Activity # 13

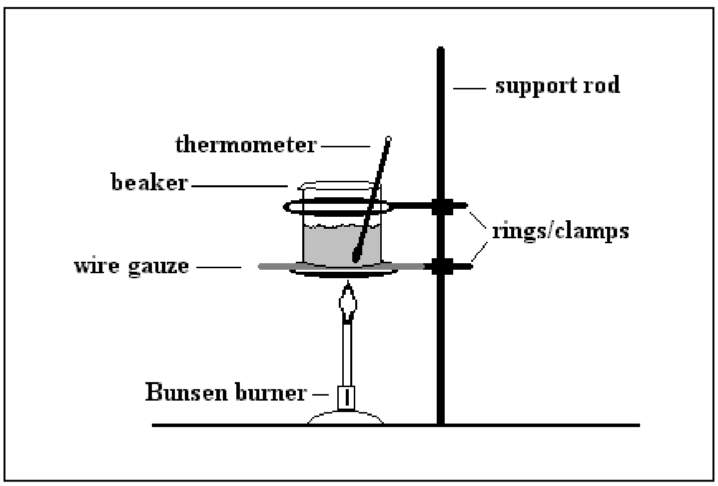
Title: Solubility Curves-Student's Copy

Purpose: to experimentally determine and graphically compare the solubility of various compounds

Materials: beaker (100 mL or larger), Bunsen burner, burner lighter, stirring rod, laboratory balance, graduated cylinder, weighing dish, Celsius thermometer, 1 solute sample per team (either NaCl or $C_{12}H_{22}O_{11}$), water, wire gauze, 2 rings/clamps, 1 support rod, clamps, chemical scoop, safety goggles, apron

Hazards/Precautions: Because of the use of open flames and heated liquids, long hair will be tied back and safety goggles and aprons will be worn.

Introduction: Each lab team will be assigned the task of determining the number of grams of a specified solute that will be required to produce a saturated solution (using 50 mL of water) at a certain temperature. The entire class's data will be compiled onto a single graph (which each person will include in the laboratory report) for analysis.



Procedure:

1. Pour **exactly 50 mL** of water into the beaker. Assemble your laboratory equipment as shown in the diagram above. (Note the second ring/clamp at the upper region of the beaker for stability!)
2. Heat (Do NOT boil!) the water in your beaker to the specified temperature. This temperature must be maintained throughout the entire activity. (This is one of the tough parts of the lab! Discuss beforehand with your lab team just how you will attempt to **maintain this constant temperature**.)
3. Determine and record the number of grams of your sample solute required to produce a saturated solution at your prescribed temperature. (Uh, oh…. another tough part! A member of your team is going to have to measure, record and tally the series of small solute masses that are added (a little at a time) to the 50 mL sample of water… to the point where no more solute can be dissolved.) Remember to stir the solution continually after each addition of solute.
4. When the saturation point is reached, the heat source can be turned off and the beaker/contents left to cool. Disposal of remaining beaker contents will be at the teacher's direction.

5. Plot the grams of solute dissolved in 100 mL of water on the Y-axis and the temperature in ^oC on the X-axis for each solute tested in class as the individual lab teams disclose the results.

Inquiry/Analysis:

1. How did you know when the saturation point was reached with your given sample?

2. If you had recorded the mass of your empty beaker at the beginning of the activity, describe how could you now check the accuracy of your calculated value for "# of grams of dissolved solute" at the specified temperature.

3. Solubility curves are usually expressed as the number of grams of solute that can be dissolved in **100 g** (or mL or cm^3) of water. What needs to be done with OUR solute masses to conform to these standards of reporting? (Remember: we only used 50 mL of water!)

4. After graphical analysis of the entire class's data, a general trend seems to be apparent. Complete the following statement: As the temperature of the solvent (water) increases, the number of grams of solute required to saturate a solution _____________ (increases, decreases, remains the same).

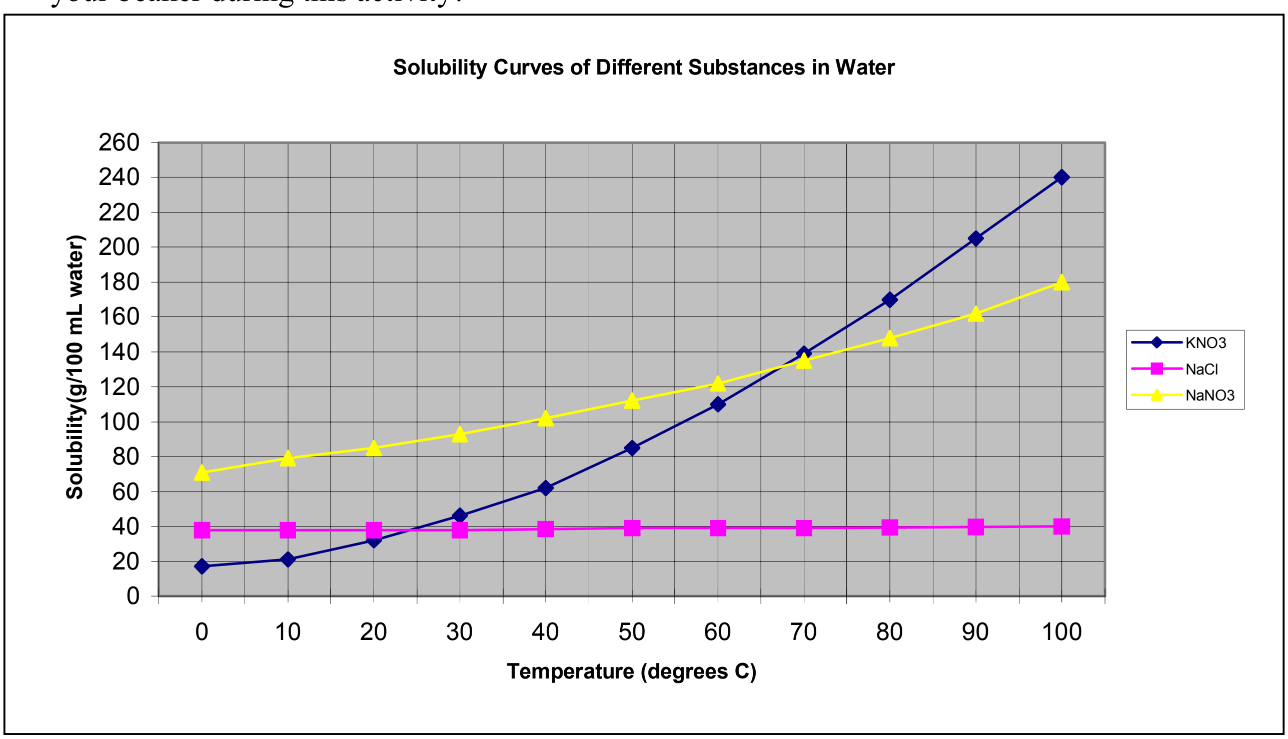
5. Do both solutes appear to be equally soluble in water?

6. What is the independent variable in this activity? (The independent variable is the factor in an experiment that is controlled by the person(s) performing the test.)

7. What is the dependent variable on your graph? (The dependent variable is the factor in an experiment that varies as a result of changes made to the independent variable.)

8. What errors would have been introduced into our lab activity if we were allowed to bring the solvent to a boil?

9. Why couldn't ice be used in an attempt to regulate and maintain a constant water temperature in your beaker during this activity?



10. The solubility of which TWO substances in the graph above is nearly the same at room temperature?
11. If the solution of potassium nitrate was cooled from 100^{oC} to $60^{o}C$, how many grams of solute would you expect to precipitate out of solution?
12. Describe a laboratory separation technique, which could be employed to recover this precipitated solute?
13. After the precipitate was separated (above), describe how you could then recover the remaining dissolved solute (in this same lab period)?
14. If equal samples of each of the saturated solutions above were collected at $45^{o}C$ and then allowed to evaporate to dryness, which would provide the greatest mass of remaining solid?