Determination of the Solubility of an Unknown Salt

Theory:

Hot tea can dissolve more sugar than iced tea, and warm water dissolves less oxygen than cold water. The maximum amount of any **solute** that can dissolve in a given amount of **solvent** is called its **solubility**, and this depends on temperature. The solubilities of gases always decreases with increasing temperature. For liquids and solids, solubilities generally increase with increasing temperature, as is the case with sugar in tea. However, there are a number of exceptions to this, two examples being cerium sulfate, $Ce_2(SO_4)_3$, and lithium carbonate, Li_2CO_3. For these ionic solids, solubility decreases with increasing temperature.

The Problem to be Investigated:

The solubility of an inorganic salt in water at various temperatures will be determined, and a solubility curve for the salt will be drawn.

The Nature of This Investigation:

A known weight of a salt will be dissolved in varying, known volumes of water, and at each concentration the solution will be allowed to cool to determine at which temperature the salt begins to crystallize out of solution. That is, for a certain weight of salt, the quantity of water will be progressively increased, and for each increase the temperature required to reach the point of saturation will be determined.

The weight of salt and the weight of water for each saturation temperature represent the concentration of a saturated solution for that temperature, and will to expressed in terms of g salt per 100 g water. A solubility curve will then be drawn by plotting the g salt per 100 g water on y-axis against saturation temperature on x axis. Chemists frequently refer to the solubility of certain compounds. What does the term solubility mean? Solubility is a quantitative means of describing the composition of solutions. Solutions may be formed by mixing various combinations of liquids, gases, and solids. In this experiment only the dissolution of a solid in a liquid will be considered. Thus, solubility as used to describe a solid dissolved in a liquid would refer to the quantitative composition of the solution. As such, the solubility of a solid may be expressed as the number of moles of solid dissolved in a liter of liquid, or as the mole fraction of the solid, or, as in this experiment, as the number of grams of solid dissolved in 100 mL of liquid. The term solute is generally applied to the constituent which is present in the solution in the least quantity, and the term solvent to the constituent present in the greatest quantity.

The preceding discussion, however, does not yet uniquely define the term solubility because solutions may be prepared which exist at an almost infinite number of compositions. Consider, for example, the preparation of a solution by the addition of sodium chloride to water. Sodium chloride may be added to the water in small quantities to produce an entire series of solutions of different compositions. The addition of sodium chloride to a given quantity of water to produce solutions of higher sodium chloride content cannot proceed indefinitely. At some point further addition of sodium chloride will not produce a solution of higher sodium chloride content. At this point no further sodium chloride will dissolve in the solution. This limiting solution composition represents what is called a saturated solution. The quantity of solute in the saturated solution is called the solubility of the particular solute in the given solvent. For sodium chloride (the solute) in water (the solvent), the solubility at 20° C is 36.0 g of sodium chloride per 100 mL of water. This means that at 20° C only 36.0 g of sodium chloride will dissolve in 100 mL of water. Obviously solutions of sodium chloride in water could be prepared at compositions less than the solubility at any given temperature. Such solutions are said to be unsaturated.

A saturated solution may be prepared by placing an excess of the solute in contact with the solvent. The solute will dissolve until the solution is saturated and the excess of solute over the saturation quantity will remain undissolved. The situation in which undissolved solid exists with solution is, however, not a static one; solute particles are continually entering the solution and dissolved solute continually reforms as the solid. The dissolution of the solid is offset by the precipitation of solid from the solution so that the quantity of undissolved solute and the composition of the solution remain constant at the saturation point.

The solubility of a solute in a solvent depends on the nature of the solute, the nature of the solvent, the pressure, and the temperature. Only the effect of temperature will be investigated in this experiment.

The effect that temperature has on the solubility of a substance is determined by the quantity of heat released or absorbed as the solute dissolves. Whether heat is liberated or absorbed depends on the energy required to disrupt the crystal structure of the solid and on the energy liberated when the solid particles interact with the solvent. If more energy is required to break particles away from the solid than is gained by virtue of the interaction of these particles with solvent, heat will be absorbed by the system, and vice versa. The dissolution of most solids is accompanied by an absorption of heat. This means that if the temperature of a saturated solution is raised, the process whereby solute particles enter the solution is favored over the precipitation of solid from solution. Thus the solubility will increase with temperature. If the dissolution of the solid releases heat, the precipitation process is favored over the dissolution process and the solubility decreases as the temperature increases.

The fluctuation of the solubility of a solute with temperature can best be represented by the curve obtained by plotting the temperature on the abscissa or x axis and the concentration in g solute per 100 g solvent on the ordinate or y axis.

Solubility-temperature curves for the variation in solubility with temperature for a few common substances in water are shown in the graph given out in class.

PROCEDURE

Part 1. Determination of Initial Saturation Temperature

1. Select a 2-hole rubber stopper that fits snugly into a test tube.

2. Carefully insert a thermometer through one hole of the rubber stopper so that the thermometer tip is near the bottom of the test tube. Through the other hole in the rubber stopper carefully insert the bent end of the coat hanger. By moving this rod up and down, or by twisting, a solution in the test tube can be agitated.

3. Place a large beaker onto a wire gauze on a ring mounted on ring stand.

4. Add to the beaker approximately 300 mL of water and begin heating the beaker and water to about 80° C.

5. From the sample given to you mass out as accurately as possible 5.00 grams of the salt. Record the actual amount used on notecard.

6. Transfer as completely as possible the salt into the test tube.

7. From a buret, dispense 3.00 mL of distilled water into the test tube. Record on the Data Sheet, to the nearest hundredth of a mL, the exact volume of distilled water added to the test tube.

8. Place the rubber stopper with the thermometer and agitator in the test tube. Be certain that the thermometer tip is close to the bottom of the test tube.

Part II. Determination of Additional Saturation Temperatures

9. Clamp the test tube to a ring stand in such a way that the test tube is suspended in the hot water bath at a height so that the liquid levels inside and outside the test tube are the same.

10. By twisting the coat hanger, stir the mixture in the test tube while heating the test tube and contents to 80°C. Keep the heating time to a minimum to avoid the possible loss of water by evaporation.

11. If the salt is not completely dissolved when the temperature of the mixture reaches 80° C, add exactly 0.50 mL of distilled water from the buret to the mixture. Record on the notecard the exact volume of distilled water added to the test tube. Again heat the test tube and contents in the hot water bath to 80° C.

12. Repeat the procedure in Step 11 until just enough distilled water has been added to dissolve all of the salt at a temperature no higher than 80° C.

13. After the salt has completely dissolved, raise the test tube and contents out of the hot water bath and allow the solution in the test tube to cool while stirring the solution continuously.

14. Observe the solution carefully. Record on the notecard the temperature at which the first crystals appear in the test tube. This temperature is called the saturation temperature. If in doubt about the exact temperature, again heat the test tube and contents to 80° C in the water bath until all of the crystals are dissolved, repeat the cooling procedure, and record the saturation temperature on the Data Sheet.

15. To the test tube and contents from Part 1, add 1.00 mL of distilled water from the buret. Record on the notecard, to the nearest hundredth of a mL, the volume of distilled water added to the test tube.

16. Heat the test tube and contents in the hot water bath until all of the crystals dissolve.

17. Remove the test tube and contents from the hot water bath and allow them to cool. Record the saturation temperature on the notecard.

18. By following the procedure in Steps 15, 16, and 17, further additions of 1.00 ml of distilled water at a time should be made and the saturation temperature determined after each addition. All data should be recorded on the notecard. If 1.00 mL increments of distilled water do not significantly change the saturation temperature, increase the quantity of distilled water to 2.00 mL/3.00 mL increments.

19. When the solution becomes quite dilute and the saturation temperature reaches room temperature, place the test tube and contents in an ice-water bath to obtain several readings between room temperature and 0° C. In all, 6 to 8 different saturation temperatures should be determined (the total volume of water should be between 12 and 20.00 mL).

Calculations:

1. Knowing the weight of salt and weight of water for each saturation temperature determined, calculate the g of salt per 100 g of water. Record these solubilities in the table on the Data Sheet.

2. Plot the solubilities found in Question 1 vs saturation temperature for each solution on the graph paper attached to this experiment.

Questions:

1. Explain why the saturation temperature is that temperature at which the first crystals appear.

2. What would be the effect on the observed saturation temperatures if some of the water were lost by evaporation? Discuss the answer.

3. Once the salt is known, determine a percent error by dividing the difference in the book value and your value by the book value and then multiplying by 100 %.

4. The dissolved oxygen content of water is very critical to marine life. Investigate the solubility of oxygen in water as a function of temperature, and couple with that an investigation of the oxygen needs of various kinds of aquatic animals. Such studies relate to "thermal pollution," a serious concern near power plants and other industries that release water into streams and rivers. Write a report on this. Pay particular attention to any such problems in your area.