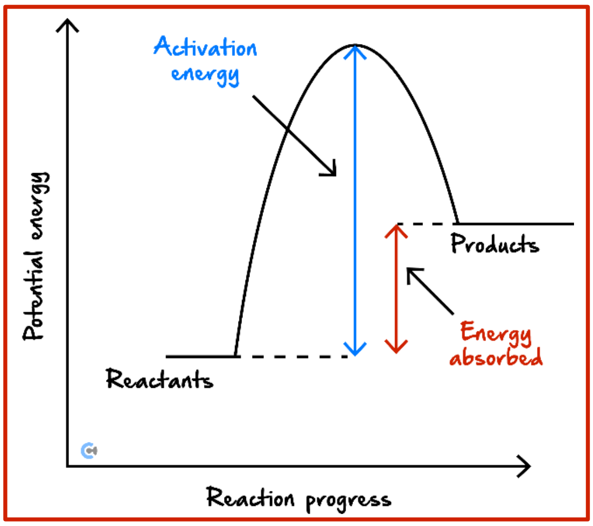
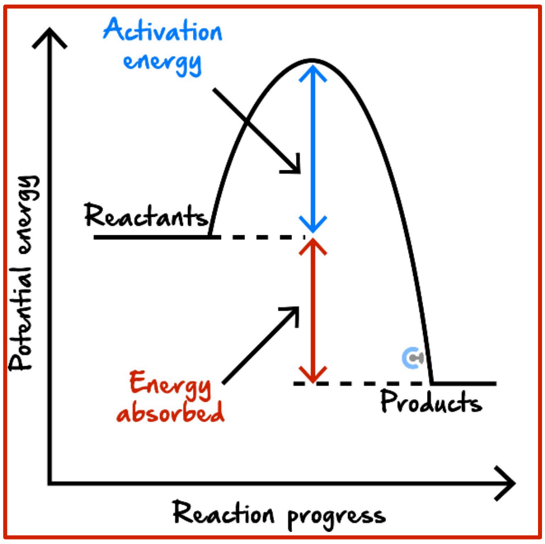
Website: contoureducation.com.au | Phone: 1800 888 300
Email: hello@contoureducation.com.au

# VCE Chemistry ¾
## Thermochemistry & Energy Calculation [0.1]
## Workshop Solutions

## Section A: Recap

**Learning Objective: [1.1.1] - Identify $\Delta H$ & $Ea$ in Endothermic/Exothermic Energy Profile Diagrams**

| Endothermic Reactions | Exothermic Reactions |
|---|---|
| [Absorb] / [Release] thermal energy. | [Absorb] / [Release] thermal energy. |
| Energy Conversion: Thermal → Chemical | Energy Conversion: Chemical → Thermal |
| Change in enthalpy ($\Delta H$) sign is: [positive] / [negative] | Change in enthalpy ($\Delta H$) sign is: [positive] / [negative] |
|  |  |

- Energy profile diagram axis labels:
  - Vertical axis: "Enthalpy ($kJ$)"
  - Horizontal axis: "Reaction progress"

Definition

**Learning Objective: [1.1.2] – Identify Differences Between Complete and Incomplete Combustion & Write Their Thermochemical Combustion Equations**

| Complete Combustion | Incomplete Combustion |
|---|---|
| Produces $CO_2$. | Produces C or CO. |
| [More] / [Less] oxygen used overall. | [More] / [Less] oxygen used overall. |

- Method of balancing: CHO
- State of fuels: Never aqueous.
- Thermochemical equations include $\Delta H$.
- The state of water in thermochemical equations is liquid when using the databook.

Definition

**Learning Objective: [1.1.3] – Apply Changing Equations to Thermochemical Equations & Energy Profile Diagrams**

- When reversing the equation, the:
  - $\Delta H$: flips
  - $E_a$: changes completely
- When multiplying the equation by some coefficient:
  - $\Delta H$ value multiplied by same number
  - $\Delta H$ units $kJ/mol \rightarrow kJ$

Definition

**Learning Objective: [1.2.1] – Apply $q = \Delta H \times n$ to Energy Released**

- To convert between mass and moles, $n = \frac{m}{Mr}$.
- To find energy released by a fuel, use the Data Book, and the formula $q = \Delta H \times n$.

Definition

## Learning Objective: [1.2.2] – Apply $\Delta H$ in $kJ/mol$, $kJ/g$ and $kJ/mL$ to Energy Calculations

- The heat of combustion $\Delta H$ value can be calculated in the following ways:

| $\Delta H\ (kJ/mol)$ | $\Delta H\ (kJ/g)$ | $\Delta H\ (kJ/mL)$ |
|---|---|---|
| $\boxed{\Delta H = \frac{q}{n}}$ | $\boxed{\Delta H = \frac{q}{m}}$ | $\boxed{\Delta H = \frac{q}{V}}$ |

- Density is used to convert between mass and volume: $\boxed{d = \frac{m}{V}}$

**Question 1 Walkthrough.**

The density of ethanol is given to be 0.790 $g/mL$. Calculate the energy released, in kilojoules, when 3.92 $L$ of ethanol is combusted at SLC.

$$m = dV = 0.790 \times 3920\ mL = 3097\ g$$
$$n = \frac{m}{Mr} = 67.3\ mol$$
$$q = \Delta H \times n = 1370 \times 67.3 =$$

Definition

## Learning Objective: [1.2.3] – Apply $q = mc\Delta T$ to Find Energy Absorbed

- To measure the energy absorbed by water, it can be calculated by $\boxed{q = mc\Delta T}$, whereby the specific heat capacity of water is $\boxed{4.18\ J\ g^{-1}\ °C^{-1}}$.

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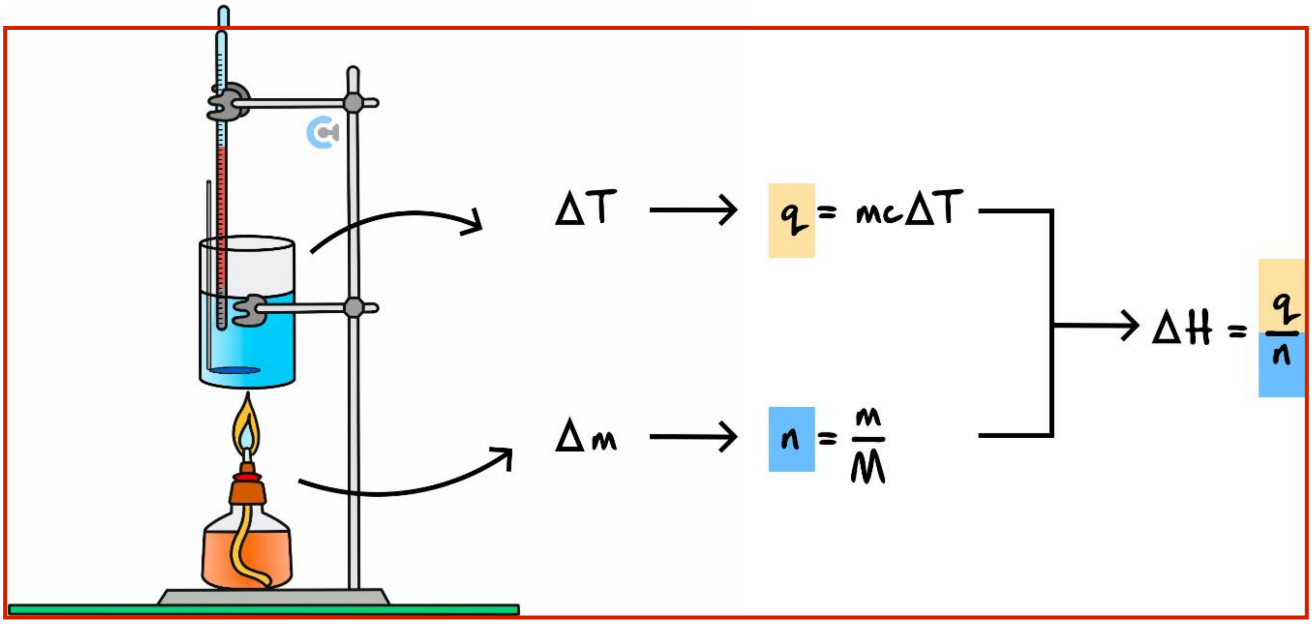
Definition

**Learning Objective: [1.2.4] - Calculate $\Delta H$ Experimentally**




**Question 2** (4 marks) **Walkthrough.**

A sample of ethanol in a spirit burner which initially weighs 198.30 $g$ undergoes complete combustion. After the combustion is complete, it is found that the spirit burner weighs 196.09 $g$. The heat energy released is used to heat 150 $mL$ of water at SLC. The temperature of the water rises to 89.60°C.

Calculate the experimental molar heat of combustion of ethanol.

$$n(C_2H_5OH) = \frac{2.21\,g}{46} = 0.048\,mol$$

$$q = mc\Delta T = 150 \times 0.997 \times 4.18 \times (89.6 - 25)$$
$$= 40382.7\,J = 40.4\,kJ$$

$$\Delta H = \frac{q}{n} = \frac{40.4\,kJ}{0.048\,mol} = 841\,kJ/mol$$

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## Section B: Warm Up (13 Marks)

INSTRUCTION: 13 Marks. 8 Minutes Writing.

**Question 3** (1 mark)

**a.** What happens to the $\Delta H$ value when the whole equation is reversed? (0.5 marks)

$\Delta H$ sign $(+/-)$ flips around.

**b.** What happens to the $\Delta H$ value which is currently in $kJ/mol$ when the whole equation is doubled? (0.5 marks)

Doubles, + it turns into just $kJ$.

**Question 4** (2 marks)

Write the thermochemical equation for the combustion of octane.

$$C_8H_{18}(l) + \frac{25}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g) \quad \Delta H = -5470\ kJ/mol$$

**Question 5** (1 mark)

Calculate the amount of energy released when 2.78 $mol$ of ethanol is combusted, providing your answer in $kJ$.

$q(C_2H_5OH) = \Delta H \times n = 1360\ kJ/mol \times 2.78\ mol$
$= 3780.8\ kJ$

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**Question 6** (1 mark)

Calculate the amount of energy released when 1.62 $kg$ of diesel is combusted.

$$q = \Delta H \times m = 45\frac{kJ}{g} \times 1620\ g = 72900\ kJ$$

**Question 7** (1 mark)

Calculate the volume of petrol, in $mL$, which must undergo complete combustion to release 1500 $kJ$ of energy.

$$V = \frac{q}{\Delta H} = 44.1\ mL$$

**Question 8 Walkthrough.**

The density of ethanol is given to be 0.790 $g/mL$. Calculate the energy released, in kilojoules, when 3.92 $L$ of ethanol is combusted at SLC.

$$m = dV = 0.790 \times 3920\ mL = 3097\ g$$
$$n = \frac{m}{Mr} = 67.3\ mol$$
$$q = \Delta H \times n = 1370 \times 67.3$$

**Question 9** (2 marks)

Calculate the amount of energy released when 5.00 $g$ of ethane is combusted at 120°C, given that the molar heat of combustion of ethane at this temperature is $-1480\ kJ/mol$.

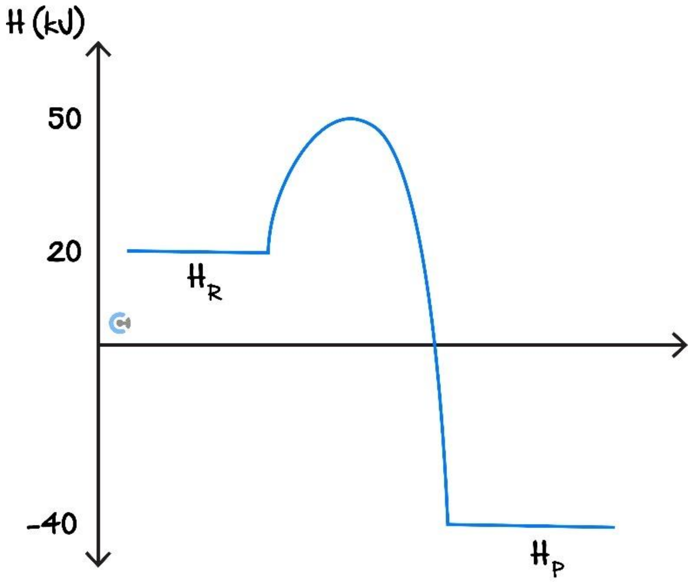
$n(C_2H_6) = \frac{m}{M} = \frac{5}{30} = 0.1667\ mol$

$q(C_2H_6) = \Delta H \times n = 1480\ kJ/mol \times 0.1667\ mol$

$\boxed{= 246.7\ kJ}$

*The following information applies to the two questions that follow.*

The diagram below represents the heat change for the reaction $A(g) + B(g) \rightarrow 2C(g)$.




**Question 10** (1 mark)

The enthalpy of reaction ($\Delta H$) for $C(g) \rightarrow 0.5A(g) + 0.5B(g)$ is:

**A.** $-120\ kJ\ mol^{-1}$

**B.** $+60\ kJ\ mol^{-1}$

**C.** $-30\ kJ\ mol^{-1}$

**D.** $+30\ kJ\ mol^{-1}$

**Question 11** (1 mark)

The activation energy for the reaction $2C(g) \rightarrow A(g) + B(g)$ is:

**A.** $30\ kJ\ mol^{-1}$

**B.** $60\ kJ\ mol^{-1}$

**C.** $90\ kJ\ mol^{-1}$

**D.** $180\ kJ\ mol^{-1}$

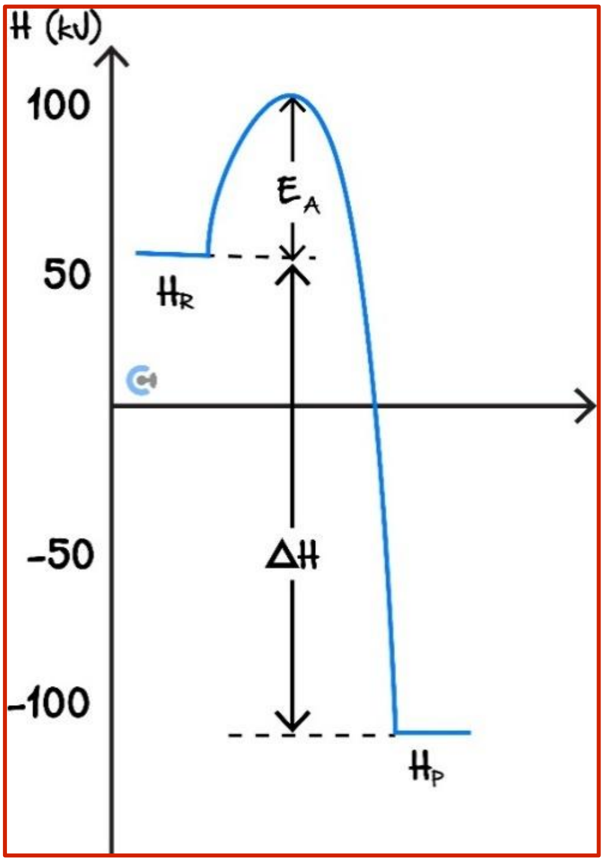
**Question 12** (3 marks)

Chlorine trifluoride, $ClF_3$, is used in the electronics industry to clean electronic circuit boards during their manufacture. It is produced by reacting chlorine and fluorine gases according to the following equation:

$$Cl_2(g) + 3F_2(g) \rightarrow 2ClF_3(g) \quad \Delta H = -160\ kJ\ mol^{-1}$$

The activation energy for the forward reaction is $50\ kJ\ mol^{-1}$.

**a.** Sketch an energy profile for the reaction on the set of axes below. Clearly mark the enthalpy of the products ($H_P$), the change in enthalpy ($\Delta H$) and the activation energy ($E_A$) of the reaction. (2 marks)



**b.** For the reaction $4ClF_3(g) \rightarrow 2Cl_2(g) + 6F_2(g)$, determine:

**i.** $\Delta H$. (0.5 marks)

$\boxed{+320\ kJ/mol}$

**ii.** $E_A$. (0.5 marks)

$\boxed{+400\ kJ/mol}$

Space for Personal Notes

## Section C: Ramping Up (11 Marks)

**INSTRUCTION:** 11 Marks. 8 Minutes Writing.

**Question 13** (3 marks)

The density of methanol is 0.791 $g/mL$. Find the volume of methanol required to release 67.3 $kJ$ of energy.

$d = \frac{m}{V} = 0.791\,g/mL$ $\Delta H = \frac{q}{m}$

$m(CH_3OH) = \frac{q}{\Delta H} = \frac{67.3\,kJ}{22.7\,kJ/g} = 2.96\,g$

$V(CH_3OH) = \frac{m}{d} = \frac{2.96\,g}{0.791\,g/mL} = 3.75\,mL$

**Question 14** (4 marks)

A sample of butane in a spirit burner which initially weighs 250.40 $g$ undergoes complete combustion. After the combustion is complete, it is found that the spirit burner weighs 248.25 $g$. The heat energy released is used to heat 200 $mL$ of water at SLC. The temperature of the water rises to 80.32°C.

Calculate the experimental molar heat of combustion of butane.

**Solutions:**

1. $m_{\text{butane}} = 250.40\,\text{g} - 248.25\,\text{g} = 2.15\,\text{g}$
2. $\Delta T = 80.32°\text{C} - 25.00°\text{C} = 55.32°\text{C}$
3. $q = 200 \cdot 4.18 \cdot 55.32 = 46,222.56\,\text{J}$
4. $q = \frac{46,222.56}{1000} = 46.22\,\text{kJ}$
5. $M_{\text{butane}} = 4(12.01) + 10(1.008) = 58.12\,\text{g/mol}$
6. $n = \frac{2.15}{58.12} = 0.0370\,\text{mol}$
7. $\Delta H_{\text{combustion}} = \frac{46.22}{0.0370} = 1,249.19\,\text{kJ/mol}$

**Final Answer:**

$$\Delta H_{\text{combustion}} = \boxed{1,249\,\text{kJ/mol}}$$

**Question 15** (4 marks)

A sample of propane is combusted in excess oxygen at standard conditions.

**a.** Write a balanced thermochemical equation for this reaction. (2 marks)

$$\boxed{C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)\ \Delta H = -2220\ kJ/mol}$$

**b.** Given that 5.70 $MJ$ of energy is released, how much propane must have been used, in $g$? (2 marks)

$$\boxed{\begin{array}{c} n = q/\Delta H = 5.70 \times 10^3/2220 = 2.57\ mol\ (1) \\ m = n \times M = 2.57 \times 44.0 = 113\ g\ (2) \end{array}}$$

Space for Personal Notes

## Section D: Getting Trickier I (9 Marks)

**INSTRUCTION:** 9 Marks. 7 Minutes Writing.

**Question 16** (4 marks)

Calculate the mass of ethanol that must be combusted to increase the temperature of 210 $mL$ of water from 10°C to 75°C, if exactly half of the heat released by this reaction is absorbed by the water.

The heat of combustion of ethanol is 1367 $kJ\ mol^{-1}$ at this temperature.

$q = m \times C \times \Delta T$

$q = 210 \times 4.18 \times 65$

$q = 57057$

$q = 57.057$ kJ

Now $\Delta H_{\text{Comb}} = \frac{q}{n}$

$n_{C_2H_5OH} = \frac{57.057}{1367}$

$n_{C_2H_5OH} = 0.04174$ mol

$\therefore m_{C_2H_5OH} = n_{C_2H_5OH} \times MM$

$\therefore m_{C_2H_5OH} = 0.04174 \times 46.068$

$\therefore m_{C_2H_5OH} = 1.923$ g

Now 50% of heat has been lost to surroundings.

$\therefore$ inital $m_{C_2H_5OH} = 2 \times 1.920$ g

$\therefore$ inital $m_{C_2H_5OH} = 3.846$ g

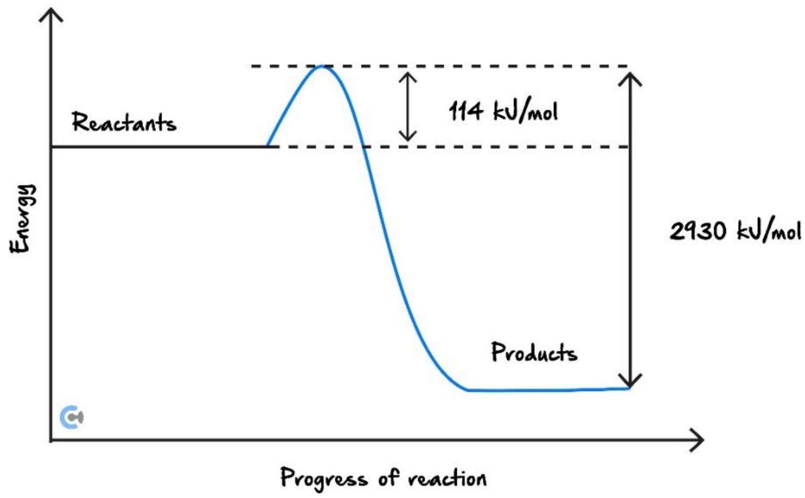
$\therefore$ inital $m_{C_2H_5OH} = 3.8$ g

**Space for Personal Notes**

**Question 17** (5 marks)

Glucose is a compound which is used in the body to store and obtain energy. Its molecular formula is $C_6H_{12}O_6$.

The following energy diagram shows the energy change associated with the combustion of glucose in $kJ\ mol^{-1}$ during cellular respiration.




**a.** Write a balanced thermochemical chemical equation for the combustion of glucose in an energy bar. (2 marks)

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \quad \Delta H = -2816\ kJ\ mol^{-1}$$

**b.** Calculate the energy released per gram of glucose. (1 mark)

$2816\ kJ$ per mole
$2816\ kJ$ per $180\ g$
Formula mass $=$ 180 glucose
Energy released per gram $= 2816//180 = 15.6\ kJ/g$
$= 15.6\ kJ/g$

**c.** An average cyclist uses 80 $kJ$ for every $km$ travelled. If the glucose content of an energy bar is 34.7 $g$, how far can an average cyclist travel on only one energy bar? (2 marks)

1 bar → 34.7 g of glucose

$E = \Delta H \times m = 34.7 \times 15.6$
$= 542.9\ kJ$

Distance $= \frac{542.9\ kJ}{80\ kJ/km}$
$= 6.79\ km$

## Section E: Getting Trickier II (8 Marks)

**INSTRUCTION:** 8 Marks. 8 Minutes Writing.

**Question 18** (4 marks)

A spirit burner containing ethanol is used to heat an insulated beaker containing 568 $g$ of water. The mass of ethanol that reacts is 1.80 $g$. A thermometer shows a temperature rise of the water from 17.4°C to 25.6°C.

Calculate the percentage of the energy transferred from the combustion of the ethanol burner to the heating of the water.

**Worked solution**

Energy transferred to water = 568 × 4.18 × 8.2 = 19 500 J = 19.5 kJ

Theoretical energy available from 1.80 g of ethanol = 1.8 × 29.6 = 53.3 kJ

Percentage efficiency of the burner is $\frac{19.5 \times 100}{53.3} = 36.5\%$

**Explanatory notes**

The energy transferred to the water can be calculated because the mass of water and the temperature change are known. Formula is $q = mc\Delta T$. The energy density figure from the data book can be used to calculate the theoretical amount of energy released by the fuel. The percentage efficiency can then be calculated from these two figures.

***Marking allocation: 3 marks***

- 1 mark for calculation of energy transfer to water
- 1 mark for theoretical energy amount
- 1 mark for efficiency calculation

Space for Personal Notes

**Question 19** (2 marks)

Find the mass of kerosene required to be completely combusted to release the same amount of energy as if 540 $mL$ of petrol were to be completely combusted.

$$q(petrol) = \Delta H \times V = 34 \frac{kJ}{mL} \times 540\ mL = 18360\ kJ$$
$$m(kerosene) = \frac{q}{\Delta H} = 496\ g$$

**Question 20** (2 marks)

Calculate the density of petrol, using Item 14 of the Data Book.

$$d = \frac{m}{V} = \frac{34}{45} = 0.756\ g/mL$$

*Let's take a BREAK!*

Space for Personal Notes

## Section F: VCAA-Level Questions I (12 Marks)

**INSTRUCTION:** 12 Marks. 0.5 Minutes Reading. 11 Minutes Writing.

**Question 21** (12 marks)

*Inspired from VCAA Chemistry 3/4 Exam 2016*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2016/2016chem-amd-w.pdf#page=38

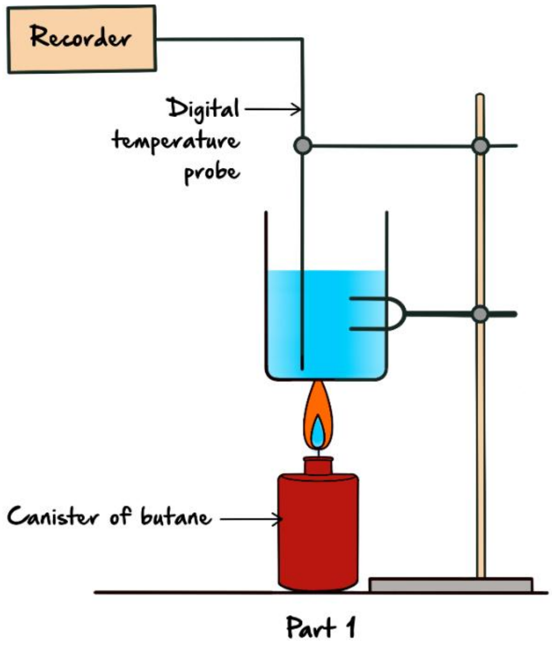
A senior Chemistry student bought a packet of Krispy Krackers from the local farmers' market. The label on the packet had no information on the energy content of the biscuits.

The student decided that he would measure the energy content of a Krispy Krackers biscuit by burning it under a can of water and measuring the temperature rise of the water.

Having performed a similar experiment in class, he knew that when the biscuit was burnt, heat energy would be lost to the environment. To increase the accuracy of the result, some butane gas from a butane canister was first burnt and the temperature rise of the water from that was measured. The heat energy released by burning butane was known, and the percentage energy loss using the equipment could be determined and adjusted to get the result for the biscuit.

The experimental setup and the results for Part 1 of the experiment are shown below:

**Part 1 – The heat content of butane**




1. Measure the initial mass of a butane canister.
2. Measure the mass of a metal can, add 250 $mL$ of water and re-weigh.
3. Set up the apparatus as in the diagram and measure the initial temperature of the water.
4. Burn the butane gas for five minutes.
5. Immediately measure the final temperature of the water.
6. Measure the final mass of the butane canister when cool.

**Results table for Part 1**

| Quantity | Measurement |
|---|---|
| Mass of empty can | 52.14 $g$ |
| Mass of can + water before combustion | 303.37 $g$ |
| Mass of butane canister before heating | 260.15 $g$ |
| Mass of butane canister after heating | 259.79 $g$ |
| Initial temperature of the water | 22.1°C |
| Final temperature of the water | 32.7°C |

**a.**

**i.** Write the balanced thermochemical equation for the complete combustion of butane. (2 marks)

| Marks | 0 | 1 | 2 | Average |
|---|---|---|---|---|
| % | 29 | 48 | 22 | 1 |

- $C_4H_{10}(g) + 6.5O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$, $\Delta H = -2874$ kJ mol$^{-1}$
- $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$, $\Delta H = -5748$ kJ mol$^{-1}$

**ii.** Calculate the amount of heat energy absorbed by the water when it is heated by burning the butane. Give your answer in kilojoules. (2 marks)

| Marks | 0 | 1 | 2 | Average |
|---|---|---|---|---|
| % | 35 | 15 | 50 | 1.2 |

$m$(water) in can = 303.37 – 52.14 = 251.23 g

$\Delta T$ of water = 32.7 – 22.1 = 10.6 °C

Energy = 4.18 J g$^{-1}$ °C$^{-1}$ × 251.23 g × 10.6 °C

= 1.113 × 10$^4$ J

= 11.1 kJ

**iii.** Calculate the experimental value of the molar heat of combustion of butane. Give your answer in $kJ\ mol^{-1}$. (2 marks)

| Marks | 0 | 1 | 2 | Average |
|---|---|---|---|---|
| % | 43 | 21 | 36 | 0.9 |

$m$(butane) reacted = 260.15 – 259.79

= 0.36 g

$n$(butane) reacted = 0.36/58.0

= 0.0062 mol

Energy released per mol $C_4H_{10}$ = 11.13 kJ/0.0062 mol

= 1.8 × 10$^3$ kJ

Molar heat of combustion = –1.8 × 10$^3$ kJ mol$^{-1}$

**iv.** Use the known enthalpy change for butane to calculate the percentage energy loss to the environment using the following relationship. (1 mark)

$$percentage\ energy\ loss\ = \frac{(theoretical\ value\ of\ \Delta H - experimental\ value\ of\ \Delta H)}{theoretical\ value\ of\ \Delta H} \times \frac{100}{1}$$

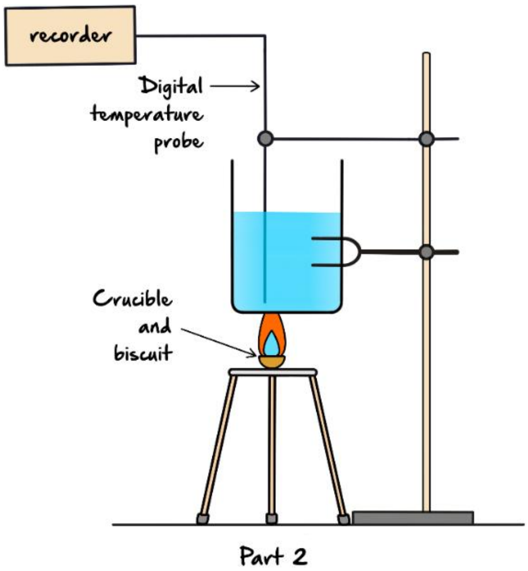
| Marks | 0 | 1 | Average |
|---|---|---|---|
| % | 52 | 48 | 0.5 |

% energy loss = [(2874 – 1.8 × 10$^3$)/2874] × 100

= 37.4%

The experimental setup and the results for Part 2 of the experiment are shown below:

**Part 2 – The heat content of a Krispy Kracker**




1. Measure the mass of a crucible, add a biscuit and re-weigh.
2. Set up the apparatus as in the diagram, using the same can of water as used in Part 1, and measure the initial temperature of the water.
3. Burn the biscuit until the flame runs out.
4. Immediately measure the final temperature of the water.
5. Measure the final mass of the crucible when cool.

**Results table for Part 2**

| **Quantity** | **Measurement** |
|---|---|
| Mass of crucible | $44.33\ g$ |
| Mass of biscuit + crucible before combustion | $46.75\ g$ |
| Mass of crucible after combustion | $44.34\ g$ |
| Mass of water (from Part 1) | $251.23\ g$ |
| Initial temperature of the water | 28.5°C |
| Final temperature of the water | 34.9°C |

**b.**

**i.** Calculate the energy content of Krispy Krackers using the data in the results table for Part 2. Give your answer in $kJ/100\ g$. (2 marks)

| Marks | 0 | 1 | 2 | Average |
|---|---|---|---|---|
| % | 52 | 17 | 30 | 0.8 |

$\Delta T$ water = 34.9 – 28.5 = 6.4 °C

Energy released into water = 4.18 J $g^{-1}$ °$C^{-1}$ × 251.3 g × 6.4 °C
= 6.7 × $10^3$ J
= 6.7 kJ

*m*(biscuit) reacted = 46.75 – 44.34
= 2.41 g

Energy per 100 g biscuit = (6.7/2.41) × 100
= 2.8 × $10^2$ kJ/100 g

**ii.** Explain why the energy content of a biscuit cannot be given in $kJ\ mol^{-1}$. (1 mark)

| Marks | 0 | 1 | Average |
|---|---|---|---|
| % | 67 | 33 | 0.4 |

A biscuit is not a pure substance/(is a mixture), so does not have a chemical formula or a molar mass.

**c.** Assume that the same percentage of heat energy loss occurred when burning the Krispy Kracker, as when the butane was burnt in Part 1. Calculate a more accurate value of the energy content of Krispy Krackers in $kJ/100\ g$. (2 marks)

| Marks | 0 | 1 | 2 | Average |
|---|---|---|---|---|
| % | 79 | 7 | 13 | 0.4 |

% energy transferred to water = 100 – 37.4 = 62.6%

2.8 × $10^2$ kJ = 62.6% of energy content

2.8 × $10^2$ = 0.626 × energy content

'Energy content' = 2.8 × $10^2$/0.626

= 4.5 × $10^2$ kJ/100 g

Space for Personal Notes

## Section G: Multiple Choice Questions (9 Marks)

**INSTRUCTION:** 9 Marks. 9 Minutes Writing.

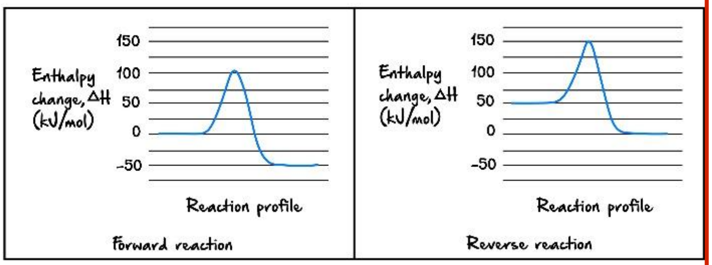
**Question 22** (1 mark)

*Inspired from VCAA Chemistry NHT Exam 2019*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/NHT/2019chem-nht-w.pdf#page=14

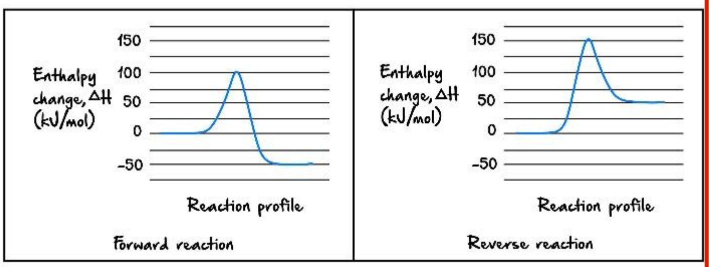
A reversible reaction has an enthalpy change, $\Delta H$, of $+50\ kJ\ mol^{-1}$ for the forward reaction.

Which one of the following pairs of energy profile diagrams, one for the forward reaction and one for the reverse reaction, represents this reaction?

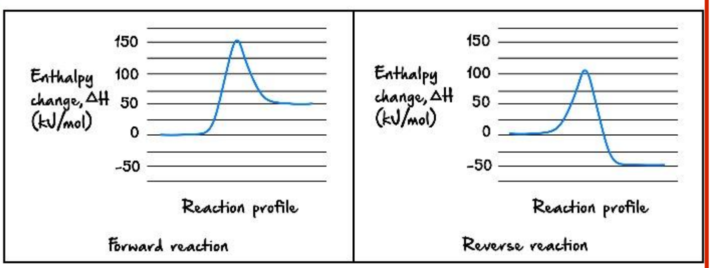
A



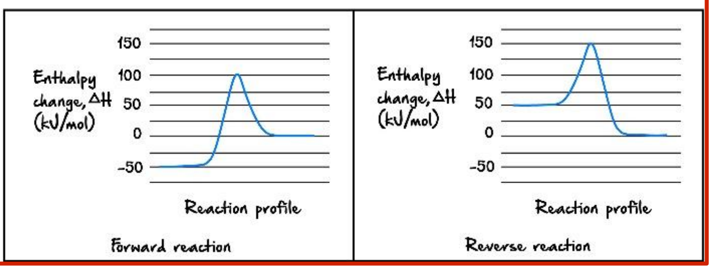

B




C




D




Forward reaction is endothermic, so reverse reaction is exothermic with delta $H = -50\ kJ/mol$, commensurate with option C.

**Question 23** (1 mark)

*Inspired from VCAA Chemistry NHT Exam 2019*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/NHT/2019chem-nht-w.pdf#page=10

0.50 $g$ of ethane, $C_2H_6$, undergoes complete combustion in a bomb calorimeter containing 200 $mL$ of water. The water temperature rises from 22.0°C to 48.5°C.

The thermochemical equation for the combustion of $C_2H_6$ using this information is:

**A.** $C_2H_6 + 5O_2 \rightarrow 2CO_2 + 3H_2O \quad \Delta H = -1330\ kJ\ mol^{-1}$

**B.** $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O \quad \Delta H = -3120\ kJ$

**C.** $C_2H_6 + 5O_2 \rightarrow 2CO_2 + 3H_2O \quad \Delta H = -1560\ kJ\ mol^{-1}$

**D.** $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O \quad \Delta H = -2660\ kJ$

**Question 24** (1 mark)

*Inspired from VCAA Chemistry Exam 2020*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2020/2020chem-w.pdf#page=13

The combustion of which fuel provides the most energy per unit of mass?

**A.** Butane ($M = 58\ g\ mol^{-1}$), which releases 49097 $MJ\ tonne^{-1}$.

**B.** Nitromethane ($M = 61\ g\ mol^{-1}$), which releases 11.63 $kJ\ g^{-1}$.

**C.** Butanol ($M = 74\ g\ mol^{-1}$), which releases 2670 $kJ\ mol^{-1}$.

**D.** Ethyne ($M = 26\ g\ mol^{-1}$), which releases 1300 $kJ\ mol^{-1}$.

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**Question 25** (1 mark)

*Inspired from VCAA Chemistry Exam 2020*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2020/2020chem-w.pdf#page=11

An experiment was carried out to determine the enthalpy of combustion of propan-1-ol. Combustion of 723 $mg$ of propan-1-ol increased the temperature of 150 $g$ of water from 22.1°C to 43.8°C.

The enthalpy of combustion is closest to:

**A.** $-17.8\ kJ/g$.

**B.** $-13.6\ kJ\ mol^{-1}$.

**C.** $-1129\ kJ\ mol^{-1}$.

**D.** $-1540\ kJ\ mol^{-1}$.

Energy added to water

$= 4.18\ J\ °C^{-1}\ K^{-1} \times m(H_2O) \times \Delta T$

$= 4.18 \times 150 \times (43.8 - 22.1)$

$= 13605.9\ J$

$= 13.6\ kJ$

$n(C_3H_8O) = 0.723\ g/60.0\ g\ mol^{-1} = 0.012\ mol$

Enthalpy of combustion

$= -(13.6059\ kJ/0.012)$

$= -1129\ kJ\ mol^{-1}$

**Question 26** (1 mark)

The complete combustion of 0.12 $mol$ of a particular fuel is found to require 0.42 $mol$ of oxygen gas.

The fuel is likely to be:

**A.** Methane.

**B.** Ethane.

**C.** Ethanol.

**D.** Butane.

Option A is incorrect, as each mole of methane would react with 2 moles of oxygen.

Option B is correct, as the ratio of fuel to $O_2$ is $0.12:0.42$ or $1:3:5$. Each fuel listed needs to be tested to see which balanced equation matches this ratio. Ethane provides a matching equation: $C_2H_6(g) + 3.5O_2(g) \rightarrow 2CO_2(g) + 3H_2O(I)$

Option C is incorrect, as each mole of ethanol would react with 3 moles of oxygen.

Option D is incorrect, as each mole of butane would react with 6.5 moles of oxygen.

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**Question 27** (1 mark)

*Inspired from VCAA Chemistry Exam 2021*
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/2021chem-w.pdf#page=6

Butane, $C_4H_{10}$, undergoes complete combustion according to the following equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

67.0 $g$ of $C_4H_{10}$ released 3330 $kJ$ of energy during complete combustion at standard laboratory conditions (SLC). The mass of carbon dioxide, $CO_2$, produced was:

**A.** 0.105 $g$

**B.** 3.18 $g$

**C.** 50.9 $g$

**D.** 204 $g$

$n(C_4H_{10})$ = 67.0 g / 58.0 g mol$^{-1}$
$n(CO_2)$ = 4 x 67.0 / 58.0
= 4.62 mol
$m(CO_2)$ = 4.62 mol x 44.0 g mol$^{-1}$
= 203 g
Alternatively:
$n(C_4H_{10})$ = 3330 kJ / 2880 kJ mol$^{-1}$
$n(CO_2)$ = 4 x 3330 / 2880
= 4.63 mol
$m(CO_2)$ = 4.63 mol x 44.0 g mol$^{-1}$
= 204 g
Alternative C is consistent with not using the $n(CO_2)/n(C_4H_{10})$ ratio indication in the combustion equation.

**Question 28** (1 mark)

Four fuels undergo complete combustion in excess oxygen, $O_2$, and the energy released is used to heat 500 $g$ of water.

Assuming that there is no energy lost to the environment, which one of the following fuels is likely to have increased the temperature of water from 32.0°C to 92.0°C?

**A.** 0.444 $g$ of hydrogen, $H_2$.

**B.** 1.96 $g$ of propane, $C_3H_8$.

**C.** 0.140 $mol$ of methane, $CH_4$.

**D.** 0.150 $mol$ of methanol, $CH_3OH$.

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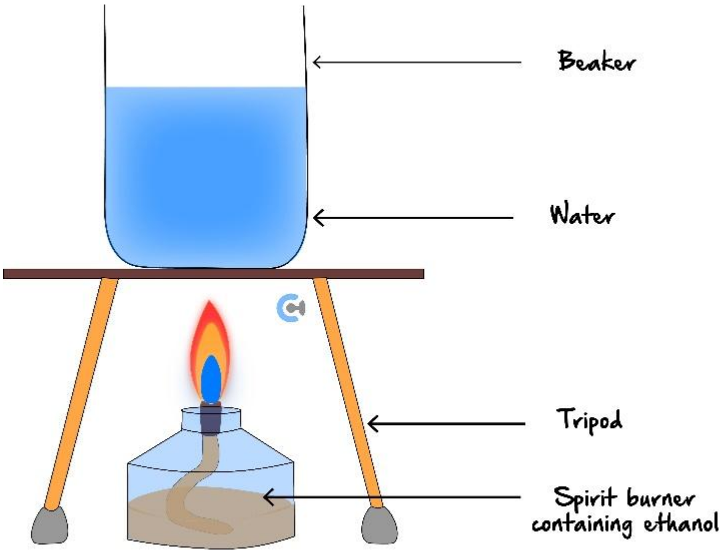
**Question 29** (1 mark)

A student set up the equipment shown in the following diagram.




The thermochemical equation for the combustion of ethanol is:

**A.** $C_2H_5OH(l) + 3.5O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -726\ kJ\ mol^{-1}$

**B.** $2C_2H_5OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \quad \Delta H = -2720\ kJ\ mol^{-1}$

**C.** $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1360\ kJ\ mol^{-1}$

**D.** $2C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1452\ kJ\ mol^{-1}$

**Question 30** (1 mark)

The student made the following experimental measuremen
Mass of water in beaker = 100.0 $g$
Increase in temperature = 10.00°C

What mass of ethanol ($M = 46\ g\ mol^{-1}$) must have unde ... re of the water? Assume that there is no heat loss.

**A.** $3.07 \times 10^{-3} g$

**B.** $1.41 \times 10^{-1} g$

**C.** $4.18\ g$

**D.** $29.5\ g$

$E_{absorbed} = mc\Delta T = 100.0 \times 4.18 \times 10.00$
$= 4180\ J$
$= 4.18\ kJ$

$n(\text{ethanol}) = \frac{E}{\Delta H} = \frac{4.18}{1360}$
$= 3.07 \times 10^{-3}\ mol$

$\Rightarrow m(\text{ethanol}) = n \times M = 3.07 \times 10^{-3} \times 46$
$= 0.141\ g$

## Section H: VCAA-Level Questions II (12 Marks)

**INSTRUCTION:** 12 Marks. 0.5 Minutes Reading. 11 Minutes Writing.

**Question 29** (12 marks)

Two fuels, ethanol and methane are being investigated.

**a.** It is found that when methane undergoes incomplete combustion to produce carbon monoxide, it releases 82.5% of the energy it would release per mole if complete combustion were to occur.

**i.** Explain why the heat of combustion for the incomplete combustion of methane is lower than the heat of combustion for the complete combustion of methane. (1 mark)

Partially oxidises instead of fully oxidising – less energy released.

**ii.** Write the thermochemical equations representing the incomplete combustion of methane to produce carbon monoxide as the only carbon-containing product. (2 marks)

$$\mathrm{CH_4(g)} + \frac{3}{2}\mathrm{O_2(g)} \rightarrow \mathrm{CO(g)} + 2\mathrm{H_2O(l)} \quad \Delta H = -734\ kJ/mol$$

**b.** Ethanol is then investigated, which is known to have a density of 0.790 $g/mL$. Some volume of ethanol is completely combusted and is used to heat 150 $mL$ of water from 20.0°C to 36.2°C.

Find the volume of ethanol which must have undergone complete combustion. (4 marks)

$$q = mc\Delta T = 150 \times 4.18 \times (36.2 - 20) = 10.16\ kJ$$
$$n(\mathrm{C_2H_5OH}) = \frac{q}{\Delta H} = 0.00741\ mol$$
$$m(\mathrm{C_2H_5OH}) = 0.341\ g$$
$$V(\mathrm{C_2H_5OH}) = 0.432\ mL$$

**c.** On a separate occasion, both ethanol and methane undergo complete combustion. Find the mass of methane required to release the same amount of energy as completely combusting 5.00 $L$ of ethanol does. (5 marks)

$$m(\mathrm{C_2H_5OH}) = dV = 0.790 \times 5000 = 3950\ g$$
$$n(\mathrm{C_2H_5OH}) = \frac{m}{Mr} = \frac{3950}{46} = 85.9\ mol$$
$$q(\mathrm{C_2H_5OH}) = \Delta H \times n = 117641\ kJ$$
$$n(\mathrm{CH_4}) = \frac{q}{\Delta H} = 132.2\ mol$$
$$m(\mathrm{CH_4}) = n \times Mr = 132.2 \times 16 = 2115\ g$$

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## Section I: Summary

*What have we learnt today?*

TIPS

Pitfalls

## Section J: Extension Questions (10 Marks)

**Question 31** (4 marks)

$E10$ fuel is a blend of unleaded petrol and ethanol, where the percentage volume/volume ($\%\ v/v$) of ethanol is 10%. Assume that the petrol component is entirely octane.

| **Fuel** | **Density** ($g\ mL^{-1}$) |
|---|---|
| Ethanol | 0.790 |
| Octane | 0.700 |

Consider a 3.00 $L$ sample of $E10$.

**a.** Calculate the volume and mass of each fuel in the sample. (2 marks)

Octane volume: $2.70\ L$ Octane mass: $1.89\ kg$

Ethanol volume: $0.30\ L$ Ethanol mass: $0.237\ kg$

**b.** Calculate the energy released from the complete combustion of the 3.00 $L$ sample. (2 marks)

**Worked solution**

Energy released by ethanol = 29.6 × 79 = 2338.4 kJ

Energy released by octane = 47.9 × 630 = 30 177 kJ

Total energy released = 2338.4 + 30 177 = 32 515.4 kJ = $3.25 \times 10^4$ kJ (to 3 sig. figures)

**Explanatory notes**

The amount of energy for each fuel is the mass of fuel multiplied by the amount of energy per gram of the fuel. The heat of combustion values are supplied in the data book.

***Mark allocation: 2 marks***

- 1 mark for the energy released from each fuel
- 1 mark for total energy provided

Multiply energy released by ethanol by 3; multiply energy released by octane by 3; and therefore, the total energy released by 3 to get $9.75 \times 10^4\ kJ$.

**Tip**

- *This question requires you to use the Chemistry data book. It is expected that you will know to do this even though the question does not direct you to the data book. You need to be familiar with the wealth of chemical information that can be found in this book.*

Sp

**Question 32** (6 marks)

Some students in a lab at standard conditions, are trying to experimentally determine the enthalpy of ethanol. Of course, the results they obtained differ from those expected from the literature.

**a.** Given the density of ethanol is 0.789 $g/mL$, and 8.00 $mL$ of ethanol raised the temperature of 250 $g$ of water by 22.7°C, calculate the percentage of heat lost to the environment. (4 marks)

$$m(\text{ethanol}) = 0.789 \times 8.00 = 6.312\ g\ (1)$$
$$E_{released} = \Delta H \times m = 29.6 \times 6.312 = 186.84\ kJ\ (2)$$
$$E_{absorbed} = mc\ \Delta T = 250 \times 4.18 \times 22.7 = 23.72\ kJ\ (3)$$
$$\%\text{energy loss} = (186.84 - 23.72)/186.84 \times 100 = 87.3\ \%\ (4)$$

**b.** Give a possible reason as to why not all of the heat was transferred to the water, and suggest a possible modification to the experiment to minimise this heat loss. (2 marks)

Some heat was absorbed by the apparatus instead of the water itself. (1)
Insulating the sides of the beaker (for example) would improve the accuracy of the experiment.

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