BUFFER SOLUTIONS - INTRODUCTION AND USES

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

Types

Acidic Buffer (pH < 7) — **weak acid + its sodium or potassium salt**
ethanoic acid — *sodium ethanoate*

Alkaline Buffer (pH > 7) — **weak base + its chloride**
ammonia — *ammonium chloride*

Biological Uses

In biological systems (saliva, stomach, and blood) it is essential that the pH stays 'constant' in order for any processes to work properly. Most enzymes work best at particular pH values.

Blood

- the pH of blood is normally about 7.4
- If the pH varies by 0.5 it can lead to unconsciousness and coma
- carbon dioxide produced by respiration can increase the acidity of blood by forming H+ ions in aqueous solution

$$\mathbf{CO_2}(aq) + \mathbf{H_2O}(aq) \rightleftharpoons \mathbf{H^+}(aq) + \mathbf{HCO_3^-}(aq)$$

- the presence of hydrogencarbonate ions in blood removes excess H^+

$$\mathbf{H^+}(aq) + \mathbf{HCO_3^-}(aq) \rightleftharpoons \mathbf{H_2CO_3}(aq) \quad \text{(equivalent to } CO_2 \text{ in water)}$$

Other Uses

Many household and cosmetic products need to control their pH values.

Shampoo Counteract the alkalinity of the soap and prevent irritation

Baby lotion Maintain a pH of about 6 to prevent bacteria multiplying

Others Washing powder

Eye drops

Fizzy lemonade

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.

	CH_3COOH(aq)	$\rightleftharpoons$	CH_3COO^-(aq)	+	H^+(aq)
relative concs.	*HIGH*		*LOW*		*LOW*

A strong acid can't be used as **it is fully dissociated and cannot remove H^+**(aq)

$$HCl(aq) \longrightarrow Cl^-(aq) + H^+(aq)$$

Adding acid Any H^+ **is removed by reacting with CH_3COO^-** ions to form CH_3COOH via the equilibrium. Unfortunately, the concentration of CH_3COO^- is small and only a few H^+ can be "mopped up". A much larger concentration of CH_3COO^- is required.

To build up the concentration of CH_3COO^- ions, sodium ethanoate is added.

Adding alkali Adds OH^- ions. Although they do not appear in the equation, they react with H^+

$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(aq)$$

Removal of H^+ from the weak acid equilibrium means that, according to Le Chatelier's Principle, more CH_3COOH will dissociate to form ions to replace those being removed.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

As the added OH^- ions remove the H^+ from the weak acid system, the equilibrium moves to the right to produce more H^+ ions. Obviously, there must be a large concentration of undissociated acid molecules to be available.

Other The concentration of a buffer solution is also important

If the concentration is too low, there won't be enough CH_3COOH and CH_3COO^- to cope with the ions added.

Summary For an acidic buffer solution one needs ...

large [CH_3COOH(aq)] - for dissociating into H^+(aq) when alkali is added

large [CH_3COO^-(aq)] - for removing H^+(aq) as it is added

This 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 CH_3COOH(aq) concentration. The sodium salt provides the large CH_3COO^-(aq) concentration.

$\therefore$ **One uses a WEAK ACID + its SODIUM OR POTASSIUM SALT**

CALCULATING THE pH OF AN ACIDIC BUFFER SOLUTION

Example 1 *Calculate the pH of a buffer solution whose [HA] is 0.1 mol dm^{-3} and [A$^-$] of 0.1 mol dm^{-3}. Assume the K_a of the weak acid HA is 2 x 10^{-4} mol dm^{-3}.*

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

re-arranging

$$[H^+_{(aq)}] = \frac{[HA_{(aq)}]\ K_a}{[A^-_{(aq)}]} = \frac{0.1 \times 2 \times 10^{-4}}{0.1} = \mathbf{2 \times 10^{-4}\ mol\ dm^{-3}}$$

$$\therefore \quad \mathbf{pH = -\log_{10} [H^+_{(aq)}]} \quad = \quad \mathbf{3.699} \quad (3.7)$$

Example 2 *Calculate the pH when 500cm^3 of 0.10 mol dm^{-3} of weak acid HX is mixed with 500cm^3 of a 0.20 mol dm^{-3} solution of its salt NaX. K_a = 4.0 x 10^{-5} mol dm^{-3}.*

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

re-arranging

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

The solutions have been mixed*; volume is now 1 dm^3* *[HX] = 0.05 mol dm^{-3}*

[X$^-$] = 0.10 mol dm^{-3}

$$\therefore \quad [H^+_{(aq)}] = \frac{0.05 \times 4.0 \times 10^{-5}}{0.1} = \mathbf{2.0 \times 10^{-5}\ mol\ dm^{-3}}$$

$$\therefore \quad \mathbf{pH = -\log_{10} [H^+(aq)]} \quad = \quad \mathbf{4.699} \quad (4.7)$$

Alkaline buffer Similar but is based on the equilibrium surrounding a weak base.

	$\mathbf{NH_3}$(aq)	+	$\mathbf{H_2O_{(l)}}$(aq)	$\rightleftharpoons$	$\mathbf{OH^-}$(aq)	+	$\mathbf{NH_4^+}$(aq)
relative concs.	*HIGH*				*LOW*		*LOW*

but one needs ; **a large conc. of OH^-(aq)** to react with any H^+(aq) added

a large conc of NH_4^+(aq) to react with any OH^-(aq) added

There is enough NH_3 to act as a source of OH^- but one needs to increase the concentration of ammonium ions by adding an ammonium salt.

Use **AMMONIA (a weak base) + AMMONIUM CHLORIDE (one of its salts)**