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 ($g \cdot L^{-1}$), 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

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 μm 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.

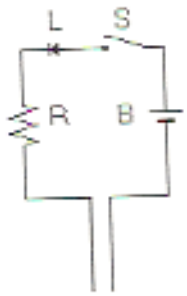
Procedure Summary

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.



Switch
LED
Battery
Resistor

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

Procedure

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 11th choice if it's in the lab and you are curious.*

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

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 <u>one at a time</u>, 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.

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. ☺

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CHEMISTRY 205

Experiment 5: SOLUBILITY *vs.* REACTIVITY

Laboratory Report

(5 Marks) **Observations and** (5 marks) **Conclusions:** **for the ten solutes you studied**

Observations:

- Appearance: colour, phase, descriptors (powder, crystalline, clear, cloudy, gas, *etc.*)
- Solubility: fully soluble, partially soluble, insoluble (suspension? settled solid?)
- Conductivity: high, low, no different than pure water (nonconductive)

Conclusions:

- Use your water solubility & conductivity observations to classify each substance by type: element *(E) vs* ionic *(I)*, polar covalent *(PC)* or nonpolar covalent *(NPC)* compound.
- For each solute:
  - Write a complete ionic equation to describe the dissolution of each solute in water.
  - If the substance did not dissolve or react, write: *reactants → NR* (*i.e.*, no reaction).
  - ***If a reaction occurred*** in acidic and/or basic solution, write a net ionic equation for it.
  - ***If a reaction occurred***, classify its type: precipitation *PPTN*, acid-base *A/B*, oxidation-reduction *O/R*.

Substance Nickel II chloride	**ID label** A
Observations	***Conclusions***
Substance appearance: green	Substance type
When added to water: light green	Dissolution equation
Conductivity: nothing	
When added to acid: Nothing (s)	Reaction equation & type
When added to base: ① become cloudy ② can't see through exc reaction	Reaction equation & type

07/11

Substance copper II nitrate **ID label** B

Observations	*Conclusions*
Substance appearance: blue	Substance type
When added to water: blue is h	Dissolution equation
Conductivity not	
When added to acid: Bubbles (s)	Reaction equation & type
When added to base: blue think (liquid)	Reaction equation & type

Substance sulphur **ID label** C

Observations	*Conclusions*
Substance appearance: yellow tabu	Substance type
When added to water: cloudy undissolve	Dissolution equation
Conductivity nothing	
When added to acid: soluble	Reaction equation & type
When added to base: Nothing	Reaction equation & type

Substance silica **ID label** D

Observations	*Conclusions*
Substance appearance: white tabu	Substance type
When added to water: chowdy undissolve	Dissolution equation
Conductivity nothing	
When added to acid: Soluble	Reaction equation & type
When added to base: Soluble	Reaction equation & type

Substance sodium carbonate	**ID label** E
Observations	***Conclusions***
Substance appearance: white like sugar	Substance type
When added to water: dissolve	Dissolution equation
Conductivity Nothing	
When added to acid: gas	Reaction equation & type
When added to base: Soluble	Reaction equation & type

Substance carbon powder	**ID label** F
Observations	***Conclusions***
Substance appearance: black	Substance type
When added to water: settle black	Dissolution equation
Conductivity low	
When added to acid: insoluble	Reaction equation & type
When added to base: stays the same	Reaction equation & type

Substance calcium sulphate	**ID label** G
Observations	***Conclusions***
Substance appearance: white powder	Substance type
When added to water: some of it dissolve	Dissolution equation
Conductivity high	
When added to acid: soluble	Reaction equation & type
When added to base: Turns cloudy	Reaction equation & type

Substance Cobalt chlorid **ID label** H

Observations	*Conclusions*
Substance appearance: pink	Substance type
When added to water: pinkish	Dissolution equation
Conductivity low	
When added to acid: hang blow change to fluch to pink	Reaction equation & type
When added to base: Nothing	Reaction equation & type

Substance Aluminum chloride **ID label** I

Observations	*Conclusions*
Substance appearance: white powder	Substance type
When added to water: fully dissolve	Dissolution equation
Conductivity high	
When added to acid: liquel turn to cloly undissolve	Reaction equation & type
When added to base: ~~solute~~ solid	Reaction equation & type

Substance zink oxid **ID label** J

Observations	*Conclusions*
Substance appearance: white	Substance type
When added to water: cludy undissive	Dissolution equation
Conductivity low	
When added to acid: sa prate the substance	Reaction equation & type
When added to base: Nothing	Reaction equation & type

Additional Questions

1. (2.5 marks) Think about the 'solubility rules' for ionic compounds you learned in class. Did your experiments agree with these rules? Explain.

2. (2.5 marks) Which types of solutes tended to be insoluble in pure water but soluble in acid or in base? Can you explain this?

CHEMISTRY 205

Experiment 6A: ANALYSIS OF AN UNKNOWN OXIDE OF COPPER

Introduction

Copper has two common oxidation (or valence) states: copper (I) or *cuprous*, Cu^{1} or Cu(I), and copper(II) or *cupric*, Cu^{2+} or Cu(II). Because of this, it is possible to obtain two oxides of copper: cuprous oxide, Cu_2O, and cupric oxide, CuO. In this experiment, you will be given one of these two oxides, and your task is to determine which one you have.

The method you will use in this experiment falls into the general class of techniques called *gravimetric analysis*. Gravimetric analysis is a method of analysis that uses the mass of product formed during a reaction as a means to investigate the nature of one of the reactants. There are typically two contexts in which gravimetric analysis is used: either (1) you know the identity of the compound but not how much of it is present in your sample, *or* (2) you know how much of it you have, but not exactly what it is! In the second case, if you happen to know that there are only a couple of potential identities for your compound, then you can likely find a simple reaction or two that would occur for both possibilities. If so, by analyzing the quantities of products and thinking about the stoichiometry of the reactions, you can determine which reactant you originally had.

In this experiment, the analytical method is based on precipitating metallic copper from a known mass of an unidentified oxide of copper. The equations for the reactions involved are shown below in the Procedure Summary. By isolating and weighing the precipitated Cu(s), the amount of oxygen that was combined with the copper in the original compound can be determined by difference. Using this information, the elemental composition of the oxide can be calculated, which will ultimately establish the identity of the oxide.

Procedure Summary

In this experiment, the unknown copper oxide (which is insoluble in water) will first be dissolved in hydrochloric acid to give a solution of copper chloride. Then the solution will be treated with metallic aluminum; this will result in the formation of aluminum chloride and the precipitation of copper metal. The two possible reaction sequences are:

$$Cu_2O(s) + 2HCl(aq) \rightarrow 2CuCl(aq) + H_2O(l)$$
$$3CuCl(aq) + Al(s) \rightarrow AlCl_3(aq) + 3Cu(s)$$

or

$$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$$
$$3CuCl_2(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

The reactions with aluminum are examples of oxidation - reduction (redox) reactions, and involve the following half reactions:

$$Al \rightarrow Al^{3+} + 3e^-$$

and

$$\textit{either} \quad Cu^+ + e^- \rightarrow Cu$$

$$\textit{or} \quad Cu^{2+} + 2e^- \rightarrow Cu$$

Both copper(I) and copper(II) have a stronger tendency to gain electrons than aluminium(III); in other words, aluminum ion is more electropositive than copper ions (which are more electronegative). For this reason, electrons will be passed from the aluminum metal to the copper ion according to one of the complete reactions shown above.

Prelaboratory Assignment

Read the procedure and answer the prelaboratory 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

- 100 mL beaker
- medicine dropper
- glass rod
- hot plate at low heat

Reagents and Materials

- copper(I) or copper(II) oxide
- 6M hydrochloric acid [Caution – corrosive]
- distilled water
- about 0.1g aluminium metal (strip)
- acetone

Procedure

NOTE: You will be working with small quantities of quite concentrated hydrochloric acid. Be careful not to get any on your skin, eyes or clothing. Work in a fume hood to avoid inhaling hydrogen chloride gas, the white fumes that diffuse out of concentrated hydrochloric acid solutions. Wipe up small spills immediately - consult your demonstrator in the case of a serious spill.

1. Weigh a clean dry 100 mL beaker (to the nearest milligram). In the beaker, weigh out between 0.15 and 0.20 g of the copper oxide that you are given. Report the weights.
2. Using a medicine dropper, carefully add about 2.5 mL of 6M hydrochloric acid. Swirl the contents of the beaker to make sure the copper oxide is thoroughly wetted by the acid. The oxide will gradually dissolve over a few minutes. Swirl the mixture occasionally and warm it at *low* heat on a hot plate.
3. When the copper oxide has completely dissolved, dilute the solution with about 2.5 mL of distilled water. Then add the aluminum. Report what happens, and specifically, how you determine that the reaction is complete.
4. When all the copper has been precipitated, fish out the remaining aluminum with a glass rod (avoid removing any of the copper with it) and place it in the waste container provided. Carefully decant off as much of the liquid in the beaker as you can with a pipette *without removing any of the copper*. (The acidic aluminum chloride solution can be flushed down the sink.) Add about 5 ml of deionized or distilled water and decant again to wash the last traces of aluminum chloride solution off the copper; do this twice.
5. Wash the wet copper twice with about 2 mL of acetone, again by decantation. This will dissolve the water in the acetone, and leave the copper wet with the more volatile acetone (boiling point = 56 °C). The remaining acetone is evaporated by heating the beaker *gently* on the hot plate.
6. Cool the beaker to room temperature and weigh it to determine the mass of copper obtained.

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CHEMISTRY 205

Experiment 6A: ANALYSIS OF AN UNKNOWN OXIDE OF COPPER

Prelaboratory Questions

1. (1.5 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.)*
2. (1.5 marks) Calculate the percentage compositions of the two possible oxides of copper.

3. (1 mark) Look up and report the colours of the two oxides of copper (include the names of the references you use to answer #3 and #4).

4. (1 mark) Look up and report the colours of the $Cu^{+}(aq)$ and $Cu^{2+}(aq)$ ions.

CHEMISTRY 205

Experiment 6B: GRAVIMETRIC DETERMINATION OF CARBONATE OR BICARBONATE

Introduction

The premise of this experiment is the same as Experiment 6A, in which you used gravimetric analysis to identify an unknown oxide of copper. In this experiment, you will identify an unknown as either sodium carbonate ("soda") or sodium hydrogen carbonate (sodium bicarbonate or "baking soda"). The method will involve the reaction of the unknown with excess hydrochloric acid, which will occur according to one of the two following reactions:

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

$$NaHCO_3(aq) + HCl(aq) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Both reactions produce the same products, but not in the same proportions. Thus, gravimetric analysis of the solid product, NaCl, can be used to determine the identity of the unknown.

Procedure Summary

A known mass of the unknown compound will be reacted with excess hydrochloric acid. The products of the reaction are carbon dioxide, sodium chloride and hydrogen chloride; these compounds and the solvent in which they will be dissolved (water) are all volatile species, except for sodium chloride. Thus, even though all the products are soluble in water, heating the solution containing them will drive off everything except the sodium chloride. After evaporating the product mixture to dryness, the sodium chloride residue will be weighed. Comparison of the mass of the starting material with the mass of sodium chloride obtained will allow for the identification of the unknown based on the reaction's observed stoichiometry.

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

- 100 mL beaker
- hot plate (on high)
- medicine dropper

Reagents and Materials

- unknown (anhydrous sodium carbonate *or* sodium hydrogen carbonate)
- 6M hydrochloric acid

Procedure

You will be working with small quantities of quite concentrated hydrochloric acid. Be careful not to get any on your skin, eyes or clothing. Avoid inhaling the vapour - work in a fume-hood. Wipe up small spills immediately - consult your demonstrator in the case of a serious spill.

1. Weigh the beaker to the nearest milligram and report its mass. Place about 0.2 g of the unknown substance in the beaker. Then weigh the beaker again. Report the mass.
2. Add two drops of the 6M hydrochloric acid to the beaker. Note what happens. Continue adding hydrochloric acid dropwise until there is no more reaction, and then add a couple more drops to be absolutely sure that an excess of acid has been added.
3. Heat the beaker on a hot plate so that the solution boils; evaporate the mixture to dryness. Be careful near the end so that the contents of the beaker do not splatter excessively (to prevent loss of product). Continue heating the beaker to ensure it and its contents are truly dry. (Any water/hydrochloric acid that condenses near the top of the beaker can be blotted up with a piece of paper towel - wear gloves! Take care not to remove any sodium chloride solution; *i.e.*, only wipe off condensate, not splattered solution.)
4. Allow the beaker to cool to room temperature and then weigh it again to determine the mass of the sodium chloride residue. The sodium chloride can be then redissolved in water and flushed down the sink.

Name: _____ Section: _____ Date: _____

CHEMISTRY 205

Experiment 6B: GRAVIMETRIC DETERMINATION OF CARBONATE OR BICARBONATE

Prelaboratory Questions

1. (1 mark) 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). *(Attach a separate page.)*
2. (1 mark) What are the boiling points of your reaction products and the solvent they will be dissolved in? [Consult a reference book like the CRC Handbook of Physics & Chemistry.] Explain why this information is critical to the design of this experiment.

3. (1 mark) "Concentrated" hydrochloric acid is made by bubbling hydrogen chloride, HCl(g), through water until the solution is saturated with the gas. All other concentrations of hydrochloric acid are ultimately made by diluting this solution, which contains 38% by weight of hydrogen chloride and has a density of 1.200 g mL^{-1}. Calculate the molarity of "conc. HCl". Show your work.

4. (1 mark) How many millilitres of 6M hydrochloric acid are required to react exactly with (a) 0.2 g of sodium carbonate, and (b) 0.2 g of sodium hydrogen carbonate? Show your calculations, including the relevant balanced chemical equations.

5. (1 mark) What is "baking soda" used for in baking? Explain how it works, and include the relevant chemical equations.