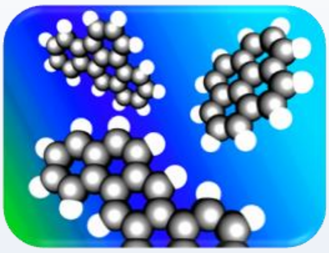
Laboratory Note Book (CH1202)

Department of Chemical Sciences



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Semester II, March 2023 - July 2023

Time: 2:30 p.m. – 5.30 p.m.

GENERAL INSTRUCTIONS

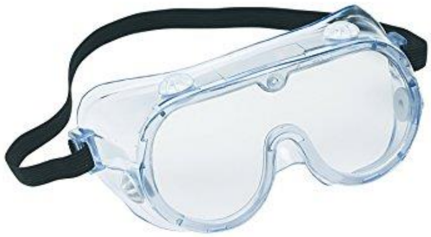
1. **Attendance** is mandatory. In case of illness etc. the student must contact the instructor and fix a schedule for making up the missed lab. All labs must be completed in order to get a passing grade.
2. **All data and results** should be recorded directly in the lab notebook. The recording should include, title of the experiment, date of experiment, working formula, data in tabulated forms, results and calculations.
3. The instructor **should sign the data** before the student leaves the lab.
4. **Graph papers and computer print-outs may be directly pasted** on the lab notebook.

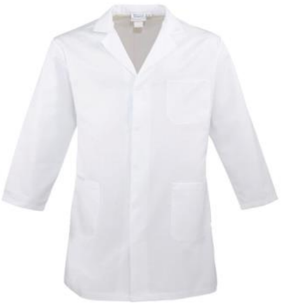
Grading:

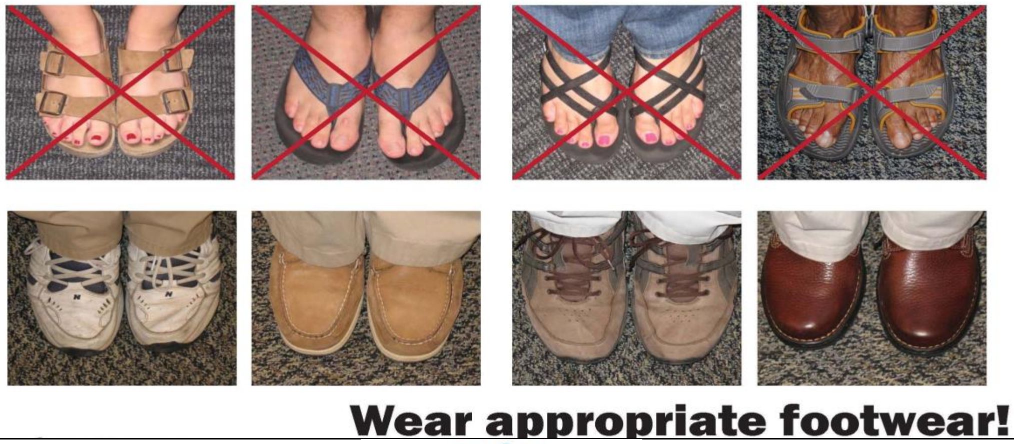
The marking scheme in the lab will be as follows:

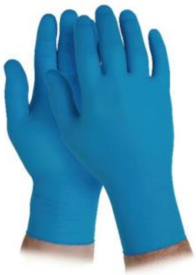
1.	Mid-semester examination (Quiz)	 30
2.	Lab notebook	 10
3.	Attendance	 10
4.	Continuous assessment by teacher	 10
5.	Final examination	 40

Mandatory!!!



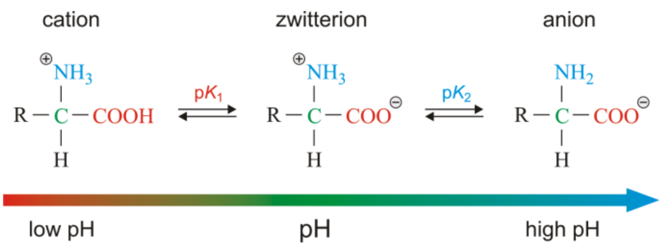


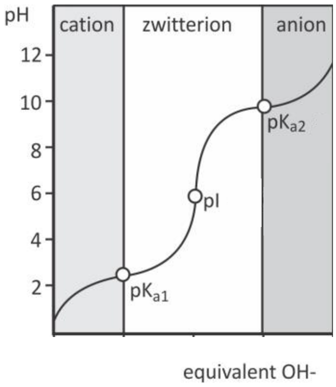

Wear appropriate footwear!



EXPERIMENT 1

DETERMINATION OF ISOELECTRIC POINT OF AN AMINO ACID





PRINCIPLE: Amino acids are molecules that contain both a base site (an -NH_2 group) and an acid site (a -COOH group). Individual amino acids differ only in the identity of the group, -R. On dissolution of an amino acid in water, the proton from the -COOH group gets transferred to the -NH_2 end of the molecule as the NH_2 group is a stronger base than -COO^- resulting a *zwitterion*. Depending on the pH of the solution the amino acids will be either in cationic form (low pH) or in anionic form (high pH).

- The pH at which the presence of these two types of ions in the same concentration is called the isoelectric point (**pI**). **At this pH, the amino acid does not migrate in an electric field**. (gel electrophoresis)
- **pI** is the pH at which the amino acid is neutral, *i.e.* the zwitterion form is dominant.

INTRODUCTION:

$$HA(aq) + H_2O = H_3O^+(aq) + A^-(aq)$$

The extent of this reaction is indicated quantitatively using the equilibrium constant, K_{eq}. The equilibrium constant is given as

$$K_{eq} = K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]} \qquad \text{............. (1)}$$

The equilibrium constant for reaction of an acid with water is usually symbolized as K_a to remind us the type of reaction being dealt with.

- The concentration of water, since present in high concentration and thus essentially a pure liquid, is not included in eq. (1).
- The strength of an acid in aqueous solution is defined in terms of the magnitude of K_a. Strong acids have K_a values larger than 1 and that of weak acids is less than 1.

The equilibrium established when a weak acid reacts with water can be explored using the following pH titration: The pH of the solution must change as the titration proceeds.

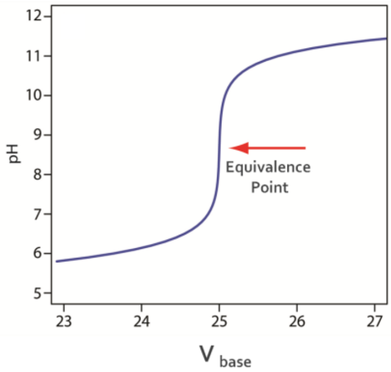
1. At the beginning of the process, before base is added, the pH of the solution is fairly low because it contains acid.
2. As titration proceeds, acid is neutralized by the added base, and pH rises.
3. Addition of base after all of the acid has been neutralized produces a basic solution, with a high pH.
4. The pH of the solution at each interval is monitored by a pH meter.

A plot of pH versus the volume of titrant added to the solution gives the so-called titration curve.

- ➢ The curve is shaped like "S". All titration curves have this characteristic shapes.

This provides the method for determining the equivalence point:

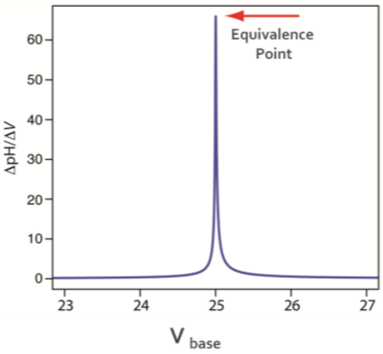
We successively add small volumes of base, measure pH after each addition, and plot the titration curve, from which we may find V_{base} at the inflection point (the equivalence point).



Moles of acid in the original aliquot is calculated as : Moles of acid = V_{base} at inflection point x M_{base}

A derivative plot needs to be created to determine the pK_a values accurately. The steps are as follows:

1. Calculate ΔpH/ΔV from the collected pH data for each addition of the titrant.
2. Plot ΔpH/ΔV against V (volume of titrant added).
3. The plot gives sharp peaks at the equivalence points corresponding to the sharp jumps in the titration plot.



In the case of amino acids, titration of the zwitterion with standard NaOH would provide the K_a value for the $-NH_3^+$ acid, which is expected to be similar to that of NH_4^+ ($pK_a = 9.25$). However, K_a

value for the -COOH group could also be determined. It is possible to generate the acid form in solution by adding a strong acid to the zwitterion. The strong acid transfers a proton to the -COO$^-$ group of the zwitterion, resulting into a cation. Titration of a solution of this cation with standard NaOH should then yield two equivalence points, one for each acid. It should thus be possible to measure both the K_a values.

APPARATUS: pH meter, beaker, burette, pipette, glass rod, spatula.

CHEMICALS: Potassium hydrogen phthalate, glycine, alanine, HCl, NaOH, phenolphthalein.

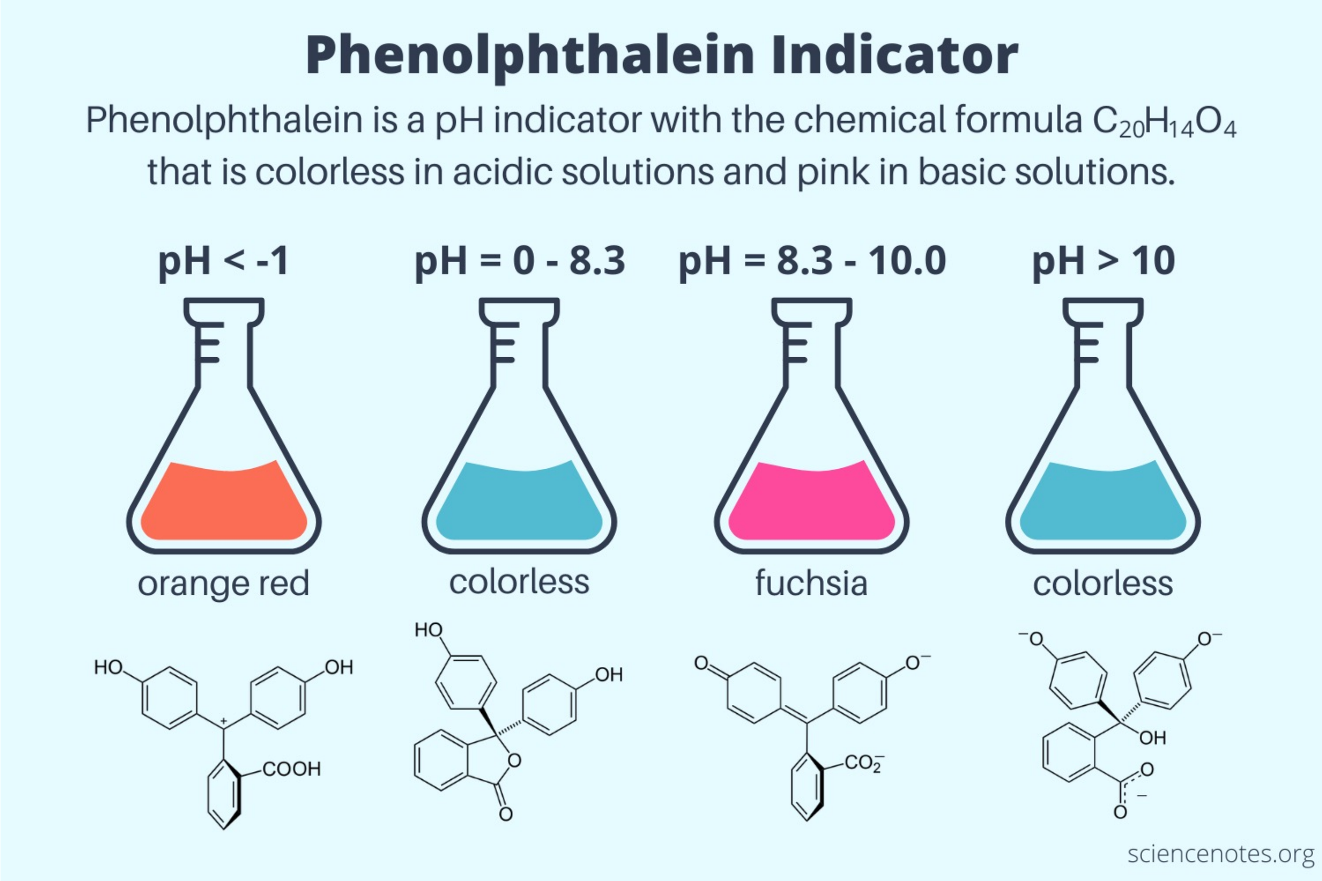
EXPERIMENTAL PROCEDURE:

i) Standardisation of NaOH solution

1) Prepare 100 mL 0.1 *M* KHP (Potassium hydrogen phthalate) solution.
2) Standardize the supplied ~0.1*M* NaOH solution against KHP solution using phenolphthalein indicator (three results).

ii) Amino acid titration and estimation of equivalence point

1) Transfer exactly 10 mL of the supplied protonated amino acid solution to a clean 100 mL beaker.
2) Add 15 mL of distilled water to the beaker so that the total volume of the amino acid solution is 25 mL.
3) To homogenize the solution, place the beaker on the top plate of a magnetic stirrer and place a 1-inch stir bar in the beaker. Rinse the pH electrode and submerge it in the solution containing protonated amino acid. Make sure that the tip of the electrode is clear of the magnetic stir bar in the beaker before starting the stirrer. The rotation rate should be reasonably fast, but not so vigorous that splashing of the solution occurs.
4) Record the initial pH of the solution. Initiate the pH titration by adding 0.5 mL of NaOH solution from burette.
5) On each addition of base solution, note the pH of the solution. Continue this addition until you find larger gaps between two subsequent pH values. This indicates approach of the equivalence point. Reduce the volume of addition of the alkali solution to 0.1 mL until you comfortably cross the sudden jump in pH, indicating the equivalence point. After the equivalence point is passed, increase each volume of addition to 0.5 mL. Repeat this process if you expect more than one equivalence points.
6) Discard the solution on completion. Rinse the pH electrode with distilled water till pH meter reading is approximately equal to that of distilled water. Leave the pH electrode in beaker of distilled water and turn the meter off.



Phenolphthalein Indicator

Phenolphthalein is a pH indicator with the chemical formula $C_{20}H_{14}O_4$ that is colorless in acidic solutions and pink in basic solutions.

pH < -1
pH = 0 - 8.3
pH = 8.3 - 10.0
pH > 10
orange red
colorless
fuchsia
colorless
HO
OH
+
COOH
HO
OH
O
O
O
O⁻
CO₂⁻
⁻O
O⁻
OH
O
O
–
sciencenotes.org

RESULTS:

Table 1. Preparation of 100 mL standard 0.1 *N* KHP solution

Weight taken (g)	Weight to be taken (g)	Strength of KHP solution

Table 2. Standardization of NaOH solution using standard KHP solution

Sl. No.	Volume of KHP (mL)	Burette reading (mL)			Average volume (mL)	Strength of NaOH solution
		Initial	Final	Difference		

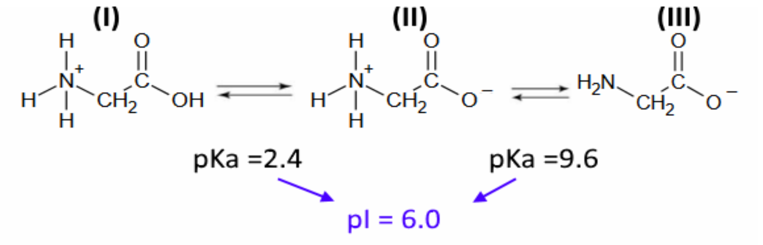
Table 3. Titration of amino acid solution using standard NaOH solution

Volume of amino acid (mL) =

Sl. No.	Volume of NaOH (mL)	pH	ΔV (mL)	ΔpH	$\Delta pH/\Delta V$

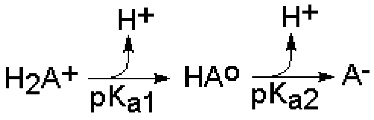
DISCUSSION:

Amino acids are more complicated than simple weak acids since amino acids have at least 2 ionizing groups. Glycine, for example, has both a carboxylic acid and an amino group that can ionize: If we dissolve the free base of glycine in pure water (ie neutral pH), it will ionize by protonating itself. The equilibrium is far to the right so most of the glycine is in the charged form called the zwitterion and glycine is still neutral because the +ve charge is netualized by the -ve charge. Glycine is always in the zwitterion form at neutral pH.



Glycine

Now if we put Glycine at an acid pH where it is fully protonated (i.e., it has all the protons bound to it which it bind), we can titrate it to reveal its 2 pK values for the alpha-carboxylic acid group and the alpha-amino group.

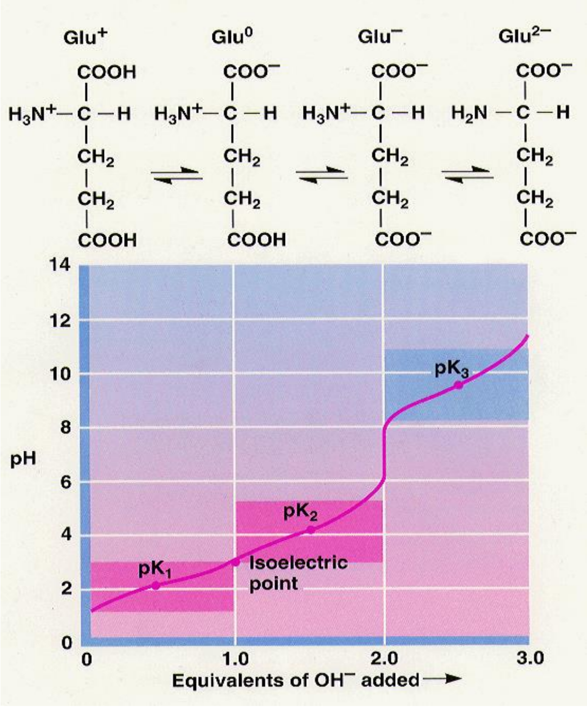
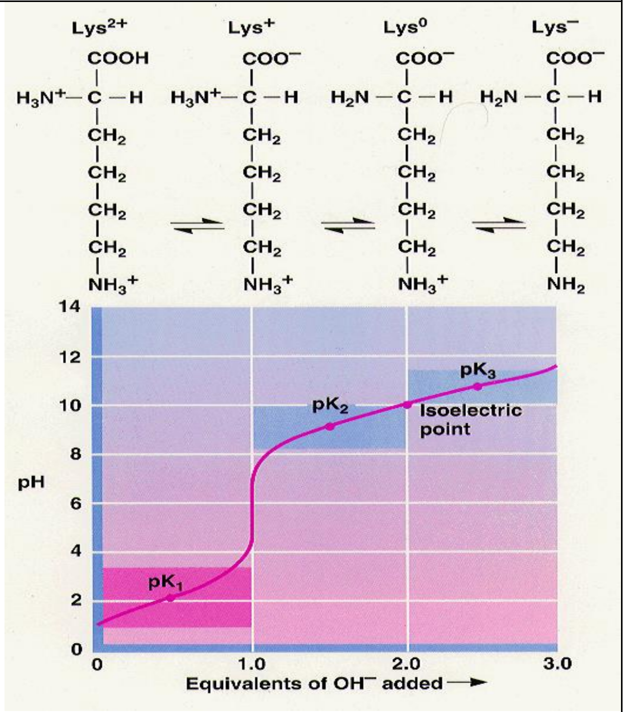


From the pK values, the pI (called the electric point or the place where Glycine has no net charge) can be calculated:

pI = (2.4 + 9.6)/2 ≈ 6 ; (pK_1 = 2.4, pK_2 = 9.6)

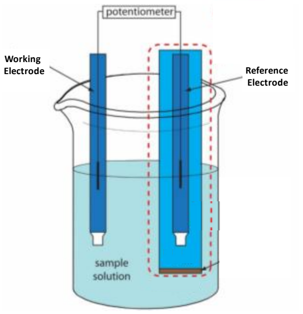
Glycine is neutral at pH 6; it has no net charge here.

Some amino acids are classified as triprotic. This is because, in addition to the ionizable protons of the α-COOH and α-NH_3 groups, they also have a dissociable proton in their R group. Although triprotic amino acids can exist as zwitterions, under physiological conditions these amino acids will be charged. If the net charge under physiological conditions is **negative**, the amino acid is classified as an **acidic** amino acid because the R group has a proton that dissociates at a pH significantly below pH 7. The remaining triprotic amino acids are classified as **basic** amino acids due to a) their having a net **positive** charge under physiological conditions and b) an R group dissociable proton with a pKa near or greater than pH 7. Titration curves for triprotic amino acids generate the same information as those for the diprotic amino acids. The pI for a triprotic amino acid can be determined graphically, although this is somewhat more challenging.

Glutamic Acid	Lysine
	

EXPERIMENT 2

Determination of the Degree of Hydrolysis and the Hydrolysis Constant by Potentiometry



PRINCIPLE: A potentiometer is used to determine the difference between the potential of two electrodes. The potential of one electrode—the working electrode—responds to the analyte's activity, and the other electrode—the reference electrode—has a known, fixed potential.

INTRODUCTION:

Anilinium hydrochloride, $C_6H_5NH_3^+Cl^-$ when dissolved in water, ionizes to form $C_6H_5NH_3^+$ and Cl^- ions, and the cation establishes the following hydrolytic equilibrium.

$$C_6H_5NH_3^+ + H_2O \rightleftharpoons C_6H_5NH_2 + H_3O^+$$

The equilibrium constant for this hydrolytic process is called the hydrolysis constant for the salt and is given by,

$$K_h = \frac{(a_{H^+} \times a_B)}{a_{BH^+}}$$

where, a_{H^+} is the activity of the free acid (H_3O^+); a_B is the activity of the free base ($C_6H_5NH_2$)· and a_{BH^+} is the activity of the unhydrolysed salt ($C_6H_5NH_3^+Cl^-$). However, for dilute solutions, we may replace activities by concentration term (why?) ; Hence,

$$K_h = [H^+] [B] / [BH^+] \qquad (1)$$

Hydrolysis constant can also be related to the dissociation constant, K_b, of the base through the ionic product of water, K_w as

$$K_h = K_w / K_b \qquad (2)$$

If c equivalents of the salt is dissolved in a litre of water, $c\alpha$ equivalents each of free base and free acid will be formed due to hydrolysis (α is the degree of hydrolysis). Thus, the pH of the solution may be related to the degree of hydrolysis as,

$$pH = -\log [H^+] = -\log (c\alpha)$$

Hence, by measuring the pH of the solution, $c\alpha$ can be calculated from which the degree of dissociation α can be obtained at a given concentration.

Also, expressing K, in terms α, using $K_h = c\alpha^2/(1-\alpha)$, the hydrolysis constant can be calculated. Substituting for K_h in equation (2) and taking $K_w = 1.0 \times 10^{-14}$ at 25°C the dissociation constant of the base, K_b can be evaluated.

APPARATUS : Potentiometer, Platinum electrode and calomel electrode.

CHEMICALS : Anilinium hydrochloride, quinhydrone,

EXPERIMENTAL PROCEDURE:

1) Prepare an N/10 aniline hydrochloride solution by dissolving appropriate quantity of the substance in distilled water (100 mL).
2) From this stock solution, dilute appropriately and get 25 mL of M/20, M/50 and M/100 solutions. Then construct the following cell:
3) Transfer the 25 mL solution to 100 mL beaker; add a pinch of Quinhydrone, stir properly to dissolve it, dip the electrodes (Pt and Calomel electrodes) in to the solution.

   ⊕ Pt|0.1 M Aniline Hydrochloride, Quinydrone **//** Calomel ⊖

4) Determine the potential of the cell. Repeat the experiment with each of the other solutions.

RESULTS:

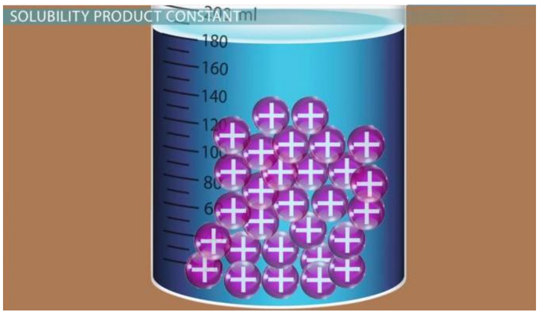
1) pH is given by pH = $(-E_{obs} + E_{QH} + E_{cal}) / 0.0591$ where $E_{QH} = 0.6996$ V and $E_{cal} = -0.242$ V (Oxidation Potential). From this relation, pH of the solution can be calculated.
2) As pH = $-\log[H^+] = -\log(c\alpha)$, pH = $-\log c - \log \alpha$, the degree of hydrolysis α can be calculated at every given concentration.
3) From α, calculate the hydrolysis constant using the relation, $K_h = c\alpha^2/(1 - \alpha)$
4) The dissociation constant, K_b can be calculated from the relation $K_b = K_w/K_h$

$C_6H_5NH_3^+Cl^-$	E_{obs}(V)	pH	α	K_h	K_b
N/10					
N/20					
N/50					
N/100					

Mean K_h = - ... , Mean K_b = ...

EXPERIMENT 3

Determination of K_{sp}, ΔG°, ΔH° and ΔS° for dissociation of $Ca(OH)_2$ in water



PRINCIPLE: Calcium hydroxide is used in a broad range of industrial, professional and consumer applications. Calcium hydroxide is used as such or in a mixture for the production of articles to be used in or for vehicles, construction, electronic apparatus, laboratories, fabrics, wood, rubber, plastics, metal, leather, chemicals, pharmaceuticals, the treatment of potable water, waste water, municipal sludge cosmetics, personal care products.

The solubility of solutes in mixed solvents is of great practical importance since many industrial process as well as laboratory procedures call for the use of solvent mixtures. The solubility of solutes in mixed solvents depends primarily on the solvation of solutes or their constituent ions by the components of solvent mixtures. Studying the thermodynamics of different salts, is important for evaluating the single ion thermodynamic parameters which help in explain the preferential solvation of the ions.

THEORY:

If we consider a reaction at equilibrium:

$$aA + bB \rightleftharpoons cC + dD$$

& the simple relationship between the equilibrium constant and the standard-state Gibbs free energy difference between products and reactants:

$$ln(\frac{a_C^c . a_D^d}{a_A^a . a_B^b}) = ln(K_{eq}) = \frac{-\Delta G^0}{RT}$$

Free Energy difference between products and reactants:

$$K_{eq} = e^{-(\frac{\Delta G^0}{RT})}$$

In a nutshell, for today's experiment, saturated solutions of calcium hydroxide at two (or more) temperatures are prepared and titrated with standardized hydrochloric acid solution. Using these titration data, the K_{sp} and ΔG ° for the dissolution of calcium hydroxide are determined. ΔG °, ΔS °, and ΔH ° for the reaction are determined from the temperature dependence.

Experiment and additional notes:

When the sparingly soluble compound calcium hydroxide is added to water, an equilibrium is established between the solid and aqueous material that can be approximated by eq 1, ignoring ion-pair formation. The equilibrium concentration of OH^- is large enough to allow its accurate determination by titration with 0.01 M HCl.

$$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$$

Aliquots of the filtered solution are titrated with standardized hydrochloric acid. The molar solubility, s, of calcium hydroxide is then determined from the calculated $[OH^-]$.

$$Ca^{2+}(aq) + 2OH^-(aq) + 2H_3O^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + 4H_2O(\ell)$$

Hence,

$$\text{moles of } Ca(OH)_2 = \tfrac{1}{2} \text{ moles of hydroxide}$$

$$\text{Molar solubility of } Ca(OH)_2\ (s) = (\text{moles of } Ca(OH)_2 \text{ in aliquot})/(\text{volume of aliquot in liters})$$

The apparent solubility product, K_{sp}, is then calculated from

$$K_{sp} = [Ca^{2+}][OH^-]^2 = 4s^3$$

The Gibbs free energy at temperature T, ΔG_T°, is related to
the equilibrium constant, K_{sp},

$$\Delta G_T^\circ = -RT\ln(K_{sp})$$

where R = 8.314 J mol^{-1} K. Since $\Delta H°$ and $\Delta S°$ do not change significantly over small temperature ranges, these quantities can be calculated from following equation by generating simultaneous linear equations at two temperatures:

$$\Delta G_T^\circ = \Delta H^\circ - T\Delta S^\circ$$

APPARATUS :

CHEMICALS :

EXPERIMENTAL PROCEDURE :

WARNING: Although sparingly soluble, calcium hydroxide is a strong base. Ingestion or contact with skin or eyes should be avoided.

1. Prepare 0.04N oxalic acid solution and standardise the supplied ~ 0.04 N NaoH solution.
2. Standardise the HCL solution provided (0.04 N)
3. A solution of calcium hydroxide will be stirring at room temperature in the laboratory. Draw off approximately 40 mL of this solution and **record the temperature**. **Filter** the solution

using a long stem funnel until you have at least 30 mL of filtrate (if the filtering is slow, consider starting step 3 below). Place 10.0mL of the filtrate in to a clean 125mL Erlenmeyer (Conical) flask and add 25mL of distilled water and a few drops of bromothymol blue indicator. **Titrate** with the standard HCl solution until the **yellow** endpoint appears. Record the volume of HCl used. **Repeat** this procedure two more times.

Table1

Conc. Of HCl--------------------- Room Temp:

Sl. No.	Volume of the filtrate	Burette reading (mL)			Average volume of HCl (mL)	Strength of Ca(OH)2 solution
		Initial	Final	Difference		
1						
2						
3						

4. Prepare a 100°C saturated calcium hydroxide solution by bringing 100mL of distilled water to a 250 mL beaker and boil. After the water has been boiling for several minutes, add about 2 g of $Ca(OH)_2$ to the water and keep it near boiling with occasional stirring until needed.
5. Bring your hot solution to a gentle boil for about two minutes, turn off the burner, **measure the temperature** and quickly draw off about 40 - 50 mL of solution. Quickly filter the solution using a clean, dry long stem funnel.
6. Place 10.0 mL of the cooled filtrate into each of three clean 125 mL Erlenmeyer (Conical) flasks and add 25 mL of distilled water and a few drops of bromothymol blue indicator. When the solution is cool, titrate with the standard HCl solution until the **yellow** endpoint appears (probably no more than 5mL of HCl will be required to reach the endpoint). Record the volume of HCl used. **Repeat** two more times.

Table2

Conc. of HCl---------------------Temp: 100°C

Sl. No.	Volume of the filtrate	Burette reading (mL)			Average volume of HCl (mL)	Strength of $Ca(OH)_2$ solution
		Initial	Final	Difference		
1						
2						
3						

7. Dispose of all used and excess chemicals in a waste container. Rinse your burette and all glassware with water before returning to its original location.

Calculations:

1. Find the **average solubility** of calcium hydroxide at each temperature by finding the hydroxide ion concentration and then converting this value to the molar solubility of $Ca(OH)_2$.
2. Find **K_{sp}** at each temperature using the average solubility value. Find **ΔG°** at each temperature using the two values of K_{sp}.
3. Find **ΔH°** and **ΔS°** using the values of K_{sp} and ΔG° at the two temperatures.

Results

Ksp	
ΔG°	
ΔH°	
ΔS°	

EXPERIMENT 4

Determination of Molecular Weight of a Polymer by Ostwald Viscometer

THEORY:

The viscosity co-efficient (η) of a liquid can be measured using Ostwald viscometer with the help of Poiseuille's equation (applicable for streamline flow of fluid).

$$\eta = \frac{\pi(\Delta p)r^4t}{8lv} \equiv \frac{\pi(h\rho g)r^4t}{8lv} \qquad \text{(Symbols have their usual meaning)}$$

If we use same viscometer (for same 'r' & 'l') for same volume of two liquids (for same 'v' & 'h') at same place (for same 'g') and temperature then

$$\eta_1 / \eta_2 = \rho_1 t_1/\rho_2 t_2$$

$$\Rightarrow \eta_1 = \eta_2 \times(\rho_1 t_1/\rho_2 t_2)$$

Hence, we can use the above equation (1) to measure the η_1 (i.e., viscosity co-efficient of liquid 1) by using the known value of η_2. So it is a relative method.

The viscosity of a polymer solution (η) is higher than that of the pure solvent (η_o) at a specified temperature and the increase in solution viscosity on dissolving the polymer in the solvent is a function of both molecular weight and concentration of the polymer solute

The ratio of the viscosity of a solution (η_s) to the viscosity of the pure solvent called relative viscosity η_r is given by

$$\eta_r = \frac{\eta_s}{\eta_0} = \frac{d_s t_s}{d_0 t_0}$$

And for solutions of low concentrations, densities of solutions, d_s are almost equal to the density of solvent d_o. Therefore,

$$\eta_r = \frac{t_s}{t_0}$$

Thus η_r is an easily measurable parameter for solutions of polymers. Increment in viscosity for any solution would be ($\eta_s - \eta_o$).

For a polymer solution, the specific viscosity, η_{sp}, is given by

$$\eta_{sp} = \frac{\eta_s - \eta_0}{\eta_0} = \frac{\eta_s}{\eta_0} - 1$$

$$\eta_{sp} = \eta_r - 1 = \frac{t_s}{t_0} - 1$$

For polymers, we go a step further and divide η_{sp} by the concentration of the solute in solution in terms of grams of solute per 100 ml of solution and we call this parameter reduced viscosity η_{red} for the solution.

$$\eta_{red} = \frac{\eta_{sp}}{c}$$

The values η_r and η_{sp} change rapidly with change in concentration but the η_{red} values of solutions change less and regularly or linearly. Thus a plot of η_{red} versus concentration (g/100ml) is very nearly a straight line (Fig. 1). For this linear relationship, we can write

$$\eta_{red} = \frac{\eta_{sp}}{c} = mc + constant$$

As for a straight line graph, m is the slope of the line and the addition constant is the intercept on the reduced viscosity axis and is called the intrinsic viscosity of the solution and is given by the symbol $[\eta]$

$$[\eta] = \lim_{c=0}\left(\frac{\eta_{sp}}{c}\right)$$

For any value of c, for the solution

$$\eta_{red} = \frac{\eta_{sp}}{c} = mc + [\eta]$$

For polymers, the intrinsic viscosity values can be obtained from the η_{red} versus c graphs. This in turn is related to molar masses (M) of straight chain polymers by a simple equation,

$$[\eta] = K\,M^a \qquad \text{[Mark-Houwink Equation].}$$

which is more valid when molar masses are above 10,000. Here K and a are constants for a given pair of polymer and the solvent at a *specified temperature*. The values of the exponent constant `a' also called the shape factor for a randomly coiled flexible linear chain polymer molecules range from 0.5 to 0.8. It is most often 0.7 but for rigid rod-like polymer molecules its values may rise to 20. The uncertainty in values of K and a make viscosity average molar masses somewhat imprecise and different from the number average and weight average molar masses. But the results still remain useful in practice.

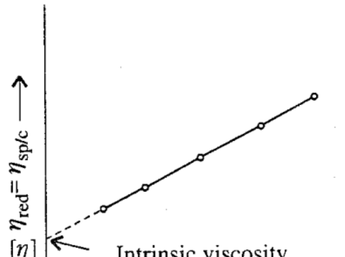


Fig. 1 Intrinsic Viscosity

APPARATUS :

CHEMICALS :

EXPERIMENTAL PROCEDURE:

1. Weigh 4.0g of dry polyvinyl alcohol on a watch glass.
2. Take about 200 mL of hot distilled water in a beaker. Gradually spread the weighted amount of polymer in small lots on the surface of hot water and stir slowly without forming bubbles of foam.
3. When the whole of the polymer quantity is thus dissolved, cover the solution and allow it to cool to room temperature.
4. Transfer the cooled solution slowly to a 250 mL volumetric flask along the side wall to minimise formation of bubbles. Thermostat the solution.
5. Rinse the beaker with small lots of distilled water and transfer these rinsings to the volumetric flask till the solution volume is made up to the mark. This is the *master solution*.
6. Prepare six other solutions whose concentrations are 80%, 60%, 40%, 30%, 20% and 10% of the *master solution*, in 50 mL volumetric flask and find the time of flow for each solution with a Ostwald viscometer.

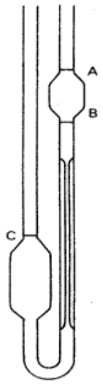


Fig. 2. Ostwald Viscometer

USING OSTWALD VISCOMETER:

1. Adjust the thermostat at certain temperature (preferably 35 °C) and record its temperature.

2. Clean the viscometer with chromic acid and wash thoroughly with distilled water. Clamp the viscometer vertically in the thermostat. Introduce a definite volume of water into the tube, by a graduated pipette, so as to fill the bend of the tube and at least half of the bulb of the wider tube. When the water in the viscometer will attain temperature of the thermostat then draw the water through the narrow tube above the upper mark (by applying manual suction through the rubber tube). Then allow the water to flow down through the capillary in the viscometer. Start the stopwatch just when the meniscus will pass the upper mark (A in Fig.2) and stop when it will pass the lower mark (B in Fig.2). Repeat it for at least three times and each time note the flow time (t_2) in second and take its mean. Take out the

viscometer, wash with acetone, dry. Clamp it again in the thermostat and introduce same volume of experimental liquid. Repeat same procedure (as of water) and note the flow time (t_1) in seconds.

Observations:

Temperature of Thermostat = °C

Solvent: Solute:

Values of constants: K= a =

S.No.	Concentration (g/100mL)	Average flow time (seconds)	η_r	η_{sp}	η_{red}

Plot the η_{red} values versus concentration (g/100 mL). Obtain the intrinsic viscosity value $[\eta]$.

Calculating molar mass

$$[\eta] = K\ M^a$$

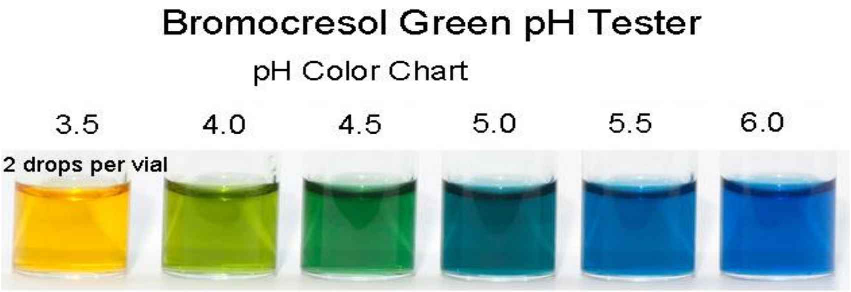
$$\log[\eta] = \log K + a \log M$$

$$\log M = \frac{\log[\eta] - \log K}{a}$$

$$M = \ Antilog \left[\frac{\log[\eta] - \log K}{a}\right]$$

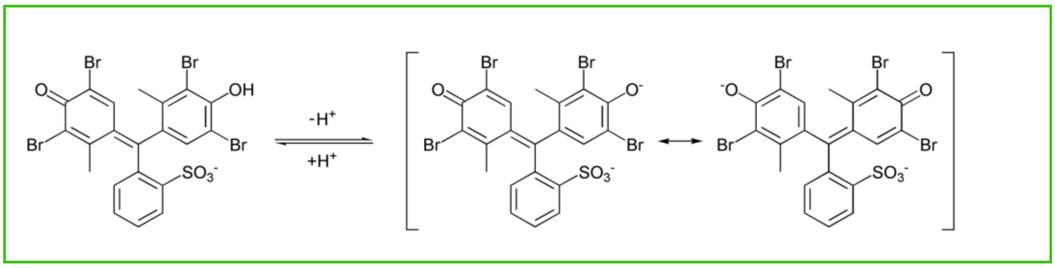
EXPERIMENT 5

DETERMINATION OF THE pK_{In} VALUE OF AN ACID-BASE INDICATOR BY SPECTROPHOTOMETRIC METHOD



PRINCIPLE: Spectrophotometric methods will be used to determine the acid dissociation constant of an acid-base indicator (Bromocresol green), the light absorption characteristics of its acid and base form. This experiment will provide you with opportunities to refine your understanding of absorption process while providing an opportunity to apply many aspects of acid-base chemistry.

INTRODUCTION: Acid–base indicators are weak acids or bases having distinctly different colours in acidic and alkaline solution, and by virtue of change of colour they indicate the end points of acid-base titrations. To illustrate this point, consider the case for Bromocresol green (an organic acid):



As shown above, this proton can be donated/or received to water to obtain a hydronium ion. If we represent the acidic form of the bromocresol green as (HIn) and the conjugate base as (In-) then the dissociation reaction looks like:

$$HIn \leftrightarrow H^{+} + In^{-} \quad \Longrightarrow \quad H_3O^{+} + In^{-}$$

The acid dissociation constant (Equilibrium constant) can be represented as:

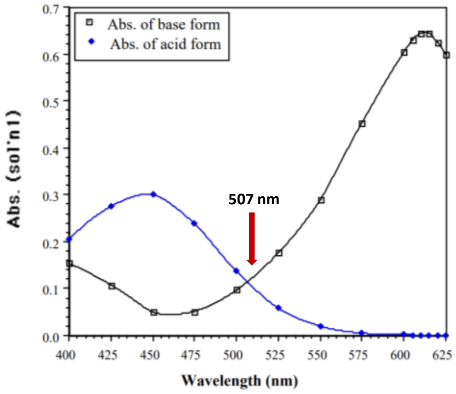
$$K_{In} = \frac{[H^{+}][In^{-}]}{[HIn]}$$

The strategy of this experiment is to adjust [H3O+] to known values using a buffer and then to measure the ratio [In-]/[HIn] spectrophotometrically. Knowing this ratio and the value of [H_3O^{+}/pH] will allow us to calculate Ka using the above equation.

- The trick then is knowing how to determine the ratio [In-]/[HIn] using light absorption measurements .

ABSORPTION SPECTRAL CHARACTERISTICS OF BROMOCRESOL GREEN:

The acid form of bromocresol green (HIn) absorbs light in a different region of the spectrum than the basic form of bromocresol green (In-). From the figure below, you may see that the two species have distinctly different values for λmax. Acid-base indicators are useful for determining pH and indicating end-points because they change color as the pH of the solution changes. It is important to note that the two solutions used to measure these absorption spectra have the same concentration of the bromocresol green indicator. We want to select a wavelength that will allow us to determine the relative concentration of each species present. One very poor choice occurs at about 507 nm, because at this wavelength both the species absorbs equally well. The best choice is the wavelength which has the largest difference in absorbance for the two species. This may be at > 565 nm where the acid form hardly absorbs light.



THEORY: The ionization equilibria of a weak acid indicator (HIn) may be represented according to,

$$HIn \leftrightarrow H^+ + In^- \qquad \text{............... (1)}$$

Acidic form *Alkaline form*

for which the ionization constant (K_{In}) in dilute solution may be defined as the concentration quotient (2)

$$K_{In} = \frac{[H^+][In^-]}{[HIn]} \qquad \text{..................... (2)}$$

where, [ ]'s represent the molar concentrations of the respective species. Transforming the equation (2) in logarithmic form one obtains,

$$pH = pK_{In} + \log \frac{[In^-]}{[HIn]} \qquad \text{................. (3)}$$

(where, $pK_{In} = -\log_{10} K_{In}$ and $pH = -\log_{10} [H^+]$ in dilute solution).

Thus, if a fixed amount of the indicator is placed in the same volume of a series of buffer solutions of different known pH values, the ratio, $[In^-]/[H_{In}]$, will increase with increase of pH. If the values of the ratio at different pH are determined by measuring the colour intensity of the indicator solutions then the pK_{In} value of the indicator can be found out if the pH of the buffer solutions is known.

If the alkaline form of the indicator (In^-) absorbs at a selected wavelength and Beer's law is obeyed in the range of concentration of the indicator used, then the absorbance (A) of the indicator solution at a particular pH will be proportional to its concentration, provided the acid form (H_{In}) does not absorb at this wavelength.

$$A = \varepsilon\,[In^-]\,l \qquad (4)$$

In a strongly alkaline solution, H_{In} is practically absent, and the absorbance (A) will correspond to the total concentration (T_{In}) of the indicator.

$$A' = \varepsilon\,[T_{In}]\,l \qquad (5)$$

Where, ε = molar extinction coefficient of In^- and l = optical path length in cm.

Mass balance equation of the indicator is,

$$T_{In} = [H_{In}] + [In^-] \qquad (6)$$

$$\therefore\ [H_{In}] = T_{In} - [In^-] \qquad (7)$$

From (5) – (4) one obtains, $\frac{(A' - A)}{\varepsilon\, l} = [H_{In}]$ (8)

$$\frac{(A)}{\varepsilon\, l} = [In^-] \qquad (9)$$

Substituting these values of HIn and In^- in equation (3) one obtains,

$$pH = pK_{In} + \log_{10}\left(\frac{A}{A' - A}\right) \qquad (10)$$

A and A′ may be measured colourimetrically. Therefore, by plotting $\log_{10}$ [A/(A′ - A)] against pH of the buffer solutions a straight line of slope =1 will be obtained, of which the intercept on the pH axis will give pK_{In} .

APPARATUS :

CHEMICALS :

EXPERIMENTAL PROCEDURE:

You will be provided with ~0.5 N NaOH and~0.5 N acetic acid. Prepare 0.5 N oxalic acid (50 mL) and follow the procedure.

1. Standardisation of NaOH (~ 0.5 N) using oxalic acid. Than standardise the acetic acid.
2. Prepare 50 mL of exact 0.4 N acetic acid (pK^H= 4.74 at 25°C) and 50 mL of exact 0.4 N NaOH solutions separately by usual procedure.
3. Take 6 hard glass test tubes of uniform dimensions and label them from 1 to 6. Prepare the following series of solutions by proper mixing (experimental pH values may be obtained from chart below, or, may be determined using a pH meter).

Test tube	Vol. of 0.4 N acetic acid (mL)	Volume of 0.4 N NaOH (mL)	Volume of Water (mL)	pH (Expt.)	A	A/(A′-A)
1	5.0	0.5	4.5	3.72		
2	5.0	1.5	3.5	4.27		
3	5.0	2.5	2.5	4.63		

4	5.0	3.5	1.5	4.99		
5	5.0	4.5	0.5	5.57		
6	0	2.5	7.5		A′ =	

4. Add a few drops of bromocresol green indicator to test tube number 6 using a dropper.
5. Set spectrophotometer at 570 nm, adjust the transmittance of water to 100%.
6. Measure the transmittance of the solution in test tube 6. If the transmittance is below 15% (i.e. Absorbance is above 0.82), take test tube 7 and add fewer number drops of the indicator to it and measure the transmittance. In this way by adjusting the number of drops of the indicator, adjust the transmittance of the alkaline form between 25 to 15 % (absorbance is above 0.60 but below 0.82) using test tube numbers 6 to 8 as required.
7. Add the same number of drops of the indicator as adjusted in step 5 to each of test tubes 1 – 5 and measure their transmittance.
8. Calculate the absorbance (A) values of solutions 1 - 5 and the absorbance (A′) of the alkaline solution of the indicator (6, 7 or 8) using the relation:
   $A = \log (100/T\ \%) = 2 - \log T$
9. Plot pH against $\log_{10} [A/(A' - A)]$ and draw the best straight line of unit slope passing through the experimental points, using the same scale for pH and $\log_{10} [A/(A' - A)]$ axis. Find pK_{In} from the intercept on the pH axis.

Table 1: Standardisation of NaOH

Sl. No	Vol. of Oxalic acid (mL)	Burette reading			Avg. Vol(mL)	Strength of NaOH
		Initial	Final	Diff.		

Table 2: Standardisation of Acetic Acid

Sl. No	Vol. of Acetic acid (mL)	Burette reading			Avg. Vol(mL)	Strength of Acetic Acid
		Initial	Final	Diff.		

CONCLUSION: pK_{In} of bromocresol green is ....

EXPERIMENT 6

Determination of the Strength of a Solution of a Strong Acid by Conductometric Titration.

PRINCIPLE: The strength of a solution of a strong acid, namely, hydrochloric acid (HCl) will be determined using a solution of a strong base, namely, sodium hydroxide (NaOH) from the change in the conductance of the solution mixture.

INTRODUCTION:

Suppose an acid is taken in a beaker and NaOH solution is gradually added to it from a burette. The reaction occurring during neutralisation is given by

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$\text{Or, } H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$$

It is evident from the above equations that as NaOH solution is gradually added, the H^+ ions, having high ionic conductance, are replaced by Na^+ having lower (ionic) conductance and hence the conductivity of the solution in the beaker gradually decreases. At the equivalence point the conductivity would be the minimum. After the equivalence point the Na^+ and OH^- ions will be accumulated in the solution that increases the conductance of the solution. If the conductances corresponding to the volume of the NaOH solution added be plotted, two straight lines having opposite slopes will be obtained. The point of intersection of the two straight lines will give the equivalence point.

The strength of the NaOH solution should be at least 5 times greater than that of the HCl solution so that the effect of the volume change on the conductance be negligible.

APPARATUS: Conductivity bridge, conductivity cell, beaker, burette, pipette.

CHEMICALS: Hydrochloric acid solution (0.05 N); Sodium hydrochloride solution (0.2 N); Oxalic acid (0.1 N); Phenolphthalein indicator; Water.

EXPERIMENTAL PROCEDURE:

1) Prepare 250 ml of approximately 0.2 N NaOH solution and standardise it by a standard solution of 0.1 N oxalic acid using phenolphthalein indicator.
2) Take NaOH in a burette.
3) Take 25 ml of the supplied acid by a pipette into a 250 ml beaker and add 125 ml water to it.
4) Place the conductivity cell in the beaker so that the electrodes are completely immersed in the acid solution.
5) Connect the cell to the conductivity bridge and measure the conductance of the solution.
6) Add NaOH solution from the burette 0.5 ml at a time in the beginning, 0.2 ml at a time near the end point point and again 0.1 ml at a time after the end point.
7) Measure the conductance of the solution after each addition of the NaOH solution.

8) Plot the conductance values against the corresponding titre values, draw the straight lines and obtain the point of intersection.

RESULTS:

Strength of NaOH solution S_1 = N

Volume of the supplied acid solution taken V_2 = 25 ml

Table 1. Titration of the supplied acid with the standardised NaOH solution.

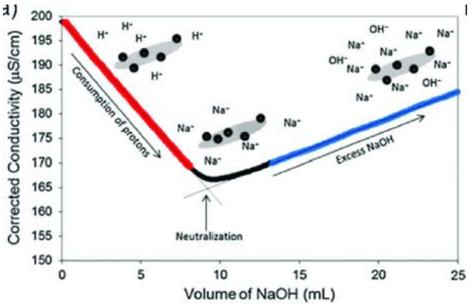
Volume of NaOH solution added (ml)	Observed conductance (ohm^{-1})	Volume of NaOH solution added (ml)	Observed conductance (ohm^{-1})

GRAPH AND CALCULATIONS:

Let the point of intersection of the straight lines obtained by plotting conductances against the titre values correspond to V_1 ml. This gives the volume of NaOH solution required to neutralise 25 ml (V_2) of the acid. Let S_2 be the strength of the supplied acid solution.

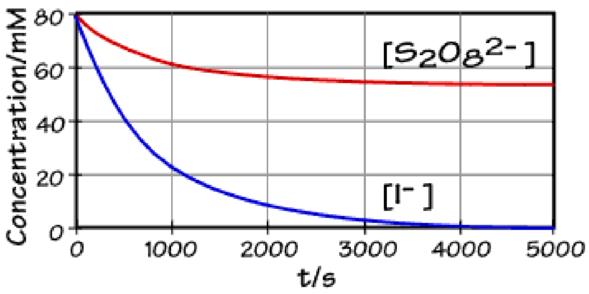
$V_1S_1 = 25 \times S_2$

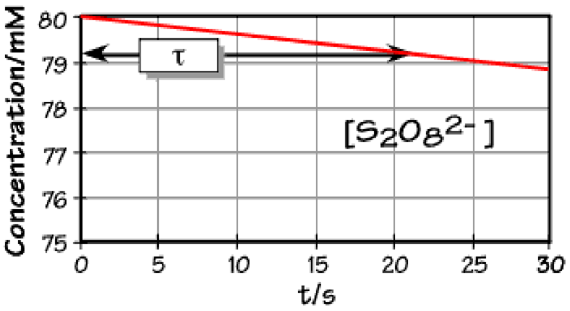
$S_2 = V_1S_1/25$ (N)



EXPERIMENT 7

Determination of Order for the Persulphate-Iodide Reaction





PRINCIPLE:

The rate law for the reduction reaction of persulphate by iodide:

$S_2O_8^{2-}(aq) + 2\,I^-(aq) \rightarrow I_2(aq) + 2\,SO_4^{2-}(aq)$ will be determined. The orders of reaction with respect to persulphate and to iodide will be experimentally obtained by measuring rates of the reaction for various concentrations of the reactants.

INTRODUCTION:

The overall oxidation of iodide ion by persulphate can be expressed as

$$2I^- + S_2O_8^{2-} \rightarrow 2SO_4^{2-} + I_2$$

Or more explicitly

$$S_2O_8^{2-} + I^- \rightarrow SO_4^{2-} + SO_4^- + I^o$$

$$SO_4^- + I^- \rightarrow SO_4^{2-} + I^o$$

$$I^o + I^o \rightarrow I_2$$

The rate of the reaction is followed by estimating the iodide formed at different time intervals, by titrating with sodium thiosulphate ($Na_2S_2O_3$) (M.W: 158.11g/mol)using starch as indicator . The volume of thiosulphate is plotted as a function of time. The initial slope of this plot gives the initial rate of the reaction. The values of initial rates obtained can be used to calculate the total order and individual orders with respect to iodide as well as persulphate ion.

APPARATUS: Burette 50 cm³, pipettes 5, 25, and 50 cm³, stoppered bottles, 250 cm³ conical flasks, 250 cm³ standard flasks, porcelain trough, porcelain tiles.

CHEMICALS: Potassium iodide solution (0.1 N); Potassium persulphate solution (0.1 N); Acetic acid (1.0 N); Sodium thiosulphate solution (0.01 N); Starch indicator, ice cold water.

EXPERIMENTAL PROCEDURE:

1) Prepare different reaction mixtures using the volumes given in the Table 1. For example, mix 1 mL of 0.1 N acetic acid(why?) and 20 mL of 0.1 N potassium iodide in a stoppered bottle.
2) Add the required amount of water (in this case 9 mL) so that the final volume is 40 mL.
3) Add 10 mL of 0.1 N potassium persulphate solution in the stoppered bottle and start the stopwatch when half the volume of persulphate is added.
4) Stir the reaction mixture and pipette out 5 mL of it into a conical flask containing ice cold water to quench the reaction (why?) and titrate the liberated iodide against thiosulphate solution using starch as indicator.
5) Repeat this every 5 minute for at least 40 minutes.
6) Carry out similar titrations with the other solution mixtures.

RESULTS:

Table 1. Typical reaction mixtures for determining the order of the reaction between iodide and persulphate ions.

Bottle no.	Volume of acetic acid(.1 N) (mL)	Volume of 0.1 N KI solution (mL)	Volume of 0.1 N $K_2S_2O_8$ solution (mL)	Volume of water in (mL)
1	1	20	10	9
2	1	10	5	24
3	1	20	5	14
4	1	10	10	19

Treatment of Data: Record all your observations systematically as follows:

- Construct separate tables for each reaction mixture as given in Table 2.
- Plot the titre value as a function of time and evaluate the initial slopes (this can be done either numerically, or graphically or by both the processes).
- The slopes obtained at the initial periods of the reaction can be taken to be initial rates. (Why are we interested in the initial rates and not the values of rates at any given time of the reaction?)

Instantaneous Rate

For most chemical reactions, the rate of the reaction tends to decrease as time passes (Figure 17.6 “Reactant Concentration vs. Time”). As the reaction proceeds, more and more of the reactant molecules are consumed to become product, which lowers the concentration of reactant molecules. The reduction in reactant concentration results in fewer effective collisions.

The decrease in reaction rate over time means that average reaction rates do not accurately represent the actual rate of reaction at all time points. **Instantaneous reaction rates**, the rate of reaction at one instant in time, can be determined from the slope of the tangent at that point in the plot of concentration vs. time. The instantaneous rate at the start of the reaction, *t* = 0, is of particular interest in kinetics and is known as the **initial rate** of the reaction.

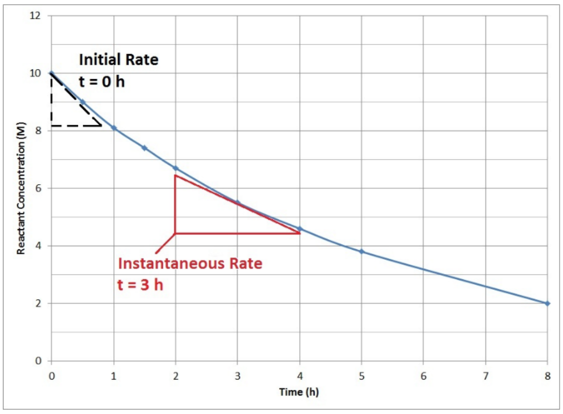
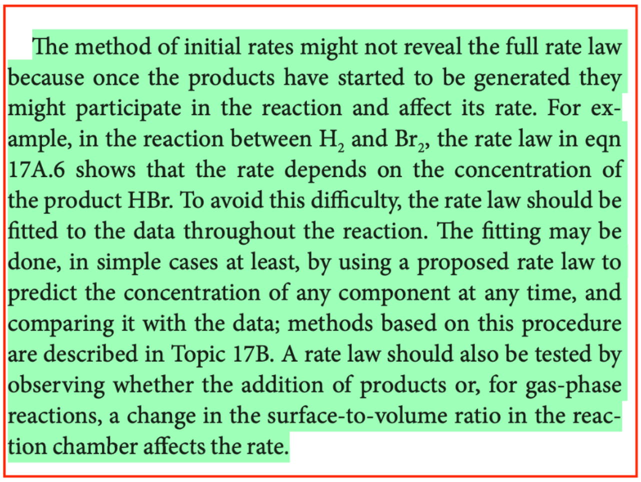


Figure 17.6 “Reactant Concentration vs. Time.” A plot of reactant concentration vs. time for a hypothetical reaction.



The method of initial rates might not reveal the full rate law because once the products have started to be generated they might participate in the reaction and affect its rate. For example, in the reaction between H_2 and Br_2, the rate law in eqn 17A.6 shows that the rate depends on the concentration of the product HBr. To avoid this difficulty, the rate law should be fitted to the data throughout the reaction. The fitting may be done, in simple cases at least, by using a proposed rate law to predict the concentration of any component at any time, and comparing it with the data; methods based on this procedure are described in Topic 17B. A rate law should also be tested by observing whether the addition of products or, for gas-phase reactions, a change in the surface-to-volume ratio in the reaction chamber affects the rate.

$$v = \frac{k_a[\mathrm{H_2}][\mathrm{Br_2}]^{3/2}}{[\mathrm{Br_2}] + k_b[\mathrm{HBr}]}$$

Table 2. Volume of thiosulphate consumed for a known aliquot of the reaction mixture at various time intervals.

Normality of thiosulphate solution taken =...............N

Bottle no.1

Time in minutes	Burette readings		Volume of thiosulphate(mL)
	Initial	Final	
0			
5			
10			
15			
20			
25			
30			
35			
40			

Bottle no.2

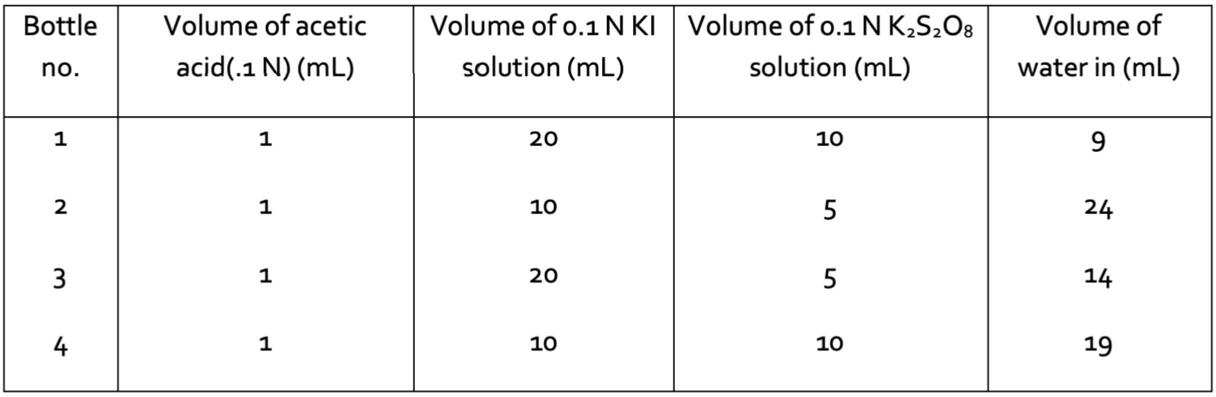
Time in minutes	Burette readings		Volume of thiosulphate(mL)
	Initial	Final	
0			
5			
10			
15			
20			
25			
30			
35			
40			

Bottle no.3

Time in minutes	Burette readings		Volume of thiosulphate(cm^3)
	Initial	Final	
0			
5			
10			
15			
20			
25			
30			
35			
40			

Bottle no. 4

Time in minutes	Burette readings		Volume of thiosulphate(cm^3)
	Initial	Final	
0			
5			
10			
15			
20			



Bottle no.	Volume of acetic acid(.1 N) (mL)	Volume of 0.1 N KI solution (mL)	Volume of 0.1 N $K_2S_2O_8$ solution (mL)	Volume of water in (mL)
1	1	20	10	9
2	1	10	5	24
3	1	20	5	14
4	1	10	10	19

Normally in a chemical reaction the rate of the reaction is proportional to the concentration of the reactants raised to the power m, where m is the order of the reaction with respect to the reactant. For the reaction between iodide and persulphate ions the rate expression can be written as

$$\text{Rate} \propto [I^-]^m [S_2O_8^{2-}]^n = k[I^-]^m [S_2O_8^{2-}]^n$$

The ratio of the initial rate values obtained for bottles 1 and 4 can be written as

$$\text{rate1/rate4} = k[100]^m [50]^n / k[50]^m [50]^n$$

(Here the volumes taken are assumed to be proportional to concentration since the total volume of the reaction mixtures is kept constant.). Therefore,

$$\text{rate1/rate4} = [2]^m \quad ; \quad \text{Or, } \log \{\text{rate1/rate4}\} = m\log 2$$

So the value of m, the order with respect to iodide can be found out. Similarly the ratio of the rates for bottles 1 and 3 can be written as

$$\{\text{rate1/rate3}\} = k[100]^m[50]^n/k[100]^m[25]^n$$

$$\text{Or } \log \{\text{rate1/rate3}\} = n\log 2$$

So the value of n, the order with respect to persulphate can be calculated. The overall order of the reaction = m+n.

Also, the ratio $\text{rate1/rate2} = k\,[100]^m[50]^n/k[50]^m[25]^n$

Or, $\log \{\text{rate1/rate2}\} = (m+n) \log 2$

Hence, the overall order of the reaction (m+n) can be calculated.

RESULTS:

Report the individual orders and overall order of the reaction. Comment how the observed reaction orders accounts for the mechanism of the reaction.

Note:

The reaction is also believed to occur in steps:

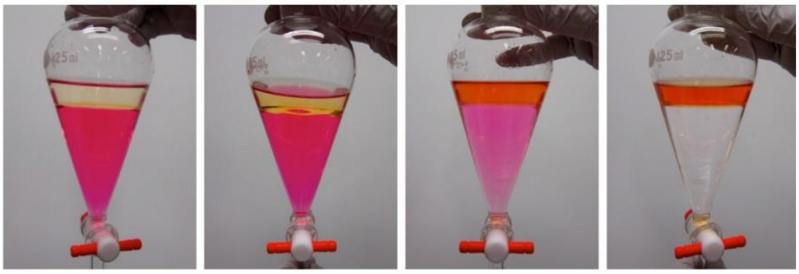
$$I^- + S_2O_8^{2-} \xrightarrow{\text{slow}} (S_2O_8I)^{3-}$$

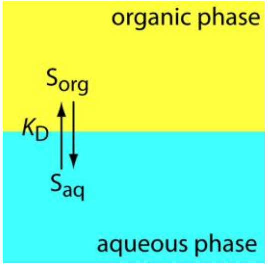
$$(S_2O_8I)^{3-} \xrightarrow{\text{fast}} 2SO_4^{2-} + I_2$$

Suggest any alternative way of studying the kinetics of the reaction between iodide and persulphate ions.

EXPERIMENT 8

STUDY OF THE DISTRIBUTION OF BENZOIC ACID BETWEEN TOLUENE & WATER





PRINCIPLE: Partition-coefficient (*P*) or distribution-coefficient (*D*) is the ratio of concentrations of a compound in a mixture of two immiscible phases at equilibrium. This ratio is therefore a measure of the difference in solubility of the compound in these two phases. The partition-coefficient generally refers to the concentration ratio of un-ionized species of compound whereas the distribution-coefficient refers to the concentration ratio of all species of the compound (ionized plus un-ionized). Most commonly, both the phases are solvents; one of the solvents is water while the second is hydrophobic. Hence the partition coefficient measures how hydrophilic ("water-loving") or hydrophobic ("water-fearing") a chemical substance is.

A liquid–liquid extraction is an important separation technique for chemical, environmental, clinical laboratories and industries. In a simple liquid–liquid extraction, the solute partitions between two immiscible phases. Because the phases are immiscible they form two layers, with the denser phase on the bottom. The solute is initially present in one of the two phases; after the extraction it is present in both phases. **Extraction efficiency**—is the percentage of solute moving from one phase to the other; and is determined by the equilibrium constant for the solute's partitioning between the phases and any other reactions involving the solute.

INTRODUCTION:

When benzoic acid (A) is distributed between toluene and water then benzoic acid is almost completely dimerised in the toluene layer but in the aqueous layer it remains mostly as single molecules without any significant dissociation. The two equilibria occur as given below:

$$A(\text{water}) \overset{K_D}{\longleftrightarrow} A\,(\text{Toluene}) \quad \text{and} \quad 2A(\text{Toluene}) \overset{K}{\longleftrightarrow} A_2\,(\text{Toluene})$$

So, $K_D = [A]_b / [A]_w$ and $K = [A_2]_b / [A]_b^2$ (where $[A]_b$ and $[A]_w$ are the concentration of benzoic acid in tolune and water repectively.

Let C_b = total molar concentration of benzoic acid in toluene layer.

& C_w = total molar concentration of benzoic acid in water layer.

So, $C_b = [A]_b + 2[A_2]_b$ & $C_w = [A]_w$

Therefore,

$C_b / C_w = ([A]_b / [A]_w) + 2([A_2]_b / [A]_w)$

$= K_D + 2 \times ([A_2]_b / [A]_b^2) \times ([A]_b^2 / [A]_w^2) \times [A]_w$

$= K_D + 2K K_D^2 C_w$ .... (1)

So, if we plot (C_b / C_w) vs. C_w, we will have a straight line with positive intercept (≡ K_D) and positive slope (≡ $2KK_D^2$). From slope and intercept we will find the K_D & K at experimental temperature.

PROCEDURE:

1. Record the room temperature before and after the experiment and average it.
2. Prepare four mixtures using following composition table

COMPOSITION TABLE

Bottle No.	Toluene (ml)	Distilled Water (ml)	Benzoic Acid (g)
I	~25	~200	1.0
II	~25	~200	1.5
III	~25	~200	2.0
IV	~25	~200	2.5

I. STANDERDISATION OF NaOH SOLUTION:

1. Prepare 100 ml ~0.1(N) oxalic acid solution by accurate weighing.

2. Prepare 500 ml 0.05(N) NaOH solution and standardise it with oxalic acid solution.

Table 1. Standardisation of NaOH solution using standard oxalicacid solution

Sl. No.	Volume of Oxalicacid (mL)	Burette reading (mL)			Average volume (mL)	Strength of NaOH solution
		Initial	Final	Difference		

II. OPTIMASATION OF OXALIC ACID USING STANDERDISED NaOH SOLUTION:

1. Titrate 5 ml Toluene layer (twice for each bottle) with NaOH solution for four bottles.
2. Titrate 50 ml aq. layer (twice for each bottle) with same NaOH solution for four bottles.

Table 2. Optimisation of benzoic acid in Tolune layer and aqueous layer using standardised NaOH solution

Sl. No.	Volume of Tolune/Aqueous layer(mL)	Burette reading (mL)			Average volume (mL)	Strength of NaOH solution
		Initial	Final	Difference		

3. Calculate C_b, C_w, (C_b / C_w) & $(\sqrt{C_b} / C_w)$ for each bottle.
4. Plot (C_b / C_w) vs. C_w and find K_D & K from the straight-line plot.

RESULTS:

K_D ....

K...................................

EXPERIMENT 9

Molecular Modelling of a Few Organic/Inorganic Molecules Using Computational Calculation

This experiment will give students an idea of visualizing molecules and show how to obtain optimized ground state structures of these molecules. A very basic theoretical knowledge will be provided before the hands on session.

EXPERIMENT 10

HOMO-LUMO Energy Optimization of a Few Organic/ Inorganic Molecules Using Computational Calculation

This experiment will discuss very brief what are the levels of theory available in Modern Quantum Chemistry Package. As such discussion needs knowledge of advanced quantum chemistry mostly we will discuss very elementary quantum chemistry. We will show how Highest Occupied Molecular Orbital (HOMO) and Lowest Unoccupied Molecular Orbital (LUMO) energy levels can be calculated for some molecules.

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