

Which arrow? → or ⇌		
<p><b>Strong Acids e.g. HCl, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub></b> Acids are proton donors. We can use HA to represent the <u>strong acid</u>. HA → H<sup>+</sup>(aq) + A<sup>-</sup>(aq) OR more correctly HA + H<sub>2</sub>O → H<sub>3</sub>O<sup>+</sup>(aq) + A<sup>-</sup>(aq) Use the → because the reaction goes to completion.</p> <ul style="list-style-type: none"> <li>All the HA has reacted with water (reaction with water is complete).</li> <li>All the HA has ionised / dissociated (turned into ions)</li> <li>There is NO HA remaining.</li> </ul> <p>The pH is &lt; 7 and is usually very low (0-2).</p>	<p><b>Salt solutions</b> These <u>can</u> require two equations. First the salt <u>dissolves</u> in water. Use the → because the dissolving process goes to completion. Then consider the ions; do none, one or two <u>react</u> with water? If one does it is because they are either weak acids or bases. (Since weak we use ⇌ here).</p>	<p><b>Strong Bases e.g. NaOH, KOH (must have OH<sup>-</sup> ion)</b> These are bases because when they dissolve in water OH<sup>-</sup>(aq) is produced. NaOH(s) → Na<sup>+</sup>(aq) + OH<sup>-</sup>(aq) Use the → because the dissolving process goes to completion.</p> <ul style="list-style-type: none"> <li>All the NaOH has dissolved in water.</li> <li>All the NaOH has ionised / dissociated (turned into ions)</li> <li>There is NO NaOH remaining.</li> </ul> <p>The pH is &gt; 7 and is usually very high (12-14).</p>
<p><b>Weak Acids e.g. HCOOH, CH<sub>3</sub>COOH, HOBr</b> Acids are proton donors. We can use HA to represent the <u>weak acid</u>. HA ⇌ H<sup>+</sup>(aq) + A<sup>-</sup>(aq) OR more correctly HA + H<sub>2</sub>O ⇌ H<sub>3</sub>O<sup>+</sup>(aq) + A<sup>-</sup>(aq) Use the ⇌ because the reaction with water does NOT go to completion; it is an equilibrium.</p> <ul style="list-style-type: none"> <li>NOT all the HA has reacted with water (reaction with water is incomplete).</li> <li>NOT all the HA has ionised / dissociated (turned into ions)</li> </ul> <p>There is MUCH HA remaining &amp; only a little H<sub>3</sub>O<sup>+</sup>(aq) + A<sup>-</sup>(aq) as equilibrium position lies to the left. The pH is &lt; 7 but not very low.</p>	<p>Salts that produce neutral solutions (pH 7) e.g. NaCl, CaCl<sub>2</sub> NaCl(s) → Na<sup>+</sup>(aq) + Cl<sup>-</sup>(aq) ; neither ion is a proton donor or acceptor so the pH is 7.</p> <p>Salts that produce solutions with pH &lt; 7 e.g. NH<sub>4</sub>Cl NH<sub>4</sub>Cl(s) → NH<sub>4</sub><sup>+</sup>(aq) + Cl<sup>-</sup>(aq) NH<sub>4</sub><sup>+</sup>(aq) + H<sub>2</sub>O(l) ⇌ NH<sub>3</sub>(aq) + H<sub>3</sub>O<sup>+</sup>(aq)</p> <p>Salts that produce solutions with pH &gt; 7 e.g. CH<sub>3</sub>COONa, NaHCO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub> NaHCO<sub>3</sub>(s) → Na<sup>+</sup>(aq) + HCO<sub>3</sub><sup>-</sup>(aq) HCO<sub>3</sub><sup>-</sup>(aq) + H<sub>2</sub>O(l) ⇌ H<sub>2</sub>CO<sub>3</sub>(aq) + OH<sup>-</sup>(aq)</p>	<p><b>Weak Bases e.g. NH<sub>3</sub>, CH<sub>3</sub>NH<sub>2</sub></b> These are bases because when they react with water OH<sup>-</sup>(aq) is produced. We can use B or B<sup>-</sup> to represent the base. B + H<sub>2</sub>O ⇌ BH<sup>+</sup>(aq) + OH<sup>-</sup>(aq) Use the ⇌ because the reaction with water does NOT go to completion; it is an equilibrium</p> <ul style="list-style-type: none"> <li>NOT all the B has reacted with water (reaction with water is incomplete).</li> <li>There is MUCH B remaining &amp; only a little BH<sup>+</sup> + OH<sup>-</sup> as the equilibrium position lies to the left.</li> </ul> <p>The pH is &gt; 7 but not very high.</p>

### Buffer solutions

A buffer solution is one which resists changes in pH when small quantities of an acid or an alkali are added to it.

#### Acidic buffer solutions

An acidic buffer solution is simply one which has a pH less than 7. Acidic buffer solutions are commonly made from a weak acid and one of its salts - often a sodium salt.

#### Alkaline buffer solutions

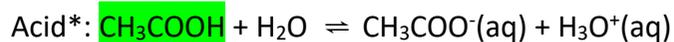
An alkaline buffer solution has a pH greater than 7. Alkaline buffer solutions are commonly made from a weak base and one of its salts.

#### How do buffer solutions work?

A buffer solution has to contain things which will remove any hydrogen ions or hydroxide ions that you might add to it - otherwise the pH will change. Acidic and alkaline buffer solutions achieve this in different ways.

### Acidic buffers e.g. CH<sub>3</sub>COOH/CH<sub>3</sub>COONa

Consist of 2 components mixed together, a weak acid and a salt containing its conjugate base



The contribution to the [CH<sub>3</sub>COO<sup>-</sup>] from the acid\* is negligible since the acid is a weak acid & equilibrium position lies to the left.

On addition of small amounts of H<sub>3</sub>O<sup>+</sup>(aq)

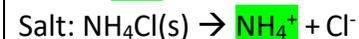
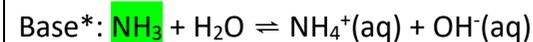
- $\text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$  ; added H<sub>3</sub>O<sup>+</sup> ions are "removed" & the pH remains the same. H<sub>3</sub>O<sup>+</sup> ions combine with the ethanoate ions to make ethanoic acid & water.

On addition of small amounts of OH<sup>-</sup>(aq)

- $\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$  ; added OH<sup>-</sup> ions are "removed" & the pH remains the same.

### Basic buffers e.g. NH<sub>3</sub>/NH<sub>4</sub>Cl

Consist of 2 components mixed together, a weak base and a salt containing its conjugate acid



The contribution to the [NH<sub>4</sub><sup>+</sup>] from the base\* is negligible since the base is a weak base & equilibrium position lies to the left.

On addition of small amounts of H<sub>3</sub>O<sup>+</sup>(aq)

- $\text{NH}_3 + \text{H}_3\text{O}^+ \rightarrow \text{NH}_4^+ + \text{H}_2\text{O}$  ; added H<sub>3</sub>O<sup>+</sup> ions are "removed" & the pH remains the same. H<sub>3</sub>O<sup>+</sup> ions combine with the ammonia molecules ions to make the ammonium ion & water.

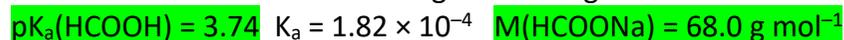
On addition of small amounts of OH<sup>-</sup>(aq)

- $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$  ; added OH<sup>-</sup> ions are "removed" & the pH remains the same

### Buffer calculation example

Calculate the mass of sodium methanoate that must be added to 100 mL of 0.861 mol L<sup>-1</sup> methanoic acid to give a solution with a pH of 3.24.

Assume there is no volume change on adding the salt.



$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

We know pH (3.24), pK<sub>a</sub> (3.74) & conc. of the acid, HCOOH (0.861 mol L<sup>-1</sup>).

Calculate [base], the [HCOO<sup>-</sup>]; the answer will be in mol L<sup>-1</sup>

From mol L<sup>-1</sup> calculate how many mol you would need for 100 mL (and not for a L, 1000 mL)

Calculate the mass you would therefore need to dissolve in 100 mL (using m = nM)