

GENERAL CHEMISTRY

(CHEM-201)

Student Name	
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Written By

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Instruction & Safety

A chemistry lab can be dangerous place. If you work with care, however you ought to survive and not eliminate your neighbors. Therefore: -

1) **ALWAYS** wear safety glasses in the lab. You may only be sitting at your bench writing in your manual, but your neighbor's distillation may explode.

2) **NEVER**, but never, allow ANY organic compound on your skin or breathe in the dust from a solid or spray from a liquid.

Some compounds are labelled as particularly dangerous. Carcinogenic ones cause cancer, some apparently harmless ones, like benzene, are very toxic indeed.

ALL ARE LETHAL IN SUFFICIENT QUANTITY.

Rubber gloves are available to protect your hands.

3) **NEVER** boil organic compounds in the open lab or allow gases to escape, use the fume cupboard.

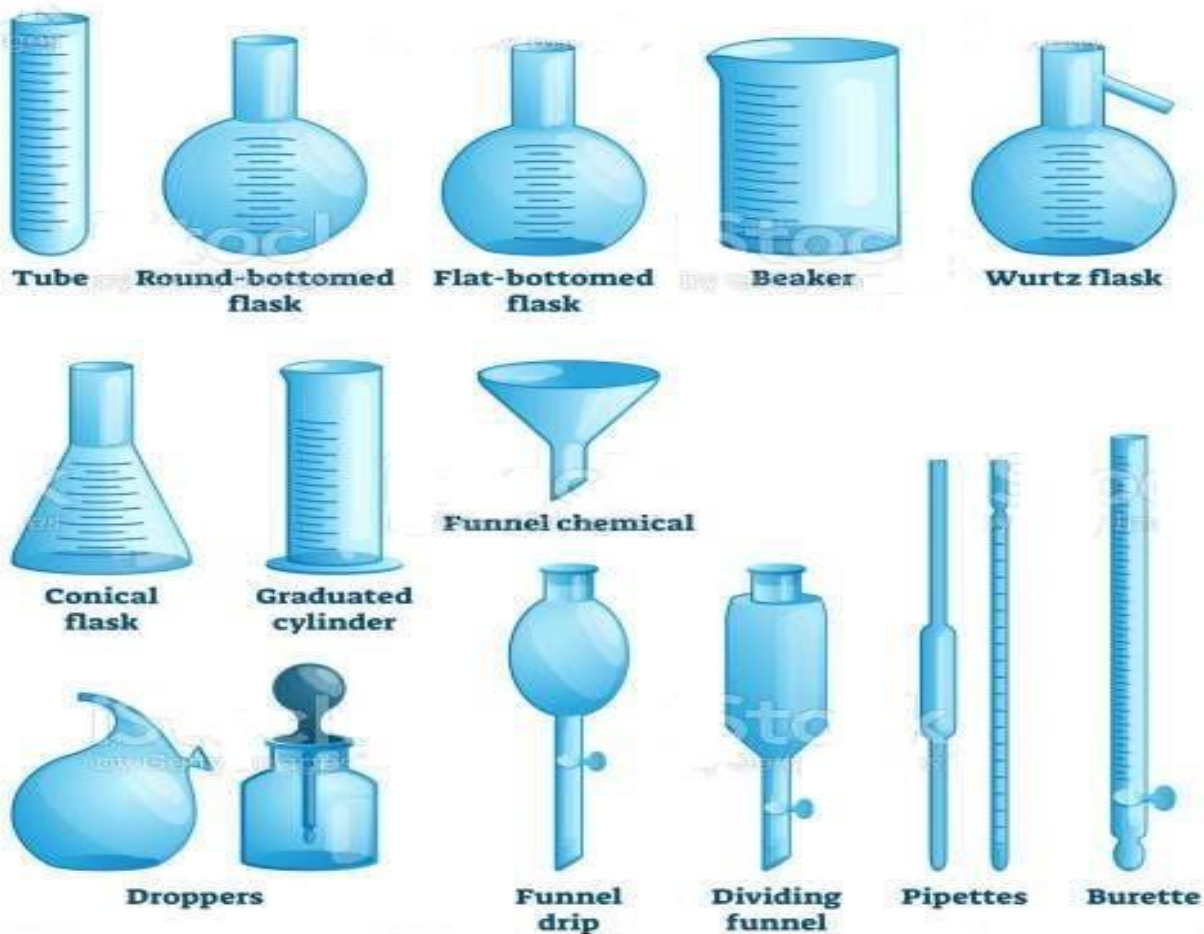
4) **BEWARE of FIRE**. Know where the fire extinguishers are. Look around you before lighting your burner. In particular, make sure no one is using ether or petrol nearby.

5) **HANDLE** glass-ware **CAREFULLY**. Use quick fit apparatus, and watch for chipped or cracked apparatus: return it to the preparation room.

6) **REPORT** all accidents to your demonstrator. You may think you are "all right" but it is not worth taking chances.

Keeping your bench clean is most important, always clean up at all times. You will have your grade reduced if you keep it in an untidy mess, the same applies to your locker.

LABORATORY GLASSWARE



Distribution of Grades

Distribution	Marks
Attendance and Reports	10
Final Exam 1 (Practical + theoretical)	10 (7+3)
Final Exam 2 (Practical + theoretical)	10 (7+3)
Total	30

Experiment 1: Determination of Liquid Density**Objective:**

To measure the density of a liquid sample (water).

Theory:

Density is defined as the mass per unit volume of a substance. Its SI unit is g/mL. The relationship is expressed as:

$$d = m / v$$

where d = density (g/mL), m = mass (g), v = volume (mL).

Apparatus and Materials:

Pipette, burette, beaker, balance, conical funnel.

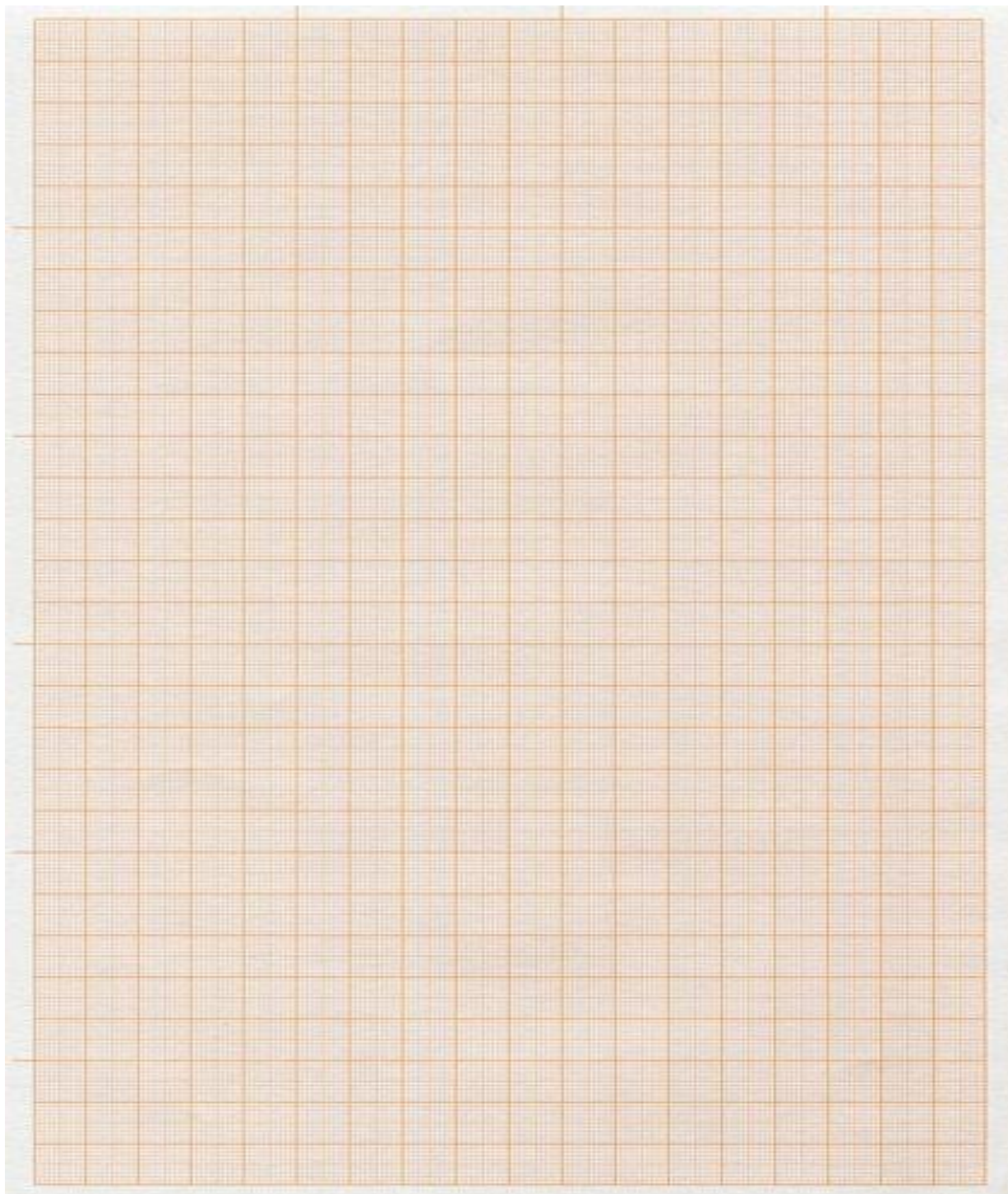
Procedure:

1. Weigh the empty beaker (m₁).
2. Using a pipette or burette, add a known volume of liquid (V) and record the total mass (m₂).
3. Repeat by adding successive known volumes and recording the corresponding masses.
4. Calculate the mass of the liquid (m = m₂ - m₁).
5. Plot mass versus volume to determine the slope, which represents the density.

Results:

16	14	13	10	8	5	3	2	Water Volume v (ml)
								Beaker Mass m ₁ (g)
								Beaker + Water Mass m ₂ (g)
								Water Mass m (g)
								Water Density d (g/ml)





Experiment 2: Introduction to Titration Experiments and Preparation of Standard Sodium Carbonate Solution

Objective:

To prepare a standard solution of sodium carbonate and perform acid-base titration.

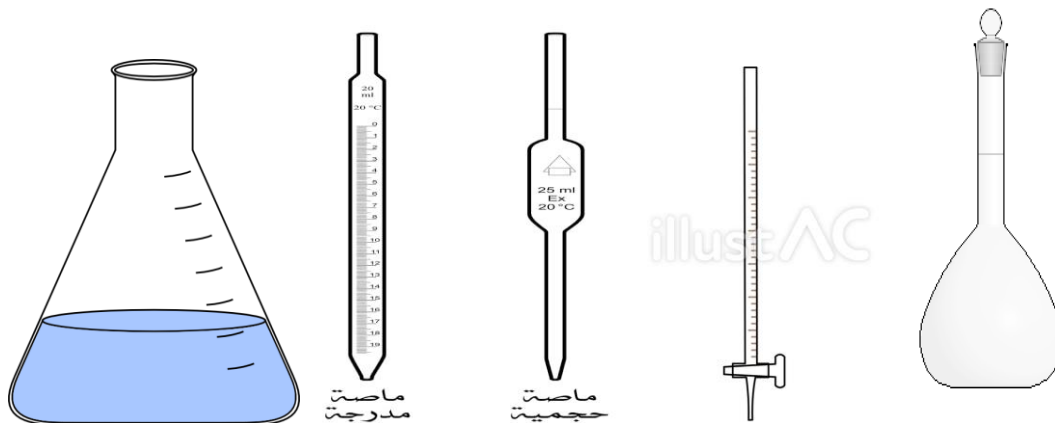
Theory:

Titration is a technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. **The equivalence point** is indicated by a sudden color change of the indicator.

A standard solution is one whose concentration is precisely known. **Indicators** are weak organic acids or bases that change color depending on the pH of the solution.

Apparatus and Materials:

Pipette, burette, conical flask, sodium carbonate, hydrochloric acid, indicators (phenolphthalein, methyl orange).



Key Equations:

$$M \times V / n = M' \times V' / n'$$

Where:

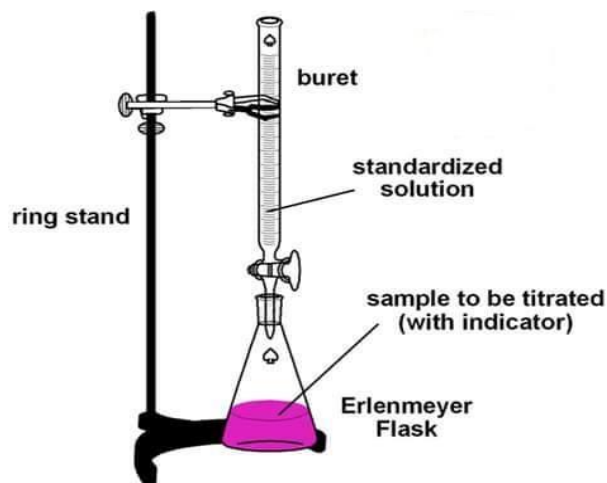
M = molarity of acid

V = volume of acid

M' = molarity of base

V' = volume of base

n, n' = stoichiometric coefficients



To convert molarity to solution strength in g/L, use the formula:

$$\text{Strength (g/L)} = \text{Molarity (mol/L)} \times \text{Molar Mass of Solute (g/mol)}.$$

Concentration of A Strong Acid = $[H^+]$

Concentration of A Strong Base = $[OH^-]$

$[H^+]$	pH	pOH
$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$	$pH + pOH = 14$	$pOH = -\log [OH^-]$

Indicators:

Indicator	The color in acidic solution	The color in basic solution	Indicator Range
Phenolphthalein (Ph.Ph)	Colorless	Pink	8 – 10
Methyl Orange (M.O)	Orange	Yellow	3 – 4

Preparation of Standard Sodium Carbonate Solution

Q: What is the weight in grams required to prepare a standard solution of sodium carbonate (Na_2CO_3) with a molarity of 0.05 M and a volume of 100 ml?

Na = 23 C = 12 O = 16

Solution

Weight in grams = molarity x molecular weight x volume in liters

$$= 0.05 \text{ mol/L} \times 106 \text{ g/mol} \times 0.1 \text{ L}$$

$$= 0.53 \text{ g}$$

❖ Procedure:

- 1- Using a balance, weigh 0.53 grams of sodium carbonate.
- 2- Place the weight in a 100 ml standard flask.
- 3- Add a little distilled water and shake until the sodium carbonate dissolves.
- 4- Continue adding distilled water to standard flask up to the mark.

Experiment 3: Determination of the Appropriate Indicator Using pH-Meter**Objective:**

To determine the suitable organic indicator for titrations using a pH-meter.

Theory:

The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration:

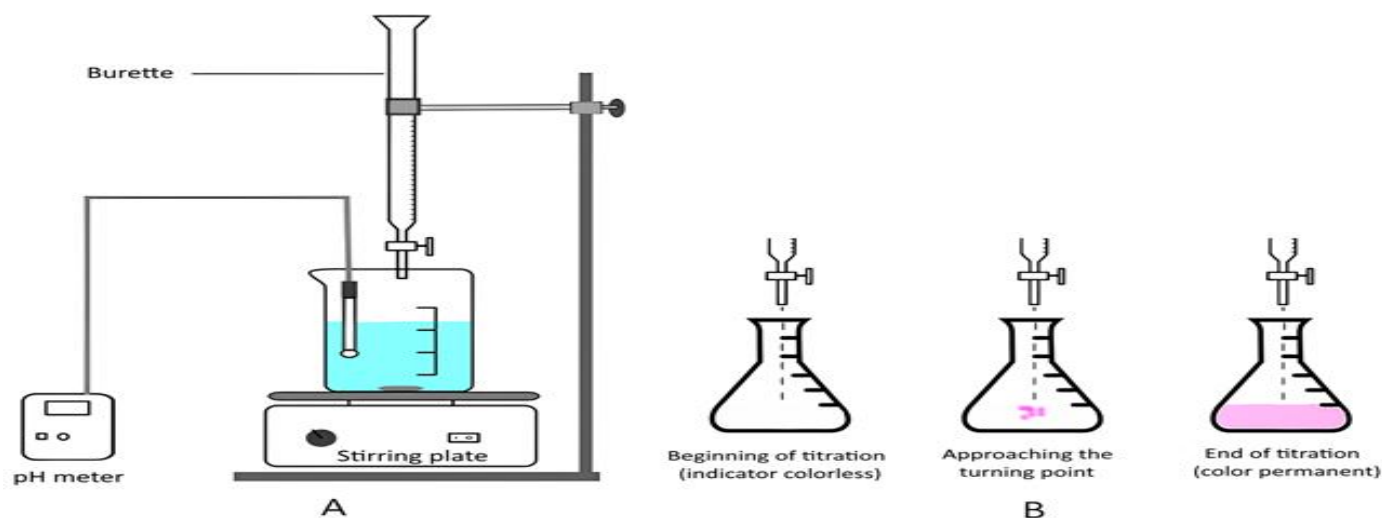
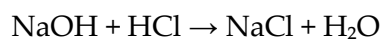
$$\text{pH} = -\log[\text{H}^+]$$

For neutral solutions, $\text{pH} = 7$; for acidic solutions, $\text{pH} < 7$; for basic solutions, $\text{pH} > 7$.

At the equivalence point of a titration, the pH changes rapidly, and the appropriate indicator is chosen based on this pH range.

Apparatus and Materials:

beaker, pipette, burette, pH meter, HCl sol, NaOH sol.

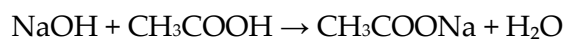
**First: Determine the organic indicator in the titration of a strong acid with a strong base: -****Procedure:**

1. Transfer **25 ml** of **HCl** solution to a 250 ml conical flask.
2. Using a burette, add the volume of sodium hydroxide **NaOH** according to the table.
3. Record the pH after each small addition of the base.
4. Plot the titration curve (pH vs. volume of base added).
5. From the titration curves, determine the appropriate indicator range.

Results:

30	26	25	24.8	24.5	24	22.5	20	15	10	5	0	NaOH (ml)
11.6	11	9.7	5.6	3.3	2.8	2.3	1.9	1.6	1.4	1.3	1.2	PH
												Student Results

Second: Determine the organic indicator in the titration of a weak acid with a strong base: -

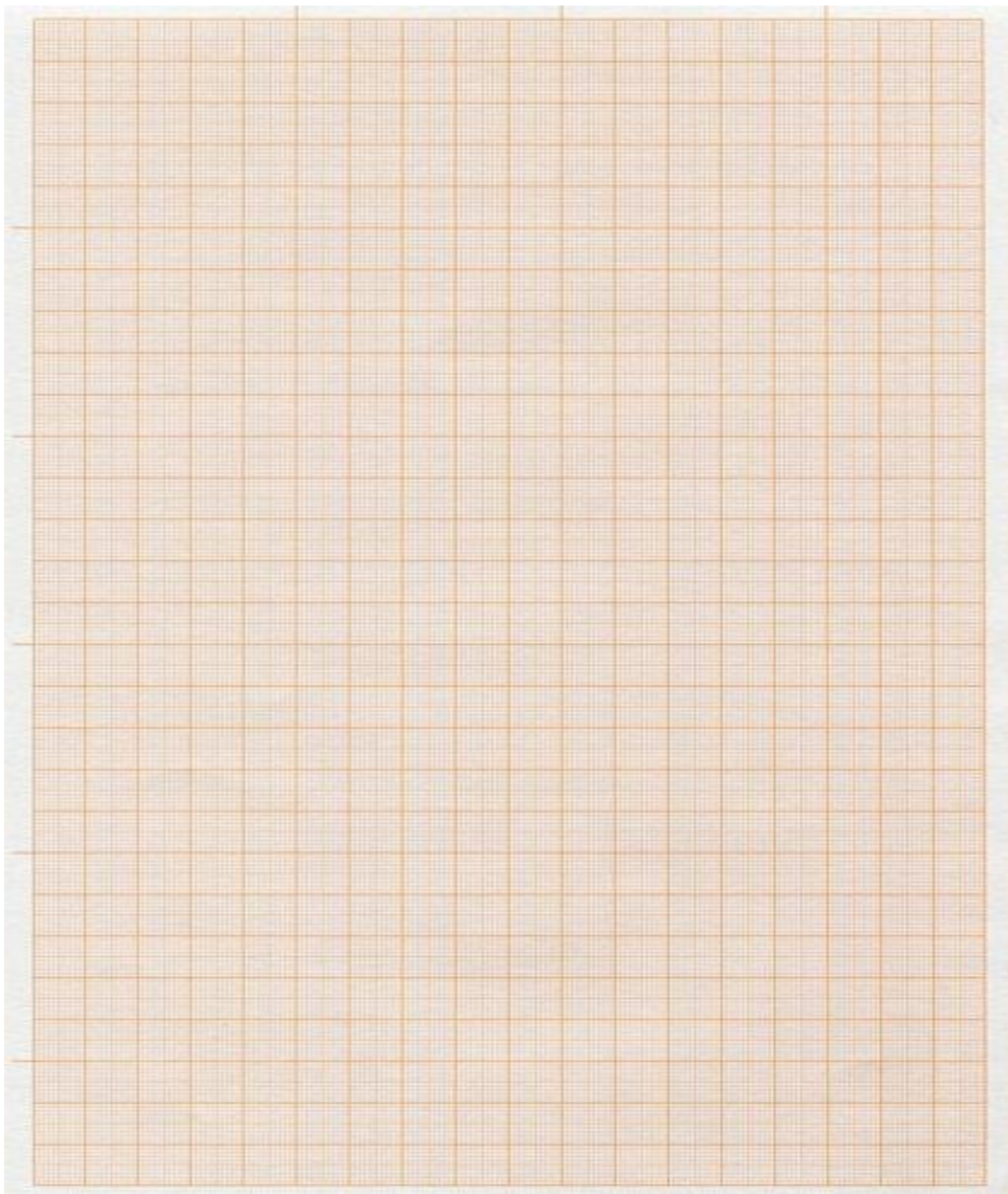
**Procedure:**

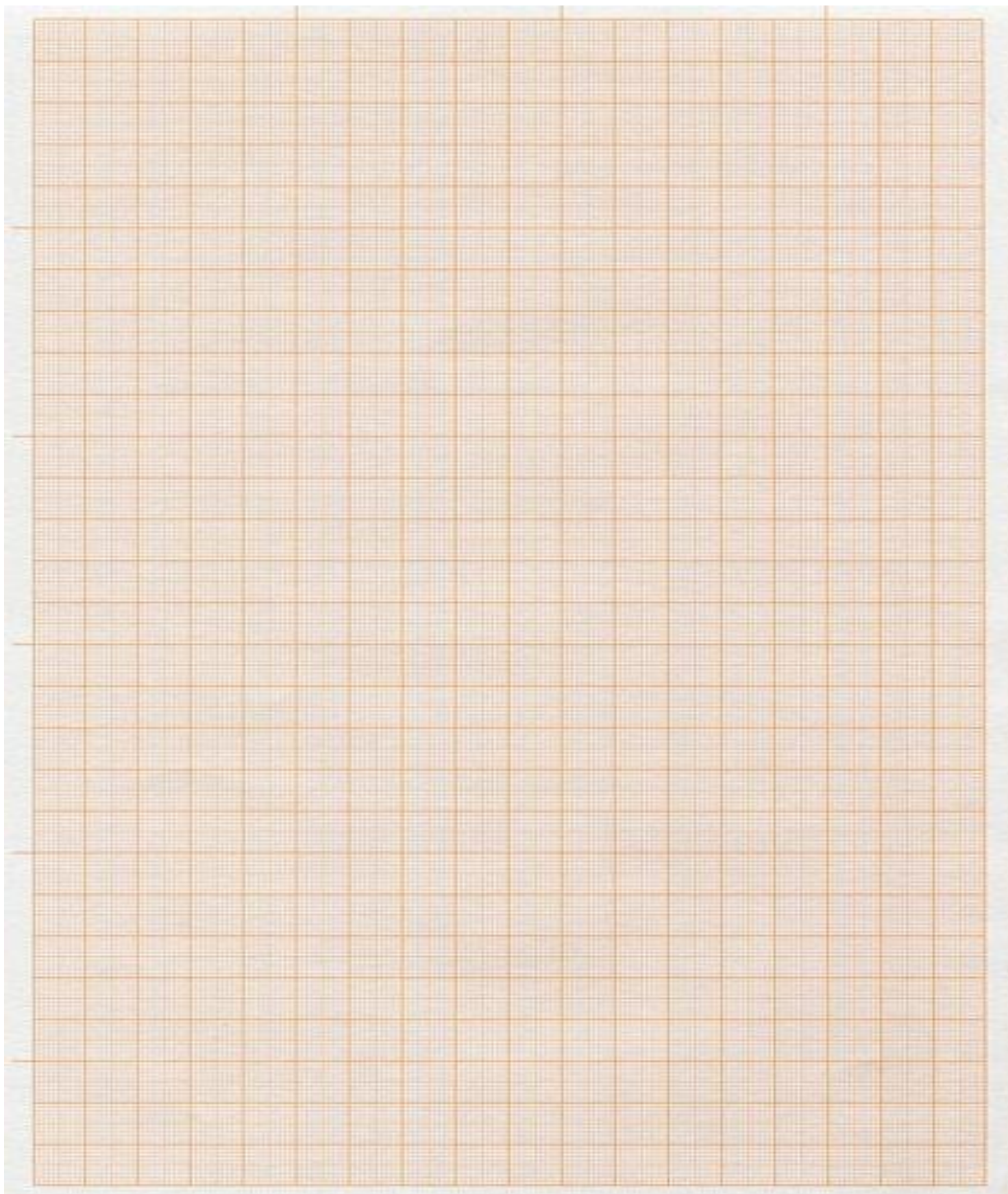
1. Transfer **25 ml** of **CH₃COOH** solution to a 250 ml conical flask.
2. Using a burette, add the volume of sodium hydroxide **NaOH** according to the table.
3. Record the pH after each small addition of the base.
4. Plot the titration curve (pH vs. volume of base added).
5. From the titration curves, determine the appropriate indicator range.

Results:

18.5	18	17	16	14	12	10	8	6	4	2	0	NaOH (ml)
10.2	9.5	6.1	5.7	5.3	5.1	4.9	4.7	4.5	4.3	4.0	3.5	PH
												Student Results

The Indicator Range for (ph.ph) is (8-10) and for (m.o) is (3-4).



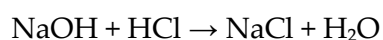


Experiment 4: Titration of NaOH with HCl

Objective:

Find the molarity and strength of a sodium hydroxide (NaOH) solution by titrating it with a solution of known concentration of hydrogen chloride (HCl).

Reaction equation:



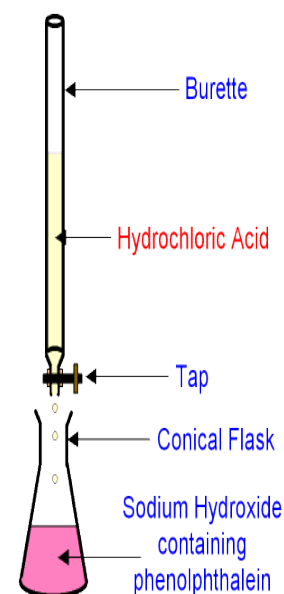
Apparatus and Materials:

Pipette, burette, conical flask, sodium hydroxide, hydrochloric acid 0.1 M indicators (phenolphthalein, methyl orange).

A: Titration of NaOH with HCl by phenolphthalein (ph.ph):

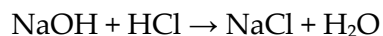
Procedure:

1. Pipette **10 mL** of **NaOH** solution into a conical flask.
2. Add 2–3 drops of indicator phenolphthalein (**Ph.Ph**).
3. Titrate against **HCl 0.1M** until the endpoint is reached (pink → colorless).
4. Repeat the titration at least three times to obtain consistent results.
5. Record the average volume of NaOH used and calculate the **molarity** and **strength** of a sodium hydroxide (NaOH) solution.



Results:

No.	Initial Reading	Final Reading	Volume	Volume Average
1	0			
2				
3				

Calculations:

From chemical equation: $n = \dots\dots\dots$ and $n' = \dots\dots\dots$

- Molarity of HCl (M)
- Volume of HCl (Average) (V)
- Volume of NaOH (M)
- Molarity of NaOH (V)

$$M \times V / n = M' \times V' / n'$$

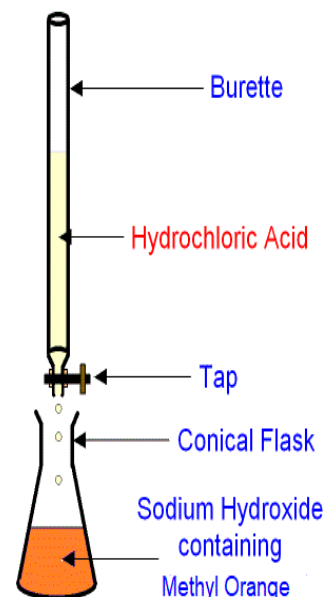
Strength NaOH = molarity NaOH $\times$ molecular weight NaOH

Concentration of NaOH = = $[\text{OH}^-]$

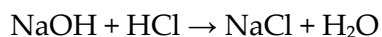
$[\text{H}^+]$	pH	pOH
$[\text{H}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$	$\text{pH} + \text{pOH} = 14$	$\text{pOH} = -\log [\text{OH}^-]$

B: Titration of NaOH with HCl by Methyl Orange (M.O):**Procedure:**

1. Pipette **10 mL** of **NaOH** solution into a conical flask.
2. Add 2–3 drops of indicator Methyl Orange (**M.O**).
3. Titrate against **HCl 0.1M** until the endpoint is reached (Orange → Yellow).
4. Repeat the titration at least three times to obtain consistent results.
5. Record the average volume of NaOH used and calculate the **molarity** and **strength** of a sodium hydroxide (NaOH) solution.

Results:

No.	Initial Reading	Final Reading	Volume	Volume Average
1	0			
2				
3				

Calculations:

From chemical equation: $n = \dots\dots\dots$ and $n' = \dots\dots\dots$

- Molarity of HCl (M)
- Volume of HCl (Average) (V)
- Volume of NaOH (M-)
- Molarity of NaOH (V-)

$$M \times V / n = M' \times V' / n'$$

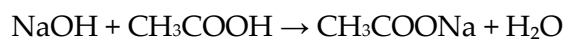
$$\text{Strength NaOH} = \text{molarity NaOH} \times \text{molecular weight NaOH}$$

Concentration of NaOH = = [OH⁻]

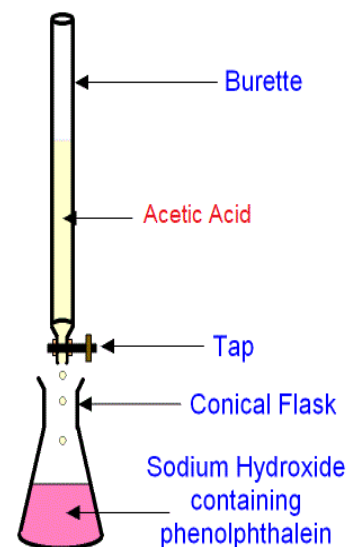
[H ⁺]	p ^H	p ^{OH}
$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$	$p^H + p^{OH} = 14$	$p^{OH} = - \log [OH^-]$

Experiment 5: Titration of NaOH with Acetic Acid (CH₃COOH)**Objective:**

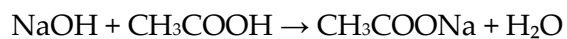
To determine the molarity and strength of acetic acid by titration with NaOH.

Reaction equation:**Procedure:**

1. Pipette **10 mL** of **NaOH 0.1 M** solution into a conical flask.
2. Add 2–3 drops of indicator phenolphthalein (**Ph.Ph**).
3. Titrate against **CH₃COOH** until the endpoint is reached (pink → colorless).
4. Repeat the titration at least three times to obtain consistent results.
5. Record the average volume of **CH₃COOH** used and calculate the **molarity** and **strength** of acetic acid (**CH₃COOH**) solution.

**Results:**

No.	Initial Reading	Final Reading	Volume	Volume Average
1	0			
2				
3				

Calculations:

From chemical equation: **n** = and **n'** =

- Molarity of NaOH (M)
- Volume of NaOH (V)
- Volume of CH₃COOH (Average) (M')
- Molarity of CH₃COOH (V')

$$\mathbf{M \times V / n = M' \times V' / n'}$$

Strength CH₃COOH = molarity CH₃COOH × molecular weight CH₃COOH

Concentration of Acetic Acid (CH₃COOH) = = [H⁺]

[OH ⁻]	p ^H	p ^{OH}
$[\text{H}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$	$\text{p}^{\text{H}} + \text{p}^{\text{OH}} = 14$	$\text{p}^{\text{OH}} = -\log [\text{OH}^-]$

Experiment 6: Titration of HCl with Sodium Carbonate (Na_2CO_3)**Objective:**

To determine the molarity and strength of HCl solution using standard sodium carbonate solution.

Reaction equation:

Step 1: $\text{HCl} + \text{Na}_2\text{CO}_3 \rightarrow \text{NaHCO}_3 + \text{NaCl}$ (Phenolphthalein endpoint) $n=1, n'=1$

Step 2: $\text{HCl} + \text{NaHCO}_3 \rightarrow \text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$ (Methyl orange endpoint)

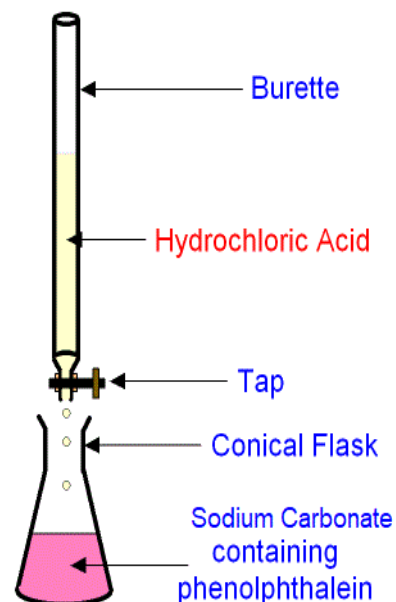
Overall: $2\text{HCl} + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$ $n=1, n'=2$

Apparatus and Materials:

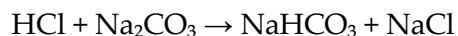
Pipette, burette, conical flask, sodium carbonate 0.05M, hydrochloric acid, indicators (phenolphthalein, methyl orange).

Titration of Na_2CO_3 with HCl by phenolphthalein (ph.ph):**Procedure:**

1. Pipette 10 mL of Na_2CO_3 (0.05M) solution into a conical flask.
2. Add 2–3 drops of indicator phenolphthalein (Ph.Ph).
3. Titrate against HCl until the endpoint is reached (pink $\rightarrow$ colorless).
4. Repeat the titration at least three times to obtain consistent results.
5. Record the average volume of HCl used and calculate the molarity and strength of a hydrochloric acid (HCl) solution.

**Results:**

No.	Initial Reading	Final Reading	Volume	Volume Average
1	0			
2				
3				

Calculations:

From chemical equation: $n = \dots\dots\dots$ and $n' = \dots\dots\dots$

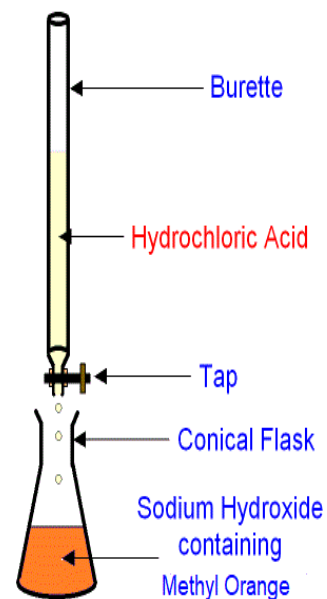
- Molarity of HCl (M)
- Volume of HCl (Average) (V)
- Volume of Na_2CO_3 (M')
- Molarity of Na_2CO_3 (V')

$$M \times V / n = M' \times V' / n'$$

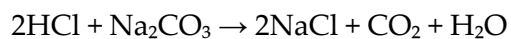
Strength HCl = molarity HCl $\times$ molecular weight HCl

Titration of Na_2CO_3 with HCl by Methyl Orange (M.O):**Procedure:**

- 1- Pipette **10 mL** of Na_2CO_3 solution into a conical flask.
- 2- Add 2–3 drops of indicator Methyl Orange (**M.O**).
- 3- Titrate against **HCl** until the endpoint is reached (orange $\rightarrow$ yellow).
4. Repeat the titration at least three times to obtain consistent results.
5. Record the average volume of **HCl** used and calculate the **molarity** and **strength** of a hydrochloric acid (**HCl**) solution.

**Results:**

No.	Initial Reading	Final Reading	Volume	Volume Average
1	0			
2				
3				

Calculations:

From chemical equation: **n** = and **n'** =

- Molarity of HCl (M)
- Volume of HCl (Average) (V)
- Volume of Na₂CO₃ (M')
- Molarity of Na₂CO₃ (V')

$$\mathbf{M \times V / n = M' \times V' / n'}$$

Strength HCl = molarity HCl × molecular weight HCl

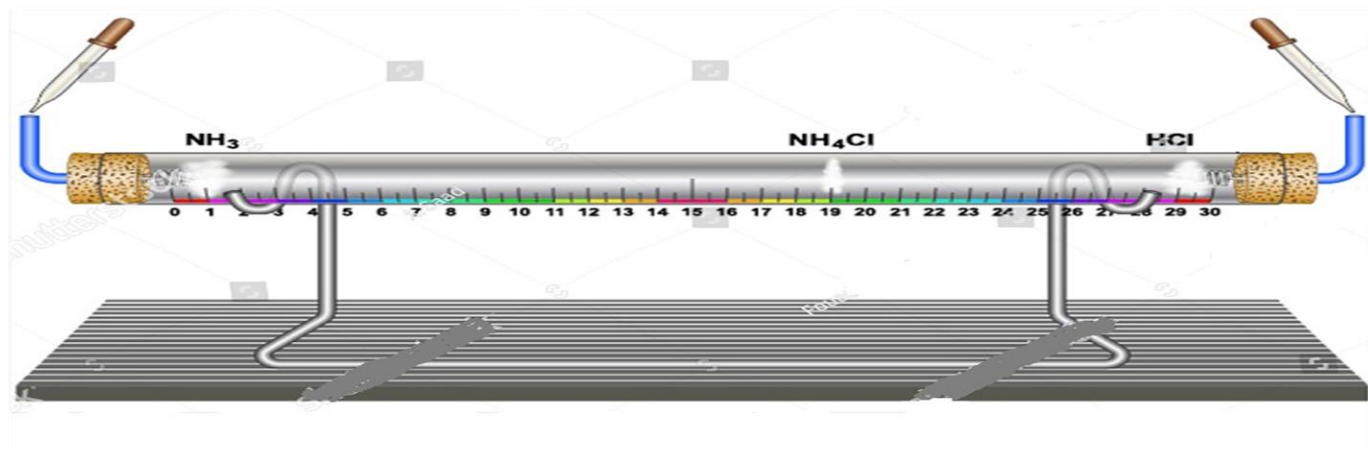
Concentration of Hydrochloric Acid (HCl) = = [H⁺]

[OH ⁻]	p ^H	p ^{OH}
$[\text{H}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$	$\text{p}^{\text{H}} + \text{p}^{\text{OH}} = 14$	$\text{p}^{\text{OH}} = -\log [\text{OH}^-]$

Experiment 7: Diffusion of Gases (Graham's Law)

Objective:

To verify Graham's law by comparing the rates of diffusion of NH_3 and HCl gases.

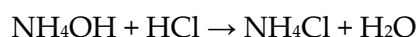


Theory:

According to **Graham's law**, the rate of diffusion of a gas is inversely proportional to the square root of its molar mass:

$$\text{Rate}_1 / \text{Rate}_2 = \sqrt{(M_2 / M_1)}$$

Chemical equation:



Apparatus and Materials:

A glass tube, 40 cm long and approximately 1 cm in diameter, two stoppers, two rapidly evaporating chemical compounds (HCl , NH_3) dissolved in a solvent.

* We have two gases, the first is NH_3 and the second is HCl , so:-

$$\frac{L_{\text{NH}_3}}{L_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}} = \sqrt{\frac{1 + 35.5}{14 + 3}} = \sqrt{\frac{36.5}{17}} = 1.44$$

So we say that the ratio of the diffusion rate of NH_3 gas is greater than the diffusion rate of HCl gas by **1.46**, and this is the **theoretical** ratio.

Procedure:

1. Take a long glass tube with cotton swabs soaked in NH_3 and HCl placed at opposite ends.
2. Allow the gases to diffuse toward each other.
3. Record the distance traveled by each gas until a white ring of NH_4Cl is formed.
4. Calculate the experimental diffusion ratio and compare with the theoretical ratio.

Results and Calculations:

The distance traveled by HCl gas (L_1)	
The distance traveled by NH_3 gas (L_2)	
Experimental ratio	
Theoretical ratio	1.44

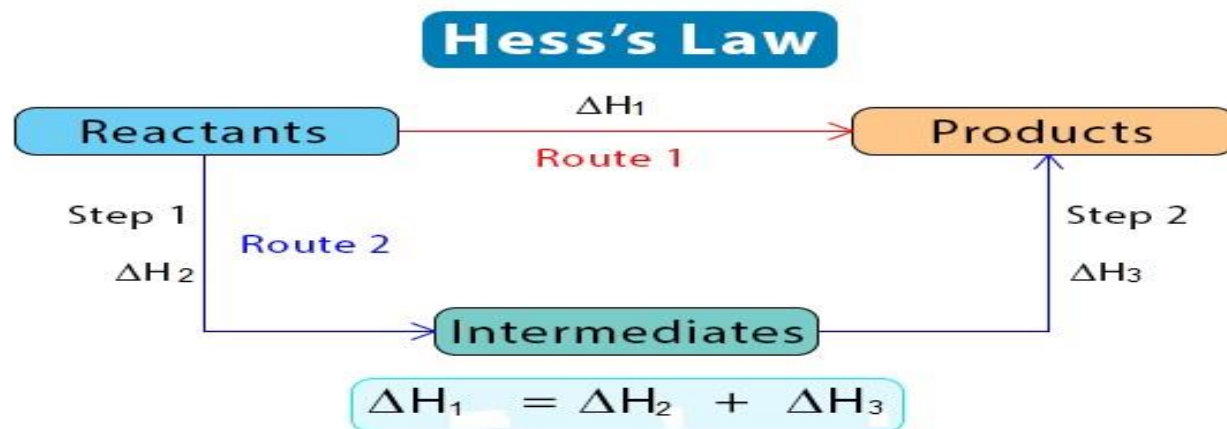
Experiment 8: Application of Hess's Law

Objective:

To verify Hess's law by determining the enthalpy change of reactions involving NaOH and HCl.

Theory:

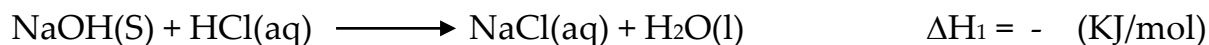
Hess's law states that the enthalpy change of a chemical reaction is independent of the pathway taken, provided the initial and final conditions are the same.



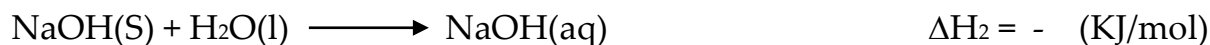
Apparatus and Materials:

Calorimeter, thermometer, (HCl) solution, (NaOH) solution, (NaOH) solid, water.

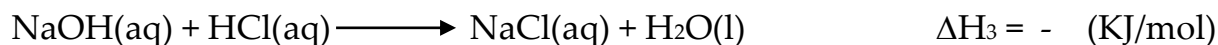
The heat equation for the first (direct) reaction:



The heat equation for the second reaction:



The heat equation for the second reaction:



Weigh the dry calorimeter tube (m_1)

Procedure of first reaction:

- 1- Place 50 ml of 0.25 M HCl solution in a calorimeter tube.
- 2- Measure the temperature of the solution inside the calorimeter tube (T_1).
- 3- Quickly weigh four granules of NaOH (S) in the range of 0.44 – 0.55 gm (**m**).
- 4- Quickly transfer the previous weight into the calorimeter tube.
- 5- Gently stir using a thermometer and record the constant temperature reached by the solution inside the calorimeter tube (T_2).
- 6- Clean and dry the calorimeter tube.

Procedure of second reaction:

1. Place 50 ml of distilled water in the calorimeter tube.
2. Measure the temperature of the solution inside the calorimeter tube (T_1).
3. Quickly weigh four granules of NaOH (S) in the range of [0.44 – 0.55 gm] (**m**).
4. Quickly transfer the previous weight into the calorimeter tube.
5. Stir gently using a thermometer and record the constant temperature reached by the solution inside the calorimeter tube (T_2).
6. Clean and dry the calorimeter tube.

Procedure of third reaction:

- 1- Place 25 ml of 0.5 M HCl solution in the calorimeter tube.
- 2- Measure the temperature of the solution inside the calorimeter tube, set to (T_1).
- 3- Place 25 ml of 0.5 M NaOH solution in the calorimeter tube.
- 4- Stir gently using a thermometer and record the constant temperature reached by the solution inside the calorimeter tube, set to (T_2).
- 5- Clean and dry the calorimeter tube.

Results and Calculations:

Weight of the calorimeter $m_1 = \dots\dots\dots$

Weight of sodium hydroxide in reaction (1) $m = \dots\dots\dots$

Weight of sodium hydroxide in reaction (2) $m = \dots\dots\dots$

Weight of the solution in all three reactions = $m_2 = 50$ grams

(Specific heat of glass) = 0.836

(Specific heat of water) = 4.18

Reaction 1	Reaction 2	Reaction 3
$T_1 =$ $T_2 =$ $\Delta T =$	$T_1 =$ $T_2 =$ $\Delta T =$	$T_1 =$ $T_2 =$ $\Delta T =$
$q_1 = m_1 \times 0.836 \times \Delta T$ $q_2 = m_2 \times 4.18 \times \Delta T$	$q_1 = m_1 \times 0.836 \times \Delta T$ $q_2 = m_2 \times 4.18 \times \Delta T$	$q_1 = m_1 \times 0.836 \times \Delta T$ $q_2 = m_2 \times 4.18 \times \Delta T$
$Q = q_1 + q_2$	$Q = q_1 + q_2$	$Q = q_1 + q_2$
Weight NaOH (m) = $n(\text{NaOH}) = \text{NaOH (m)} / M_{\text{wt}} \text{NaOH}$	Weight NaOH (m) = $n(\text{NaOH}) = \text{NaOH (m)} / M_{\text{wt}} \text{NaOH}$	Volume NaOH (V) = $n(\text{NaOH}) = M(\text{NaOH}) \times V(\text{NaOH})$
$\Delta H = Q / n(\text{NaOH})$	$\Delta H = Q / n(\text{NaOH})$	$\Delta H = Q / n(\text{NaOH})$

Experiment 9: Determination of Reaction Order

Objective:

To study the effect of concentration on the rate of reaction between sodium thiosulfate and hydrochloric acid.

Reaction equation:



Apparatus and Materials:

Beaker (150 ml); Beaker (50 ml); Two burettes; Stopwatch; White paper; HCl solution; $\text{Na}_2\text{S}_2\text{O}_3$ solution.

First: Change the concentration of $\text{Na}_2\text{S}_2\text{O}_3$ while the concentration of HCl remains constant:

Procedure:

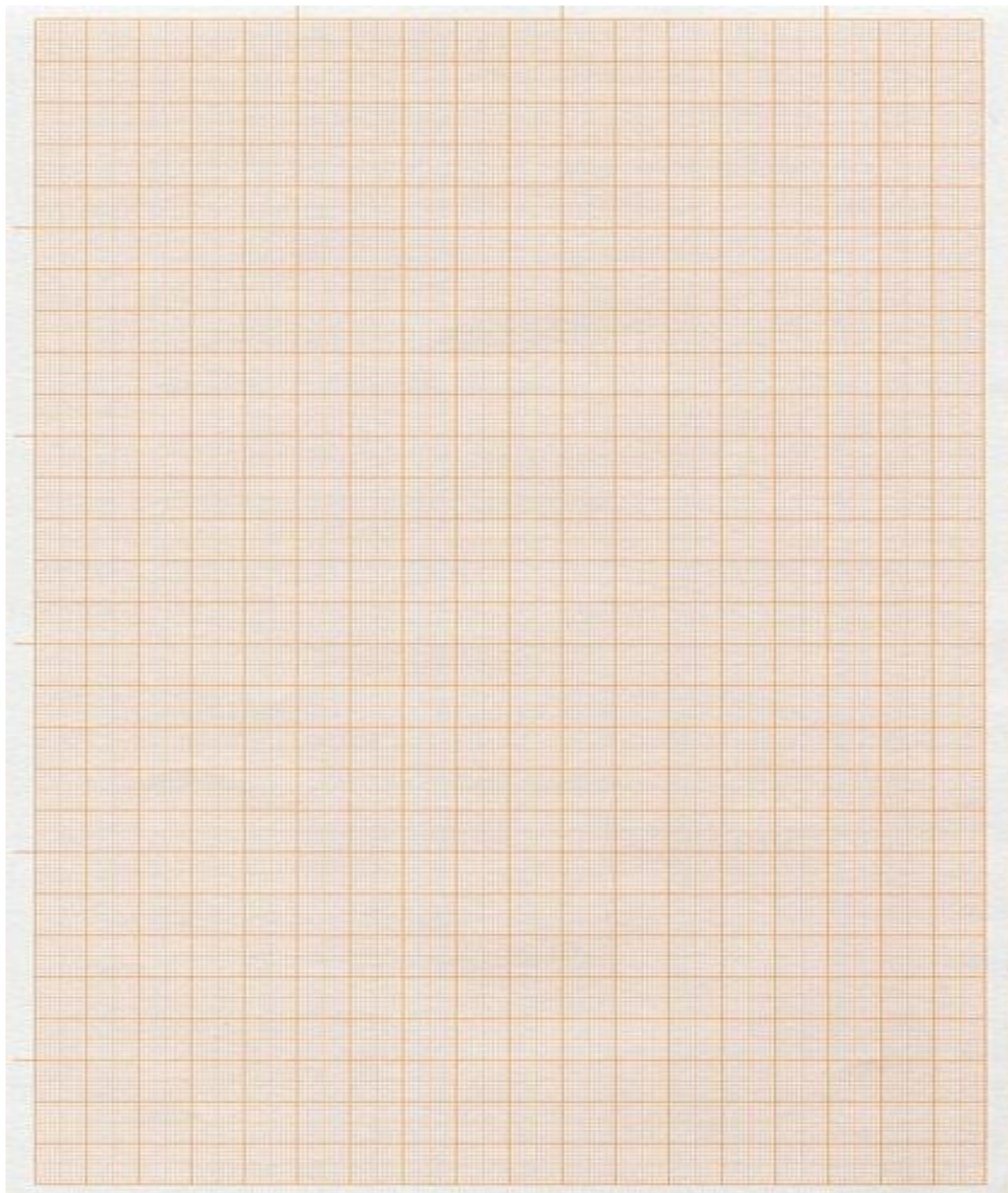
1. Fill the two burettes with the acid and thiosulfate solutions.
2. Draw a small circle in the center of the white paper, shade it in with ink, and place the beaker over it.
3. Prepare the clock.
4. Place 25 ml of the thiosulfate solution in the beaker.
5. Take 4 ml of the acid solution into the small beaker.
6. Pour the acid from the small beaker while simultaneously pressing the clock to begin timing.
7. Stop the clock the moment the circle is completely hidden and record the time.
8. Repeat steps (2) to (4) according to the results table (1).

Result Table (1):

Log [Na ₂ S ₂ O ₃]	Log 1/t	Time (min)	V (HCl) (ml)	V (H ₂ O) (ml)	V (Na ₂ S ₂ O ₃) (ml)	[Na ₂ S ₂ O ₃]
			4 ml	0	25	
				5	20	
				10	15	
				15	10	
				20	5	

$$\log \frac{1}{t} = \log k' + n \log [\text{Na}_2\text{S}_2\text{O}_3]$$

By drawing the relationship between **log [Na₂S₂O₃]** after dilution (see the table) on the x-axis and $\log \frac{1}{t}$ on the y-axis, we obtain a straight line whose slope is (**n**) the order of the reaction.



Second: Change the concentration of HCl while the concentration of Na₂S₂O₃ remains constant:

Procedure:

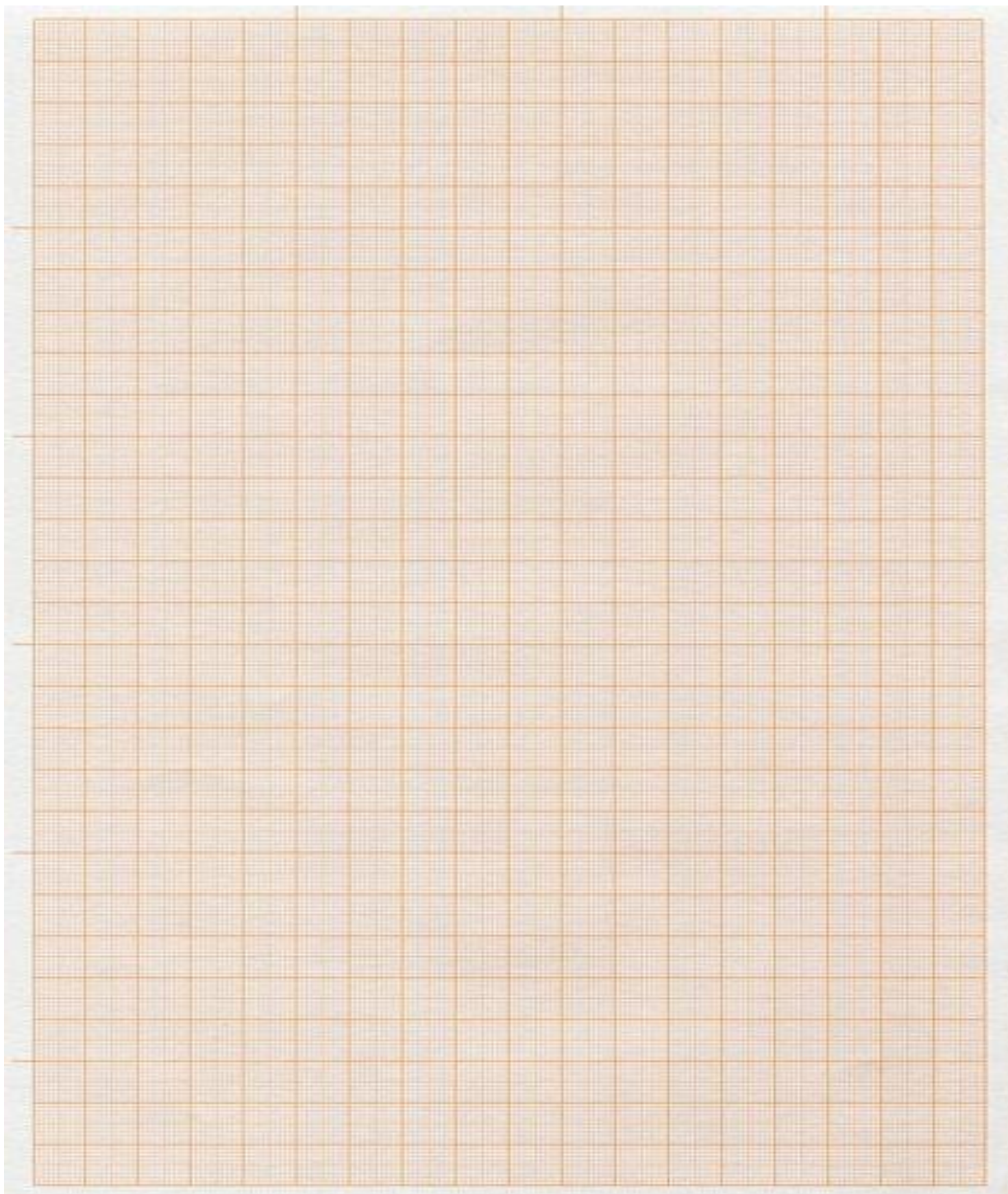
1. Fill the two burettes with the acid and thiosulfate solutions.
2. Draw a small circle in the center of the white paper, shade it in with ink, and place the beaker over it.
3. Prepare the clock.
4. Place 25 ml of the acid solution in the beaker.
5. Take 4 ml of the thiosulfate solution into the small beaker.
6. Pour the thiosulfate from the small beaker while simultaneously pressing the clock to begin timing.
7. Stop the clock the moment the circle is completely hidden and record the time.
8. Repeat steps (2) to (4) according to the results table (2).

Result Table (2):

Log [HCl]	Log 1/t	Time (min)	V (Na ₂ S ₂ O ₃) (ml)	V (H ₂ O) (ml)	V (HCl) (ml)	[HCl]
			4 ml	0	25	
				5	20	
				10	15	
				15	10	
				20	5	

$$\log \frac{1}{t} = \log k' + m \log [\text{HCl}]$$

By drawing the relationship between **log [HCl]** after dilution (see the table) on the x-axis and $\log \frac{1}{t}$ on the y-axis, we obtain a straight line whose slope is (**m**) the order of the reaction. **Overall reaction order = n + m**



Experiment 10: Molecular Weight Determination by Depression of Freezing Point

Objective:

To determine the molecular weight of a non-volatile solute using the freezing point depression method.

Theory:

$$\Delta T = K_f \times m$$

where ΔT is the freezing point depression, K_f is the cryoscopic constant, and m is the molality.

From this, molecular weight can be calculated as:

$$M_{wt} = (K_f \times m_2 \times 1000) / (\Delta T \times m_1)$$

Apparatus and Materials:

Calorimeter; thermometer; glass rod; ice; beaker (250 ml); burette; balance; **unknown A**; **unknown B**.

Procedure:

- 1 – Place 25 ml of distilled water using a burette into the calorimeter.
- 2 – Weigh 5 g of unknown A and place it inside the calorimeter. Then, dissolve it completely.
- 3 – Close the calorimeter, insert the thermometer, and place the calorimeter inside the beaker filled with ice.
- 4 – Wait until the first part of the solution inside the calorimeter freezes, then record the freezing point.

Results:

	A مجهول	A مجهول
Solvent Mass (m_1)	25	25
Solute Mass (m_2)	5	5
Solvent Freezing point (T_1)	0	0
Solvent Freezing point (T_2)		
$\Delta T = (T_2 - T_1)$		
Molecular Weight (M_{wt})		

$$M_{wt} = K_f \cdot (m_2 \times 1000 / \Delta T \times m_1)$$

$$K_f = 1.86 \text{ } ^\circ\text{C/mol}$$