COLLEGE OF SCIENCES
CHEMISTRY DEPARTMENT

CHEM 101
LABORATORY OF GENERAL
CHEMISTRY (1)

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1444 - 2022

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.

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

We should keep in mind that "*NO report is presented NO experiment is done*". 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.

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

All 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.

These experiments are written here exactly as they were tested and carried out.

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

Our 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

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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:
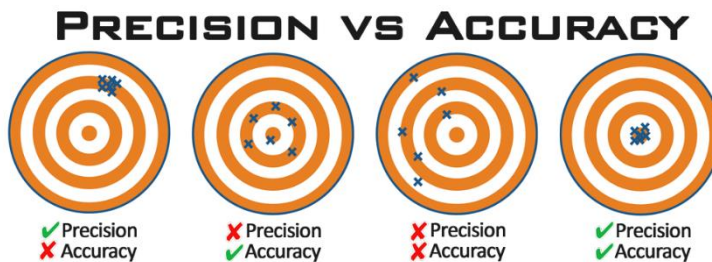
Student number:

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 :

$$\Delta x = |x - x_0| \qquad\qquad (\textbf{x}) : \textbf{measured value}$$

$$(\textbf{x}_0) : \textbf{ actual value}$$

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

$$Error(\%) = \frac{\Delta x}{x_0} \ X \ 100$$

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

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

.

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:**

$$\Delta m = |\ \overline{m} - m_{actual}\ |$$
consider $_{dwater} = 1$ g/ml then,
m_{actual} = amount of water taken in experiment.

- **Determining of m%**

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

- **Determining the accuracy of measurement:**

$$\text{Accuracy in m value} = \overline{m} \pm \Delta m$$

Equipment and reagents
Balance.   25 or 50 mL graduated cylinder
Beaker   Burette

Procedure
1. Weight an empty beaker and record the mass as $m_{empty\ beaker}$
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 $m_{beaker\ and\ water}$
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 $\bar{m}$, the experimental error (Δm), the percent error (m%), and the Accuracy of the mass values using Graduated cylinder and Burette.

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

A- Graduated cylinder

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

$\bar{m} =$

$\Delta m =$

$m\% =$

Accuracy =

B- Burette

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

$\bar{m} =$

$\Delta m =$

$m\% =$

Accuracy =

- State which is more accurate the graduated cylinder or the graduated burette.

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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:

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

- 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

First method

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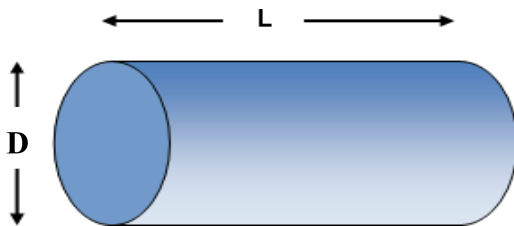
Procedure
1. Using a proper an accurate ruler, measure the length (L) and the diameter of the specimen (massive cylinder) (D).
2. Weigh the mass (m) of your specimen (massive cylinder).

Results and calculations
1. Report your measurements as follows:

L	D	m

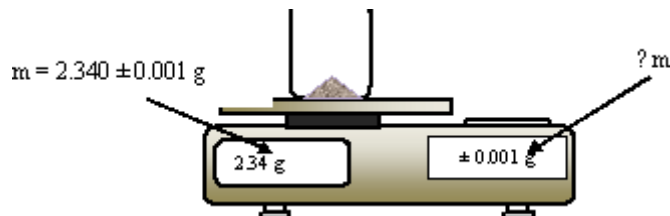2. Calculate the volume of your specimen (massive cylinder).

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

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

$m = 2.340 \pm 0.001$ g

234 g   ± 0.001 g   ? m

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

6   7   cm

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

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Second method

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
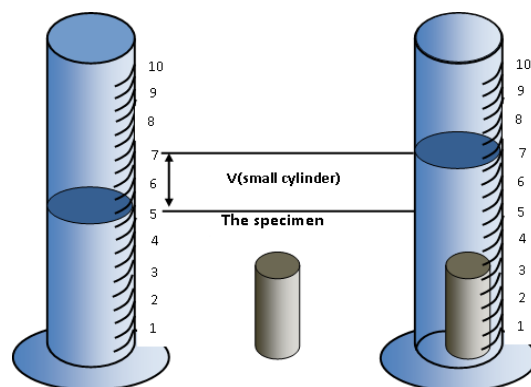
1. Report your measurements as follows:

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 , ΔV₁ and ΔV₂ as follows:
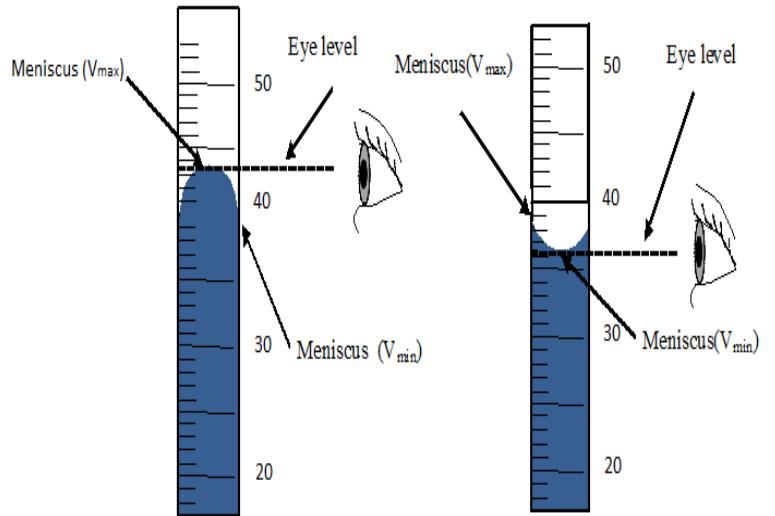
Δm is evaluated exactly as described in the first method.

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

$\Delta V_{graduated\ cylinder}$ 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:

ΔV₁	ΔV₂	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: $$\text{Molarity} = \frac{n \text{ solute}}{V_{\text{solution (L)}}}$$

Where n is number of moles of solute and $S_{\text{olution (L)}}$ 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:

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

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.(**record it as m$_{NaCl}$**)
2) <u>Transfer this mass to a 1</u>00-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:

$m_{Nacl} = 3\ g$

$V = 1000\ cm^3 = 1.000\ dm^3$

(a) (b) (c) (d) (e) (f)
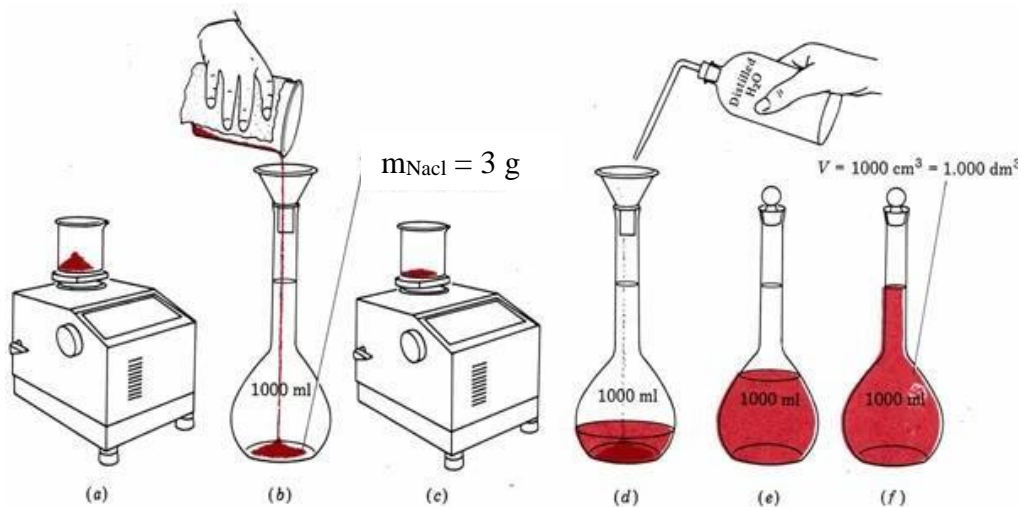
Results and calculations

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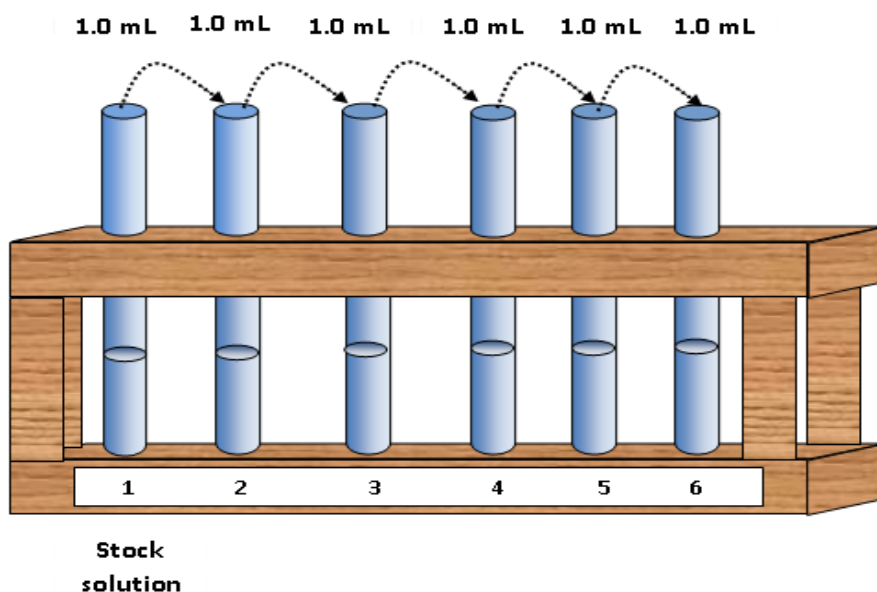
$$M_{conc} \times V_{initial} = M_{dil} \times V_{total}$$

EXPERIMENT (4):

PREPARING SOLUTIONS BY SERIAL DILUTION

Theoretical information

- 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 in expement 8)**

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.
2) Tabulate the results of your experiment and calculations:

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

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

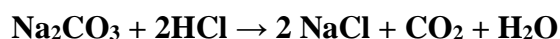$M_{dil \text{ (Test tube 3)}} =$

EXPERIMENT (5):

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

$$Na_2CO_3 + 2HCl \rightarrow 2\ NaCl + CO_2 + H_2O$$

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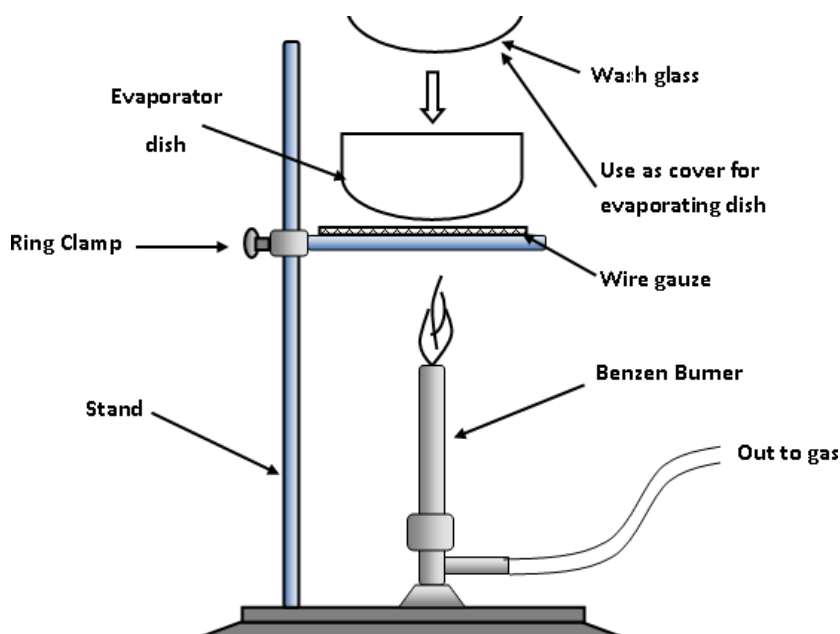
7. Carefully heat the solution in the covered evaporating dish with a Bunsen burner in order to remove the water generated in the reaction (as well as any excess HCl)

**(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. **The limiting reactant**

- 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.

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

$$Mw_{Na2CO3} = (23x2) + 12 + (16x3) = 106 \ g/mol$$

$$n_{Na2CO3} =$$

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

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

$$n_{HCl} =$$

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

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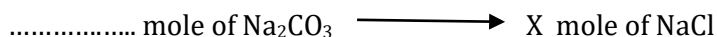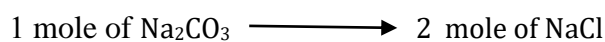
    So the limiting reactant is ……….. .

2.The yield percentage

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

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

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

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

$$1 \text{ mole of } Na_2CO_3 \longrightarrow 2 \text{ mole of NaCl}$$

$$................. \text{ mole of } Na_2CO_3 \longrightarrow X \text{ mole of NaCl}$$

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

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

$$m_{NaCl, theorical} = n_{NaCl, theorical} \times Mw_{NaCl}$$

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

$m_{NaCl, theorical} =$

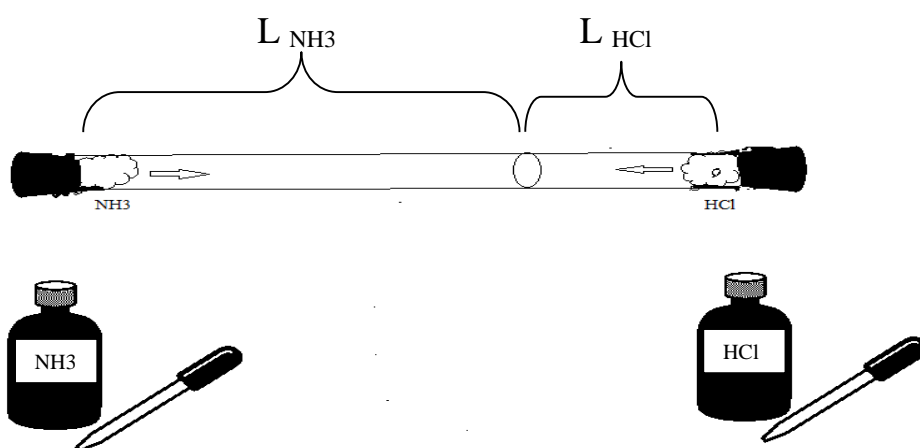- The yield percentage of NaCl can be calculated by the following equation:

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

EXPERIMENT (6):

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 :

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

Question;

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

EXPERIMENT (7):

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 :

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

- 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
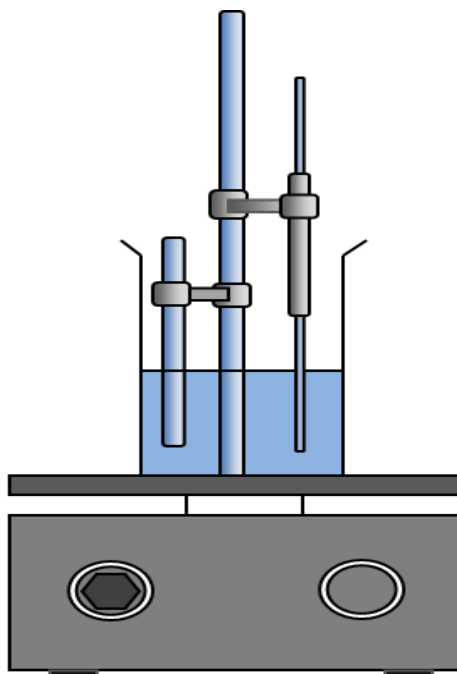
4. Prepare a boiling water bath around 95 ° C.
5. Immerse the thermometer in the hot-water bath (see figure).
6. Fill the test tube with 3 ml of volatile liquid .
7. Submerge your test tube; shake it gently until all the liquid vaporized.
8. Watch the thermometer and record the temperature of the boiling 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:

m₁	m₂	m₃	T	P

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

$$m_{gas} = m2 - m1 =$$

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

d$_{water}$ = 1 g/ml

$$m_{water} = m3 - m1 \qquad\qquad V = \frac{m_{water}}{density} =$$

$$=$$

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 (8):

DETERMINATION OF THE TRANSFERED THERMAL ENERGY: ENDOTHERMIC AND EXOTHEMIC

Objectives

Measurement of the amount of Heat (Q) change because of mixing two substances.

Theoretical information

- In this experiment, you will measure the amount of heat involved two mixtures (solid-liquid and liquid-liquid).using the following relationship:

$$q = m \times \rho \times (T_{final} - T_{initial})$$

Where:

m is the mass, in grams, of the solid or the liquid substance.

ρ is the specific heat (J /g.°C).

$T_{initial}$ is the initial temperature of substance .

T_{final} is the initial temperature of substance after heat transferred .

- When two substances having different temperatures were mixed together in an adiabatic container, the substance that has the lower temperature gains the energy and its sign will be positive ($q_{Endothermic}$) and that having the higher temperature loses the same energy and its sign will be negative ($Q_{Exothermic}$).

$$q_{Endo} = m_1 \times \rho_1 \times (T_{final} - T_{initial})$$

$$q_{Exo} = m_2 \times \rho_2 \times (T_{final} - T_{initial})$$

$$q_{Endo} = - q_{Exo}$$

Materials and equipment

Water at room temp.	Glass or metal rod
250 –mL beaker	Balance
Thermometers	Graduated cylinder
Styro-foam cup	Water bath

Procedure:

Part one *(NOTE:* **Considering the density of water is 1.0 g/cm³,)**

1. Weigh out 25 g of tap water and put it in a Styro-foam cup. Record its mass as m_{water}.
2. Measure the water temperature, record it as $T_{initial\ water}$.
3. Put the glass or metal rod in the hot water bath for 15 minutes.
4. Measure the hot water bath temperature, record it as $T_{initial,rod}$.
5. Take the rod outside and insert it immediately in the styro-foam cup that contain water.
6. Shake and immediately record the new temperature as T_{final}.

Results and calculations

1. Tabulate your results as follows:

m_{water}	$T_{initial,water}$	m_{rod}	$T_{initial,rod}$	T_{final}

2. Calculate the heat quantity changes, in the units of Joule, as follows:

- First: Heat gained by water:

$$q_{water} = m_{water} \times \rho_{water} \times (T_{final} - T_{initial,water})$$

$\rho_{substance}$ J/g °C
ρ_{water} = 4.184
ρ_{glass} = 0.836
ρ_{iron} = 0.452

- Second: Heat lost by the rod:

$$q_{rod} = m_{rod} \times \rho_{rod} \times (T_{final} - T_{initial,rod})$$

Part two

1. In 250 mL beaker, add 25 g of tap water , measure its temperature and record it as $T_{initial,water}$.
2. Read the temperature of hot water bath and record it as $T_{initial,Hot\ water}$
3. From hot water bath take 25 g of hot water and add it to the beaker in the 1st step. Record your temperature as T_{final}.

Results and calculations

1. Tabulate your results as follows:

m_{water}	$m_{hot\ water}$	$T_{initial\ water}$	$T_{initial,hot\ water}$	T_{final}

$$\rho_{water} = \rho_{hot\ water} = 4.184 \text{ J/g °C}$$

2. Calculate the heat quantity changes as follows:

- First: Heat gained by water in the units of Joule:

$$q_{water} = m_{water} \times \rho_{water} \times (T_{final} - T_{initial,water})$$

- Second: Heat lost by hot water in the units of Joule:

$$q_{hot\ water} = m_{hot\ water} \times \rho_{hot\ water} \times (T_{final} - T_{initial,\ hot\ water})$$

EXPERIMENT (9):

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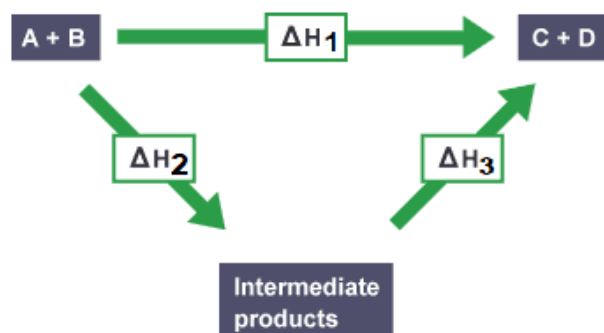
- In this experiment, you will measure and compare the change of Enthalpy involved in the following three reactions:

  REACTION 1:
  $NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ ΔH_1

  REACTION 2:
  $NaOH(s) + H_2O(l) \rightarrow NaOH(aq)$ ΔH_2

  REACTION 3:
  $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ ΔH_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 **m$_{glass}$**.
- Assemble your styro-foam calorimeter cup as in the figure .

Thermometer

Stirrer

Cover

Reaction mixture
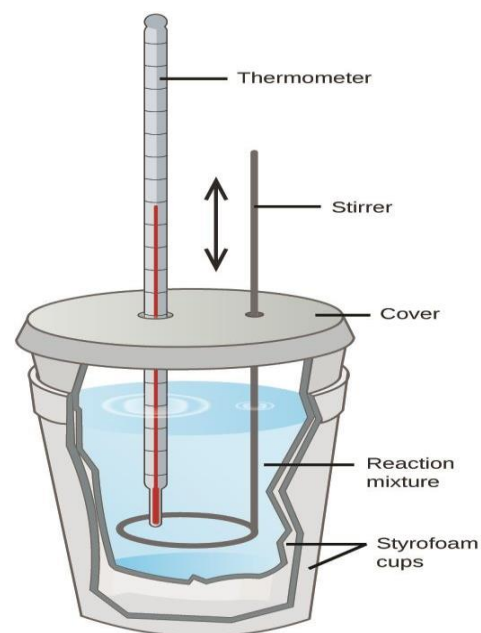
Styrofoam cups

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 **T$_1$**.
2. Accurately weight about 0.25 g of solid NaOH ,Record the exact mass as m$_{1,NaOH}$
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 **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 **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 **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 **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 **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}			
$\Delta H = \dfrac{-Q}{n}$			

Proof the Hess's law.

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

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:

$$\Delta T_f = K_f \times molality$$

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

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

Procedure

1. Add exactly 10 g (10 mL) of deionized water in the test tube. Record it as $m_{Solvent}$.
2. Weigh a 2 g of Solute. Record your precise mass as m_{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 $T_{solution}$.

Results and calculation

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

	$m_{Solvent}$	m_{Solute}	$T_{solvent}$	$T_{solution}$	$\Delta T_f = T_{solvent} - T_{solution}$
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)

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