Which arrow?

→ or ⇌

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>.

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

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

Use the → 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>.

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

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

Use the ⇌ 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.

Salt solutions

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).

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

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

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

$NH_4Cl + aq \rightarrow NH_4^+ + Cl^-$

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

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

$NaHCO_3 + aq \rightarrow Na^+ + HCO_3^-$

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

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

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

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

Use the → 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 ⇌ 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

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_3COOH/CH_3COONa
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: $CH_3COONa + aq \rightarrow CH_3COO^-(aq) + Na^+(aq)$

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.

Basic buffers e.g. NH_3/NH_4Cl
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: $NH_4Cl + aq \rightarrow NH_4^+ + Cl^-$

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)