g mol}^{-1}} = 0.0214 \text{ mol}$$

$$\text{Number of moles of CuSO}_4 = 0.500 \text{ mol dm}^{-3} \times \frac{100.0}{100.0} \text{ dm}^3 = 0.0500 \text{ mol}$$

According to the equation, 1 mole of Mg reacts with 1 mole of CuSO₄.

In this reaction, 0.0214 mole of Mg reacted with 0.0214 mole of CuSO₄.

Thus, CuSO₄ was in excess. Mg was the limiting reactant. (1)

c) Assuming that the mass and specific heat capacity of the reaction mixture were 101.0 g and 4.2 J g⁻¹ K⁻¹ respectively, calculate the enthalpy change of the reaction, in kJ mol⁻¹.

$$\begin{aligned} \text{Amount of heat} &= m \times c \times \Delta T = 101.0 \text{ g} \times 4.2 \text{ J g}^{-1} \text{ K}^{-1} \times 22.6 \text{ K} \\ &= 9\,590 \text{ J} = 9.59 \text{ kJ} \quad (1) \end{aligned}$$

$$\text{Enthalpy change of reaction} = \frac{-9.59 \text{ kJ}}{0.0214 \text{ mol}} = -448 \text{ kJ mol}^{-1} \quad (1)$$

∴ the enthalpy change of the reaction is -448 kJ mol⁻¹.



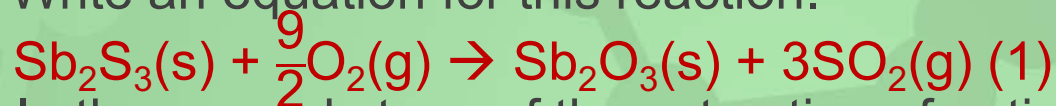
Topic Exercise (p.77)



19 Antimony is a solid element that is used in industry. The method used for the extraction of antimony depends on the grade of the ore.

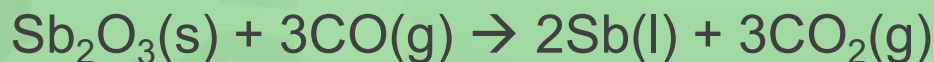
- a) In the first stage of the extraction of antimony from a high-grade ore, antimony(III) sulphide is roasted in air to convert it into antimony(III) oxide (Sb_2O_3) and sulphur dioxide.

Write an equation for this reaction.



- b) In the second stage of the extraction of antimony from a high-grade ore, antimony(III) oxide is reacted with carbon monoxide at high temperature.
- i) Use the standard enthalpy changes of formation in the following table and the equation given below the table to calculate a value for the standard enthalpy change for this reaction.

	$\text{Sb}_2\text{O}_3(\text{s})$	$\text{CO}(\text{g})$	$\text{Sb}(\text{l})$	$\text{CO}_2(\text{g})$
ΔH_f^\ominus (kJ mol ⁻¹)	-705	-111	+20	-394





Topic Exercise (p.77)

b) i) $\Delta H_r^\ominus = \text{sum of } \Delta H_f^\ominus [\text{products}] - \text{sum of } \Delta H_f^\ominus [\text{reactants}]$
 $= 2 \times \Delta H_f^\ominus [\text{Sb(l)}] + 3 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] - \Delta H_f^\ominus [\text{Sb}_2\text{O}_3(\text{s})] - 3 \times \Delta H_f^\ominus [\text{CO}(\text{g})]$
 $= [2(+20) + 3(-394) - (-705) - 3(-111)] \text{ kJ mol}^{-1} (1)$
 $= -104 \text{ kJ mol}^{-1} (1)$
 $\therefore \text{the standard enthalpy change of the reaction is } -104 \text{ kJ mol}^{-1}.$

ii) Suggest why the value for the standard enthalpy of formation of liquid antimony, given in the above table, is NOT zero.

Antimony is NOT in its standard state. / The standard state of antimony is solid. (1)



Topic Exercise (p.77)

20 Ethyne is a gaseous hydrocarbon with molecular formula C_2H_2 .



- a) Suggest why the enthalpy change of formation of $C_2H_2(g)$ CANNOT be determined directly by experiment.
- b) Hess's law can be used to find enthalpy changes which CANNOT be determined directly by experiment. State Hess's law.
- c) Based on the enthalpy changes of combustion ΔH_c of $C_2H_2(g)$, $C(\text{graphite})$ and $H_2(g)$ to construct an enthalpy change cycle and applying Hess's law can give the enthalpy change of formation of $C_2H_2(g)$.
- i) Draw, with labels, this enthalpy change cycle.
- ii) The standard enthalpy changes of combustion ΔH_c^\ominus of $C_2H_2(g)$, $C(\text{graphite})$ and $H_2(g)$ are given below:



Topic Exercise (p.77)

	ΔH_c^\ominus (kJ mol ⁻¹)
C ₂ H ₂ (g)	-1 300
C(graphite)	-394
H ₂ (g)	-286

- I) State the standard conditions for 'standard enthalpy change'.
- II) Calculate the standard enthalpy change of formation of C₂H₂(g).
(HKDSE, Paper 1B, 2017, 7)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



Topic Exercise (p.77)

21 Nonane (C_9H_{20}) is one of the components in the fuel kerosene.



a) Using nonane as an example, state the meaning of the term 'standard enthalpy change of combustion' with the aid of a chemical equation.

The standard enthalpy change of combustion of a substance, ΔH_c^\ominus , is the enthalpy change when one mole of the substance is completely burnt in oxygen under standard conditions. (1)



b) State the standard conditions for 'standard enthalpy change'.

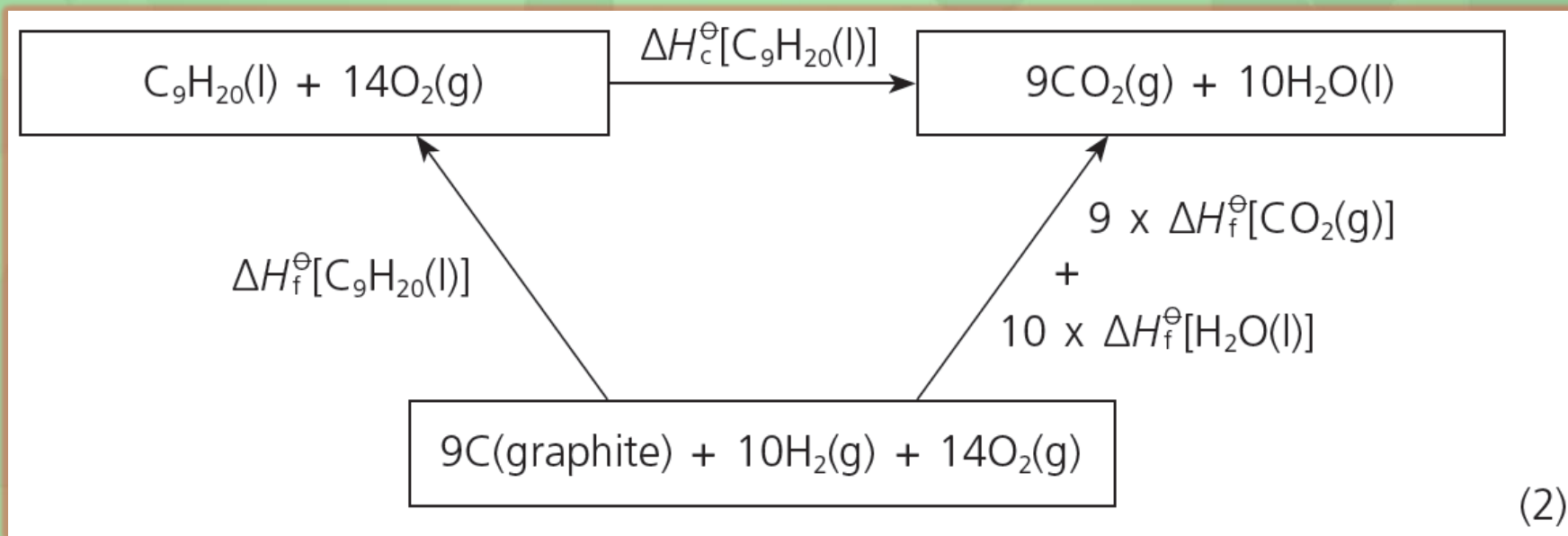
298 K

1 atmosphere (1)



Topic Exercise (p.77)

- c) Based on the enthalpy changes of formation ΔH_f^\ominus of $\text{C}_9\text{H}_{20}(\text{l})$, $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$ to construct an enthalpy change cycle and applying Hess's law can give the enthalpy change of combustion of $\text{C}_9\text{H}_{20}(\text{l})$.
- i) Draw, with labels, this enthalpy change cycle.





Topic Exercise (p.77)

- ii) Use the values given in the table below to calculate the standard enthalpy of combustion of nonane.

Substance	ΔH_f^\ominus (kJ mol ⁻¹)
C ₉ H ₂₀ (l)	-275
CO ₂ (g)	-394
H ₂ O(l)	-286

$$\begin{aligned}\Delta H_c^\ominus [\text{C}_9\text{H}_{20}(\text{l})] &= 9 \times \Delta H_f^\ominus [\text{CO}_2(\text{g})] + 10 \times \Delta H_f^\ominus [\text{H}_2\text{O}(\text{l})] - \Delta H_f^\ominus [\text{C}_9\text{H}_{20}(\text{l})] \\ &= [9(-394) + 10(-286) - (-275)] \text{ kJ mol}^{-1} \quad (1) \\ &= -6\,131 \text{ kJ mol}^{-1}\end{aligned}$$

∴ the standard enthalpy change of combustion of nonane is -6 131 kJ mol⁻¹.



Topic Exercise (p.77)



22 Calcium oxide reacts with dilute hydrochloric acid according to the equation below.



In an experiment to investigate this reaction, the following procedure was carried out.

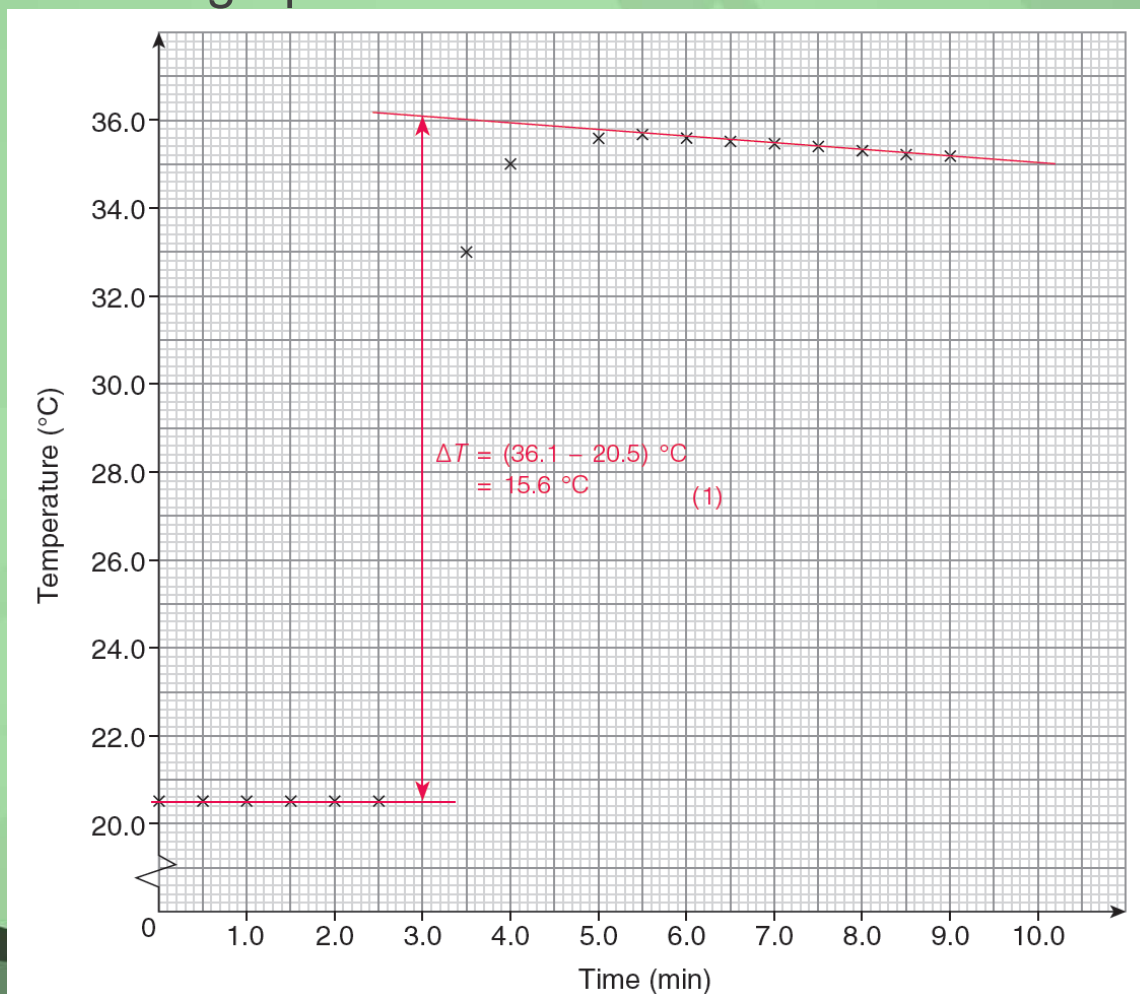
- 50.0 cm³ of 2.00 mol dm⁻³ hydrochloric acid (excess) were placed in a polystyrene foam cup.
- The temperature of the acid was measured every 30 seconds for 2.5 minutes.
- At 3.0 minutes, 1.46 g of calcium oxide powder were added to the acid and stirred the reaction mixture.
- The temperature of the resulting solution was measured every 30 seconds for a further of 6.0 minutes.



Topic Exercise (p.77)

The graph below shows the results obtained.

- a) i) What is the greatest temperature rise of the resulting solution as estimated from the graph?





Topic Exercise (p.77)

ii) Given that:

Mass of the resulting solution = 51.5 g

Specific heat capacity of the resulting solution = $4.2 \text{ J g}^{-1} \text{ K}^{-1}$

Calculate the enthalpy change for the quantities in this experiment.

Enthalpy change for quantities in experiment

$$= m \times c \times \Delta T = 51.5 \text{ g} \times 4.2 \text{ J g}^{-1} \text{ K}^{-1} \times 15.6 \text{ K}$$

$$= 3\,370 \text{ J} = 3.37 \text{ kJ (1)}$$

iii) Calculate the enthalpy change of reaction when 1 mole of calcium oxide reacts with dilute hydrochloric acid.

(Relative atomic masses: O = 16.0, Ca = 40.1)

$$\text{Molar mass of CaO} = (40.1 + 16.0) \text{ g mol}^{-1} = 56.1 \text{ g mol}^{-1}$$

$$\text{Number of moles of CaO used} = \frac{1.46 \text{ g}}{56.1 \text{ g mol}^{-1}} = 0.0260 \text{ mol}$$

$$\begin{aligned} \text{Enthalpy change of reaction when 1 mole of CaO reacts} &= \frac{-3.37 \text{ kJ}}{0.0260 \text{ mol}} \\ &= -130 \text{ kJ mol}^{-1} \text{ (1)} \end{aligned}$$



Topic Exercise (p.77)

b) The standard enthalpy change of the reaction between calcium oxide and hydrochloric acid is -197 kJ mol^{-1} .

Suggest TWO reasons why the enthalpy change of reaction calculated in (a)(iii) is different from this value.

Any two of the following:

- Heat loss to the surroundings (1)
- The apparatus absorbs heat (1)
- Impure calcium oxide / calcium oxide absorbs moisture from the air (1)

c) Using the information given below, calculate the standard enthalpy change of thermal decomposition of calcium carbonate.

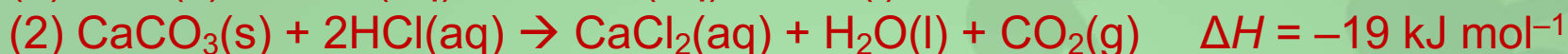
Standard enthalpy change of reaction of CaO(s) and HCl(aq) = -197 kJ mol^{-1}

Standard enthalpy change of reaction of $\text{CaCO}_3\text{(s)}$ and HCl(aq) = -19 kJ mol^{-1}

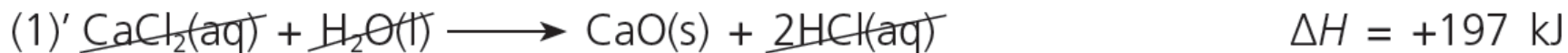


Topic Exercise (p.77)

Enthalpy changes of the following processes are known:



Reverse equation (1) to give equation (1)'. Add together the equations as shown below:



According to Hess's law:

Standard enthalpy change of thermal decomposition of $\text{CaCO}_3 = [+197 + (-19)] \text{ kJ mol}^{-1}$

(1)

$= +178 \text{ kJ mol}^{-1}$ (1)

\therefore the standard enthalpy change of thermal decomposition of calcium carbonate is $+178 \text{ kJ mol}^{-1}$.



Topic Exercise (p.77)



23 The enthalpy change of formation (ΔH_f) of $\text{Mg}(\text{OH})_2(\text{s})$ can be determined indirectly from the enthalpy change of formation of $\text{H}_2\text{O}(\text{l})$ and the enthalpy changes of *Reactions (1)* and *(2)* below.



An experiment as outlined below was carried out to determine the enthalpy change of *Reactions (1)*:

120.0 cm³ of 1.10 mol dm⁻³ HCl(aq) (excess) was placed in a polystyrene foam cup. The temperature of the acid in the cup was measured with a thermometer at half-minute intervals.

Right at the third minute, 3.45 g of $\text{Mg}(\text{OH})_2(\text{s})$ were added to the cup. The mixture in the cup was then stirred with the thermometer and its temperature was measured for an additional 5 minutes.

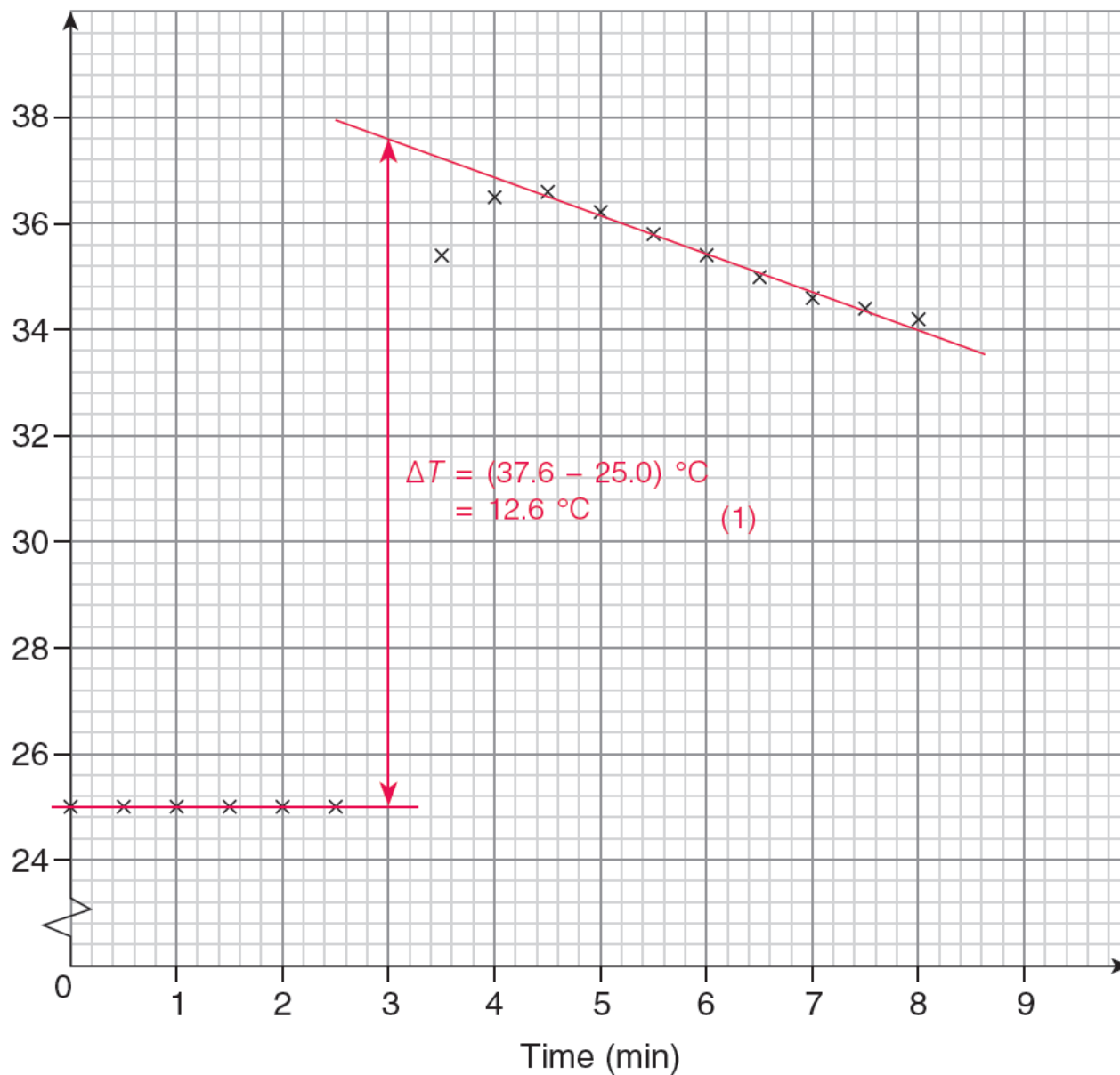
The recordings of temperature were shown in the graph below:

a) i) Deduce the greatest temperature change of the reaction mixture from the graph.



To

Temperature (°C)





Topic Exercise (p.77)

a) ii) Calculate the enthalpy change of *Reactions (1)*, in kJ mol^{-1} , under the conditions of the above experiment.

(Assume that the heat capacity of the polystyrene foam cup is negligible, and that the heat capacity of the solution is 512 J K^{-1} .)

(Relative atomic masses: $\text{H} = 1.0$, $\text{O} = 16.0$, $\text{Mg} = 24.3$)

$$\begin{aligned}\text{Enthalpy change for quantities in experiment} &= 512 \text{ J K}^{-1} \times 12.6 \text{ K} \\ &= 6\,450 \text{ J} = 6.45 \text{ kJ (1)}\end{aligned}$$

$$\text{Molar mass of Mg(OH)}_2 = (24.3 + 2 \times 16.0 + 2 \times 1.0) \text{ g mol}^{-1} = 58.3 \text{ g mol}^{-1}$$

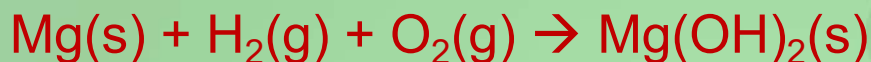
$$\text{Number of moles of Mg(OH)}_2 = \frac{3.45 \text{ g}}{58.3 \text{ g mol}^{-1}} = 0.0592 \text{ mol}$$

$$\text{Enthalpy change of Reaction (1)} = \frac{-6.45 \text{ kJ}}{0.0592 \text{ mol}} = -109 \text{ kJ mol}^{-1}$$



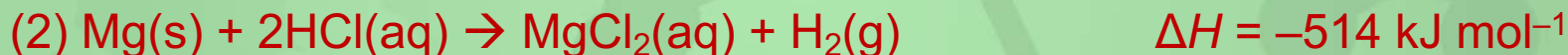
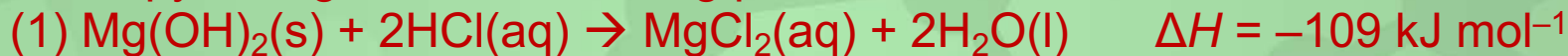
Topic Exercise (p.77)

- b) i) Write the equation which corresponds to the enthalpy change of formation of $\text{Mg(OH)}_2(\text{s})$.

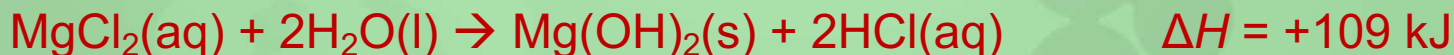


- ii) Given that under the same conditions, the enthalpy change of *Reactions (2)* is -514 kJ mol^{-1} , and the enthalpy change of formation of $\text{H}_2\text{O(l)}$ is -286 kJ mol^{-1} , calculate ΔH_f of $\text{Mg(OH)}_2(\text{s})$.

Enthalpy changes of the following processes are known:



Reverse equation (1) to give equation (1)'.



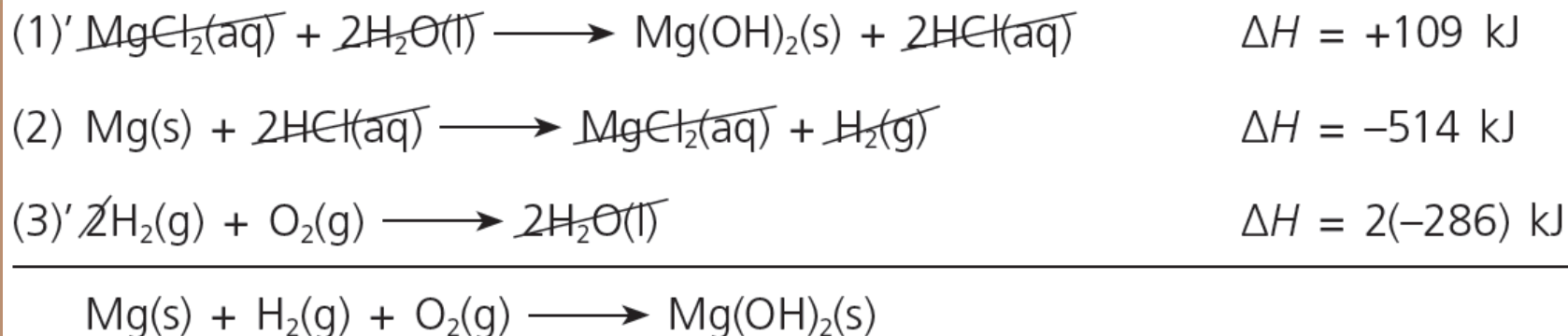
Multiply equation (3) by 2 to give equation (3)'.



Add together the equations as shown below:



Topic Exercise (p.77)



According to Hess's law:

$$\Delta H_f[\text{Mg}(\text{OH})_2(\text{s})] = [(+109) + (-514) + 2(-286)] \text{ kJ mol}^{-1} \quad (1)$$

$$= -977 \text{ kJ mol}^{-1} \quad (1)$$

$\therefore \Delta H_f$ of $\text{Mg}(\text{OH})_2(\text{s})$ is -977 kJ mol^{-1} .



Topic Exercise (p.77)

24 A thermometric titration was carried out as outlined below.



- 25.0 cm³ of dilute ethanoic acid were placed in a polystyrene foam cup and its temperature measured.
- 5.00 cm³ of 1.24 mol dm⁻³ aqueous ammonia were added to the acid, the mixture stirred and the temperature measured again.
- Further 5.00 cm³ portions of aqueous ammonia were added, followed by temperature measurement, until a total of 50.0 cm³ of aqueous ammonia had been added.

The graph shows the results obtained.

- a) Explain why the temperature of the liquid mixture rose until the maximum temperature was reached.

The reaction was exothermic. / Heat is released until the reaction was complete. (1)

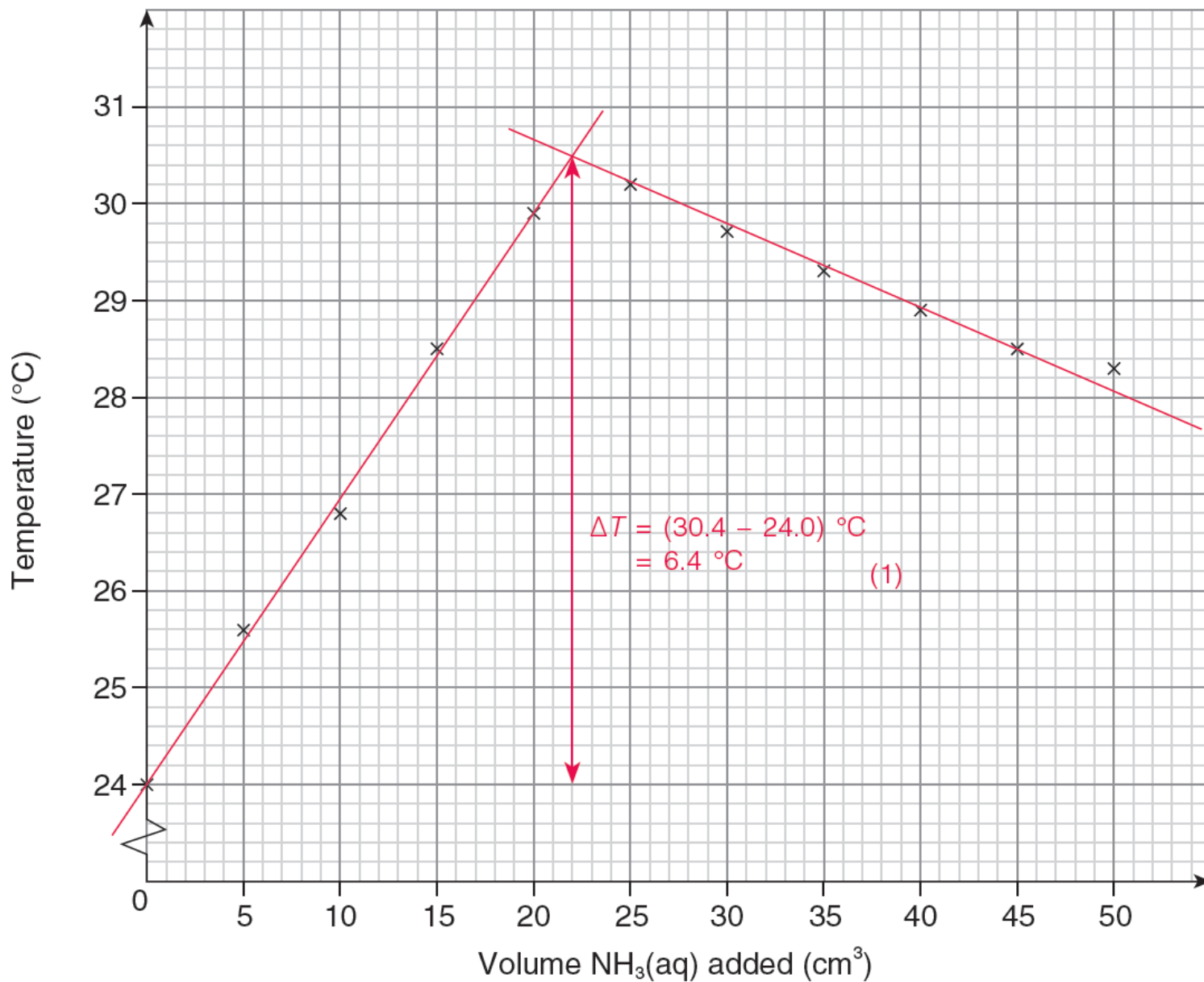
- b) Explain why the temperature of the liquid mixture fell on adding continuously the aqueous ammonia after the maximum temperature had been reached.

No further reaction occurred. The excess aqueous ammonia cooled the solution. (1)

- c) From the graph, find

- i) the greatest temperature rise; and **(Refer to the graph)**
- ii) the volume of aqueous ammonia required to neutralise the acid.

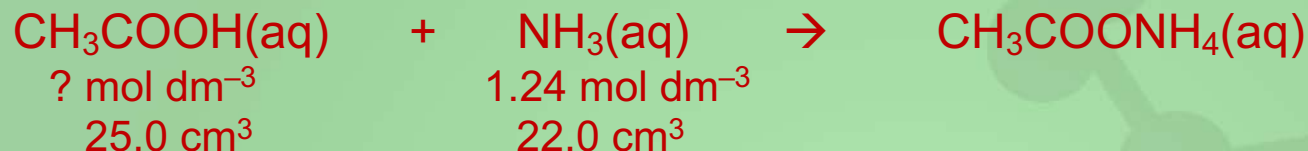
22.0 cm³ (1)





Topic Exercise (p.77)

d) Use your value from (c)(ii) to calculate the concentration of the ethanoic acid.



$$\begin{aligned} \text{Number of moles of NH}_3 \text{ in } 22.0 \text{ cm}^3 \text{ solution} &= 1.24 \text{ mol dm}^{-3} \times \frac{22.0}{1\,000} \text{ dm}^3 \\ &= 0.0273 \text{ mol (1)} \end{aligned}$$

According to the equation, 1 mole of CH_3COOH requires 1 mole of NH_3 for neutralisation.

i.e. number of moles of CH_3COOH in 25.0 cm^3 solution = 0.0273 mol

$$\begin{aligned} \text{Concentration of ethanoic acid} &= (0.0273 \text{ mol}) / \left(\frac{25.0}{1\,000} \text{ dm}^3 \right) \\ &= 1.09 \text{ mol dm}^{-3} \quad (1) \end{aligned}$$



Topic Exercise (p.77)

- e) Calculate the enthalpy change of neutralisation between ethanoic acid and aqueous ammonia.

(Assume that the density of the liquid mixture was 1.00 g cm^{-3} and that its specific heat capacity was $4.2 \text{ J g}^{-1} \text{ K}^{-1}$)

$$\text{Volume of reaction mixture} = (25.0 + 22.0) \text{ cm}^3 = 47.0 \text{ cm}^3$$

$$\text{Mass of reaction mixture} = 47.0 \text{ g}$$

$$\begin{aligned} \text{Amount of heat released} &= m \times c \times \Delta T = 47.0 \text{ g} \times 4.2 \text{ J g}^{-1} \text{ K}^{-1} \times 6.4 \text{ K} \\ &= 1\,260 \text{ J} = 1.26 \text{ kJ (1)} \end{aligned}$$

$$\text{Number of moles of } \text{NH}_3 \text{ reacted} = 0.0273 \text{ mol}$$

$$\text{Enthalpy change of neutralisation} = \frac{-1.26 \text{ kJ}}{0.0273 \text{ mol}} = -46.2 \text{ kJ mol}^{-1} \text{ (1)}$$

\therefore the enthalpy change of neutralisation between ethanoic acid and aqueous ammonia is $-46.2 \text{ kJ mol}^{-1}$.



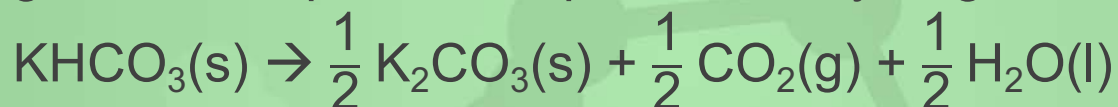
Topic Exercise (p.77)



25 You are provided with common laboratory apparatus, and the following chemicals:

- solid potassium carbonate;
- solid potassium hydrogencarbonate; and
- 2 mol dm⁻³ hydrochloric acid.

Outline how you would perform an experiment to determine the enthalpy change of decomposition of potassium hydrogencarbonate.



Briefly describe the calculations involved.

Allow a known mass of potassium carbonate to react with a known volume of 2 mol dm⁻³ hydrochloric acid (excess) in a polystyrene foam cup. Measure the maximum temperature change of the resulting solution. (1)

Calculate

- the amount of heat involved in the reaction;
 - number of moles of potassium carbonate reacted;
 - enthalpy change of reaction between 1 mole of potassium carbonate and hydrochloric acid.
- (1)



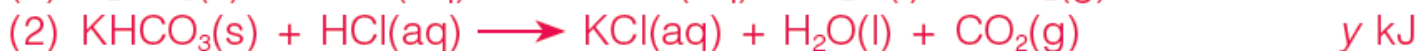
Topic Exercise (p.77)

Allow a known mass of potassium hydrogencarbonate to react with the same volume of 2 mol dm^{-3} hydrochloric acid (excess) in a polystyrene foam cup. Measure the maximum temperature change of the resulting solution. (1)

Calculate

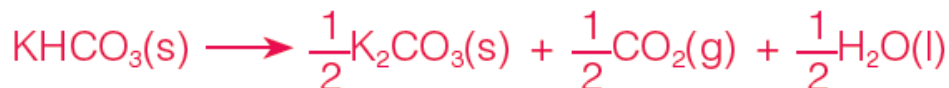
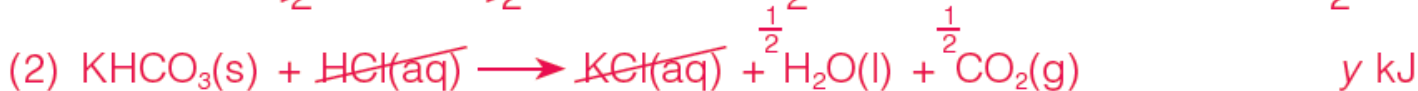
- the amount of heat involved in the reaction;
- number of moles of potassium hydrogencarbonate reacted;
- enthalpy change of reaction between 1 mole of KHCO_3 and HCl .

The enthalpy changes of the following processes are known:



Multiply equation (1) by $\frac{1}{2}$ and reverse it to give equation (1)'.

Add together the equations as shown below:



According to Hess's law:

$$\text{Enthalpy change of decomposition of } \text{KHCO}_3(\text{s}) = \left(-\frac{x}{2} + y\right) \text{ kJ mol}^{-1}$$

Communication mark

(2)

(1)