CALORIMETRY: DETERMINATION OF HEATS OF SOLUTION

By:

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Division #1101

Group: Thursday AM A

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1. CALCULATIONS

1. Calculate the heat capacity of the calorimeter.

TRIAL 1:

ΔT Calorimeter⁄cool water= Tf - T1i

= 34.5 - 19.1

= 15.4°C

ΔT Warm = Tf - T2i

34.5 - 55.4

= -20.9°C

q cal = **-**[(15*4.184*15.4) + (20*4.184*-20.9)]

= -(966.504 - 1748.912)

= -(-782.408)

= 782. 408

C cal = q cal / ΔT Calorimeter⁄cool water

C cal = 782.408/15.4

= 50.80571429 J°C

= 50.8 J°C

TRIAL 2:

ΔT Calorimeter⁄cool water= Tf - T1i

= 37.8 - 20.0

= 17.8°C

ΔT Warm = Tf - T2i

37.8 - 57.0

= -19.2°C

q cal = **-**[(15*4.184*17.8) + (20*4.184*-19.2)]

= -(1117.128 - 1606.656)

= -(-489.528)

= 489.528

C cal = 489.528/17.8

= 27.50157303 J°C

= 27.5 J°C

TRIAL 3:

ΔT Calorimeter⁄cool water= Tf - T1i

= 34.8 - 20.1

= 14.7°C

ΔT Warm = Tf - T2i

34.8 - 57.0

= -22.2°C

q cal = -[(15*4.184*14.7) + (20*4.184*-22.2)]

= -(-935.124)

= 935.124

C cal = 935.124/14.7

= 63.61387755 J°C

= 63.6 J°C

Average heat capacity of the three trials:

Avg. Heat Capacity = (50.8 J°C + 27.5 J°C + 63.6 J°C)/3

= 47.3 J°C

Therefore, the average heat capacity of the calorimeter is approximately 47.3 J°C.

2. Using the heat capacity of the calorimeter, calculate the molar heat of solution of sodium acetate and sodium acetate trihydrate for each mass that was carried out. Tabulate the values.

Sodium Acetate:

$NaC_2H_3O_2$

Given mass: **1.50 g**

MM: 82.03379 g/mol

n = m/MM

n = 1.50/82.03379

= 0.018285148 moles

= 0.0183 moles

Sodium Acetate Trihydrate:

$NaC_2H_3O_2 \cdot 3H_2O$

= $C_2H_9NaO_5$

Given mass: **1.51 g**

MM: 136.07963 g/mol

n = m/MM

n = 0.011096444 moles

= 0.0111 moles

$q_{solution}$(or $\Delta H_{solution}$) = $-[(m_{solvent/solute} \cdot \rho_{solvent/solute} \cdot \Delta T_{solvent/solute}) + (C_{cal} \cdot \Delta T_{cal})]$

Sodium acetate: 21.6 - 20 = 1.6 C

q solution = -[(36.5 * 4.184 * 1.6) + (47.3 J°C * 1.6)]

q solution = -320.0256

-320.0256/0.018285148 moles = -17501.94201 J

= -17.50194201 kJ/mol

= -17.5 kJ/mol

Sodium acetate trihydrate: 18 - 19.6 = -1.6 C

q solution = -[(36.51 * 4.184 * -1.6) + (47.3 *-1.6)]

q solution = 320.092544

325.692544/0.011096444 moles

=28.84640737 kJ/mol

=28.8 kJ/mol

For each mass, I'll be finding the amount in moles and substituting the temperature changes accordingly to get the following values:

Sodium Acetate - Mass	**Temperature Change**	**Moles**	**Heat Capacity**
1.50 g	1.6	0.018285148 moles	**-17.5 kJ/mol**
2.01 g	2.9	0.024502098 moles	**-23.9 kJ/mol**
2.50 g	3.3	0.030475246 moles	**-22.1 kJ/mol**

3.03 g	4	0.036935999 moles	**-22.4 kJ/mol**

Sodium Acetate Trihydrate- Mass	**Temperature Change**	**Moles**	**Heat Capacity**
1.51 g	-1.6	0.011096444 moles	**28.8 kJ/mol**
2.01 g	-2	0.014770763 moles	**27.4 kJ/mol**
2.52 g	-2.1	0.018518568 moles	**23.2 kJ/mol**
3.05 g	-3	0.022413347 moles	**27.6 kJ/mol**

3. Calculate an average value for each molar heat of solution and compare them with the accepted values. Calculate the precision; relative error of your result.

Accepted molar heat of solution for Anhydrous sodium acetate: -17.3kJ/mol

Average molar heat from lab: -21.5 kJ/mol

Accepted molar heat of solution for Sodium acetate trihydrate: 19.7 kJ/mol

Average molar heat from lab: 26.8 kJ/mol

Anhydrous Sodium Acetate

Relative spread for Anhydrous sodium acetate:

(-23.9 + 17.5 / -21.5) * 1000

= 298 ppt

Relative error:

= (-21.5 + 17.3/ -17.3)*100

= 24.3%

Sodium Acetate Trihydrate

Relative spread for Sodium Acetate trihydrate:

(28.8 - 23.2/26.8)* 1000

= 209 ppt

Relative error:

=(26.8 - 19.7 / 19.7)*100

= 36.0%

Percent uncertainty:

TRIAL #1:

Original Equation:

qsolution(or ΔHsolution) = −[(msolvent⁄solute · ρsolvent⁄solute · ΔTsolvent⁄solute) + (Ccal · ΔTcal)]

% Uncertainty = [(0.2/15.0mL)+(0.2/19.1) + (0.2/20.0mL) + (0.2/34.5) + 0.2/19.1)]*100

=0.050072843 J/g·C

=0.05007 J/g·C

= 5.01%

TRIAL #2:

% Uncertainty = [(0.2/15.0mL)+(0.2/20.0) + (0.2/20.0mL) + (0.2/37.8) + 0.2/20.0)]*100

=0.048624338 J/g·C

=0.04862 J/g·C

= 4.86%

TRIAL #3:

% Uncertainty = [(0.2/15.0mL)+(0.2/20.1) + (0.2/20.0mL) + (0.2/34.8) + 0.2/20.1)]*100

=0.048980957 J/g·C

=0.04898 J/g·C

= 4.90%

Trial #:	% Uncertainty
1	5.01%
2	4.86%
3	4.90%

Average of the three trials:

Average % uncertainty = (5.01% + 4.86% + 4.90%) / 3

= 4.92%

Average uncertainty in J/g·C:

(0.05007 J/g·C + 0.04862 J/g·C + 0.04898 J/g·C) / 3

= 0.049223333 J/g·C

= 0.04922 J/g·C

2. DISCUSSION

Around one paragraph, discussing why the results were higher OR lower than the actual values (-17.32 kJ/mol for sodium acetate and 19.66 kJ/mol for sodium acetate trihydrate)

My results were higher than the actual values for Anhydrous sodium acetate, along with Sodium acetate trihydrate. The accepted molar heat of solution for Anhydrous sodium acetate is -17.3kJ/mol, and the value obtained from this lab for the average molar heat is -21.5 kJ/mol.

This reaction is exothermic, hence the negative sign. Energy is released when the solute is dissolved. The accepted molar heat of -17.3 kJ/mol is less than the value obtained from the lab, -21.5 kJ/mol as a larger quantity of heat is released in the lab due to error to be discussed in more detail, hence the increase. The sign is negative, and in this case, it just means that energy is being released. A larger number behind the negative sign indicates more heat being released in the exothermic reaction.

As for Sodium Acetate Trihydrate, the accepted molar heat of solution is 19.7kJ/mol, whereas the average molar heat obtained from the lab and my calculations is 26.8 kJ/mol. In this reaction, heat is absorbed and an endothermic reaction occurs as a result. An increase of heat absorbed is observed between the accepted molar heat value of 19.7 kJ/mol compared to the lab derived value of molar heat, 26.8 kJ/mol. This difference is relatively large, primarily due to lab errors.

The following lab errors result in higher values than the actual values:

1. Anhydrous Sodium acetate absorbs water very easily, and a source of error could be that it wasn't weighed out and transferred quick enough into the calorimeter. This increases the weight of it, when in reality the additional mass is from the absorbed water.
2. There was a slight increase in mass for Anhydrous Sodium Acetate. Instead of weighing out 1.50 grams, 1.51 grams was measured out instead which tampered the result. Again, an increase in quantity of substance means an increase in heat absorbed to dissolve the substance.
3. According to the CHEM1101 lab manual (Archer et al., 2020), another source of error would be not fully drying the calorimeter and thermometer before each trial. When drying, the plastic tube should've been removed to dry all of the water stuck in the cavity that holds the plastic tube. Again, this would increase the mass of the substance, thus increasing the molar heat of solution as a result as more heat is produced in the exothermic reaction with a greater mass, and absorbed to dissolve a greater mass of substance in an endothermic reaction.

4.

Account for the sign of the heat of solution obtained for sodium acetate and for its trihydrate in terms of the theory presented in the introduction to this experiment. Explain why a substance with an endothermic heat of solution actually dissolves.

In this lab, the heat of solution obtained for sodium acetate was negative; an exothermic reaction took place where heat was released. The sign for molar heat, q reaction and Enthalpy changes, ΔH, were negative for sodium acetate, thus indicating an exothermic reaction; a heat generating process. The calorimeter would get hotter after the reaction is complete.

Sodium Acetate:

$NaC_2H_3O_2 \rightarrow NaC_2H_3O_2$ + Energy

Sodium Acetate Trihydrate on the other hand has a positive value for molar heat, q reaction and enthalpy changes, ΔH which indicates an absorption of heat by the reactants, with heat absorbed as a result. The calorimeter would get colder after the reaction is complete.

Sodium Acetate Trihydrate:

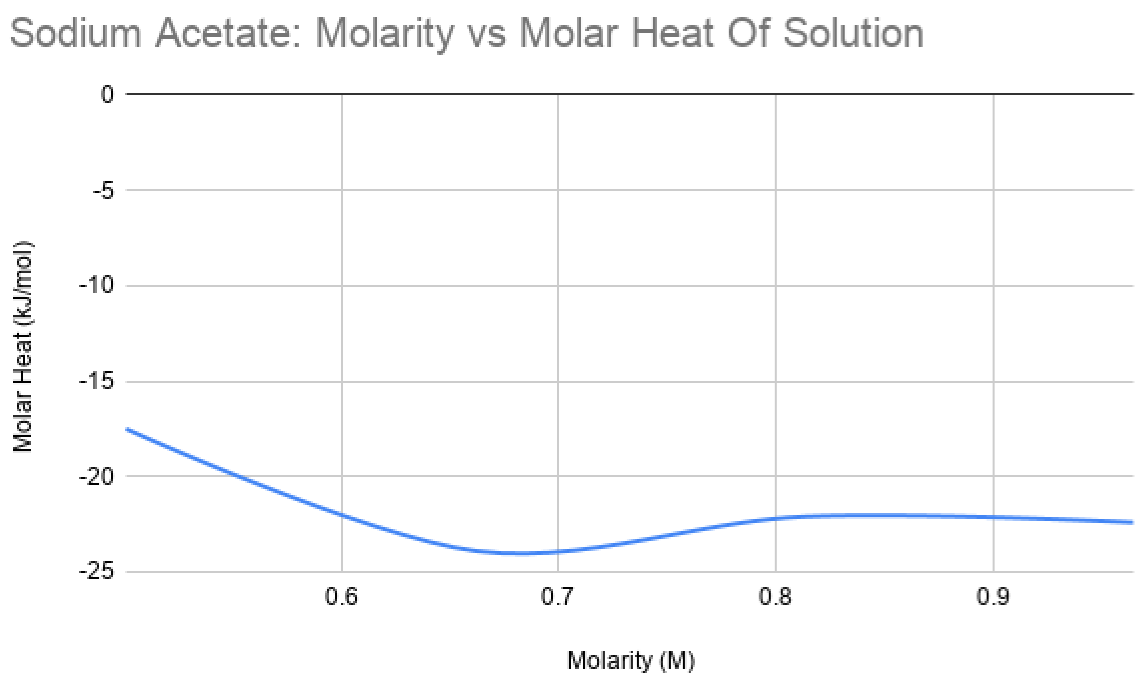
$NaC_2H_3O_2 + 3H_2O$ + Energy $\rightarrow C_2H_9NaO_5$

As for why a substance with an endothermic heat of solution dissolves, it is because the net energy from breaking and forming bonds results in heat being absorbed into the system as the solute dissolves. Ions are stabilized in the solution by their interaction with water molecules. These ions are said to be hydrated. When the temperature is increased, the solubility of the solute is increased as well as a result.

5.

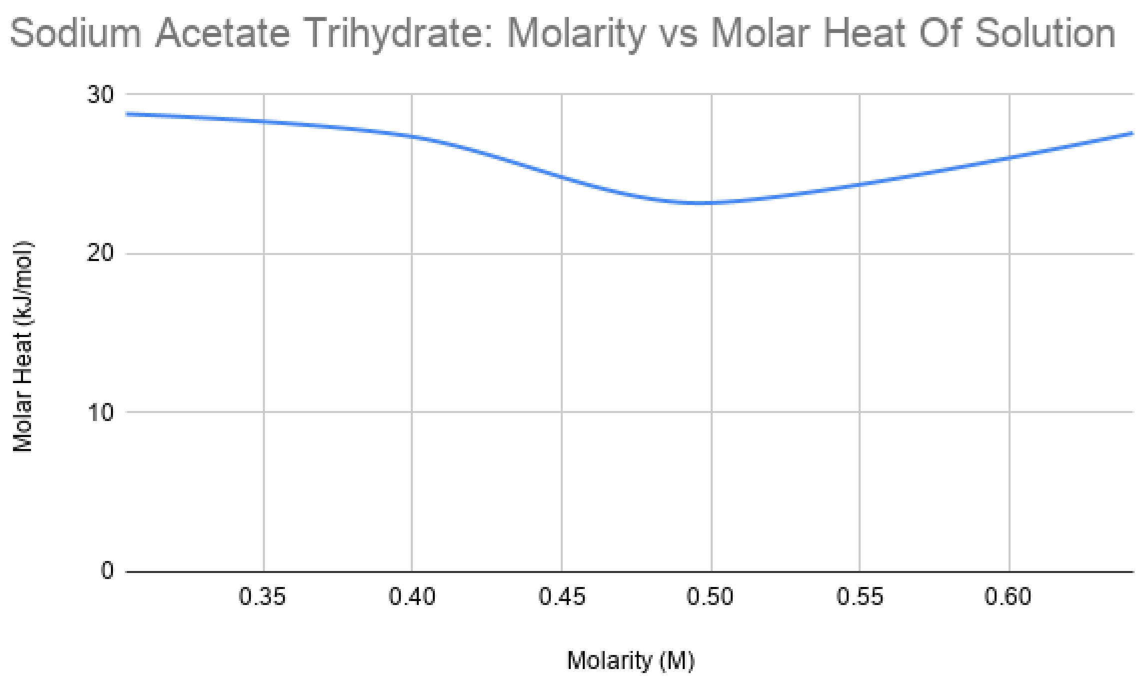
Sodium Acetate Data:

Moles	Volume	Molarity	Molar Heat
0.018285148 moles	0.0365 L	0.500962958 M	-17.5 kJ/mol
0.024502098 moles	0.03701 L	0.662039935 M	-23.9 kJ/mol
0.030475246 moles	0.0375 L	0.812673226 M	-22.1 kJ/mol
0.036935999 moles	0.0383 L	0.964386396 M	-22.4 kJ/mol



Sodium Acetate Trihydrate Data:

Moles	Volume	Molarity	Molar Heat
0.011096444 moles	0.03651 L	0.303928895 M	28.8 kJ/mol
0.014770763 moles	0.03701 L	0.399101945 M	27.4 kJ/mol
0.018518568 moles	0.03752 L	0.493563432 M	23.2 kJ/mol
0.022413347 moles	0.03805 L	0.641612273 M	27.6 kJ/mol



There seems to be an error in my data, as my graphs look very odd. I made sure that the calculations were all correct, along with the right value used to obtain them from the experiment. There is a concave dent in both graphs for both solutions, and a point of equilibrium is reached at and after 0.80 M for Sodium Acetate, whereas the graph for Sodium Acetate trihydrate does not show a point of equilibrium but a concave curve along the positive x axis. It fails to demonstrate a proper endothermic reaction graph as the values are not accurate.

In the earlier calculations, despite an increase in concentration, there was a sudden spike for sodium acetate that messed the data up at 2.01 grams, the trend was that the exothermic data was getting larger in the negative direction, and then it jumped from -17.5 kJ/mol to -23.9 kJ/mol for 2.01 grams, and then back to -22.1 kJ/mol and values similar to it. The spike there is why my graph looks weird, and is due to an accidental increase of solute which increased the exothermic energy released as discussed in the lab errors section.

This same trend is noticed for sodium acetate trihydrate, it starts at a high value for the first piece of data, then decreases for the next to masses and spikes up on the final trial's mass.

Conclusion:

In conclusion, my results were higher than the ideal values for molar heat capacity. The average heat capacity of my calorimeter is 47.3 J/g·C, the average molar heat of solution for sodium acetate is -21.5 kJ/mol, and for sodium acetate trihydrate is 26.8 kJ/mol. The relative error for sodium acetate was 24.3% and the relative spread was 298 ppt. For sodium acetate trihydrate, it was 36.0%, and 209 ppt. Lastly, the total average uncertainty for this experiment is 0.04922 J/g·C, or in percentage, 4.92%.

Bibliography:

Archer, D.W., Burk, R.C., White, C.A., Wolff, P.A., Levac, S., and Mesnic, N., CHEM1101 Chemistry for Engineers, Carleton University, 2020-2021.