A background image of various laboratory glassware including flasks, beakers, and test tubes, some containing liquids, set against a blurred, colorful bokeh background of blue, purple, and pink lights.

Chapter 4

Theories of Acids and Bases (Acid-Base Equilibria)

CHAPTER ANALYSIS



FOCUS

- One of the major chapters
- Focus more on calculations
- 7 **key** concepts



EXAM

- Always tested



WEIGHTAGE

- Heavy weightage

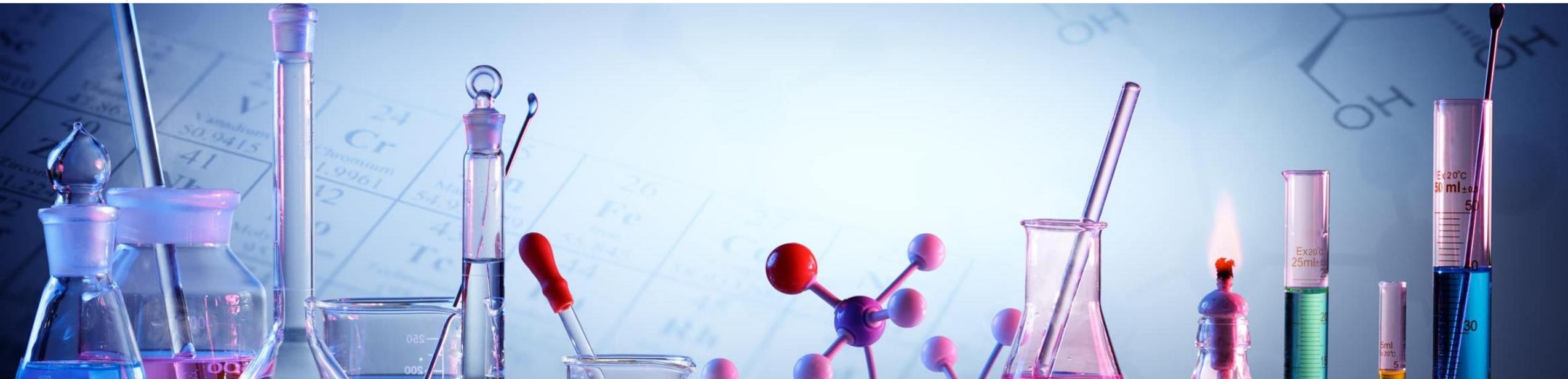


KEY CONCEPT

Introduction to Acid & Base

Ionic Product of Water (K_w)

Dissociation Constant of Acid & Base



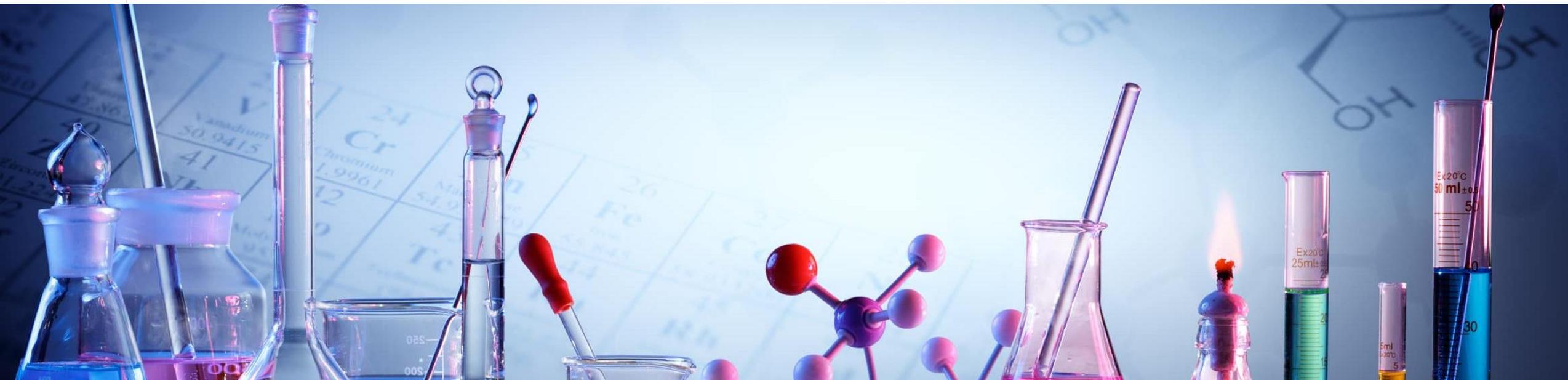
KEY CONCEPT

Methods to Calculate pH of Acid/Base

Methods to Calculate pH of a Salt Solution

Buffer Solutions

Titration Curves



Introduction to Acid & Base

ARRHENIUS THEORY

- **Arrhenius Acid** dissolves in water to **produce H^+ (aq)/ H_3O^+ (aq) ions**
- **Arrhenius Base** dissolves in water to **produce OH^- (aq) ions**

BRONSTED-LOWRY THEORY

- A **Bronsted Acid** is any species that **donates a proton H^+**
- A **Bronsted Base** is any species that **accepts a proton H^+**

LEWIS THEORY

- **Lewis Acid** **accepts electron pair** to form a dative covalent bond.
- **Lewis Base** **donates electron pair** to form a dative covalent bond.

Introduction to Acid & Base

CONJUGATE ACID-BASE PAIR

- The **product formed from an acid** is known as the **conjugate base** of the acid and the **product formed from a base** is known as the **conjugate acid** of the base.
- The **stronger the acid**, the **weaker its conjugate base** is
- The **stronger the base**, the **weaker its conjugate acid** is

TYPES OF ACID

- Monoprotic or **monobasic acids** can donate only one H^+ per molecule.
- **Polyprotic acids** have more than one H^+ per molecule. They dissociate in water to form H^+ in a **stepwise manner**.

TYPES OF BASE

- Monoprotic or **monoacidic bases** can accept only one H^+ per molecule.
- **Polyprotic base** can accept more than one H^+ per molecule. They accept H^+ in a **stepwise manner**.

Introduction to Acid & Base

STRENGTH OF ACID & BASE

- **Strong** Acid/Base dissociates **completely** in aqueous solution.
- **Weak** Acid/Base dissociates **partially** in aqueous solution.

Note: \rightarrow is use to represent the dissociation of strong acid/base while \rightleftharpoons is use to represent the dissociation of weak acid/base

DEGREE OF DISSOCIATION

$$\alpha = \frac{[\text{H}_3\text{O}^+]_{\text{equilibrium}}}{[\text{acid}]_{\text{initial}}}$$

- The **greater the degree of dissociation**, the **stronger the acid/base** is
- For strong acid/base, $\alpha = 1$
- For weak acid/base, $0 < \alpha < 1$

IONIC PRODUCT OF WATER (K_w)

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \text{ mol}^2 \text{ dm}^{-6}$$

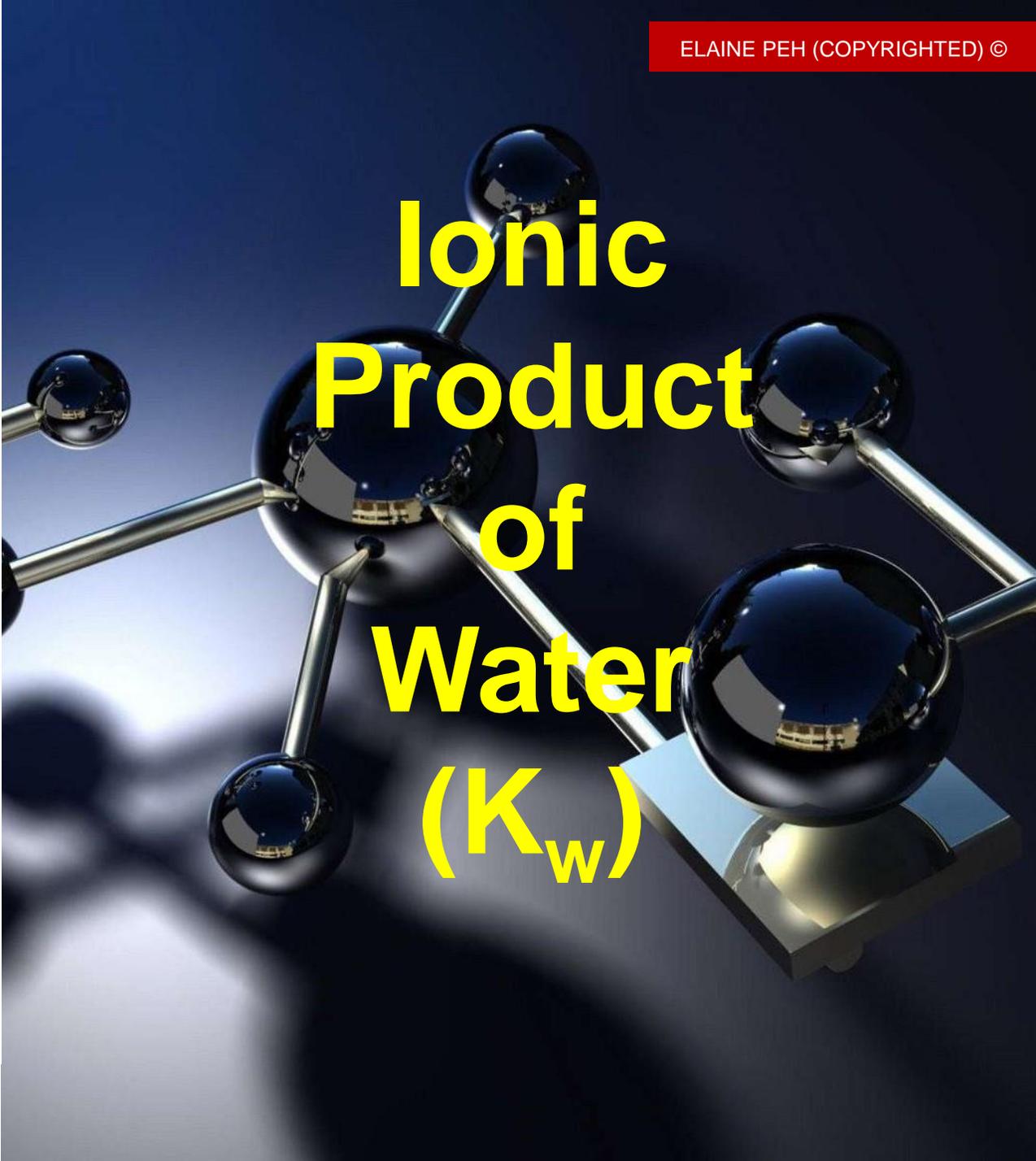
- Water undergoes **auto-ionisation** when a water molecule donates a proton to another water molecule

VARIATION OF K_w WITH TEMPERATURE



- When **temperature increase**, the **equilibrium will shift to the right** to absorb heat, **producing more H_3O^+ and OH^- ions**.
- Hence both $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ increases, which result in an increase in K_w .
- pH of water also decreases** when **temperature increase**, however this does not mean that water is becoming more acidic.
- Water remains neutral** as there is still an equal concentration of H_3O^+ and OH^- ions.

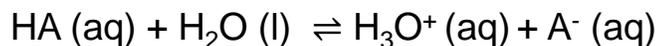
Note: Linking K_w with Chapter 9: Chemical Equilibria



**Ionic
Product
of
Water
(K_w)**

Dissociation Constant of Acid & Base

ACID DISSOCIATION CONSTANT (K_a)



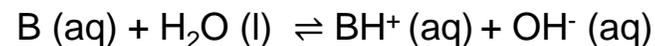
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

- Units = mol dm⁻³
- It **measures** the **extent of dissociation** of a weak acid
- The **larger the K_a** , the **greater degree of dissociation**, hence the **stronger the acid**
- Concentration of Water/Solvent are excluded from the equilibrium expression

$$\text{p}K_a = -\log [K_a]$$

The **smaller the $\text{p}K_a$** , the **larger the K_a** , the **stronger the acid**

BASE DISSOCIATION CONSTANT (K_b)



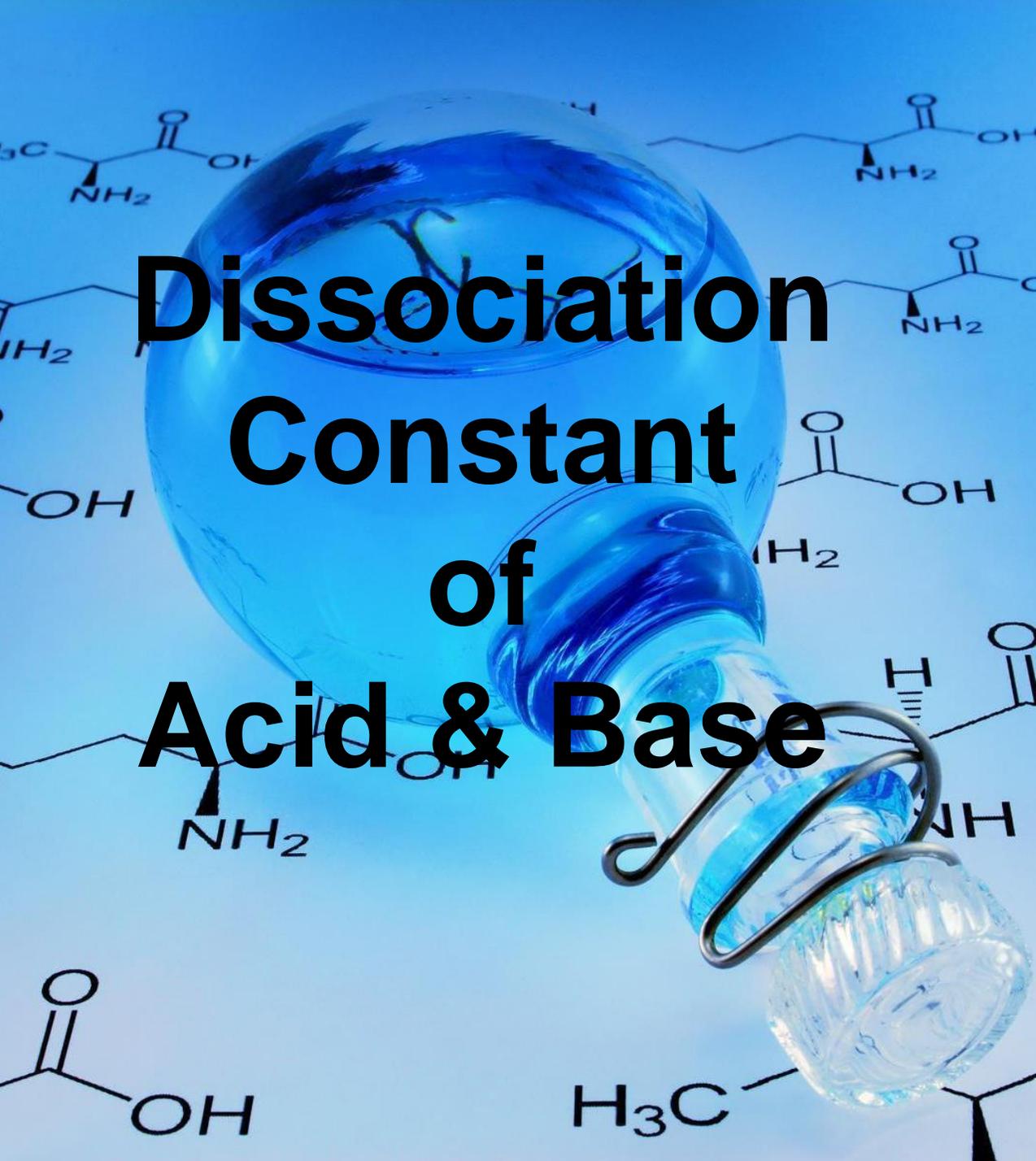
$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

- Units = mol dm⁻³
- It **measures** the **extent of dissociation** of a weak base
- The **larger the K_b** , the **greater degree of dissociation**, hence the **stronger the base**
- Concentration of Water/Solvent are excluded from the equilibrium expression

$$\text{p}K_b = -\log [K_b]$$

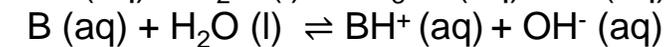
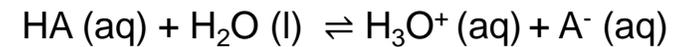
The **smaller the $\text{p}K_b$** , the **larger the K_b** , the **stronger the base**

Note: $\text{p}K_a$ and K_a is the BEST indicator of the strength of the acid/base as it measure the degree of dissociation of the acid/base and it is a constant at constant temperature and does not vary with concentration.



Dissociation Constant of Acid & Base

RELATIONSHIP BETWEEN K_a , K_b AND K_w



$$K_a \times K_b = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \times \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

$$= [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

$$K_a \times K_b = K_w = 10^{-14}$$

$$\text{p}K_a + \text{p}K_b = \text{p}K_w = 14$$

STEP 1: Determine if it is an Acid or Base

- For acid, find pH using $\text{pH} = -\log [\text{H}_3\text{O}^+]$
- For base, find pH using $\text{pOH} = -\log [\text{OH}^-]$

To obtain pH from pOH, using the pH scale at 25°C: $\text{pH} = 14 - \text{pOH}$

STEP 2: Determine if it is strong or weak acid/base

- **Strong Acid/Base** dissociates **completely** in aqueous solution.

For a strong **monoprotic acid**, $[\text{acid}] = [\text{H}_3\text{O}^+]$

For a strong **diprotic acid**, $2x[\text{acid}] = [\text{H}_3\text{O}^+]$

For a strong **monoprotic base**, $[\text{base}] = [\text{OH}^-]$

For a strong **diprotic base**, $2x[\text{base}] = [\text{H}_3\text{O}^+]$

- **Weak Acid/Base** dissociates **partially** in aqueous solution.

We need to make use of K_a/K_b

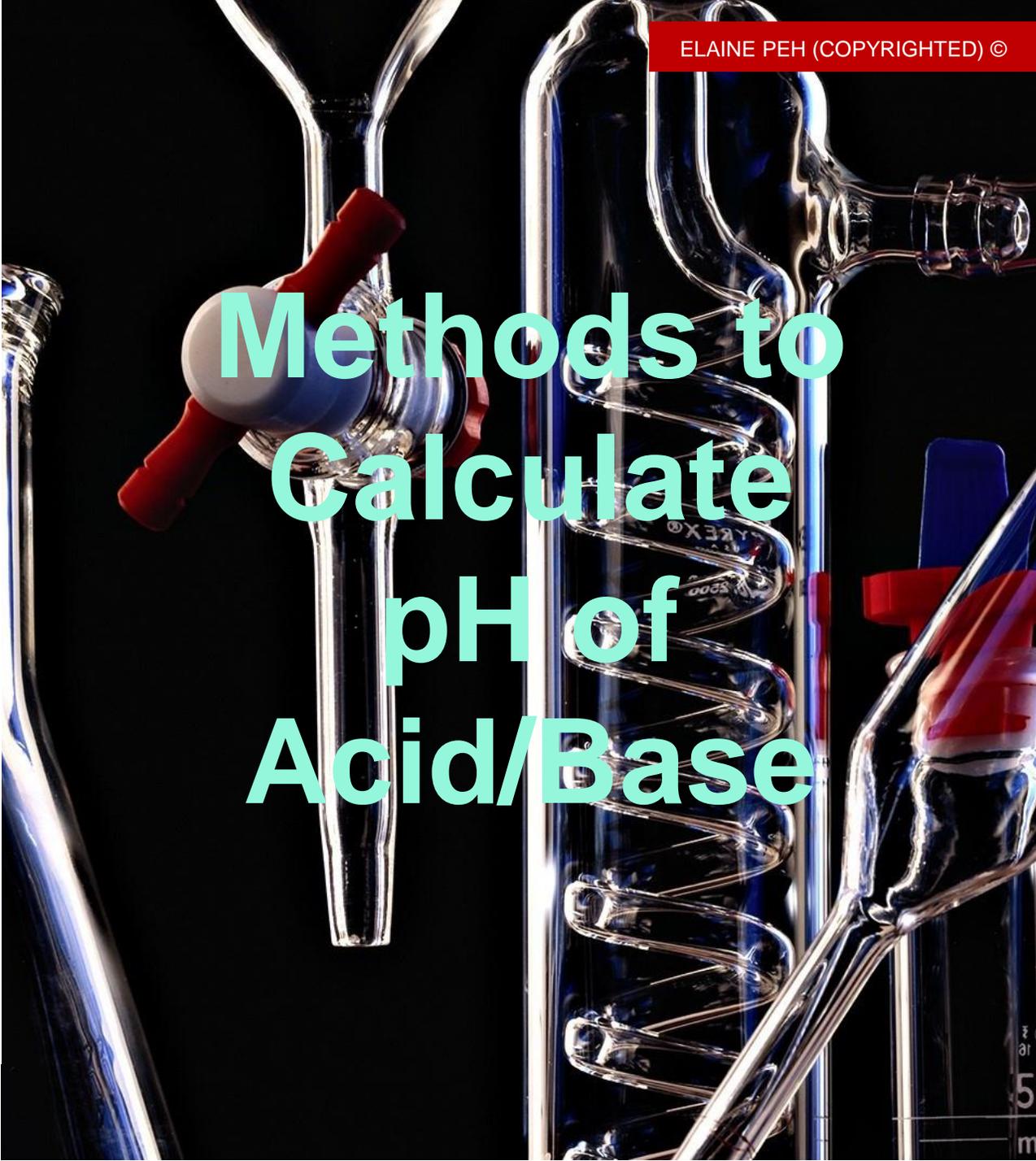
Step 2a: Write the reaction equation for the dissociation of the weak acid/base

Step 2b: Complete the I.C.E Table

Note: Recap the writing of I.C.E Table in Chapter 9: Chemical Equilibria

Step 2c: Obtain the $[\text{H}_3\text{O}^+]/[\text{OH}^-]$ from K_a/K_b

Note: For polyprotic acid/base, there will be more than one K_a/K_b . To determine which K_a/K_b to use, always refer back to the starting reagent in the reaction equation.

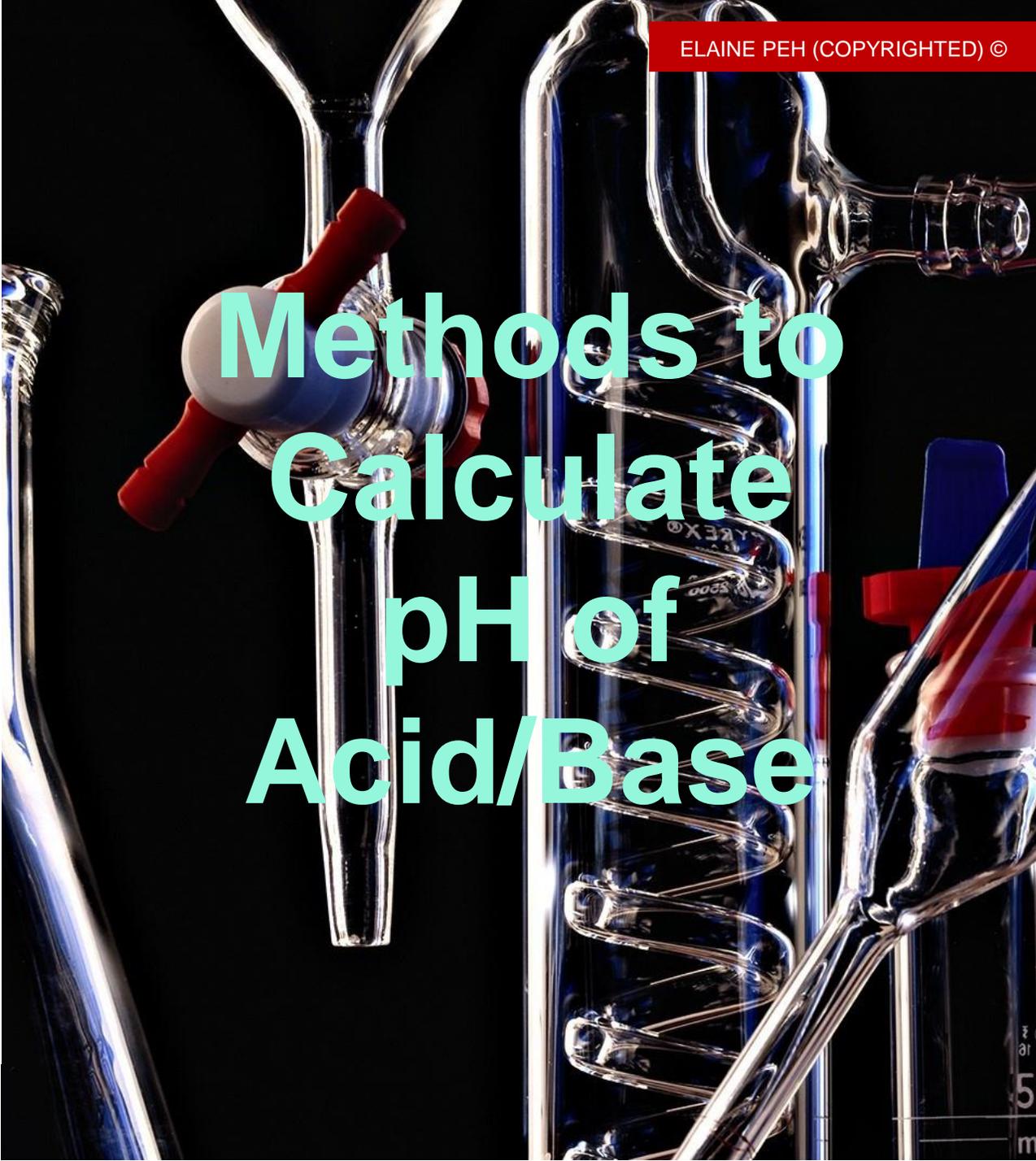


Methods to Calculate pH of Acid/Base

STEP 3: Determine if the contribution of $[\text{H}_3\text{O}^+]$ / $[\text{OH}^-]$ by the auto-ionisation of water is negligible or significant

Auto-ionisation of water produces $[\text{H}_3\text{O}^+] = 10^{-7}$ and $[\text{OH}^-] = 10^{-7}$ (at 25°C)

- For **high concentration** of acid/base the **contribution** of $[\text{H}_3\text{O}^+]$ / $[\text{OH}^-] = 10^{-7}$ from **auto-ionisation** of water is **negligible**.
- For **very diluted** acid/base the **contribution** of $[\text{H}_3\text{O}^+]$ / $[\text{OH}^-] = 10^{-7}$ from **auto-ionisation** of water is **significant**
-



Methods to Calculate pH of Acid/Base

Methods to Calculate pH of a Salt Solution

STEP 1: Determine if it is an acidic or basic salt

- Acidic salt is formed from a strong acid and weak base
- Basic salt is formed from a weak acid and strong base
- For acidic salt, find pH using $\text{pH} = -\log [\text{H}_3\text{O}^+]$
- For basic salt, find pH using $\text{pOH} = -\log [\text{OH}^-]$
- Neutral salt $\text{pH} = 7$

STEP 2: Use the weak acid/base method to find the pH/pOH

- Step 2a: Write the reaction equation for the dissociation of the weak acid/base
- Step 2b: Complete the I.C.E Table
- Step 2c: Obtain the $[\text{H}_3\text{O}^+]/[\text{OH}^-]$ from K_a/K_b

BUFFER SOLUTION

A buffer solution is a solution in which its pH remains **almost unchanged** on the **addition of a small amount** of acid or base.

It consist of a weak acid/base and its conjugate base/acid

ACIDIC BUFFER

An **acidic buffer** solution is made up of a **weak acid** and its **conjugate base**.

BASIC BUFFER

A **basic buffer** solution is made up of a **weak base** and its **conjugate acid**.

Buffer Solutions



USING OF EQUATIONS TO ILLUSTRATE THE BUFFERING EFFECTS

Acidic buffer (HA, A⁻)

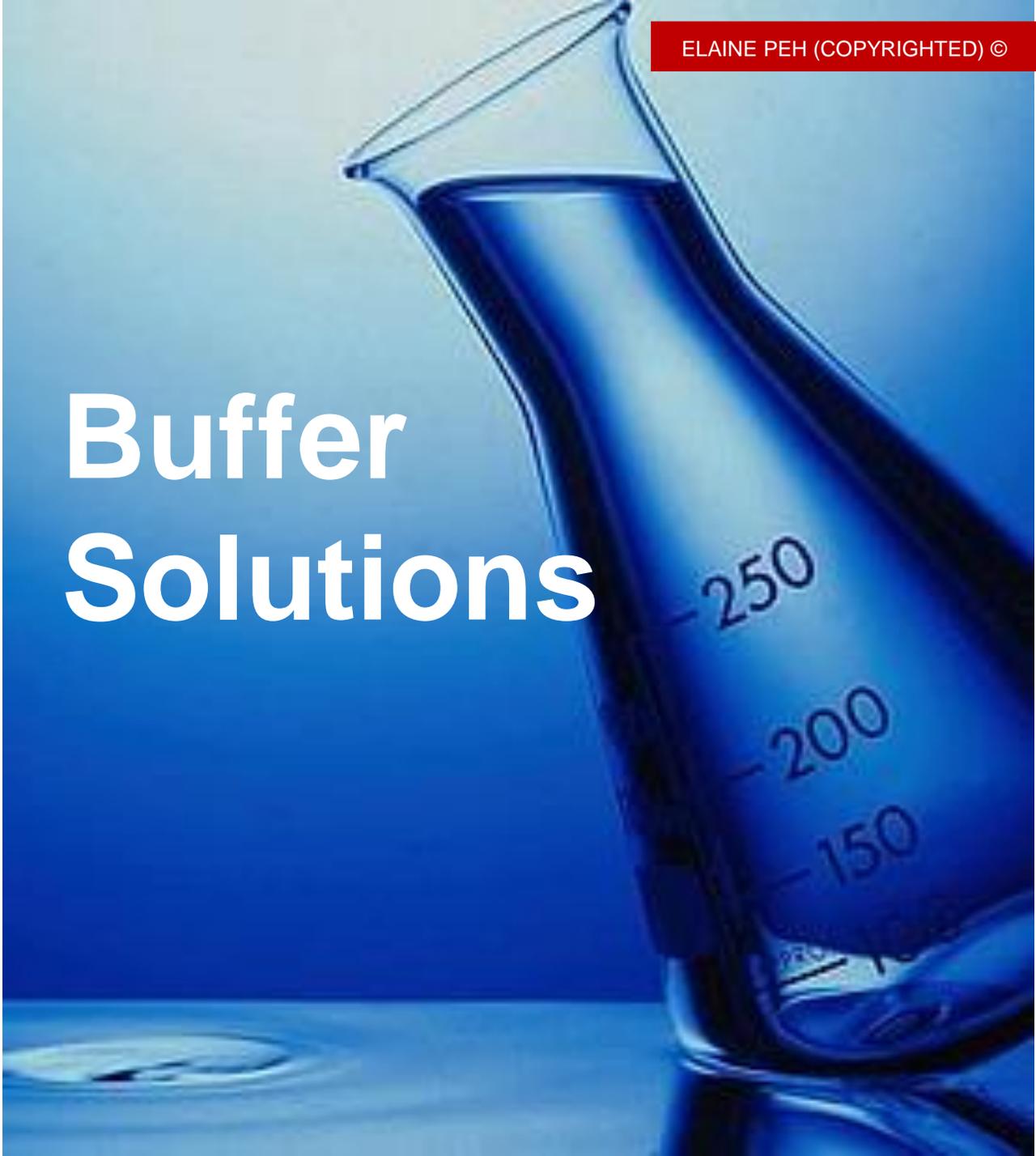
- When small amount of acid is added: $A^- + H_3O^+ \rightarrow HA + H_2O$
- When small amount of base is added: $HA + OH^- \rightarrow A^- + H_2O$

Basic buffer (B, BH⁺)

- When small amount of acid is added: $B + H_3O^+ \rightarrow BH^+ + H_2O$
- When small amount of base is added: $BH^+ + OH^- \rightarrow B + H_2O$

Note: when writing equations, make sure that the products formed from the reaction must already be present in the buffer solution

Buffer Solutions



Buffer Solutions

USING OF WORDS TO ILLUSTRATE THE BUFFERING EFFECTS

Acidic buffer (HA, A⁻)

- When small amount of acid is added: The **addition of H₃O⁺** to the buffer solution will **react with the basic conjugate base, A⁻**, that is present in the solution. The **formation of HA is negligible** as compared to the **large reservoir of HA** already present in the solution. Hence the [H₃O⁺] in the buffer solution does not increase much and the pH of the solution is maintained.
- When small amount of base is added: The **addition of OH⁻** to the buffer solution will **react with the acid, HA**, present in the solution. The **formation of A⁻ is negligible** as compared to the **large reservoir of A⁻** already present in the solution. Hence the [OH⁻] in the buffer solution does not increase much and the pH of the solution is maintained.

USING OF WORDS TO ILLUSTRATE THE BUFFERING EFFECTS

Basic buffer (B, BH⁺)

- When small amount of acid is added: The **addition of H₃O⁺** to the buffer solution will **react with the base, B**, present in the solution. The **formation of BH⁺ is negligible** as compared to the **large reservoir of BH⁺** already present in the solution. Hence the [H₃O⁺] in the buffer solution does not increase much and the pH of the solution is maintained.
- When small amount of base is added: The **addition of OH⁻** to the buffer solution will **react with the acidic conjugate acid, BH⁺**, present in the solution. The **formation of B is negligible** as compared to the **large reservoir of B** already present in the solution. Hence the [OH⁻] in the buffer solution does not increase much and the pH of the solution is maintained.

COMMONLY TESTED EXAMPLE: Buffer in Blood

- Buffer system in blood: Carbonic acid, H_2CO_3 , and bicarbonate, HCO_3^-
- When pH of blood drops: $\text{HCO}_3^- + \text{H}_3\text{O}^+ \rightarrow \text{H}_2\text{CO}_3 + \text{H}_2\text{O}$
- When pH of blood increase: $\text{H}_2\text{CO}_3 + \text{OH}^- \rightarrow \text{HCO}_3^- + \text{H}_2\text{O}$

CALCULATION OF PH OF BUFFER SOLUTIONS

Step 1: Determine if it is an acidic or basic buffer solution

- Using the Henderson-Hasselbalch equation:
- $\text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$
- $\text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{base}]}$

Step 2: Calculate the concentration of the salt and the acid/base in the solution AFTER the reaction (if any) has complete

Buffer Solutions

MAXIMUM BUFFER CAPACITY

This occurs when $[\text{acid}]/[\text{base}]$ is equals to the $[\text{salt}]$.

- $\text{pH} = \text{pK}_a$ & $\text{pOH} = \text{pK}_b$

Titration Curves



INDICATOR

- They are substances that change colour according to the pH of the solution.

Indicator	Acid Colour (at pH below the pH range)	Alkaline Colour (at pH above the pH range)	pH Range
Methyl Orange	Red	Yellow	3.1 – 4.4
Litmus	Red	Blue	5.0 – 8.0
Phenolphthalein	Colourless	Red	8.3 – 10.0

TERMINOLOGIES

- **Titrant** is the **solution of known concentration** that is added to another solution of unknown concentration. It is **found in the burette**.
- **Equivalence point** of a titration is reached when the reactants have just neutralised each other according to the **stoichiometric ratio** given by the balanced equation. This is a **theoretical value**.
- **End point** of a titration is reached **when the indicator changed its colour** when 1 drop of excess titrant is added. This is an **experimental value**.

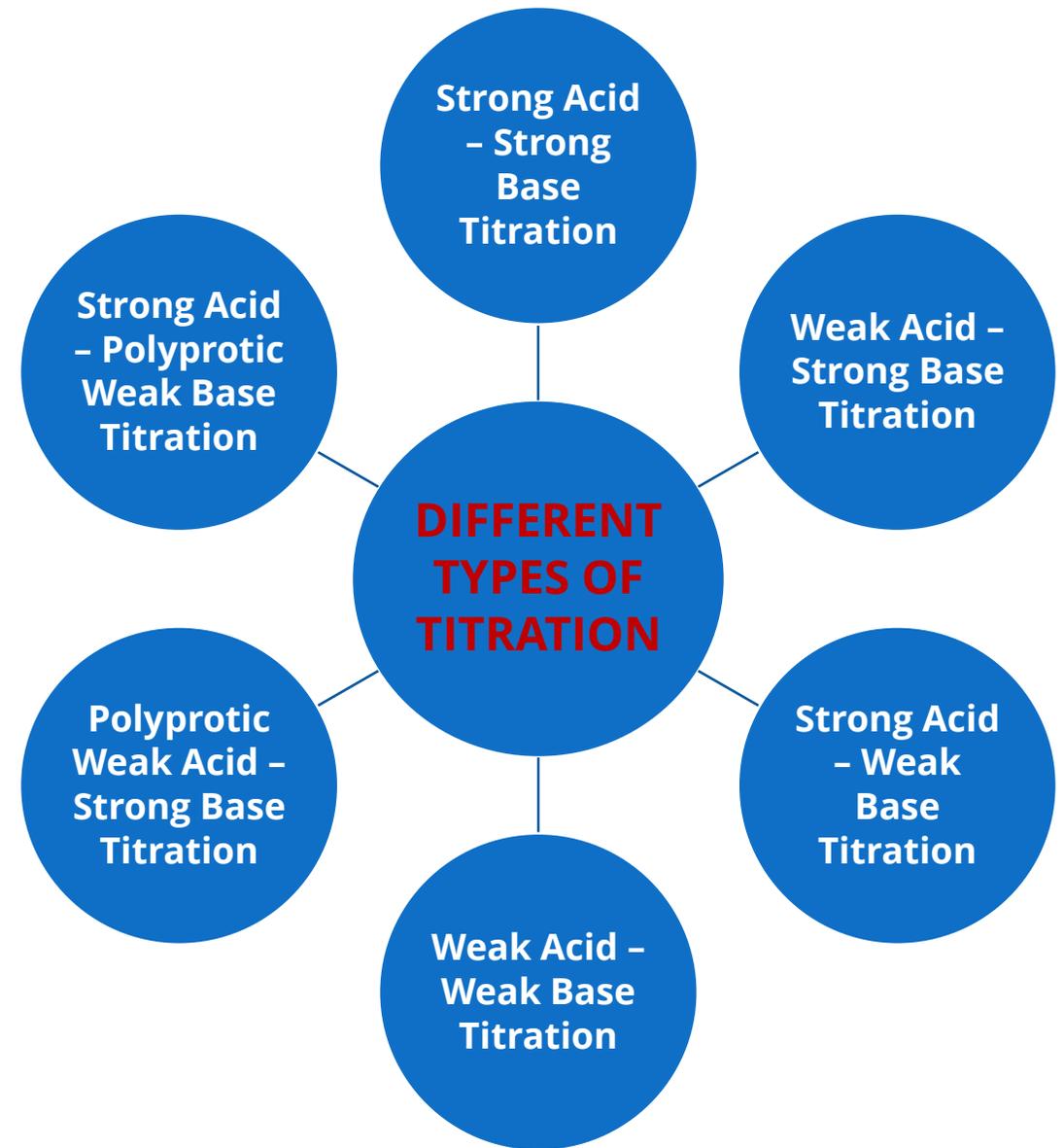
Titration Curves



COMMONLY TESTED POINTS DURING TITRATION

- **Initial pH:**
 - Use the methods to calculate pH of acid/base
- **pH at equivalence point**
 - Use the methods to calculate pH of a salt solution
- **pH at half-neutralisation point**
 - Determine if it is a buffer or not
 - Use the pH formula for a buffer solution if it is a buffer solution
 - If it is not a buffer solution, use the method to calculate pH of acid/base
- **pH after equivalence point**
 - Determine if it is a buffer or not
 - Use the pH formula for a buffer solution if it is a buffer solution
 - If it is not a buffer solution, use the method to calculate pH of acid/base
- **Suitable indicator used**
 - Determine the type of titration first to determine the equivalence point

Titration Curves

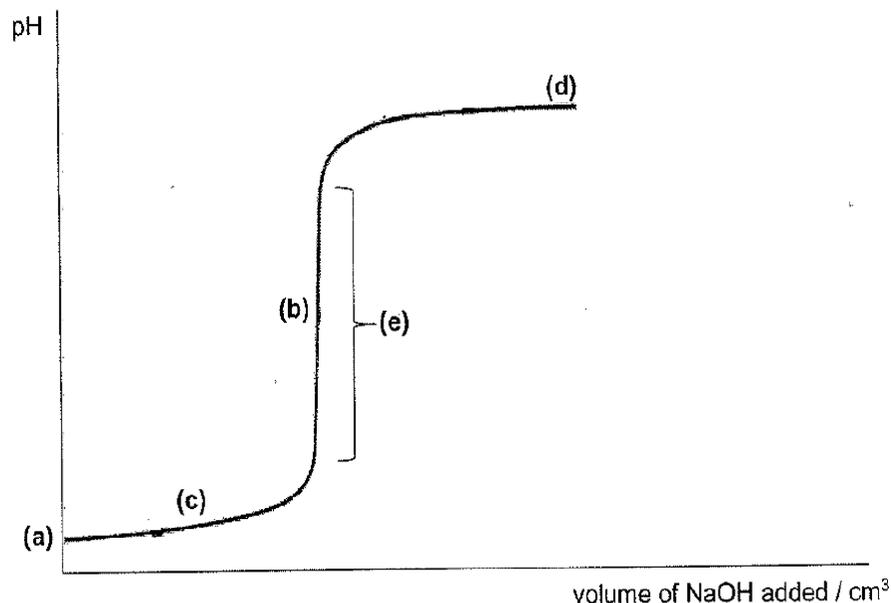


Titration Curves

1. Strong Acid – Strong Base Titration

- Salt formed: neutral
- Marked pH change: 4-10
- Common indicator used: Any indicator is suitable
- No buffering regions

Example: $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
 25 cm³ of 0.1 mol dm⁻³ HCl in the conical flask titrated with 0.1 mol dm⁻³ NaOH in the burette:



(a) **Initial pH** (when 0.00 cm³ of NaOH added)
 $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.1) = 1$

(b) **pH and volume at equivalence point (mid-point of the vertical section)**
 amount of HCl in the conical flask
 $= 0.1 \times (25.0 \div 1000) = 0.00250 \text{ mol}$
 amount of NaOH required = 0.00250 mol
 volume of NaOH required =
 $(0.00250 \div 0.1) \times 1000 = 25.00 \text{ cm}^3$

pH at equivalence point = 7 (as there is no salt hydrolysis)

(d) **pH after equivalence point** (when 50.00 cm³ of NaOH added)

amount of NaOH added = $0.1 \times (50.00 \div 1000)$
 $= 0.00500 \text{ mol}$
 amount of excess NaOH = $0.00500 - 0.00250$
 $= 0.00250 \text{ mol}$
 total volume = $25.0 + 50.00 = 75.0 \text{ cm}^3$
 $[\text{NaOH}] = 0.00250 \div (75.0 \div 1000)$
 $= 3.3333 \times 10^{-3} \text{ mol dm}^{-3} = [\text{OH}^-]$
 $\text{pOH} = -\log [\text{OH}^-] = -\log (3.3333 \times 10^{-3})$
 $= 1.50$
 $\text{pH} = 14 - \text{pOH} = 14 - 1.50 = 12.5$

(c) **pH and volume at half-neutralisation point**

volume at half-neutralisation point = 12.50 cm³

amount of NaOH added = $0.1 \times (12.50 \div 1000)$
 $= 0.00125 \text{ mol}$
 amount of HCl left in the conical flask =
 $0.00250 - 0.00125 = 0.00125 \text{ mol}$
 total volume at half-neutralisation point
 $= 25.0 + 12.50 = 37.5 \text{ cm}^3$
 $[\text{HCl}] = 0.00125 \div (37.5 \div 1000)$
 $= 3.3333 \times 10^{-3} \text{ mol dm}^{-3} = [\text{H}_3\text{O}^+]$
 $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (3.3333 \times 10^{-3}) = 1.5$

(e) **suitable indicator**

- phenolphthalein (working range: 8.3 – 10.0)
 end-point: colourless to pink
- methyl orange (working range: 3.1 – 4.4)
 end-point: red to orange

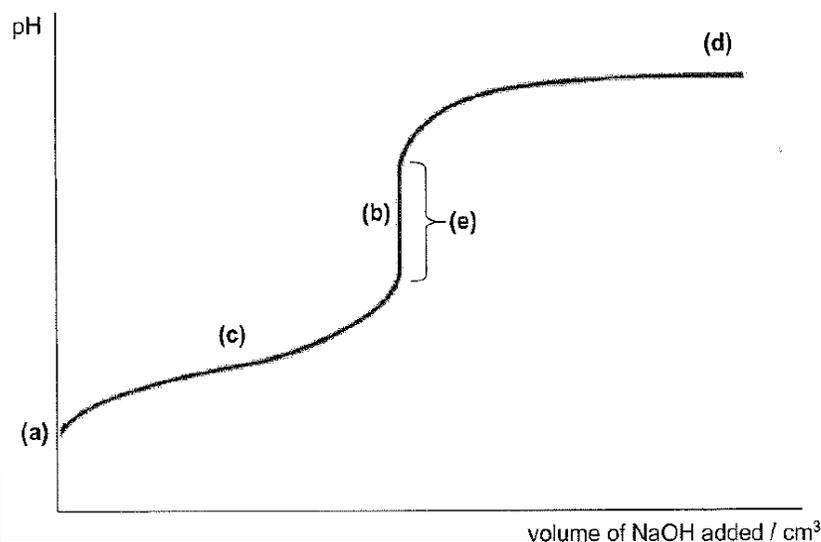
Titration Curves

2. Weak Acid – Strong Base Titration

- Salt formed: basic
- Marked pH change: 7.5-10.5
- Common indicator used: Phenolphthalein
- **Buffering region** can be found **before/after** the equivalence point (depending on which reagent is in the conical flask)

Example:

$\text{CH}_3\text{COOH (aq)} + \text{NaOH (aq)} \rightarrow \text{CH}_3\text{COONa (aq)} + \text{H}_2\text{O (l)}$
 25 cm³ of 0.1 mol dm⁻³ CH₃COOH in the conical flask titrated with 0.1 mol dm⁻³ NaOH in the burette (K_a of CH₃COOH = 1.74×10^{-5} mol dm⁻³):



(a) Initial pH (when 0.00 cm³ of NaOH added)
 $[\text{H}_3\text{O}^+] = \sqrt{K_a \times c} = 1.3190 \times 10^{-3} \text{ mol dm}^{-3}$
 $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.3190 \times 10^{-3}) = 2.9$

(c) pH and volume at half-neutralisation point

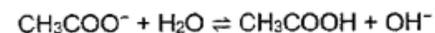
volume at half-neutralisation point = **12.50 cm³**

amount of NaOH added = $0.1 \times (12.50 \div 1000) = 0.00125 \text{ mol}$
 amount of CH₃COOH left in the conical flask = $0.00250 - 0.00125 = 0.00125 \text{ mol}$
 amount of CH₃COO⁻ formed in the conical flask = 0.00125 mol
 $[\text{acid}] = [\text{salt}] \rightarrow$ maximum buffering capacity.
 $\text{pH} = \text{p}K_a = -\log(1.74 \times 10^{-5}) = 4.8$

(b) pH and volume at equivalence point (mid-point of the vertical section)
 amount of CH₃COOH in the conical flask = $0.1 \times (25.0 \div 1000) = 0.00250 \text{ mol}$
 amount of NaOH required = 0.00250 mol
 volume of NaOH required = $(0.00250 \div 0.1) \times 1000 = 25.00 \text{ cm}^3$

(d) pH after equivalence point (when 50.00 cm³ of NaOH added)
 $\text{pH} = 12.5$
 (similar calculation as SA-SB)

pH at equivalence point > 7 because the conjugate base (anion) undergoes hydrolysis to give OH⁻



amount of CH₃COO⁻ = 0.00250 mol
 $[\text{CH}_3\text{COO}^-] = 0.00250 \div (50.0 \div 1000) = 0.0500 \text{ mol dm}^{-3}$
 $K_b(\text{CH}_3\text{COO}^-) = K_w \div K_a(\text{CH}_3\text{COOH}) = 5.75 \times 10^{-10} \text{ mol dm}^{-3}$
 $[\text{OH}^-] = \sqrt{K_b \times c} = 5.3619 \times 10^{-6} \text{ mol dm}^{-3}$
 $\text{pH} = 14 - \text{pOH} = 8.7$

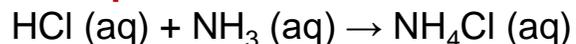
(e) suitable indicator
 phenolphthalein (working range: 8.3 – 10.0)
 end-point: colourless to pink

Titration Curves

3. Strong Acid – Weak Base Titration

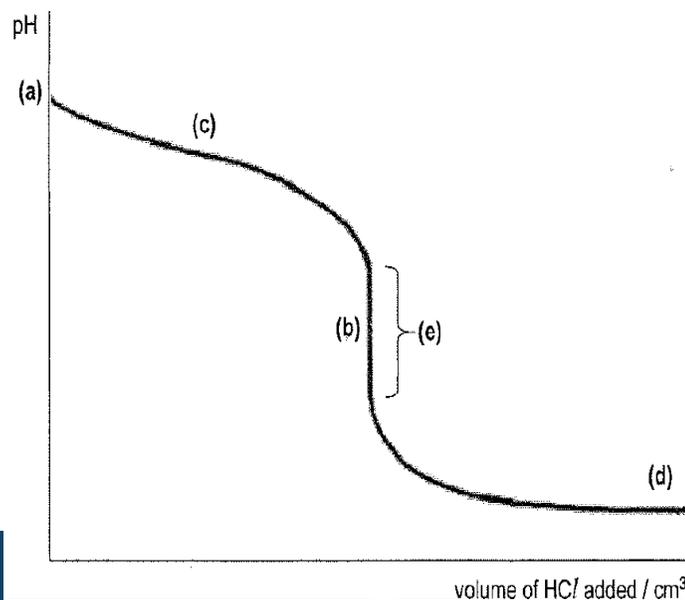
- Salt formed: acidic
- Marked pH change: 3.5-6.5
- Common indicator used: Methyl Orange
- **Buffering region** can be found **before/after** the equivalence point (depending on which reagent is in the conical flask)

Example:



25 cm³ of 0.1 mol dm⁻³ NH₃ in the conical flask titrated with 0.1 mol dm⁻³ HCl in the burette

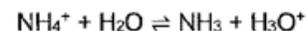
(K_b of NH₃ = 1.78 × 10⁻⁵ mol dm⁻³):



(a) **Initial pH** (when 0.00 cm³ of HCl added)
 $[[\text{OH}^-] = \sqrt{K_b \times c} = 1.3341 \times 10^{-3} \text{ mol dm}^{-3}$
 pH = 14 – pOH = **11.1**

(c) **pH and volume at half-neutralisation point**
 volume at half-neutralisation point = **12.50 cm³**
 amount of HCl added = 0.1 × (12.50 ÷ 1000) = 0.00125 mol
 amount of NH₃ left in the conical flask = 0.00125 mol
 amount of NH₄⁺ formed in the conical flask = 0.00125 mol
 [base] = [salt] → maximum buffering capacity.
 pOH = pK_b = –log(1.78 × 10⁻⁵) = 4.7
 pH = 14.0 – 4.7 = **9.3**

(b) **pH and volume at equivalence point (mid-point of the vertical section)**
 amount of NH₃ in the conical flask = 0.00250 mol
 amount of HCl required = 0.00250 mol
 volume of HCl required = **25.00 cm³**
 pH at equivalence point < 7 because the conjugate acid (cation) undergoes hydrolysis to give H₃O⁺



amount of NH₄⁺ = 0.00250 mol
 $[\text{NH}_4^+] = 0.00250 \div (50.0 \div 1000) = 0.0500 \text{ mol dm}^{-3}$
 $K_a(\text{NH}_4^+) = K_w \div K_b(\text{NH}_3) = 5.62 \times 10^{-10} \text{ mol dm}^{-3}$
 $[\text{H}_3\text{O}^+] = \sqrt{K_a \times c} = 5.3009 \times 10^{-6} \text{ mol dm}^{-3}$
 pH = **5.3**

(d) **pH after equivalence point** (when 50.00 cm³ of HCl added)
 amount of HCl added = 0.00500 mol
 amount of excess HCl = 0.00250 mol
 total volume = 25.0 + 50.00 = 75.0 cm³
 $[\text{HCl}] = 0.00250 \div (75.0 \div 1000) = 3.3333 \times 10^{-3} \text{ mol dm}^{-3} = [\text{H}_3\text{O}^+]$
 pH = **1.5**

(e) **suitable indicator**
 methyl orange (working range: 3.1 – 4.4)
 end-point: red to orange

Titration Curves



5. Titration with polyprotic acid/base:

- Neutralisation takes place in more than one step as the polyprotic acid/base dissociate in water to form H^+ / accepts H^+ in a stepwise manner
- For a dibasic weak acid or diacidic weak base, the titration curve will have two equivalence point and two buffering regions.
- Two different indicator will be required depending on the pH of the two equivalence points.

Practice Questions

Question: Find the pH of a 0.1 mol dm^{-3} solution of HNO_3 .

Answer: $\text{pH} = 1.0$

Question: Find the pH of a 0.1 mol dm^{-3} solution of NaOH .

Answer: $\text{pH} = 13.0$

Question: Explain why the pH of a solution containing $1 \times 10^{-8} \text{ mol dm}^{-3} \text{ HNO}_3$ is 6.96 and not 8.

Answer: As the $[\text{H}_3\text{O}^+]$ from the acid is only $1 \times 10^{-8} \text{ mol dm}^{-3}$ which is much lower than $[\text{H}_3\text{O}^+]$ from water (10^{-7}), hence we need to take into consideration the contribution of $[\text{H}_3\text{O}^+]$ by the auto-ionisation of water.

Question: Given that the acid dissociation constant for ethanoic acid is $1.74 \times 10^{-5} \text{ mol dm}^{-3}$, find the pH of a 0.1 mol dm^{-3} solution of ethanoic acid.

Answer: $\text{pH} = 2.88$

Practice Questions

Question: Malonic acid ($\text{HO}_2\text{CCH}_2\text{CO}_2\text{H}$) is a type of dicarboxylic acid with a $\text{pK}_{\text{a}1} = 2.85$ and $\text{pK}_{\text{a}2} = 5.70$. (a) Calculate the pH of a 0.1 mol dm^{-3} malonic acid.

Answer: $\text{pH} = 1.92$

Question: Find the pH of a 0.1 mol dm^{-3} solution of H_2SO_4 .

Answer: $\text{pH} = 0.7$

Question: The pH of a solution of HCl is 3.30, calculate the concentration of HCl in the solution.

Answer: $[\text{HCl}] = 5.01 \times 10^{-4} \text{ mol dm}^{-3}$

Question: Calculate the pH of the solution formed by dissolving 1.00 g of sodium oxide in 500 cm^3 of water.

Answer: $\text{pH} = 12.8$.

Practice Questions

Question: Calculate the K_a value of a weak acid HA given that the pH of a 0.05 mol dm^{-3} solution is 4.5.

Answer: $K_a = 2 \times 10^{-8} \text{ mol dm}^{-3}$

Question: Calculate the pH and pK_a value of a weak acid given that it 1% dissociated for a 0.2 solution.

Answer: $pH = 2.69$, $pK_a = 4.69$

Question: Given that the acid dissociation constant for HCN is $4.9 \times 10^{-10} \text{ mol dm}^{-3}$, find the pH of a 0.2 mol dm^{-3} NaCN solution.

Answer: $pH = 8.6$

Question: Calculate the pH of 0.06 mol dm^{-3} CH_3COONa , K_a of $\text{CH}_3\text{COOH} = 1.80 \times 10^{-5} \text{ mol dm}^{-3}$.

Answer: $pH = 8.76$

Practice Questions

Question: Calculate the pH of 0.03 mol dm⁻³ NH₄Cl, K_b of NH₃ = 1.78 x 10⁻⁵ mol dm⁻³.

Answer: $pH = 5.39$

Question: Given that the K_a value of ethanoic acid is 1.76 x 10⁻⁵ mol dm⁻³, calculate the (a) K_a value of its ethanoate ion and (b) the pH of a 0.1 mol dm⁻³ solution of ethanoate ion.

Answer:

(a) $K_a = 5.68 \times 10^{-10}$ mol dm⁻³

(b) $pH = 8.88$

Question: Calculate the pH of the solution made by mixing 10.0 cm³ of 0.1 mol dm⁻³ NaOH and 15.0 cm³ of 0.1 mol dm⁻³ CH₃COOH given that the pK_a of CH₃COOH is 4.75.

Answer: $pH = 5.05$

Question: Calculate the pH of a buffer solution containing 0.2 mol dm⁻³ lactic acid and 0.3 mol dm⁻³ sodium lactate with $K_a = 1.38 \times 10^{-4}$ mol dm⁻³.

Ans: $pH = 4.04$

Practice Questions

Question: Calculate the pH of the solution made by mixing 60.0 cm³ of 0.1 mol dm⁻³ NH₃ and 40.0 cm³ of 0.1 mol dm⁻³ HCl given that the pK_b of NH₃ is 4.74.

Answer: $pH = 8.96$

Question: Calculate the pH of the solution made by mixing 10.0 cm³ of 0.1 mol dm⁻³ NaOH and 20.0 cm³ of 0.05 mol dm⁻³ H₂SO₄.

Answer: $pH = 1.48$

Question: An acidic buffer solution can be obtained by dissolving 2.00 g of CH₃COONa in 500 cm³ of 0.01 mol dm⁻³ ethanoic acid, K_a of ethanoic acid = 1.84 × 10⁻⁵ mol dm⁻³. (a) Calculate the pH of the buffer solution formed (b) Calculate the mass of sodium ethanoate that needs to be dissolved in 500 cm³ of 0.01 mol dm⁻³ ethanoic acid to produce a buffer solution of pH 4.5.

Answer:

(a) $pH = 5.42$

(b) 0.236 g

Practice Questions

When sodium carbonate reacts with hydrochloric acid, the neutralisation occurs in two stages:

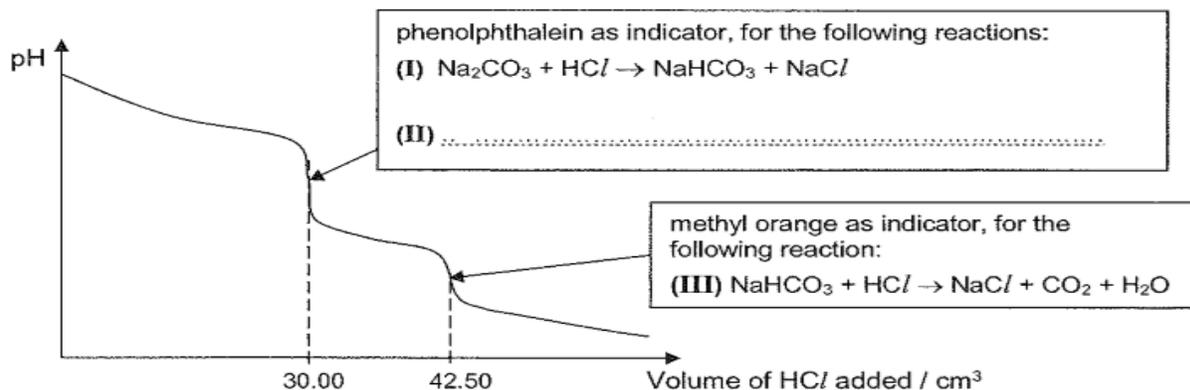
Question:



In an experiment, 25.0 cm^3 of a solution containing sodium carbonate and sodium hydroxide was titrated against $0.100 \text{ mol dm}^{-3}$ hydrochloric acid using phenolphthalein as indicator. The volume of acid just needed for the pink colour to be discharged was noted.

Methyl orange indicator was then added and the titration was continued. The volume of acid that is just required for the methyl orange to change from yellow to orange colour was noted.

The titration curve below summarises the results of the experiment.



Answer:

(a) $1.25 \times 10^{-3} \text{ mol}$



(c) $0.0700 \text{ mol dm}^{-3}$

- Calculate the amount, in moles, of NaHCO_3 that reacted with HCl in reaction **(III)**
- Write the equation for reaction **(II)** in the box provided in the diagram.
- Calculate the concentration of Na_2CO_3 and the concentration of NaOH in the solution, giving your answers in mol dm^{-3} .

Test yourself!

- (a) show understanding of, and apply the Arrhenius theory of acids and bases
- (b) show understanding of, and apply the Bronsted-Lowry theory of acids and bases, including the concept of conjugate acids and conjugate bases
- (c) show understanding of, and apply the Lewis theory of acids and bases (including non-aqueous system, e.g. reaction between BF_3 and NH_3)

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