

Richard A. Marino  
Partners: Sterling and Connor  
CHM113  
Freezing Point Depression: Lauric Acid  
10 Nov 2014

## Purpose

When substances freeze and solidify, the molecules form a rigid, and often very closely packed, dense solid. During the freezing process, additional chemical bonds are formed, which increases the rigidity, and brings the molecules closer together. This often results in a solid that is more dense than the liquid. Water is an exception, as the hydrogen bonds that form in the water push the molecules further apart than they are in liquid form (resulting in a less dense solid).

If impurities are present in the solution, they make it more difficult for bonds to form, requiring more energy to be removed from the substance to form these bonds. This lowers the freezing point. In this lab, we examine the effects of impurities in a solution lowering the melting point by measuring the freezing point of Lauric Acid in three solutions -- pure, with 0.75 grams of Benzoic Acid per 8 grams of Lauric Acid, and with 1.5 grams of Benzoic Acid per 8 grams of Lauric Acid. More Benzoic Acid caused lower freezing point.

# Procedure

In a prior lab, we determined the freezing point of Lauric Acid to be 43.2 °C. We inserted test tubes of **0.750 g BA/8.00 g LA** and **1.50 g BA/8.00 g LA** into hot water baths to melt them.

After the test tubes were melted, we removed each one individually, and inserted it into a room temperature water bath to cause it to begin freeze. We inserted the temperature probe at the same time, and stirred the contents of the test tube continuously. To aid in the cooling process, we occasionally stirred the water in the bath, to prevent hot spots. Once the temperature stabilized for a period (due to hitting the freeze point), we added ice chips to the bath to more quickly remove energy.

# Equations

This fairly simple experiment used only two equations, the freezing point depression equation,

$$\Delta T = K_f m$$

in which  $\Delta T$  is the change in the freezing point,  $K_f$  is the freezing point depression constant for the solvent, and  $m$  is the molality of the solution.

We also used the molality formula:

$$\text{Molality} = \text{moles solute} / \text{mass solvent}$$

# Data

We already established two pieces of data that were required; the melting point of pure Lauric Acid, and the Molecular Weight of Benzoic Acid.

MW Benzoic Acid: 122.12 [g/mol] ( $C_7H_6O_2$ )

Freezing Point of Pure Lauric Acid: 43.2 °C

We also determined our freezing points, and changes in the aforementioned:

	<u>0.750 [g BA] / 8.00 [g LA]</u>	<u>1.50 [g BA] / 8.00 [g LA]</u>
Freezing Point	41.28 °C	37.80 °C
$\Delta T$	1.92 °C	5.40 °C

For 0.750 grams of BA / 8.00 grams of LA:

molality = moles solute / mass solvent.

mass of BA [g] / molecular weight [g/mol] = 0.750 [g] / 122.12 [g/mol] = 0.006142 [mol].

molality = 0.006142 [moles] / 0.008 [kg] = 0.76769 [moles / kg]

$\Delta T / m = K_f \rightarrow 1.92 [^{\circ}\text{C}] / 0.76769 [\text{moles} / \text{kg}] = 2.501 [^{\circ}\text{C}/\text{mol}]$

For 1.50 grams of BA / 8.00 grams of LA:

mass of BA [g] / molecular weight [g/mol] = 1.50 [g] / 122.12 [g/mol] = 0.012283 [mol].

molality = 0.012283 [moles] / 0.008 [kg] = 1.53538 [moles / kg]

$\Delta T / m = K_f \rightarrow 5.40 [^{\circ}\text{C}] / 1.53538 [\text{moles} / \text{kg}] = 3.517 [^{\circ}\text{C}/\text{mol}]$

Average  $K_f = 3.009 [^{\circ}\text{C}/\text{mol}]$

# Discussion

We determined the average Freezing point depression constant for our Benzoic Acid to be **3.009** [ $^{\circ}\text{C}/\text{mol}$ ]. We noticed during the experiment that precipitate had formed in the test tubes during melting. This occurred in both solutions, but especially in the 1.50 gram solution. We believe this affected our results significantly, as there were pre-existing solids in the solution while we were trying to freeze it, which would further lower our freezing point.

In general, our graphs were fairly stable.

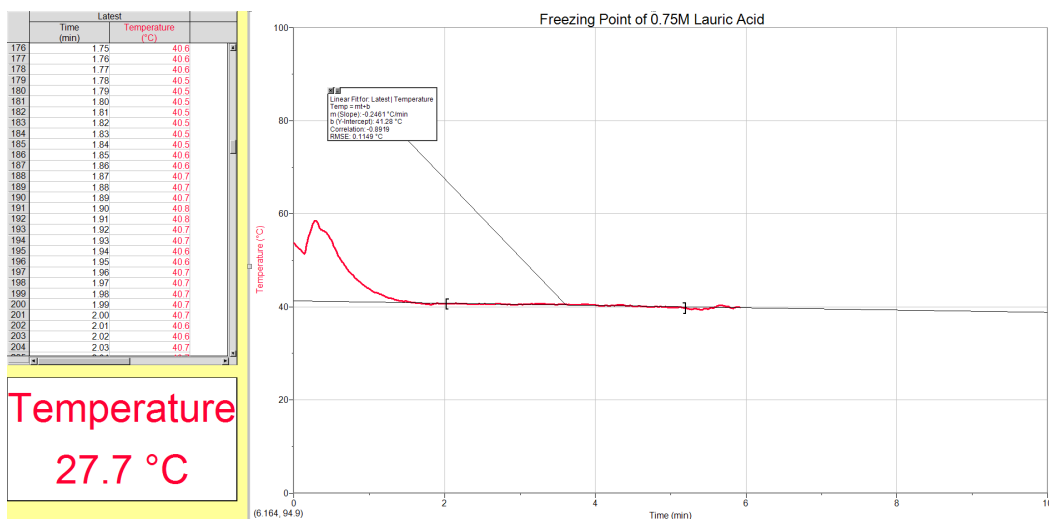
# Conclusions

In this lab, we attempted to calculate the freezing point depression constant of benzoic acid. We did this by measuring the freezing points of different concentrations of the benzoic acid in solution of lauric acid. As the freezing point depression constant is a constant, the freezing point temperature should decrease linearly with increases in concentration. This did not appear to be the case, as doubling the concentration resulted in more than doubling the change in freezing point in our case. We believe this error is due to precipitate being present in the solution while we attempted to freeze it. Consequently, we did not get a constant value for our freezing point depression constant. We presume the precipitate came from prior contamination of the solutions.

The experiment was useful in demonstrating how this process could be performed.

# Appendix

## 0.750 g BA / 8.00 g LA



## 1.50 g BA / 8.00 g LA

