

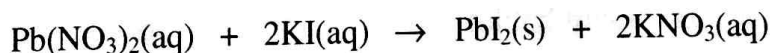
## CHEMISTRY 205

**Experiment 4: PREPARATION OF SOME METAL HALIDES****Experiment 4A: PREPARATION OF LEAD(II) IODIDE****Introduction**

When ionic compounds (*i.e.*, salts) are dissolved in water, the cations and anions dissociate from each other and become surrounded by water molecules, or *solvated*. The process of dissolution involves disrupting the attractive forces between ions, and replacing them with attractions between ions and solvent molecules – because there are so many forces involved, it is very difficult to predict exactly how soluble any particular ionic compound will be! By making careful observations, chemists have devised general "solubility rules" to help us remember which types of salts dissolve readily in water and which do not. Generally, alkali metal (group I) salts, ammonium salts and nitrate salts are always highly soluble, and so are *most* halide and sulfate salts (*e.g.*,  $\text{NaNO}_3$ ,  $\text{NaCl}$ ,  $\text{NaOH}$ ). Conversely, salts containing hydroxide, sulfide, carbonate or phosphate are usually insoluble (*e.g.*, limestone,  $\text{CaCO}_3$ ; a component of rust,  $\text{Fe}(\text{OH})_3$ ), but it is important to note that the word "insoluble" describes a range of solubilities. "Slightly soluble" compounds will dissolve a little (*e.g.*,  $\text{Ca}(\text{OH})_2$ ), while "sparingly soluble" salts are almost entirely insoluble (*e.g.*,  $\text{Fe}(\text{OH})_3$ ). In Chem 206, you will learn how to treat solubility quantitatively, but the general solubility rules you will find in your textbook are essential to remember.

NOTE: There are some critical exceptions to the rules; for example, silver and lead halide salts are actually highly insoluble. The solubility rules are only useful when you remember the key exceptions as well.

In this experiment, you will perform a simple precipitation reaction in which you form an insoluble lead halide. Two solutions of carefully selected soluble salts will be mixed together, and an insoluble ionic compound will precipitate out as a solid, leaving soluble spectator ions in the overlying *supernatant* solution. You will isolate the solid product by filtration and measure its mass, then based on the quantities of starting materials used, you will calculate your percent yield. The reaction equation is:



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**Procedure Summary**

A solution of lead(II) nitrate will be added to a solution of potassium iodide. The solid lead(II) iodide formed will be collected by filtration, dried and weighed on the filter paper. The product will also be scraped off the filter paper and weighed again to illustrate how much yield is routinely lost when a product is isolated.

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**Prelaboratory Assignment**

Read the procedure and answer the preliminary questions before coming to the lab period. Your demonstrator will inspect and collect in your prelab before you are permitted to begin the experiment -- keep a copy of it for yourself, and have the TA sign your receipt record.

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**Materials**

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**Apparatus**

- lab coat & safety glasses
- 50 mL beaker
- 100 mL beaker
- 10 mL graduated transfer pipette (use a rubber bulb to fill)
- 2 mL volumetric pipette (use a rubber bulb to fill)
- glass rod with rubber policeman
- 125 mL Erlenmeyer (conical) flask
- glass funnel
- filter paper
- small piece of bent wire, or a paperclip
- Pasteur pipette with rubber bulb
- spatula

**Materials**

- 0.250 M lead(II) nitrate solution [Caution:  $\text{Pb}^{2+}$  is toxic; avoid contact]
- 0.250 M potassium iodide solution
- deionized water
- acetone (squirt bottle)

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**Procedure**

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1. Using a 10 mL graduated transfer pipette (see below), measure out exactly 7.0 mL of 0.250 M potassium iodide solution and transfer it into a clean 100 mL beaker. Describe the liquid's colour and whether it is clear (*i.e.*, transparent) or turbid (*i.e.*, cloudy).

**Using a graduated transfer pipette:** Pour ~10 mL of solution into a clean dry small beaker – this way you avoid placing the pipette directly into the stock bottle, which could contaminate the stock solution. Do NOT rinse the pipette with water before using it; if it is already wet, rinse it first with a small volume of the solution you wish to measure.

**To measure your solution:** Draw the solution into the pipette using your rubber bulb until it is above the 7.0 mL mark. Then quickly remove the bulb and place your index finger over the top to stop the liquid running out. Allow the level to drop to the 7.0 mL line by letting air to leak past your finger. Then keep your finger firmly in place. Carefully lift the pipette from the solution, and touch the pipette tip to the side of the beaker to remove any drops clinging to the outside of the pipette. Drain the liquid into your clean, dry 100 mL beaker, touching the pipette tip to the inside of the beaker to help draw the liquid out.

**NOTE:** Use the same technique for volumetric pipettes, except for this: with a volumetric pipette, do NOT force out the small droplet that hangs inside the pipette's tip, because the pipette has been calibrated to take this little droplet into account.

2. Using a 2 mL volumetric pipette, measure out exactly 2.0 mL of 0.250 M lead(II) nitrate solution – use the same technique as above, but less volume so you don't waste solution. Note your observations of this solution (while it is in your beaker!). Carefully add the lead nitrate solution from the pipette to the beaker containing the potassium iodide solution. Thoroughly mix the reagents by gently swirling the beaker, being careful not to spill any of the mixture. Note your observations of what happened when you mixed the reagents, and the appearance of both the precipitated solid and the supernatant solution.

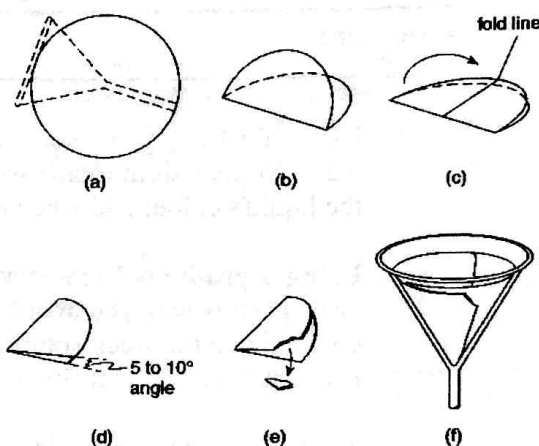
The next steps describe how to go about collecting the precipitated  $\text{PbI}_2$  by gravity filtration.

3. First: prepare the filter paper. Obtain a large circle of filter paper with a radius just slightly smaller than the height of the glass funnel's cone.

#### Exp.4: SYNTHESIS

4. The paper must now be folded into the shape of a cone so it can sit inside the funnel. Follow the directions below, and refer to the diagrams (a) – (f) at the right:

- Fold the paper carefully in half.
- Fold it in half again, but not quite perfectly. You want the two straight edges to be slightly offset, with a 5-10° angle between them. [This will make the cone fit more snugly into the funnel.]
- Tear off a small corner of the smaller section.
- Open up the larger section of the folded paper to make a cone.



[Diagram taken from N.E. Griswold, H.A. Neidig, J.N. Spencer, C.L. Stanitski. *Laboratory Handbook for General Chemistry*, 2<sup>nd</sup> Ed. Brooks/Cole Thomson Learning: 2002, p.46.]

5. Tare the empty balance and weigh the dry folded filter paper (without the funnel!). Record its mass on your data sheet; you will need this later to calculate the mass of your precipitate.
6. Place the filter paper into the funnel. Dampen it with a few milliliters of deionized water and make sure the paper sticks evenly to the glass. Next, place the filter into the top of a clean 125 mL Erlenmeyer flask. To keep the funnel from forming a tight seal in the mouth of the flask, place a small piece of bent wire (or a bit of rolled-up paper towel) over the rim of the flask's mouth for the funnel to rest upon (this leaves an air vent).
7. Gently swirl the beaker containing the precipitated lead(II) iodide to suspend the solid, and carefully pour the suspension into the filter. Use the rubber policeman (glass rod with rubbery tip like a kitchen spatula) to scrape as much of the solid as possible onto the filter. Some  $\text{PbI}_2$  will remain adhering to the beaker.
8. While gravity pulls the liquid through the filter, make observations of the filtrate that drips into the collecting flask. When all the liquid has drained through, transfer the filtrate from the collecting flask back to the beaker. Use the rubber policeman to dislodge as much  $\text{PbI}_2$  as possible from the glass, then pour this suspension back onto the filter to collect the rest of the precipitate. When all the liquid has drained again, discard the filtrate in the waste container provided.
9. The solid collected on the filter paper will contain  $\text{PbI}_2$  with traces of unreacted reagents and spectator ions. To wash these impurities away, the solid must be rinsed with a very small volume of deionized water: use a Pasteur pipette (eyedropper) to apply 2 mL of deionized water dropwise evenly over the filter paper and solid. Note the appearance of the filtrate now – the solubility of  $\text{PbI}_2$  is very low in pure water, but small amounts of it will dissolve while washing it. Once the water has drained, discard the filtrate.

#### Exp.4: SYNTHESIS

10. Use a squirt bottle to apply a few milliliters of acetone evenly over the filter and solid. [This will help the solid dry more quickly, because water adhered to the solid will dissolve in the acetone and be rinsed away. Mostly acetone will be left adhering to the solid, and acetone evaporates much more quickly than water.] When the acetone has passed through the filter, discard the acetone filtrate into the waste container provided for acetone (located in the fume hood).
11. Label a clean, dry 100 mL beaker with your name. Carefully lift the filter paper containing your product out of the funnel, and rest it inside the beaker so it will not fall over. Place the beaker inside the fume hood to dry; it will likely take over 30 minutes. The paper and solid should both look dry before you continue.

WHILE WAITING: perform Experiment 4B in its entirety.

12. When your product is dry: tare an empty balance and weigh the dry filter paper and  $\text{PbI}_2$ . Calculate the total mass of lead(II) iodide collected. Use this mass to calculate your percent yield of product.
13. Weigh a clean dry watch glass. Using a spatula, scrape as much of the  $\text{PbI}_2$  as possible off the filter paper and onto the watch glass without scraping bits of paper off too. You will notice that it is not possible to remove all of the solid from the porous paper. Weigh the watch glass with your product, then discard your product into the lead waste container provided.
14. Clean and put away all of your glassware and other equipment, and make sure your work area is clean and tidy before having your TA sign your data sheets.

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## CHEMISTRY 205

## Experiment 4A PREPARATION OF LEAD(II) IODIDE

## Laboratory Report

## Observations, Data and Calculations

1. (2 marks) Describe what happened when you added the lead(II) nitrate to the potassium iodide solution. Include initial observations of the reagent solutions.

Both are transparent. change in color  $\rightarrow$  yellow precipitate formed.

2. (3 marks)

Volume of lead(II) nitrate solution used	20 mL
Moles of lead(II) ions	0.0005 moles
Volume of potassium iodide solution used	<del>8.0 mL</del> 7.0 mL
Moles of iodide ions	0.00175 moles
Theoretical yield of lead(II) iodide (moles)	0.0005 moles
Mass of dry filter paper	<del>0.792 g</del>
Mass of filter paper plus dry lead(II) iodide	1.054 g
Mass of lead(II) iodide prepared	0.262 g
Moles of lead(II) iodide prepared: "total yield"	0.00056 moles
Percent total yield of lead(II) iodide	113%

Phanir  
24/10



Mass of empty watch glass	21.484 g
Mass of watch glass plus lead(II) iodide	21.676 g
Mass of lead(II) iodide isolated	0.192 g
Moles of lead(II) iodide isolated: "isolated yield"	$1.085 \times 10^{-4}$ moles
Percent isolated yield of lead(II) iodide	21.44%
Moles of lead(II) iodide lost during isolation	
Percent of total yield lost during isolation of product	

### Conclusions

- (2 marks) Chemists always report the yield (as a %) when they perform a reaction, because it shows how efficient the reaction is as a method of preparing the desired product. When would it be important to report both the total yield and the isolated yield of a reaction? Did it matter very much for your reaction in this experiment? Explain.
- (3 marks) Imagine you wanted to perform an experiment similar to this one, but you only have access to solutions of lead(II) nitrate and potassium sulfate.
  - Would a precipitation reaction occur if you mixed these two solutions? Explain.
  - Write a balanced molecular equation to describe the outcome of mixing these solutions.

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#### Exp.4: SYNTHESIS

- 2c. What would be the yield of solid product (if one forms) if you used solutions with the same concentrations and volumes as you used in this experiment? (7 mL of 0.250M potassium sulfate and 2 mL of 0.250 lead(II) nitrate) Show your calculations.
3. (5 marks) In your experiment, one of the reagents was limiting and the other was in excess. Therefore, after the lead(II) iodide precipitated from the solution, there were spectator ions left in solution, and also the unreacted excess reagent. Assuming complete precipitation of lead (II) iodide (*i.e.*, 100% yield), calculate the concentrations of EACH type of ion ( $\text{Pb}^{2+}$ ,  $\text{K}^+$ ,  $\text{NO}_3^-$ ,  $\text{I}^-$ ) present in the solution after the  $\text{PbI}_2$  precipitated out. Show your calculations.



## Experiment 4B: SYNTHESIS OF MANGANESE(II) CHLORIDE

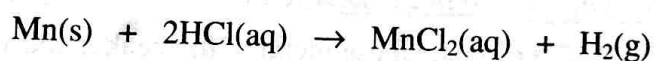
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### Introduction

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This experiment illustrates another class of reactions: oxidation-reduction, or *redox*, reactions. Oxidation and reduction are processes that involve the transfer of electrons from one species to another, and therefore always occur together. Electrons are transferred from the reducing agent (which becomes oxidized) to the oxidizing agent (which becomes reduced). You should recall from class that chemists keep track of electrons by assigning imaginary charges called *oxidation numbers* to the atoms in substances; when a species is oxidized, its oxidation number increases, and conversely when a species is reduced, its oxidation number decreases.

In this experiment you will perform a simple redox reaction between a metal and an acid. Certain metallic elements will dissolve in acidic solutions because they can be oxidized by  $H^+$ ; the  $H^+$  becomes reduced to elemental hydrogen by removing electrons from the metal. For example, when manganese metal is reacted with hydrochloric acid, the metal dissolves to form the salt, manganese(II) chloride:



You will perform and observe the above redox reaction, measure the mass of your product and calculate the percentage yield of the reaction.

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### Procedure Summary

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A sample of manganese will be dissolved in concentrated hydrochloric acid. Excess hydrochloric acid and water will be evaporated to obtain solid manganese(II) chloride.

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### Prelaboratory Assignment

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Read the procedure and answer the pre laboratory questions before coming to the lab period. Your demonstrator will inspect and collect in your prelab before you are permitted to begin the experiment - keep a copy of it for yourself.

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## Materials

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### Apparatus

- hot plate
- 100 mL beaker
- desiccator
- tongs for handling the hot beaker
- rubber gloves

### Materials

- manganese metal
- 10 M hydrochloric acid [**Caution – corrosive**]

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## Procedure

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1. Weigh the clean, dry 100 mL beaker to the nearest milligram.
2. Place 0.05 to 0.10 g of manganese metal in the beaker and weigh it again.
3. Working in a fume-hood, carefully add 30 drops (approx 1.5 mL) of 10 M hydrochloric acid using the Pasteur pipette provided.
4. Wait for the manganese to dissolve and hydrogen evolution to cease, about 5 minutes. Notice, as the reaction proceeds, whether the tube is becoming hotter (an exothermic reaction), or colder (an endothermic reaction).
5. Place the beaker on the hot plate. The liquid will begin to evaporate.
6. When the contents of the beaker are dry (you should be left with a solid) note the appearance of the product, allow the beaker to cool for at least 5 minutes in a desiccator, and then weigh it. Note the mass. *Do not weigh the beaker hot* - you will obtain an incorrect weight because of convection currents in the air, and you may even damage the balance.
7. Return the beaker to the hot plate and heat it for another 3 to 5 minutes, cool it again in the desiccator, and reweigh it. Note the mass. If the mass does not agree with the previous mass within 3 mg, repeat the heating, cooling and weighing, until it does. Note the final mass.
8. Dispose of the manganese chloride in the container provided, after redissolving it in a little water. (*Manganese is an essential metal in trace quantities, but is toxic in greater amounts and should not be dumped down the sink.*)

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## CHEMISTRY 205

## Experiment 4B SYNTHESIS OF MANGANESE(II) CHLORIDE

## Laboratory Report

## Observations, Data and Calculations

1. (2 marks) Describe your observations of the action of the hydrochloric acid on the manganese.

gas released, heat ~~was~~ released (exothermic reaction)

2. (3 marks)

Mass of empty beaker	31.901 g
Mass of beaker plus manganese	31.977 g
Mass of manganese used	0.076 g
Moles of manganese used	0.001383 g
Theoretical yield of manganese(II) chloride (moles)	0.001383 g
Mass of empty beaker (same as above)	31.901 g
Mass of beaker manganese(II) chloride	32.056 g
Mass of manganese(II) chloride prepared	0.155 g
Moles of manganese(II) chloride prepared	0.00123 mole
Percent yield of manganese(II) chloride	88.99%

*Phuanti*  
24/10

**Conclusions**

1. (3 marks) Write the oxidation half reaction and reduction half reaction involved in the redox reaction between manganese and acid, and show how they add up to give the overall reaction equation given in the introduction. Label each species with its oxidation number.
2. What effect on your calculated yield would the following experimental errors have had?
  - a. (1 mark) Incomplete evaporation of the excess hydrochloric acid.
  - b. (1 mark) Excessive heating of the product to cause decomposition (say to manganese(II) oxide).
  - c. (1 mark) Cooling the product in the open air (especially on a humid day).
3. (2 marks) Which of the above errors are the most probable?
4. (2 marks) Suppose you spilled a bit of the 10 M hydrochloric acid on the bench. What are the appropriate steps you might be asked to take to clean it up?

## CHEMISTRY 205

## Experiment 5: SOLUBILITY vs. REACTIVITY

## Introduction

**SOLUTIONS:** A solution is a mixture of pure substances that appears to be *homogeneous* on a macroscopic scale. One component is usually present as the dominant component, called the *solvent*. The other minor components (there is often only one) are considered to be *dissolved* in the solvent and are called the *solutes*. There are many kinds of solutions. The table below gives a number of examples and indicates the normal state of the solution and of its components:

Various types of solutions			
Example	State of Solution	Solution's Components	
		State of Solute	State of Solvent
Air, natural gas	Gas	Gas	Gas
Gasoline, antifreeze	Liquid	Liquid	Liquid
Brass, pewter	Solid	Solid	Solid
Soda water, "ammonium hydroxide"	Liquid	Gas	Liquid
Brine, sugar solution	Liquid	Solid	Liquid
Hydrogen in palladium	Solid	Gas	Solid
Tooth amalgam (Ag/Sn/Cu/Zn & Hg), sodium amalgam (Na & Hg)	Solid	Liquid	Solid

When the solute is present in relatively small amounts, we refer to the solution as *dilute*. If it is present as a relatively large percentage of the solution we describe the solution as a *concentrated* solution. If no more of the solute can be dissolved at a particular temperature, the solution is said to be *saturated*. In more quantitative terms, we express *concentration* in convenient units such as grams per litre ( $\text{g}\cdot\text{L}^{-1}$ ), weight percent (w/w %), parts per million (ppm), moles per litre (M) *etc.*

**SUSPENSIONS:** It is important to note that some mixtures may appear homogeneous when viewed with the unaided eye, but when examined under a microscope they are seen to contain more than one phase. These are not accurately called solutions and should be referred to as *suspensions*. Suspensions often appear cloudy or *turbid* in appearance. When turbidity is not obvious, equipment more sophisticated than the naked eye can be used to show that a suspension *scatters light*, which a true solution does not do; solutions always appear *clear* or transparent, although they can be colourless like water or coloured like filtered fruit juice. Suspensions are given special names based on the nature of the suspended material and the medium in which it is



## Exp.5: SOLUBILITY

suspended. When a solid is suspended in a liquid, it is technically called a *sol*, although often it is simply referred to as a suspension. When both phases in a suspension are liquids, that is, when one liquid is suspended as tiny droplets within another liquid, then the mixture is called an *emulsion*; an example is milk. Liquid suspended in a gas is called an *aerosol*; a common example is fog. Smoke is an example of a suspension of tiny particles of solid material (soot) in a gas (air). Indefinitely stable suspensions, known as *colloids* or *colloidal dispersions*, are very commonly encountered. The cloudy brown water seen in many rivers is typically a colloidal suspension of tiny clay (solid) particles on the order of 1  $\mu\text{m}$  in diameter.

**BACK TO SOLUTIONS & SOLUBILITY:** When a substance dissolves in a solvent, the fundamental particles of the solute (atoms, molecules or ions) break away from each other. As an example, let's consider the dissolution of a sugar cube in a glass of water. In the sugar cube, each sugar molecule has other sugar molecules as its nearest neighbours. When the sugar cube is placed in the water, sugar molecules at the solid's surface begin interacting with water molecules, and eventually they are pulled off of the solid and become completely surrounded by solvent molecules. Thus, interactions between molecules (and/or ions) are responsible for substances dissolving in a solvent. Dissolution is a physical process, not a chemical process: the molecules (and/or ions) themselves *remain intact*, but the environment around them changes.

There is a basic guiding principle based on this, which states that "like dissolves like". A substance composed of ions or polar molecules will tend to dissolve in a polar solvent, while nonpolar substances (*e.g.*, oils) will tend to dissolve in nonpolar solvents (*e.g.*, paint thinner). Thus, chemists can look at the molecular structure of a compound and get an idea of what solvents it might dissolve in. In fact, tests of solubility are often among the first tests performed on a new substance; the solubility of a substance in various solvents is a fundamental property of a substance, like melting point, colour, density *etc.*, and can be used to characterize it. In this experiment you will attempt to dissolve various substances in water, in acidic solution, and in basic solution (all of which are polar solvents, since they are water-based). The effect of temperature, stirring, and the state of aggregation of the solid solutes will also be studied.

When ionic compounds dissolve in polar solvents, the oppositely charged ions separate from one another to move freely throughout the solution, which enables the solution to conduct electricity. Solutions containing only molecules or atoms (*i.e.*, no ions) do not conduct electricity significantly. Different ionic compounds separate into ions in water to a greater or lesser extent, and the conductivity of the resulting solution will vary accordingly. You will employ a simple conductivity testing device to qualitatively investigate this phenomenon.

**SOLUBILITY vs. REACTION:** Interactions between solutes and solvents are not always "innocent". Some substances dissolve because they undergo a chemical reaction with the solvent. It is important to keep in mind some of the signs that suggest chemical reaction, such as changes in colour and evolution of gas. The evolution of heat is more difficult to interpret, because simple dissolution itself is often accompanied by temperature changes.

In this experiment, you will encounter some cases of chemical reaction that lead to dissolution. Remember that you have learned three types of chemical reaction: precipitation reactions, acid-base reactions, and oxidation-reduction reactions. You should determine which type of reaction is occurring whenever you suspect a chemical reaction has caused a solute to dissolve.



### Procedure Summary

A range of solutes will be tested with various common solvents to see what happens. In your observations, you should note whether a solution (or an emulsion, suspension, *etc.*) is formed and how quickly. Don't forget to note the initial appearances of the substances! Also note any dependency the dissolution process appears to have on temperature and particle size, and note the effect of stirring or not stirring. Be sure to make detailed observations:

- Are there colour changes that might indicate a chemical reaction rather than simple dissolution?
- If the mixture/solution seems to be evolving a gas, is it boiling, or is the gas perhaps the result of a chemical reaction, or is it just the release of air trapped in the solute?

### Pre-Laboratory Assignment

Read the procedure and answer the pre-laboratory questions before coming to the lab. Your demonstrator will inspect and collect in your prelab before you are permitted to begin the experiment - keep a copy of it for yourself, and have the TA sign your receipt record.

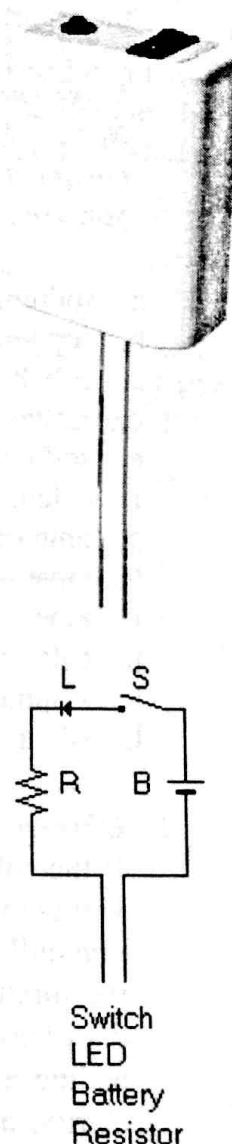
### Materials

#### Apparatus

- 24-well test plate (CLEAN!)
- Pasteur or Beral pipettes
- Conductivity tester (see diagrams on the right)
- small glass stir rods
- 40 small glass vials

#### Reagents

- See the lists in the procedure section



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**Procedure**


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**NOTE:** As you follow the following procedures, be sure to note your results in the appropriate data tables in the lab report.

You will be working with small quantities of quite concentrated hydrochloric acid and sodium hydroxide. Be careful not to get any on your skin, eyes or clothing. Wipe up small spills immediately - consult your demonstrator in the case of a serious spill. In the steps that follow, one drop *means* one drop!

1. **Prepare the vials:** Place a *very small* quantity of each of the substances you are given to study into a clean dry sample vial. (You will be given TEN of the substances listed below.) Some of the substances are expensive - use just enough to see easily. Place each vial in a well in the test plate (or test tube rack). *Make sure you note into which well you placed each substance!* (The sample well cells might be marked 1 - 6 and A - D already. If not, label the vials with a wax pencil. A space has been left beside each substance on the report form for you to record whatever ID label you used.)

**Make sure your set of substances is diverse - for example, don't test all the chlorides!**

*Note: Some choices have been crossed off the list because they give subtle or complicated results, but you could try one of them as an 11<sup>th</sup> choice if it's in the lab and you are curious.*

- |   |  |
|---|--|
| a. sodium bicarbonate                     | m. sodium carbonate  |
| b. copper(II) nitrate (hydrate)           | n. calcium carbonate                                       |
| c. cobalt (II) chloride (hydrate)         | o. calcium sulphate (hydrate)                              |
| d. nickel(II) chloride (hydrate)          | p. copper(II) sulphate (hydrate)                           |
| e. sodium chloride                        | q. <del>potassium permanganate (use 1 small crystal)</del> |
| f. calcium chloride                       | r. zinc metal (powder)                                     |
| g. ammonium chloride                      | s. tin metal (powder, if possible)                         |
| h. <del>silver chloride (expensive)</del> | t. magnesium (use ¼" max, if ribbon)                       |
| i. zinc oxide                             | u. <del>iron wire (degreased, use ½" max)</del>            |
| j. calcium oxide                          | v. carbon powder (charcoal)                                |
| k. sulphur                                | w. sucrose   |
| l. silica                                 | x. <del>iodine</del>                                       |

2. **Observe the substances and test water-solubility:** Note the appearance of the substances. If the solid consists of large crystals, gently crush them (tap with a glass stir rod) in the vial before proceeding. Then, working on the samples one at a time, use a pipette to add *one* millilitre of water (~ 25 drops from a Pasteur pipette) to each solid in its vial and stir the mixture with a glass stir rod. In each case, note the extent to which the substance dissolves (completely, partially or not at all). If the substance appears not to dissolve, try adding a few more drops of water. If a solution forms, note its colour. Also note its clarity: a true solution will be clear (transparent), whereas a suspension of undissolved fine solid particles will be cloudy.

After each test, rinse off the stir rod *thoroughly* and dry it with paper towel before moving to the next test.

## Exp.5: SOLUBILITY

3. **Test conductivity:** Using the conductivity tester, immerse both probe electrodes of the tester simultaneously in the liquid in the vial, and flip on the switch briefly. Note whether the red light shows, and how brightly. [Before starting, you should check what the maximum brightness would be by pressing the probe electrodes together and switching the tester on *for a moment only*. Be careful, of course, not to let the electrodes touch each other while testing the solutions.]

After each test, rinse off the probe electrodes *thoroughly* with distilled water and check that they are clean by testing the conductivity of a small sample of distilled water in a vial. The red light should not come on because pure water is a poor conductor. If the water has been contaminated by residues from the electrodes, the light may show. If this happens, clean the electrodes again (and replace the test sample of water).

4. **Test each substance for reaction with acid and base:** Place another small sample of each substance in a sample vial with 1 mL water (as in steps 1-2), so that you now have a pair of vials for each substance. (*You might not have enough vials to do them all at once*). For each substance, label one vial as "acid-test" and the other vial as "base-test". Now, work with each substance one at a time, to directly compare their acid vs base reactivity:

Using a pipette, add 2-3 drops of 6 M hydrochloric acid (*caution!*) to the "acid-test" vial and gently stir the mixture. Make sure the drops mix into the water and don't stick to the side of the vial. Note any indications of dissolution or reaction as appropriate. Next, add 2-3 drops of 6 M sodium hydroxide (*caution!*) to the "base-test" vial and gently stir the mixture. Again, note any indications of dissolution or reaction.

5. **Neutralize the solutions and clean up:**

To the "acid-test" vials: add 2-3 drops of 6 M NaOH to each tube.

To the "base-test" vials: add 2-3 drops of 6 M HCl to each tube.

Stir the solutions and pour the contents of each vial into the waste container provided. Use a wash bottle to rinse out any remaining solids into the waste container. Return the clean (but wet) vials to the stockroom.

Finally, rinse off any spatulas used and wipe off your bench. ☺

Name: \_\_\_\_\_ Section: \_\_\_\_\_ Date: \_\_\_\_\_

**CHEMISTRY 205****Experiment 5: SOLUBILITY vs. REACTIVITY****Prelaboratory Questions**

- (2 marks) Provide a *maximum* half-page summary of the experimental *procedure*. This can take the form of a flowchart (recommended) or a list of the main steps in the procedure (remember to both précis and paraphrase). (*Space provided on the next page.*)
- (1.5 marks) Look up (in your textbook, library reference books, or on the web) the chemical formulae for the following common substances and match them to the substances listed in the procedure. Indicate if the match is not exact.

Limestone	
Lime	
Gypsum	
Rock salt	
Sugar	
Baking soda	
Soda	
Sand	

- (1.5 marks) Look up the formulae of those substances in the list in Part 1 of the procedure which are *NOT* listed in question 2.

Name	Formula		Name	Formula