CHAPTER – 4 NOTES

ACID – BASE CHEMISTRY

SOCH BADLO BY MAK

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4.1. THE pH :

"pH is the negative logarithm of the H^+ ion concentration in a solution.**"**

Mathematically : $pH = -\log[H^+]$

OR

"pH is **logarithm (base 10)** of the reciprocal of hydrogen ion concentration."

Mathematically :

$$pH = \log\frac{1}{[H]^+}$$

$$H^+ + H_2O \rightarrow H_3O^+$$

$$pH = \log\frac{1}{[H3O]^+}$$

$$pH = -\log[H_3O]^+$$

pOH: "The concentration of hydroxide ions in a solution can be expressed in terms of pOH."

Mathematically :

$$pOH = -\log[OH]^-$$

Introduction of pH:
• Introduced in 1909 by Danish biochemist **S.P.L. Sorensen**.
• Symbol **pH**:
Letter **p** from German word *potenz* (meaning power or exponent of 10).
In Latin, "pH" is said to mean *pondus hydrogenii* (weight of hydrogen).

Historical context:
1909: Sørensen published a paper in Biochem Z discussing effect of H^+ ions on enzyme activity.
• He coined the term pH and defined it as:

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

1924: Sørensen updated the definition, stating pH is a function of the activity (effective concentration) of H^+ ions, not just concentration.

$$\mathbf{pH = -\log a\{H^+\}}$$

where $a\{H^+\}$ = activity of hydrogen ions.

Relation between pOH and pH:

$$pKw = pH + pOH$$

At 25° $\quad pH + pOH = 14$

4.2. THE pH SCALE:

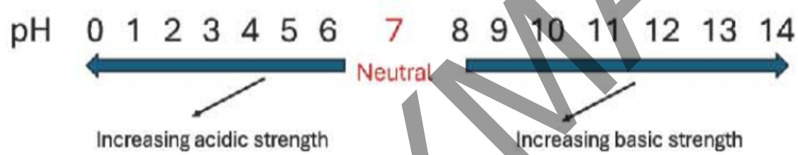
Purpose: Numerical scale showing acidic or alkaline strength of a solution.

Range: 0 to 14.

Acids: pH < 7.

Bases/Alkalis: pH > 7.

Neutral: pH = 7.



Interpretation:

Lower pH → more acidic.

Higher pH → more alkaline.

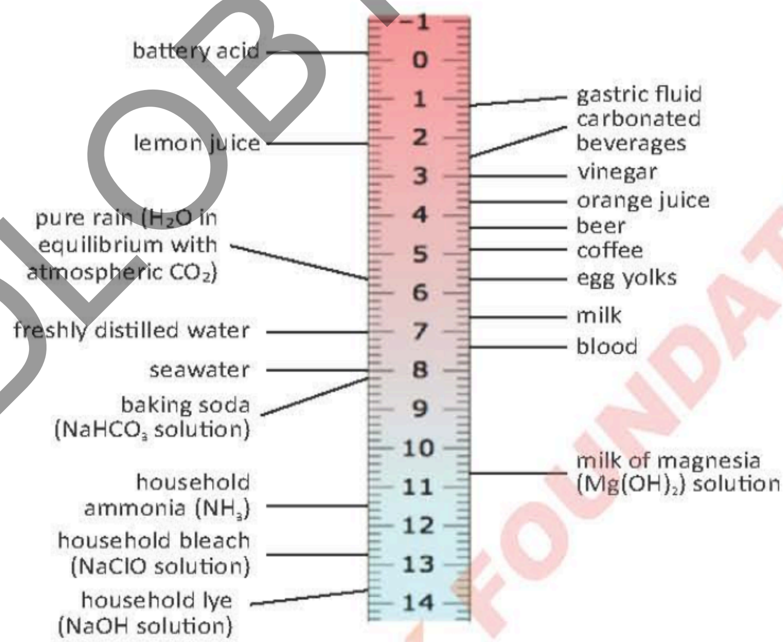
Example:

pH 0 → $[H_3O^+] = 1$ M.

pH 14 → $[OH^-] = 1$ M.

Typical range of $[H^+]$ in solutions:

Between 1 M (pH 0) and 10^{-14} M (pH 14).



Example – 4.1:

Find the pH of a solution of 0.002M of HCl.

Data: $[H^+] = 2 \times 10^{-3}$ M

Required : pH = ?

Solution : As we know,

$$pH = -\log [H^+]$$

$$pH = -\log (2 \times 10^{-3})$$

$$\mathbf{pH = 2.70}$$

<u>Example – 4.2:</u>

If moist soil has a pH of 7.84, what is the H^+ concentration of the soil solution?

Data : pH = 7.84

Required : H^+ concentration of soil solution = $[H^+]$ = ?

Solution : As we know

$$pH = -\log [H^+]$$

$$7.84 = -\log [H^+]$$

$$-7.84 \text{ (antilog)} = [H^+]$$

$$\mathbf{[H^+] = 1.45 \times 10^{-8}}$$

<u>CONCEPT ASSESSMENT EXERCISE – 4.1</u>

Q1). What is the pH of a solution of 2g pure H_3PO_4 per dm^3 of solution?

Data : Mass of H_3PO_4 = 2g

Volume of solution = 1 dm^3

Required : pH = ?

Solution : Molar mass of H_3PO_4 = 3 + 31 + (16) 4

Molar mass of H_3PO_4 = 98 g/mol

Now, to find moles:

Moles in solution = $\frac{mass}{molar\ mass}$

Moles in solution = $\frac{2}{98}$

Moles in solution = 0.0204 mol

Reaction : $H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$

	Initial	change	equilibrium
H_3PO_4	0.0204	$-x$	$0.0204 - x$
H^+	0	$+x$	x
$H_2PO_4^-$	0	$+x$	x

$$Ka = \frac{[H^+][H2PO4]^-}{[H3PO4]}$$

$$Ka = \frac{x \cdot x}{0.0204 - x}$$

$$Ka = \frac{x^2}{0.0204}$$ ∵ x is small compared to 0.0204

∵ Ka = 7.5 x 10^{-3}

$$7.5 \times 10^{-3} = \frac{x^2}{0.0204}$$

$7.5 \times 10^{-3} (0.0204) = x^2$

$x^2 = 1.53 \times 10^{-4}$

$$\sqrt{x^2} = \sqrt{1.53 \times 10^{-4}}$$

$x = 0.0124$

As we know,

pH = – log [H^+]

pH = – log (0.0124)

pH = 1.91

Molarity = 0.0204M

Q2). Calculate the concentration of hydrogen ion (H^+) in a solution of sulphuric acid having pH of 1.5.

Solution). pH = – log [H^+]

1.5 = – log [H^+]

– 1.5 (antilog) = [H^+]

[H^+] = 0.031 mol dm^{-3}

4.3 IONIC PRODUCT OF WATER AND CALCULATION OF pH AND pOH.

The product of the concentration of $H^+$and OH^-ions in pure water at room temperature (298 K) is:

$$Kw = [H]^+[OH]^-$$

$$Kw = [H_3O]^+[OH]^-$$

• Kw is the ionic product or dissociation constant of water.

Unit of Kw: $mol^2\ dm^{-6}$

• In pure water, the concentrations of H^+ and OH^- ions = 1×10^{-7}

• Value of Kw at room temperature: 1×10^{-14}

• pKw formula = pKw = – log Kw

$$pKw = -\log(1 \times 10^{-14})$$

$$pKw = 14$$

Relation of Kw with pKa and pKb :

$$pKa + pKb = pKw = 14 \text{ (at 298K)}$$

Relation of pH and pOH:

$$pH + pOH = pKw = 14 \text{ (at 298 K)}$$

• If the pKa value of an acid is known, the pKb value of its conjugate base can be found.

<u>Example – 4.3</u>

If the concentration of NaOH in a solution is 2.5 × 10^{-4} M, what is the concentration of H_3O^+ ions at 25°C?

Data : Concentration of NaOH = 2.5 x 10^{-4}M at 25°C

Required : $[H_3O]^+$ = ?

Solution : As we know ,

$$Kw = [H_3O^+][OH]^{--}$$

$$\frac{Kw}{[OH]} = [H_3O]^+$$

$$[H_3O]^+ = \frac{1 \times 10^{-14}}{2.5 \times 10^{-4}}$$

$$\mathbf{[H_3O]^+ = 4 \times 10^{-19}\ M}$$

<u>Example – 4.4</u>

Calculate the pH value of 0.001 mol dm^{-3} solution of NaOH at 25°C.

Data : $[OH^-] = 0.001\ mol\ dm^{-3}$

Required : pH = ?

Solution : As we know,

$$Kw = [H_3O^+][OH^-]$$

$$\frac{Kw}{[OH]} = [H_3O^+]$$

$$[H_3O^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-13}}$$

$$[H_3O^+] = 1 \times 10^{-11}$$

$$pH = -\log [H_3O^+]$$

$$pH = -\log (1 \times 10^{-11})$$

$$\mathbf{pH = 11}$$

<u>Example – 4.5</u>

Calculate the pH of 5 x 10^{-5}M of solution of NaOH

Solution) As we know,

$$pOH = -\log[OH]$$

$$pOH = -\log (5 \times 10^{-5})$$

$$pOH = 4.30$$

To find pH,

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

pH = 14 – 4.30

pH = 9.70

<u>CONCEPT ASSESSMENT EXERCISE – 4.2</u>

Q1). The concentration of hydroxide ion in a given solution of slaked lime [$Ca(OH)_2$] is 0.001M. Calculate the concentration of hydrogen ion in it.

Data : [OH^-] = 0.001M

Required : [H^+] = ?

Solution : As we know,

Kw = [OH^-][H^+]

$$\frac{Kw}{[OH]} = [H^+]$$

$$[H^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-3}}$$

[H^+] = 1 x 10^{-11}

Q2). An aqueous solution contains 2 x 10^{-3} M of hydrogen ions [H^+]. Calculate pOH of this solution.

Data : [H^+] = 2 x 10^{-3} M

Required : pOH = ?

Solution : As we know,

pH = – log [H^+]

pH = – log (2 x 10^{-3})

pH = 2.69

To find pOH :

pH + pOH = 14

2.69 + pOH = 14

pOH = 14 – 2.69 = **11.31M**

4.4 pH TITRATION CURVES

• To find the concentration of an unknown solution, **titration** is used.

Titration: "Titration is a method where a solution of known concentration (titrant) is added from a burette into a solution of unknown concentration (analyte) in a conical flask."

Titrant : In burette, a known concentration called titrant.

Analyte : In conical flask unknown concentration called analyte.

Indicator : An indicator is used to show when the reaction is complete (endpoint).

Endpoint : The endpoint is when the number of moles of acid equals the moles of base this is the equivalence point.

Equivalence point : Mid point in the graph indicates that the amount of titrant is equal to the amount of analyte is called equivalence point.

Neutralization point: when OH ions react with H ions.

• There are different pH titration curves depending on the strength of the acids and bases used.

a. Strong Acid and Strong Alkali pH Titration Curve :

Analyte: Hydrochloric acid (HCl), 1.0 mol dm^{-3}

Titrant: Sodium hydroxide (NaOH), 1.0 mol dm^{-3}

Indicator : Phenolphthalein (pH 8.3–10) or Bromothymol blue (pH 6.0–7.6). Colour changes are sharp due to the steep vertical section of the curve.

Titration Process:

• Acid (HCl) is placed in a conical flask.

• Base (NaOH) is added from a burette in small volumes.

• Initially, solution has only H^+ ions, so pH is low (1– 2).

• As NaOH is added, OH^- ions neutralize H^+ ions, pH rises slowly.

• Near equivalence point, a small addition of NaOH causes a sharp rise in pH.

Reaction:

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

(H^+ from acid reacts with OH^- from base to form water)

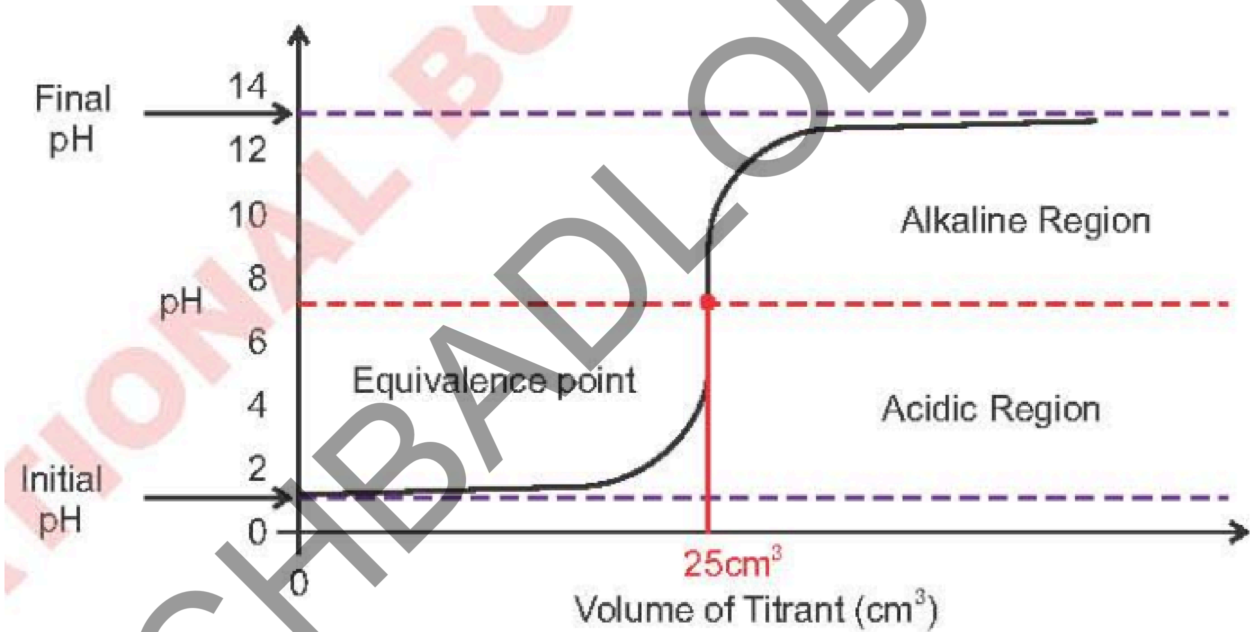
Equivalence Point:

- Volume: 25 cm^3 of NaOH to neutralize 25 cm^3 of HCl.
- pH at equivalence point: 7 (neutral, because strong acid + strong base).

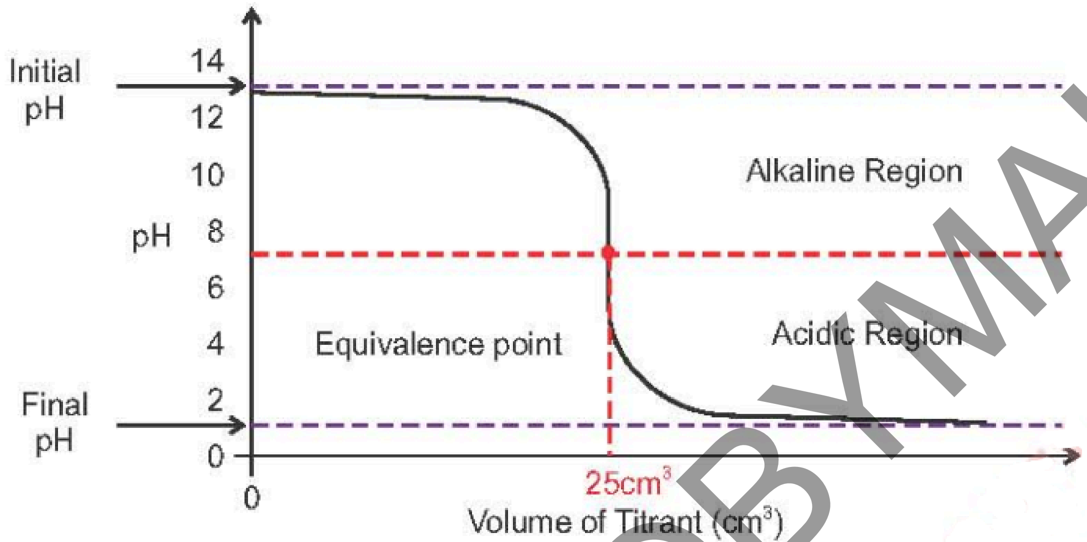
Exceeding Equivalence Point:

- Extra NaOH increases OH^- concentration.
- pH rises above 7 into alkaline range (13–14).

pH titration curve of 1 mol dm^{-3} HCl (25 cm^3) with NaOH :



pH titration curve of 1 mol dm^{-3} NaOH (25 cm^3) with HCl:



pH Titration Curve: HCl added to NaOH

Shape: Same as NaOH titrated with HCl, flipped.

Initial pH: High (13–14).

Final pH: Low (1).

Equivalence point: pH = 7.

b. Strong Acid and Weak Alkali pH Titration Curve

Analyte: Strong acid (HCl) in conical flask (unknown)

Titrant: Weak alkali (NH_3) in burrette (known)

Indicator: Methyl orange (changes from red → yellow around acidic pH)

Titration process:

- Initially there will be only H^+ ion present in conical flask, pH (1– 2) HCl
- As weak alkali is added, pH increases slightly because H^+ ions react with NH_3
- pH change is gradual until near equivalence point

Reaction:

$$HCl + NH_3 \rightarrow NH_4^+ + Cl^-$$

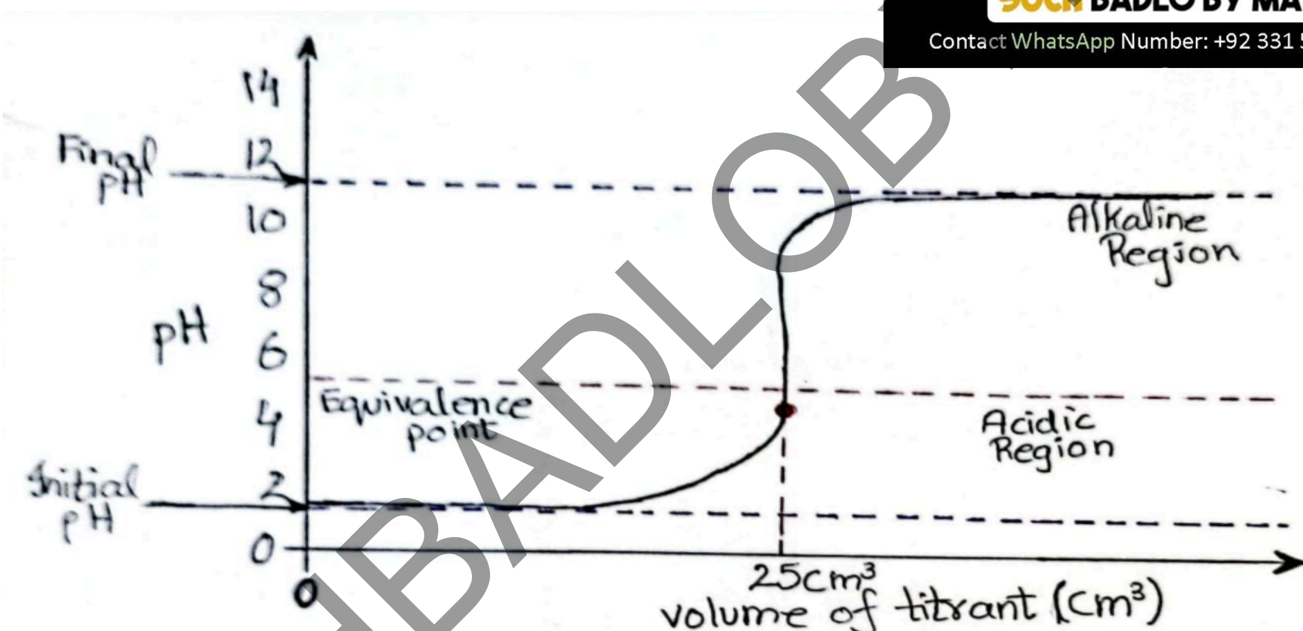
The change in pH is not that much until volume added reaches an equivalence point.

Equivalence point: pH 5.5 (acidic) because NH_4^+ is a weak acid

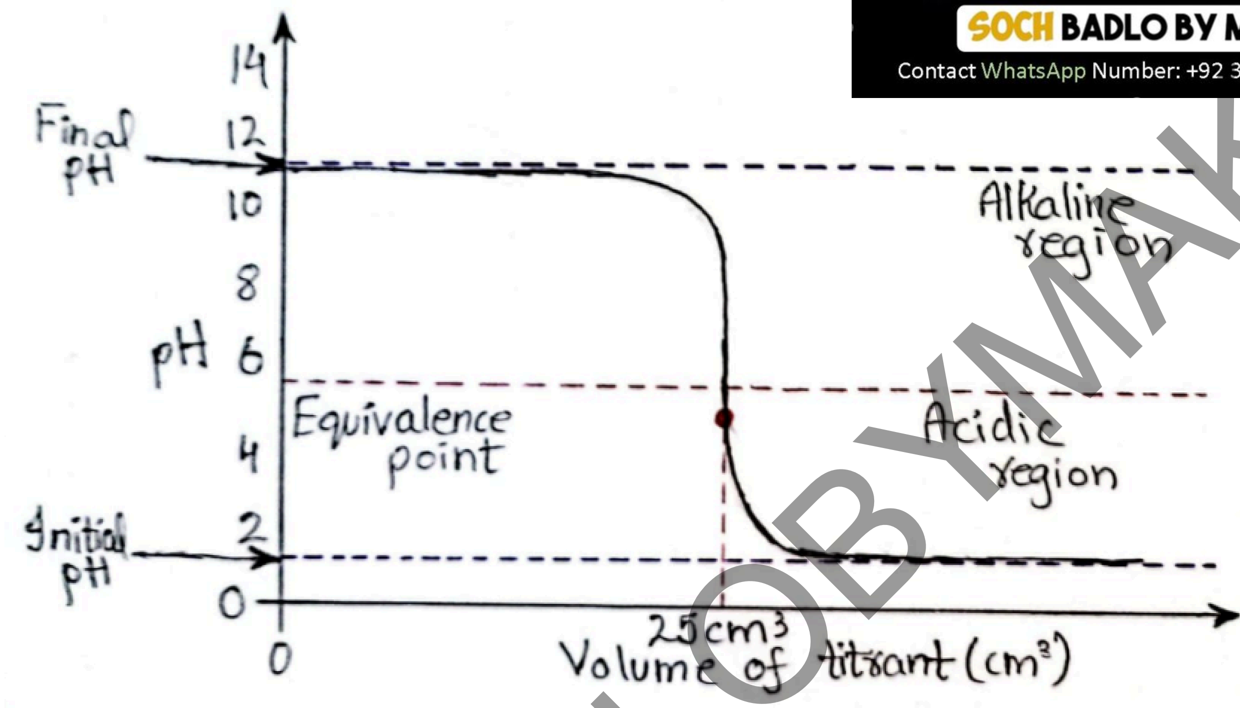
$$H^+ + NH_3 \rightarrow NH_4^+$$

Exceeding Equivalence point: pH rises above 7, but not as high as strong alkali because NH_3 is weak.

pH titration curve of 1 mol dm^{-3} HCl (25 cm^3) with NH_3:



pH titration curve of 1 mol dm^{-3} NH_3 (25 cm^3) with HCl :



- Here, weak alkali is taken at conical flask, and strong acid at burrette.
- pH titration curve for strong acid added to weak base has the same shape.
- Equivalence point will be 5.5 pH

c. Weak acid and strong alkali pH titration curve:

Analyte: Weak acid (CH_3COOH) in conical flask (unknown)

Titrant: Strong alkali (NaOH) in burrette (known)

Indicator: Phenolphthalein (colorless → pink in basic range)

Titration process:

- Initial pH: 2–3 (weak acid in conical flask)
- As NaOH is added, pH rises slowly because H^+ ions react with OH^- to form water
- Near equivalence point, pH increases steeply

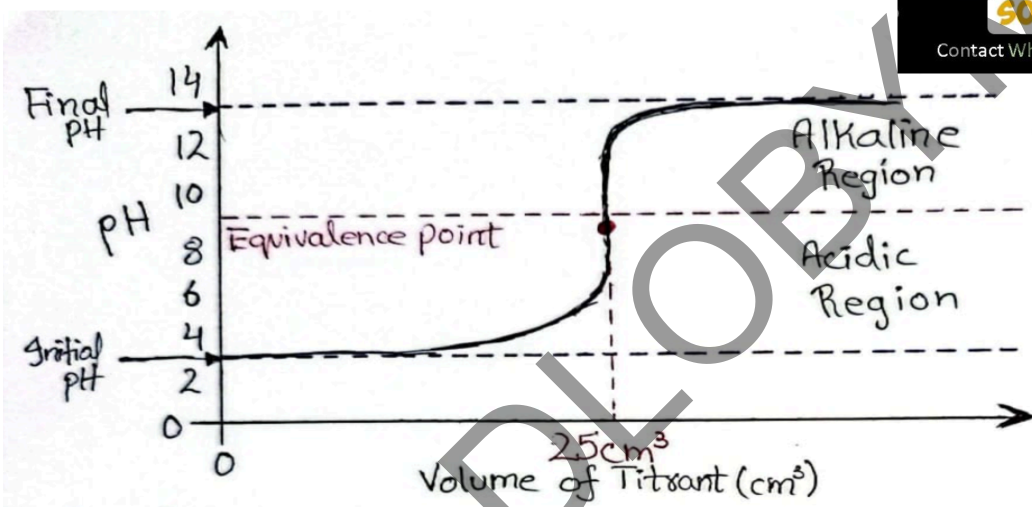
Reaction:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

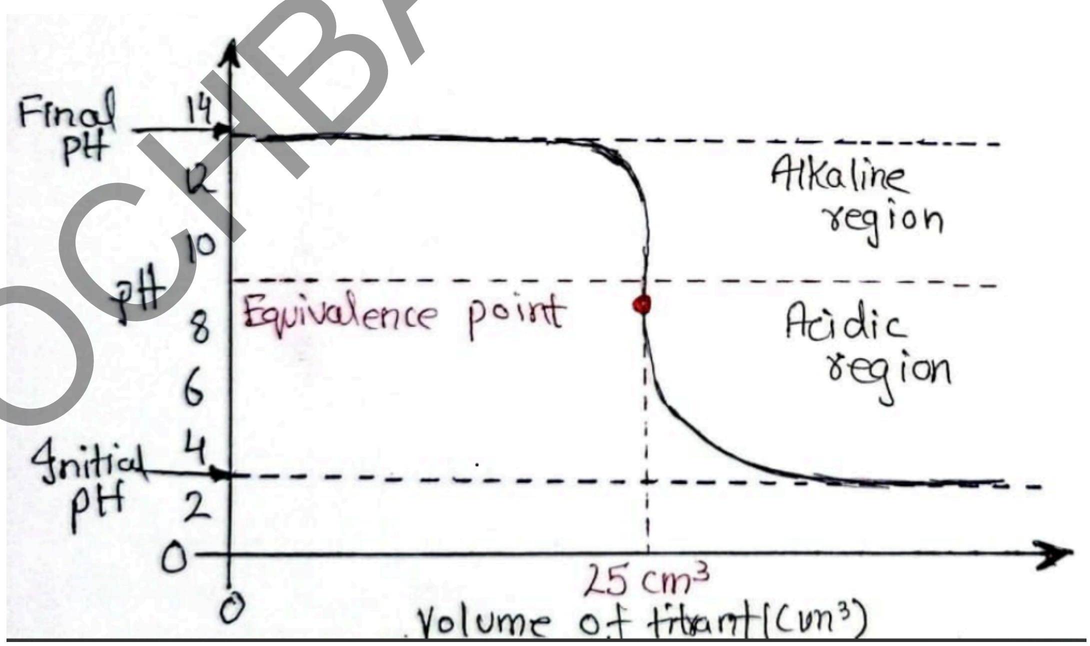
Equivalence point: pH 9 (slightly basic) because CH_3COO^- is a relatively strong base

Exceeding Equivalence point: pH increases up to 13–14 (excess NaOH present)

pH titration curve of a strong base added to weak acid :



titration curve of a weak acid added to a strong base:



• Here, weak acid is taken in burrette and strong alkali is taken in conical flask.

• pH titration curve for a weak acid added to a strong alkali has the same shape.

• The equivalence point will be 9 pH.

d. Weak acid and weak alkali pH titration curve :

Analyte: Weak acid (CH_3COOH)

Titrant: Weak alkali (NH_3)

Titration process:

• Initial pH: 2–3 (H^+ ions from weak acid present)

• As weak alkali is added, pH increases gradually

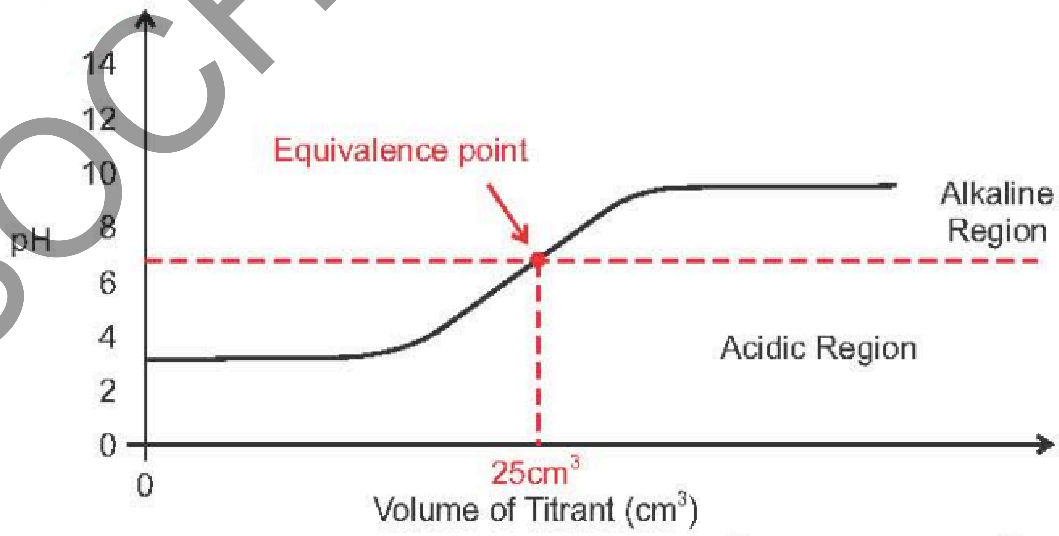
• No steep vertical section in the curve

Reaction:

$$CH_3COOH + NH_3 \rightarrow CH_3COONH_4$$

Equivalence point: pH depends on strengths of acid and base; shown as a **"point of inflexion,"** not a sharp jump.

Exceeding Equivalence point: pH rises slowly; final pH < strong alkali range



CONCEPT ASSESSMENT EXERCISE – 4.3

Q1). In a titration it is found that 25cm³ of 0.1M solution of NaOH is neutralized with 19 cm³ of HCl of unknown concentration, calculate concentration of given HCl solution:

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

Data :

M_1 = 0.1M (NaOH)

V_1 = 25 cm^3 (NaOH)

V_2 = 19 cm^3 (HCl)

Required : Concentration of HCl = M_2 = ?

Solution : According to formula:

$M_1V_1 = M_2V_2$ (for 1:1 reaction)

$(0.1)(25) = M_2\,(19)$

$2.5 = M_2\,(19)$

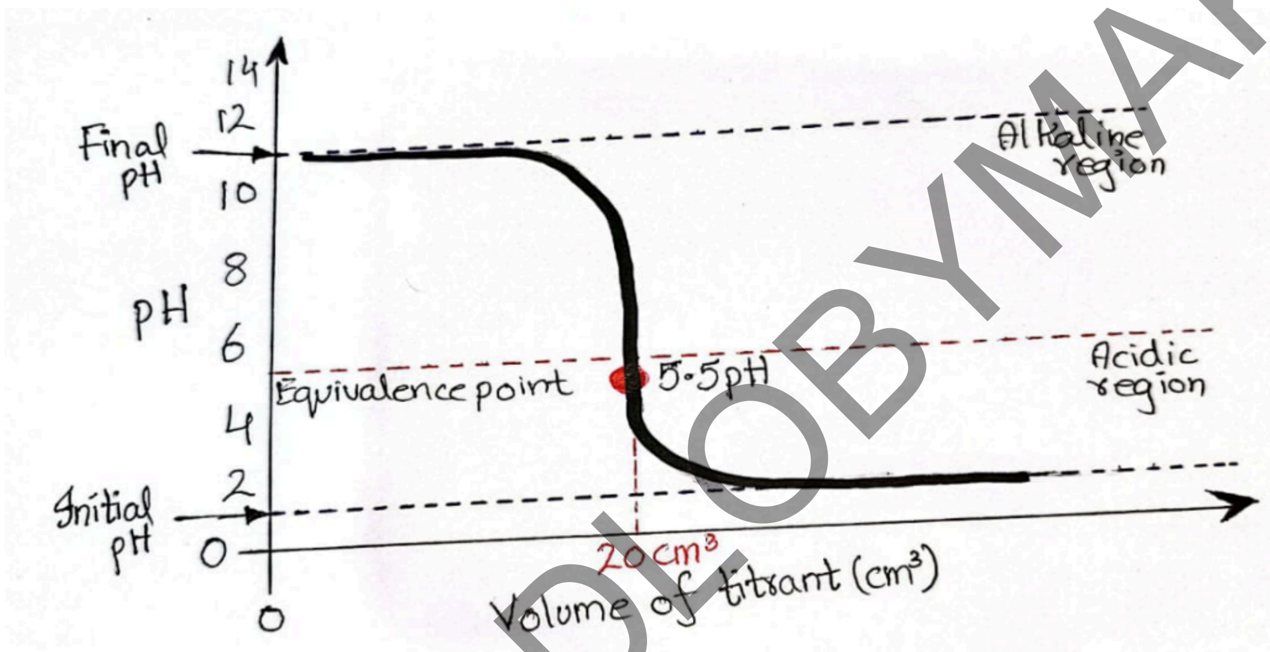
$\frac{2.5}{19} = M_2$

M_2 = 0.1315M

Q2). Draw pH titration curve when 20cm^3 of HCl from burette is added to 20 cm^3 of aqueous solution of NH_3 present in conical flask.

Given: Strong acid in burette (HCl) and Weak base in conical flask (NH_3)

pH titration curve when HCl is added to NH_3 (20 cm^3):



- **'Equivalence point'** will be **5.5 pH**
- **'Initial pH'** will be between **1– 2 pH**
- **'Final pH'** will be between **10 – 12 pH**

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1. Multiple Choice Questions (MCQs)

i. The value of the ionic product of water
a) depends on volume of water
b) depends on temperature
c) changes by adding acid or alkali
d) always remains constant

ii. A base when dissolved in water yields a solution with a hydroxyl ion concentration of 0.05 mol dm^{-3}. The solution is
a) Basic
b) Acidic
c) Neutral
d) either b or c

iii. pH scale was introduced by.

a) Arrhenius
b) Sorensen
d) Lewis
d) Lowry

iv. pH of solution is defined by expression

a) $\log[H^+]$
b) $\log\left[\frac{1}{H^+}\right]$
c) $\frac{1}{\log[H^+]}$
d) $\frac{1}{-\log[H^+]}$

v. The pH of a 10^{-3} M HCl solution at 25°C if it is diluted 1000 times, will be

a) 3
b) zero
c) 5.98
d) 6.02

vi. How many dm^3 of water must be added to 1 dm^3 an aqueous solution of HCl with a pH of 1 to create an aqueous solution with pH of 2?

a) 0.1 dm^3
b) 0.9 dm^3
c) 2.0 dm^3
d) 9.0 dm^3

vii. What is the approximate pH of a 1×10^{-3} M NaOH solution?

a) 3
b) 11
c) 7
d) 1×10^{-11}

viii. Calculate the pOH of a solution at 25°C that contains 1×10^{-10} M of hydronium ions, i.e. H_3O^+

a) 4.0
b) 9.0
c) 1.0
d) 7.0

ix. The pH value of a 10 M solution of HCl is

a) less than 0
b) equal to 0
c) equal to 1
d) equal to 2

x. Which of the following has the highest pH?

a) $\frac{M}{4}\ KOH$
b) $\frac{M}{4}\ NaOH$
c) $\frac{M}{4}\ NH_4OH$
d) $\frac{M}{4}\ Ca(OH)_2$

xi. Which of the following statements are correct?

(i) $Kw = [H^+][OH^-] = 10^{-14}\ mol^2dm^{-6}$ at 298K
(ii) At 298K $[H^+] = [OH^-] = 10^{-7}$
(iii) Kw does not depend upon temperature
(iv) Molarity of pure water = 55.55 M

a) (i), (ii) and (iii)
b) (i), (ii) and (iv)
c) (i) and (iv)
d) (ii) and (iii)

EXERCISE SHORT ANSWER QUESTIONS.

Q1). Calculate H^+ ion concentration of a solution prepared by dissolving 4g of NaOH (Atomic weight if Na = 23 amu) in 1000 cm^3 of solution?

Data :

Mass of NaOH = 4g

Molar Mass of NaOH = 40 g/mol

Volume = 1000 cm^3 = 1 dm^3

Required : $[H^+]$

Solution : According to formula:

$$\text{Moles} = \frac{mass}{molar\ mass}$$

$$\text{Moles} = \frac{4}{40}$$

Moles = 0.1 mol

Now, $pOH = -\log [OH^-]$

$pOH = -\log (0.1)$

$pOH = 1$

$\because$ pH + pOH = 14 $pH = 14 - pOH$

$pH = 14 - 1$

$pH = 13$

Now to find $[H^+]$: As we know,

$pH = -\log [H^+]$

$13 = -\log [H^+]$

-13 (antilog) = $[H^+]$

$\mathbf{[H^+] = 1 \times 10^{-13}\ M}$

Q2). Calculate the pH of 0.005 molar solution of H_2SO_4.

Data : $[H^+] = 0.005$ M

Required : pH = ?

Solution : According to formula.

$pH = -\log [H^+]$

$pH = -\log [0.005]$

pH = 2.30

Q3). Calculate the pH of the following compounds :

(i) 10^{-4} M KOH

Sol). $pOH = -\log [OH^-]$

$pOH = -\log (10^{-4})$

$pOH = 4$

Now to find pH,

$pH + pOH = 14$

$pH = 14 - pOH$

$pH = 14 - 4$

pH = 10

(ii) 10^{-10} M HCl

Sol). $pH = -\log [H^+]$

$pH = -\log (10^{-10})$

pH = 10

(iii) 10^{-10} M KOH

Sol). $pOH = -\log [OH^-]$

$pOH = -\log (10^{-10})$

$pOH = 10$

Now to find pH,

$pH + pOH = 14$

$pH = 14 - pOH$

$pH = 14 - 10$

pH = 4

(iv) 10^{-4} M HCl

Sol). $pH = -\log [H^+]$

$pH = -\log (10^{-4})$

pH = 4

Q4). 100cm^3 of 0.04 M HCl aqueous solution is mixed with 100cm^3 of 0.02M NaOH solution. Calculate the pH of the resulting solution.

Data : Volume of HCl = V_1 = 100 cm^3 = 0.100 L

Molarity of HCl = M_1 = 0.04 M

Volume of NaOH = V_2 = 100 cm^3 = 0.100 L ,

Molarity of NaOH = M_2 = 0.02 M

Required: pH of resulting solution

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

As we know,

$M_1V_1 = (0.04)(0.1) = 0.004$

$M_2V_2 = (0.02)(0.1) = 0.002$

Excess HCl = 0.004 – 0.002 = 0.002 mol

Total volume = 0.1 + 0.1 = 0.2 L

$$[H^+] = \frac{excess\ moles}{total\ volume}$$

$$[H^+] = \frac{0.002}{0.2}$$

$[H^+] = 0.01$ M

According to pH formula :

$pH = -\log [H^+]$

$pH = -\log (0.01)$

pH = 2

Q5). Equal volume of three acid solutions of pH 3, 4, and 5 are mixed in a vessel. What will be the H^+ ion concentration in the mixture?

Data : pH = 3,4,5

Required: $[H^+]$ = ?

Solution : According to formula

$pH = -\log [H^+]$

Put pH = 3 $\quad -3\ (\text{antilog}) = [H^+]$

$[H^+] = 0.001$ M

Put pH = 4 $\quad -4\ (\text{antilog}) = [H^+]$

$[H^+] = 0.0001$ M

Put pH = 5 $\quad$ – 5 (antilog) = $[H^+]$

$[H^+] = 0.00001$ M

Taking average,

$$[H^+] = \frac{0.001 + 0.0001 + 0.00001}{3}$$

$$[H^+] = \frac{0.00111}{3}$$

$$\mathbf{[H^+] = 3.7 \times 10^{-4}}$$

Q6). A 20 cm^3 sample of 0.2 mol dm^{-3} $NH_{3(aq)}$ was titrated with 0.1 $moldm^{-3}$ HCl. On the following axes sketch how the pH changes during this titration. Mark clearly where the end point occurs.

Ans). $NH_3 + HCl \rightarrow NH_4Cl$

Data : $M_1 = NH_3 = 0.2$

$V_1 = NH_3 = 20\ cm^3$

$M_2 = HCl = 0.1$

$V_2 = HCl = ?$

According to formula,

$M_1V_1 = M_2V_2$ (for 1:1 reaction)

By putting values,

$(0.2)(20) = (0.1)V_2$

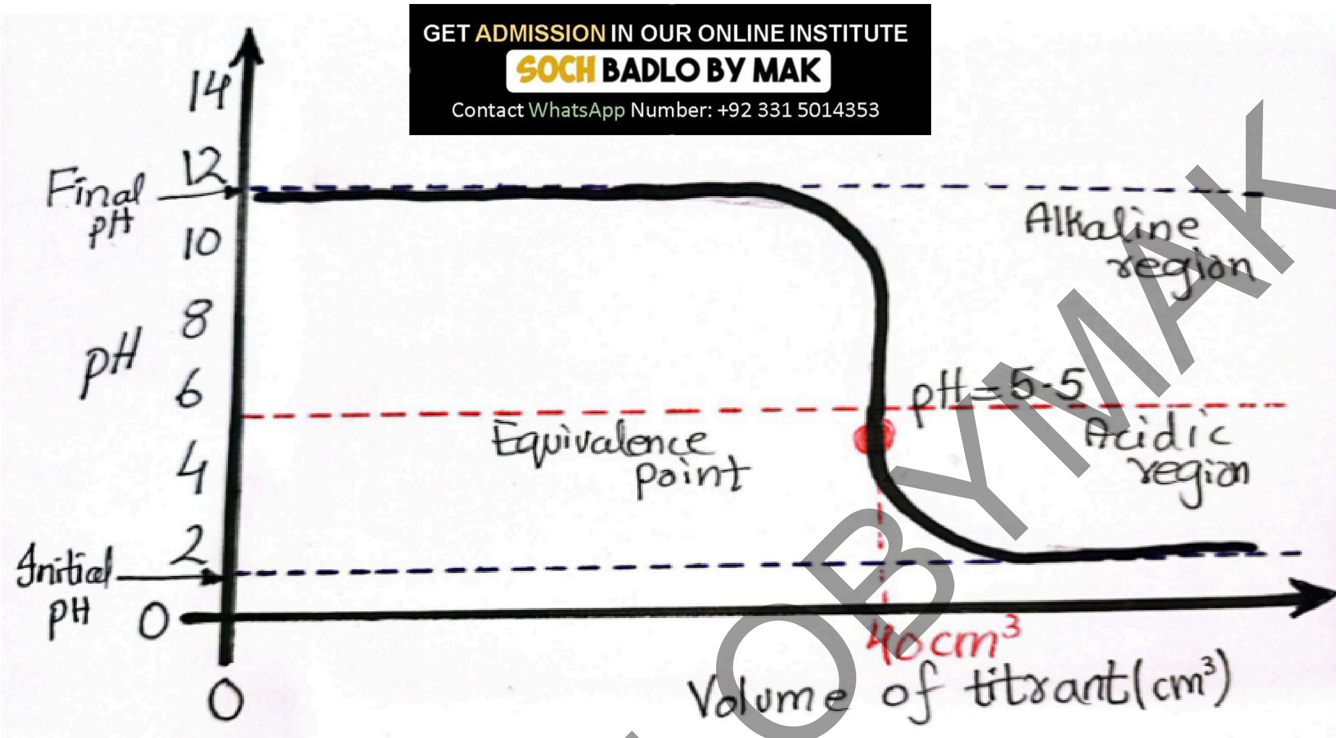
$4 = V_2\ (0.1)$

$$\frac{4}{0.1} = V_2$$

$V_2 = 40cm^3$

- So, the volume of added titrant will be $40cm^3$.
- At Burette strong acid HCl (titrant) is taken
- At Conical flask weak base NH_3 (analyte) is taken.

pH titration curve for strong acid(HCl) added to weak base(NH_3) 40cm^3:



- **Equivalence point** will be **5.5 pH**
- **Initial pH** will be between **1–2.**
- **Final pH** will be between **10–12.**

SLO BASED QUESTIONS:

Q1). Prove that pH + pOH = 14 at 25°C.

Sol). Ionization constant of water:

$[H^+][OH^-] = Kw$ $\because$ $Kw = 10^{-14}$

$[H^+][OH^-] = 10^{-14}$

Taking log on both sides:

$\log[H^+][OH^-] = \log 10^{-14}$ $\because$ $\log a.b = \log a + \log b$

$\log[H^+] + \log[OH^-] = -14 \log 10$ $\because$ $\log x^n = n \log x$

$\log[H^+] + \log[OH^-] = -14 \times 1$ $\because$ $\log 10 = 1$

$\log[H^+] + \log[OH^-] = -14$

Multiplying both sides with –1

$(-\log[H^+]) + (--\log[OH^-]) = -(--14)$ $\because$ $pH = -\log[H^+]$ and $pOH = -\log[OH^-]$

pH + pOH = 14

Hence proved.

Q2). Prove that pKa + pKb = 14.

Sol). As we know,

$Ka \times Kb = Kw$ $\because Kw = 10^{-14}$

$Ka \times Kb = 10^{-14}$

Taking log on both sides:

$\log Ka \times Kb = \log 10^{-14}$ $\because \log a \times b = \log a + \log b$

$\log Ka + \log Kb = -14 \log 10$

$\log Ka + \log Kb = -14$

Multiplying both sides with –1.

$(--\log Ka) + (--\log Kb) = -(--14)$ $\because pKa = -\log Ka$ and $pKb = -\log Kb$

pKa + pKb = 14

Hence proved.

Q3). Prove that pKa + pKb = pKw.

Sol). As we know,

$$Ka \times Kb = Kw$$

Applying log on both sides,

$$\log Ka \times Kb = \log Kw \qquad \because \log a \times b = \log a + \log b$$

$$\log Ka + \log Kb = \log Kw$$

Multiplying both sides by –1,

$$(-- \log Ka) + (-- \log Kb) = - \log Kw$$

$\because$ pKa = – log Ka, pKb = – log Kb, pKw = – log Kw

$$\mathbf{pKa + pKb = pKw}$$

Hence proved.

Q4). prove that Ka $\propto \frac{1}{Kb}$

Sol). Reaction for an acid:

$$HA \rightleftharpoons H^+ + A^-$$

$$Ka = \frac{[H^+][A^-]}{[HA]} \text{ ______(i)}$$

Reaction of a base :

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

$$Kc = \frac{[HA][OH^-]}{[A^-][H2O]}$$

$$Kc[H_2O] = \frac{[HA][OH^-]}{[A^-]} \text{ ______(ii)}$$

Multiplying equation (i) and (ii)

$$\because Kc[H_2O] = Kb$$

$$Ka \times Kb = \frac{[H^+][A^-]}{[HA]} \times \frac{[HA][OH^-]}{[A^-]}$$

$Ka \times Kb = [H^+][OH^-]$ $\because Kw = [H^+][OH^-]$

$Ka \times Kb = Kw$

$Ka = \frac{Kw}{Kb}$ $\because Kw = \text{constant}$

$$\textbf{Ka} \propto \frac{1}{Kb}$$

Hence proved.

Q5). What is the purpose of pH titration curves?

Ans). The steepest part of the curve shows where the acid and base have completely neutralized each other. This helps us know exactly how much titrant is needed.

Indicators change color at specific pH ranges.

By looking at the pH at the equivalence point on the curve, we can choose the best indicator.

Curve shape tells us if the acid/base is strong or weak.

→ Strong acid + strong base: sharp vertical jump at pH 7

→ Weak acid + strong base: gradual rise, equivalence point > 7

→ Strong acid + weak base: equivalence point < 7

In weak acid/base titrations, curves show buffer zones where pH changes very slowly.

Q6). Phenolphthalein is pink in alkaline solution and colourless in acidic solution. During a titration, the conical flask contains a neutralised solution at the end point. Why does the phenolphthalein indicator become colourless instead of remaining pink?

Ans). At the end point, the solution in the conical flask is only just neutralised. This means there is no excess alkali present to maintain the pink colour of phenolphthalein. Since phenolphthalein is colourless in acidic and neutral solutions, its colour disappears when the last drop of acid neutralises the base.

Q7). Differentiate between End point and Equivalence point.

Difference :

End point	Equivalence point
"**The end point** is the stage where the **indicator changes colour**, usually when there is a slight **excess of OH^- ions** (so the pH jumps just **above 7**, into the **basic range**)."	"**The equivalence point** is the stage on the titration curve **where the acid and base have completely neutralised each other**."
On the pH titration curve, the end point **corresponds to the point where the chosen indicator undergoes its colour change**.	On the pH titration curve, **the equivalence point corresponds to the sharp inflection.**
The end point is the **final observed pH of the solution, detected by the indicator's colour change**.	The pH at the equivalence point **depends on the type of acid–base titration: ≈7 (strong acid + strong base), >7 (weak acid + strong base), <7 (strong acid + weak base).**
The end point is a value that **indicates neutralisation is achieved to a visible level.**	The equivalence point is the **true chemical neutralisation point where the moles of H^+ = moles of OH^-.**

Q8). What is the use of a pipette?

Ans). A pipette is used to accurately measure and transfer a fixed volume of a liquid, usually the solution of known concentration, into the conical flask during a titration. It ensures precision and reliability in quantitative analysis.

Q9). Why is an indicator used?

Ans). An indicator is used in titration to show the point at which neutralisation has occurred. Since acids and bases are colourless in dilute solutions, an indicator helps to visually detect the end point by changing colour, ensuring accuracy in determining the unknown concentration.

Q10). What is the use of a burette and conical flask?

Ans). Burette : "A burette is used to deliver a measured volume of the solution of known concentration (the titrant) into the conical flask.
It allows controlled addition, drop by drop, until the end point is reached."

Conical flask : "A conical flask is used to contain a fixed volume of the solution of unknown concentration (the analyte). The solution from the burette reacts with the analyte until neutralisation is achieved."

Q11). Explain why the product of $[H^+]$ and $[OH^-]$ is always constant at a given temperature.

Ans). The product of $[H^+]$ and $[OH^-]$ is always constant at a given temperature because the self-ionisation of water is an equilibrium process. At constant temperature, the equilibrium constant (Kw) remains fixed, so $[H^+][OH^-] = Kw$.

Q12). Why does Kw increase with temperature?

Ans). The dissociation of water is an endothermic process, so increasing temperature shifts equilibrium towards more ionisation, hence Kw increases.

Q13). What is a neutralization point?

Ans). "The neutralisation point is the stage in a titration at which an acid has completely reacted with a base (or vice versa) to form salt and water."
At the neutralisation point, the number of moles of **H^+ ions = OH^- ions.**