BUFFER SOLUTIONS

Buffer solutions

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Buffer solutions

Before you start it would be helpful to…

- **know that weak acids and bases are only partly ionised in solution**
- **be able to calculate pH from hydrogen ion concentration**
- **be able to construct an equation for the dissociation constant of a weak acid**

Buffer solutions - Brief introduction

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

Acidic Buffer	**(pH < 7)**	**made from**	**a weak acid**	**+**	**its sodium or potassium salt**
			ethanoic acid		sodium ethanoate
Alkaline Buffer	**(pH > 7)**	**made from a**	**weak base**	**+**	**its chloride**
			ammonia		ammonium chloride

Uses

- Standardising pH meters
- Buffering biological systems (eg in blood)
- Maintaining the pH of shampoos

Buffer solutions - uses

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

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.
e.g. If the pH of blood varies by 0.5 it can lead to unconsciousness and coma

Most enzymes work best at particular pH values.

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

Shampoo **Buffer solutions counteract the alkalinity of the soap and prevent irritation**

Baby lotion **Buffer solutions maintain a pH of about 6 to prevent bacteria multiplying**

Others **Washing powder, eye drops, fizzy lemonade**

Buffer solutions - action

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

NB 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, which dissociates completely.

$$CH_3COO^-Na^+(aq) \longrightarrow CH_3COO^-(aq) + Na^+(aq)$$

relative concs. LOW HIGH HIGH

.

Buffer solutions - action

We have now got an equilibrium mixture which contains a high concentration of ***both*** the undissociated weak acid, $CH_3COOH(aq)$, and its conjugate base, $CH_3COO^-(aq)$.

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relative concs. HIGH LOW LOW

$$CH_3COO^-Na^+(aq) \longrightarrow CH_3COO^-(aq) + Na^+(aq)$$

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Adding alkali

This adds OH^- ions

These react with the small concentration of H^+ ions: $H^+(aq) + OH^-(aq) \quad H_2O(l)$

Removal of H^+ from the weak acid equilibrium means that more CH_3COOH will dissociate to form ions to replace those being removed.

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

As the added OH^- ions remove the H^+ from the weak acid system, the equilibrium moves to the right to produce more H^+ ions.

(There needs to be a large concentration of undissociated acid molecules to be available)

Buffer solutions - ideal concentration

The concentration of a buffer solution is 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 a buffer solution one needs ...

large [$CH_3COOH(aq)$] - for dissociating into $H^+(aq)$ when alkali is added
weak acid (equilibrium shifts to the right)

large [$CH_3COO^-(aq)$] - for removing $H^+(aq)$ as it is added
conjugate base (equilibrium shifts to the left)

This situation can't exist if only acid is present; 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.

One uses a WEAK ACID + its SODIUM OR POTASSIUM SALT

Alkaline buffer solutions - action

Alkaline buffer

Very similar but is based on the equilibrium surrounding a weak base; AMMONIA

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

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)

Calculating the pH of an acidic buffer solution

Calculate the pH of a buffer whose [HA] is 0.1 mol dm^{-3} and [A^{-}] of 0.1 mol dm^{-3}. The K_a of the weak acid HA is 2×10^{-4} mol dm^{-3}

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

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re-arrange

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

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from information given

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

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

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$$pH = -\log_{10}[H^+(aq)] = 3.699$$

Calculating the pH of an acidic buffer solution

Calculate the pH of the solution formed when $500cm^3$ of 0.1 mol dm^{-3} of weak acid HX is mixed with $500cm^3$ of a 0.2 mol dm^{-3} solution of its salt NaX. $K_a = 4 \times 10^{-5}$ mol dm^{-3}.

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

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

therefore $[HX]$ = 0.05 mol dm^{-3} and

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

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Substituting

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

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