COLLEGE OF SCIENCES
CHEMISTRY DEPARTMENT

CHEM 101
LABORATORY OF GENERAL CHEMISTRY (1)

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1445 - 2023

CONTENTS AND EXPERMINTS TIMETABLE

INTRODUCTION

<u>Our</u> aims in this course are:

1) Recognizing the importance of safe lab practices.
2) Knowing all safety rules required to perform experiments in a safe environment.
3) Identifying basic equipment and glassware used in any chemistry laboratory.
4) Using equipment, glassware and chemicals in the safe and the proper way.
5) Realizing that chemical changes and its scientific phenomenon are present in our daily life at homes, streets everywhere and every time. Thus, every materialized things, changes and behaviors can scientifically be explained and understood.

<u>To</u> achieve these aims, we must recognize the following:

1) Because chemistry has developed largely through experiments, the study of chemistry is increased by laboratory experiences beside the principles of chemistry discussed in the classrooms.
2) The laboratory environment and its lifestyle differ considerably from the outside world and much more structured than our lifestyle at homes, or even in the classrooms.
3) Using chemicals, glassware and equipment doesn't follow our personal styles, emotions or wishes but it solely follows conventional rules that must be followed and obeyed.
4) Chemistry is the science of our life (food, transportation, buildings, medicines, clothes, batteries …etc.)

<u>This</u> laboratory course is designed in such way to be executed as if we have never encounter a chemistry laboratory and never had any preliminary experiences with chemicals, laboratory tools and equipment.

<u>This</u> course will develop an appreciation and respect to "Chemistry Laboratories" and will uncover the proper and conventional regulations and protocols.

<u>Safety</u> of persons inside and nearby laboratory is the primary concern that shouldn't be neglected or underestimated at any time and at any circumstances.

<u>Chemical</u> experiments can be carried out successfully only if we obey laboratory's rules and follow its procedures.

<u>This</u> laboratory course also will help us to understand the theoretical bases of the experiments and how they verify and explain theoretical findings.

<u>We</u> should keep in mind that "*<u>NO report is presented NO experiment is done</u>*". This simply means that any chemical experiment has to be reported in writing. The report of an experiment is a presentation of the idea of doing the experiment, how it has been done, what materials and equipment are used, which properties are measured and what are the values of these measurements, how calculations, finally what is the final findings of the experiment.

<u>Our</u> efforts and work can be evaluated easily. We can score the highest possible value as long as we work with attention and care.

<u>All</u> of the experimental works we will encounter here are and approved by the Chemistry Department at King Saud University as they are consistent, with the B SC program of the this department. They were all carried out successfully in one of the general chemistry laboratories in this department.

<u>These</u> experiments are written here exactly as they were tested and carried out.

<u>We</u>, greatly, thank the head of the Chemistry Department at King Saud University Prof. Zeid Alothman, the general chemistry course coordinator.

<u>Our</u> great appreciation must be delivered to teacher assistants and chemistry technicians: **Sultan Saad Almadhhi, Nabil Mohammed Alsahli, Mishary, Khaled Aldaghash, Mohammed Abdullah Alkhathran, Abdullah A. Adam, Suliaman Mohammed Alhmoud and Rajeh Theeb Alotaibi** whose their concerted and mutually supportive efforts made the best in testing the validity of all experiments. What we see here could have not been seen as it is without their sincere efforts.

Ahmad Alowais and Taieb Aouak
20 RABI'II 1439 – 07 JANUARY 2018

<u>Edited by Ahmed Awaji , August 2023</u>

BASIC PRINCIPLES OF LABORATORIESSAFETY RULES AND HOUSEKEEPING RULES

A) Safety rules

Work in a laboratory should be follow this safety rules with no exception:

1. **Do not violate safety rules** that given here or safety instructions given by your laboratory instructor.

2. **Locate the safety equipment**. Find the eyewash fountains, safety showers, fire extinguishers, fire blankets, first aid kit, and all of what exits to be used in case of emergency.

3. **Protect your eyes**. Wear goggles all times. If you need eyeglasses, it must be worn under goggles. You should not wear contact lenses unless allowed by your laboratory instructor.

4. **Tie long hair back**. If it is long, that will keep your hair out of burner flames and harmful liquids.

5. **Wear shoes that cover all of your feet**. Broken glass on the laboratory floor is too common. Your feet will need more protection than that afforded by open- toed shoes or sandals.

6. **Wear clothes that cover most of your body**. Good clothing can be protected with a laboratory coat. Remember lab coat is required in the chemistry labs. You may not be allowed to work in the lab without a lab coat.

7. **Do not eat or drink in the laboratory.** Foods and drinks are susceptible to contamination by chemicals that could cause a serious harm to your health. Therefore, food and drinks are not allowed in chemistry labs.

8. **Do not taste any chemical. (TASTING CHEMICALS IS STRICTLY FORBIDDEN AND PROHIBITED).** Some chemicals are very harmful and some are deadly substances. Never think or try to taste any chemical in chemistry labs.

9. **Do not smell chemicals directly.** Do not get your nose close to chemicals. Use your hand to waft the odor to your nose.

10. **Do not pipet by mouth.** Use a rubber suction bulb or special pipette filler.

11. **Do not put flammable substances near flame.** Many chemicals are flammable and cause a sudden fire break out. Never have flammable substances near flame.

12. **Do not engage in games in the laboratory.** Laboratory is a place for chemistry work; it is not a playing space. Avoid playing any games inside your lab.

13. **Do not do or watch unauthorized experiments.** In lab, just perform a specified task with your instructor. Do not do any other experiment.

14. **Do not work in the laboratory in the absence of your laboratory instructor or his/her authorized representative.** Your instructor is the only responsible person in the lab and you are not allowed to stay in the lab if the instructor is absent.

15. **Use a fume hood when required**. Some chemicals are volatile so we do the experiment that has volatile chemical in the fume hood.

16. **Handle glass tubing with care.** When tubing (including glass thermometers) is to be inserted through a rubber stopper, the tubing must be lubricated with water or glycerol. Hold the tubing with a cloth or a paper towel near the end that will be inserted, and use a twisting motion during insertion.

17. **Be aware of your neighbors.** Are they obeying the safety rules? A neighbor's accident may not hurt him/her, but it may hurt you badly. Report any unsafe behavior of your neighbor to the lab instructor.

18. **Wash your hands before leaving the laboratory.** Always put in your mind that chemical may have contaminate your hands. Therefore, wash your hands before you leave the lab.

19. **Tell your instructor about any accident or a spill immediately no matter how minor it is.** Never ignore spills or accidents. Even if you think they are minor, report them to your instructor promptly.

20. **<u>AGAIN</u>: Do not violate safety rules**. Never violate safety rules given here or safety instructions given by your laboratory instructor.

B) Housekeeping rules

Good housekeeping in the laboratory will lead to pleasant surroundings. In addition, it will provide a safe work site in which you may be assured that chemicals are not contaminated.

1. Clean up broken glass immediately with a broom and a dustpan. Do not use your hands. Special containers are available for disposal.

2. Clean up solid and liquid spills immediately, but only after checking with your laboratory instructor about possible safety hazards.

3. Do not pour any chemical into a sink without authorization. Often, disposal bottles will be provided.

4. Take containers to the stock of chemicals. Never bring stock chemicals to your stand.

5. Read the label on a bottle carefully. Is it the correct chemical?

6. Do not insert a pipette or medicine dropper into a stock bottle. Avoid contamination by pouring the liquid into one of your glassware before taking a sample.

7. Use special care with stoppers or tops of stock bottles. Do not allow them to pick up contamination. Your laboratory instructor may provide additional instructions for handling the stoppers or tops found in your laboratory.

8. Take no more of a chemical than the experiment requires.

9. Never return an unused chemical to a stock bottle. You must assume that the chemical is contaminated. It must be discarded.

10. Set up your glassware and apparatus away from the edge of the laboratory bench.

11. Follow any other housekeeping rules given by your laboratory instructor.

I have read, I understand, and I will follow the above safety rules.

Student name:

Student number:

Signature:

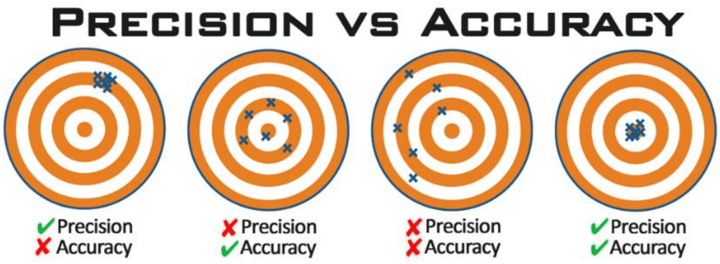
Date

EXPERIMENT (1):

EVALUATION OF ERROR, ACCURACY AND PRECISION

Accuracy is closeness of the measurements to a specific value.
Precision is the closeness of the measurements to each other.

The significant figures displayed the differences between precision and accuracy



- Experimental error (Δx) is defined as the absolute value of the difference between :

$$\Delta x = |x - x_0|$$

(x) : measured value

(x_0) : actual value

- The ***percent error*** is the ratio of the experimental error (Δx) to the actual value (x) multiplied by 100.

$$Error(\%) = \frac{\Delta x}{x_o} X\ 100$$

- The ***precision*** of a measurement is a measure of the reproducibility of a set of measurements, and is recorded as:

$$\textbf{Precision} = \mathbf{x \pm \Delta x}$$

.

Part one: Evaluation of percent error

In This part, a substance will be weighed and the percent error will be calculated.

Equipment and reagents

Balance
Beaker
Known weight Object (x_0)

Procedure

1. Zero the balance.
2. Weigh the object. Record this mass as **x**.

x_0	x

Results and calculation

Calculate the experimental error (Δ**x),** Error % and precision

Experimental error (Δ**x) =**

Error % =

Precision =

Part two: Determination of accuracy and precision

In this part, the average mass $\overline{m}$ the experimental error in mass (Δm), the percent error of the measured mass (m %), and the accuracy of mass measurement will be determined as follows:

- **Determining of $\overline{m}$**

$$\overline{m} = \frac{m_1+m_2+m_3+\cdots+m_n}{n}$$

, **n** is the number of measurements.

- **Determining of Δm:**

$$\Delta m = | \ \overline{\boldsymbol{m}} - \boldsymbol{m_{actual}} \ |$$

consider d_{water} = 1 g/ml then,
$\boldsymbol{m_{actual}}$ = amount of water taken in experiment.

- **Determining of m%**

$$\boldsymbol{m}\ \% = \frac{\boldsymbol{\Delta m}}{\boldsymbol{m_{actual}}}\ \boldsymbol{x\ 100}$$

- **Determining the accuracy of measurement:**

Accuracy in m value = $\overline{\boldsymbol{m}}$ ± Δm

Equipment and reagents

Balance. 25 or 50 mL graduated cylinder
Beaker Burette

Procedure

1. Weight an empty beaker and record the mass as $\mathbf{m}_{\text{empty beaker}}$
2. Add 25 mL of water measured carefully by a graduated cylinder into the beaker.
3. Weight the beaker with the added water and record the mass as $\mathbf{m}_{\textbf{beaker and water}}$
4. Empty the water from the beaker and repeat B, C steps two more times.
5. Repeat all of the above steps using a dry 50 mL- graduated burette.

Results and Calculation

Calculate the average $\overline{m}$, the experimental error (Δm), the percent error (m%), and the Accuracy of the mass values using Graduated cylinder and Burette.

$m_{\text{empty beaker}}$ =

$m_{\text{water}} = m_{\text{beaker and water}} - m_{\text{empty beaker}}$

A- Graduated cylinder

	m_1	m_2	m_3
$m_{\text{beaker and water}}$			
m_{water}			

$\overline{m}$ =

Δm =

m% =

Accuracy =

B- Burette

	m_1	m_2	m_3
$m_{\text{beaker and water}}$			
m_{water}			

$\overline{m}$ =

Δm =

m% =

Accuracy =

- **S**tate which is more accurate the graduated cylinder or the graduated burette.

EXPERIMENT (2)

DETERMINATION OF THE DENSITY USING DIFFERENT METHODS

Objectives

- Determination volume by two different methods.
- Using measured volumes and masses to calculate densities.
- Using the relationship between mass, volume, and density to find desired unknown quantities.
- Evaluating results using error analysis.

Theoretical information

- The density is defined as the ratio of mass, m, to volume, V:

$$\boxed{\boldsymbol{d = \frac{m}{V}}}$$

- Usually, densities units are g/ml or g/cm^3 (1 ml = 1 cm^3).
- Masses are measured on electronic balances.
- Volume is measured in many ways.
- In this experiment, you will measure masses and volumes to determine density.

Materials and equipment

A specimen (massive cylinder)
Ruler
100 mL-graduated cylinder
Balance

<u>**First method**</u>

Determination of density directly by calculation of volume and weighing mass Of a geometric specimen

If the specimen in uniform shape:

- We can take its diminution and apply it in suitable relation to find its volume.
- The mass obtain directly weighting the specimen on balance.
- Find the density by Applying on its law.

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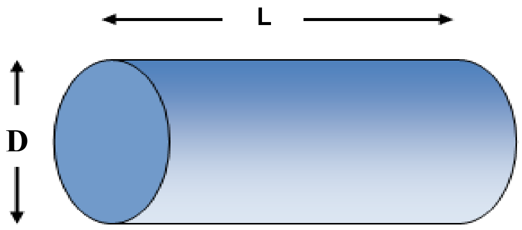


Procedure

1. Using a proper an accurate ruler, measure the length (L) and the diameter of the specimen (massive cylinder) (D).
2. Weigh the mass (m) of your specimen (massive cylinder).

Results and calculations

1. Report your measurements as follows:

L	D	m

2. Calculate the volume of your specimen (massive cylinder).

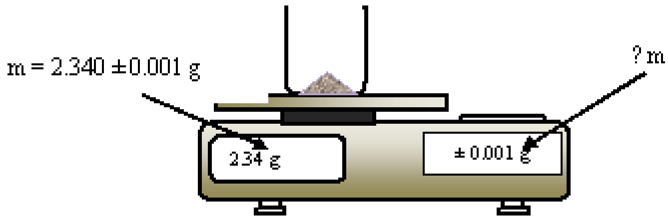
$$Volume = \frac{\pi D^2 L}{4} =$$

3. Calculate the density of your specimen (massive cylinder).

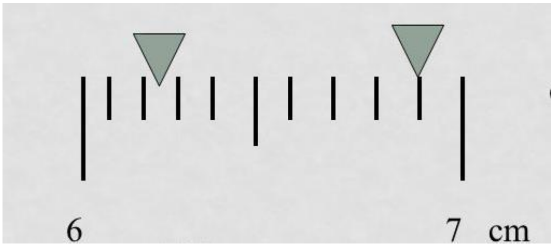
$$Density = \frac{m}{V} =$$

Evaluate the values of Δm , ΔL and ΔD as follows:

Δm is the error occurred during the weighing on the balance. The figure below shows how you find the value of **Δm** is taken as follows



ΔL and **ΔD** are the errors occurred during the measurement of the length and the diameter of the cylinder using a ruler.
The figure below shows how you find the value of **ΔL** and **ΔD**



The above figure shows two measurements.
The correct reading of the first is 6.2 cm ± 0.1 cm
The correct reading of the second is 6.8 cm ± 0.1 cm

4. Tabulate your errors of measurements:

ΔL	ΔD	Δm

5. Calculate the error in the density (Δdensity), and its accuracy:

$$\Delta Density = \pm density \left| \frac{\Delta L}{L} + \frac{\Delta D}{D} + \frac{\Delta m}{m} \right|$$

The accurate density = density ± Δdensity

<u>Second method</u>

Determination of density by displacement of water

Theoretical information

When volume of an object cannot be calculated by a mathematical equation because of its irregular shap, or desired to determine the volume without using the mathematical equation, it can be determined by water displacement.

Procedure

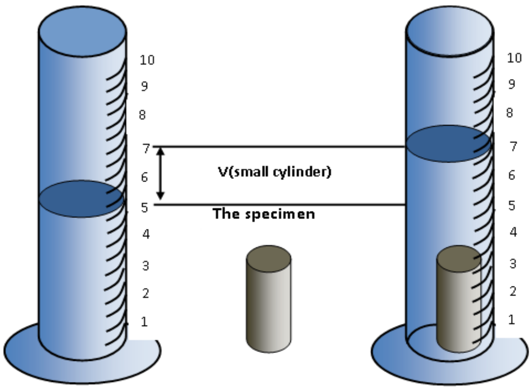
1. Pour water into the graduated cylinder. Record the exact volume as V1.
2. Place your object inside the graduated cylinder.
3. Record the new volume as V2.

(If object is the same object used in the first method, do not weigh it again)

Results, calculations and comparison

1. Report your measurements as follows:

V1	V2	m



2. Calculate the volume of your specimen :

$$V = V2 - V1$$

3. Calculate the density of specimen :

$$Density = \frac{m}{V}$$

Evaluate the values of Δm , ΔV_1 and ΔV_2 as follows:

Δm is evaluated exactly as described in the first method.

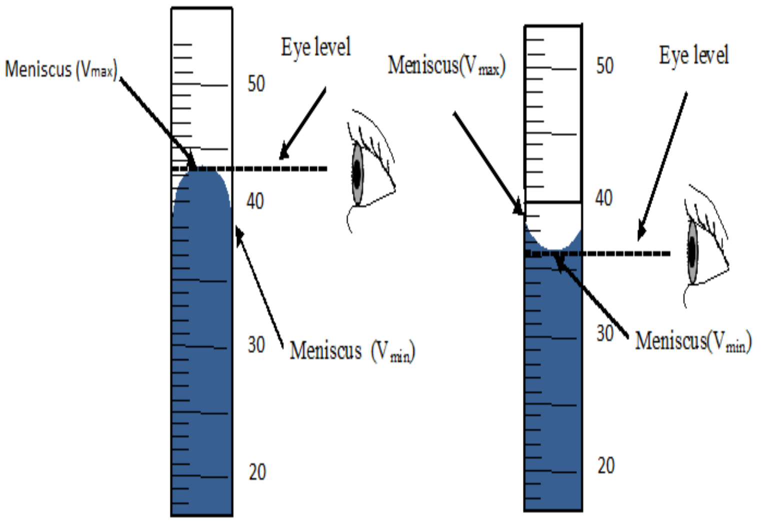
$\Delta V_1 = \Delta V_2 = \Delta V_{\text{graduated cylinder}}$

$\Delta V_{\text{graduated cylinder}}$ is written on the top of the cylinder.

If not written, it can be evaluated by using the following equation:

$$\Delta V_{graduated\ cylinder} = \frac{V_{max} - V_{min}}{2}$$

As is described in the scheme



4. Tabulate the values of experimental errors:

ΔV_1	ΔV_2	Δm

5. Calculate the error in the density (Δdensity), and its accuracy:

$$\Delta Density = \pm density \left| \frac{\Delta V1}{V1} + \frac{\Delta V2}{V2} + \frac{\Delta m}{m} \right|$$

The accurate of density = density ± Δdensity

EXPERIMENT (3):

PREPARATION OF A SOLID/LIQUID SOLUTION

Part one: Dissolving a solid solute in a solvent

Theoretical information

- ***The concentration of the solution***: is the amount of solute that is dissolved in a given quantity of solvent.
- A diluted solution contains only a small amount of solute in a given amount of solution, unlike concentrated solution contains a large amount of solute.
- Chemists mostly use molarity as concentration unit.
- ***The molarity*** M of solution: is the number of moles of solute in one liter of solution.
-

Calculated as follows: $\boxed{\text{Molarity} = \dfrac{\mathbf{n\ solute}}{\mathbf{V_{solution\ (L)}}}}$

Where n is number of moles of solute and $S_{olution\ (L)}$ is the volume of solution in liter

- If n_{solute} is not known, it can be calculated from its mass "m_{solute}" and its molar mass "Mw_{solute}" as follow:

$$\boxed{n_{solute} = \frac{m_{solute}}{Mw_{solute}}}$$

Strength of Solution (S) : The amount of solute dissolved in one liter of solution .

$$\boxed{S = \ Molarity\ x\ M_w}$$

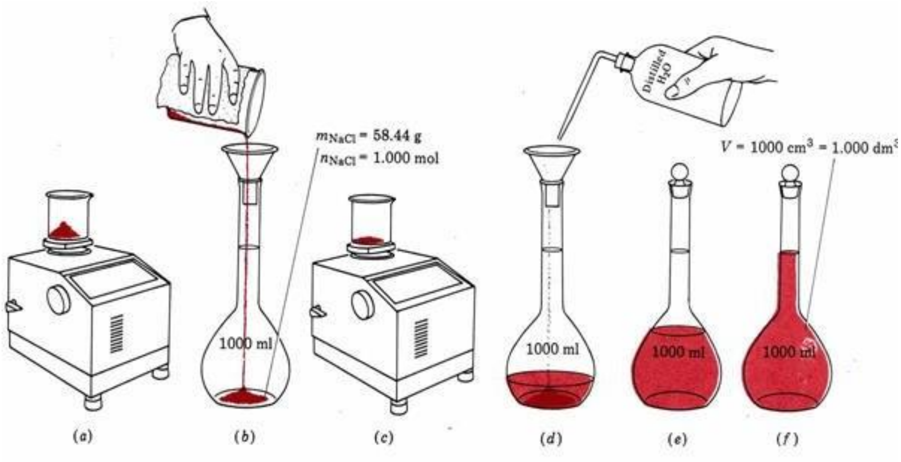
Equipment and Reagents

Balance	Small beaker
100-mL volumetric flask	Funnel
Wash bottle	Graduated Cylinder
Distillated water	Sodium chloride

Procedure

1) In a small beaker, weigh a mass of 3.00 g of NaCl.(**<u>record</u> it as m_{NaCl}**)
2) Transfer this mass to a 100-mL volumetric flask.
3) Rinse beaker and funnel with a small amount of distillated water into the flask.
4) Repeat the previous step two more timesto make sure the no residual salt .
5) Cap the flask and invert it several times to dissolve the solid.
6) Add water until the liquid is just below the etched line on the neck of the flask.
7) Bring the water to the line by adding the last few drops of water drop-by-drop using a wash bottle.
8) Cap and invert the flask several times to ensure proper mixing.
9) The final solution volume is 100 mL and contains a mass of NaCl equals to m_{Nacl} .

The preparing of 100 ml of an aqueous solution containing 3 g of solid NaCl as it showing figure below:



Results and calculations

1) Tabulate your results

m_{NaCl}	$V_{solution}$

2) Calculate the molarity of your solution.

Part 2: Preparing a solution from another solution by dilution

Theoretical information

- Laboratory stock solutions are usually concentrated.
- We will use the previous solution that already prepared in part 1 as a stock solution.

Equipment and Reagents

Graduated Cylinder — Wash bottle
100-mL volumetric flask — Distillated water
0.5 mol/L sodium chloride solution.

Procedure

1) Using a graduated cylinder transferred 20 mL of the 0.5 M NaCl solution into
   a 100- mL volumetric flask.
2) Add water until just below the etched line on the neck of the flask.
3) Bring the water to the line by adding the last few drops of water drop-by-drop using a wash bottle.
4) Cap and invert the flask several times to ensure proper mixing.
5) The final solution is diluted NaCl solution.

The figure below shows the preparing of a diluted solution from a concentrated solution.

(a) A volume (V_s) containing the desired moles of solute (M_s) is measured from a stock solution of known concentration.

(b) The measured volume of stock solution is transferred to a second volumetric flask.

(c) The measured volume in the second flask is then diluted with solvent up to the volumetric mark [$(V_s)(M_s) = (V_d)(M_d)$].

Results and calculations

1) Tabulate your results as follows:

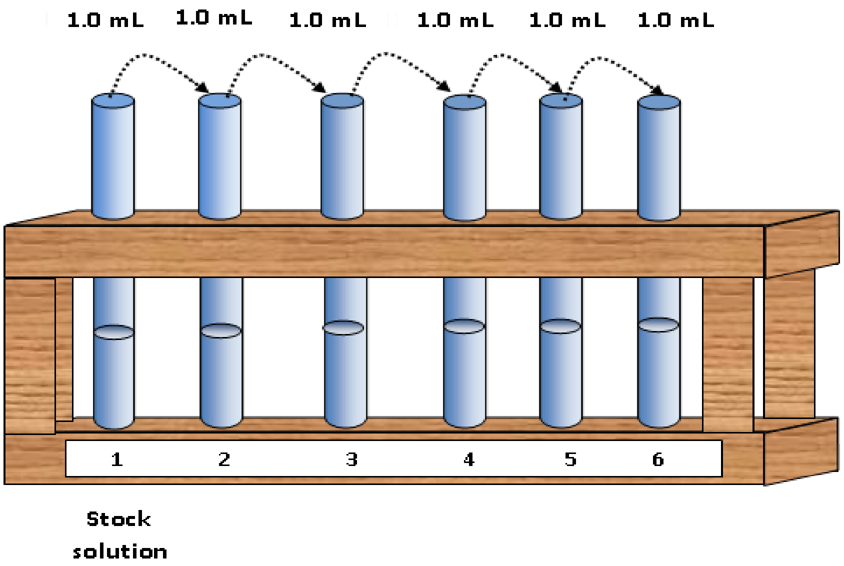
M_{conc}	$V_{initial}$	M_{dil}	V_{total}

2) Calculate the molarity of your diluted solution

$M_{conc} \times V_{initial} = M_{dil} \times V_{total}$

Part 3 : PREPARING SOLUTIONS BY SERIAL DILUTION

- Serial dilution is a series of dilutions are conducted with constant dilution factor.
- Chemist resort to serial dilution when they need to prepare very low concentration solution and there is no suitable pipette that help them to take required volume.
- The source of dilution material for each step comes from the diluted material of the previous step.
- Dilution continues until the desired concentration is reached.



To calculate the concentration of the diluted solution in the successive dilution process (M_{dil}),

$$M_{dil} = \frac{M_{conc} \times V_{initial}}{V_{total}}$$

Equipment and Reagents

10.0-mL pipette
Three test tubes in a rack.
25 ml-Graduated Cylinder
A stock solution (0.1 mol/L Nacl solution). **(We can Prepared solution in part 2**

Procedure

1. Place 3 clean and dry test tubes in a test tube rack and number them.
2. Using 10 mL-volumetric pipette transfer precisely 1.0 mL of NaCl solution 0.1 M .
3. Add water until reach line of the 10 mL in the graduated cylinder.
4. Pour this solution into the **test tube number 1**.
5. Rinse the volumetric pipette with water.
6. From the solution in test tube number number 1, transfer 1.0 mL into the graduated cylinder.
7. Add water until reach line of the 10 mLin the graduated cylinder.
8. Transfer this solution to the **test tube number 2**.
9. Repeat the steps from 5 to 8 to prepare the remaining solutions labeled as **test tube number 3**.

Results and calculations

1) Using the dilution equation, calculate the molarity of the three solutions.

$M_{dil\ (Test\ tube\ 1\ )}$ =

$M_{dil\ (Test\ tube\ 2\ )}$ =

$M_{dil\ (Test\ tube\ 3\ )}$ =

EXPERIMENT (4):
TITRATION AND DETERMINATION THE SUITABLE INDICATOR

What is a Titration?

A titration is an analytical procedure used to determine the accurate concentration of a sample by reacting it with a standard solution.

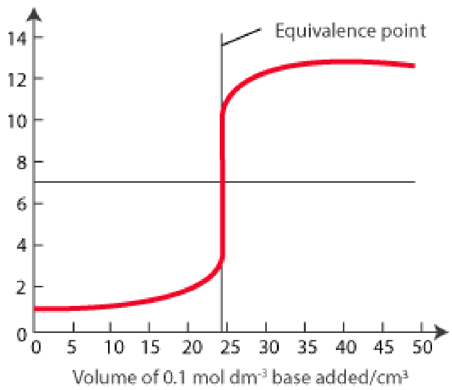
<u>End Point</u> – the pH at which the indicator changes color in solution .

<u>Equivalence Point</u> – the point at which the acid has completely reacted with or been neutralized by the base. (mol of H^+ = mol of OH^-)

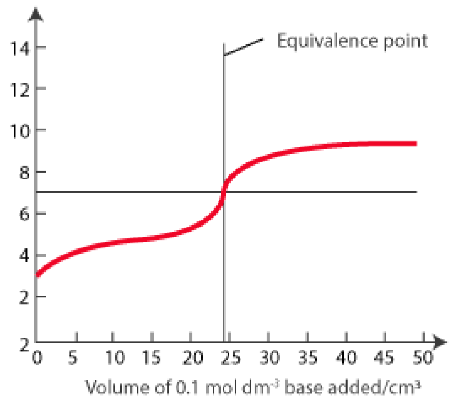
<u>Indicator</u> - a substance (weak acid) that has distinctively different colors in acidic and basic media.

Depending on the pH range of indicator we can determine the suitable indicator for each titration as we see in the graph below.

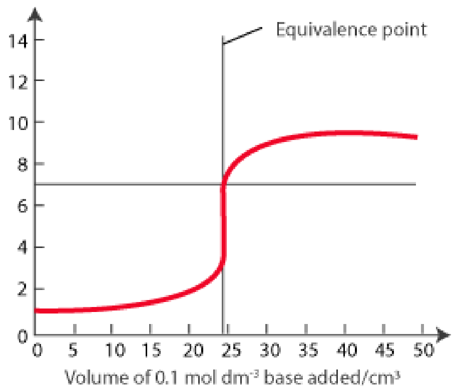
Name of indicator	Color at lower pH	pH range	End-point	Color at higher pH
phenolphthalein **(Ph.Ph)**	colorless	8.2 - 10.0	9.3	pink/violet
methyl orange **(M.O)**	red	3.2 - 4.4	3.7	yellow



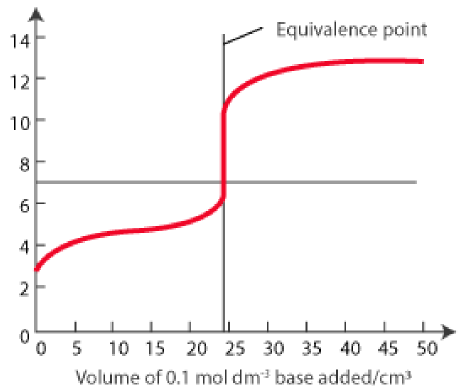
(Strong acid and strong base)



(Strong acid and weak base)



(Weak acid and weak base)



(Weak acid and strong base)

Determination The Molarity of NaOH by titration with HCl

Neutralization reactions involve the reaction of an acid and a base to produce a salt (ionic compound) and water.

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Purpose:
To determine the concentration of an unknown solution of HCl by titrating with
A standard solution of of NaOH (0.1 mol/L) .

This process involves a solution of known concentration (the titrant or standard solution) delivered from a burette into the unknown solution (the analyte) until the substance being consumed.

Materials:

50-mL Burette	250-mL conical flasks	funnel
250-mL beaker	10-mL pipette	**(** Ph.ph) indicator
HCL(aq) (0.1 mol/L)		NaOH (aq) (unknown concentration)

Procedure

1. Rinse the burette with H_2O. (Check valve and seal)
2. Rinse with <10 ml of NaOH .
3. Fill burette with NaOH at 0.00 ml.record as ini. Read.
4. Obtain exactly 10mL of HCl solution in the flask
5. Add 2-3 drops of indicator to the flask.
6. Add the NaOH from the burette, slowly, dropwise, to the HCl sample in the Flask . Keep one hand on the valve and the other hand constantly swirling the flask.
7. Stop when the color of indicator change .
8. Record the volume you reach as a final reading .
9. Repeate the titration two more times (step 4 - 8).

<u>IMPORTANT NOTES</u>

1- Intial reading for run 2 and run 3 is the final reading for the previous run .

2- You will likely NOT be successful the first time you try this, as you have no idea how many mL of HCl it will take to neutralize your unknown acid sample.

3- Do a rough titration, adding ~1 mL of NaOH at a time to find the approximate volume needed for neutralization , Once you have an idea of the volume at which the colour will change, then repeat the titration until you have 2 “good” results.

4- Acceptable titration practice requires two results that are within 0.1 ml of each other)

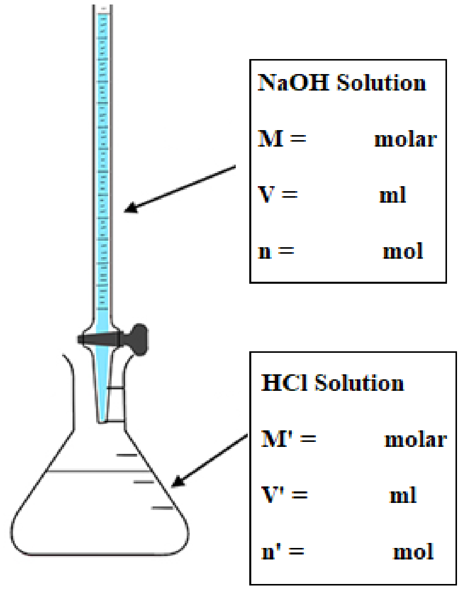
5- Drain the burette (discard unused solutions in waste container), and rinse the burette and other glassware with water and put it back on their boxes.

Results and calculations

<u>Titration using ph.ph</u>

	Init.Read	Final Read	Volume	Average
Run 1				
Run 2				
Run 3				

M : molarity
V : volume
n : no. of moles



1- Calculate the concentration of NaOH :

$$\frac{\mathbf{M\ .V}}{\mathbf{n}} = \frac{\mathbf{M'.\ V'}}{\mathbf{n'}}$$

2- Calculate the Strength of NaOH **Solution**

$$S = M_{NaOH} x\ M_w$$

Mw NaOH= 23 + 16 + 1 = 40 g/mol

Calculate the PH , POH of HCl and NaOH solutions

	HCl solution	**NaOH solution**
$[H^+]$		
$[OH^-]$		
PH		
POH		

PH = - log $[H^+]$
POH = - log $[OH^-]$
PH + POH = 14
$[H^+]$ x $[OH^-]$ = 1 x 10^{-14}

EXPERIMENT (5)
Determination The Molarity of HCl by titration with Na2CO3

Sodium carbonate reacts with hydrochloric acid according to the following equation:

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H2O$$

In other words, to neutralize all the carbonate, two equivalent of HCl should be used .

When one equivalent of HCl is added to the carbonate it is transformed into bicarbonates.
the pH of the solution changes form 11.5 (alkaline) to 8.3. which is happen at same zone of ph.ph.

$$Na_2CO_3 + HCl \longrightarrow NaHCO_3 + NaCl \qquad \text{ph.ph.}$$

When another equivalent of HCl is added to the solution of bicarbonate, complete neutralization takes place , The pH of solution changes from 8.3 to 3.8, which is near enough to zone of M.O.

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2 \qquad \text{M.O.}$$

In this titration when ph.ph is used the volume of acid used will be equivalent to half of the carbonate, but when M.O is used the volume of acid used will be equivalent to all carbonate. So ,when we comine the two reaction equation we will result with :

$$Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + CO_2 + H_2O$$

Name of indicator	Color at lower pH	pH range	End-point	Color at higher pH
phenolphthalein **(Ph.Ph)**	colorless	8.2 - 10.0	9.3	pink/violet
methyl orange **(M.O)**	red	3.2 - 4.4	3.7	yellow

Materials:

50-mL Burette
250-mL conical flasks
Funnel
10-mL pipette
250-mL beaker
Ph.ph indicator
M.O indicator
Na_2CO_3 (aq) (0.05 M)
HCL(aq) (unknown concentration)

Procedure:

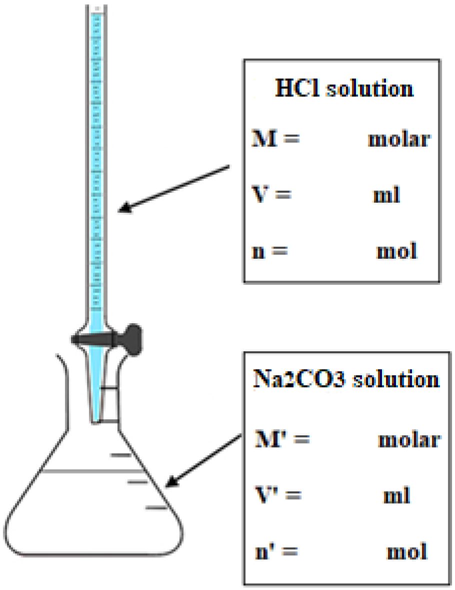
1. Rinse the burette with H2O. then rinse it with <10 ml of HCl. (Discard in waste container)
2. Fill burette with HCl at 0.00 ml.record as ini. Read.
4. Obtain exactly 10.0 mL of Na_2CO_3 solution in the flask.
5. Add 2-3 drops of ph.ph indicator to the flask.
6. Add the HCl from the burette, slowly, dropwise, to the Na_2CO_3 in the Flask .
   Keep one hand on the valve and the other hand constantly swirling the flask.
7. Stop when the indicator change its color , record the volume you reach as a final reading
8. Repeate the titration two more times (step 4 - 8).
9. Using M.O indicator , repeate titration (step 2-8) .

Results and calculations

M : molarity
V : volume
n : no. of moles

A :Titration using ph.ph

	Init.Read	Final Read	Volume	Average
Run 1				
Run 2				
Run 3				

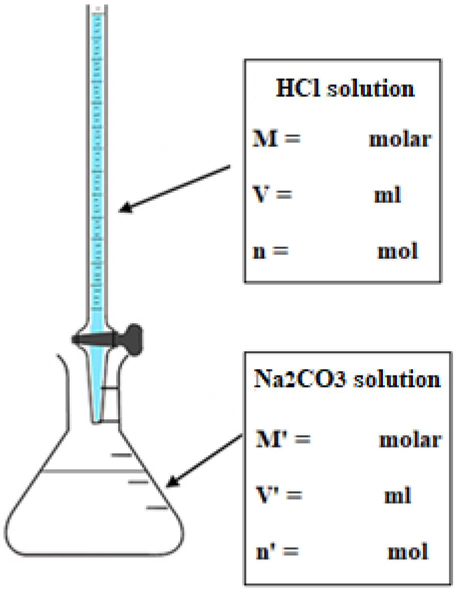


1- Calculate the concentration of HCl :

$$\frac{\mathbf{M\ .V}}{\mathbf{n}} = \frac{\mathbf{M'.\ V'}}{\mathbf{n'}}$$

B :Titration using M.O

	Init.Read	Final Read	Volume	Average
Run 1				
Run 2				
Run 3				



2- Calculate the concentration of HCl :

$$\frac{\mathbf{M\ .V}}{\mathbf{n}} = \frac{\mathbf{M'.\ V'}}{\mathbf{n'}}$$

3- Calculate the Average concentration of HCl solution :

$$\boldsymbol{M_{HCl}} = \frac{\boldsymbol{M_{ph.ph} + M_{M.O}}}{\mathbf{2}} =$$

4- Calculate the Strength of HCl Solution **:**

$$S = M_{HCl} x\ M_w$$

Mw HCl = 1 + 35.5 = 36.5 g/mol

EXPERIMENT (6):

REACTION STOICHIOMETRY: DETERMINATION OF THE LIMITING REACTANT AND YIELD PERCENTAGE

Objectives

The objectives of this experiment are the determination of:

1. The limiting reactant.
2. The percentage of the yield.

Theoretical information

$$\mathbf{Na_2CO_3 + 2HCl \rightarrow 2\ NaCl + CO_2 + H_2O}$$

The balancing coefficients indicate that there is a 1:2 mole ratio between Na_2CO_3 and HCl. This means that for every one mole of sodium carbonate reacts with two moles of HCl and produce two moles of NaCl.

Safety

- Be careful when handling the hydrochloric acid, it can cause chemical burns to the skin.
- If any acid spills on you, rinse immediately under running water for 15 minutes and report the accident to your instructor.
- Acid spills may also can be neutralized using sodium bicarbonate solution.
- Be sure to exercise appropriate caution when using the Bunsen burner and handling hot equipment.

Materials and equipment

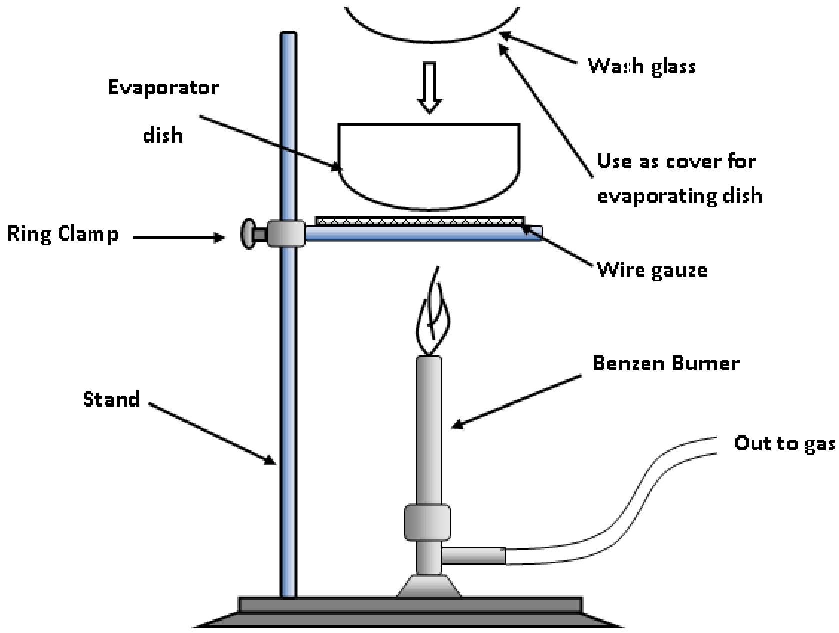
- Sodium carbonate (Na_2CO_3),
- Hydrochloric acid Solution HCl (1 M)
- Balance.
- Evaporating dish.
- Watch glass (to fit as a cover for the evaporating dish),
- Stand and ring clamp and wire gauze.
- 10 mL-pipette
- small beaker
- Bunsen burner or hot plate

Procedure

1. Measure and record the mass of your clean dry evaporating dish with watch glass. Record this mass as **$m_{INITIAL}$**.
2. Carefully weigh 0.3 – 0.4 g of Na_2CO_3 to the evaporating dish. Record the exact mass as $m_{(reactant)}$.
3. Using your dropper pipette, obtain exactly 10.0 mL of the 1.0 molar HCl (aq). Record the exact volume as **V_{HCl}**.
4. Add hydrochloric acid drop by drop to the sodium bicarbonate in the evaporating dish.
5. Carefully mix the reactants after every 4-5 drops of HCl.

(The reaction will be complete by noticing once no more bubbles are appeared)

6. Assemble the stand, ring clamp and wire gauze As shown in the figure below :



7. Carefully heat the solution in the covered evaporating dish with a Bunsen burner in order to remove the water generated in the reaction (as well as any excess HCl)

(**Flame should be adjusted to a appropriate temperature and wafted under the evaporating dish constantly)**

8. Continue heating until the contents are completely dry. Note that the watch glass cover should also be dry!
9. Allow the evaporating dish to cool to room temperature.
10. Measure the mass of the evaporating dish + watch glass + residue (NaCl). Record the exact mass as **Total product**.

Results and calculations

Tabulate your experimental results:

m_{Na2CO3}	V_{HCl}	$M_{\ HCl}$

(Molar masses (g.mol^{-1}) : H = 1.008, C = 12.01, O = 16, Na = 22.99, Cl = 35.45)

1. <u>The limiting reactant</u>

- To determine which of the reactants is the limiting reactant, you have to know the number of moles that used in the experiment ,which is ($n_{Na_2CO_3}$) and (n_{HCl}) , then divided by the coefficient of each reactants in equation.

$$n_{Na2CO3} = \frac{m_{Na2CO3}}{Mw_{Na2CO3}}$$

$$\boxed{\begin{aligned} &Mw_{Na2CO3} \\ &= (23x2) + 12 \\ &+ (16x3) \\ &= 106\ g/mol \end{aligned}}$$

$$n_{Na2CO3} =$$

$$\frac{n_{Na2CO3}}{1} =$$

$n_{HCl} = M_{\ HCl} \times V_{HCl}$

$n_{HCl} =$

$$\frac{n_{HCl}}{2} =$$

- The reactant which gives the lowest quotient is the limiting reactant.

  So the limiting reactant is ……….. .

2,The yield percentage

- Calculate the mass of NaCl produced (m_{NaCl}):

m_{INITIAL}	$m_{\text{TOTAL product}}$

$m_{\text{NaCl, actual}} = m_{\text{TOTAL(product)}} - m_{\text{INITIAL}}$

- The consumed amount of the limiting reactant will be used stoichiometrically to calculate the amount of NaCl that should be produced theoretically.

1 mole of $Na_2CO_3 \longrightarrow$ 2 mole of NaCl

……………. mole of $Na_2CO_3 \longrightarrow$ X mole of NaCl

X mole of NaCl ($n_{\text{NaCl, theorical}}$) =

Determined ($m_{\text{NaCl, theorical}}$) using the following equation:

$m_{\text{NaCl, theorical}} = n_{\text{NaCl, theorical}}$ x Mw_{NaCl}

$\boxed{Mw_{NaCl} = 23 + 35.5 = 58.5\ g/mol}$

$m_{\text{NaCl, theorical}} =$

- The yield percentage of NaCl can be calculated by the following equation:

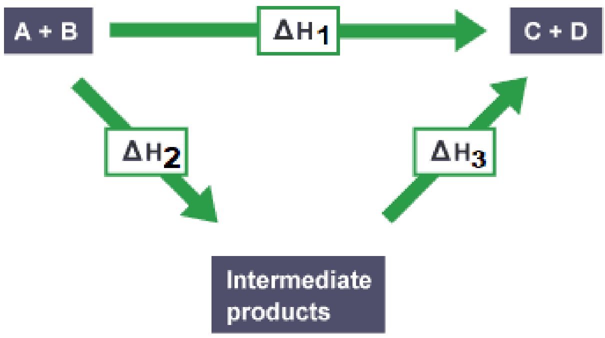
$$Nacl\ yeild\ \% = \frac{m_{Nacl,actual}}{m_{Nacl,theorical}}\ x\ 100$$

EXPERIMENT (7):

ENTHALPY OF REACTION: HESS'S LAW

Theoretical information

Hess's law states that at constant temperature and pressure, enthalpy (ΔH) of any thermal chemical reaction (absorbed or release) will remain constant irrespective what path taken to obtain the reaction (one step or multiple steps)



- In this experiment, you will measure and compare the change of Enthalpy involved in the following three reactions:

REACTION 1:
$NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ $\quad\Delta H_1$

REACTION 2:
$NaOH(s) + H_2O(l) \rightarrow NaOH(aq)$ $\quad\Delta H_2$

REACTION 3:
$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ $\quad\Delta H_3$

- It is clear that because reaction 1 is the sum of the reactions 2 and 3 :
$(\Delta H_1) = (\Delta H_2 + \Delta H_3)$

Objectives

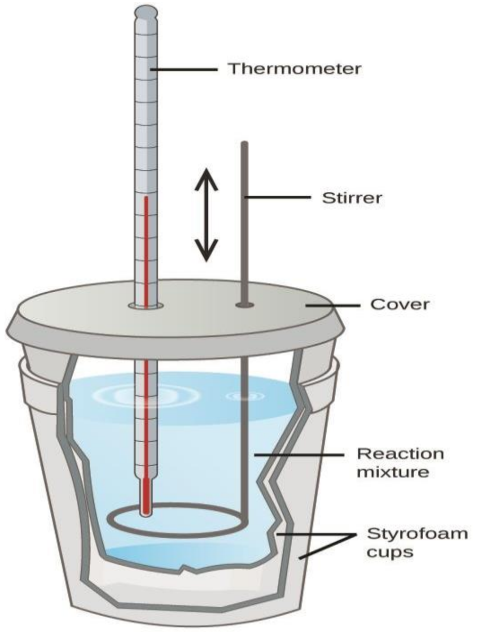
- Measuring the reactions' enthalpy and verifying Hess's Law.

Materials

Solid Sodium hydroxide (NaOH) .	Graduated cylinder
0.50 mol/L sodium hydroxide solution	Small beaker
0.50 mol/L Hydrochloric acid solution	Thermometer
0.25 mol/L Hydrochloric acid solution	Styro-foam cup calorimeter.

Procedure

- Weigh the glass tube and record its mass as $\mathbf{m_{glass}}$.
- Assemble your styro-foam calorimeter cup as in the figure .



REACTION 1:

1. Pour 25 mL of (0.25 M, HCl) into your styro-foam cup calorimeter. Stir carefully until a constant temperature is reached. Measure and record this temperature as $\mathbf{T_1}$.
2. Accurately weight about 0.25 g of solid NaOH ,Record the exact mass as $m_{1,\mathbf{NaOH}}$
3. Place the solid NaOH into styro-foam cup calorimeter which contin HCl solution, stir gently to insure that the solid is completely dissolved. Record the highest temperature reached as $\mathbf{T_2}$.
4. Discard the solution safely and rinse the cup thoroughly with water.

REACTION 2:

1. Pour 25 mL of cool distilled water into your styro-foam cup calorimeter. Stir carefully until a constant temperature is reached, record this temperature as $\mathbf{T_1}$.
2. Accurately weigh about 0.25 g of solid sodium hydroxide, Record the exact mass as $m_{2,NaOH}$.
3. Place the solid NaOH into styro-foam cup calorimeter which contin water.
4. Stir gently to insure that the solid is completely dissolved. Record the highest temperature reached as $\mathbf{T_2}$.
5. Discard the solution safely and rinse the cup thoroughly with water.

REACTION 3:

1. Pour 12.5 mL of (0.50 M , HCl) into your styro-foam cup calorimeter. Stir carefully until a constant temperature is reached, record it as $\mathbf{T_1}$.
2. Wash graduated cylinder and accurately take 12.5 mL of (0.50 M, NaOH)
3. Pour the NaOH solution into styro-foam cup calorimeter which contin HCl solution.
4. Stir the mixture gently and record the highest temperature reached as $\mathbf{T_2}$.
5. Discard the solution and rinse the cup thoroughly with water.

Results and calculations:

$m_{calorimeter}$	$m_{solution}$	$m_{1,NaOH}$	$m_{2,NaOH}$	ρ_{water} J/g °C	ρ_{glass} J/g °C
				4.18	0.836

- Tabulate your results as follows:

	Reaction 1	Reaction 2	Reaction 3
t_1			
t_2			
$\Delta T = t_2 - t_1$			
$q_{Solution}$ $q_{sol} = m_{sol} \times \rho_{water} \times \Delta T$			
$q_{calorimeter}$ $q_{cal} = m_{cal} \times \rho_{glass} \times \Delta T$			
$Q = q_{sol} + q_{cal}$			
n_{NaOH}			
$\Delta H = \frac{-Q}{n}$			

Proof the Hess's law.

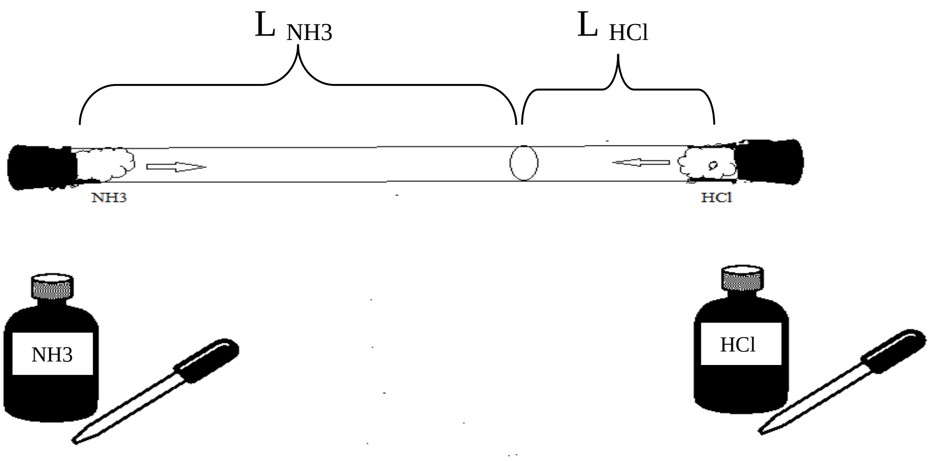
$(\Delta H_1) = (\Delta H_2 + \Delta H_3)$

EXPERIMENT (8):

GRAHAM'S LAW OF GAS DIFFUSION

Graham's law states that a gas will effuse at a rate inversely proportional to the square root of its molecular mass under same conditions of temperature and pressure.

$$\frac{L_1}{L_2} = \frac{\sqrt{M_{w2}}}{\sqrt{M_{w1}}} = \frac{\sqrt{d_2}}{\sqrt{d_1}}$$



Reaction Equation**:**

Ghraham Law for this reaction is:

$$\frac{L_{NH3}}{L_{HCl}} = \frac{\sqrt{Mw_{HCl}}}{\sqrt{Mw_{NH3}}} = \frac{\sqrt{d_{HCl}}}{\sqrt{d_{NH3}}}$$

Materials and equipment

Glass tube (40 cm x 1 cm)   Two stoppers

HCl and NH3 solutions.   Cotton

Procedure:

1- Put the glass tube in horizontal position as in diagram
2- Insert the cotton in the ends of glass tube.
3- At the same time, inject equal amount of each solution in the cotton (one in each side) and close them quickly by stoppers.
4- After few minutes, you will observe the formation of white smoke inside the glass tube, mark it with pen.
5- Measure the distance moved by each gas (from center of the cotton to the white smoke).

Results:

1- Distance moved by HCl gas (L_{HCl}) = cm.
2- Distance moved by NH3 gas (L_{NH3}) = cm.

Molar masses (g/mol): H =1 , N =14 , C =12 , Cl =35.5

Calculation :

1- The theoretical ratio between the molar masses of the two gases (Y) :

2- The measured ratio between the molar masses of the two gases (X) :

3- Error percentage :

$$Error\ \% = \left|\frac{Y - X}{Y}\right| X100$$

<u>Question;</u>

Unknown gas faster two times than methane (CH4), calculate its molar mass?

EXPERIMENT (9):

DETERMINATION THE MOLAR MASS OF VOLATILE LIQUID

Objectives

- Application of the gas law.
- Observation of evaporation and condensation.
- Measurements of certain physical properties of a gaseous substance.
- Determination of the molar mass of an unknown substance.

Theoretical information

- Volatile liquids have low boiling point, so they evaporate very easily at room temperature .
- Volatile substances are usually composed of nonpolar molecules, which easily vaporize to the gaseous state , then the common method to determine its molar mass is to use the ideal gas law:

$$PV = nRT$$

- If volume, pressure, and temperature of the gas are measured, the numbers of moles (n) can be calculated :

$$n = \frac{P\ X\ V}{R\ X\ T}$$

- Molar mass (M) can be evaluated using the following equation:
-

$$Mw = \frac{m}{n} \quad \text{,then} \quad Mw = \frac{m_{gas}.R.T}{P.V}$$

Materials and equipment

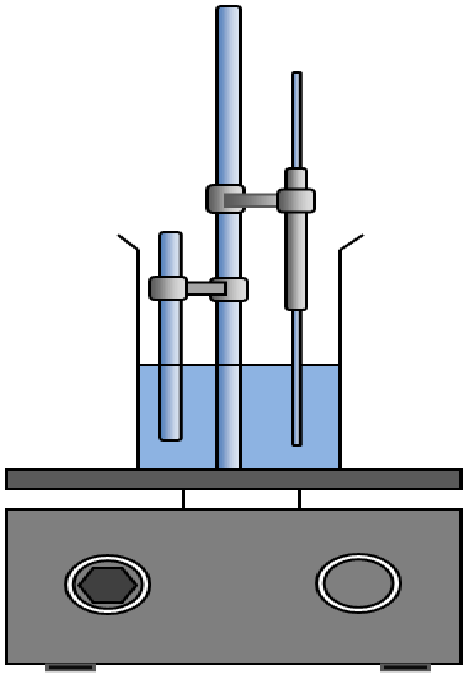
Unknown volatile Liquid	Holder
13 cm × 1 cm-test tube	Needle
Aluminum foil	Ice bath
A conventional barometer	Water bath

Procedure

1. Trim a piece of aluminum foil , that it just covers the top of a test tube.
2. Use a needle to make a small hole in the center of the foil.
3. Measure the mass of the test tube and the foil. Record it as **m_1.**
   (Notice: You can replace test tube with 25 ml conical flask)

Noting that, the best results are obtained when the test tube containing the sample is completely submerged in the water bath to just below the foil cap.

Consider using tall form beakers for your water baths



4. Prepare a boiling water bath around 95 ^{o} C.
5. Immerse the thermometer in the hot-water bath (see figure).
6. Fill the test tube with 3 ml of volatile liquid .
7. Submerge your test tube; shake it gently until all the liquid vaporized.
8. Watch the thermometer and record the temperature of the boiling water bath
   In Kelvin unit as **T**.
9. Using a test-tube holder, transfer the test tube quickly to the ice water bath.
10. Leave the test tube to cool for about one minute, then remove it and dry it completely.
11. Measure the mass of the test tube and the aluminum foil. Record it as **m_2**.
12. Rinse out the test tube and fill it to the top with tap water and cover it with aluminum foil .measure the mass and record it as **m_3**.
13. Record the room’s barometric pressure. Record it as **P.**

Results and calculations

Report your measurements as follows:

m_1	m_2	m_3	T	P

1- Calculate the mass of the unknown gas (***it equals the mass of the condensed vapor***), $\mathbf{m_{gas}}$, as follows:

$$m_{gas} = m_2 - m_1 =$$

2-Calculate the volume of the test tube (**it equals the volume of gas**) **V:**

d_{water} =1 g/ml

$$m_{water} = m_3 - m_1$$

$$V = \frac{m_{water}}{density} =$$

$$=$$

You can measure the volume of water by pouring it into suitable graduated cylinder and record the volume of the test tube (V) .

3-Calculate the molar mass of the unknown substance, Mw, as follows :

$$Mw = \frac{m_{gas}.R.T}{P.V}$$

EXPERIMENT (10):

DETERMINATION THE MOLAR MASS OF AN UNKNOWN SUBSTANCE BY FREEZING POINT DEPRESSION

Objectives

Using the Raoul's relationship between freezing point depression and the molality of the solution to determine:

1- The molar mass of an unknown compound .

2- The $K_{f,solvent}$

Theoretical information

- The freezing point depression is one of four important *colligative* properties of ideal solutions (freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure increase).

- The depression in the freezing point, ΔT_f, is:

$$\Delta T_f = T_{solvent} - T_{solution}$$

- Where:

$T_{f,\ solvent}$ = the freezing temperature of solvent (water).

$T_{f,\ solution}$ = the freezing temperature of solution

Note that ΔT_f is positive because the freezing temperature of the solution is lower than the pure solvent, and the word "*depression*" implies the negative sign.

- The molality is given by the following equation:

$$Molality = \frac{n_{solute(mole)}}{m_{solvent(kg)}} = \frac{m_{solute}(g) x 1000}{M_{W\ (solute)}\left(\frac{g}{mol}\right) x m_{solvent}(g)}$$

- The magnitude of ΔT_f is proportional to the molality of the dissolved solute:

$$\Delta T_f = K_f \times \text{molality}$$

$$\Delta T_f = K_f x \frac{m_{solute} x 1000}{M_{W\ (solute)}\left(\frac{g}{mol}\right) x m_{solvent}(g)}$$

Where K_f is the proportionality constant, it called the *molal freezing point constant*, *which* depends on ***the nature of solvent***.

Materials and equipment

250 ml Beaker
A wide test tube
A suitable thermometer
Glucose
Unknown substance
Balance
Ice
Deionized water
Salt

Procedure

1. Add exactly 10 g (10 mL) of deionized water in the test tube. Record it as $\mathbf{m_{Solvent}}$.
2. Weigh a 2 g of Solute. Record your precise mass as **Solute**.
3. Add this mass to the test tube. **(Make sure that all of the solute is completely dissolved)**
4. Prepare ice bath by filling a 250 ml-beaker with layer of ice and salt consecutively.
5. Insert test tube in the ice bath and stir the solution with metal stirrer until you see the ice formation occur.
6. Take the test tube out of the ice, measure the freezing point of the solution. Record this temperature as $\mathbf{T_{solution}}$.

Results and calculation

- Present the results of your experiment in the following tables:

	$\mathbf{m_{Solvent}}$	$\mathbf{m_{Solute}}$	$\mathbf{T_{solvent}}$	$\mathbf{T_{solution}}$	$\Delta T_f = \mathbf{T_{solvent}} - \mathbf{T_{solution}}$
Unknown					
Glucose					

A: Determination of the molar mass of the Unknown substance (K_f = 1.86 °C.molal^{-1})

$$M_{W\ (unknown)} = \frac{K_f\ x\ m_{solute} x 1000}{\Delta T_f\ x\ m_{solvent}}$$

B: Determination the $K_{f,\ water}$ in Glucose *solutions ($M_{W\ (glucose)}$ = 180.16 g/mol)*

$$K_f = \frac{\Delta T_f x M_{W\ (glucose)}\ x m_{solvent}}{m_{solute} x 1000}$$