A Level · Cambridge (CIE) · Chemistry

43 questions

Hess's Law and Enthalpy

Mark scheme

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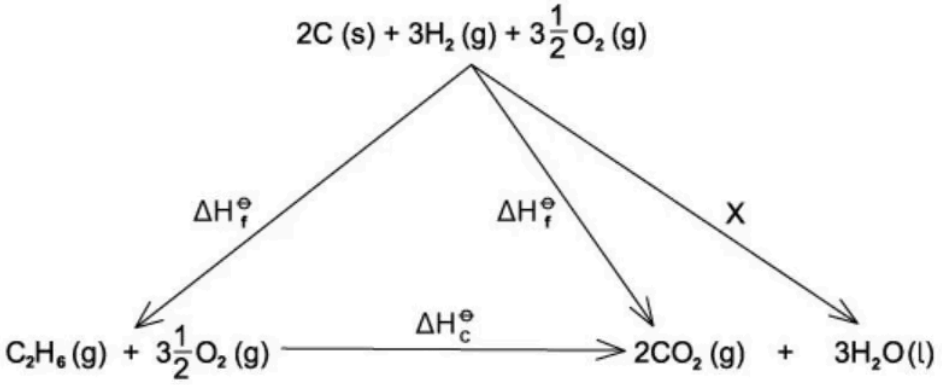
Total Marks **/136**

1 Enthalpy changes that are difficult to measure directly can often be determined using Hess' Law to construct an enthalpy cycle.

Which enthalpy change is indicated by X in the enthalpy cycle shown?



A. +1 x Enthalpy of formation of water

B. -1 x Enthalpy of formation of water

C. +3 x Enthalpy of formation of water

D. -3 x Enthalpy of formation of water

(1 mark)

2 (a) Standard enthalpy changes include enthalpy of formation, enthalpy of combustion and enthalpy of neutralisation.

i) Define *enthalpy of formation.*

[1]

ii) Write an equation that represents the enthalpy of formation of ammonia, NH_3.

[1]

(2 marks)

(b) State what is meant by *standard* conditions in enthalpy measurements.

(1 mark)

(c) Which of the following equation(s) can represent the enthalpy of neutralisation?

A. H_2SO_4 (aq) + 2NaOH (aq) → Na_2SO_4 (aq) + 2H_2O (l)

B. H_2SO_4 (aq) + Na_2CO_3 (aq) → Na_2SO_4 (aq) + H_2O (l) + CO_2 (g)

C. HCl (aq) + NaOH (aq) → NaCl (aq) + H_2O (l)

Explain your answer.

(2 marks)

3 Which equation below can represent both an enthalpy change of formation and combustion?

A. CH_4 (g) + 2O_2 (g) → CO_2 (g) + 2H_2O (l)

B. 2Na (s) + ½O_2 (g) → Na_2O (s)

C. HCl (aq) + NaOH (aq) → NaCl (aq) + H_2O (l)

D. C (s) + O_2 (g) → CO_2 (g)

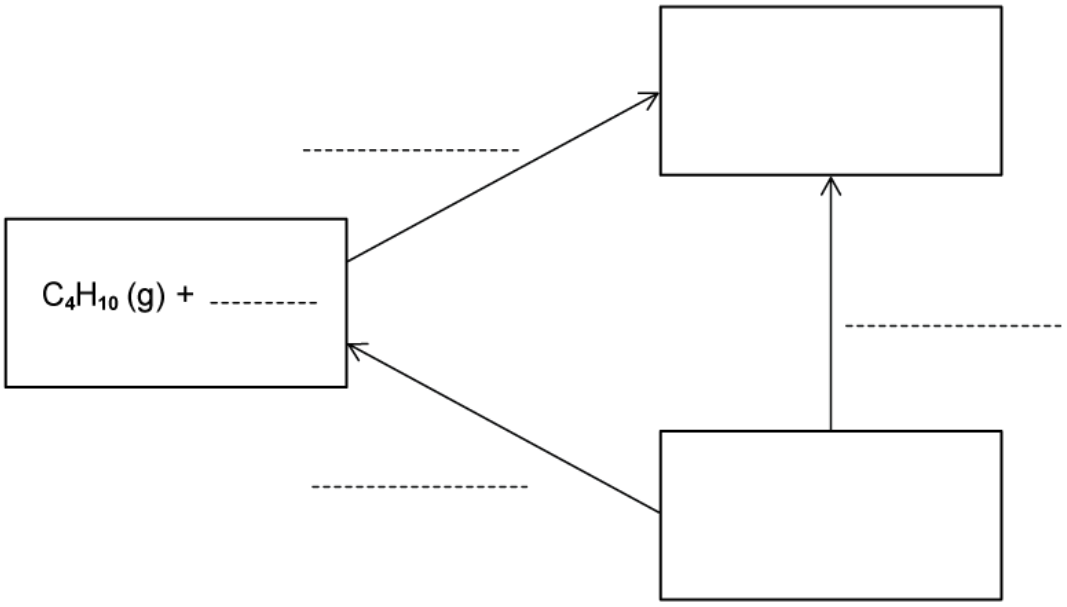
(1 mark)

4 (a) Complete the following Hess' Law energy cycle relating butane to enthalpies of formation and combustion.

- enthalpy of formation, $\Delta H^{\ominus}_f$
- enthalpy of combustion, $\Delta H^{\ominus}_c$

On your diagram:

- include the relevant species in the two empty boxes,
- label each enthalpy change with its appropriate symbol,
- complete the remaining two arrows showing the correct direction of enthalpy change



(4 marks)

(b) Use the data in Table 1.1 to calculate the enthalpy of combustion of butane, ΔH^{θ}_{c}

Table 1.1

Substance	**Enthalpy of formation, ΔH^{θ}_{f} / kJ mol^{-1}**
CO_2 (g)	-393.5
C_4H_{10} (g)	-125
H_2O (g)	-242

ΔH^{θ}_{c} (C_4H_{10}) = ...kJ mol^{-1}

(1 mark)

(c) Using the values given in **Table 1.2**, calculate the enthalpy of combustion of butane.

Table 1.2

Bond	**Bond Energy / kJmol^{-1}**
C-C	348
C-H	410
C=O	743
H-O	463
O=O	495

ΔH^{θ}_{c} (C_4H_{10}) = ...kJ mol^{-1}

(1 mark)

(d) Suggest, with a reason, which of the two values of enthalpy of combustion from parts **(b)** and **(c)** is likely to be more accurate.

(2 marks)

5 (a) Propane gas is commonly used as a fuel for outdoor cooking. It can be produced in a number of ways, including from the addition reaction of propyne gas with hydrogen. Propyne has the formula $CHCCH_3$ and includes a triple bond.

Construct an equation for the formation of propane from propyne. Include state symbols.

(2 marks)

(b) Table 2.1 lists the relevant enthalpy of combustion data for the formation of propane from propyne.

Table 2.1

	Enthalpy of combustion, $\Delta H^{\ominus}_c$ / kJ mol^{-1}
Hydrogen	-285.8
Propane	-2220
Propyne	-1940

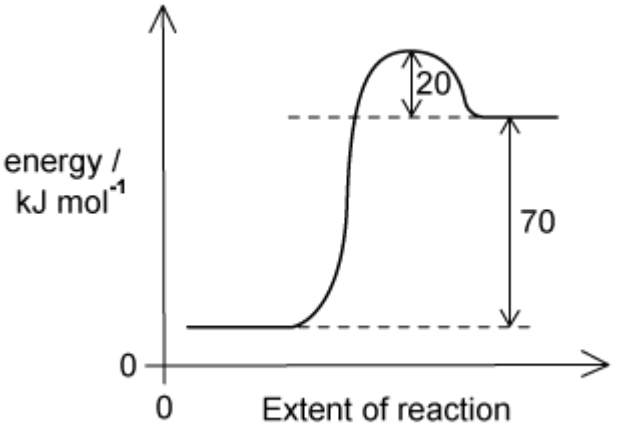
Calculate the enthalpy change in kJ mol^{-1} for the formation of propane from propyne. Show all working.

(3 marks)

(c) Suggest why a calculated bond enthalpy value is often different to the data book value.

(1 mark)

6 The reaction pathway for a reversible reaction is shown below:



Which statement is correct?

A. The activation energy of the reverse reaction is +90 kJ mol^{-1}

B. The activation energy of the forward reaction is +20 kJ mol^{-1}

C. The activation of the reverse reaction is +20 kJ mol^{-1}

D. The enthalpy change of forwards reaction is - 70 kJ mol^{-1}

(1 mark)

7 (a) Freon 21 is a hydrochlorofluorocarbon that was used as a propellant and refrigerant but later withdrawn from use because of its effect on ozone depletion. The formula is $CHCl_2F$.

Freon 21 is produced by reacting hydrogen fluoride with trichloromethane in the following reaction:

$$CHCl_3\ (g) + HF\ (g) \rightarrow CHCl_2F\ (g) + HCl\ (g)$$

Use the data in Table 2.1 to calculate the enthalpy of reaction, ΔH_r, for the formation of $CHCl_2F(g)$.

compound	enthalpy change of formation, ΔH_f / kJmol^{-1}
$CHCl_3(g)$	–103.2
$CHCl_2F(g)$	–284.1
HF(g)	–273.3
HCl(g)	–92.3

Table 2.1

enthalpy change of reaction, ΔH_r = ... kJ mol^{-1}

(2 marks)

(b) Using Fig. 2.1, draw a labelled reaction pathway diagram for part (a) and include the activation energy.

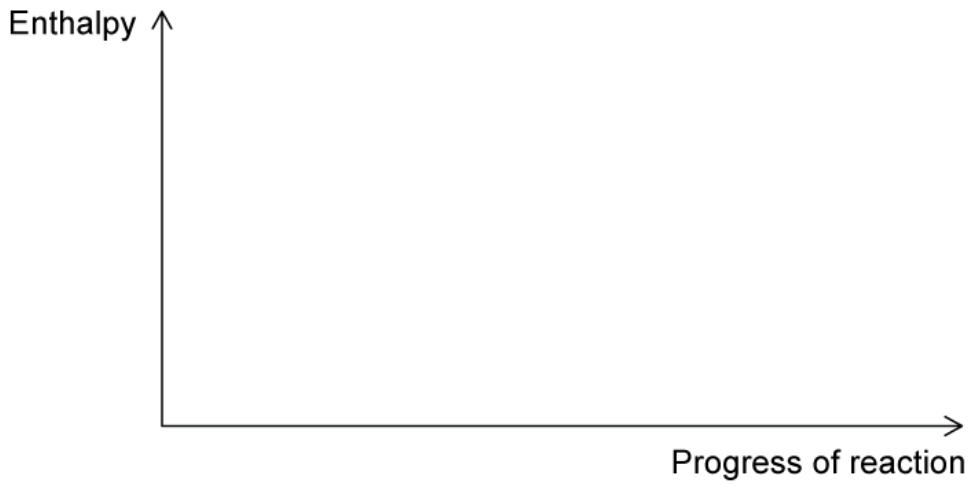


Fig. 2.1

(2 marks)

(c) The reaction pathway diagram has a feature known as the **transition state**. Explain what this means.

(1 mark)

8 Which is the correct equation to calculate the enthalpy of reaction using enthalpy of formation data?

A. $\Delta H_r = \Sigma \Delta H_{f\ products} + \Sigma \Delta H_{f\ reactants}$

B. $\Delta H_r = \Sigma \Delta H_{f\ products} - \Sigma \Delta H_{f\ reactants}$

C. $\Delta H_r = \Sigma \Delta H_{f\ reactants} - \Sigma \Delta H_{f\ products}$

D. $\Delta H_r = \Sigma \Delta H_{f\ reactants} + \Sigma \Delta H_{f\ products}$

(1 mark)

9 (a) Enthalpy changes of combustion can be determined using calorimetry or calculated using Hess cycles.

Fig. 3.1 shows the equipment required to determine the enthalpy of combustion of 2-methylpropan-2-ol, C_4H_9OH.

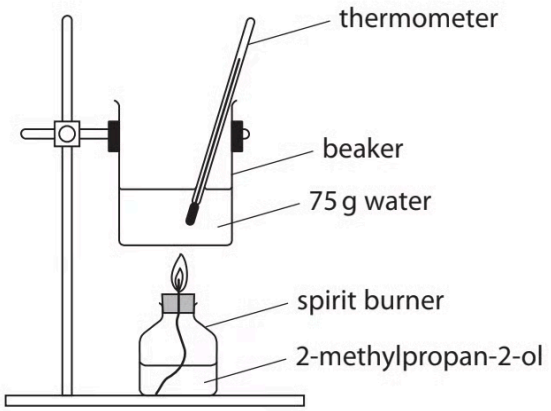


Fig. 3.1

Table 3.1 shows the results of the experiment.

Table 3.1

	Initial	Final	Change
Mass of spirit burner / g	267.35	266.78	
Temperature of water / °C	19.5	65.3	

Calculate the enthalpy change of combustion, ΔH_c, of 2-methylpropan-2-ol. Include the sign and units.

(5 marks)

(b) The standard enthalpy change of combustion, ΔH_c^{θ} of 2-methylpropan-2-ol can be calculated using the data shown in Table 3.2.

Table 3.2

Compound	**ΔH_f^{θ} / kJ mol^{-1}**
2-methylpropan-2-ol	–359
carbon dioxide	–394
water	–286

i) State why oxygen does not have a value for ΔH_f^{θ}.

[1]

ii) Complete Fig. 3.2, showing the Hess' Law energy cycle for the combustion of 2-methylpropan-2-ol.

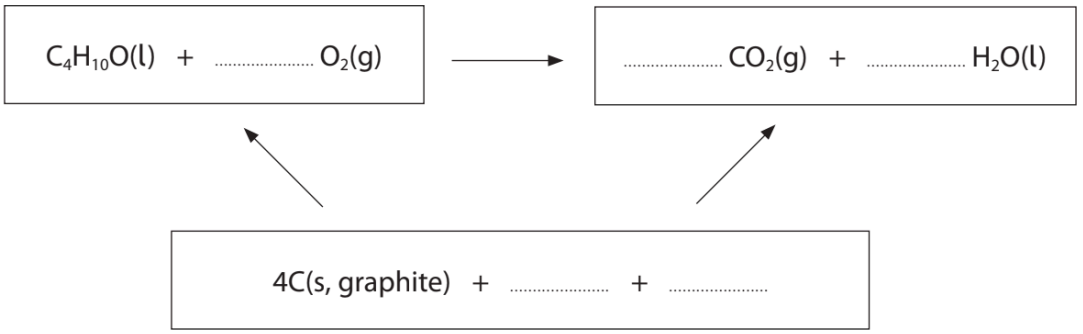


Fig. 3.2

[2]

iii) Calculate the standard enthalpy change of combustion of 2-methylpropan-2-ol.

[2]

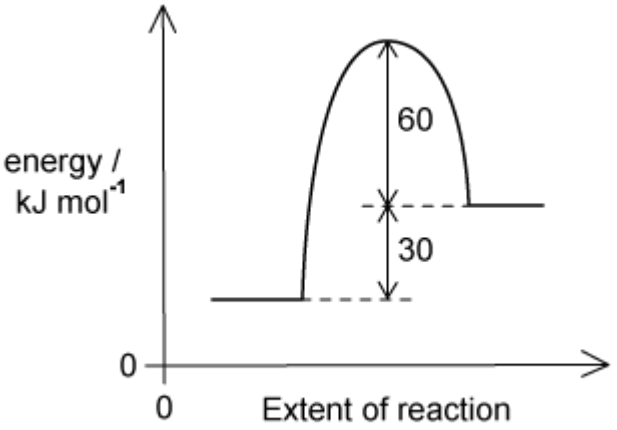
(5 marks)

(c) The value for ΔH_c^{θ} obtained in part **(a)** is much less exothermic than ΔH_c^{θ} calculated in **(b)(iii)**.

Suggest **two** reasons for this, other than non-standard conditions.

(2 marks)

10 The reaction pathway for a reversible reaction is shown below.



Which statements are correct?

1. The forward reaction is endothermic.
2. The enthalpy change for the backward reaction is –30 kJ mol^{-1}.
3. The activation energy for the forward reaction is +90 kJ mol^{-1}.

A. 1 only

B. 1 and 2

C. 2 and 3

D. 1, 2 and 3

(1 mark)

11 (a) Define the term average bond enthalpy.

(2 marks)

(b) Give the formula for calculating the standard enthalpy change of reaction, ΔH_r, using bond energies.

(1 mark)

(c) Use the data given in Table 3.1 to calculate the enthalpy change for the hydration of ethene to form ethanol.

CH_2CH_2 (g) + H_2O (g) → CH_3CH_2OH (l)

Table 3.1

Bond	**Bond energy / kJ mol^{-1}**
C-C	346
C=C	614
C-H	414
C-O	358
O-H	463

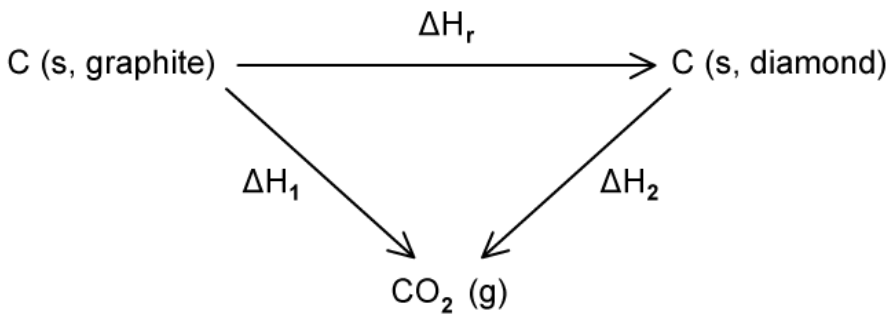
(3 marks)

(d) Use Table 3.2 to calculate the following enthalpy change for the reaction of ethane and bromine to form 1,2-dibromoethane and hydrogen bromide.

Bond	Bond enthalpy / kJ mol^{-1}
C-C	346
C-H	414
C-Br	285
H-Br	366
Br-Br	193

(4 marks)

12 Hess's Law can be used to calculate the enthalpy change for reactions that are difficult to measure experimentally, such as the conversion of graphite to diamond.



Which equation shows the correct application of Hess's law to calculate the enthalpy change for the conversion of graphite to diamond?

A. $\Delta H_r = \Delta H_1 + \Delta H_2$

B. $\Delta H_r = \Delta H_1 - \Delta H_2$

C. $\Delta H_r = \Delta H_2 - \Delta H_1$

D. $\Delta H_r = \Delta H_1 \times \Delta H_2$

(1 mark)

13 Propanol has the molecular formula C_3H_8O. Use the following information to calculate the enthalpy change for the formation of propanol.

The enthalpy change of combustion of hydrogen is –286 kJ mol^{-1}

The enthalpy change of combustion of carbon is –394 kJ mol^{-1}

The enthalpy change of combustion of propanol is – 1816 kJ mol^{-1}

A. -1816 kJ mol^{-1}

B. -510 kJ mol^{-1}

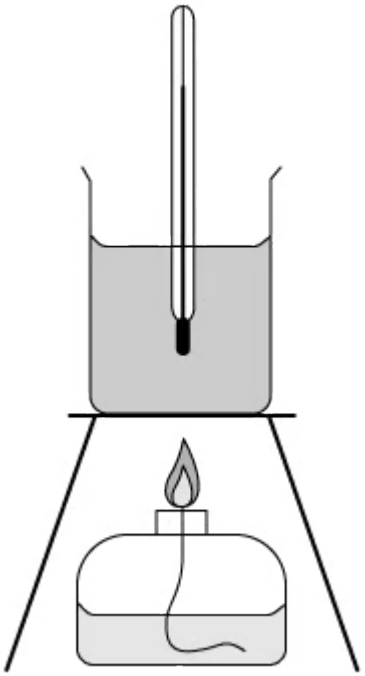
C. +510 kJ mol^{-1}

D. +1136 kJ mol^{-1}

(1 mark)

14 A student carried out an experiment to determine the enthalpy change for the combustion of ethanol.

The following results were obtained by the student. The specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$.



Start temperature of the water	21 °C
Final temperature of the water	54 °C
Mass of alcohol burner before burning	259.75 g
Mass of alcohol burner after burning	259.18 g
Mass of glass beaker plus water	150.00 g
Mass of glass beaker	50.0 g

How much of the heat energy produced by the burning of ethanol went into the water?

A. 6897 J

B. 20691 J

C. 22572 J

D. 13794 J

(1 mark)

15 (a) Define the term *standard enthalpy of combustion,* $\Delta H^{\ominus}_{c}$.

(3 marks)

(b) Construct an equation for the complete combustion of propanol, $CH_3CH_2CH_2OH$ (l).

(2 marks)

(c) Construct a Hess's Law cycle for the complete combustion of propanol.

(3 marks)

(d) Use the enthalpy of formation data given in Table 4.1, to calculate the enthalpy change of the reaction, $\Delta H^{\ominus}_{r}$.

Table 4.1

	$CH_3CH_2CH_2OH$ (l)	O_2 (g)	CO_2 (g)	H_2O (l)
$\Delta H^{\ominus}_f$ (kJ mol^{-1})	-303	0	-393.5	-285.8

(3 marks)

16 (a) Propane reacts with chlorine to form chloropropane.

C_3H_8 (g) + Cl_2 (g) → C_3H_7Cl (g) + HCl (g)

i) Use bond energies from Table 1.1 to calculate the enthalpy change for this reaction.

Table 1.1

Bond	**Bond Energy**
C-H	410
C-Cl	340
Cl-Cl	242
H-Cl	431

Include a sign in your answer.

enthalpy change = ... kJ mol^{-1}

[3]

ii) State the conditions needed for this reaction to occur.

[1]
(4 marks)

(b) Using Fig. 1.1, construct a labelled reaction pathway diagram for the reaction in part **(a)**

including the activation energy.

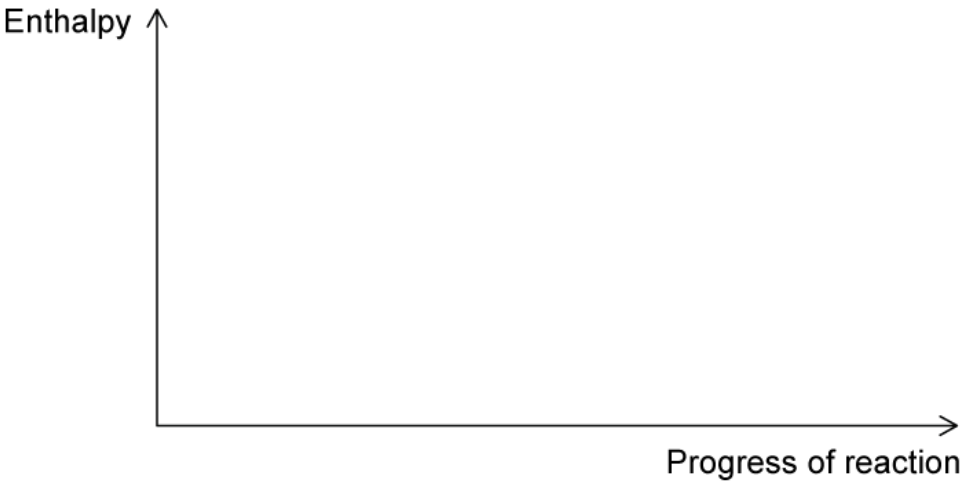


Fig. 1.1

(3 marks)

(c) Ethane and chlorine will also react together under the same conditions.

State and explain the difference in enthalpy change for the two reactions

(2 marks)

(d) Propane will react with bromine in a similar reaction to part **(a)**.

Suggest, with a reason, whether the sum of the bonds broken would be larger or smaller compared to the reaction between propane and chlorine.

(2 marks)

17 PCl_5 (g) dissociates as follows.

$$PCl_5\ (g) \rightarrow PCl_3\ (g) + Cl_2\ (g)$$

The yield of the products increases as the temperature is raised from 80^{o}C to 110^{o}C.

Which row correctly describes the formation of PCl_3 (g)?

	Shape of PCl_3 molecule	The reaction is
A	Pyramidal	Endothermic
B	Pyramidal	Exothermic
C	Trigonal	Endothermic
D	Trigonal	Exothermic

(1 mark)

18 (a) Urea, $CO(NH_2)_2$, is a naturally occurring substance which can be hydrolysed with water to form ammonia and carbon dioxide.

The standard enthalpy changes of formation of water, urea, carbon dioxide and ammonia (in aqueous solution) are given below in Table 1.1

Table 1.1

compound	**$\Delta H^{\theta}f$ / kJ mol^{-1}**
H_2O (l)	–287.0
$CO(NH_2)_2$ (aq)	–320.5
CO_2 (g)	–414.5
NH_3 (aq)	–81.0

i) Write an equation for the reaction.

[1]

ii) Construct a simple energy cycle for the hydrolysis of urea.

[2]

(3 marks)

(b) Use these data to calculate the standard enthalpy change for the hydrolysis of urea.

(1 mark)

(c) Calculate the enthalpy of combustion of urea, given the following equation and the data in Table 1.1.

$$2CO(NH_2)_2\,(s) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 2N_2\,(g) + 4H_2O\,(l)$$

(1 mark)

19 (a) The apparatus shown in Fig. 2.1 can be used to determine the enthalpy of combustion of butan-1-ol, C_4H_9OH (M_r = 74.12 $gmol^{-1}$).

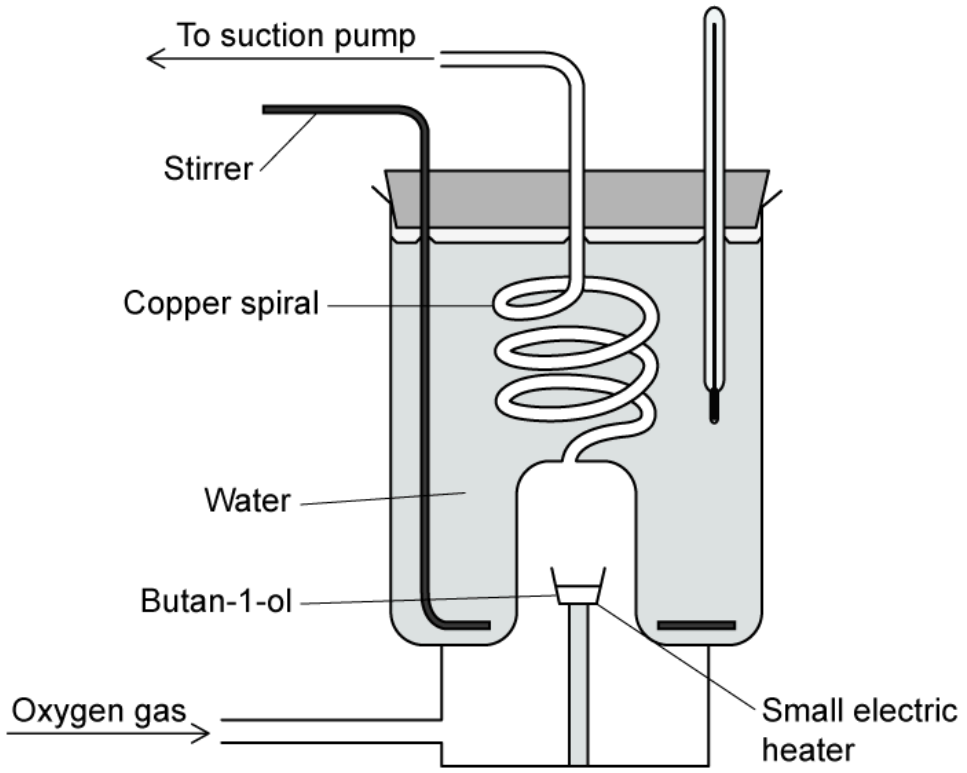


Fig. 2.1

i) Write an equation to represent the enthalpy of combustion of butan-1-ol.

[2]

ii) Suggest the purpose of the copper spiral and small electric heater in the apparatus.

[2]
(4 marks)

(b) An experiment was carried out and the follow measurements were recorded in Table 2.1

Table 2.1

Mass of butan-1-ol / g	2.20 g	Initial temperature of water / ^{o}C	22.5
Volume of water / cm^3	875 cm^3	Final temperature of water / ^{o}C	25.0

(Specific heat capacity of water = 4.18 J $g^{-1}K^{-1}$)

Use the data to calculate the enthalpy of change, *q*, for the reaction. Give your answer to two significant figures.

q = ... kJ

(2 marks)

(c) Determine the enthalpy of combustion of butan-1-ol using Table 2.1 and your answer to part **(b)**.

Give your answer to three significant figures.

ΔH = ... kJ mol^{-1}

(3 marks)

20 The first stage in the industrial production of nitric acid from ammonia can be represented by the following equation.

$$4NH_3\,(g) + 5O_2\,(g) \rightleftharpoons 4NO\,(g) + 6H_2O\,(g)$$

Using the following standard enthalpy change of formation data, what is the value of the standard enthalpy change ΔH_f^{θ} for this reaction?

Compound	ΔH_f^{θ} /kJ mol^{-1}
NH_3 (g)	-51.3
NO (g)	+92.2
H_2O(g)	-239.6

A. + 863.6 kJ mol^{-1}

B. - 1601.2 kJ mol^{-1}

C. - 863.6 kJ mol^{-1}

D. - 1274.0 kJ mol^{-1}

(1 mark)

21 (a) Two common oxides of nitrogen are nitrogen monoxide, NO, and nitrogen dioxide, NO_2.

i) Complete Table 3.1 to show the oxidation number of nitrogen in each compound.

Table 3.1

Compound	**NO**	**NO_2**
oxidation number of N		

[1]

ii) Write equations for the formation of NO_2 by:

- The reaction of N_2 with O_2
- The reaction of NO with O_2

[2]

(1 mark)

(b) Molecules of NO_2 can be formed by the reaction between N_2 and O_2. The bond between the N and O atoms in NO_2 is a double covalent bond.

The enthalpy change of reaction for this reaction is +497 kJ mol^{-1}. Calculate the bond enthalpy, in kJ mol^{-1}, of the N=O bond. Use relevant data from Table 3.2.

Table 3.2

Bond	**Bond energy / kJ mol^{-1}**
N≡N	941
O=O	495

Bond enthalpy of the N=O bond = ... kJ mol^{-1} [2]

(3 marks)

(c) The boiling points of carbon dioxide and nitrogen dioxide are -78.5 ^{o}C and 21 ^{o}C respectively.

Suggest a reason for the difference.

(2 marks)

(d) Write an equation for the enthalpy of formation of nitrogen monoxide, NO.

(1 mark)

22 Hydrogen atoms bond covalently to Iodine atoms to form hydrogen iodide as shown in the equation below:

$$H_2\,(g) + I_2\,(g) \rightarrow 2HI\,(g)$$

Which statement best describes what is meant by the average HI bond enthalpy?

A. The energy stored in a covalent bond.

B. The energy required to break one covalent bond in the gas phase.

C. The energy required to break one mole of the HI bonds in the gas phase.

D. The energy released when two atoms form a covalent bond.

(1 mark)

23 Titanium occurs naturally as the mineral rutile, TiO_2. One possible method of extraction of titanium is to reduce the rutile by heating with carbon.

$$TiO_2\,(s) + 2C\,(s) \rightarrow Ti\,(s) + 2CO\,(g)$$

The standard enthalpy changes of formation of TiO_2 (s) and CO (g) are –890 kJ mol^{-1} and –105 kJ mol^{-1} respectively.

What is the standard enthalpy change of the extraction of titanium?

A. + 450 kJ mol^{-1}

B. + 680 kJ mol^{-1}

C. - 450 kJ mol^{-1}

D. - 660 kJ mol^{-1}

(1 mark)

24 (a) Propane gas is used widely as a fuel which can be be used in camping gas stoves. The enthalpy of combustion for propane is -2219.2 kJ mol^{-1}.

What is the minimum mass of fuel is needed to bring 150 cm^3 of water at 21 °C to its boiling point?

Show your working.

..... g

(3 marks)

(b) Suggest why more propane may be required than calculated in part **(a)**.

(1 mark)

(c) Using Fig. 5.1, construct a labelled reaction pathway diagram for the reaction in part **(a)** including the activation energy.

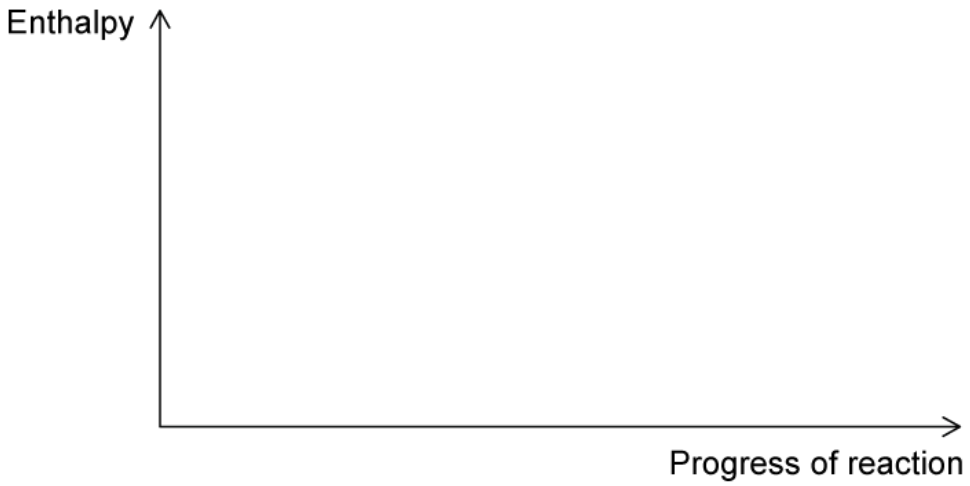


Fig. 5.1

(3 marks)

25 A student mixed 30.0 cm^3 of 0.0250 mol dm^{-3} potassium hydroxide solution with 30.0 cm^3 of 0.0250 mol dm^{-3} nitric acid. The temperature rose by 0.50 ^oC. Assume no heat was lost to the surroundings.

The final mixture had a specific heat capacity of 4.18 J cm^{-3} K^{-1}.

What is the molar enthalpy change for the reaction?

A. -84.0 kJ mol^{-1}

B. -83.6 kJ mol^{-1}

C. 167.2 kJ mol^{-1}

D. -167.2 kJ mol^{-1}

(1 mark)

26 The combustion of ethanol (C_2H_5OH) is increasingly being used to fuel cars.

The standard enthalpy change of formation of carbon dioxide is –382 kJ mol^{-1}.

The standard enthalpy change of formation of water is –275 kJ mol^{-1}.

The standard enthalpy change of formation of ethanol is −266kJ mol^{-1}.

What is the standard enthalpy change of combustion of ethanol?

A. - 1367 kJ mol^{-1}

B. - 1323 kJ mol^{-1}

C. - 391 kJ mol^{-1}

D. - 948 kJ mol^{-1}

(1 mark)

27 The equations below show the formation of sulfur oxides from sulfur and oxygen.

$$S\ (s) + O_2\ (g) \rightarrow SO_2\ (g) \quad \Delta H_f^{\theta} = -297\ kJ\ mol^{-1}$$

$$S\ (s) + 1½\ O_2\ (g) \rightarrow SO_3\ (g) \quad \Delta H_f^{\theta} = -395\ kJ\ mol^{-1}$$

What is the enthalpy change of reaction, ΔH_r^{θ}, of $2SO_2\ (g) + O_2\ (g) \rightarrow 2SO_3\ (g)$?

A. + 98 kJ mol^{-1}

B. - 98 kJ mol^{-1}

C. + 196 kJ mol^{-1}

D. - 196 kJ mol^{-1}

(1 mark)

28 (a) The enthalpy change of solution for ammonium chloride can be measured using calorimetry. 12.04 g of ammonium chloride is dissolved in 125.0 cm^3 of water at 19.5 °C.

The enthalpy of solution of ammonium chloride is +15.1 kJ mol^{-1}. Determine the energy change, in J, when 12.04 g of ammonium chloride is dissolved in 125.0 cm^3 of water.

(2 marks)

(b) Use your answer to part (a) to determine the change in temperature, in °C, when the ammonium chloride is dissolved.

(1 mark)

(c) Use your answer to part (b) to determine the final temperature, in °C, of the solution during the reaction.

(1 mark)

29 An experiment was carried out to determine the approximate value for the molar enthalpy change of neutralisation.

75 cm^3 of 3.00 mol dm^{-3} hydrochloric acid was placed in a polystyrene beaker of negligible heat capacity. Its temperature was recorded, and then 75 cm^3 of 3.00 mol dm^{-3} potassium hydroxide at the same temperature was quickly added, and the solution was stirred.

The temperature rose by 14 °C. The resulting solution may be considered to have a specific heat capacity of 4.18 J g^{-1} K^{-1}.

Calculate the molar enthalpy change of neutralisation.

A. -4876.7 J mol^{-1}

B. -39 013.3 J mol^{-1}

C. -39.0 J mol^{-1}

D. -11 843.3 J mol^{-1}

(1 mark)

30 (a) Determine the enthalpy of hydrogenation of propene using the data in Table 2.1.

Table 2.1

Bond	Bond enthalpy / kJ mol^{-1}
H-H	435
C-H	413
C-C	347
C=C	619

Enthalpy of hydrogenation, ΔH_r = ... kJ mol^{-1}

(3 marks)

(b) Use the data in Table 2.1 to suggest, with a reason, whether the polymerisation of propene is exothermic or endothermic

(2 marks)

(c) Carbon, hydrogen and ethene each burn exothermically in an excess of air.

$C\,(s) + O_2\,(g) \rightarrow CO_2\,(g)$	$\Delta H^{\theta}_c = -393.7$ kJ mol^{-1}
$H_2\,(g) + ½O_2\,(g) \rightarrow H_2O\,(l)$	$\Delta H^{\theta}_c = -285.9$ kJ mol^{-1}
$C_2H_4\,(g) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 2H_2O\,(l)$	$\Delta H^{\theta}_c = -1411.0$ kJ mol^{-1}

Use the data to calculate the standard enthalpy change of formation, ΔH^{θ}_f, in kJ mol^{-1}, of ethene at 298 K.

$$2C\,(s) + 2H_2\,(g) \rightarrow C_2H_4\,(g)$$

ΔH^{θ}_f = ... kJ mol^{-1}

(2 marks)

31 The standard enthalpy change ΔH^{θ} for the below reaction is -1776 kJ.

$$2N\ (g) + 6F\ (g) \rightarrow 2NF_3\ (g)$$

What is the bond energy of the N–F bond?

A. - 592

B. + 592

C. + 296

D. - 296

(1 mark)

32 The complete combustion of ethyne, C_2H_2, is shown in the equation below.

$$H{-}C{\equiv}C{-}H + 2½\ O{=}O \rightarrow O{-}H{-}H + 2O{=}C{=}O$$

Using the average bond enthalpies given in the table, what is the enthalpy change of combustion of ethyne?

Bond	Average bond enthalpy / kJ mol^{-1}
C−H	390
C≡C	870
O=O	489
C=O	790
O−H	510
O−C	290

A. + 1307.5 kJ mol^{-1}

B. - 1307.5 kJ mol^{-1}

C. + 1390 kJ mol^{-1}

D. - 1390 kJ mol^{-1}

(1 mark)

33 A student calculated the standard enthalpy change of formation of propane, C_3H_8, using a method based on standard enthalpy changes of combustion.

He used correct values for the standard enthalpy change of combustion of propane (–2220 kJ mol^{-1}) and hydrogen (–286 kJ mol^{-1}) but he used an incorrect value for the standard enthalpy change of combustion of carbon. He then performed his calculation correctly. His final answer was –158 kJ mol^{-1}.

What did he use for the standard enthalpy change of combustion of carbon?

A. -1234 kJ mol^{-1}

B. -411.3 kJ mol^{-1}

C. -2084 kJ mol^{-1}

D. -694.7 kJ mol^{-1}

(1 mark)

34 Given the following enthalpy changes,

$I_2\,(s) \rightarrow I_2\,(g)$ $\quad \Delta H^{\ominus} = +38$ kJ mol^{-1}

$I_2\,(g) + 3Cl_2\,(g) \rightarrow 2ICl_3\,(s)$ $\quad \Delta H^{\ominus} = -214$ kJ mol^{-1}

What is the correct value for ΔH_f^{θ} of iodine trichloride, $ICl_3\,(s)$?

A. –214 kJ mol^{-1}

B. –176 kJ mol^{-1}

C. –88 kJ mol^{-1}

D. +176 kJ mol^{-1}

(1 mark)

35 Using the following information:

$$CO\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H^{\theta} = -283\ kJ\ mol^{-1}$$

$$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow H_2O\ (l) \qquad \Delta H^{\theta} = -286\ kJ\ mol^{-1}$$

$$H_2O\ (g) \rightarrow H_2O\ (l) \qquad \Delta H^{\theta} = -44\ kJ\ mol^{-1}$$

What is the enthalpy change, ΔH^{θ}, for the following reaction?

$$CO_2\ (g) + H_2\ (g) \rightarrow CO\ (g) + H_2O\ (g)$$

A. -41 kJ mol^{-1}

B. +41 kJ mol^{-1}

C. -525 kJ mol^{-1}

D. +525 kJ mol^{-1}

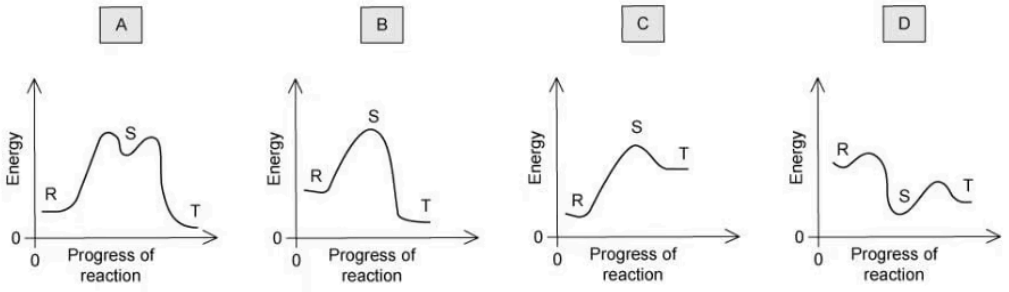
(1 mark)

36 Compound *R* into compound *T*, it was found that the reaction proceeded by way of compound *S*, which could be isolated. The following steps were involved.

$R \rightarrow S$; ΔH is positive

$S \rightarrow T$; ΔH is negative

Which reaction profile fits these data?



(1 mark)

37 The table below discusses three types of enthalpy change:

'+' means that this type of standard enthalpy change can only have positive values.

'–' means that this type of standard enthalpy change can only have negative values.

'+ / –' means that either positive or negative values are possible.

Which row is correct?

	Formation	Combustion	Neutralisation
A	+	+	+ / –
B	+ / –	+	+ / –
C	+ / –	–	–
D	–	–	+

(1 mark)

38 Iodine trichloride, ICl_3, is made by reacting iodine with chlorine.

$$I_2\,(s) + Cl_2\,(g) \rightarrow 2ICl\,(s)\,; \quad \Delta H^{o} = +14\ kJ\ mol^{-1}$$

$$ICl\,(s) + Cl_2\,(g) \rightarrow ICl_3\,(s)\,; \quad \Delta H^{o} = -88\ kJ\ mol^{-1}$$

By using the data above, what is the enthalpy change of the formation for solid iodine trichloride?

A. $-162\ kJ\ mol^{-1}$

B. $-81\ kJ\ mol^{-1}$

C. $-74\ kJ\ mol^{-1}$

D. $-60\ kJ\ mol^{-1}$

(1 mark)

39 In the gas phase, phosphorus pentachloride can be thermally decomposed into gaseous phosphorus trichloride and chlorine.

$$PCl_5 \rightarrow PCl_3 + Cl_2$$

The table below gives the relevant bond energies found in these compounds.

Bond	Bond energy / kJ mol^{-1}
P – Cl (in both chlorides)	328
Cl – Cl	241

What is the enthalpy change in the decomposition of the reaction?

A. – 415 kJ mol^{-1}

B. + 415 kJ mol^{-1}

C. + 95 kJ mol^{-1}

D. – 95 kJ mol^{-1}

(1 mark)

40 In a calorimetric experiment 2.50 g of a fuel is burnt in oxygen. 30 % of the energy released during the combustion is absorbed by 500 g of water, the temperature of which rises from 25 °C to 68 °C.

The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.

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

A. 25 284 J

B. 63 210 J

C. 119 827 J

D. 301 000 J

(1 mark)

41 Which equation correctly shows how the bond energy for the covalent bond Y-Z can be calculated by dividing ΔH by n?

A. $n\ YZ\ (g) \rightarrow n\ Y\ (g) + \frac{n}{2}\ Z_2\ (g)$

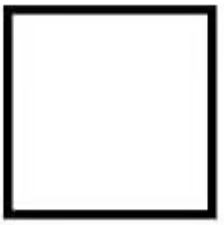
B. $Z\ (g) + Y\ Z_{n-1}\ (g) \rightarrow YZ_n\ (g)$

C. $2\ YZ_n\ (g) \rightarrow 2\ YZ_{n-1}\ (g) + Y_2\ (g)$

D. $YZ_n\ (g) \rightarrow Y\ (g) + nZ\ (g)$

(1 mark)

42 The diagram shows the skeletal formula of cyclobutane.



The enthalpy change of formation of cyclobutane is +75.1 kJ mol^{-1}, and the enthalpy change of atomisation of graphite is +712 kJ mol^{-1}.

The bond enthalpy of C–H is 390 kJ mol^{-1} and of H–H is 429 kJ mol^{-1}.

What is the average bond enthalpy of the C–C bond in cyclobutane, rounded to the nearest whole number?

A. 236 kJ mol^{-1}

B. 315 kJ mol^{-1}

C. 342 kJ mol^{-1}

D. 700 kJ mol^{-1}

(1 mark)

43 Some bond energy values are listed below.

Bond	Bond energy / kJ mol^{-1}
Br – Br	194
Cl – Cl	247
C – H	412
C – Cl	338

These bond energy values relate to the following four reactions:

W	$Br_2 \rightarrow 2Br$
X	$2Cl \rightarrow Cl_2$
Y	$CH_3 + Cl \rightarrow CH_3Cl$
Z	$CH_4 \rightarrow CH_3 + H$

What is the correct order of enthalpy changes of the above reactions from most negative to most positive?

A. Y → Z → W → X

B. Z → W → X → Y

C. Y → X → W → Z

D. X → Y → Z → W

(1 mark)

Answers

1 The correct answer is C because:

- The enthalpy of formation is defined as the enthalpy change when 1 **mole** of a compound is formed from its elements in their standard states and under standard conditions.
- X does show H_2O (l) being formed from H_2 (g) and O_2 (g) however, **3 moles** is produced hence x 3 enthalpy of formation of water.
- Using the Hess cycle, we are travelling in the **same direction** as the arrows in our alternative route (via the enthalpy of formation then combustion of C_2H_6 and water), therefore it is **positive.**

(1 mark)

2 (a) i) The definition of *enthalpy of formation is:*

- The enthalpy change when 1 mole of a compound is formed from its elements under standard conditions; [1 mark]

ii) The equation for the enthalpy formation of ammonia is:

- $½N_2$ (g) + $1½H_2$ (g) → NH_3 (g); [1 mark]

[Total: 2 marks]

- The three underlined terms in part (i) are essential to scoring the mark
- **Careful:** This equation often trips people up
  - You will be accustomed to seeing this equation with integers as the coefficients (1,3,2)
  - This would be incorrect for an equation that describes enthalpy of formation as it does not produce 1 mole of the product

(2 marks)

(b) What is meant by *standard* conditions in enthalpy measurements:

- The temperature must be 298 K
  AND
  The pressure must be 101 kPa; [1 mark]

[Total: 1 mark]

- You need to know the standard conditions
- By definition, the word standard is describing the conditions so the answer requires 298 K and 101 kPa

(1 mark)

(c) The equation(s) which can represent the enthalpy of neutralisation are:

- **B** and **C;** [1 mark]
- (Because) the enthalpy of neutralisation produces 1 mole of water from an acid and base; [1 mark]

[Total: 2 marks]

- The identity of the acid and base are irrelevant as long as only one mole of water is produced in the reaction
- Don't be put off by reactions with carbonates and CO_2 as a product – it is still a neutralisation reaction
- An acid–base reaction does not necessarily have to be neutralisation
  - For example, when ammonia reacts with hydrogen chloride, the reaction is acid–base, but not neutralisation:
    - $HCl + NH_3 \rightleftharpoons NH_4Cl$

(2 marks)

3 The correct answer is D because:

- The enthalpy of formation is the enthalpy change when one mole of a compound is formed **from its elements** in their standard states and under standard conditions
    - The only reactions that starts with the relevant elements in their standard states is **B** and **D**
- The enthalpy of combustion is defined as the enthalpy change when **one mole of a substance** reacts completely with oxygen under standard conditions
    - The only reaction that shows **one** mole of a substance reacting with oxygen is **D**

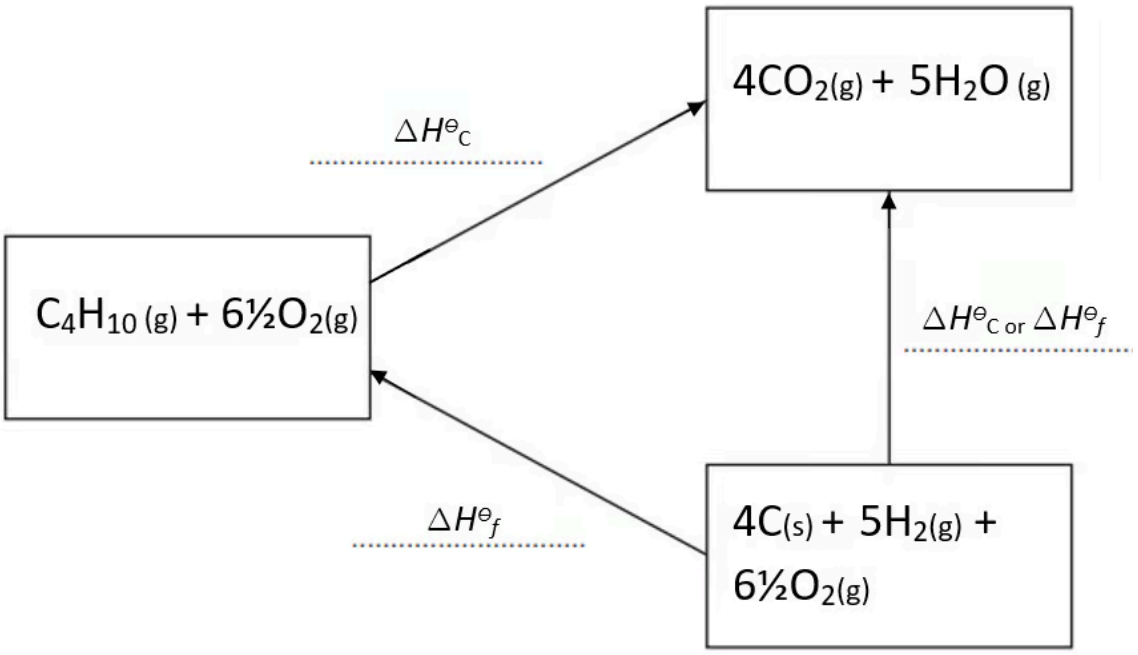
A is incorrect as this does represent formation because the products are not being formed from the elements

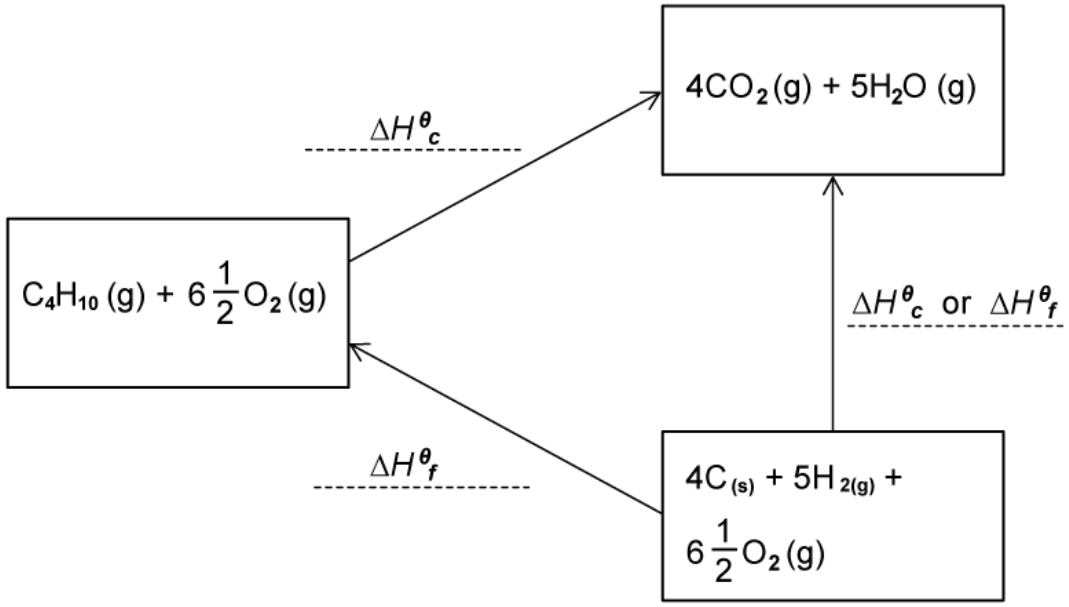
B is incorrect as this does represent the formation because one mole of a compound is formed, but it does not represent combustion as 2 moles of sodium are reacting

C is incorrect as this reaction does not represent formation or combustion, and many of the compounds are not in their standard states

(1 mark)

4 (a) The Hess' Law cycle for butane





- Correct formulae in the boxes; [1 mark]
- Correct balancing/ coefficient in the boxes; [1 mark]

- Correct enthalpy changes labelled; [1 mark]
- Both arrows shown in the correct direction; [1 mark]

[Total: 4 marks]

- You must include the state symbols for the substance as they are part of the standard enthalpy change definitions
- The formation of oxides is the same as the combustion of elements, so either term can be used

(4 marks)

(b) Calculating the enthalpy of combustion of butane, ΔH^{θ}_{c}:

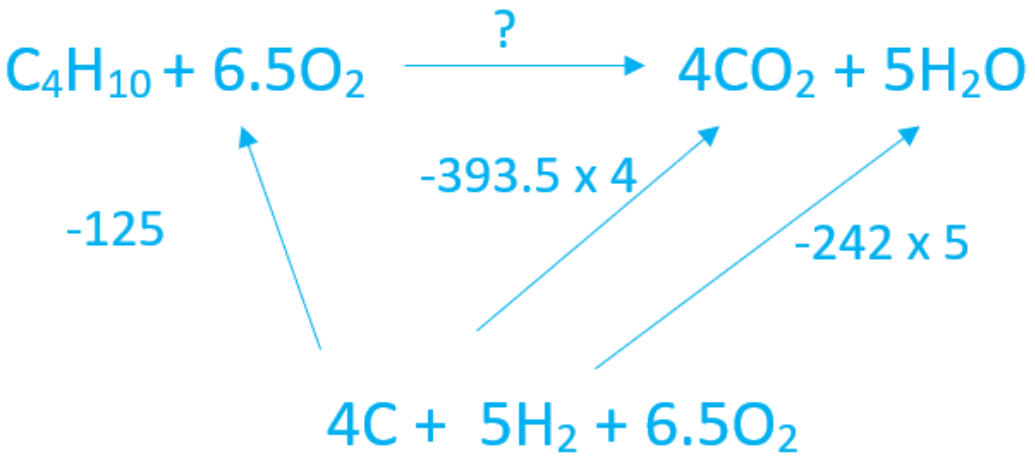
- $\Delta H^{\theta}_{c} = -\Delta H^{\theta}_{f}[C_4H_{10}(g)] + 4\Delta H^{\theta}_{f}[CO_2(g)] + 5\Delta H^{\theta}_{f}[H_2O(g)]$; [1 mark]
- $\Delta H^{\theta}_{c} = -(-125) + 4(-393.5) + 5(-242) = -2659\ kJmol^{-1}$; [1 mark]

[Total: 2 marks]

- The enthalpy of combustion equation is:

$C_4H_{10} + 6.5O_2 \rightarrow 4CO_2 + 5H_20$

- It can help to draw a Hess cycle with the values in:



- To find the unknown go around the cycle anticlockwise, add if the arrow(s) follows the direction, subtract if it goes against the direction
- Be careful with your signs– it is a good idea to separate the sign from the mathematical operation by using brackets around the terms

(1 mark)

(c) The enthalpy of combustion of butane is:

Bonds Broken / $kJmol^{-1}$	Bonds Formed / $kJmol^{-1}$
3 x C-C = 3 x 348 = 1044 10 x C-H = 10 x 410 = 4100 6.5 x O=O = 6.5 x 495 = 3217.5	8 x C=O = 8 x 743 = 5944 10 x O-H = 10 x 463 = 4630
Total = +8361.5 [1 mark]	Total = -10574 [1 mark]

- ΔH_r = +8361.5-10574 = -2212.5 kJ ; [1 mark]

[Total: 3 marks]

Sometimes it is easier to set out bond energy calculations like a balance sheet

(1 mark)

(d) Of the two values of enthalpy of combustion from parts **(b)** and **(c)** which is likely to be more accurate:

- The value from part (b) / -2659 kJ; [1 mark]

- The calculation in (b) uses data from experimentally measured values;
  AND
  The bond energies in part (c) are an average taken from a range of values in a range of similar compounds; [1 mark]

[Total: 2 marks]

(2 marks)

5 (a) The equation for the formation of propane from propyne is:

$CHCCH_3\ (g) + 2H_2\ (g) \rightarrow CH_3CH_2CH_3\ (g)$

- Correct species
  AND
  Correct balancing; [1 mark]

- Correct state symbols; [1 mark]

[Total: 2 marks]

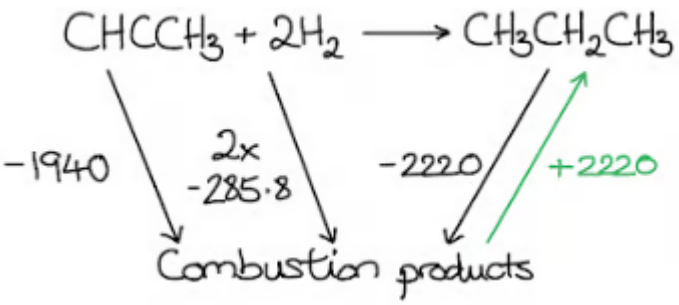
- You do not specifically need to know the alkyne homologous series

- You should be able to apply your knowledge and understanding to use the information given in an exam question

(2 marks)

(b) To calculate the enthalpy change for the formation of propane from propyne:

Hess cycle method:



- Correct cycle; [1 mark]

- $\Delta H^{\ominus}_r$ = -1940 + (2 x -285.8) + 2220; [1 mark]

- $\Delta H^{\ominus}_r$ = -291.6 (kJ mol^{-1}); [1 mark]

Equation method:

- Enthalpy change of reaction, $\Delta_r H^\ominus = \sum \Delta H^\ominus_c$ reactants - $\sum \Delta H^\ominus_c$ products; [1 mark]

- $\Delta H^\ominus_r$ = [(-1940) + (2 x -285.8)] - [(-2220)]; [1 mark]

- $\Delta H^\ominus_r$ = -291.6 (kJ mol^{-1}); [1 mark]

[Total: 3 marks]

- You can either use a Hess cycle or the equation for these calculations

(3 marks)

(c) A calculated bond enthalpy value is often different to the data book value because:

- (The data book value) is an average value which is derived from a number of different compounds; [1 mark]

[Total: 1 mark]

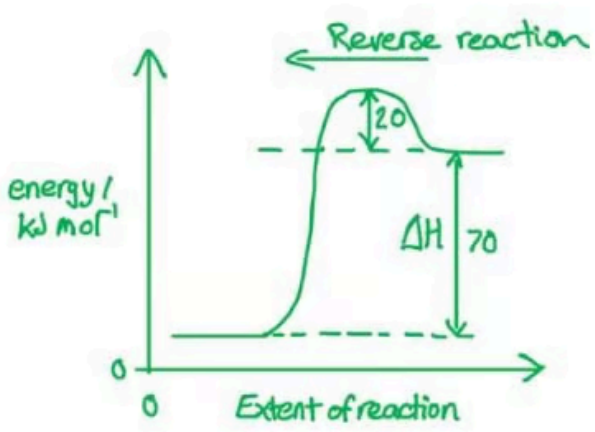
- This is a very common question and the answer is always the same, but students often forget to revise this and lose the mark

(1 mark)

6 The correct answer is C because:

- The activation energy can be measured on a reaction pathway by measuring the energy difference from the energy level to the top of the 'hump'.

- **If you look at the reversible reaction** The reaction is going from right to left here. Measuring the energy level on the right to the top of the hump which is +20 kJ mol^{-1}.



A is incorrect as this is the activation energy plus the enthalpy change (the difference between the two energy levels) see diagram.

B is incorrect as to measure the activation energy, you must measure the difference from the energy level to the top of the hump. Going in the forward direction you would, therefore, do 70 kJ mol^{-1} + 20 kJ mol^{-1} = +90 kJ mol^{-1}.

D is incorrect as the forward reaction is endothermic and hence must have a positive sign.

(1 mark)

7 (a) To calculate the enthalpy of reaction, ΔH_r, for the formation of $CHCl_2F(g)$:

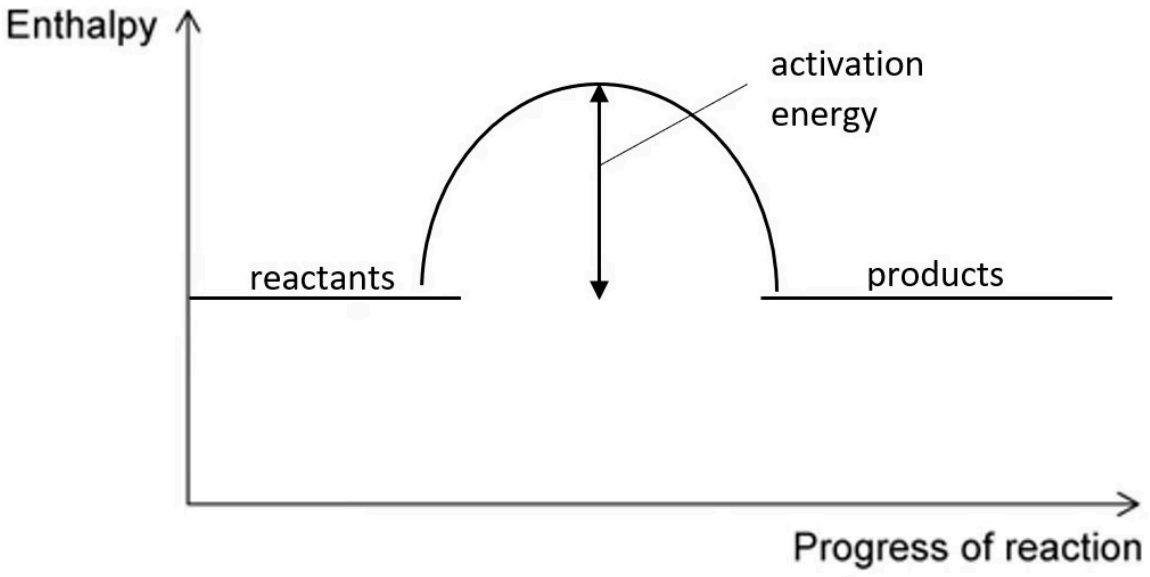
- $\Delta H_r = (-284.1) + (-92.3) - (-103.2) - (-273.3)$; [1 mark]
- 0.1 (kJ mol^{-1}); [1 mark]

[Total: 2 marks]

- To solve this problem you need to use the formula:
  - $\Delta H_r = \sum \Delta H_{f\ products} - \sum \Delta H_{f\ reactants}$
- Be careful not to mix up the signs of the energy terms with the mathematical operations
  - **Tip:** Use brackets around the terms
- If there are coefficients >1 in the balanced equation, don't forget to multiply the energy terms by the coefficient

(2 marks)

(b) A reaction pathway diagram for part (a) with activation energy labelled is:



- Correct diagram with the correction positions of reactants and products ; [1 mark]
- Activation energy correctly labelled; [1 mark]

[Total: 2 marks]

- You can use the formulae of the reactants and products instead of the labels 'reactants' and 'products'
- The enthalpy of reaction is close to zero, so the two horizontal lines should be at an almost identical height
- If the lines are far apart then you will not score the mark as you have not applied correctly the information calculated in part (a).

(2 marks)

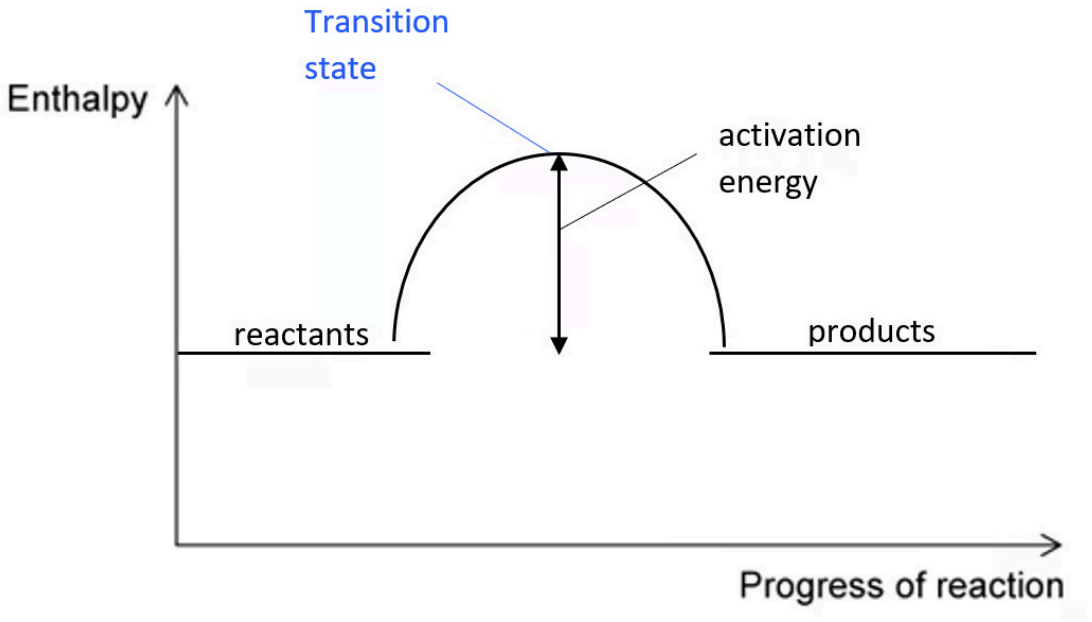
(c) Explain what is meant by the transition state:

- The transition state is a stage during the reaction at which chemical bonds are partially broken and formed; [1 mark]

[Total: 1 mark]

- The transition state is very unstable – it cannot be isolated and is higher in energy than the reactants and products

- It is located at the very top of the curve on the reaction pathway diagram.



(1 mark)

8 The correct answer is B because:

- To calculate the enthalpy change of reaction using enthalpy of formation data, it is always products - reactants:
  - $\Delta H_r = \Sigma \Delta H_{f\ products} - \Sigma \Delta H_{f\ reactants}$

A is incorrect as the formation enthalpies are being added together

C is incorrect as the products and reactants are the wrong way around

D is incorrect as the products and reactants are the wrong way around **AND** the formation enthalpies are being added together

(1 mark)

9 (a) To calculate the enthalpy change of combustion, ΔH_c, of 2-methylpropan-2-ol:

- Mass change = 0.57 g
  AND
  Temperature change = 45.8 ^{o}C; [1 mark]

- q = 14358 J; [1 mark]

- M_r of 2-methylpropan-2-ol = 74.0; [1 mark]

- Moles of 2-methylpropan-2-ol = 0.0077027 mol
  OR
  7.7027 x 10^{-3} mol; [1 mark]

- ΔH = -18641000 / -1.8641 x 10^{6} J mol^{-1}
  OR
  ΔH = -1864.1 / -1900 kJ mol^{-1}; [1 mark]

[Total: 5 marks]

- Table 3.1 enables you to calculate the mass of 2-methylpropan-2-ol used and the temperature change of the reaction
  - **Careful:** The mass of 2-methylpropan-2-ol is **not** the mass to be used in the $q = mc\Delta T$ equation

- Using $q = mc\Delta T$:
  - q is to be calculated
  - m = 75 g, as this is the mass of the water being heated
  - c = 4.18 J g^{-1} K^{-1}
  - ΔT = 65.3 - 19.5 = 45.8 ^{o}C

- So, the $q = mc\Delta T$ calculation is:
  - q = 75 x 4.18 x 45.8 = 14358 J

- M_r of 2-methylpropan-2-ol = (4 x 12.0) + (9 x 1.0) + (16.0 + 1.0) = 74.0

- Moles of 2-methylpropan-2-ol = $\frac{\text{mass}}{M_r} = \frac{0.57}{74.0}$ = 0.0077027 mol

- So, the $\Delta H = \frac{q}{n}$ calculation is:
  - $\Delta H = \frac{14358}{0.0077027}$

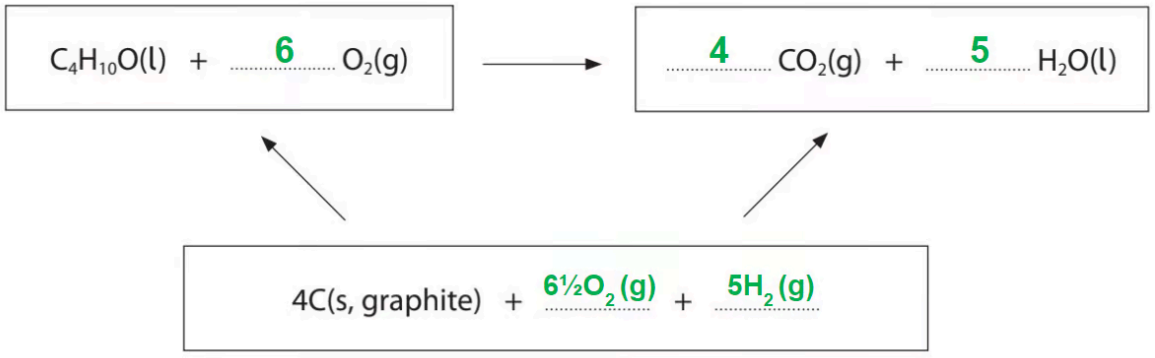
- ΔH = −18641000 / −1.8641 x 10^6 J mol^{-1} **OR** ΔH = −1864.1 / −1900 kJ mol^{-1}
    - **Remember:** This is an exothermic reaction so the value should be negative, as it is releasing heat energy

(5 marks)

(b) i) Oxygen does not have a value for ΔH_f^θ because:

- Oxygen is an element
  AND
  In its standard state; [1 mark]

ii) The complete Hess cycle is:



- Balanced reactants **AND** products; [1 mark]
- Balanced elements with state symbols; [1 mark]

iii) The standard enthalpy change of combustion of 2-methylpropan-2-ol is:

- ΔH = − ΔH_f^θ reactants + ΔH_f^θ products
  OR
  ΔH = − (−359) + (−394 x 4) + (−286 x 5); [1 mark]
- = -2647 kJ mol^{-1}; [1 mark]

[Total: 5 marks]

- The standard enthalpy change of formation of an element in its standard state is always 0 as no change occurs
- The combustion products of an alcohol will be carbon dioxide and water
- **Remember:** Enthalpy of combustion is about the complete combustion of **one mole** of a chemical
    - Ensure that you balance the equations carefully as correctly balancing oxygen for this reaction will give a fraction, i.e. 6½ or $^{13}/_{2}$
- When constructing an equation to calculate the enthalpy change of combustion, use the arrows in the cycle to help you:
    - Two arrows start from $4C + 6½O_2 + 5H_2$ and two arrows point to $4CO_2 + 5H_2O$
    - This means there are two routes to get from $4C + 6½O_2 + 5H_2$ to $4CO_2 + 5H_2O$
    - Direct route:
        - The enthalpy change of this route = $4 \times \Delta H_f(CO_2) + 5 \times \Delta H_f(H_2O) = 4 \times (-394) + 5 \times (-286) = -3006$
    - Indirect route, via $C_4H_{10}O + 6O_2$:
        - The enthalpy change of this route = $\Delta H_f(C_4H_{10}O) + \Delta H_c(C_4H_{10}O) = -359 + \Delta H_c(C_4H_{10}O)$
    - The enthalpy change of both routes will be the same:
        - $-359 + \Delta H_c(C_4H_{10}O) = -3006$
        - $\Delta H_c(C_4H_{10}O) = -3006 + 359 =$ **-2647 kJ mol^{-1}**
- It is very easy to get the calculation the wrong way around and end up with a positive number, so don't rush and write down your workings meticulously so as not to make an error in rearranging
- You will lose a mark if you don't get the sign correct

(5 marks)

(c) **Two** reasons for the value for ΔH_c^{θ} obtained in part **(a)** being much less exothermic than ΔH_c^{θ} calculated in **(b)(iii)** are:

Any **two** from:

- Heat / energy is lost (in the enthalpy experiment) **OR** The heat capacity of the beaker has not been included; [1 mark]
- Incomplete combustion (of alcohol); [1 mark]
- Some alcohol / fuel evaporates; [1 mark]
- No stirrer so temperature is not uniform; [1 mark]

[Total: 2 marks]

- These are typical reasons why the experimental values are less exothermic than the calculated values which you should know

(2 marks)

10 **The correct answer is D because:**

- The reactants have less energy than the products, so energy is taken into the reaction and therefore endothermic.
- The enthalpy change of the backward reaction (going from products to reactants is equal to -30 kJ mol^{-1}.
- The activation energy for the forward reaction is the total energy needed to go from reactants to products
    - In this case it is +90 kJ mol^{-1}.

(1 mark)

11 (a) The definition of average bond enthalpy is:

- The energy needed to break one mole of bonds; [1 mark]
- In a gaseous molecule averaged over similar compounds; [1 mark]

[Total: 2 marks]

- The average bond enthalpy term is the average amount of energy needed to break a specific type of bond, measured over a wide variety of different gaseous molecules
- It is essentially the average of all the bond dissociation enthalpies for a specific type of bond

(2 marks)

(b) The formula for calculating the standard enthalpy change of reaction, ΔH_r, using bond energies is:

- ΔH_r = enthalpy change for bonds broken + enthalpy change for bonds formed
  OR
  $\Delta H_r = \sum \Delta H$ Bonds broken + $\sum \Delta H$ Bonds made; [1 mark]

[Total: 1 mark]

- Bond energies are used to find the ΔH_r of a reaction when this cannot be done experimentally

(1 mark)

(c) To calculate the enthalpy change for the hydration of ethene:

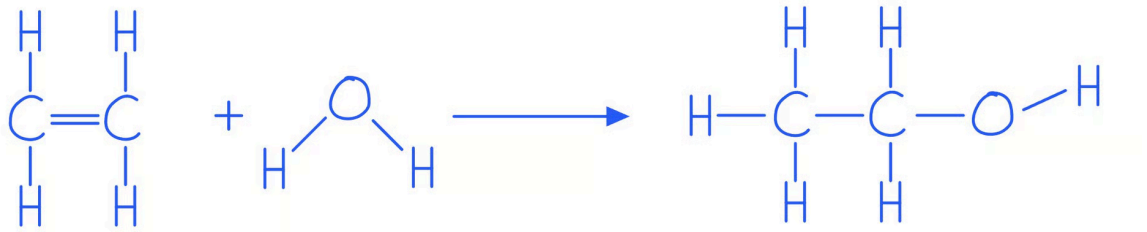
- ∑bonds broken (endothermic)
  E(C=C) = 614
  E(C-H) x 4 = 1656
  E(O-H) x 2 = 926
  Total = 3196 kJ; [1 mark]
- ∑bonds made (exothermic)
  (C-C) = 346
  E (C-H) x 5 = 2070

(C-O) = 358
(O-H) = 463
Total = -3237 kJ; [1 mark]

- ΔH_r = 3196 + (-3237) = -41 kJ mol^{-1}; [1 mark]

[Total: 3 marks]

- It can be really helpful to draw out the structures of the reactants and products when calculating enthalpy change using bond enthalpies as it is easier to see the number of each type of bond



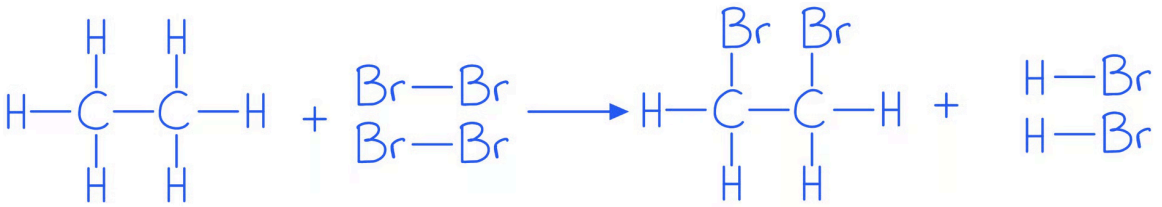
(3 marks)

(d) To calculate the enthalpy change for the reaction of ethane and bromine:

- $CH_3CH_3 + 2Br_2 \rightarrow CH_2BrCH_2Br + 2HBr$; [1 mark]
- ∑bonds broken (endothermic) E(C-C) = 346 E(C-H) x 6 = 2484 E(Br-Br) x 2 = 386 Total = 3216 kJ; [1 mark]
- ∑bonds formed (exothermic) E(C-C) = 346 E(C-H) x 4 = 1656 E(C-Br) x 2 = 570 E(H-Br) x 2 = 732
- Total = -3304 kJ; [1 mark]
- ΔH_r = -88 kJ mol^{-1}; [1 mark]

[Total: 4 marks]

- Again, drawing out the correct structures will help here given that there are two molecules of bromine and hydrogen bromine



(4 marks)

12 The correct answer is B because:

- To calculate ΔH_r we have to:
  - Travel along the ΔH_1 arrow in the same direction
    - This means that we are adding ΔH_1
  - Then travel along the ΔH_2 arrow in the opposite direction
    - This means that we have to reverse the arrow and, therefore, take away ΔH_2
  - This leads to a final equation of $\Delta H_r = \Delta H_1 - \Delta H_2$

A is incorrect as ΔH_1 and ΔH_2 are being added together

C is incorrect as ΔH_1 and ΔH_2 are the wrong way around

D is incorrect as the only time that we would be multiplying for enthalpy calculations is to work with the stoichiometry of the reaction

(1 mark)

13 The correct answer is B because:

- Hess's law states that the overall enthalpy change will be the same regardless of the route taken.
- **Known quantities from the question**
  - $\Delta H_1 = (3 \times -394) + (4 \times -286) = -2326$
  - $\Delta H_2 = -1816$
- **Calculate ΔH_f**
  - $\Delta H_f = \Delta H_1 - \Delta H_2$
  - $\Delta H_f = -2326 - (-1816)$
  - $\Delta H_f = -510$ kJ

(1 mark)

14 The correct answer is D because:

- **The energy transferred as heat is given by the equation; $q = mc\Delta T$**
  - Where q is heat transferred (J), m is the mass of water (g), c is the specific heat capacity (J g^{-1} $^{\circ}C^{-1}$), ΔT is the temperature change ($^{\circ}C$).
- **Known quantities:**
  - Mass of water; 150 – 50 = 100 g
  - Temperature change; 54 – 21 = 33 $^{\circ}C$
  - Specific heat capacity of water; 4.18 J g^{-1} $^{\circ}C^{-1}$
- **Substitution into the equation:**
  - $q = mc\Delta T = 100 \times 4.18 \times 33 = 13794$ J

(1 mark)

15 (a) The *standard enthalpy of combustion,* $\Delta H^{\ominus}_{c}$ is:

- (The) enthalpy change / energy change when one mole of a substance; [1 mark]
- Is completely burnt / reacts completely in oxygen; [1 mark]
- With all reactants and products in their standard states
  OR
  With all reactants and products under standard conditions **OR**
  With all reactants and products in their normal states at 298K and 100 kPa; [1 mark]

[Total: 3 marks]

- Be careful not to confuse products and reactants in thermodynamic definitions
- Enthalpy of formation is defined according to the production of one mole of product, whereas enthalpy of combustion is based on one mole of a reactant / substance

(3 marks)

(b) The equation for the complete combustion of propanol is:

$2CH_3CH_2CH_2OH\ (l) + 9O_2\ (g) \rightarrow 6CO_2\ (g) + 8H_2O\ (l)$

OR

$CH_3CH_2CH_2OH\ (l) + 4.5O_2\ (g) \rightarrow 3CO_2\ (g) + 4H_2O\ (l)$

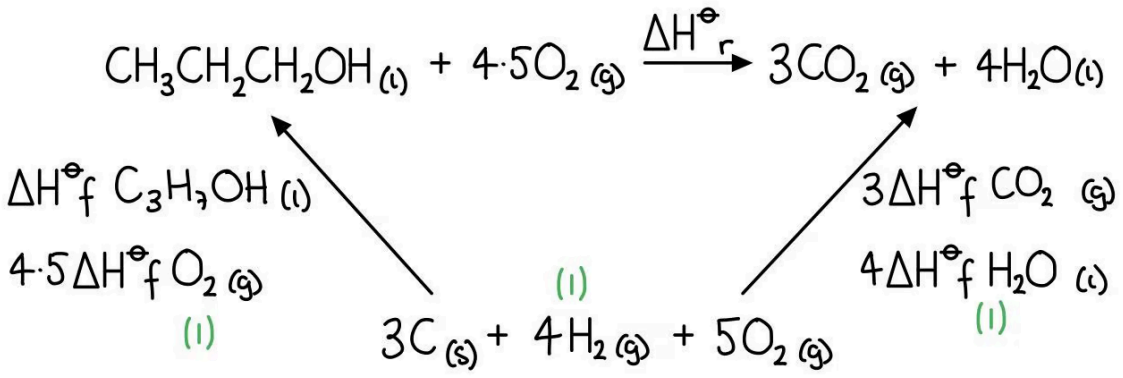
- Correct formula **AND** balancing of reactants; [1 mark]
- Correct formula **AND** balancing of products; [1 mark]

[Total: 2 marks]

- A common mistake is to leave out the oxygen for the combustion equation, remember complete combustion requires oxygen
- If you are not using fractions to balance the equation, then you must double the number of moles of the other chemicals.
- Double-check your balancing once you have written your equation
    - Many chemical calculations work from a balanced symbol equation, therefore, if your equation is incorrect then any subsequent calculations will contain errors

(2 marks)

(c) The Hess's Law cycle for the complete combustion of propanol is:



- Correct elements in standard states
  AND
  Correct balancing; [1 mark]

- Correct direction of arrow from elements in standard states to the reactants
  AND
  Correct data; [1 mark]

- Correct direction of arrow from elements in standard states to the products
  AND
  Correct data; [1 mark]

[Total: 3 marks]

- **Remember:** Arrows must be in the right direction, so the reactants and products are formed from the elements in their standard states so the arrows should point upwards in this case

- To calculate a value for ΔH_r^{θ}

  - $\Delta H^{\ominus}_r = \Sigma \Delta H^{\ominus}_f$ products – $\Sigma \Delta H^{\ominus}_f$ reactants

- Then apply:

  - If you follow the direction of the arrow, you add the quantity

  - If you go against the arrow, you subtract the quantity

(3 marks)

(d) To calculate the enthalpy change for the reaction, $\Delta H_r^{\ominus}$:

- $\Delta H^{\ominus}_r = \Sigma \Delta H^{\ominus}_f$ products - $\Sigma \Delta H^{\ominus}_f$ reactants; [1 mark]
- $\Delta H^{\ominus}_r$ = [(3 x -393.5) + (4 x -285.8)] - (-303); [1 mark]
- $\Delta H^{\ominus}_r$ = (-2323.7) - (-303.0) = -2020.7 kJ mol^{-1}; [1 mark]

[Total: 3 marks]

- In the question, $\Delta H^{\ominus}_f$, data has been given so the equation used to calculate the enthalpy change is:
  - $\Delta H^{\ominus}_r = \Sigma \Delta H^{\ominus}_f$ products – $\Sigma \Delta H^{\ominus}_f$ reactants
- Take note of the data that you have been given as that will dictate the calculation that will be used

(3 marks)

16 (a) i) The enthalpy change for this reaction is:

- Sum of the bonds broken:

E(C-H) = 410

E(Cl-Cl)= 242

Total = +652 kJ; [1 mark]

- Sum of the bonds formed:

E(C-Cl) = 340

E(H-Cl)= 431

Total = -771 kJ; [1 mark]

- Enthalpy change= 652-771 = -119 kJ; [1 mark]

ii) The conditions needed for this reaction to occur are:

- UV light / sunlight/ high temperature; [1 mark]

[Total: 4 marks]

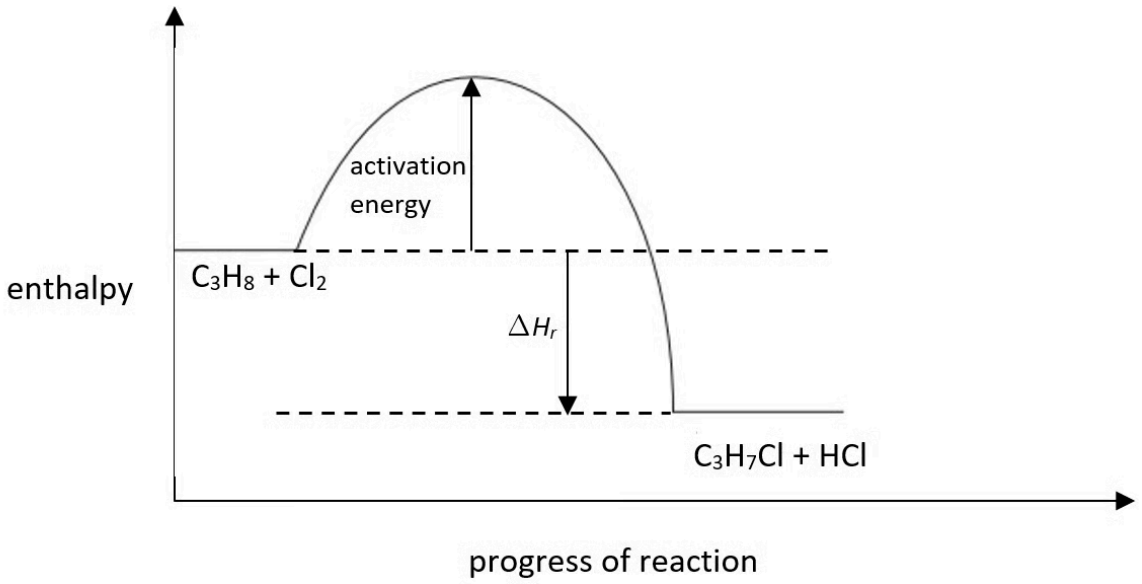
- Always have a methodical approach to bond energy calculations and break it down into steps
- Be careful to check if more than one of the same bond is broken or formed
- Remember that bonds broken have a positive enthalpy change because energy is needed to break the bonds
- Find the enthalpy change of reaction using:

Enthalpy change = Σ bonds broken + Σ bonds made

(4 marks)

(b) A labelled reaction pathway (energy level) diagram for the reaction in part (a) including

the activation energy:



- Correct location of reactants and products; [1 mark]
- Upward arrow labelled as activation energy; [1 mark]
- Downward arrow labelled as ΔH_r / enthalpy change; [1 mark]

[Total: 3 marks]

- There is no need to add the values for the enthalpy changes in part a), just the relative starting and finishing positions of the reactants and products

(3 marks)

(c) The difference in enthalpy change for the two reactions is:

- There is no / little difference in enthalpy change for the reactions; [1 mark]
- Because the same number and type of bonds are broken and formed in both reactions; [1 mark]

[Total: 2 marks]

- Bond energies are average values obtained from a range of similar compounds
- There may be very slight differences in bond energy between compounds as the presence of other atoms and bonds could have a small influence on bond strength
- For this reason you could argue that there may be a small difference between the halogenation of propane and the halogenation of ethane

(2 marks)

(d) Whether the sum of the bonds broken would be larger or smaller with bromine instead of chlorine:

- Smaller; [1 mark]
- Bromine is a larger atom than chlorine;
  AND
  The Br-Br bond is (longer and) weaker than Cl-Cl; [1 mark]

[Total: 2 marks]

- You should know the trends in atomic radius for the halogens which allows you to predict the difference in bond length between halogen molecules
- Shorter covalent bonds are stronger than longer bonds

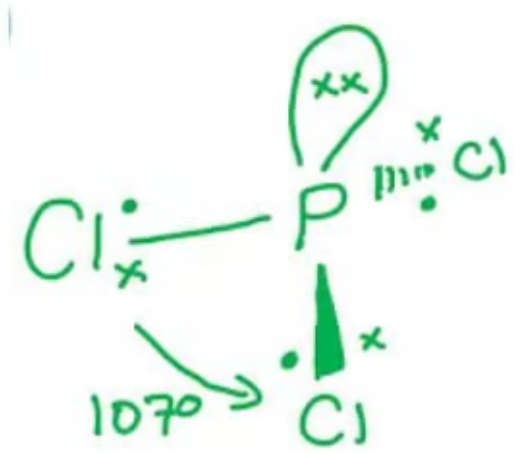
(2 marks)

17 The correct answer is A because:

- As more heat energy is applied to PCl_5 (g) the extent of dissociation has increased. This means the reaction must be endothermic.

- The valence shell of the phosphorous atom in PCl_3 (g) contains 8 electrons arranged in 3 bonded pairs with the 3 chlorine atoms and 1 lone pair from the phosphorous atom (phosphorous is in group 5).

- There is equal repulsion between bonding pairs but greater repulsion from the lone pair (2.5^o) to the bonded pairs. Hence a pyramidal shape.



B is incorrect as the reaction is endothermic not exothermic.

C is incorrect as a trigonal shape is only possible when there are 3 bonded pairs and no lone pairs.

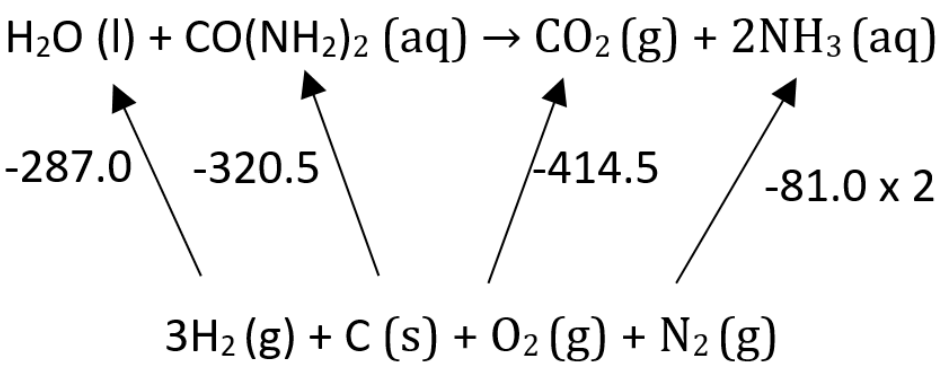
Dis incorrect as the reaction is endothermic, not exothermic. A trigonal shape is only possible when there are 3 bonded pairs and no lone pairs.

(1 mark)

18 (a) i) The equation for the hydrolysis of urea by water is:

- H_2O (l) + $CO(NH_2)_2$ (aq) → CO_2 (g) + $2NH_3$ (aq); [1 mark]

ii) A simple energy cycle for the hydrolysis of urea is:



- Arrows pointing upwards; [1 mark]
- Correct substances, coefficients and enthalpy change labels; [1 mark]

[Total: 3 marks]

- Since these are enthalpies of formation the arrows must point from the elements in their standard states to the compounds
- Pay attention to multiples of the enthalpies changes or you may lose the mark
- Make sure you have the correct state symbols that match the standard state of the element as the cycle uses enthalpies of formation which by definition must start from the elements in their standard states

(3 marks)

(b) Calculating the standard enthalpy change for the hydrolysis of urea:

- ΔH_r = 607.5 - 576.5 = 31 kJ mol^{-1}; [1 mark]

[Total: 1 mark]

- **Remember:** As you go around the cycle using Hess's Law, add the terms if you follow the arrow and subtract the terms if you go against the arrow
    - Using brackets around the terms helps to avoid careless calculation errors
- The workings are: ΔH_r = -(-287.0) - (-320.5) + (-414.5) + (-81.0 x 2)

(1 mark)

(c) Calculating the enthalpy of combustion of urea:

$\Delta H_c = \sum \Delta H_f \, products - \sum \Delta H_f \, reactants$

- ΔH_c = (-829) + (-1148) - (-641) = -1336 kJ mol^{-1}; [1 mark]

[Total: 1 mark]

- The trick here is realising the method to solve this question is the same as the previous question, i.e. using the enthalpies of formation to solve the problem
- The relevant enthalpies of formation are:
  - Oxygen and nitrogen are both 0 kJ mol^{-1} as they are elements
  - ΔH_f for $CO(NH_2)_2$ is −320.5 kJ mol−1
    - **BUT** there are 2 moles used in the balanced equation, so 2 x −320.5 = −641 kJ mol^{-1}
  - ΔH_f for CO_2 is −414.5 kJ mol−1
    - **BUT** there are 2 moles used in the balanced equation, so 2 x −414.5 = −829 kJ mol^{-1}
  - ΔH_f for H2O is −287 kJ mol−1
    - **BUT** there are 4 moles used in the balanced equation, so 4 x −287 = −1148 kJ mol^{-1}
- Instead of using a cycle, you can use the relationship:
  - $\Delta Hr = \sum \Delta H_f \, products - \sum \Delta H_f \, reactants$ **OR** $\Delta H_c = \sum \Delta H_f \, products - \sum \Delta H_f \, reactants$
    - Since the reaction is a combustion reaction
- ΔHr = (−414.5 x 2) + (−287.0 x 4) − (−320.5 x 2)
- ΔHr = (−829) − (−1148) − (−641)
- ΔHr = −1336 kJ mol^{-1}

(1 mark)

19 (a) i) An equation to represent the complete combustion of butan-1-ol:

C_4H_9OH (l) + $6O_2$ (g) → $4CO_2$ (g) + $5H_2O$ (l)

- Correct products; [1 mark]
- Correct balancing of reactants and products; [1 mark]

ii) The purpose of the copper spiral and small electric heater in the apparatus:

- The copper spiral ensures complete heat exchange with the surrounding water / conducts heat from the waste gases (combustion products) to the water; [1 mark]
- To ignite (and vaporise) the butan-1-ol; [1 mark]

[Total: 4 marks]

- Be careful when you balance combustion reactions of alcohols as it is very easy to miss the oxygen in the alcohol itself
- Although you may not have seen this apparatus before you should have performed or seen an enthalpy of combustion experiment and be able to use your knowledge and apply it in unfamiliar circumstances
- You must have the idea that the alcohol is burning– only stating that it is vaporised by the heater would not be enough to score the mark as vaporising alone doesn't include combustion

(4 marks)

(b) The enthalpy of change, q, for the reaction is:

- $q = mc\Delta T$;

OR

q = 875 x 4.18 x 2.5; [1 mark]

- q = 9,143.75 J = 9,100 J = 9.1 kJ to 2 sig figs; [1 mark]

[Total: 2 marks]

- Since you are given the unit in the answer space make sure your convert to kilojoules or you may be penalised since this is a relatively simple calculation to perform
- Failing to round to two significant figures would also cost you a mark

(2 marks)

(c) The enthalpy of combustion of butan-1-ol is:

- $\Delta H = \frac{q}{n}$; [1 mark]
- $n = \frac{2.20\ \text{g}}{74.12\ \text{gmol}^{-1}} = 0.02968\ \text{mol}$; [1 mark]
- $\Delta H = \frac{9143.75\ \text{J}}{0.02968\ \text{mol}}$ = 308,077.83 = -308 kJ mol^{-1} ; [1 mark]

[Total: 3 marks]

- Don't round off a calculation until the end
  - If you used 9,100 J from part **(b)** you would get -307 kJ mol^{-1} as your final answer and would lose a mark
- A correct final answer without working would get full marks
- Failure to include the minus sign will cost you a mark

(3 marks)

20 The correct answer is C because:

- The **standard enthalpy of formation** (ΔH_f^θ) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
- For elements in the standard state, the enthalpy of formation is zero; therefore, O_2 is zero.
- **Equation:**
  - $\boldsymbol{\Delta H_f = \Sigma \Delta H_f \text{ (products)} - \Sigma \Delta H_f \text{ (reactants)}}$
- **Known quantities:**
  - Reactants:
    - $4NH_3 = (4 \times -51.3) = -205.2$
    - $5O_2 = (5 \times 0) = 0$
    - Total = -205.2 + 0 = -205.2
  - Products:
    - $4NO = (4 \times 92.2) = 368.8$
    - $6H_2O = (6 \times -239.6) = -1437.6$
    - Total = 368.8 - 1437.6 = 1068.8
- **Substitution into the equation:**
  - ΔH_f = -1068.8 + 205.2 = -863.6 kJ

For this type of question, make sure that you show how you calculated the totals for the reactants and products. This is because you may still gain marks even if you get the wrong answer

(1 mark)

21 (a) i) The completed table to show the oxidation number of nitrogen in each compound:

Compound	NO	NO_2
oxidation number of N	+2; [1 mark]	+4; [1 mark]

ii) The equations for the formation of NO_2:

- $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g)$; **OR** $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$; [1 mark]
- $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$; **OR** $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$; [1 mark]

[Total: 3 marks]

- Remember that to form an uncharged compound the oxidation numbers have to balance to zero
- For NO
  - Oxidation number of oxygen atom = −2
  - NO has no charge so the total oxidation number is 0
  - Oxidation number of N = +2
- For NO_2
  - Oxidation number of each oxygen atom = −2
  - For two oxygen atoms = 2 x (−2) = −4
  - NO_2 has no charge so the total oxidation number is 0
  - Oxidation number of N = +4
- State symbols are not necessary to get the marks for the equations

(1 mark)

(b) The bond enthalpy, in kJ mol^{-1}, of the N=O bond:

Bonds Broken / kJ mol^{-1}	Bonds Formed / kJ mol^{-1}
2 x O=O = 2 x 495 = 990 1 x N≡N = 1 x 941 = 941	4 x N=O = 4 x ? = 4(N=O)
Total = +1931 [1 mark]	Total = -4(N=O)

- ΔH_r = +1931-4(N=O) = +497 kJ mol^{-1} ; [1 mark]
- Bond enthalpy of the N=O bond = $\frac{1434}{4} = 358.5 \text{ kJ mol}^{-1}$; [1 mark]

[Total: 3 marks]

- The reaction equation is

$N_2 (g) + 2O_2 (g) \rightarrow 2NO_2$

- From this we can see that 1 N≡N and two O=O bonds are broken and 4 N=O bonds are formed
- The complete workings for the last step are here

ΔH_r = +1931–4(N=O) = +497 kJ mol^{-1}

–4(N=O) = +497–1931

4(N=O) = 1434

Bond enthalpy of the N=O bond = $\frac{1434}{4} = 358.5 \text{ kJ mol}^{-1}$

- Bond energies are usually expressed without a sign as they can be positive or negative

(3 marks)

(c) Suggest a reason for the difference in boiling points between carbon dioxide and nitrogen dioxide:

- The van der Waals/ intermolecular forces are (much) stronger in NO_2 / weaker in CO_2; [1 mark]

ONE from

- NO_2 has permanent dipole-dipole attractions (between the molecules) but CO_2 does not ; [1 mark]
- NO_2 is polar but CO_2 is not; [1 mark]
- In CO_2 the molecule is symmetrical and the dipoles cancel out / there is no net dipole; [1 mark]

[Total: 2 marks]

- This question shows how much shape influences intermolecular forces and boiling point
- The C=O bond is actually more polar than the N=O bond as there is a greater difference in electronegativity between the atoms, but the presence of a lone pair on N makes it bent linear, whereas CO_2 is linear
- The linear shape causes the dipoles to cancel out in CO_2, so there is no net dipole and hence it only has weaker dispersion(London) forces

(2 marks)

(d) An equation for the enthalpy of formation of nitrogen monoxide, NO:

- $\frac{1}{2}N_2\,(g) + \frac{1}{2}O_2\,(g) \rightarrow NO\,(g)$; [1 mark]

[Total: 1 mark]

- Be careful– an easy trap to fall into!
- Enthalpy of formation is defined as the formation of *1 mole* of a compound from its elements in their standard states and weaker candidate would write it as:

$N_2\,(g) + O_2\,(g) \rightarrow 2NO\,(g)$

- State symbols must be included to score the mark as they are part of the definition of enthalpy of formation

(1 mark)

22 The correct answer is C because:

- This is the exact definition of the average bond enthalpy and is applied specifically to the HI (g) bond, which is what the question asked.
- Key points from the definition to remember are:
  - The molecule must be in **the gas phase.**
  - Has to apply to **one mole.**
  - **Always endothermic** as it's the 'energy required'.

A is incorrect as this is the wrong terminology. Whilst you can picture in your mind the energy is stored, it's important to remember that the definition is very specific. '**Energy required to break**..' are the words to use. It is also an incomplete definition. Fails to touch on all points listed.

B is incorrect as the definition is for 1 mole of bonds rather than just a unitless number. You must learn the definition word for word.

D is incorrect as the average bond enthalpy relates to the **energy required not released.**

(1 mark)

23 The correct answer is B because:

- The **standard enthalpy of formation** ($\Delta H_f^{\ominus}$) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
- Elements in the standard state have an $\Delta H_f^{\ominus}$ of zero; therefore, C and Ti are zero.
- **Equation:**
  - **ΔH_f = $\Sigma\Delta H_f$ (products) - $\Sigma\Delta H_f$ (reactants)**
- **Known quantities:**
  - Reactants:
    - TiO_2 = −890
    - 2C = (2 × 0) = 0
    - Total = −890 + 0 = −890
  - Products:
    - 2CO = (2 × −105) = −210
    - Ti = 0
    - Total = −210 + 0 = −210
- **Substitution into the equation:**
  - ΔH_f = −210 − (−890) = 680 kJ

For this type of question, make sure that you show how you calculated the totals for the reactants and products. This is because you may still gain marks even if you get the wrong answer

Using brackets around negative numbers is a good way to ensure that you don't lose a negative symbol by mistake, e.g. (−890)

(1 mark)

24 (a) The mass of propane required is:

- $q = mc\Delta T$ calculation 150 x 4.18 x 79 = 49533 J mol^{-1} **OR** 49.533 kJ mol^{-1} ; [1 mark]
- Moles of propane = $\frac{49.533}{2219.2}$ = 0.0223....; [1 mark]
- Mass of propane = 0.023 x 44.0 = 0.98 g; [1 mark]

[Total: 3 marks]

- This is another way of asking a $q = mc\Delta T$
- You have all the information to calculate q
  - Remember to use the mass of water that has been boiled
- The more challenging part is to then calculate the number of moles of propane used
  - q is calculated in J, so you must convert it to kJ before dividing it by (–)2219.2 which is the enthalpy of combustion of propane
  - Enthalpy of combustion is the energy released when 1 mole of the compound or element burns completely in a plentiful supply of oxygen
  - But we don't have one mole of propane, we have 0.0223.... moles
  - Therefore to find out the mass in grams, you must multiply by the M_r of propane, 44.0

(3 marks)

(b) More propane is required than calculated in part (a) because:

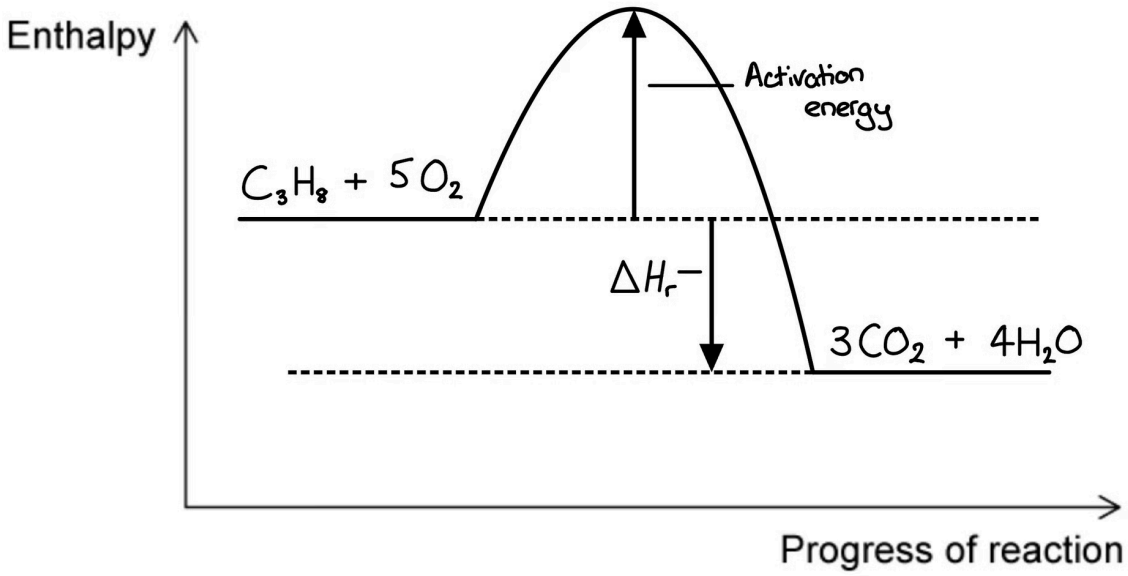
- Some of the heat / energy is lost to the surroundings; [1 mark]

[Total: 1 mark]

- This is a practical example so you should consider the experiments that you have done previously to help you answer this question
    - Heat loss to the surroundings is a common error in calorimetry experiments

(1 mark)

(c) The completed diagram is:



- Correct location of and stoichiometry of reactants and products; [1 mark]
- Upward arrow labelled as activation energy; [1 mark]
- Downward arrow labelled as ΔH_r / enthalpy change / ΔH_c; [1 mark]

[Total: 3 marks]

- The combustion of propane is an exothermic reaction, so the reactants must be higher in energy than the products
- Your reactants and products should be written as a balanced equation

(3 marks)

25 The correct answer is D because:

- **The energy transferred as heat is given by the equation; $q = mc\Delta T$**
  - Where q is heat transferred (J), m is the volume of water (cm^3), c is the specific heat capacity (J cm^{-3} K^{-1}), ΔT is the temperature change (˚C).
- **Known quantities:**
  - Volume of solution; 30 + 30 = 60 cm^3
  - Temperature change; 0.5 ˚C
    - Change in temperature does not need to be converted into Kelvin
  - Specific heat capacity; 4.18 J cm^{-3} K^{-1}
- **Substitution into the equation:**
  - $q = mc\Delta T = 60 \times 4.18 \times 0.5 = 125.4$ J
- **Divide this answer by the moles** as the question asked for the molar enthalpy change.
  - Note: When given the concentrations of two solutions, it usually would be the limiting reagent that is used to calculate the molar enthalpy.
- Here, both solutions react in a 1:1 molar ratio and are of the same concentration so we can just use the moles of one of them.
- **Inputting into the formula**
  - $n = \text{conc} \times \frac{\text{vol}}{1000} = 0.025 \times \frac{30}{1000} = 0.00075$ mol

- **Divide the energy released by the moles to get the energy per mol**
    - $\frac{q}{n} = \frac{125.4}{0.00075}$ = -167 200 J mol^{-1}
- **Convert into kJ and add a minus sign as the reaction is exothermic (as the temperature rose)**
    - **-167.2 kJ mol^{-1}**

IMPORTANT: As this is an acid/ base reaction, the molar enthalpy change can also be expressed as ΔH_{neut}. The molar enthalpy change will be calculated per mole of water formed. The numerical answer will be the same.

A is incorrect as 30 cm^3 was used to calculate *q* and not the total volume of the solutions. Remember that the heat energy is spread out across the total volume of both solutions. This is what gives rise to the total temperature change.

B is incorrect as 30 cm^3 was used to calculate *q* **AND** the value of 4.2 was used for the specific heat capacity of the solution - you should be using the specification value which is 4.18.

C is incorrect as the reaction is exothermic as the temperature rises. Not endothermic.

(1 mark)

26 The correct answer is B because:

- The **standard enthalpy of formation** ($\Delta H_f^{\ominus}$) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
- Elements in the standard state have an $\Delta H_f^{\ominus}$ of zero; therefore, O_2 is zero.
- **Equation:**
  - **ΔH_f = $\Sigma\Delta H_f$ (products) - $\Sigma\Delta H_f$ (reactants)**
- **Known quantities:**
  - Reactants:
    - C_2H_5OH = −266
    - $3O_2$ = (3 × 0) = 0
    - Total = −266 + 0 = −266 kJ
  - Products:
    - $2CO_2$ = (2 × −382) = −764
    - $3H_2O$ = (3 × −275) = −825
    - Total = (−764) + (−825) = −1589 kJ
- **Substitution into the equation:**
  - ΔH_f = (−1589) + (−266) = −1323 kJ

For this type of question, make sure that you show how you calculated the totals for the reactants and products. This is because you may still gain marks even if you get the wrong answer

Using brackets around negative numbers is a good way to ensure that you don't lose a negative symbol by mistake, e.g. (−890)

(1 mark)

27 The correct answer is D because:

- **Equation: $\Delta H_r = \Sigma \Delta H_f$ (products) - $\Sigma \Delta H_f$ (reactants)**
- **Known quantities:**
  - Reactants: ΔH_f^{θ} $2SO_2$ = (–297 × 2) kJ mol^{-1}
- Products: ΔH_f^{θ} $2SO_3$ = (–395 × 2) kJ mol^{-1}
- **Substitution into the equation:**
  - ΔH_r = -790 + 594 = -196 kJ

(1 mark)

28 (a) The energy change, in J, when 12.04 g of ammonium chloride is dissolved in 125.0 cm^3 of water:

- $n(NH_4Cl) = \frac{12.04 \text{ g}}{53.5 \text{ g mol}^{-1}}$ = 0.225 (mol); [1]
- q = 0.225 mol x 15 100 J mol^{-1} = 3 397.5 (J); [1]

[Total: 2 marks]

- You need to work 'backwards' from the ΔH_{sol} given in the question to work out the energy change when 12.04 g or 0.225 mol of ammonium chloride is used
- This will give you a value in kJ mol^{-1} for q, but the question was asked for q in J

(2 marks)

(b) To determine the value in temperature, in °C, when the lithium chloride is dissolved:

- $\Delta T = \frac{q}{mc} = \frac{3397.5 \text{ J mol}^{-1}}{125.0 \text{ cm}^3 \text{ x } 4.18 \text{ J K}^{-1} \text{ mol}^{-1}}$ = 6.5 (°C); [1]

[Total 1 mark]

- Rearrange the $q = mc\Delta T$ equation to give you the temperature change

(1 mark)

(c) To determine the final temperature, in °C, of the solution during the reaction:

- 19.5 °C - 6.5 °C = 13.0 °C; [1]

[Total: 1 mark]

- **The positive sign on the enthalpy of solution value tells you that the reaction is endothermic so the temperature will decrease**
    - **In other words, you must subtract the enthalpy change from the initial temperature**

(1 mark)

29 The correct answer is B because:

- The **standard enthalpy change of neutralisation** (ΔH_{neut}^{θ}) is the enthalpy change when one mole of water is formed by the reaction of an acid with alkali under standard conditions.
- **The energy transferred as heat is given by the equation; $q = mc\Delta T$**
    - Where q is heat transferred (J), m is the volume of water (cm^3), c is the specific heat capacity (J cm^{-3} K^{-1}), ΔT is the temperature change (˚C).
- **Known quantities from the question**
    - The total volume of the solution = 150 cm^3
    - Specific heat capacity of the solution = 4.18 J g^{-1} K^{-1}
    - Temperature change = 14 ˚C
- **Substitution into the equation:**
    - $q = mc\Delta T$ = 150 × 4.18 × 14 = 8778 J
- **Calculating moles:**

- **Note:** When given the concentrations of two solutions normally It would be the limiting reagent that is used to calculate the molar enthalpy
- Here, both solutions react in a 1:1 molar ratio and are of the same concentration so we can just use the moles of one of them
- n = conc x $\frac{\text{vol}}{1000}$ = 3 x 0.075 = 0.225

- **Calculating molar enthalpy change of neutralisation:**
- $\Delta H = -\frac{q}{n} = -\frac{8778}{0.225}$ = 39 013.3 J mol^{-1}

(1 mark)

30 (a) To determine the enthalpy of hydrogenation of propene:

- Hydrogenation equation: $CH_3CH=CH_2 + H_2 \rightarrow C_3H_8$; [1 mark]

Calculation option 1:

- ΔH_r = (C-C) + (C=C) + 6(C-H) +(H-H) -2(C-C) -8(C-H); [1 mark]
- ΔH_r = (347) + (619) + 6(413) + 435 - 2(347) -8(413) = -119 kJ mol^{-1}; [1 mark]

OR

Calculation option 2:

- ΔH_r = (C=C) + (H-H) - (C-C) - 2(C-H); [1 mark]
- ΔH_r = (619) + (435) - (347) -2(413) = -119 kJ mol^{-1}; [1 mark]

[Total: 3 marks]

- For this question, you need to know the structure of propene and that hydrogenation involves a reaction with hydrogen
- Always look for shortcuts to save you time
  - E.g. if the same bond is broken and formed, then there is no need to include it in the calculation

(3 marks)

(b) Suggest, with a reason, whether the polymerisation of propene is exothermic or endothermic:

- One C=C bond is broken and two C-C bonds are formed; **OR**
  One C=C bond is is weaker / 619 $kJmol^{-1}$ than two C-C bonds / 2 x 347 = 694 $kJmol^{-1}$; [1 mark]

- There is a net output of energy / the reaction is exothermic; [1 mark]

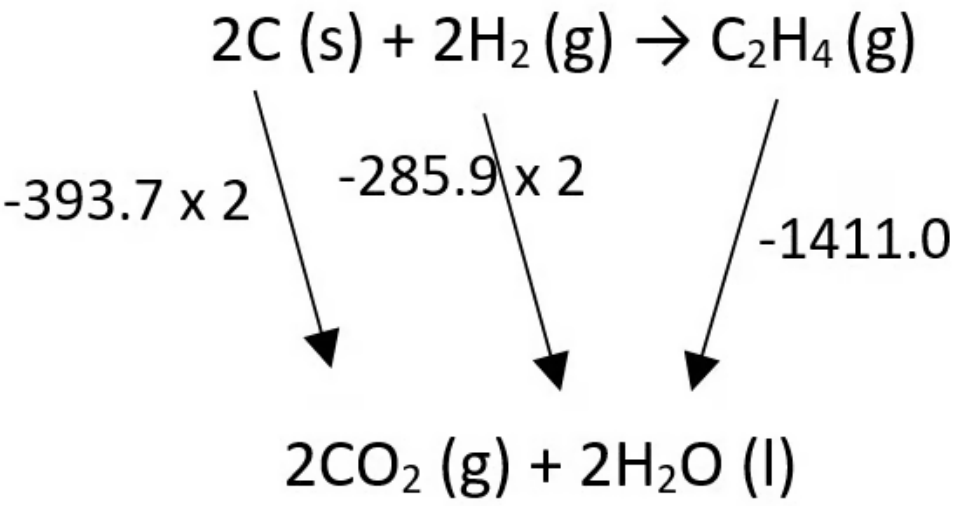
[Total: 2 marks]

- The polymerisation of propene would be:
  - $nCH_3CH=CH_2 \rightarrow n\ -[-CH(CH_3)CH_2-]-$
- For *n* molecules of propene:
  - *n* C=C bonds are broken
  - 2*n* C–C bonds are formed
- The net energy change is the equivalent of breaking one C=C bond and forming two C–C bonds, which is exothermic

(2 marks)

(c) To calculate the standard enthalpy change of formation of ethene:

- Correctly labelled cycle



OR

ΔH^{θ}_{f} = 2(–393.7) + 2(–285.9) – (–1411.0); [1 mark]

- ΔH^{θ}_{f} = + 51.8 kJ mol^{-1}; [1 mark]

[Total: 2 marks]

- The arrows must point downwards in the cycle and have the correct coefficients
- The wrong sign for –1411 would lose the first mark

(2 marks)

31 The correct answer is D because:

- The **standard enthalpy of reaction** (ΔH^{θ}), is the enthalpy change when the reactants shown in the stoichiometric equation react to give products under standard conditions
    - **Remember:** The reactants and products must be in their standard state.
- There are 6 N-F bonds in the product

- Therefore, $\frac{-1776}{6}$ = –296 kJ

(1 mark)

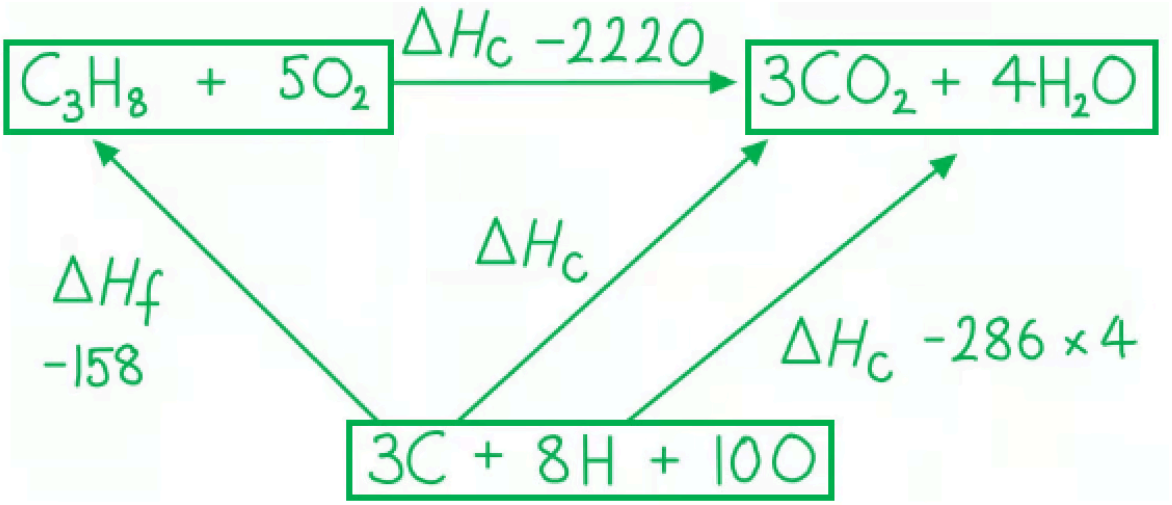
32 The correct answer is B because:

- The **standard enthalpy change of combustion** ($\Delta H_c^{\ominus}$) is the enthalpy change when one mole of a substance is burnt in excess oxygen under standard conditions. The reactants and products must be in their standard states.
- Enthalpy changes of combustion are always exothermic.
- **Equation:**
  - ΔH = Σ(bonds broken) - Σ(bonds formed)
- **Bonds broken:**
  - This is endothermic, and therefore all figures are positive.
  - 2 × C-H = 390 × 2 = 780
  - C≡C = 870
  - 2.5 O=O = 2.5 × 489 = 1222.5
  - Sum of bonds broken = 2872.5
- **Bonds formed:**
  - This is exothermic, and therefore all figures are negative.
  - 2 × O-H = -510 × 2 = -1020
  - 4 C=O = -790 × 4 = -3160
  - Sum of bonds broken = −4180
- ΔH = Σ(bonds broken) - Σ(bonds formed)
  - ΔH = 2872.5 - 4180 = -1307.5 kJ

(1 mark)

33 The correct answer is B because:

- The standard enthalpy change of combustion (ΔH_c^{θ}) is the enthalpy change when one mole of a substance is burnt in excess oxygen under standard conditions. The reactants and products must be in their standard states
- Using the balanced chemical equation for the combustion of propane, the following Hess cycle can be established



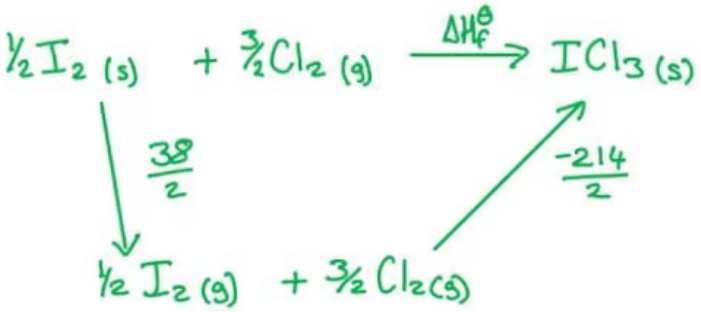
- The enthalpy of combustion of propane is given as -2220 kJ mol^{-1}
- The enthalpy of combustion of hydrogen is given as -286 kJ mol^{-1}
  - **Careful:** There are 4 moles of hydrogen so this means that it is 4 x -286 kJ mol^{-1}
- The question asks for the enthalpy of combustion of carbon that has been used
  - **Careful:** There are 3 moles of carbon so this means that it is 3 x ΔH_c (carbon)
- The question states that the final answer achieved by the student for the enthalpy of formation of propane is -158 kJ mol^{-1}
- Assuming that the enthalpy of combustion of carbon is Y, then Hess' law can be applied:
  - –2220 = –(–158) + 3Y + (4 x –286)

- This rearranges and evaluates to:
    - 3Y = –2220 – 158 – (4 x –286)
    - Y = $\frac{-2220-158-(4\times-286)}{3}$
    - Y = $\frac{-1234}{3}$
    - Y = -411.3 kJ mol^{-1}

(1 mark)

34 The correct answer is C because:

- The **standard enthalpy of formation** (ΔH_f^θ) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
- Hess's law states that the overall enthalpy change will be the same regardless of the route taken as long as the reactants and products remain the same.
- **Draw a Hess cycle showing the known quantities from the question**
    - Note that the equations given in the question are double the moles used in the equation below, so both enthalpy changes are halved.

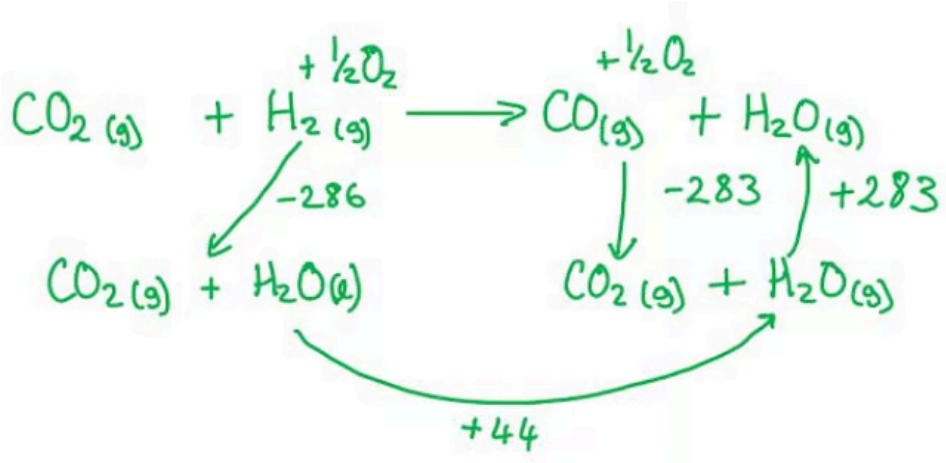


- **Calculate the total using the new pathway**
    - $\Delta H_r = \frac{38}{2} + \frac{-214}{2}$ = -88 kJ mol^{-1}

(1 mark)

35 The correct answer is B because:

- The **standard enthalpy of formation** (ΔH_f^{θ}) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
- Hess's law states that the overall enthalpy change will be the same regardless of the route taken as long as the reactants and products remain the same.
- **Draw a Hess cycle showing the known quantities from the question.**

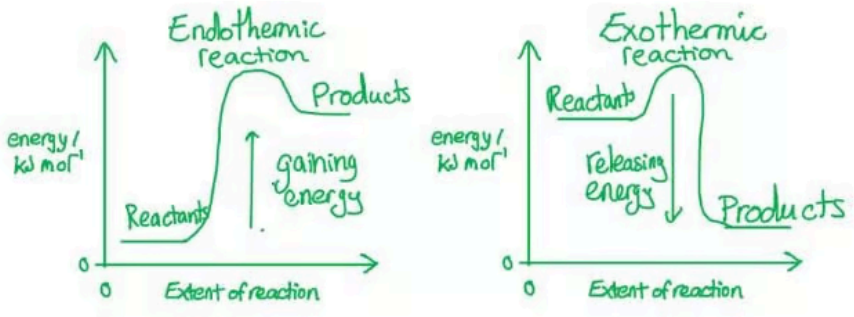


- **Add the enthalpy changes for the new pathway.**
  - ΔH= -286 + 44 + 283 = 41kJ mol^{-1}

(1 mark)

36 The correct answer is A because:

- The reaction R → S is endothermic as ΔH is positive, so it is gaining energy
- The reaction S → T is exothermic as ΔH is negative, so it is releasing energy
- The energy profile diagrams for exothermic and endothermic reactions are shown below:



- When the two reactions are combined the energy increases then decreases as seen in option **A**

(1 mark)

37 The correct answer is C because:

- The **standard enthalpy of formation** (ΔH_f^{θ}) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.
    - Enthalpy changes of formation can be exothermic or endothermic (+ / –).
- The **standard enthalpy change of combustion** (ΔH_c^{θ}) is the enthalpy change when one mole of a substance is burnt in excess oxygen under standard conditions. The reactants and products must be in their standard states.
    - Enthalpy changes of combustion are always exothermic (–).
- The **standard enthalpy change of neutralisation** (ΔH_{neut}^{θ}) is the enthalpy change when one mole of water is formed by the reaction of an acid with alkali under standard conditions.
    - Enthalpy changes of neutralisation are always exothermic (–).

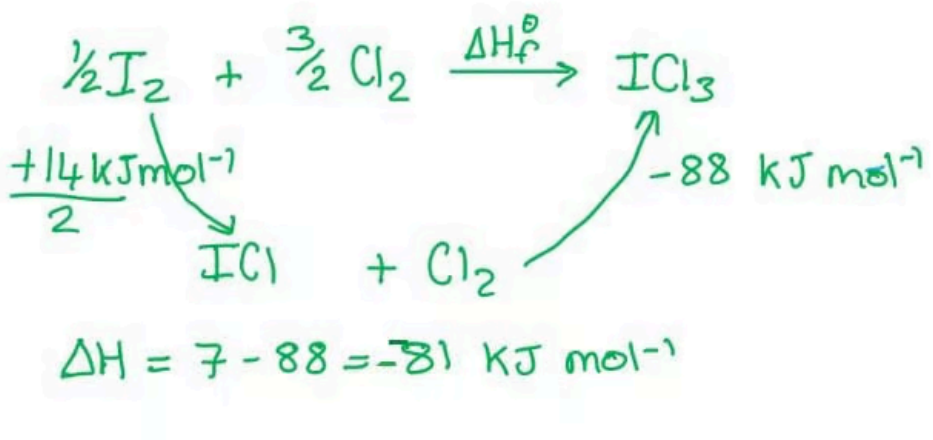
(1 mark)

38 The correct answer is B because:

- The enthalpy change of formation ($\Delta H_f^{\ominus}$) is the enthalpy change when one mole of a compound is formed from its elements under standard conditions. The reactants and products must be in their standard states.

- To answer this question construct a Hess cycle and use the information given in the question to show the energy of the alternate routes.

- Remember that the definition is for one mole of the compound, so the equation needs to be balanced to produce one mole of ICl_3.



(1 mark)

39 The correct answer is B because:

- **Equation: ΔH = Σ bonds broken - Σ bonds formed**
- **Bonds broken:**
  - This is endothermic, and therefore all figures are positive
  - 5 × P - Cl = 328 × 5 = 1640
- **Bonds formed:**
  - This is exothermic, and therefore all figures are negative
  - 3 × P - Cl = -328 × 3 = -984
  - Cl - Cl = -241
  - Sum of bonds broken = −1225
- **ΔH = Σ bonds broken - Σ bonds formed**
  - ΔH = 1640 - 1225 = 415 kJ

(1 mark)

40 The correct answer is C because:

- **The energy transferred as heat is given by the equation; $q = mc\Delta T$**
  - Where q is heat transferred (J), m is the mass of water (g), c is the specific heat capacity (J g^{-1} $^{\circ}C^{-1}$), ΔT is the temperature change ($^{\circ}C$)
- **Known quantities:**
  - Mass of water; 500 g
  - Temperature change; 68 – 25 = 43 $^{\circ}C$
  - Specific heat capacity of water; 4.18 J g^{-1} $^{\circ}C^{-1}$
- **Substitution into the equation:**

- q = mcΔT = 500 × 4.18 × 43 = 89870 J
- We can see from the question that this represents 30% of the energy released by the combustion of the fuel.
- **The total energy released:**
  - 89870 × $\frac{100}{30}$ = 299566.7 J
- The question asks for the total energy released per gram of fuel burnt, 2.5 g of fuel was used in the experiment
- **Energy per gram of fuel:**
  - x = $\frac{299566.7}{2.5}$ = 119 827 J

(1 mark)

41 The correct answer is D because:

- We can see from the question that it is asking for an equation that matches the expression.
  - Y-Z = $\frac{\Delta H}{n}$
- Where *n* is the number of Y−Z bonds broken.
- **Equation: Δ*H* = ∑ bonds broken - ∑ bonds formed**
- To make this simpler, we can use an example to help visualise the question.
  - $CH_4 \rightarrow C + 4H$
- In this case, there are 4 bonds broken and no bonds formed.
- In CH_4 there are 4 C−H bonds, which would be *n* in the equation, from this there are 4 hydrogen atoms, also *n* in the equation.
- Substitution of Y and Z into this equation would give:
  - $YZ_n \rightarrow Y + nZ$

- This matches option **D**

A is incorrect as *n* in the reactant side would represent the number of moles of the ZY molecule present; this will not tell us the number of bonds in the molecule.

B is incorrect as the molecule YZ is on both sides of the equation and will not enable us to calculate the Z-Y bond energy.

C is incorrect as there are a different number of atoms on either side of the equation so not a balanced equation.

(1 mark)

42 The correct answer is C because:

- **Write out the balanced symbol equation:**
  - $4C\ (s) + 4H_2\ (g) \rightarrow C_4H_8\ (g)$
- **Known quantities:**
  - C (graphite) → C (g) ΔH^o = +712 kJ/mol
  - ΔH_f^θ of cyclobutane = 75.1 kJ/mol
  - ΔH^θ = H-H = 429 kJ/mol
  - ΔH^θ = C-H = 390 kJ/mol
- **Equation: ΔH = Σ bonds broken - Σ bonds formed**
  - 75.1 = Σ bonds broken - Σ bonds formed
  - 75.1= [4(712) + 4(429)] - [8(390) + 4(x)]
- The equation can then be re-arranged for x
  - 75.1= 4564 - [3120 + 4x]
  - 75.1 = 1444 + 4x
  - -1368.9 = 4x
  - x = 342 kJ mol^{-1}

(1 mark)

43 The correct answer is C because:

- **Equation: ΔH = Σ bonds broken - Σ bonds formed**
- **Calculate the ΔH for each reaction:**
  - ΔH $Br_2 \rightarrow 2Br = 194 - 0 = 194$
  - ΔH $2Cl \rightarrow Cl_2 = 0 - 247 = -247$
  - ΔH $CH_3 + Cl \rightarrow CH_3Cl = 0 - 338 = -338$
  - ΔH $CH_4 \rightarrow CH_3 + H = 412 - 0 = 412$
- The correct order from most negative to positive is
  - $Y \rightarrow X \rightarrow W \rightarrow Z$

(1 mark)

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1. Enthalpy changes that are difficult to measure directly can ...
2. Standard enthalpy changes include enthalpy of formation, ...
3. Which equation below can represent both an enthalpy cha...
4. Complete the following Hess' Law energy cycle relating but...
5. Propane gas is commonly used as a fuel for outdoor cookin...
6. The reaction pathway for a reversible reaction is shown bel...
7. Freon 21 is a hydrochlorofluorocarbon that was used as a p...
8. Which is the correct equation to calculate the enthalpy of r...
9. Enthalpy changes of combustion can be determined using c...
10. The reaction pathway for a reversible reaction is shown b...
11. Define the term average bond enthalpy.
12. Hess's Law can be used to calculate the enthalpy change f...
13. Propanol has the molecular formula C3H8O. Use the follo...
14. A student carried out an experiment to determine the ent...
15. Define the term standard enthalpy of combustion, $\Delta H\ominus c$.
16. Propane reacts with chlorine to form chloropropane. C3H...
17. PCl5 (g) dissociates as follows. PCl5 (g) → PCl3 (g) + Cl2 (g) ...
18. Urea, CO(NH2)2, is a naturally occurring substance which ...
19. The apparatus shown in Fig. 2.1 can be used to determine ...
20. The first stage in the industrial production of nitric acid fro...
21. Two common oxides of nitrogen are nitrogen monoxide, ...
22. Hydrogen atoms bond covalently to Iodine atoms to form ...
23. Titanium occurs naturally as the mineral rutile, TiO2. One ...
24. Propane gas is used widely as a fuel which can be be used...
25. A student mixed 30.0 cm3 of 0.0250 mol dm-3 potassium ...
26. The combustion of ethanol (C2H5OH) is increasingly being...
27. The equations below show the formation of sulfur oxides f...
28. The enthalpy change of solution for ammonium chloride c...
29. An experiment was carried out to determine the approxi...
30. Determine the enthalpy of hydrogenation of propene usin...
31. The standard enthalpy change $\Delta H\theta$ for the below reaction ...
32. The complete combustion of ethyne, C2H2, is shown in th...