

BUFFER SOLUTIONS - INTRODUCTION AND USES

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

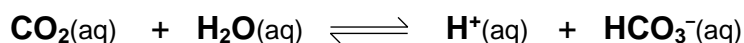
<i>Types</i>	Acidic Buffer (pH < 7)	weak acid + its sodium or potassium salt
		<i>ethanoic acid</i> + <i>sodium ethanoate</i>
	Alkaline Buffer (pH > 7)	weak base + its chloride
		<i>ammonia</i> + <i>ammonium chloride</i>

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



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



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

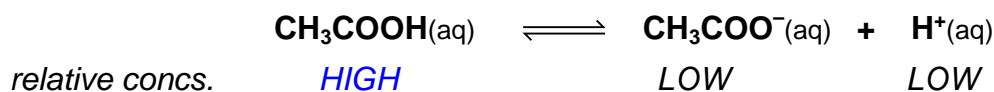
Others Washing powder

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.



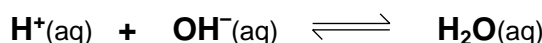
A strong acid can't be used as **it is fully dissociated and cannot remove H⁺(aq)**



Adding acid Any H⁺ is removed by reacting with CH₃COO⁻ ions to form CH₃COOH via the equilibrium. Unfortunately, the concentration of CH₃COO⁻ is small and only a few H⁺ can be "mopped up". A much larger concentration of CH₃COO⁻ is required.

To build up the concentration of CH₃COO⁻ ions, sodium ethanoate is added.

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



Removal of H⁺ from the weak acid equilibrium means that, according to Le Chatelier's Principle, more CH₃COOH will dissociate to form ions to replace those being removed.



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 CH₃COOH and CH₃COO⁻ to cope with the ions added.

Summary For an acidic buffer solution one needs ...

large [CH₃COOH(aq)] - for dissociating into H⁺(aq) when alkali is added

large [CH₃COO⁻(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 CH₃COOH(aq) concentration. The sodium salt provides the large CH₃COO⁻(aq) concentration.

∴ 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 $2 \times 10^{-4} \text{ mol dm}^{-3}$.

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

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

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

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 . $K_a = 4.0 \times 10^{-5} \text{ mol dm}^{-3}$.

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

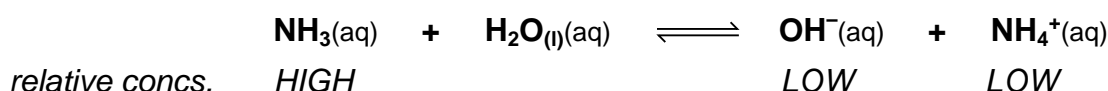
$$\text{re-arranging} \quad [H^+_{(aq)}] = \frac{[HX_{(aq)}] K_a}{[X^-_{(aq)}]}$$

The solutions have been mixed; volume is now 1 dm^3 $[HX] = 0.05 \text{ mol dm}^{-3}$
 $[X^-] = 0.10 \text{ mol dm}^{-3}$

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

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

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



but one needs ; **a large conc. of $\text{OH}^-(\text{aq})$** to react with any $\text{H}^+(\text{aq})$ added
a large conc of $\text{NH}_4^+(\text{aq})$ to react with any $\text{OH}^-(\text{aq})$ 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)