





## Action of acidic buffer solution

### (i) Addition of small amount of acid.

- addition of  $H^+$  ions shifts the position of equilibrium to the left because  $H^+$  ions combine with  $CH_3COO^-$  ions to form more  $CH_3COOH$  until equilibrium is re-established.
- the large reserve supply of  $CH_3COO^-$  ensures that the concentration of  $CH_3COO^-$  ions in solution does not change significantly.
- the large reserve supply of  $CH_3COOH$  ensures that the concentration of  $CH_3COOH$  molecules in solution does not change significantly.
- so the pH remains almost unchanged.

### (ii) Addition of small amount of Alkali

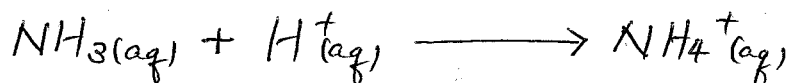
- the added  $OH^-$  ions combine with  $H^+$  ions to form water.
- this reduces the  $H^+$  ion concentration.
- the position of equilibrium shifts to the right.
- $CH_3COOH$  molecules ionise to form more  $H^+$  and  $CH_3COO^-$  ions until equilibrium is re-established.
- the large reserve supply of  $CH_3COOH$  ensures that the concentration of  $CH_3COOH$  molecules in solution does not change significantly.
- the large reserve supply of  $CH_3COO^-$  ensure that the concentration of  $CH_3COO^-$  ions in solution does not change significantly.
- so the pH remains almost unchanged.



## Action of alkaline buffer solution

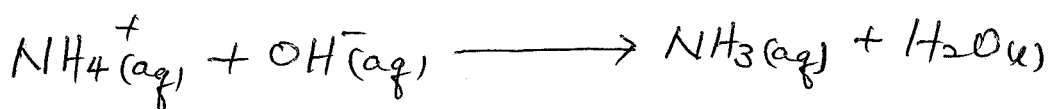
(i) Addition of small amount of acid.

The added  $H^+$  ions will react with the  $NH_3$  to form  $NH_4^+$



(ii) Addition of small amount of alkali

The added  $OH^-$  ions will react with the  $NH_4^+$  to form water and  $NH_3$



## Exercise

A mixture of  $0.500 \text{ mol dm}^{-3}$  aqueous ammonia and  $0.500 \text{ mol dm}^{-3}$  ammonium chloride acts as a buffer solution.

a. Explain how this buffer solution minimises changes in pH on addition of:

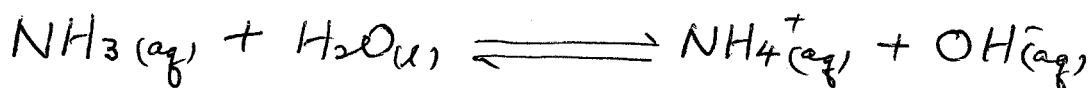
i) dilute hydrochloric acid.

ii) dilute sodium hydroxide.

b. Explain why dilute aqueous ammonia alone will not act as a buffer solution.

## Workings.

a. The equilibrium exists in the buffer system is :



i) When small amount of hydrochloric acid is added, the additional  $\text{H}^+$  ions combine with the  $\text{OH}^-$  ions in the equilibrium mixture to form  $\text{H}_2\text{O}$ .

The position of equilibrium shifts to the right.

Because there are relatively high concentrations of ammonia (base) and ammonium ions (conjugate acid) present compared with the concentration of added  $\text{H}^+$  ions, the pH remains almost unchanged.

ii) When small amount of sodium hydroxide is added, the additional  $\text{OH}^-$  ions shift the position of equilibrium to the left.

More ammonia and water is formed.

Because there are relatively high concentrations of ammonia and ammonium ions present compared with the concentration of added  $\text{OH}^-$  ions, the pH remains almost unchanged.

b. Ammonia is a weak base. The equilibrium lies well over to the left.

There are not enough  $\text{NH}_4^+$  ions in the equilibrium mixture to remove added  $\text{OH}^-$  ions.

## Uses of buffer solutions

Buffers are vital in biological systems.

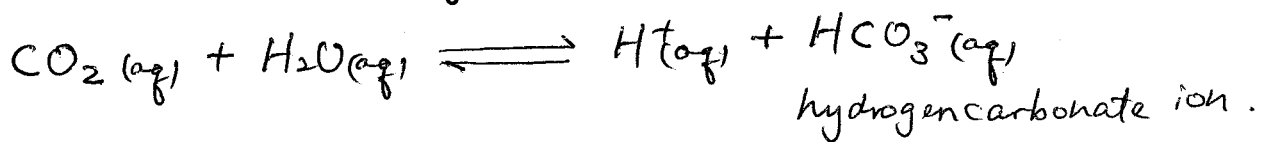
In humans, the pH of the blood is kept between 7.35 and 7.45.

There are a number of different buffers in the blood:

- hydrogen carbonate ions,  $\text{HCO}_3^-$ .\*
- haemoglobin and plasma proteins.
- dihydrogen phosphate ( $\text{H}_2\text{PO}_4^-$ ) and hydrogen phosphate ( $\text{HPO}_4^-$ ) ions.

The cells in our body produce carbon dioxide as a product of aerobic respiration (the oxidation of glucose to provide energy).

Carbon dioxide combines with water in the blood to form a solution containing hydrogen ions.



This reaction is catalysed by enzyme carbonic anhydrase.

When the blood passed through the small blood vessels around our lungs, hydrogen carbonate ions are rapidly converted to carbon dioxide and water.

The carbon dioxide escapes into the lungs.

The production of  $H^+$  ions, if left unchecked, would lower the pH of the blood and cause 'acidosis'.

This may disrupt some body functions and eventually lead to coma.

The equilibrium between carbon dioxide and hydrogencarbonate is the most important buffering system in the blood.

If  $[H^+]$  increases:

- the position of equilibrium shifts to the left.
- $H^+$  ions combine with  $HCO_3^-$  ions to form carbon dioxide and water until equilibrium is restored.
- this reduces  $[H^+]$  in the blood and helps keep the pH constant.

If  $[H^+]$  decreases:

- the position of equilibrium shifts to the right.
- some carbon dioxide and water combine to form  $H^+$  and  $HCO_3^-$  ions until equilibrium is restored.
- this increases  $[H^+]$  in the blood and helps keep the pH constant.



## Other uses of buffer solutions

- industrial processes eg. electroplating, the manufacture of dye and in the treatment of leather.
- culturing and selection of bacteria.
- make sure that pH meters record the correct pH.

## Exercise

a. One of the buffers in blood plasma is a mixture of dihydrogenphosphate ( $\text{H}_2\text{PO}_4^-$ ) ions and hydrogenphosphate ( $\text{HPO}_4^{2-}$ ) ions.

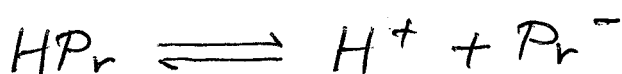
(i) Identify the conjugate acid and base in this buffer.

(ii) Write a balanced equation for the equilibrium between these two ions.

b. Some proteins in the blood can act as buffers.

The equation below shows a simplified equilibrium equation for this reaction.

(Pr = protein)

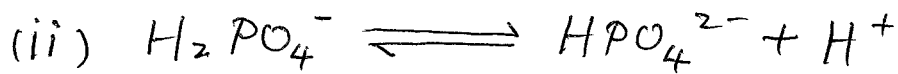


Explain how this system can act as a buffer to prevent the blood getting too acidic.

## Workings

a. (i) conjugate acid =  $\text{H}_2\text{PO}_4^-$

base =  $\text{HPO}_4^{2-}$



b. The addition of hydrogen ions shifts the position of equilibrium to the left.

$\text{Pr}^-$  (the deprotonated form of the protein) combines with the extra hydrogen ions to form  $\text{HPr}$  (the protonated form of the protein) until equilibrium is re-established.

As long as the added  $\text{H}^+$  ions are not excessively high, the reserve supply of  $\text{Pr}^-$  able to convert them to  $\text{HPr}$ .

The protein buffer system can help to maintain the pH of the blood almost constant.