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Calorimetry Lab Report

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1 Introduction

1.1 Purpose of Calorimetry

Calorimetry is a scientific experiment that analyses the heat exchange from an object (a system, process, substance, or solution, etc.) to the surroundings. The basis behind calorimetry derives from the first law of thermodynamics [1], which states that energy cannot be created or destroyed. Instead, energy is converted from one form to another.

In calorimetry, an isolated system is created from which an object can be inserted. The thermal energy emitted from the object is absorbed into the surrounding system until a thermal equilibrium is reached. The temperature changes are recorded, and the results can be used to calculate the energy change in the system. As this is an isolated system, it is assumed that all the heat energy from the object goes into the calorimeter and is not transferred out of the system. However, environmental losses will always occur to some degree, as making a completely efficient system is impossible.

Calorimetry is particularly useful for calculating the thermal energy exchange from a reaction, such as a chemical or combustion reaction. It is used across different scientific disciplines, such as environmental sciences, bioanalytical chemistry, and material analysis sciences.

1.2 Hypothesis

A calorimeter can accurately determine the enthalpy change of a heated strip of aluminium cooling in a body of water and the neutralisation reaction between HCl (Hydrochloric acid) and NaOH (Sodium Hydroxide), which should approximately be $-57.9 \text{ kJ mol}^{-1}$ [2].

1.3 Variables

The calorimetry measurements will be performed in three parts, as detailed in Appendix A. The first part will measure the enthalpy change of water in the calorimeter to calculate the heat capacity of the calorimeter. The second

measurement will calculate the specific heat capacity of a strip of aluminium, and the final measurement will calculate the molar heat of reaction of HCl and NaOH.

The dependent variable for these will be the temperature change in the calorimeter, measured in °C, which is used to calculate the heat released from the object using equation 1.

The controlled variables for the calorimetry process include the lab's temperature and pressure, the materials and build of the calorimeter, the solutions' molar concentration, and the mass of the aluminium sample.

2 Results

2.1 Calculations

2.1.1 Equations

Equation 1 is used to calculate the amount of heat energy lost or gained (Q) from a substance by multiplying the mass of the substance (m) in grams by the specific heat capacity of the substance (c) in J/kg·°C, and the change in temperature (ΔT) in °C.

$$Q = m \times c \times \Delta T$$

Equation 1

Equation 2 shows the heat lost by an object, in this case, an aluminium metal strip (Q_{metal}). According to the first law of thermodynamics, the thermal energy from the metal will be transferred to its surroundings, in this case, the water (Q_{water}) and the calorimeter ($Q_{calorimeter}$).

$$-(Q_{metal}) = Q_{water} + Q_{calorimeter}$$

Equation 2

Combining Equations 1 and 2 gives the following equation:

$$\begin{aligned} m_{metal} \times c_{metal} \times \Delta T_{metal} \\ = -(m_{water} \times c_{water} \times \Delta T_{water} + c_{calorimeter} \times \Delta T_{calorimeter}) \end{aligned}$$

Equation 3

For calculating the heat of reaction for a chemical process, Equations 2 and 3 can be combined, and the heat exchanged by the metal (Q_{metal}) is substituted by the heat exchanged by the reaction ($Q_{reaction}$):

$$Q_{reaction} = -(m_{water} \times c_{water} \times \Delta T_{water} + c_{calorimeter} \times \Delta T_{calorimeter})$$

Equation 4

Equation 5 is the same calculation as equation 2, but for hot water instead of aluminium.

$$-(Q_{hot-water}) = Q_{cold-water} + Q_{calorimeter}$$

Equation 5

Equation 5 is required to calculate the heat capacity of the calorimeter ($C_{calorimeter}$) shown in equation 6:

$$Q_{calorimeter} = c_{calorimeter} \times \Delta T_{calorimeter}$$

Equation 6

Equation 7 is used for calculating the molar heat of reaction ($q_{reaction}$), as the heat of reaction ($Q_{reaction}$) per mole of the limiting reactant ($n_{reactant}$):

$$q_{reaction} = \frac{Q_{reaction}}{n_{reactant}}$$

Equation 7

The number of moles of reactant ($n_{reactant}$) can be calculated using equation 8, where $V_{reactant}$ is the volume of the limiting reactant and $C_{reactant}$ is the concentration of the limiting reactant.

$$n_{reactant} = V_{reactant} \times C_{reactant}$$

Equation 8

Equations 7 and 8 can be combined to form the following equation:

$$q_{\text{reaction}} = \frac{Q_{\text{reaction}}}{V_{\text{reactant}} \times C_{\text{reactant}}}$$

Equation 9

2.1.2 Results and Calculations

2.1.2.1 Measuring heat capacity of the calorimeter

Table 1 – The measurements taken from the first part of the calorimetry, measuring the heat capacity of the calorimeter.

T_{hot}	$m_{\text{hot water}}$	T_{cold}	$m_{\text{cold water}}$	T_{max}	$\Delta T_{\text{hot water}}$	$\Delta T_{\text{cold water}}$	$\Delta T_{\text{calorimeter}}$
60°C	50g	24.3°C	50g	44.0°C	-16.0°C	19.7°C	19.7°C

$Q_{\text{hot water}}$ needs to be calculated first using equation 1.

$$Q_{\text{hot water}} = m \cdot c \cdot \Delta T = 50\text{g} \times 4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \cdot (-16.0^\circ\text{C})$$

$$Q_{\text{hot water}} = -3344\text{J}$$

Repeat for cold water.

$$Q_{\text{cold water}} = m \cdot c \cdot \Delta T = 50\text{g} \times 4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \cdot 19.7^\circ\text{C}$$

$$Q_{\text{cold water}} = 4117.3\text{J}$$

These can then be inserted into equation 5:

$$-(-3344J) = 4117.3J + Q_{calorimeter}$$

Rearrange the equation to solve.

$$Q_{calorimeter} = -773.3J$$

Using equation 6, we can calculate the heat capacity of the calorimeter.

$$-773.3J = c_{calorimeter} \times 19.7^{\circ}\text{C}$$

Rearrange to solve

$$c_{calorimeter} = \frac{-773.3J}{19.7^{\circ}\text{C}} \approx -39.25J/^{\circ}\text{C}$$

The heat capacity of the calorimeter is calculated to be a negative number, which isn't physically possible; this indicates that there has been an error in the process of the calorimetry. This will be discussed further in section 3.1; due to this, it would not be possible to calculate the rest of the measurements. Therefore, to overcome this issue, the heat capacity of the calorimeter will be recalculated using a new value for T_{max} , 41.33 °C. This value would give a corrected heat capacity value of 20.0 J/°C, which aligns more closely with expectations.

2.1.2.2 Measuring the specific heat capacity of a metal

Table 2 - measurements from the second part of the calorimetry

T_{metal}	m_{metal}	T_{water}	m_{water}	T_{max}	ΔT_{metal}	ΔT_{water}	$\Delta T_{calorimeter}$
110°C	18.55g	21.1°C	47.5g	27.2°C	-82.8°C	6.1°C	6.1°C

Table 3 – Measurements of the temperature of the calorimeter over time.

<i>Time (s)</i>	<i>T_{water} (°C)</i>	<i>Time (s)</i>	<i>T_{water} (°C)</i>	<i>Time (s)</i>	<i>T_{water} (°C)</i>
0	21.5	100	27.1	200	26.6
10	23.3	110	27.0	210	26.5
20	24.9	120	27.0	220	26.5
30	25.9	130	27.0	230	26.5
40	26.6	140	26.9	240	26.4
50	26.9	150	26.8	250	26.3
60	27.1	160	26.8	260	26.3
70	27.1	170	26.8	270	26.3
80	27.2	180	26.7	280	26.2
90	27.1	190	26.6	290	26.2

To calculate the heat capacity of the aluminium sample, we can input the measured numbers into equation 3 (and our corrected calorimeter heat capacity value).

$$18.55g \times C_{metal} \times (-82.8^{\circ}\text{C}) = -(47.5g \times 4.18 \frac{\text{J}}{\text{g} \cdot ^{\circ}\text{C}} \times 6.1^{\circ}\text{C} + 20.0\text{J}/^{\circ}\text{C} \times 6.1^{\circ}\text{C})$$

$$C_{metal} \approx 0.87\text{J}/\text{g} \cdot ^{\circ}\text{C}$$

2.1.2.3 Measuring the molar heat of reaction of HCl with NaOH

T_{start}	T_{end}	V_{HCl}	V_{NaOH}	C_{HCl}	C_{NaOH}	n_{HCl}	n_{NaOH}
19.2°C	26.8°C	50ml	52ml	1.0M	1.0M	0.050mol	0.052mol

Using equation 4 we can calculate the value for $Q_{reaction}$.

$$Q_{reaction} = -(102g \times 4.18 \frac{J}{g \cdot ^\circ C} \times [26.8^\circ C - 19.2^\circ C] + \frac{20.0J}{^\circ C} \times [26.8^\circ C - 19.2^\circ C])$$

$$Q_{reaction} = -3392.34J$$

Substituting this into equation 7 will provide us with the molar heat of reaction.

$$q_{reaction} = \frac{-3392.34J}{0.050mol} = -67.85kJ/mol$$

3 Discussion

3.1 Key Findings and Errors

The findings from the calorimetry measurements indicate a substantial error in either the method or a potential fault in the equipment. The heat capacity of any object cannot be less than 0, as this would indicate an object capable of inputting energy into a system without taking it from something else, which, according to the first law of thermodynamics, is impossible. The most likely cause of the error is faulty readings from the thermometer, whether from an actual problem with the equipment or human error from trying to read the thermometer during the experiment. A solution to this would be to perform the calorimetry with a thermocouple instead, minimising the chance of human error.

The results for the aluminium sample are very close to what would be expected, but the molar heat of reaction is a bit more exothermic than anticipated. This could be due to potential instrument error or an incorrect measurement of the solution's volume.

3.2 Bomb Calorimeters

In the OME industry, the most common use of calorimetry is with bomb calorimeters. These are high-pressure vessels that are designed to measure the heats of combustion. The “bomb” is a pressurised vessel that is submerged within a tank of water which is continually stirred. The bomb itself contains an oxygen atmosphere from which combustible matter, typically C,H,O,N explosive compounds, are ignited. The bomb calorimeter has been in use since the 19th century and is still in use in OME industries today as an effective tool for measuring the heat of combustion, it has also extended beyond OME to other fields such as environmental studies.

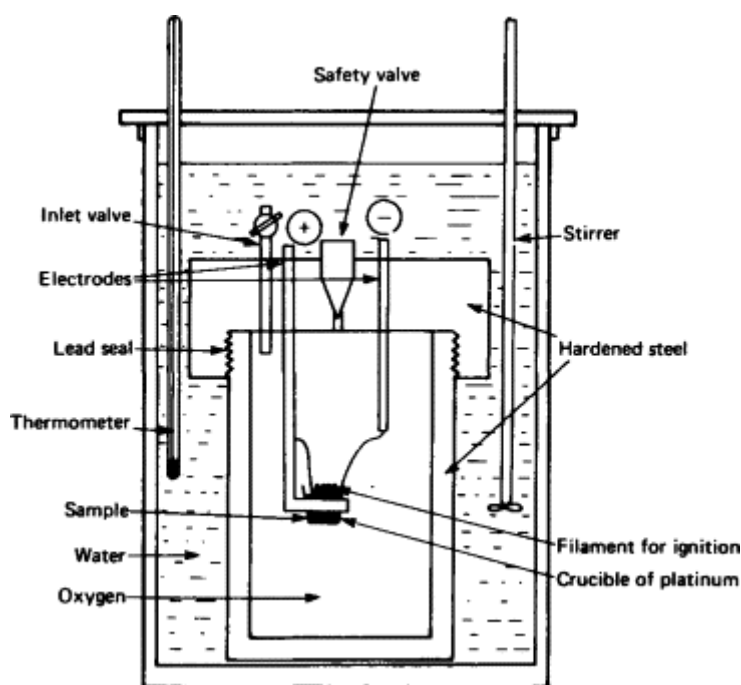


Figure 1 – [3] A diagram of a traditional bomb calorimeter.

4 Conclusion

The calorimeter is an effective tool for calculating heat capacity and the heat energy of reactions; however, with poor practices and faulty equipment, it can be very easy to acquire false or impossible results. Despite this, even a simple calorimeter has the potential to provide an accurate result, as shown with the aluminium sample. Eliminating human error where possible will reduce the

likelihood of acquiring incorrect data. Due to the error in calculating the heat capacity of the calorimeter, this calorimetry experiment did not meet the hypothesis; further experiments should be conducted to gain conclusive data.

References

1. K. S. Schmitz, Physical chemistry: Concepts and Theory. Elsevier, 2016.
2. "enthalpy change of neutralisation."
<https://www.chemguide.co.uk/physical/energetics/neutralisation.html>
3. J. O. Bird and P. J. Chivers, Newnes Engineering and Physical Science Pocket book. Butterworth-Heinemann, 1993.

Appendix A (Materials and Methodology)

A.1 Materials and Equipment

- 2 Styrofoam Cups and lid
- Thermometer
- Stirring Rod
- Clamp and Stand
- Glass Beaker
- Bunsen Burner
- Aluminium metal strip
- 1.0 M NaOH
- 1.0 M HCl

A.2 Methodology

The following methodology was rewritten from the provided data sheet [x].

1. A calorimeter was assembled using two Styrofoam cups, one placed inside the other; the lid had two holes punched through which were large enough to fit a stirring rod and a thermometer tightly.
2. 50.0 ml of water was added to the calorimeter, and the temperature was recorded (T_{cold}). Separately, another 50.0 ml of water was heated in a beaker till it reached a temperature of 60°C, the temperature of the hot water was recorded (T_{hot}).
3. The hot water from the beaker was added to the calorimeter, and the lid was sealed tightly. The water in the calorimeter was mixed thoroughly until the maximum equilibrium temperature was reached. The maximum temperature of the mixture was recorded (T_{max}), and the recorded temperatures can be used to calculate the calorimeter's heat capacity using equations 1, 5, and 6.
4. A sample of aluminium was weighed, recording its mass as m_{metal} .
5. A water bath was set up using a beaker and a Bunsen burner. The water bath was heated until boiling, and the aluminium sample was added. The water bath was left to boil for 5 minutes, allowing the aluminium to reach

thermal equilibrium with the boiling water. The temperature of the water bath was then recorded (T_{hot}).

6. As the water bath was being set up, 30.0 ml of water was added to the calorimeter, and its temperature was recorded (T_{cold}).
7. The aluminium sample was transferred from the water bath to the calorimeter, the mixture was stirred, and the temperature was recorded at 10-second intervals. The maximum temperature of the calorimeter was recorded as T_{max} .
8. The calorimeter and apparatus were reset, 52.0 ml of 1.0 M NaOH was placed in the calorimeter and 50.0 ml of HCl was poured into a beaker.
9. The two solutions were left for a few minutes until both reached room temperature, which was recorded as T_{start} . The HCl was then added to the NaOH in the calorimeter, and the maximum temperature reached during the neutralisation reaction is recorded as T_{end} .