

## 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<sup>+</sup> ions in aqueous solution



- the presence of hydrogencarbonate ions in blood removes excess H<sup>+</sup>



*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

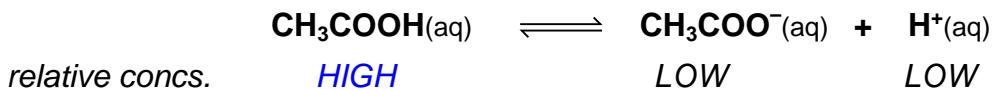
*Others* Washing powder

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.



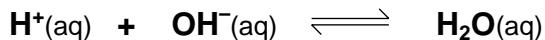
A strong acid can't be used as **it is fully dissociated and cannot remove H<sup>+</sup>(aq)**



**Adding acid** Any H<sup>+</sup> is removed by reacting with CH<sub>3</sub>COO<sup>-</sup> ions to form CH<sub>3</sub>COOH via the equilibrium. Unfortunately, the concentration of CH<sub>3</sub>COO<sup>-</sup> is small and only a few H<sup>+</sup> can be "mopped up". A much larger concentration of CH<sub>3</sub>COO<sup>-</sup> is required.

**To build up the concentration of CH<sub>3</sub>COO<sup>-</sup> ions, sodium ethanoate is added.**

**Adding alkali** Adds OH<sup>-</sup> ions. Although they do not appear in the equation, they react with H<sup>+</sup>



Removal of H<sup>+</sup> from the weak acid equilibrium means that, according to Le Chatelier's Principle, more CH<sub>3</sub>COOH will dissociate to form ions to replace those being removed.



As the added OH<sup>-</sup> ions remove the H<sup>+</sup> from the weak acid system, the equilibrium moves to the right to produce more H<sup>+</sup> 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<sub>3</sub>COOH and CH<sub>3</sub>COO<sup>-</sup> to cope with the ions added.

**Summary** For an acidic buffer solution one needs ...

**large [CH<sub>3</sub>COOH(aq)] - for dissociating into H<sup>+</sup>(aq) when alkali is added**  
**large [CH<sub>3</sub>COO<sup>-</sup>(aq)] - for removing H<sup>+</sup>(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<sub>3</sub>COOH(aq) concentration. The sodium salt provides the large CH<sub>3</sub>COO<sup>-</sup>(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  $0.1 \text{ mol dm}^{-3}$  and  $[A^-]$  of  $0.1 \text{ mol dm}^{-3}$ . Assume the  $K_a$  of the weak acid HA is  $2 \times 10^{-4} \text{ mol dm}^{-3}$ .

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

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

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

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

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

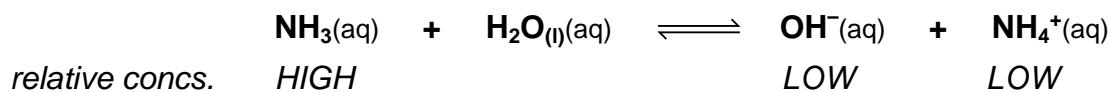
$$\text{re-arranging } [H^+_{(aq)}] = \frac{[HX_{(aq)}] K_a}{[X^-_{(aq)}]}$$

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

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

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

**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  $\text{NH}_3$  to act as a source of  $\text{OH}^-$  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)**