

Effect of Molarity on Specific Heat Capacity

Introduction

While cooking pasta, my mother always adds a pinch of salt to the water. Initially, I assumed that she added salt to flavor the water so that it can be absorbed by the pasta. However, while studying states of matter, I realized another reason for the addition of salt. Salt increases the boiling point of water, which results in the water having a higher temperature and allows the pasta to cook better. The increased boiling point is because of boiling point elevation, which occurs when solute particles take up the required surface area for boiling of the solvent to take place.

Based on a general trend, the boiling point is seen to be inversely proportional to the specific heat capacity. Water and iron are examples which fit this trend. Therefore, it can be said that the addition of salt also results in decreasing the specific heat capacity of water, and this trend would follow as more amount of salt is added.

This situation made me ponder whether a similar pattern exists with other salts and leads to my research question, “**How does the molarity of magnesium sulfate heptahydrate solution ($MgSO_4 \cdot 7H_2O(aq)$) affect its specific heat capacity?**”. Magnesium sulfate was chosen over other salts because of its comparatively high solubility.

Background Information

Specific Heat Capacity

Specific heat capacity is defined as the amount of heat energy required to raise the temperature of a unit mass of a substance by one kelvin or degree celsius. The following formula is used for the calculation of the specific heat capacity:

$$Q = mc\Delta T \dots\dots\dots \text{Equation 1}$$

(where Q is heat energy in joules [J], m is mass in grams [g], c is specific heat capacity in joules per grams and celsius [$J \cdot g^{-1} \cdot ^\circ C^{-1}$], specific heat capacity of water is $4.1819 J \cdot g^{-1} \cdot ^\circ C^{-1}$, and ΔT is temperature change in celsius [$^\circ C$])

Specific heat capacity is a physical property and is also defined as the measure of heat energy stored by material per kilogram and kelvin or celsius. Therefore, specific heat capacity is seen to be directly proportional to the amount of heat energy stored.

A factor which affects the specific heat capacity of a substance is the bonding between the particles of a substance. Since specific heat capacity is the amount of heat energy required to raise the temperature of a substance, a stronger bond between the particles would result in a higher value of specific heat capacity as larger amount of energy would be required to increase the average random kinetic energy of the particles of the substance (increase the temperature).

Specific Heat Capacity of Ionic Compounds

Water (H_2O), a polar solvent, consists of two slightly positive hydrogen atoms and a slightly negative oxygen atom due to the high difference in electronegativity between the two elements. Therefore, when ionic compounds are dissolved in water, the individual ions of the ionic compound dissociate and are attracted to the positive and negative hydrogen and oxygen atoms respectively.

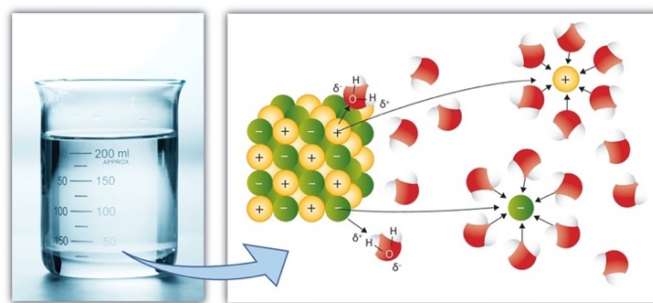


Figure 1:- Dissociation of ionic compounds in water¹

Molarity

Molarity is defined as the number of moles of solute per liter of solution.² Molarity is also known as molar concentration and is seen as a way to express the concentration of a solution:

$$\text{Concentration} = \frac{\text{Number of moles}}{\text{Liter of solution}}$$

Molarity is represented as M with units moles per liter (mol. l^{-1}).

Hypothesis

I hypothesize that as the molarity of the magnesium sulfate heptahydrate solution would increase, its specific heat capacity would decrease. This is because given that the chosen salt (magnesium sulfate) is an ionic compound, it would dissociate in water and form dipole-dipole bonds with the hydrogen and oxygen atoms. As more salt is added, more of these dipole-dipole bonds would form and would result in the breaking of the hydrogen bonds between water molecules. Since hydrogen bonds are stronger than the dipole-dipole bonds and specific heat capacity is the amount of heat energy required to raise the temperature (increase the average random kinetic energy) of a substance, weaker bonds would result in a lower specific heat capacity.

Methodology

To carry out this experiment, first, different molar solutions of magnesium sulfate heptahydrate are created. Then, using the method of mixture, the specific heat capacity of each solution is determined. A polystyrene cup, covered with a thermocol lid on the top, is used in the method of mixture to ensure minimum heat loss to the surroundings. In the end, a graph displaying the results is plotted.

¹ Ball, David et al. Introduction To Chemistry: General, Organic, And Biological. 1st ed., Unnamed Publisher, 2011, pp. 9.3. Accessed 15 Oct 2019.

² "Molarity, Molality, Or Normality? (A Quick Review)". *Carolina.Com*, 2018, <https://www.carolina.com/teacher-resources/Interactive/molarity-molality-or-normality-quick-review/tr10800.tr>. Accessed 19 Oct 2019.

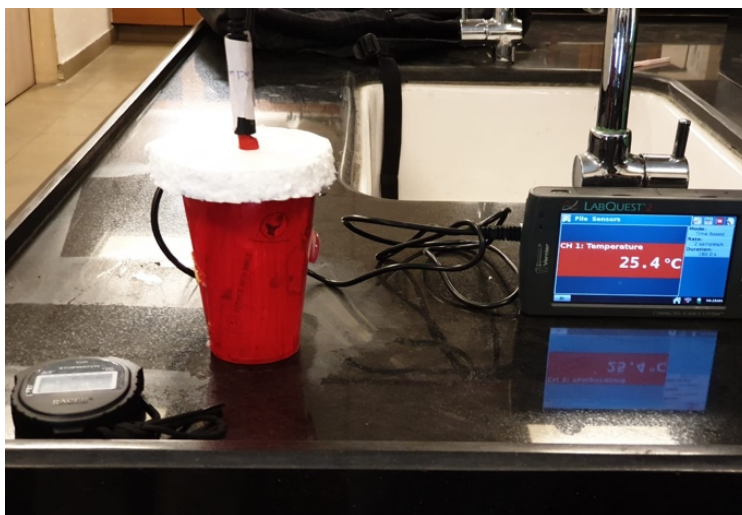


Figure 2:- Experimental setup

Variables

Independent Variable

Molarity of magnesium sulfate heptahydrate solution ($M/mol.l^{-1}$): The molarity is measured in moles per liter ($mol.l^{-1}$) and is varied by dissolving different amounts of magnesium sulfate in 150 g of distilled water. Dissolving different amounts of salt results in different number of moles of that salt and ultimately changes the molarity of the solution. A top pan balance with an uncertainty of $\pm 0.01 g$ will be used to measure the mass of magnesium sulfate. The following molarity values are investigated: 0.2 M, 0.4 M, 0.6 M, 0.8 M, 1.0 M, 1.2 M.

Dependent Variable

Specific heat capacity of magnesium sulfate heptahydrate solution ($c_{mg}/J.g^{-1}.^{\circ}C^{-1}$): Rearranging equation 1,

$$c = \frac{Q}{m\Delta T} \dots \dots \dots \text{Equation 2}$$

Equation 2 is used to calculate the specific heat capacity of the magnesium sulfate solution. Mass of the solution will be measured using a Top Pan Balance and the change in temperature will be measured in celsius ($^{\circ}C$) using a temperature probe connected to a data logger with an uncertainty of $\pm 0.1^{\circ}C$.

Control Variables

1. Surrounding temperature: Heat is known to transfer from systems with higher temperature to systems with lower temperature. Therefore, given that the solution would have a lower temperature than the surroundings, heat may flow from the surroundings to the solution. This transfer of heat would ultimately increase the temperature of the solution and result in an anomaly in the calculated specific heat capacity. Now, as the heat transfer rate from the surroundings to the solution is dependent on the temperature difference between the two, a possible method to negate the effect of such heat transfer is to ensure constant temperature of the surroundings. To this end, the experiment was conducted in the lab with closed windows, and the surrounding temperature was measured with a temperature probe and found to be $27.9^{\circ}C$.

2. Water purity: Specific heat capacity of water is affected by factors such as the existence of impurities, minerals, etc. To ensure that the specific heat capacity of the magnesium sulfate solution is measured accurately, distilled water was employed for all trials.
3. Duration of each trial: To ensure accuracy in the calculated specific heat capacity, the time taken for each solution to reach thermal equilibrium should be controlled. Therefore, the duration of each trial (after mixing) was set as 40 s.
4. Mass of distilled water and solution: Varying mass of distilled water and magnesium sulfate solution for each trial would result in differing amounts of heat energy required to raise the temperature. Furthermore, given that the duration of each trial is held constant, different quantities of water and magnesium sulfate solution for each trial would make it unfeasible to calculate the specific heat capacity. Therefore, for each trial, the mass of distilled water and magnesium sulfate heptahydrate solution was taken as 30 g.
5. Specific heat capacity of the polystyrene cup: Since the magnesium sulfate heptahydrate solution would have a higher temperature than distilled water, there would be an increase in the temperature of water when it is mixed with the solution. In such a case, if the specific heat capacity of the polystyrene cup is not controlled, the heat absorbed by the cup would vary which would ultimately affect the specific heat capacity of the magnesium sulfate solution. By ensuring that the same polystyrene cup is used, its specific heat capacity was controlled.
6. Pressure: A change in pressure conditions affects the specific heat capacity of a substance. Therefore, if pressure is not controlled, the values of specific heat capacity of differing concentrations of magnesium sulfate solution would be affected by a variable (pressure) which is not the focus of the investigation. Pressure was controlled by conducting the experiment in a lab with constant ambient temperature and shut windows.

List of Apparatus

1. Cutter
2. Distilled water
3. Erlenmeyer flask ($\pm 0.05 \text{ cm}^3$)
4. Magnesium sulfate heptahydrate
5. Magnetic stirrer
6. Polystyrene cup
7. Spatula
8. Stopwatch ($\pm 0.01 \text{ s}$)
9. Temperature Probe attached to Data Logger, LabQuest2 $\times 2$ ($\pm 0.1^\circ\text{C}$)
10. Thermocol sheet
11. Top pan balance ($\pm 0.01 \text{ g}$)
12. Watch glass

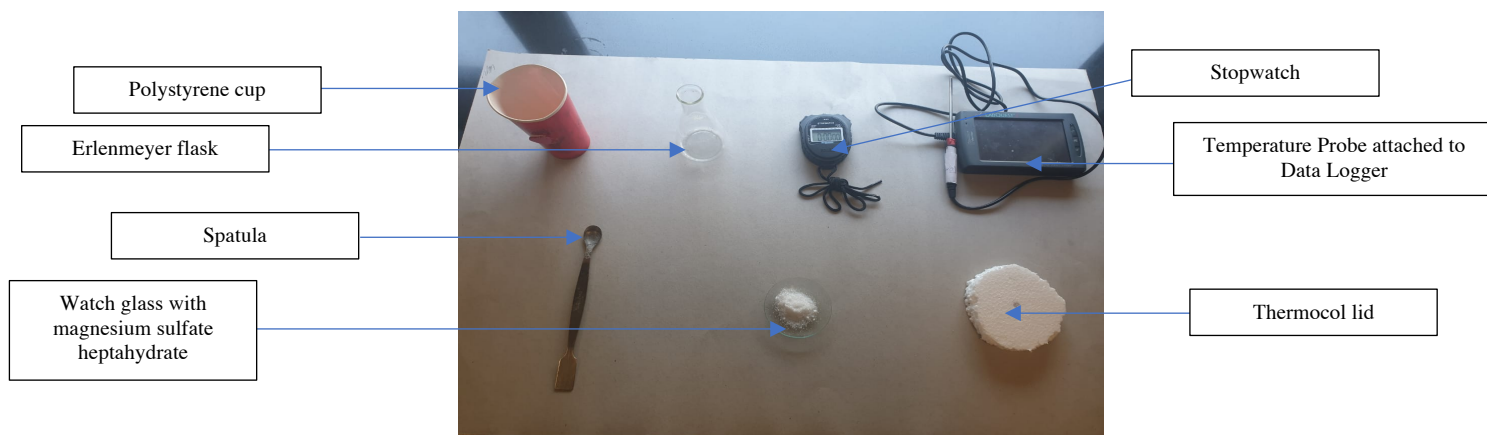


Figure 3:- Apparatus used in the experiment

Safety Precautions

1. Wear appropriate safety apparel such as safety glasses, masks, and gloves as magnesium sulfate is seen as an eye and respiratory irritant. Furthermore, magnesium sulfate is seen to be harmful if absorbed through the skin, hence it is important to wear a lab coat while conducting the experiment.
2. Immediately clean spilt magnesium sulfate heptahydrate solution as it may act as a skin irritant and an unnecessary nuisance.

Procedure

Preparing appropriate molar magnesium sulfate heptahydrate solutions

1. The top pan balance is set to zero and 150 g of distilled water is taken in an erlenmeyer flask kept on top of the balance.
2. To create a 0.2 M solution of magnesium sulfate heptahydrate, the following calculations are done to determine the appropriate mass of magnesium sulfate to be used

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution}} \dots \dots \dots \text{Equation 3}$$

Now, density is expressed as

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \dots \dots \dots \text{Equation 4}$$

Equation 4 can be rearranged to obtain volume of solution in equation 3

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

As we have taken 150 g of distilled water and the density of distilled water at 20°C is 0.99823 g.ml⁻¹

$$\text{Volume} = \frac{150}{0.99823} = 150.26 \text{ ml} \approx 150 \text{ ml}$$

Now, number of moles of a substance can be expressed as

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}} \dots\dots\dots \text{Equation 5}$$

Equation 5 can be rearranged and substituted in equation 3 to obtain the amount of magnesium sulfate for a 0.2 M solution

$$0.2 = \frac{\text{Mass} / \text{Molar Mass}}{150} \dots\dots\dots \text{Equation 6}$$

Rearranging equation 6 and substituting the molar mass of magnesium sulfate

$$\text{Mass} = \frac{0.2 \times 150 \times 246.47}{1000} = 7.39 \text{ g} \approx 7 \text{ g}$$

(we divide by 1000 to convert ml to l)

3. After setting the top pan balance to zero, the amount of magnesium sulfate obtained in step 2 is measured by transferring the salt to a watch glass kept on top of the balance. A spatula is used to transfer the salt from the container to the watch glass.
4. Steps 1-3 are then repeated to prepare the remaining molar solutions of magnesium sulfate heptahydrate (0.4 M, 0.6 M, 0.8 M, 1.0 M, 1.2 M).

Creating the thermocol lid

1. A thermocol sheet is taken and the polystyrene cup is placed inverted on-top of the sheet.
2. Using a pencil, a circle, slightly bigger than the rim of the cup, is drawn on the thermocol sheet.
3. Using a cutter, the circle drawn on the thermocol sheet is cut.
4. To ensure that the temperature probe can pass through the lid and be immersed in the solution, an appropriately sized hole was cut for the temperature probe in the thermocol lid using a cutter.

Method of mixture

1. The temperature probe is connected to the data logger and is setup.
2. The top pan balance is set to zero and 30 g of cool distilled water (around 12°C) is taken in the polystyrene cup kept on top of the balance.
3. The temperature probe is then placed in the polystyrene cup to measure the temperature of water. This is done to ensure that the temperature of distilled water is constant before the addition of the magnesium sulfate heptahydrate solution. The final and initial temperatures of the distilled water are recorded.
4. After setting the balance to zero, 30 g of 0.2 M magnesium sulfate heptahydrate solution is taken in an erlenmeyer flask kept on top of the top pan balance. The temperature of the solution is then measured with another temperature probe for 40 s and recorded.
5. Next, the solution obtained in the previous step is added to distilled water in the polystyrene cup and the cup is covered with the thermocol lid.
6. The temperature probe is inserted through the hole in the thermocol lid and the temperature is then measured for 40 s. The final temperature (when thermal equilibrium is attained) is noted.

7. Steps 2-6 are repeated for the remaining solutions. For each concentration, the specific heat capacity is calculated five times to reduce the effect of random error on the outcome of the investigation.

Data Collection and Processing

Raw Data

Molarity, $M/\text{mol. l}^{-1}$	Temperature, $T/^{\circ}\text{C}$ ($\pm 0.1^{\circ}\text{C}$)											
	Trial 1				Trial 2				Trial 3			
	T_{wi}	T_{wf}	T_{mgi}	T_{mgf}	T_{wi}	T_{wf}	T_{mgi}	T_{mgf}	T_{wi}	T_{wf}	T_{mgi}	T_{mgf}
0.2	13.7	20.3	25.2	18.1	11.6	17.8	24.9	18.3	13.9	20.4	24.3	18.0
0.4	10.9	16.7	25.5	19.1	11.4	17.1	25.7	19.9	12.3	18.3	25.3	18.9
0.6	12.2	18.3	24.7	18.5	12.7	19.1	25.7	18.9	12.5	18.2	24.1	17.8
0.8	11.9	19.0	25.9	17.8	12.9	19.5	25.4	18.1	13.1	19.4	25.2	17.9
1.0	11.2	18.0	25.6	18.1	11.8	17.7	26.0	19.1	12.4	18.6	25.7	18.4
1.2	13.5	19.9	25.3	17.8	11.9	18.5	25.1	17.2	13.4	19.3	24.7	17.6

Molarity, $M/\text{mol. l}^{-1}$	Temperature, $T/^{\circ}\text{C}$ ($\pm 0.1^{\circ}\text{C}$)							
	Trial 4				Trial 5			
	T_{wi}	T_{wf}	T_{mgi}	T_{mgf}	T_{wi}	T_{wf}	T_{mgi}	T_{mgf}
0.2	11.7	17.8	24.9	18.9	12.6	19.1	24.3	17.8
0.4	12.9	19.9	25.1	17.8	13.0	18.9	25.4	19.5
0.6	11.9	18.1	24.4	18.0	12.1	19.0	24.2	16.3
0.8	13.1	19.5	25.7	18.7	12.3	18.5	25.2	18.5
1.0	10.9	16.8	26.1	19.3	12.6	19.2	25.9	18.4
1.2	12.3	18.6	24.3	17.0	12.4	18.2	24.9	17.9

Table 1:- Raw data table displaying temperature change of five trials

*In table 1, T_{wi} and T_{wf} represent the initial and final temperature of distilled water respectively. Similarly, T_{mgi} and T_{mgf} represent the temperature of the magnesium sulfate solution before and after being added to distilled water respectively.

Data Processing

Apparatus Uncertainty

Apparatus	Uncertainty
Temperature Probe connected to Data Logger (LabQuest2)	$\pm 0.1^{\circ}\text{C}$
Top pan balance	$\pm 0.01 \text{ g}$

Table 2:- Uncertainties of apparatus used

Determining specific heat capacity of magnesium sulfate heptahydrate solution

Trial 1 of 0.2 M has been used as an example.

Temperature change of distilled water (ΔT_w) can be calculated as

$$T_{wf} - T_{wi}$$

$$\Delta T_w = 20.3 - 13.7 = 6.6^\circ\text{C}$$

Using uncertainty of the temperature probe, we can calculate the uncertainty of ΔT_w

$$\begin{aligned}\pm\Delta T_w &= 0.1 + 0.1 = \pm 0.2^\circ\text{C} \\ \therefore \Delta T_w &= 6.6 \pm 0.2^\circ\text{C}\end{aligned}$$

Using *equation 1*, heat energy gained by distilled water (Q_w) can be calculated as

$$Q_w = m_w \cdot c_w \cdot \Delta T_w$$

(where m_w represents the mass of distilled water)

$$Q_w = \frac{4.1819 \times 30.00 \times 6.6}{1000} = 0.828 \text{ kJ}$$

(we divide by 1000 so as to convert from J to kJ)

Using the uncertainty of mass and temperature, as determined above, we can determine the uncertainty in the heat gained by distilled water

$$\begin{aligned}\pm Q_w &= \left(\frac{0.2}{6.6} + \frac{0.01}{30.00} \right) \times 0.828 = \pm 0.025 \text{ kJ} \approx \pm 0.02 \text{ kJ} \\ \therefore Q_w &= 0.83 \pm 0.02 \text{ kJ}\end{aligned}$$

Temperature change of the magnesium sulfate heptahydrate solution (ΔT_{mg}) can be calculated as

$$\begin{aligned}\Delta T_{mg} &= T_{mgi} - T_{mgf} \\ \Delta T_{mg} &= 25.2 - 18.1 = 7.1^\circ\text{C}\end{aligned}$$

Using the uncertainty of the temperature probe, we can calculate the uncertainty of ΔT_{mg}

$$\begin{aligned}\pm\Delta T_{mg} &= 0.1 + 0.1 = \pm 0.2^\circ\text{C} \\ \therefore \Delta T_{mg} &= 7.1 \pm 0.2^\circ\text{C}\end{aligned}$$

From First Law of Thermodynamics

Heat gained by distilled water (Q_w) = Heat lost by magnesium sulfate solution (Q_{mg})

$$\begin{aligned}\therefore Q_{mg} &= 0.83 \pm 0.02 \text{ kJ} = m_{mg} \cdot c_{mg} \cdot \Delta T_{mg} \\ \text{(where } m_{mg} \text{ represents the mass of magnesium sulfate solution)}\end{aligned}$$

Using values from above

$$0.83 = 30.00 \times c_{mg} \times 7.1$$

$$\begin{aligned}\frac{0.83}{30.00 \times 7.1} \times 1000 &= c_{mg} = 3.90 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1} \\ \text{(Multiply by } 1000 \text{ to write in } J \text{ from } kJ)\end{aligned}$$

Using uncertainty values of Q_{mg} , m_{mg} , ΔT_{mg} to calculate uncertainty of c_{mg}

$$\pm c_{mg} = \left(\frac{0.02}{0.83} + \frac{0.01}{30.00} + \frac{0.2}{7.1} \right) \times 3.90 = \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

$$\therefore c_{mg} = 3.9 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

Since the specific heat capacity is calculated for 5 trials, we can calculate the average specific heat capacity (c_{mgavg})

$$c_{mg2} = 3.9 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

$$c_{mg3} = 4.3 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

$$c_{mg4} = 4.2 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

$$c_{mg5} = 4.2 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

The average then becomes,

$$\frac{3.9 + 3.9 + 4.3 + 4.2 + 4.2}{5} = 4.1 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

Using the maximum and minimum value of the specific heat capacities, we can calculate the uncertainty in c_{mgavg}

$$\pm c_{mgavg} \frac{4.3 - 3.9}{2} = \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

$$\therefore c_{mgavg} = 4.1 \pm 0.2 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$$

The above calculations are repeated for the remaining molar solutions of magnesium sulfate.

Processed Data

Molarity, M/mol.l ⁻¹	Trial 1		Trial 2	
	Specific Heat Capacity $c_{mg}/\text{J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$	Uncertainty of Specific Heat Capacity $\pm c_{mg}/\text{J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$	Specific Heat Capacity $c_{mg}/\text{J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$	Uncertainty of Specific Heat Capacity $\pm c_{mg}/\text{J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$
0.2	3.9	0.2	3.9	0.2
0.4	3.8	0.2	4.1	0.2
0.6	4.1	0.2	3.9	0.2
0.8	3.66	0.17	3.8	0.2
1.0	3.78	0.19	3.6	0.2
1.2	3.56	0.18	3.50	0.17

Molarity, M/mol.l ⁻¹	Trial 3		Trial 4	
	Specific Heat Capacity $c_{mg}/J.g^{-1}.^{\circ}C^{-1}$	Uncertainty of Specific Heat Capacity $\pm c_{mg}/J.g^{-1}.^{\circ}C^{-1}$	Specific Heat Capacity $c_{mg}/J.g^{-1}.^{\circ}C^{-1}$	Uncertainty of Specific Heat Capacity $\pm c_{mg}/J.g^{-1}.^{\circ}C^{-1}$
0.2	4.3	0.2	4.2	0.2
0.4	3.9	0.2	4.0	0.2
0.6	3.8	0.2	4.1	0.2
0.8	3.61	0.19	3.8	0.2
1.0	3.56	0.19	3.6	0.2
1.2	3.47	0.19	3.61	0.19

Molarity, M/mol.l ⁻¹	Trial 5		Average Specific Heat Capacity $c_{mgavg}/J.g^{-1}.^{\circ}C^{-1}$	Uncertainty of Average Specific Heat Capacity $\pm c_{mgavg}/J.g^{-1}.^{\circ}C^{-1}$
	Specific Heat Capacity $c_{mg}/J.g^{-1}.^{\circ}C^{-1}$	Uncertainty of Specific Heat Capacity $\pm c_{mg}/J.g^{-1}.^{\circ}C^{-1}$		
0.2	4.2	0.2	4.1	0.2
0.4	4.2	0.2	4.0	0.2
0.6	3.4	0.2	3.9	0.4
0.8	3.9	0.2	3.75	0.15
1.0	3.69	0.19	3.65	0.11
1.2	3.5	0.2	3.53	0.07

Table 3:- Processed data table displaying specific heat capacity of five trials and the average specific heat capacity

Graphing Data

From table 3, the graph of molarity and its respective average specific heat capacity can be plotted.

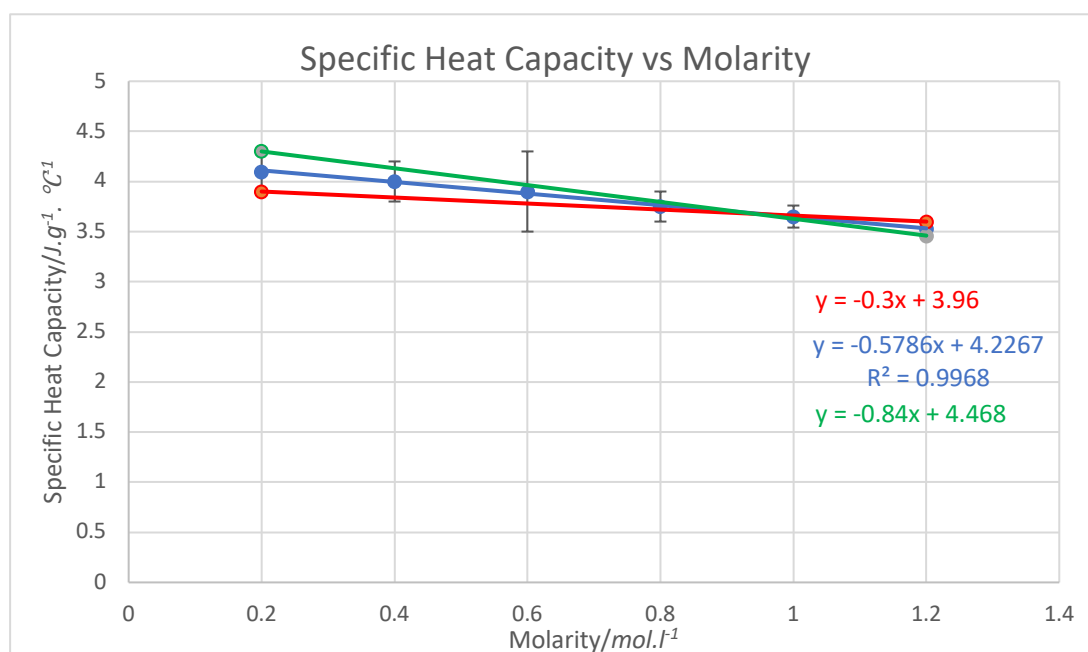


Figure 4:- Graph displaying average specific heat capacity vs molarity where blue represents the best-fit line, red represents the max-slope, and green indicates the min-slope

Error Calculation from Graph

Slope Uncertainty:

$$\text{Slope Uncertainty } (\pm s) = \frac{|Max\ Slope - Min\ Slope|}{2}$$
$$\pm s = \frac{|-0.3 - (-0.84)|}{2} = \pm 0.27\ J.l.g^{-1}.\text{ }^{\circ}\text{C}^{-1}.mol^{-1} \approx \pm 0.3\ J.l.g^{-1}.\text{ }^{\circ}\text{C}^{-1}.mol^{-1}$$

Accuracy:

The methodology of the investigation was employed to determine the specific heat capacity of water. The obtained value was then compared to the literature value of specific heat capacity of water to determine the accuracy of the investigation.

$$\text{Percent Error} = \frac{|Experimental\ value - Literature\ value|}{Literature\ Value} \times 100$$
$$\text{Percent Error} = \frac{|4.3 - 4.1819|}{4.1819} \times 100 = 2.82\%$$

Conclusion and Evaluation

Conclusion

From the gathered and processed data, it can be seen that the hypothesis holds to be valid. *Figure 4* shows that as the molarity of magnesium sulfate (independent variable) increases, its specific heat capacity (dependent variable) decreases. That is, there exists a negative correlation between the independent and dependent variables. Using the slope of the graph displayed in *figure 4*, the change in the average specific heat capacity for a corresponding change of 0.1 M in the molarity can be determined as follows

$$\frac{\Delta y}{\Delta x} = \frac{\Delta c_{mgavg}}{\Delta M} = -0.5786 \text{ (as slope of the best fit line in figure 4 is } -0.5786\ J.l.g^{-1}.\text{ }^{\circ}\text{C}^{-1}.mol^{-1})$$
$$\frac{\Delta c_{mgavg}}{0.1} = -0.5786 \text{ (as calculating for an increase in molarity by 0.1 M)}$$
$$\Delta y = -0.058\ J.l.g^{-1}.\text{ }^{\circ}\text{C}^{-1}$$

Hence, the specific heat capacity of the magnesium sulfate solution reduces by 0.058 J.l.g⁻¹.°C⁻¹ for an increase in molarity of 0.1 M. However, since the graph does not pass through the origin, a directly proportional relationship with a negative constant of proportionality does not exist. Given that a high correlation exists between the independent and dependent variable ($R^2 = 0.9968$), the aim of the investigation to determine a relation between the molarity of magnesium sulfate and its specific heat capacity has been fulfilled.

Evaluation

Despite the aim of the investigation being fulfilled, there were random and systematic errors in the experiment which limited its scope. Vertical error bars in the graph displayed in *figure 4*, graph not passing through the origin, and percentage uncertainty of 2.82% act as evidence for the existence of systematic and random errors.

Systematic Errors

1. Amount of data: In the investigation, only six molar solutions of magnesium sulfate were tested. Testing a smaller number of molarities increases the impact outliers could have on the outcome of the investigation and lead to the prediction of an inaccurate correlation. By involving a larger number of molarities in the investigation, a more accurate correlation could be predicted and further investigated.
2. Loss of thermal energy to surroundings: Despite the polystyrene cup being covered with a thermocol lid, there was some loss of heat energy to the surroundings. The reason for this loss

is the hole which was made in the thermocol lid to insert the temperature probe. This energy loss is evident from the temperature difference between the surroundings and the temperature of the solution. This error results in an underestimation of the thermal energy and overestimation of the specific heat capacity. The effect of this error on the investigation was explored by placing 30 g of distilled water at 13.9°C, in the lab with a room temperature of 27.9°C for 100 s. Due to the transfer of heat energy, the final temperature of distilled was recorded to be 14.1°C. As the obtained temperature change was quite less compared to the temperature change obtained in the investigation, the effect of this error was deemed to be small. Regardless, the effect of this error could be reduced by cutting a more appropriate hole in the lid and inserting the temperature probe through the hole by covering it with woolen insulation to prevent loss of thermal energy through the hole.

3. Temperature Probe: The temperature probe used was one of the major contributors of uncertainty to the investigation. Despite the temperature being recorded to an uncertainty of $\pm 0.1^\circ\text{C}$, propagation of uncertainties increased the uncertainty for the temperature change to $\pm 0.2^\circ\text{C}$. To place this in context, the typical temperature change in the investigation was 6.5°C and therefore,

$$\text{Percentage Uncertainty} = \frac{0.2}{6.5} \times 100 = 3.08\%$$

Using a temperature probe which reports values to a higher precision would reduce the effect of such uncertainty on the outcome of the investigation. Since temperature change is used to calculate the specific heat capacity, obtaining specific heat capacity values which are spread out from one another would also reduce the effect of this uncertainty. Such values of specific heat capacity would be obtained if more molar solutions of magnesium sulfate are explored.

4. Enthalpy of dilution: Depending on the ratio in which the magnesium sulfate solution is diluted, it would take in some heat energy. The result is some amount of thermal energy being absorbed by the magnesium sulfate solution and some being transferred to distilled water. This leads to the thermal energy lost by the solution being underestimated. The effect of this error on the investigation was explored by mixing 30 g of distilled water to 30 g of 1.2 M magnesium sulfate solution (as 1.2 M was the largest molarity investigated) with both being at the same temperature and left for 100 s. After 100 s, the data logger displayed no difference between the initial and final temperature, and hence, the effect of this error on the investigation was insignificant. Although, based on the equipment used, this error is insignificant, using a temperature probe which reports values to a higher precision would allow for the enthalpy of dilution to be measured and accounted for in the error calculations.

Random Errors

1. Number of trials: In the investigation, only five trials were conducted for each molarity due to the paucity of time. This results in the outliers in the data having a larger impact on the results on the investigation. This is also represented in *figure 4* through the vertical error bars. The precision of the average specific heat capacity could have been increased by conducting a larger number of trials for each molarity.

Further scope of investigation

This investigation can be further extended to determine a relation between specific heat capacity and solubility of a salt. By carrying out this investigation and plotting graphs of specific heat capacity vs molarity of different salts (sodium sulfate, etc), the rate of decrease (slope) of specific heat capacity as molarity increases for different salts can be explored. This would ultimately allow us to determine how a salt's solubility has an effect on its specific heat capacity.

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