IONIC EQUILIBRIUM
LESSON OUTCOME

- EXPLAIN THE THEORY OF ACID & BASES
- IDENTIFY ACID, BASES CONJUGATE ACID & CONJUGATE BASES
- Calculating pH for acid & bases
- Determination of degree @ % dissociation/ionization.
Theory of Acids and bases

Three theories used to explain acids and bases

- Bronsted-Lowry Theory
- Arrhenius Theory
- Lewis Theory
Arrhenius Theory

- **Acid**
  - Dissociates in water to produce hydrogen ions (H$^+$)
- **Base**
  - Dissociates in water to produce hydroxide ions, (OH$^-$)
- **Examples:**
  - HNO$_3$ (aq) $\rightarrow$ H$^+$ (aq) + NO$_3^-$ (aq)
  - NaOH (aq) $\rightarrow$ Na$^+$ (aq) + OH$^-$ (aq)
- **The limitations of Arrhenius theory**
  - Restricted to aqueous solution
  - Does not account for the basicity of substances, especially ammonia
Brønsted – Lowry Theory

- Acid
  - Is a proton donor (a substance that can transfer $\text{H}^+$)
- Base
  - Is a proton acceptor
- So, an acid-base reaction is a proton-transfer reaction
- The theory can explain the acidity and basicity of weak acid and base. This compound partially dissociates in water. Thus, the symbol ($\leftrightarrow$) is used to show that the reaction is reversible. When a Bronsted –Lowry acid is placed in water, it donates a proton to the solvent. Hence water act as Bronsted –Lowry base.
In the reverse reaction, the hydroxonium ion acts as a proton donor (acid), and the A⁻ ion acts as a proton acceptor (base).

Example of Bronsted Lowry acids

- CH₃COOH
- CH₃COOH + H₂O ↔ H₃O⁺ + CH₃COO⁻
A Bronsted–Lowry base such as ammonia accepts a proton from the solvent (water), which acts as a Bronsted–Lowry acid.

\[
\text{NH}_3 \quad + \quad \text{H}_2\text{O} \quad \leftrightarrow \quad \text{OH}^- \quad + \quad \text{NH}_4^+
\]

In the reverse reaction, the hydroxide ion acts as a proton acceptor (base), and the NH\(_4^+\) ion acts as a proton donor (acid).

Substances such as water, ions such as hydrogen sulphate ion are both Bronsted Lowry acids and bases. They are known as ampholytes and show amphiprotic behaviour.

\[
\text{HSO}_4^- \quad + \quad \text{H}_2\text{O} \quad \leftrightarrow \quad \text{H}_3\text{O}^+ \quad + \quad \text{SO}_4^{2-}
\]

\[
\text{HSO}_4^- \quad + \quad \text{H}_2\text{O} \quad \leftrightarrow \quad \text{OH}^- \quad + \quad \text{H}_2\text{SO}_4
\]
Exercises

- For each of the following identify the acid, base, conjugate acid and conjugate base:

  - $\text{HCl} + \text{H}_2\text{PO}_4^- \iff \text{H}_3\text{PO}_4 + \text{Cl}^-$
  - $\text{C}_6\text{H}_5\text{OH} + \text{H}_2\text{O} \iff \text{C}_6\text{H}_5\text{O}^- + \text{H}_3\text{O}^+$
  - $\text{NH}_2\text{OH} + \text{H}_2\text{O} \iff \text{NH}_2\text{OH}_2 + \text{OH}^-
  - $\text{NH}_3 + \text{HSO}_4^- \iff \text{NH}_4^+ + \text{SO}_4^{2-}$
Lewis Theory

- Lewis defined an acid as an electron pair acceptor and a base as an electron pair donor.

\[
\begin{align*}
H^+ + \text{OH}^- &\rightarrow H-O-H \\
\text{acid} &\quad \text{base}
\end{align*}
\]

- Example of reaction between ammonia, $\text{NH}_3$ and boron trifluoride, $\text{BF}_3$

\[
\text{BF}_3 + : \text{NH}_3 \rightarrow F_3\text{B} : \text{NH}_3
\]
Relative Strength of Bronsted – Lowry Acids and Bases

• **Degree of Dissociation**
  
  • A strong acid will have a higher tendency to dissociate to form the hydroxonium ion as compared to a weaker acid. Hence the higher the degree of dissociation of acids, the stronger the acid.

  • The acids in order of increasing strength is:
    \[
    \text{CH}_3\text{COOH} < \text{HNO}_2 < \text{HSO}_4^- < \text{HCl}
    \]

  • The strength of bases depends on their ability to form the OH\(^-\) ions in aqueous solution. A strong bases will have a higher ‘degree of dissociation’ than weak bases

  • The bases in order of increasing strength is:
    \[
    \text{NO}_2^- < \text{CH}_3\text{COO}^- < \text{NH}_3
    \]
Acid/Base Dissociation Constant ($K_a$ and $K_b$)

- Consider a weak monoprotic acid, HA, which dissociates partially in water:
  \[ HA + H_2O \rightleftharpoons H_3O^+ + A^- \]

  At equilibrium,
  \[ \frac{[H_3O^+][A^-]}{[HA]} = K_a \]

  - The higher the value of $K_a$ (that is the equilibrium position is more towards the right hand side) the stronger the acid

- Weak base, B, which dissociates partially in water:
  \[ B + H_2O \rightleftharpoons HB^+ + OH^- \]

  At equilibrium,
  \[ \frac{[HB^+][OH^-]}{[B]} = K_b \]

  - The higher the value of $K_b$, the stronger the base
**pK\textsubscript{a} and pK\textsubscript{b}**

- Is another way of expressing the acid/base dissociation constant

\[
pK_a = -\log K_a \\
pK_b = -\log K_b
\]

- The larger the value of \(K_a\) or \(K_b\), the smaller the value of \(pK_a\) or \(pK_b\).

- **Exercise:**
  - Calculate the \(pK_b\) for the bases.

<table>
<thead>
<tr>
<th>Acid</th>
<th>NH\textsubscript{3}</th>
<th>CH\textsubscript{3}COO</th>
<th>NO\textsubscript{2}^-</th>
</tr>
</thead>
<tbody>
<tr>
<td>(K_b/\text{mol dm}^{-3})</td>
<td>(1.8 \times 10^{-5})</td>
<td>(5.6 \times 10^{-10})</td>
<td>(2.2 \times 10^{-11})</td>
</tr>
</tbody>
</table>

- Arrange the bases in order of increasing strength
Ostwald Dilution Law and Dissociation Constant

- Consider a weak acid, HA, of concentration C \text{ mol dm}^{-3} \text{ and degree of dissociation of } \alpha

\[
\begin{array}{c|c|c|c}
\text{Initial} & \text{HA} & \text{H}^+ & \text{A}^- \\
\hline
C & 0 & 0 \\
\text{Equilibrium} & C(1-\alpha) & C\alpha & C\alpha \\
\end{array}
\]

At equilibrium: \( Ka = \frac{(C\alpha)(C\alpha)}{C(1-\alpha)} \)

HOWEVER, \( \alpha \) is very small, hence \( 1-\alpha \approx 1 \)

\( Ka = c \alpha^2 \) \hspace{1cm} Since \( H^+ = C\alpha \)

\[
\alpha = \sqrt{\frac{Ka}{C}}
\]

For weak base

\( B + \text{H}_2\text{O} \leftrightarrow \text{HB}^+ + \text{OH}^- \)

\[
[\text{OH}^-] = \sqrt{K_bC}
\]
Ionic Product of Water

• The ion product of water
  • Hydrogen ion (H\textsuperscript{+}) is key = It indicates the acidity @ basicity of the solution
  • Equilibrium constant for the autoionization of water

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

\[ K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} \]

• Since, the degree of dissociation of water is extremely small, the concentration of water can be considered constant

\[ K_c[H_2O] = K_w = [\text{H}^+][\text{OH}^-] \]

\[ K_w = (1.0 \times 10^{-7} \text{ mol dm}^{-3})(1.0 \times 10^{-7} \text{ mol dm}^{-3}) = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6} \]
The ionic product of water can be expressed in terms of pKw, where

\[ pK_w = -\log K_w \]

\[ = -\log(1.0 \times 10^{-14}) = 14 \]

From the expression:

\[ K_w = [H^+][OH^-] \]

Another expression can be derived by taking the (- log) of the K_w expression

\[ pK_w = POH + pH \]

Thus, if the pOH of the solution is known, its pH value can be calculated from the above relationship.
Measurement of acidity and basicity

- The pH of solution is defined as *negative logarithm of the hydrogen ion concentration (in mol/L)*

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

solution is

- Neutral if, \([H^+] = [OH^-] ; \text{pH} = 7\)
- Acidic if, \([H^+] > [OH^-] ; \text{pH} < 7\)
- Basic if, \([H^+] < [OH^-] ; \text{pH} > 7\)

- The pOH of solution is defined as *negative logarithm of the hydroxide ion concentration (in mol/L)*

\[ \text{pOH} = -\log [OH^-] \]

\[ \text{pH} + \text{pOH} = 14.00 \]
Exercise

• Calculate the pH of the following solutions at 298 K

a) 0.10 mol dm$^{-3}$ NaOH
b) 0.50 mol dm$^{-3}$ NH$_3$ (K$_b$ = 1.8x10$^{-5}$ mol dm$^{-3}$)
c) 1.5 mol dm$^{-3}$ Ba(OH)$_2$ \{Ba(OH)$_2$ is a strong base\}
d) 2.0 mol dm$^{-3}$ H$_2$SO$_4$

• Answers

a) 13 \hspace{1cm} b) 11.48 \hspace{1cm} c) 14.48 \hspace{1cm} d) – 0.602
Calculate the value of both pH and pOH of the following solution.

a) 0.020 M HCl

b) 0.0050 M NaOH

c) A blood specimen containing $7.2 \times 10^{-8}$ mole/L of hydrogen ion. Is the blood specimen slightly acidic or basic?

d) 0.00035 M Ba(OH)$_2$ where this compound has to be considered to be 100% dissociated.
Relationship between \( pK_a \), \( pK_b \) and \( pK_w \)

- Consider an aqueous of ethanoic acid:
  \[
  \text{CH}_3\text{COOH} + \text{H}_2\text{O} \quad \leftrightarrow \quad \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \\
  K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} 
  \]

- Consider an aqueous solution containing the ethanoate ion (which is the conjugate base of \( \text{CH}_3\text{COOH} \)):
  \[
  \text{CH}_3\text{COO}^- + \text{H}_2\text{O} \quad \leftrightarrow \quad \text{CH}_3\text{COOH} + \text{OH}^- \\
  K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^- ]} 
  \]
Multiplying $K_a$ and $K_b$:

$$K_a \times K_b = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \times \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_2\text{COO}^-]}$$

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

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

Taking $-\log$ throughout:

$$-\log K_a + - \log K_b = - \log K_w$$

Hence, for an acid/base conjugate pair:

$$pK_a + pK_b = pK_w$$
example

• Calculate the pH of 0.10 mol dm\(^{-3}\) solution of sodium ethanoate. (Ka For CH\(_3\)COOH= 1.8\(\times\)10\(^{-5}\) moldm\(^{-3}\))

**Answer**

\[
\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^- \\
\]

\([\text{OH}^-] = \text{K}_b \times C\]

\[\text{K}_b \text{ for CH}_3\text{COO}^- = 1.0 \times 10^{-14} = 5.55 \times 10^{-10} \text{ mol dm}^{-3}\]

\[1.8 \times 10^{-5}\]

\[\text{[OH}^-] = 5.55 \times 10^{-10} \times 0.1 = 7.45 \times 10^{-6}\]

\[\text{pOH} = \log 7.45 \times 10^{-6} = 5.13\]

\[\text{pH} = 14 - 5.13\]

\[= 8.87\]
Acid-Base Titration

- **Purpose**: to determine the volume of two solutions (acid and base) which will exactly neutralize each other.

- Most acids, bases and salts are colorless in water. Therefore, an indicator is necessary to pin-point the end point of the titration.

**Acid-Base Indicators**

- Substances which change color according to the pH of the solution to which they are added.

- **examples**

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Color in acid</th>
<th>Color in base</th>
</tr>
</thead>
<tbody>
<tr>
<td>phenolphthalein</td>
<td>colorless</td>
<td>purple</td>
</tr>
<tr>
<td>Methyl orange</td>
<td>red</td>
<td>yellow</td>
</tr>
<tr>
<td>Methyl red</td>
<td>yellow</td>
<td>red</td>
</tr>
</tbody>
</table>
Equivalent point and End point of Acid-Base Titration

• The equivalent point is when the amount of acid and base present exactly neutralises one another. At this point, neither acid nor alkali are present in excess. The solution consists of salt and water only.

• The end point of titration is when the indicator changes color.

• The indicator must change color at the equivalent point, not before or after.
**Suitable Indicator**

<table>
<thead>
<tr>
<th>Titration</th>
<th>Suitable Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strong acid + Weak base</td>
<td>Methyl Orange</td>
</tr>
<tr>
<td>Weak acid + Strong base</td>
<td>Phenolphthalien</td>
</tr>
<tr>
<td>Strong acid + Strong base</td>
<td>Methyl Orange or Phenolphthalien</td>
</tr>
<tr>
<td>Weak acid + Weak base</td>
<td>No suitable indicator</td>
</tr>
</tbody>
</table>
LESSON OUTCOME

- IDENTIFY TYPE OF SALTS
- CALCULATE pH of salts
- Define and example of buffer
- Identify types of buffer
- Calculate pH of buffer
- Solve the problem for buffer system
Neutral Solutions:

Salt formed from a strong acid and a strong base is dissolved in water.

Example: \( \text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} \)

Salts containing an alkali metal or alkaline earth metal ion (except \( \text{Be}^{2+} \)) \textbf{and} the conjugate base of a \textbf{strong} acid (\textit{e.g.} \( \text{Cl}^-, \text{Br}^-, \text{and NO}_3^- \)).

\[
\text{NaCl} (s) + \text{H}_2\text{O} \rightarrow \text{Na}^+ (aq) + \text{Cl}^- (aq)
\]

\( \text{Na}^+ \) and \( \text{Cl}^- \) are not hydrolysed by water. There is no production of extra \( \text{H}^+ \) or \( \text{OH}^- \) ions. Hence, the solution is \textbf{neutral}.

Basic Solutions:

Salts derived from \textbf{a strong base and a weak acid}.

\[
\text{NaCH}_3\text{COOH} (s) + \text{H}_2\text{O} \rightarrow \text{Na}^+ (aq) + \text{CH}_3\text{COO}^- (aq)
\]

\[
\text{CH}_3\text{COO}^- (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{CH}_3\text{COOH} (aq) + \text{OH}^- (aq)
\]
Acidic solution.

An example is:

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

**Weak base Strong acid**

In water, ammonium chloride dissociates completely into ammonium ions and chloride ions:

\[
\text{NH}_4\text{Cl (s)} \overset{\text{H}_2\text{O}}{\rightarrow} \text{NH}_4^+ (aq) + \text{Cl}^- (aq)
\]

The ammonium ion, \(\text{NH}_4^+\), acts as a Bronsted-Lowry acid and donates a proton to water:

\[
\text{NH}_4^+ + \text{H}_2\text{O(aq)} \leftrightarrow \text{NH}_3 (aq) + \text{H}_3\text{O}^+ (aq)
\]

On the other hand, the \(\text{Cl}^-\) ions do not react with water:

\[
\text{Cl}^- (aq) + \text{H}_2\text{O (l)} \leftrightarrow \text{HCl (aq)} + \text{OH}^- \quad (\text{this reaction does not take place})
\]

Because \(\text{HCl (a strong acid)}\) formed must dissociate completely to \(\text{H}^+\) and \(\text{Cl}^-\) ions.

Salts with small, highly charged metal cations (e.g. \(\text{Al}^{3+}\), \(\text{Cr}^{3+}\), and \(\text{Be}^{2+}\)) and the conjugate base of a strong acid.

\[
\text{Al(H}_2\text{O)}_6 (aq)^{3+} \leftrightarrow \text{Al(OH)(H}_2\text{O)}_5 (aq)^{2+} + \text{H}^+ (aq)
\]
Acid Hydrolysis of Al$^{3+}$

\[ \text{Al(H}_2\text{O)}_6^{3+} + \text{H}_2\text{O} \rightarrow \text{Al(OH)(H}_2\text{O)}_5^{2+} + \text{H}_3\text{O}^+ \]
Acid-Base Properties of Salts

Solutions in which both the cation and the anion hydrolyze:

- $K_b$ for the anion $> K_a$ for the cation, solution will be basic
- $K_b$ for the anion $< K_a$ for the cation, solution will be acidic
- $K_b$ for the anion $\approx K_a$ for the cation, solution will be neutral
# Acid-Base Properties of Salts

<table>
<thead>
<tr>
<th>Type of Salt</th>
<th>Examples</th>
<th>Ions That Undergo Hydrolysis</th>
<th>pH of Solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cation from strong base; anion from strong acid</td>
<td>NaCl, KI, KNO₃, RbBr, BaCl₂</td>
<td>None</td>
<td>≈ 7</td>
</tr>
<tr>
<td>Cation from strong base; anion from weak acid</td>
<td>CH₃COONa, KNO₂</td>
<td>Anion</td>
<td>&gt; 7</td>
</tr>
<tr>
<td>Cation from weak base; anion from strong acid</td>
<td>NH₄Cl, NH₄NO₃</td>
<td>Cation</td>
<td>&lt; 7</td>
</tr>
<tr>
<td>Cation from weak base; anion from weak acid</td>
<td>NH₄NO₂, CH₃COONH₄, NH₄CN</td>
<td>Anion and cation</td>
<td>&lt; 7 if $K_b &lt; K_a$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>≈ 7 if $K_b \approx K_a$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>&gt; 7 if $K_b &gt; K_a$</td>
</tr>
<tr>
<td>Small, highly charged cation; anion from strong acid</td>
<td>AlCl₃, Fe(NO₃)₃</td>
<td>Hydrated cation</td>
<td>&lt; 7</td>
</tr>
</tbody>
</table>
Hydrolysis

- Remembering the following equations must hold simultaneously

\[ K_w = \left[ H^+ \right] \times \left[ OH^- \right] \]

\[ K_a = \frac{\left[ H^+ \right] \times \left[ Ac^- \right]}{\left[ HAc \right]} \] \text{ and } \[ K_b = \frac{\left[ NH_4^+ \right] \times \left[ OH^- \right]}{\left[ NH_4OH \right]} \]

- Also, it can be shown that

\[ K_h = \frac{K_w}{K_a \times K_b} \]

\[ pK_h = pK_w - pK_a + pK_b \]
Buffer System

Is a solution of a weak acid and its salt to form **acidic buffer solution** or weak base and its salt to form **basic buffer solution**.

A buffer solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.
Acidic Buffer Solution

• With pH between 4 and 7
• Can be prepared from a weak acid and its conjugate base.
• Example; an aqueous mixture of ethanoic acid and sodium ethanoate.

\[
\begin{align*}
\text{CH}_3\text{COONa} \text{ (aq)} & \rightarrow \text{CH}_3\text{COO}^- \text{ (aq)} + \text{Na}^+ \text{ (aq)} \\
\text{CH}_3\text{COOH} \text{ (aq)} & \rightarrow \text{CH}_3\text{COO}^- \text{ (aq)} + \text{H}^+ \text{ (aq)}
\end{align*}
\]

When a little acid is added, the above equilibrium will shift to the left to mop up the added of H\(^+\) ion

\[
\text{CH}_3\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{COOH}
\]

When a little base is added, it is removed by the following reaction

\[
\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}
\]

Added H\(^+\) is mopped up by the conjugate base

Added OH\(^-\) is mopped up by the acid
**Basic Buffer Solution**

- With pH between 7 and 10
- Can be prepared from a weak base and its conjugate acid.
- Example: an aqueous mixture of ammonia and ammonium chloride

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

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

When a little base is added, the above equilibrium will shift to the left to mop up the added OH\(^-\) ion

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

When a little acid is added, it is removed by the following reaction

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

Added OH\(^-\) is mopped up by the conjugate acid

Added H\(^+\) is mopped up by the base
Exercise

- An aqueous mixture of carbonic acid and sodium hydrogen carbonate can function as a buffer solution. Write a balanced equation to show how the system works.
- Answer:

\[
\begin{align*}
H_2CO_3(aq) + H_2O(l) & \leftrightharpoons H_3O^+ + HCO_3^- \\
NaHCO_3 & \rightarrow Na^+ + HCO_3^-
\end{align*}
\]

On the addition of a little acid:

\[
H^+ + HCO_3^- \rightarrow H_2CO_3
\]

On the addition of a little base:

\[
OH^- + H_2CO_3 \rightarrow HCO_3^- + H_2O
\]
pH of Buffer Solutions

Consider an aqueous mixture of a weak acid, HA and its salt MA.

\[ \text{MA} \xrightarrow{100\%} \text{M}^+ + \text{A}^- \]
\[ \text{HA} \leftrightarrow \text{H}^+ + \text{A}^- \]

Applying the equilibrium law to the above system

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

The [A\textsuperscript{-}] at equilibrium can be assumed to be the concentration of MA which dissociates 100%.

The [HA] at equilibrium can be assumed to be equal to the original concentration of HA (since the degree of dissociation is very small).
• Rearranging the equation

\[ [H^+] = K_a \frac{[\text{Acid}]}{[\text{salt}]} \]

• Taking \(-\log\) throughout:

\[ \text{pH} = pK_a - \log \frac{[\text{Acid}]}{[\text{salt}]} \]

• Similarly for a basic buffer solution:

\[ [\text{OH}^-] = K_b \frac{[\text{base}]}{[\text{salt}]} \]

• Taking \(-\log\) throughout:

\[ \text{pOH} = pK_b - \log \frac{[\text{base}]}{[\text{salt}]} \]
Examples

• Calculate the pH of buffer solution consisting of 0.10 M CH₃COOH and 0.25 M CH₃COONa. (Kₐ for CH₃COOH = 1.8 x 10⁻⁵ M)

• Answer

Using: pH = pKₐ - \log \frac{[Acid]}{[salt]}

\begin{align*}
&= 4.74 - \log 0.1 = 5.14 \\
&\quad - \log 0.25
\end{align*}

• 4.28g of ammonium chloride was added to 250 cm³ of 0.50 mol dm⁻³ ammonia solution. Calculate the pH of the resulting solution. (Kₜ for ammonia = 1.8 x 10⁻⁵ mol dm⁻³)

• Answer = 9.45
Calculating the pH of a buffer solution

Calculate the pH of a buffer whose [HA] is 0.1 mol dm\(^{-3}\) and [A\(^-\)] of 0.1 mol dm\(^{-3}\).

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

re-arrange

\[
[H^+(aq)] = \frac{[HA(aq)] \times K_a}{[A^- (aq)]}
\]

from information given

\[
[A^-] = 0.1 \text{ mol dm}^{-3}
\]
\[
[HA] = 0.1 \text{ mol dm}^{-3}
\]

If the \(K_a\) of the weak acid HA is 2 \times 10^{-4} \text{ mol dm}^{-3}.

\[
[H^+(aq)] = \frac{0.1 \times 2 \times 10^{-4}}{0.1} = 2 \times 10^{-4} \text{ mol dm}^{-3}
\]

\[
\text{pH} = -\log_{10} [H^+(aq)] = 3.699
\]
Calculate the pH of the solution formed when 500 cm$^3$ of 0.1 mol dm$^{-3}$ of weak acid HX is mixed with 500 cm$^3$ of a 0.2 mol dm$^{-3}$ solution of its salt NaX. $K_a = 4 \times 10^{-5}$ mol dm$^{-3}$.

$$K_a = \frac{[H^+(aq)] [X^-(aq)]}{[HX(aq)]}$$

re-arrange $$[H^+(aq)] = \frac{[HX(aq)] K_a}{[X^-(aq)]}$$

The solutions have been mixed; the volume is now 1 dm$^3$

therefore $$[HX] = 0.05 \text{ mol dm}^{-3} \text{ and } [X^-] = 0.10 \text{ mol dm}^{-3}$$

Substituting $$[H^+(aq)] = \frac{0.05 \times 4 \times 10^{-5}}{0.1} = 2 \times 10^{-5} \text{ mol dm}^{-3}$$

$$pH = -\log_{10} [H^+(aq)] = 4.699$$
Preparation of Buffer Solution

1. Choose a suitable weak acid (or weak base) whose $\text{pK}_a$ (or $\text{pK}_b$) is almost equal to the required pH of the buffer solution
2. Calculate the ratio of the concentration of the acid (or base) to that concentration of the salt
3. Add the required amount of salt to the acid (or base) as determined by step 2 above.
1. Explain how you would prepare a buffer solution of pH = 5.2

Answer

Ethanoic acid with $pK_a = 4.74$ is chosen.

Using: $\text{pH} = pK_a - \log \frac{[\text{Acid}]}{[\text{salt}]}$

\[
5.2 = 4.74 - \log \frac{[\text{Acid}]}{[\text{salt}]} \quad \Rightarrow \quad \frac{[\text{Acid}]}{[\text{salt}]} = 0.35
\]

- Dissolve a mixture containing ethanoic acid and sodium ethanoate in the mole ratio 0.35:1 in 1 dm$^3$ of water.
- Alternatively, prepare a 1.0 moldm$^{-3}$ solution of ethanoic acid and sodium ethanoate and mix the two solutions together in a volume ratio of 0.35:1.
1. Explain how you would prepare a buffer solution of pH = 9.5
Answer:
[base] = 1.27
[salt]
Theory and application of buffer system
Maintaining the pH of blood

- Are important in the proper functioning of cells and blood.
- Help maintain the pH of blood close to 7.4.

A change in the pH of the blood affects the uptake of oxygen and cellular processes.

→ Human blood must be maintained at a pH of 7.40 ± 0.05 for the rest of the body biochemistry to function. This is achieved with a buffer made from CO$_2$:

$$\text{CO}_2(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq})$$
Theory and application of buffer system

Chemistry In Action: Maintaining the pH of Blood

Red blood cells in a capillary

Tissues

<table>
<thead>
<tr>
<th>Capillary</th>
<th>Tissues</th>
</tr>
</thead>
<tbody>
<tr>
<td>Erythrocyte</td>
<td>HbO₂⁻ + H⁺</td>
</tr>
<tr>
<td>CO₂ → CO₂ + H₂O → H₂CO₃ → H⁺ + HCO₃⁻</td>
<td></td>
</tr>
<tr>
<td>O₂ + HHb</td>
<td></td>
</tr>
<tr>
<td>Capacitor</td>
<td></td>
</tr>
</tbody>
</table>

Lungs

<table>
<thead>
<tr>
<th>Capillary</th>
<th>Lungs</th>
</tr>
</thead>
<tbody>
<tr>
<td>Erythrocyte</td>
<td>HbO₂⁻ + H⁺</td>
</tr>
<tr>
<td>CO₂ → CO₂ + H₂O → H₂CO₃ → H⁺ + HCO₃⁻</td>
<td></td>
</tr>
<tr>
<td>O₂ + HHb</td>
<td></td>
</tr>
<tr>
<td>Capacitor</td>
<td></td>
</tr>
</tbody>
</table>

Plasma

<table>
<thead>
<tr>
<th>Capillary</th>
<th>Plasma</th>
</tr>
</thead>
</table>

HCO₃⁻
THE END