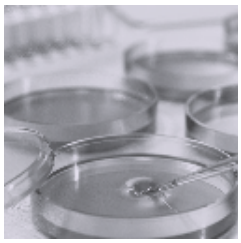


TABLE OF CONTENTS

LAB INFORMATION

EXPERIMENTAL PROCEDURES

Please note that several supplementary lab materials will be provided online by your course instructor. Individual labs will make reference to these materials as needed. This has been done to reduce the amount of paper used in the lab manual, and thus the cost of the manual for you.

Name: _____

Lab Partner: _____

Room Number: _____

Section: _____

Locker Number: _____

Locker Combination: _____ - _____ - _____

TA Name: _____

TA's Email: _____

Also, take a picture of this information to have it handy for any emergency.

GENERAL LABORATORY RULES

Be considerate of the other students in your laboratory section and try to share the space and equipment in a safe and courteous way. Disturbances such as running, loud noises, and practical jokes can lead to serious accidents and must be avoided. Any student who refuses to comply with requests from a TA or who is repeatedly involved in improper lab conduct will be asked to leave the laboratory and will not be allowed to finish lab work for that experiment. Continued and repeated violations of laboratory rules may result in expulsion from the course.

Do not begin laboratory work if your TA has not finished lecturing. Also, please be aware of those around you so that you will be prepared if an accident happens. PLEASE notify your TA immediately if there is an accident or broken glass! Also, **YOU SHOULD ALWAYS WEAR YOUR SAFETY GLASSES IN LAB DURING EACH AND EVERY LABORATORY EXPERIMENT DAY**, regardless of what procedures you are currently performing; someone near you may be doing something hazardous. Continue wearing safety equipment even though you have finished with your own experiment.

PLEASE CLEAN UP YOUR LAB BENCHES WHEN YOU HAVE FINISHED WITH YOUR WORK. You must leave the laboratory in as good of condition as when lab started. This includes returning all supplies to their proper place and cleaning all areas in which you worked. You may lose points on your laboratory report if you do not leave your laboratory space clean.

Waste chemicals must be disposed of properly. For most of the experiments there will be special containers for the various materials, and your lab TA will tell you the proper way to dispose of each waste. Never pour potentially toxic waste down the drain! If you are uncertain if something is safe to pour down the drain, PLEASE ask your TA.

SAFETY

You must follow good safety practices while in the laboratory in order to prevent accidents. Our department has a very good laboratory safety record—it is extremely unlikely that there will be a serious accident if you learn and follow some simple precautions. These rules are for YOUR SAFETY!

- Before you are allowed to do any laboratory work, you will have to pass a laboratory safety quiz with a grade of 60% or higher. It will be online and available for you to complete from home. If you fail to achieve a 60% or higher on the online quiz, the lab coordinator will give you a paper quiz. Once passed you will be allowed to attend lab. **Note: The paper quiz *will not* replace your original lab safety quiz grade.**
- You must wear approved safety glasses that have protection from all directions. Safety Glasses must be worn whenever anyone is doing laboratory work in the room. ABSOLUTELY NO EXCEPTIONS TO THIS RULE! This is by far the most important safety precaution that you can take in ANY laboratory environment. Almost any injury to the body that could happen in the laboratory will heal, except eye injuries. Your eyes are absolutely irreplaceable and priceless: YOU NEED TO PROTECT THEM AT ALL TIMES!

NOTES >>>

- You must wear shoes with a closed toe and closed heel in the laboratory at all times. Closed-toe and closed-heel shoes can protect your feet from broken glass on the floor, and chemicals that are spilled on the floor that can be absorbed through the skin. If you elect to wear sandals or other open-toe shoes during the rest of your day, then PLEASE PLAN TO BRING A PAIR OF CLOSED-TOE AND CLOSED-HEEL SHOES WITH YOU TO USE DURING YOUR LAB EXPERIMENTS! You will be asked to leave the lab if you do not comply with this rule.
- You will need an approved laboratory coat or apron to work in the labs. This coat or apron provides an important layer of protection in case of a chemical spill.
- You must wear pants or long bottom clothing when working in the lab. Your TA will go over the length during the safety presentation at the beginning of the semester.
- Long hair should be secured in a bun or ponytail (if short enough) and your hair and/or jewelry should never dangle onto the lab benches and experimental areas. The lab coat or apron should fit and not be too large. (The lab aprons can be adjusted.)
- Learn in advance the location of fire extinguishers, safety showers, eyewash fountains, and other emergency equipment so that they may be used quickly when needed. Your TA will point out the location of these items.
- Never taste any chemical in the laboratory. Always use a rubber bulb to draw liquid into a pipet. You should never eat or drink while in the lab; food may easily become contaminated.
- Move carefully and deliberately in the lab; avoid bumping anyone. This is especially important when you are carrying chemicals or glassware to and from the stockroom. Students with disabilities that might make it difficult for them to handle chemicals and equipment safely must consult the lab coordinator before laboratory work begins.
- ALWAYS WASH YOUR HANDS AND YOUR ARMS THOROUGHLY AFTER YOU FINISH YOUR LAB WORK. It is possible for you to get chemicals on your hands without realizing it. Wash your hands before you touch your face to avoid getting chemicals in your eyes or mouth.
- Wipe the lab bench area where you worked with a paper towel moistened with water and a drop of detergent.

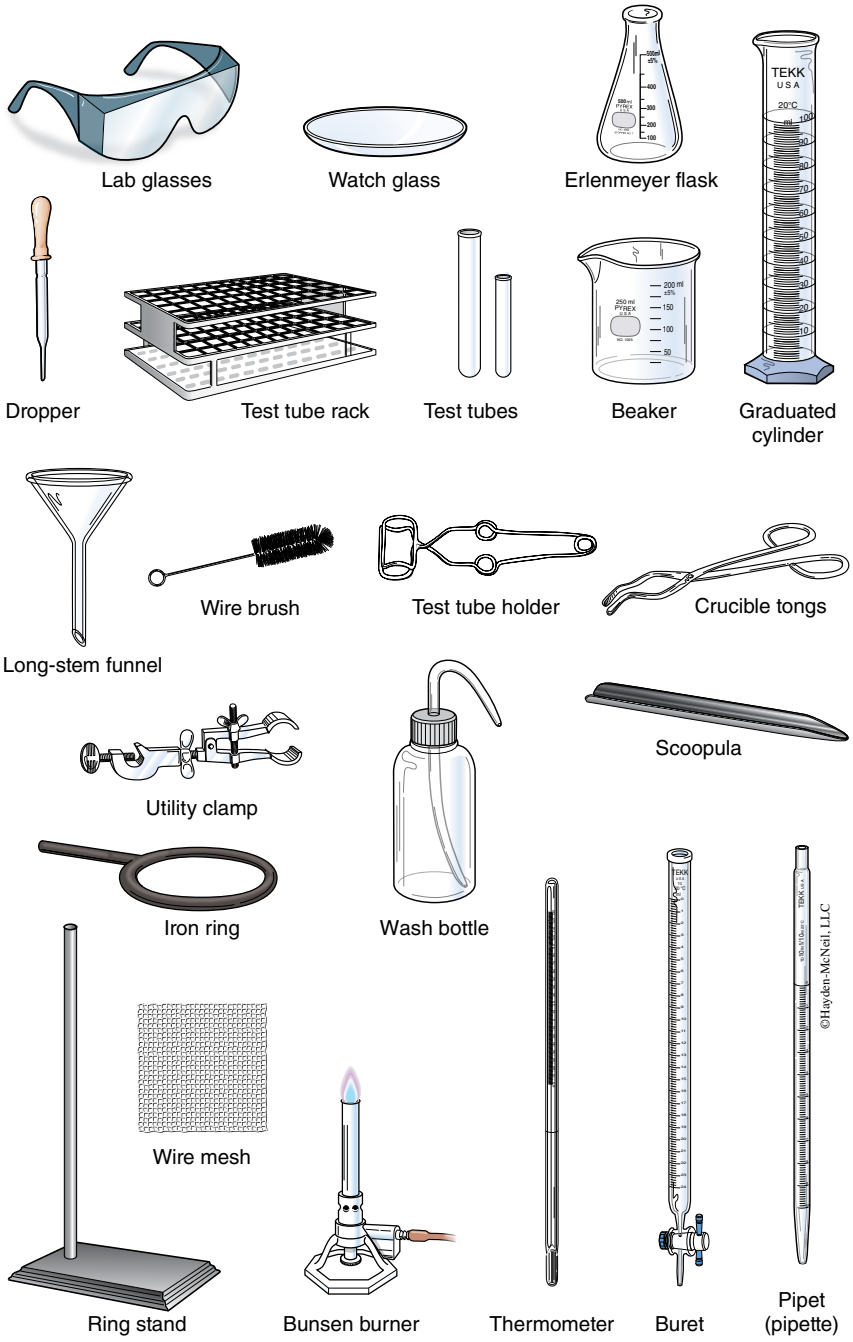
NOTES >>>

FIRST AID

- If you get chemicals in your eyes, immediately wash them with a large amount of water. The eyelids should be pulled back while the eye is thoroughly washed with water. An eyewash fountain is available in every laboratory. However, you can use the closest sink/water source for this. Someone else should notify the TA of the accident IMMEDIATELY.
- Many chemicals will injure or burn the skin. Also, many chemicals can be absorbed into the body through the skin. If you get significant amounts of any chemical on your body or clothes, wash with large amounts of soap and water and notify your TA.
- If your clothes begin to burn, remove burning clothing immediately OR immediately roll on the floor to try to extinguish the flames. Other students should help by patting the area with their hands and by wrapping a fire blanket (or jacket. etc.) tightly around the burning area. As soon as the flames are extinguished, use water to quickly cool the skin to minimize the injury.
- Report all injuries (cuts, burns, etc.), no matter how minor, to your TA immediately. A safety report must be filled out for every accident or safety incident.

NOTES >>>

COMMON LABORATORY EQUIPMENT



NOTES >>>

DIRECTIONS FOR THE LABORATORY NOTEBOOK

The laboratory notebook is essential in industrial and academic research labs. It is the primary record of experimental results and provides a means for communicating these results between co-workers. It also plays an essential role in legal issues such as establishing patent rights to an invention or resolving questions of scientific misconduct. One of the goals of the Chem 200/202 lab is for you to begin to learn how to keep a proper scientific notebook as well as to report the results of experiments in an effective and professional manner.

The notebook used for this class contains 40 bound pages, each with a duplicate page (marked "copy") that provides a direct copy of what you write on the original page. The notebook is in the back portion of this lab manual.

At the end of each lab period you should sign and date the bottom of each lab page you used that period. Then, tear off the copy pages and turn them in to your TA. The copy pages will be the official record of your laboratory work and will be used to resolve any disputes that may arise. Make sure you press hard when writing in your laboratory notebook, so your information transfers to the carbon copy that your TA will need to read.

Several times during the semester, your TA may collect your notebook for grading. They will expect to see the following:

TABLE OF CONTENTS

The first page of your notebook should be reserved for a table of contents. This is where you list each experiment and the pages in your notebook where the results for that experiment are recorded.

PRE-LAB

Before each experiment you should prepare a pre-lab page that contains the following:

- The title of the experiment.
- A brief statement about the purpose of the experiment (in your own words).
    - Why am I doing this experiment?
- An outline of the procedures to be used (in your own words).
- A fun fact about some part of the experiment (i.e., techniques, equipment, solutions being used, etc.).
- A description of any safety hazards and the appropriate precautions that should be used (in your own words).

An example of a proper pre-lab is given on page 10.

Your TA will collect the non-copy page of the pre-lab at the beginning of the lab period when the experiment is to be performed. If the page is not complete before you come to lab, 5 points will automatically be deducted from your score.

NOTES

EXPERIMENTAL RESULTS

All experimental results and other pertinent facts should be recorded directly in the lab notebook with a pen, not a pencil. You should use ink because it is a legal document that should not be altered. If you do need to change any information, you should cross it out and initial it, then very clearly make the correction; even for a single number. Never tear pages from the notebook, or the notebook becomes invalid. Writing results on a separate piece of paper then copying them into the notebook is not proper technique and won't be allowed. If your TA sees you using loose paper, they will confiscate it and you will have to start the experiment again. Given this, perfect neatness is not expected. However, your writing should be legible and the data properly labeled. Your goal should be to put in enough information that someone familiar with the experiment can look at your notebook pages and, without talking to you, figure out what is what.

For example, in the eighth experiment you will be using an instrument called a STAR spectrometer to directly see the colors of the radiation emitted by individual transition in the gas. The experimental data will consist of colors of the spectrum, the wavelengths, and the intensities of the light seen. This means that in your notebook you should have a column labeled "Colors of the Spectrum," and next to it a column of numbers labeled "Wavelength," and next to it a column labeled "Intensities." The information should be written neatly enough that the data can be read and the reader can tell which color is associated with which wavelength and intensity. Note that it is also important for you to include units with your numbers. If all the units are the same this can be done in the column headings.

In addition to any quantitative data, all other facts relevant to your experiment should also be recorded in your notebook as observations. This include the names of any lab partners and the role they played in the procedure, any changes to the procedure from what was recorded in your pre-lab, information on sample preparation (dilutions, amounts weighed out, etc.), and any problems that occurred.

NOTES >>>

DIRECTIONS FOR THE EXPERIMENT REPORTS

A formal, type-written report for each laboratory experiment will be required. These will be due at the beginning of the lab period one week from the date that the experiment is finished. Late lab reports will receive an incomplete.

A sample lab report is given on pages 11–12. Your lab report should follow this general format:

PROCEDURE

In the beginning of your lab report, there should be a procedure section where you refer back to the notebook and list any changes to the procedure from your Pre-lab.

EXPERIMENTAL DATA

The experimental data comes next. In some cases it will be appropriate to present your data in a table or tables, in other cases the data should be presented in graphs. The appropriate choice will be indicated in the experiment description. Units for numbers should be indicated in the column headings of a table and the axis labels of the graph. Graphs must be produced using a computer program such as Excel. This program is generally available on the campus computers and it is to your advantage to learn how to use it. You can find many helpful videos on YouTube on how to prepare an Excel graph or you may go to the Help Room, GMCS-212, to ask a TA for further help. **Additional tips on graphs are presented on pages 13–14.**

CALCULATED RESULTS

After the experimental data comes a summary of any calculated results. In most cases these should be presented in table format.

SAMPLE CALCULATIONS

Following the calculated results, there should be a section where you write down sample calculations for each type of calculation you need to do for one trial. If you have more of the SAME calculations you could just show the results. These may be handwritten.

DISCUSSION/QUESTIONS

At the end of your lab report comes a discussion section. The discussion section should begin with a sentence that concisely summarizes your results as well as answer the questions at the end of each experiment. For example, let's say the experiment involved measuring the density of pennies from 1987. A concise summary of your results would be something like:

> *In this experiment, I determined the density of the 1987 pennies was 7.9 g/mL.*

This should be followed by a discussion of the quality of this result. First, are the results reasonable based on your understanding of the chemistry involved? Why or why not? Continuing the previous example, you might want to say:

> *The density of pennies from 1987 was 7.9 g/mL. which is a reasonable value since the penny is a mixture of copper (d = 8.92 g/mL) and zinc (d = 7.14 g/mL).*

NOTES >>>

In many cases you will have made multiple measurements and report the average results. If the different measurements give similar values this should give you greater confidence in the precision of your results, but if they vary considerably this would be a cause of concern you should mention in your discussion.

> *Furthermore, two separate measurements gave density values that were quite close, 7.9 g/mL and 8.1 g/mL, suggesting that my results are precise.*

In some cases, you may be able to compare your results to literature value. For example, if I obtained a density of a 1987 penny, you can look up the density for that year in your text, so the sentence could be:

> *However, the text lists the density of the 1987 penny as 6.8–7.6 g/mL, which is almost +0.3 more than the value I determined. This indicates that despite the consistency of our values, the result is not very accurate, since my value is not in the range for the density of a 1987 penny.*

Next, you should discuss possible sources of error in the experiment and how this may have affected your results. Of course, if your results are way off, as in the density of pennies example, the first thing you should do is re-check your calculation! If you are still way off, you will want to pay particular attention to the error discussion in your report. However, even if your results are quite good, you will still need to include an error discussion because there will always be some error. To think about the error, what you want to do is to go over exactly what you did in the experiment. Each time you made a measurement, there will be an error associated with that measurement. For example, suppose you have to measure out a certain volume of liquid—even if you were very careful, how accurately could you have done that with the equipment you used? ±5 mL, ±1 mL, ±0.2 mL? All the experiments involve reading a number off a scientific instrument. Even if the instrument is working perfectly, the number will be ±1 in the last digit. The instrument also will need to be calibrated correctly, What if this was not done properly, whether because of an instrument problem or operator error or the calibration solution was the wrong concentration, how will this affect your results? In addition to errors associated with making measurements, errors will also result from contamination due to dirty glassware, improper handling of stock solutions and so on.

An OK discussion error will list the possible sources of error for the particular experiment being done. A good discussion of error will indicate how the error affects the measurements. For example, an OK error discussion for the density of pennies example might be:

> *The fact that my density of the 1987 pennies is much larger than the text value indicates that there were sources of error in this experiment. Possible sources of error could be the pennies were not dry enough, there could have been dirt on the pennies that lead to the density being higher, or I lost water in the 100 mL graduated cylinder by not properly sliding the pennies into the water.*

NOTES >>>

A good error discussion might be:

> *The fact that my density of the 1987 pennies is so much larger than the text value indicates that there were large sources of error in this experiment. The reason we got a higher density value could be the pennies may have dirt and other particles from being used for currency for so many years. Another reason the density of the pennies could be high was that we did not dry the pennies enough. Not allowing the pennies to properly dry allowed for the mass to increase. We had great difficulty in making sure we did not "drop" the pennies into the 100 mL graduated cylinder. We could have possibly lost volume due to not properly sliding the pennies into the glassware. Given this, I believe the pennies having been handled by previous people and quite possibly having dirt on them was the greatest source of error in this experiment.*

Included in your discussion, should be the questions that you will be asked for each experiment. Your TA might want you to footnote or number the answers to the questions in your discussion. This method is to make it easier for your TA to identify which question you are referring to in your discussion.

NOTES >>>

EXAMPLE OF PRE-LAB

MOLAR MASS OF AN UNKNOWN ACID

Pre-lab (In the Notebook)

PURPOSE

The purpose of the experiment is to determine the molar mass of an unknown acid by titration with a standardized base.

PROCEDURE

After obtaining a solid sample of an unknown acid, approximately 3.2 g will be distributed among three 125-mL Erlenmeyer flasks. Using analytical techniques, the precise mass of acid in each flask will be determined by recording the mass of each flask before and after the addition of approximately 1/4 the total acid sample.

Each of the three acid samples in the flasks will be dissolved in DI water (approximately 25 mL) and 2–3 drops of phenolphthalein will be added prior to titration.

Each dissolved acid sample will be titrated with the sodium hydroxide solution standardized from the previous experiment: $[OH^-] = 0.1078$ M.

FUN FACT

The technique of titration is used in the process of biodiesel to help remove the fatty acids in water vegetable oil (WVO) that would normally react to make soap instead of biodiesel. A sample titration process is done to help determine the acidity of a sample of WVO so the rest of the batch can be neutralized properly.

SAFETY

Sodium Hydroxide, NaOH, is a strong base and may cause severe chemical burns. Avoid contact with skin and eyes. Given that the formula of the acid is unknown, the same safety precautions should be applied to the acid. As always, goggles, lab apron/coat, pants, and closed-toe shoes will be worn at all times.

NOTES >>>

EXAMPLE OF EXPERIMENT REPORT

Iwanna A. Realbad
Lab Partner: Bea S. Okwithme 3/03/2014

CHEM 200, EXPERIMENT 6
MOLAR MASS OF AN UNKNOWN ACID

PROCEDURE

See the pre-lab report on page 25 of my laboratory notebook for an outline of the general procedure. In the actual experiment, the following changes to the written procedure were made: 1/3 of the sample was placed in each of the three flasks instead of the 1/4 and the sample was dissolved in approximately 40 mL of DI water instead of 25 mL.

EXPERIMENTAL DATA

Sodium hydroxide standardized from the previous experiment: $[OH^-] = 0.1078$ M

Trials:	Trial One:	Trial Two:	Trial Three:
Mass of flask and acid (g)	102.580	103.022	101.952
Mass of empty flask (g)	101.329	102.085	100.946
Mass of acid in flask (g)	1.251	0.937	1.006
Final buret reading (mL)	23.70	18.02	36.80
Initial buret reading (mL)	0.13	0.51	18.02
Volume of base added (mL)	23.57	17.51	18.78

CALCULATED RESULTS

Trials:	Trial One:	Trial Two:	Trial Three:
Moles of acid (mol)	2.541×10^{-3}	1.888×10^{-3}	2.024×10^{-3}
Molar Mass of acid (g/mol)	492.3	496.3	497.0
Average Molar Mass (g/mol)	**495.2**		
Standard Deviation	**± 2.5**		

SAMPLE CALCULATIONS

(May be Handwritten)

For Trial One:

Mass of acid in flask:

$$102.580 - 101.329 = \mathbf{1.251\ g}$$

Volume of base added:

$$23.70 - 0.13 = \mathbf{23.57\ mL}$$

Moles of acid:

$$23.57\text{ mL NaOH} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.1078\text{ mol NaOH}}{1\text{ L}} \times \frac{1\text{ mol of acid}}{1\text{ mol of NaOH}} = \mathbf{2.540846 \times 10^{-3}\ mol}$$

NOTES >>>

Molar Mass of acid:

$$\frac{1.251\,\text{g}}{2.54084610^{-3}\,\text{mol}} = \mathbf{492.325856\ g/mol}$$

Average Molar Mass:

$$\frac{492.3 + 496.3 + 497.0}{3} = \mathbf{495.2\ g/mol}$$

Standard Deviation:

$$\sqrt{\frac{(492.3-495.2)^2 + (496.3-495.2)^2 + (497.0-495.2)^2}{3-1}}$$

$$= \mathbf{2.\underline{5}3574467}$$

Answer:

495.2 ± 2.5 g/mol

DISCUSSION/QUESTIONS

The molar mass of the unknown acid was determined to be 495.2 ± 2.5 g/mol. This is a reasonable value for a molar mass of a molecular compound. In addition, all three trials gave quite similar numbers, making my three trials very precise among one another.

However, although no major problems were encountered during this experiment, I expect that the true molar mass is somewhat higher than what I have determined. Because the phenolphthalein endpoint is so faint and comes so quickly, it is very easy to add too much sodium hydroxide to the flask. When too much base is used in the titration, the calculated molar mass of the acid is too low (the molar mass of the acid is inversely proportional to the moles of base used).

Given that my determination of the acid's molar mass is probably too low, I believe that my third trial is very close to being correct. I was able to stop this titration at the desired "ultra-faint" shade of pink. In contrast, I believe that my first trial is probably the most incorrect given the dark color of the phenolphthalein when I stopped the titration.

There is some chance that the concentration of the sodium hydroxide can react with carbon dioxide (in the air) to produce sodium bicarbonate. If this happened it would mean that the true molar mass of the unknown acid is even higher than I have determined it to be.[1] If the hydroxide solution is less concentrated than I determined it to be in (due to the reaction with the air) then I will have used too much of it in the molar mass titrations.[2] The effect on the molar mass calculations is the same as overshooting the endpoint.[3] Great care was taken to keep the bottle of sodium hydroxide tightly capped at all times and to use the solution in the buret as quickly as possible. For these reasons, I believe that the value used for the concentration of sodium hydroxide was correct.

NOTES >>>

1. Refers to Question 1
2. Refers to Question 2
3. Refers to Question 3

GRAPHING DATA

WHY ARE GRAPHS IMPORTANT?

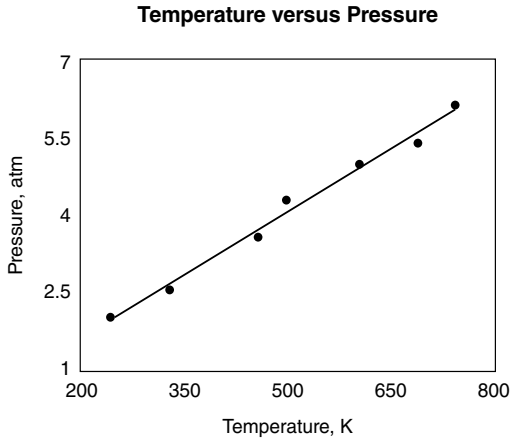
Frequently, data are collected as pairs of numbers, such as the pressure of a gas at several values of temperature. It is natural to present such data in a table, but it is often more useful to present the data as a "scatter graph." With a graphical presentation it is often easier to see patterns in the data, which might not be obvious from the table. You may also spot data that doesn't seem to "fit" with the rest of the data, and such data should be suspected of being an error. The best thing to do in such a case is to collect more data in that region.

LINEAR GRAPHS

Suppose that the following data are collected for a constant amount of gas at a constant volume. From Table 1, one might **guess** that there is a linear relationship between the temperature and pressure, but there is **no doubt** of that conclusion when the graph is seen. The table and the graph contain exactly the same information, but they serve different purposes. The data can be more accurately read from the table, but the relationship between data points can be more easily seen in the graph.

Table 1.

Temperature, K	Pressure, atm
233	2.1
321	2.6
453	3.6
497	4.3
604	5.0
688	5.4
743	6.1



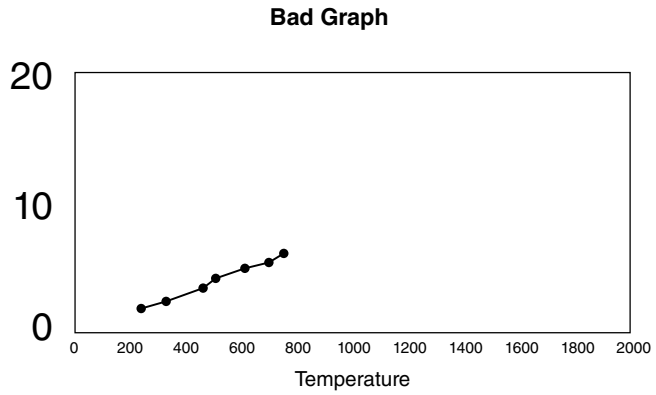
The graph shown above is a good example of a "good" graph. There are several features that make it so:

- The scales of the x- and y-axis are such that the data points are spread across the entire graph.
- Title is labeled as well as both x- and y-axis are labeled and the labels include units.
- The font size for the axis scale and labels is appropriately sized so the numbers and letters are easy to read and are well proportioned relative to the size of the graph.
- The data points are plotted as individual points and lines **do not** connect the points. This is called an **"x, y scatter" plot in Excel**. Always choose this type of plot if you use Excel to graph your data. Never choose a "line" plot even if you expect the plot to be linear. (The "line" plot assumes equal spacing between points on the x-axis and this is generally not correct.)

NOTES >>>

- The line that is shown is the best fit linear line through the data points (a "trend line" in Excel). **Note:** It does not connect the points.

Below, the same data is shown in a "bad" graph.



- The features that make this a "bad" graph are as follows:
- The scales of the x- and y-axis are not properly set so all the data is crunched together and most of the graph space is empty.
- There is no label for the y-axis.
- The x-axis has a label, but no units are given.
- The fonts are the wrong sizes: the font used for the y-axis scale is way too big and the x-axis is way too small.
- Lines connect the data points.

NOTES >>>

ERROR ANALYSIS

Most measurements are neither exact nor precise. Exceptions include: exact counts (e.g., the number of students in a room) and numbers, which are exact by definition (the speed of light is defined as exactly equal to 299,792,458 m/s). Some numbers (such as π) can be calculated to as many significant figures as we wish, and many conversion factors (such as 1000 mL/L) are also exact. However, almost all numbers with which you will work are uncertain to some extent.

ERRORS IN MEASUREMENT

For any quantitative measurement, there is a correct or "true" value. We expect that any measured values reported in the literature are "true" values and when the measurements are reproduced in more than one laboratory, we become even more confident of the accuracy of that measurement. If we know the "true" value, we can report our error as:

$$\text{Error} = (\text{measured value}) - (\text{"true" value})$$

Other quantities, which are sometimes reported, are the relative error—the error divided by the measured value, and the percent error—the relative error expressed as a percentage (multiplied by 100).

If we make one measurement and find that the error is significant, we would say that there is a discrepancy between our result and the literature value and such a discrepancy could be caused by a number of reasons:

1. You are a novice and you made some mistakes in the measurement or the analysis.
2. Your apparatus is not calibrated correctly and there is systematic error resulting from this.
3. The "true" value is actually incorrect!
4. The error is a random fluctuation.

If the first point applies to you, checking both the procedure and the calculations is important. In fact, even experienced scientists can make mistakes and know that they have to continuously check their work.

The second point can be corrected by assuring that your instrument is well calibrated. This involves making a measurement on a standard sample, that is, a sample for which very careful and reproducible measurements have been made. An experiment, which shows how systematic error in volumetric measurements can be reduced through calibration of volumetric glassware, is included in this manual.

It is always possible that the "true" value is incorrect and if you are very confident that you have not made any mistakes and your instruments are properly calibrated, point three should be considered as a possibility. Scientists sometimes disagree in their results and if the measurement is very important, a controversy will eventually right itself by re-evaluation of the results by many people.

The fourth possibility for error: random fluctuations in the environment (such as temperature fluctuations or vibrations) can be reduced by developing higher quality instruments and being more careful, but, ultimately, random error will always exist to

NOTES >>>

some extent. The best method for reducing any residual random error is to take multiple measurements. If the error is truly random, the set of measurements will deviate randomly around the "true" value. Then, the average will be the best estimate of that value and the more measurements that are done the closer the average will be to the "true value" (assuming no systematic error). The range of values obtained with several measurements is used to estimate the uncertainty of the result and this range can be mathematically expressed by one value, the standard deviation, σ. The standard deviation is given by the following equation:

$$\sigma = \sqrt{\frac{\sum_i (x_i - \overline{x})^2}{n-1}} \text{ where } \overline{x} = \frac{\sum_i x_i}{n}$$

Where x_i is the value of a particular measurement and $\overline{x}$ is the average of "n" measurements. To use the equation for σ, first calculate the average and then subtract the average from the value of each measurement. Fortunately, most calculators are able to calculate the average and standard deviation of a set of numbers with a function key. Try out your calculator with these measured values for the melting point of naphthalene: 80.6°C, 80.7°C, 81.2°C. You should determine that $\overline{x} = 80.8$°C and σ = 0.3°C. The final result would be written as 80.8 ± 0.3°C. This means that we are very confident that the "true" value is between 80.5°C and 81.1°C. The mean (with the bar on top of it) is the sum of: (80.6 + 80.7 + 81.2) divided by THREE, which equals 80.8. Since we only have three sample numbers (80.6, 80.7, and 81.2), therefore n = 3. Sigma (σ) is the Greek symbol for standard deviation, so your calculations would look like this:

$$\sigma = \sqrt{\frac{(80.6 - 80.8)^2 + (80.7 - 80.8)^2 + (81.2 - 80.8)^2}{3 - 1}}$$

$$\sigma = \sqrt{\frac{(-0.2)^2 + (0.1)^2 + (0.4)^2}{2}}$$

$$\sigma = \sqrt{\frac{(0.04) + (0.01) + (0.16)}{2}}$$

$$\sigma = \sqrt{\frac{0.21}{2}} = 0.324$$

$\sigma = 0.3$

(Rounded to one significant figure)

If there is no "true" value available with which to compare our results, our work constitutes a new contribution to science. Before it is published we want to assure that all errors are reduced as much as possible. Experience gained in trying to reduce errors on known samples is critical to this process.

ESTIMATING UNCERTAINTY FROM SINGLE MEASUREMENTS

When the number of interest is the result of a simple measurement, such as the temperature using a mercury or alcohol thermometer, the uncertainty can often be estimated satisfactorily by how easily the instrument can be read. Suppose that the marks on the thermometer are one degree apart. It is almost always possible to estimate the

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temperature to better than ±1°C because you can estimate what fraction of mercury or alcohol between the two nearest marks. This method is called INTERPOLATION: estimating a number to a precision not shown by the division marks on the thermometer. If the marks are widely spaced, then you may be able to reliably estimate the temperature to within ±0.1 degree, but if the marks are very closely spaced, you may be able to only estimate the temperature to within ±0.5 degrees. This requires judgment and experience, and you are not expected to estimate the uncertainty of a measurement the same as the TA (or anyone else!) would.

CALCULATIONS

Suppose that you need to do a calculation that involves two or more numbers, each of which has some uncertainty. How do the individual uncertainties affect the calculation? Take the example of calculating the pressure of a gas from the measurement of its volume and temperature, assuming that the number of moles of gas is known exactly. The uncertainties in the temperature and volume were estimated during the experiment.

$$P = \frac{nRT}{V}$$

Where:

$T = 300.0 \pm 0.8\ K$

$V = 20.00 \pm 0.06\ L$

$N = 1.000 \pm 0.001\ mol$

$R = 0.0820578\ \frac{L \cdot atm}{mol \cdot K}$

The constant R is an accurately measured literature value and we assume that the error is negligible. The best estimate of the pressure is:

$$P = \frac{(1.000\ mol)\left(0.0820578 \frac{L \cdot atm}{mol \cdot K}\right)(300.0\ K)}{20.00\ L} = 1.231\ atm$$

The rules for determining the number of significant digits were used to limit our answer to four significant digits. Reporting the number in this way suggests that the uncertainty in the result is ± 0.001 atm. In the following calculations, we will show that even though this method is simplistic, it does give a good estimate of the uncertainty.

Let's look at the effect of the uncertainty in the temperature, reported as ± 0.8 K, on the final pressure. First, calculate the pressure again with a temperature of 300.8 K, which is the highest value that we think the temperature could have been during the experiment.

$$P = \frac{(1.000\ mol)\left(0.0820578 \frac{L \cdot atm}{mol \cdot K}\right)(300.8\ K)}{20.00\ L} = 1.234\ atm$$

The pressure differs from the original calculation by:

$$|1.231 - 1.234| = 0.003\ atm$$

Since the uncertainty in the temperature can also affect the pressure in the other direction (the lowest value that we think the temperature could be is 299.2 K), the

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effect of the uncertainty in the temperature on the uncertainty in the pressure would be ± 0.003 atm.

Similarly, we can calculate the pressure with the highest and lowest values that we think the volume could have been, 20.06 L and 19.94 L, respectively. From this, we find that the effect of the uncertainty in the volume on the uncertainty in the pressure would be ± 0.004 atm.

The sum of the uncertainties to the pressure caused by both effects could be used to estimate the uncertainty:

$$P = 1.231 \pm (0.003 + 0.004) = 1.231 \pm 0.007 \text{ atm}$$

This is not a bad estimate when there are only two variables with uncertainty. However, the sum always overestimates the uncertainty since errors from both the temperature and the volume are not likely to occur as the maximum in the same direction. A better estimate of the uncertainty in the pressure is given by using the square root of the sum of squares of each error:

$$\sqrt{(0.003)^2 + (0.004)^2} = 0.005 \text{ atm}$$

Therefore, we should report our result as $P = 1.231 \pm 0.005$ atm. Although 0.005 is larger than the uncertainty suggested by the number of significant digits in the result, it is the right order of magnitude. So taking care of significant digits is a good way to estimate the uncertainty in the result of a calculation.

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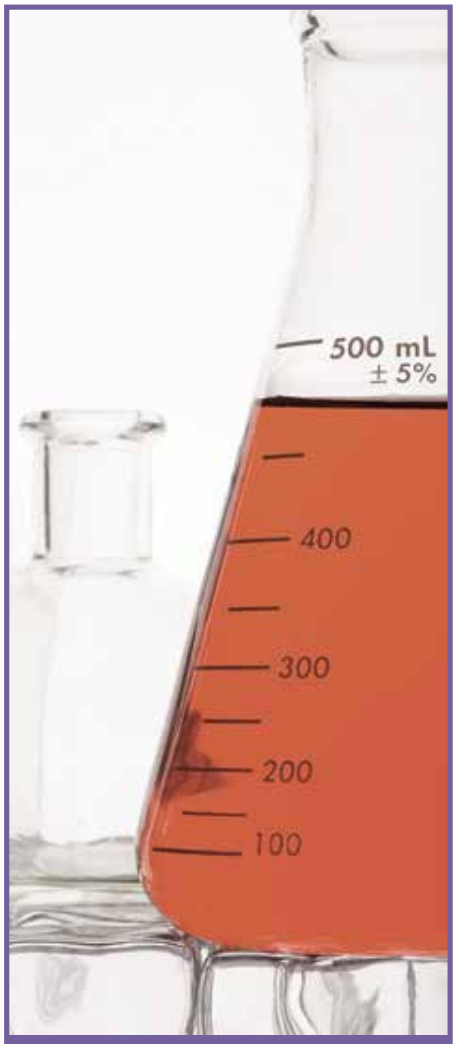
USE OF VOLUMETRIC EQUIPMENT

I. PURPOSE

In this experiment, you will learn about the different types of laboratory glassware: when to use them and how precise they are. Furthermore, you will learn techniques for accurately measuring volumes using graduated cylinders, burets and pipets. You will also learn how to calibrate volumetric equipment to minimize systematic error and how to estimate the magnitude of random error.



II. BACKGROUND

You will learn that the most accurate instrument you have in the laboratory is the analytical balance. To prepare solutions of specific concentration we need to add specific grams of a substance to a specific volume of a liquid, usually water. Most laboratory glassware and plastic-ware used to perform all type of laboratory manipulation and experiments is not specifically manufactured to contain accurate volumes of liquid. Preparing solutions of accurate concentration and making dilutions requires the use of volumetric glassware, which is manufactured with calibration marks that specify exact volumes.

A. USE OF VOLUMETRIC GLASSWARE

Volumetric glassware is available with different levels of accuracy—more accurate glassware is usually much more expensive. Your TA will show you examples of the common types of volumetric glassware. Here we will describe the specification, proper cleaning, and use of some volumetric glassware items.

GLASSWARE SPECIFICATIONS

Common glassware items such as beakers and flasks may or may not have calibration marks on them. Some of the glassware is calibrated **To Deliver (TD)**—the indicated amount when transferred to a separate container.

Examples of to deliver glassware are: burets, pipets, and most graduated cylinders. Other glassware, such as: volumetric flasks and some graduated cylinders are calibrated **To Contain (TC)**—the indicated amount. Sometimes the temperature at which the item has been calibrated will be indicated.

However, any volumetric glassware can be calibrated to improve accuracy. It should be noted that items that are calibrated to deliver the indicated amount would only be correct for the type of fluid used for the calibration. For example, a volumetric buret will deliver a different amount of water than honey! Since density of any substance changes with temperature, for best accuracy, the volumetric glassware should be used at the specified temperature, indicated on the glassware.

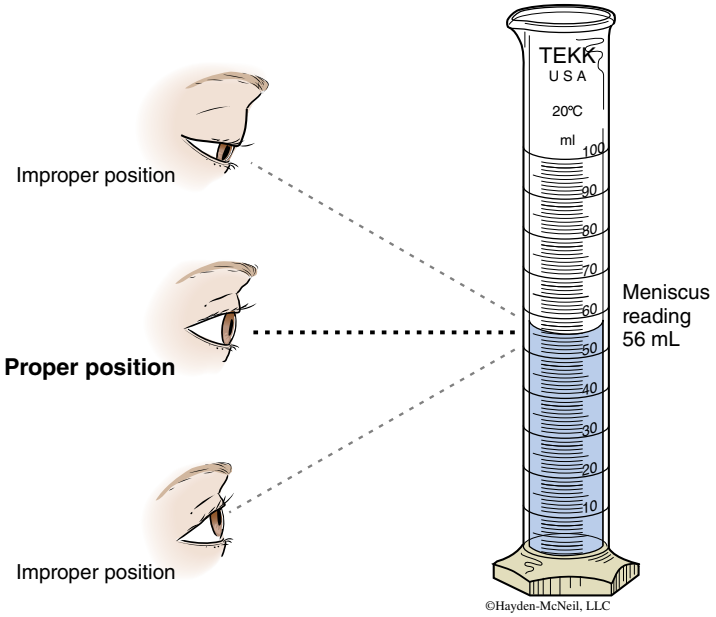


Figure 1-1. *Parallax error*

READING VOLUMES WITH CALIBRATION MARKS

When the level of the liquid is being read on the scale (buret or graduated cylinder), or adjusted to a calibration mark (pipet or volumetric flask), it is important that the glassware be *vertical* and the volume read from the bottom of the curved liquid surface (called the **meniscus**) at eye level (see Figure 1-1). The meniscus, can be seen more easily by holding a piece of paper or a paper that is specifically made for this purpose, called a **reading card**, behind the glassware. You might want to try reading a buret or other items both correctly (eye level) and incorrectly (another angle) to see how much difference it makes.

CLEANING GLASSWARE WITH DEIONIZED WATER (DI)

DI water is water that is free of contaminant ions, but it is expensive to produce; we manufacture it in the CSL building. This difference in cost, between DI water and tap water, is about the same as bottled water to tap water. Part of your locker equipment, the wash bottle, which should be filled only with DI water, will be used to rinse your equipment. If you contaminate the wash bottle, all of your experimental results may be

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wrong; keep it uncontaminated and do not use somebody else's wash bottle. You never know if they fill it up with DI or tap water or something else. Your wash bottle is the most important part of your laboratory equipment for the aforementioned reasons.

CLEANING GLASSWARE

In order to obtain the most accurate measurements with volumetric glassware, it is especially important that the glassware be clean. To test whether glassware is clean, fill it with deionized (DI) water and let it all drain out. If it is clean, the water should form a continuous film on the inside surface. If it is not clean, the water will form droplets on the walls. Once glassware is clean, it will remain clean if it is thoroughly rinsed with deionized water after each use. Keeping the glassware clean after each lab is part of what is practiced in every laboratory called "*Good Laboratory Practice*" (GLP) to ensure that experiment results are not influenced by contamination.

If glassware needs to be cleaned, first fill it with soapy tap water (do not use too much soap since the soap will be hard to remove and may interfere with your reactions) and let it sit for a few minutes, then rinse it several times with tap water. Soft brushes can also be used, but only when the glassware is extremely dirty (we need to conserve water). Give your glassware a final rinse with DI water.

LABELING GLASSWARE

If you are using several identically-sized containers for your solutions, they will all look the same unless you label them. Most glassware has white opaque spots that can be marked with a pencil. Pencil is great for this because it can easily be removed with an eraser. Don't use ballpoint pens or marking pens as they are difficult to remove. Besides pencil, Sharpie can be used on the glassware. To remove sharpie you will need a little bit of ethanol (rubbing alcohol) to remove the mark.

GLASSWARE AND THE ANALYTICAL BALANCE

To calibrate and to prepare solutions, especially standard solutions, you need to use the analytical balance. This is a very sensitive instrument and you need to know how to use it to obtain good results. Most of our balances have a glass cover, this is to protect the balance from air drafts and to avoid balance fluctation. That is why it is important to close the doors (if possible) when you weigh any sample. Also, it is very important to maintain them and use the same balance to reduce variations. If you have any solid spill, use a brush to clean up the mess. If you do not do this, when you place your container with your sample, this solid may be attached to the container and the mass value won't be as accurate. Be very gentle when placing any object on the pan since these balances can break or lose the calibration very easily.

B. RANDOM AND SYSTEMATIC ERRORS IN MEASUREMENTS

Read the section on *Error Analysis* in the laboratory manual. In this experiment, you will be making multiple measurements in order to calculate the errors introduced in the experimental procedures. Review the definitions of the terms ***systematic*** and ***random*** errors in your chemistry book. The best way to minimize random error is to work as carefully as possible and to take multiple measurements. The "best value" will be the average of the measurements and the magnitude of the random error is given

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by the standard deviation. Formulas for calculating the average and standard deviation are given in the laboratory manual section on *Error Analysis*.

All scientific calculators are capable of rapidly performing these calculations; it would be worth your effort to learn how your calculator can do these calculations for you. If you are using the suggested calculator for this course, the Casio fx-300ms-plus, the manual can be found online at ***http://bit.ly/1nOeXNM***. Most other manuals can be found by searching for the specific brand and model of calculator.

Systematic error cannot be reduced by multiple measurements. Instead, instruments must be calibrated by measuring known quantities. Higher quality volumetric glassware will generally be more accurate, but regardless of the quality, improve the accuracy by calibrating the volume, and reducing a possible source of systematic error. This can be done by measuring the mass of a sample of water; since the density of water is known very accurately, the volume can be accurately determined from the mass. If done carefully, this measured volume can be taken as the "true value" to be used in any experiment requiring a very precise volume measurement.

C. PERCENT DIFFERENCE

Scientists will compare their results with actual or "true values." There are several types of comparisons that call for a different type of analysis, i.e., percent difference, percent error, and percent yield. In this experiment we will be focusing on percent difference. **Percent difference** is applied when comparing two experimental quantities, E_1 and E_2, neither of which can be considered the "true" value. The percent difference is the absolute value of the difference over the average times 100.

$$\text{Percent Difference} = \frac{|E_1 - E_2|}{\frac{1}{2}(E_1 + E_2)} \times 100$$

Equation 1-1. *Percent Difference*

III. SAFETY PRECAUTIONS

- *Safety glasses are always required* as long as anyone in the lab is still performing laboratory work!
- *An apron or lab coat and closed-toe/closed-heel shoes are required.*
- *No food or drinks in the lab at any time.*
- Always check your glassware carefully for broken edges that can cut you.
- To reduce glass breakage, set up a support stand with a buret clamp before you bring the buret and the pipet. It is very common that students place pipets on the counters and they roll off onto the floor and break.
- There is no hazardous waste for this experiment.

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IV. MATERIALS AND REAGENTS

QT	Equipment (communal equipment)	QT	Equipment (from your locker)
1	25 mL volumetric pipet	1	50 mL beaker
1	50 mL buret	1	150 mL beaker
1	Pipet bulb	1	400 mL beaker
NO HAZARDOUS WASTE		3	125 mL Erlenmeyer Flask
		1	100 mL graduated cylinder
		1	10 mL graduated cylinder
		1	Medicine dropper

V. EXPERIMENTAL PROCEDURE

You will work **individually** on this experiment, however you will need to share some glassware, please plan ahead.

You will need to check out one 50 mL buret, one 25 mL volumetric pipet, and one pipet bulb from your TA. **You must return all three at the end of the lab period in a clean condition.**

PLEASE SAVE WATER!

A. USING A PIPET

1. Before beginning the experiment, fill a 400 mL beaker with DI water (free of contaminated ions) and set it aside so that it has time to equilibrate to room temperature. You will use this water in parts A, B, C and D of the experiment.

2. Clean your 100 mL graduated cylinder, pipet and buret as described in the background section before using them in this experiment.

3. Your TA will demonstrate how to use a pipet bulb to fill the pipet. Remember:

    - Never use your mouth to suck liquid into a pipet!

    - Never set the pipet where it can roll off the counter; use a buret clamp and support stand to hang it vertically when not in use.

4. You should make a few practice runs by filling and draining the pipet with DI water from a 150 mL beaker. First, look at the print on the top neck of the pipet to check the volume, the temperature of its calibration, and whether it is a TC or TD type and write this information in your laboratory notebook. If you are using it to transfer liquid, it should be a TD type. Also note where the calibration mark (a thin line) is located on the top neck portion, and check the tip on the bottom to make sure it is not chipped.

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5. **Follow the steps below for filling and dispensing fluid from your pipet**:

    a. Squeeze the air out of the bulb.

    b. While still holding the squeezed bulb: hold the pipet firmly in one hand, near the top, while inserting the bulb. **Do not shove the bulb onto the pipet.** Once inserted, always hold the pipet vertically. **Do not let go of the squeezed bulb.** If you do you will let out all the air and will have to re-squeeze the bulb.

    c. Immerse the tip of the pipet into your beaker of DI water a few milliliters above the bottom of the beaker. Using one hand to hold the bulb and the top part of the pipet, use the other hand to grab the beaker and bring it to eye level of the mark of the pipet. To fill the pipet, slowly release your grip of the bulb. Fill the pipet until just above the calibration mark. **DO NOT allow the water to get drawn into the bulb.**

    d. Remove the pipet tip from the DI water **before** adjusting the volume in the pipet. Now, remove the bulb and place your finger on the top part of the pipet, quickly. Deliver any excess of water by slowly twisting your finger on top of the pipet and stop when the bottom of the meniscus is just on the top of the calibration mark. **Always have the meniscus at eye level when measuring the volume.** You are aiming to get the **bottom of the meniscus** to be at the **very top of the calibration mark line**.

    e. When the volume has been properly adjusted carefully move the pipet (always holding it vertically) to the receiving container. Do not shake or jerk the pipet as you may cause some water to fall out, altering your volume.

    f. To deliver the water to the container, place the tip of the pipet against the side of the container while holding the pipet vertical; place the tip against the inside of the beaker and then remove your finger from the top of the pipet and let the liquid flow out of the pipet. *Some water will remain in the tip, do not try to get that water out, this volume is calibrated for in the TD designation.*

    g. When you need to refill the bulb of the pipet it is best to remove the pipet before squeezing the bulb in order to prevent spraying out liquid from the pipet.

6. During your practice with the pipet observe whether or not droplets of water remain on the inside of the pipet after delivering the water to a container. If droplets are present this is a sign that the inside of the pipet is not sufficiently clean. To clean the pipet, follow the glassware cleaning instructions at the beginning of this experiment (draw the soapy water up into the pipet using the pipet filler). The rinsing with soapy water may need to be repeated a couple of times; if after several repetitions of the cleaning you are still seeing droplets inside the pipet speak with your TA.

7. Take three clean, dry (at least outside) 125 mL Erlenmeyer flasks and identify them (e.g., 1, 2, and 3) with a pencil. Weigh each flask on the analytical balance and record the mass in your laboratory notebook.

8. Using the DI water, that you set aside at the beginning of the experiment in your 400 mL beaker, use the 25 mL volumetric pipet and deliver exactly 25 mL of

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DI water into flask #1. Weigh the flask again with water, and record the mass in your laboratory notebook. Repeat this procedure to pipet 25 mL of DI water into flasks 2 and 3.

9. Use your experimental data to calculate the mass of water delivered to each flask. If the masses differ by more than 0.05 g, repeat steps 7–8 to obtain results with less variation. You may add additional 25 mL volumes of water to flasks already containing water; however, you need to take the mass of the flask and any water already in it as the "mass of the flask" for your calculations. Alternately, you can empty the flask and dry the inside with a rolled up paper towel. Water inside the flask is not important as long as it does not change; what it is important is that the flask is dry in the outside so its weight does not change.

10. When you have completed these measurements, dry the flasks with paper towels and allow them to air dry. You will need them dry to complete part C.

11. Using your experimental data, calculate the volume of water delivered to each flask using the denisty value for water from Table 1-1 that is closest to the current room temperature. Remember to write the density of the water and your current room temperature in your data table.

12. Calculate the average and the standard deviation of the three volume measurements. Finally, calculate the absolute difference and percent difference between the average volume and the expected volume. Enter this information in your laboratory notebook.

B. USING A GRADUATED CYLINDER

1. Set the graduated cylinder on a flat surface to read the volume. Avoid entrapped air bubbles by slowly pouring liquids down the side of the cylinder. Read the volume as shown in Figure 1-1, interpolating the volume to the nearest $1/10^{th}$ of the volume marks (0.1 mL in this case).

2. First, weigh the cleaned and dried graduated cylinder and use this value for all your calculations (and your partner's since this is not going to be completely dry when he/she does this part). Then fill the cylinder up to the 100 mL mark with water; add the last drops with a dropper so that the bottom of the meniscus is as close to the the 100 mL mark as possible. Make sure there are no drops stuck to the outside of the container.

3. Tare the balance, and reweigh the graduated cylinder plus water, and record the data in your laboratory notebook. Empty the cylinder and repeat step 2 twice to obtain three consistent values of mass. The masses that you measure should differ by no more than 0.500 g. Repeat the process if necessary to obtain 3 consistent results.

4. Calculate the average and the standard deviation of the three volume measurements. Finally, calculate the absolute difference and percent difference between the average volume and the expected volume. Enter this information in your laboratory notebook.

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C. USING A BURET

1. Set up a support stand and buret clamps. Place the buret in the clamps and check the position—it should be vertical to your eye; this is important to read the volume accurately. Place a 125 mL Erlenmeyer flask below the buret and position the height of the buret so that the tip is just below the lip of the flask.

2. The burets you will use in this lab will have Teflon valves (sometimes called stopcocks). These valves are secured by two plastic rings that screw together. The threads should be finger tight so that the buret does not leak when filled.

3. First you will need to prepare the buret for use by rinsing with a small quantity of the solution you are using, which in this case will be DI water.

    a. Pour 15–20 mL of DI water into a clean 50 mL beaker.

    b. Place the buret in a buret clamp, close the buret valve, and place a waste beaker beneath the buret.

    c. Pour about 10 mL of DI water into the buret from the 50 mL beaker; a funnel can be used to avoid spilling if required.

    d. Remove the buret and with a circular motion cover all the surface area of the buret with solution. Also, rotate the valve completely around several times to drive the entrapped air from the valve and tip; there should be no visible air bubbles in the tip as they will keep you from obtaining accurate volume measurements.

    e. Pour in another 10 mL of solution and repeat the process. Drain out most of the rinsing DI water. This ensures that the buret is completely covered with the titrant, in this case water.

    f. Finally, fill the buret with DI water to above the 1.00 mL mark by pouring in more DI water. Remember, since you are making a difference measurement (final volume minus initial volume), **it is unnecessary to fill the buret to the 0.00 mL mark.**

4. The buret needs to be set up close to the bench edge to facilitate its use and the meniscus of the liquid should be eye level. You should read the buret by estimating to the nearest $1/10^{th}$ of the distance between the calibration marks. For example, if the calibration marks are every 0.1 mL you should try to estimate the reading to the nearest 0.01 mL. However, the reading will likely be accurate to only about 0.02 mL. Be sure to record both the initial and the final volume since the delivered volume is the difference of the two volume measurements.

5. The valve is very precise, and with care, you can deliver a single drop or even a fraction of a drop of liquid. A drop hanging from the tip can be transferred to the receiver by touching it to the tip. Using a waste beaker as a receiver, practice turning the valve slowly until you can controllably deliver a single drop of liquid from the buret.

6. Take the three clean and dry 125 mL Erlenmeyer flasks used in Part A, weigh them again on the analytical balance. Record the masses in your laboratory notebook as mass of flask.

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7. If necessary refill your 50 mL buret with DI water, to a volume between 1.00 and 10.00 mL. Read the volume to the nearest 0.01 mL and record this volume in your laboratory notebook as the initial buret reading.

8. Now, place one of your tared flasks under the buret and transfer about 35 mL of DI water. Read the final volume to the nearest 0.01 mL and record this volume in your laboratory notebook as the final buret reading.

9. Repeat step 7 and 8 twice more, dispensing the water into the remaining two flasks. It isn't necessary to deliver the same volume to each flask, but aim for 30–35 mL. Record the initial and final buret volumes (to the nearest 0.01 mL) in your laboratory notebook.

10. Weigh your Erlenmeyer flasks on the analytical balance, making sure to tare the balance between measurements. Record these masses in your laboratory notebook as mass of flask with water.

11. Use your masses and volumes of water for each trial to calculate the density of the water of each trial. Be sure to show a sample calculation (with units) of this step in your laboratory notebook.

12. Using the temperature of the room, look up the theoretical value of the density for water from Table 1-1. Make sure you write the theoretical density and the current room temperature in your data table. Finally, calculate the percent difference between your experimental average density and the theoretical value (density value from Table 1-1).

D. THE VOLUME OF ONE DROP OF WATER

Sometimes medications from the pharmacy are put into a bottle with a medicine dropper in the cap, and the instructions call for you to take a certain number of drops per day. Perhaps you would like to know more precisely how much medication this really is. Similarly, in the laboratory you sometimes need to deliver a small amount of water or other liquid to a solution when there is not a great need for accuracy. You have a dropper (commonly called a medicine dropper) in your locker drawer for this purpose. In this part of the experiment, you will measure the volume of one drop of water from this dropper. For accuracy in this determination, you will dispense 30 drops of water and measure its average mass.

1. Weigh a clean, dry 50 mL beaker, and record the mass in your laboratory notebook.

2. Using the DI water that you set aside at the beginning of the experiment, add 30 drops of water to the beaker using your medicine dropper. When using the dropper hold it vertically to ensure reproducible drop formation and results. Weigh the beaker with water, and record the mass.

3. Record the room temperature and the density of water at this temperature in your laboratory notebook. Next, calculate the average mass of one drop of water from the medicine dropper. Finally, calculate the volume of one drop of water using the density value for water from Table 1-1 that is the closest to the current room temperature. Record the information in your laboratory notebook. Be sure to show your calculations, with proper units in your report.

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VI. DISCUSSION/QUESTIONS

1. **Discussion:** Using the results from the Part A, Part B, and Part C compare and comment on the accuracy and precision of your volume measurements using a pipet, graduated cylinder, and buret. Which one gave more accurate measurements? Which one gave more precise measurements?

2. **Discussion:** How do your results discussed in question 1 reflect systematic errors in the measurements? How do your results reflect random errors in the measurements?

3. What mass of water would be delivered from your 25 mL pipet if the solutions were at 50°C? The density of water is 0.9881 g/mL at 50.0°C.

4. Was the density of water measured with the buret in Part C accurate and/or precise? Explain your answer. *Hint: Use Table 1-1 to find the theoretical value density of water at a certain temperature.*

5. If you were to use a 25 mL pipet or a 100 mL graduated cylinder to make further measurements, how would you know what value of the volumes to use in your calculation? Explain how the calibration procedure reduces systematic error.

Table 1-1. *Density of water at different temperatures.*

Temperature (°C)	d_{H_2O} (g/mL)	Temperature (°C)	d_{H_2O} (g/mL)
18.0	0.99860	22.0	0.99777
18.5	0.99850	22.5	0.99865
19.0	0.99841	23.0	0.99754
19.5	0.99830	23.5	0.99742
20.0	0.99820	24.0	0.99730
20.5	0.99809	24.5	0.99716
21.0	0.99799	25.0	0.99704
21.5	0.99788	25.5	0.99690

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