Q:1 [M/J 2002 (2)]

2 Ethanol, C_2H_5OH, is a most important industrial chemical and is used as a solvent, a fuel and an intermediate in large scale organic synthesis.

Ethanol is prepared industrially by the reaction of ethene and steam in the presence of a catalyst.

$$C_2H_4(g) + H_2O(g) \longrightarrow C_2H_5OH(g)$$

The standard enthalpy change of the reaction can be determined by using the standard enthalpy changes of combustion, $\Delta H_c^\ominus$, at 298 K.

	$\Delta H_c^\ominus$ / kJ mol^{-1}
$C_2H_4(g)$	–1411
$C_2H_5OH(l)$	–1367

(a) Calculate the standard enthalpy change for the following reaction.

$$C_2H_4(g) + H_2O(l) \longrightarrow C_2H_5OH(l)$$

[2]

(b) (i) Define the term *standard enthalpy change of combustion*.

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....

....

(ii) Explain why the state symbols for water and ethanol given in the equation in **(a)** have been changed from those quoted in the industrial process.

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(iii) Write the equation for the complete combustion of ethanol.

....

[4]

(c) Ethanol is miscible with water because of hydrogen bonding between molecules of ethanol and water. Draw a diagram, including dipoles, to show the hydrogen bonding between a molecule of ethanol and a molecule of water.

[2]

Q:2 [MJ 2006 (2: d)]

(d) The equation for the complete combustion of ethyne is given below.
Use appropriate bond energy data from the *Data Booklet* to calculate a value for the enthalpy change of combustion of ethyne.

$$C_2H_2(g) + \tfrac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(g)$$

[3]

(e) The value for the standard enthalpy change of combustion of ethyne is $-1300\,kJ\,mol^{-1}$.

(i) Define the term *standard enthalpy change of combustion*.

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(ii) Explain why your answer to **(d)** does not have the same value as the standard enthalpy change of combustion.

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....
[3]

Q:3 [M/J 2010 - 23 (1)]

1 Hydrazine, N_2H_4, can be used as a rocket fuel and is stored as a liquid. It reacts exothermically with oxygen to give only gaseous products.

The enthalpy change of a reaction such as that between hydrazine and oxygen may be calculated by using standard enthalpy changes of formation.

(a) Define the term *standard enthalpy change of formation*, $\Delta H_f^{\ominus}$.

....

....

.... [3]

(b) Hydrazine reacts with oxygen according to the following equation.

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

(i) Use the data in the table to calculate the standard enthalpy change of this reaction.

compound	$\Delta H_f^{\ominus}$/kJ mol^{-1}
$N_2H_4(l)$	50.6
$H_2O(g)$	–241.8

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

(ii) Although the above reaction is highly exothermic, hydrazine does not burn spontaneously in oxygen.
Suggest a reason for this.

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....

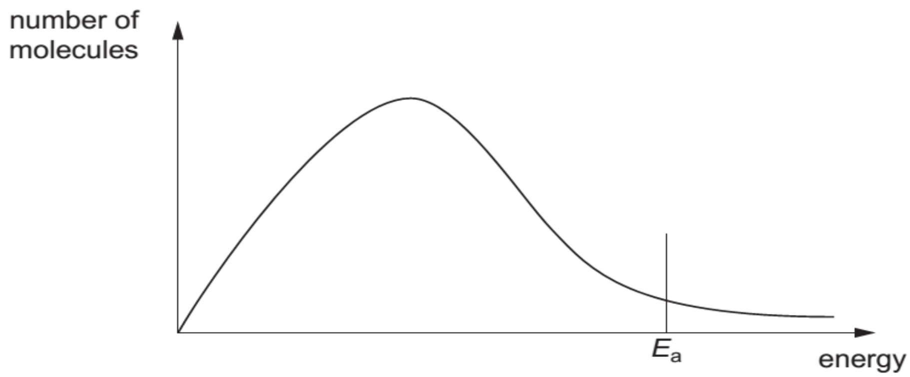
(iii) Suggest why using hydrazine as a rocket fuel could be regarded as being 'environmentally friendly'.

....

....
[4]

Q:4 [MJ 2010 – 21 (2)]

2 The diagram below shows, for a given temperature T, a Boltzmann distribution of the kinetic energy of the molecules of a mixture of two gases that will react together, such as nitrogen and hydrogen.
The activation energy for the reaction, E_a, is marked.



(a) On the graph above,

(i) draw a new distribution curve, **clearly labelled T′**, for the same mixture of gases at a higher temperature, **T′**;

(ii) **mark clearly, as H**, the position of the activation energy of the reaction at the higher temperature, **T′**.

[3]

(b) Explain the meaning of the term *activation energy*.

....

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.... [2]

Q:5 [M/J 2011 – 21 (5 :d)]

(d) The standard enthalpy change of combustion of C_2H_2, $\Delta H^{\ominus}_{c}$, is $-1300\,kJ\,mol^{-1}$ at 298 K.

Values of relevant standard enthalpy changes of formation, $\Delta H^{\ominus}_{f}$, measured at 298 K, are given in the table.

substance	$\Delta H^{\ominus}_{f}$ / kJ mol^{-1}
$CO_2(g)$	–394
$H_2O(l)$	–286

(i) Write balanced equations, with state symbols, that represent

the standard enthalpy change of combustion, $\Delta H^{\ominus}_{c}$, of C_2H_2, and

....

the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, of C_2H_2.

....

(ii) Use the data above and your answer to **(i)** to calculate the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, of C_2H_2.
Show clearly whether the standard enthalpy change of formation of C_2H_2 has a positive or negative value.

[6]

Q:6 [M/J 12 – 21 (3)]

3 Methanol, CH_3OH, is considered to be a possible alternative to fossil fuels, particularly for use in vehicles.

Methanol can be produced from fossil fuels and from agricultural waste. It can also be synthesised from carbon dioxide and hydrogen.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of formation of carbon dioxide.

equation ....

definition ....

....

.... [3]

(b) Relevant $\Delta H_f^\ominus$ values for the reaction that synthesises methanol are given in the table.

compound	$\Delta H_f^\ominus$ / kJ mol^{-1}
$CO_2(g)$	–394
$CH_3OH(g)$	–201
$H_2O(g)$	–242

(i) Use these values to calculate $\Delta H^\ominus_{reaction}$ for this synthesis of methanol.

Include a sign in your answer.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

$\Delta H^\ominus_{reaction}$ = ...kJ mol^{-1}

(ii) Suggest **one** possible environmental advantage of this reaction. Explain your answer.

....

....

[5]

Q:7 [M/J 2012 – 22 (2)]

2 Alcohols such as methanol, CH_3OH, are considered to be possible replacements for fossil fuels because they can be used in car engines.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of combustion, $\Delta H^{\ominus}_{c}$, for methanol at 298 K.

equation ....

definition ....

....

.... [3]

Methanol may be synthesised from carbon monoxide and hydrogen. Relevant $\Delta H^{\ominus}_{c}$ values for this reaction are given in the table below.

compound	$\Delta H^{\ominus}_{c}$ / kJ mol^{-1}
$CO(g)$	–283
$H_2(g)$	–286
$CH_3OH(g)$	–726

(b) Use these values to calculate $\Delta H^{\ominus}_{reaction}$ for the synthesis of methanol, using the following equation. Include a sign in your answer.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

$\Delta H^{\ominus}_{reaction}$ = ...kJ mol^{-1}

[3]

Q:8 [M/J 2012 – 23 (3)]

3 With the prospect that fossil fuels will become increasingly scarce in the future, many compounds are being considered for use in internal combustion engines. One of these is DME or dimethyl ether, CH_3OCH_3. DME is a gas which can be synthesised from methanol. Methanol can be obtained from biomass, such as plant waste from agriculture.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of combustion, ΔH°_c, for DME at 298 K.

equation ....

definition ....

....

.... [3]

(b) DME may be synthesised from methanol. Relevant enthalpy changes of formation, ΔH°_f, for this reaction are given in the table below.

compound	ΔH°_f / kJ mol^{-1}
$CH_3OH(l)$	–239
$CH_3OCH_3(g)$	–184
$H_2O(l)$	–286

Use these values to calculate $\Delta H^\circ_{reaction}$ for the synthesis of DME, using the following equation. Include a sign in your answer.

$$2CH_3OH(l) \rightarrow CH_3OCH_3(g) + H_2O(l)$$

$\Delta H^\circ_{reaction}$ = ... kJ mol^{-1}
[3]

Q:9 [M/J 2013 – 21 (2 :c)

(c) The standard enthalpy changes of formation of $NH_3(g)$ and $H_2O(g)$ are as follows.

$NH_3(g)$, $\Delta H_f^\ominus = -46.0\ \text{kJ mol}^{-1}$ $H_2O(g)$, $\Delta H_f^\ominus = -242\ \text{kJ mol}^{-1}$

Use these data and the value of $\Delta H^\ominus_{reaction}$ given below to calculate the standard enthalpy change of formation of NO(g).
Include a sign in your answer.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \qquad \Delta H^\ominus = -906\ \text{kJ mol}^{-1}$$

[4]

Q:10 [M/J 2013 – 23 (1)]

1 Carbon disulfide, CS_2, is a volatile, flammable liquid which is produced in small quantities in volcanoes.

(a) The sequence of atoms in the CS_2 molecule is sulfur to carbon to sulfur.

(i) Draw a 'dot-and-cross' diagram of the carbon disulfide molecule.
Show outer electrons only.

(ii) Suggest the shape of the molecule and state the bond angle.

shape ....

bond angle ....
[3]

(b) Carbon disulfide is readily combusted to give CO_2 and SO_2.

(i) Construct a balanced equation for the complete combustion of CS_2.

....

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

....

....

....
[3]

(c) Calculate the standard enthalpy change of formation of CS_2 from the following data. Include a sign in your answer.

standard enthalpy change of combustion of CS_2 = $-1110\ \text{kJ mol}^{-1}$

standard enthalpy change of formation of CO_2 = $-395\ \text{kJ mol}^{-1}$

standard enthalpy change of formation of SO_2 = $-298\ \text{kJ mol}^{-1}$

[3]

Q:11 [O/N 2002 (4: e)

(e) Urea, $CO(NH_2)_2$, is a naturally occurring substance which can be hydrolysed with water to form ammonia according to the following equation.

$$H_2O(l) + CO(NH_2)_2(aq) \longrightarrow CO_2(aq) + 2NH_3(aq)$$

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

compound	$\Delta H_f^{\ominus}$/kJ mol^{-1}
$H_2O(l)$	−287.0
$CO(NH_2)_2(aq)$	−320.5
$CO_2(aq)$	−414.5
$NH_3(aq)$	−81.0

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

[2]

Q: 12 [O/N 2003 (3)]

3 (a) (i) What is meant by the *standard enthalpy change of formation*, $\Delta H^{\ominus}_f$, of a compound? Explain what is meant by the term *standard*.

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....

(ii) Write an equation, with state symbols, for the $\Delta H^{\ominus}_f$ of water.

....

(iii) Explain why the $\Delta H^{\ominus}_f$ for water is identical to the standard enthalpy change of combustion of hydrogen.

....

....[4]

(b) When calcium is placed in water, aqueous calcium hydroxide is formed and hydrogen is given off.

(i) Write the equation for the reaction of calcium with water.

....

(ii) When 1.00 g of calcium is placed in 200 g of water, the temperature increases by 12.2 °C when the reaction is completed. The specific heat capacity of water, *c*, is 4.2 $J\,g^{-1}\,K^{-1}$.

Calculate the heat released in the experiment.

(iii) Calculate the standard enthalpy change of reaction in $kJ\,mol^{-1}$ for your equation in **(b)(i)**.

[4]

(c) **(i)** State *Hess' Law*.

....

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....

(ii) Use Hess' Law and your result in **(b)(iii)** to calculate the $\Delta H_f^{\ominus}$ of $Ca(OH)_2$ (aq). You also need the $\Delta H_f^{\ominus}$ of water which is $-286\,kJ\,mol^{-1}$.

[4]

(d) Calculate the volume of hydrogen, measured at room temperature and pressure, liberated in the experiment described in **(b)(ii)**.

[2]

Q:13 [O/N 2004 (1: d)

(d) (i) Explain how enthalpy changes, ΔH values, for covalent bonded molecules can be calculated from bond energies.

....

....

(ii) Use bond energies from the *Data Booklet* to calculate ΔH for the following dissociation.

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

[3]

Q:14 [O/N 2005 (2)]

2 Carbon disulphide, CS_2, is a volatile, stinking liquid which is used to manufacture viscose rayon and cellophane.

(a) The carbon atom is in the centre of the CS_2 molecule.

Draw a 'dot-and-cross' diagram of the carbon disulphide molecule.

Show outer electrons only.

[2]

(b) Suggest the shape of the molecule and give its bond angle.

shape ....

bond angle .... [2]

(c) Explain the term *standard enthalpy change of formation*, $\Delta H_f^{\ominus}$.

....

....

.... [3]

(d) Calculate the standard enthalpy change of formation of CS_2 from the following data.

standard enthalpy change of formation of SO_2 = $-298\ \text{kJ mol}^{-1}$

standard enthalpy change of formation of CO_2 = $-395\ \text{kJ mol}^{-1}$

standard enthalpy change of combustion of CS_2 = $-1110\ \text{kJ mol}^{-1}$

[3]

Q: 15 [O/N 2006 (4: d,e)

The unsaturated hydrocarbon **Z** is obtained by cracking hexane and is important in the chemical industry.

The standard enthalpy change of combustion of **Z** is –2059 kJ mol^{-1}.

(d) Define the term *standard enthalpy change of combustion*.

....

.... [2]

When 0.47 g of **Z** were completely burnt in air, the heat produced raised the temperature of 200 g of water by 27.5 °C.

(e) **(i)** Calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the relative molecular mass of **Z**.

[4]

Q: 16 [O/N 2007 (1: e)

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

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H_c^{\ominus} = -393.7\ \text{kJ mol}^{-1}$$

$$H_2(g) + ½O_2(g) \rightarrow H_2O(l) \qquad \Delta H_c^{\ominus} = -285.9\ \text{kJ mol}^{-1}$$

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \qquad \Delta H_c^{\ominus} = -1411.0\ \text{kJ mol}^{-1}$$

Use the data to calculate the standard enthalpy change of formation, $\Delta H_f^{\ominus}$, in kJ mol^{-1}, of ethene at 298 K.

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

$\Delta H_f^{\ominus}$ = ... kJ mol^{-1}
[3]

Q: 17 [O/N 2008 (2: c)

(c) **(i)** Define the term *standard enthalpy change of formation.*

....

....

....

(ii) Use the data below to calculate the standard enthalpy change of formation of ketene.

	$\Delta H^{\ominus}$/kJ mol^{-1}
standard enthalpy change of formation of CO_2	–395
standard enthalpy change of combustion of H_2	–286
standard enthalpy change of combustion of $CH_2{=}C{=}O$	–1028

[6]

Q:18 [O/N 2009 – 21 (3: a, c)]

3 Alkanes such as methane, CH_4, undergo few chemical reactions. Methane will, however, react with chlorine but not with iodine.

Relevant standard enthalpy changes of formation for the reaction of methane with chlorine to form chloromethane, CH_3Cl, are given below.

	$\Delta H_f^{\ominus}$/kJ mol^{-1}
CH_4	–75
CH_3Cl	–82
HCl	–92

(a) **(i)** Use the data to calculate $\Delta H^{\ominus}_{reaction}$ for the formation of CH_3Cl.

$$CH_4 \;+\; Cl_2 \;\rightarrow\; CH_3Cl \;+\; HCl$$

(ii) The corresponding reaction with iodine does **not** take place.

Use bond energy data from the *Data Booklet* to calculate a 'theoretical value' for $\Delta H_{reaction}$ for the following equation.

$$CH_4 + I_2 \rightarrow CH_3I + HI$$

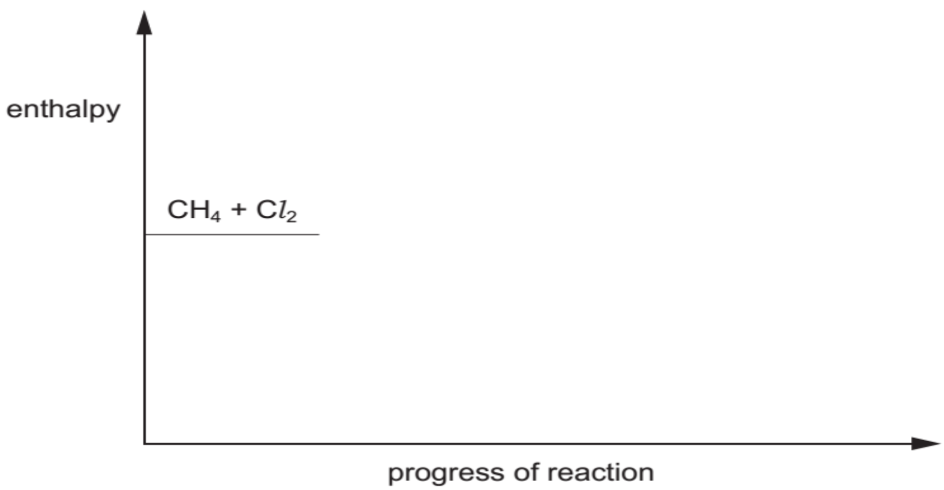
(iii) Suggest why this reaction does **not** in fact occur.

....

....

[5]

(c) The energy of activation for the formation of CH_3Cl is 16 kJ mol^{-1}.
Use this figure and your answer to **(a)(i)** to complete the reaction pathway diagram below showing the formation of CH_3Cl from CH_4 and Cl_2.
Show clearly the intermediate organic species and the final products.
Indicate on your sketch the relevant enthalpy changes and their values.



[4]

Q: 19 [O/N 2010 – 22 (3: d,e,f)]

The unsaturated hydrocarbon, **E**, is obtained by cracking hexane and is important in the chemical industry.

The standard enthalpy change of combustion of **E** is $-2059\,kJ\,mol^{-1}$.

(d) Define the term *standard enthalpy change of combustion*.

....

.... [2]

When 0.47 g of **E** was completely burnt in air, the heat produced raised the temperature of 200 g of water by 27.5 °C. Assume no heat losses occurred during this experiment.

(e) **(i)** Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the relative molecular mass, M_r, of **E**.

[4]

(f) Deduce the molecular formula of **E**.

[1]

Q: 20 [O/N 2011 – 22 (3)]

3 For some chemical reactions, such as the thermal decomposition of potassium hydrogencarbonate, $KHCO_3$, the enthalpy change of reaction cannot be measured directly.

In such cases, the use of Hess' Law enables the enthalpy change of reaction to be calculated from the enthalpy changes of other reactions.

(a) State Hess' Law.

....

....

.... [2]

In order to determine the enthalpy change for the thermal decomposition of potassium hydrogencarbonate, two separate experiments were carried out.

experiment 1

30.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid (an excess) was placed in a conical flask and the temperature recorded as 21.0 °C.
When 0.0200 mol of potassium carbonate, K_2CO_3, was added to the acid and the mixture stirred with a thermometer, the maximum temperature recorded was 26.2 °C.

(b) (i) Construct a balanced equation for this reaction.

....

(ii) Calculate the quantity of heat produced in **experiment 1**, stating your units.
Use relevant data from the *Data Booklet* and assume that all solutions have the same specific heat capacity as water.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of K_2CO_3.
Give your answer in kJ mol^{-1} and include a sign in your answer.

(iv) Explain why the hydrochloric acid must be in an excess.

....

.... [4]

experiment 2

The experiment was repeated with 0.0200 mol of potassium hydrogencarbonate, $KHCO_3$.
All other conditions were the same.
In the second experiment, the temperature fell from 21.0 °C to 17.3 °C.

(c) **(i)** Construct a balanced equation for this reaction.

....

(ii) Calculate the quantity of heat absorbed in **experiment 2**.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of $KHCO_3$.
Give your answer in kJ mol^{-1} and include a sign in your answer.

[3]

(d) When $KHCO_3$ is heated, it decomposes into K_2CO_3, CO_2 and H_2O.

$$2KHCO_3 \rightarrow K_2CO_3 + CO_2 + H_2O$$

Use Hess' Law and your answers to **(b)(iii)** and **(c)(iii)** to calculate the enthalpy change for this reaction.
Give your answer in kJ mol^{-1} and include a sign in your answer.

[2]

Q:21 [O/N 2013 – 21 (5: d)

(c) Propane and butane have different values of standard enthalpy change of combustion.

Define the term *standard enthalpy change of combustion*.

....

....

.... [2]

(d) A 125 cm^3 sample of propane gas, measured at 20 °C and 101 kPa, was completely burnt in air.
The heat produced raised the temperature of 200 g of water by 13.8 °C.
Assume no heat losses occurred during this experiment.

(i) Use the equation $pV = nRT$ to calculate the mass of propane used.

(ii) Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(iii) Use the data above and your answers to **(i)** and **(ii)** to calculate the energy produced by the burning of 1 mol of propane.

[5]

Q:22 [O/N 2013 – 23 (2: c, d)

(c) Define the term *standard enthalpy change of combustion*.

....

....

.... [2]

(d) A 1.00 cm^3 sample of $C_{14}H_{30}$ was completely burnt in air.
The heat produced raised the temperature of 250 g of water by 34.6 °C.
Assume no heat losses occurred during this experiment.
The density of $C_{14}H_{30}$ is 0.763 g cm^{-3}.

(i) Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the energy produced by the combustion of 1 mol of $C_{14}H_{30}$.

[5]