

St. Andrew's Junior College H1 Chemistry 2022 Lecture Notes Theories of Acids and Bases

Assessment Objectives

Candidates should be able to:

- (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 bases;
- (c) explain qualitatively the differences in behaviour between strong and weak acids and bases in terms of the extent of dissociation;
- (d) explain the terms pH; K_a ; K_b ; K_w [The relationship $K_w = K_a K_b$ is not required];
- (e) calculate [H⁺(aq)] and pH values for strong acids, and strong bases;
- (f) explain the choice of suitable indicators for acid-base titrations, given appropriate data, in terms of the strengths of the acids and bases;
- (g) (i) explain how buffer solutions control pH;
 (ii) describe and explain the uses of buffers, including the role of H₂CO₃/HCO₃⁻ in controlling pH in blood.

Lecture Outline:

- 1. Theories of acids and bases
- 2. Autoionisation of water
- 3. Dissociation of weak acid and bases
- 4. Acid-Base Properties of Salt Solutions
- 5. Indicators for acid-base titrations
- 6. Monoprotic Acid-Base Titration Curves
- 7. Buffer solutions

References:

- 1. Chemistry by Silberberg
- 2. Chemistry by Raymond Chang
- 3. Chemistry for Advanced level by Peter Cann
- 4. Chemistry in context by Hill and Holman
- 5. A level Chemistry by Ramsden

1. THEORIES OF ACIDS AND BASES

Candidates should be able to:

- show understanding of, and apply the Arrhenius theory of acids and bases
- show understanding of, and apply the Brønsted–Lowry theory of acids and bases, including the concept of conjugate acids and conjugate bases

1.1 Arrhenius theory of acids and bases

An Arrhenius acid is a substance that produces H⁺ in aqueous solution.

An Arrhenius base is a substance that produces OH^{-} in aqueous solution.

E.g.: $HNO_3(l) + H_2O(l) \longrightarrow NO_3^{-}(aq) + H_3O^{+}(aq)$

Note: H_3O^+ is hydronium ion, which is the same as $H^+(aq)$

 $NaOH(s) + aq \longrightarrow Na^{+}(aq) + OH^{-}(aq)$

• An acid-base reaction involves the formation of salt and water

1.2 Brønsted–Lowry theory of acids and bases

A Brønsted–Lowry acid is a proton (H⁺) donor.

A Brønsted–Lowry base is a proton (H⁺) acceptor.



- An acid–base reaction involves the transfer of a proton from the acid to the base.
- When a Brønsted–Lowry acid (HA) loses a proton, the resulting product (A⁻) is called the <u>conjugate base</u> of HA.
- When a Brønsted–Lowry base (B) accepts a proton, the resulting product (BH⁺) is called

the conjugate acid of B.

• A conjugate acid-base pair differs by a H⁺.

Examples of conjugate pairs:

HCl / Cl [−]	NH_3/NH_4^+
H₂O / OH⁻	H_2O / H_3O^+
CH3COOH / CH3COO-	CH ₃ NH ₂ /CH ₃ NH ₃ ⁺

Exercise	1		
Identify the acid, b	base, conjugate acio	d and conjugate b	ase in the following reactions :
a) CH₃COOH	I + H ₂ O 🚤	- CH₃COO-	+ H ₃ O ⁺
acid	base	conjugate base	conjugate acid
b) CH₃COOH	I + HC <i>l</i> →	► CH ₃ COOH ₂ ⁺	+ C <i>l</i> ⁻
base	acid	conjugate acid	conjugate base
c) HNO₃	+ H₂O →	NO ₃ ⁻	+ H ₃ O ⁺
acid	base c	onjugate base	conjugate acid
d) HNO₃ +	$H_2SO_4 \longrightarrow$	HSO4 ⁻	+ H ₂ NO ₃ ⁺
base	acid c	onjugate base	conjugate acid

∽ Self–Check: Q1

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	Arrhenius	Brønsted-Lowry
Acids	produces H ⁺ in water	donates H⁺
Bases	produces OH ⁻ in water	accepts H ⁺

Candidates should be able to:

• Explain qualitatively the differences in behavior between strong and weak acids and bases in terms of the extent of dissociation.

1.3 Strength of Acids and Bases



Candidates should be able to:

• explain pH and apply them in calculations

1.4 The pH scale

Concentration of H⁺ can be used as a measure of acidity and alkalinity of a solution.

However, [H⁺] can be very small values. Hence, we can express these values as pH.

pH = - <i>log</i> 10 [H ⁺]	or	$[H^+] = 10^{-pH}$	[H ⁺] in mol dm ⁻³
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A solution with <u>low pH</u> has a <u>high [H⁺]</u>, and is more acidic than a solution with higher pH.

Similarly,

 $pOH = -log_{10} [OH^{-}]$

Note: The logarithm used are to the base 10 (not to the base e), so make sure when doing calculations, you press the log or lg button on your calculator (not the ln button)

1.5 Basicity of acid

- HCl and CH₃COOH are monoprotic (monobasic) acids as each acid can only lose one H⁺.
- H₂SO₄ is a **diprotic** (dibasic) acid as it can lose **two H**⁺.

 $H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^$ conjugate base $HSO_4^- + H_2O \longrightarrow H_3O^+ + SO_4^{2-}$ conjugate base

• Conversely for bases, NaOH is a monoacidic base and Ca(OH)₂ is a diacidic base.

Candidates should be able to:

- explain K_w and apply them in calculations
- calculate [H⁺(aq)] and pH values for strong acids and strong bases

2. AUTOIONISATION OF WATER

Water dissociates very slightly to give ions.

 $H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$

When $[H^+] = [OH^-]$, the solution is neutral. When $[H^+] > [OH^-]$, the solution is acidic. When $[H^+] < [OH^-]$, the solution is alkaline. We name the equilibrium constant for the autoionisation of water as ionic product of water, **K**w.

Kw = [H⁺] [OH⁻] Note: [H₂O] is not reflected in K_w

 $pK_w = -log_{10}K_w = -log_{10}([H^+][OH^-])$ $= - \log_{10}[H^+] - \log_{10}[OH^-]$ $pK_w = pH + pOH$ $pH = pK_w - pOH$

Note: Derivation of pK_w is not required

At 25 °C

Kw	= $1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$	in Data Booklet
р <i>К</i> _w	= 14	
р <i>К</i> w	= pH + pOH	
14	= pH + pOH	
pН	= 14 – pOH	
Note	: Derivation of pK _w is not require	d

In pure water at 25°C,

 $H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$

 $[H^+] = [OH^-] = 1.00 \text{ x } 10^{-7} \text{ mol } \text{dm}^{-3}$ $pH = pOH = - lg (1.00 \times 10^{-7}) = 7.0$

The autoionisation of water molecules is an endothermic process.

 $H_2O(l) \longrightarrow H^+(aq) + OH^-(aq) \Delta H > 0$

 $K_w = \frac{kf}{kb}$. When temperature increases, both k_f and k_b increases. Since the forward reaction is endothermic and is favoured, k_f increases more than k_b . Hence, <u>K_w increases when</u> temperature increases.

Table 1: Variation of ionic pr	oduct of water, K _w ,	with temperature
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Temperature / °C	$K_{\rm w}$ / mol ² dm ⁻⁶
20	0.68 x 10 ⁻¹⁴
30	1.47 x 10 ⁻¹⁴
40	2.92 x 10 ⁻¹⁴

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K _w	= [H⁺] [OH⁻]
рН	= p <i>K</i> _w – pOH
K _w	= 1.00 x 10 ⁻¹⁴ mol ² dm ⁻⁶ at 25°C
рН	= 14 – pOH at 25°C



[2018 P1 Q14]

The ionic product of water, K_w , is affected by temperature.

Temperature / °C	$K_{\rm w} \ge 10^{-14} / {\rm mol}^2 {\rm dm}^{-6}$
10	0.293
40	2.92

Which statement describes what happens as the temperature of water is increased from 10 °C to 40 °C?

A pH of water decreases and $[H^+] = [OH^-]$

B pH of water decreases and [H⁺] is greater than [OH⁻]

C pH of water increases and $[H^+] = [OH^-]$

D pH of water increases and $[H^+]$ is less than $[OH^-]$

Ans: A

 $H_2O(l)$ \longrightarrow $H^+(aq) + OH^-(aq)$ At 40°C $K_w = [H^+][OH^-] = 2.92 \times 10^{-14}$ $[H^+] = [OH^-]$ $[H^+]^2 = 2.92 \times 10^{-14}$ $[H^+] = 1.709 \times 10^{-7} \text{ mol dm}^{-3}$ pH = 6.77

pH of pure water at 40°C is less than pH 7, but it DOES NOT mean that the water is not neutral! A neutral solution has pH 7 only at 25 °C. As K_w changes with temperature, pH of a neutral solution can change with temperature. Hence, as long as [H⁺] = [OH⁻], a solution is neutral.



Steps in Calculating pH of Strong Acid/Base

- 1. Write equation to illustrate complete dissociation (if necessary).
- 2. Determine $[H^+]$ for acids (or $[OH^-]$ for bases)
- 3. Determine pH or pOH using

Acids	Bases
pH = – log ₁₀ [H ⁺]	pOH = – log₁₀ [OH⁻]
	pH = 14 – pOH at 25°C

Exercise 3
Calculate the pH of a) 0.0100 mol dm ⁻³ sulfuric acid
b) 0.00250 mol dm ⁻³ aqueous sodium hydroxide
a) $H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$
0.0100 mol dm ⁻³ 0.0200 mol dm ⁻³
pH = - lg (0.0200) = 1.70
b) NaOn \rightarrow Na + On
0.00250 moi am ² 0.00250 moi am ²
pOH = -lg(0.00250)
= 2.602
$pH = pK_w - pOH$
= 14 - 2.602
= 11.4

Exercise 4

The concentration of OH^- ions in a certain household ammonia cleaning solution is 0.0025 mol dm⁻³. Calculate the pH of the solution.



∽ Self–Check: Q2 – 5

Candidates should be able to:

• explain the terms K_a; K_b.

3 DISSOCIATION OF WEAK ACIDS AND BASES

3.1 Dissociation constants of acids/bases, Ka / Kb

I. Acid dissociation constant, K_a for a weak acid

For a weak monobasic (monoprotic) acid, the dissociation can be represented as follows:

 $HA(aq) \longrightarrow H^+(aq) + A^-(aq)$

We name the equilibrium constant for the dissociation of an acid its **acid dissociation** constant, K_{a} .

$$K_{a} = \frac{[H_{3}O^{+}]_{eqm} [A^{-}]_{eqm}}{[HA]_{eqm}}$$
Units: mol dm⁻³
pK_a = -log₁₀ K_a

Note: $[H_2O]$ is ignored in K_a as it is present in large amount and its concentration remains almost constant

• *K*_a is <u>temperature</u> dependent.

Table 2: Dissociation constants of acids in water at 25 °C

Acids	Formula	K _a / mol dm⁻³	р <i>К</i> а	
Ethanoic acid	CH ₃ COOH	1.74 x 10 ^{−5}	4.75	stronger acid
Phenol	C ₆ H₅OH	1.28 x 10 ⁻¹⁰	9.89	weaker acid

• K_a and pK_a gives a good measurement of the strength of an acid, as it indicates the <u>extent</u> to which the <u>acid is dissociated</u>. The <u>higher the K_a </u> or the <u>lower the pK_a </u>, the <u>stronger the acid</u>.

Which of the following best indicates the strength of an acid?

(1) *K*_a (2) pH

 K_a best indicates the strength of an acid. It <u>does not vary with concentration of acids</u>. At the same temperature, a <u>stronger acid</u> would have a <u>higher K_a </u>, indicating that it has a <u>larger extent</u> <u>of dissociation of H⁺</u>.

pH only indicates the <u>concentration of $H^+(aq)$ in solution</u>. It <u>varies with concentration of the acid</u>. A weak acid of high concentration can give a lower pH than a stronger acid of lower concentration.

Eg: pH of 0.1 mol dm $^{-3}$ aqueous CH_3COOH (weak acid) $\,\approx 2.9$

pH of 10^{-5} mol dm⁻³ aqueous HCl (strong acid) = 5.0

pH can only be used to compare the acid strength of two acids when they have the same concentration.

Eg: pH of 0.1 mol dm $^{-3}$ aqueous CH_3COOH (weak acid) $\,\approx 2.9$

Polybasic (polyprotic) acids, such as H₃PO₄, have more than one dissociation constant.

Dissociation	Equilibrium	K _a / mol dm ⁻³	р <i>К</i> а
K _{a1}	$H_3PO_4(aq)+H_2O(I) \longrightarrow H_3O^+(aq)+H_2PO_4^-(aq)$	7.52 x 10 ^{−3}	2.12
K _{a2}	$H_2PO_4^{-}(aq)+H_2O(l) = H_3O^{+}(aq)+HPO_4^{2-}(aq)$	6.23 x 10 ^{−8}	7.21
K _{a3}	$HPO_4^{2-}(aq)+H_2O(l) \longrightarrow H_3O^+(aq)+PO_4^{3-}(aq)$	2.20 x 10 ⁻¹³	12.7

Table 3: Dissociation constants of phosphoric acid, H₃PO₄ (triprotic)

Why is $K_{a3} < K_{a2} < K_{a1}$ in a polyprotic acid? It becomes more <u>difficult</u> to lose a <u>proton</u> from a <u>negatively charged ion</u>due to the <u>stronger electrostatic forces of attraction</u> between the anion and proton.

II. Base dissociation constant, K_b

Consider a weak base B in water:

 $B(aq) + H_2O(l) \longrightarrow BH^+(aq) + OH^-(aq)$

 $\mathbf{K}_{\mathbf{b}} = \frac{[\mathbf{BH}^+]_{eqm} [\mathbf{OH}^-]_{eqm}}{[\mathbf{B}]_{eqm}} \qquad \text{Units: mol dm}^{-3}$ $\mathbf{p}\mathbf{K}_{\mathbf{b}} = -\log_{10}\mathbf{K}_{\mathbf{b}}$

• *K*_b is <u>temperature</u> dependent.

Table 4: Dissociation constants of	bases in water at 25 °C
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Bases	Formula	$K_{\rm b}$ / mol dm ⁻³	р <i>К</i> ь	
Methylamine	CH ₃ NH ₂	4.54 x 10 ^{−4}	3.34	stronger base
Ammonia	NH₃	1.74 x 10 ^{−5}	4.76	weaker base

K_b and pK_b gives a good measurement of the strength of a base. The <u>higher the K_b</u> or the <u>lower the pK_b</u>, the <u>stronger the base</u>.



weak acid		weak base	
Ka	$\thickapprox \frac{\left[\mathrm{H}^{+}\right]^{2}}{\left[\mathrm{HA}\right]_{\mathrm{initial}}}$	$\boldsymbol{K}_{b} \approx \frac{[OH^{-}]^{2}}{[B]_{initial}}$	

∽ Self–Check: Q6

3.2 Degree of dissociation, α

The degree of dissociation, α , of a substance is defined as the fraction of substance that has undergone dissociation at equilibrium.

 $\alpha = \frac{\text{amount dissociated}}{\text{total initial amount}}$

For (monoprotic) strong acids / bases which dissociate completely, $\alpha = 1$

For (monoprotic) weak acids / bases which dissociate partially, $\alpha << 1$



Self–Check: Q7 and Q8



Differences between strong / weak vs concentrated / diluted acids

https://www.youtube.com/watch?v=RE3CKkkMIjo

4. ACID-BASE PROPERTIES OF SALT SOLUTIONS

- Salts can from solutions that are neutral, acidic or basic.
 - If the ions in the salt undergo <u>hydrolysis</u>, they will form acidic or alkaline solutions. The salt is termed acidic salt or basic salt.

If the ions in the salt do not undergo hydrolysis, the solution remains neutral. The salt is termed neutral salt.

• Neutral Salt of a Strong Base and Strong Acid

 $NaOH + HCl \longrightarrow NaCl + H_2O$

NaCl in water form hydrated ions:

NaCl \longrightarrow Na⁺(aq) + Cl⁻ (aq)

Na⁺ and Cl⁻ do not undergo hydrolysis in water. Thus, the resultant solution is neutral.

• Basic Salt of a Strong Base & Weak Acid

NaOH + CH₃COOH \longrightarrow CH₃COO⁻Na⁺ + H₂O

CH₃COONa in water form hydrated ions:

 $CH_3COONa \longrightarrow CH_3COO^{-}(aq) + Na^{+}(aq)$

Na⁺ does not hydrolyse

 CH_3COO^- is a <u>conjugate base</u> of weak acid CH_3COOH . CH_3COO^- partially hydrolyses in water to give OH^- . Thus, the resultant solution is <u>alkaline</u>. $CH_3COO^- + H_2O \longrightarrow CH_3COOH + OH^-$ A *K*_b expression can be written for this hydrolysis.

 $K_{\rm b} = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$

• Acidic Salt of a Strong Acid and Weak Base

 $HCl + NH_3 \longrightarrow NH_4Cl$

NH₄C*l* in water form hydrated ions:

 $NH_4Cl \longrightarrow NH_4^+(aq) + Cl^-(aq)$

Cl- does not hydrolyse

 NH_4^+ is a <u>conjugate acid</u> of weak base NH_3 . NH_4^+ partially hydrolyses in water to give H_3O^+ .

Thus, the resultant solution is <u>acidic</u>.

 NH_4^+ + H_2O \longrightarrow NH_3 + H_3O^+

A K_a expression can be written for this hydrolysis.

$$\kappa_{a} = \frac{[NH_{3}][H_{3}O^{+}]}{[NH_{4}^{+}]}$$

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Type of acids and bases mixed	Type of salt formed	Example	Solution formed
Strong base + strong acid	Neutral salt	NaC/	Neutral
Strong base + weak acid	Basic salt	CH₃COO⁻Na⁺	Alkaline
Strong acid + weak base	Acidic salt	NH₄⁺C <i>l</i> ⁻	Acidic

∽ Self–Check: Q9

Candidates should be able to:

• explain the choice of suitable indicators for acid-base titrations, given appropriate data in terms of the strengths of the acids & bases

5. INDICATORS

5.1 Indicators for acid–base titrations

Most indicators for acid-base titrations are themselves weak acids or bases that exhibit different colours in their dissociated and undissociated forms.

For example, let the formula of methyl orange be HIn At equilibrium,

At low pH, [H⁺] is high, the equilibrium position lies to the <u>left</u>. [HIn] >> [In⁻], colour observed: <u>red</u>

At high pH, [H⁺] is low, the equilibrium position lies to the <u>right</u>. [In⁻] >> [HIn] colour observed: <u>yellow</u>

End point of titration

The end point of a titration occurs when the indicator changes colour.

In the case of methyl orange, the end point occurs when it changes colour from red to orange, or yellow to orange, when the initial solution is acidic or basic respectively.

Choice of indicator

The choice of an indicator for a particular titration is based on its <u>pH range for colour change</u> (working range), which must <u>coincide with the region of rapid pH change in the titration curve</u> (vertical section of the titration curve). (See titration curve in next section)

Indicator	Co	lour	pH range for colour change /	
	Acid	Base	working range	
Methyl orange	Red	Yellow	3.2 – 4.4	
Screened methyl orange	Violet	Green	3.2 – 4.4	
Bromothymol blue	Yellow	Blue	6.0 - 7.6	
Phenolphthalein	Colourless	Pink	8.2 – 10.0	
Thymolphthalein	Colourless	Blue	9.3 – 10.5	

Table 5: Some common Acid–base Indicators

Titration	Type of salt formed	Indicator
strong acid – strong base	Neutral salt	Any indicator
weak acid – strong base	Basic salt	Phenolphthalein
		Thymolthalein
strong acid – weak base	Acidic salt	methyl orange
		screened methyl orange

6. MONOPROTIC ACID-BASE TITRATION CURVES

Consider a titration carried out by adding a base/acid from a burette to an acid/base in a conical flask.

The change in pH throughout the titration can be recorded using a pH meter as shown below.



If the pH is plotted against volume of base added, a titration curve will be obtained. Several possible types of titration curve are obtained.

- o Titration of strong acid against strong base / strong base against strong acid
- o Titration of weak acid against strong base / strong base against weak acid
- o Titration of strong acid against weak base / weak base against strong acid

6.1 Strong Acid – Strong Base titration

Consider a titration of 25.0 cm³ of 1.00 mol dm⁻³ HCl with 1.00 mol dm⁻³ NaOH.



Initial pH (before titration)

The conical flask contains only strong acid.

Eg:

pH = 0

When some strong acid is neutralised (before equivalence point)

- PH increases as the acid in the conical flask is gradually neutralised by the added base.
- The solution in the conical flask contains <u>remaining strong acid</u> and salt formed. The remaining strong acid determines the pH.

Eg: When 24.90 cm³ of 1.00 mol dm⁻³ NaOH is added to 25.0 cm³ of 1.00 mol dm⁻³ HCl

	NaOH(aq)	$+$ HCl(aq) \rightarrow	$NaCl(aq) + H_2O(l)$		
Initial amt / mol	0.0249	0.025			
Amt after neutralisation/ mol	0	0.025 - 0.0249 = 0.0001			
$[H^+] = [HCl] = \frac{0.0001}{49.9/1000} = 2.004 \times 10^{-3} \text{ mol dm}^{-3}$					
pH = 2.70					

At equivalence point

- The number of moles of base added just neutralise the number of moles of acid present.
- The solution in the conical flask contains only the neutral salt and water formed. Hence,
 pH = 7 at 25°C (from the dissociation of water at 25°C).

When excess strong base is added (after equivalence point)

All the strong acid in the conical flask has been completely neutralised at equivalence point. As excess strong base is still added, the solution in the conical flask contains <u>excess strong</u> <u>base</u> and the salt previously formed. The excess strong base determines the pH of the solution.

Eg: When 25.10 cm³ of 1.00 mol dm⁻³ NaOH is added to 25.0 cm³ of 1.00 mol dm⁻³ HCl,

NaOH(aq) +

HCl (aq) \longrightarrow NaCl (aq) + H₂O(l)

```
      Initial amt / mol
      0.0251
      0.025

      Amt after neutralisation/ mol
      0.0251 - 0.025 = 0.0001
      0

      [OH^-] = [NaOH] = \frac{0.0001}{0.0251 + 0.025} = 1.996 \times 10^{-3} \text{ mol dm}^{-3}
      pOH = -lg (1.996 \times 10^{-3})

      = 2.699
      pH = 14 - 2.699

      = 11.3
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The pH changes rapidly near the equivalence point from pH = 2.70 when 24.90 cm³ of NaOH was added, to pH = 11.3 when 25.10 cm³ of NaOH was added.

Suitable Indicator:

Any indicator which changes colour within the region of rapid pH change in the titration curve is suitable: <u>methyl orange / screened methyl orange / phenolphthalein /</u> <u>thymolphthalein</u>.



Note: HCl (aq) is in the burette. NaOH (aq) is in the conical flask.

6.2 Weak Acid – Strong Base titration



Consider a titration of 25.0 cm³ of 1.00 mol dm⁻³ CH₃COOH with 1.00 mol dm⁻³ NaOH.

<u>Initial pH</u>

The conical flask contains only weak acid.

Before equivalence point

The solution in the conical flask contains remaining weak acid and its conjugate base from the salt formed.

At equivalence point

- The number of moles of base added just neutralise the number of moles of acid present.
- The solution in the conical flask contains the <u>basic salt</u> and water formed.

After equivalence point

All the weak acid in the conical flask has been completely neutralised at equivalence point. As excess strong base is still added, the solution in the conical flask contains <u>excess strong</u> <u>base</u> and the salt previously formed. The excess strong base determines the pH of the solution.

Eg: When 25.10 cm³ of 1.00 mol dm⁻³ NaOH is added to 25.0 cm³ of 1.00 mol dm⁻³ CH₃COOH

	CH₃COOH	+ NaOH>	CH₃COONa	+ H ₂ O
Initial amt / mol	0.025	0.0251	0	
Amt after neutralisation/ mol	0	0.0251 - 0.025 = 0.0001	0.025	

 $[OH^{-}] = [NaOH] = \frac{0.0001}{0.0251+0.025} = 1.996 \times 10^{-3} \text{ mol dm}^{-3}$ pOH = 2.699 pH = 14 - 2.699 = 11.3

Suitable Indicator

Phenolphthalein / thymolphthalein

Consider a titration of 25.0 cm³ of 1.00 mol dm⁻³ NaOH with 1.00 mol dm⁻³ CH₃COOH.

 $CH_3COOH(aq) + NaOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$

Strong base only



weak acid + salt

6.3 Strong Acid – Weak Base titration



Consider a titration of 25.0 cm³ of 1.00 mol dm⁻³ HCl with 1.00 mol dm⁻³ NH₃.

Initial pH

The conical flask contains only strong acid.

Before equivalence point

The solution in the conical flask contains remaining strong acid and the salt formed. The remaining strong acid determines the pH of the solution.

At equivalence point

- The number of moles of base added just neutralise the number of moles of acid present.
- The solution in the conical flask contains the acidic salt and water formed.

After equivalence point

All the acid in the conical flask has been completely neutralised at equivalence point. As excess weak base is still added, the solution in the conical flask contains excess weak base and its conjugate acid from the salt previously formed.

Suitable Indicator

Methyl orange/screened methyl orange

Consider a titration of 25.0 cm³ of 1 mol dm⁻³ NH₃ with 1.00 mol dm⁻³ HCl.

 $HCl(aq) + NH_3(aq) \longrightarrow NH_4Cl(aq)$



Candidates should be able to:

- explain how buffer solutions control pH
- describe & explain their uses, including the role of H_2CO_3/HCO_3^- in controlling the pH in blood

7. BUFFER SOLUTIONS

7.1 Definition of buffer

A buffer solution is a solution whose <u>pH remains almost unchanged when a small amount of</u> <u>H⁺ or OH⁻ is added to it</u>. It maintains the pH by removing the added H⁺ or OH⁻. In general there are two types of buffer:

- a) Acidic buffer A solution of a weak acid and its conjugate base *E.g.* CH_3COOH and CH_3COONa
- b) Basic buffer- A solution of a weak base and its conjugate acid
 - *E.g* NH_3 and NH_4Cl

7.2 Action of a buffer

a) Acidic buffer

Consider a mixture of CH₃COOH – <u>weak acid</u> CH₃COONa – contains its conjugate base, CH₃COO⁻

On adding a small amount of H⁺

 H^+ + $CH_3COO^ \longrightarrow$ CH_3COOH

The small amount of added H⁺ is removed by the large amount of CH_3COO^- in the buffer. On adding small amount of OH^-

 $OH^- + CH_3COOH \longrightarrow CH_3COO^- + H_2O$

The small amount of added OH^- is removed by the <u>large amount of CH_3COOH </u> in the buffer.

Therefore pH remains almost unchanged.

b) Basic buffer

Consider a mixture of

NH₃ - weak base

NH₄Cl - contains its conjugate acid NH₄⁺

On adding a small amount of OH⁻

 $OH^- + NH_4^+ \longrightarrow NH_3 + H_2O$

The small amount of added OH^- is removed by the large amount of NH_4^{\pm} in the buffer. On adding a small amount of H^+

 $H^+ + NH_3 \longrightarrow NH_4^+$

The <u>small amount of added H⁺</u> is removed by the <u>large amount of NH_3 in the buffer</u>. Therefore, pH <u>remains almost unchanged</u>.

7.3 Uses of Buffer Solutions

Buffer solutions are used in situations where it is necessary to maintain pH. In many biological systems, a change in pH can have a great effect on the functioning of a cell. Therefore, they are usually buffered.

Blood buffer system of H₂CO₃ and HCO₃⁻

The pH of human blood must be maintained within a narrow range of 7.35 - 7.45. The buffer in blood consists of a mixture carbonates (HCO₃^{-/}H₂CO₃).

The cells in our body produce CO_2 as a product of aerobic respiration. CO_2 dissolves with water in the blood to form carbonic acid, which dissociates partially to give hydrogen carbonate.

 $H_2CO_3 \longrightarrow HCO_3^- + H^+$

H₂CO₃/HCO₃⁻ forms a buffer system in blood.

When $[H^+]$ in the blood increases (e.g. from the lactic acid which is produced during exercise), HCO_3^- reacts with H^+ , decreasing the $[H^+]$ and helps keep the pH almost constant.

 H^+ + $HCO_3^ \longrightarrow$ H_2CO_3

When [H⁺] in the blood decreases (e.g. from hyperventilation), the position of equilibrium shifts right, increasing the [H⁺] and helps keep the pH almost constant.

 $H_2CO_3 + OH^- \longrightarrow HCO_3^- + H_2O$

When $[OH^-]$ in the blood increases, H_2CO_3 reacts with OH^- , decreasing the $[OH^-]$ and helps keep the pH almost constant.

The blood buffer system

https://www.youtube.com/watch?v=r6UAEbhRXNI



∽ Self–Check: Q10