Thermochemistry: Enthalpy of Combustion for Fuels

Purpose

The purpose of this experiment is to determine and compare the enthalpies of combustion of several fuels using combustion calorimetry.

Learning Objectives

Gain experience with the fundamentals of experimental calorimetry

Measure and compare the enthalpy of combustion for several different fuels

Equipment

- calorimeter
- fuel burners

Chemicals

- Hexane
- One of the following: Methanol, Ethanol, Isopropyl alcohol, *n*-Propyl alcohol, Methyl *t*-butyl ether, methyl oleate

Introduction

Thermochemistry and Reaction Enthalpies

The topics of thermochemistry, enthalpy, enthalpy changes, and heat are covered in your text. Very simply;

- **Thermochemistry** is the relationship between chemical behavior and energy.
- **Heat**, q, is energy that flows in or out of a system because of a temperature difference.
- **Enthalpy**, H, is a quantity closely related to the internal energy of a chemical system, with an adjustment for pressure and volume (usually a small adjustment).
- **Enthalpy change**, ΔH, is the difference between the final and initial enthalpies of a system that undergoes some sort of change.

Enthalpy changes occurring during chemical reactions are important both because of their fundamental relationship to why reactions occur, and because energy produced by (or required by) chemical reactions have many important practical applications (fuels, dietetics, etc.).

Enthalpy and Heat

The definition of enthalpy is written so that when a chemical change occurs under conditions of constant pressure, the enthalpy change of the chemical system is identical to the heat absorbed by the system: $\Delta H = q_\mathrm{p}$ (heat absorbed at constant pressure). Enthalpy changes associated with reactions are often referred to simply as "heats of reaction" (or "heats of combustion" for combustion reactions). Sign conventions are such that when a process absorbs heat (an **endothermic** process), q and ΔH are positive, and when a process releases heat (an **exothermic** process), q and ΔH are negative. This is commonly referred to as the "selfish convention" and may not match what you learn in another science class. It's referred to as "selfish" because if you put yourself in the place of the system (whatever is burning or reacting), ask if you would be receiving energy or giving it away. If you receive it, then just like your bank account with money, you would be *increasing* your enthalpy (+). The reverse is true: if you give away heat, that would be reducing your enthalpy (-).

Heat, q, and enthalpy change, ΔH, are really the same for constant pressure reactions. However, we will be discussing heat (enthalpy changes) on three different scales for this lab: for the amount of fuel actually burned, for one gram of fuel burned, and for one mole of fuel burned. We will use the following conventions to help distinguish these:

Table TC.1: Conventions

Symbol	Meaning	Units
q_reaction	the heat of a reaction for the amount of fuel actually burned	kJ
Q_g	the heat of reaction for one gram of fuel burned	kJ/g
ΔH_mol	the heat of reaction for one mole of fuel burned	kJ/mol

These are actually all enthalpy changes (each based on heat of reaction at constant pressure, q_P). We are using different symbols here just to help differentiate their specific meanings—as carried out, per gram, and per mole.

Combustion Calorimetry

A calorimeter is a device used to measure energy changes as chemical changes occur. The reaction is carried out in some sort of insulated container so that, ideally, no heat is exchanged with the calorimeter's surroundings. For an exothermic process, all heat released by the process is absorbed by the materials that make up the calorimeter's interior, causing its temperature to rise. Similarly, for an endothermic process, heat absorbed by the reaction comes from the calorimeter's interior, causing its temperature to drop. If the heat capacity—the amount of energy needed to raise or lower the temperature of the calorimeter (and its contents) by precisely 1 °C—is known, then the temperature change can be directly related to the heat released or absorbed by the chemical process. Details of the calculations will be provided later.

Measurements of energy changes for combustion reactions are very important, but pose experimental difficulties. Combustion reactions often produce significant quantities of hot gases that rapidly escape, and accurate calorimetry requires that energy from these gases be captured by the calorimeter. The best design for combustion calorimeters is the "bomb" design (see Figure TC.1), where the sample is held in a heavy-walled sealed vessel (bomb) charged with pure oxygen under pressure. An electrical ignition system starts the reaction, and the sample burns rapidly because of the oxygen atmosphere. These devices must be designed to withstand high pressures—reactions occurring in a pressurized pure oxygen atmosphere are essentially explosions. All of the hot gases are held inside the calorimeter's insulated "jacket", so no heat is lost.

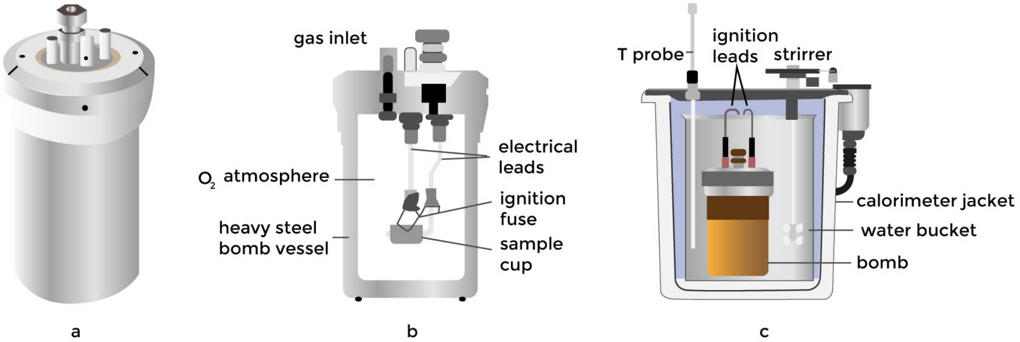


Figure TC.1: Parr Model 1108 Oxygen Bomb for Calorimetry: a) Photograph of oxygen bomb; b) Diagram showing components for oxygen gas inlet, electrical feed-throughs for ignition leads, fuse, and sample cup; c) The bomb unit is placed in a container of water which in turn is placed in a large, well-insulated calorimeter "jacket." The fuel is ignited electrically, and the temperature change of the water is monitored. Adapted from www.parrinst.com.

This type of calorimeter gives excellent results, but it does have three major drawbacks for application to general chemistry laboratories:

- they are expensive (well over $5,000 each)
- their operation is time consuming—you'd probably only complete 1-2 runs during the course of a 2-hour lab period.
- because it increases in pressure while the combustion occurs, this is not measuring the constant-pressure energy change (q_p).

Simple Combustion Calorimeter

A simple combustion calorimeter will be used for this lab. It will provide results that are sufficiently accurate to allow comparisons between different fuels.

Figure TC.2 shows the basics of the calorimeter design. The main components of the device (heat exchanger and vessel) are made of aluminum, an excellent heat conductor. The calorimeter vessel also contains a measured amount of water.

When a fuel burner is placed under the calorimeter, hot gases from the combustion rise and are directed through a channel in the middle of the heat exchanger, and around the outside of the vessel. Heat transmits efficiently from the hot gases into the water, and the insulation reduces the amount of heat lost to the outside. The temperature probe allows continuous measurement of temperature. Key measurements are the initial and final temperatures, and the initial and final masses of the fuel burner.

Before the calorimeter may be used to measure the enthalpy change for a fuel's combustion, it must be calibrated. Specifically, the calorimeter's heat capacity, $C_{\text{calorimeter}}$, must be determined experimentally. The simplest way to do this is to burn a fuel whose enthalpy of combustion is known to find the relationship between heat of reaction and temperature rise. For this lab, hexane, C_6H_{14}, a simple hydrocarbon, will be used for calibration. For hexane, Q_g is –44.8 kJ/g.

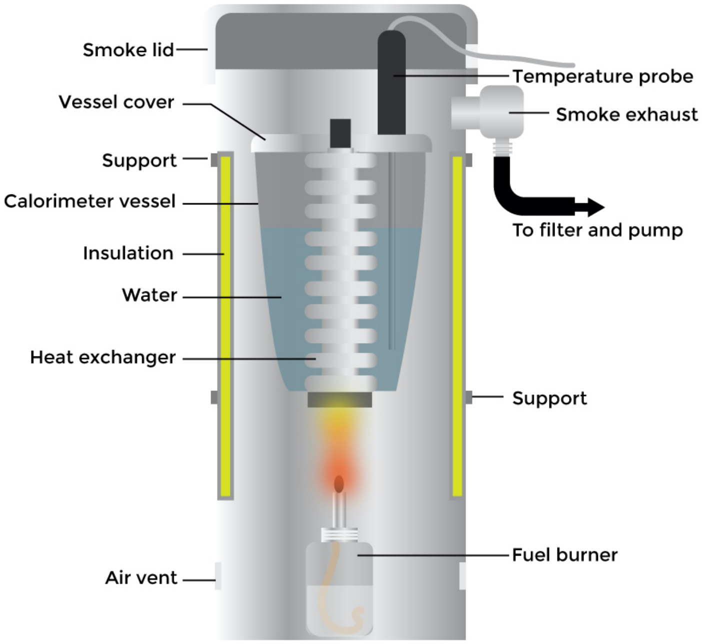


Figure TC.2: Combustion Calorimeter used for this lab

Key Calculation

Details of calculations will be provided in the Data Analysis section, so only the basics will be presented here. In each experimental run, you will measure m_{burned}, the mass of fuel burned, and ΔT, the calorimeter's temperature rise. The key equation is shown in equation TC.1:

$$q_{\text{reaction}} = -C_{\text{calorimeter}} \times \Delta T \qquad \text{(Equation TC.1)}$$

In the calibration step, q_{reaction} will be known from the amount of hexane fuel burned, and you will measure ΔT, so you can calculate $C_{\text{calorimeter}}$.

For the other fuels, you will measure ΔT and m_{burned}, and then calculate q_{reaction} from the equation above. Notice the negative sign in the calculation: when the temperature rises, ΔT is positive, and q_{reaction} is negative, corresponding to an exothermic process.

Once you know q_{reaction} for the fuel, you can calculate ΔH_{mol} and Q_g for the fuel, based on the relationships depicted in Figure TC.3.

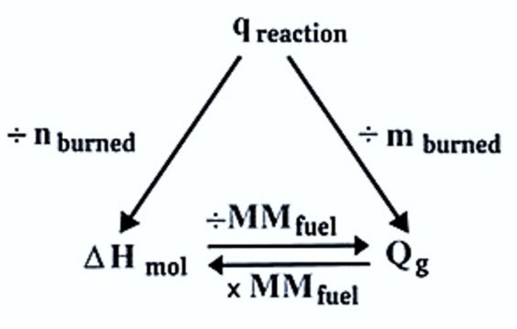


Figure TC.3: ΔH_{mol} and Q_g Relationships

Fuels

Fuels are important for some obvious reasons: when they are burned, they provide the energy needed to produce electricity, power vehicles, and heat buildings. A special concern for vehicle fuels is that they must be carried with the vehicle, and refueling must be convenient and safe. Fuels that are “energy dense” are preferred, and both *gravimetric* and *volumetric* energy density are important.

Gravimetric energy density is energy provided per kilogram (or per gram, pound, etc.). If a fuel has a high gravimetric energy density, a smaller mass of fuel is needed. Note that for “Q_g” values are equivalent to gravimetric energy densities.

Volumetric energy density is energy provided per liter (or per milliliter, gallon, etc.). If a fuel has a high volumetric energy density, a smaller volume of fuel (fewer gallons) is needed.

Other fuel characteristics such as how “clean” they burn and how easily they ignite (octane and cetane ratings – see box on the next page) are also important. For this lab, we will focus on gravimetric energy density, and also on energy per mole.

Octane and Cetane Ratings

> *Message: Don’t try to interchange gasoline and diesel fuel! Read below to understand the difference in how they operate.*

Octane Rating

For gasoline engines, the ability of a fuel to resist “preignition” when compressed is critical. Gasoline engines utilize an ignition system designed to fire the engine’s spark plugs at specific times in the engine’s cycle. However, during the engine cycle, the air-fuel mixture is compressed, causing the mixture to heats up and may cause poorer grade fuels to ignite spontaneously. This results in “knocking” (or “pinging”), leads to poor engine performance, and can also damage the engine. Higher octane fuels resist preignition—essentially, they are harder to ignite and *require* a spark. Most modern engines are designed to run on fuels with an octane rating a bit below 90. The only engines that require higher octane fuel are high compression engines. For regular engines designed for lower octane fuels, *there is absolutely no advantage to using higher octane fuels!* Higher octane fuels do *not* provide higher power (but they cost more).

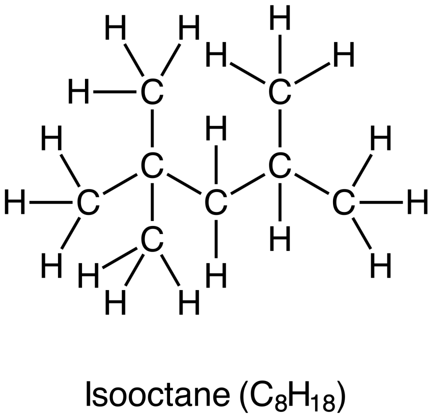


Figure TC.4: Isooctane C_8H_{18}

Octane ratings are based on comparisons to “isooctane” (systematic name: 2,2,4-trimethylpentane, Figure TC.4), a specific form of C_8H_{18}, which is assigned an octane rating of 100. Note that isooctane’s carbon chain is “branched” instead of having all carbon compounds in a straight line. Aromatic compounds, compounds with benzene-like structures, also have high octane ratings, as does ethanol. Linear hydrocarbons (all of the carbons in a straight line, without branches) have poor octane characteristics, and n-heptane, the straight-chain seven-carbon hydrocarbon, is assigned an octane rating of zero.

Cetane Rating

Cetane ratings apply to diesel fuels, and in many ways, are the *opposite* of octane ratings. Diesel engines are initially charged with air only, and undergo much higher compression than would occur in a gasoline engine. Thus, the diesel fuel is injected into air at a very high temperature, causing it to ignite upon injection. Hence, the critical characteristic of diesel fuels is that they ignite spontaneously at high temperatures. Cetane rating is essentially a measure of how easily a fuel ignites, just the opposite of the octane rating for gasoline fuels.

Cetane ratings measured relative to “cetane”, a *linear* (continuing string of C–C bonds, without branches) form of $C_{16}H_{34}$ hydrocarbon (Figure TC.5, which is assigned a rating of 100. Characteristics that provide high octane ratings—branches in the carbon chains, aromatic structures—usually tend to lead to very poor cetane ratings.

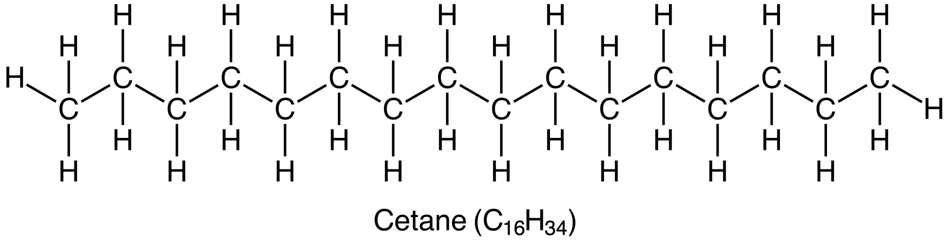


Figure TC.5: Cetane $C_{16}H_{34}$

Procedure

The fuels used in this laboratory are flammable, and the procedure requires open flames. Be careful about what you wear for this lab. Clothing with loose sleeves and/or made from readily ignitable synthetic fabrics should be avoided.

Safety Precautions

- Special care should be taken to avoid spilling any fuel. If a fuel does spill, inform your instructor and make sure everyone around you is aware of the spill, and that there are no sources of ignition (e.g., open flames) nearby before you start cleaning up the spill.
- The only bottles of flammable solvent should be the ones that are part of the fuel lamps. Remove any other containers of flammable solvent, including acetone wash bottles, from your work area.
- Because the lab procedure utilizes open flames, be particularly careful about materials that can ignite, such as papers (lab papers, paper towels, etc.).
- In a well-lit room, small flames are not always readily visible. Always take a good look at fuel burners to see if they are lit before handling them.
- Hot objects tend to look the same as cold objects—be cautious when you are about to pick up any fuel burner that has been used recently.

Equipment and Software Familiarization

Calorimeter

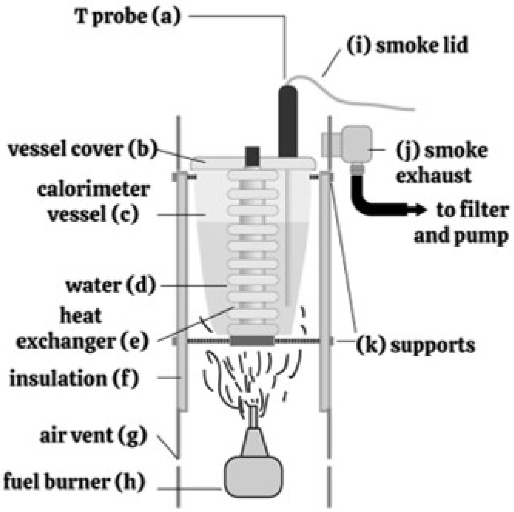


Figure TC.6: Combustion Calorimeter

Figure TC.6 shows a fully labeled schematic of the calorimeter design. Use Figure TC.6 and the directions below to familiarize yourself with the apparatus.

- Note that the cap fits *loosely* over the calorimeter body.
- A filter and air pump, not shown in the figure but attached at (k) in the figure, help minimize emission of soot into the lab.

Fuel Burners

Fuel burners for hexane, C_6H_{14}, (the calibration standard) and several other fuels will be provided. A few different types of burners will be used, depending on the physical characteristics of the fuel. Figure TC.7 shows a generic image of a burner.

Burner wicks may need to be adjusted periodically, to a length of about 3–6 mm (on the shorter side is usually better, see Figure TC.7).

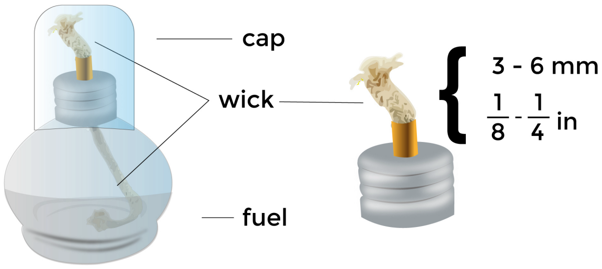


Figure TC.7: Fuel Burner

- Do not let wicks burn down to the metal stem.
    - If wicks need to be lengthened, use tweezers.
- After any type of adjustment, light the burner and let it burn for 20–30 seconds.

- Burners may be extinguished by placing the cap back over the wick.
- Keep the cap on the burner except when it is in use.
  - The cap should be on the burner when it is weighed.
  - Remove the cap right before you light the burner.
  - Put the cap back on to extinguish the burner.

Setting up the Calorimeter

1. Be sure that your data collection device (Lab Quest) is set up according to your instructor's directions.

2. Pour 350 ml deionized water (graduated cylinder measurements) into the calorimeter vessel. (Use fresh water for each run. Dispose of the water appropriately.)

3. Turn on the air pump (connected through the filter).

4. Make sure the glass stirring road is in place, you must stir continually during the entire duration of the run.

5. Make sure the temperature probe is positioned correctly.

6. Weigh the hexane burner with its cap in place and record this on your report sheet. Use the same balance for the initial and final readings for any run.

Start Temperature Monitoring

7. Connect the Temperature Probe to LabQuest and choose 'New' from the 'File Menu'.

8. On the Meter screen, tap 'Rate'. Change the data-collection rate to 0.5 sample/second (interval of 2 seconds/sample) and the data-collection length to 600 seconds. Data collection will last 10 minutes. Select 'OK'.

9. Monitor temperature (in °C) for about 30 seconds, and record the initial temperature of the water, t_1, in your data table.

10. Light the burner (remove its cap just before lighting) and insert it under the calorimeter.

11. Watch the temperature on the monitor. It should start to rise after a few seconds of heating. Continue heating until either;
    a. the temperature has risen by at least 10 °C OR
    b. 5 minutes of heating time has elapsed (from LabQuest first bullet point), whichever comes sooner.

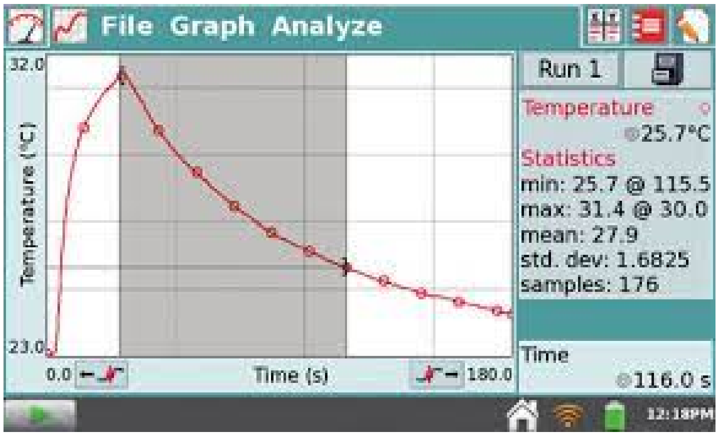
12. Continue stirring the water until the temperature stops rising.

13. Continue to monitor the temperature for about 1 more minute after removing the burner, then click the 'Stop' icon to terminate data collection. Record this maximum temperature, t_2.

14. Remove the burner and promptly extinguish and recap it. Data collection will stop after 10 minutes (or stop before 10 minutes if temperature has risen at least 10 °C).

15. Record the final mass of the burner with its cap in place.

Preliminary Data Evaluation



Temperature
25.7°C
Statistics
min: 25.7 @ 115.5
max: 31.4 @ 30.0
mean: 27.9
std. dev: 1.6825
samples: 176
Time
116.0 s

Figure TC.8: LabQuest screen showing an initial temperature, then temperature increase, followed by a slow cooling. The peak of the temperature is where you should record the final temperature.[a]

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[a] Image from LabQuest Quickstart guide (LQ2_QSG_1203).

16. Calculate m_{burned}, the difference between the initial and final mass of the burner.

17. A typical T versus time plot is shown in Figure TC.8.

    a. Tap-and-drag over the portion of the plot before heating was started, then choose ‘Statistics’ from the ‘Analyze’ menu. Select the check box to select a data column. The statistics information is then displayed in the panel to the right of the graph. The mean (average) temperature value will be your value of T_{init}.
    b. Tap-and-drag over the portion of the plot that includes the peak after heating was stopped (see shaded section (b) in Figure 10), then choose ‘Statistics’ from the ‘Analyze’ menu. Select the check box to select a data column. The statistics information is then displayed in the panel to the right of the graph. The mean (average) temperature value will be your value of T_{final}.
        i. Calculate $\Delta T = T_{\text{final}} - T_{\text{init}}$
        ii. Print the plot.

18. Calculate m_{burned} (the difference between the initial and final mass of the burner).

19. Calculate $\Delta T/m_{\text{burned}}$.

20. Repeat the process (steps 5–18) with the hexane burner.

    a. Before starting, clear the data display (but make sure you printed the previous plot!).
    b. Be sure to replace the water in the calorimeter before starting the next run.
    c. Tap on the ‘filing cabinet’ icon to start a new run.

21. Compare the $\Delta T/m_{\text{burned}}$ ratios for the two runs to see if they agree within 10% (i.e., the percent difference is $\leq$ 10%).

    a. Percent difference calculations are covered in Appendix D, with the key calculation shown below. R1 and R2 are the two measured results.
    b. If the two ratios agree, you can proceed to the measurement of another fuel.
    c. If they do not, make additional runs until two ratios agree to within 10%.

> Note: Here is a shortcut for figuring out if the percent difference is less than 10%. (This is NOT a percent difference calculation, just a useful estimate.) Estimate the two values' average and divide that by 10. If this is more than the difference between the two values, then the % difference is less than 10%.
>
> For example, if the ratios are 26.8 and 29.4 °C/g, their average is 28.1 °C/g.
>
> One tenth of this is 2.8 °C/g.
>
> The difference between the two values is only 2.6 °C/g, so their difference is less than 10%.

22. Repeat process (steps 5–20) using a burner with a different fuel. You should obtain two replicate runs for which the T/m_{burned} ratios that agree with each to within 10%.
    a. The ratio does not need to agree with the ratio found for the hexane run!
    b. If the ratios differ by more than 10%, make additional runs until two results agree to within 10%.

23. Overall, you should have two data sets (hexane and one other fuel), each representing at least two replicate measurements for which T/m_{burned} agree within 10%. Thus: at least four trials total.

Cleanup Checklist

- Pour out water from the calorimeter. (The calorimeter does not need to be dried.)
- Make sure all fuel burners are extinguished and capped.
- Moisten a paper towel with soapy water and wipe down any soot on the outside of the calorimeter.
- Use soap and water (and paper towels as needed) to clean out the inside of the smoke cap and the upper portion of the inside of the calorimeter.
- Organize supplies in your work area, and wipe down the bench top if there were any water spills.

Data Analysis

General Rules

- For the additional fuel, select your best two runs, i.e., if you made more than two runs for a given fuel, use the two that had the closest agreement.
- Include units in all calculations.
- Keep track of significant figures in both the data and the calculations.
  - T_{init} and T_{final} values should be recorded to one decimal place (e.g., 26.3 °CC/g), and so ΔT values should also be calculated to one decimal (e.g., $\Delta T = 11.2$ °C).

- All masses should have been recorded to milligram precision (e.g., 56.235 g), and so the mass burned, m_{burned}, should always be expressed to milligram precision (e.g., 0.748 g).
- Overall, because ΔT values will normally have three significant figures, and m_{burned} will have three (or more) significant figures, final results should be expressed to three significant figures.

Calibration Calculation

Calculating Heat Produced by Hexane

1. Hexane has a known enthalpy of combustion per gram (Q_g) of -44.8 kJ/g.

2. Multiplying the value of m_{burned} by this value gives the heat produced by burning the hexane, q_{reaction}. For example, if $m_{\text{burned}} = 0.641$ g, the heat of reaction is;

$$q_{\text{reaction}} = Q_g \times m_{\text{burned}} = \text{-44.8 kJ/g} \times 0.641\text{ g} = \text{-28.7 kJ}$$

Calculating the Heat Capacity of the Calorimeter

3. The heat of reaction is related to the calorimeter's heat capacity and the temperature change as follows:

$$q_{\text{reaction}} = -C_{\text{calorimeter}} \times \Delta T \qquad \text{(Equation TC.1 revisited)}$$

Use this equation along with the q_{reaction} value calculated in step 1 and the measured value of ΔT to find $C_{\text{calorimeter}}$.

Example: If q_{reaction} is −28.7 kJ and ΔT is equal to 11.8 °C, what is the heat capacity of the calorimeter, $C_{\text{calorimeter}}$? Rearrange equation TC.1 to solve for $C_{\text{calorimeter}}$, and substitute in the given values.

$$C_{\text{calorimeter}} = -\frac{q_{\text{reaction}}}{\Delta T} = \frac{-28.7\text{kJ}}{11.8\,^{\circ}\text{C}} = 2.43\text{kJ}/^{\circ}\text{C}$$

4. Carry out this calculation for both calibration runs.

Average and % Difference

5. Calculate the average value of $C_{\text{calorimeter}}$ for the two runs.

6. Calculate the percent difference between the two values of $C_{\text{calorimeter}}$.

7. Use the average value of $C_{\text{calorimeter}}$ for all subsequent calculations.

Calculations for Other Fuels

Carry out the following calculations for each fuel:

1. For each run of fuel, calculate the heat of the reaction from the average value of $C_{\text{calorimeter}}$ and ΔT using equation TC.1.

2. Calculate the molecular mass of the fuel (MM_{fuel}).

3. For each run of fuel, calculate the moles of fuel burned, n_{burned}, by dividing the mass of fuel burned by its molecular mass.

$$nburnn_{\text{burned}}(\text{mol}) = \frac{m_{\text{burned}}(\text{g})}{MM_{\text{fuel}}(\text{mol}} \qquad \text{(Equation TC.2)}$$

4. Next, for each run, calculate ΔH_{mol}, the enthalpy change per mole of fuel burned by dividing q_{reaction} by n_{burned}.

$$\Delta H_{\text{mol}}(\text{kJ/mol}) = \frac{q_{\text{reaction}}(\text{kJ})}{n_{\text{burned}}(\text{mol})} \qquad \text{(Equation TC.3)}$$

deltaH

5. You can also calculate Q_g, the enthalpy change per gram of fuel burned. There are two completely equivalent ways of doing this: You can divide ΔH_{mol} by the fuel's molecular mass or divide q_{reaction} by m_{burned};

$$QgQ_g = \frac{\Delta H_{\text{mol}}}{MM_{\text{fuel}}} \quad OR \quad Q_g = \frac{q_{\text{reaction}}}{m_{\text{burned}}} \qquad \text{(Equation TC.4)}$$

The two methods should give you the same value (within normal round-off errors). Pick one or the other – you do not need to do both. Note that Q_g, the enthalpy change per gram of fuel, is an expression of the fuel's gravimetric energy density—how well a gram of the fuel provides heat.

6. Find the average value of ΔH_{mol} for the fuel, and then calculate the % difference between the values of ΔH_{mol} (for the same fuel).

7. Write the balanced chemical equation for combustion of the fuel in O_2. The only products should be $H_2O(g)$ (water vapor does not condense within the calorimeter) and $CO_2(g)$. Adjust the coefficients in the balanced equation so that there is only one mole of fuel. Coefficient for O_2 may be fractional, e.g, 8½.

8. Given values of ΔH_f° for the fuel and products (see table below), calculate $\Delta H_{\mathrm{reaction}}$ for the combustion reaction. The calculations are reviewed later.

9. Calculate the % error in your experimental value (ΔH_{mol}), treating the ΔH value calculated from ΔH_f° values as the "true" value. Note that since the $\Delta H_{\mathrm{reaction}}$ value calculated from ΔH_f° values is for one mole of the fuel, you can view its units as kJ/mol.

Comparisons

You will be asked to compare the fuels on both Q_g and ΔH_{mol} bases, and to values for other types of fuels.

Table TC.2: Enthalpies of Formation (ΔH_f°) for Fuels and Related Substances

Fuel	Formula	ΔH_f° (kJ/mol)
Hydrocarbons		
Acetylene (*g*)	C_2H_2	+227
Butane (*g*)	C_4H_{10}	-126
Benzene (*l*)	C_6H_6	+83
n-hexane (*l*)	C_6H_{14}	+198
n-Octane (*l*)	C_8H_{18}	-250
iso-Octane (*l*)	C_6H_6	-259
n-Dodecane (*l*)	$C_{12}H_{26}$	-351
Cetane (*l*)	$C_{16}H_{34}$	-456
Oxygenated Fuels		
Methanol (*l*)	CH_4O	-239
Ethanol (*l*)	C_2H_6O	-278
Isopropyl alcohol (2-propanol) (*l*)	C_3H_8O	-318
n-Propyl alcohol (*l*)	C_3H_8O	-303
Methyl t-butyl ether (MTBE) (*l*)	$C_5H_{12}O$	-314
Methyl oleate (*c*)	$C_{19}H_{36}O_2$	-735

Other: Carbon dioxide, $CO_2(g)$, −394 kJ/mol; Water vapor, $H_2O(g)$, −242 kJ/mol; Oxygen, $O_2(g)$, 0 kJ/mol

Notes: An “n” before a name indicates a “normal” straight chain structure (all of the carbons are linked in a row, with no branches). Generally assume that a structure is “n” unless some other designation is given, e.g., treat “hexane” as “n-hexane.”

Table TC.3: Major Components of Common Fuels

Fuel	**Components**	**Treat as**
Biodiesel	methyl esters of fatty acids	methyl oleate
Coleman fuel	n-hydrocarbons, average 8 C	n-octane
Diesel fuel	n-hydrocarbons, average 12 C	n-dodecane
Gasoline	hydrocarbon mix, average 8 C	isooctane
Kerosene	n-hydrocarbons, average 12 C	n-dodecane
Lamp oil	n-hydrocarbons, average 12 C	n-dodecane
Natural gas	low molecular mass hydrocarbons	methane

Review of Enthalpy of Formation Calculations

This topic is covered in your general chemistry text, but the essential equations are below:

$$\Delta H^\circ_{\text{reaction}} = \sum n_{\text{p}} \Delta H^\circ_{\text{f}}(products) - \sum n_{\text{p}} \Delta H^\circ_{\text{f}}(reactants) \qquad \text{(Equation TC.5)}$$

For these reactions, the $\Delta H_{\text{reaction}}$ corresponds to ΔH_{mol} for the combustion of your fuel (assuming you've balanced the chemical equation with 1 mole of fuel as the reactant). Add together the $\Delta H^\circ_{\text{f}}$ values for the products (CO_2 and H_2O), accounting for the moles of each, and then subtract the $\Delta H^\circ_{\text{f}}$ values for your fuel and the oxygen. For example, the balanced chemical equation for the combustion of benzene, C_6H_6, is:

$$C_6H_6(l) + 7.5\,O_2(g) \longrightarrow 6\,CO_2(g) + 3\,H_2O(g) \qquad \text{(Reaction TC.1)}$$

The calculation is:

$$\begin{aligned}
\Delta H_{\text{reaction}} &= \left(6 \times \Delta H^\circ_{\text{f}}(CO_2(g)) + 3 \times \Delta H^\circ_{\text{f}}(H_2O(g))\right) \\
&\quad - \left(1 \times \Delta H^\circ_{\text{f}}(C_6H_6(l)) + 7.5 \times \Delta H^\circ_{\text{f}}(O_2(g))\right) \\
&= (6 \times (-394) + 3 \times (-242)) - (1 \times (83) + 7.5 \times (0)) \\
&= -3173 \text{ kJ}
\end{aligned}$$

Keep in mind, in the calculations above, the stoichiometric coefficients (6, 3, 1, and 7.5) are actually moles and the $\Delta H^\circ_{\text{f}}$ values have units of kJ/mol, so the resulting unit is kJ. This is the ΔH value for one mole of the fuel.

Name: ______________________

Section: __________ Date: __________

Report Sheet:

Thermochemistry: Enthalpy of Combustion for Fuels

Standardization of Calorimeter with Hexane Burner

> Note: There are spaces here to record data for three runs of each burner. **Select your *two closest* runs to enter into Labflow.**

Report Table TC.1: Mass and Temperature Readings with Hexane Burner

	Run 1	**Run 2**	**Run 3**
Initial Mass Burner (g)			
Initial Temperature H_2O (°C)			
Final Temperature H_2O (°C)			
Final Mass Burner (g)			

Alcohol Fuel Data Collection

Number code for your alcohol: __________

> Note: There are spaces here for three runs of each burner. **Select your *two closest* runs to enter into Labflow.**

Report Table TC.2: Mass and Temperature Readings with Second Fuel Burner

	Run 1	**Run 2**	**Run 3**
Initial Mass Burner (g)			
Initial Temperature H_2O (°C)			
Final Temperature H_2O (°C)			
Final Mass Burner (g)			