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VCE Chemistry ¾

Energy Calculations & Delta H [1.2]

Homework Solutions

Homework Outline:

Compulsory	Pg 2-Pg 11
Supplementary	Pg 12-Pg 19

Section A: Compulsory

Sub-Section [1.2.1]: Apply $\mathbf{q = Delta\ H \times n}$ to Energy Released

Question 1 (2 marks)

a. Find the amount of energy released during the complete combustion of 12.5 mol of methanol. (1 mark)

$$\boxed{q = \Delta H \times n = 726 \times 12.5 = 9075\ kJ}$$

b. Determine the amount of energy released by the complete combustion of 7.45 kg of ethanol (C_2H_5OH). (1 mark)

2a) $q = 7450g \times 29.6 = 220.52 kJ \approx 220 kJ$ (3.s.f)

Question 2 (2 marks)

Decane undergoes combustion according to the following thermochemical equation:

$$C_{10}H_{22}(l) + \frac{31}{2}O_2(g) \rightarrow 10CO_2(g) + 11H_2O(l);\ \Delta H = -4240\ kJ\ mol^{-1}$$

a. Determine the amount of energy released by 10.8 g of decane. (1 mark)

$n(C_{10}H_{22}) = \frac{10.8g}{((12\times10)+22)} = 0.07605\ mol$

$q(C_{10}H_{22}) = \Delta H \times n = 0.07605 mol \times 4240 = 322\ kJ$

b. 127 kJ of energy was released by the combustion of decane. Determine the amount in mol of decane combusted. (1 mark)

$n = \frac{q}{\Delta H} = \frac{127\ kJ}{4240 kJ/mol} = 0.02995\ mol$

$= 3.00 \times 10^{-2}\ mol$

Question 3 (2 marks)

a. Given that a combustion engine that uses ethanol produces 25.5 kJ, calculate the amount (in mol) of ethanol that would have been inputted to the engine, considering there is no energy loss. (1 mark)

$n(C_2H_5OH) = \frac{q}{\Delta H} = \frac{25.5 kJ}{1360 kJ/mol} = 0.01875 mol$

$= 0.0188 mol$

b. A particular sample of a mystery fuel was being examined. It was found that 0.88 mol of that mystery fuel was combusted to produce 1266 kJ of energy. Determine the molar heat of combustion of the mystery fuel. (1 mark)

$\Delta H = \frac{q}{n} = \frac{1266 kJ}{0.88 mol} = 1438.6 kJ/mol$

$= 1.4 \times 10^3 kJ/mol$

c. A mixture of 2.00 moles of methanol ($CH_3OHCH_3OHCH_3OH$) and 4.00 moles of propane ($C_3H_8C_3H_8C_3H_8$) is combusted. Calculate the total energy released in kilojoules.

c) please correct the formula of methanol (CH_3OH) and propane(C_3H_8) and the solution is

Energy released from methanol $= 2 \times -726 = -1452$ kJ

Energy released from propane $= 4 \times -2220 = -8880$ kJ

Total energy released $= -1452 + (-8880) = -10332$ kJ

Space for Personal Notes

Sub-Section [1.2.2]: Apply Delta H in kJ/mol, kJ/g and kJ/mL to Energy Calculations

Question 4 (2 marks)

a. 15.9 kJ of energy was released when a particular amount of propane was burnt in the presence of an unlimited supply of oxygen. Calculate the amount, in moles, of propane that was burnt.

$m = \frac{q}{\Delta H} = \frac{15.9\,kJ}{49.7\,kJ/g} = 0.320g$ (3 s.f)

b. Diesel is a fuel being investigated.

i. A sample of 1.00 kg of diesel is completely combusted. Find the energy in kilojoules released. (1 mark)

$$q = \Delta H \times m = 45\ kJ/g \times 1000\ g = 45000\ kJ$$

ii. A sample of 1.00 L of diesel is completely combusted. Find the energy in kilojoules released. (1 mark)

$$q = \Delta H \times V = 37\ kJ/mL \times 1000\ mL = 37000\ kJ$$

Question 5 (3 marks)

a. A sample of octane releases 156.0 kJ of energy. Calculate the mass of the octane that underwent combustion. (1 mark)

ai) $m = \frac{q}{\Delta H} = \frac{156.0\,kJ}{47.9\,kJ/g} = 3.260g$ (4 s.f)

b. A beaker of propan-1-ol underwent combustion. It was found a 14.29 g sample released 500 kJ of energy. Calculate the heat of combustion of propan-1-ol in kJ/g . (1 mark)

$$(i) \quad \Delta H = \frac{q}{m} = \frac{500 kJ}{14.29 g} = 35.0 kJ/g \quad (3.s.f)$$

c. A sample of natural gas is combusted, whereby 590 kJ of energy is released. Find the volume, in L, of natural gas combusted. (1 mark)

$$V = \frac{q}{\Delta H} = \frac{590}{0.0035} = 168571\ mL = 169\ L$$

Question 6 (3 marks)

a. Given that the density of propane is 0.257 g/mL, find the amount of energy released by the complete combustion of 10.0 L of propane (C_3H_8). (2 marks)

$$m(C_3H_8) = 0.257 g/mL \times 10000 mL = 2570 g$$
$$q(C_3H_8) = 2570 g \times 50.5 kJ/g = 1.30 \times 10^5 kJ$$

b. Calculate the volume of butan-1-ol required to release 1.77 kJ of energy, given that the density of butan-1-ol is 0.745 g/mL . (1 mark)

$$m = \frac{q}{\Delta H} = \frac{1.77 kJ}{37.8 kJ/g} = 0.046825 g$$
$$V = \frac{m}{d} = \frac{0.046825 g}{0.745 g/mL} = 0.06285 mL = 6.28 \times 10^{-2} mL$$

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Sub-Section [1.2.3]: Apply $q = mc\Delta T$ to Find Energy Absorbed

Question 5 (1 mark)

Calculate the amount of energy required to heat up 10.2 g of water by 5.0°C.

$q = mc\Delta T = 10.2 \times 4.18 \times 5 = 213J$

Question 6 (1 mark)

It was found that 1.15×10^4 J heated up a sample of water by 20.5°C. Determine the mass of water that must have been heated up.

$q = mc\Delta T \therefore m = \frac{q}{c\Delta T} = \frac{1.15 \times 10^4 \times 1000}{4.18 \times 20.5}$

$= 134g$

Question 7 (2 marks)

In an experiment, 200.0 g of water is heated from 22.0°C to 28.0°C. The water absorbs 5004.0 J of energy. Find the experimental specific heat capacity of water.

1. **Determine the specific heat capacity of water:**

   Given:

   - $Q = 5004.0\,\text{J}$
   - $m = 200.0\,\text{g}$
   - $\Delta T = 28.0°\text{C} - 22.0°\text{C} = 6.0\,\text{K}$

   Note: conversion to K is not necessary

   Specific heat capacity:

$$c = \frac{Q}{m \times \Delta T} = \frac{5004.0\,\text{J}}{200.0\,\text{g} \times 6.0\,\text{K}} = \boxed{4.17\ J/g\ K}$$

Space for

Sub-Section [1.2.4]: Calculate Delta H Experimentally

Question 8 (3 marks)

A scientist is exploring the capabilities of biodiesels as a potential replacement for fossil fuels in the future. She uses a sample of methyl stearate ($C_{19}H_{38}O_2$) which is placed in a spirit burner and used to heat up a metal can with 250 mL of water.

The mass of the spirit burner across the experiment was monitored and the following values were found.

The initial mass of the spirit burner: 288.05 g
The final mass of the spirit burner: 279.51 g

a. Calculate the amount of methyl stearate (mol) that underwent combustion. (1 mark)

$m(C_8H_{18}) = 0.703 g/mL \times 20000 mL$
$= 14060 g$
$n(C_8H_{18}) = \frac{14060}{12\times8+18} = 123.33 mol$
$q(C_8H_{18}) = \Delta H \times n = 5470 \times 123.33 mol = 674633.33 kJ$
$= 6.75 \times 10^5 kJ$

b. The temperature of the water was found to increase from 25.0°C to 37.5°C. Find the amount of energy absorbed by the water in kJ. (1 mark)

$q = 1.5 kJ$
$\Delta H = \frac{q}{n} \Rightarrow n = \frac{q}{\Delta H}$
$n(C_8H_{18}) = \frac{1.5}{5470} = 0.00027 mol$
$m(C_8H_{18}) = 0.00027 \times (12\times8+18) = 0.03126 g$
$d = \frac{m}{v} \Rightarrow v = \frac{m}{d}$
$v(C_8H_{18}) = \frac{0.03126 g}{0.703 g/mL} = 0.044 mL$

c. Hence or otherwise, determine the heat of combustion in $kJ\ mol^{-1}$ of methyl stearate. (1 mark)

$\Delta H = \frac{q}{n} = \frac{13.0 kJ}{0.0287 mol} = 452.96 kJ/mol$
$= 453 kJ/mol$

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Question 9

A sample of 1.250 g of octane (C_8H_{18}) undergoes complete combustion in a spirit burner. The evolved energy is used to heat up an aluminium can containing 250 mL of water. The temperature of the water rises from 15.7°C to 22.2°C. Calculate the heat of combustion of octane in kJ/mol.

$q = mc\Delta T = 250 \times 4.18 \times (22.2 - 15.7)$
$= 6792.5\text{ J}$
$= 6.79\text{ kJ}$

$n(C_8H_{18}) = \frac{1.250\text{ g}}{(12\times8)+(18\times1)} = 0.010965\text{ mol}$

$\Delta H(C_8H_{18}) = \frac{q}{n} = \frac{6.79\text{ kJ}}{0.010965\text{ mol}} = 622.935\text{ kJ/mol}$
$= 623\text{ kJ/mol}$

Question 10 (4 marks)

A sample of 2.00 g of propane was completely burnt, and the energy released was used to heat a volume of water. The temperature of the water increases from 28.0°C to 43.0°C.

a. Calculate the energy released during the combustion of propane, in kilojoules. (2 marks)

The energy released is calculated as:

$$\text{Energy released} = \text{Moles of propane} \times \Delta H_{\text{combustion}}$$

$$\text{Moles of propane} = \frac{\text{Mass of propane}}{\text{Molar mass of propane}} = \frac{2.00\text{ g}}{44.1\text{ g/mol}} = 0.04535\text{ mol}$$

$$\text{Energy released} = 0.04535\text{ mol} \times (-2220\text{ kJ/mol}) = -100.68\text{ kJ}$$

b. Find the volume of water, in mL, which was heated up, assuming 100% energy transfer. (2 marks)

The energy absorbed by the water is given by:

$$q = mc\Delta T$$

Rearranging to find the mass of water:

$$m = \frac{q}{c\Delta T}$$

Where:

$\Delta T = 35.0°\mathrm{C} - 20.0°\mathrm{C} = 15.0°\mathrm{C}$

$c = 4.18\,\mathrm{J/g^{\circ}C}$

Convert q to joules:

$$q = -100.68\,\mathrm{kJ} = -100680\,\mathrm{J}$$

$$m = \frac{100680\,\mathrm{J}}{4.18\,\mathrm{J/g^{\circ}C} \times 15.0°\mathrm{C}} = 1605.75\,\mathrm{g}$$

The density of water is $1.00\,\mathrm{g/mL}$, so the volume is equal to the mass:

$$\text{Volume of water} = 1605.75\,\mathrm{mL}$$

Sub-Section: The 'Final Boss'

Question 11 (10 marks)

An experiment was conducted to investigate the energy changes during the combustion of ethanol in a spirit burner.

a. Write the thermochemical equation for the complete combustion of ethanol. (2 marks)

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1370\,\text{kJ/mol}$$

b. Given that the density of ethanol is 0.789 g/mL, find the heat of combustion of ethanol in kJ/mL. (2 marks)

$m_{ethanol} = d \times V$

$= 0.789\,g/mL \times 1\,mL$

$= 0.789\,g$

$n_{ethanol} = \frac{0.789\,g}{46.0\,g/mol}$

$= 0.01715\,mol$

$\therefore$ 1 mL ethanol $= 0.01715$ mol ethanol $\therefore$ 0.01715 mol/mL

$\therefore$ Heat of Combustion (kJ/mL) $= \Delta H \times \frac{n(\text{ethanol})}{mL}$

$= \left(\frac{-1370\,kJ}{mol}\right)\left(\frac{0.01715\,mol}{mL}\right)$

$= -23.48\,kJ/mL$

c. During the experiment, 12.35 g of ethanol in a spirit burner is combusted and used to heat 100.0 mL of water at SLC. After the experiment is completed, the final temperature of the water is 47.8°C. Find the experimental molar heat of combustion of ethanol. (4 marks)

1. Energy absorbed by water:

$$\text{Mass of water} = \text{Volume} \times \text{Density} = 100.0\,\text{mL} \times 1.00\,\text{g/mL} = 100.0\,\text{g}$$

$$\Delta T = \text{Final temperature} - \text{Initial temperature} = 47.8°\text{C} - 25.0°\text{C} = 22.8°\text{C}$$

$$q = mc\Delta T = 100.0\,\text{g} \times 4.18\,\text{J/g}°\text{C} \times 22.8°\text{C} = 9530.4\,\text{J}$$

Convert to kJ:

$$q_{\text{water}} = \frac{9530.4\,\text{J}}{1000} = 9.53\,\text{kJ}$$

2. Moles of ethanol burned:

$$n = \frac{\text{Mass burned}}{\text{Molar mass}} = \frac{12.35\,\text{g}}{46.0\,\text{g/mol}} = 0.2685\,\text{mol}$$

3. Experimental molar heat of combustion:

$$\Delta H_{\text{exp}} = \frac{q_{\text{water}}}{n} = \frac{9.53\,\text{kJ}}{0.2685\,\text{mol}} = -35.5\,\text{kJ/mol}$$

Answer: The experimental molar heat of combustion of ethanol is $-35.5\,\text{kJ/mol}$.

d. Calculate the energy transferred to the water per mL using the total energy absorbed and the volume of water. (2 marks)

1. Total energy absorbed by the water:

$$q_{\text{water}} = 9.53\,\text{kJ}$$

2. Volume of water:

$$V_{\text{water}} = 100.0\,\text{mL}$$

3. Energy transferred per mL:

$$\text{Energy per mL} = \frac{q_{\text{water}}}{V_{\text{water}}} = \frac{9.53\,\text{kJ}}{100.0\,\text{mL}} = 0.0953\,\text{kJ/mL}$$

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Section B: Supplementary

Sub-Section [1.2.1]: Apply $q = \text{Delta H} \times n$ to Energy Released

Question 12 (2 marks)

a. How much energy (in kilojoules) is released when 3.50 moles of hydrogen are combusted completely? (1 mark)

$$\text{Energy released} = 3.50\,\text{moles} \times (-286\,\text{kJ/mol}) = -1001.0\,\text{kJ}$$

b. A sample contains 1.25 moles of carbon (graphite). How much energy (in kilojoules) is released when the carbon is combusted? (1 mark)

$$\boxed{493\ kJ}$$

Question 13 (2 marks)

Calculate the energy released when 10.0 grams of glucose ($C_6H_{12}O_6$) are combusted.

1. **Moles of glucose:**

$$\text{Moles of glucose} = \frac{\text{Mass of glucose}}{\text{Molar mass of glucose}} = \frac{10.0\,\text{g}}{180.0\,\text{g/mol}} = 0.05556\,\text{mol}$$

2. **Energy released:**

$$\text{Energy released} = \text{Moles of glucose} \times \Delta H_{\text{combustion}} = 0.05556\,\text{mol} \times (-2840\,\text{kJ/mol}) = -157.78\,\text{kJ}$$

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Question 14 (3 marks)

A sample contains 12.0 g of butane and 8.0 g of hydrogen gas. Calculate the total energy released when both are combusted.

1. **Moles of butane:**

$$\text{Moles of butane} = \frac{\text{Mass of butane}}{\text{Molar mass of butane}} = \frac{12.0\,\text{g}}{58.0\,\text{g/mol}} = 0.2069\,\text{mol}$$

2. **Energy released by butane:**

$$\text{Energy released (butane)} = \text{Moles of butane} \times \Delta H_{\text{combustion}} = 0.2069\,\text{mol} \times (-2880\,\text{kJ/mol}) = -595.86\,\text{kJ}$$

3. **Moles of hydrogen:**

$$\text{Moles of hydrogen} = \frac{\text{Mass of hydrogen}}{\text{Molar mass of hydrogen}} = \frac{8.0\,\text{g}}{2.0\,\text{g/mol}} = 4.0\,\text{mol}$$

4. **Energy released by hydrogen:**

$$\text{Energy released (hydrogen)} = \text{Moles of hydrogen} \times \Delta H_{\text{combustion}} = 4.0\,\text{mol} \times (-286\,\text{kJ/mol}) = -1144.0\,\text{kJ}$$

5. **Total energy released:**

$$\text{Total energy released} = \text{Energy released (butane)} + \text{Energy released (hydrogen)} = -595.86\,\text{kJ} + (-1144.0\,\text{kJ}) = -1739.86\,\text{kJ}$$

Question 15 (4 marks)

A 50/50 by-mass fuel mixture of ethanol and methane is combusted completely. If the total mass of the mixture is 20.0 g, calculate the total energy released.

1. **Mass of each fuel:**
Since the mixture is 50/50 by mass, each fuel contributes half the total mass:

$$\text{Mass of ethanol} = \text{Mass of methane} = \frac{20.0\,\text{g}}{2} = 10.0\,\text{g}$$

2. **Moles of ethanol:**

$$\text{Moles of ethanol} = \frac{\text{Mass of ethanol}}{\text{Molar mass of ethanol}} = \frac{10.0\,\text{g}}{46.0\,\text{g/mol}} = 0.2174\,\text{mol}$$

3. **Energy released by ethanol:**

$$\text{Energy released (ethanol)} = \text{Moles of ethanol} \times \Delta H_{\text{combustion (ethanol)}} = 0.2174\,\text{mol} \times (-1370\,\text{kJ/mol}) = -297.83\,\text{kJ}$$

4. **Moles of methane:**

$$\text{Moles of methane} = \frac{\text{Mass of methane}}{\text{Molar mass of methane}} = \frac{10.0\,\text{g}}{16.0\,\text{g/mol}} = 0.625\,\text{mol}$$

5. **Energy released by methane:**

$$\text{Energy released (methane)} = \text{Moles of methane} \times \Delta H_{\text{combustion (methane)}} = 0.625\,\text{mol} \times (-890\,\text{kJ/mol}) = -556.25\,\text{kJ}$$

6. **Total energy released:**

$$\text{Total energy released} = \text{Energy released (ethanol)} + \text{Energy released (methane)} = -297.83\,\text{kJ} + (-556.25\,\text{kJ}) = -854.08\,\text{kJ}$$

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Sub-Section [1.2.2]: Apply Delta H in kJ/mol, kJ/g and kJ/mL to Energy Calculations

Question 16 (1 mark)

A generator burns 15.0 mL of petrol to produce energy. Calculate the total energy released during this process.

$\boxed{510\ kJ}$

Question 7 (4 marks)

a. The density of butan-1-ol is 0.745 g/mL, calculate the amount of energy that will be released by 1.50 L of butan-1-ol given its molar heat of combustion is 37.8 kJ/g. (2 marks)

$m(C_4H_9OH) = 1500 mL \times 0.745 g/mL = 1117.5 g$

$q(C_4H_9OH) = 37.8 kJ/g \times 1117.5 g = 4.22 \times 10^4 kJ$

b. A mixture of 3.00 g of ethane and 2.00 g of carbon is combusted. Calculate the total energy released in kJ. (2 marks)

4. **Total Energy Released by Ethane and Carbon (Intermediate Question 4):**
   - Energy released by ethane: $-156.0\,\mathrm{kJ}$
   - Energy released by carbon: $-65.67\,\mathrm{kJ}$
   - **Total energy released:**

$$-221.67\,\mathrm{kJ}$$

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Question 8 (3 marks)

The density of octane is 0.703 g/mL under SLC.

a. Calculate the amount of energy released by a canister containing 20.0 L of octane when it undergoes complete combustion. (1 mark)

$\boxed{\begin{array}{l} m(C_8H_{18}) = d \times V = 0.703\ g/mL \times 20000\ mL = 14060\ g \\ q(C_8H_{18}) = \Delta H \times m = 47.9\ kJ/g \times 14060\ g = 6.73 \times 10^5\ kJ \end{array}}$

b. Determine the volume of octane required to release 1.5 kJ of energy. (2 marks)

$\boxed{\begin{array}{l} n(C_8H_{18}) = \frac{q}{\Delta H} = \frac{1.5\ kJ}{5460\ kJ/mol} = 2.74 \times 10^{-4}\ mol \\ V(C_8H_{18}) = n \times V_m = 2.74 \times 10^{-4}\ mol \times 24.8 = 6.81 \times 10^{-3}\ L \end{array}}$

Question 17 (3 marks)

A ship uses 50.0 kg of kerosene during a trip.

a. Calculate the total energy released from combusting the kerosene in megajoules. (1 mark)

$\boxed{q(H_2) = \Delta H \times m = 46 \times 50000 = 23000\ MJ}$

b. Calculate the volume, in litres, of kerosene used during the trip. (2 marks)

$\boxed{\begin{array}{l} mL(kerosene) = \frac{q\ kJ}{\Delta H\ kJ/mL} = \frac{2300000\ kJ}{37\ kJ/mL} \\ = 62162.1621\ mL \\ = 62.2\ L \end{array}}$

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Sub-Section [1.2.3]: Apply $q = mc\Delta T$ to Find Energy Absorbed

Question 18 (1 mark)

Calculate the amount of heat energy, in kJ, required to heat up a kettle containing 1.5 L of water from 25°C to 100°C.

$q = mc\Delta T = 1500 \times 0.997 \times 4.18 \times (100 - 25)$
$= 468839.25\ J$
$\approx 468.8\ kJ$
$= 4.7 \times 10^2\ kJ$

Question 19 (1 mark)

If 21.7 kJ of energy was inputted to a pot of water to increase the temperature from 14.0°C to 44.4°C, find the volume of water which was present.

$m = \frac{q}{c\Delta T} = \frac{21700}{4.18 \times (44.4 - 14)} = 170.769\ g$
$= 171\ g$

Question 20 (3 marks)

A sample of water with a volume of 500 mL at SLC absorbs 20.9 kJ of energy. Find the final temperature of the water.

$$q = mc\Delta T$$

$$\Delta T = \frac{q}{mc}$$

$$\Delta T = \frac{20,900}{500 \times 4.18}$$

$$\Delta T = \frac{20,900}{2090} = 10.0^\circ\text{C}$$

$$T_{\text{final}} = T_{\text{initial}} + \Delta T$$

$$T_{\text{final}} = 25 + 10 = 35.0^\circ\text{C}$$

Sub-Section [1.2.4]: Calculate Delta H Experimentally

Question 21 (3 marks)

A particular spirit burner contained a sample of an unknown fuel. It is used to heat up a beaker containing 150 mL of water. The following data was obtained:

- The initial mass of the spirit burner: 100 mL
- The final mass of the spirit burner: 98.5 mL
- The initial temperature of water: 25°C
- The final temperature of water: 37°C

a. Calculate the volume of the unknown fuel that underwent combustion. (1 mark)

$100 - 98.5 = 1.5 mL$

b. Determine the amount of energy absorbed by the water, in kJ. (1 mark)

$q = mc\Delta T = (150 \times 4.18 \times (37-25))$

$= 7524 J$

$= 7.52 kJ$

c. Calculate the heat of combustion of the unknown fuel in $kJ\ mol^{-1}$. (1 mark)

$n = \frac{V}{V_m} = \frac{0.0015}{24.8} = 6.05 \times 10^{-5} mol$

$\Delta H = \frac{q}{n} = \frac{7.52 kJ}{6.05 \times 10^{-5}} = 124330 kJ/mol$

$= 1.24 \times 10^5 kJ/mol$

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Question 22 (4 marks)

A burner containing ethanol (C_2H_5OH) is used to heat 200.0 g of water. The water temperature rises from 20.0°C to 45.0°C. The initial mass of the burner and ethanol is 120.50 g, and the final mass is 119.75 g. Calculate the experimental heat of combustion of ethanol in kJ/mol.

$$q = mc\Delta T$$

$$q = 200.0 \times 4.18 \times 25.0 = 20,900 \text{ J}$$

$$\text{Mass burned} = 120.50 - 119.75 = 0.75 \text{ g}$$

$$\Delta H = \frac{q}{\text{mass burned}} \times \text{molar mass}$$

$$\Delta H = \frac{20,900}{0.75} \times 46.07$$

$$\Delta H = 1283.82 \text{ kJ/mol}$$

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Question 23 (5 marks)

A student uses ethanol (C_2H_5OH) to heat 300.0 g of water. The water's temperature increases from 22.0°C to 52.0°C. The initial volume of ethanol is 25.0 mL, and the final volume is 22.5 mL. The density of ethanol is 0.789 g/mL.

Calculate the experimental heat of combustion of ethanol in kJ/mol.

Step-by-Step Calculation:

Step 1: Calculate the heat absorbed by the water

$$q = mc\Delta T$$

$$q = 300.0 \times 4.18 \times 30.0 = 37,620\,\text{J}$$

Step 2: Calculate the volume of ethanol burned

$$\text{Volume burned} = 25.0 - 22.5 = 2.5\,\text{mL}$$

Step 3: Convert the volume of ethanol burned to mass

$$\text{Mass burned} = \text{Volume burned} \times \text{Density}$$

$$\text{Mass burned} = 2.5 \times 0.789 = 1.9725\,\text{g}$$

Step 4: Calculate the experimental heat of combustion

$$\Delta H = \frac{q}{\text{mass burned}} \times \text{molar mass}$$

$$\Delta H = \frac{37,620}{1.9725} \times 46.07$$

$$\Delta H = 878.66\,\text{kJ/mol}$$

Space

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