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Project #1- Freezing Point Depression Constant for Acetic Acid

Introduction:

The purpose of this lab was to experimentally determine the colligative property; freezing point depression constant, for acetic acid. Colligative properties are properties of solutions that depend on the number of molecules in a given amount of solvent and not on properties such as size or mass of the molecules. Colligative properties include: vapor pressure; boiling point; depression of freezing point and osmotic pressure. Freezing point is the equilibrium between a liquid and solid phase and is lowered by the presence of a solute compared to a pure solvent. The solute particles cannot enter the solid phase, hence, less molecules participate in the equilibrium so equilibrium is achieved at a lower temperature at which the rate of freezing becomes equal at the rate of solidifying. Because an impurity is being added the freezing temperature of the pure acetic acid would be higher than when acetylsalicylic acid is added to the solution, when the molecules of the solvent freeze they form crystals of pure solvent. At the freezing point of pure solution there was a equilibrium of molecules joining and detaching from the solid but when a solute is added to the solution in the crystals remain as pure solvent so the solute cannot join. The solute comes into to contact with the crystal because it is part of the solution. When the solute is added the solvent cannot come into contact with the crystals and join the solid as fast as before because of the presence of the solute. This means the freezing point is lowered. The freezing point depression constant can be solved for if the change in the freezing point after solute is added, mass of the solute and solvent in the solution and the molar mass of the solute are known. Freezing point depression is used by mixing compounds and changing their properties by changing the concentration of solute to solvent. Freezing point depression is used in anti-freeze fluid, homemade ice cream and in technical applications to avoid freezing.

Procedure

1. MicroLAB is set up to measure the temperature change with temperature in Celsius on the Y axis and time in seconds on the X axis.
2. A clean beaker and test tube were weighed and mass recorded
3. a few mL acetic acid were added to test tube, weighed with beaker and mass was recorded.
4. Add 10 drops acetylsalicylic acid to test tube, weighed this and beaker and recorded mass.
Trial #1 acetylsalicylic acid weight = .255 Trial #2 acetylsalicylic acid weight = .397 Trial #3 acetylsalicylic acid weight = .383
5. Placed test tube in ice water and stirred until solution was super-cooled and could freeze.
Recorded temperature after super-cooling was complete and temperature had leveled.
6. Input data into formula $K_f = [T_f(\text{MM solute})(\text{kg solvent})]/\text{g solute}$ and solve for K_f .
7. Repeat twice and find average K_f value.

Data and Calculations:

Experimentally determined freezing point temperature for pure acetic acid: 14.73 degrees C.

See attached graphs for thermistor data

Attachment 1: untitled1 acetic acid freezing point

Attachment 2: FP Acetic Trial #1

Attachment : FP Acetic Trial #2

Attachment 4: FP Acetic Trial #3

K_f for acetic acid:

	Trial #1	Trial #2	Trial #3
Mass of test tube and beaker	99.717g	99.721g	99.244g
Mass of test tube, beaker & solvent	103.110g	103.486g	104.282g
Mass of test tube, beaker, solvent & solute	103.365g	103.883g	104.665g
Freezing temperature of solution	12.88C	12.92C	9.79C
Calculated mass of solvent	3.393g	3.765g	5.042g
Calculated mass of solvent (kilograms)	.003393	.003765	.005042
Calculated mass of solute	.255g	.397g	.383g
Calculated change in T_f	1.85C	1.81C	1.79C

Calculated K_f (degrees C/m)	4.45	3.01	4.26
Average calculated K_f for acetic acid without trial #2	4.35		
Average calculated K_f for acetic acid with trial #2	3.90		

Trial #2 was not used when calculating average value because its results were so far off it was the outlier.

$$\Delta T = K_f [(\text{grams solute} / \text{MM solute}) / (\text{kg solvent})]$$

K_f = change in T_f (MM solute) (kg solvent) / (grams solute) where the MM of acetylsalicylic acid is 180.157 g/mol.

Trial #1:

$$K_f = 1.85(180.157)(.003393) / (.255) = 4.43$$

To calculate the change in the freezing temperature the freezing point temperature for pure acetic acid minus freezing point temperature of solution.

Conclusion and Data Analysis:

The purpose of this lab was to experimentally determine the colligative property; freezing point depression constant, for acetic acid. For this lab acetylsalicylic acid was used as the solute to measure the freezing point depression for acetic acid. The freezing point depression constant is highly effected by the concentration of the solvent and solute and whether or not the solute is completely dissolved or not will greatly effect the calculations. The higher the concentration of solvent and the lower the concentration of solute the closer the solutions freezing temperature will be to the pure solvents freezing temperature. By adding mass of solute with a far higher molar mass will in fact not influence the freezing point depression of the solvent as much as adding a solute with less mass but a far lower molar mass. Again, the concentration of the solute was what was important when determining freezing point depression.

After all freezing temperatures had been measured, solutes and solvents weighed, and molar mass of acetylsalicylic acid determined then the freezing point depression constant was determined using the average value found experimentally. Trial #2 was thrown out and not used for the final calculation because this trial was an outlier; it may have had such bad results because the amount of solute was so large in the solution that it did not fully dissolve. So only a percent of the solute was able to affect the

freezing point of the acetic acid resulting in a very off end result. The average calculated from the other trials was a freezing point depression constant (K_f) for acetic acid of 4.35. The known K_f for acetic acid is 3.9. The calculated average K_f was 3.9. This may have been caused because the first and third trial was slightly high and the second trial was very low. One possible cause of error was that the temperature may not have been at 25 degrees Celsius and that is what the accepted value was calculated for. Also the solution underwent supercooling so it may have not been its exact freezing points. Other modes of error could be slight discrepancies weighing the solute and solvent which would change results. The experimentally obtained K value had an 11% error so the lab was successful.