

GENERAL CHEMISTRY LAB SAFETY

The chemistry laboratory can be a place of discovery and learning. However, by the very nature of laboratory work, it can be a place of danger if proper common-sense precautions aren't taken. While every effort has been made to eliminate the use of explosive, highly toxic, and carcinogenic substances from the experiments which you will perform, there is a certain unavoidable hazard associated with the use of a variety of chemicals and glassware. You are expected to learn and adhere to the following general safety guidelines to ensure a safe laboratory environment for both yourself and the people you may be working near. Students who fail to follow all safety rules may be asked to leave the lab or suffer grading penalties.

ATTIRE

- a) Safety goggles must be worn at all times while in the laboratory. This rule must be followed whether you are actually working on an experiment or simply writing in your lab notebook.
- b) Closed toe shoes and long pants must be worn in the lab. Sandals and shorts are not allowed.
- c) Long hair must be tied back when using open flames.

CONDUCT

- a) Eating, drinking, and smoking are strictly prohibited in the laboratory.
- b) No unauthorized experiments are to be performed.
- c) Never taste anything. Never directly smell the source of any vapor or gas; instead by means of your cupped hand, waft a small sample to your nose. Do not inhale these vapors but take in only enough to detect an odor if one exists.
- d) Always wash your hands before leaving lab.
- e) Learn where the safety and first-aid equipment is located. This includes fire extinguishers, fire blankets, and eye-wash stations.
- f) Notify the instructor immediately in case of an accident.
- g) Proper Handling of Chemicals and Equipment
- h) Consider all chemicals to be hazardous unless you are instructed otherwise.
- i) Know what chemicals you are using. Carefully read the label *twice* before taking anything from a bottle.
- j) Excess reagents are never to be returned to stock bottles. If you take too much, dispose of the excess.
- k) Many common reagents, for example, alcohols and acetone, are highly flammable. Do not use them anywhere near open flames.
- l) Always pour acids into water. If you pour water into acid, the heat of reaction will cause the water to explode into steam, sometimes violently, and the acid will splatter.
- m) If chemicals come into contact with your skin or eyes, flush immediately with copious amounts of water and consult with your instructor.
- n) Never point a test tube or any vessel that you are heating at yourself or your partner, it may erupt like a geyser.
- o) Dispose of chemicals properly. Waste containers will be provided and their use will be explained by your Lab Assistance. Unless you are explicitly told otherwise, assume that only water may be put in the lab sinks.
- p) Clean up all broken glassware immediately and dispose of the broken glass properly.
- q) Never leave burners unattended. Turn them off whenever you leave your workstation. Be sure that the gas is shut off at the bench rack when you leave the lab.
- r) Beware of hot glass. it looks exactly like cold glass.

COMMONLY USED LABORATORY EQUIPMENT



- | | |
|---|-------------------|
| A. Beaker | I. Pipette filler |
| B. Erlenmeyer Flask | J. Test tube |
| C. Graduated cylinder | K. Dropper |
| D. Wash bottle | L. Funnel |
| E. Burette | M. Spatula |
| F. Ring stand | N. Test tube rack |
| G. Burette clamp | O. Crucible tongs |
| H. Graduated pipette / volumetric pipette | P. Crucible |

Figure 1.1: Commonly used laboratory equipment

GENERAL INSTRUCTIONS

- a) Before starting any experiment, have a clear idea of what is to be done by studying carefully the details of the experiment. Plan your work so as to maximize laboratory time.
- b) Record all details, such as accurate weights and volumes, color changes, or evolution of gases, on a notebook. Do not use scraps of paper to do this.
- c) Keep your working bench clean and dry. Store glasswares that are not immediately required, place dirty glasswares in a plastic bowl if they are to be cleaned later, and throw solid wastes in the proper waste basket.
- d) Carry all operations involving unpleasant and noxious materials in a properly functioning fume cupboard.

CLEANING OF GLASSWARES

It is a good habit to clean all glasswares immediately after use as the identity of the contaminant is known and thus easier to remove.

- a) Use a test-tube brush and a commercial washing powder or liquid detergent.
- b) Rinse thoroughly with tap water, and finally with distilled water. Drain off the water as completely as possible.
- c) Use a warm air blower or a hot oven (100-120°C for about 1 hour) to dry the glassware, placing them upside down. Glassware to be used immediately may be rinsed with a little industrial spirit or acetone before drying.

CHEMISTRY LABORATORY TECHNIQUES

TRANSFERRING MATERIALS

- a) Always bring a CLEAN, DRY container and CLEAN, DRY transfer tool with you to get any chemicals.
- b) Take the correct amount of chemical from a container but NEVER put any back once you have taken it to your lab station.
- c) Put the lid back on any chemical container when you are finished with it.
- d) NEVER use the same piece of equipment or apparatus for different chemicals. Failure to do this could result in contamination of one or more stock bottles or jars. Always start with clean, dry containers for each separate chemical.
- e) Pour liquids slowly to avoid splashing.
- f) In transferring liquids from one container to another at your lab station, use a stirring rod laid across the top of one container to direct the flow of the liquid into the other. This technique is especially useful when you are separating a liquid from a solid (decanting).

ANALYTICAL BALANCE

Many experiments require accurate weights of solid chemicals. An analytical balance is used for this purpose. Most laboratories are equipped with digital electronic balances and/or single pan analytical balances. These balances are accurate to 0.001 and 0.0001 g respectively. General rules when using an analytical balance:

- a) Before weighing, make sure the reading is at zero. After weighing, press "tares" or turn the dial so that the reading is back at zero.
- b) Use a weighing bottle with a stopper. Do not use pieces of paper.
- c) Do not weigh corrosive or hot chemicals.
- d) Keep the balance clean all the time.

MEASURING VOLUMES OF LIQUIDS

- Always read the volume of a liquid in a container so your eye level is at the same level as the surface of the liquid.
- When the volume of water and water-solutions are read in a container, you will notice that there is a curved surface at the top of the liquid. This is called the meniscus. Always read the bottom of the meniscus (figure 1.2).
- The following lab equipment will be used for dispensing, measuring or mixing liquids during the year:

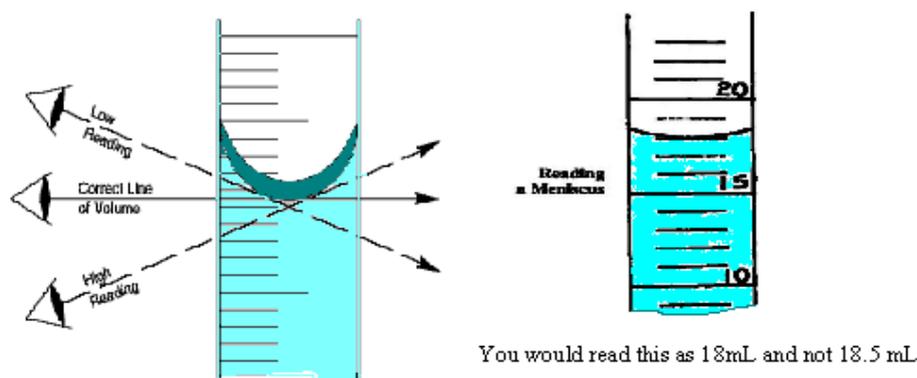


Figure 1.2: Proper technique for reading a burette or a graduated cylinder

1. GRADUATED CYLINDERS

- These are designed to measure volumes of liquids accurately.
- No chemicals are to be mixed in a graduated cylinder.
- A graduated cylinder should never be heated or placed in an oven.
- Different size graduated cylinders have different "sensitivities"; that is, different degrees of accuracy. Selection of the proper graduated cylinder depends not only on the total volume you need to measure but the accuracy that you need as well.

2. BEAKERS AND FLASKS

- Only approximate volumes are measured in beakers and flasks.
- Beakers and flasks are primarily used for mixing and heating chemicals.

3. BURETTES

- A burette is used like a graduated cylinder in that a specific volume can be accurately measured.
- The burette is held by a ring stand and the liquid is dispensed by a small valve at the bottom.
- Burettes that have not been cleaned properly are major sources of error in labs requiring their use.

4. VOLUMETRIC FLASKS

- These long-necked flasks are used for preparing a specific volume (e.g. 1.00 liter, 2.00 liters) of a solution. There is a single line etched on the neck for this particular volume only.

5. PIPETTE

- i. Pipettes are narrow, hollow tubes that are used to draw a specific, small volume of liquid with a squeeze bulb. Typically, the capacity of a pipette is 5.0 or 10.0 milliliters.
- ii. Pipettes should never be used for stirring solutions.
- iii. NEVER use your mouth to draw a solution into a pipette!

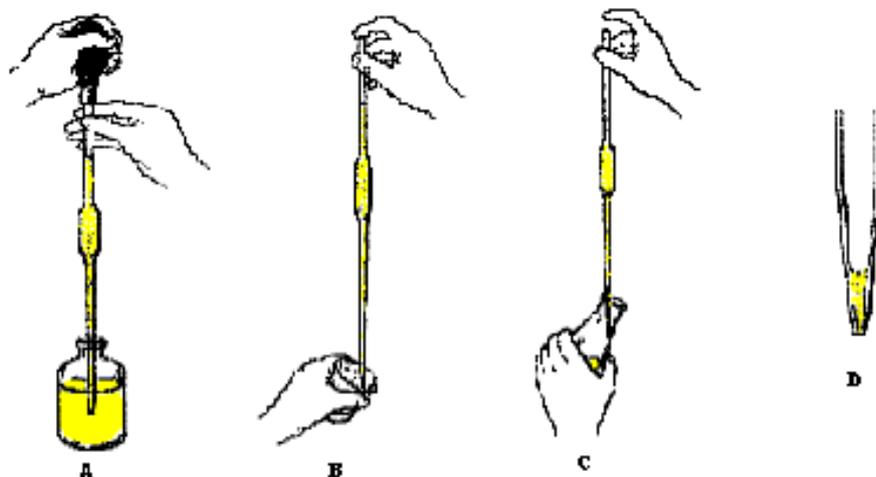


Figure 1.3: Proper technique for using pipette

- A. Suction by a pipette filler
- B. Drain to graduation mark touching off adhering drop
- C. After discharge, allow a 20 seconds drainage period. Touch the tip to the side of the flask.
- D. The liquid remaining in the tip of the pipette should not be blown out. The pipette was calibrated for this amount to remain.

FILTRATION AND DRYING

- a) When separating a solid from a liquid, filtration can be used. The liquid phase will go through the filter paper while the solid phase will not.
- b) The filter paper is folded in half TWICE (in quarters) so that it will open up and fit into a funnel (either glass or plastic).
- c) The filter paper can be moistened with water from a wash bottle so it will adhere to the sides of the funnel.
- d) The funnel containing the filter paper should be placed in a filtering rack (with flask underneath) or in the mouth of a flask.
- e) Never add so much liquid that it overflows the side of the filter paper. Some solid material could end up in the flask.
- f) As solid material collects in the filter paper, the rate of flow may slow down. This is normal.
- g) Once all of the solid/liquid mixture to be filtered has been transferred to the filter paper, wash the solid material, using a wash bottle, with distilled water several times.
- h) The filter paper containing the separated, washed solid, should be dried. This can be accomplished by removing the filter paper from the funnel and placing it in a clean, dry, pre-weighed beaker. The beaker containing the filter paper can then be air-dried or placed in an oven.

EVALUATION OF EXPERIMENT DATA

The chemist scientist of today and future therefore needs to have thorough grounding in experiment techniques, in how to acquire data, and in how to evaluate the data collected. Only with knowledge of how to evaluate collected data can any significance be placed on the experimental measurements. Precision and accuracy are two essential concepts in the evaluation of data.

In making measurements it is important to recognize and if possible, estimate sources of error. Experimental error may be classified as either systematic or random. In systematic error the cause usually is detectable and can be corrected. This type of error causes measurements to be consistently higher or lower than actual value. These include errors present in the system itself, error present in the measuring device, personal error and gross error such as incorrect recording data or miscalculation. Random errors are more related to experimental uncertainty than accuracy. Their sources cannot be identified and are beyond our control. They also tend to fluctuate in random fashion about a measured value.

PRECISION

Precision is a quantitative measure of the reproducibility of experimental measurements how well repeated measurements of the same quantity agree with one another. Precision is frequently measured in terms of the average deviation, which is determined by following

- determine the average value from a series of measurements
- determine the deviation of each measurement from the average value
- determine the average of the deviations without regard to sign

ACCURACY

The agreement of experimental measurements with the accepted value of a quantity is measured in terms of the error. The error is the difference between the values of quantity as measured and the accepted value of the same quantity:

$$\text{Error} = \text{measured value} - \text{accepted value}$$

When the error in a measurement is put on a relative basis it becomes more useful and is known as the relative error. The error divided by the accepted value and multiplied by 100%. This is usually known as the *percent error*.

$$\text{Percent error} = \frac{\text{Error}}{\text{Accepted value}} \times 100\%$$

EXAMPLE: in the determination of the concentrations of an unknown acid by titration with standard base, four measurements were made: 0.1025 M, 0.1018 M, 0.1020 M, and 0.1024 M. Compute the average value and the average deviation. Assume that the accepted value of the unknown acid molarity is 0.1014 M as determined by several different experienced experimenters using sound technique. Compute the error and percent error for determination of the molarity of an unknown acid.

SOLUTION: the average value is computed by summing the four measurements and dividing by four. This yields an average value of 0.1022 M. The individual deviations of each measurement from the average value are 0.0003 (0.1025 M), 0.0004 (0.1018 M), 0.0002 (0.1020 M) and 0.0002 (0.1024 M). The sum of the four deviations 0.0011, divided by four yields the average deviation, 0.00028. Since this deviation represents an uncertainty in the measurements, the molarity of the unknown acid is not precisely 0.1022 M, but ranges from 0.1019 M to 0.1025 M and should therefore be reported with the average deviation included, that is 0.1022 ± 0.0003 M.

Further, to make the measurement of precision more useful, the average deviations are put on the percentage basis by determine the relative average deviation. This is the average deviation divided by the the average value and multiplied by 100%.

$$\text{Relative average deviation} = \frac{\text{Average deviation}}{\text{Average of measurements}} \times 100\%$$

$$\text{Relative average deviation} = \frac{0.00028}{0.1022} \times 100\% = 0.27\%$$

A relative average deviation of 0.27% or better (that is, smaller) is typical expected precision value for an acid-base titration. In general, however, the precision of an experiment varies with the technique and/or apparatus used. A number of variables built into the method of the experiment can affect its precision. In the method tested over long period of time, the results (average value) should not only agree very well with one another (have good precision) but also agree very closely with the accepted value (have high accuracy)

Since the error is the difference between the measured value and accepted value, in this case it is $0.1022 - 0.1014 = + 0.0008$. From this error, the percent error is calculated.

$$\text{Percent error} = \frac{+ 0.0008}{0.1014} \times 100\% = + 0.8\%$$

The only significant of the sign of the error (+ or -) is that the measured value is greater or smaller than the accepted value. In the example above, the percent error, which is measures the accuracy of the experiment, is larger than the precision. This is an indication of existence of systematic error and can be corrected. If all systematic errors have been eliminated, the accuracy should be comparable to the precision of the experiment that measures random error. Thus the accuracy of the experiment is related to the precision (random error) but is also related to possible systematic error.

PROPAGATION OF ERROR AND DEVIATIONS

Often the final result of an experiment will be a combination of the measurements of several different items. The errors (or deviations) in all of those measurements are combined to produced the error (or deviation) of the result. If two measurements are added or subtracted, the error of the result is the sum of the absolute error of the two measurements. If two measurements are multiplied or divided, the error of the result is the sum of the relative error of the measurements. Identical rules to the propagation of deviation.

GRAPHS

Graphs are used to present the data in picture from so that they can be more readily grasped by the reader. Occasionally a graph is used to follow a trend. Notice that the best smooth curve is drawn through the data points. This is not the same as connecting the dots, all data points will not fall on the line. Often, a graph is used to show how well data fit a straight line. The line drawn may either be visually estimated.

- The graph must have a title. Title your graph. Make sure you describe what is represented on the horizontal and vertical axes.
- The scale you select should be easy to work with. Such as a 1 square represents 1, 2, 5, or 10. Or perhaps 5 squares represent a unit of 1, 10, or a multiple of ten, such as 20,100, or 0.1. Never change the scale along an axis. However, you may use a different scale for each axis.
- Select a scale that uses the greatest portion of the paper. Don't go overboard here. The axes must be both numbered and labeled.

- d) Using a ruler draw the coordinate axes on the graph paper.
- e) Put "tick" marks on the axes to show the measurements.
- f) Plot your points and then draw a best fit curve or line. Do not connect the dots.

UNIT OF MEASUREMENT

The International System of Units (abbreviated SI from the French Le Système international d'unités) is the modern form of the metric system. It is the world's most widely used system of units, both in everyday commerce and in science.

The international system of units consists of a set of units together with a set of prefixes. The units of SI can be divided into two subsets. There are seven base units. Each of these base units is nominally dimensionally independent. From these seven base units several other units are derived.

Name	Symbol	Quantity
meter	m	length
kilogram	kg	mass
second	s	time
ampere	A	electric current
kelvin	K	thermodynamic temperature
mole	mol	amount of substance
candela	cd	luminous intensity

A prefix may be added to units to produce a multiple of the original unit. All multiples are integer powers of ten. For example, kilo- denotes a multiple of a thousand and milli- denotes a multiple of a thousandth hence there are one thousand millimetres to the metre and one thousand meters to the kilometer. The prefixes are never combined: a millionth of a kilogram is a milligram not a microkilogram.

Name	Symbol	Factor	Name	Symbol	Factor
tera-	T	10^{12}	deci-	d	10^{-1}
giga-	G	10^9	centi-	c	10^{-2}
mega-	M	10^6	milli-	m	10^{-3}
kilo-	k	10^3	micro-	μ	10^{-6}
hecto-	h	10^2	nano-	n	10^{-9}
deca-	da	10^1	pico-	p	10^{-12}

SI derived units are part of the SI system of measurement units and are derived from the seven SI base units

 <p>UNIVERSITI TEKNIKAL MALAYSIA MELAKA</p>	<p>No. Dokumen: TB/MMK/ DACS 1232</p>	<p>No. Isu./Tarikh 1/1-8-2008</p>
<p>CHEMISTRY Experiment 1: Usage and Calibrate of Lab Glassware Equipment</p>	<p>No. Semakan/Tarikh 1/1-8-2016</p>	<p>Jumlah Mukasurat 3</p>

OBJECTIVE:

To calibrate glassware such as burette, pipette, volumetric flask, beaker, and measuring cylinder and list possible errors that could occur in an experiment.

LEARNING OUTCOMES

After conducting this experiment, you should be able to:

1. State the reasons why laboratory equipments must be calibrated prior to use;
2. Calibrate glassware's; e.g. burette, pipette and volumetric flasks; and
3. List the possible errors that could occur in an experiment.

INTRODUCTION:

Chemists use a variety of glassware to measure the volume of chemicals. The specific type of glassware used in any situation depends on how accurately or precisely the volume needs to be known. An error is a bound on the precision and accuracy of the result of a measurement.

These can be classified into two types: **random error** and **systematic error**. Random error is caused uncertainty and carelessness in the measurement apparatus, where as systematic errors is occur when the equipments used are imperfect or faulty and were not calibrated before use. These errors could be minimized by checking and calibrating the equipments prior to use.

Glassware is used for a wide variety of functions which include volumetric measuring, storing chemicals or samples, mixing or preparing solutions or other mixtures. For the **burette**, **pipette** and **volumetric flask** are made to determine specific volume of a solution at a specific temperature. Therefore, it is important to determine the type of glassware that will be used and calibrate it. Meanwhile, there are types of glassware are used for the approximate measurement of volume likes **reagent bottles**, **beakers**, **conical flasks**, and **graduated cylinders**. Normally they are used to prepare solutions with an approximate concentration.

There are three general methods commonly employed to calibrate glassware. These are as follows: **Direct calibration:** A volume of water delivered by a burette or pipette, or contained in a volumetric flask, is obtained directly from the weight of the water and its density (refer to Table 1.1). **Indirect calibration:** Volumetric glassware can be calibrated by comparison of the mass of water it contains or delivers at a particular temperature with that of another vessel which had been calibrated directly. The volumes are directly related to the masses of water. This method is convenient if many pieces of glassware are to be calibrated. **Relative calibration:** It is often necessary to know only the volumetric relationship between two items of glassware without knowing the absolute volume of either.

This experiment will acquaint you with the precision, accuracy, and tolerance of measurements made using a beaker with calibration marks, a graduated cylinder, a burette, and a pipette. The objective of this experiment is to determine the accuracy and precision of your volumetric glassware, so that you can chose the proper instrument for our next analytical experiments

Table 1.1: Density of Water at Various Temperatures

Temperature, (°C)	Density, (g/mL)	Temperature, (°C)	Density, (g/mL)
23	0.9966	28	0.9955
24	0.9964	29	0.9952
25	0.9962	30	0.9949
26	0.9959	32	0.9944
27	0.9957	33	0.9941

LAB EQUIPMENT

100 mL volumetric flask
50 mL burette
20 mL pipette
150 mL beaker
100 mL beaker

100 mL measuring cylinder
analytical balance
funnel
dropper
wash bottle

CHEMICALS AND MATERIALS

Distilled water

EXPERIMENTAL PROCEDURE

Part A: Calibrate Burette

1. Wash, clean and dry 100 mL beaker (label A), and record its weight.
2. Clean the burette with distilled water.
3. By using funnel (avoid spilling) to pour the distilled water into the burette to a point the mark for 0.00 volume; (make sure the portion of the burette below the stopcock is filled or without any air bubbles).
4. Transfer 50 mL water from burette into beaker (label A).
5. Repeat step 3 and 4, by adding distilled water into beaker (label A) from burette until the volume reaches 100 mL.
6. Record the weight of water and beaker for 100 mL volume. Calculate the density of water (in g/mL).
7. Compare value of density was measured with standard density of water at ambient temperature.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (\text{eq. 1.1})$$

8. Compare this calculated density to its actually density (via temperature) using the percent error formula shown below:

$$\frac{|\text{Theoretical} - \text{Experiment}|}{\text{Theoretical}} \times 100\% = \% \text{Error} \quad (\text{eq. 1.2})$$

Part B: Calibrate Pipette

1. Rinse, clean and dry the 100 mL beaker (label A) has been used, and records its weight.
2. Rinse the 20 mL pipette with distilled water.
3. Fill the pipette by using pipette bulb until the distilled water is reach pipettes level line
4. Transfer 20 mL water from pipette into beaker (label A). Withdraw the last drop of water by touching it to the glass surface for about 15 seconds.
5. Repeat step 3 and 4, by adding distilled water into beaker (label A) from pipette until the volume reaches 100 mL.
6. Record the weight of water and beaker for 100 mL volume. Calculate the density of water (in g/mL).
7. Compare value of density was measured with standard density of water at ambient temperature.

Part C: Calibrate Volumetric Flask

1. Rinse, clean and dry the 100 mL beaker (label A) has been used, and records its weight.
2. Rinse and clean the 100 mL volumetric flask with distilled water.
3. Pour the distilled water into the volumetric flask until approximately $\frac{1}{2}$ from the marked level by using a funnel.
4. Use a clean dropper to add the water until it reaches the required level. Position your eyes parallel to the meniscus level.
5. Transfer 100 mL water from volumetric flask into beaker (label A).
6. Record the weight of water and beaker for 100 mL volume. Calculate the density of water (in g/mL).
7. Compare value of density was measured with standard density of water at ambient temperature.

Part D: Calibrate Measuring Cylinder

1. Rinse, clean and dry the 100 mL beaker (label A) has been used, and records its weight.
2. Rinse and clean the 100 mL measuring cylinder with distilled water.
3. Pour the distilled water into the measuring cylinder until it reaches the 100 mL level. Position your eyes parallel to the meniscus level.
4. Transfer 100 mL water from measuring cylinder into beaker (label A).
5. Record the weight of water and beaker for 100 mL volume. Calculate the density of water (in g/mL).
6. Compare value of density was measured with standard density of water at ambient temperature.

Part E: Calibrate Beaker

1. Rinse, clean and dry the 100 mL beaker (label A) has been used, and records its weight.
2. Rinse and clean the 150 mL beaker with distilled water.
3. Pour the distilled water into the beaker until it reaches the 100 mL level. Position your eyes parallel to the meniscus level.
4. Transfer 100 mL water from 150 mL beaker into 100 mL beaker (label A).
5. Record the weight of water and beaker for 100 mL volume. Calculate the density of water (in g/mL).
6. Compare value of density was measured with standard density of water at ambient temperature.

 UTeM اوتيمرسني تېكنيكل مليسيا ملاك UNIVERSITI TEKNIKAL MALAYSIA MELAKA	UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DACS 1232	No. Isu./Tarikh 1/7-7-2008
CHEMISTRY Experiment 2: Density of Liquid and Solid		No. Semakan/Tarikh 1/7-7-2016	Jumlah Mukasurat 3

OBJECTIVE:

To calculate the density of given substances from measured values of volume and mass

LEARNING OUTCOMES

After conducting this experiment, you should be able to:

1. Known that type of technique to determine the density of substance
2. Calculate the density of an unknown liquid
3. Calculate the density of an irregularly shaped solid.

INTRODUCTION:

Density is a physical property of liquids and solids. We can define density (d) as the amount of mass is given volume (eq 2.1). The unit of density is g/cm^3 or g/mL . For a homogeneous object, the formula mass/volume may be used. To determine the density of a liquid, the volume may be measured directly but the mass may use an indirect technique called **weighing by difference** (figure 2.1). The density of a solid material, we can measure mass directly using balance but the volume especially an irregular object cannot to be found directly. However, its can found indirectly from the amount of a liquid it displaces. This technique is called **volume displacement** (figure 2.2).

$$density = \frac{mass}{volume} \quad (\text{eq 2.1})$$

After collecting the experiment data, we can calculate density by dividing the mass by the volume. However, most important calculated value was the proper unit. The density of liquids and solids is usually expressed in gram per millimeter (g/mL) or gram per cubic centimeter (g/cm^3). Since $1 \text{ mL} = 1 \text{ cm}^3$, the numerical value for density in g/mL and g/cm^3 is identical. For example, the density of water may be expressed as 1.00 g/mL or 1.00 g/cm^3 .

Weighing by difference, the technique of the mass of the chemical was calculated by subtracting the weight of an empty container from the total weight of the container and chemical. Meanwhile, volume displacement technique based on Archimedes principle that volume of solid (that sinks) is equal with displaces an amount of water. It is can be calculated by the difference of volume before and volume after an object is immersed in an amount of water.

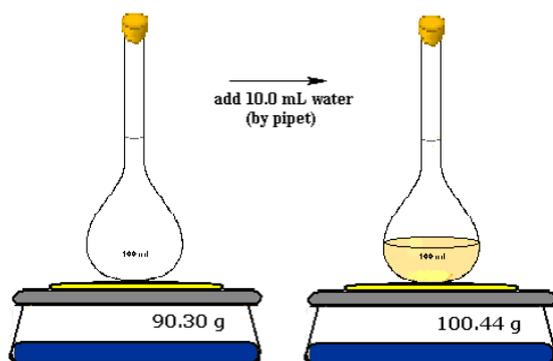


Figure 2.1: Weighing By Difference

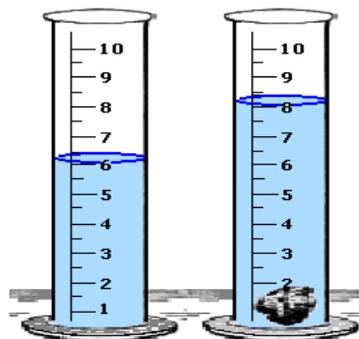


Figure 2.2: Volume Displacement

LAB EQUIPMENT

10 mL graduated cylinder
 50 mL burette
 100 mL beakers
 Retort stands and clamps

Filter funnel
 Dropper
 Wash bottle

CHEMICALS AND MATERIALS

Distilled water
 Unknown liquid
 Irregularly shaped solid

EXPERIMENTAL PROCEDURE

Part A: Calibration of Burette and Graduated Cylinder

1. Rinse, clean and dry the 10.0 mL graduated cylinder.
2. Clean the burette with distilled water.
3. By using funnel (avoid spilling) to pour the distilled water into the burette. Read the initial level and its need not start at 0.00 mL (make sure the portion of the burette below the stopcock is filled or without any air bubbles).
4. Carefully, transfer amount of water into graduated cylinder from the burette until to a point the mark for 2.0 mL of cylinder. Read the final level and record these data and calculate the volume of water.
5. Repeat step 4, by continues adding of water from the burette into cylinder for the point of 4.0 mL, 6.0 mL, 8.0 mL and 10.0 mL.
6. From these data, plot a graph volume of graduated cylinder vs. volume of burette. These calibration graph will be use for experiment Part B and Part C

Part B: Density of an Unknown Liquid

1. Rinse, clean and dry the 10.0 mL graduated cylinder (used in Part A)
2. Weigh the cylinder on an analytical balance. Record its mass to the nearest 0.01 g.
3. Obtain about 15 mL of an unknown liquid in a 100 mL beaker, and record the unknown label.
4. By using a dropper, add between 2 to 4 mL of the unknown liquids to the cylinder. Record the volume to the nearest 0.1 mL.
5. Weigh the liquid and cylinder on the same balance used for the first weighing, and again records the mass to the nearest 0.01 g.
6. To validate your experimental technique, and repeat step 4 and 5 by using the same liquid in set of volume of the unknown liquid between 5 to 7 mL and between 8 to 10 mL.
7. Calculate the density of the unknown liquid from your mass and volume (from the cylinder and calibration graph) data.
8. Return the liquid to its container when you finish with it.

Part C: Density of an Irregularly Shaped Solid

1. Rinse, clean and dry the 10.0 mL graduated cylinder (used in Part A).
2. Add about 4 to 5 mL of water to the cylinder and record the volume to the nearest 0.1 mL.
3. Obtain an irregularly shaped object and record its identity label. Spread the solid onto a towel and check to see if it is dry. Dry it with a towel, if necessary
4. Weight this object as accurately as possible and record it mass.
5. Tilt the graduated cylinder to about 40° angle and carefully slide the solid down into the water (the solid should be completely submerged in water and not exceed the capacity of the cylinder).
6. Gently tap the cylinder to remove any air bubbles adhering to the solid).
7. Record the new volume to the nearest 0.1 mL. The change in volume gives the volume of the solid.
8. To validate your experimental technique, and repeat all step by using the same object (must clean and dry) in set of volume between 5 to 6 mL and between 6 to 7 mL of the water to the cylinder.
9. Calculate the density of the unknown liquid from your mass and volume (from the cylinder and calibration graph) data.

	UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DACS 1232	No. Isu./Tarikh 1/7-7-2008
CHEMISTRY Experiment 3: Preparation and Standardization of Solution		No. Semakan/Tarikh 1/7-7-2016	Jumlah Mukasurat 3

OBJECTIVE:

To prepared and determine the concentration of solution

LEARNING OUTCOMES

After conducting this experiment, you should be able to:

1. How to prepare the standard solution and dilute the solution
2. Determine the concentration of solution in various units

INTRODUCTION:

In chemistry, concentration is the measure of how much of a given substance there is mixed with another substance. This can apply to any sort of chemical mixture, but most frequently the concept is limited to homogeneous solutions, where it refers to the amount of solute in a substance

A solution generally consists of a solute dissolved in a solvent. To work quantitatively with a solution we must have ways to express their concentrations. The most useful expressions are molarity (M) density (g/L or g/mL) and percent weight (%w/w). A solution is prepared in a volumetric flask to ensure that a precise final volume is obtained.

Molarity = moles of solute per liters of solution

Density = weight of solute per volume of solution

Percent weight = (weight of solute per weight of solution) x 100%

A **standard solution** is prepared from pure ($\geq 99\%$) substance and the exact concentration of the solution is known. Some solutions need to be standardized because their concentrations vary over time due to chemical reactions such as oxidation.

To lower the concentration of solution, we use the process of **dilution**, or adding more solvent to a solution. In this process, the relative numbers of solute and solvent particles change. Adding more solvent increase the number of solvent particles and increases the volume. For a given amount of the original solution, the number of solute stays the same, but they are spread out through a greater volume, so their concentration is less.

If we dilute a solution of known molarity, we can calculate the molarity of diluted solution using the formula given by.

$$M_1V_1 = M_2V_2 \quad (\text{eq. 3.1})$$

where M_1 and V_1 the molarity and volume of the solution before dilution, M_2 and V_2 are the molarity volume of the solution after dilution

Titration is a common laboratory method of quantitative/chemical analysis which can be used to determine the concentration of a known reactant. Because volume measurements play a key role in titration, it is also known as *volumetric analysis*. It was process of determining the concentration of one substance in solution by reacting it with a solution of another substance that has a known concentration. In a typical acid-base titration, a pipette is used to precisely measure a volume of a

solution of unknown concentration into a conical flask, followed by the addition of an indicator solution.

An indicator solution is a dye that has one color in an acidic solution and another in a basic solution. A standard solution is that added from the burette until some visual effect, such as color change, signals that the reaction is complete. The solution in the burette is referred to as the titrant.

In this experiment, you will be required to prepare a standard solution of oxalic acid. The oxalic acid will be used to standardize a solution of sodium hydroxide of unknown concentration. The chemical reaction between oxalic acid and sodium hydroxide occurs as follows:



The molar equivalence of the reactants can be calculated using the formula given by

$$\frac{M_a V_a}{a} = \frac{M_b V_b}{b} \quad (\text{eq. 3.3})$$

Where M_a and V_a the molarity and volume of the acid, M_b and V_b are the molarity and the volume of the base; a and b are the stoichiometric constants respectively.

LAB EQUIPMENT

Small and large beakers	20 mL pipette
100 mL volumetric flask	Filter funnel
50 mL burette	Dropper
Retort stands and clamps	Glass rod
250 mL conical flask	

CHEMICALS AND MATERIALS

Oxalic acid crystal ($H_2C_2O_4 \cdot 2H_2O$)	Thymol blue indicator solution
NaOH solution (already prepared in the lab)	Distilled water

EXPERIMENTAL PROCEDURE

Part A: Preparation of solution

1. Weigh approximately 0.35 – 0.40 g of oxalic acid in a small beaker. Record the exact weight of the acid.
2. Dissolved the acid in some distilled water and stirred by using glass rod.
3. Use funnel transfer all acid solution from beaker into a 100 mL volumetric flask
4. Add distilled water into volumetric flask until the diluted solution fills the flask to the marked level.
5. Placed the stopper in the flask and inverted the flask for several times to thoroughly mix the solution.
6. Calculate the concentration of solution in g/mL, g/L and molarity

Part B: Dilution of NaOH solution

1. Use a beaker to obtain approximately 30 mL of NaOH solution (already prepared for you in the lab)
2. Pipet 20 mL of NaOH from beaker into a 100 mL volumetric flask
3. Add distilled water into volumetric flask until the diluted solution fills the flask to the marked level.
4. Placed the stopper in the flask and inverted the flask for several times to thoroughly mix the solution.

Part C: Titration

1. Set up the equipment for titration. Use a beaker to obtain approximately 80 mL of NaOH solution (that have prepared from **Part B**)
2. Fill the burette with the oxalic acid that you have prepared from **Part A** (use a funnel). Record the initial reading on the burette.
3. Pipet 20 mL of NaOH into a conical flask and add 2 - 5 drops of Thymol Blue. Titrate the NaOH solution until the color of the solution changes from blue to a permanent light yellow. Record the final reading.
4. Repeat step 2 and 3 two more times. Determine the average amount of oxalic acid used.
5. Calculate the concentration of diluted NaOH and standard NaOH (already prepared for you in lab)

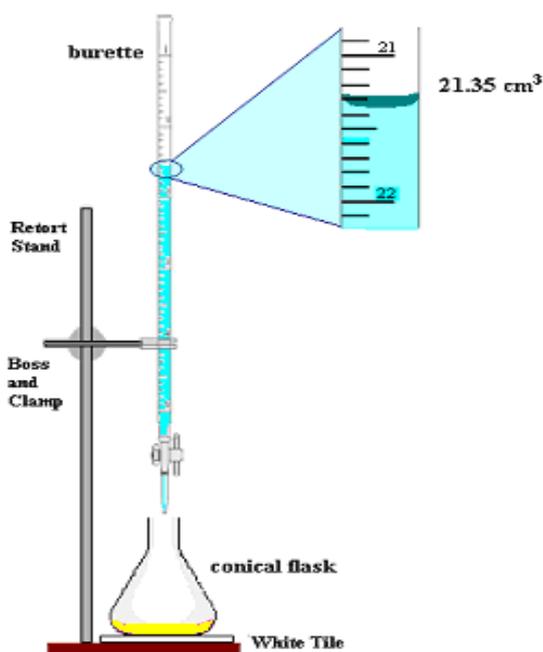


Figure 3.1: Experimental setup for titration process

	UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DACS 1232	No. Isu./Tarikh 1/7-7-2008
CHEMISTRY Experiment 4: Vinegar Analysis		No. Semakan/Tarikh 1/7-7-2016	Jumlah Mukasurat 2

OBJECTIVE:

To determine the percent by mass of acetic acid in vinegar.

LEARNING OUTCOMES

After conducting this experiment, you should be able to:

1. Calculate the concentration of acetic acid in vinegar
2. Determine the percent by mass of acetic acid in vinegar

INTRODUCTION:

Acetic acid is the chemical compound responsible for the characteristic odor and sour taste of vinegar. Household vinegar is a 4 – 5 % (by mass) acetic acid, CH_3COOH , solution (4% is the minimum federal standard). Generally, caramel flavoring and coloring are also added to make the product aesthetically more appealing. In this experiment you are going to analyze a sample of vinegar to find the percent-by-mass of acetic acid in the vinegar.

Volumetric analysis is a technique that employs the measurement of volumes to determine quantitatively the amount of a substance in solution. In any reaction between two or more species, the reaction equation will show the stoichiometric ratio of reacting species. In a volumetric analysis, if one of these species is present in known molar concentration, then by taking a fixed volume of one solution and progressively adding the other solution, it is possible to find a point at which complete reaction of the substances has occurred. This is called the *equivalence point*. The incremental process is called *titration* and enables the concentration of a solution to be determined from the ratio of reacting volumes, the stoichiometric equation and the one known concentration.

The titration techniques is the method used for determine of the percent by mass of acetic acid in vinegar. A measured mass of vinegar is titrated to the phenolphthalein or thymol blue endpoint with a measured volume of a standardized sodium hydroxide solution. As the volume and molar concentration of the standardized NaOH solution are known, the moles of NaOH used for the analysis are also known.

The moles of CH_3COOH are calculated from the balanced equation:



The mass of CH_3COOH in the vinegar is calculated from the measured moles of CH_3COOH neutralized in the reaction and its molar mass, 60.05 g/mol:

$$\text{Mass}(g) \text{ of } \text{CH}_3\text{COOH} = \text{mol } \text{CH}_3\text{COOH} \times 60.05 \text{ g/mol } \text{CH}_3\text{COOH} \quad (4.2)$$

Finally, the percent by mass of CH_3COOH in vinegar is calculated:

$$\% \text{ by mass of } \text{CH}_3\text{COOH} = \frac{\text{mass (g) of } \text{CH}_3\text{COOH}}{\text{mass (g) of vinegar}} \times 100\% \quad (4.3)$$

LAB EQUIPMENT

Volumetric flask 100 mL,
Pipette 20 mL,

Burette 50 mL,
Conical flask 125 or 250 mL

CHEMICALS AND MATERIALS

Standardized NaOH solution,
Sample of vinegar

Thymol blue indicator

EXPERIMENTAL PROCEDURE:

1. Clean 100 mL volumetric flask, and measure its mass.
2. Add the 10 mL of the vinegar to a volumetric flask with previously measured mass.
3. Measure the combined mass of the volumetric flask and sample of vinegar.
4. Dilute a concentrated solution with fill the volumetric flask to the 100 mL mark using distilled water, gently shake the solution.
5. In 125 or 250 mL Erlenmeyer flask, pipette 20 mL of the diluted acid solution (from volumetric flask). Add 5 drop of thymol blue indicator solution.
6. Rinse a clean 50 mL burette with the standardized NaOH solution; fill the burette with NaOH solution. Adjust the volume to the zero mark and record the initial volume.
7. Titrate the sample solution with NaOH solution until the solution just turns blue and the blue color persists for 30 seconds. Read and record the final volume of NaOH titrant in the burette.
8. Repeat the titration procedure above (step 5 to 7), 2 more times.

 UNIVERSITI TEKNIKAL MALAYSIA MELAKA	UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DACS 1232	No. Isu./Tarikh 1/7-7-2008
CHEMISTRY Experiment 5: Boyle's Law		No. Semakan/Tarikh 1/7-7-2016	Jumlah Mukasurat 2

OBJECTIVE:

The primary objective of this experiment is to determine the relationship between the pressure and volume of a confined gas.

INTRODUCTION:

The gas we use will be air, and it will be confined in a syringe connected to a Pressure Sensor (see Figure 5.1). When the volume of the syringe is changed by moving the piston, a change occurs in the pressure exerted by the confined gas. This pressure change will be monitored using a Pressure Sensor. It is assumed that temperature will be constant throughout the experiment. Pressure and volume data pairs will be collected during this experiment and then analyzed. From the data and graph, you should be able to determine what kind of mathematical relationship exists between the pressure and volume of the confined gas. Historically, this relationship was first established by Robert Boyle in 1662 and has since been known as Boyle's law.

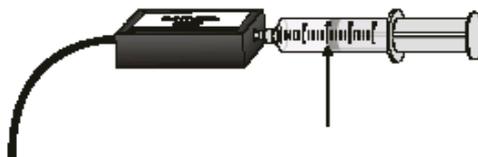


Figure 5.1: Experimental setup for Boyle's Law

LAB EQUIPMENT

Windows PC Vernier computer interface
Logger Pro

Vernier Gas Pressure Sensor
20-mL gas syringe

CHEMICALS AND MATERIALS

Air

EXPERIMENTAL PROCEDURE:

1. Prepare the Pressure Sensor and an air sample for data collection.
 - a. Plug the Pressure Sensor into Channel 1 of the computer interface. With the 20-mL syringe disconnected from the Pressure Sensor, move the piston of the syringe until the front edge of the inside black ring (indicated by the arrow in Figure 5.1) is positioned at the 10.0 mL mark.
 - b. Attach the 20-mL syringe to the valve of the Pressure. Newer Vernier Gas Pressure Sensors have a white stem protruding from the end of the sensor box—attach the syringe directly to the white stem with a gentle half-turn.

2. Prepare the computer for data collection.
 - a. Prepare the computer for data collection by opening the *Experiment 6* folder from *Chemistry with Computers*. Then open the experiment file that matches the sensor you are using.
 - b. On the Graph window, the vertical axis has pressure scaled from 0 to 250 kPa. The horizontal axis has volume scaled from 0 to 20 mL.
3. Click Collect to begin data collection.
4. Collect the pressure vs. volume data. It is best for one person to take care of the gas syringe and for another to operate the computer.
 - a. Move the piston to position the front edge of the inside black ring (see Figure 5.2) at the 5.0-mL line on the syringe. Hold the piston firmly in this position until the pressure value stabilizes.
 - b. When the pressure reading has stabilized, click Keep. Type "5.0" in the edit box. Press the ENTER key to keep this data pair. Note: You can choose to redo a point by pressing the ESC key (after clicking Keep, but before entering a value). Continue the procedure for volumes of 7.5, 10.0, 12.5, 15.0, 17.5, and 20.0 mL.
 - c. Click Stop when you have finished collecting data.
5. In your data table, record the pressure and volume data pairs displayed in the Table window (or, if directed by your instructor, print a copy of the Table window).
6. Examine the graph of pressure vs. volume. Based on this graph, decide what kind of mathematical relationship you think exists between these two variables, direct or inverse. To see if you made the right choice:
 - a. Click the Curve Fit button,
 - b. Choose Variable Power ($y = Ax^n$) from the list at the lower left. Enter the value of n in the Degree/Exponent edit box that represents the relationship shown in the graph (e.g., type "1" if direct, "-1" if inverse). Click Try Fit.
 - c. A best-fit curve will be displayed on the graph. If you made the correct choice, the curve should match up well with the points. If the curve does not match up well, try a different exponent and click Try Fit again. When the curve has a good fit with the data points, then click OK.
7. Once you have confirmed that the graph represents either a direct or inverse relationship, print a copy of the Graph window, with the graph of pressure vs. volume and its best-fit curve displayed. Enter your name(s) and the number of copies you want to print.

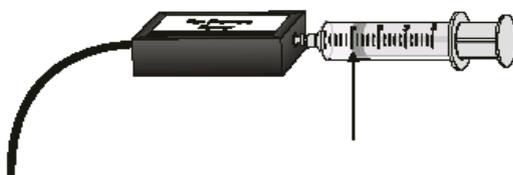


Figure 5.2: Experimental setup for Boyle's law

 UTeM اونیورسیتی تکنیکل ملیسیا ملاک UNIVERSITI TEKNIKAL MALAYSIA MELAKA	UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DACS 1232	No. Isu./Tarikh 1/7-7-2008
CHEMISTRY Experiment 6: Heat of Reaction of Carbonates		No. Semakan/Tarikh 1/7-7-2016	Jumlah Mukasurat 2

OBJECTIVE:

To calculate the heat of reaction of carbonates compound

INTRODUCTION:

Chemical reactions are always accompanied by energy changes. A chemical reaction that releases heat energy to the environment is said to be exothermic. One that absorbs heat from the environment is said to be endothermic. The heat change of a system at constant pressure is called the heat of reaction or enthalpy change. The enthalpy change of a system is given by symbol ΔH . The heat content of a system which undergoes an exothermic reaction is negative because the heat content is reduced.

Calorimetry is the basic technique for studying any process where heat energy is of interest. Calorimeters are designed to be well-insulated, so no heat is gained from or lost to the surroundings. The basic strategy in calorimetry is to use a temperature change and a heat capacity to determine the heat of reaction using the following equation:

$$q_{soln} = (m \times s \times \Delta T) + (C_p \times \Delta T) \quad (6.1)$$

where, q_{soln} is the heat of reaction of the solution

m is the total masses of the reaction

s is the specific heat capacity of the final solution

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

C_p is the heat capacity of calorimeter

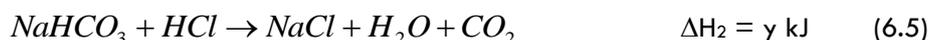
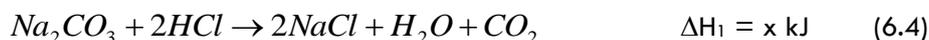
The heat capacity of the calorimeter, C_p is often ignored such that equation 6.1 may be simplified as follows:

$$q_{soln} = (m \times s \times \Delta T) \quad (6.2)$$

The molar enthalpy of a reaction, ΔH_{soln} is negative of the heat of reaction, q_{soln} per mole of limiting reagent, n

$$\Delta H_{soln} = \frac{-q_{soln}}{n} \quad (6.3)$$

The objective of this experiment is to determine the heat of reaction (q_{rxn}) and calculate the molar enthalpy ΔH of following reactions:



The calculations are simplified by assuming the value of s as equal to $4.18 \text{ J/g } ^\circ\text{C}$ and the density of the final solution as equal to 1.0 g/mL . The heat capacity of the calorimeter can be ignored.

LAB EQUIPMENT

Thermometer
20 mL pipette
50 mL beaker
100 mL beaker
200 mL beaker

Glass stirrer
Small towel
Large filter paper or cardboard (calorimeter covering)

CHEMICALS AND MATERIALS

Hydrochloric acid, HCl 2.0 M
Sodium carbonate, Na_2CO_3
Sodium bicarbonate, NaHCO_3

EXPERIMENTAL PROCEDURE:

1. Construct a simple calorimeter (see figure 6.1) using the 100 mL and 200 mL beakers and a small towel. Be sure to dry the inner beaker thoroughly before use.
2. Using the inner beaker, weigh $1.80 - 2.20 \text{ g}$ of Na_2CO_3 . Record the exact weight of the solid. Place the inner beaker inside the larger beaker as shown in the figure.
3. Pipet 20 mL of the HCl solution into another beaker and measure initial temperature, T_{initial} (to one decimal place)
4. Pour the acid into the calorimeter and immediately replace the calorimeter covering. Stir the solution and determine the highest temperature (if temperature rises) or lowest temperature (if the temperature decreases). Record this temperature as the final temperature of the reaction, T_{final} .
5. Thoroughly clean the inner beaker (of the calorimeter) and repeat step 2 – 4, this time weight $2.5 - 3.0 \text{ g}$ of NaHCO_3 .
6. Calculate ΔT , q_{soln} and ΔH_1 and ΔH_2

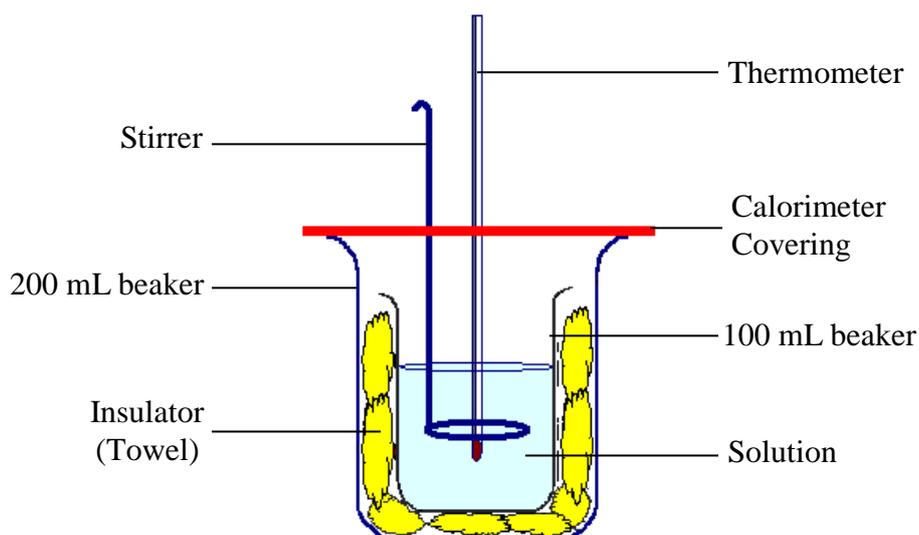


Figure 6.1: A Simple Calorimeter

**LAB REPORT
EXPERIMENT 1**

Usage and Calibrate of Lab Glassware Equipment

Student Name:	Course:	Section:
Instructor:	Table No.	Date:

A. DATA AND RESULT

Water temperature (°C):

	<i>measured volume (ml)</i>	<i>beaker - A (gram)</i>	<i>beaker – A with water (gram)</i>	<i>true volume (gram)</i>	<i>density of water (g/ml)</i>	<i>percent error (%)</i>
Burette
	
	
Pipette
	
	
Volumetric flask
	
	
Measuring cylinder
	
	
Beaker
	
	

Average result:	Average percent error	Rank of most precise equipment
Burette
Pipette
Volumetric flask
Measuring cylinder
Beaker

LAB REPORT
EXPERIMENT 2
Density of Liquid and Solid

Student Name:	Course:	Section:
Instructor:	Table No.	Date:

A. DATA AND RESULT

Part A: Calibration of Burette and Graduated Cylinder

<i>Volume of graduated cylinder (mL)</i>	2.0	4.0	6.0	8.0	10.0
<i>Volume of burette (mL)</i>					

1. Sketch the graph of volume graduated cylinder vs volume of burette.

Part B: Density of an Unknown Liquid

Liquid label:

	<i>Trial 1</i>	<i>Trial 2</i>	<i>Trial 3</i>
Weighing Result:			
<i>Mass of graduated cylinder (g)</i>
<i>Total mass: graduated cylinder + liquid (g)</i>
<i>Mass of liquid (g)</i>
Measuring of volume:			
<i>Volume of liquid in graduated cylinder (mL)</i>
<i>Volume of liquid from calibration graph (mL)</i>
Calculating of density:			
<i>Density of liquid</i>
<i>Average density</i>		

Part C: Density of an Irregularly Shaped Solid

Solid object label:

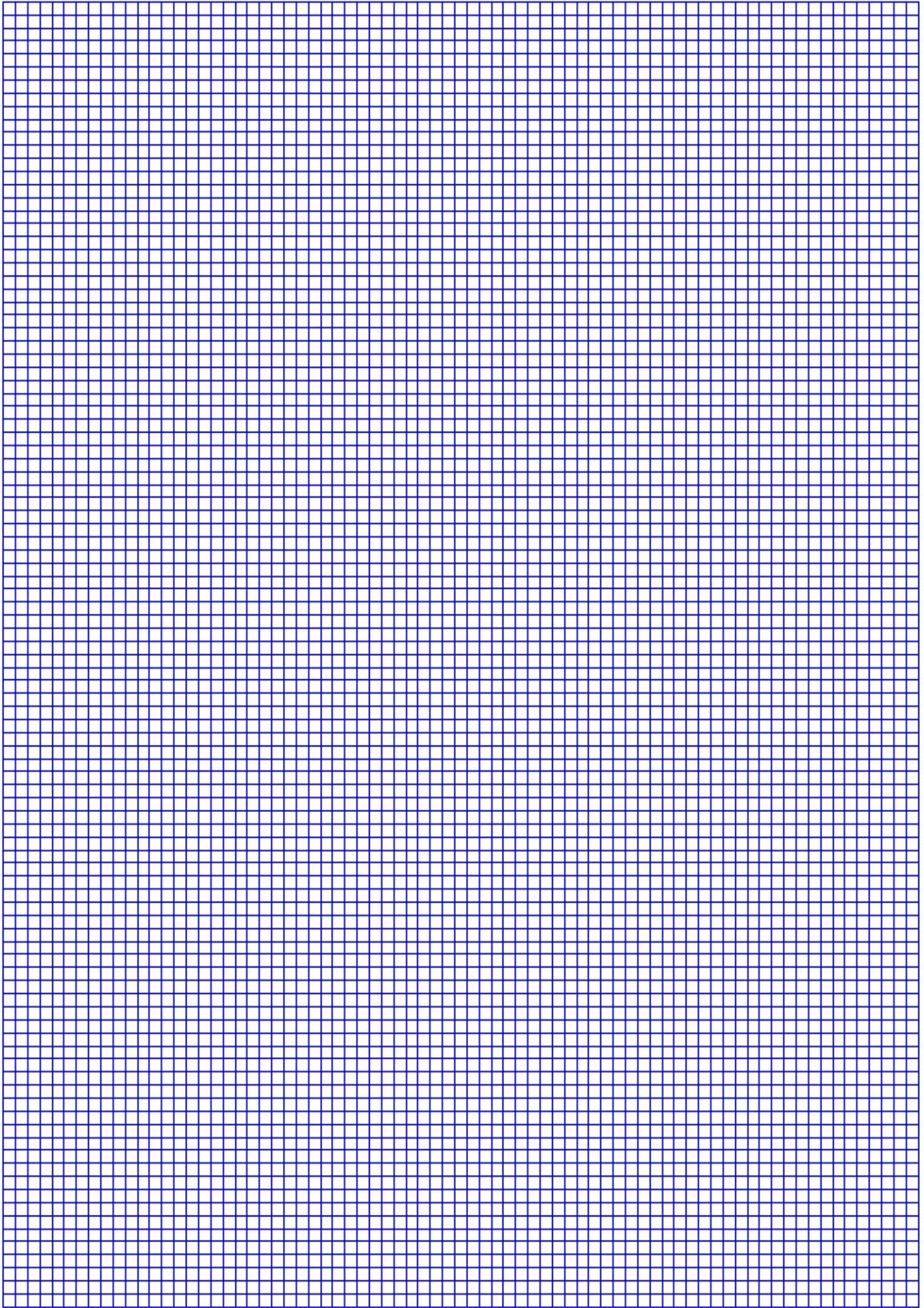
	Trial 1	Trial 2	Trial 3
Weighing Result:			
Mass of solid object(g)
Measuring of volume:			
Volume of water in graduated cylinder(mL)
Total volume :water+ solid object (mL)
Volume of solid object (mL)
Volume of solid object from calibration graph (mL)
Calculating of density:			
Density of solid object
Average density		

B. QUESTIONS

1. Calculate the density of a rectangular solid which has a mass 25.71g. It is 2.30 cm long, 4.01 cm wide and 1.82 cm high.

2. The density of pure gold is 19.3 g/mL. A 'gold' bracelet has a mass of 48.2 g and displaces 3.00 mL of water. Is the bracelet pure gold?

3. The mass of the Earth is 5.98×10^{24} kg and average diameter is 12700 km. assuming the earth is spherical, what is the average density of our planet in g/cm³



**LAB REPORT
EXPERIMENT 3**

Preparation and Standardization of Solution

Student Name:	Course:	Section:
Instructor:	Table No.	Date:

A. DATA AND RESULT

Part A: Preparation of solution

Preparation of solution result:

Mass of beaker(g) :

Total mass: beaker + oxalic acid (g) :

Mass of oxalic acid (g) :

Molar mass of oxalic acid(g/mol) :

Volume of oxalic acid solution (mL) :

1. Calculate the concentration of solution in g/mL, g/L and molarity

Part B: Dilution of NaOH solution

Dilution of NaOH solution **result:**

Initial volume of NaOH solution (mL) :

Final volume of NaOH solution (mL) :

Part C: Titration

Molarity of standard oxalic acid (M) :

Volume of NaOH solution for titration (mL) :

Titration:	Trial 1	Trial 2	Trial 3
Initial reading (mL)
Final reading (mL)
Volume of oxalic acid(mL)
Average of used oxalic acid (mL)		

B. QUESTIONS

1. Calculate the molarity of diluted NaOH and standard NaOH (already prepared for you in lab)

LAB REPORT
EXPERIMENT 4
Vinegar Analysis

Student Name:	Course:	Section:
Instructor:	Table No.	Date:

A. DATA AND RESULT

Concentration NaOH Solution (M)

Vinegar Sample (brand)

Volume of Vinegar (mL)

Mass of volumetric flask (g)

Mass of volumetric flask + vinegar sample (g)

Mass of vinegar sample (g)

Titration:	Trial 1	Trial 2	Trial 3
Initial reading (mL)
Final reading (mL)
Volume of NaOH (mL)
Average of used NaOH (mL)		

B. QUESTIONS

1. Calculate molar concentration of acid acetic in the vinegar sample after titration

2. Calculate the percent by mass (% w/w) of acid acetic in this vinegar?

3. Compare with minimum federal standard, calculate the percent of differences of vinegar?
Give comment.

4. Why vinegar sample must to be diluted before titration?

LAB REPORT
EXPERIMENT 5

Boyle's Law: Pressure-Volume Relationship in Gases

Student Name:	Course:	Section:
Instructor:	Table No.	Date:

A. DATA AND RESULT

Data and calculation:

<i>Volume (mL)</i>	<i>Pressure (kPa)</i>	<i>Constant, k</i>

B. QUESTIONS

1. If the volume is *doubled* from 5.0 mL to 10.0 mL, what does your data show happens to the pressure? Show the pressure values in your answer.

2. If the volume is *halved* from 20.0 mL to 10.0 mL, what does your data show happens to the pressure? Show the pressure values in your answer.

7. What experimental factors are assumed to be constant in this experiment?
8. One way to determine if a relationship is inverse or direct is to find a proportionality constant, k , from the data. If this relationship is direct, $k = P/V$. If it is inverse, $k = P \times V$. Based on your answer to Question 4, choose one of these formulas and calculate k for the seven ordered pairs in your data table (divide or multiply the P and V values). Show the answers in the third column of the Data and Calculations table.
9. How *constant* were the values for k you obtained in Question 8? Good data may show some minor variation, but the values for k should be relatively constant.
10. Using P , V , and k , write an equation representing Boyle's law. Write a verbal statement that correctly expresses Boyle's law.

2. Calculate the enthalpy reaction

3. Which reaction is endothermic or exothermic? Write the thermochemical equation.

4. Student performs a calibration experiment to determine the heat capacity of the calorimeter. The student mixes 25.0 mL of warm water at 41.2 °C and 35.0 mL of cool water at 15.1 °C. The final temperature of the mixture is 19.8 °C. Assume that the specific heat of water is 4.184 J/g °C and that the density of water is 1.00 g/mL. Calculate the heat capacity of the calorimeter.