Crescent School
SCH3U name_____

Acids and Bases: Titration #1
Determination of [NaOH]
by Microtitration with HCl of Known Concentration

The purpose of this experiment is to determine the concentration of an NaOH solution by exactly neutralizing a given volume of HCl(aq) of a known concentration with NaOH(aq).

We need to add the stoichiometric amount of NaOH(aq) to the HCl(aq)—no more, no less.

This procedure is called a **titration**, and will be done on a microscale.

The balanced chemical equation for the neutralization of NaOH(aq) with HCl(aq) is

HCl(aq) + NaOH(aq) → _____

What colour are the reactants and products in this reaction? _____

What potential problem does this present? _____

How can we circumvent this problem? _____

Phenolphthalein (weird spelling) is an acid-base indicator.

Phenolphthalein is *clear* in *acidic* solution; bright *pink* in *basic* solution.

What colour do you expect phenolphthalein to be at the exact point of neutralization—the *equivalence* (or stoichiometric) point of this reaction? _____

We will add a small amount of phenolphthalein to our reacting system to indicate when the acid has been *exactly* neutralized.

What colour change do you think is easier to detect: (circle correct answer)
clear to pale pink – or – pale pink to clear?

Given your answer, to what solution—NaOH or HCl—will you add the phenolphthalein?
_____

Equipment for the microtitration:

- rxn vessel: bottom of clear plastic 2-L pop bottle (contains 5 "wells")
- acid and base dispensers: 2 plastic Beral pipets (unit of volume = _____)
- stirrer: a plastic fork with all but one tine removed
- watch glass or petrie dish
- phenolphthalein solution
- HCl(aq) of known concentration (today, [HCl] = _____ mol/L)
- NaOH(aq) of unknown concentration
- tissue to clean the fork between trials

Sample Calculation:

Do we need to know the volume (mL) of 1 drop? _____

Consider the following:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$$

0.10 mol/L	? mol/L
10 d (d = drop)	19 d

What did we assume about the vol. of a drop of HCl compared to a drop of NaOH? ____

How could we carry out a microtitration using different-sized droppers for the acid and base? _____

Do the microtitration. Use 15 d of HCl in each well. Carry out 5 microtitrations; average the number of drops of NaOH. Do 1 set of <u>calculations below</u> to find [NaOH]. Don't forget to write the balanced chemical equation**. [NaOH] = _____ mol/L**

Trial	# drops HCl	# drops NaOH
1		
2		
3		
4		
5		

—fin—

Crescent School
SCH3U name: _____

Acids and Bases: Titration #2
Determination of the Concentration of Acetic acid, CH_3COOH, in (mol/L)
and % (m/v), in Commercial Vinegar

White vinegar is claimed to be 5.0% (m/v). In this experiment you will titrate an acetic acid solution (vinegar) against a NaOH solution of known concentration to a phenolphthalein endpoint. Report your results as [acetic acid], in mol/L and % (m/v) to verify this claim.

Balanced equation for the reaction:

____ $CH_3COOH(aq)$ + ____ $NaOH(aq)$ → ____ _____ + ____ _____

[NaOH] = _____ mol/L

Tabulate Data Here:

Trial #	Start volume NaOH (mL)	Final volume NaOH (mL)	Net volume NaOH (mL)	colour/description of end point
1				
2				
3				
4				
5				
6				

Average volume of NaOH solution used: _____ mL (exclude volumes marked with *)

. . . over

Lab Report maximum length: **1 page**, word-processed. Include:

LAB REPORTS LONGER THAN 1 PAGE, 12 POINT FONT, WILL NOT BE MARKED.

Descriptive Title

Abstract — Three sentences at most: Purpose; *general* method; results. The brief introduction on the other side of this sheet may be helpful. {C, 3}

Data Table — Do not include start/ending volume—only net volumes and average volume of NaOH. Make a proper table, please. {C, 1}

Calculations — Include a balanced chemical equation; calculate [CH_3COOH] in both mol/L and in % (m/v).

Post-Lab Questions

1. How did the value for the percent acetic acid in vinegar that you determined compare with that claimed by the manufacturer? Calculate the % error; comment on how you could improve your result. {C, 2}

2. While carrying out a titration, explain why is it permissible to use a wash bottle to rinse down the inside of the erlenmeyer flask, or to "shoot" a hanging drop of NaOH off the buret tip. {C, 1}

3. Explain why you must rinse the buret with the NaOH solution to be used in the experiment before filling it with the same NaOH solution. {C, 1}

4. a) Why do we titrate into an erlenmeyer flask instead of into a beaker? {C, 1}

   b) Why is it okay to have some distilled water in the erlenmeyer flask before you pipet the acid solution into the flask? {C, 1}

—fin—

Crescent School
SCH3U

Titration #3: Determination of the Molar Mass of an Unknown Solid Acid

In this experiment you will determine the molar mass of an unknown solid acid by titration with NaOH of known concentration to a phenolphthalein end point.

Acid given: HA or H_2A or H_3A (circle one)

Balanced chemical reaction of acid with NaOH:

_____

[NaOH] = _____ mol/L

ALL OF THE POWDERED ACID MUST BE TRANSFERRED TO THE FLASK; DISSOLVE THE ACID IN WATER BEFORE YOU BEGIN TITRATING.

Trial #	mass of unknown acid (g)	initial Buret Vol (mL)	final Buret Vol (mL)	Net Volume NaOH (mL)	colour/description of end point	Calculated molar mass of unknown acid (g/mol)
1						
2						
3						
4						

Average the resultant molar mass values as appropriate—see quality of end point.

Lab Report—as with previous lab: 1 page maximum; word processed
Descriptive Title, Abstract, Data Table, Sample calculation including balanced chemical equation, Conclusion

Post-Lab Questions

1. If HA = 204 g/mol; H_2A = 126 g/mol; H_3A = 192 g/mol, determine the % error in the molar mass that you determined. {PS, 2}
% error = absolute value of [(your result – accepted value)/accepted value] * 100%

2. How—specifically—would the calculated molar mass of the unknown acid change if all of the solid acid was not transferred to the erlenmeyer flask used in the titration? {C, 2}

—fin—

Crescent School
SCH3U name_____

Stoichiometry: Limiting and Excess Reagent, or "Who's the Boss?"

You will carry out the following chemical reaction with the indicated quantities.

$2Al(s) + 3CuCl_2(aq) \rightarrow 3Cu(s) + 2AlCl_3(aq)$ (equation 1)
0.25g 0.51g

NOTE:

- Al is in the form of aluminium foil
- $CuCl_{2(aq)}$ is light blue in colour
- $Cu_{(s)}$ is deposited as a spongy brown precipitate
- $AlCl_{3(aq)}$ is a clear, colourless, solution

Pre-lab Prediction

What do you expect to observe after the reaction is finished? What colour will the solution be? Will there be any solid present? What colour will it be? Feel free to carry out any calculations.

Safety Notes

- Do not hold onto the beaker during the reaction. It can get hot.
- Wear safety glasses.

Procedure

To start the reaction, add some distilled water to the beaker containing the $CuCl_2$ and the Al foil, as directed by the teacher. Don't add too much water. You may stir the solution with a glass rod.

Disposal

Dispose of chemicals in the designated waste containers only. Do not pour anything down the drain.

Post-lab Questions

To be answered immediately after the reaction is carried out.

1. State your observations by answering the following questions:
   a) What colour was the solution of $CuCl_2$ *before* the Al foil was added?

   b) What colour was the solution **(not** the solid in the solution) after the reaction was complete? That is, did all of the $CuCl_2$ react?

   c) Based on the colour, what is the solid formed in the solution?

   d) Did all of the Al foil react? How do you know?

2. a) Based on your observations and the answers to post-lab question #1 above, which reactant was initially present in excess amount, the copper (II) chloride or the aluminium?

   b) How does this observation compare with your **Pre-Lab Prediction**? If possible, propose an explanation. Carry out calculations to support this. If you're stuck, proceed to question #3.

3. Reconsider the balanced chemical equation for the reaction performed, along with the quantities of each reactant used.

$2Al_{(s)} + 3CuCl_{2(aq)} \rightarrow 3Cu_{(s)} + 2AlCl_{3(aq)}$
0.25g 0.51g

a) Convert the number of grams of Al into moles. Do the same for $CuCl_2$.

b) From the balanced equation (see above), note that the ratio of aluminium to copper (II) chloride is

2 mol Al : 3 mol $CuCl_2$

From your calculations in part (a) above, state the *actual* mol : mol ratio of Al : Cu used in this experiment. Fill in the blanks below.

_____ mol Al : _____ mol $CuCl_2$

Rewrite this ratio in the form of

_____ mol Al : 3.00 mol $CuCl_2$

c) Based on the above:
   i. which reactant is present in excess?

   ii. which reactant is limiting?

d) Is this consistent with your empirical observations? Explain briefly.

e) Is this consistent with your **pre-lab prediction**? Explain briefly.

Further Questions:

4. Based on your answers in question #3 above, calculate:
   a) The number of moles of Cu metal formed in the reaction;

   b) The mass of Cu metal formed in the reaction;

   c) The number of moles and mass of unreacted Al remaining at the end of the reaction.

5. What mass of $CuCl_2 \cdot 2H_2O$ is required to furnish 0.51 g of *anhydrous* $CuCl_2$? (Hint: Look up hydrated ionic compounds.)

—*fin*—

Crescent School
SCH3U name:_____

Determination of the Thickness of Aluminium Foil

Purpose: The purpose of this experiment is to determine the thickness of a sheet of Aluminium foil in an experiment of your own design. The reference materials and equipment/chemicals in our laboratory are available.

Pre-lab Questions:

1. Is this experiment qualitative or quantitative? Explain. (C)

2. Can an ordinary ruler alone be used to measure the thickness of Al foil? Explain. (C)

3. What is required to complete the experiment? (I) Think in terms of:
   - equipment
   - chemicals/supplies
   - reference materials.

4. In point form, state the procedure. You may be able to come up with more than one method. (I)

Crescent School
SCH3U name: _____

Experiment:
Determination of the Thickness of Aluminium Foil

Post-lab Questions:

1. Using your empirical data, determine the thickness, in millimetres, of a sheet of Al foil. Provide a neat, well organized solution below. No rough work on this sheet. (I)

2. Comment on your result. Is it reasonable? If you used more than one method, in which one do have more faith? (C)

—fin—

Crescent School
SCH3U name_____

Calculations with Measured Quantities

The mass of several plums is as follows:

58.3 g
75 g
53.5 g
55.1 g
65.24 g

Calculate the average mass of a plum.

—fin—

Crescent School
SCH3U name: _____

Observation of a Burning Candle

Introduction

You will observe a burning candle to study some of physical and chemical processes associated with combustion.

Pre-lab Questions (Answer in the space provided, before you begin the experiment. You may consult your textbook or another resource as required.)

1. Aside from wax (fuel) state two requirements for combustion of candle wax. {K/U, 2}

_____ _____

2. What are the units for the following? For example, the *rate* of walking = metres per second or m/s or $m \cdot s^{-1}$

a) eating crackers _____

b) a snail moving _____

c) What is common to all units of rate? _____

Materials

large candle, 3”x5” index card or piece of cardboard, 250 mL beaker, wire mesh, matches wooden splint, access to electronic balance, bunsen burner, sparker, tongs

Safety

Wear safety glasses; blazer off; tuck your tie into your shirt as demonstrated. This is standard procedure for all labs.

Procedure

Part 1 Teacher Demonstrations

a) The catalytic combustor consists of a 2.5 cm diameter glass tube fitted with a ceramic honeycomb. Observe the catalytic combustor in use. Record your observations below:

b) Observe as the teacher heats solid candle wax in a beaker.

What happens to the wax as it is heated? _____

What happens when a lighted splint is held above the melted, hot wax? _____

You will make use of this observation later on in this lab.

c) What do you observe when a large beaker is inverted over a burning candle? _____

What do you observe when a lighted splint is inserted into this inverted beaker?

_____

From these observations, write a balanced chemical equation for the *complete combustion* of wax, a hydrocarbon, whose average formula is represented by $C_{25}H_{52}$.

_____

Now write a balanced chemical equation, in terms of n, for the *complete combustion of a hydrocarbon* with formula of C_nH_{2n+2}.

1 $C_nH_{2n+2}(g)$ + _____ $O_2(g)$ → _____

Part 2. The burning candle

a) Light the larger candle. Drip some wax onto a piece of cardboard and stand the candle in the melted wax. Hold the candle upright until it can stand alone.

b) Place the candle and stand on the electronic balance and record its initial mass.

c) Light the candle; let it burn for 1 minute or so. Then record the mass of the candle and stand every 15 seconds for ca 3 minutes. Record your data Table 1.

Table 1. Mass data for burning of a candle (after 1 minute "warm-up")

time (seconds)	mass (g)
0	
30	
60	
90	

d) Using EXCEL, plot the mass of the candle (y-axis) vs. time (x-axis). The teacher will demonstrate. Provide a descriptive title for the graph, label axes (with units). Draw a line of best fit on the graph; include equation of the trend line and the R^2 value. Attach the graph.

What does the shape of the line tell you about the rate of combustion over time?

_____

How does the mass of uncombusted wax affect the rate of combustion?

_____

Part 3. The 250-mL beaker

a) Lower a clean, dry 250-mL beaker, right side up, into the candle flame, and hold it there for a moment. The candle flame should just touch the bottom of the beaker.

What is the substance that forms on the bottom of the beaker? _____

b) Is the combustion that occurs in a candle complete or incomplete (circle one)?

Part 4. Re-light the candle without touching the wick

a) Light the candle and allow it to burn for 1 minute.

b) Hold a wooden splint in the candle until it catches fire.

c) Blow out the candle with a short puff of breath. Immediately hold the burning splint about 1 cm from the wick in the smoke rising from the candle without touching the candle. Observe and record what happens. _____
_____

d) Based on your observation, what is combusting when a candle is burning (circle one):

solid wax liquid wax gaseous wax the wick

e) Is it possible to light the bottom of a candle? Explain. _____
_____

Part 5. The Wire Mesh

a) Light the candle. Holding the (room temperature) wire mesh horizontally using tongs, slowly lower the mesh over top of the flame of your candle.

b) Record what happens. _____

c) Why do you suspect the mesh stopped the flame from burning? _____

_____

d) Light the **Bunsen burner**. Holding the wire mesh horizontally using tongs, slowly lower the mesh over top of the flame of the Bunsen burner, until the mesh is directly over the inner blue cone of flame of the Bunsen burner.

e) Hold still and observe any patterns that form on the mesh. What shape do you see in the mesh?

_____

f) What does this tell you about the horizontal cross-sectional shape of the Bunsen burner flame?

_____

g) Place the **hot** wire mesh back over the burning candle. Observe what happens.
Based on this, what effect does the temperature of the mesh have on the candle flame?

_____
_____

—fin—

Crescent School
SCH3U name: _____

Data Analysis: Measurement, Significant Figures, Precision, Accuracy

[ref: Chem 13 News, Feb 2008, p 15]

Accuracy represents how close a measurement is to the accepted value.

Percentage error is defined as

$$\% \text{ error} = \frac{|\text{measured value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

Precision has two meanings:

1. How close a group of measurements of the same thing are to each other. For example, if you throw five darts at a dartboard and they all land close to each other, the throws are precise. Note: high precision doesn't necessarily mean high accuracy; your closely-grouped darts may be well away from the bull's eye.

2. The number of significant digits, or decimal places in a measurement. For example, a ruler graduated in mm is more precise than a ruler graduated in cm.

Below are some student-determined data for the density of ethanol, CH_3CH_2OH, aka beverage alcohol.

Descriptive Title of Table: _____

Measurements (@ 20°C) ↓	**Trial**				
	1	2	3	4	5
mass of graduated cylinder + ethanol (g)	47.78	55.22	65.96	73.27	124.29
mass of empty graduated cylinder (g)	43.43	43.43	43.43	43.43	43.43
mass of ethanol (g)	4.35	11.79	22.53	29.84	80.86
volume of ethanol (mL)	5.0	15.0	27.9	37.3	100.0
density of ethanol ($g \cdot mL^{-1}$)					

Analysis Questions

1. a) Write a suitable title for the data table in the space provided above the table. {C, 1}

   b) What was the smallest scale division (graduation) on the graduated cylinder used to obtain these data? {K/U, 1}

   _____

2. Write the calculated values for density of ethanol in the final row of the table. Pay close attention to significant figures. {PS, 3}

3. In terms of definition #2 of precision, circle the most precise value for the density of ethanol. {K/U, 1}

4. Using spreadsheet software, plot the following graphs on the **same set of axes**:
a) mass of (graduated cylinder + ethanol) (y-axis) versus volume ethanol (x-axis);
b) mass of ethanol only (y-axis) versus volume ethanol (x-axis).

Plot an "X-Y (Scatter)"—as it is called in EXCEL—not a line graph for each. Label each axis appropriately; include proper units. Provide a *descriptive* title to the graph—NOT simply *y* versus *x*.
For each graph, add a *trendline*. Include on the graph the equation of each line in the form of y = mx + b; include the "R^2" value associated with each line.

The graph should fit on one page. Staple it to this assignment.{C, 10}

5. a) How do the slopes of each line compare? _____ {K/U, 1}
b) What property of ethanol is represented by the slope of each line? {K/U, 1}

_____

c) With reference to the graph of mass of (graduated cylinder + ethanol) versus volume ethanol, what is the significance of the y-intercept?
_____ {K/U, 1}

d) In a perfect world, what should the y-intercept value be for the graph of mass of ethanol versus volume ethanol? _____ {K/U, 1}

6. a) Using all five trials, determine the average value for the density of ethanol; pay close attention to sig figs. Answer below. {PS, 3}

b) How does the value determined above compare to the slope of the lines?

_____ {K/U, 1}

7. The accepted value of the density of ethanol at 20^{o}C is 0.789 g/mL. Calculate the % error of the average density of ethanol. {PS, 2}

—fin—

Crescent School
SCH3U

Determination of the Molar Mass of Propane by Downward Displacement of Water

NB. DO NOT TURN THIS PAGE OVER UNTIL YOU ARE ASKED TO DO SO.

The purpose of this activity is to determine the molar mass of propane, _____ .
After the experiment, sketch the apparatus below, to the right of Hank Hill:

"Molar mass is something we talk about
a lot over at Strickland Propane, I tell you what."

We will collect propane from a compressed gas cylinder by downward displacement of water.

How can we use the ideal gas law, **PV = nRT** to determine the molar mass of propane?

Aside from the downward displacement apparatus and the gas cylinder, what equipment will be required? What data must we record?

Let's do the collection . . .

Explain why the propane that we have collected is contaminated with water vapour. Think about the method that we used.

How can we know the total pressure of the gas(es) inside the collection vessel?

The vapour pressure of water (at the temperature of the experiment) must be subtracted from the total pressure of the gases.

Temperature (^{o}C)	Pressure (kPa)
19	2.20
20	2.33
21	2.49
22	2.81
23	2.99
24	3.17
25	3.36

This table is also on page 558, Table 12.4 and is on our data sheet.
Let's do the calculations . . .

Determine the percentage error of our empirically determined value for the molar mass of propane, C_3H_8.

Questions

1. a) What two criteria must be met such that a gas can be collected by d.d. of water?

_____

_____

b) Why must the water level in the gas collection vessel be the same as in the surrounding beaker when recording the volume of gas collected?

c) When a gas is collected by downward displacement of water, what adjustments to any calculations must we make?

2. Calculate the mass of dry hydrogen in 750 mL of moist hydrogen collected over water at 25.0^{o}C and 99.5 kPa. The water level inside and outside the collection vessel are equal. [Answer: 0.0585 g]

3. A 296 mL sample of oxygen is collected over water at 23^{o}C on a day when the barometric pressure is 750.3 mmHg. The water level inside and outside the collection vessel are equal. What volume would the dry oxygen occupy at 48^{o}C under a pressure of 101.3 kPa? [Answer: 307 mL]

4. A 5.42 L sample of a gas was collected over water on a day when the temperature was 24.0^{o}C and the barometric pressure was 706 mm Hg. The water level inside and outside the collection vessel were equal. The dry sample of gas had a mass of 5.60 g. What is the molar mass of the dry gas? [Answer: 28.0 g/mol]

—fin—

Crescent School
SCH3U name_____

Demonstration: The Combustion of Propane
or

_____

In this demonstration, we'll look at the complete combustion of propane, C_3H_8. Begin by balancing the chemical equation that represents this reaction:

$$1\ C_3H_8(g) + \_\_\_\_\_\ O_2(g) \rightarrow \_\_\_\_\_\ CO_2(g) + \_\_\_\_\_\ H_2O(g)$$

propane

NOTE:

- source of O_2 is air, which is 21% O_2
- combustion will occur in a clear 2-L pop bottle
- propane gas will be collected by downward displacement of water, as illustrated by the teacher
- watch the reactions; complete the table below

trial	**volume C_3H_8 reacted, mL (assume SATP)**	**observations**	**explanation (think in terms of the relative amounts of propane and air)**
1	40		
2	80		
3	160		

Re-write the balanced chemical equation for the combustion of propane:

$1\ C_3H_8(g) + \_\_\_\_\_\ O_2(g) \rightarrow \_\_\_\_\_\ CO_2(g) + \_\_\_\_\_ H_2O(g)$

propane

ie 1 mol C_3H_8 : _____ mol O_2

or 1 volume C_3H_8 : _____ volume O_2

(air is 21% O_2)

or 1 volume C_3H_8 : _____ volume air

Based on this ratio, calculate the optimal volume (mL) for propane and for air, assuming a *total* volume of 2-L. Do preliminary calculations on scrap paper. Your neat, annotated answer goes below.

Go back to the beginning of this handout; write another descriptive title for this activity.
[reference: Zhilin, D.M.; *J Chem Educ*, 2012, **89**, 649-651.]

—fin—

Crescent School
SCH3U name _____

Determination of Molar Mass of Acetone (C_3H_6O) on a Jumbo-Scale

Mr J. carried out following experiment in our lab to determine the molar mass of acetone. You will be doing a smaller scale version of this experiment in an upcoming lab—and on the laboratory exam. A demonstration follows . . .

Equipment/supplies

- 2-L Erlenmeyer flask with Al foil cap (pin-hole in cap)
- 2 plastic wastepaper/recycling baskets—big enough to hold a 2-L flask
- large capacity balance
- some acetone (a liquid at room temperature)
- thermometer
- barometer
- kettle

Procedure

- Weigh clean, dry 2-L flask with Al cap.
- Add few mL acetone, replace Al cap, swirl to "coat" insides of flask with acetone. There should be no more than 1 mL of liquid acetone at the bottom of the flask after swirling.
- Submerge as far as possible, without wetting the Al cap, in hot water (>60^{o}C) til all of the acetone vapourizes.
- Record the temperature of the hot water bath immediately after you remove the flask. This is the temperature of the acetone vapour in the flask. (The vapour is in *thermal equilibrium* with the hot water.)
- Condense the acetone vapour by submerging flask in cool water. Again, do not wet the Al cap.
- Dry the flask.
- Re-weigh the flask.

Data

mass of condensed acetone	4.44 g
temp of hot water bath	62 ^{o}C
P_{atm} (classroom barometer)	757 mmHg
Volume of flask	2.12 L

Use these data, and whatever else you require, to calculate the molar mass of acetone with % error. You can probably fit your calculations somewhere on this page.

—fin—

Crescent School
SCH3U

Demonstration: Examination of the Reactivity Trends in Alkali Metals

We carried out the reaction of each of Li, Na, K with water in order to

- determine the products of the reaction using various chemical tests;
- examine the trends in reactivity for the alkali metals

The reaction of lithium with water is represented by the balanced chemical equation:

$$\mathbf{2\ Li(s)\ +\ 2\ HOH(l)\ \rightarrow\ 2\ LiOH(aq)\ +\ H_2(g)}$$

Similarly for Na and K.

<u>We used the following tests/observations to verify the products:</u>

- The aqueous solution of LiOH conducted electricity, indicating the presence of dissolved ions. As a control, we tested a solution of LiOH and obtained the same result. Similarly for NaOH, KOH.

- We collected the hydrogen gas by downward displacement of water and conducted a "pop" test. The H_2 was not collected when Na and K were reacted with water. Reactions were too vigorous; the hydrogen could have combined with oxygen in the air and exploded.

- A few drops of phenolphthalein (note the weird spelling) turned the solution pink, indicating the presence of hydroxide, OH^- ions and a basic solution. As a control, we tested a solution of LiOH and obtained the same result. Similarly for NaOH, KOH.

- A red-coloured flame test on the resultant solution indicated that Li^+ ions were present. (Na produced a yellow flame; potassium's flame was violet.) Controls were also carried out.

- The reaction of Li with water was pretty tame; Na reacted much faster; K reacted so vigorously that the hydrogen ignited, producing a violet-coloured flame.

—fin—

Crescent School
SCH3U name _____

Review Questions—Measurement and Significant Figures

1. Someone says to you: "There were about 50,000 people at the concert." What level of confidence do you have in this number? Rewrite this number in scientific notation to properly reflect your confidence in this value. {MC,1} _____

2. a} Why do counted quantities have an infinite number of significant figures associated with them? {C, 1}_____

   b} What is assumed about the right-most digit in any measured quantity? {K/U, 1}

   _____

   c} What rule must you follow when taking measurements with an analogue scale? {K/U,1}_____

   d} How do we handle uncertainty when using a measuring device with a digital readout? {K/U, 1}

   _____

3. How much water is in the graduated cylinder illustrated below? Read the bottom of the meniscus. {K/U,2}

4. Indicate the number of significant figures and underline the estimated digit in each of the following: {K/U, 1 mark each}
   a) The railing is 12.865 m long. _____
   b) The volume of the car's gasoline tank is 48.7 L _____
   c) The mass of cement is 55.73 kg. _____

5. a} Rewrite the measured quantity 0.00460 g in sci. notation. {C,1} _____
   b) Explain why the three left-most zeros "disappear" when this number is re-written in scientific notation. {C,1} _____

      _____

   c) Explain why the right-most zero does not disappear when this measured quantity is re-written in scientific notation. {C,1} _____

      _____

d) In your own words explain the term *accuracy*. Illustrate your answer with an example. {C,1}

_____

_____

6. a} In your own words explain the two meanings of the term *precision*. Illustrate each with an example. {C,2}

i. _____

ii. _____

b} A chemist carries out the same experiment five times to determine the volume of a sodium hydroxide solution required to neutralize an acid sample. She obtains the values 12.84 mL, 12.28 mL, 12.30 mL, 12.27 mL, 12.29 mL. What average volume of sodium hydroxide solution should be reported? Show calculations. {PS, 2}

7. A student wishes to determine the density of aluminium using several chunks of the metal. The student masses each piece to be 15.7g; 16.2g; 9.15g. He puts all of the pieces into a water displacement can, the displaced water occupying a volume of 14 mL.
What density will the student obtain? {PS, 3}

8. A 9.76 g sample of table sugar is placed in a 25.00 mL flask. The flask is then topped up to the 25.00 mL mark with benzene$_{(l)}$. The sugar and benzene have a total mass of 26.31 g. The sugar does not dissolve in the benzene. If the density of benzene is 0.879 g/mL, what is the density of sugar? {PS, 5}

—fin—

Crescent School
SCH3U

names: _____
_____

Thought Lab: Determination of the Molar Mass of Two Unknown Gases

In this activity you will use empirical data to determine the molar mass of two gases. Together with other information, you will be able to identify each gas.

The answers to #1 and #2 will help with the rest of this assignment. You must show your calculations, but no rough work on this paper.

1. Air is composed of 78% N_2; 21% O_2; 1% Ar (v/v). Determine the average molar mass of air. {PS, 3}

average molar mass of air = _____

2. Two identical flasks at the same T and P each contain a different gas. How do the number of moles of gas in each flask compare? {K/U, 1}

_____

Calculation of Molar Mass of an Unknown Gas:

3. Gas #1

→ NOT COLLECTED BY DOWNWARD DISPLACEMENT OF WATER

mass of "empty" flask + lid (flask contains air*)	450.64 g
mass of same flask + lid + unknown gas (no air)	452.05 g
volume of flask	2.28×10^3 mL
P_{atm}	750.1 mmHg
room temperature	24^oC

Use these data, determine the molar mass of Gas #1. Hints: Use your answers to Q 1 & 2 above. Your first calculation should be to find the number of moles of air in the flask. (Remember: 760 mmHg = 101.3 kPa; $R = 8.31\ L \cdot kPa \cdot mol^{-1} \cdot K^{-1}$; $0^oC = 273K$)

molar mass of Gas #1: _____

Additional Information on Gas #1

- density = 1.83 g/L @ SATP
- clear, colourless, odourless
- does not support combustion
- the solid compound sublimates at room temperature.

4. Identity of Gas #1 is _____

5. Gas #2

→ NOT COLLECTED BY DOWNWARD DISPLACEMENT OF WATER

mass of "empty" flask + lid (flask contains air)	505.90 g
mass of same flask + lid + unknown gas (no air)	504.83 g
volume of flask	2.27×10^3 mL
P_{atm}	750 mmHg
room temperature	23^oC

Use these data, determine the molar mass of Gas #1. Hints: Use your answers to Q 1 & 2 above. Your first calculation should be to find the number of moles of air in the flask. (Remember: 760 mmHg = 101.3 kPa; R = 8.31 $L \cdot kPa \cdot mol^{-1} \cdot K^{-1}$; 0^oC = 273K)

6. molar mass of Gas #2: _____

7. Explain how the following information supports the fact Gas #2 is natural gas (aka Bunsen burner gas) ? {C, 4}

- less dense than air
- clear and colourless
- flammable when mixed with O_2
- a small amount of sulfur-containing compound is added to the gas to give it a disagreeable odour; ie. the sample of gas analyzed was not 100% pure

8. a) If natural gas is (v/v);

CH_4	94.5 %
C_2H_6	3.0%
C_3H_8	0.2%
N_2	1.6%
CO_2	0.7 %

Find the average molar mass of natural gas. {PS, 3}

average molar mass of natural gas: _____

b) Calculate the percentage error in the molar mass of our natural gas, compared to the above. {PS, 2}

9. The carbon dioxide was obtained from dry ice ($CO_{2(s)}$); the methane from our gas taps. How did we get a large flask full of each gas, uncontaminated with air or water vapour? {C, 4}

Acknowledgement: The concept, data collection and analysis by Andy Arrowsmith, class of 2006. Andy supplied the dry ice; he carried out research in our Chemistry lab.

—fin—

Crescent School
SCH3U name: _____

Empirical Determination of the Molar Volume of $H_2(g)$ and of $O_2(g)$ @ ca. SATP

Review of the Mole Concept

The mol is a unit of _____ {K/U, 1}.

Mol Quantities {K/U, 4}
1.00 mol contains 6.02×10^{23} _____;

For an element such as He or K or Fe,

1.00 mol contains 6.02×10^{23} _____.

For a molecular compound such as H_2O,

1.00 mol H_2O contains 6.02×10^{23} _____ .

For an ionic compound such as NaCl,

1.00 mol NaCl contains 6.02×10^{23} _____.

The Avogadro constant, 6.02×10^{23} particles*mol^{-1}, is determined (circle all that apply):
a) empirically b) theoretically c) qualitatively d) quantitatively {K/U, 2}

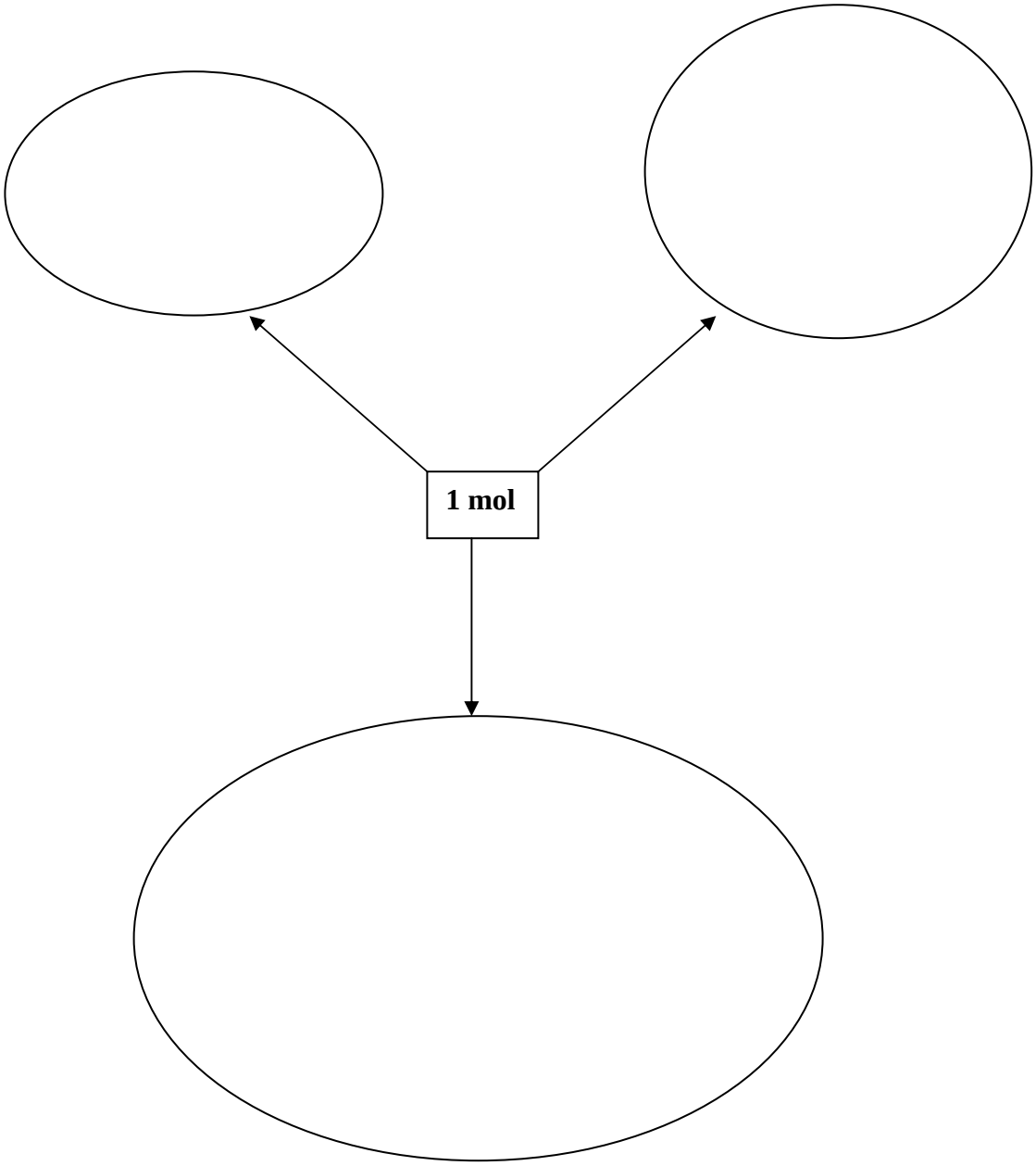
Molar Mass
The mass of 1 mol of an element or compound can be calculated using data from the

_____ . {K/U, 1}

Molar Volume of a Gas
We also know that the volume of 1 mol of any gas at SATP (25^oC, 100kPa) is 24.8 L.

Pre-lab Preparation

1. In the space below, construct a "map" of the mol concept, incorporating three different aspects of the mole. Your class notes may be helpful. {K/U, 3}



2. Consider a sample of 37.5 g of $CH_4(g)$ at SATP. Convert this amount to

a) moles (Show your calculations; pay attention to sig figs; include proper units.)
{PS, 3 x 2 = 6}

Ans: _____

b) number of molecules

Ans: _____

c) volume at SATP

Ans: _____

Purpose of the Experiment

In this experiment you will determine the molar volume of H_2 and O_2 under conditions that are very close to SATP. We will separately prepare H_2 and O_2 and collect each of these gases by <u>downward displacement of water</u>. This technique will be demonstrated by the teacher. The apparatus is illustrated below:

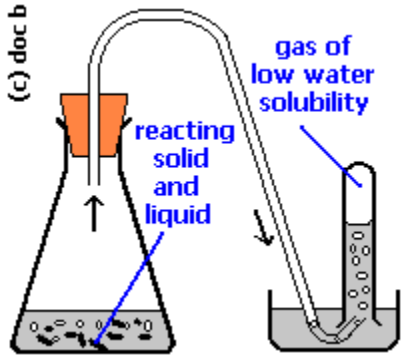


figure 1. Apparatus for collection of gas by downward displacement of water.

A. Preparation of hydrogen gas, $H_2(g)$

You will produce hydrogen gas according to the unbalanced chemical equation:

1 Mg(s) + ____HCl(aq) → ____$MgCl_2$(aq) + ____H_2(g) (eq'n 1)

A measured amount of magnesium will be reacted with excess hydrochloric acid, HCl(aq), and the hydrogen collected by downward displacement of water. We will measure and record the volume of H_2(g) collected.

3. Eq'n 1 above is an example of a _____ reaction. {K/U, 1}

4. Balance eq'n 1 by filling in the blanks above. {PS, 1}

5. Refer to the balanced chemical form of eq'n 1. If the coefficients in the balanced chemical reaction refer to moles, how many moles of H_2(g) will be produced for every mole of Mg that reacts? {PS, 2}

   1 mol Mg produces _____ mol H_2(g), or

   x mol Mg produces _____ mol H_2(g)

6. We use an <u>excess</u> of HCl(aq) in this experiment. Explain why this makes things easier. {C, 2}

7. When 2.47 g of Mg reacts completely with excess HCl(aq), 2.3 L of H_2 was collected by downward displacement of water. Use these data and your answers to the above questions to calculate the molar volume of H_2 in L/mol. Pay close attention to sig figs; include proper units throughout your calculation. {PS, 3}

molar volume of $H_2(g)$ = _____

8. Complete the left hand column of the data table for Part A below in preparation for the experiment.

B. TEACHER DEMONSTRATION: Preparation of oxygen gas, $O_2(g)$

You will react household bleach (5.0 % (m/v) sodium hypochlorite, NaOCl(aq)) with 10% (m/v) hydrogen peroxide, H_2O_2 according to the balanced chemical equation

$$H_2O_2(aq) + NaOCl(aq) \rightarrow O_2(g) + H_2O(l) + NaCl(aq) \qquad \text{(eq'n 2)}$$

The O_2 produced will be collected by downward displacement of water and its volume measured and recorded.

9. As the reaction proceeds, the mass of the reactant solution will decrease as O_2 is given off. If 1.3 g of O_2 is produced and is found to occupy 983 mL, calculate the molar volume of O_2 in L/mol. Pay close attention to sig figs; include proper units throughout your calculation. {PS, 3}

molar volume of $O_2(g)$ = _____

10. We will use 100 mL of bleach (an excess amount) in a 500 mL flask and several mL of 10% H_2O in a small vial. We will carefully lower the vial into the flask without mixing the solutions. Fill in the data table for Part B as the teacher completes the demonstration.

Materials

Each pair of students requires:

- access to an electronic balance
- 2 x 250 mL erlenmeyer flask with one-holed stopper and some glass tubing through the hole
- ca. 1m rubber hose attached to the glass tube above
- pneumatic trough
- retort stand and clamp
- 50 mL 1 mol/L HCl
- ca. 0.20 g Mg ribbon
- sand paper or emery paper to clean Mg ribbon
- 250 mL erlenmeyer flask in which to collect the $H_2(g)$ in part A
- 250 mL flask to collect the $O_2(g)$ in part B
- 100 mL fresh household bleach (in a 500 mL flask fitted with one-holed stopper)
- ca. 10 mL of 3% $H_2O_2(aq)$ in a test tube (or ca 3 mL 10% H_2O_2 diluted to ca. 10 mL)
- 2 small pieces paper towel

Safety Notes

- wear safety goggles and an apron throughout the experiment
- if any hydrochloric acid, bleach or hydrogen peroxide gets on your hands, wash with plenty of water

The teacher will demonstrate how to carry out Part A; Part B is a demonstration.
We will pool our data at the end of the experiment.

Record data for Part A here: $Mg + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Record data for Part B here: $H_2O_2(aq) + NaOCl(aq) \rightarrow O_2(g) + H_2O(l) + NaCl(aq)$

Data Analysis & Post-Lab Questions

11. a) Use your data from Part A, and what you learned from completing the pre-lab questions, to determine the molar volume of $H_2(g)$ in L/mol. {PS, 3}

molar volume of $H_2(g)$ _____

b) Use your data from Part B, and what you learned from completing the pre-lab questions, to determine the molar volume of $O_2(g)$ in L/mol. {PS, 3}

molar volume of $O_2(g)$ _____

12. a) How do the molar volumes of $H_2(g)$ and $O_2(g)$ that you determined compare to each other? {C, 1}

b) According to what you know about the mole concept as it relates to gases, should the molar volume of $H_2(g)$ and $O_2(g)$ be the same? Explain briefly. {C, 2}

c) Calculate the <u>average</u> molar volume of $H_2(g)$ and $O_2(g)$. {PS, 2}

d) The accepted molar volume of any gas is 24.8 L at SATP (25^oC, 100 kPa). Calculate the percentage error of your average molar volume of H_2 and O_2. {PS, 2}

e) Were the conditions under which we worked close to SATP? {C, 1}

13. This question relates to Part A of this experiment. If you react 0.20 g of Mg with 0.100 L of HCl(aq) whose concentration is 1.0 mol HCl per litre of solution (1.0 mol/L), prove that the HCl is present in an excess amount. Make sure to use the balanced chemical equation. {PS, 3}

—fin—

Crescent School
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Isolation of Acetone from Nail Polish Remover by Distillation; Microscale Determination of Boiling Point of Acetone; Determination of Molar Mass of Acetone by the Dumas Method

References
Szafran, Z; Pike, RM; Foster, JC; Microscale General Chemistry Laboratory with Selected Macroscale Experiments; J Wiley and Sons, New York, 1993, p 59 & p 155

Introduction
Nail polish remover typically contains acetone, CH_3COCH_3.

You will begin the experiment by distilling nail polish remover in order to collect pure acetone. This separation will leave any higher-boiling ingredients, such as the colouring agent behind.

You will determine the boiling point of the distillate—acetone—using a unique, and highly accurate microscale method.

Finally, you will determine the molar mass of acetone using the Dumas Method. This involves vapourizing excess acetone in a container of known volume, at a known temperature and pressure, followed by condensing the gaseous acetone in a cool water bath. The mass of the liquefied acetone corresponds to the mass of the acetone vapour previously in the flask. Using the Ideal Gas Law, PV = nRT, you can determine the number of moles of acetone vapour in the flask. Since you know the mass of vapour that was in the flask, and the corresponding number of moles, you can calculate the molar mass (g/mol) of acetone.

Purpose
1. To distill acetone-containing nail polish remover;
2. To measure the boiling point of acetone using a unique, microscale method;
3. To determine the molar mass of acetone.

Safety Precautions
Wear safety glasses and a lab coat/apron throughout the experiment. If anything gets on your skin, flush with plenty of cold water. Advise the teacher of any broken glass, spills, or mishaps.

Materials

For the class:
One or more distillation apparatuses, set up in the fumehood.
50 mL acetone-containing nail polish remover;
several centigram electronic balances;
several electric kettles.

Each group requires:
2 - 600 mL beakers;
micro test tube; ca 1 cm long glass capillary tube with one end sealed;
2 small elastics (orthodontic);
clean, dry, glass "Everfresh" juice bottle with pinhole in the metal screw-on cap.
100 mL graduated cylinder.

Disposal

Dispose of chemicals as directed by the teacher—nothing down the sink.

Prelab Questions:

1. a) bp of acetone (from Merck Index or Internet) = _____
   b) calculated molar mass of acetone, C_3H_6O, = _____

2. You must complete the handout entitled "Determination of Molar Mass of Acetone on a Jumbo Scale", on the next page.

Determination of Molar Mass of Acetone (C_3H_6O) on a Jumbo-Scale

Mr J carried out following experiment in our lab to determine the molar mass of acetone. You will be doing a smaller scale version of this experiment in an upcoming lab—and on the laboratory exam. A demonstration follows . . .

Equipment/supplies

- 2-L Erlenmeyer flask with Al foil cap (pin-hole in cap)
- 2 plastic wastepaper/recycling baskets—big enough to hold a 2-L flask
- large capacity balance
- some acetone (a liquid at room temperature)
- thermometer
- barometer
- kettle

Procedure

- Weigh clean, dry 2-L flask with Al cap.
- Add few mL acetone, replace Al cap, swirl to "coat" insides of flask with acetone. There should be no more than 1 mL of liquid acetone at the bottom of the flask after swirling.
- Submerge as far as possible, without wetting the Al cap, in hot water (>60^{o}C) til all of the acetone vapourizes.
- Record the temperature of the hot water bath immediately after you remove the flask. This is the temperature of the acetone vapour in the flask. (The vapour is in *thermal equilibrium* with the hot water.)
- Condense the acetone vapour by submerging flask in cool water. Again, do not wet the Al cap.
- Dry the flask.
- Re-weigh the flask.

Data

mass of condensed acetone	4.44 g
temp of hot water bath	62.0 ^{o}C
P_{atm} (classroom barometer)	757.0 mmHg
Volume of flask	2.12 L

Use these data, and whatever else you require, to calculate the molar mass of acetone with % error. You can probably fit your calculations somewhere on this page or on the previous page.

BOTH LAB PARTNERS: READ THE ENTIRE PROCEDURE BEFORE YOU BEGIN.

Procedure

1. Distillation

Several students can share the same distillation apparatus. Review your notes on the distillation that you did at the beginning of the course.

- Fill the ‘stil pot no more than ¾ with nail polish remover.
- Add a boiling stone.
- Have the teacher approve your distillation apparatus before you begin.
- Discard the first few mL of distillate in the organic chemical waste beaker provided.
- Record the temperature of the thermometer in the ‘stil head as the distillate is collected.
- Collect a further 5 mL of distillate in a clean test tube labeled with your name.

Temperature of ‘stil head thermometer during collection: _____

2. Microscale Boiling Point Determination of Distillate

Lab Partner #1

- Using a clean pipet, add pure acetone to a height of about 1.5 cm to the mini test tube.
- Place a 1 cm microcapillary tube, open end down, in the above tube containing the acetone as shown in figure 1. The microcapillary tube must be leaning against the walls of the test tube.
- With an elastic band, fasten a thermometer to the test tube, with the bulb of the thermometer even with the bottom of the test tube.

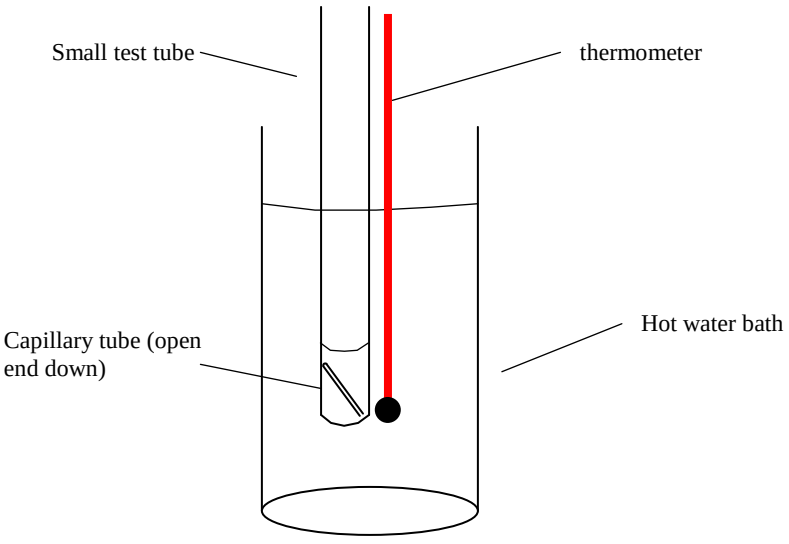


Figure 1. Assembly for micro boiling point determination

Lab Partner #2

- Prepare a hot water bath at 60^{o}C (no lower), using hot water from the kettle mixed with cold tap water if necessary.

Both Partners

- Place the mini test tube/thermometer assembly in the hot water bath as shown in figure 1.
- Within seconds you should see a steady stream of bubbles (acetone vapour) exiting the open end of the microcapillary. This indicates boiling.
- As the temperature decreases, the bubbling rate will decrease.
- At this point one student should keep an eye on the bubbling while the other watches the thermometer. Remember to estimate between the two smallest scale divisions.
- **The boiling point of the sample is reached when the last bubble exits the microcapillary.**

Experimentally determined boiling point of distillate: _____

- Detach the mini test tube, with elastics, from the thermometer and place the test tube in the beaker provided in the fumehood for later disposal by the teacher.

3. Determination of Molar Mass of the Distillate

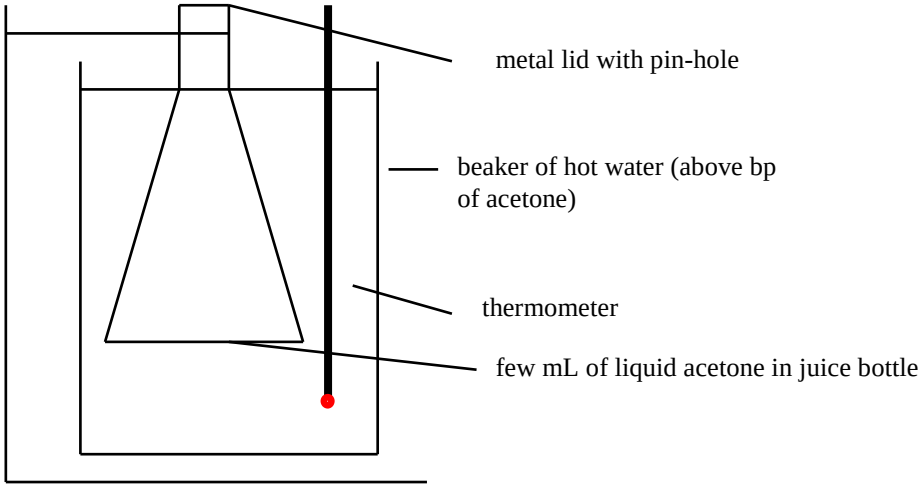


Figure 2. Assembly for determination of molar mass

- Record the mass of a clean, dry bottle, including lid.
- Remove the lid from the bottle; add a few mL of pure acetone; swirl to "coat" the inside of the bottle with acetone. There should be no more than 1 mL of liquid acetone at bottom of bottle. Recap the bottle.
- In a 600 mL beaker, prepare an approximately 60^{o}C warm water bath using water from the kettle and cold tap water, if necessary. (The water temperature must be at least above the bp of acetone.)
- In another 600 mL beaker, prepare a similarly-sized cold tap water bath.
- Place a thermometer in the hot water bath.
- In the next steps, do not let water touch the lid of the flask.
- With your hand, immerse the bottle, to just below the top in the hot water bath, as shown in figure 2, until the liquid acetone vapourizes.
- When the acetone has vapourized, record the temperature of the water bath—this is the temperature of the vapour at this point in the experiment.
- Quickly transfer the flask to the cold water bath. Allow to cool for several minutes.
- Remove the flask from the water bath. Dry it, and re-weigh.

Record Experimental Data Here (there are more spaces than you'll need)

Post Lab Questions

1. Write an appropriate abstract for this experiment below. {C, 3}

2. a) How did the temperature at the ‘stil head, as you collected the distillate, compare with the boiling point of the distillate that you determined? {K/U, 1}

b) How does the boiling point that you obtained compare with the accepted boiling point for acetone? That is, what % error did you obtain? {I, 2}

3. Use the data you collected in part 3 to determine the molar mass of acetone (g/mol). Pay attention to units throughout the calculations. No rough work on this sheet. {I, 4}

experimentally determined molar mass of acetone = _____.

4. a) The formula for acetone is C_3H_6O. What was your percentage error in the molar mass? {I, 2}

b) How could you re-do this experiment, using different equipment in our lab, to make it more accurate? Hint: Consult pre-lab questions. {I, 1}

—fin—

Crescent School
SCH3U name_____

Types of Chemical Reactions—something to think about:

Required Background: Types of Chemical Reactions (synthesis, decomposition, single displacement, double displacement, combustion)

Question: When Na metal is added to $CuSO_4(aq)$, what reaction should occur?

Let's try the reaction. List observations below:

1. What, specifically, is present in an aqueous solution of $CuSO_4$?

2. What is produced during the reaction of Na with $CuSO_4(aq)$?

3. What happens to what is produced during the reaction?

4. What remains after the reaction?

Write three chemical equations to explain your observations. Use these to answer the initial question.

5. _____

6. _____

7. _____

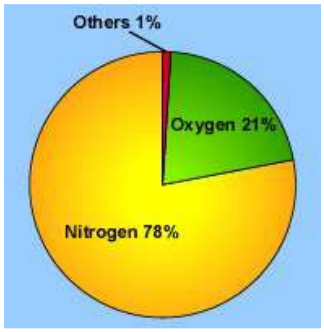
—fin—

Crescent School
SCH3U name _____

A Determination of the Percentage Oxygen (v/v) in Air

[references: JCE Editorial Staff, *J Chem Educ*, 2001, **78**, 512A; Martins, GF, 1987, *J Chem Educ*, **64**, 809; Birk, JP; McGrath, SK, *J Chem Educ*, 1981, **58**, 804.]

Air is not pure oxygen. Its composition, in terms of % by volume, is illustrated as follows:



In this investigation, you will make use of the rusting of iron to determine the % O_2 (v/v) in air. Iron reacts with moist $O_2(g)$ to produce a mixture of hydrated iron (III) oxides, $Fe_2O_3 \cdot nH_2O$ and iron (III) oxide-hydroxide, $FeO(OH) \cdot Fe(OH)_3$.

To keep things simple, and with no loss of accuracy in the experiment, we'll represent the rusting of iron with the equation:

$$4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s) \qquad \text{(eqn 1)}$$

Steel wool is a virtually pure iron, with a high surface area.

The experimental apparatus is illustrated in figure 1. We will use a lot less steel wool—not much is required.

Each pair of students requires:
2 - 100 mL glass graduated cylinder (the second graduated cylinder can be shared between two groups)
grease pencil
small, loosely packed "ball" of steel wool, about 0.3 g will be more than enough, to fit inside the graduated cylinder
1- L beaker filled with tap water
tongs
stirring rod
access to some acetone (in the fumehood)
ca 20 mL of 50% vinegar ([acetic acid] ≈ 0.4 mol/L)
several paper towels

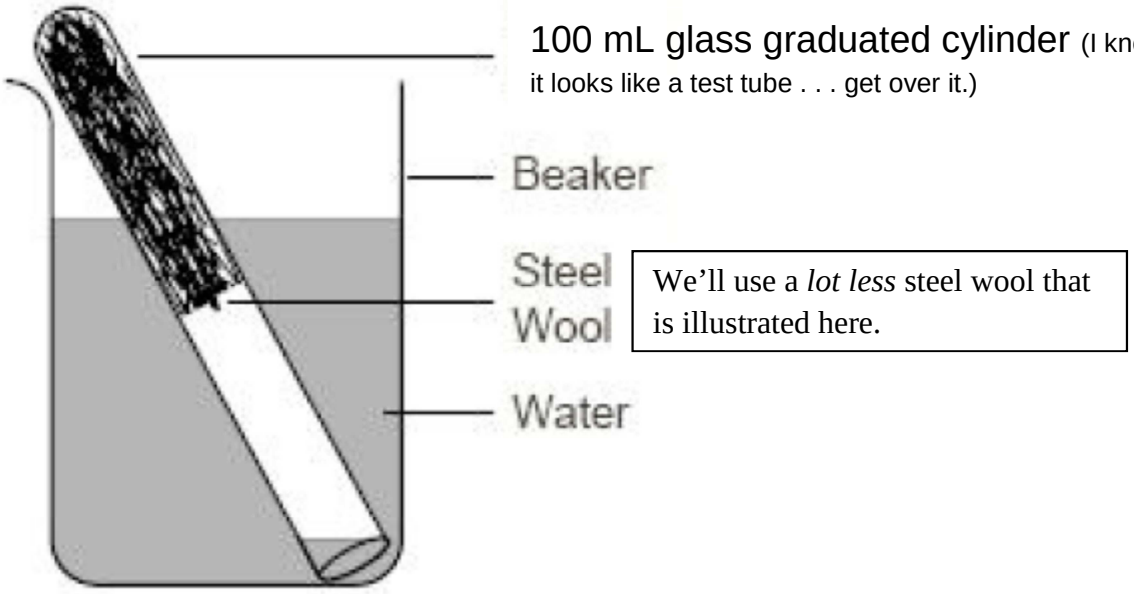


Figure 1. Experimental set up

Brief Procedure

NB. This does not tell you what measurements need to be taken, or when to take them. See—and complete—the pre-lab questions.

- Perform this step in the fume hood. Start by washing the piece of steel wool you were given with a bit of acetone. This will remove any oil from the steel wool. Press the steel wool dry between paper towels.
- "Fluff" the steel wool to increase its volume, so it will fit snuggly in the bottom of the graduated cylinder.
- Hold the steel wool with tongs and place it in a 50% vinegar solution for a few seconds. Remove the steel wool from the vinegar solution, shake off the excess vinegar—no need to dry it—and carefully, so as not to compact the "ball" of steel wool, use the stirring rod to move the steel wool to the bottom of the graduated cylinder.
- Invert the graduated cylinder in the beaker of water as illustrated in Figure 1.
- Wait for the reaction between the steel wool and the oxygen in the air to finish.

Table 1. Experimental Data (There may be more cells than you require.)

Pre-lab Questions

1. a) Why wash the steel wool with acetone? {K/U, 1}
b) Why would the manufacturer of steel wool put a thin layer of oil on it prior to packaging? {K/U, 1}

2. What do you expect to happen to the water level in the graduated cylinder as the reaction proceeds? {K/U, 1}

3. a) Roughly speaking, how long do you expect this reaction to take—a few seconds, a few minutes, a couple of hours, overnight, a week? Provide some justification for your answer.
b) How will you know—visually—when the reaction is over? {K/U, 1}

4. a) Should the amount of iron used be limiting or excess? Explain with the aid of eqn 1 above. {PS, 2}
b) Explain why the steel wool need *not* be weighed before the reaction. {K/U, 1}

5. a) Is the volume of the steel wool used in this experiment an important thing to know? Explain briefly. {C, 3}
b) Explain how you could determine the volume of the steel wool used. {K/U, 2}
c) Parts (a) and (b) aside, explain why can we ignore the volume occupied by the steel wool. {C, 3}

6. Complete Table 1 to indicate which data must be recorded before and after the reaction. {K/U, 2}

Post-Lab Questions

1. The usual lab report, maximum 1 page. (Descriptive title, abstract, data, calculations, answers to questions, conclusion.)

2. Use the data you obtained in the experiment to determine the %(v/v) O2 in air. Pay attention to the proper use of significant figures. {PS, 3}

3. a) If the following side reaction were to occur during this investigation, explain *how* it would it affect the %(v/v) of $O_2(g)$ in the air that you determined.

$$Fe(s) + 2\ CH_3COOH(aq) \rightarrow Fe(CH_3COO)_2(aq) + H_2(g)$$

b) Is it possible that this reaction occurred in your experiment? Explain briefly. {C, 2}

—fin—

Crescent School
SCH3U name_____

Quiz: Reactivity of Alkali Metals with Water-b

total marks = _____
12

This quiz is based on the demonstration of the reaction of Li, Na, K with water.

1. Write the unbalanced chemical equation for the reaction of potassium with water. Be sure to include the physical states as appropriate: (s), (l), (g), (aq). {K/U, 4}

2. A positive "pop" test proves the existence of hydrogen gas. Write the unbalanced chemical equation that corresponds to the pop test itself, NOT to the production of hydrogen gas. Be sure to include the physical states as appropriate: (s), (l), (g), (aq). {K/U, 4}

3. What did we do to prove that the resultant solution(s) were basic? Be sure to include the corresponding observation. {K/U, 2}

4. List the order of reactivity (with water) of the alkali metals we tested. List the least reactive alkali metal on the left. {K/U, 2}

—fin—

Crescent School
SCH3U name:_____

Assignment: Determination of the Atomic Radius of an Aluminium Atom

This assignment builds on your previous determination of the thickness of a piece of Aluminium foil. You will need either:
a) the measurements and the calculated thickness of the foil—among other things—to determine the radius of one aluminium atom, or
b) the density and mole concept ← preferred method.

Assume that an aluminium atom is spherical. To make it easier, think of the sphere sitting inside a cube. Stack the cubes.

Provide a diagram with your answer.

State your answer for the radius of an aluminium atom in picometres (pm),
where 1 pm = 10^{-12}m.

1. Show your calculations, with units. Pay attention to significant digits.

2. Look up the accepted value of the radius of an Al atom. State the % error in the value you determined. State your source.

—fin—

Crescent School
SCH3U name _____

Let's Take the Fun Out of Rum & Coke

Those of us over the age of 19, may, from time to time, enjoy a mixed drink. In Canada, rum mixed with cola has proven to be a popular libation. Julian, of *Trailer Park Boys*, swears by it.

a) Calculate the concentration, in mol/L, of sodium ions, sucrose, ethanol, and caffeine in 200 mL of a rum and coke beverage prepared by mixing 1 part rum with 4 parts coca-cola? Assume volumes are additive.

b) Calculate the ppm concentration of sodium ions in this beverage. What simplifying assumption must be made to be able to solve this problem? Why is it okay to make this assumption?

Coca-Cola contains 30 mg sodium (as Na^+) per 250 mL 30 g sugar (assume sucrose $C_{12}H_{22}O_{11}$) per 250 mL 34.5 mg caffeine ($C_8H_{10}N_4O_2$) per 355 mL	Rum contains 40% (v/v) ethanol, C_2H_5OH (density of ethanol = 0.79 $g \cdot mL^{-1}$)

On this sheet, provide a neat, well organized answer. Do rough work elsewhere.

—fin—

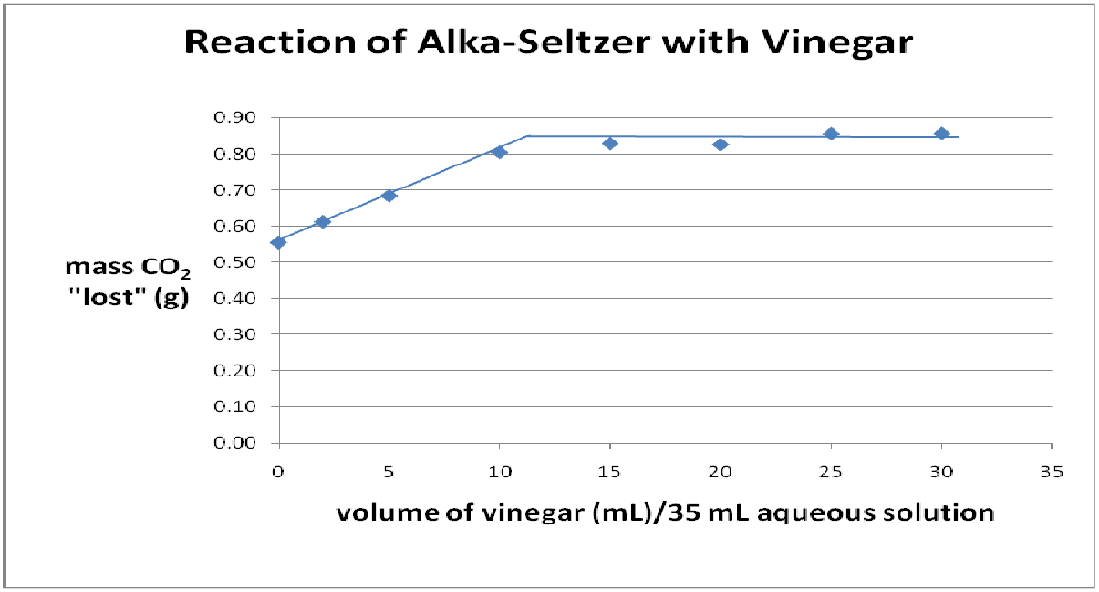
Answers: a) $[Na^+]$ = 0.0042 mol/L; [sucrose] = 0.28 mol/L; [caffeine] = 0.000401 mol/L; [ethanol] = 1.4 mol/L

b) [Na+] = 97 ppm. Assume density of the beverage is 1.0 g/mL (same as water); this allows us to equate 1000 mL of the drink with 1000 g. To 2 sig figs this is a reasonable assumption.

Crescent School
SCH3U name _____

Determination of the Water Solubility of CO_2 and of CH_4

Introduction
We recently did some research on the reaction of Alka-Seltzer tablets with aqueous solutions of vinegar. A graph of our class data is reproduced below.



You calculated that each A-S tablet contains about 1.6 g of $NaHCO_3$, while the manufacturer claims 1.9 g. We attributed the lower amount of $NaHCO_3$ to the solubility of CO_2 in water—the solvent for the acetic acid (vinegar). Let's see if this is true.

In this experiment you will determine the solubility of carbon dioxide in water. By comparison, you'll also measure the water solubility of methane (CH_4).

Materials

- pneumatic trough
- retort stand and clamp
- 3 – 500 mL Erlenmeyer flasks
- 2 – 2-L beakers
- 1000 mL graduated cylinder
- 2 rubber stopper to fit two of the 500 mL flasks
- one-holed rubber stopper fitted with a rubber hose to fit the other 500 mL flask
- another rubber hose attached to gas jet (source of methane)
- vinegar (60 mL)
- 2.5 g $NaHCO_3$ (baking soda) wrapped in some paper towel
- Saran Wrap
- grease pencil

General Instructions

- The CO_2 will be prepared by reacting $NaHCO_3$ with excess vinegar.
- The CO_2 will be collected by downward displacement of water until it *completely* fills a 500 mL erlenmeyer flask.
- Stopper this flask, and transfer it—inverted—to a 2-L beaker that contains 1000 mL of water, accurately measured. Then remove the stopper from the inverted flask, allowing the CO_2 to be in contact with the water. The teacher will demonstrate.
- Cover the entire apparatus with Saran wrap. Label the beaker with your initials.
- You will examine the beaker next class to determine what volume (mL) of CO_2 dissolved in the water.
- For comparison, you will also measure the water solubility of natural gas, which we will assume is pure methane, CH_4, in a similar manner. Get the methane from the gas tap.

Pre-Lab Questions (Answer on a separate sheet.)

1. a) Why is it necessary to measure the *total* volume of a 500 mL erlenmeyer flask?
   b) How do you propose to measure the *total* volume of a 500 mL erlenmeyer flask using the equipment provided?

2. Why is it necessary to have a *known* volume of water, 1000 mL in this case, in the large beaker?

3. On the data tables on the following page, indicate what needs to be recorded today and what needs to be recorded next class.

4. Write the balanced chemical equation for the reaction of vinegar and baking soda.

Data Collection

Record data in the tables below. There may be more spaces than you need.

1. For the solubility of CO_2

2. For the solubility of CH_4

Post-Lab Questions (Answer on a separate sheet.)

1. Use your data to calculate the solubility of CO_2 in water in units of mol/L, g/L, mL(SATP)/L.

2. Use your data to calculate the solubility of CH_4 in water in units of mol/L, g/L, mL(SATP)/L.

3. a) The Merck Index (11 edition, entry 1816) states solubility of CO_2 to be 80 mL CO_2(g)/100 mL of water at 20^{o}C. Calculate the percentage error of your experimentally determined value.
   b) How do you expect the solubility of CO_2 in water at 10^{o}C and at 60^{o}C to compare to its solubility at room temperature?
4. The Merck Index (11th edition, entry 5863) states the solubility of methane to be 3.5 mL CH_4/100 mL H_2O at 17^{o}C. Calculate the percentage error of your experimentally determined solubility of CH_4.
5. Qualitatively explain why CH_4 has such a low water solubility.
6. Write a suitable *abstrac*t for this experiment.

—fin—

Crescent School
SCH3U name _____

Visible Spectrophotometry and the Beer Lambert Law

Pre-lab questions due: _____
Graphs due: _____
Post-lab questions due: _____

Purpose

The purpose of this laboratory is to:

- learn how to use the Spectronic 20 (the Spec 20), a spectrophotometer;
- record the visible-region spectrum of solution of strawberry Kool-Aid, and to determine its wavelength of maximum absorption, λ_{max}.
- determine the relationship between the absorbance of light and concentration of a solute; to "discover" the Beer-Lambert Law
- use a spreadsheet to perform calculations and plot graphs
- prepare and use a calibration curve based on the Beer-Lambert Law

Introduction

This experiment deals with the absorbance of light by solutes in aqueous solution.
Coloured compounds absorb visible light—that's why they are coloured. Conversely, compounds that dissolve in water to produce clear solutions do not absorb visible radiation—they transmit virtually 100% of the visible-region light shone through them.

You will be given an aqueous solution of Kool-Aid crystals, prepared at 1.12 g/L in a small beaker.
In part A you will:

i. use the Spec 20 to analyse the absorption of light by the solution from 400 nm to 700 nm.
ii. determine the wavelength at which the solute, the Kool-Aid, absorbs the most light (λ_{max}).

In part B you will determine the relationship between the absorbance of light at the λ_{max} of a solution as a function of the concentration of the solute.

Background Information

<u>What is Light?</u>

Light is a form of electromagnetic radiation. Sir Isaac Newton (1642-1727) postulated that light consisted of particles—now called photons. In the early twentieth century, scientists began to consider light as a wave. Today, both models of light are used, depending on the application. This is known as Wave-Particle Duality of light.

In this experiment, we will consider light as a wave. Light waves make up a small part of what is known as electromagnetic radiation. Radio and TV transmissions are electromagnetic radiation; so are microwaves, infrared light, ultraviolet light and X-rays.

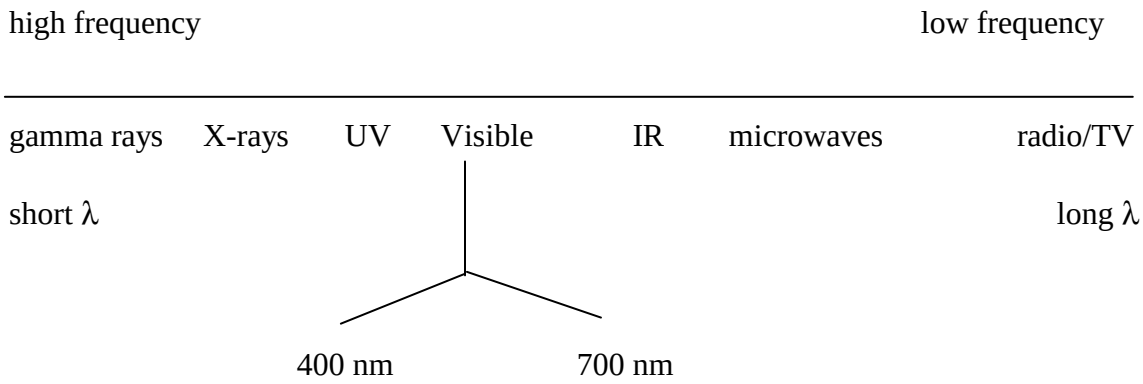
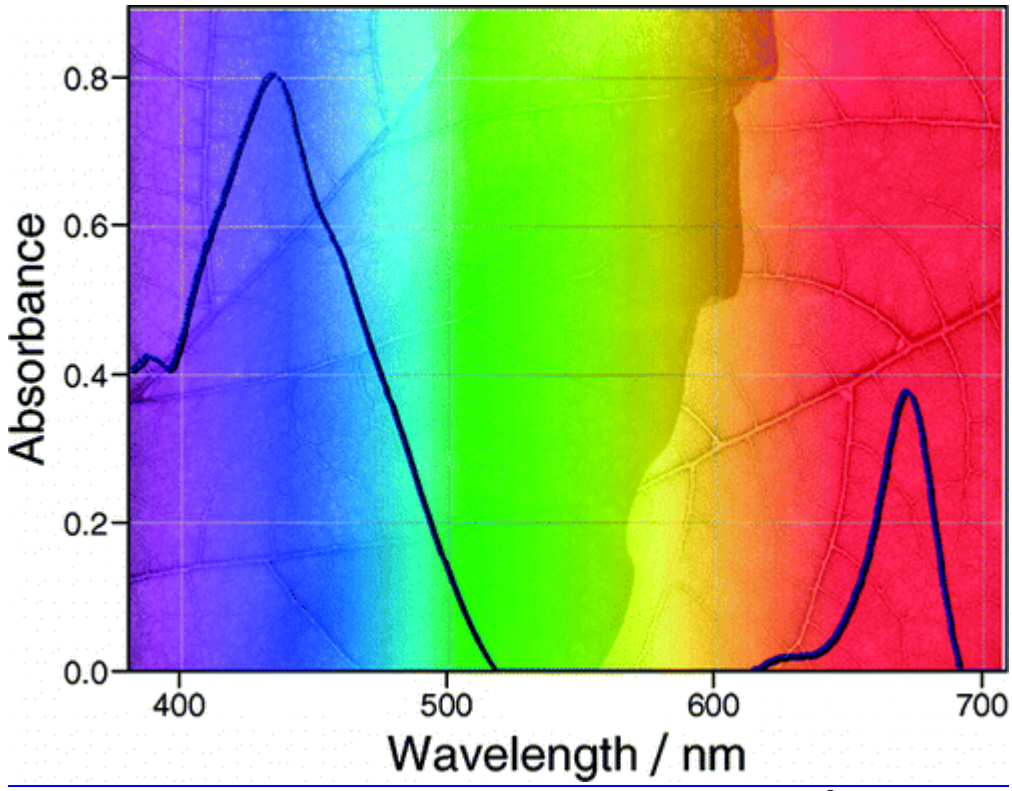


Figure 1. The Electromagnetic Spectrum. Visible light—light that the human eye can perceive—is only a small portion of the electromagnetic spectrum.



All electromagnetic waves travel at the speed of light, 3.00 x 10^8 m/s. The frequency, f, and the wavelength, λ, are related by the universal wave equation

$$v = f \cdot \lambda \qquad \text{(eqn 1)}$$

The transmittance of light, T, by a solution is defined as the fraction of the incident radiation transmitted by the solution.

$$T = P_t/P_o \qquad \text{(eqn 2)}$$

where P_t is the power of light *transmitted* by the solution and P_o is the power of the light that is *incident* on the solution. (Units of power, P, are W; transmittance has no units, since it is a ratio of two power values.)

Transmittance is sometimes given as a percentage.

$$\%T = 100\% * T \qquad \text{(eqn 3)}$$

Alternatively, we may wish to consider the amount of light *absorbed* by a solution. While it may seem reasonable that the absorbance, A, and transmittance, T, of light by a given solution are reciprocals of one another, this is not the case. They are related by the equation

$$\mathbf{A = -\log_{10}(\%T/100)} \qquad \text{(eqn 4)}$$

(For a derivation of this equation, dust off your calculus and see, for example, D A Skoog and D M West, *Fundamentals of Analytical Chemistry*, 3rd ed, Holt, Rinehart and Winston, New York, 1976, p 505, or any advanced analytical chemistry text.)

Schematic Diagram of a Spectrophotometer—a Spec 20

The light source in a Spec 20 is a tungsten lamp. This lamp gives off all wavelengths of visible light. To get a light of a specific wavelength—monochromatic light—a wavelength separator such as a prism or diffraction grating is used. This is called a *monochromator*.

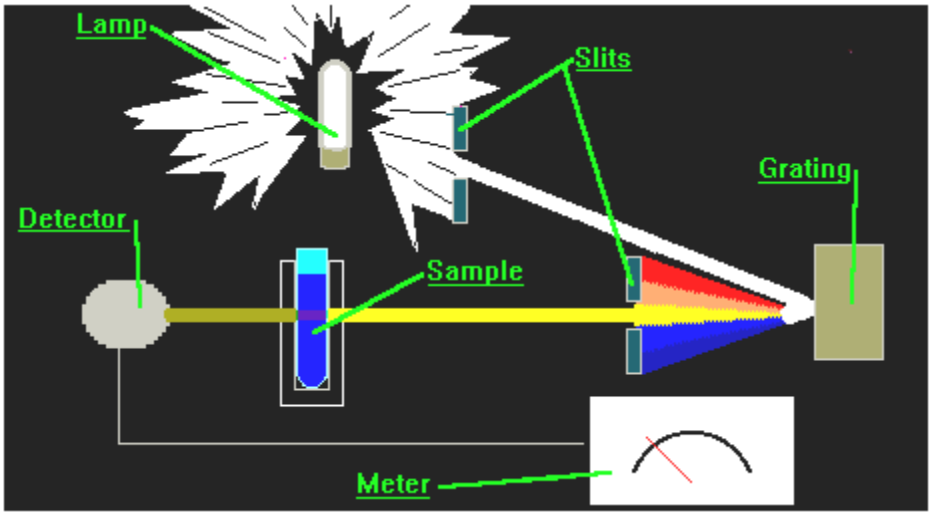


Figure 2. Schematic of a single beam spectrophotometer such as a Spec 20.

In Part A of this experiment we will examine the absorbance of visible light by a solution of strawberry Kool-Aid solution. Since the Spec 20 operates at only one wavelength at a time, we will "scan" the 400 nm to 700 nm region in 20 nm increments. *For reasons that will be clear to you later in the experiment, use 10 nm increments in the region 480 nm to 540 nm.*

Pre-lab Questions (Answer on a separate sheet.)

1. Prepare a set of EXCEL tables and graphs as follows:
   a) For Part A: Table 1 with columns for λ (nm) , %T, A
   Graph 1: A (y axis) versus λ (x axis)

   You will record your data directly into the EXCEL spreadsheet you prepared.

   b) For Part B: Table 2 with columns for %T, A, [Kool-Aid]
   Graph 2: A versus [Kool-Aid]

2. What is the mathematical relationship between the % Transmittance of a solution and the Absorbance of light by the solution?

3. Read the procedure for Parts A and B. Then visit the website
   http://www.wellesley.edu/Biology/Concepts/Html/analogspec20instructions.html

   for an animated tutorial on how to use the Spec 20.

   a) You have a, say, 6.24 g/L solution of a Kool-Aid. Explain, with the aid of a calculation, how you would prepare 10.00 mL of a 2.34 g/L solution. You have a medicine dropper and a 10.00 mL graduated cylinder—the poor man's volumetric flask).

4. Why is it necessary to calibrate the Spec 20 before use?

5. Refer to step #4 of the Part A procedure below. After the cuvette is rinsed with distilled water, why is it necessary to subsequently rinse it with a few mL of the Kool-Aid solution before the filling the cuvette with the same Kool-Aid solution?

Part A: Det'n of the Absorbance Spectrum of Your Kool-Aid Solution

1. The Spec 20 will be warmed up for 15 minutes prior to class to stabilize the light source.
2. Look carefully at the scales on the Spec 20. Why do we record %T—and subsequently convert it to A—when we could simply record A right off the scale? (If you're using the digital machine, record %T anyway and let your spreadsheet convert it to A.)
3. Use the wavelength control knob on the top right of the Spec 20 to set the wavelength to 400 nm. This is the first wavelength that will be used. Some Spec 20s have a filter adjustment located at the lower left of the front of the machine. If so, set this for the range 340 - 599 nm. When wavelengths greater than 599 nm are selected, adjust the filter accordingly.
4. Rinse a clean cuvette with some distilled water. Now rinse the same cuvette with a few mL of your Kool-Aid solution. Now fill the cuvette with your Kool-Aid solution. Fill another clean cuvette with distilled water.

5. With the sample compartment empty, adjust the left-hand knob on the front of the Spec 20 to read 0%Transmittance, or 0%T. (When the sample compartment is empty, the "gate" is closed. No light passes through it, hence 0% transmittance.)
6. Wipe the cuvette containing distilled water with a tissue. Insert it into the sample compartment, lining up the white vertical line on the cuvette with the raised mark at the front of the sample compartment.
   If the cuvettes are always inserted this way, any imperfections in the glass will constitute a *systematic*, rather than a *random* error. (Why is this important?)
7. Adjust the right-hand knob on the front of the Spec 20 to read 100%T. This allows the Spec 20 to ignore any absorbance of light due to the solvent—distilled water in this case.

You must repeat the calibration steps (#5—7) every time a new wavelength is selected.

8. Place the cuvette containing your Kool-Aid solution in the sample compartment; close the lid. Read and record the value of %T, estimating the decimal place. (Why estimate the decimal place?)
9. Repeat step 7 for wavelengths 420 nm to 700 nm in 20 nm increments. Use 10 nm increments between 480 nm and 520 nm.
10. Using a spreadsheet program, convert %T values to absorbance, A. Plot a graph of A versus λ for strawberry Kool Aid. Identify the wavelength of maximum absorbance of light, λ_{max}. Check this value with the teacher before proceeding to Part B.

Part B: Determination of the Relationship Between Absorbance of Light and Concentration of the Solute: the Beer-Lambert Law.

11. Set the λ_{max} , determined in Part A, on the Spec 20. Calibrate the Spec 20 at this λ as in steps 5—7 above.
12. Record the %T of your Kool-Aid solution at this λ. Once again, enter this into a speadsheet program and convert it to A. Use these data to plot a graph of A (y axis) versus [Kool-Aid] (g/L) on the x axis.
13. You must now record the %T (and hence A) for several diluted solutions of Kool-Aid of known concentration. Use the equation $M_cV_c = M_dV_d$, a 10 mL graduated cylinder and a medicine dropper to perform these dilutions.

NB The concentration of the Kool-Aid stock solution is 1.12 g/L.

14. a) Draw the line of best fit (or better yet, let the spreadsheet do it) on this graph of A versus [Kool-Aid] in units of g/L.

    b) Determine the slope of the best-fit line (or let the spreadsheet do it) and write the equation of the line.

Post-Lab Discussion

Bring your graphs to our next class for a useful post-lab discussion. This will help you answer the post-lab questions.

name _____

Visible Spectrophotometry and the Beer Lambert Law

Post-Lab Questions (Answer neatly in the space provided.)

1. Write the Beer-Lambert Law. You must be able to explain this equation and how to use it. {C, 1}

2. Use your graph of A vs [Kool-Aid] (square brackets around a chemical name or formula refers to the concentration of that substance, in the specified units) to determine the value of ε, the molar extinction coefficient, for strawberry Kool-Aid] at its λ_{max}. Include proper units for ε. {K/U, 3}

3. Imagine that you have a solution of strawberry Kool-Aid of unknown concentration. Explain how your graph of A vs [Kool-Aid] can be used, along with a Spec 20, to determine the concentration of the given strawberry Kool-Aid solution? {K/U, 2}

4. a) Remember that mass (g) of a given compound can be converted to moles, provided that the molar mass of the compound is known. That means that for a compound of known molar mass, concentration units of g/L can be used interchangeably with mol/L. Why did we express the concentration of the Kool-Aid solution in g/L rather than mol/L? (C)

   b) Why do chemists prefer using moles, rather than grams, whenever possible? (C)

5. A certain blue dye (F D & C Blue #1) has an ε value of 1.38×10^5 $mol^{-1} \bullet L \bullet cm^{-1}$ at its λ_{max} of 630 nm. A sample of the dye is analysed in a Spec 20 at 630 nm in a cuvette of diameter 1.25 cm. The transmittance is recorded to be 38.6%. What is the concentration, mol/L, of the dye? (PS)

—fin—

Crescent School
SCH3U name_____

Let's get started: Analysis of a Mixture

Since this is the first activity in 11 Chemistry, we will begin with a pre-lab discussion. Take notes; record answers to pre-lab questions as appropriate.

Introduction and Pre-Lab Questions

It has been suggested that the sugar (sucrose, $C_{12}H_{22}O_{11}(s)$) packets available in coffee shops and cafeterias may, in some cases, be "cut" with baking soda (sodium hydrogen carbonate, also called sodium bicarbonate, $NaHCO_3(s)$).

1. Why might sugar—in a sugar packet—be mixed with baking soda?

_____

You will be given three 1.0 g samples, each labelled with an identification number, which you will record.

2. Why is it necessary to record the identification number of each sample?

_____

Your job, should you choose to accept it, is to identify each sample, and to explain your reasoning.

3. a) Explain why you are not allowed to taste the samples that you are given in the lab?

_____

b) Why is it okay to taste a sample obtained from a cafeteria?

_____

Materials/Equipment

- 3–1.0 g samples, each labelled with an identification number: pure sucrose; pure $NaHCO_3$; ca 50/50 mix of sucrose/$NaHCO_3$.
- phenolphthalein indicator in dropping bottle
- conductivity apparatus
- some aqueous acetic acid, $CH_3CO_2H(aq)$—commonly called vinegar
- Beral pipet
- access to an electronic balance
- distilled water
- beakers, etc

Useful information:

- Phenolphthalein (note the spelling) is an acid-base indicator: it is clear in acidic solution; pink in solutions of pH > 8
- Solutions that contain dissolved ions conduct electricity
- Vinegar reacts with baking soda to produce carbon dioxide gas, and other products.

4. Identify sucrose and $NaHCO_3$ as either ionic or covalent compounds.

---

5. On scrap paper, complete question 6 now. Then . . .
a) Based on your predicted results, do you expect this investigation to be qualitative or quantitative? Explain briefly.
b) Will the addition of phenolphthalein and a test of electrical conductivity give you enough information to unambiguously identify each sample? Explain briefly.

6. On a separate sheet, prepare a neatly drawn table to record the tests that you plan to do. Include ID numbers of each sample and a separate column for *predicted* observations as well as *actual* observations.

7. Explain why reading the post-lab questions before beginning the investigation is a good idea?

Get busy.

Post-lab Questions (Answer—neatly—on a separate page; no rough work will be marked.)
8. Identify each unknown you were given. *Briefly* explain your reasoning.

9. What are your take-aways from this investigation? Namely . . .
a) Was the investigation qualitative or quantitative?
b) Write a balanced chemical equation for the reaction of vinegar with baking soda. Include physical states of each reactant and product—(s, l, g, aq).
c) Write a note on how you used a conductivity meter and phenolphthalein indicator to help you in this investigation.
d) What role, if any, did the vinegar play in this investigation?
e) Imagine: You're sitting in McDonald's with your friends, who are not chemists. The McDonald's people aren't stupid: they won't add baking soda to the sugar. Yet someone at your table insists that they do. Explain how you could prove that there is no baking soda mixed with the sugar. Remember—you're sitting in McDonald's.

—fin—

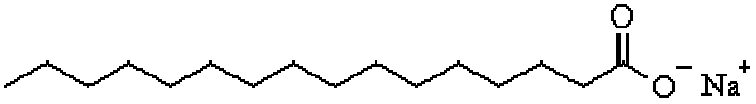
Crescent School
11 Chemistry (SCH3U) name: _____

Preparation and Testing of Soap

references: J Chem Ed Classroom Activity #14, J Chem Ed, 76, 2, Feb 1999, p 192B
http://waltonfeed.com/old/soap/soaptabl.html
Hill and Petrucci, General Chemistry, 6th edition, Prentice Hall, Upper Saddle River, NJ, 1999

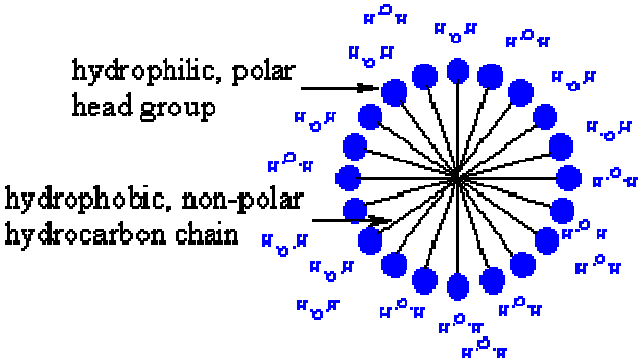
Introduction
Download (Green Room) and view the Power Point presentation on soap before completing this lab.

Soap is a special type of molecule. It has a long, non-polar hydrocarbon tail and an ionic head.



non-polar hydrocarbon 'tail' polar 'head '

Soap molecules surround a droplet of oil by dissolving their non polar tails in the oil. The polar head dissolves in the surrounding water forming a *micelle*. This is how soap *emulsifies* fat or oil.



A soap micelle.

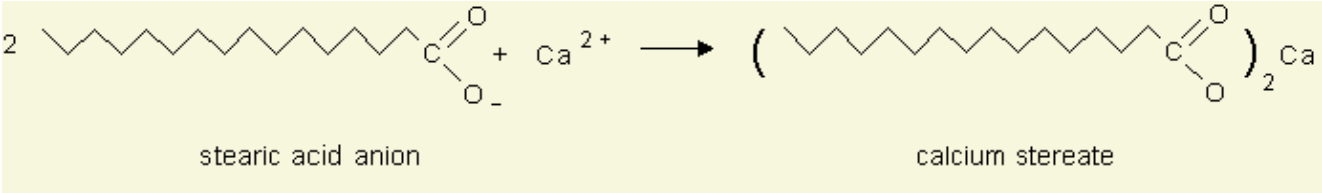
The *emulsification* of fats and oils (non-polar) in water (polar) allows fats and oils to be washed away.

In this experiment you will prepare a simple soap using coconut oil and a concentrated solution of sodium hydroxide, NaOH(aq). (Sodium hydroxide can be purchased at the hardware store. Ask for *lye*.)

The chemical reaction involved in the synthesis will be discussed further in AP Chemistry.

Hardness of Water
Hard water has a high mineral content. This content usually consists of high levels of metal cations, mainly calcium (Ca^{2+}) and magnesium (Mg^{2+}).

Calcium ions (and Mg^{2+} ions) react with soap anions (eg stearate anions) to form sparingly soluble "soap scum", in the form of calcium stearate. (Apologies for the spelling of calcium stearate below—downloaded from internet; not editable)



Soap scum reduces the efficiency of soap. It can also leave a residue on clothing. Calcium stearate can clog pipes and leave a nasty-looking bathtub ring as well.

<u>Materials</u>

- electronic balance; hot plate
- alcohol thermometer; stirring rod
- 10 mL graduated cylinder; 2 – 50 mL beakers
- 20 g coconut oil in a 50 mL beaker; 20 mL of 6.0 mol/L NaOH in 50 mL beaker
- piece of plastic wrap (10 cm x 10 cm)
- small sample of coconut oil-based soap to use in tests, if these tests are to be done on the same day as the soap preparation
- four test tubes containing ca 5 mL of each: distilled water, tap water, “hard water, “soft” water—in a test tube rack
- universal indicator in dropping bottle

Teacher Demonstration:

Both dishwashing detergent and automatic dishwasher powder will be added to samples of distilled water, tap water, “hard water, “soft” water. All of the observations will be discussed.

Safety Precautions

- Wear safety goggles and a lab apron throughout the experiment.
- Do not allow the electrical cord from the hotplate to dangle over the edge of the bench. Listen carefully to the teacher's instructions.
- If 6.0 mol/L NaOH solution gets on your skin, wash immediately with plenty of cold water. Notify the teacher.
- Be careful not to push the stirrer through the bottom of the styrofoam cup when you are stirring.

Procedure: Preparation of a Simple Coconut Oil-Based Soap

1. Label a small styrofoam coffee cup with your initials.
2. At the same time, place the 20 g of coconut oil, in a 50 mL beaker, and the 50 mL beaker containing 20 mL of 6.0 mol/L NaOH using a hot plate. When these samples have reached ca. 45^{o}C +/- 5^{o}C, carefully pour both into the styrofoam cup, with stirring.
3. Stir the mixture constantly for about 20 minutes until it has the consistency of cold honey. Take note of what you see. If the mixture does not thicken after fifteen minutes of stirring, let it sit for about three minutes. Then stir it for a few more minutes.
4. Lay some plastic wrap directly on top of the soap, in the styrofoam cup. Leave the soap in the location suggested by the teacher. We will examine—and use—the hardened soap next class.

Testing of Soap (and Detergent—Teacher demonstration) Performance in Different "Types" of Water

5. Test a few flakes of the coconut oil-based soap in ca 5 mL samples of water, as outlined below. For pH testing, use two drops of universal indicator in each sample.

Label the observation table below as outlined in Post-Lab #1.

	distilled water	tap water	"hard" water	"soft" water	pH
coconut oil-based soap					
dishwashing "liquid"					
automatic dishwasher powder					

Crescent School
11 Chemistry (SCH3U) name: _____

Preparation and Testing of Soap

Post-Lab Questions

1. Do the observations bear out your expectations regarding the performance of soap and of detergent in "hard" water? Explain briefly, including a chemical equation. {C, 2 + 1}

2. Given that kitchen drains are frequently clogged with grease or oil, explain why commercial drain cleaners such as Drāno contain NaOH as the active ingredient. Include, at a minimum, a word equation in your answer. {C, 2 + 1}

3. Briefly explain how soap works to emulsify fats and oils. Provide a diagram of a *micelle* in your answer. Be as specific as possible in your diagram. {C, 2 + 1}

—fin—

Crescent School
SCH3U name_____

Separation of a Heterogeneous Mixture by Physical Means

Objective: To separate a mixture of _____ .

Experimental Procedure: Design your own procedure and then carry out the separation. Any necessary equipment/supplies will be available.

Note: A technique called *decanting* may be helpful here. A demonstration will be given.

Pre-lab Questions

1. a) What is a *heterogeneous* mixture? (C)

   b) What is a *homogeneous* mixture? (C)

2. What physical properties do you plan to make use of in this separation? (K/U)

3. Write a step-by-step procedure, or flow chart, for the separation. (C)

Crescent School
SCH3U name_____

Activity - Separation of a Heterogeneous Mixture by Physical Means

Post-Lab Questions

1. Which component of the mixture did you separate first? Explain. (C)

2. Why is it important not to add too much water to the mixture? (C)

3. a) At what point in the experiment did you decant? (C)

   b) What advantage is there to decanting? (K/U)

—fin—

Crescent School
SCH3U

A Simple Distillation: Purification of Salt Water

Introduction

Salt water is an aqueous solution of table salt (sodium chloride, NaCl). More specifically, salt water consists of aqueous sodium ions, $Na^+(aq)$, and aqueous chloride ions, $Cl^-(aq)$.
One way to purify salt water is to distil it. This involves boiling the salt water and condensing the steam.

A chemical reaction will be used to verify the purity of the distillate.

Purpose: The purpose of this experiment is to purify a salt water solution by distillation, and to test that the distillate is free of salt.

Note:

- Aqueous solutions of sodium chloride (NaCl(aq)) and of silver nitrate $AgNO_3(aq)$) are clear and colourless.
- Silver chloride, AgCl, a white solid, is insoluble in water.
- Make sure that several boiling stones are present in the 'still pot. They are required for even, controlled boiling.

Safety: Wear safety goggles. If you get silver nitrate on your skin, immediately rinse with plenty of cold water.

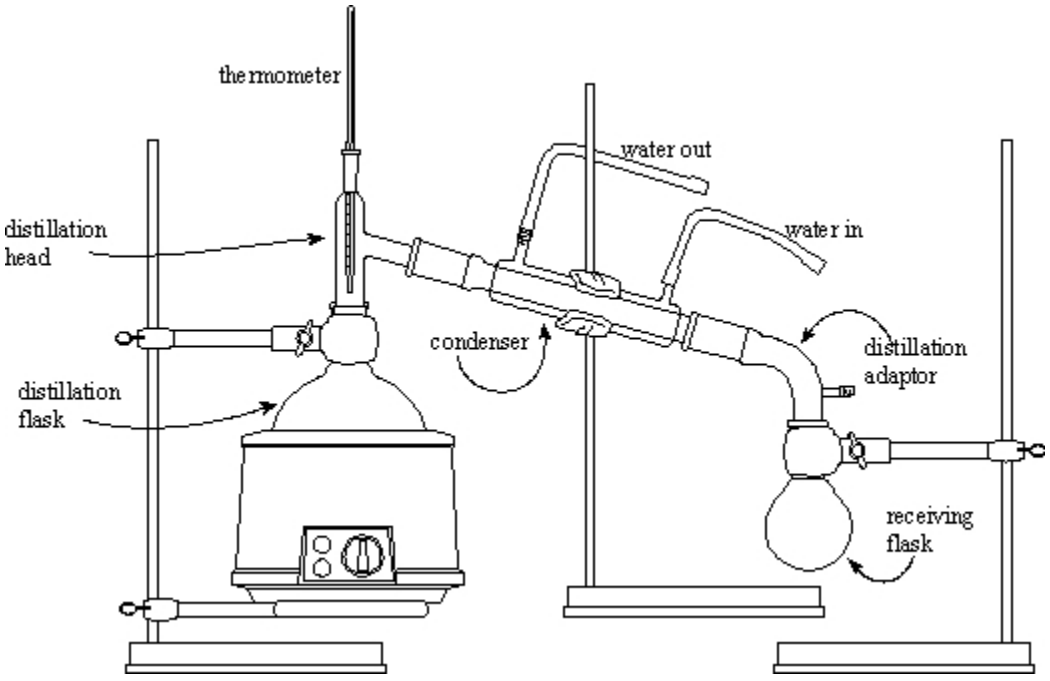


figure 1. a simple distillation apparatus

Procedure

1. Testing for the presence of chloride ions (present in NaCl) in water.

Control: In a well plate or test tube, add a three drops of silver nitrate solution, $AgNO_3(aq)$ to a few mL of distilled water. Record your observation. Dispose of this solution in the appropriately labelled chemical waste beaker.

Positive Test: Add a three drops of silver nitrate solution, $AgNO_3(aq)$ to a few mL of salt water, NaCl(aq). Record you observation. Dispose of this solution in the appropriately labelled chemical waste beaker.

Here's an explanation of the positive test. Aqueous silver ions (from silver nitrate solution) react with aqueous chloride ions (from sodium chloride solution) to form a white precipitate of silver chloride, AgCl as follows:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaCl(aq)$$

white precipitate

2. Assemble the distillation apparatus. Make sure that several boiling stones (boiling chips or boileezers) are present in the 'still and that cooling water is (slowly) flowing through the condenser. Check with the teacher before lighting the burner.

- ! Discard the first 5 mL of water collected.
- ! Collect several more mL of water. Test for the presence of chloride ions with silver nitrate solution. Note your results.
- ! Carefully note the temperature during the distillation. Remember to estimate between the two smallest scale divisions.

Class Demonstration

The teacher will set up a sample of boiling water and boiling salt water. Compare these boiling points to the one shown on the thermometer during the distillation. (Answer Post-Lab # 2 now.)

Pre-lab Questions:

1. What is the purpose of the condenser {K/U, 1}?

2. Provide a definition for the terms *aqueous* and *precipitate*. You may consult the glossary of your chemistry textbook, or the internet. (C)

3. Use the internet to find the meaning of *boiling point elevation*. (K/U)

—fin—

Crescent School
SCH3U name _____

A Simple Distillation: Purification of Salt Water

Post-Lab Questions

1. What two pieces of evidence suggested that the distillation was successful? {K/U, 2}

2. a) What temperature was recorded for pure boiling water in the teacher demonstration? _____
   What temperature was recorded for boiling salt water, NaCl(aq), in the teacher demonstration? _____
   What temperature did you record *during* your distillation? _____
   b) Is the temperature recorded during the distillation the same as that of boiling salt water? Explain briefly; think about *boiling point elevation*. {C, 2}

   c) **In one sentence**, explain the significance of the temperature reading on the thermometer during a distillation. Do not refer specifically to your distillation, but to any distillation. Use the word *distillate* (i.e. the liquid that distils over.) in your answer.{C, 2}

3. Explain why distillation involves a *physical* separation. {C, 2}

4. Why it is possible that a homogeneous mixture of two liquids might **not** be amenable to separation by distillation? {C, 2}

5. Explain how you would distill a sample of water containing methanol (b.p. 65^{o}C) and table salt, NaCl. That is:
   a) Which component of the given solution will distil first? {K/U, 1}

   b) What will remain in the distillation pot (i.e. the flask that contains the solution to be distilled) at the end of the distillation? {K/U, 1}

6. Why must boiling stones be used in the distillation flask? {K/U, 1}

7. State one practical application of distillation other than the purification of salt water that you did in this lab. {MC, 1}

—fin—

Crescent School
SCH3U name: _____

A Few Chemical Tests

In this activity, you will carry out several chemical tests:

1. Burning splint test for the presence of $CO_2(g)$.
2. "Pop" test for hydrogen gas ($H_2(g)$.
3. Glowing splint test for the presence of $O_2(g)$.
4. Addition of $AgNO_3(aq)$ to a solution to test for the presence of NaCl.

Materials
Each pair of students requires
wash bottle with distilled water
250 mL beaker to use as test tube holder
small piece of masking tape to label test tubes
4 empty test tubes
10 mL of 1.0 mol/L hydrochloric acid (HCl(aq))
1 cm piece of magnesium (Mg)
1 labelled test tube containing 5 mL salt water
1 test tube containing a tiny bit of $MnO_2(s)$
250 mL beaker containing ca 2 g of baking soda ($NaHCO_3$)
25 mL vinegar ($HC_2H_3O_2(aq)$)
bunsen burner with igniter
wooden splint
5 mL salt water in small labelled beaker

Procedure
1. Add a few mL of vinegar (acetic acid, $HC_2H_3O_2(aq)$) to the baking soda ($NaHCO_3(s)$) in the beaker provided. The following represents the chemical reaction that occurs:

$HC_2H_3O_2(aq)$	+	$NaHCO_3(s)$	→	$NaC_2H_3O_2(aq) + CO_2(g) + H_2O(l)$
acetic acid		sodium hydrogen carbonate		sodium acetate

When the reaction has subsided, insert a *burning* splint above the solution. Record your observations, with an explanation, in the space below.

2. In a test tube, add several mL of 1.0 mol/L HCl(aq) to a piece of magnesium ribbon.
Loosely cover the mouth of the tube with your thumb to contain the gas.

$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Insert a burning splint into the mouth of the test tube. Record your observations, with an explanation, in the space below.

3. Add a few mL of hydrogen peroxide ($H_2O_2(aq)$) solution to the manganese IV oxide ($MnO_2(s)$ powder in the test tube provided. You will observe the reaction corresponding to:

$$2\ H_2O_2(aq) \xrightarrow{MnO_2} 2\ H_2O(l) + O_2(g)$$

After the reaction has subsided, insert a *glowing* splint. Record your observations, with an explanation, in the space below.

4. Obtain a few mL of each of three water samples: distilled water, tap water, salt water (NaCl(aq)). To each sample, add two or three drops of silver nitrate ($AgNO_3(aq)$) solution. Record your observations, with an explanation, in the space below. The following chemical equation may help.

$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Post-Lab Questions

1. Which of the above, #1 – 4, represent *chemical* reactions? How do you know? {C, 2}

2. What would you expect to see if a burning splint is inserted into a sample of pure O_2? {MC, 1}

3. Write a balanced chemical equation to predict what would occur when $AgNO_3(aq)$ is added to a solution of potassium chloride, KCl(aq). Be sure to indicate the physical state (solid (s), aqueous (aq), liquid (l), gas(g)) of each reactant and product.{PS, 2}

4. a) Identify the chemical reactions in steps 2, 3, 4 as synthesis, decomposition, single displacement, double displacement or combustion. More than one description may apply in one of these reactions. {K/U, 3}

	Reaction type
step 2. $Mg(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$	_____
step 3. $2\,H_2O_2(aq) \xrightarrow{MnO_2} 2\,H_2O(l) + O_2(g)$	_____
step 4. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$	_____

b) What is the role of MnO_2 in the reaction in step 3? {K/U, 1}

—fin—

Crescent School
SCH3U name: _____

Getting Started: A Quantitative Relationship in Popcorn Making

[reference: Fantini, JL; Fuson, MM; Evans*, TA; *J Chem Ed*, vol 83, March 2006, p 414]

In this experiment, you will investigate the correlation between the mass of an un-popped kernel of popcorn with the mass of the popped kernel, called the flake.

Each pair of students requires:

- 250 mL erlenmeyer flask with one-holed rubber stopper and clamp attached
- bunsen burner and sparker
- 5 kernels of popping corn
- glass rod
- electronic balance

Procedure

1. Read these instructions; fill in the column headings in Table 1 on the back page; provide a descriptive title for the table.
2. Record mass of kernel in Table 1.
3. Place massed kernel in erlenmeyer flask, replace one-holed rubber stopper.
4. Ignite bunsen burner—ca 4 cm tall flame (not too vigourous).
5. Grab flask by the clamp; move flask continuously over the flame—be vigilant—until kernel pops.
6. Remove flask from flame; *immediately* invert flask to get the flake away from the heat.
7. Remove stopper to obtain the flake. Record mass of the flake.
8. Repeat for all the kernels you are given.
9. When you're done, record your data in the teacher's spreadsheet.

Data Analysis

Use either your data, or the class data, if posted on our website, to do the following:

1. With EXCEL, prepare a scatter plot–not a line graph–of flake mass (y-axis) versus kernel mass (x-axis).
2. Label axes; include proper units. Provide a descriptive title for the graph, not simply "flake mass vs kernel mass. "
3. Include only a line of best fit, called a *trendline*. Include the equation of the *trendline* and the R^2 value. Put your name on the graph, attach it to this handout. {C, 7}

. . . continued

4. a) In the space below, rewrite the equation of your *trendline* using "flake mass" and "kernel mass" instead of "y" and "x". {C, 1}

   flake mass = _____

   b) The R^2 value from your graph = _____

   In your own words, what does the R^2 value represent?

   _____

   What R^2 value corresponds to a prefect correlation?

   _____

   c) Look at the value of the y-axis intercept. Taking into account experimental uncertainty, what do you think the *real* value of the y-axis intercept should be? {K, 1}

   d) When popcorn is popped, water in the kernel is vapourized. From your graph, what is the average percentage by mass, to two significant figures, of water in a kernel of popping corn? {K, 1}

   e) If we could somehow engineer *uber*-popcorn, whose kernel mass is 5.0 g, what, according to your graph, would be the expected mass of the flake? {PS, 2}

5. What did you learn from this activity? Answer in concise sentences. {C, 2}

Table 1 _____

trial		
1		
2		
3		
4		
5		

—fin—

Crescent School
AP Chemistry name:_____

You're in "Hot" Water!

Here's the situation:
You have 100 mL of isotopically labelled water, $^{3}H_2O$. That is, the water is made with the radioactive isotope of hydrogen, tritium, ^{3}H. Compounds that contain radioactive isotopes are called "hot".

You pour this 100 mL of "hot" water into the ocean and allow it to thoroughly mix.

Question:
After the "hot" water mixes thoroughly with the ocean water, you remove 100 mL of ocean water. How many molecules of $^{3}H_2O$ do you expect there to be in this 100 mL sample? (Ignore any radioactive decay of the tritium.)

Solution:

1. First, decide *what* information is required to solve the problem. Plan a strategy of how you would solve the problem. Also, decide *where* you might look for this information. You are not required to solve the problem at this stage, but simply decide what is required and plan a strategy.

2. As part of a class discussion, ask whatever questions you think are necessary. The teacher may or may not answer your question.

3. For homework, solve the problem. Be sure to pay attention to significant digits. Do rough work on scrap paper. Put your neat, well-organized, step-by-step solution on the back of this sheet. DO NOT SUBMIT YOUR ROUGH WORK.

   [reference: D. Diemente, J. Chem. Ed., **77**, 8, p1010 (August 2000)]

—fin—

Crescent School
SCH3U name_____

Determination of the Simplest Formula of Zinc Chloride

Introduction
In this experiment you will react a piece of zinc metal with a small excess of concentrated hydrochloric acid, HCl(aq) according to the unbalanced chemical equation:

$$Zn(s) + HCl(aq) \rightarrow Zn_xCl_y + H_2(g)$$

Your job, should you choose to accept it, is to carry out this experiment in such a way as to determine the values of *x* and *y* in Zn_xCl_y.

> Before you continue, read the pre-lab questions.
> This will help you understand what you need to do.

Each pair of students requires:
access to fume hood
access to centigram electronic balance
electric hot plate
125 erlenmyer flask
ca 1 gram Zn
ca 15 mL conc (12 mol/L) HCl(aq)
dropper for the acid
latex gloves

Safety Precautions: Eye protection, latex gloves, lab apron. The entire reaction, except for weighing, must be carried out in the fume hood.

General Procedure
It is up to you to decide what measurements need to be taken. We'll have a pre-lab discussion.

Add concentrated HCl(aq) to the Erlenmeyer flask containing the zinc. Swirl the flask to speed the reaction. Once the reaction is complete, dry the zinc chloride solution, which will contain some unreacted HCl(aq), until you obtain a white, or off-white powder. That said, your challenges include the fact that zinc chloride is *deliquescent* (see pre-lab questions). Further, in the drying, it is important not to melt the zinc chloride.

Pre-Lab Questions

1. What type of chemical reaction occurs in this experiment? _____ {K/U, /1}

2. a) What does *deliquescent* mean? {C, /2}

   b) What will you do, procedure-wise, to best deal with the fact that zinc chloride is deliquescent? {K/U, /1}

3. State one reason why this **experiment should be carried out in the fume hood**.
   Hint: Hydrochloric acid is prepared by bubbling hydrogen chloride gas, HCl(g) into water, according to:

   $$HCl(g) + H_2O(l) \rightarrow HCl(aq)$$

4. Explain why this experiment is *quantitative*. {C /2}

5. Why is it suggested that you *gently* heat the flask as the zinc reacts with HCl(aq)? {K/U, /1}

6. Complete the left hand column in the following data table. There may be more cells in the table than you need. Provide an appropriate descriptive title in the space below.

{PS, /4}

Descriptive title: _____

These data must be entered into the teacher's spreadsheet before you leave the lab.

7. In point form, write a procedure for this experiment. {PS, /6}

Post-Lab Questions

8. Use your data to calculate the simplest formula of zinc chloride, Zn_xCl_y. {PS, 5}

9. Explain why simplest formula is also known as the empirical formula. {C, 2}

10. a) How does the formula that you obtained compare with the accepted formula for zinc chloride? {C, /1}

b) For each part below, you need to back up your answer. Provide a sample calculation—do a "what if" calculation. How would the x : y ratio in the formula change if:

i. The zinc chloride obtained was not sufficiently dried before it was weighed? Provide a sample calculation—do a "what if" calculation. Answer on a separate sheet. {PS, /4}

ii. The zinc chloride was sufficiently dried, but was left on the lab bench for, say one hour, before it was weighed? Provide a sample calculation—do a "what if" calculation. Answer on a separate sheet. {PS, /4}

iii. Not all of the zinc metal reacted. Provide a sample calculation—do a "what if" calculation. Answer on a separate sheet. {PS, /4}

11. Now that we've made zinc chloride, how could we recover the zinc as the pure metal? Your answer should consist <u>only</u> of a balanced chemical equation, indicating the physical state of all reactants and products. {K/U, 5}

—fin—