ACIDS & BASES - IONIC EQUILIBRIA

Acid-base theories

1. LEWIS
   - acid: electron pair acceptor
   - base: electron pair donor

   e.g. \( \text{H}_3\text{N}: \rightarrow \text{BF}_3 \rightarrow \text{H}_3\text{N}^+ \text{BF}_3^- \) see co-ordinate bonding

2. BRØNSTED-LOWRY
   - acid: proton donor
   - base: proton acceptor

   \[ \text{HCl} \rightarrow \text{H}^+\text{(aq)} + \text{Cl}^-\text{(aq)} \]
   \[ \text{NH}_3\text{(aq)} + \text{H}^+\text{(aq)} \rightarrow \text{NH}_4^+\text{(aq)} \]

Q.1 Classify the following according to Lewis theory and Brønsted-Lowry theory.

   - \( \text{H}_3\text{O}^+ \)
   - \( \text{BF}_3 \)
   - \( \text{H}_2\text{SO}_4 \)
   - \( \text{H}_2\text{O} \)
   - \( \text{NH}_4^+ \)
   - \( \text{CH}_3\text{NH}_2 \)

   B-L
   Lewis

Conjugate systems

- Acids are related to bases: \( \text{ACID} \leftrightarrow \text{PROTON} + \text{CONJUGATE BASE} \)
- Bases are related to acids: \( \text{BASE} + \text{PROTON} \leftrightarrow \text{CONJUGATE ACID} \)

For an acid to behave as an acid, it must have a base present to accept a proton...

\[ \text{HA} + \text{B} \leftrightarrow \text{BH}^+ + \text{A}^- \]

Q.2 Classify all the species in the following equations as acids or bases.

- \( a) \quad \text{HSO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{SO}_4^{2-} \)
- \( b) \quad \text{CH}_3\text{COOH} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^- \)
- \( c) \quad \text{CH}_3\text{COO}^- + \text{H}_2\text{O} \leftrightarrow \text{CH}_3\text{COOH} + \text{OH}^- \)
THE STRENGTH OF ACIDS

**Strong acids** completely dissociate (split up) into ions in aqueous solution

- **e.g.** HCl  \( \text{---->} \)  \( \text{H}^+\text{(aq)} \)  +  \( \text{Cl}^-\text{(aq)} \)  
- **e.g.** HNO\(_3\)  \( \text{---->} \)  \( \text{H}^+\text{(aq)} \)  +  \( \text{NO}_3^-\text{(aq)} \)  
- **e.g.** H\(_2\)SO\(_4\)  \( \text{---->} \)  2\( \text{H}^+\text{(aq)} \)  +  \( \text{SO}_4^{2-}\text{(aq)} \)

**MONOPROTIC** 1 replaceable H

**DIPROTIC** 2 replaceable H's

**Weak acids** partially dissociate into ions in aqueous solution  
- **e.g.** ethanoic acid  \( \text{CH}_3\text{COOH} \text{---->} \text{CH}_3\text{COO}^-\text{(aq)} \)  +  \( \text{H}^+\text{(aq)} \)

*Theory*

When a weak acid dissolves in water an equilibrium is set up

\[
\text{HA}_{(aq)} + \text{H}_2\text{O}_{(l)} \text{---->} \text{A}^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}
\]

The weaker the acid  
- the less it dissociates  
- the more the equilibrium lies to the left.

The relative strengths of acids can be expressed as \( K_a \) or \( pK_a \) values (see later).

The **dissociation constant** for the weak acid HA is

\[
K_a = \frac{[\text{H}^+\text{(aq)}][\text{A}^-\text{(aq)}]}{[\text{HA}_{(aq)}]} \text{ mol dm}^{-3}
\]

THE STRENGTH OF BASES

**Strong bases** completely dissociate into ions in aqueous solution  
- **e.g.** NaOH  \( \text{---->} \)  \( \text{Na}^+ \)  +  \( \text{OH}^-\text{(aq)} \)

**Weak bases** partially react to give ions in aqueous solution  
- **e.g.** ammonia (see below)

When a weak base dissolves in water an equilibrium is set up

\[
\text{NH}_3\text{(aq)} + \text{H}_2\text{O}_{(l)} \text{---->} \text{NH}_4^+\text{(aq)} + \text{OH}^-\text{(aq)}
\]

as in the case of acids  

it is more simply written

\[
\text{NH}_3\text{(aq)} + \text{H}^+\text{(aq)} \text{---->} \text{NH}_4^+\text{(aq)}
\]

The weaker the base  
- the less it dissociates  
- the more the equilibrium lies to the left.

The relative strengths of bases can be expressed as \( K_b \) or \( pK_b \) values.
HYDROGEN ION CONCENTRATION

Introduction
- hydrogen ion concentration determines the acidity of a solution
- hydroxide ion concentration determines the alkalinity
- for strong acids and bases the concentration of ions is very much larger than their weaker counterparts which only partially dissociate.

pH
hydrogen ion concentration can be converted to pH:  \[ \text{pH} = -\log_{10} \left( \frac{[H^+(aq)]}{\text{mol dm}^{-3}} \right) \]
to convert pH into hydrogen ion concentration \[ [H^+(aq)] = \text{antilog} (-\text{pH}) \]

pOH
An equivalent calculation for bases converts the hydroxide ion concentration to pOH \[ \text{pOH} = -\log_{10} \left( \frac{[OH^-}(aq)]}{\text{mol dm}^{-3}} \right) \]

Ionic Product of Water... \( K_w \)

Formula
Despite being covalent, water conducts electricity to a very small extent.

This is due to the slight ionisation...

\[ \text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \]
or

\[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq) \]

Applying the equilibrium law to the second we get \[ K_c = \frac{[H^+(aq)] [OH^-(aq)]}{[\text{H}_2\text{O}(l)]} \]

\[ \text{[ ] is the equilibrium concentration in mol dm}^{-3} \]

As the dissociation is small, the water concentration is very large compared with the dissociated ions and any changes to its value are insignificant; its concentration can be regarded as constant.

This “constant” is combined with the dissociation constant \( K_c \) to get a new constant \( K_w \).

\[ K_w = [H^+(aq)] [OH^-}(aq)] \text{ mol}^2 \text{ dm}^{-6} \]
\[ = 10^{-14} \text{ mol}^2 \text{ dm}^{-6} \text{ (at 25°C)} \]

Variation
The value of \( K_w \) varies with temperature because it is based on an equilibrium.

<table>
<thead>
<tr>
<th>Temperature / °C</th>
<th>0</th>
<th>20</th>
<th>25</th>
<th>30</th>
<th>60</th>
</tr>
</thead>
<tbody>
<tr>
<td>( K_w ) / 10^{-14} \text{ mol}^2 \text{ dm}^{-6}</td>
<td>0.11</td>
<td>0.68</td>
<td>1.0</td>
<td>1.47</td>
<td>5.6</td>
</tr>
</tbody>
</table>
The relationship between pH and pOH

Because H⁺ and OH⁻ ions are produced in equal amounts when water dissociates, their concentrations will be the same.

\[ [H^+] = [OH^-] = 10^{-7} \text{ mol dm}^{-3} \]

- take the equation for Kₘₜₜ
  \[ [H^+] [OH^-] = 10^{-14} \text{ mol}^2 \text{ dm}^{-6} \]
- take logs of both sides
  \[ \log[H^+] + \log[OH^-] = -14 \]
- multiply by minus
  \[ -\log[H^+] - \log[OH^-] = 14 \]
- change to pH and pOH
  \[ \text{pH} + \text{pOH} = 14 \]

(at 25°C)

N.B. As they are based on the position of equilibrium and that varies with temperature, the above values are only true if the temperature is 25°C (298K)

Neutral solutions may be regarded as those where \([H^+] = [OH^-]\).
Therefore a neutral solution is pH 7 only at a temperature of 25°C (298K)

The value of Kₘₜₜ is constant for any aqueous solution at the stated temperature

<table>
<thead>
<tr>
<th>[H⁺]</th>
<th>1 10⁻¹ 10⁻² 10⁻³ 10⁻⁴ 10⁻⁵ 10⁻⁶ 10⁻⁷ 10⁻⁸ 10⁻⁹ 10⁻¹⁰ 10⁻¹¹ 10⁻¹² 10⁻¹³ 10⁻¹⁴</th>
</tr>
</thead>
<tbody>
<tr>
<td>[OH⁻]</td>
<td>10⁻¹⁴ 10⁻¹³ 10⁻¹² 10⁻¹¹ 10⁻¹⁰ 10⁻⁹ 10⁻⁸ 10⁻⁷ 10⁻⁶ 10⁻⁵ 10⁻⁴ 10⁻³ 10⁻² 10⁻¹ 1</td>
</tr>
<tr>
<td>pH</td>
<td>0  1  2  3  4  5  6  7  8  9  10  11  12  13  14</td>
</tr>
</tbody>
</table>

strongly acidic – weakly acidic – neutral – weakly alkaline – strongly alkaline

Q.3 Convert the following pH values to [H⁺] ; 13 7.5 3.21 -0.6993
Convert the following [H⁺] values to pH ; 0.01 2.5 x 10⁻⁴ 1.1 x 10⁻¹³

BUFFER SOLUTIONS - Introduction

**Definition**
“Solutions which resist changes in pH when small quantities of acid or alkali are added.”

**Types**
- **Acidic Buffer** (pH < 7) made from a weak acid + its sodium or potassium salt ethanoic acid sodium ethanoate
- **Alkaline Buffer** (pH > 7) made from a weak base + its chloride ammonia ammonium chloride

**Uses**
- Standardising pH meters
- Buffering biological systems
CALCULATING THE pH AND pOH OF STRONG ACIDS AND BASES

- This is relatively easy because the species have completely dissociated
- Only needs to know the original concentration of the acid or base

**Example 1** Calculate the pH of 0.1M hydrochloric acid.

\[
\text{HCl (a strong monoprotic acid) is fully dissociated. \quad HCl \rightarrow H^+ (aq) + Cl^- (aq)}
\]

The \([H^+]\) is therefore the same as the original concentration of HCl i.e. 0.1M.

\[
pH = - \log_{10} \; [H^+] = - \log_{10} (10^{-1}) = 1
\]

ANS. 1

**Example 2** Calculate the pH of 0.001M sodium hydroxide.

\[
\text{Sodium hydroxide (a strong base) is fully dissociated. \quad Na^+OH^- \rightarrow Na^+ (aq) + OH^- (aq)}
\]

The \([OH^-]\) is therefore the same as the original concentration of NaOH i.e. 0.001M.

\[
pOH = - \log_{10} \; [OH^-] = - \log_{10} (10^{-3}) = 3
\]

\[
\text{and} \quad pH = 14 - pOH = 14 - 3 = 11
\]

ANS. 11

**Q.4** Calculate the pH and pOH of the following solutions.

- a) \(HCl; \; 0.1M, \; 0.5M\)

- b) \(H_2SO_4; \; 0.1M, \; 0.5M\)

- c) \(KOH; \; 0.1M\)

- d) \(NaOH; \; 2M, \; 0.0005M\)

- e) The solution remaining when 30 cm\(^3\) of 0.100M \(NaOH\) has been added to 20 cm\(^3\) of 0.200M \(HCl\)

- f) The solution remaining when 24.9 cm\(^3\) of 0.100M \(NaOH\) has been added to 25 cm\(^3\) of 0.100M \(HCl\)
CALCULATING THE pH AND pOH OF WEAK ACIDS AND BASES

• can’t be calculated by just knowing the concentration.
• need to know the extent of the ionisation (from $K_a$) and the original concentration.

The dissociation constant for a weak acid ($K_a$)

A weak monobasic acid (HA) dissociates in water thus.

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

Applying the equilibrium law we get

$$K_c = \frac{[H_3O^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}][H_2O_{(l)}]}$$

$[\ ]$ is the equilibrium concentration in mol dm$^{-3}$

For a weak acid (little dissociation) in dilute solution, the concentration of water is large compared with the dissociated ions and any changes to its value are insignificant so its concentration can be regarded as constant.

Combine this “constant” with the dissociation constant ($K_c$) to get a new one ($K_a$).

$$K_a = \frac{[H_3O^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}][H_2O_{(l)}]}$$

A simpler way to write it all out is

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

The dissociation constant $K_a$ is then

$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]}$$

The weaker the acid
• the less it dissociates
• the fewer ions you get
• the smaller $K_a$

The stronger the acid, the more the equilibrium lies to the right ... $K_a$ will increase.

pKa
• very weak acids have very small $K_a$ values
• it is easier to compare the strength as p$K_a$ values

The conversion is carried out thus...
- the units of $K_a$ are mol dm$^{-3}$

To convert pH into hydrogen ion concentration

$$K_a = \text{antilog} (-pK_a)$$
Calculating the pH of a weak acid

**Theory**

A weak monobasic acid (HA) dissociates in water

\[ \text{HA}(aq) \rightleftharpoons \text{H}^+(aq) + \text{A}^- (aq) \]

the dissociation constant (\(K_a\)) is

\[ K_a = \frac{[\text{H}^+(aq)] [\text{A}^- (aq)]}{[\text{HA}(aq)]} \text{ mol dm}^{-3} \]

**Assumptions**

The equation shows that, on dissociation, ions are formed in equimolar amounts.

\[ [\text{H}^+(aq)] = [\text{A}^- (aq)] \]

As the acid is weak, the dissociation is small.

The equilibrium concentration of HA can be approximated to be its original value.

the equation can be re-written ...

\[ [\text{H}^+(aq)]^2 = K_a [\text{HA}(aq)] \]

and

\[ [\text{H}^+(aq)] = \sqrt{K_a [\text{HA}(aq)]} \]

The pH can then be calculated ...

\[ \text{pH} = - \log_{10} [\text{H}^+(aq)] \]

**Q.6** Calculate the pH of the following solutions of weak acids . . .

a) 0.1M monobasic \((K_a = 2 \times 10^{-4} \text{ mol dm}^{-3})\)

b) 0.01M monobasic acid \((K_a = 7.5 \times 10^{-3} \text{ mol dm}^{-3})\)

**Q.5** Write out expressions for \(K_a\) for the following weak acids . . .

\(\text{HF}\) \hspace{1cm} \(\text{HCN}\)

\(\text{CH}_3\text{COOH}\) \hspace{1cm} \(\text{C}_6\text{H}_5\text{COOH}\)
ACID - BASE INDICATORS

General

Many indicators are weak acids and partially dissociate in aqueous solution

\[ \text{HIn}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{In}^-\text{(aq)} \]

\[ \text{red} \quad \leftrightarrow \quad \text{blue} \]

The un-ionised form (HIn) is a different colour to the anionic form (In\(^-\)).

Apply Le Chatelier’s Principle to predict any colour change

\[ \text{[ ] is the equilibrium conc. in mol dm}^{-3}\]

Example

In acid - increases [H\(^+\)] - equilibrium moves to the left to give red undissociated form

In alkali - increases [OH\(^-\)] - although OH\(^-\) ions don’t appear in the equation they remove H\(^+\) ions to form water.

- equilibrium will move to the right to produce a blue colour

Choice

- Must have an easily observed colour change.
- Must change quickly in the required pH range over the addition of ‘half’ a drop of reagent

**examples**

<table>
<thead>
<tr>
<th>pH</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
<th>11</th>
<th>12</th>
<th>13</th>
<th>14</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl Orange</td>
<td>pink</td>
<td>change</td>
<td>yellow</td>
<td></td>
<td></td>
<td></td>
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<td></td>
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<td></td>
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<td></td>
</tr>
<tr>
<td>Litmus</td>
<td>red</td>
<td>change</td>
<td>blue</td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>colourless</td>
<td>change</td>
<td>red</td>
<td></td>
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</tr>
</tbody>
</table>

Choice

To be useful, an indicator must... change over the “vertical” section of the curve where there is a large change in pH for the addition of a very small volume.

The indicator used depends on the pH changes around the end point - the indicator must change during the ‘vertical’ portion of the curve -

**Methyl Orange**

**Phenolphthalein**
**pH Curves**

**Types**

There are **four possible types of acid-base titration**; each has a characteristic curve.

- strong acid (HCl) v strong base (NaOH)
- weak acid (CH₃COOH) v strong alkali (NaOH)
- strong acid (HCl) v weak base (NH₃)
- weak acid (CH₃COOH) v weak base (NH₃)

In the examples below, alkali (0.1M) is added to 25cm³ of acid (0.1M).

*The end points need not be "neutral" due to the phenomenon of salt hydrolysis*

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**Q.7**

- Why can’t indicators be used for a weak acid - weak base titration?

- What alternative methods can be used?

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**Q.8**

Sketch and explain the pH curves you get when...

- 0.1M hydrochloric acid is added to 0.1M sodium carbonate
- 0.1M sodium hydroxide is added to 0.1M ethanedioc acid

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**Other curves**

Some titrations have more than one ‘vertical’ portion

- acid v. carbonate
- NaOH v. diprotic acids
CALCULATING THE pH OF ACID/ALKALI MIXTURES

The method depends on whether there are weak or strong acids and alkalis and which is in excess.

1. Calculate initial moles of H⁺ and OH⁻
2. Calculate which one is in excess
3. Calculate its concentration in the combined solution
4. Convert concentration to pH

**Example** Calculate the pH after 20 cm³ of 0.1M HCl is added to 25 cm³ of 0.1M NaOH

1. original moles of H⁺ = 0.1 x 20/1000 = 2 x 10⁻³ moles
2. original moles of OH⁻ = 0.1 x 25/1000 = 2.5 x 10⁻³ moles
   moles of excess OH⁻ = 5 x 10⁻⁴
3. final volume (20 + 25) = 45 cm³ = 0.045 dm³
4. [OH⁻] = 5 x 10⁻⁴ / 0.045 = 0.0111 mol dm⁻³
   pOH = 1.95
   pH = 14 - 1.95 = 12.05

**WEAK ACID**

**EXCESS**

**STRONG BASE**

1. Calculate initial moles of acid and alkali
2. Calculate the excess moles of OH⁻
3. Calculate the OH⁻ concentration in the combined solution
4. Convert concentration to pH

**Example** Calculate the pH after 22 cm³ of 0.1M CH₃COOH is added to 25 cm³ of 0.1M NaOH

1. original moles of H⁺ = 0.1 x 22/1000 = 2.2 x 10⁻³ moles
2. original moles of OH⁻ = 0.1 x 25/1000 = 2.5 x 10⁻³ moles
   moles of excess OH⁻ = 3 x 10⁻⁴
3. final volume (22 + 25) = 47 cm³ = 0.047 dm³
4. [OH⁻] = 3 x 10⁻⁴ / 0.047 = 6.38 x 10⁻³ mol dm⁻³
   pOH = 2.20
   pH = 14 - 2.20 = 11.80

**EXCESS**

**WEAK ACID**

**STRONG BASE**

1. Calculate initial moles of acid and alkali
2. Calculate the excess moles of acid
3. Calculate the moles of anion formed (same as the alkali used up)
4. Use the value of Kₐ for the weak acid to calculate the value of [H⁺]
5. Convert concentration to pH

**Example** Calculate the pH after 20 cm³ of 0.1M KOH is added to 25 cm³ of 0.1M CH₃COOH

1. original moles of CH₃COOH = 0.1 x 25/1000 = 2.5 x 10⁻³ moles
2. original moles of KOH = 0.1 x 20/1000 = 2.0 x 10⁻³ moles
3. excess moles CH₃COOH = 5.0 x 10⁻⁴
3. moles of CH₃COO⁻ formed = moles of H⁺ used = 2.0 x 10⁻³
4. Kₐ for CH₃COOH = [H⁺][CH₃COO⁻] / [CH₃COOH] = 1.7 x 10⁻⁵ at 25°C
   \[\frac{1.7 \times 10^{-5} \times 5 \times 10^{-4}}{2.0 \times 10^{-3}} = 4.25 \times 10^{-6}\text{ mol dm}^{-3}\]
5. pH = -log [H⁺] = 5.37

YOU ONLY NEED TO PUT IN THE MOLAR RATIO (NOT THE CONCENTRATIONS) BECAUSE THE
BUFFER SOLUTIONS - Action

**Acid buffer** It is essential to have a weak acid for an equilibrium to be present so that ions can be removed and produced. The dissociation is small and there are few ions.

\[
\text{CH}_3\text{COOH}_{(aq)} \rightleftharpoons \text{CH}_3\text{COO}^-_{(aq)} + \text{H}^+_{(aq)}
\]

relative concs.  
HIGH  LOW  LOW

*Adding acid* Any \( \text{H}^+_{(aq)} \) is removed by reacting with \( \text{CH}_3\text{COO}^- \) ions to form \( \text{CH}_3\text{COOH} \). Unfortunately, the concentration of these ions is small and only a few \( \text{H}^+_{(aq)} \) can be “mopped up”.

To build up the concentration of \( \text{CH}_3\text{COO}^- \) ions, sodium ethanoate is added.

*Adding alkali* This adds \( \text{OH}^- \) ions. Although they do not appear in the equation, they react with \( \text{H}^+ \) ions

\[
\text{H}^+_{(aq)} + \text{OH}^-_{(aq)} \rightleftharpoons \text{H}_2\text{O}(l)
\]

Removal of \( \text{H}^+ \) from the system means that, according to Le Chatelier’s Principle, more \( \text{CH}_3\text{COOH} \) will dissociate to form ions to replace those being removed.

**Summary** For an acidic buffer solution one needs ...

large \( [\text{CH}_3\text{COOH}_{(aq)}] \) to be available to dissociate into \( \text{H}^+_{(aq)} \) when alkali is added

large \( [\text{CH}_3\text{COO}^-_{(aq)}] \) to be able to remove \( \text{H}^+_{(aq)} \) as it is added

This situation can’t exist if only acid is present so a mixture of the acid and salt is used.

The weak acid provides the equilibrium and the large \( \text{CH}_3\text{COOH}_{(aq)} \) concentration.

The sodium salt provides the large \( \text{CH}_3\text{COO}^-_{(aq)} \) concentration.

\[ \therefore \text{ One uses ethanoic acid (a weak acid) + one of its salts (sodium ethanoate) } \]

**Alkaline buffer** Very similar but is based on the equilibrium surrounding a weak base e.g. ammonia

\[
\text{NH}_3_{(aq)} + \text{H}_2\text{O}(l) \rightleftharpoons \text{OH}^-_{(aq)} + \text{NH}_4^+_{(aq)}
\]

relative concs.  
HIGH  LOW  LOW

but one needs ;  

a large conc. of \( \text{OH}^-_{(aq)} \) to react with any \( \text{H}^+_{(aq)} \) added

a large conc of \( \text{NH}_4^+_{(aq)} \) to react with any \( \text{OH}^-_{(aq)} \) added

as with the acid buffer one needs to greatly increase the concentration of ammonium ions by adding a solution of an ammonium salt.

\[ \text{One uses ammonia (a weak base) + one of its salts (ammonium chloride) } \]

Adding acid

Adding alkali
SALT HYDROLYSIS

Introduction  Many salts dissolve in water to produce solutions which are not neutral. This is because the ions formed react with the hydroxide and hydrogen ions formed when water dissociates. There are four distinct systems. All dissociated ions are aqueous ions. When mixed, the ions of strong acids and bases remain apartions of weak acids and bases associate.

Salts of strong acids and strong bases

Sodium chloride completely dissociates
Na⁺ Cl⁻ ——> Na⁺ + Cl⁻

Water partially dissociates
H₂O ⇌ H⁺ + OH⁻

Pairs of ions stay apart as they are from strong acids and bases.

\[ [\text{H}^+] = [\text{OH}^-] \quad \text{NEUTRAL} \]

Salts of strong acids and weak bases

Ammonium chloride completely dissociates
NH₄⁺ Cl⁻ ——> Cl⁻ + NH₄⁺

Water partially dissociates
H₂O ⇌ H⁺ + OH⁻

Some ammonium and hydroxide ions associate to form ammonia (a weak base) and water.

\[ [\text{H}^+] > [\text{OH}^-] \quad \text{ACIDIC} \]

Salts of weak acids and strong bases

Sodium ethanoate completely dissociates
CH₃COO⁻ Na⁺ ——> CH₃COO⁻ + Na⁺

Water partially dissociates
H₂O ⇌ H⁺ + OH⁻

Some ethanoate and hydrogen ions associate to form ethanoic acid (a weak acid).

\[ [\text{OH}^-] > [\text{H}^+] \quad \text{ALKALINE} \]

Salts of weak acids and weak bases

Ammonium ethanoate completely dissociates
CH₃COO⁻ NH₄⁺ ——> CH₃COO⁻ + NH₄⁺

Water partially dissociates
H₂O ⇌ H⁺ + OH⁻

Some ethanoate and some hydrogen ions associate to form ethanoic acid and some ammonium ions combine with some hydroxide ions to produce ammonia. The pH depends on the relative values of the two dissociation constants but the solution will be . . .

APPROXIMATELY NEUTRAL