Freezing Point Depression

Calculations

\[ \Delta T_f = iK_fm \]

\( \Delta T_f \) is the freezing point depression (decrease in freezing point)

\( i \) = The van't Hoff factor accounts for dissociation of solute in solution.

\( i \) = 1 for this experiment (for ionic solutes \( i \) depends on the number of ions that are produced on dissociation)

\( K_f \) = cryoscopic constant

\( K_f \) = 3.90 °C-kg/mol for this experiment

\( m \) = molality (moles solute/1000 g solvent)

\( m \) = the unknown

Rearrange the formula:

\[ m = \frac{\Delta T_f}{iK_f} \]

1. Determine the freezing point of pure Lauric Acid (the solvent)
2. Determine the freezing point depression (\( \Delta T_f \)) for each solute sample (1 g, 2 g, 3 g)

\( \Delta T_f \) is 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 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 (C₈H₆Cl₂) = 147.00 g/mol
- naphthalene (C₁₀H₈) = 128.17 g/mol
- stearic acid (aka octadecanic acid) (C₁₈H₃₆O₂) = 284.47 g/mol
- camphor (C₁₀H₁₆O) = 152.23 g/mol
- cetyl alcohol (aka 1-hexadecanol) (C₁₆H₃₀OH) = 242.44 g/mol

Example for Camphor

1. Freezing point of Lauric Acid = 46.0 °C
2. \( \Delta T_f \) 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