

## Experiment 2 Solubility

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Date:

### Key Objectives

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1. Terminology - solute, solvent, solution, miscible, immiscible
2. Effect of particle size, temperature, concentration and agitation on rate of dissolving
3. Solubility units/measurements
4. Calculating Enthalpy of Solution
5. Use of ratio/proportions and linear equations to solve problems

### Discussion

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A complete discussion of solutions can be found in OER Textbook Chapter 11.

**Solutions** are mixtures of two or more substances. The simplest solution is composed of a **solute** dissolved in a **solvent**. The solute is the substance that is least abundant in the solution, while the solvent is the substance that is most abundant. Solutions are commonly named after the solute, for instance when 10.0 g of potassium chloride is dissolved in 100. mL of water, we would refer to the solution as a potassium chloride solution.

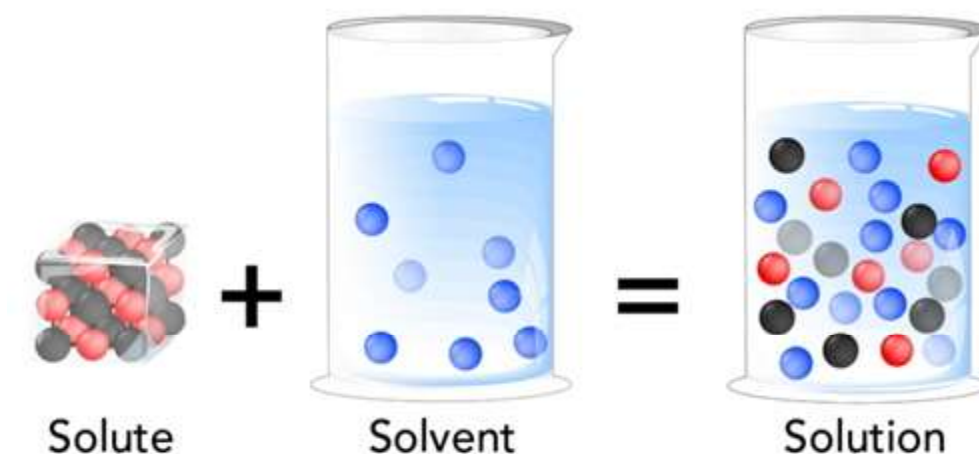


Figure 2.1: A solute is dissolved in a solvent to create a solution. credit: unknown (modified by author)

Mixtures are homogeneous, the solute is uniformly distributed throughout the solvent and will not form separate phases. Mixtures may have variable compositions, meaning the amount of solute dissolved in the solvent can vary. The particles of a solution are generally small molecules or ions, larger molecules tend to make colloids.

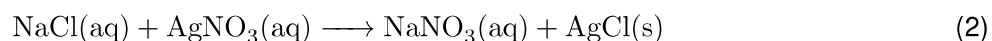
## Experiment 2 Solubility

Solutions can be formed from any combination of the normal phases of matter (solid, liquid, and gas), with the most commonly encountered types of solution being solids in liquids or liquids in liquids.

Original state of solute	Solvent	Examples
gas	gas	air; natural gas; oxygen-acetylene mixture used in welding
gas	liquid	carbonated drinks; water in rivers and lakes containing oxygen
gas	solid	hydrogen in platinum
liquid	gas	water vapour in air; gasoline-air mixture
liquid	liquid	alcohol in water; antifreeze in water
liquid	solid	amalgams, such as mercury in silver
solid	gas	mothballs in air
solid	liquid	sugar in water; table salt in water; amalgams
solid	solid	alloys, such as the copper-nickel alloy used to make coins

Figure 2.2: Nine different types of solutions. credit: unknown

Solutions are important in chemistry because many reactions will only occur at a meaningful rate in solutions. For instance solid sodium chloride and silver nitrate will not react, while a solution of sodium chloride and silver nitrate will react instantly to produce a precipitate.



The rate of a reaction is directly proportional to the number of molecules that can interact, and how much kinetic energy each molecule has, thus reactions between solids are slow (very few molecules can interact, and low energy) compared to those between solutions (large numbers of molecules can interact, higher energy). We will study the kinetics of a reaction in a future lab!

## Solids in Liquids

The rate at which a solid solute dissolves in a liquid solution depends on a number of factors. Solids can only dissolve at the surface that is in contact with the solution, thus factors which effect this can cause a solid to dissolve at different rates. We will examine each of these factors in turn:

- Particle size
- Temperature
- Concentration of solute
- Agitation or stirring.

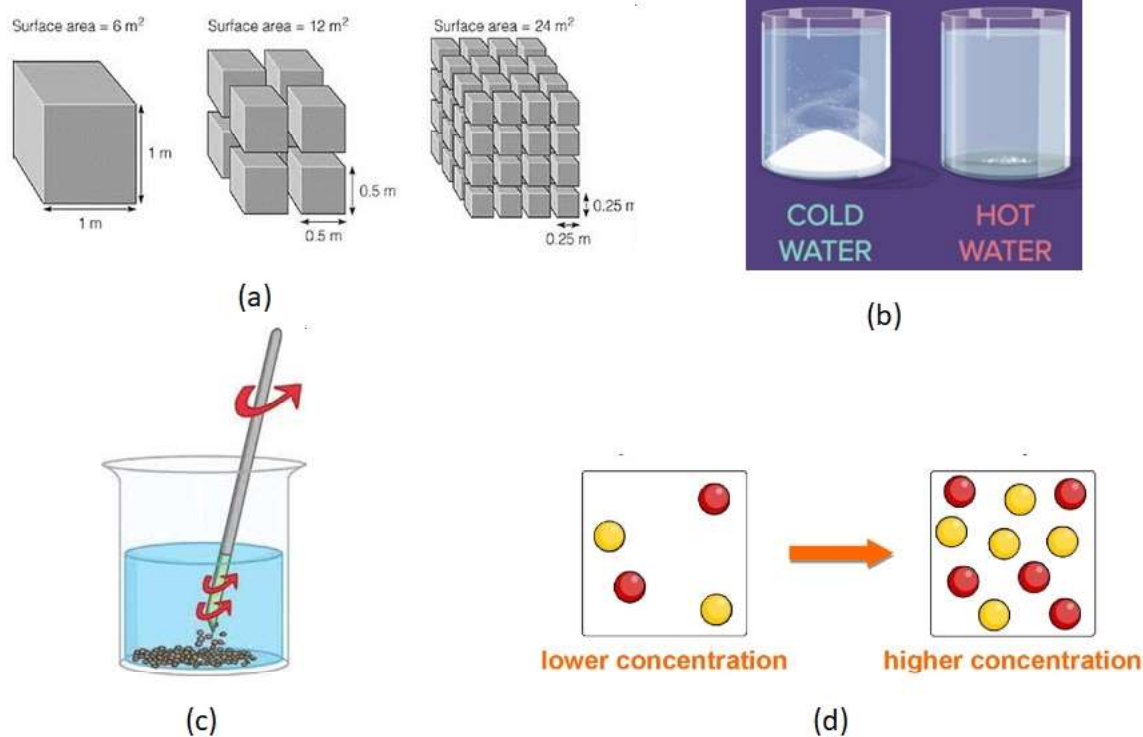


Figure 2.3: 4 factors affecting rate of dissolving (a) particles size/surface area, (b) temperature (c) stirring (d) concentration credit: unknown

**Solubility** is defined as the maximum amount of solute that will dissolve in a specific amount of solvent under stated conditions (normally temperature). Solubility is often expressed qualitatively as being **soluble** (will form an aqueous solution), **slightly soluble** (small amounts will form an aqueous solution), or **insoluble** (the solute will not dissolve in the solution). Solubility can also be expressed quantitatively as the grams of solute that will dissolve in 100 grams of water. Table 2.1 shows the solubility data for four different ionic salts in water.

## Experiment 2 Solubility

Temperature (°C)	KCl (g)	NaCl (g)	KBr (g)	BaCl <sub>2</sub> (g)
0.0	27.6	35.7	53.5	31.6
10.	31	35.8	59.5	33.3
20.	34	36	65.2	35.7
30.	37	36.3	70.6	38.2
40.	40.	36.6	75.5	40.7
50.	42.6	37	80.2	43.6
60.	45.5	37.3	85.5	46.6
70.	48.3	37.8	90	49.4
80.	51.1	38.4	95	52.6
90.	54	39	99.2	55.7
100.	55.6	39.8	104	58.8

Table 2.1: Solubility Data for Four Salts (g solute/100 g water).

## Liquids in Liquids

To qualitatively describe the solubility of two liquids we use the term **miscibility**. A **miscible** solution is one in which the liquids are soluble in each other and will mix completely, while the term **immiscible** is used to describe two liquids that are insoluble in each other, thus will not mix, and will show a visible line separating the two liquids (or phases). An example of liquids that are miscible would be ethanol and water (beer), while the phrase “oil and water don’t mix” describes two substances that are immiscible.



(a)



(b)

Figure 2.4: (a) Example of two miscible liquids mixing. (b) Example of two immiscible liquids that don't mix and clearly show a phase boundary between the liquids. credit: unknown

## Concentration - Qualitative Descriptions

The **concentration** of a solution is a measure of how much solute is dissolved in a solution. It can be given both qualitatively and quantitatively.

Qualitatively a **dilute** solution contains relatively small amounts of solute in a given volume, while a **concentrated** solution contains relatively large amounts of solute in a given volume of solution. The terms are not well defined, and are normally only used to distinguish between two solutions.

A solution can also be described as saturated, unsaturated or supersaturated. A **saturated** solution contains the maximum possible amount of solute dissolved in a solution at a given temperature and pressure. For example a solution consisting of 42.6 grams of KCl dissolved in 100 grams of water at 50 °C would be a saturated solution. A saturated solution is an **equilibrium** solution in which the rate at which solute dissolves is equal to the rate at which solute precipitates from the solution.



A **unsaturated** solution is one in which less than the maximum amount of solute is dissolved in the solution. For example at 50 °C any amount of KCl less than 42.6 grams dissolved in 100 grams of water would be unsaturated.

A **supersaturated** solution contains more than the maximum amount of solid dissolved in the solution. This may seem like a contradiction of the term “maximum amount”, but a supersaturated solution is an inherently unstable state, and will form a saturated solution (with excess solid precipitating out) if disturbed. An example of this would be making Jello, a solution of Jello in water is made at a high temperature and then allowed to cool, as the solution cools, the solubility of the Jello decreases, a supersaturated solution forms, and if disturbed will form a solid.



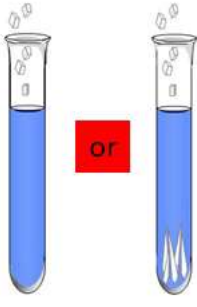
Unsaturated (U)	Saturated (S)	Supersaturated (SS)
<ul style="list-style-type: none"> <li>○ [Solute] &lt; Max</li> <li>○ Clear Solution</li> </ul>	<ul style="list-style-type: none"> <li>○ [Solute] = Max</li> <li>○ Clear Solution + tiny amount solid</li> <li>○ Equilibrium</li> </ul>	<ul style="list-style-type: none"> <li>○ [Solute] &gt; Max</li> <li>○ Clear Solution</li> <li>○ Inherently unstable</li> </ul>
		
(a)	(b)	(c)

Figure 2.5: The difference between unsaturated, saturated and supersaturated solutions. credit: author

## Concentration - Quantitative Descriptions

Quantitatively the amount of solute dissolved in solution can be expressed in a number of different ways depending on what is being described. The most common expressions are given below.

**Mass-percent solution** is defined as the mass of solute divided by the mass of the solution. For example, a 10% solution of NaCl in water contains 10 grams of NaCl in 90 grams of water, or 10 grams of NaCl in 100 grams of solution.

$$\text{Mass Percent} = \left( \frac{\text{grams solute}}{\text{grams solute} + \text{grams solvent}} \right) * (100) = \left( \frac{\text{grams solute}}{\text{grams solution}} \right) * (100) \quad (4)$$

**Mass per 100 grams solvent** is similar to Mass-percent, but is expressed in terms of 100 grams of solvent not solution. Thus, the 10 mass-percent solution described above, the Mass per 100 grams solvent would be 11 grams NaCl per 100 grams of H<sub>2</sub>O.

$$\left( \frac{10 \text{ grams NaCl}}{90 \text{ grams H}_2\text{O}} \right) * (100) = 11 \text{ grams} = \frac{11 \text{ grams NaCl}}{100 \text{ grams H}_2\text{O}} \quad (5)$$

**Molarity** is the most common quantitative expression of concentration used in chemistry. Molarity is defined as the moles of solute per liter of solution. The concentration of 10 grams of NaCl in 90 grams of water discussed above would produce a 1.90 M solution of NaCl.

$$\text{Molarity} = M = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}} \quad (6)$$

## Enthalpy of Solution

When a substance dissolves in a solution there is an associated change in the Gibbs Free Energy ( $\Delta G$ ), Enthalpy ( $\Delta H$ ) and Entropy ( $\Delta S$ ). A complete discussion can be found in OER Chapter 11.1 and 11.2. The change in entropy is difficult to measure, so we will focus instead on the change in the Enthalpy.

The enthalpy change gives insight into the Intermolecular Forces (IMF) between the molecules (in both the solute and solvent as they dissolve. For a review of IMF's read OER 10.1 or Experiment 20 - Intermolecular Forces. When a substance dissolves there are three types of interactions between the molecules that must be considered:

1. solute-solute interactions - one must break the Intermolecular Forces (IMF) or bonds between the molecules in order to separate them, this process requires energy (is endothermic) and  $\Delta H$  is positive.
2. solvent-solvent interactions - one must break the IMF between solution molecules in order to allow the solvent particles to dissolve, this process also requires energy (and is endothermic) and  $\Delta H$  is positive.

3. solvent-solute interactions - When solvent molecules interact (via IMF) energy is released (exothermic), and  $\Delta H$  is negative.

Depending on the strength of the interactions, the solution may increase in temperature (exothermic, think of hot packs) or decrease in temperature (endothermic, think of cold packs).

To calculate the energy change review OER 9.1-9.2. The enthalpy change of the water is given by Equation 7 where  $m$  = mass of solution,  $s$  = specific heat (we will assume it is the same as pure water since it is a dilute solution), and  $\Delta T$  = change in temperature.

$$q = m \cdot s \cdot \Delta T \quad (7)$$

The heat gained or lost by the water is due to the Enthalpy of Solution ( $\Delta H_{soln}$ ) due to mixing of the solute and solvent. The Enthalpy of Solution can be calculated from Equation 8 where  $n$  = moles (or mass) of solute. More information can be found in OER Textbook 11.2.

$$q = n \cdot \Delta H_{soln} \quad (8)$$

The two equations (and the chemical system) are coupled in that the heat gained or lost by the water is lost or gained due to the mixing of solute and solvent molecules. We can then combine the equations, being careful to note that the left side will use the mass of the solution (solute + solvent), and use the heat capacity of the solution, while we will use the moles of solute on the right hand side (instead of mass). Additionally we will add a negative sign to account for the heat gained/lost being opposites.

$$\begin{aligned} \text{Solution (water + solute)} &= - \text{Enthalpy of Solution} \\ m \cdot s \cdot \Delta T &= -n \cdot \Delta H_{soln} \end{aligned} \quad (9)$$

One can visualize the process of Solution Formation as shown in Figure 2.6.

## Experiment 2 Solubility

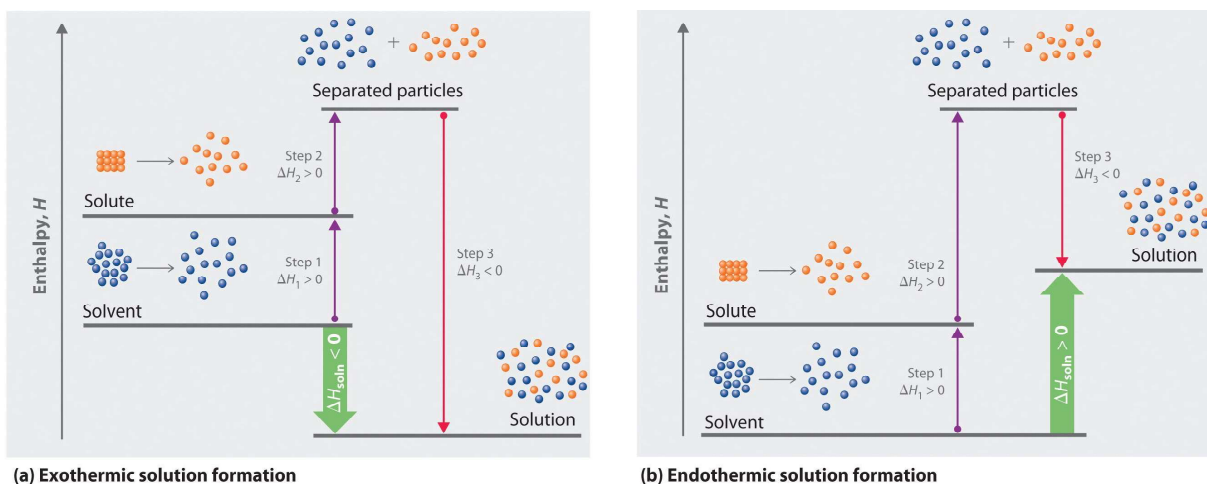


Figure 2.6: Dissolving process for an (a) exothermic process and (b) endothermic process. credit: "Factors Affecting Solution Formation" by LibreTexts is licensed under CC BY-NC-SA.

## Procedure

The experiment is broken apart into several sections. The sections may be completed in any order. The first and last parts take the longest and should be started at the beginning of the lab period, one each day.

### Determining the Concentration of a Saturated Solution

In this section you will determine the concentration of KCl in a saturated solution and compare it to the theoretical value. Make all measurements to the highest precision possible. Perform only one trial, use the second trial only if you make a mistake in trial 1. The method used in this section is referred to as "heating to a constant mass", and involves heating the sample multiple times to ensure that all of the water is removed.

1. Prepare a water bath using a 400 mL beaker 3/4 full of water on which you will place an evaporating dish. Before you heat the water, make sure the evaporating dish will fit on top of the beaker (and won't fall in).
2. Record the temperature of the room.
3. Weigh an empty evaporating dish. Obtain 5-6 mL of saturated potassium chloride solution and pour it into the dish. Weigh the dish + potassium chloride solution. Record all masses on your data table.
4. Place the evaporating dish on the beaker of water. Boil the solution until all of the potassium chloride solution has evaporated to dryness (it should take 30-45 minutes). You may need to add water to the beaker (**DO NOT** let it boil dry, you will shatter your beaker and receive a zero for this lab). You may want to periodically break the crust formed over the solution of potassium chloride to speed evaporation up.
5. While evaporating perform the other sections of the lab.



6. When all of the liquid in the potassium chloride solution in the evaporating dish has evaporated cool the dish.
7. When cool, wipe any excess water off the bottom of the dish with a paper towel. Place the dish on the wire gauze and heat gently for an additional 1-2 minutes to evaporate any last traces of water. Do not heat strongly because your sample might splatter.
8. Allow your sample to cool and weigh it.
9. Reheat your sample on the wire gauze for an additional 1-2 minutes, allow the sample to cool and reweigh it.
10. When finished allow the dish to cool, dissolve the remaining potassium chloride in water, and dispose of it in the sink.

### Solubility versus Temperature (Comparing Saturated and Unsaturated Solutions)

1. Label 4 weighing boats or papers and weigh out the following: 1.0 g NaCl, 1.4 g NaCl, 1.0 g  $\text{NH}_4\text{Cl}$  and 1.4 g  $\text{NH}_4\text{Cl}$ .
2. Add 1.0 grams of NaCl and  $\text{NH}_4\text{Cl}$  to **separate** test tubes labeled A and B. Add 5.0 mL of distilled water to each. Mix each test tube for 2 minutes. Record your observations by circling the appropriate answers.
3. Add 1.4 grams of NaCl to test tube A, and  $\text{NH}_4\text{Cl}$  to test tube B. Mix each test tube for 4 minutes. Record your observations.
4. Place test tube A and B into a beaker of boiling water. Stir occasionally (caution, hot things are hot). Note the results after 5 minutes.
5. Remove the tubes and run under tap water for 2 minutes. Let stand for 5 minutes. Record your observations.
6. Dispose of the solutions in the sink.

### Relative Solubility of a Solute in Two Solvents

1. Add 2 mL of heptane and 5 mL of water to a test tube. Mix vigorously (as demonstrated in class) the contents for 30 seconds. Record your results. Wait 2 minutes, and record your results.
2. To the previous test tube add 1 mL of saturated iodine-water solution. Note the color of each layer. Mix the contents for 30 seconds. Record your results. Wait 2 minutes, and record your results.
3. Dispose of the contents of the test tube in the waste bottle marked "Volatile Wastes".

## Experiment 2 Solubility

### Miscibility of Liquids

1. Obtain 3 dry test tubes.
2. In test tube 1 mix 1 mL of kerosene and 5 mL of isopropyl alcohol.
3. In test tube 2 mix 2 mL of kerosene and 5 mL of water.
4. In test tube 3 mix 5 mL of water and 2 mL of isopropyl alcohol.
5. Mix each of the three test tubes for 15 seconds. Note which pairs are miscible and which are not.
6. Dispose of the contents of the test tubes in the waste bottle marked "Volatile Wastes".

### Rates of Dissolving - Particle Size

1. Obtain two dry test tubes. Fill the first test tube with 0.5 grams of fine NaCl. Fill the second tube with 0.5 g of coarse (Rock Salt or Tablets) NaCl.
2. Add 10 mL of water to each test tube.
3. Stir both test tubes at the same time, noting the number of seconds each takes to fully dissolve. (You should not have to shake the tubes for more than 2 minutes, if you do inform your instructor).
4. Dispose of both solutions in the sink.

### Rates of Dissolving - Temperature

1. Weigh out two 0.5 gram samples of fine NaCl.
2. Boil 200 mL of water on a hot-plate.
3. Fill one beaker with 50. mL room temperature water and a second beaker with 50. mL of boiling water.
4. Add the 0.5 gram sample of NaCl to each beaker and record the time required for each to dissolve completely. Record your results.
5. As soon as the crystals are dissolved in the hot water, take the beaker containing the hot solution, and slowly tilt it back and forth. Observe the layer of denser salt solution on the bottom. Repeat this step with the cold solution.
6. Dispose of both solutions in the sink.

### Enthalpy of Solution

1. Obtain two styrofoam coffee cups, a temperature probe (we will not use the thermometers as shown in Figure 2.7), a cardboard lid, and a stirring rod, and a 50 mL graduated cylinder.
2. Weight your empty calorimeter (the two styrofoam cups), and record the result.

3. Measure out approximately 50.0 mL of room temperature distilled water using the graduated cylinder, pour it in the calorimeter and record the mass of the water used.
4. Place the temperature probe in the water and wait 5 minutes for the water/probe to reach equilibrium.
5. While waiting for the temperature probe to reach equilibrium set up the data recording parameters to take data every 5 seconds for 10 minutes.
6. Weigh out 0.1 moles of the substance to be tested. Record the mass
7. Start taking data with the temperature probes, after 2-3 measurements, add the substance being tested to the calorimeter, quickly place the lid on the calorimeter and begin slowly stirring the solution.
8. Collect data until the temperature is done changing (it should take less than 5 minutes).
9. Record the Initial and Final temperature readings.
10. Save your data in case you need to plot it or check it again.
11. Empty the solution down the drain.
12. Repeat the measurement for the listed substances in the Results section.

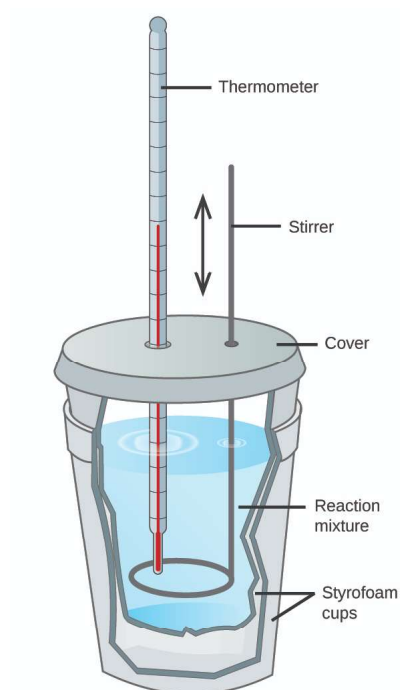


Figure 2.7: Calorimeter setup. credit: <https://courses.lumenlearning.com/openstax-chemistry/chapter/calorimetry/>

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Date: \_\_\_\_\_

## Results

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### Determining the Concentration of a Saturated Solution

	<b>Trial 1</b>	<b>Trial 2</b>
1. Room Temperature:	_____	_____
2. Mass of evaporating dish:	_____	_____
3. Mass of evaporating dish + KCl solution:	_____	_____
4. Mass of evaporating dish + dry KCl after 1 <sup>st</sup> heating:	_____	_____
5. Mass of evaporating dish + dry KCl after 2 <sup>nd</sup> heating:	_____	_____
6. Mass of solution. (show calc.)	_____	_____
7. Mass of the KCl (solute) (show calc.):	_____	_____
8. Mass of H <sub>2</sub> O (solvent) (show calc.):	_____	_____
9. Mass percent of KCl in solution (show calc.):	_____	_____
10. Grams of KCl per 100 grams water (experimental) (show calc.):	_____	_____
11. Grams of KCl per 100 grams water (theoretical) (Explain):	_____	_____
12. Percent Error in grams KCl per 100 grams water (show calc.):	_____	_____
13. In determining the concentration of potassium chloride in a saturated solution you recorded the temperature of the room. Why?		

Experiment 2 Solubility

**Solubility versus Temperature (Comparing Saturated and Unsaturated Solutions)**

Experiment	NaCl	NH <sub>4</sub> Cl
1.0 g X + 5 mL water	Dissolved: ( yes / no ) Saturated / Unsaturated	Dissolved: ( yes / no ) Saturated / Unsaturated
1.0 g X + 5 mL water + 1.4 g X	Dissolved: ( yes / no ) Saturated / Unsaturated	Dissolved: ( yes / no ) Saturated / Unsaturated
2.4 g X + 5 mL water + heat	Dissolved: ( yes / no ) Saturated / Unsaturated	Dissolved: ( yes / no ) Saturated / Unsaturated
2.4 g X + 5 mL water + cooled	Dissolved: ( yes / no ) Saturated / Unsaturated	Dissolved: ( yes / no ) Saturated / Unsaturated

1. At room temperature what can you say about the solubility of NaCl? Be as specific as you can (include numbers).
2. At room temperature what can you say about the solubility of NH<sub>4</sub>Cl? Be as specific as you can (include numbers).
3. At high temperature (boiling) what can you say about the solubility of NaCl? Be as specific as you can (include numbers).
4. At high temperature (boiling) what can you say about the solubility of NH<sub>4</sub>Cl? Be as specific as you can (include numbers).

### **Relative Solubility of a Solute in Two Solvents**

1. Sketch a picture of the test tube containing heptane + water after 30 seconds and after two minutes (label your pictures).
2. Which liquid is denser? Explain (based on your experimental observations).
3. Sketch a picture of the test tube containing heptane + water + iodine after 30 seconds and after two minutes (label your pictures, including colors).
4. In which solvent is iodine more soluble? Explain (based on your experimental observations).
5. Does your experimental results match with our general rule of "like dissolves like"? Explain.

## Experiment 2 Solubility

### Miscibility of Liquids

1. Which liquid pairs tested were miscible?
2. Which liquid pairs tested were immiscible?
3. From the previous section how would you classify the mixture of heptane and water? Explain.

### Rates of Dissolving - Particle Size

1. Time required for fine salt crystals to dissolve: \_\_\_\_\_
2. Time required for coarse salt crystals to dissolve: \_\_\_\_\_
3. The amount of salt, the volume of water, and the temperature of the systems were kept constant between the two test tubes, only particle size was varied. What trend was observed? Explain why the trend occurs.

### Rates of Dissolving - Temperature

1. Time required for fine salt crystals to dissolve in the hot solution: \_\_\_\_\_
2. Time required for fine salt crystals to dissolve in the cold solution: \_\_\_\_\_
3. The amount of salt, the volume of water, and the particle size of the salt particles were kept constant between the two test tubes, only temperature was varied. What trend was observed? Explain why the trend occurs.



## Enthalpy of Solution

Measurement	Trial 1 $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$	Trial 2 $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$	Trial 1 $\text{NH}_4\text{NO}_3$	Trial 2 $\text{NH}_4\text{NO}_3$
Moles Solute (mol):	0.1	0.1	0.1	0.1
Grams Solute (calc) (g):				
Grams Solute (used) (g):				
Mass of Calorimeter (g):				
Mass of Calorimeter + Water (g):				
Mass of Water used (g):				
Initial Temperature ( $^{\circ}\text{C}$ ):				
Final Temperature ( $^{\circ}\text{C}$ ):				
Change in Temperature ( $^{\circ}\text{C}$ ):				
Calculated Enthalpy of Solution (kJ/mol):				
Accepted Enthalpy of Solution (kJ/mol):				
Percent Error:				

In the space below show any necessary calculations needed (at a minimum show the calculation for the Enthalpy of Solution).

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1. Average Enthalpy of Solution for  $\text{CaCl}_2$ : \_\_\_\_\_
2. Average Enthalpy of Solution for  $\text{NH}_4\text{NO}_3$ : \_\_\_\_\_
3. Which compound gave a positive Enthalpy of Solution? Explain.
4. Which compound gave a negative Enthalpy of Solution? Explain.
5. Explain how both substances dissolved spontaneously yet one released energy and one absorbed energy.
6. What can one conclude about the sign of the Entropy of Solution for  $\text{CaCl}_2$  dissolving? Explain.
7. Which compound is likely to have the larger change in Entropy? Explain.
8. Your results are likely to not be very accurate due to several experimental limitations. List two assumptions made in experiment that could impact the accuracy of your results.

## Post Lab Questions

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Using the solubility data in Table 2.1 answer the following questions.

1. What is the percentage by mass of NaCl in a saturated solution of sodium chloride at 50.0°C. Explain. 1. \_\_\_\_\_
2. Calculate the solubility of potassium bromide (by interpolation) at 23°C. Hint: Assume that the solubility increases by an equal amount for each degree between 20.0°C, and 30.0°C. Explain. 2. \_\_\_\_\_
3. A saturated solution of barium chloride at 40.0°C contains 100. grams of water. How much additional barium chloride can be dissolved by heating the solution to 70.0°C? Explain. 3. \_\_\_\_\_
4. A saturated solution of barium chloride at 30.0°C contains 25.0 grams of water. How many grams of barium chloride is dissolved in the water? Explain. 4. \_\_\_\_\_
5. A saturated solution of barium chloride at 30.0°C contains 150. grams of water. How many additional grams of barium chloride can be dissolved by heating the solution to 60.0°C? Explain. 5. \_\_\_\_\_

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6. A solution of KCl is saturated at 50 °C.

(a) How many grams of solute are dissolved in 100. grams of water? 6(a) \_\_\_\_\_

(b) What is the total mass of the solution? Explain. 6(b) \_\_\_\_\_

(c) What is the mass percent of this solution? Explain. 6(c) \_\_\_\_\_

(d) If the solution is heated to 100 °C, how much more KCl can be dissolved in the solution without adding more water? Explain. 6(d) \_\_\_\_\_

(e) If the solution is saturated at 100 °C, and then cooled to 30 °C, how many grams of solute will precipitate out? Explain. 6(e) \_\_\_\_\_

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## Prelab Questions

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1. In the discussion section, what class of reaction occurred in equation (2)?
2. What does the term homogeneous mean?
3. Using Excel, graph the data in Table 2.1 showing the solubility of the four salts as a function of temperature. Attach the graph to the back of the pre-lab. Include a curve fit, the equation of the line, and the  $R^2$  value for each of the sets of data.
4. Using the graph **qualitatively** what amount of KCl that will dissolve in 100 grams of water at 95°C. Explain.
5. Using the graph **quantitatively** determine the amount of  $\text{BaCl}_2$  that will dissolve in 100 grams of water at 43°C. Explain.
6. Prove that a 10-mass percent solution of NaCl is the same as a 1.90 M solution of NaCl.

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