

University of Wales Trinity Saint David

An Investigation into Calorimetry

Heat Capacity and Specific Heat

Jamie Hamilton-Swingewood (2308342)
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Introduction

In this experiment we determine the heat capacity of the calorimeter in addition to the specific heat capacity of an unknown metal sample. The primary hypothesis tested during this experiment is the first law of thermodynamics, also known as the law of conservation of energy. The law states that a systems internal energy (E) is equal to the difference of heat exchange (Q) and the work done (W) by the system^[1], for example a gas will expand or do work in addition to heating up when we give it external heat energy. In relation to the calorimetry experiment, this means that the energy lost from the hot piece of metal will be gained by the cooler water, and so the water heats up and metal sample cools. The calorimeter is to be considered ideal, in which it will lose no heat energy to its surroundings. From this the sub-hypothesis to be tested is that the specific heat capacity of our metal sample will be readily calculated by using equation {1.3} below where Q is the heat transferred, m is the mass of the substance, c is the specific heat capacity, and ΔT is the change in temperature.

$$\Delta Q = \Delta U + \Delta W \text{ {1.1}}$$

$$\Delta W = 0 \text{ (providing no work is done),} \quad \Delta Q = \Delta U = mc\Delta T \text{ {1.2}}$$

$$\therefore \Delta Q = mc_p\Delta T \text{ {1.3}}$$

Or more accurately it can be represented as the integral,

$$\int_{T_1}^{T_2} mc_p(T)dT = mc_p(T_2 - T_1), \quad \text{{1.4}}$$

The specific heat (c) of a substance is simply the heat capacity (C) per unit mass and is a measure of how insensitive matter is to the exchange of energy. Material with a higher specific heat will therefore require more energy for I give a mass to cause a particular temperature increase. In water for example, 1 kilogramme is raised by 1 degree Kelvin per 4186 joules of energy added provided that there is no phase transition occurring^[2]. It is worth noting that the heat capacity of a substance can change depending on pressure or volume. If a system is closed and has constant volume then its heat capacity will change as a function of internal energy with respect to temperature change of the system. Resulting in C_v for heat capacity at constant volume.

$$C_v = \frac{\partial U}{\partial T} \text{ {2.1}}$$

Therefore, we arrive at equation 1.3 again, $dU=C_vdT$, with some brief algebra. Likewise for heat capacity at constant pressure, C_p , follows the same form as constant volume but for an ideal gas is now a function of the change in enthalpy of a system with respect to temperature once more. This is because the change in heat dQ is equal to the change in enthalpy dH in uniform pressure.

$$C_p = \frac{\partial H}{\partial T} \text{ {2.2}}$$

Now we can define our systems energy at constant volume or pressure given the number of mols (n) of a substance^[3].

$$Q = n \int_{T_1}^{T_2} C_p dT, \text{ or } Q = n \int_{T_1}^{T_2} C dT \text{ {2.3}}$$

Results

Heat Capacity of Calorimeter.

The following are results from

$$\begin{aligned}T_{Hot} &= 60^{\circ}\text{C}, & m_{Hot} &= 50\text{g}, & \Delta T_{Hot} &= -21^{\circ}\text{C} \\T_{Cold} &= 23^{\circ}\text{C}, & m_{Cold} &= 50\text{g}, & \Delta T_{Cold} &= +16^{\circ}\text{C} \\T_{Max} &= 60^{\circ}\text{C}, & \Delta T_{Calorimeter} &= +16^{\circ}\text{C}\end{aligned}$$

Using eqn {1.3} we can determine first, the heat lost by the hot water, then the heat gained by the cold water and subsequently determine the heat capacity (H_c) of our calorimeter.

$$\Delta Q = mc\Delta T \text{ {1.3}}$$

$$\Delta Q_{Hot} = 50 \times 4.186 \times (-21) = -4395.3\text{J} \text{ {1.3.1}}$$

$$\Delta Q_{Cold} = 50 \times 4.186 \times 16 = 3348.9\text{J} \text{ {1.3.2}}$$

$$\Delta Q_{Calorimeter} = -\Delta Q_{Hot} - \Delta Q_{Cold} \text{ {1.3.3}}$$

$$\Delta Q_{Calorimeter} = 4395.3 - 3348.9 \text{ {1.3.4}}$$

$$\Delta Q_{Calorimeter} = 1046.4\text{J} \text{ {1.3.5}}$$

$$C_{Calorimeter} = \frac{\Delta Q_{Calorimeter}}{\Delta T_{Calorimeter}} = \frac{1046.4}{16} = 65.4\text{J/K}^{-1} \text{ {1.3.6}}$$

Specific Heat Capacity of Metal Sample.

$$\begin{aligned}T_{metal} &= 84^{\circ}\text{C}, & m_{metal} &= 18\text{g}, & \Delta T_{metal} &= -52^{\circ}\text{C} \\T_{water} &= 24^{\circ}\text{C}, & m_{water} &= 30\text{g}, & \Delta T_{water} &= +8^{\circ}\text{C} \\T_{Max} &= 32^{\circ}\text{C}, & \Delta T_{Calorimeter} &= +8^{\circ}\text{C}\end{aligned}$$

Again using equation 1.3,

$$Q_{Water} = 30 \times 4.186 \times 8 = 1004.6\text{J} \text{ {1.3.7}}$$

$$\Delta Q_{Cal} = C_{Cal} \times \Delta T_{water} = 65.4 \times 8 = 523.2\text{J} \text{ {1.3.8}}$$

$$-\Delta Q_{metal} = \Delta Q_{water} + \Delta Q_{Cal} \text{ {1.3.9}}$$

$$\Delta Q_{metal} = -1004.6 - 523.2 = -1527.8\text{J} \text{ {1.3.10}}$$

$$c_{Pmetal} = \frac{\Delta Q_{metal}}{m\Delta T_{metal}} \text{ {1.3.11}}$$

$$c_{Pmetal} = \frac{-1527.8}{18 \times -52} \text{ {1.3.12}}$$

$$c_{Pmetal} = 1.63\text{J/g}^{-1}\text{K}^{-1}$$

Discussion

from the experimental results we can see that a heat transfer has occurred between our samples of water in the first instance allowing us to determine the heat capacity of the calorimeter. In the second instance heat was absorbed by the water from the metal, causing the temperature of the metal to cool and the water to increase until they found equilibrium at 32°C and so the specific heat capacity of the metal sample was found to be $1.63\text{Jg}^{-1}\text{K}^{-1}$. The crux of this experiment lies in the application of the first law of thermodynamics in that the energy of the system cannot be created or destroyed and so it must be transferred either from one form to another or from one substance to another. In analysing this heat exchange we have been able to satisfy the previously outlined hypothesis gaining insight to the thermodynamic properties of the metal sample. Although it was assumed that the calorimeter was ideal in practicality there will have been heat loss to the surroundings this would be compounded by the inaccuracies of measuring with a thermometer by eye as a source of potential error and is an intrinsic uncertainty that one must be aware of. These factors can be mitigated by using more sophisticated equipment such as wireless digital thermometers or indeed a closed bomb (or vessel) calorimeter. Completely sealing off the heat exchange and not needing a human to read a thermometer thus heavily mitigating these potential errors.

Relevance in OME

Calorimetry is used widely throughout the study of OME and is particularly prevalent in assessing the power of novel research explosives or identification unknown samples of explosives. By measuring the energy released during a deflagration or detonation reaction calorimetry can provide the necessary quantitative data to inform scientists and engineers of the characteristics of an explosive. Thus, allowing for desired explosives to be tuned for specific applications. Calorimetry is also used to assess the quality, safety and state of decomposition of an explosive, as when degraded or mixed with binders the thermal properties of a given substance can be vastly different from its pure unaltered state. Particularly of note in industry is closed bomb calorimetry, which is a closed vessel of constant volume. In which a combustion or detonation reaction will occur inside the vessel. The energy is then released via highly conductive walls of the 'closed bomb' to heat up a known mass of water and this change in temperature is very accurately measured allowing for the calculation of energy released^[4].

Conclusion

In conclusion the experimental findings support the hypothesis laid out within the framework of the first law of thermodynamics. That is the heat gained by the water is equal to that of the heat lost by the metal and so energy is conserved, the laws of thermodynamics and fabric of the universe remain intact.

The heat capacity of the calorimeter was found to be 65.4J K^{-1} .

The specific heat capacity of the metal sample was found to be $1.63\text{Jg}^{-1}\text{K}^{-1}$.

Future research

To use a wider array of materials and masses to see the impact that other substances have on the transfer of energy. Will highly ordered materials like ionic crystals or semiconductors have more favourable heat transfer properties to say polymers?

Bibliography

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- [3]A. A. Carter, Chemical Thermodynamics. Meredith College: LibreTexts, 2020, pp. 12.1.8 to 12.1.13.
- [4]E. Adriana Camarillo and H. Flores, “Determination of the energies of combustion and enthalpies of formation of nitrobenzenesulfonamides by rotating-bomb combustion calorimetry,” Journal of Chemical Thermodynamics/The journal of chemical thermodynamics, vol. 42, no. 3, pp. 425–428, Mar. 2010, doi: <https://doi.org/10.1016/j.jct.2009.10.003>.

Appendix

Materials and method

- Styrofoam cup calorimeter (known as the cup)
 - Water
 - Thermometer
 - Digital scale
 - Metal sample
 - Stirring rod
 - Glass beaker
 - Heat source
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1. 50.0 ml of water added to the cup and its temperature taken.
 2. 50ml of water added to beaker and heated to 80°C.
 3. When the hot water reached 60°C it was added to the cup.
 4. The water mix was then stirred thoroughly and allowed to briefly rest as the temperature came to equilibrium.
 5. The equilibrium temperature was then recorded.
 6. Using equation 1.3 and 1.3.6 the Heat Capacity of the cup was calculated.
 7. Now the metal sample was weighed, and this was recorded.
 8. A hot water bath was set up and the metal sample was placed in the hot water bath.
 9. The hot water bath was maintained near boiling for 5 minuets and then the heat was taken off for a further 2 minutes.
 10. Now 30.0ml of water was added to the empty cup and its temperature recorded.
 11. The temperature of the metal and hot water bath should be the same so record the temperature of the hot water and quickly add the metal into the calorimeter.
 12. Measure the temperature of the water in the calorimeter as heat is exchanged between it and the metal.
 13. The water and metal sample in the cup will eventually come to equilibrium when this was reached the temperature was recorded.
 14. Using equation 1.3 and 1.3.12 the specific heat capacity of the metal was calculated.