

## Introduction:

Energy has always intrigued me and it is something we constantly rely on. From the power that drives our vehicles to the heat that fuels our daily cooking energy is all around us. As our world faces increasingly severe consequences from the overuse of fossil fuels, the need to find alternative energy sources has become urgent. Throughout my IBDP journey, there has been a strong emphasis on the role we play in addressing global issues particularly climate change. However in class the opportunities to explore the real world applications of energy are limited. For this reason, I wanted to use this IA to investigate a practical aspect of energy specifically, the efficiency and energy output of alternative fuels like alcohols.

Recently, the impact of global warming has become more personal and pressing for me, especially living in Singapore, where rising temperatures and extreme weather patterns are becoming more noticeable each year. In recent years Singapore has experienced record breaking heatwaves, with 2023 marking one of the hottest years on record. These high temperatures have not only made daily life more uncomfortable but have also raised concerns about urban heat islands, increasing energy consumption for cooling, and the strain on public health and infrastructure. These visible effects of climate change have made me more aware of how our choices of energy sources play a role in shaping the planet's future. This growing awareness motivated me to investigate how different fuels such as alcohols can be compared in terms of energy output and environmental impact and thus chose to focus my investigation on alcohol combustion. Alcohols are considered relatively clean burning fuels and are among the most renewable options available. With various types of alcohols used across industries and transportation I was curious to compare their enthalpies of combustion to determine which type provides the most energy efficiently.

**Research Question:** What is the effect of increasing the number of carbon atoms in an alcohol on its enthalpy of combustion measured by the change in temperature of 100ml of water while combusting methanol, ethanol, propanol and butanol?

## Background Information:

Alcohols are organic compounds with one or more hydroxyl (-OH) functional groups. Alcohols can be classified as primary, secondary and tertiary depending on the position of the hydroxyl group. The hydroxyl group is attached to a carbon with at least two hydrogen atoms in a primary alcohol, to a carbon with only one hydrogen atom in a secondary alcohol and to a carbon with no hydrogen atoms in a tertiary alcohol.

The general formula for an alcohol is  $C_nH_{2n+1}OH$

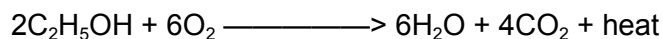
In a primary alcohol, the carbon which carries the hydroxyl group is only attached to one alkyl group (An alkyl group is a group such as methyl,  $CH_3$ , or ethyl,  $CH_3CH_2$ . These are groups containing chains of carbon atoms which may be branched. Alkyl groups are given the general symbol R). In each case there is only one linkage to an alkyl group from the  $CH_2$  group holding the -OH group. There is an exception to this. Methanol,  $CH_3OH$ , is counted as a primary alcohol even though there are no alkyl groups attached to the carbon with the -OH group on it.

Examples:

Name	Molecular formula	Structural formula
methanol	CH <sub>3</sub> OH	<pre>       H         H — C — O — H               H           </pre>
ethanol	C <sub>2</sub> H <sub>5</sub> OH	<pre>       H   H             H — C — C — O — H                   H   H           </pre>
propanol	C <sub>3</sub> H <sub>7</sub> OH	<pre>       H   H   H                 H — C — C — C — O — H                       H   H   H           </pre>
butanol	C <sub>4</sub> H <sub>9</sub> OH	<pre>       H   H   H   H                     H — C — C — C — C — O — H                           H   H   H   H           </pre>

Alcohols are highly flammable and can undergo combustion. Combustion is a reaction in which a substance reacts with oxygen to give carbon dioxide (CO<sub>2</sub>), water (H<sub>2</sub>O) and heat energy. All combustion reactions are exothermic (If more energy is released when the bonds are made than is required to break bonds, energy will be released into the environment).

Example:



The chemical energy involved in reaction is called enthalpy change. For a given substance the enthalpy can be thought of as a kind of energy store that provides or accepts in the form of heat. The symbol for enthalpy of a substance is written as H. For a chemical reaction the heat released or absorbed is a measure of the difference in energy between reactants and products (under the same conditions of temperature and pressure).

$$\Delta H = H(\text{products}) - H(\text{reactants})$$

Since an enthalpy change represents the heat transferred in a chemical reaction, the appropriate SI unit is the joule (J). In an exothermic reaction, the products have lower energy

than the reactants and the reaction releases energy in the form of heat to the surroundings. Therefore,  $\Delta H$  is negative ( $\Delta H < 0$ ) for an exothermic reaction.

To calculate the energy transferred from the combustion of alcohols to the water this equation will be used:  $q = m \cdot c \cdot \Delta T$

Where:

$m$  is the mass of water in the beaker

$c$  is the specific heat capacity of water ( $4.18 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$ )

$\Delta T$  is the increase in the temperature of water.

The experimental enthalpy of combustion of an alcohol ( $\Delta H$ ) will be calculated dividing the energy transferred to the water ( $q$ ) by the moles of alcohol burned in the reaction ( $n$ ):

$$\Delta H = - q / n$$

**Hypothesis:** As the number of carbon atoms in primary alcohols increases, the total number of bonds also increases. Thus, more energy is required to break these bonds during combustion. This results in a higher enthalpy of combustion with increasing carbon chain length following the trend: Methanol < Ethanol < Propanol < Butanol.

**Variables:**

<b>Independent variable</b>	The type of primary alcohol used	The alcohols Methanol, Ethanol, Propan-1-ol, Butan-1-ol have increasing carbon atoms which affect the enthalpy of combustion. This is the variable I choose to change.
<b>Dependent Variable</b>	The enthalpy of combustion (kJ/mol)	This is the variable that is calculated for the experiment. The change in the water's temperature and change in the mass of the alcohol in the spirit burner will be measured. These values will be used to find the heat energy ( $Q$ ) and subsequently enthalpy change ( $\Delta H$ ), will be calculated in $\text{kJ mol}^{-1}$ . This is the main value measured in the experiment and it shows how much energy is released when one mole of alcohol is

		burned
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<b>Control variable</b>	Volume of water heated in the beaker	The volume of water was kept constant at 100 ml of distilled water for each trial. This ensured that the same amount of heat from the burning alcohol was absorbed by the water, making comparisons fair
	The type of beaker used	A standard 100 cm <sup>3</sup> glass beaker was used throughout the experiment to maintain consistency in heat transfer and minimize differences in heat loss to the surroundings
	Distance between the spirit burner with alcohol and the beaker containing water	The distance between the spirit burner and the beaker was fixed at approximately 3 cm using a clamp and measured with a ruler. This helped keep the flame's effect on the beaker uniform in all trials.
	The time period the alcohol in the spirit burner is allowed to burn	The alcohol in each spirit burner was allowed to burn for exactly 2 minutes, ensuring that each fuel had the same time to transfer energy to the water
	The type and purity of alcohol burned	While doing repeats, the same purity of alcohol from the lab was used and measured, to ensure that the results were only affected by the number of carbon atoms and not by differences in the alcohol used.

#### Apparatus and Materials:

- 4 spirit burners each containing 25 cm<sup>3</sup> of the following pure alcohols:
  - Methanol

- Ethanol
- Propan-1-ol
- Butan-1-ol
- 100 ml of distilled water (per trial)
- Electronic balance (precision:  $\pm 0.001$  g)
- Measuring cylinder (precision:  $\pm 0.5$  ml)
- Volumetric pipette (precision:  $\pm 0.03$  ml)
- Labquest temperature probe with inbuilt stopwatch (precision:  $\pm 0.2^\circ\text{C}$ )
- 100 ml beaker
- Tripod stand
- Clamp stand
- Match box

### Method:

1. Using a volumetric pipette (precision:  $\pm 0.03$  ml) measure exactly 25 cm<sup>3</sup> of each pure alcohol (methanol, ethanol, propan-1-ol, and butan-1-ol) and transfer it into separate clean spirit burners.
2. Measure and record the initial mass of each spirit burner using an electronic balance (precision of  $\pm 0.001$  g)
3. Using a measuring cylinder (precision:  $\pm 0.5$  ml), measure 100 ml of distilled water for each trial and pour it into a 100 ml glass beaker.
4. Place the beaker on a tripod stand, and position the spirit burner beneath it. Adjust the clamp stand so that the distance between the wick of the burner and the base of the beaker is 3 cm.
5. Using a temperature probe with an inbuilt stopwatch (precision:  $\pm 0.2^\circ\text{C}$ ), measure and record the initial temperature of the water.
6. Ignite the spirit burner and allow the alcohol to burn for exactly 2 minutes. Start the stopwatch immediately after ignition and monitor the water temperature throughout. The temperature probe will monitor and log the temperature change over exactly 2 minutes.
7. After 2 minutes, extinguish the flame carefully and immediately reweigh the spirit burner using the electronic balance to determine the final mass.
8. Repeat the above steps for all four alcohols using the same setup and control variables for 3 trials each
9. Calculate the temperature change ( $\Delta T$ ) of the water and the change in mass ( $\Delta m$ ) of each spirit burner. Use these values to determine the energy released ( $Q$ ) using the equation  $Q = mc\Delta T$  and then calculate the enthalpy change using the equation  $H = -Q/n$

### Safety:

1. I wore safety goggles and lab coats to protect against burns, splashes and broken glass (in case of breaking of beaker)
2. I conducted the experiment in a well ventilated lab to avoid accumulation of alcohol vapors
3. I kept all flammable materials away from the workspace to prevent accidental ignition.
4. I used a clamp stand to maintain a fixed 3 cm distance between the flame and the beaker reducing the risk of contact burns

5. I measured and transferred alcohols carefully using a volumetric pipette and measuring cylinder to prevent spills
6. I ignited spirit burners using a matchstick at arms length keeping my hands and face away from the flame
7. Ensured all burners were capped immediately after use to prevent alcohol evaporation and vapor exposure
8. Kept a fire extinguisher and fire blanket nearby in case of emergency
9. I allowed hot equipment to cool before handling to avoid burns.
10. Performed the experiment on a stable, flat surface to prevent tipping or accidental contact with the flame



**Raw Data Table:**

Alcohol	Trial	Initial Mass of Burner ( $\pm 0.001$ g)	Final Mass of Burner ( $\pm 0.001$ g)	Initial Temp. ( $\pm 0.2$ °C)	Final Temp. ( $\pm 0.2$ °C)
<b>Methanol</b>	1	19.651	17.790	24.6	39.8
	2	16.730	15.850	25.1	45.6
	3	16.610	14.840	24.7	61.5
<b>Ethanol</b>	1	18.740	17.050	24.7	43.3
	2	18.270	17.600	25.4	48.8
	3	18.060	17.210	24.9	54.1
<b>Propanol</b>	1	17.650	17.130	25.1	34.6
	2	16.990	16.390	25.2	48.0
	3	17.600	16.930	25.5	52.5
<b>Butanol</b>	1	19.560	18.670	24.8	57.2
	2	21.110	20.060	25.3	52.3
	3	20.870	19.800	25.3	59.5

**Data Processing and Calculations:**

**Table 1: Calculations of Mass of Alcohol Burnt ( $\Delta m$ )**

Alcohol	Trial	Mass Difference (Initial – Final) (g)	Uncertainty	Average Mass (g)	Average Uncertainty	Final Mass Burnt (g)
<b>Methanol</b>	1	$19.651 - 17.790 = 1.861$	$0.001 + 0.001 = \pm 0.002$	$(1.861 + 0.880 + 1.770)/3 = 1.504$	$\pm 0.002$	$1.50 \pm 0.002$
	2	$16.730 - 15.850 = 0.880$				
	3	$16.610 - 14.840 = 1.770$				
<b>Ethanol</b>	1	$18.740 - 17.050 = 1.690$	$0.001 + 0.001 = \pm 0.002$	$(1.690 + 0.670 + 0.850)/3 = 1.070$	$\pm 0.002$	$1.07 \pm 0.002$
	2	$18.270 - 17.600 = 0.670$				
	3	$18.060 - 17.210 = 0.850$				

<b>Propanol</b>	<b>1</b>	17.650 – 17.130 = 0.520		(0.520+0.600 +0.670)/3 =0.597		0.60 ± 0.002
	<b>2</b>	16.990 – 16.390 = 0.600				
	<b>3</b>	17.600 – 16.930 = 0.670				
<b>Butanol</b>	<b>1</b>	19.560 – 18.670 = 0.890		(0.890+1.050 +1.070)/3 =1.003		1.00 ± 0.002
	<b>2</b>	21.110 – 20.060 = 1.050				
	<b>3</b>	20.870 – 19.800 = 1.070				

**Table 2: Temperature Change (ΔT) for All Alcohols**

Alcohol	Trial	Initial Temp (°C)	Final Temp (°C)	ΔT (°C)	Uncertainty	Average ΔT (°C)	Avg Uncertainty	Final ΔT (°C)
<b>Methanol</b>	1	24.6	39.8	15.2	0.2 + 0.2 = ±0.4	24.17	±0.4	24.2 ±0.4
	2	25.1	45.6	20.5				
	3	24.7	61.5	36.8				
<b>Ethanol</b>	1	24.7	43.3	18.6		23.73		23.7 ±0.4
	2	25.4	48.8	23.4				
	3	24.9	54.1	29.2				
<b>Propanol</b>	1	25.1	34.6	9.5		19.77		19.8 ± 0.4
	2	25.2	48.0	22.8				
	3	25.5	52.5	27.0				
<b>Butanol</b>	1	24.8	57.2	32.4		31.20		31.2 ± 0.4



	2	25.3	52.3	27.0				
	3	25.3	59.5	34.2				

**Table 3: Energy Absorbed by Water (Q) for All Alcohols**

**Formula:  $Q=mc\Delta T$  ( $c = 4.18 \text{ J/g}^\circ\text{C}$ )\***

**$m=100\text{g}$  (uncertainty  $\Delta m=\pm 0.5\text{g}$ )**

$$\Delta Q = Q \times \left( \frac{\Delta m}{m} + \frac{\Delta T}{T} \right)$$

Alcohol	Trial	$\Delta T$ ( $^\circ\text{C}$ )	Q (kJ)	Uncertainty ( $\pm\text{kJ}$ )	Average Q (kJ)	Avg. Uncertainty ( $\pm\text{kJ}$ )	Final Q (kJ)
<b>Methanol</b>	1	15.2	$100 \times 4.18 \times 15.2$ /1000=6.35	$6.35 \times (0.5/100 + 0.4/15.2) = 0.19$	$(6.35 + 8.57 + 15.38)/3 = 10.07$	$(0.19 + 0.27 + 0.25)/3 = 0.24$	$10.07 \pm 0.24$
	2	20.5	$100 \times 4.18 \times 20.5$ /1000=8.57	$8.57 \times (0.005 + 0.026) = 0.27$			
	3	36.8	$100 \times 4.18 \times 36.8$ /1000=15.38	$15.38 \times (0.005 + 0.011) = 0.25$			
<b>Ethanol</b>	1	18.6	$100 \times 4.18 \times 18.6$ /1000=7.77	$7.77 \times (0.005 + 0.4/18.6) = 0.21$	$(7.77 + 9.78 + 12.20)/3 = 9.92$	$(0.21 + 0.22 + 0.23)/3 = 0.22$	$9.92 \pm 0.22$
	2	23.4	$100 \times 4.18 \times 23.4$ /1000=9.78	$9.78 \times (0.005 + 0.017) = 0.22$			
	3	29.2	$100 \times 4.18 \times 29.2$ /1000=12.20	$12.20 \times (0.005 + 0.014) = 0.23$			
<b>Propanol</b>	1	9.5	$100 \times 4.18 \times 9.5$ /1000=3.97	$3.97 \times (0.005 + 0.0421) = 0.19$	$(3.97 + 9.53 + 11.29)/3 = 8.26$	$0.19 + 0.21 + 0.22)/3 = 0.21$	$8.26 \pm 0.21$
	2	22.8	$100 \times 4.18 \times 22.8$ /1000=9.53	$9.53 \times (0.005 + 0.0175) = 0.21$			
	3	27.0	$100 \times 4.18 \times 27.0$ /1000=11.29	$11.29 \times (0.005 + 0.0148) = 0.22$			
<b>Butanol</b>	1	32.4	$100 \times 4.18 \times 32.4$ /1000 =13.54	$13.54 \times (0.005 + 0.0123) = 0.23$	$(13.54 + 11.29 + 14.29)/3 = 13.04$	$(0.23 + 0.22 + 0.24)/3 = 0.23$	$13.04 \pm 0.23$
	2	27.0	$100 \times 4.18 \times 27.0$ /1000=11.29	$11.29 \times (0.005 + 0.0148) = 0.22$			
	3	34.2	$100 \times 4.18 \times 34.2$ /1000=14.29	$14.29 \times (0.005 + 0.0117) = 0.24$			

**Table 4: Moles of Alcohol Burnt (n) for All Alcohols**

$$\Delta n/n = \frac{\Delta m}{m} + \frac{\Delta \text{molar mass}}{\text{molar mass}}$$

$\Delta \text{molar mass} = 0.01 \text{ g/mol}$

Alcohol	Molar Mass (g/mol)	Mass Burnt (g)	n (mol)	Uncertainty ( $\pm$ mol)	Final n (mol)
Methanol	32.04	1.50 $\pm$ 0.002	1.50 / 32.04 = 0.0468	0.0468*(0.002/1.50)+(0.01/32.04)= 0.000077	0.0468 $\pm$ 0.00008
Ethanol	46.07	1.07 $\pm$ 0.002	1.07 / 46.07 = 0.0232	0.0232*(0.002/1.07)+(0.01/46.07)= 0.000047	0.0232 $\pm$ 0.00005
Propanol	60.10	0.60 $\pm$ 0.002	0.60 / 60.10 = 0.00998	0.00998*(0.002/0.60)+(0.01/60.10) = 0.000035	0.00998 $\pm$ 0.00004
Butanol	74.12	1.00 $\pm$ 0.002	1.00 / 74.12 = 0.0135	0.0135*(0.002/1.00)+(0.01/74.12)= 0.000029	0.0135 $\pm$ 0.00003

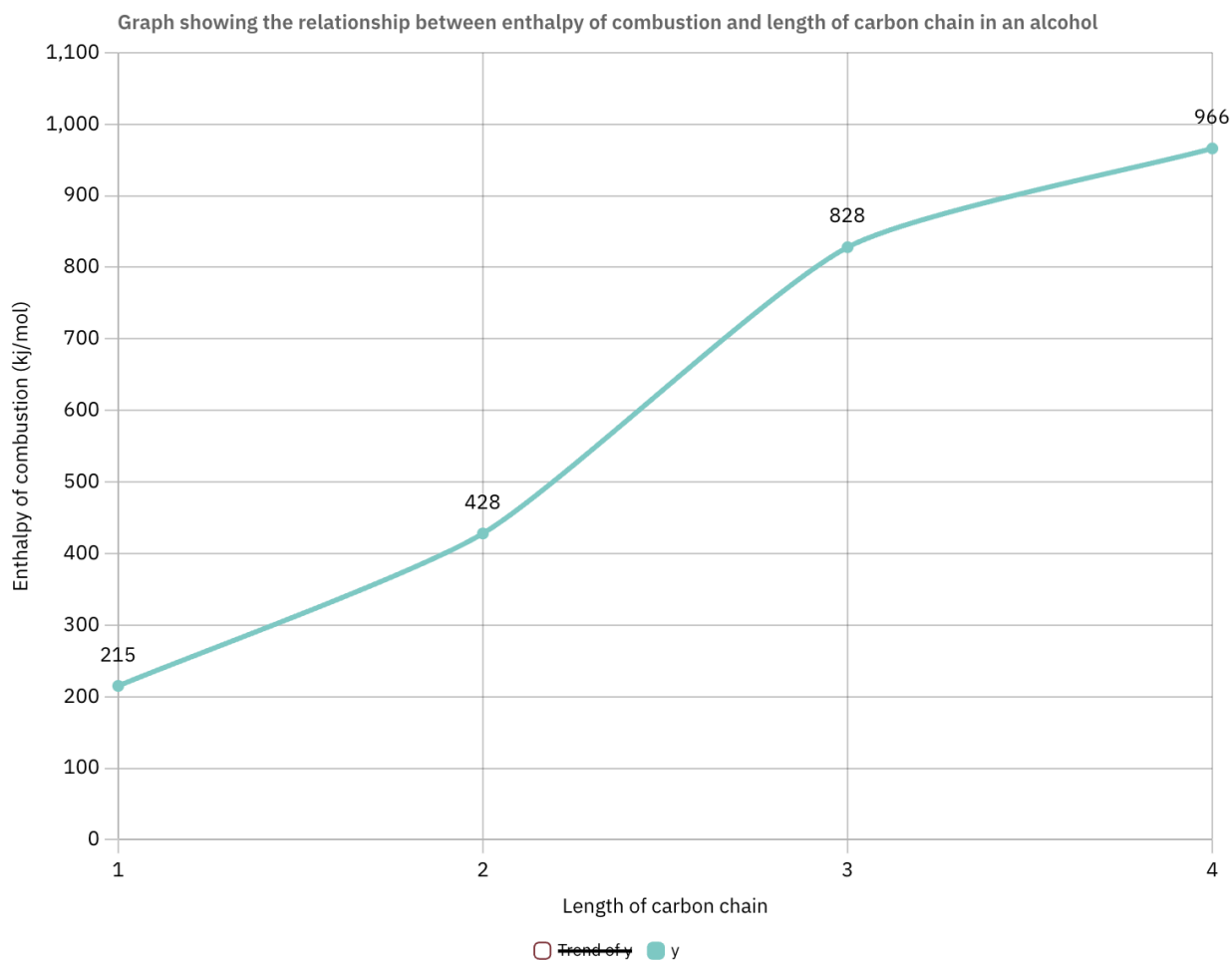
**Table 5: Enthalpy of Combustion ( $\Delta H$ ) for All Alcohols**

Formula:  $H = -q/n$

$$\Delta H = H \times \left( \frac{\Delta Q}{Q} + \frac{\Delta n}{n} \right)$$

Alcohol	Final. Q (kJ)	Final. n (mol)	$\Delta H$ (kJ/mol)	Uncertainty ( $\pm$ kJ/mol)	Final $\Delta H$ (kJ/mol)
Methanol	-10.07	0.0468	-10.07 / 0.0468 = -215.2	$215.2 \times (0.24/10.07 + 0.00006/0.0468) = 5.4$	-215 $\pm$ 5
Ethanol	-9.92	0.0232	-9.92 / 0.0232 = -427.6	$427.6 \times (0.22/9.92 + 0.00004/0.0232) = 10.2$	-428 $\pm$ 10
Propanol	-8.26	0.00998	-8.26 / 0.00998 = -827.7	$827.7 \times (0.21/8.26 + 0.00003/0.00998) = 23.5$	-828 $\pm$ 24
Butanol	-13.04	0.0135	-13.04 / 0.0135 = -965.9	$965.9 \times (0.23/13.04 + 0.00003/0.0135) = 19.2$	-966 $\pm$ 19

## Graph:



**Conclusion:** The graph clearly demonstrates a positive linear relationship between the number of carbon atoms in a primary alcohol and its enthalpy of combustion. As the carbon chain length increases from methanol (1 carbon) to butanol (4 carbons) the enthalpy of combustion increases from 215 kJ/mol for methanol to 966 kJ/mol for butanol. This trend aligns with theoretical expectations as the longer chain alcohols contain more bonds, particularly C-C and C-H bonds which release more energy while combustion.

The trend observed  $\text{Methanol} < \text{Ethanol} < \text{Propanol} < \text{Butanol}$  supports the principle that more energy is released when more bonds are formed in the combustion products ( $\text{CO}_2$  and  $\text{H}_2\text{O}$ ). The relatively constant gradient of approximately 250.3 kJ/mol per carbon atom between each successive alcohol suggests a near linear increase in enthalpy of combustion. This indicates that each additional  $\text{CH}_2$  group contributes a relatively consistent amount of energy to combustion showing the proportionality between carbon chain length and combustion energy.

This pattern occurs because each  $\text{CH}_2$  group introduces additional C-H and C-C bonds that release energy when broken and new bonds are formed with oxygen during combustion. Since the bond enthalpies of these groups are fairly consistent across alcohols, the energy released per added carbon atom remains relatively uniform and causes a linear increase in total enthalpy of combustion.

## Evaluation:

Source: NIST Chemistry WebBook

Alcohol	Experimental Enthalpy of Combustion (kJ/mol)	Theoretical Enthalpy of Combustion (kJ/mol)
Methanol	215	-726
Ethanol	428	-1367
Propanol	828	-2021
Butanol	966	-2674

There is discrepancy between my experimental and theoretical values of enthalpy of combustion. The measured values are lower than standard literature values. This deviation is due to several systematic and random errors that occurred during the experiment.

One of the greatest sources of error was heat loss to the surroundings. Although the alcohols were combusted directly beneath the beaker filled with water a substantial portion of the energy released was lost to the surrounding air and the beaker, which was not insulated. This inefficiency in heat transfer could be reduced by using a more insulated apparatus, such as a bomb calorimeter, to better contain and measure the released energy.

Another systematic error likely arose from incomplete combustion, which may have occurred due to the use of an old wick due to unavailability of new wicks in the school lab. This was sometimes indicated by soot deposits on the base of the beaker. Incomplete combustion results in the production of carbon or carbon monoxide instead of carbon dioxide releasing less energy than complete combustion. Ensuring clean wick and conducting the experiment in a well ventilated area would help promote complete combustion. Additionally, alcohol evaporation between the weighing of the burner and ignition could have affected the accuracy of the mass loss measurements, leading to underestimation of the actual fuel combusted. This could be minimized by reducing setup time and working in a draft free environment.

Random errors also influenced the outcome. While a fixed flame to beaker distance of 3 cm was maintained using a clamp stand, minor inconsistencies during setup could have led to slight variations in heat transfer efficiency. Furthermore, while a Labquest temperature probe ( $\pm 0.2^{\circ}\text{C}$ ) was used to monitor the temperature change digitally small fluctuations or lag in probe response could have impacted the precision of the temperature readings.