BENCHMARK

VCE Chemistry Assessments

Chemistry

Solutions

Benchmark Open SAC 3.1: Fuels compared

- Writing time is **55 minutes**

Approved materials

- One scientific calculator

Materials supplied

- Data Book

Instructions

- Answer all questions in the spaces provided.
- Write using blue or black pen.

Students are **not** permitted to bring mobile phones or other electronic devices into the room.

Contents pages

Questions (4 questions, 40 marks) 2–9

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Questions

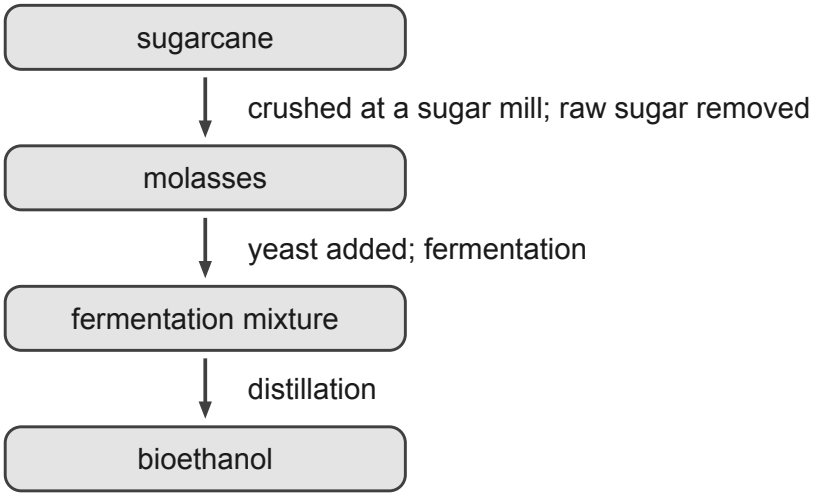
Instructions

- Answer **all** questions in the spaces provided.
- Write your responses in English.
- Give simplified answers to all numerical questions, with an appropriate number of significant figures; unsimplified answers will not be given full marks.
- Show all working in your answers to numerical questions; no marks will be given for an incorrect answer unless it is accompanied by details of the working.
- Ensure that chemical equations are balanced and that the formulas for individual substances include an indication of state, for example, $H_2(g)$, NaCl(s).
- Unless otherwise indicated, the diagrams in this book are **not** drawn to scale.

Question 1 (8 marks)

Most buses in Australian cities operate on diesel or natural gas, which are both fossil fuels. In Mackay, Queensland, two buses have been trialled on a fuel consisting mainly of ethanol. This fuel is made from sugarcane at a distillery in the nearby town of Sarina. A researcher is comparing this renewable fuel with alternative transport fuels.

The distillery manufactures its ethanol from molasses, a by-product of raw sugar production. The main steps in the production of the fuel are shown below.



a. Ethanol made from sugarcane is classified as a renewable fuel, but diesel is not.
Explain this difference. 2 marks

① Ethanol is made from sugarcane, which is grown again each season, so it is replaced by natural processes within a relatively short period of time.

② Diesel is made from crude oil, a fossil fuel that formed over millions of years, so it is not replaced within a relatively short period of time.

b. The sugars in molasses are converted into glucose before fermentation. Calculate the maximum mass, in kg, of ethanol that can be produced by the fermentation of 25.0 kg of glucose. 2 marks

① $n(C_6H_{12}O_6) = \frac{m}{M} = \frac{25.0 \times 10^3}{180.0} = 138.88\ldots$ mol

② $n(C_2H_5OH) = 2 \times n(C_6H_{12}O_6) = 2 \times 138.88\ldots = 277.77\ldots$ mol, so
$m(C_2H_5OH) = n \times M = 277.77\ldots \times 46.0 = 1.28 \times 10^4$ g = 12.8 kg (to 3 s.f.)

c. Fermentation stops when the mixture contains about 15% ethanol by volume, because the yeast cells die at higher ethanol concentrations. Explain why the fermentation mixture is then distilled. 2 marks

① Distillation separates the ethanol from the water, which makes up most of the mixture, so that the product can be used as a fuel.

② Ethanol has the lower boiling point, so it vaporises more readily and the condensed vapour is much richer in ethanol.

d. Write a balanced thermochemical equation for the complete combustion of ethanol. Include states. 2 marks

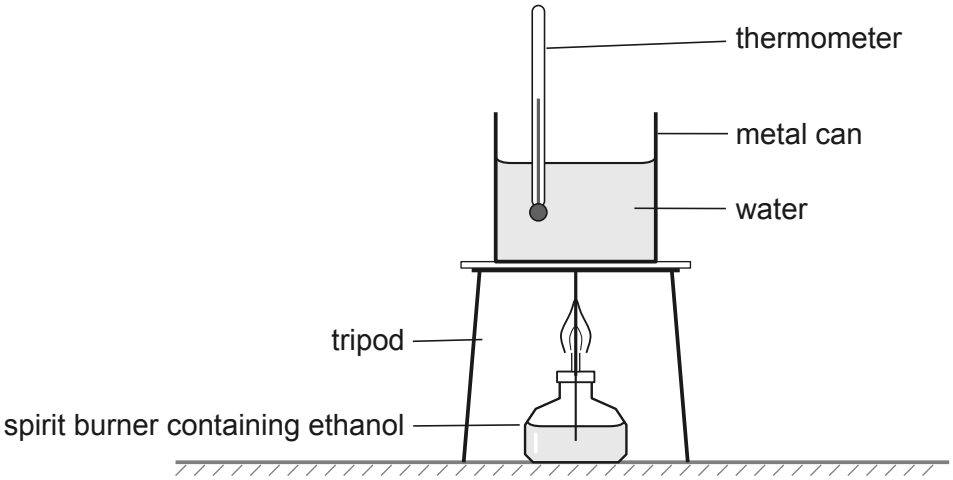
① $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$, balanced, with correct states

② $\Delta H = -1367$ kJ (accept -1367 kJ mol^{-1})

Question 2 (12 marks)

The researcher wants to measure the energy released when a sample of the fuel from the distillery is combusted. She places the fuel in a spirit burner and uses the flame to heat water in a metal can, as shown below.

She records the initial and final temperatures of the water, together with the initial and final masses of the spirit burner.



Her results are recorded in the table below. Assume the fuel is pure ethanol.

Measurement	Result
mass of water in the can (g)	200.0
initial temperature of the water (°C)	21.4
final temperature of the water (°C)	45.8
initial mass of the spirit burner (g)	152.36
final mass of the spirit burner (g)	150.96

a. Calculate the heat energy, in kJ, absorbed by the water. 2 marks

① $\Delta T = 45.8 - 21.4 = 24.4$ °C

② $q = mc\Delta T = 200.0 \times 4.18 \times 24.4 = 20\,398.4$ J = 20.4 kJ (to 3 s.f.)

b. Calculate the experimental molar heat of combustion of ethanol, in kJ mol^{-1}. 2 marks

① $m(C_2H_5OH) = 152.36 - 150.96 = 1.40$ g, so $n(C_2H_5OH) = \frac{m}{M} = \frac{1.40}{46.0} = 0.030434\ldots$ mol

② molar heat of combustion $= \frac{q}{n} = \frac{20.3984}{0.030434\ldots} = 670$ kJ mol^{-1} (to 3 s.f.; accept ΔH = –670 kJ mol^{-1}; consequential on part a.)

c. Calculate the energy transformation efficiency of this method, as the percentage of the chemical energy released by the fuel that is absorbed by the water. 2 marks

① efficiency $= \frac{\text{experimental molar heat of combustion}}{\text{Data Book molar heat of combustion}} \times 100$

② $= \frac{670.23\ldots}{1367} \times 100 = 49.0\%$ (accept 49.1%, from 29.7 kJ g^{-1}; consequential on part b.)

d. The researcher sees a black solid on the base of the can. Write a balanced equation for the incomplete combustion of ethanol that forms this solid and water. Include states. 2 marks

① $C_2H_5OH(l) + O_2(g) \rightarrow 2C(s) + 3H_2O(l)$, correct formulae, balanced

② correct states (accept $H_2O(g)$)

e. Explain, in terms of bond breaking and bond making, why the combustion of ethanol is exothermic. 2 marks

① Energy is absorbed to break the bonds in the reactants, ethanol and oxygen.

② More energy is released when the new bonds in the products, CO_2 and H_2O, form, so overall energy is released.

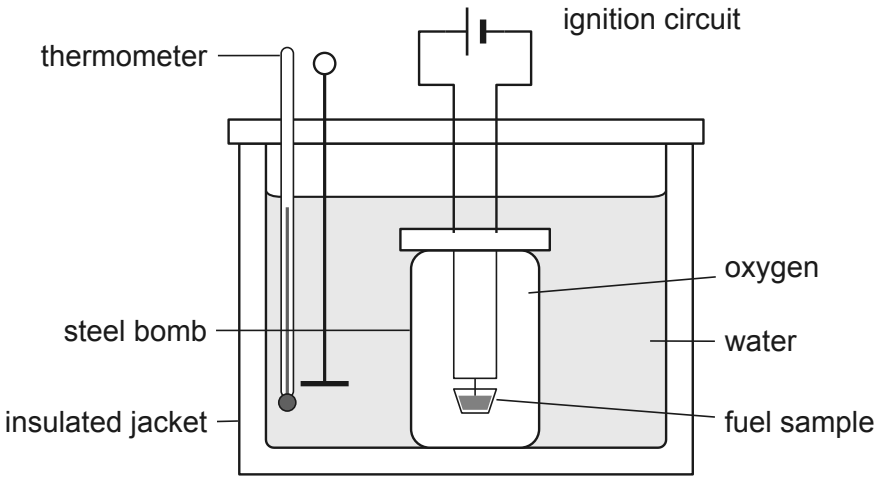
f. Other than incomplete combustion, identify two reasons why the efficiency calculated in part **c.** is so low. 2 marks

① Much of the heat released is lost to the surroundings, such as the air around the flame and the can.

② Some heat is absorbed by the metal can and the tripod rather than by the water (accept: some ethanol evaporates from the wick without being combusted).

Question 3 (9 marks)

Buses with diesel engines can also operate on biodiesel, which is manufactured from vegetable oils and animal fats. The researcher requires a more accurate experimental value for the energy released by combustion. She determines the heat of combustion of a sample of biodiesel in the bomb calorimeter shown below.



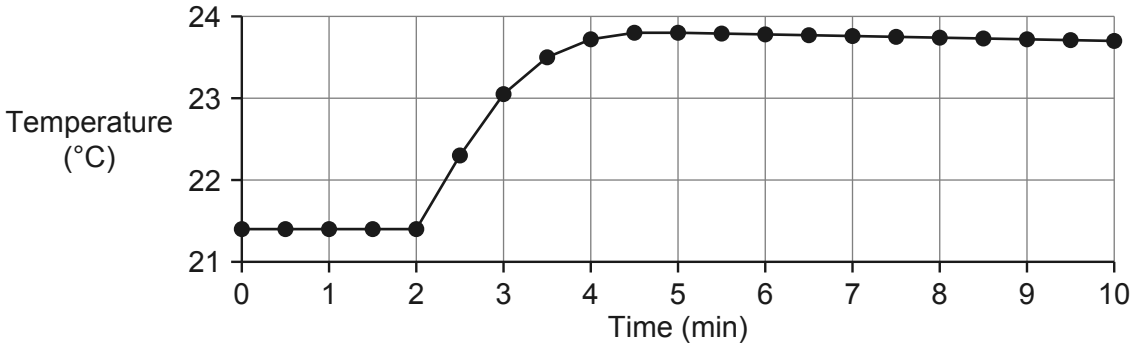
Before using the calorimeter, she must determine its calibration factor. She therefore combusts 1.012 g of benzoic acid. Benzoic acid is a standard that releases 26.43 kJ g^{-1} of energy. The temperature of the water rises by 2.708 °C.

a. Calculate the calibration factor of the calorimeter, in kJ $°C^{-1}$. 2 marks

① $E = m \times 26.43 = 1.012 \times 26.43 = 26.747\ldots$ kJ

② $CF = \frac{E}{\Delta T} = \frac{26.747\ldots}{2.708} = 9.88$ kJ $°C^{-1}$ (to 3 s.f.)

She then combusts 0.650 g of the biodiesel in the same calorimeter, which contains the same volume of water. The sample is ignited at 2.0 min. She records the temperature of the water every 0.5 min, and her readings are shown in the graph below.



b. Use the graph to determine the temperature change of the water. 1 mark

① $\Delta T = 23.80 - 21.40 = 2.40$ °C (accept 2.38 to 2.42 °C)

c. Calculate the heat of combustion of this fuel, in kJ g^{-1}. 2 marks

① $q = CF \times \Delta T = 9.8770\ldots \times 2.40 = 23.705\ldots$ kJ

② heat of combustion $= \frac{q}{m} = \frac{23.705\ldots}{0.650}$

= 36.5 kJ g^{-1} (to 3 s.f.; accept 36.2 to 36.8 kJ g^{-1}, or 36 kJ g^{-1}, consequential on parts a. and b.)

d. Explain why the bomb calorimeter gives a heat of combustion much closer to the accepted value than the method used in Question 2. 2 marks

① The fuel is combusted in excess oxygen inside a sealed bomb, so the combustion is complete.

② The bomb sits in water inside an insulated jacket and the calorimeter is calibrated, so little heat is lost to the surroundings and the heat absorbed by every part of the calorimeter is accounted for.

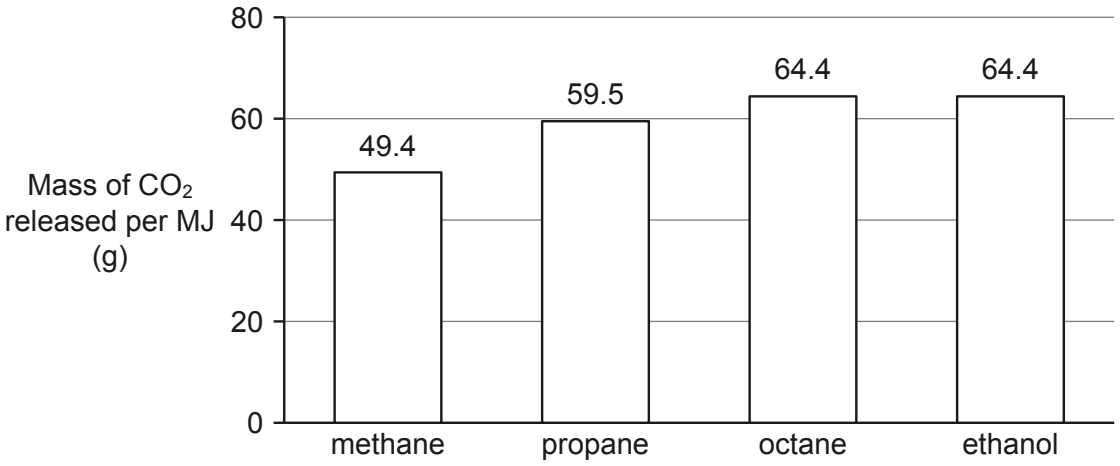
e. Suppose the calorimeter had contained 100 mL more water during the second trial than during the calibration. Explain the effect that this difference would have on the heat of combustion that the researcher calculated. 2 marks

① With more water, more energy is needed to raise the temperature by 1 °C, so the true calibration factor during the trial is larger than the value from part **a.**

② The same energy gives a smaller temperature rise, so $q = CF \times \Delta T$, using the smaller calibration factor, is too low: the calculated heat of combustion is lower than the true value.

Question 4 (11 marks)

The researcher then compares the quantity of carbon dioxide released by different transport fuels. For each fuel, she uses its molar enthalpy of combustion to calculate the mass of CO_2 released when combustion produces 1 MJ of energy. Her results are shown in the bar chart below.



Diesel is a mixture of hydrocarbons. It can be modelled as hexadecane, $C_{16}H_{34}(l)$, which has a molar enthalpy of combustion of –10 699 kJ mol^{-1}.

a. Calculate the mass of CO_2, in g, released per 1.00 MJ of energy when hexadecane undergoes complete combustion. 3 marks

① $n(C_{16}H_{34}) = \dfrac{\text{energy released}}{\text{energy released per mole}} = \dfrac{1.00 \times 10^3}{10\,699} = 0.093466\ldots$ mol

② $n(CO_2) = 16 \times n(C_{16}H_{34}) = 16 \times 0.093466\ldots = 1.4954\ldots$ mol

③ $m(CO_2) = n \times M = 1.4954\ldots \times 44.0 = 65.8$ g (to 3 s.f.)

b. Methane releases less CO_2 per MJ than any of the other fuels shown. Explain why, with reference to the composition of the fuels. 2 marks

① Methane has the highest ratio of hydrogen atoms to carbon atoms, 4 : 1.

② A larger share of its energy is released by forming water rather than carbon dioxide, so less CO_2 forms for each MJ released.

c. The bus manufacturer states that its research indicates these ethanol buses "could reduce carbon by up to 90 per cent compared to traditional fuel", meaning diesel. Evaluate this claim. In your answer, refer to the bar chart, your answer to part **a.** and the production of bioethanol. 3 marks

① Per MJ, the CO_2 from combusting ethanol (64.4 g) is almost the same as from diesel (65.8 g), so the reduction cannot come from the combustion itself.

② The CO_2 that this ethanol releases was recently absorbed from the atmosphere by sugarcane during photosynthesis, so its net emissions are much lower than those of fossil diesel.

③ Fossil fuels are still used to grow, transport, ferment and distil the crop, so the net reduction is less than 100%: the claim is reasonable only for net emissions, and "up to 90%" is a best case.

d. Some buses run on natural gas, which is mainly methane. In a test, 2.00 L of methane is ignited with 3.50 L of oxygen gas at SLC. The reaction continues until the limiting reactant is consumed, and combustion is complete. Calculate the energy released, in kJ. 3 marks

① $n(CH_4) = \frac{V}{V_m} = \frac{2.00}{24.8} = 0.080645\ldots$ mol and $n(O_2) = \frac{V}{V_m} = \frac{3.50}{24.8} = 0.14112\ldots$ mol

② $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$: $n(O_2$ needed$) = 2 \times 0.080645\ldots = 0.16129\ldots$ mol, more than is present, so oxygen gas is limiting

③ $n(CH_4$ reacting$) = \frac{0.14112\ldots}{2} = 0.070564\ldots$ mol, so energy $= n \times 890 = 0.070564\ldots \times 890$ $= 62.8$ kJ (to 3 s.f.)

End of questions

Additional writing space

Additional writing space

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