Q1. (a) The table below contains some mean bond enthalpy data.

Bond	H–O	O–O	O=O
Mean bond enthalpy/kJ mol^{-1}	463	146	496

The bonding in hydrogen peroxide, H_2O_2, can be represented by H–O–O–H. Use these data to calculate the enthalpy change for the following reaction.

$$H_2O_2(g) \rightarrow H_2O_2(g) + \tfrac{1}{2} O_2(g)$$

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(3)

(b) The standard enthalpy of formation, $\Delta H_f^{\ominus}$ for methane, is –74.9 kJ mol^{-1}. Write an equation, including state symbols, for the reaction to which this enthalpy change applies.

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(2)

(c) The enthalpy changes for the formation of atomic hydrogen and atomic carbon from their respective elements in their standard states are as follows.

$\tfrac{1}{2} H_2(g) \rightarrow H(g)$ $\quad \Delta H^{\ominus} = +218$ kJ mol^{-1}

$C(s) \rightarrow C(g)$ $\quad \Delta H^{\ominus} = +715$ kJ mol^{-1}

(i) By reference to its structure, suggest why a large amount of heat energy is required to produce free carbon atoms from solid carbon.

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(ii) Parts (b) and (c) give enthalpy data for the formation of $CH_4(g)$, H(g) and C(g). Use these data and Hess's Law to calculate the value of the enthalpy change for the following reaction.

$$CH_4(g) \rightarrow C(g) + 4H(g)$$

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(iii) Use your answer from part (c)(ii) to calculate a value for the mean bond enthalpy of a C–H bond in methane.

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(5)
(Total 10 marks)

Q2. (a) Define the term *standard enthalpy of combustion*, $\Delta H_c^{\ominus}$

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(3)

(b) Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

Bond	C == C	C—C	C—H	O == O	O == C	O—H
Mean bond enthalpy/ kJ mol^{-1}	612	348	412	496	743	463

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(3)

(c) State why the standard enthalpy of formation, $\Delta H_f^{\ominus}$, of oxygen is zero.

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(1)

(d) Use the data from the table below to calculate a more accurate value for the standard enthalpy of combustion of propene.

Compound	$C_3H_6(g)$	$CO_2(g)$	$H_2O(g)$
Standard enthalpy of formation, $\Delta H_f^{\ominus}$/ kJ mol^{-1}	+20	–394	–242

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(3)

(e) Explain why your answer to part (b) is a less accurate value than your answer to part (d).

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(2)
(Total 12 marks)

Q3. (a) Explain the meaning of the terms *mean bond enthalpy* and *standard enthalpy of formation*.

Mean bond enthalpy .....

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Standard enthalpy of formation .....

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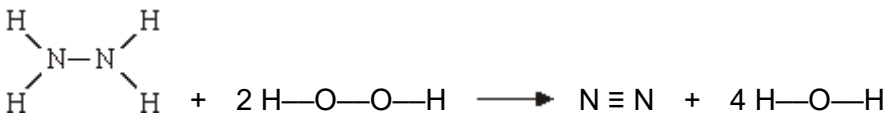
(5)

(b) Some mean bond enthalpies are given below.

Bond	N–H	N–N	N≡N	H–O	O–O
Mean bond enthalpy/kJ mol^{-1}	388	163	944	463	146

Use these data to calculate the enthalpy change for the following gas-phase reaction between hydrazine, N_2H_4, and hydrogen peroxide, H_2O_2



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(3)

(c) Some standard enthalpies of formation are given below.

	$N_2H_4(g)$	$H_2O_2(g)$	$H_2O(g)$
$\Delta H_f^{\ominus}$ /kJ mol^{-1}	+75	–133	–242

These data can be used to calculate the enthalpy change for the reaction in part (b).

$$N_2H_4(g) + 2H_2O_2(g) \rightarrow N_2(g) + 4H_2O(g)$$

(i) State the value of $\Delta H_f^{\ominus}$ for $N_2(g)$.

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(ii) Use the $\Delta H_f^{\ominus}$ values from the table to calculate the enthalpy change for this reaction.

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(4)

(d) Explain why the value obtained in part (b) is different from that obtained in part (c)(ii).

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(1)

(Total 13 marks)

Q4. Use the information below to answer this question.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^{\ominus} = -394 \text{ kJ mol}^{-1}$$

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

$$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g) \qquad \Delta H^{\ominus} = -126 \text{ kJ mol}^{-1}$$

The standard enthalpy of combustion of butane, in kJ mol^{-1}, is

A −2880

B −2590

C −806

D −554

(Total 1 mark)

Q5. This question is about the reaction given below.

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

Enthalpy data for the reacting species are given in the table below.

Substance	CO(g)	$H_2O(g)$	$CO_2(g)$	$H_2(g)$
$\Delta H_f^{\ominus}$ / kJ mol^{-1}	−110	−242	−394	0

The standard enthalpy change for this reaction of carbon monoxide and steam is

A +42 kJ mol^{-1}

B −42 kJ mol^{-1}

C +262 kJ mol^{-1}

D −262 kJ mol^{-1}

(Total 1 mark)

Q6. Methanol, CH_3OH, is a convenient liquid fuel.

(a) An experiment was conducted to determine the enthalpy of combustion of liquid methanol. The energy obtained from burning 2.12 g of methanol was used to heat 150 g of water. The temperature of the water rose from 298 K to 362 K. (The specific heat capacity of water is 4.18 J K^{-1} g^{-1})

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

(ii) Use the data above to calculate a value for the enthalpy of combustion of one mole of liquid methanol.

(7)

(b) Methanol can be synthesised from methane and steam by a process that occurs in two stages.

Stage 1 $CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g)$ $\Delta H^{\ominus} = +206$ kJ mol^{-1}

Stage 2 $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^{\ominus} = -91$ kJ mol^{-1}

(i) Explain why, in *Stage 1*, a higher yield of hydrogen and carbon monoxide is **not** obtained if the pressure is increased.

(ii) *Stage 2* is carried out at a compromise temperature of 500K. By considering what would happen at higher and lower temperatures, explain why 500 K is considered to be a compromise for *Stage 2*.

(5)

(c) The standard enthalpies of combustion of carbon monoxide and of hydrogen are –283 kJ mol^{-1} and –286 kJ mol^{-1}, respectively. Use these data and the enthalpy change for *Stage 2* to calculate a value for the standard enthalpy of combustion of gaseous methanol.

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(Total 15 marks)