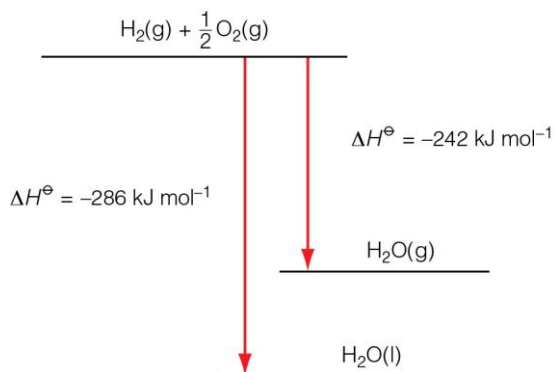


## Pages 259–261 Exam practice questions

- 1 a) Insulated [1] polystyrene beaker/container [1], 0–50 °C thermometer. [1]
- b) i)  $\text{Mg(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{MgSO}_4\text{(aq)} + \text{Cu(s)}$  [1]
- ii) Energy transferred =  $50 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 5.0 \text{ K}$  [1]  
 $= 1045 \text{ J}$  [1]
- iii) Any two of:
- the density of the solution =  $1.0 \text{ g cm}^{-3}$
  - the specific heat capacity of the solution =  $4.18 \text{ J g}^{-1} \text{ K}^{-1}$
  - the mass of magnesium added is ignored. [2]
- iv) Amount of  $\text{CuSO}_4$  used =  $\frac{50}{1000} \text{ dm}^3 \times 0.040 \text{ mol dm}^{-3} = 0.0020 \text{ mol}$  [1]
- Energy transferred for 1 mol  $\text{CuSO}_4 = \frac{1045 \text{ J}}{0.0020 \text{ mol}}$  [1]  
 $= 522\,000 \text{ J mol}^{-1}$
- Enthalpy change =  $-520 \text{ kJ mol}^{-1}$  (2 s.f.) [1]
- 2 a) Energy needed =  $500 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 80 \text{ K} = 167 \text{ kJ}$  [1]
- b) Energy from burning butane =  $2876 \text{ kJ mol}^{-1}$
- Amount of butane burnt to heat the water =  $167 \text{ kJ} \div 2876 \text{ kJ mol}^{-1} = 0.0581 \text{ mol}$ . [1]
- c) Volume of butane needed =  $0.0581 \text{ mol} \times 24 \text{ dm}^3 \text{ mol}^{-1} = 1.4 \text{ dm}^3$  (2 s.f.) [1]
- d) The calculation assumes that combustion of butane to carbon dioxide and water is complete [1], and that all the energy from the flame heats the water and not the container or the surroundings. [1]
- 3 a)  $\text{CH}_3\text{COONa(aq)} [1] + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$  [1]
- b) To avoid any spillage/loss of reactants due to the effervescence of  $\text{CO}_2\text{(g)}$ . [1]
- c) Energy change =  $50 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 8.0 \text{ K}$  [1]  
 $= 1672 \text{ J}$  [1]
- d) Amount of ethanoic acid =  $\frac{50}{1000} \text{ dm}^3 \times 1.0 \text{ mol dm}^{-3} = 0.050 \text{ mol}$  [1]
- e) 0.050 mol ethanoic acid absorbs 1672 J
- Therefore 1 mol ethanoic acid absorbs  $\frac{1672 \text{ J}}{0.050 \text{ mol}} = 33\,400 \text{ J mol}^{-1}$  [1]
- Standard enthalpy change of reaction =  $+33\,000 \text{ J mol}^{-1}$  or  $+33 \text{ kJ mol}^{-1}$  [2]  
 ([1] for 33, [1] for sign and units)
- f) Two significant figures because the data for the volume of acid, the concentration of acid and the temperature rise are given to 2 significant figures. [1]

- 4 a) Upper level [1]; lower levels with labelled axes. [2]



- b) Enthalpy change =  $+44 \text{ kJ mol}^{-1}$  [2]  
 ([1] for sign, [1] for the value)
- 5 a) The standard enthalpy change of combustion of a substance,  $\Delta_c H^\ominus$ , is the enthalpy change when one mole of the substance [1] burns completely in oxygen [1] under standard conditions. [1]
- b)  $\text{CH}_3\text{CH}_2\text{CH}_3(\text{l}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$  [2]  
 (combustion products [1]; balancing [1])
- c) The standard enthalpy change of formation of a compound,  $\Delta_f H^\ominus$ , is the enthalpy change when one mole of the compound [1] forms from its elements under standard conditions [1] with the elements and the compound in their standard (stable) states. [1]
- d)  $3\text{C}(\text{s, graphite}) + 3\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{CH}_2\text{CHO}(\text{l})$  [2]  
 (elements in correct states [1], balancing [1])
- 6 a)  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}(\text{s}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{g})$  [2]  
 ([1] for balanced equation, [1] for state symbols)
- b) Endothermic [1], because bonds must be broken between  $\text{CaSO}_4$  and  $\text{H}_2\text{O}$  molecules [1] and no new bonds are formed.
- c) Because it is difficult to determine just how much energy is required to decompose a known mass of the hydrate. The salt has to be heated to a high temperature at which it is impossible to measure the energy changes. [1]
- d)  $\Delta_r H^\ominus = \Delta_f H^\ominus[\text{CaSO}_4(\text{s})] + 2\Delta_f H^\ominus[\text{H}_2\text{O}(\text{l})] - \Delta_f H^\ominus[\text{CaSO}_4 \cdot 2\text{H}_2\text{O}(\text{s})]$  [2]  
 $= (-1434 \text{ kJ mol}^{-1}) + (2 \times -286 \text{ kJ mol}^{-1}) - (-2023 \text{ kJ mol}^{-1})$  [1]  
 $= +17 \text{ kJ mol}^{-1}$  [2]  
 ([1] for 17, [1] for sign and units)

## 7 a) Reaction 1

$$\Delta_r H = \Delta_f H^\ominus[\text{CO}_2(\text{g})] - \Delta_f H^\ominus[\text{SnO}_2(\text{s})] \quad [1]$$

$$= (-394 \text{ kJ mol}^{-1}) - (-581 \text{ kJ mol}^{-1})$$

$$= +187 \text{ kJ mol}^{-1} \quad [2]$$

([1] for 187, [1] for sign and units)

## Reaction 2

$$\Delta_r H = 2\Delta_f H^\ominus[\text{CO}(\text{g})] - \Delta_f H^\ominus[\text{SnO}_2(\text{s})] \quad [1]$$

$$= (2 \times -110 \text{ kJ mol}^{-1}) - (-581 \text{ kJ mol}^{-1})$$

$$= -220 + 581 = +361 \text{ kJ mol}^{-1} \quad [2]$$

([1] for 361, [1] for sign and units)

## b) Both reactions are endothermic requiring an input of energy using fuels.

The energy input and therefore fuel used in Reaction 1 is less than Reaction 2 [1], so on just this basis Reaction 1 is more economic. [1]

8 a) i) Box top right:  $2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$  [1]

$$\text{Box below: } 2\text{C}(\text{s}) + 3\text{H}_2(\text{g}) + 3.5\text{O}_2(\text{g}) \quad [1]$$

ii) The standard enthalpy change of formation of a compound is the enthalpy change when one mole of the compound [1] forms from its elements [1] under standard conditions with the elements and the compound in their standard (stable) states. [1]

$$\text{iii) } \Delta H_1 = \Delta_f H^\ominus [\text{C}_2\text{H}_5\text{OH}(\text{l})] = -277 \text{ kJ mol}^{-1} \quad [1]$$

$$\Delta H_2 = 2\Delta_f H^\ominus [\text{CO}_2(\text{g})] + 3\Delta_f H^\ominus [\text{H}_2\text{O}(\text{l})] \quad [1]$$

$$= (2 \times -394 \text{ kJ mol}^{-1}) + (3 \times -286 \text{ kJ mol}^{-1}) = -1646 \text{ kJ mol}^{-1} \quad [1]$$

$$\Delta_c H[\text{ethanol}] = -\Delta H_1 + \Delta H_2 \quad [1]$$

$$= -(-277 \text{ kJ mol}^{-1}) + (-1646 \text{ kJ mol}^{-1}) = -1369 \text{ kJ mol}^{-1} \quad [2]$$

([1] for 1369, [1] for sign and units)

$$\text{b) i) } \text{Energy transferred} = 150 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 14.8 \text{ K} \quad [1]$$

$$= 9279.6 \text{ J} \quad [1]$$

$$\text{ii) } \text{Amount of ethanol burnt} = 0.898 \text{ g} \div 46 \text{ g mol}^{-1} = 0.01952 \text{ mol} \quad [1]$$

$$\text{Energy transferred per mole of ethanol burnt} = 9279.6 \text{ J} \div 0.01952 \text{ mol}$$

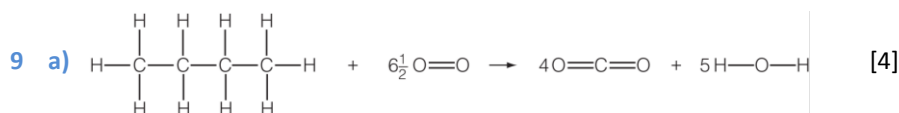
$$= 475390 \text{ J mol}^{-1}$$

$$= 475 \text{ kJ mol}^{-1} \text{ (3 s.f.)}$$

$$\Delta_c H[\text{ethanol}] = -475 \text{ kJ mol}^{-1} \quad [2]$$

iii) The experimental method is very crude and greatly underestimates the enthalpy change for the combustion of ethanol. Most of the energy from the flame is not transferred to the water but heats up the rest of the apparatus and the surrounding air. [1]

Enthalpy changes of formation in tables of data are derived from enthalpy changes measured by much more sophisticated apparatus such as a bomb calorimeter. [1]



([1] for each molecule)

b) *Bonds broken*

3 C—C

10 C—H

$6\frac{1}{2}$  O=O [1]

*Bonds formed*

8 C=O

10 H—O [1]

c) Enthalpy change =  $+(3 \times 347 \text{ kJ mol}^{-1}) + (10 \times 413 \text{ kJ mol}^{-1}) + (6\frac{1}{2} \times 498 \text{ kJ mol}^{-1})$  [1]

$-(8 \times 805 \text{ kJ mol}^{-1}) - (10 \times 464 \text{ kJ mol}^{-1})$  [1]

$= +8408 \text{ kJ mol}^{-1} - 11\,080 \text{ kJ mol}^{-1}$

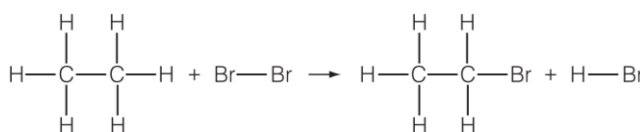
$= -2672 \text{ kJ mol}^{-1}$  [2]

([1] for 2672, [1] for sign and units)

d) The bond enthalpies values are averages for a range of molecules and so are not exactly the right values for all the molecules in this reaction. [1]

Bond enthalpy values apply to molecules in the gas phase. Water is a liquid at 298 K and the energy change for water condensing from a gas to a liquid alters the value at 298 K. [1]

10 a) *Reaction 1*



*Bonds broken*

1 C—H

1 Br—Br

*Bonds formed*

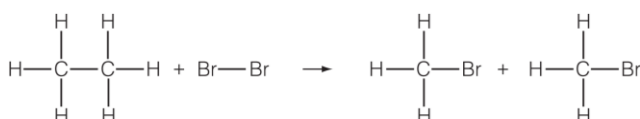
1 C—Br

1 H—Br [1]

Enthalpy change =  $+413 \text{ kJ mol}^{-1} + 193 \text{ kJ mol}^{-1} - 290 \text{ kJ mol}^{-1} - 366 \text{ kJ mol}^{-1}$  [1]

$= -50 \text{ kJ mol}^{-1}$  [1]

*Reaction 2*



*Bonds broken*

1 C—C

1 Br—Br

*Bonds formed*

2 C—Br

[1]

Enthalpy change =  $+347 \text{ kJ mol}^{-1} + 193 \text{ kJ mol}^{-1} - (2 \times 290 \text{ kJ mol}^{-1})$  [1]

$= -40 \text{ kJ mol}^{-1}$  [1]

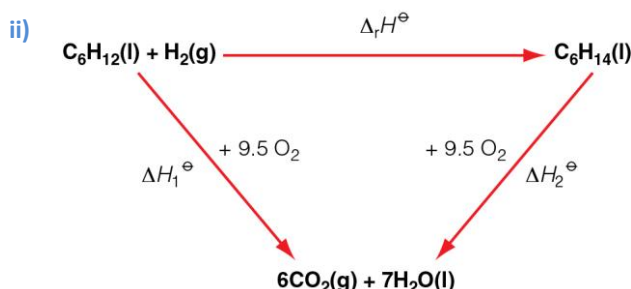
b) Any two good points in an argument [2] such as:

- The enthalpy change for Reaction 1 is more negative than for Reaction 2, which might explain why it is the one that is observed to happen.
- However the difference is small. As explained in Section 2.6 of Student Book 1, the enthalpy change for a reaction is not a reliable guide to the preferred direction of change. Other factors are involved.
- Sometimes it is the reaction that goes faster that is preferred even when an alternative is more exothermic. In this instance, the energy required to break a C–H bond in Reaction 1 is greater than the energy required to break a C–C bond in Reaction 2. Easier breakage of the C–C bond in Reaction 2 could suggest that this reaction is likely to happen more easily.

c) Mean bond enthalpies do not always tally with the bond strengths in specific compounds. [1]  
Mean bond enthalpies are calculated assuming substances are gases whereas bromine in both reactions and C<sub>2</sub>H<sub>5</sub>Br in Reaction 1 are liquids. [1]

11 a) i)  $\text{C}_6\text{H}_{12}(\text{l}) + \text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_{14}(\text{l})$  [1]

A metal catalyst is used, usually one of nickel, palladium or platinum. [1]



Cycle [1] Balanced combustion products [1]; adding in 9.5 O<sub>2</sub> twice. [1]

iii)  $\Delta H_1 = \Delta_c H^\ominus[\text{C}_6\text{H}_{12}(\text{l})] + \Delta_c H^\ominus[\text{H}_2(\text{g})]$

$$= -4003 \text{ kJ mol}^{-1} + (-286 \text{ kJ mol}^{-1}) = -4289 \text{ kJ mol}^{-1} \quad [1]$$

$$\Delta H_2 = \Delta_c H^\ominus[\text{C}_6\text{H}_{14}(\text{l})] = -4163 \text{ kJ mol}^{-1} \quad [1]$$

$$\Delta_r H = \Delta H_1 - \Delta H_2 = -4289 \text{ kJ mol}^{-1} - (-4163 \text{ kJ mol}^{-1}) = -126 \text{ kJ mol}^{-1} \quad [1]$$

b) For all the alkanes the reaction is essentially the same. Each alkene has a C=C double bond.

One of the two C–C bonds breaks and an H–H bond breaks. Two new C–H bond form. [1] The only difference in each case is the number of unreactive carbon and hydrogen atoms in the alkene. [1] The molecules are all very similar and so the bond enthalpies have almost the same value. [1]

12 a) i) This is important in many forms of transport but especially in aviation because the fuel has to be lifted. [1]

ii) Energy given out per gram from burning methanol:

$$726 \text{ kJ mol}^{-1} \div 32 \text{ g mol}^{-1} = 22.7 \text{ kJ g}^{-1} \quad [1]$$

Energy given out per gram from burning octane:

$$5470 \text{ kJ mol}^{-1} \div 114 \text{ g mol}^{-1} = 47.5 \text{ kJ g}^{-1} \quad [1]$$

b) i) This is important wherever the space for storing the fuel is limited, for example on a motorcycle.

ii) Energy given out per  $\text{cm}^3$  from burning methanol

$$= (726 \text{ kJ mol}^{-1} \times 0.793 \text{ g cm}^{-3}) \div 32 \text{ g mol}^{-1}$$

$$= 18.0 \text{ kJ cm}^{-3} [1]$$

Energy given out per  $\text{cm}^3$  from burning octane

$$= (5470 \text{ kJ mol}^{-1} \times 0.703 \text{ g cm}^{-3}) \div 114 \text{ g mol}^{-1}$$

$$= 33.7 \text{ kJ cm}^{-3} [1]$$

c) This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning. Assess the quality of the answer taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make in the answer:

- Octane provides more energy per gram or per  $\text{cm}^3$ .
- All the atoms in the fuel react with oxygen when it burns, unlike methanol which includes oxygen atoms in its molecules.
- For aviation there is a clear benefit to using the hydrocarbon fuel.
- However, octane is obtained from a non-renewable resource
- Octane requires much more oxygen for complete combustion (over 8 times as much per mole) so a much larger volume of air has to be supplied to burn the fuel.
- Methanol has the potential to become a more sustainable fuel because it can be made from a wide range of renewable resources, including many wastes.