Freezing Point Depression

Calculations

$\Delta T_f = iK_f m$

ΔT_f is the freezing point depression (decrease in freezing point)

i = The van't Hoff factor accounts for dissociation of solute in solution.
for this experiment the i = 1
(for ionic solutes i depends on the number of ions that are produced on dissociation)

K_f = cryoscopic constant
= *3.90 °C-kg/mol for this experiment*

m = molality (moles solute/1000 g solvent)
this is the unknown

Rearrange the formula:

$m = \Delta T_f / iK_f$

1. Determine the freezing point of pure Lauric Acid (the solvent)
2. Determine the freezing point depression (ΔT_f) for each solute sample (1 g, 2, g, 3 g)
   Subtract the temperature at which the graph levels off from the freezing point of lauric acid
3. Divide freezing point depression by iK_f, (3.90 °C-kg/mol)
4. Divide by 100 to convert the molality to moles/10 grams.
   (You used 10 g lauric acid as the solvent)
   This will the give you the number of moles in the number of grams that was used.
5. Divide the # of grams by the number of moles to get the molar mass.
   # of grams is 1 g, 2 g, 3 g
6. Determine the average molar mass from the 3 trials
7. Determine percent error for the average ((actual – observed) / actual)) x 100) using the known molar masses of
   - *p*-dichlorobenzene ($C_6H_4Cl_2$) = *147.00 g/mol*
   - naphthalene ($C_{10}H_8$) = *128.17 g/mol*
   - stearic acid (*aka* octadecanic acid) ($C_{18}H_{36}O_2$) = *284.47 g/mol*
   - camphor ($C_{10}H_{16}O$) = *152.23 g/mol*
   - cetyl alcohol (*aka* 1-hexadecanol) ($C_{16}H_{30}OH$) = *242.44 g/mol*

Example for Camphor

1. Freezing point of Lauric Acid = 46.0 °C
2. ΔT_f for 1 g camphor = 3.0 °C
3. 3.0 °C/3.90 °C-kg/mol = 0.77 mol/kg
4. 0.77 mol/kg / 100 = 0.0077 mol/10 g
5. 1 g / 0.0077 mol = 129.9 g
6. Repeat for 2 g and 3 g samples and determine average molar mass
7. Determine percent error using the average value: ((actual – observed) / actual)) x 100