Strong Acids e.g. HCl, HNO_3, H_2SO_4

Acids are proton donors. We can use HA to represent the <u>strong acid</u>.

$HA \rightarrow H^+(aq) + A^-(aq)$ OR more correctly

$HA + H_2O \rightarrow H_3O^+(aq) + A^-(aq)$

Use the $\rightarrow$ because the reaction goes to completion.

- All the HA has reacted with water (reaction with water is complete).
- All the HA has ionised / dissociated (turned into ions)
- There is NO HA remaining.

The pH is < 7 and is usually very low (0-2).

Weak Acids e.g. HCOOH, CH_3COOH, HOBr

Acids are proton donors. We can use HA to represent the <u>weak acid</u>.

$HA \rightleftharpoons H^+(aq) + A^-(aq)$ OR more correctly

$HA + H_2O \rightleftharpoons H_3O^+(aq) + A^-(aq)$

Use the $\rightleftharpoons$ because the reaction with water does NOT go to completion; it is an equilibrium.

- NOT all the HA has reacted with water (reaction with water is incomplete).
- NOT all the HA has ionised / dissociated (turned into ions)

There is MUCH HA remaining & only a little $H_3O^+(aq) + A^-(aq)$ as equilibrium position lies to the left.

The pH is < 7 but not very low.

Which arrow?

$\rightarrow$ or $\rightleftharpoons$

Salt solutions

These <u>*can*</u> require two equations.

First the salt <u>*dissolves*</u> in water. Use the $\rightarrow$ 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 $\rightleftharpoons$ here).

Salts that produce neutral solutions (pH 7) e.g. NaCl, $CaCl_2$

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$; neither ion is a proton donor or acceptor so the pH is 7.

Salts that produce solutions with pH < 7 e.g. NH_4Cl

$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$

$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

Salts that produce solutions with pH > 7 e.g. CH_3COONa, $NaHCO_3$, Na_2CO_3

$NaHCO_3(s) \rightarrow Na^+(aq) + HCO_3^-(aq)$

$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

Strong Bases e.g. NaOH, KOH (must have OH^- ion)

These are bases because when they dissolve in water $OH^-(aq)$ is produced.

$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

Use the $\rightarrow$ because the dissolving process goes to completion.

- All the NaOH has dissolved in water.
- All the NaOH has ionised / dissociated (turned into ions)
- There is NO NaOH remaining.

The pH is > 7 and is usually very high (12-14).

Weak Bases e.g. NH_3, CH_3NH_2

These are bases because when they react with water $OH^-(aq)$ is produced. We can use B or B^- to represent the base.

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

Use the $\rightleftharpoons$ because the reaction with water does NOT go to completion; it is an equilibrium

- NOT all the B has reacted with water (reaction with water is incomplete).
- There is MUCH B remaining & only a little BH^+ + OH^- as the equilibrium position lies to the left.

The pH is > 7 but not very high.

Buffer solutions	Acidic buffers e.g. CH_3COOH/CH_3COONa	Basic buffers e.g. NH_3/NH_4Cl
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.	Consist of 2 components mixed together, a weak acid and a salt containing its conjugate base Acid*: $CH_3COOH + H_2O \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$ Salt: $CH_3COONa(s) \rightarrow CH_3COO^-(aq) + Na^+(aq)$ The contribution to the $[CH_3COO^-]$ from the acid* is negligible since the acid is a weak acid & equilibrium position lies to the left. On addition of small amounts of $H_3O^+(aq)$ • $CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$; added H_3O^+ ions are "removed" & the pH remains the same. H_3O^+ ions combine with the ethanoate ions to make ethanoic acid & water. On addition of small amounts of $OH^-(aq)$ • $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$; added OH^- ions are "removed" & the pH remains the same.	Consist of 2 components mixed together, a weak base and a salt containing its conjugate acid Base*: $NH_3 + H_2O \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ Salt: $NH_4Cl(s) \rightarrow NH_4^+ + Cl^-$ The contribution to the $[NH_4^+]$ from the base* is negligible since the base is a weak base & equilibrium position lies to the left. On addition of small amounts of $H_3O^+(aq)$ • $NH_3 + H_3O^+ \rightarrow NH_4^+ + H_2O$; added H_3O^+ ions are "removed" & the pH remains the same. H_3O^+ ions combine with the ammonia molecules ions to make the ammonium ion & water. On addition of small amounts of $OH^-(aq)$ • $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$; added OH^- 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^{-1} methanoic acid to give a solution with a pH of 3.24.
Assume there is no volume change on adding the salt.
$pK_a(HCOOH) = 3.74$ $K_a = 1.82 \times 10^{-4}$ $M(HCOONa) = 68.0$ g mol^{-1}

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

We know pH (3.24), pK_a (3.74) & conc. of the acid, HCOOH (0.861 mol L^{-1}).
Calculate [base], the $[HCOO^-]$; the answer will be in mol L^{-1}
From mol L^{-1} 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)