Faculty of Mechanical Engineering
Universiti Teknikal Malaysia Melaka

CHEMISTRY
DMCU1233

Laboratory Experiment and Tutorial

Imran Syakir Mohamad
Mohd Haizal Mohd Husin

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 Assistant. 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
B. Erlenmeyer Flask
C. Graduated cylinder
D. Wash bottle
E. Burette
F. Ring stand
G. Burette clamp
H. Graduated pipette / volumetric pipette

I. Pipette filler
J. Test tube
K. Dropper
L. Funnel
M. Spatula
N. Test tube rack
O. Crucible tongs
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 glassware that are not immediately required, place dirty glassware 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 0.0001g 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

a) 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.
b) 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).
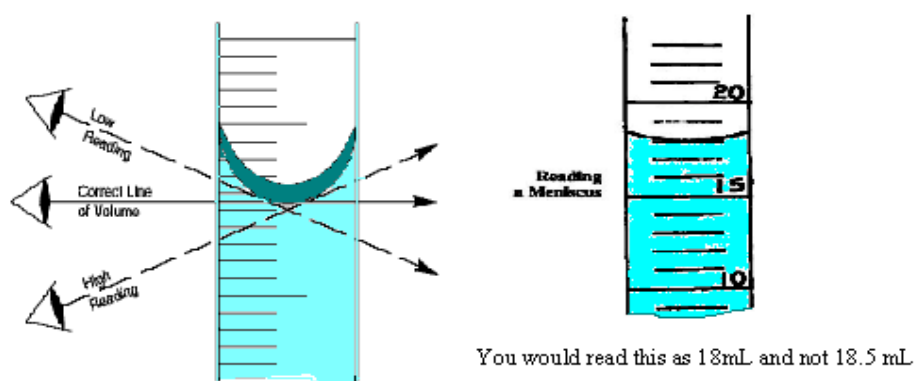c) The following lab equipment will be used for dispensing, measuring or mixing liquids during the year:

You would read this as 18mL and not 18.5 mL

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

1. GRADUATED CYLINDERS
   i.   These are designed to measure volumes of liquids accurately.
   ii. No chemicals are to be mixed in a graduated cylinder.
   iii. A graduated cylinder should never be heated or placed in an oven.
   iv. 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
   i.   Only approximate volumes are measured in beakers and flasks.
   ii. Beakers and flasks are primarily used for mixing and heating chemicals.

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

4. VOLUMETRIC FLASKS
   i. 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!
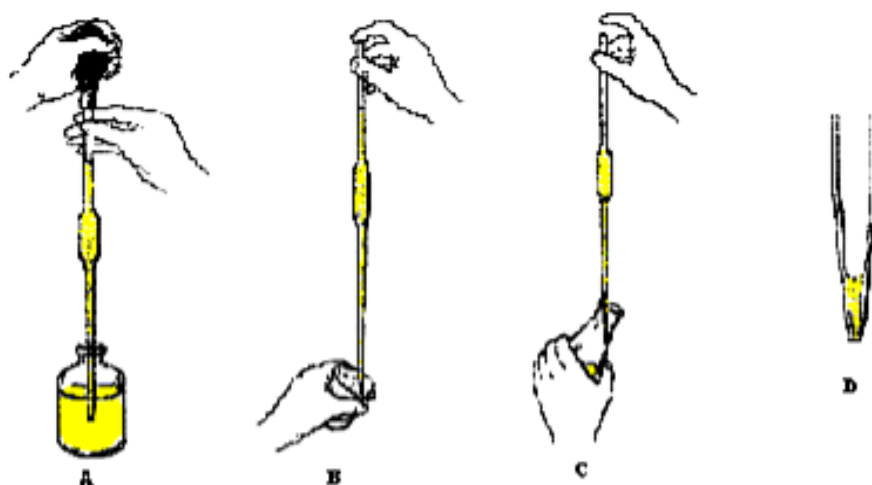

Figure 1.3: Proper technique for using pipette

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 that 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 mare 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 measured 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
 a. determine the average value from a series of measurements
 b. determine the deviation of each measurement from the average value
 c. 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 term of the error. The error is the difference between the values of quantity as measured and the accepted value of the same quantity:

Error = measured value – accepted value

When the error in a measurement is put on a relative basis it become 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*.

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

EXAMPLE: in the determine 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 the 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 average value and multiplied by 100%.

$$Re\mathit{lative}\quad average\quad deviation = \frac{Average\quad deviation}{Average\quad of\quad measurements} x \quad 100\%$$

$$Re\mathit{lative}\quad average\quad deviation = \frac{0.00028}{0.1022} x \quad 100\% = 0.27\%$$

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.

$$Percent\quad error = \frac{+0.0008}{0.1014} x \quad 100\% = + \quad 0.8\%$$

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 produce 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.

a) The graph must have a title. Title your graph. Make sure you describe what is represented on the horizontal and vertical axes.
b) 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.
c) 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

Derived units with special names

Physical Quantity	Unit	Symbol	SI base units (other units)
Frequency	hertz	Hz	s^{-1} (1/s)
Energy	joule	J	$m^2 \cdot kg \cdot s^{-2}$ (N·m)
Force	Newton	N	$m \cdot kg \cdot s^{-2}$ (m·kg/s^2)
Pressure	Pascal	Pa	$m^{-1} \cdot kg \cdot s^{-2}$ (N/m^2)
Electric charge	coulomb	C	A·s
Temperature	Celsius	°C	K ($t_{°C} = t_K - 273.15$)
Volume	Liter	L	dm^3 (L)

Other quantities and units

Physical Quantity	Name	Symbol	Expression in terms SI base units
area	square meter	m^2	m^2
volume	cubic meter	m^3	m^3
speed, velocity	meter per second	m/s	$m \cdot s^{-1}$
acceleration	meter per second squared	m/s^2	$m \cdot s^{-2}$
density	kilogram per cubic meter	kg/m^3	$kg \cdot m^{-3}$
heat capacity	joule per kelvin	J/K	$kg \cdot m^2 \cdot s^{-2} \cdot K^{-1}$
molar heat capacity	joule per kelvin mole	J/(K·mol)	$kg \cdot m^2 \cdot s^{-2} \cdot K^{-1} \cdot mol^{-1}$

Conversion of units refers to **conversion factors** between different units of measurement for the same quantity.

Mass unit
453.6 grams = 1 pound (lb)

1 metric ton = 1000 kg

1 kilogram = 1000 gram = 2.205 pound (lb)

1 ounce = 28.35 grams

1 gram = 1000 milligram

Volume Unit
1 liter (L) = 1000 millimeter (mL) = 1000 cm^3

1 cubic foot = 28.316 L

1 milliliter (mL) = 0.001 L = 1 cm^3

1 gallon = 4.546 09 L

1 liter (L) = 1 dm^3 = 0.001 m^3

Length unit
1 mile = 1.609 kilometer = 5280 feet

2.54 cm = 1 inch

1 meter (m) = 100 cm = 39.37 inches

1 yard = 36 inches = 0.914 m

1 millimeter = 0.001 m = 0.1 cm

1 Angstrom (Å) = 1.0 x 10^{-8} cm = 0.10 nm

Energy
1 joule = 10^7 ergs

1 calorie = 4.184 joule

UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DMCU1233	No. Isu./Tarikh 1/1-9-2019
CHEMISTRY **Experiment 1: Usage and Calibrate of Lab Glassware Equipment**	No. Semakan/Tarikh 1/1-9-2019	Jumlah Mukasurat 4

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 equipment 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, whereas systematic errors is occur when the equipment used are imperfect or faulty and were not calibrated before use. These errors could be minimized by checking and calibrating the equipment 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 measured (by using Equation 1.1) with standard density of water at ambient temperature.

$$density = \frac{mass}{volume}$$ (Eq. 1.1)

8. Compare this calculated density to its actually density (via temperature) using the percent error formula shown in Equation 1.2:

$$\frac{|Theoretical - Experiment|}{Theoretical} x \ 100\% = \%Error \qquad \text{(Eq. 1.2)}$$

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 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 ½ 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 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 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 measured with standard density of water at ambient temperature.

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): ……………………

	Measured volume (ml)	Beaker A (g)	Beaker A with water (g)	True volume (g)	Density of water (g/mL)	Percent error (%)
Burette						
Pipette						
Volumetric flask						
Measuring cylinder						
Beaker						

Average result:

	Average percent error	*Rank of most precise equipment*
Burette	…………..…………..	…………..…………
Pipette	…………….………..	…………….………
Volumetric flask	……………….…....	……………………..
Measuring cylinder	……………………..	……………………..
Beaker	……………………..	……………………..

B. QUESTIONS

1. How do the finding compare to your expectations? Explain

2. What is the difference in terms of usage for?

   a. volumetric flask

   b. burette

   c. measuring cylinder

3. Why is that the last drop of water or solution should not be blown out during the transfer using a pipette?

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 (Equation 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, it can found indirectly from the amount of a liquid it displaces. This technique is called **volume displacement** (Figure 2.2).

$$density = \frac{mass}{volume}$$ (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 mL = 1 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.
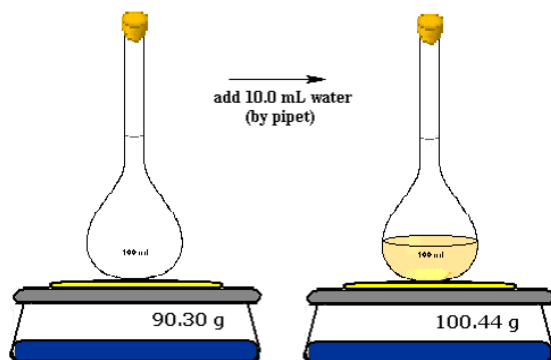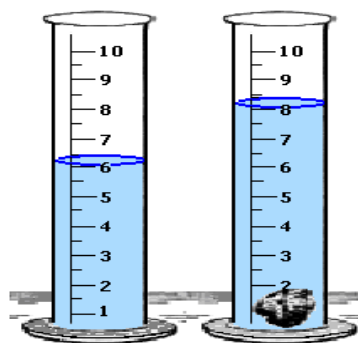
Figure 2.1: Weighing By Difference

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 used 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. Till 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.

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

Volume of graduated cylinder (mL)	2.0	4.0	6.0	8.0	10.0
Volume of burette (mL)					

1. Sketch the graph of volume graduated cylinder (x-axis) vs volume of burette (y-axis).

Part B: Density of an Unknown Liquid

Liquid label: ………………………

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

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 25.71 g. 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 x 10^{24} kg and average diameter is 12700 km. assuming the earth is spherical, what is the average density of our planet in g/cm^3
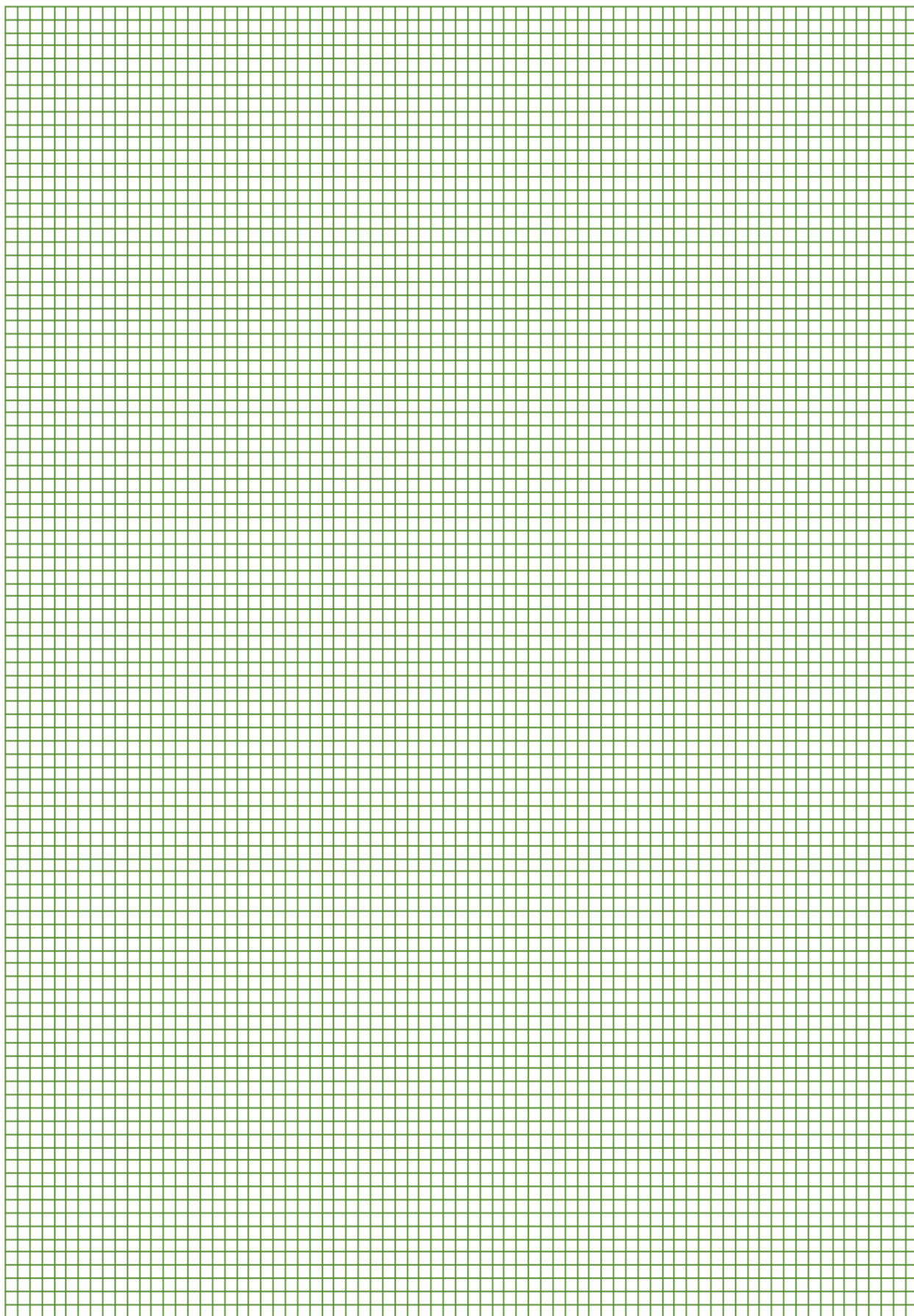
UNIVERSITI TEKNIKAL MALAYSIA MELAKA	No. Dokumen: TB/MMK/ DMCU1233	No. Isu./Tarikh 1/1-9-2019
CHEMISTRY **Experiment 3: Preparation and Standardization of Solution**	No. Semakan/Tarikh 1/1-9-2019	Jumlah Mukasurat 3

OBJECTIVE:

To prepared and determine the concentration of solution

LEARNING OUTCOMES

After conducting this experiment, you should be able to:
1. 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.

$$M_iV_i = M_fV_f \qquad \text{(Eq. 3.1)}$$

where M_i and V_i the molarity and volume of the solution before dilution, M_f and V_f 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:

$$H_2C_2O_4 \bullet 2H_2O + 2NaOH \rightarrow Na_2C_2O_4 + 2H_2O \qquad \text{(Eq. 3.2)}$$

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

$$\frac{M_aV_a}{a} = \frac{M_bV_b}{b} \qquad \text{(Eq. 3.3)}$$

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 ($H_2C_2O_4 \bullet 2H_2O$)	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)
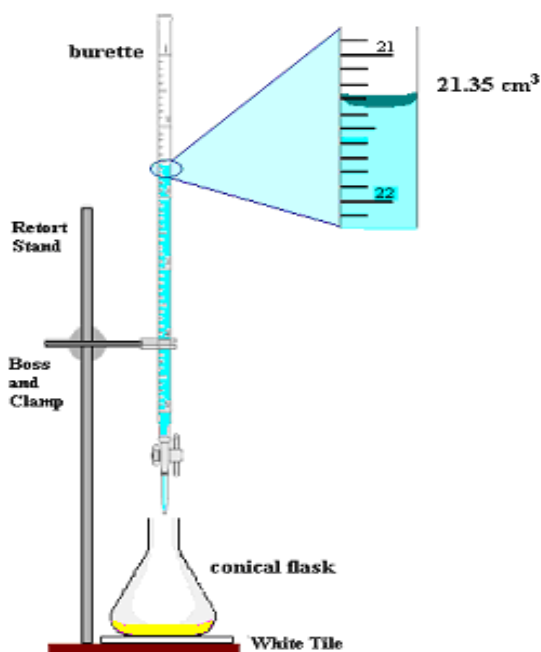

Figure 3.1: Experimental setup for titration process

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)

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 are 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:

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow NaCH_3CO_{2(aq)} + H_2O_{(l)} \qquad \text{(Eq. 4.1)}$$

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:

$$Mass(g)\ of\ CH_3COOH = mol\ CH_3COOH \times 60.05\ g/mol\ CH_3COOH \qquad \text{(Eq. 4.2)}$$

Finally, the percent by mass of CH₃COOH in vinegar is calculated:

$$\% \, by \; mass \quad of \; CH_3COOH = \frac{mass \;(g)\; of \; CH_3COOH}{mass \;(g)\; of \; vinegar} \times 100\%$$
(Eq. 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:

2. Clean 100 mL volumetric flask, and measure its mass.
3. Add the 10 mL of the vinegar to a volumetric flask with previously measured mass.
4. Measure the combined mass of the volumetric flask and sample of vinegar.
5. Dilute a concentrated solution with fill the volumetric flask to the 100 mL mark using distilled water, gently shake the solution.
6. 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.
7. 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.
8. 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.
9. Repeat the titration procedure above (step 5 to 7), 2 more times.

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?

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.
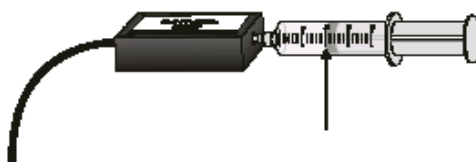
Figure 5.1: Experimental setup for Boyle's Law

LAB EQUIPMENT

Windows PC Vernier computer interface   Vernier Gas Pressure Sensor
Logger *Pro*   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.
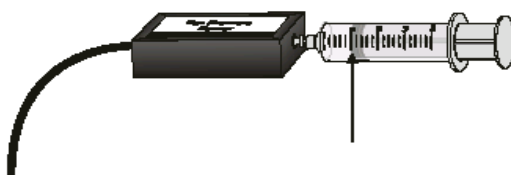

Figure 5.2: Experimental setup for Boyle's law

LAB REPORT

EXPERIMENT 5

Boyle's Law: Pressure-Volume Relationship in Gases

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

A. DATA AND RESULT

Volume (mL)	Pressure (kPa)	Constant, k

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.

3. If the volume is *tripled* from 5.0 mL to 15.0 mL, what does your data show happened to the pressure? Show the pressure values in your answer.

4. From your answers to the first three questions *and* the shape of the curve in the plot of pressure versus volume, do you think the relationship between the pressure and volume of a confined gas is direct or inverse? Explain your answer.

5. Based on your data, what would you expect the pressure to be if the volume of the syringe was increased to 40.0 mL? Explain or show work to support your answer.

6. Based on your data, what would you expect the pressure to be if the volume of the syringe was decreased to 2.5 mL? Explain or show work to support 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 **x** 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, writes an equation representing Boyle's law. Write a verbal statement that correctly expresses Boyle's law.

Chapter 1

1. Give an example for each of the following terms:
   a. Matter   b. Substance   c. Mixture

2. Give an example of homogeneous mixture and an example of heterogeneous mixture.

3. Explain the difference between a physical property and chemical property.

4. As the following statements describe chemical and physical property.
   a. Water boils below 100°C on top of a mountain.
   b. Lead is denser than aluminum.
   c. Sugar tastes sweets.

5. Give the SI units for expressing the following:
   a. Length   b. volume   c. mass   d. time   e. energy   f. temperature

6. Write the numbers presented by the following prefixes:
   a. mega-   b. kilo-   c. deci-   d. centi-   e. milli-   f. micro-   g. nano-   h. pico-

7. An oversea flight leaves New York in the late afternoon and arrives in London 8.50 hours later. The airline distance from New York to London is about 5.6×10^3 km. What is the average speed of the plane in kilometers per hours?   (658.8 km/h)

8. Express normal body temperature, 98.60°F in °C and Kelvin.   (37.0°C, 310.15 K)

9. Carry out the following conversions.
   a. 76.3 mm to meters   (0.0763 m)
   b. 482.2 in^3 to cubic centimeters   (7902 cm^3)
   c. 476 cm^2 to square inches   (73.8 in^2)

10. An average adult has 5.2 L of blood, what volume of blood in m^3?   (5.2×10^{-3} m^3)

11. A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb), what is this mass in milligrams (mg)?   (3.78×10^4 mg)

12. To determine the density of ethyl alcohol, a student pipettes a 5.00 mL sample into an empty flask weight 15.246 g, he finds that the mass of the flask + ethyl alcohol =19.171 g. Calculate the density of ethyl alcohol.   (0.785 g/mL)

13. The density of ethanol, a colorless liquid is 0.798 g/mL. Calculate the mass of 17.4 mL of the liquid?   (13.9 g)

14. Carry out the following conversions
    a. 90.0 km/h to meters per second.   (25.0 m/s)
    b. 4.17 g/s to kilograms per hour.   (15.0 kg/h)
    c. 1.39 ft/s to kilometer per hour.   (1.53 km/h)

15. The commonly accepted measurement now used by dietary specialist in assessing whether a person is overweight is the body mass index (BMI). BMI is based on a person's weight and height. It is the mass, in kilograms, divided by the square of the height in meters, which is, expressed in kg/m^2. Generally speaking, if the BMI exceeds 25, a person considered overweight. What is the BMI of a person is 69.0 inches tall and weight 158 lb?   (23.3 kg/m^2)

Chapter 2

1. Compare the properties of the following: proton, neutron and electron? Describe the structure of the atom?

2. What is the difference between an atom and a molecule?

3. Explain the meaning of each term in the symbol $_Z^A X$

4. Give the number of protons, neutrons and electrons in each of the following species
    a) $^{17}_8O$ b) $^{199}_{80}Hg$ c) $^{200}_{80}Hg$ d) $^{14}_7N^{3-}$ e) $^{54}_{26}Fe^{2+}$ f) $^{31}_{15}P^{3-}$ g) $^{130}_{53}I^-$ h) $^{107}_{47}Ag^+$

5. Write the names and symbols for four elements in each of the following categories:
    a) nonmetal   b) metal   c) metalloid

6. Define with two examples, the following terms:
    a) alkali metals   b) alkaline earth metals   c) halogens   d) noble gases

7. Give an example of each of the following:
    a) a monatomic   b) a monatomic anion   c) a polyatomic cation   d) a polyatomic anion

8. Identify the following as elements or compounds: NH_3, N_2, S_8, NO, CO, CO_2, H_2, SO_2

9. Give number of protons, neutrons and electrons for each of the following common ions: Na^+, Ca^{2+}, Al^{3+}, Fe^{2+}, I^-, F^-, S^{2-}, O^{2-}, N^{3-}, K^+, Mg^{2+}, Fe^{3+}, Br^-, Mn^{2+}, C^{4-}, and Cu^{2+}

10. Define molecular formula and empirical formula. What are the similarities and differences between the empirical formula and molecular formula of a compound?

11. What are the empirical formulas of the following compounds:
    a) C_2N_2 b) C_6H_6 c) C_9H_{20} d) P_4O_{10} e) B_2H_6 f) Al_2Br_6 g) $Na_2S_2O_4$

12. Explain why the formula HCl can represent two different chemical systems?

13. Name the following compounds:
    a) HBr   b) $Al(OH)_3$ c) HNO_2 d) Fe_2O_3 e) $CaCl_2$ f) Li_3N
    g) NH_4NO_2 h) H_2O i) $KMnO_4$ j) $Na_2S_2O_4$ k) $K_2Cr_2O_7$ l) $TiCl_4$

14. Write the formulas for the following compounds:
    a) potassium chloride   b) ammonia   c) barium iodide   d) hydrogen bromide
    e) sodium chlorate   f) titanium (IV) oxide   g) nitric acid   h) hydrobromic acid

15. Predict the formula and name of a binary compound formed from the following element:
    a) Na and H   b) B and O   c) Al and F   d) Na and S
    e) F and O   f) Sr and Cl

Chapter 3

1. There are three naturally occurring isotopes of neon. Their percent abundances are: neon-20 (90.51%), neon-21 (0.27%), neon-22 (9.22%). Calculate weighted average atomic mass of neon.   (20.19 g/mol)

2. The atomic masses of ^6_3Li and ^7_3Li are 6.0151 amu and 7.0160 amu, respectively. Calculate the neutral abundances of these two isotopes. The average atomic mass of Li is 6.941 amu.   (92.5%, 7.5%)

3. Calculate the formula masses of:
   a) Li_2O b) $BaCl_2$ c) $Mg(NO_3)_2$ d) Al_2Br_6 e) $K_2Cr_2O_7$ f) Ba(OH)   g) $NaHCO_3$ h) $CaSO_4.2H_2O$

4. Calculate
   i. the mass, in gram, of 0.155 mol C_3H_8
   ii. the mass, in milligrams, of 2.45×10^{-4} mol of ethane, C_2H_6
   iii. the number of moles of C_4H_{10} in a 165 kg sample
   iv. the number of moles of phosphoric acid in a 760 mg sample.

5. Calculate
   a. The number of moles of aluminum in cube of the metal 5.5 cm an edge (d = 2.70 g/cm³)
   b. The volume occupied by 1.38 mol carbon tetrachloride, a liquid with a density of 1.59 g/mL

6. Phenol has composition 76.57% C, 6.43% H and 17.00% O. Determine its empirical formula.

7. Resorcinol, used in manufacturing resins drugs and other products, is 65.44% C, 5.49% H, and 29.06% O by mass. Its molecular mass is 110 g/mol. What is its molecular formula?

8. Balance the following equation:
   a. PCl_5 (s) + H_2O (l) → H_3PO_4 (aq) + HCl (aq)   c) Mg_3N_2 (s) + H_2O (l) → $Mg(OH)_2$ (s) + NH_3 (g)
   b. PbO (s) + NH_3 (g) → Pb (s) + N_2 (g) + H_2O (l)   d) Fe (s) + O_2 (g) (s) → Fe_2O_3 (s)

9. All alkali metals react with water to produce hydrogen gas and the corresponding alkali metal hydroxide. A typical reaction is that between lithium and water.
   $$Li (s) + H_2O (l) → LiOH (aq) + H_2 (g) \text{ (not balanced)}$$
   a. How many moles of H_2 will be formed by the complete reaction of 6.23 moles of Li with water?
   b. How many grams of H_2 will be formed by the complete reaction of 80.57 g of Li with water?
   c. How many grams of LiOH will be formed by the complete reaction of 80.57 g of Li with water?

10. Urea [$(NH_2)_2CO$] is prepares by reacting ammonia with carbon dioxide:
    $$NH_3 + CO_2 → (NH_2)_2CO + H_2O \text{ (not balanced)}$$
    In one process 637.2 g of NH_3 are allowed to react with 1142 g of CO_2.
    a. Which of the reactants is limiting reagent?
    b. Calculate, the mass of $(NH_2)_2CO$ formed? And excess reagent (in grams) is left at the end of the reaction?

11. What is the molarity of an 85.0 mL ethanol (C_2H_5OH) solution containing 1.77 g of ethanol?

12. The label of a stock bottle of aqueous ammonia indicates that the solution 28% NH_3 by mass and has density of 0.898 g/mL. Calculate the molarity of solution.

13. Describe how you would prepare 5.00×10^2 mL of 1.75 M H_2SO_4 solution starting with an 8.61 M stock solution of H_2SO_4

14. How many milliliter of a 0.610 M NaOH solution are needed to neutralize 20.0 mL of a H_2SO_4 solution?
    $$NaOH + H_2SO_4 → Na_2SO_4 + H_2O \text{ (not balanced)}$$

Chapter 4

1. Describe the four quantum numbers used to characterize an electron in an atom.

2. An electron in an atom is in the $n = 3$ quantum level. List the possible values of l and ml that it can have.

3. What is electron configuration? Describe the role that the Pauli Exclusion Principle and Hund's rule play in writing the configuration of elements.

4. Explain the meaning of the symbol $4d^6$

5. How many electrons are permitted in each of the following subshells
    a. 2s   b. 4d   c. 6p   d. fifth shell   e. 5f

6. Write detailed electron configurations and an orbital diagram for the following atoms and ions:
   a. K   b. Mn^{2+} c. P   d. O^{2-} e. C   f. N
    g. Fe^{2+} h. Fe^{3+} i. Ca^{2+} j. Li   k. Cu   l. Cu^+
    m. B   n. Ne   o. Cl   p. Cl^- q. Na   r. Na^+

7. What neutral atom is presented by each of the following configurations
    a. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
    b. $1s^2\ 2s^2\ 2p^1$
    c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

8. Explain the meaning of diamagnetic and paramagnetic.

9. Define the magnetic properties (paramagnetic or diamagnetic) for the following atoms.
    a. K   b. Mn^{2+} c. P   d. C   e. Li   f. N
    g. Fe^{2+} h. Fe^{3+} i. Ca^{2+} j. Cu   k. Sr   l. Ar

10. Which of the following elements are metals, nonmetals or metalloids?
    a. As   b. Xe   c. Fe   d. Li   e. B   f. Cl
    g. Ba   h. P   i. I   j. Si   k. Sr   l. Ar

11. What is representative element? Give names and symbols of for four representative elements.

12. A neutral atom of certain element has 20 electrons. (a) Write the ground-state electron configuration of the element, (b) classify the element, and (c) determine whether the atoms of these elements are diamagnetic or paramagnetic.

13. Write the outer electron configurations for the
    a. alkali metals
    b. alkaline earth metals
    c. halogens
    d. noble gases

14. Specify the group of the periodic table in which each of the following elements is found:
    a. $[Ne]\ 3s^1$
    b. $[Ne]\ 3s^2\ 3p^3$
    c. $[Ne]\ 3s^2\ 3p^6$
    d. $[Ar]\ 4s^2\ 3d^8$

15. Group the species that are isoelectronic: Be^{2+}, F^-, Fe^{2+}, N^{3-}, He, S^{2-}, Co^{3+}, Ar

Chapter 5

1. Use the second member of each group from Group 1A to Group 7A to show that the number of valance electrons on an atom of the element is the same as its group number.

2. Use Lewis dot symbol to show the formation of aluminum oxide (Al_2O_3)

3. Explain what an ionic bond is? And name five metals and five nonmetals that are very likely to form ionic compounds

4. Show the formulation of ionic bonds between these following atoms. What are the name and formula of the compound that results
   a. Barium and iodine   b. Aluminum and oxygen

5. What is Lewis's contribution to our understanding of the covalent bond?

6. Show the formation of covalent bonds between these following atoms. What are the name and formula of the compound that result
   a. 1 carbon, 3 hydrogen and 1 chlorine   b. 1 nitrogen and 3 hydrogen

7. Write Lewis structures for the following molecules:
   a. ICl   b. PH_3 c. BF_3 d. CO_3^{2-} e. N_2H_4

8. Determine whether each of the following is polar or non polar
   a. SO_2 b. BF_3 c. CO_2

9. Give an example for each type of intermolecular forces.
   a. dipole-dipole interaction   b. dipole-induced dipole interaction   c. ion-dipole interaction
   d. dispersion forces   e. van der Waals forces

10. Arrange the following sets of atoms in the expected order of increasing electronegativity, that is from lowest to highest electronegativity value
    a. Cl, Mg, Si, Al   b. P, As, N, Sb   c. As, Se, Br

11. Indicate the polar nature ($\delta+$ and $\delta-$) of the following bond.
    a. Br --- Cl   b. H --- Cl   c. I ---Cl   d. Cl ---F

12. Which of the following can form hydrogen bonds with water? CH_3OCH_3, CH_4, HF, CHOOH

13. Which of the following species are capable of hydrogen bonding among themselves? H_2S, C_6H_6, CH_3OH

14. Diethyl ether has a boiling point of 34.5°C and 1-butanol has a boiling point of 117°C:
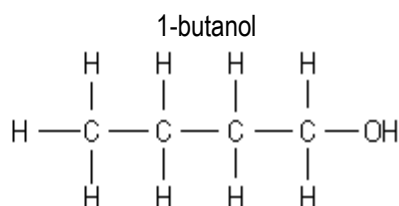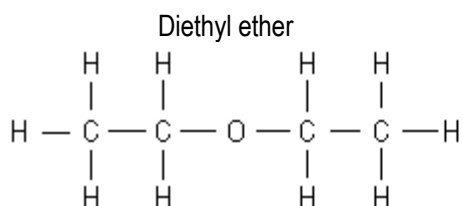
Both of these compounds have the same number and types of atoms. Explain the difference in their boiling points

Chapter 6

1. What is a phase change? Name all possible changes that can occur among the vapor, liquid, and solid phases of a substance.

2. List the physical characteristic of gas. Name five element and five compounds that exist as gases at room temperature.

3. What is the pressure in atmospheres of a gas that supports a column of mercury to a height of 808 mm
   (1.06 atm)

4. A 4.0 L sample of gas at 700 torr pressure is changed at constant temperature until its pressure is 1.315789 atm. What is its new volume? (2.80 L)

5. A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. Calculate the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL.

6. A 6.50 L sample of gas is warmed at constant pressure from 2°C to 77°C. What is its final volume
   (8.27 L)

7. A sample of carbon monoxide gas occupies 3.20 L at 125°C. Calculate the temperature at which the gas will occupy 1.54 L if the pressure remains constant.

8. Sulfur hexafluoride, SF_6 is a colorless, odorless, very unreactive gas. Calculate the pressure exerted by 1.82 moles of gas in steel vessel of volume 5.43 L at 69.5°C

9. What volume will 7.00 g of O_2 occupy at STP? (4.91 L)

10. If 0.00206 kg of a gas occupies 3.33 L at 17°C and 700 torr. What is the molar mass of the gas?
    (16.0 g/mol)

11. Calculate the density of carbon dioxide in grams per liter (g/L) at 752 mmHg and 55°C.

12. What volume does the oxygen gas liberated by heating $3x10^{-4}$ kg of $KClO_3$ until it completely decomposition to KCl and O_2 occupy? The gas is collected at STP (0.0827 L)

13. Ethanol (C_2H_5OH) burn in air.
    $$C_2H_5OH \ (l) + O_2 \ (g) \rightarrow \ CO_2 \ (g) + H_2O \ (l)$$
    Balance the equation and determine the volume of air in liters at 35.0°C and 790 mmHg required to burn 227 g of ethanol. Assume that air is 21.0 percent O_2 by volume.

14. Explain why liquids unlike gases, are virtually in compressible.

15. What is surface tension? What is the relationship between intermolecular forces and surface tension?

16. What is viscosity? What is the relationship between intermolecular forces and viscosity?

17. Describe and give examples of the following types of crystals.
    a. ionic crystals   b. covalent crystals   c. molecular crystals   d. metallic crystals

18. Explain why diamond is harder than graphite. Why graphite is an electrical conductor but diamond is not?

19. What is an amorphous solid? How does it differ from crystalline solid?

Chapter 7

1. Define and give an example of each the following terms: thermal energy, chemical energy, potential energy, kinetic energy, law of conservation of energy.

2. Explain the following terms: thermochemistry, system, surroundings, open system, closed system, isolated system, exothermic process, and endothermic process.

3. Calculate the heat evolved when 266 g of white phosphorus (P_4), burn in air according to the equation.
   $$P_4 (s) + 5\ O_2 (g) \rightarrow P_4O_{10} (s) \qquad\qquad \Delta H = -3013\ kJ$$
   (-6.47 x 10³ kJ)

4. How much heat, does it take to raise the temperature of 225 g of water from 25.0°C to 100°C. (Specific heat of water is 4.18 J g^{-1} K^{-1})
   (70.5 kJ)

5. A quantity of 1.922 g of methanol (CH_3OH) was burned in a constant-volume bomb calorimeter. Consequently the temperature of water rose by 4.20°C. If the quantity of water surrounding the calorimeter was exactly 2000 g and the heat capacity of calorimeter was 2.02 kJ/K, calculate the molar heat of combustion of CH_3OH.
   (-727 kJ/mol)

6. Synthesis gas is a mixture of carbon monoxide and hydrogen that is used to synthesis a variety of organic compounds, such as methanol. One reaction for producing synthesis gas is show here:
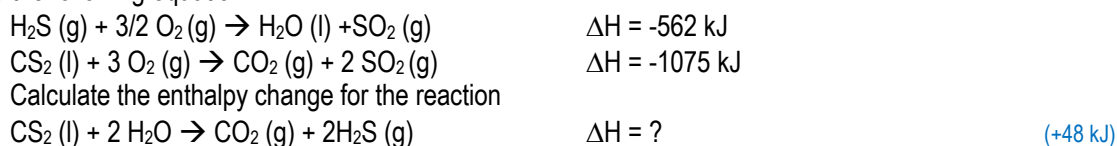
   $$3\ CH_4 (g) + 2\ H_2O (l) + CO_2 (g) \rightarrow 4\ CO (g) + 8\ H_2 (g)$$
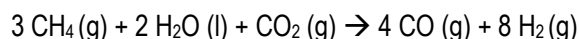
   Calculate the standard enthalpy change for this reaction if the standard enthalpies of formation
   $CH_4 (g) = -74.81$ kJ mol^{-1}, $H_2O (l) = -285.8$ kJ mol^{-1}, CO2 (g) = -393.5 kJ mol^{-1} and CO (g) = -110.5 kJ mol^{-1}
   (+747.5 kJ)

7. Ethylene derived from petroleum, is used to make ethanol for use as fuel or solvent, the reaction is
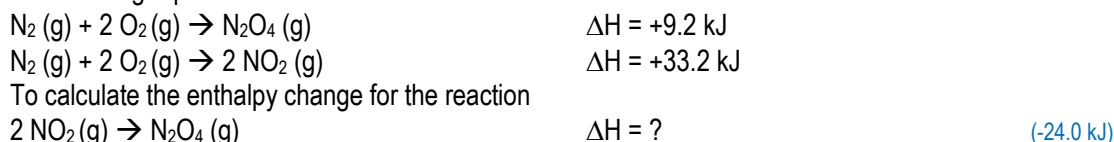 $$C_2H_4 (g) + H_2O (l) \rightarrow CH_3CH_2OH (l)$$
   Calculate $\Delta H°$ for this reaction, when the standard enthalpies of formation $C_2H_4 = 52.3$ kJ/mol, $H_2O = -285.8$ kJ/mol, and $CH_3CH_2OH = -277.7$ kJ/mol
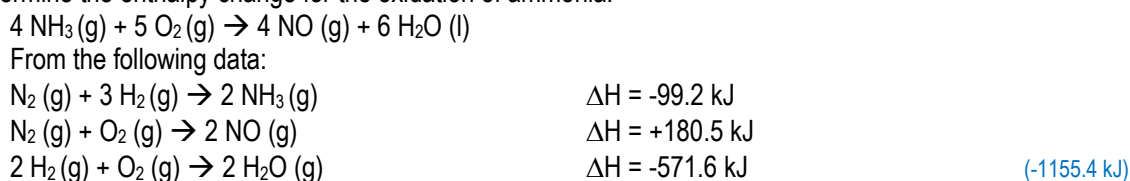   (-44.2 kJ)

8. Use the following equation:
$H_2S (g) + 3/2\ O_2 (g) \rightarrow H_2O (l) + SO_2 (g)$	$\Delta H = -562$ kJ
$CS_2 (l) + 3\ O_2 (g) \rightarrow CO_2 (g) + 2\ SO_2 (g)$	$\Delta H = -1075$ kJ

   Calculate the enthalpy change for the reaction
   $CS_2 (l) + 2\ H_2O \rightarrow CO_2 (g) + 2H_2S (g)$ $\Delta H = ?$ (+48 kJ)

9. Use the following equation:
$N_2 (g) + 2\ O_2 (g) \rightarrow N_2O_4 (g)$	$\Delta H = +9.2$ kJ
$N_2 (g) + 2\ O_2 (g) \rightarrow 2\ NO_2 (g)$	$\Delta H = +33.2$ kJ

   To calculate the enthalpy change for the reaction
   $2\ NO_2 (g) \rightarrow N_2O_4 (g)$ $\Delta H = ?$ (-24.0 kJ)

10. Determine the enthalpy change for the oxidation of ammonia.
    $4\ NH_3 (g) + 5\ O_2 (g) \rightarrow 4\ NO (g) + 6\ H_2O (l)$
    From the following data:
$N_2 (g) + 3\ H_2 (g) \rightarrow 2\ NH_3 (g)$	$\Delta H = -99.2$ kJ
$N_2 (g) + O_2 (g) \rightarrow 2\ NO (g)$	$\Delta H = +180.5$ kJ
$2\ H_2 (g) + O_2 (g) \rightarrow 2\ H_2O (g)$	$\Delta H = -571.6$ kJ

    (-1155.4 kJ)

LAB ASSESSMENT

Code/Subject: **DMCU 1233 CHEMISTRY**

Title of Experiment: _____

Section/Group   : _____

		Very Weak	**Weak**	**Modest**	**Good**	**Excellent**
		0	**1**	**2**	**3**	**4**
PROCEDURE	**Ability to choose and use equipment**	Unable to identify and mishandling the equipment	Able to identify but unable to use equipment	Minor mistakes while using equipment	No mistakes while using equipment, but not comfortable	No mistakes & comfortable with the equipment
WORK PROCESS	**Ability to construct and understanding of experiment**	Work process is missing or unclear	Major work process are missing	Some work process are missing	Most of the work process are followed correctly	Work process are followed correctly and in order
	Ability to show data	All data are missing or incorrect	Most data are missing or incorrect	Most data are correctly taken, no missing data	All data are taken but minor data are incorrect	All data are correctly taken
OUTCOME	**Completion of the experiment**	Experiment not completed.	Some of minor part of the experiment is complete.	All minor part of the experiment is complete.	Some major part of the experiment is complete.	Experiment is 100% complete.
DISCIPLINE	**Ability to perform lab and safety regulations**	Major flaws to criteria, > 10 minutes late	Minor flaws to criteria, < 10 minutes late	Minor flaws to dress code and safety	Conform to dress code, minor flaw to safety	Conform to all criteria

PERIODIC TABLE OF THE ELEMENTS

1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.94	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.085	15 P Phosphorus 30.974	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.630	33 As Arsenic 74.922	34 Se Selenium 78.97	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium [97]	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.60	53 I Iodine 126.904	54 Xe Xenon 131.293
55 Cs Cesium 132.905	56 Ba Barium 137.327	* 57 - 70	71 Lu Lutetium 174.967	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.084	79 Au Gold 196.997	80 Hg Mercury 200.592	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [209]	85 At Astatine [210]	86 Rn Radon [222]
87 Fr Francium [223]	88 Ra Radium [226]	** 89 - 102	103 Lr Lawrencium [262]	104 Rf Rutherfordium [267]	105 Db Dubnium [270]	106 Sg Seaborgium [269]	107 Bh Bohrium [270]	108 Hs Hassium [270]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [281]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [293]	118 Og Oganesson [294]

*Lanthanide series	57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.242	61 Pm Promethium [145]	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.045
**Actinide series	89 Ac Actinium [227]	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium [237]	94 Pu Plutonium [244]	95 Am Americium [243]	96 Cm Curium [247]	97 Bk Berkelium [247]	98 Cf Californium [251]	99 Es Einsteinium [252]	100 Fm Fermium [257]	101 Md Mendelevium [258]	102 No Nobelium [259]