

Bethlehem Central High School
Delmar, New York
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How to Keep a Laboratory Notebook

In science, as in most endeavors, the facility for making accurate observations and recording them in comprehensible fashion is of central importance. The bridge between laboratory observations and subsequent understanding of them is the observer's notebook.

Individual scientists, after some experience with this process, develop methods of record keeping which are uniquely suited to their purposes and inclinations. A criterion one should apply to one's notebook keeping is that another investigator with a reasonable degree of familiarity with the field should be able to reproduce your work using your notebook as a guide.

As a starting point for your record keeping you may find the following observations and suggestions on notebooks useful.

- (1) Good notebook writers acquire their ability through thoughtful practice not inheritance.
- (2) Most useful notebooks have the following characteristics:
 - a) A full name and address of the owner inside or on the front cover.
 - b) An index of experiments near the front.
 - c) Numbered and initialed pages.
 - d) A title for each experiment with a date and a full signature of the investigator.
 - e) A consistent and clear organization or format for the recording of data and observations.
 - f) Concise and legible handwriting.
- (3) Promptness in recording experimental observations in the notebook enhances their accuracy. Ideally the notebook is a record of what you observed and not a record of what you thought you should have observed or of what you observed colored by your expectations. This difficult aspect of recording observations calls for the utmost in personal honesty. If you succeed in being honest, you will find your reasoning and understanding of what you are doing will be continually challenged as you seek the whys and wherefores underlying the observations.
- (4) The notebook should be a bound hard-covered book. Spiral or loose-leaf notebooks are not recommended, and the keeping of experimental observations on loose sheets or paper scraps is not acceptable. If corrections are necessary, these should be made with additional notes rather than erasures.
- (5) All data, calculations, notes and doodles pertinent to the experiment should be in the notebook. A convenient and highly recommended way of accomplishing this end while preserving order is to use the left hand of the pages of the notebook for calculations, tables of physical constants that pertain to the experiment and that are imported from sources such as the CRC, rough copies of chromatograms and/or any other instrument tracings, flowcharts or other schematic drawings, safety cautions, and other reminders, and to use the right hand pages for the actual record of the experiment.
- (6) The right hand side of your notebook should be concise, legible, and have clear and labeled sections. You should skip a line or two in between the sections.
 - a) A title for the experiment with the date and your full signature.
 - b) A brief statement of the objective of the experiment in a section clearly titled "Purpose".
 - c) During the experiment, you should record what you did and what you observed at the time in a section clearly titled "Procedures and Data" (you may want to keep rough notes and data on the left side and then to write a coherent record of what you did soon after, including data and observations, on the right side). Do not rewrite the procedure from the lab handout. Instead tell, in the briefest possible way, what materials you handled and what data and/or observations you obtained, so that someone else familiar with the subject could reproduce your work.

A laboratory notebook should never be confused with a formal laboratory report which is a much more polished piece of work and which contains an introductory section that includes background information about the experiment as well as a final section on the discussion and conclusion of the results.

How to Write a Formal Laboratory Report

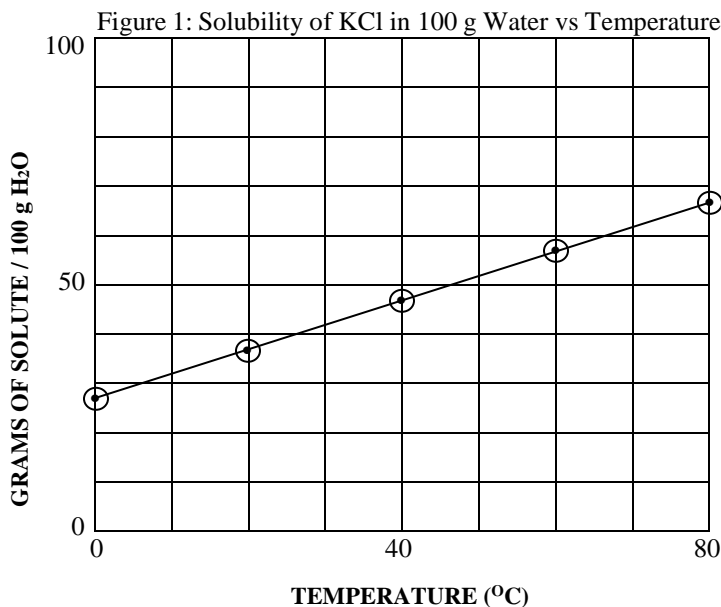
Although there are many different and specific ways to write a formal lab report, there are certain general guidelines that are common to all. For this class, use the following guidelines.

- Make your report concise. Type your report. On the first page, use the class heading and title your report.
- Arrange your report logically. Use the following format and subtitle each section with these headings:
 - Introduction: You should have a paragraph which introduces and gives background to the experiment. The last sentence of the introduction section should have a clear statement of the purpose or purposes of the experiment.
 - Materials and Methods: You may list the materials (equipment and supplies) used. You should write the procedure in paragraph form. Be concise, but include enough detail so that anyone could repeat the experiment as you performed it. Use the passive voice in this section of the report.
 - Results and Discussion: In the “Results and Discussion” section, you are to talk about your data and observations. For each and every Table and Figure you include, you must refer to it somehow and someplace in this section. In other words, “walk” the reader through your results. What do your Tables “tell” you that you want the reader to be aware of? Are there any observations or any other information that you want the reader to know about that would help explain the data in the Tables? Did you use any equations to produce results from data? What were the equations? What do your graphs “tell” you that you want the reader to be aware of? Does the plotted data suggest a trend or pattern, what is that trend or pattern, and what might the trend or pattern mean?

Data and observations should be placed in numbered and titled tables.

Temperature (°C)	Mass of KCl/100 g H ₂ O (g)
0	27
20	37
40	47
60	57
80	67

- Graphs should be inserted in this section. Graphs are numbered and titled. All graphs and any other diagrams are referred to as **figures**.



- Conclusion: In this section, you wrap up the report. Did you or did you not accomplish your goal or goals. Did you find something that was unusual or unexpected? In your conclusion, you may pose a question or problem for a future experiment. For example: *“The data suggest that the solubility of KCl in water is directly related to temperature between 0°C and 80°C. (1) Would this relationship continue from 80°C up to 100°C? (2) It would be interesting to determine whether the solubility of other salts in water also have a direct linear relationship with respect to temperature.”*

Laboratory Safety: Precautions and Practices

- 1) Before coming to the laboratory, read the through the experiment and note any safety warnings.
- 2) Complete any *“Questions for Thought”* questions in the introduction portion of the experiment.
- 3) When necessary, write out lab procedures in your notebook BEFORE your come to lab. **ANY LAB PROCEDURE THAT YOU DEVELOP MUST BE APPROVED BY YOUR TEACHER BEFORE YOU BEGIN.**
- 4) Shoes must be worn in the laboratory. No open toed shoes are allowed. Avoid wearing overly bulky or loose fitting clothes. Remove dangling jewelry. Long hair should always be pulled back.
- 5) Wear a laboratory apron and safety goggles at all times while performing lab.
- 6) Know the locations of the eye wash stations, the safety showers, the fire extinguisher, and the fire blanket, and know how to use them.
- 7) Be aware of all the exits from the laboratory.
- 8) Do not chew gum, eat food, or drink in the laboratory. Do not apply make-up or lip balm in the laboratory. Never taste any chemicals and keep your hands away from your face when handling chemicals. Wash your hands after handling chemicals.
- 9) Listen to all directions given by your teacher during the pre-lab discussion period.
- 10) Make sure know how to properly and safely use all laboratory equipment during an experiment.
- 11) Conduct only assigned experiments and only when your teacher is present.
- 12) Read and understand each step of the experiment completely before performing it. Ask for help from your teacher if you have any questions or concerns about a step in the procedure.
- 13) Keep your workspace clear of any nonessential materials.
- 14) Check all glassware for chips or cracks before using them and notify the teacher if you discover any.
- 15) When using a Bunsen burner be sure that long hair is pulled back and that loose or flowing clothing is secure. Never leave a lit Bunsen burner unattended.
- 16) Never use a Bunsen burner when a flammable liquid is being used.
- 17) Read the label on any chemicals or solutions that are being used to be sure that you have the correct chemical or solution.
- 18) Notify the teacher immediately of any spills, breakage, burns, cuts, or any other accidents.
- 19) To avoid contamination, never return chemicals or solutions to their containers after you have dispensed them.
- 20) When diluting an acid, always pour the acid slowly into water while stirring to dissipate the heat.
- 21) When heating a liquid in a test tube, turn the mouth of the test tube away from yourself and others.
- 22) Return all materials to their proper place at the end of each experiment.
- 23) Properly dispose of all unused chemicals or solutions at the end of the experiment. Your teacher will instruct you as to the proper disposal.
- 24) Leave your work area clean at the end of each experiment.
- 25) Wash your hands with soap and water at the end of each experiment.

Lab-1: Graphing in Microsoft Excel

Introduction:

In many chemical and physical changes, one characteristic changes as a result of a change in another characteristic. Data concerning changing characteristics is collected and usually placed in a table for analysis. However, it is very difficult to visualize the magnitude of the dependence of one variable upon another by looking at a table. A more descriptive way of displaying data is by the use of a graph.

Why data is graphed:

When an experiment generates many measurements, it can be difficult to make sense of or see patterns in all the information. Graphing the results of an experiment involving two variables helps to make the relationship of the variables more obvious. A graph can help analyze data and clarify patterns.

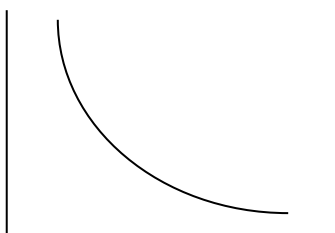
How to plot a scientific graph:

Scientific graphs show the relationship between experimental variables. One variable is called the independent variable. The independent variable is the one that the experimenter changes in the experiment. The independent variable is plotted on the horizontal, or "x", axis of the graph. The second variable, called the dependent variable, is plotted on the vertical, or "y", axis of the graph. The dependent variable is the one that responds to changes in the independent variable. The independent variable is the variable that is usually known and can usually be determined (preset) before the start of the experiment. The independent variable is INDEPENDENT of the experiment BECAUSE it can usually be determined before the start of the experiment. The dependent variable is the variable that is unknown and must be determined by running the experiment. The dependent variable is DEPENDENT on the experiment BECAUSE it can only be determined by running the experiment.

How to title a scientific graph:

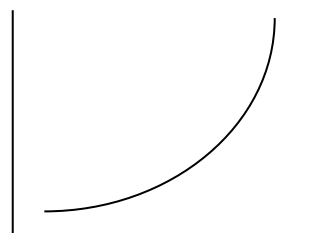
Scientific graphs are titled using the titles of the axes. When titling a graph you mention the y-axis first (the dependent variable) and then you mention the x-axis (the independent variable).

Data is often manipulated in order to produce a straight-line graph. This is called the linearization of data. Linearized data makes it easy to confirm the suspected relationship of the initial data. Linearized data with R^2 values equal to 1 are desired. Linearized data with R values equal to +1 or -1 are desired



This graph shows an inverse (indirect) relationship.
To linearize this graph try:

$$y \text{ vs } 1/x \quad \text{or} \quad y = 1/x^2 \quad \text{or} \quad \ln y = x$$



This graph shows a direct relationship.
To linearize this graph try:

$$y \text{ vs } x^2 \quad \text{or} \quad y = x^3 \quad \text{or} \quad y = \ln x$$

R is a statistical measure called the **correlation coefficient**. **R** represents the strength and direction of the relationship between the independent and dependent variables. **R** = +1 indicates the strongest direct relationship between the two variables. **R** = -1 indicates the strongest inverse relationship between the two variables. **R** = 0 indicates the no relationship between the two variables.

R² is a statistical measure called the **coefficient of determination**. **R²** represents the percentage of the variation in the dependent variable that is explained by the regression line (the linear trendline). In Data Set-1, **R²** = 1 which means that 100% of the change in pressure is explained by the change in temperature. In Data Set-2, **R²** = 0.9974 which means that 99.74% of the change in the mass of KCl dissolved in 100 g of water is explained by the change in temperature. In other words, **R²** measures the strength of the relationship between the independent and dependent variables.

Purpose: To learn to graph data in Microsoft Excel.

BELOW IS A PROCEDURE FOR GRAPHING IN MICROSOFT EXCEL 2013.

AT HOME, SOME OF YOU MAY HAVE MICROSOFT EXCEL 2010, MICROSOFT EXCEL 2007, OR EVEN MICROSOFT EXCEL 2003. EACH IS SLIGHTLY DIFFERNT FROM MICROSOFT EXCEL 2013. IF YOU HAVE MICROSOFT EXCEL 2010 OR MICROSOFT EXCEL 2003 AT HOME, THERE ARE PROCEDURES FOR THESE VERSIONS LATER IN THIS WORKSHEET. HOWEVER, IF YOU HAVE MICROSOFT EXCEL 2007 AT HOME, YOU WILL NEED TO FIGURE IT OUT BY STUDYING BOTH OF THE PROCEDURES FOR MICROSOFT EXCEL 2010 AND MICROSOFT EXCEL 2003.

Procedure-1: [*using Microsoft Excel 2013*]

- 1) Open Microsoft Excel 2013. In column A, fill in the data for the independent variable. In column B, fill in the data for the dependent variable. *In a two column data table, it is traditional that the data for the independent variable is in the left hand column and that the data for the dependent variable is in the right hand column.*
- 2) Highlight the data in the data table and then click on “Insert” at the top of the screen.
- 3) In the “Charts” section at the top of the screen, find and click on the icon for “Insert Scatter (X,Y) or Bubble Chart” and then click on the first icon under scatter.
- 4) Click on the big + on the upper right of the chart to reveal “Chart Elements” and then click on “Axes”, “Axis Titles”, “Chart Title”, and “Gridlines”. Hold mouse arrow on the word “Gridlines” until a grey arrow appears on the right. Click on arrow and then click on all four boxes.
- 5) Click on “Chart Title” and title your graph using the format: *Dependent Variable Name versus Independent Variable Name.*
- 6) Then label your axes. *Include units in parentheses after your label.* Example: Pressure (atm), Mass (g), Time (min). For temperature, write the unit as (oC). To superscript the “o”, highlight the “o” symbol and then click on “Home” at the top of the screen. Then click on small arrow under the “Font” section and then click on “superscript” and then OK.
- 7) *At this point, your graph is unfinished and is going to need some formatting.*
- 8) Lengthen your graph by dragging down the bottom until the graph appears as 8.5 by 11 inch page (all graph will be printed “in portrait”).
- 9) To manually scale the x-axis, right click on any number of the x-axis and then click on “Format Axis”. Then in the “Bounds” section, type in the minimum value and the maximum value for the x-axis.
- 10) You will have to then type in values for the major and minor increments. **You may need to adjust these major and minor units so that the entire x-axis begins and ends where you want it to! The “Minor unit” should be $\frac{1}{2}$, $\frac{1}{4}$, $\frac{1}{5}$, or $\frac{1}{10}$ of the “Major unit”.**
- 11) You may need to click “axis value” under the section “Vertical axis crosses” to set the value to be the same as the minimum value in the “Bounds” section.
- 12) Repeat for the y-axis.

If you want to connect-the-points: (if not, skip to step-14)

- 13) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. Then click on the “paint can” icon and then click on “solid line”. Under “Line”, adjust style and weight and change color to black. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

14) To plot a best-fit curve, click on one of the plotted data points and then click on “Add Trendline”. For most graphs, you are going to want a linear trendline. If so, click on “Linear” and then also click on “Display Equation on chart” and “Display R-squared value on chart” and then move the equation and R-squared value to the up and to the right of the graph title. **Your graph should now be done!**

15) *When you want to print your graphs, click on the graph near the title, then click on “File” at the top of the screen, and then click on “Print”. Under “Settings” change the “orientation” from “Landscape” to “Portrait”. Make sure your graph looks ok on the page. Then click “Print”.*

Data Set-1: Pressure versus Temperature of a Gas.

Temperature (°C)	Pressure (atm)
-19.0	0.89
13.1	1.00
47.0	1.12
95.3	1.29

*Manually scale the temperature axis from -25.0°C to 100.0°C.

*Manually scale the pressure axis from 0.8 atm to 1.4atm.

*Use a linear trendline. Notice that R^2 is 1 (ie. a perfect fit!).

Data Set-2: Mass of KCl per 100 g Water versus Temperature.

Temperature (°C)	Mass of KCl per 100 g H ₂ O (g)
10.0	33.3
20.0	35.8
30.0	37.2
40.0	39.5
50.0	42.0
60.0	44.0
70.0	46.0
80.0	48.7
90.0	50.0

*Let the program scale the temperature axis. Manually scale the mass axis from 25.0 g to 55.0 g.

*Use a linear trendline. Notice that R^2 is 0.9974 (ie. an almost perfect fit!).

Data Set-3: Volume versus Pressure of a Gas.

Pressure (mmHg)	Volume (mL)
626.	26.0
690.	23.5
720.	22.5
785	20.7
845	19.2
880.	18.5

*Manually scale the pressure axis from 600 mmHg to 900 mmHg.

*Manually scale the volume axis from 15 mL to 30 mL.

*Use a linear trendline. Notice that the R^2 is 0.9874 (ie. a poorer fit!).

Delete linear trendline.

*Now use the power trendline. Notice that the equation is $y = 16072.7138x^{-0.9984}$.

The exponent is close to -1, which means the relationship of the data is y vs $1/x$ (an inverse relationship)

LINEARIZING DATA

You are now going to linearize the data from Data Set-3. Calculate the inverse pressures for the pressures in Data Set-3. Round your values to the nearest hundred thousandths place (the nearest 0.00001) and record the values in the data table below.

Data Set-4: Volume versus Inverse Pressure of a Gas.

Inverse Pressure (mmHg ⁻¹)	Volume (mL)
	26.0
	23.5
	22.5
	20.7
	19.2
	18.5

*Manually scale the inverse pressure axis from 0.00110 mmHg⁻¹ to 0.00170 mmHg⁻¹.

*Manually scale the volume axis from 15 mL to 30 mL.

*Use linear trendline. Notice that the R² is 0.9997 (ie. an almost perfect fit!).

GRAPHING TWO DATA SETS ON THE SAME GRAPH

- 1) In column A, fill in the independent variable for the first data set. In column B, fill in the dependent variable for the first data set. In column C, fill in the independent variable for the second data set. In column D, fill in the dependent variable for the second data set.
- 2) Highlight the data for the first data set in the data table and then click on “Insert” at the top of the screen.
- 3) In the “Charts” section at the top of the screen, find and click on the icon for “Insert Scatter (X,Y) or Bubble Chart” and then click on the first icon under scatter. A graph should appear.
- 4) Click on the big + on the upper right of the new graph to reveal “Chart Elements” and then click on “Axes”, “Axis Titles”, “Chart Title”, “Gridlines” **AND** “Legend”.
- 5) While still in this section, hold mouse arrow on the word “Gridlines” until a grey arrow appears on the right. Click on arrow and then click on all four boxes.
- 6) The first data set will be referred to as Series1. Right Click on “Series1” and then click on “Select Data” and then click on “Edit”. Type in a series name for the first data set (Series1) and then click OK, and then click OK again.
- 7) Right Click on “Series1” and then click on “Select Data” and THEN click on the “Add” button. Type a name in for the second data set (Series2).
- 8) NOW click in the space right below “Name” called “X Values” and THEN highlight the data from column C. Some *gobbledygook* should appear in the “X Values” space!
- 9) Now delete the {1} from the space below “X Values” called “Y Values” and THEN highlight the data from column D. Some *gobbledygook* should appear in the “Y Values” space! NOW click OK. *At this point, both data sets should be graphed but your graph is going to need some formatting.*
- 10) Lengthen your graph by dragging down the bottom until the graph appears square.
- 11) Click on “Chart Title” and title your graph using the format: *Dependent Variable Name versus Independent Variable Name.*
- 12) Click on “Axis Title” for the x-axis and type in a label for the x-axis axis including units. For example: Pressure (atm), Mass (g), Time (min). For temperature, write the unit as (oC). To superscript the “o”, highlight the “o” symbol and then click on “Home” at the top of the screen. Then click on small arrow under the “Font” section and then click on “superscript” and then OK.

- 13) Repeat for the y-axis.
- 14) To manually scale the x-axis, right click on any number of the x-axis and then click on “Format Axis”. Then in the “Bounds” section, type in the minimum value and the maximum value for the x-axis.
- 15) You will have to then type in values for the major and minor increments. **You may need to adjust these major and minor units so that the entire x-axis begins and ends where you want it to! The “Minor unit” should be $\frac{1}{2}$, $\frac{1}{4}$, $\frac{1}{5}$, or $\frac{1}{10}$ of the “Major unit”.**
- 16) You may need to click “axis value” under the section “Vertical axis crosses” to set the value to be the same as the minimum value in the “Bounds” section.
- 17) Repeat for the y-axis.
- 18) Right click on one of the plotted data points in the first data set and then click on “Format Data Series”. A window will appear at the right. Click on the Paint Can” icon, then on “Marker”, and then click on “Marker Options”. Under “Marker Options” you can change the size and/or customize the style of your data points by clicking on “Built in”. Now you change the size and/or customize the style of you data points.
- 19) **ALL DATA POINTS SHOULD BE BLACK!** Right click on one of the plotted data points in the first data set and then click on “Format Data Series”. A window will appear at the right. Click on the Paint Can” icon, then on “Marker”. Under the “Fill” section change the color to black. Under the “Border” section change the color to black.

If you want to connect-the-points: (if not, skip to step-21)

- 20) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. At the very top, click on the pencil icon called “outline” and then change the color to black. To adjust the weight of the line click on the pencil icon called “outline” and then near the bottom you can change the weight of the line. Repeat for the other data set. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

- 21) To plot a best-fit curve, click on one of the plotted data points and then click on “Add Trendline”. A window will appear at the right. For most graphs, you are going to want a linear trendline. If so, click on “Linear”. At the very bottom of the window, click on “Display equation on chart” as well as “Display R-squared value on chart” ” and then move the equation and R-squared value to the up and to the right of the graph title. Repeat for the other data set. **Your graph should now be done!**

Data Set-5: Temp versus Time: Heating and Cooling Curves.

Melting Point Determination (Heating Curve)		Freezing Point Determination (Cooling Curve)	
Time (min)	Temp (°C)	Time (min)	Temp (°C)
0.00	25.0	0.00	65.0
0.50	30.0	1.00	56.5
1.25	40.5	2.00	54.0
2.50	51.0	3.00	53.5
3.25	53.0	4.00	52.5
4.50	54.0	5.00	49.0
6.00	60.0	6.00	40.0

***Manually scale the time axis from 0.00 min to 6.00 min and manually scale the temperature axis from 20.0°C to 70.0°C.**

***Connect the points for the data in each set and notice where the two graphs cross.**

Procedure-2: [using Microsoft Excel 2010]

- 1) Open Microsoft Excel 2010. In column A, fill in the data for the independent variable. In column B, fill in the data for the dependent variable. *In a two column data table, it is traditional that the data for the independent variable is in the left hand column and that the data for the dependent variable is in the right hand column. (See Data Set-1 on the next page.)*
- 2) Highlight the data in columns A and B. Click on “Insert” at the top of the screen. Next click on the “Scatter” icon found at the top of the screen and then click on the drop down choice that looks exactly like the “Scatter” icon. *This should produce a scatter graph of the data.*
- 3) You are now in a tab called “Chart Tools”. Click on the “layout-1” icon found under the “Design” sub-tab of the “Chart Tools” tab.
- 4) To title your graph and label your axes, click on the textbox for each and type in the graph title or axis name. For titles, use the format: *Dependent Variable Name versus Independent Variable Name*. For axes *include units in parentheses after your label*. Example: Pressure (atm), Mass (g), Time (min). For temperature, initially write the unit as (oC). To superscript the “o”, highlight the “o” and then click on the “Home” tab. Then click on the drop down button in the lower right of the “Font” section. Under “Effects”, you can click on “Superscript”. Then click OK.
- 5) To add in gridlines, click on the “Layout” tab found under the “Chart Tool” tab and then click on the “Gridlines” tab. Then click on “Primary Horizontal Gridlines” and then click on “Major and Minor Gridlines” for the x-axis. Repeat for the y-axis by clicking on “Primary Vertical Gridlines” and then click on “Major and Minor Gridlines”.
- 6) Lengthen your graph by dragging down the bottom until the graph appears square.
- 7) If you are graphing only one set of data on your graph, click on and then delete the box to the right of the graph, “Series 1”.
- 8) To manually scale the x-axis, right click on the x-axis and then click on “Format Axis”. Then click “Fixed” on ALL four options: “Minimum”, “Maximum”, “Major unit”, and “Minor unit”. Also click on “Axis Value” under “Vertical axis crosses”. THEN, manually fill in the values for these options. *The value for “Axis Value” under “Vertical axis crosses” must be the same as “Minimum”.* **The “Minor unit” should be $\frac{1}{2}$, $\frac{1}{4}$, $\frac{1}{5}$, or $\frac{1}{10}$ of the “Major unit”.** Repeat for the y-axis.

If you want to connect-the-points: (if not skip to step-10)

- 9) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. Under “Line Color”, click “Solid line” and then change the color to black. Click “Close” when done. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

- 10) To plot a best-fit curve, click on one of the plotted data points and then click on “Add Trendline”. For most graphs, you are going to want a linear trendline, which is the default trendline.
- 11) In the “Format Trendline” popup box click on “Display equation on chart” as well as “Display R-squared value on chart” and then close. Then move the equation and R-squared value to the up and to the right of the graph title. **Your graph should now be done!**
- 12) *When you want to print your graphs, click on the graph near the title and then click on print preview. Click on “set-up” and change the “orientation” from “Landscape” to “Portrait”. Then click OK. Then click “Print”.*
- 13) Now create graphs for Data Set-2, Data Set-3, and Data Set-4.

Data Set-1: Pressure versus Temperature of a Gas.

Temperature (°C)	Pressure (atm)
-19.0	0.89
13.1	1.00
47.0	1.12
95.3	1.29

*Manually scale the temperature axis from -25.0°C to 100.0°C .

*Manually scale the pressure axis from 0.8 atm to 1.4atm.

*Use a linear trendline. Notice that R^2 is 1 (ie. a perfect fit!).

Data Set-2: Mass of KCl per 100 g Water versus Temperature.

Temperature (°C)	Mass of KCl per 100 g H ₂ O (g)
10.0	33.3
20.0	35.8
30.0	37.2
40.0	39.5
50.0	42.0
60.0	44.0
70.0	46.0
80.0	48.7
90.0	50.0

*Let the program scale the temperature axis. Manually scale the mass axis from 25.0 g to 55.0 g.

*Use a linear trendline. Notice that R^2 is 0.9974 (ie. an almost perfect fit!).

Data Set-3: Volume versus Pressure of a Gas.

Pressure (mmHg)	Volume (mL)
626.	26.0
690.	23.5
720.	22.5
785	20.7
845	19.2
880.	18.5

*Manually scale the pressure axis from 600 mmHg to 900 mmHg.

*Manually scale the volume axis from 15 mL to 30 mL.

*Use a linear trendline. Notice that the R^2 is 0.9874 (ie. a poorer fit!). Delete linear trendline.

*Now use the power trendline. Notice that the equation is $y = 16072.7138x^{-0.9984}$. The exponent is close to -1, which means the relationship of the data is y vs 1/x (an inverse relationship)

LINEARIZING DATA

You are now going to linearize the data from Data Set-3. Calculate the inverse pressures for the pressures in Data Set-3. Round your values to the nearest hundred thousandths place (the nearest 0.00001) and record the values in the data table below.

Data Set-4: Volume versus Inverse Pressure of a Gas.

Inverse Pressure (mmHg ⁻¹)	Volume (mL)
	26.0
	23.5
	22.5
	20.7
	19.2
	18.5

*Manually scale the inverse pressure axis from 0.00110 mmHg⁻¹ to 0.00170 mmHg⁻¹.

*Manually scale the volume axis from 15 mL to 30 mL.

*Use linear trendline. Notice that the R^2 is 0.9997 (ie. an almost perfect fit!).

GRAPHING TWO DATA SETS ON THE SAME GRAPH (SEE DATA BELOW)**Data Set-5: Temp versus Time: Heating and Cooling Curves.**

Melting Point Determination (Heating Curve)		Freezing Point Determination (Cooling Curve)	
Time (min)	Temp (°C)	Time (min)	Temp (°C)
0.00	25.0	0.00	65.0
0.50	30.0	1.00	56.5
1.25	40.5	2.00	54.0
2.50	51.0	3.00	53.5
3.25	53.0	4.00	52.5
4.50	54.0	5.00	49.0
6.00	60.0	6.00	40.0

*Manually scale the time axis from 0.00 min to 6.00 min and manually scale the temperature axis from 20.0°C to 70.0°C.

*Connect the points for the data in each set and notice where the two graphs cross.

- FIRST DATA SET:** Fill in the data for the independent variable in column A and fill in the data for the dependent variable in column B. **SECOND DATA SET:** Fill in the data for the independent variable in column C and fill in the data for the dependent variable in column D.
- Highlight the data in columns A and B. Click on “Insert” at the top of the screen. Click on the “Scatter” icon found at the top of the screen and then click on the choice that looks exactly like the “Scatter” icon. *This should produce a scatter graph of the first data set.*
- Click on the “layout-1” icon found in “Chart Layouts” of the “Design” sub-tab of the “Chart Tools” tab. Temporary titles for the graph and axes will appear.
- To title your graph and label your axes, click on the textbox for each and type in the graph title or axis name. For titles, use the format: *Dependent Variable Name versus Independent Variable Name*. For axes *include units in parentheses after your label*. Example: Pressure (atm), Mass (g), Time (min). For temperature, initially write the unit as (oC). To superscript the “o”, highlight the “o” and then click on the “Home” tab. Then click on the drop down button in the lower right of the “Font” section. Under “Effects”, you can click on “Superscript”. Then click OK.
- To add in gridlines, click on the “Layout” tab found under the “Chart Tool” tab and then click on the “Gridlines” tab. Then click on “Primary Horizontal Gridlines” and then click on “Major and Minor Gridlines” for the x-axis. Repeat for the y-axis by clicking on “Primary Vertical Gridlines” and then click on “Major and Minor Gridlines”.
- Lengthen your graph by dragging down the bottom until the graph appears square.
- Since you will soon be adding a second data set to you graph, you will need to name each data set. Right click on area of the scatter graph near the box labeled “Series1”. Then click on “Select Data”. Rename Series1 by clicking on the “Edit” button and then rename the series in the box labeled “Series name”. Then click OK.
- Now to add the second at set, click on the “Add” button. Name the second series of data in the box labeled “Series name”. Then click on the box labeled “Series X values” and then highlight the data from column C. Some *gobbledygook* should appear in the “Series X values” box.
- In the box below labeled “Series Y Values”, keep the equal sign but delete the {1} and THEN highlight the data from column D. Some *gobbledygook* should appear in the “Series Y Values” space! Then click OK. Then click OK again. You should now have both sets of data on the graph.

- 10) Right click on the box with the two series names and then click on “Top” found under “Legend Position:” and then close. The names of the two data sets should now be under the title.
- 11) To manually scale either axis, follow the directions in step-8 on the first page.

If you want to connect-the-points: (if not skip to step-13)

- 12) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. Under “Line Color”, click “Solid line” and then change the color to black. Click “Close” when done. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

- 13) To plot a best-fit curve, right click on a plotted data points and then click on “Add Trendline”. For most graphs, you are going to want a linear trendline. If so, click on “Linear” and then click OK.
- 14) Now right click on the newly graphed trendline and then click on “Format Trendline”. Under “Options”, click on “Display equation on chart” as well as “Display R-squared value on chart” and then move the equation and R-squared value to the up and to the right of the graph title. **Your graph should now be done!**

IF YOU HAVE MICROSOFT EXCEL 2007 AT HOME, YOU CAN FIGURE IT OUT BY STUDYING THE PROCEDURES FOR BOTH MICROSOFT EXCEL 2010 AND MICROSOFT EXCEL 2003.

Procedure-3: [using Microsoft Excel 2003]

- 1) Open Microsoft Excel 2003. In column A, fill in the data for the independent variable. In column B, fill in the data for the dependent variable. *In a two column data table, it is traditional that the data for the independent variable is in the left hand column and that the data for the dependent variable is in the right hand column. (See Data Set-1 on the next page.)*
- 2) Highlight the data in the data table and then click on the “Chart Wizard” icon in the tool bar. *If the icon is not in the tool bar, it can be found in “Insert” at the top of the screen.*
- 3) On Step 1 of 4, click on XY (Scatter) and then click on next. **Skip** Step 2 of 4 by clicking on next.
- 4) On Step 3 of 4, click on “Titles”. Title your graph using the format: *Dependent Variable Name versus Independent Variable Name*. Then label your axes. *Include units in parentheses after your label*. Example: Pressure (atm), Mass (g), Time (min). For temperature, write the unit as (oC). You will be able to superscript the “o” in step 9.
- 5) Still on Step 3 of 4 under “Gridlines”, click on major and minor gridlines for both the x-axis and y-axis. The click on next. **Skip** step 4 of 4 by clicking finish.
- 6) *At this point, your graph is unfinished and is going to need some formatting.*
- 7) Lengthen your graph by dragging down the bottom until the graph appears square.
- 8) If you are graphing only one set of data on your graph, click on and then delete the box to the right of the graph, “Series 1”.
- 9) If either of your axes includes the degree symbol “o”, highlight the “o” symbol and then right click on it. Then click on “Format Axis Title”. Under “effect”, click on “superscript” and then OK.

10) To manually scale the x-axis, right click on the x-axis and then click on “Format Axis”. Then click on the “Scale” tab. Under “auto”, **uncheck** ALL five options: “Minimum”, “Maximum”, “Major unit”, “Minor unit”, and “Value (Y) axis crosses at”. THEN, manually fill in the values for these options. *The value for “Value (Y) axis crosses at” must be the same as “Minimum”.* **The “Minor unit” should be $\frac{1}{2}$, $\frac{1}{4}$, $\frac{1}{5}$, or $\frac{1}{10}$ of the “Major unit”.** Then click OK. Repeat for the y-axis. Then click OK. Repeat for the y-axis.

11) Right click somewhere on the gray shaded area of the graph avoiding a gridline and then click on “Format Plot Area”. Under “Border”, click on “Automatic” and under “Area”, click on “None”. Click OK when done.

If you want to connect-the-points: (if not skip to step-13)

12) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. Under “Line”, click “custom” and then adjust style and weight and change color to black. Click OK when done. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

13) To plot a best-fit curve, click on one of the plotted data points and then click on “Add Trendline”. For most graphs, you are going to want a linear trendline. If so, click on “Linear” and then click OK.

14) Now right click on the newly graphed trendline and then click on “Format Trendline”. Under “Options”, click on “Display equation on chart” as well as “Display R-squared value on chart” and then move the equation and R-squared value to the up and to the right of the graph title. **Your graph should now be done!**

15) *When you want to print your graphs, click on the graph near the title and then click on print preview. Click on “set-up” and change the “orientation” from “Landscape” to “Portrait”. Then click OK. Then click “Print”.*

16) Now create graphs for Data Set-2, Data Set-3, and Data Set-4.

Data Set-1: Pressure versus Temperature of a Gas.

Temperature (°C)	Pressure (atm)
-19.0	0.89
13.1	1.00
47.0	1.12
95.3	1.29

*Manually scale the temperature axis from -25.0°C to 100.0°C.

*Manually scale the pressure axis from 0.8 atm to 1.4atm.

*Use a linear trendline. Notice that R² is 1 (ie. a perfect fit!).

Data Set-2: Mass of KCl per 100 g Water versus Temperature.

Temperature (°C)	Mass of KCl per 100 g H ₂ O (g)
10.0	33.3
20.0	35.8
30.0	37.2
40.0	39.5
50.0	42.0
60.0	44.0
70.0	46.0
80.0	48.7
90.0	50.0

*Let the program scale the temperature axis. Manually scale the mass axis from 25.0 g to 55.0 g.

*Use a linear trendline. Notice that R² is 0.9974 (ie. an almost perfect fit!).

Data Set-3: Volume versus Pressure of a Gas.

Pressure (mmHg)	Volume (mL)
626.	26.0
690.	23.5
720.	22.5
785	20.7
845	19.2
880.	18.5

- *Manually scale the pressure axis from 600 mmHg to 900 mmHg.
- *Manually scale the volume axis from 15 mL to 30 mL.
- *Use a linear trendline. Notice that the R^2 is 0.9874 (ie. a poorer fit!). Delete linear trendline.
- *Now use the power trendline. Notice that the equation is $y = 16072.7138x^{-0.9984}$. The exponent is close to -1, which means the relationship of the data is y vs 1/x (an inverse relationship)

LINEARIZING DATA

You are now going to linearize the data from Data Set-3. Calculate the inverse pressures for the pressures in Data Set-3. Round your values to the nearest hundred thousandths place (the nearest 0.00001) and record the values in the data table below.

Data Set-4: Volume versus Inverse Pressure of a Gas.

Inverse Pressure (mmHg ⁻¹)	Volume (mL)
	26.0
	23.5
	22.5
	20.7
	19.2
	18.5

- *Manually scale the inverse pressure axis from 0.00110 mmHg⁻¹ to 0.00170 mmHg⁻¹.
- *Manually scale the volume axis from 15 mL to 30 mL.
- *Use linear trendline. Notice that the R^2 is 0.9997 (ie. an almost perfect fit!).

GRAPHING TWO DATA SETS ON THE SAME GRAPH (SEE DATA BELOW)

Data Set-5: Temp versus Time: Heating and Cooling Curves.

Melting Point Determination (Heating Curve)		Freezing Point Determination (Cooling Curve)	
Time (min)	Temp (°C)	Time (min)	Temp (°C)
0.00	25.0	0.00	65.0
0.50	30.0	1.00	56.5
1.25	40.5	2.00	54.0
2.50	51.0	3.00	53.5
3.25	53.0	4.00	52.5
4.50	54.0	5.00	49.0
6.00	60.0	6.00	40.0

- *Manually scale the time axis from 0.00 min to 6.00 min and manually scale the temperature axis from 20.0°C to 70.0°C.
- *Connect the points for the data in each set and notice where the two graphs cross.

- 1) In column A, fill in the independent variable for the first data set. In column B, fill in the dependent variable for the first data set. In column C, fill in the independent variable for the second data set. In column D, fill in the dependent variable for the second data set.
- 2) Highlight the data in columns A and B in the data table and then click on the chart wizard icon in the tool bar. *If the icon is not in the tool bar, it can be found in "Insert" at the top of the screen.*

- 3) On Step 1 of 4, click on XY (Scatter) and then click on next.
- 4) On Step 2 of 4, click on “Series”. The first data set will be referred to as Series1. Type a name in for the first data set (Series1).
- 5) THEN hit the “Add” button under “Series”. Type a name in for the second data set (Series2). NOW, click in the space right below “Name” called “X Values” and THEN highlight the data from column C. Some *gobbledygook* should appear in the “X Values” space!
- 6) Now delete the {1} from the space below “X Values called “Y Values” and THEN highlight the data from column D. Some *gobbledygook* should appear in the “Y Values” space! The click on next.
- 7) On Step 3 of 4, click on “Titles”. Title your graph using the format: *Dependent Variable Name versus Independent Variable Name*. Then label your axes. *Include units in parentheses after your label*. Example: Pressure (atm), Mass (g), Time (min). For temperature, write the unit as (oC). You will be able to superscript the “o” in step 13.
- 8) Still on Step 3 of 4 under “Gridlines”, click on major and minor gridlines for both the x-axis and y-axis. The click on next. Skip step 4 of 4 by clicking finish.
- 9) *At this point, your graph is unfinished and is going to need some formatting.*
- 10) Lengthen your graph by dragging down the bottom until the graph appears square.
- 11) Right click on the box with the series names and then click on “Format Legend”. Next click on the “Placement” tab and then click on “Top”. The box should now be under the title.
- 12) To manually scale the x-axis, right click on the x-axis and then click on “Format Axis”. Then click on the “Scale” tab. Under “auto”, **unclick** ALL five options: “Minimum”, “Maximum”, “Major unit”, “Minor unit”, and “Value (Y) axis crosses at”. THEN, manually fill in the values for these options. *The value for “Value (Y) axis crosses at” must be the same as “Minimum”*. **The “Minor unit” should be $\frac{1}{2}$, $\frac{1}{4}$, $\frac{1}{5}$, or $\frac{1}{10}$ of the “Major unit”**. Then click OK. Repeat for the y-axis. Then click OK. Repeat for the y-axis.
- 13) Right click on one of the plotted data points in the first data set and then click on “Format Data Series”. You should be in the tab “Patterns”. Under “Marker” you can change the size and/or customize the style of you data point. **ALL DATA POINTS SHOULD BE BLACK!** Change both foreground and background to black. Click OK when done. Repeat with the second data set. Click OK when done.
- 14) Right click somewhere on the gray shaded area of the graph avoiding a gridline and the click on “Format Plot Area”. Under “Border”, click on “Automatic” and under “Area”, click on “None”. Click OK when done.

If you want to connect-the-points: (if not skip to step-16)

- 15) To connect-the-points, right click on one of the plotted data points from a data set and then click on “Format Data Series”. Under “Line”, click “custom” and then adjust style and weight and change color to black. Click OK when done. Click OK when done. **Your graph should now be done!**

If you want a best-fit curve (do not connect the points):

- 16) To plot a best-fit curve, click on one of the plotted data points and then click on “Add Trendline”. For most graphs, you are going to want a linear trendline. If so, click on “Linear” and then click OK.
- 17) Now right click on the newly graphed trendline and then click on “Format Trendline”. Under “Options”, click on “Display equation on chart” as well as “Display R-squared value on chart” and then move the equation and R-squared value to the up and to the right of the graph title. **Your graph should now be done!**

Lab-2: Accuracy and Precision

Introduction:

Since all measurements contain an estimated digit, all measurements some contain uncertainty (error). Scientists try to limit the uncertainty (error) as much as possible but they cannot eliminate it. There are three main reasons for uncertainty in measurements.

a) Reasons for Uncertainty (Error):

i) Instrumental Error:

All measuring instruments have error. The more sensitive and precise the instrument is, the lower the amount of error will be.

ii) Observer Error:

An instrument is only as good as the person using it! Persons who have more experience and who take more precautions will generally record measurements with less error.

iii) Procedural Error:

Measurements can have error due to faulty experimental procedure.

b) How to Measure Error:

In an experiment, it is important to be able to state the level of confidence of one's data. This is done through statistical analysis. Often this analysis is quite involved. However, for this course we will use two simple tests for error.

i) A test for Accuracy:

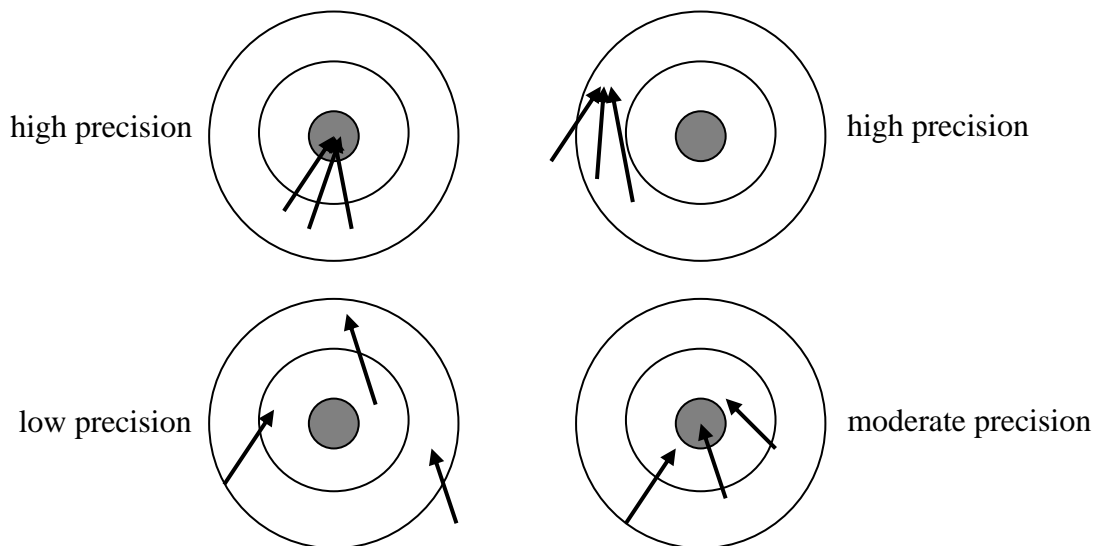
Accuracy measures how close a measured value is to the accepted value. A **test for accuracy** uses the equation for Percent (%) Error

$$\% \text{ Error} = \frac{\text{Measured Value} - \text{Accepted Value}}{\text{Accepted Value}} \times 100$$

% Error can be positive or negative! In this course, a % Error of less than |5%| will represent high accuracy. A % Error between |5%| and |10%| will represent moderate accuracy. A % Error greater than |10%| will represent low accuracy.

ii) A test for Precision:

Precision measures how close several measured trials are to one another. The calculations to determine precision are complicated and will not be used in this course. Instead, we will use the "eyeball test" – do the data appear closely bunched?



Graduated cylinders are not always the most accurate instruments. The graduated markings are printed on by machine and can have an error range of $\pm 5\%$ or more. However, since you will simply want to use the graduated cylinder to measure “change in volume”, the absolute accuracy of the markings themselves is not a problem as long as the increments of the markings are accurate (or accurate enough). It is known, that the laboratory burets accurately dispense a known amount of liquid. In this investigation, you will use a buret to test the accuracy of a graduated cylinder.

Good technique is needed when measuring volume. Be sure to read the bottom of the meniscus in both the buret and the graduated cylinder at eye level.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Obtain a 50-mL buret and stand. Obtain a 100-mL graduated cylinder.
- 2) Using a 100-mL beaker, fill the 50-mL buret above the zero mark. Let out some of the water in order to “charge” the value and tip of the buret. Make sure your buret reads zero.
- 3) Dispense exactly 10-mL of water to the graduated cylinder from the buret. Read and record graduated cylinder.
- 4) Dispense exactly a second 10-mL of water to the graduated cylinder from the buret (the buret should now read 20-mL). Read and record graduated cylinder.
- 5) Repeat until all 40-mL of water has been dispensed from the buret into the graduated cylinder.
- 6) Repeat the entire procedure twice more.

Trial 1		Trial 1		Trial 1		Average	
Volume of Water Dispensed by the buret (mL)	Volume of Water in the Graduated Cylinder (mL)	Volume of Water Dispensed by the buret (mL)	Volume of Water in the Graduated Cylinder (mL)	Volume of Water Dispensed by the Buret (mL)	Volume of Water in the Graduated Cylinder (mL)	Volume of Water Dispensed by the Buret (mL)	Volume of Water in the Graduated Cylinder (mL)
10.0		10.0		10.0		10.0	
20.0		20.0		20.0		20.0	
30.0		30.0		30.0		30.0	
40.0		40.0		40.0		40.0	

In this lab, you will be making a graph of *Average Volume of Water in the Graduated Cylinder versus Volume of Water Dispensed by the Buret*. A graph’s title is always in the form of *Y-axis versus X-axis*. Thus when you create your graph, the graduated cylinder volume will be on the Y-axis and the buret volume will be on the X-axis. Include the linear fit equation as well as the **coefficient of determination (R^2)** value. Print your graph **in portrait!**

Lab-4: The Empirical Formula of a Compound

Introduction:

A compound is a substance that is composed of two or more elements bonded in a definite mole ratio (proportion). A specific compound will always contain the same elements in the same definite ratio (proportion). This is known as the **Law of Constant Composition**. By mass, the definite ratio is given as percentages. Water is a compound of hydrogen and oxygen with the definite ratio of 2 moles hydrogen to 1 mole oxygen. By mass, the ratio is 11.2% hydrogen to 89.8% oxygen for water. Hydrogen peroxide is a compound of hydrogen and oxygen with the definite ratio of 1 mole hydrogen to 1 mole oxygen. By mass, the ratio is 5.9% hydrogen to 94.1% oxygen for hydrogen peroxide.

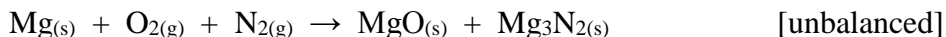
The Law of Constant Composition, among others, convinced John Dalton of the atomic nature of matter and led him to outline his atomic theory in these four postulates:

- (1) Elements are composed of tiny particles called atoms.
- (2) Atoms of the same element are exactly alike and have the same set of properties, but the atoms of one element are different from the atoms of all other elements.
- (3) Atoms of an element cannot be created, destroyed, or changed into atoms of another element.
- (4) Compounds are formed when atoms of two or more elements are chemically combined in a fixed small whole number ratio.

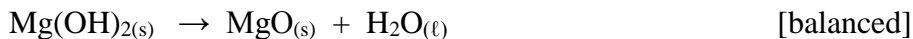
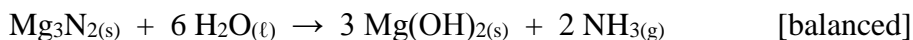
A compound's empirical formula is the formula showing the simplest (most reduced) ratio of the elements in the compound. For water (H₂O), the empirical formula is H₂O. For hydrogen peroxide (H₂O₂), the empirical formula is HO. For glucose (C₆H₁₂O₆), the empirical formula is (CH₂O). If the masses of the elements (or the percent by mass values of the elements) in a compound are known, the empirical formula for the compound can be easily calculated.

In 1778, Lavoisier concluded that combustion was a reaction of oxygen in the air with a sample of matter. He realized that the substance burned gained mass, the same mass lost from the surrounding air. A great deal of chemical knowledge has been amassed by using simple combustion experiments conducted with crucibles, burners, and balances.

Magnesium reacts vigorously when heated in the presence of air. The magnesium-oxygen reaction is energetic enough to allow some of the magnesium to react with nitrogen from the air. Although there is a higher percentage of N₂ gas in the air than O₂ gas, the O₂ gas is more reactive and magnesium oxide forms in greater amount than magnesium nitride.



The magnesium nitride can be easily converted to magnesium oxide by adding water, which converts it to magnesium hydroxide and ammonia gas. If this mixture is heated to dryness, the ammonia gas is driven off and the magnesium hydroxide is converted to magnesium oxide.



A compound's empirical formula may or may not be its actual chemical formula. A compound's actual chemical formula is called its molecular formula. Additional information besides percent composition is needed in order to determine a compound's molecular formula.

Purpose: To determine the empirical formula for a compound of magnesium and oxygen.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Obtain a preheated porcelain crucible and cover. Mass the porcelain crucible and cover.
_____g
- 2) Obtain a sample of magnesium (≈ 6 in. cut into small pieces) and place it in the crucible.
- 3) Record several observations of the pure magnesium.

- 4) Mass the crucible, cover, and magnesium sample. _____g
- 5) Determine the mass of the magnesium sample. _____g
- 6) Nestle the crucible securely in the clay triangle. Set the lid *slightly* off-center on the crucible to allow air to enter but to prevent the magnesium oxide to escape.
- 7) Holding the Bunsen burner, heat the bottom of the crucible gently for about 1 minute and then set the burner down and heat the crucible strongly for ten minutes. ***If you see white smoke, you are heating the crucible too strongly.***
- 8) Turn off the Bunsen burner and allow the crucible, cover, and contents to cool ***on the ring stand*** for 5 minutes. *Obtain a 50 mL beaker with about 10 mL of distilled water and a dropper pipette.*
- 9) Carefully remove the cover and dropper in just enough distilled water to cover the solid. Replace the cover and heat crucible and contents ***VERY gently*** until no sound can be heard and then and only then heat the crucible and contents strongly for 5 additional minutes. **The contents must be dry!**
- 10) Carefully remove the crucible, cover, and contents to the wire ceramic pad to cool (≈ 5 minutes).
- 11) Mass the crucible and contents. _____g
- 12) Record several observations of the contents of the crucible.

- 13) Determine the mass of the magnesium oxide sample. _____g
- 14) Determine the mass of the oxygen in the magnesium oxide sample. _____g
- 15) Based on your values for the masses of the magnesium and oxygen in the magnesium oxide sample, determine the empirical formula of the magnesium oxide. **SHOW WORK BELOW.**

The empirical formula is _____

Lab-5: Paper Chromatography

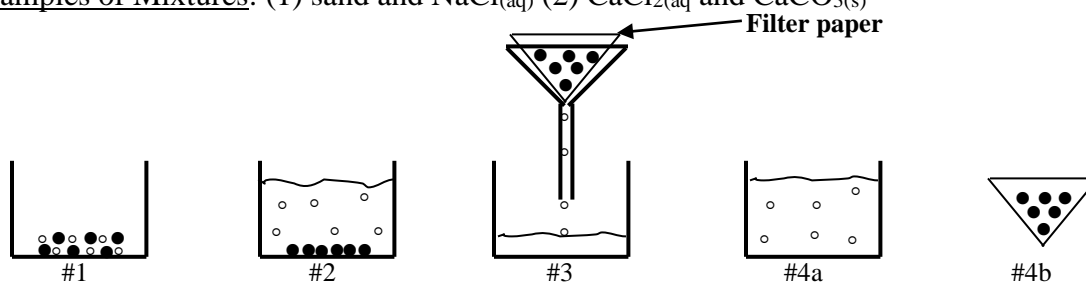
Introduction:

Mixtures can be separated into their individual component parts by various methods. Differences in physical properties such as mass, particle size, molecular polarity, boiling and freezing points, and solubility determine the best method to be used when separating a particular mixture.

1) Filtration:

Filtration is used to separate a water soluble solid from a water insoluble solid (#1). A mixture of the two solids is placed into a beaker and water is added. The water soluble solid dissolves into the solvent (#2). A piece of filter paper is placed in a funnel and the mixture is poured through the funnel where the filter traps the water insoluble solid while allowing the dissolved water soluble solid to pass through with the water (#3). The substances are now separated (4a & 4b)

Examples of Mixtures: (1) sand and $\text{NaCl}_{(aq)}$ (2) $\text{CaCl}_{2(aq)}$ and $\text{CaCO}_{3(s)}$

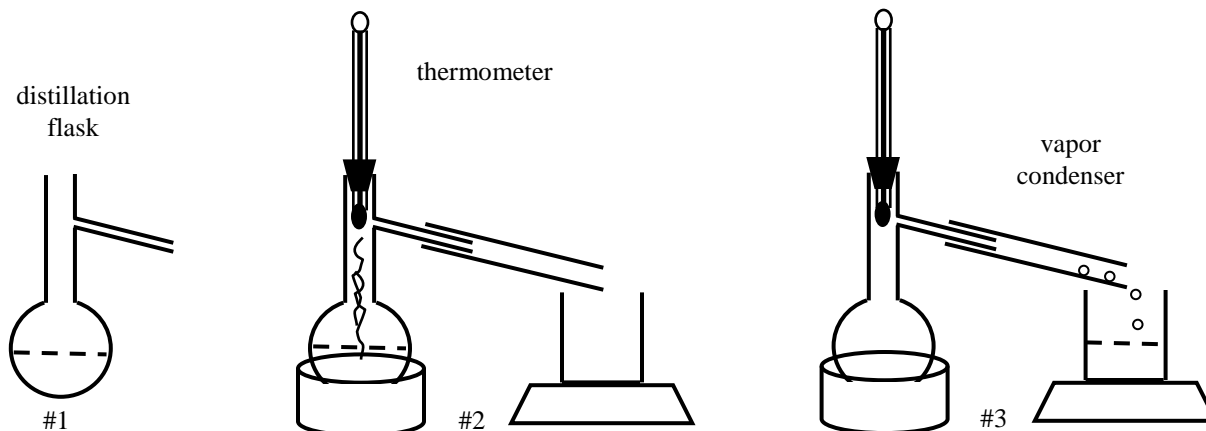


2) Distillation:

Distillation is used to separate a soluble solid solute from a liquid solvent (#1). The solution is boiled. The liquid solvent evaporates leaving the now solid solute behind (#2). The solvent's vapor is collected, cooled and condensed back to a now pure liquid (#3). ***This is how distilled water is made!***

Distillation can also be used to separate a soluble liquid solute from a liquid solvent (#1). The solution is boiled. The liquid with the lower boiling point evaporates leaving the liquid with the higher boiling point behind (#2). The vapor of the lower boiling point liquid is collected, cooled and condensed back to a now pure liquid (#3). This liquid-liquid distillation only works if the two liquids' boiling points are different by at least 5°C . The thermometer is used to know when the first liquid has been separated. As the first liquid is being separated the temperature remains constant. When there is a sudden rise in temperature, this means that the second liquid is now evaporating and it is time to switch the collection beaker. ***This is how distilled liquors are made!***

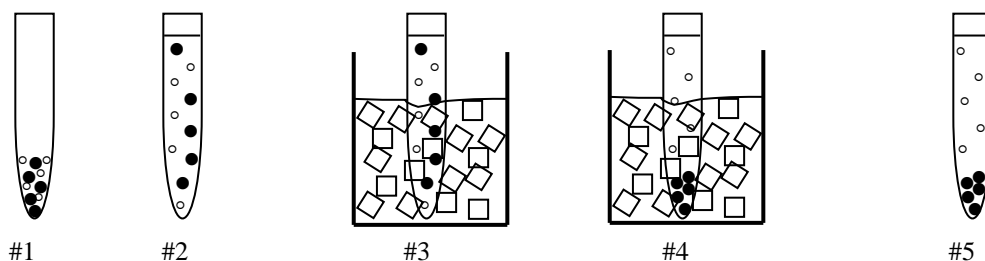
Examples of Mixtures: (1) salt water, (2) alcohol and water [$\text{C}_2\text{H}_5\text{OH}_{(aq)}$]



3) **Crystallization:**

Crystallization is used to separate two solids whose solubility in a cold solvent differs greatly (#1). Both solids are dissolved together in an appropriate liquid solvent (#2). The solution is cooled below the saturation point of the solid of lower solubility (#3). Crystals of the solid of lower solubility begin to form (#4). The mixture can then be filtered (#5).

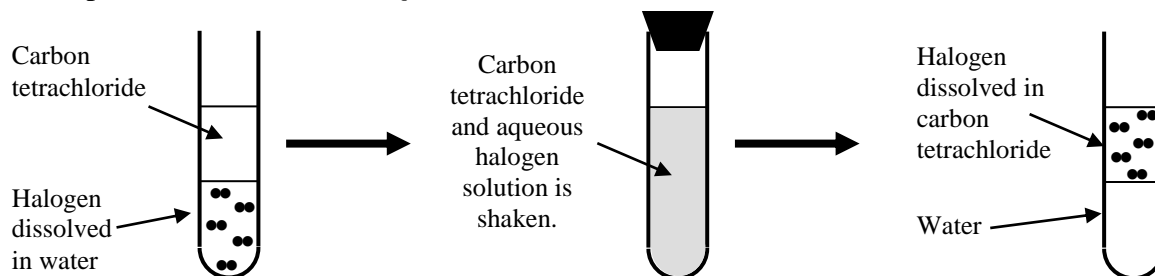
Examples of Mixtures: (1) $KI_{(s)}$ & $KNO_{3(s)}$



4) **Extraction:**

Extraction is a separation method in which a component or a mixture of components is attracted differently to a pair of immiscible solvents. In a typical procedure, a component or a mixture of components is blended with a solvent that extracts (dissolves) soluble or semi-soluble components. This extract is then added to a second solvent that is immiscible with the first solvent. Any component or components in the first extract that is more soluble in the second solvent (due to differences in molecular polarity) is extracted (dissolved) into the second solvent.

Examples of Mixtures: (1) $I_{2(aq)}$ & $H_2O_{(l)}$

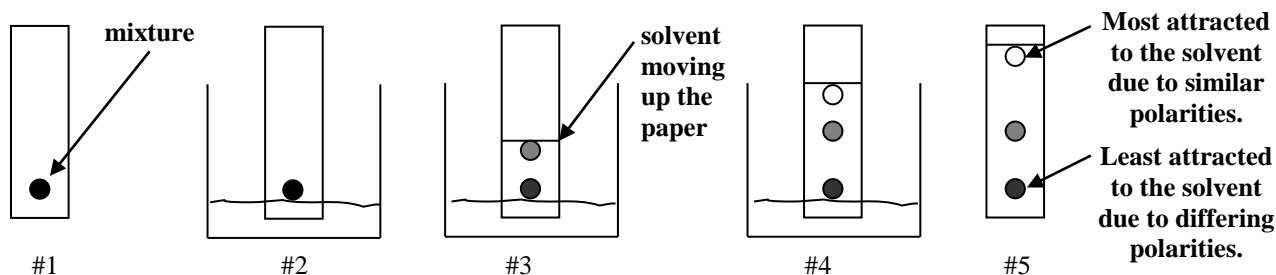


5) **Chromatography:**

Chromatography is a separation method in which different components of a mixture are attracted differently to a solid surface (*the stationary phase*) or to a liquid or a gas solvent (*the mobile phase*). The component of the mixture that has the most similar polarity with the solvent will move right along with the solvent and the component of the mixture that has the least similar polarity with the solvent will be left behind on the solid surface.

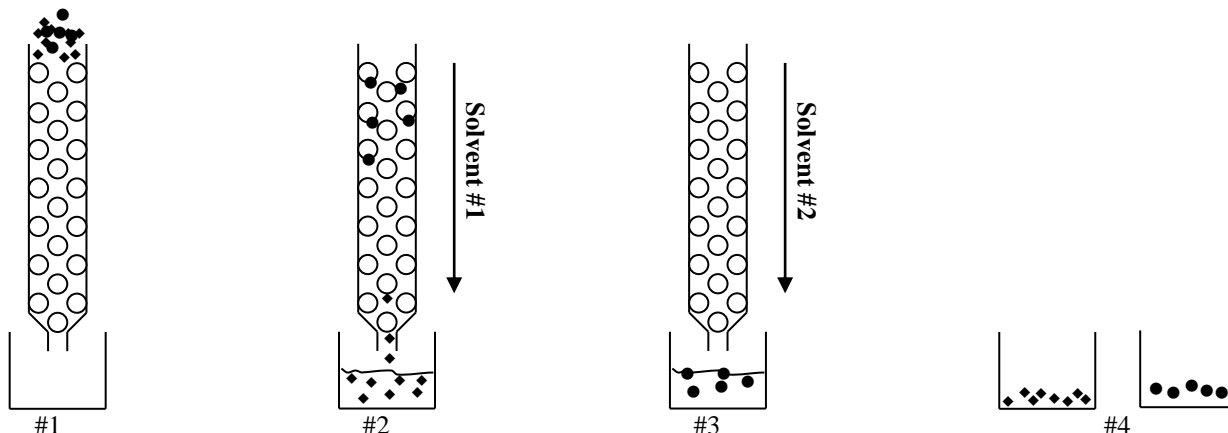
i) **Paper Chromatography:**

A mixture is applied to the bottom of a piece of chromatography paper (#1). The bottom of the paper is suspended into the solvent (#2). The solvent begins to move up the paper and the mixture components begins to separate based on their different polarities (#3, #4, & #5)



ii) Column Chromatography:

A mixture is placed at the top of the column which is packed with a special material (#1). A first solvent is sent through the column which readily carries one of the substances in the mixture (but not the other) down the column due to similarities in their polarities (#2). A second solvent is then sent through the column which then readily carries the second substance down the column due to similarities in their polarities (#3). The solvent is allowed to evaporate from the two separated fractions to reveal crystals of the separated pure substances (#4).

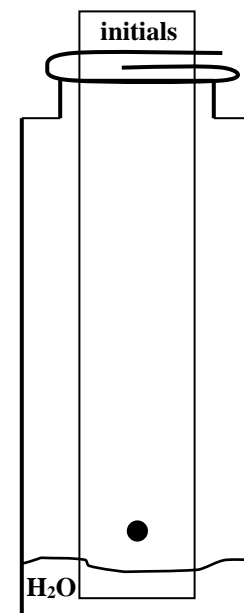


Purpose:

To separate a mixture of substances using paper chromatography.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Each of you will work individually and produce your own chromatogram.
- 2) Obtain one graduated cylinder, one paper clip, one small bottle, one metric ruler, and one piece of chromatography paper.
- 3) with a dot of water-soluble green ink.
- 4) ***Put your three initials on the top of your paper*** (on the opposite end from the green ink dot).
- 5) Place the paper clip across the paper near the end opposite the green ink dot. Insert the paper, dot end first, into the bottle. Adjust the paper clip so the paper hangs down into the bottle **just** off the bottom - without touching the bottom or the sides of the bottle.
- 6) Remove the paper (be careful NOT to move the paper clip). Carefully pour 10 mL of water into the bottle. Reinsert the paper into the bottle. The very bottom of the paper should now be in the water **but not the green ink dot itself!**
- 7) **Remove your goggles and apron and return to your classroom seat.**
- 8) During this time, begin reading the article from “Scientific American” written by William H. Stein and Stanford Moore in 1951 entitled “Chromatography”. Both Stein and Moore received the Nobel Prize in Chemistry in 1972.
- 9) **When told, return to your lab station and put on your goggles and apron.**



- 10) Carefully remove your chromatogram and lay it on a piece of paper towel to dry. **CAREFULLY REMOVE THE PAPER CLIP.**
- 11) Using a pencil, draw a line across the paper to show the water line.
- 12) Using a metric ruler, measure the distance ...

... the **water** traveled from the initial dot: _____ cm.

... the **yellow** traveled from the initial dot: _____ cm.

... the **blue** traveled from the initial dot: _____ cm.

- 13) Chromatography can be used to determine an unknown by determining the constant called R_f .

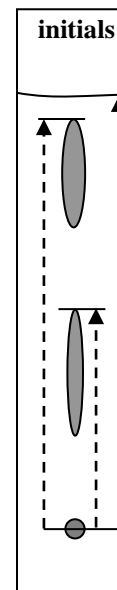
$$R_f = \frac{\text{the distance traveled by a substance}}{\text{the distance traveled by the solvent}}$$

This ratio is constant for pure substances, thus it can be used to identify an unknown.

What is the R_f value for the yellow ink in your chromatogram? _____

What is the R_f value for the blue ink in your chromatogram? _____

- 14) Cleanup and return all materials.
- 15) You will need to finish reading the article "Chromatography".



Lab-6: Quantitative Analysis of a Chemical Reaction

Introduction:

The Law of Conservation of Matter states that “matter can be neither created nor destroyed.” Thus the total mass of the reactants must be equal to the total mass of the products. However, this is only true for physical and chemical reactions.

In a balanced equation for a chemical reaction, the coefficients of the substances represent the relative number of moles of each substance in the reaction. In the laboratory, the determination of the masses of substances used or produced can be used to find the number of moles of each and, if successful, the coefficients for the substances in the equation.

Purpose:

Write a balanced equation for the reaction that occurs between solid iron metal and an aqueous solution of copper (II) chloride dihydrate.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Clean the surface of the iron sample with emery cloth to remove any rust or dirt.
- 2) Determine the mass of a clean watch glass to the nearest 0.0001 g. _____ g
- 3) Determine the mass of the iron sample and watch glass to the nearest 0.0001 g.
_____ g
- 4) Determine the mass of a clean 250-mL beaker to the nearest 0.0001 g. _____ g
- 5) Pour about 40.0 mL of 0.300 M copper salt solution into a 50-mL the beaker. Place the clean iron sample into the clean massed 250-mL beaker and add the 40-mL of 0.300 M copper salt solution.
- 6) Allow the beaker to stand undisturbed for about 20 minutes. Every few minutes carefully stir the beaker. (*Complete the “Questions for Thought” on the back.*)
- 7) After about 20 minutes, record any evidence of a chemical reaction.

- 8) *Carefully* remove the iron sample from the solution and hold it above the solution. *Carefully* rinse any remaining solution or product off the iron sample and into the beaker, using steady jet of distilled water from a wash bottle. Make sure that the iron sample is completely rinsed.
- 9) Place the iron sample onto the watch glass and put it into the drying oven for 15 minutes to dry. **Go to step 12.**
- 10) Determine the mass of the iron sample to the nearest 0.0001 g. _____ g
(*The sample may be heavier because it has some of the solid reaction product clinging to it.*)
- 11) Completely wash and dry the iron sample. Then again place the iron sample on the watch glass and put it into a drying oven to completely dry. Determine (later) the mass of the iron sample to the nearest 0.0001 g. _____ g (*The sample may now be lighter than before because some of the solid reaction product was cleaned off of it.*)
- 12) Carefully and slowly decant the liquid portion of the contents of the beaker in order to leave behind the solid reaction product. Carefully rinse the remaining solid reaction product with distilled water using the wash bottle and decant once more. Repeat twice more or until all remaining reactant solution or product solution has been removed. Place the beaker in a spot to dry as indicated.
- 13) After the solid reaction product has been completely dried in the drying oven, determine (later) the mass of the beaker and solid reaction product. _____ g
- 14) Observe the solid reaction product. _____

Questions for Thought:

- 1) Write the chemical formula for solid copper (II) chloride dihydrate.
- 2) When pure solid iron reacts with aqueous copper (II) chloride, what possible aqueous iron product(s) may form? Write the chemical formula(s).
- 3) When pure solid iron reacts with aqueous copper (II) chloride, what solid product(s) may form? Write the chemical formula(s).
- 4) What evidence might you observe to support your answer to question-3?
- 5) Using your answer(s) from question-2 and question-3, write a balanced equation(s) to represent the chemical reaction(s) that may take place when pure solid iron reacts with aqueous copper (II) chloride.
- 6) Classify the chemical reaction(s) you wrote in question-5
- 7) Based on the Activity Series of Metals in Aqueous Solution found in your textbook (Table 4.5), would you expect the equation(s) you wrote in question-5 to be spontaneous. Explain.

Lab-7: Qualitative Analysis of a Solution

Introduction:

The process of determining the composition of a sample of matter by conduction chemical tests is called qualitative analysis. An Unknown solution can be subjected to a series of chemical and physical tests and the results are logically analyzed in order to determine the composition of the solution. The tests can include: flame testing, conductivity testing (electrolyte or nonelectrolyte), and reagent testing. Reagent testing uses a number of known solutions (reagents) to determine the identity of an unknown solution.

In reagent testing, one can look for the formation of a precipitate. Knowledge of the solubility rules and/or an analysis scheme helps to detect the presences of various ions in solution. Solutions of compounds containing Ag^{1+} , Pb^{2+} , or Hg_2^{2+} ions can be detected by adding HCl and observing for the precipitation of white crystals.

In reagent testing, one can look for the evolution of a gas. Solutions of compounds containing NH_4^+ ions can be detected by adding a strong base and smelling for the presence of $\text{NH}_3(\text{g})$. Solutions of compounds containing CO_3^{2-} ions can be detected by adding a strong acid and observing for the presence of $\text{CO}_2(\text{g})$.

In reagent testing, one can look for a specific color change due to the formation of a complex ion. Solutions of compounds containing Cu^{2+} ion turn from light blue to dark blue when $\text{NH}_3(\text{aq})$ is added. Solutions of compounds containing Fe^{3+} ions turns from orange to dark red when $\text{SCN}^-(\text{aq})$ is added.

Your knowledge of the solubility rules and the flame test colors is essential for qualitative analysis.

Ions that form Soluble Compounds	Exceptions	Ions that form Insoluble Compounds	Exceptions
Group-1 cations & NH_4^{1+}		CO_3^{2-}	Group-1 cations & NH_4^{1+}
NO_3^{1-}		S^{2-}	Group-1 cations & NH_4^{1+}
$\text{C}_2\text{H}_3\text{O}_2^{1-}$ [$\text{CH}_3\text{COO}^{1-}$]		PO_4^{3-}	Group-1 cations & NH_4^{1+}
HCO_3^{1-}		CrO_4^{2-}	Group-1 cations & NH_4^{1+} , & Mg^{2+} , Ca^{2+}
ClO_3^{1-} & ClO_4^{1-}			
Cl^{1-} , Br^{1-} , I^{1-}	Ag^{1+} , Pb^{2+} , Hg_2^{2+}	OH^{1-}	Group-1 cations & NH_4^{1+} , & Ca^{2+} , Sr^{2+} , Ba^{2+}
SO_4^{2-}	Ag^{1+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+}		

Flame Test Colors for Some Alkali Metal Ions & Alkaline Earth Metal Ions						
Metal Ion	Na^{1+}	Li^{1+}	K^{1+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
Color	Dark Yellow	Red (Crimson)	Violet (Lilac)	Orange-Red	Red-Scarlet	Yellowish Green

Purpose:

In this experiment, you will determine the identity of two sets of unknown solutions. The dropper bottles in the first set of unknown solutions are labeled Q, R, and S. Each bottle contains one of the following solutions: 1.0M $\text{Pb}(\text{NO}_3)_2$, 1.0M NaCl , and 1.0M K_2CO_3 . The dropper bottles in the second set of unknown solutions are labeled X and Y. Each bottle contains one of the following solutions: 1.0M AgNO_3 and 1.0M BaCl_2 .

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

Part I: Caution, AgNO_{3(aq)} will stain your skin black. Wash your hands frequently.

- 1) Combine a drop of solution Q and a drop of solution X. Record your observation in Data Table 1.
- 2) Combine a drop of solution Q and a drop of solution Y. Record your observation in Data Table 1.
- 3) Based on your observations, determine the identity if solution Q (_____).
- 4) Combine a drop of solution Q and a drop of solution R. Record your observation in Data Table 1.
- 5) Combine a drop of solution Q and a drop of solution S. Record your observation in Data Table 1.
- 6) Based on your observations, determine the identity if solution R (_____) and solution S (_____).

Data Table 1			
Combination of Solutions Q and X	Combination of Solutions Q and Y	Combination of Solutions Q and R	Combination of Solutions Q and S

Part II:

- 7) Using your conclusions from Part I, determine and write a procedure to identify solution X and solution Y by using one or more of the identified solutions from Part I. Be sure to describe the observations that would allow you to distinguish between the solutions X and Y.

- 8) Perform your procedure and record your observations in Data Table 2, which you will make below (neatly... “Mr. Reed neatly”).

- 9) Solution X is _____ and solution Y is _____.

Part III:

10) Obtain a nichrome loop and 50 mL of distilled water.

11) Dip the loop into the solution of Q. Place the wet loop into the Bunsen burner flame for a few seconds. Repeat as many times as necessary to observe flame color. Record color below. Rinse the loop very well in the distilled water and repeat the process with the other solutions.

Q: _____ R: _____ S: _____

X: _____ Y: _____

Lab-8: Heat of Neutralization

Introduction:

Every chemical change is accompanied by a change in energy usually in the form of heat. If heat is evolved, the reaction is exothermic, and if heat is absorbed, the reaction is endothermic. The energy change of a reaction that occurs at constant pressure is called the heat of reaction or the enthalpy of reaction (ΔH_R). This quantity of heat is measured experimentally by allowing the reaction to occur in a calorimeter. In this experiment, you will determine the heat of neutralization when an acid and a base react to form 1 mole of water. In a perfect calorimeter, heat is exchanged only between the reaction and the calorimeter's water. Technically, some heat may be absorbed the calorimeter. Alas all calorimeters exchange some heat with its environment. This amount of heat is called the calorimeter's heat capacity (the amount of heat required to raise its temperature 1°C). We are going to "pretend" that our calorimeter is a perfect calorimeter.

Purpose:

You will determine the **molar heat of neutralization** of 1.00 M HCl and 1.00 M NaOH in kilojoules.

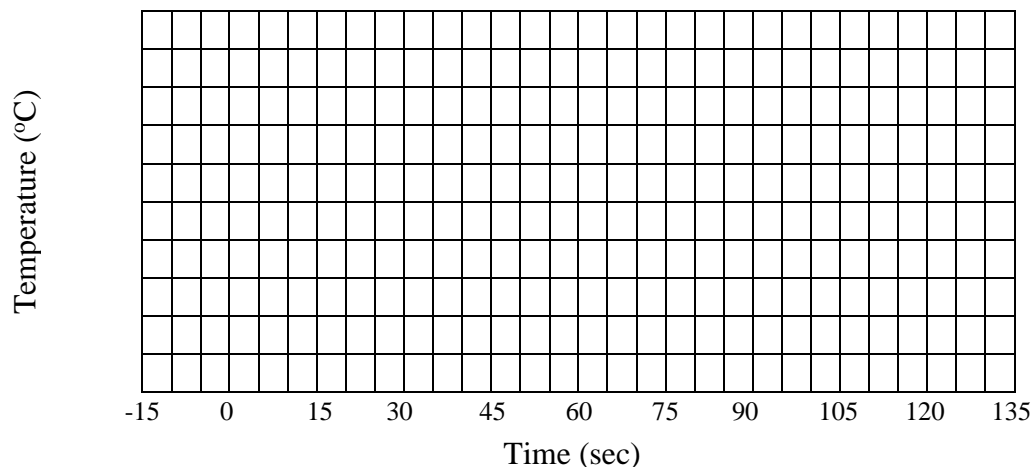
Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Carefully measure 50.0 mL of 1.0 M NaOH in a 100-ml graduated cylinder and pour it into the calorimeter. Leaving the thermometer out, place the top on the calorimeter.
- 2) Rinse the graduated cylinder with some tap water. Then rinse the graduated cylinder with some - about 10 mL - of the 1.0 M HCl. Completely dry the graduated cylinder (*as best as possible*) with a paper towel drying wand. Carefully measure 50.0 mL of the 1.0 M HCl. Leaving the HCl in the graduated cylinder measure the initial temperature of the acid _____°C.
- 3) Rinse the thermometer with tap water and dry. Insert the thermometer into the calorimeter's top and measure the initial temperature of the base _____°C. **(The two initial temperatures should not differ by more than 0.5°C. If they do ask for assistance)**
- 4) **Reread entire step before commencing.** Carefully add all of the acid to the base in the calorimeter being careful NOT to splash any acid on the upper sides of the calorimeter. Stir the mixture gently with the thermometer observing the temperature every 15-seconds for the next 2 minutes.

_____ , _____ , _____ , _____ , _____ , _____ , _____ , _____ , _____ .
15 s 30 s 45 s 60 s 75 s 90 s 105 s 120 s 135 s

- 5) Construct a Temperature. vs Time plot to extrapolate the instantaneous ΔT value. The temperature at -15 seconds and at 0 seconds will be the initial temperature of the two solution before mixing.

Show Mr. Reed the plot.



- 6) Record the final temperature for the water mixture from your Temperature vs Time plot.
_____ °C.
- 7) Calculate the temperature change of the water mixture _____ °C.
- 8) Assuming you now only have 100.0 g of water, calculate the heat gained (in joules) by the water.
 $q_{\text{water}} = m_{\text{water}} \cdot C_{\text{water}} \cdot \Delta T_{\text{water}}$

- 9) Using your values from step 8 determine the heat released by the neutralization reaction.

$$\Delta H_R = \text{_____ J}$$

10) Calculate

- (a) the number of moles of HCl reacted.

$$\text{Moles}_{\text{HCl}} = M_{\text{HCl}} \cdot V_{\text{liters}}$$

- (b) Calculate the number of moles of NaOH reacted.

$$\text{Moles}_{\text{NaOH}} = M_{\text{NaOH}} \cdot V_{\text{liters}}$$

- 11) Using your values from step 10, calculate the number of moles of water produced.

- 12) Calculate the **molar** (per mole of water) heat of neutralization in **kilojoules per mole**.

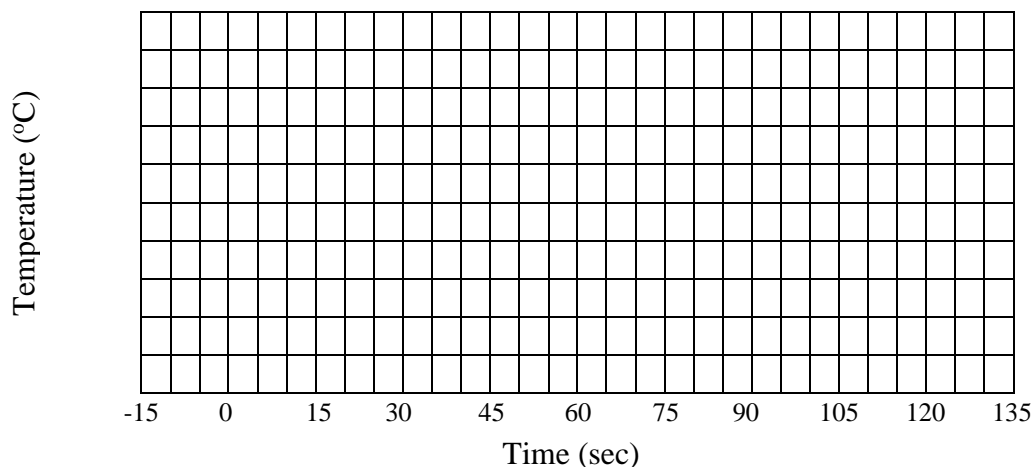
Repeat the procedure to collect data for a 2nd trial.

Clean and dry both the calorimeter and thermometer very well! WASH! WASH! DRY! DRY!

- 13) Carefully measure 50.0 mL of 1.0 M NaOH in a 100-ml graduated cylinder and pour it into the calorimeter. Leaving the thermometer out, place the top on the calorimeter.
- 14) Rinse the graduated cylinder with some tap water. Then rinse the graduated cylinder with some - about 10 mL - of the 1.0 M HCl. Dry the graduated cylinder (*as best as possible*) with a paper towel drying wand. Carefully measure 50.0 mL of the 1.0 M HCl. Leaving the HCl in the graduated cylinder measure the initial temperature of the acid _____°C.
- 15) Rinse the thermometer with tap water and dry. Insert the thermometer into the calorimeter's top and measure the initial temperature of the base _____°C. **(The two initial temperatures should not differ by more than 0.5°C. If they do ask for assistance)**
- 16) **Reread entire step before commencing.** Carefully add all of the acid to the base in the calorimeter being careful NOT to splash any acid on the upper sides of the calorimeter. Stir the mixture gently with the thermometer observing the temperature every 15-seconds for the next 2 minutes.

_____ , _____ , _____ , _____ , _____ , _____ , _____ , _____ , _____
15 s 30 s 45 s 60 s 75 s 90 s 105 s 120 s 135 s

- 17) Construct a Temp. vs Time plot to extrapolate the instantaneous ΔT value and extrapolate the data as Mr. Reed showed you how in the first trial.



- 18) Record the final temperature for the water mixture from your Temperature vs Time plot.
_____°C.
- 19) Calculate the temperature change of the water mixture _____°C.
- 20) Assuming you now only have 100.0 g of water, calculate the heat gained (in joules) by the water.
 $q_{\text{water}} = m_{\text{water}} \cdot C_{\text{water}} \cdot \Delta T_{\text{water}}$

- 21) Using your values from step 20 determine the heat released by the neutralization reaction.

$\Delta H_R = \text{_____ J}$

22) Calculate ...

(a) the number of moles of HCl reacted.

$$\text{Moles}_{\text{HCl}} = M_{\text{HCl}} \cdot V_{\text{liters}}$$

(b) Calculate the number of moles of NaOH reacted.

$$\text{Moles}_{\text{NaOH}} = M_{\text{NaOH}} \cdot V_{\text{liters}}$$

23) Using your values from step 22, calculate the number of moles of water produced.

24) Calculate the **molar** (per mole of water) heat of neutralization in **kilojoules per mole**.

25) Write the "molecular" neutralization equation for the reaction between HCl and NaOH.

26) Write the net ionic equation for the reaction that occurs between HCl and NaOH.

Lab-9: The Law of Dulong and Petit

Introduction:

The first direct approach to the determination of atomic masses was proposed in 1819 by two Frenchmen, Pierre Dulong and Alexis Petit. They assumed that the amount of heat required to raise the temperature of one atom of a solid element by 1°C should be independent of the type of atom ... since atoms are atoms! In their words, "*the atoms of all simple bodies have the same capacity for heat.*" Since one mole of every element contains the same number of atoms, a fixed amount of heat should then be required to raise the temperature of one mole of a solid at constant pressure by 1°C. This amount of heat is called the **molar heat capacity** (C_p) and Dulong and Petit set this value at about 25 J/mol°C.

$$C_p = \Delta H/\Delta T \approx 25 \text{ J/mol}^\circ\text{C}$$

The Law of Dulong and Petit is expressed most simply in terms of the property called **specific heat** (c_p). Specific heat is the amount of heat needed to raise the temperature of 1 gram of a substance at constant pressure by 1°C.

$$c_p = q/m\Delta T$$

By the Law of Dulong and Petit, the product of the molar mass (MM or M) of a solid element multiplied by its specific heat (c_p) is approximately equal to the **molar heat capacity constant** of 25 J/mol°C.

$$M \times c_p \approx 25 \text{ J/mol}^\circ\text{C}$$

thus ...

$$M \approx (25 \text{ J/mol}^\circ\text{C}) \div (c_p)$$

This was a relatively crude approach but remember that it was 1819!

Purpose:

In this experiment you will determine the specific heat of an unknown metallic element. Using your experimental value for the unknown's specific heat and the Law of Dulong and Petit, you will approximate the molar mass for the unknown element.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

DATA TABLE ON BACK!

- 1) **STUDENT-1:** Find and record the mass of the unknown metallic element. Give metal to student-2.
- 2) **STUDENT-2:** Place enough water in a large beaker to cover the unknown metallic element [**do not put the unknown metallic element in the water at this time**]. Bring water to a boil.
- 3) **STUDENT-1:** Mass the empty calorimeter. Fill the calorimeter with JUST ENOUGH tap water to completely cover the unknown metallic element. Carefully read and record the mass of the calorimeter and the water. **DO NOT SPILL THE WATER!**
- 4) **STUDENT-1:** Read and record the initial temperature of the water in the calorimeter.
- 5) **STUDENT-2:** Carefully place the unknown metallic element in the boiling water with the tongs. Allow several minutes for the boiling water and the unknown metallic element to reach the same temperature. While touching the unknown metallic element with the thermometer's bulb, read and record the initial temperature of the hot unknown metallic element in the calorimeter (*which is also the temperature of the boiling water*).
- 6) **STUDENT-2:** Carefully and quickly transfer the unknown to the calorimeter. **Don't splash water!** Carefully stir the water and record the final temperature of the water in the calorimeter (which is also the final temperature of the unknown). **This will occur in 60 seconds.**
- 7) Repeat steps 2 – 6 twice more. **Empty, dry, and use new water in the calorimeter each time!**
- 8) Calculate the specific heat of the unknown metallic element for each trial and then the average specific heat for unknown metallic element.
- 9) Using the **average** calculated specific heat values and the Law of Dulong and Petit, determine the molar mass of the unknown metallic element.

Sample	Trial	Unknown Metallic Element's Mass (g)	Initial Temp of the unknown Metal (°C) (hot)	Mass of the Calorimeter without the Top (g)	Mass of the Calorimeter and Water (g)	Initial Temp of Water in the Calorimeter (°C)	Final Temp of Water & Metal in the Calorimeter (°C)
Unknown Metallic Element	Trial 1						
	Trial 2						
	Trial 3						

*To find the change in temperature (ΔT), you always subtract initial temperature from final temperature.

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

*To find the specific heat of the unknown, you need to first find the heat absorbed (q_{water}) by the water:

$$q_{\text{water}} = m_{\text{water}} \cdot c_{\text{water}} \cdot \Delta T_{\text{water}}$$

*The heat absorbed by the water (q_{water}) will be equal to the heat lost by the unknown metal (q_{metal}).

$$q_{\text{metal}} = -q_{\text{water}}$$

*So now you can find the specific heat of the unknown metallic element:

$$q_{\text{metal}} = m_{\text{metal}} \cdot c_{\text{metal}} \cdot \Delta T_{\text{metal}}$$

*Substituting one equation into the other produces the equation below:

$$m_{\text{metal}} \cdot c_{\text{metal}} \cdot \Delta T_{\text{metal}} = -m_{\text{water}} \cdot c_{\text{water}} \cdot \Delta T_{\text{water}}$$

$$c_{\text{metal}} = \frac{-m_{\text{water}} \cdot c_{\text{water}} \cdot \Delta T_{\text{water}}}{m_{\text{metal}} \cdot \Delta T_{\text{metal}}}$$

Trial 1

Trial 2

Trial 3

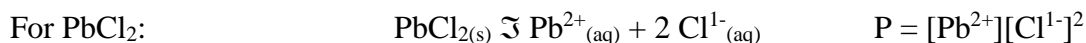
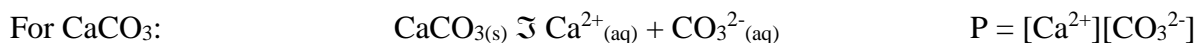
Lab-10: Precipitation Titrations

Introduction:

The process of a reaction between an observed volume of solution of unknown concentration with an observed volume of solution of known concentration (a ***standard solution***) is called ***titration***. When pairs of ions, that compose very insoluble ionic compounds, appear in solution together, precipitation generally occurs. If the concentration of one of the ions is known, then the amount of the other ion (necessary for “complete precipitation”) can be determined through precipitation titration.

The concentration of the ions needed to trigger precipitation is governed by the ionic compound’s K_{sp} value. The K_{sp} for a compound represents the maximum value that the ion-product (P) may be for a pair of ions in solution.

A solution’s ion-product value (P) is determined by multiplying the molar concentration of the cation by the molar concentration of the anion each raised to a power equal to its coefficient in the solubility equation.



To determine whether precipitation will occur, it is necessary to calculate the ion-product value (P) for a particular solution and then to compare the solution’s P value to the compound’s K_{sp} value.

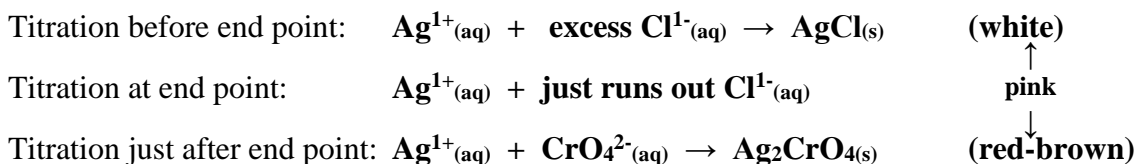
- ☞ If the solution’s $P <$ the compound’s K_{sp} , then no precipitation will occur.
- ☞ If the solution’s $P >$ the compound’s K_{sp} , then precipitation will occur.
- ☞ If the solution’s $P =$ the compound’s K_{sp} , then the solution is just saturated.

Questions For Thought:

- 1) If a solution contains $[\text{Ca}^{2+}] = 8.1 \times 10^{-5} \text{ M}$ and $[\text{CO}_3^{2-}] = 1.6 \times 10^{-5} \text{ M}$, then will precipitation occur if the K_{sp} for CaCO_3 is 4.5×10^{-9} ?
 - a) Write the solubility equation. _____
 - b) Determine the value of P.

- 2) If a solution contains $[\text{Pb}^{2+}] = 9.5 \times 10^{-2} \text{ M}$ and $[\text{Cl}^{-}] = 7.3 \times 10^{-2} \text{ M}$, then will precipitation occur if the K_{sp} for PbCl_2 is 1.7×10^{-5} ?
 - a) Write the solubility equation. _____
 - b) Determine the value of P.

The chloride ion concentration of a solution may be determined by titrating a sample with a silver nitrate solution of known concentration. A white precipitate of silver chloride forms. The indicator used to signal to end point is a dilute solution of potassium chromate. When all the chloride ions have reacted (precipitated), any excess silver ions added will react with the chromate ions to form a red-brown precipitate of silver chromate. This procedure is known as **Mohr's method**.



Purpose:

In this laboratory activity, you will determine the molar concentration of chloride ions in a water sample [*Mohr's Method*] by titrating it with a silver nitrate solution of known concentration (**0.10 M**).

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Create a data table in the space below in the "Data and Observations" section.
- 2) Obtain about 30 mL of "seawater" in a 50 mL beaker. Pipette 20 mL of the "seawater" into a 100 mL graduated cylinder. Dilute the "seawater" up to the 100 mL mark.
- 3) Fill the burets with the diluted "seawater" and the 0.10 M silver nitrate solution. Read the initial volumes of the burets. **Caution:** Exposure to $\text{AgNO}_3_{(\text{aq})}$ will turn skin black. Wash your hands frequently.
- 4) Put about 10 mL (*it doesn't have to be exactly 10 mL - just about 10 mL*) of the diluted "seawater" in the Erlenmeyer flask. Add 1 mL of chromate indicator to the Erlenmeyer flask.
- 5) Begin adding drops of the silver nitrate solution to the Erlenmeyer flask. Swirl the Erlenmeyer flask after every few drops of silver nitrate solution that are added. Continue to add silver nitrate solution until the solution turns pink (a mixture of white and red-brown). *If you add too much of the silver nitrate solution and the solution turns dark red-brown, then add some more diluted "seawater" until the pink color is obtained.*
- 6) Read the final volumes for the diluted "seawater" and silver nitrate solutions and record them in the data table.
- 7) Wash and very carefully "shake dry" the Erlenmeyer flask.
- 8) Repeat steps 2 - 5 two more times.

Data and Observations: Create your data table here!

Lab-11: Activity Series of Selected Metals

Introduction:

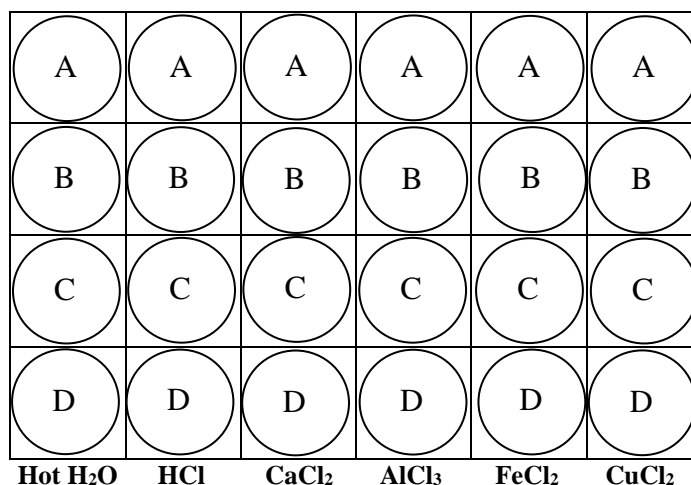
An activity series is a ranking of elements based on their ease of reaction. In the case of metals, an activity series is based on their ease of oxidation and the metals are listed from the most reactive to the least reactive. When metals are oxidized, their atoms lose one or more electrons. An activity series can be used to predict whether or not a particular reaction will or will not occur. As a general rule, metallic ions will spontaneously be displaced from solution by another metal higher on the activity series. Hydrogen gas is considered with the metals in the metallic activity series.

Purpose:

To rank, from highest to lowest, the activity of the four pure metals **based on your observations**.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Obtain 7 pieces of metal-A, metal-B, metal-C, and metal-D. **HANDLE WITH FORCEPS ONLY!**
- 2) Place a small piece of each metal in the wells as shown in the diagram below. The 7th piece of each metal is to be used as a control.



- 3) In column-1, add enough drops of hot distilled water to almost fill each well. Observe each well for any evidence of a reaction and the level of that reaction. ***You may have to re-observe several times.***

- 4) In column-2 add enough drops of 1.0 M HCl to almost fill each well. **Caution: 1.0 M HCl is caustic and corrosive.** Observe each well for any evidence of a reaction and the level of that reaction. ***You may have to re-observe several times.***

- 5) In column-3, add enough drops of 0.10 M CaCl_2 to almost fill each well. Observe each well for any evidence of a reaction and the level of that reaction. *You may have to re-observe several times.*

- 6) In column-4, add enough drops of 0.10 M AlCl_3 to almost fill each well. Observe each well for any evidence of a reaction and the level of that reaction. *You may have to re-observe several times.*

- 7) In column-5, add enough drops of 0.10 M FeCl_2 to almost fill each well. Observe each well for any evidence of a reaction and the level of that reaction. *You may have to re-observe several times.*

- 8) In column-6, add enough drops of 0.10 M CuCl_2 to almost fill each well. Observe each well for any evidence of a reaction and the level of that reaction. *You may have to re-observe several times.*

- 9) Carefully rinse your well plate in the dilute baking soda bath. Make sure all the metal samples are dumped into this bath.

10) Carefully rinse your well plate with hot/warm water in the sink in the prep room.

11) Thoroughly clean the well plate. Use a cotton swab to scrub clean each well in the well plate and dry the well plate with paper towel.

12) Tentatively rank, from highest to lowest, the activity of the four metals **based on your observations.**

_____ > _____ > _____ > _____

Lab-12: Molar Mass of an Unknown Volatile Liquid

Introduction:

Volatile liquids evaporate at a high rate at room temperature and thus have high vapor pressures. When heated to a high temperature, a volatile liquid will quickly evaporate and expand to fill the container. The amount of vapor that will fill the container depends on the container's volume, the temperature of the vapor, and the prevailing atmospheric pressure. Any excess vapor will be expelled from the container. Once equilibrium between the amount of vapor in the container and the prevailing conditions has been reached, the system can cool, the vapor can condense and its mass can be determined. Using the density form of the universal gas law ($dRT = MMP$), it is possible to determine the molar mass (MM) of an unknown volatile liquid, where d is density (g/L), R is 0.0821 L·atm/mol·K, T is temperature (K), and P is pressure (atm).

This procedure will work with any volatile liquid since equal numbers of vapor molecules will occupy the same volume under the same conditions of temperature and pressure. This is known as Avogadro's Law (or Avogadro's Hypothesis):

$$n \propto \frac{PV}{T}$$

This eventually led to the combined gas law:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

The combined gas law eventually led to the universal gas law:

$$n = \frac{PV}{RT}$$

or

$$PV = nRT$$

Since molar mass (MM) is equal to the ratio of mass (m) to moles (n), then moles (n) is equal to the ratio of mass (m) to molar mass (MM).

$$n = m/MM$$

Substituting in the universal gas law, one gets...

$$MMPV = mRT$$

Dividing both sides by volume, one gets...

$$MMP = \frac{mRT}{V}$$

Since the ratio of m/V is equal to density, the equation becomes...

$$dRT = MMP$$

or

$$MM = \frac{dRT}{P}$$

Purpose:

In this experiment, you will determine the molar mass of an unknown volatile liquid.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!**Part-1:**

- 1) On a hot plate, begin heating at the maximum setting a 400-mL beaker with 200-mL of water.
- 2) Obtain and **record the atmospheric pressure in Data Table-1**. Obtain a large white trough and half fill with tap water.
- 3) Obtain two 250-mL Erlenmeyer flasks and a stopper w/3cm glass tube. Record your trial-1 flask # _____. Record your trial-2 flask # _____.
- 4) Mass the Erlenmeyer flasks with stopper. **Record masses in Data Table-1**.
- 5) Remove the stopper, and add 10 dropper full of the unknown volatile liquid to the Erlenmeyer flask and then re-stopper the Erlenmeyer flask. **Allow the flask and liquid to sit for several minutes**.
- 6) CAREFULLY place the flask into the beaker of hot water. Bring the water to a boil. Boil for 3 minutes. Observe the system. Once you are sure all the liquid has been evaporated, **record the temperature of vapor in the flask in Data Table-1 (think)**.
- 7) CAREFULLY place your thumb over the glass tube. With your thumb securely over the glass tube place the flask into the trough to cool for one minute.
- 8) When the flask has cooled and the vapor has condensed carefully and complete dry the outside of the flask and stopper.
- 9) Mass the flask (with the unknown volatile liquid inside) and the stopper. **Record mass in Data Table-1**
- 10) Repeat steps 5 to 9 with the second flask.

Part-2:

- 11) Rinse out the flask and CAREFULLY add enough water to the two flasks to bring the level of the water just up to the bottom of the stopper. Use the dropper pipette to get the water level just right. Re-dry to outside of the flask and stopper. Mass the flasks (filled with water) and the stopper. **Record the masses of the flasks with the water in Data Table-2. 1 g of water \approx 1 mL of water.**
- 12) Remove water from the flask and completely dry the flask and rubber stopper. Repeat Part-2.

Data Table-1	Trial-1		Trial-2	
Atmospheric Pressure	atm		atm	
Mass of Empty Flask and Stopper	g		g	
Temperature of Vapor in Flask	°C	K	°C	K
Mass of Flask, Stopper, and Condensed Vapor	g		g	
Mass of Condensed Vapor	g		g	

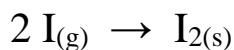
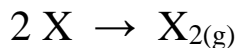
Data Table-2	Trial-1	Trial-2
Mass of Flask & Water up to Bottom of Stopper	g	g
Mass of Water up to Bottom of Stopper	g	g
Free Volume of Flask	mL	mL

Lab-13: Activity Series of the Halogens

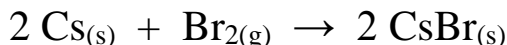
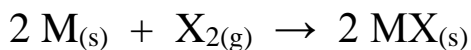
Introduction:

Pure halogens are highly reactive elements that exist as diatomic molecules. At room temperature, fluorine and chlorine are gases, bromine is liquid and iodine is solid. Pure fluorine gas is light yellow, pure chlorine gas is yellow-green, pure liquid bromine is reddish-brown, and pure solid iodine is dark violet [so dark the crystals appear black]. All halogens have strong odors.

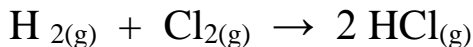
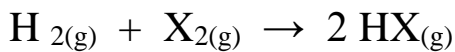
halogens atoms have high chemical activity due to having a valence that is one electron away from a full octet. Pure halogen atoms (X) react with themselves to form diatomic molecules (X₂).



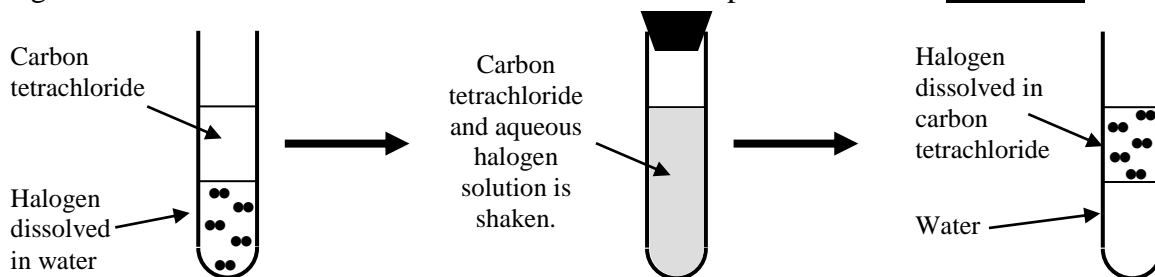
Pure halogens (X₂) easily react with alkali metals (M) to form extremely stable ionic compounds called halide salts. In fact, halogen means – “*salt former*”.



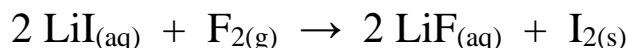
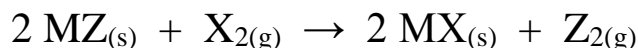
Halogens also react with hydrogen to form gaseous halides. When dissolved in water, most gaseous halides form strong acids (HCl_(aq), HBr_(aq), and HI_(aq)). Aqueous HF is a “strong” weak acid.



Pure halogens are sparingly soluble in water which is a polar solvent. However, pure halogens are more soluble in less polar solvents such as isopropanol and very soluble in nonpolar solvents such as carbon tetrachloride or mineral oil. If an aqueous solution of a halogen is mixed with carbon tetrachloride, the halogen will be drawn into the carbon tetrachloride. This process is called **extraction**.



Pure halogens can undergo **spontaneous** single replacement reactions with other halide salts.



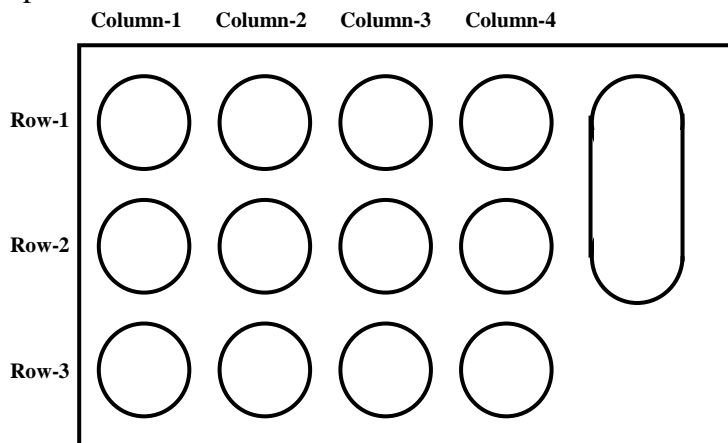
The reaction will only be spontaneous if the more active halogen is the pure element and the less active halogen is in the halide salt.

Purpose:

To rank, from highest to lowest, the activity of the pure halogens - chlorine, bromine, and iodine - based on your experimental observations.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

1) Set up a well plate as shown below.



2) In **row-1**, half fill each well with bromine water (*pure bromine dissolved in water*).

Observe the color of bromine water. _____

3) In **row-2**, half fill each well with chlorine water (*pure chlorine dissolved in water*).

Observe the color of chlorine water. _____

4) In **row-3**, half fill each well with iodine water (*pure iodine dissolved in water*).

Observe the color of iodine water. _____

5) **LEAVE COLUMN-1 AS A CONTROL.**

6) In **column-2**, using the spatula add a small amount (lentil size) of solid NaBr crystals each well.

7) One by one, stir each well with a toothpick.

8) After a few minutes, observe and record the final color of the mixtures in column-2.

Row-1 _____ Did a color change occur? _____

Row-2 _____ Did a color change occur? _____

Row-3 _____ Did a color change occur? _____

9) In **column-3**, using the spatula add a small amount (lentil size) of solid NaCl crystals each well.

10) One by one, stir each well with a toothpick.

11) After a few minutes, observe and record the final color of the mixtures in column-3.

Row-1 _____ Did a color change occur? _____

Row-2 _____ Did a color change occur? _____

Row-3 _____ Did a color change occur? _____

12) In **column-4**, using the spatula add a small amount (lentil size) of solid NaI crystals each well.

13) One by one, stir each well with a toothpick.

14) After a few minutes, observe and record the final color of the mixtures in column-4.

Row-1 _____ Did a color change occur? _____

Row-2 _____ Did a color change occur? _____

Row-3 _____ Did a color change occur? _____

15) **CAREFULLY PLACE THE WELL PLATE INTO THE BIN OF SOAPY WATER.**

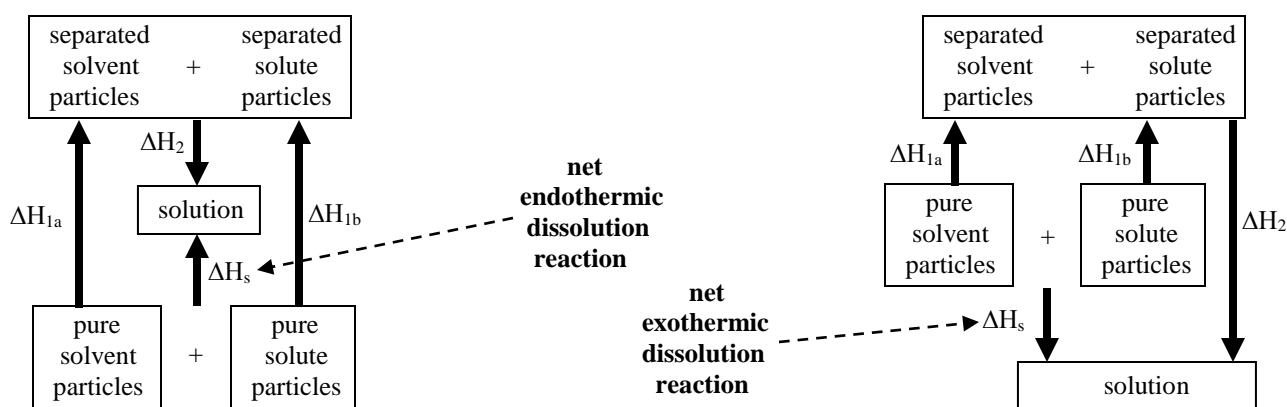
Lab-14: Heat of Solution

Introduction:

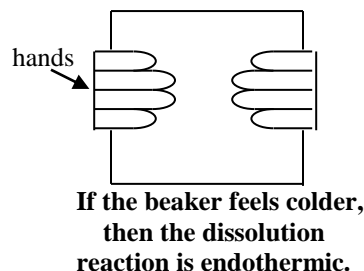
Dissolution (dissolving) reactions are examples of physical changes. When a substance (the solute) dissolves in a second substance (the solvent) a solution is formed. Solutions are homogeneous mixtures. Solutions in which the solvent is water are called aqueous solutions. Some solid substances absorb heat as they dissolve in water. Such reactions are endothermic. Some solid substances release heat as they dissolve in water. Such reactions are exothermic.

Dissolution can be thought of as taking place in two stages. In the first stage, the solute particles must be separated from each other [$\Delta H_{1b} = (+)$] and the solvent particles must be separated from each other [$\Delta H_{1a} = (+)$]. This stage is the endothermic stage since the breaking of attractions (chemical bonds or intermolecular forces) requires the absorption of energy. In the second stage, the solvent particles must surround and attract the solute particles [$\Delta H_2 = (-)$]. This stage is the exothermic stage since the forming of attractions (intermolecular forces) requires the release of energy.

If the energy absorbed in stage-1 is greater than the energy released in stage-2, then the overall heat of solution is positive [$\Delta H_s = (+)$] and the dissolution (dissolving) reaction is endothermic. However, if the energy absorbed in stage-1 is less than the energy released in stage-2, then the overall heat of solution is negative [$\Delta H_s = (-)$] and the dissolution (dissolving) reaction is exothermic.

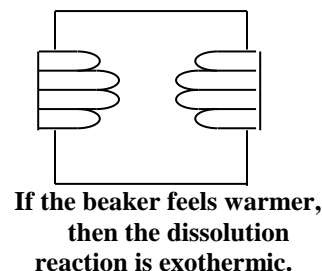


An easy qualitative method to determine whether a dissolution reaction is endothermic or exothermic is to hold a beaker of water while dissolving a substance into the water. If the beaker of water begins to feel colder, then the dissolution reaction is endothermic. If the beaker of water begins to feel warmer, then the dissolution reaction is exothermic.



The Law of Conservation of Energy states that energy cannot be created nor destroyed so that if something is losing energy, something else must be gaining that exact amount of energy.

ENERGY LOST = ENERGY GAINED!



A precise quantitative method for determining the amount of heat absorbed or released during the dissolution of a substance in water is calorimetry.

The water's temperature rises by 30°C.
Where did the heat come from?

A substance dissolves in 500g of water. If the water's temperature increases 30°C, then the water just absorbed 62700 J of heat from the substance [$q = mC\Delta T$]

The water's temperature falls by 15°C.
Where did the heat go to?

A substance dissolves in 800g of water. If the water's temperature decreases 15°C, then the substance just absorbed 50160 J of heat from the water [$q = mC\Delta T$]

Lab-15: Catalysis

Introduction:

A catalyst is a substance that increases the rate of a chemical reaction without itself being permanently altered. A catalyst can also decrease the rate of a chemical reaction. These kinds of catalysts are called **inhibitors**. Unless otherwise stated, the term catalyst usually refers to a substance that increases the rate of a chemical reaction.

A catalyst operates by lowering the activation energy required for the reaction. **Activation energy** is the minimum amount of energy that the reacting particles must initially absorb in order to begin rearranging themselves to become the products. The catalyst does this by providing a location for the reaction to occur but with an alternative pathway (an alternative reaction mechanism) of lower energy. A reaction's **mechanism** is the series of steps by which reacting particles rearrange themselves to become the products.

Since a catalyst is not permanently altered during the chemical reaction, the catalyst can be completely recovered at the end of the reaction. A catalyst is never used up! Remember, during a chemical reaction reactants are used up and products are produced!

Let's consider the following hypothetical reaction: $2 \text{X}_2\text{Z} \rightarrow 2 \text{X}_2 + \text{Z}_2$

Questions for Thought:

- 1) What is(are) the reactant(s)? _____
- 2) What is(are) the product(s)? _____
- 3) What type of chemical reaction is this? _____

Let's say that this reaction is very slow and occurs with the following step mechanism:



Questions for Thought:

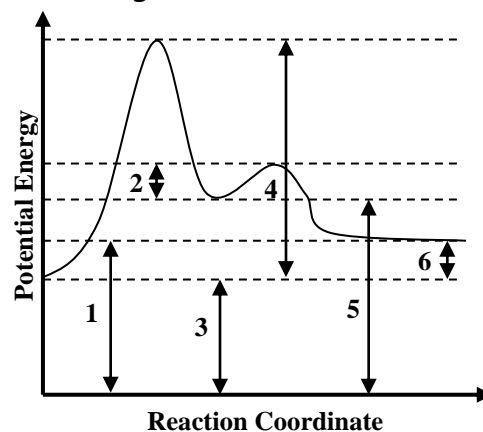
- 4) What is X_2Z_2 ? _____
- 5) What is the net reaction? _____

Let's say that the potential energy diagram for this reaction is given below:

Questions for Thought:

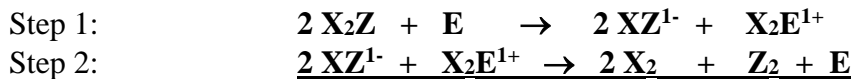
- 6) Is this overall reaction endothermic or exothermic?

- 7) Select the number on the diagram that indicates the
 - a) 1st activation energy: _____
 - b) potential energy of the reactants: _____
 - c) overall heat of reaction: _____
 - d) potential energy of the intermediate: _____
 - e) 2nd activation energy: _____
 - f) potential energy of the products: _____



- 8) Which step in the uncatalyzed reaction mechanism is the slow step? _____ Based on your answer, write the rate law for the reaction.

Let's consider introducing the catalyst, "E", into the system and let's say that the catalyzed reaction is now faster and occurs with the following two step mechanism:



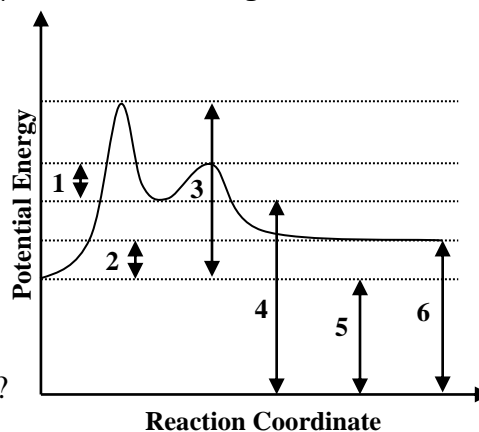
Questions for Thought:

- 9) How many moles of catalyst, **E**, is introduced in Step 1? _____
- 10) How many moles of catalyst, **E**, is produced by Step 2? _____
- 11) What is XZ^{1-} ? _____
- 12) What is X_2E^{1+} ? _____
- 13) What is the net reaction? _____
- 14) Why doesn't **E** appear in the net reaction? _____

Let's say that the potential energy diagram for the catalyzed reaction is given below:

Questions for Thought:

- 15) Select the number on the diagram that indicates the
 - a) 1st activation energy: _____
 - b) potential energy of the reactants: _____
 - c) overall heat of reaction: _____
 - d) potential energy of the intermediate: _____
 - e) 2nd activation energy: _____
 - f) potential energy of the products: _____
- 16) Is the heat of reaction for the catalyzed reaction more, less, or the same as the heat of reaction for the uncatalyzed reaction?



- 17) Which step in the catalyzed reaction mechanism is the slow step? _____ Based on your answer, write the rate law for the reaction.

The presence of the catalysts has created an alternative mechanism, which has lower activation energy. Notice that the catalyst does not appear in the net reaction equation, nor is it consumed by the reaction (1 mole of **E** was added at the beginning of the reaction and 1 moles of **E** was left at the end of the reaction). In a reaction equation, a catalyst can usually be identified as a substance that appears at the very beginning of the reaction and again at the very end of the reaction.

Notice that in both the uncatalyzed reaction and in the catalyzed reaction the intermediates were produced in the first step of the mechanism and used up in the second step of the mechanism.

- 18) What is the overall reaction order of the uncatalyzed reaction? _____
- 19) What is the overall reaction order of the catalyzed reaction? _____
- 20) Does an increase in the concentration of the catalyst change the reaction of the reaction? **Explain.**

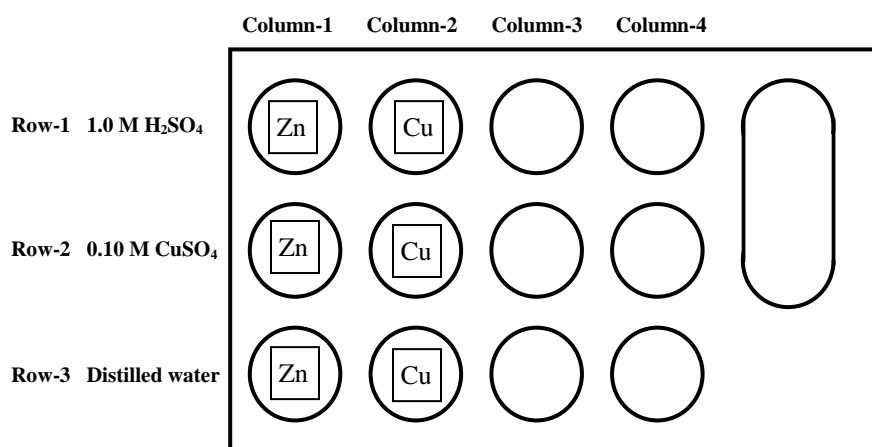
Purpose: To study the uncatalyzed and catalyzed reaction between zinc and aqueous sulfuric acid.

Procedure: **GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!**

- 1) Obtain a well plate, and put 6 pieces of zinc and 4 pieces of copper in the large oval well.
- 2) Obtain the three dropper bottles (1.0 M H_2SO_4 , 0.10 M CuSO_4 , and distilled water).
- 3) Obtain a pair of forceps.
- 4) Obtain two pieces of paper towel.
- 5) Obtain a watch glass.
- 6) Obtain a magnifying glass.
- 7) Obtain a 50 mL beaker half filled with distilled water. *This is to rinse your forceps in.*
- 8) **Wear goggles and apron for the entire lab!**

Part-A:

- 1) **ALMOST** fill the *first two* wells of each row in the well plate according to the diagram below:



- 2) Simultaneously place one piece of zinc into each well in column-1. Observe the wells with the magnifying glass. You may remove the zinc from the well and place it on a piece of paper towel using the forceps but rinse and dry the forceps each time using the 50 mL beaker of distilled water and one piece of paper towel. (*Note the speed of any changes as well as the changes themselves.*)

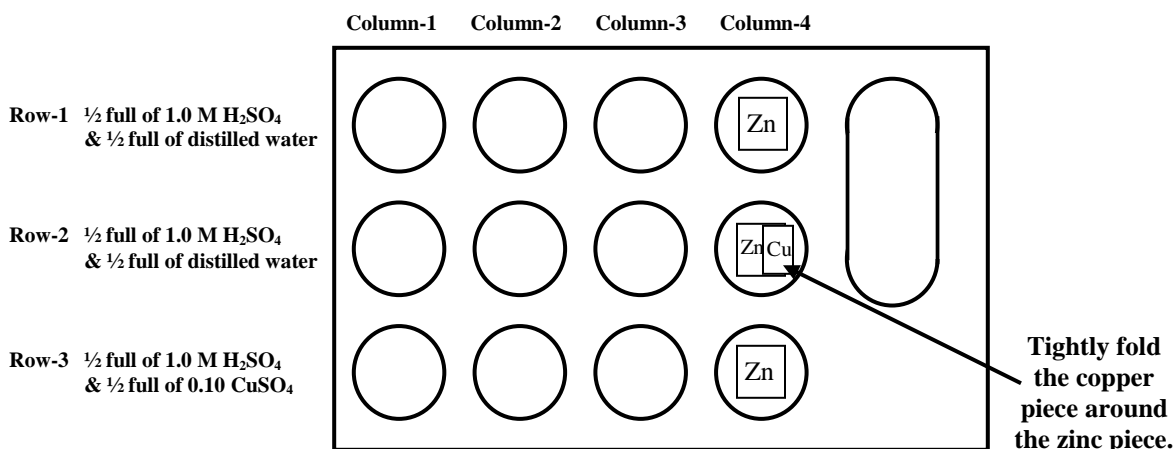
Observations: _____

- 3) Simultaneously place one piece of copper into each well in column-2. Observe the wells with the magnifying glass. You may remove the copper from the well and place it on a piece of paper towel using the forceps but rinse and dry the forceps each time using the 50 mL beaker of distilled water and one piece of paper towel. (*Note the speed of any changes as well as the changes themselves.*)

Observations: _____

Part-B:

- 4) Fill the wells of **Column-4** in the well plate according to the diagram below: **NOTE: NOT ALL OF THE WELLS ARE SET UP THE SAME! ROW-3 IS DIFFERENT!**



- 5) Mass the watch glass. _____ g
- 6) Mass the watch glass with the piece of copper on it. _____ g
Determine the mass of the copper. _____ g
- 7) Carefully fold the piece of copper in half while wrapping it around one of the pieces of zinc. **Be sure to press the copper very tightly around the piece of zinc.**
- 8) Simultaneously place the metal pieces into each well of Column-4 as shown above.
- 9) Observe the wells with the magnifying glass. You may remove the metals from their wells and place them on a piece of paper towel using the forceps but rinse and dry the forceps each time using the 50 mL beaker of distilled water and one piece of paper towel. (*Note the speed of any changes as well as the changes themselves. In which do the bubbles appear the fastest? Slowest?*)

Observations: _____

- 10) Remove the piece of copper wrapped zinc from the well of Row-2, Column-4. Carefully remove the copper. Rinse the piece of copper in distilled water. **Completely** dry the piece of copper and then mass it on the watch glass. _____ g Determine the mass of the copper. _____ g Within reason, did the mass of the copper change? _____

- 11) Write the reaction that takes place between the zinc and the sulfuric acid.
- _____

- 12) Write the reaction that takes place between the zinc and the copper (II) sulfate.
- _____

Lab-16: Chemical Kinetics

Introduction:

A chemical reaction occurs when reacting particles (the reactants) rearrange themselves to form new particles (the products). Bonds between initial atom groupings must be broken, the atoms must be rearranged, and new bonds between new atom groupings must be formed.

In order for a chemical reaction to occur, the reacting particles (the reactants) must collide with sufficient energy and in a proper orientation (angle) in space. This is known as the *Collision Theory*. The rate of a chemical reaction is affected by (1) the frequency and (2) the effectiveness of these particle collisions.

The frequency of particle collisions can be affected by (1) the speed of the reacting particles (which is related to temperature) and (2) the concentration of the reacting particles.

Questions for Thought:

- 1) What affects the speed of a group of particles in an aqueous solution?

- 2) What affects the concentration of a group of particles in an aqueous solution?

- 3) State the relationship between the speed of the reacting particles and the frequency of particle collisions.

- 4) State the relationship between the concentration of the reacting particles and the frequency of particle collisions.

Effective particle collisions are ones that result in change. **NOT ALL COLLISIONS ARE EFFICTIVE!** Effective particle collisions can be affected by (1) the energy of the collision between reacting particles and (2) the ability to get a proper orientation (angle) of the collision between reacting particles.

Questions for Thought:

- 5) What affects the energy of a collision between reacting particles?

- 6) What affects the ability to get a proper orientation (angle) of a collision between reacting particles?

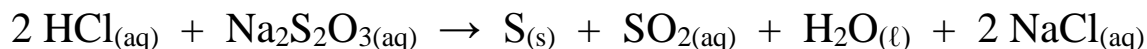
- 7) State the relationship between the energy of a collision between reacting particles and the effectiveness of the collision.

- 8) State the relationship between the ability to get a proper orientation of a collision between reacting particles and the effectiveness of the collision.

Other factors such as **the nature of the reactants** and **the reaction mechanism** can also affect the rate of a reaction. *Reactions in which few (or no) bonds need to be broken between the atoms of the reactants before rearrangement can occur are usually faster than reactions in which many bonds need to be broken before rearrangement can occur.* But reactions in which many bonds (usually covalent) need to be broken between the atoms of the reactants before rearrangement can occur are usually slower. Also, reactions with many steps in their reaction mechanism tend to be slower than reactions with few steps. In a reaction, the slowest step in the mechanism is called the **rate-determining step**. Simple reactions between aqueous ions are usually fast because the reactants are already separated into free moving ions. Complex reactions between aqueous ions can be slower if there are more complicated rearrangements between the atoms and if there are numerous steps in the reaction mechanism. A reaction such as photosynthesis: $6 \text{CO}_2 + 6 \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2$ is very slow because of the complex rearrangement of the atoms. A **Catalyst** can increase the rate of a reaction by lowering the activation energy.

Determining the Rate Law of a “Sulfur Clock” Reaction:

Purpose: To determine the rate law for the reaction between hydrochloric acid, HCl, and sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.



The rate law will be written in the form of ...

$$\text{Rate} = k[\text{HCl}]^m[\text{Na}_2\text{S}_2\text{O}_3]^n$$

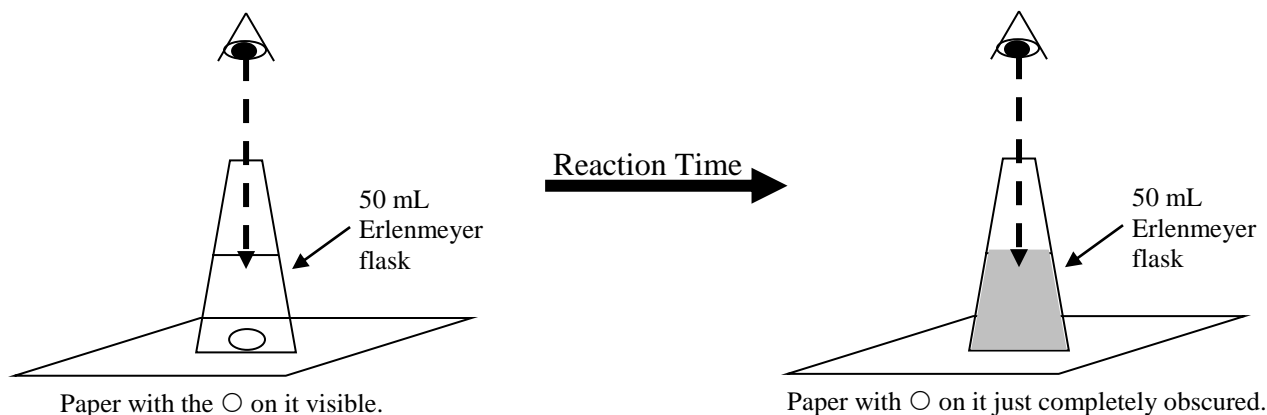
In Part-A, the HCl concentration will be held constant while the $\text{Na}_2\text{S}_2\text{O}_3$ concentration is varied. In Part-B, the HCl concentration will be varied while the $\text{Na}_2\text{S}_2\text{O}_3$ concentration is held constant. Reaction times will be measured by monitoring the appearance of sulfur.

As solid sulfur is produced in the reaction, the reaction mixture will become clouded with yellow precipitate. The reaction time will be measured by noting the time at which it is no longer possible to see through the mixture.

The time recorded is the time it takes for a certain concentration of sulfur [S] to appear which obscures the \bigcirc on the paper below the Erlenmeyer flask. Thus the average rate of the reaction is equal to ...

$$\text{Rate} = \frac{[\text{S}]}{\text{Reaction Time}}$$

It is important that you are consistently observing the moment in time it takes to just obscure the \bigcirc from view.

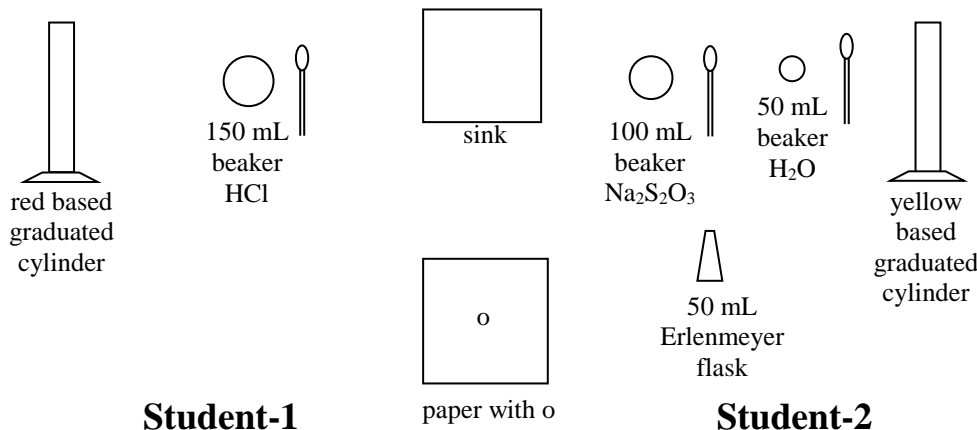


Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Obtain 50 mL of **1.00 M HCl** in the 150 mL beaker. (**CAUTION: HCl IS CORROSIVE TO EYES. WEAR YOUR GOGGLES AT ALL TIMES.**)
- 2) Obtain 50 mL of **0.30 M Na₂S₂O₃** in the 100 mL beaker.
- 3) Obtain 25 mL of distilled water in the 50 mL beaker.

Part A:

- 4) Set up your lab station as shown by the diagram below:



- 5) **Student-1:** Put 4 mL of the 1.0 M HCl solution into the red based graduated cylinder and wait for student-2 to complete step 6.
- 6) **Student-2:** Put 3 mL of the 0.30 M Na₂S₂O₃ solution into the yellow based graduated cylinder. Then CAREFULLY add 3 mL of distilled water into the same yellow based graduated cylinder until the total volume is 6 mL. Pour this diluted Na₂S₂O₃ solution into the 50 mL Erlenmeyer flask and swirl to thoroughly mix the distilled water and the Na₂S₂O₃ solution. Put the 50 mL Erlenmeyer flask onto the paper over the ○.

Student-2 you are going to time the reaction with the stop watch. Start the time when student-1 adds the HCl solution to the 50 mL Erlenmeyer flask. Stop the time the moment that you can no longer see the ○ when looking straight down into the Erlenmeyer flask. Record the time, in seconds, in the *Experiment-1/Trial-1* section of the data table below.

- 7) **Student-1:** Quickly but CAREFULLY add the 4 mL of the HCl solution into the 50 mL Erlenmeyer flask and swirl once to thoroughly mix the diluted Na₂S₂O₃ solution and the HCl solution. MAKE SURE THAT THE ERLENMEYER FLASK IS OVER THE ○ ON THE PAPER.

Student-1: When the reaction is over, empty the contents of the Erlenmeyer flask down the sink. Thoroughly wash and dry the Erlenmeyer flask.

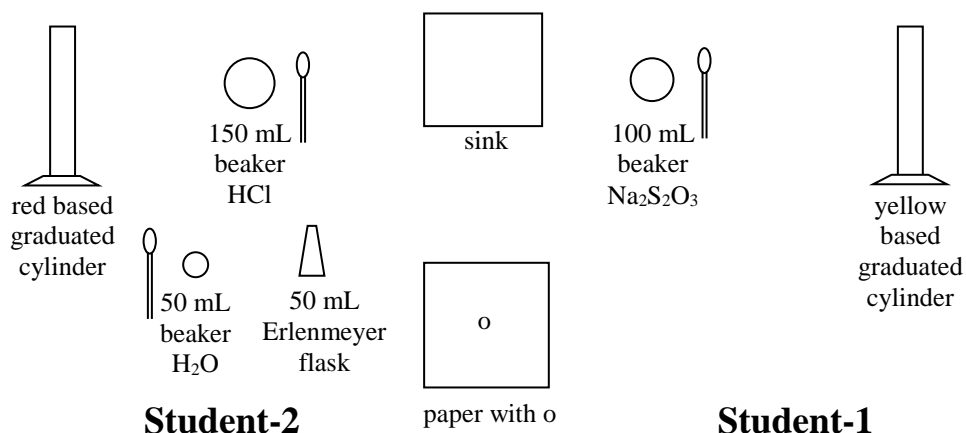
- 8) Repeat steps 5, 6 and 7 two more times for *Experiment-1*.
- 9) Repeat steps 5, 6 and 7 three times for *Experiment-2*. Notice the different volumes used.

Experiment	Student 1	Student 2		Reaction Time in Seconds (<i>to the nearest tenth</i>)			
	mL HCl	mL H ₂ O	mL Na ₂ S ₂ O ₃	Trial 1	Trial 2	Trial 3	Average
1	4.0	3.0	3.0				
2	4.0	0.0	6.0				

It is important that you are consistently observing the moment in time it takes to just obscure the ○ from view. This is why student-1 and student-2 must switch places in order that the same student who timed the reactions in Part-A times the reactions in Part-B.

Part B:

10) Set up your lab station as shown by the diagram below: **NOTICE THE CHANGES BELOW!**



11) **Student-1:** Put 4 mL of the 0.30 M $\text{Na}_2\text{S}_2\text{O}_3$ solution into the yellow based graduated cylinder and wait for student-1 to complete step 13.

12) **Student-2:** Put 3 mL of the 1.0 M HCl solution into the red based graduated cylinder. Then CAREFULLY add 3 mL of distilled water into the same red based graduated cylinder until the total volume is 6 mL. Pour this diluted HCl solution into the 50 mL Erlenmeyer flask and swirl to thoroughly mix the distilled water and the HCl solution. Put the 50 mL Erlenmeyer flask onto the paper over the ○.

Student-2 you are going to time the reaction with the stop watch. Start the time when student-1 adds the $\text{H}_2\text{S}_2\text{O}_3$ solution to the 50 mL Erlenmeyer flask. Stop the time the moment that you can no longer see the ○ when looking straight down into the Erlenmeyer flask. Record the time, in seconds, in the *Experiment-4/Trial-1* section of the data table below.

13) **Student-1:** Quickly but CAREFULLY add the 4 mL of the $\text{Na}_2\text{S}_2\text{O}_3$ solution into the 50 mL Erlenmeyer flask and swirl once to thoroughly mix the diluted HCl solution and the $\text{Na}_2\text{S}_2\text{O}_3$ solution. MAKE SURE THAT THE ERLENMEYER FLASK IS OVER THE ○ ON THE PAPER.

Student-1: When the reaction is over, empty the contents of the Erlenmeyer flask down the sink. Thoroughly wash and dry the Erlenmeyer flask.

14) Repeat steps 12, 13 and 14 two more times for ***Experiment-3***.

15) Repeat steps 12, 13 and 14 three times for ***Experiment-4***. **Notice the different volumes used.**

Experiment	Student 2		Student 1	Reaction Time in Seconds (<i>to the nearest tenth</i>)			
	mL HCl	mL H_2O	mL $\text{Na}_2\text{S}_2\text{O}_3$	Trial 1	Trial 2	Trial 3	Average
3	3.0	3.0	4.0				
4	6.0	0.0	4.0				

Data Analysis:

In order to complete the data table on the next page, you must determine the initial concentrations of the two reactants at the moment of mixing by using the dilution equation, $M_cV_c = M_dV_d$, and solve for M_d .

Recall that the HCl solution was initially 1.0 M and that the $\text{Na}_2\text{S}_2\text{O}_3$ solution was initially 0.30 M and that the final volume for both solutions was 10.0 mL. **SHOW WORK BELOW.**

Experiment-1: HCl $\text{Na}_2\text{S}_2\text{O}_3$

Experiment-2: HCl $\text{Na}_2\text{S}_2\text{O}_3$

Experiment-3: HCl $\text{Na}_2\text{S}_2\text{O}_3$

Experiment-4: HCl $\text{Na}_2\text{S}_2\text{O}_3$

In order to complete the data table below, you must determine the average rate of the reaction for each experiment. The average rate of reaction is equal to the molar concentration of the sulfur produced that just obscured the \bigcirc divided by the reaction time in seconds.

$$\text{Rate} = \frac{[\text{S}]}{\text{Reaction Time}}$$

However, in these experiments it is not possible to know that actual molar concentration of the sulfur at the moment that the \bigcirc was obscured from view. But whatever the molar concentration was, it was the same for all experiments. Thus a relative average reaction rate can be determined by the equation ...

$$\text{Relative Rate} = \frac{1}{\text{Reaction Time}}$$

Complete the data table below.

	Experiment	HCl (M)	Na ₂ S ₂ O ₃ (M)	Relative Rate (s ⁻¹)
Part A	1			
	2			
Part B	3			
	4			

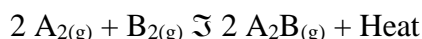
Based on the data above, INFORMALLY determine the reaction orders for HCl and Na₂S₂O₃ and then write the rate law for the reaction. Remember that this is REAL data and that going from a relative rate of 0.019 s⁻¹ to 0.035 s⁻¹ is an approximate doubling of the relative rate and that going from a relative rate of 0.011 s⁻¹ to 0.035 s⁻¹ is an approximate tripling of the relative rate. **IF YOU NEED ADVICE, ASK.** *The reaction orders will be zero order, first order, second order, or possibly something in between.*

The reaction is _____ order in HCl and _____ order in Na₂S₂O₃.

Lab-17: Le Chatelier's Principle

Introduction:

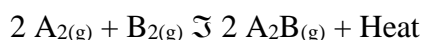
Many chemical reactions are reversible - reactants produce products and then the products produce the reactants back again. In a closed system, a state of equilibrium is eventually reached - reactants produce products as fast as the products produce the reactants back again. Once equilibrium is reached, the concentrations of the reactants and products remain constant. Consider the system at equilibrium:



As fast as two A_2 particles react with one B_2 particle to make two A_2B particles, somewhere else in the system two A_2B particles are reacting with each other to produce two A_2 particles and one B_2 particle back again. The rate of the forward reaction is equal to the rate of the reverse reaction! As long as this situation continues, the concentrations of A_2 , B_2 , and A_2B in the system will not change!

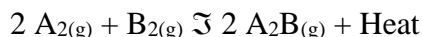
To make a change in the concentrations of A_2 , B_2 , and A_2B in the system, it is necessary to disturb the system. The disturbance will cause a temporary imbalance between the rates of the forward and reverse reactions. When something happens that disturbs a system at equilibrium, chemists say that the system has been subjected to a "stress"! There are three types of stresses that can disturb a system at equilibrium:

- 1) A change in concentration: If the concentration of one or more of the substances taking part in the reaction changes, the equilibrium will be disturbed. Consider the system at equilibrium:



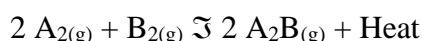
If some more A_2 is added to the system then the forward reaction can suddenly proceed faster than the reverse reaction (greater chance of collisions between A_2 and B_2). More A_2B will be produced faster than A_2 and B_2 . This situation will last only so long because the amount of A_2 and B_2 will become scarce slowing down the forward reaction. Simultaneously, the amount of A_2B will become more abundant speeding up the reverse reaction. At some point in time the two reactions will again proceed at the same rate and a new equilibrium will be established! However, at this new equilibrium there will be more A_2B than before and less A_2 and B_2 .

- 2) A change in temperature: If the temperature of the equilibrium system changed the equilibrium is disturbed. Consider the system at equilibrium:



If the temperature of the system is increased, then the reverse reaction can suddenly proceed faster than the forward reaction. The reverse reaction is endothermic and the presence of more heat will increase the rate of an endothermic reaction more than it will increase the rate of an exothermic reaction. More A_2 and B_2 will be produced faster than A_2B . This situation will last only so long because the amount of A_2B will become scarce slowing down the reverse reaction. Simultaneously, the amount of A_2 and B_2 will become more abundant speeding up the forward reaction. At some point in time the two reactions will again proceed at the same rate and a new equilibrium will be established! However, at this new equilibrium there will be more A_2 and B_2 than before and less A_2B .

- 3) A change in volume: If the volume of the equilibrium system changed the equilibrium is disturbed. Consider the system at equilibrium:



If the volume on the system is decreased then the forward reaction can suddenly proceed faster than the reverse reaction because the system's decreased volume will create a greater likelihood of collisions between two A_2 particles and a B_2 particle than between two A_2B particles! This situation will last only so long because the amount of A_2 and B_2 will become scarce slowing down the forward reaction. Simultaneously, the amount of A_2B will become more abundant speeding up the reverse reaction. At some point in time the two reactions will again proceed at the same rate and a new equilibrium will be established! However, at this new equilibrium there will be more A_2B than before and less A_2 and B_2 .

In the very late 19th century, Henri Louis Le Chatelier developed his principle by studying equilibrium systems. He came up with the following rule: **When a system at equilibrium is stressed, then the system will respond in such a way as to relieve (reduce) the stress.**

- ☞ If the concentration of a substance is increased, then the equilibrium will "shift" in the direction that will decrease it and vice versa. **An increase in a substance favors the reaction that uses it up!**
- ☞ If a system's temperature is increased (because more heat is added to the system), then the equilibrium will "shift" in the direction of the endothermic reaction (because some of the added heat will be absorbed) and vice versa. **An increase in temperature favors the endothermic reaction!**
- ☞ If a gaseous system's volume is decreased (pressure is increased), then the equilibrium will "shift" in the direction that will produce fewer moles of gas (because fewer moles of gas lowers pressure) and vice versa. **A decrease in volume favors the reaction that produces fewer moles of gas!**

Purpose:

In this laboratory activity, you will study an aqueous Co²⁺ equilibrium system and how it responds when subjected to various stresses. $Co^{2+}_{(aq)} = [Co(H_2O)_6]^{2+}_{(aq)}$

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Put 50 mL of water in a 50 mL Erlenmeyer flask and place on a hot plate. Heat water so that it is hot but not boiling. This is now the "**Hot Water Bath**".
- 2) Put 50 mL of water in a 100 mL beaker and add ice. This is now the "**Cold Water Bath**".
- 3) Set up a test tube rack with eight test tubes as shown below.



- 4) Obtain 10 mL of the 1 M CoCl₂ solution in a 50 mL Erlenmeyer flask. The color is _____.
- 5) Quarter-fill test tube-1 with the 1 M cobalt (II) chloride solution. ***This is control solution-1.***
- 6) Obtain 20 mL of 12 M HCl in a 50 mL Erlenmeyer flask.
- 7) Add just enough 12 M HCl drop-by-drop to the 1 M cobalt (II) chloride solution in the 50-mL Erlenmeyer flask to turn it to dark blue.
- 8) Quarter-fill test tube-8 with the dark blue solution. ***This is control solution-2.***
- 9) Obtain 25 mL of distilled water in a 50 mL beaker. Add just enough water, drop-by-drop, into the 50 mL Erlenmeyer flask of the dark blue solution to turn it to violet.
- 10) Quarter-fill test tube-2 with the violet solution. ***This is control solution-3.***
- 11) Quarter-fill test tube-3, test tube-4, test tube-5, test tube-6, and test tube-7, with the violet solution. The equation below represents the equilibrium system of the violet solution with the point of equilibrium in the middle (*neither shifted far to the left nor shifted far to the right*).



Now that you have five test tubes (#3 - #7) with the equilibrium somewhere in the middle, devise 4 tests to stress the system. You may ask for a solution that has not been mentioned! Write a procedure and include what color changes you expect to see. **Have your procedure approved.** Record your results.

12) _____

13) _____

14) _____

15) _____

Lab-18: Hydrolysis of Salts

Introduction:

A **salt** is an ionic compound that is produced by an acid and a base. The **hydrolysis** of a salt is a chemical reaction in which a salt's cation or anion reacts chemically with water to produce an acidic or basic solution.

A **salt** can be classified as acidic, basic, or neutral depending on how it affects the pH of the water.

- ☞ A salt that lowers the pH of water is classified as acidic.
- ☞ A salt that raises the pH of water is classified as basic.
- ☞ A salt that does not change the pH of water is classified as neutral.

A **salt** can also be classified as acidic, basic, or neutral depending on its "parents".

- ☞ A salt that is produced from a strong acid and a weak base is classified as acidic.
- ☞ A salt that is produced from a weak acid and a strong base is classified as basic.
- ☞ A salt that is produced from a strong acid and a strong base is classified as neutral.
- ☞ A salt that is produced from a weak acid and a weak base is classified based on the K_a and K_b values of the weak acid and weak base. If the $K_a > K_b$, then the salt is acidic. If the $K_a < K_b$, then the salt is basic. If the $K_a = K_b$, then the salt is neutral.

Some salts hydrolyze when dissolved in water and some salts do not hydrolyze when dissolved in water.

- ☞ Neutral salts of strong acids and strong bases **DO NOT hydrolyze** (a chemical reaction) when dissolved in water. They simply dissolve (a physical reaction) in the water.
- ☞ Acidic salts, basic salts, and neutral salts of weak acids and weak bases **DO hydrolyze** (a chemical reaction) when dissolved in water.

☞ **Neutralization:** Acid + Base \rightarrow Salt + Water

☞ **Hydrolysis:** Salt + Water \rightleftharpoons Acid + Base

Questions for Thought:

- 1) What is the K_a of $\text{HC}_2\text{H}_3\text{O}_2$? _____
- 2) What is the K_b of NH_3 (NH_4OH)? _____
- 3) Based on each salt's parent acid and base, determine whether the salt is acidic, basic, or neutral.

Salt	Parent Acid	Parent Base	Type of Salt
SnCl_2			
FeSO_4			
$\text{Sr}(\text{NO}_3)_2$			
Na_2SO_3			
Li_2CO_3			
$\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$			
KBr			

You're determining the type of salt based on theory.

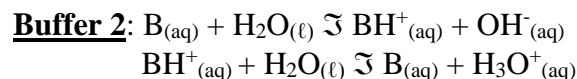
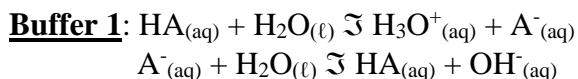
Purpose: To determine the pH of various salt-water solutions in order to classify the salts as acidic, basic, or neutral.

Lab-19: Buffers

Introduction:

Your blood must be maintained at pH 7.35 to 7.45 for you to stay healthy. However, chemical reactions taking place in your body are continuously pumping a stream of hydrogen ions into your blood. Your body maintains the proper blood pH, in spite of the hydrogen ions, due to blood buffers. Buffered solutions maintain a relatively constant pH when limited amounts of acid or base are added to them.

Buffers usually consist of solutions of a weak acid (HA) and a salt of its conjugate base (LiA, NaA, KA, etc.) or solutions of a weak base (B) and a salt of its conjugate acid (BHCl, BHBr, BHI, BHNO₃, etc.) The salts used in a buffered solution should be those that are created by a weak acid and a strong base or by a weak base and a strong acid.

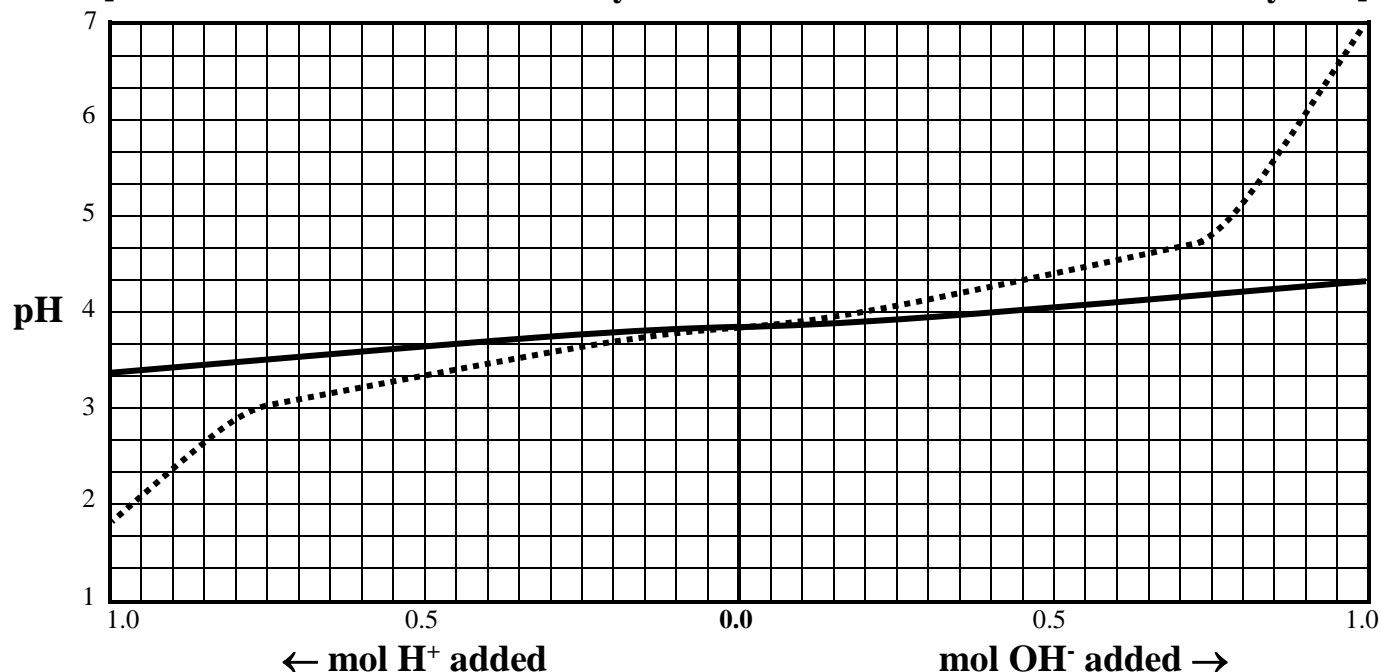


Buffered systems (solutions) are best when the concentration of the weak acid and salt of its conjugate base are roughly equal or when the concentration of the weak base and salt of its conjugate acid are roughly equal. In fact, when the concentrations of the weak acid and salt of its conjugate base are exactly equal, then the pH of the system (solution) is equal to the weak acid's pK_a. Also, when the concentrations of the weak base and salt of its conjugate acid are exactly equal then the pH of the system (solution) is equal to the weak base's pK_b.

Buffered systems (solutions) do not have an unlimited capacity to resist the effects of adding an acid or a base. A buffered system's capacity to resist changes in pH is directly related to the initial concentrations of weak acid and conjugate base (or weak base and conjugate acid). *The greater that the initial concentrations of the weak acid and conjugate base are (or the weak base and conjugate acid), the greater the system's capacity to resist changes in pH will be. SEE GRAPH BELOW.*

Lactic Acid – Lactate Ion Buffer systems:

[..... 1.0M HLac & 1.0M Lac¹⁻ System & — 2.0M HLac & 2.0M Lac¹⁻ System]



Notice that the 2 M lactic acid & 2 M lactate ion buffer system (solid line) resists changes in pH much better than the 1 M lactic acid & 1 M lactate ion buffer system (dotted line) when the same amount of acid or base is added.

The pH of a buffer system can be calculated by using the Henderson-Hasselbach equation.

For buffers that consist of a weak acid and a salt of its conjugate base, the equation is...

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^{1-}]}{[\text{HA}]}$$

For buffers that consist of a weak base and a salt of its conjugate acid, the equation is...

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{HB}^{1+}]}{[\text{B}]}$$

Technically, the ratios $[\text{A}^{1-}]/[\text{HA}]$ and $[\text{HB}^{1+}]/[\text{B}]$ are molarity ratios. However, since they are ratios and since the volume will be the same for both molarities in the ratio, then only the moles of the weak acid and the moles of the salt of its conjugate base (or the moles of the weak base and the moles of the salt of its conjugate acid) need to be in the ratio. It's a short cut!

Questions for Thought:

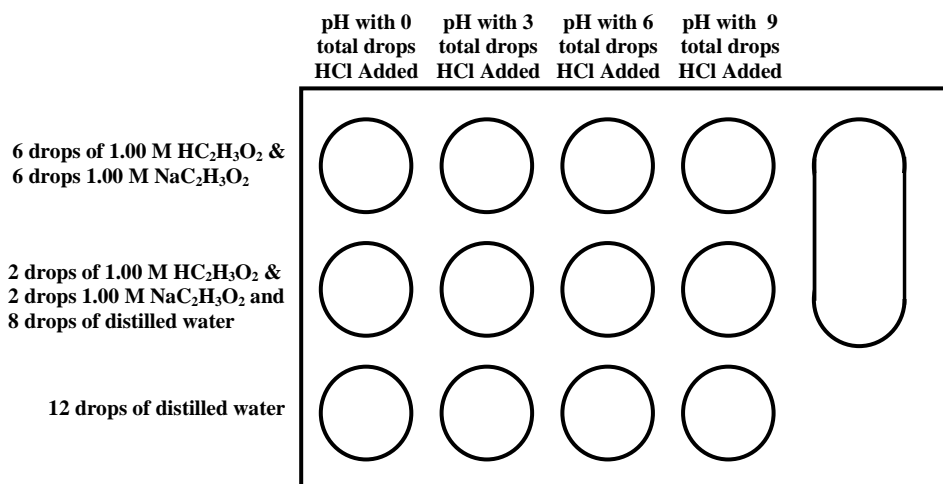
- 1) What is the pH of a buffer system that is composed of 0.0450 M H_2CO_3 and 0.0550 M Na_2CO_3 ? (K_a for $\text{H}_2\text{CO}_3 = 4.2 \times 10^{-7}$)
- 2) What is the pH of a 500 mL buffer system that is composed of 3.00 g NaHCO_3 and 3.50 g Na_2CO_3 ? (K_a for $\text{HCO}_3^{1-} = 4.8 \times 10^{-11}$)
- 3) A $\text{NH}_3\text{-NH}_4\text{Cl}$ buffer solution is to be prepared with a pH of 9.00. What must be the value of the $[\text{HB}^{1+}]/[\text{B}]$ ratio? (K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$)
- 4) What is the pH of the buffer system in question-2 after the addition of 6 mL of 1.00 M NaOH?

Purpose: To measure the pH changes that occur when acids and bases are added to buffered and unbuffered distilled water.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

Part-I:

- (1) Put on your goggles and aprons.
- (2) Obtain...
 - 1-clean well plate,
 - 2-dropper pipettes,
 - 1-dropper bottle with 1.0 M HCl, CAUTION: 1.0 M HCl IS CAUSTIC AND CORROSIVE!**
 - 1-50 mL beaker half filled with 1.00 M HC₂H₃O₂
 - 1-50 mL beaker half filled with 1.00 M NaC₂H₃O₂,
 - 1-50 mL beaker half filled with distilled water,
 - 1-roll of pHHydrion paper,
 - 1-pair of forceps,
 - Several sheets of paper towel.
- (3) Make 12 1-inch pieces of pHHydrion paper.
- (4) Set up the well plate as shown below. **BE CAREFUL WHEN ADDING THE DROPS. IF YOU MAKE A MISTAKE YOU WILL NEED TO REDO THAT STEP LATTER.**



- (5) Once the well plate is set up, **one at a time** dip a piece of pHHydrion paper into each well. Use the paper to gently stir the mixture. Leave the pHHydrion paper in the well for 10 to 20 seconds. Remove the pHHydrion paper to a piece of paper towel and read the pH to the nearest ½. Record the pH in the appropriate well above. Repeat until all the pH's of the wells have been read and recorded.
- (6) Wash and dry the well plate. **If you need to repeat any of the experiments, do it now.** Wash and dry the well plate again.

Question for Thought: Using the Henderson-Hasselbalch equation, calculate the **theoretical** initial pH for both buffer systems. Remember that each solution dilutes the other.

$$\text{pH} = 4.74 + \log\left[\frac{[\text{NaC}_2\text{H}_3\text{O}_2]}{[\text{HC}_2\text{H}_3\text{O}_2]}\right]$$

Part-II:

(1) **Put on your goggles and aprons.**

(2) Obtain...

1-clean well plate,

2-dropper pipettes,

1-dropper bottle with 1.0 M NaOH, CAUTION: 1.0 M NaOH IS CAUSTIC AND CORROSIVE!

1-50 mL beaker half filled with 1.00 M HC₂H₃O₂

1-50 mL beaker half filled with 1.00 M NaC₂H₃O₂

1-50 mL beaker half filled with distilled water,

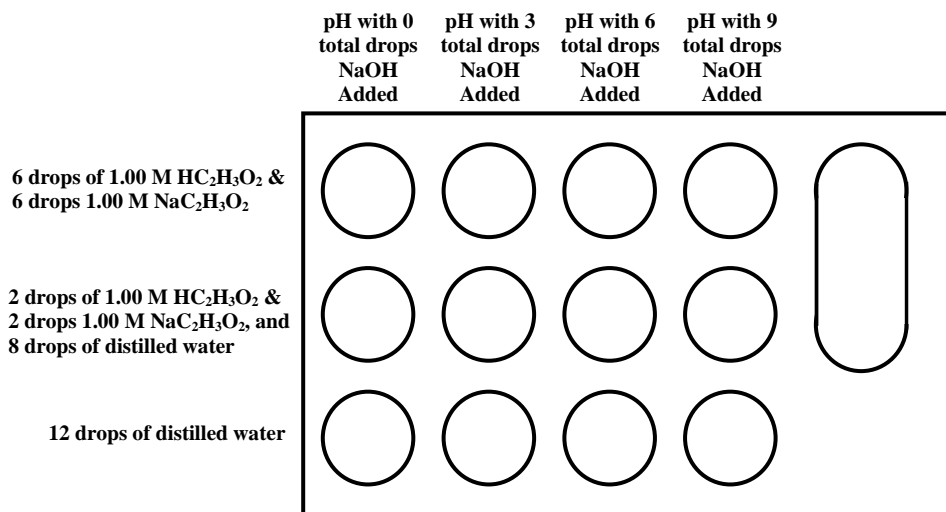
1-roll of pHyrion paper,

1-pair of forceps,

Several sheets of paper towel.

(3) Make 12 1-inch pieces of pHyrion paper.

(4) Set up the well plate as shown below. **BE CAREFUL WHEN ADDING THE DROPS. IF YOU MAKE A MISTAKE YOU WILL NEED TO REDO THAT STEP LATTER.**



(5) Once the well plate is set up, **one at a time** dip a piece of pHyrion paper into each well. Use the paper to gently stir the mixture. Leave the pHyrion paper in the well for 10 to 20 seconds. Remove the pHyrion paper to a piece of paper towel and read the pH to the nearest ½. Record the pH in the appropriate well above. Repeat until all the pH's of the wells have been read and recorded.

(6) Wash and dry the well plate. **If you need to repeat any of the experiments, do it now.** Wash and dry the well plate again.

Question for Thought: If both buffers started out at approximately the same pH, then why was one more effective at resisting changes in pH than the other?

Lab-20: K_{sp} of a Slightly Soluble Salt

Introduction:

From everyday life experiences you are familiar with several substances that are very soluble in water. Table sugar [$C_{12}H_{22}O_{11}$] and table salt [$NaCl$] are examples of soluble substances. Many other substances, such as limestone [$CaCO_3$] and gasoline are quite insoluble in water. Solubility and insolubility are relative terms, however. Almost all substances have at least a slight solubility in water. Once a substance achieves its maximum solubility in a sample of water [saturates a sample of water], a state of dynamic equilibrium is established between the dissolved substance and its undissolved solid. At equilibrium, the rate of dissolution of ions from the solid equals the rate of precipitation of the ions from the solution.



The product of the equilibrium ion concentrations raised to their molar dissolution coefficients is a temperature specific constant that is called the **solubility product constant** [K_{sp}].

The equilibrium expression for $AgCl$ is... $K_{sp} = [Ag^{1+}][Cl^{1-}]$

The equilibrium expression for $Mg(OH)_2$ is... $K_{sp} = [Mg^{2+}][OH^{1-}]^2$

The equilibrium expression for $Ca_3(PO_4)_2$ is... $K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

The greater the value of an ionic solid's solubility product constant [K_{sp}], the more soluble the ionic solid is in pure water. *See the generalizations below:*

- ☞ An ionic solid with a K_{sp} value greater than 1 is considered to be soluble in pure water.
- ☞ An ionic solid with a K_{sp} value greater than or equal to 1×10^{-4} but less than 1 is generally considered to be slightly soluble in pure water.
- ☞ An ionic solid with a K_{sp} value less than 1×10^{-4} is generally considered to be insoluble in pure water.

In this experiment, you will gravimetrically determine the K_{sp} of the slightly soluble salt – Li_2CO_3 . *A gravimetric determination is one that uses weight measurements.*

Questions for thought:

- 1) What is the equilibrium dissolution-precipitation reaction for the saturated Li_2CO_3 system?

- 2) What is the equilibrium expression for the saturated Li_2CO_3 system above? *Remember that only saturated solutions are at equilibrium.*

$$K_{sp} = \underline{\hspace{10em}}$$

For this experiment you will be making several assumptions:

- (1) *In a saturated solution of a slightly soluble salt, the volume of the solution is the same as the volume of the water itself.*
- (2) *1.0 gram of water is equal to 1.0 mL of water.*

You will determine the saturated solution's molarity by determining the mass of the solution and then by determining the mass of the solid alone after all the water has been allowed to evaporate as well as the mass of the water that evaporated. Once you know the mass of the salt, you can determine the moles of salt present in the original solution. Once you know the mass of the water, you can determine the volume in liters of the original solution.

Purpose: To experimentally determine the K_{sp} of the slightly soluble salt - lithium carbonate.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

Day -1

- 1) Obtain a clean 50 mL beaker, a clean 100 mL beaker, a clean 250 mL Erlenmeyer flask, a plastic funnel, and a piece of round filter paper.
- 2) **Record the number of the 50 mL beaker.** _____
- 3) Mass the empty 50 mL beaker. _____ g Record which scale you used. _____
- 4) Add between 40 mL and 50 mL of saturated lithium carbonate solution into the clean 100 mL beaker. Filter the saturated lithium carbonate solution into the 250 mL Erlenmeyer flask.
- 5) Add no more than 30 mL of a filtered saturated solution of lithium carbonate to the 50 mL beaker.
- 6) Mass the 50 mL beaker with the lithium carbonate solution. _____ g
- 7) Place the 50 mL beaker with the solution in the location indicated by the teacher.
- 8) Wash and dry the 400 mL beaker, the 250 mL Erlenmeyer flask, and the funnel and return.

It is necessary to wait for all of the water to evaporate from the saturated solution. Best results are obtained if this process occurs at room temperature over several weeks.

Day -2

- 9) Mass the beaker with the solid lithium carbonate. _____ g
- 10) Place the 50 mL beaker with the solid lithium carbonate in the location indicated by the teacher.
- 11) Determine the mass of the lithium carbonate that was present in your sample of saturated solution.
_____ g
- 12) Determine the mass of the water that was present in your sample of saturated solution.
_____ g
- 13) Determine the volume of the water that was present in your sample of saturated solution.
(Assume that the density of water is 1 g/mL.)
_____ L
- 14) Determine the molarity of the lithium carbonate that was present in the saturated solution.
(Assume that the volume of the solution is equal to the volume of the water.)
_____ M
- 15) Determine the K_{sp} value for the lithium carbonate at room temperature.

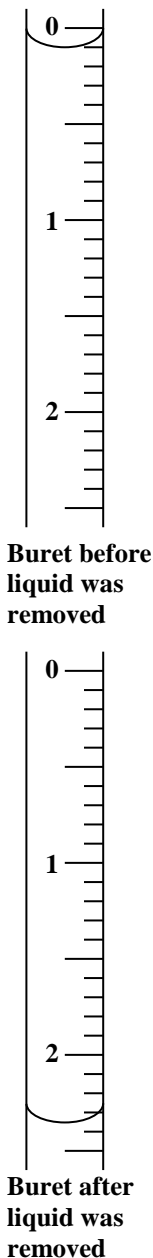
The K_{sp} is _____

Lab-21: Acid-Base Titration

Introduction: Titration is a laboratory technique that is used to determine the concentration of a solution by reacting it with a standard solution. A standard solution is a solution of known concentration. An indicator is usually used to signal the end of the reaction which is called the end point – the point in the titration when the indicator begins to change color. When an appropriate indicator is used the end point will correctly signal the equivalence point – the point in the titration when the moles of H^+ are equal to the moles of OH^- .

It is important to choose an appropriate indicator when conducting an acid-base titration. The chart below shows six common acid-base indicators and the type of reactions for which they are appropriate.

Type of Reaction	pH at Equivalence Point	Appropriate Acid-Base Indicator	Color Change And pH Range of Color Change
strong acid - weak base	Below 7	methyl orange	red to orange (3.2-4.4)
		bromocresol green	yellow to blue (3.8-5.4)
strong acid - strong base	Equal to 7	bromthymol blue	yellow to blue (6.0-7.6)
		Litmus	red to blue (5.5-8.2)
weak acid - strong base	Above 7	phenolphthalein	colorless to pink (8.2-10)
		thymol blue	yellow to blue (8.0-9.6)



A **Buret** is a piece of equipment that is used for titration. It is a long graduated cylinder with a valve at the bottom. Unlike a normal graduated cylinder, which is used to measure the volume of liquid put in, a buret is used to measure the volume of liquid removed. Also, a buret is incremented from the top to the bottom (see diagram).

Questions for Thought:

- 1) What is the volume of the solution in the buret before any liquid was removed?
_____ mL
- 2) What is the volume of the solution in the buret after some liquid was removed?
_____ mL
- 3) What is the volume of the solution in the buret that was removed? _____ mL

The following equation can be used for acid-base titration data: $\rho_A M_A V_A = \rho_B M_B V_B$ where ρ_A is the number of H^+ per unit acid, M_A is the molarity of acid, V_A is the volume (in liters) of the acid, ρ_B is the number of OH^- per unit base, M_B is the molarity of base, and V_B is the volume (in liters) of the base.

The data collected from a weak acid-strong base titration can be used to determine the K_a value of the weak acid. *The pH of the acid-base mixture halfway to neutralization is equal to the weak acid's pK_a .*

Purpose:

- I) To determine the molar concentration of an acetic acid [$HC_2H_3O_2$] solution by titrating it with a base [$NaOH$] of known concentration (**0.150 M**).
- II) To plot the curve for the titration of acetic acid and sodium hydroxide and use the curve to estimate the K_a value of acetic acid.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

Part I: Put on your goggles and aprons.

- (1) Remove the burets from the stand and add about 5 mL of the acid to the buret with the red valve. Using a small wad of paper towel over the open end, hold the buret horizontal and rinse the buret back and forth with the acid solution. Drain the buret using the valve and place the buret on the stand. Repeat the process with the base solution and the buret with the blue valve.
- (2) Fill the buret with the red valve with HCl and the buret with the blue valve with NaOH. Drain a small amount from each buret to charge the valves. Record the initial volumes of each buret.
- (3) Put about 10 mL (*it doesn't have to be exactly 10 mL - just about 10 mL*) of the acid in the Erlenmeyer flask. Put a drop of phenolphthalein in the acid in the Erlenmeyer flask.
- (4) Begin adding drops of base into the Erlenmeyer flask. Swirl the Erlenmeyer flask after every few drops of base are added. Continue to add base until the solution turns the palest of pink. *If you add too much of the base and the solution turns dark pink, then add some more acid and then begin adding more base until the palest of pink is obtained.*
- (5) Read the final volumes for the acid and base solutions and record them in the data table
- (6) Wash and very carefully "shake dry" the Erlenmeyer flask.
- (7) Repeat steps 3 - 6 two more times.

Data and Observation:

Part I: Table 1: Volume of Acid and Base Used during Titration

	Trial 1		Trial 2		Trial 3	
	HC ₂ H ₃ O ₂	NaOH	HC ₂ H ₃ O ₂	NaOH	HC ₂ H ₃ O ₂	NaOH
Initial Volume (mL)						
Final Volume (mL)						
Volume Used (mL)						

Part II: Put on your goggles and aprons.

- (1) Put exactly 10.0 mL of the acid in the 50-mL beaker. Use a strip of universal pH paper to determine the initial pH of the acid and record it in the data table.
- (2) Begin adding base into the 50-mL beaker at ½ mL amounts. *Read the new volume for the base and record it in the data table.* Use a strip of universal pH paper (about ½ in long) to determine the initial pH of the acid and record it in the data table.
- (3) Repeat step-3 until 9 mL of NaOH has been added.
- (4) **Clean up and leave the lab station as you found it!**

Data and Observation:

Part II: Initial pH of acetic acid solution (before adding any NaOH). _____

Table 2: pH of solution during Titration

Total Milliliters of NaOH Added	pH	Total Milliliters of NaOH Added	pH
1)		10)	
2)		11)	
3)		12)	
4)		13)	
5)		14)	
6)		15)	
7)		16)	
8)		17)	
9)		18)	

Lab-22: Redox Titration

Introduction:

- Redox reactions typically go to completion.
- A redox titration is a procedure that can be used to determine the concentration of a reactant.
- The goal is to find the moles of titrant needed to bring the reaction to completion. Then, using stoichiometry, the goal is to find the moles of the unknown.
- Often, no outside indicator is needed for a redox titration because one of the reactants will serve as the indicator. If an extra drop of titrant is added at the titration's endpoint [when the reactants are at their stoichiometric ratio], then a color change would be observed because the extra added titrant has nothing with which to react.
- In this experiment, a purple/pink permanganate solution will be added to a colorless oxalate solution. At first all of the added pink permanganate is completely reacted by the colorless oxalate to produce all colorless products. Eventually, all the oxalate is reacted and the pink color remains. This indicates that the titration has gone too far. Thus back titration is required until a **pale yellow** is achieved – **END POINT!**
- This reaction is slow at room temperature. You will need to create a hot water bath in which to heat the oxalate solution before you begin to titrate it with the potassium permanganate solution. Do this by half-filling a 250-mL Erlenmeyer flask with water and place it on the hot plate. Heat the water so that it is hot but not boiling ($\approx 90^{\circ}\text{C}$). The oxalate solution must be kept warm throughout the titration. *Thus, you may need to reheat the test tube part way through the titration.*
- **This reaction has one oddity!** One of the products, aqueous manganese (II) ions, is a catalyst for the reaction. At the start, there is no aqueous manganese (II) ions present and the disappearance of the pink color is slow. But after each addition of permanganate ions, the disappearance of the pink color accelerates due to the presence of more and more aqueous manganese (II) ions.

Questions for Thought:

- 1) What is the formula for potassium permanganate? _____
- 2) What is the oxidation number for the manganese in the permanganate ion? _____
- 3) What is the formula for sodium oxalate? _____
- 4) What is the oxidation number for the carbon in the oxalate ion? _____
- 5) What is the formula for manganese (II) ions? _____

Purpose: To determine the molar concentration of a potassium permanganate solution (based on three trials of data) by titrating it with a solution of 0.00200 M sodium oxalate that has been acidified with sulfuric acid.

Materials: A large test tube (25 mm \times 200 mm), a 250-mL Erlenmeyer flask, a hot plate, two burets with buret stand, two 50-mL beakers (to hold the 0.00200 M sodium oxalate solution and the unknown potassium permanganate solution), and paper towel (to make drying wands for the large test tube).

- 6) The reaction between the potassium permanganate and the **acidified** sodium oxalate produces manganese (II) ions and carbon dioxide. Assuming that the potassium and the sodium ions are spectator ions, write a balanced net ionic equation for the redox reaction that takes place.

Keep in mind that this is a redox reaction that is taking place in an acidified environment.

YOU MUST COMPLETE THIS STEP BEFORE YOU CAN MOVE ON TO STEPS 2 – 4!

- 7) Using the volumes used, calculate the moles of **oxalate** (0.00200 M) titrated in each of the trials.

Trial 1

Trial 2

Trial 3

8) Using stoichiometry, calculate the moles of **permanganate** used in each of the trials.

Trial 1

Trial 2

Trial 3

9) Using the volumes used, calculate the molar concentration of the **potassium permanganate** solution in each of the trials.

Trial 1

Trial 2

Trial 3

10) The average molar concentration of the potassium permanganate solution is _____M

Lab-23: Reactions of Acids and Bases

Introduction:

Acids and bases are two classes of compounds with distinct properties. Some of these properties are directly observable while other properties are instead based on an interpretation of the facts.

A Compound's formula can usually indicate whether it is an acid or base.

Compound	How to Recognize an Acid or Base Compound from its Formula
Acids	<ul style="list-style-type: none"> Acid formulas usually start with hydrogen (H). (i.e. HCl, H₂SO₄) Organic acid formulas will end with the carboxyl group, (COOH) (i.e. CH₃COOH)
Bases	<ul style="list-style-type: none"> Ionic base formulas start with a metal or a (+)P.A.I. and end with the (-)P.A.I. hydroxide (OH⁻). (i.e. NaOH, Ba(OH)₂) Covalent base formulas can have the general formula R-NH₂, where R is a <u>single hydrogen atom</u> or a <u>group of carbon and hydrogen atoms</u>. (i.e. NH₃, CH₃CH₂NH₂)

Below are some directly observable properties of acids and bases.

Properties of Acids based on Direct Observation	Properties of Bases based on Direct Observation
1) Acids taste sour. (not a safe test!)	1) Bases taste bitter. (not a safe test!)
2) Acids react with many metals to produce H _{2(g)} .	2) Bases feel slippery.
3) Acids react with carbonate salts and hydrogen carbonate (bicarbonate) salts to produce CO _{2(g)} .	3) Bases react with ammonium salts to produce NH _{3(g)} .
4) Acids are caustic and corrosive.	4) Bases are caustic and corrosive.
5) Acids turn blue litmus red, etc.	5) Bases turn red litmus blue, etc.
6) Aqueous acid solutions will conduct electricity.	6) Aqueous base solutions will conduct electricity.
7) Acids neutralize bases.	7) Bases neutralize acids.

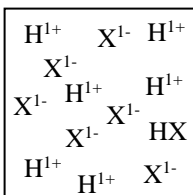
Below are some properties of acids and bases based on interpretation of facts.

Properties of Acids based on Interpretation	Properties of Bases based on Interpretation
1) Acids provide free moving ions when dissolved in water. Acids are electrolytes.	1) Bases provide free moving ions when dissolved in water. Bases are electrolytes.
2) Acids ionize in water. Strong acids ionize nearly 100% in H₂O whereas weak acids do not.	2) Ionic bases will dissociate in water whereas covalent bases ionize in water.
3) Acids produce <u>hydronium ions</u> (H ₃ O ⁺) as the only positive ions in water according to the Arrhenius Theory . [H ₃ O ⁺ (aq) ≈ H ⁺ (aq)]	3) Bases produce <u>hydroxide ions</u> (OH ⁻) as the only negative ions in water according to the Arrhenius Theory .
4) Acids are proton (H ⁺) donors according to the Bronsted-Lowry Theory .	4) Bases are proton (H ⁺) acceptors according to the Bronsted-Lowry Theory .
5) Acids are electron (e ⁻) acceptors according to the Lewis Theory .	5) Bases are electron (e ⁻) donors according to the Lewis Theory .

In water, nonmetallic binary oxides act as acids. In water, metallic binary oxides act as bases.

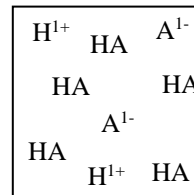
Binary Oxides	<ul style="list-style-type: none"> Nonmetallic oxides will ionize and form acidic solutions when dissolved in water. Metallic oxides will ionize and form basic solutions when dissolved in water.
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A beaker contains the **STRONG acid** HX dissolved in water. Notice that almost 100% of the HX molecules ionized as they dissolved in the water.



Many ions

A beaker contains the **WEAK acid** HA dissolved in water. Notice that very few of the HA molecules ionized as they dissolved in the water.



Few ions

Purpose:

In this laboratory activity, you will observe the reactions of a strong acid with several metals and with a carbonate and a hydrogen carbonate. You will identify any gas emitted by these reactions. You will observe the reaction of calcium oxide with water and its effect on the pH of the water. You will observe the reaction of carbon dioxide with water and its effect on the pH of the water. You will observe the reaction of a strong base with an ammonium salt. You will identify the gas emitted by this reaction. You will observe the reaction of HCl gas with NH₃ gas. You will write the net ionic equations for all observed reactions.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Set up a test tube rack with test tubes as shown in the diagram below.



- 2) Quarter fill test tube-1 with 0.10 M HCl. Quarter fill test tube-2 with 0.10 M HCl. Quarter fill test tube-3 with 0.10 M HCl. Quarter fill test tube-4 with 0.10 M HCl. Quarter fill test tube-5 with distilled water. Quarter fill test tube-6 with distilled water. Quarter fill test tube-7 with 0.10 M NaOH. Quarter fill test tube-8 with 0.10 M NH₃ and then stopper the test tube.
- 3) In test tube-1, place a piece of copper. Stopper test tube for 1 minute and observe. Carefully remove the stopper and hold a lit wooden splint just inside the mouth of the test tube. Observe.

- 4) In test tube-2, place a piece of zinc. Stopper test tube for 1 minute and observe. Carefully remove the stopper and hold a lit wooden splint just inside the mouth of the test tube. Observe.

- 5) In test tube-3, place a pea sized amount of sodium carbonate (washing soda). Stopper test tube for 1 minute and observe. Carefully remove the stopper and hold a lit wooden splint just inside the mouth of the test tube. Observe.

- 6) In test tube-4, place a pea sized amount of sodium hydrogen carbonate (baking soda). Stopper test tube for 1 minute and observe. Carefully remove the stopper and hold a lit wooden splint just inside the mouth of the test tube.

- 7) In test tube-5, carefully add several drops of bromthymol blue into the water. Observe. Place a pea sized amount of calcium oxide. Stopper test tube for several minutes. Observe.

- 8) In test tube-6, carefully add several drops of bromthymol blue into the water. Observe. Using a straw, carefully blow into the test tube for several minutes. Observe.

- 9) In test tube-7, place a pea sized amount of ammonium chloride. Stopper test tube for 1 minute and observe. Carefully remove the stopper and waft toward your nose. Observe.

- 10) Move test tube-8 and hold its mouth right next to the mouth of test tube-1. Remove the stopper from test tube-8 and observe.

11) Write the net ionic equation for the reaction that occurred in step-3. If no reaction occurred, then write "*No reaction occurred.*"

12) Write the net ionic equation for the reaction that occurred in step-4. If no reaction occurred, then write "*No reaction occurred.*"

13) Write the net ionic equation for the reaction that occurred in step-5. If no reaction occurred, then write "*No reaction occurred.*"

14) Write the net ionic equation for the reaction that occurred in step-6. If no reaction occurred, then write "*No reaction occurred.*"

15) Write the net ionic equation for the reaction that occurred in step-7. If no reaction occurred, then write "*No reaction occurred.*"

16) Write the net ionic equation for the reaction that occurred in step-8. If no reaction occurred, then write "*No reaction occurred.*"

17) Write the net ionic equation for the reaction that occurred in step-9. If no reaction occurred, then write "*No reaction occurred.*"

18) Write the net ionic equation for the reaction that occurred in step-10. If no reaction occurred, then write "*No reaction occurred.*"

Lab-24: Spectrophotometry

Introduction:

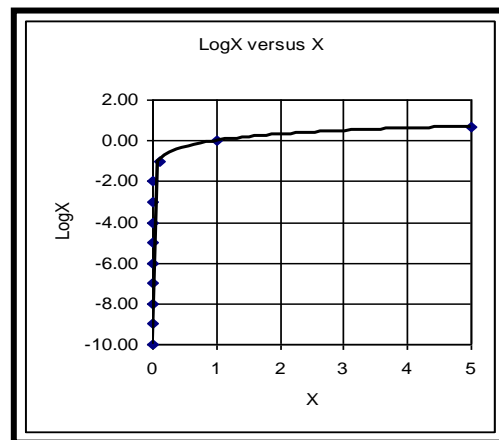
Many compounds absorb light from regions of the electromagnetic spectrum. A **spectrophotometer** is a device designed to determine the wavelength of light that a compound absorbs. When an aqueous solution of a substance is placed in the light path of a spectrophotometer, the sample may absorb all of the light, some of the light, or none of the light. When a sample does not absorb any light at a specific wavelength, it is then said to transmit 100% of the light. Light absorption occurs at wavelengths whose energy corresponds to the energy necessary to cause electronic excitations of atoms, ions, or molecules. **Absorbance** is defined as:

$$A = 2 - \log \%T$$

where **A** is **absorbance** and **T** is the **percentage of light transmitted** through the sample.

Questions for Thought:

- 1) What is the log of 100? _____
- 2) What the absorbance of a sample if the transmittance is 100%? _____
- 3) What is the log of 0? This is a problem since the answer on the calculator is “*error*”! So what is the absorbance of a sample if the transmittance is 0%?
 - a) Based on the graph to the right, what is the log of X as X approaches zero? ($\log X = ??$ as $X \rightarrow 0$)
 - b) Based on your answer above, what is the absorbance of a sample if the transmittance is 0%?



The absorbance of light at a specific wavelength depends on the material in the sample, the concentration of the sample, and path length of the light. This is known as the **Beer-Lambert Law** or simply **Beer's Law**.

$$A = abc$$

This is a linear, $y = mx + b$, equation where **A** is the **absorbance** (a dimensionless value), **a** is the **molar absorptivity constant** ($L/mol \cdot cm$), **b** is the **path length** (cm), and **c** is **molar concentration** (mol/L). The **molar absorptivity constant**, which is sometimes called the molar extinction coefficient, is not only specific to the substance but also to the temperature and wavelength.

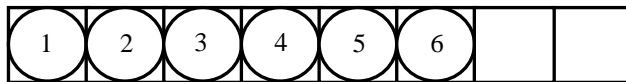
Spectrophotometry is useful for two reasons. First, the visible absorption spectrum of a substance (a continuous plot of a substance's absorbance values in the visible spectrum) is a unique characteristic of that substance and thus can be used in the identification of an unknown substance. Second, the absorbance values are directly proportional to the molar concentration at a specific wavelength and path length and thus absorbance can be used to determine the concentration of a sample. The most common way that absorbance is used to determine the concentration of an unknown solution is to plot the absorbance values of a substance at various known concentrations and then to interpolate the concentration using the unknown's absorbance value.

Purpose:

To plot a standard absorption curve at a specific wavelength for a known solution of KMnO_4 and then to use the curve to determine the molar concentration of an unknown solution of KMnO_4 . To determine the molar absorptivity constant for a KMnO_4 solution at a specified wavelength.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

- 1) Obtain 40 mL of the known KMnO_4 solution and 40 mL of distilled water in 50 mL beakers.
- 2) Fill buret-1 with the KMnO_4 solution **of known** concentration and fill buret-2 with distilled water.
- 3) Below, record the concentration of the KMnO_4 solution **of known** concentration in g/L.
- 4) Set up a test tube rack with cuvettes as shown in the diagram below.



Please be careful. Each cuvette costs about \$10!

- 5) Add 5 mL of distilled water into cuvette-1.
 - 6) Add 1 mL of the KMnO_4 solution **of known** concentration into cuvette-2 and then add 4 mL of distilled water into the cuvette-2. Stopper cuvette-2 and then invert to mix.
 - 7) Add 2 mL of the KMnO_4 solution **of known** concentration into cuvette-3 and then add 3 mL of distilled water into the cuvette-3. Stopper cuvette-3 and then invert to mix.
 - 8) Add 3 mL of the KMnO_4 solution **of known** concentration into cuvette-4 and then add 2 mL of distilled water into the cuvette-4. Stopper cuvette-4 and then invert to mix.
 - 9) Add 4 mL of the KMnO_4 solution **of known** concentration into cuvette-5 and then add 1 mL of distilled water into the cuvette-5. Stopper cuvette-5 and then invert to mix.
 - 10) Add 5 mL of the KMnO_4 solution **of known** concentration into cuvette-6.
 - 11) Wipe off any fingerprints from all test tubes with lens paper (*fingerprints with absorb light and thus give false absorbance readings*).
-
- 12) Below, record the wavelength of light being used in the spectroscope.
 - 13) Set the spectroscope so that it measures transmittance. Place cuvette-1 into the spectroscope. Adjust the spectroscope so that it reads 100% transmittance. Set the spectroscope so that it now measures absorbance. Read the absorbance value of cuvette-1 (it should be zero!). Record below.
 - 14) One by one, place cuvette-2 through cuvette-6 into the spectroscope. Read the absorbance values. Record below.
-
- 15) Empty and dry cuvette-1. Half-fill cuvette-1 with the solution of KMnO_4 **of unknown** concentration.
 - 16) Place cuvette-1 into the spectroscope and read the absorbance values. Record below.
 - 17) Empty and thoroughly rinse the cuvettes and burettes.
 - 18) Create a standard absorption curve in Microsoft Excel: **Absorbance versus Concentration (M)**

Data and Observations:

- 1) The concentration of the known solution of KMnO_4 is _____ g/mL.
- 2) The wavelength of light being used is _____ nm
- 3) Table 1: Absorbance Values

Test Tube	Percentage of Known KMnO_4	Concentration (g/L)	Concentration (M)	Absorbance Value
1	0%			
2	20%			
3	40%			
4	60%			
5	80%			
6	100%			

- 4) The absorbance of the unknown solution of KMnO_4 is _____

Lab-25: The Electrolytic Cell

Introduction:

The passage of an electric current through a pure liquid substance or a solution is accompanied by a chemical reaction at the electrodes. Oxidation occurs at the anode and reduction occurs at the cathode. An electric source pulls electrons from the anode and sends them to the cathode. This set-up is known as an electrolytic cell. In electrolytic cells, the electrode connected to the negative pole of the electric source is the cathode and the electrode connected to the positive pole of the electric source is the anode.

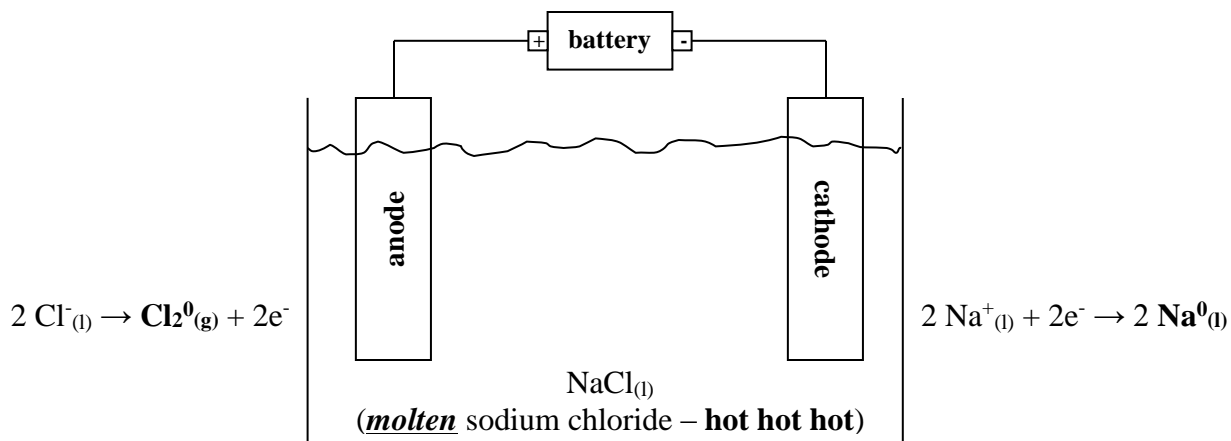
The quantitative relationship between the redox reaction and the electric current is $q = It$, where q is the charge (coulombs), I is the current (amperes), and t is the time (seconds).

The quantitative relationship between the redox reaction and the voltage of the electric source is $q = E \div V$, where q is the charge (coulombs), E is the energy (joules), and V is the voltage (volts).

The quantitative relationship between the redox reaction and the coulombs of charge is $q = n\mathcal{F}$, where q is the charge (coulombs), n is number of mole of electrons exchanged in the redox reaction (moles), and \mathcal{F} is Faraday's constant (96500 C/e⁻).

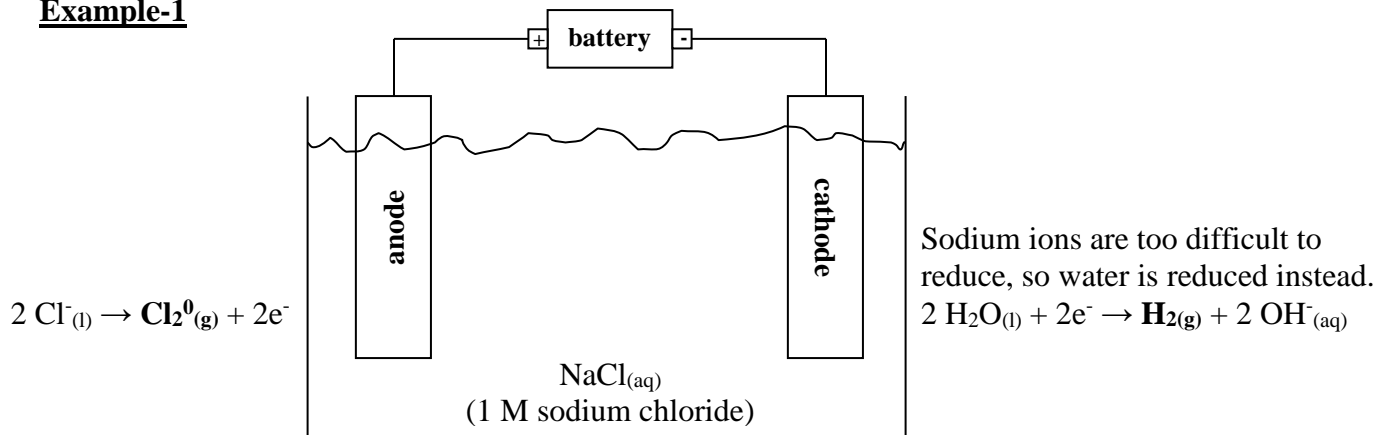
☞ *Electrolytic cells are used to produce a reaction that is otherwise nonspontaneous.*

☞ The decomposition reactions of many pure binary salts are brought about by an electrolytic cell. When used for this purpose, the electrodes are often made of graphite and are inactive electrodes which means they only serve as the site for reduction and oxidation reactions.



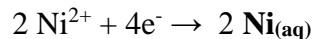
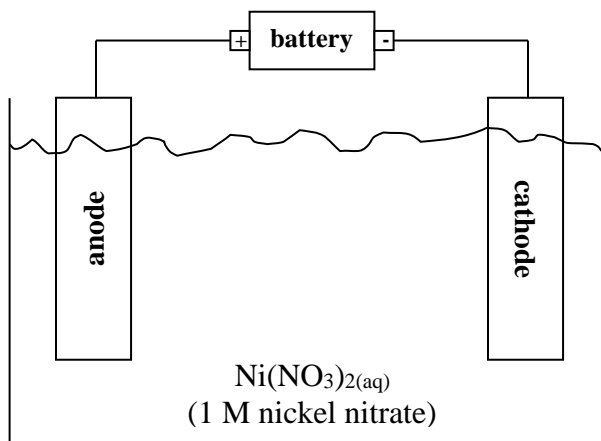
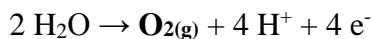
☞ The decomposition reactions of many aqueous salt solutions are brought about by an electrolytic cell. When used for this purpose, the cation and/or the anion of the salt is not always part of the redox reaction. Sometimes water is reacting at one or both of the electrodes.

Example-1

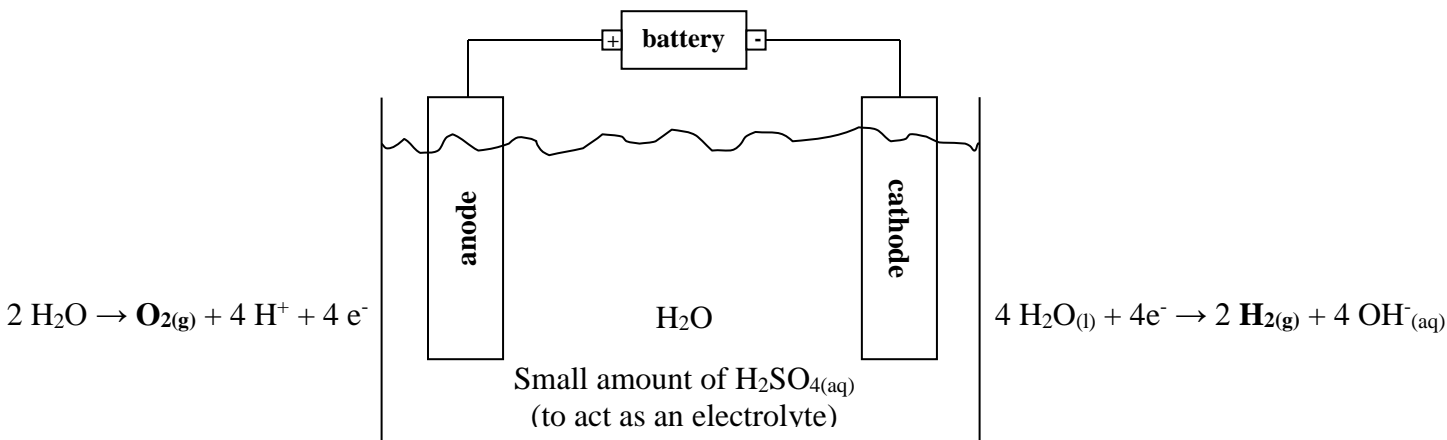


Example-2

Nitrate ions are too difficult to oxidize, so water is oxidized instead.

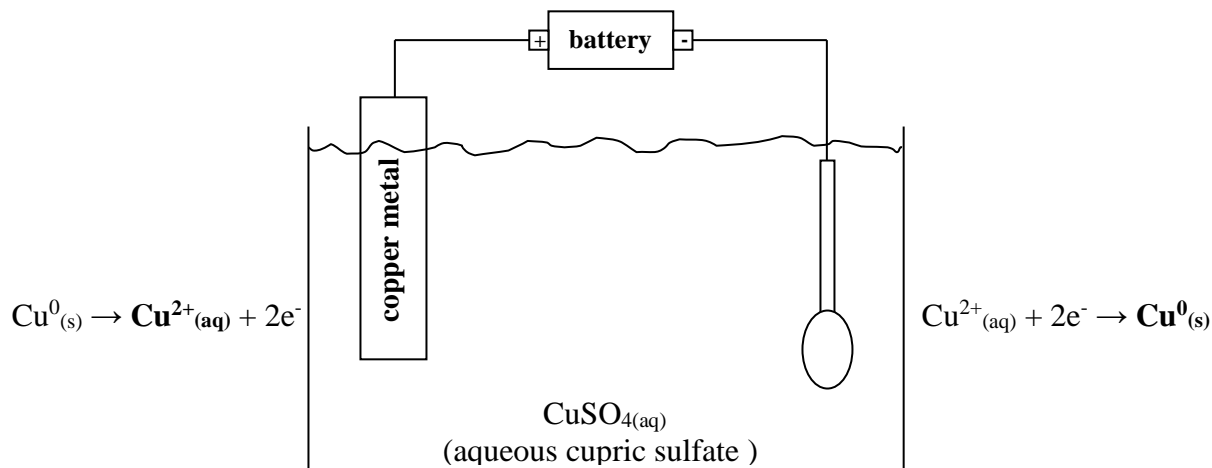


☞ The decomposition of water is brought about by an electrolytic cell. When used for this purpose, a small amount of an electrolyte is needed to conduct the electric current.



☞ *Electrolytic cells are also used as electroplating apparatuses.*

A thin coating of metal is deposited on a surface usually to enhance the beauty of the surface or to protect the surface from corrosion. In electroplating, the surface to be plated is set up as the cathode in the electrolytic cell and a piece of the plating metal is used as the anode. When used for this purpose, the electrodes are active which means they take part in the reaction.



When electroplating a surface with a metal, there are some basic rules which must be followed:

- ☞ **The object to be plated must always be the cathode in the electrolytic cell.**
- ☞ **The cathode must always be connected to the negative pole of the electric source.**
- ☞ **The anode in the electrolytic cell must be made from the plating metal.**
- ☞ **The electrolyte used in the electrolytic cell must contain the cation of the plating metal.**

Purpose:

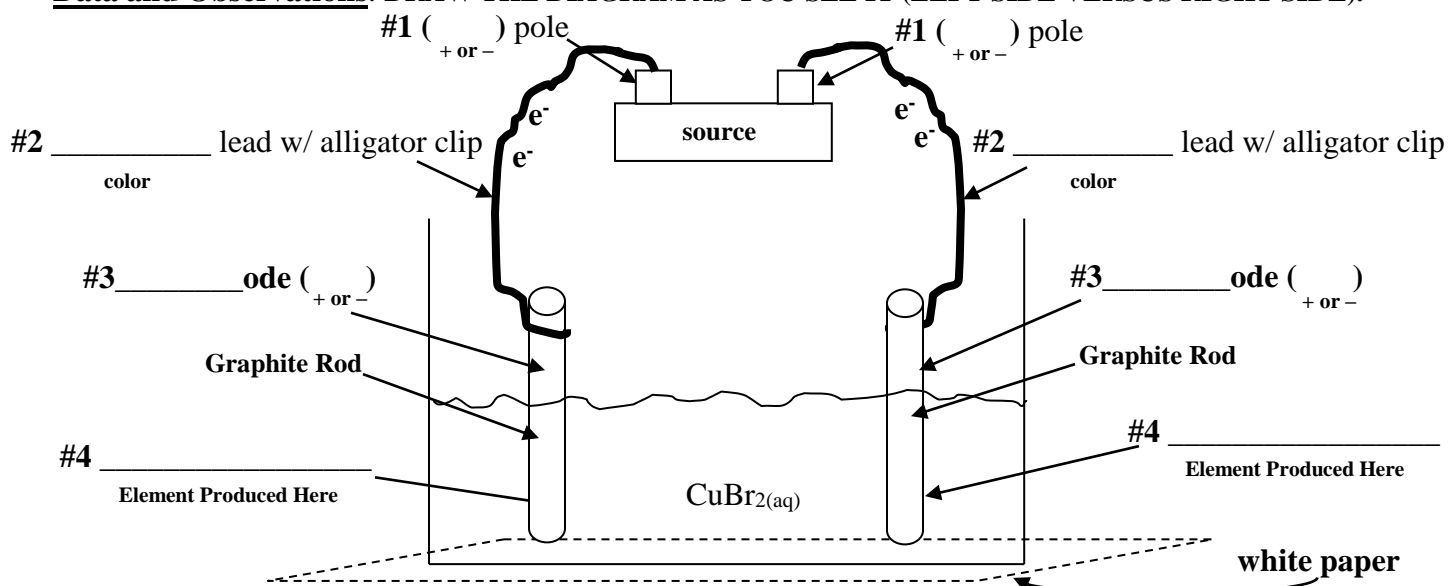
In this laboratory activity, you construct an electrolytic cell and you will undertake the electrolysis of an aqueous salt solution. You will also construct an electroplating apparatus and you will electroplate a penny with silver.

Procedure: WEAR GOGGLES AND APRON. WASH YOUR HANDS FREQUENTLY.

Part-I: An Electrolytic Cell: The Electrolysis of Aqueous Copper (II) Bromide

- 1) Obtain an electric source, a set of red and a black leads with alligator clips, two graphite rods (two mechanical pencil leads), and a 150 mL beaker with 20.0 mL of 1.0 M CuBr_2 solution.
- 2) Connect the red and black leads with alligator clips to the two poles of the electric source (red to red and black to black). Set the voltage of the electric source to 3.0 volts. **DO NOT TURN ON ELECTRIC SOURCE YET!**
- 3) Connect one of the graphite rods to the black lead with alligator clips and the other graphite rod to the red lead with alligator clips. **DO NOT ALLOW THE GRAPHITE RODS TO TOUCH.**
- 4) Based on the set up of your electrolytic cell, complete #1, #2, and #3 on the diagram below.
- 5) Carefully smell (wafting) the CuBr_2 solution before electrolysis. Record observation below in #5.
- 6) Turn on electric source and allow the electrolytic cell to run for 5 minutes. **OBSERVE.**
- 7) After 5 minutes, turn off the electric source. Observe and record below any color changes to the CuBr_2 solution.
- 8) Carefully lift out the electric leads and release the graphite rods onto a piece of paper towel. Observe and record below in #5 any color changes to the graphite rods.
- 9) Based on your observations, complete #4 on the diagram below.
- 10) Carefully smell (wafting) the CuBr_2 solution after electrolysis. Record observation below in #5.
- 11) Rinse and dry the 100 mL beaker. Throw out graphite rods. Return all other materials.

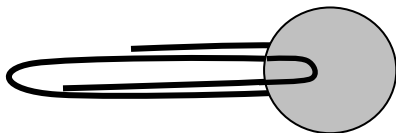
Data and Observations: DRAW THE DIAGRAM AS YOU SEE IT (LEFT SIDE VERSUS RIGHT SIDE).



#5 (Additional Observations): _____

Part-II: An Electrolytic Cell: The Electroplating of a Penny with Silver

- 1) Put on your goggles and apron.
- 2) Obtain a clean acid-washed penny.
- 3) Mass the penny. _____ g
- 4) Attach the penny to the paper clip as shown below.



- 5) Connect the paper clip with the penny to one end of the black lead with alligator clip and then connect the other end of the black lead with alligator clip to the negative pole of the electric source.
- 6) Obtain a silver strip with emery cloth. Clean the silver strip with the emery cloth.
- 7) Connect the silver strip to one end of the red lead with alligator clip and then connect the other end of the red lead with alligator clip to the positive pole of the electric source.
- 8) Set the voltage of the electric source to 3.0 volts. **DO NOT TURN ON YET!**
- 9) Half fill the small plastic container with the silver nitrate/potassium cyanide solution. **BEWARE. DO NOT TOUCH THIS SOLUTION. WASH YOUR HANDS FREQUENTLY.**
- 10) Turn on electric source. Hold the penny and the silver strip into the silver nitrate/potassium cyanide solution for 60 seconds. **PLEASE, DO NOT PUT THE ALLIGATOR CLIPS INTO THE SILVER NITRATE SOLUTION!** Turn off electric source.
- 11) Turn on the faucet to a slow gentle flow of water. **Without touching** the wet penny/paper clip or silver strip, **rinse** them in flowing water and then drop them onto a piece of paper towel, and **carefully pat** them both dry. **WASH YOUR HANDS.**
- 12) Carefully remove the paper clip and mass the penny (**BE SURE THAT THE PENNY IS DRY**).

Trial-1: Mass the penny. _____ g

- 13) Reattach the penny to a new paper clip. Re-clean the silver strip with the emery cloth. Repeat steps 7 to 12 twice more.

Trial-2: Mass the penny. _____ g Trial-3: Mass the penny. _____ g

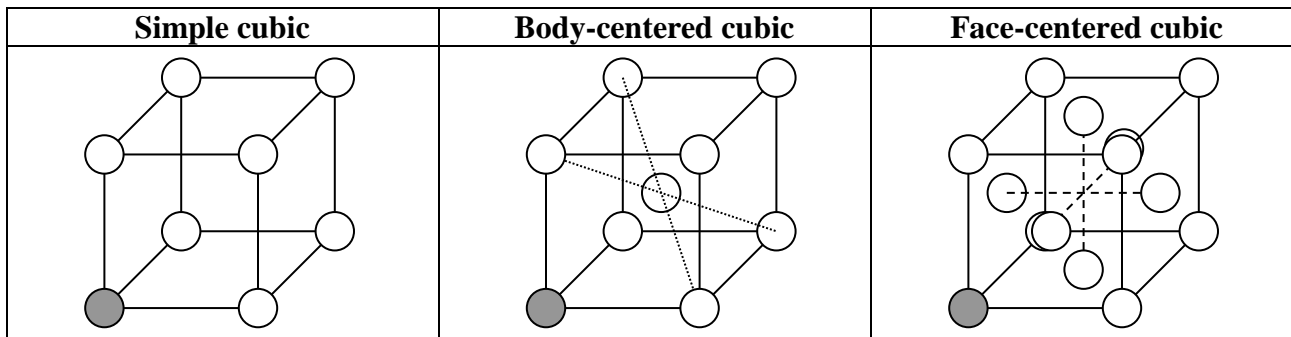
- 14) Throw out all use paper clips.
- 15) Retie and return the set of red and a black leads with alligator clips.
- 16) Return the used silver nitrate/potassium cyanide solution. **DO NOT POUR DOWN THE SINK!**
- 17) Using silver polish, clean your penny.
- 18) Wipe down your lab area with a wet paper towel. **WASH YOUR HANDS.**

Lab-26: Unit Cells: The Crystalline Structure of Metallic Solids

Introduction:

Pure metals can be thought of as spherical metallic cations held together by a sea of mobile electrons. The metallic cation spheres pack together in special and reproducible patterns. This arrangement of atoms is called the **crystal lattice**. The simplest basic repeating arrangement in the crystal lattice is called the **unit cell**. In a unit cell, the **coordination number** is the total number of metallic cation spheres that are in contact with a any given metallic cation sphere.

There are fourteen different types of unit cells, but here you are only going to study three different types: *simple cubic unit cells*, *body-centered cubic unit cells*, and *face-centered cubic unit cells*.



In each unit cell, only $\frac{1}{8}$ of the spheres at the corners are actually in the cell. The remaining $\frac{7}{8}$ of the sphere belongs to seven other unit cells. In each unit cell, only $\frac{1}{2}$ of the spheres on the face are actually in the cell. The other half of the sphere belongs to one other unit cell. In each unit cell, 100% of any sphere actually inside the unit cell belongs to that unit cell.

For simple cubic unit cells, there is a total of only 1 sphere inside the unit cell: $8 \times \frac{1}{8} = 1$.

For body-centered unit cells, there are a total of 2 spheres inside the unit cell: $[8 \times \frac{1}{8}] + 1 = 2$.

For face-centered unit cells, there are a total of 4 spheres inside the unit cell: $[8 \times \frac{1}{8}] + [6 \times \frac{1}{2}] = 4$.

The three types of cubic unit cells have characteristic coordination numbers. Simple cubic unit cells have a coordination number of 6. Body-centered unit cells have a coordination number of 8. Face-centered unit cells have a coordination number of 12.

The length of the edge (s) of each unit cell is related to the radius (r) of the spheres that are present.

For simple cubic unit cells, the relationship is: $s = \frac{4r}{\sqrt{4}}$

For body-centered unit cells, the relationship is: $s = \frac{4r}{\sqrt{3}}$

For face-centered unit cells, the relationship is: $s = \frac{4r}{\sqrt{2}}$

The general equation below relates the edge (s) of any unit cell to the radius (r) of the spheres present:

$$s = \left(\sqrt{\frac{2c\#}{3}} \right) (r) \quad \text{where } c\# \text{ is the unit cell's coordination number.}$$

A unit cell's volume (V_{cell}) is equal to its edge (s) cubed: $V_{\text{cell}} = s^3$. This volume represents the volume filled by matter (the spheres) as well as the volume of the void (the empty space in between the spheres).

A sphere's volume (V_{sphere}) is related to its radius (r): $V_{\text{sphere}} = \frac{4\pi r^3}{3}$

The three types of cubic unit cells have characteristic percents of the cube filled by matter. Simple cubic unit cells are 52.36% filled by matter. Body-centered cubic unit cells are 68.02% filled by matter. Face-centered cubic unit cells are 74.04% filled by matter.

These percentages are equal to...
$$\frac{\# \text{ spheres per cell} \times V_{\text{sphere}}}{V_{\text{cell}}} \times 100$$

The greater the percentage of the unit cell that is filled by matter, the greater the efficiency is of the packing of the unit cell. Of the three types of cubic unit cells, face-centered cubic unit cells have the highest packing efficiency. Pure metals such as aluminum, copper, lead, nickel, platinum, and silver have face-centered cubic unit cells. Pure metals such as chromium, iron, niobium, and vanadium have body-centered cubic unit cells. Oddly, plutonium is the only pure metal that has a simple cubic unit cell structure!

The density of a metal (in g/cm³) is related its type of unit cell.

$$d = \frac{[\text{Molar Mass (g/mol)}] \times [\# \text{ of Atoms per Unit Cell (\# atoms)}]}{[\text{Volume of unit cell (cm}^3\text{)}] \times [\text{Avogadro's Number (\# atoms/mol)}]}$$

Purpose: To build a unit cell and to verify the relationship between the atomic radius and the length of the edge of the unit cell. To verify the percent of the unit cell volume filled by atoms.

Procedure:

Part I: Determining the edge length of unit cells based on sphere radius.

- 1) Measure the diameter of one of the spheres and use it to find the radius of the spheres.

Diameter - _____ cm Radius - _____ cm

- 2) Calculate the edge length of a simple cubic unit cell based on the sphere radius.

_____ cm

- 3) Calculate the edge length of a body-centered cubic unit cell based on the sphere radius.

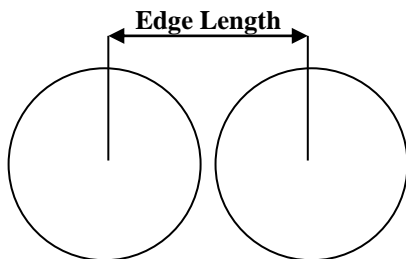
_____ cm

- 4) Calculate the edge length of a face-centered cubic unit cell based on the sphere radius.

_____ cm

Part II: Determining the edge length of unit cells based on direct measurement.

- 5) Build each cubic unit cell and measure the edge length of each cubic unit cell. Be sure to measure from the center of sphere to sphere when measuring the edge length of each cubic unit cell. **SEE BELOW.**



The edge length of a simple cubic unit cell: _____ cm

The edge length of a body-centered cubic unit cell: _____ cm

The edge length of a face-centered cubic unit cell: _____ cm

Part III: Determining the percent of unit cell volume filled by atoms in for each of the unit cells.

- 6) Calculate the volume of a simple cubic unit cell based on the direct measurement of the cell edge.

_____ cm³

- 7) Calculate the volume of a body-centered cubic unit cell based on the direct measurement of the cell edge.

_____ cm³

- 8) Calculate the volume of a face-centered cubic unit cell based on the direct measurement of the cell edge.

_____ cm³

- 9) Calculate the volume of a sphere, based on the radius you determined from the direct measurement of the sphere diameter.

_____ cm³

- 10) Based on the number of spheres in a simple cubic unit cell, calculate the volume taken up by those spheres.

_____ cm³

- 11) Based on the number of spheres in a body-centered cubic unit cell, calculate the volume taken up by those spheres.

_____ cm³

- 12) Based on the number of spheres in a face-centered cubic unit cell, calculate the volume taken up by those spheres.

_____ cm³

- 13) Based on your values from step-6 and step-10, calculate percent of volume filled by spheres in a simple cubic unit cell.

_____ %

- 14) Based on your values from step-7 and step-11, calculate percent of volume filled by spheres in a body-centered cubic unit cell.

_____ %

- 15) Based on your values from step-8 and step-12, calculate percent of volume filled by spheres in a face-centered cubic unit cell.

_____ %

Lab-27: Freezing Point Depression and Molar Mass Determination

Introduction:

Solutions are homogeneous mixtures that contain two or more substances. The major component is called the solvent, and the minor component is called the solute. Since the solution is composed primarily of solvent, the physical properties of a solution resemble those of the solvent. Some of these properties, called colligative properties, are independent of the nature of the solute and depend only on the concentration of the solute.

Freezing point depression is one of the colligative properties. A solution will freeze below the normal freezing point of the pure solvent in direct proportion to the molal concentration of the solute. This relationship is expressed by the equation:

$$\Delta T_f = K_f m i$$

where ΔT_f is the lowering of the freezing point, K_f is the molal freezing point constant of the solvent, m is the molal concentration of the solute, and i is the van't Hoff factor of the solute. The van't Hoff factor for a nonelectrolyte solute is always equal to one ($i = 1$) because one mole of a nonelectrolyte will form one mole of dissolved solute particles. The van't Hoff factor for an electrolyte solute is ideally equal to the number of ions that compose the electrolyte. However, the actual van't Hoff factor for an electrolyte is always somewhat less than its ideal value. This is due to ionic interactions that exist in the solution which cause cations and anions to behave as a single dissolved particle instead of separate dissolved particles. This effect becomes more pronounced as the concentration of the solute in the solvent increases.

A practical application of the freezing point depression effect is addition of antifreeze into the water in car radiators. The function of a car's radiator is to cool the engine. The alcohol, ethylene glycol - $C_2H_4(OH)_2$, is antifreeze. Early in the automobile industry, engineers needed to find something to add to the water in the radiator that would greatly lower the freezing point without damaging the radiator itself. An alcohol such as ethylene glycol is a good choice for several reasons. (1) It is a nonelectrolyte. (2) It is fairly nonvolatile (has a fairly low vapor pressure). (3) It is highly miscible in water. (4) It is non-corrosive to metal.

A second practical application of the freezing point depression effect is addition of salt to icy sidewalks and roads. Traditionally, halite ($NaCl$) was used. When spread onto icy sidewalks or roads, it begins to dissolve into the water and lower the water's freezing point. Thus the ice melts and stays liquid even below zero degrees celsius. Halite is an electrolyte with an ideal van't Hoff factor of two. One mole of Halite ideally produces two moles of ions when dissolved in water. Halite's solubility in water is fairly constant, even at temperatures at or below zero degrees celsius.

Today, a product known as supermelt ($CaCl_2$) is used on many sidewalks. Supermelt is an electrolyte with an ideal van't Hoff factor of three. One mole of supermelt ideally produces three moles of ions when dissolved in water. Mole for mole, supermelt lowers the freezing point more than halite.

A third practical application of the freezing point depression effect is ability to determine an unknown substance's molar mass. Using a known amount of a known solvent, you can determine the molar mass of a known amount of an unknown solute by measuring the freezing point depression of the known solvent.

One drawback to this method of determining molar mass is the relatively small change in the freezing point and the limitation of the thermometer. It is necessary to form an almost saturated solution and then to make the thermometer readings very carefully.

Purpose:

In this experiment, you will determine the molar mass of an unknown sugar.

Procedure: GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!

Part 1: DETERMINING THE FREEZING POINT OF THE DISTILLED WATER

- 1) Obtain supermelt in a 600 mL beaker.
- 2) Obtain a plastic container. Fill the container half-full with ice and then pour the supermelt over the ice to create an ice bath with a temperature below minus 8 degrees celsius. [**thermometer #1**] *Add more ice if necessary.*
- 3) Measure out 30 mL of distilled water into a graduated cylinder and then add it to the large test tube.
- 4) Determine the freezing point temperature of distilled water. [**thermometer #2**] *This occurs when "slushy" ice crystals appear in the test tube. _____°C*

Part 2: DETERMINING THE FREEZING POINT DEPRESSION OF THE SUGAR-WATER SOLUTION

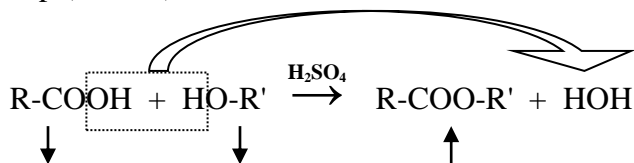
- 1) Obtain and mass a second large dry test tube and 250 mL beaker. _____ g
- 2) Add about 8 g of unknown sugar to test tube [**about 2 level spoonfuls**]. It need not be exact but it should be close.
- 3) Mass the sugar, test tube, and 250 mL beaker. _____ g
(The mass of sugar in the test tube should be somewhere between 8g and 9g.)
- 4) Measure out 30 mL of distilled water into a graduated cylinder.
- 5) Insert the thermometer into the test tube with the sugar. Wiggle the thermometer down into the bottom of the test tube.
- 6) Add the 30 mL of water to the test tube with sugar. **IMMEDIATELY**, begin stirring the sugar water mixture. **DO NOT ALLOW THE SUGAR TO SIT OR IT WILL HARDEN. KEEP STIRRING. THIS MAY TAKE 5 TO 8 MINUTES OF STIRRING TO DISSOLVE ALL OF THE SUGAR.**
- 7) Mass sugar, water, test tube, and 250 mL beaker. _____ g
- 8) Make sure ice bath has a temperature below minus 8 degrees celsius. [**thermometer #1**]
- 9) Determine the freezing point temperature of sugar-water solution. [**thermometer #2**] *This occurs when "slushy" ice crystals appear in the test tube. _____°C*
- 10) Thoroughly wash all equipment. Sugar and supermelt require a lot of water to rinse off. Return all equipment.
- 11) Molality is defined as the moles of _____ per the kilograms of _____!
- 12) Based on your data, determine the molar mass of the unknown sugar. **SHOW WORK BELOW.**

The molar mass is _____ g/mol

Lab-28: Esterification

Introduction:

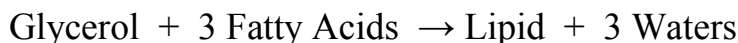
Esterification is the process by which an organic (a carboxylic) acid and an alcohol react to produce an ester and water. Many esterification reactions are catalyzed by sulfuric acid (H_2SO_4) which is a good dehydrating agent. Esterification is an example of dehydration synthesis. The water produced by esterification is formed when the "H" from the alcohol's hydroxyl group (OH) combines with the "OH" from the acid's carboxyl group (COOH).



The general formula for an ester is RCOOR' or $\text{R-COO-R}'$ where the first part (RCO) is contributed by the organic acid and the second part (OR') is contributed by the alcohol. Thus an ester's formula is typically written with the organic acid part first and the alcohol part last. In the IUPAC system, esters have two names. The first name ends in "yl" and names the alcohol part of the ester. The last name ends in "oate" and names the organic acid part of the ester.

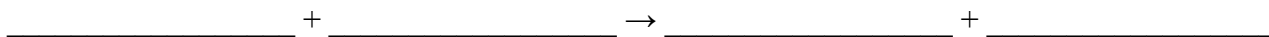
Esters made from alcohols and organic acids of low molecular mass are compounds that have agreeable fruity or minty odors and flavors. Many odors in flowers come from esters. For these reasons, natural and synthetic esters are used in the perfume, candy, and soft drink industries.

Lipids (fats, oils, and waxes) are esters of the alcohol - glycerol - and three fatty acids. A fatty acid is simply a long-chain organic acid in which the "R-group" is usually at least 11 carbon atoms long.



Questions for Thought:

1) Complete the equation below for an esterification reaction by filling the blanks with words.



2) Place a **box** around the "OH" and a **circle** around the "H" that will form the water during esterification.



3) Write the molecular formulas for the esters that would be produced from the reactions in question 2.

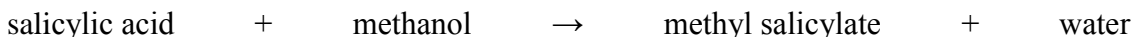


4) Write the IUPAC names for the esters written in question 3.



5) Why are a few small drops of sulfuric acid often used when producing an ester from the reaction of an organic acid and an alcohol?

- 6) The structure for salicylic acid can be found on page 1088 in your textbook. **Neatly**, write a balanced equation using **structural formulas** for the formation of methyl salicylate.



Purpose: To produce an ester and study several of its characteristic properties.

Procedure: **GOGGLES AND APRONS ARE TO BE WORN AT ALL TIMES!**

- 1) Put on your goggles and aprons.
- 2) Create a hot bath using 150 mL of water in a 250 mL beaker. Turn hot plate to high.
- 3) Obtain a sample of salicylic acid on a watch glass. Observe (phase, color, texture, and odor) the salicylic acid and record those observations in the appropriate space below (#1).
- 4) Obtain a sample of methanol (≈ 10 mL) in a small beaker. Observe (phase, color, and odor) the methanol in the small beaker and record those observations in the appropriate space below (#2).
- 5) Put the salicylic acid in the 125 mL Erlenmeyer flask. Add 2 dropper pipettes full of the methanol into the Erlenmeyer flask to cover and to dissolve the salicylic acid. Swirl the Erlenmeyer flask to dissolve the salicylic acid. **Very carefully** add two drops of concentrated sulfuric acid. Carefully swirl the Erlenmeyer flask to mix the contents.
- 6) Place your Erlenmeyer flask in the hot water bath and heat the contents until a white solid or light brown liquid appears. Using the crucible tongs, IMMEDIATELY remove the Erlenmeyer flask from the hot plate and allow the contents to cool. Scrape out the contents of the Erlenmeyer flask (if possible) and place it on the watch glass. Observe (phase, color, texture - solid, and odor) the material on the watch glass and record those observations in the appropriate space below (#3). **WASH YOUR HANDS!**
- 7) Use the remaining methanol to clean out the Erlenmeyer flask. Leave your station as you found it.

Data and Observations:

- 1) Observe and describe the salicylic acid before mixing and heating it with the methanol.

- 2) Observe and describe the methanol before mixing and heating it with the salicylic acid.

- 3) Observe and describe the product of the salicylic acid and methanol mixture after heating.

