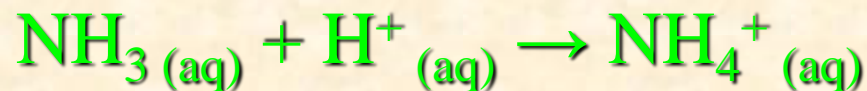
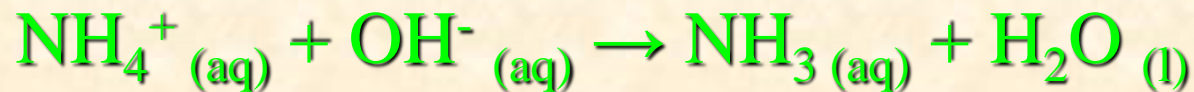


# Buffers

- Example: Ammonia is combined with its salt,  $\text{NH}_4\text{Cl}$ , in sol'n:
- If **acid** is added to this solution, ammonia reacts with the  $\text{H}^+$  :

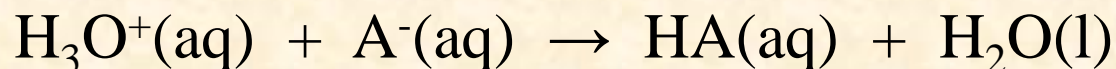


- If a **base** is added to this solution, the  $\text{NH}_4^+$  from the dissolved salt will react with the  $\text{OH}^-$  :

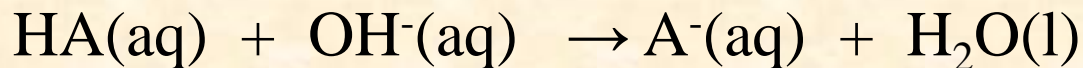


# BUFFER SOLUTION

- When an acid is added, the conjugate base converts the excess  $\text{H}_3\text{O}^+$  ion into its acid (conjugate base removes excess  $\text{H}_3\text{O}^+$ )



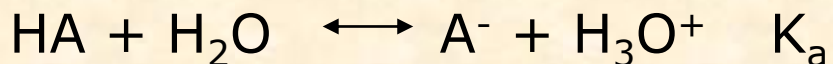
- When a base is added, the acid converts the excess  $\text{OH}^-$  ion into its conjugate base and water (acid removes excess  $\text{OH}^-$  ion)



- These reactions go to completion (large equilibrium constants)

# pH calculations of buffer solutions

Buffer consisting of a weak acid and its salt with a strong base



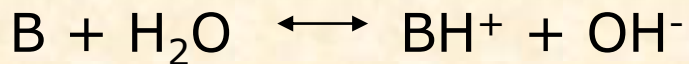
Henderson – Hasselbalch equation

$$\text{pH} = \text{pK}_a + \log\left[\frac{\text{A}^-}{\text{HA}}\right]$$

HA – weak acid

A<sup>-</sup> – conjugated base

Buffer consisting of a weak base and its salt with a strong acid



$$\text{pOH} = \text{pK}_b + \log\left[\frac{\text{BH}^+}{\text{B}}\right]$$

B – weak base

BH<sup>+</sup> – conjugated acid

# The Henderson-Hasselbalch Equation



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$
$$[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HA}]}{[\text{A}^-]}$$

$$-\log[\text{H}_3\text{O}^+] = -\log K_a - \log\left(\frac{[\text{HA}]}{[\text{A}^-]}\right)$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

# HENDERSON-HASSELBALCH EQUATION

In general

$$\text{pH} = \text{pK}_a + \log\left(\frac{C_b}{C_a}\right) \quad \text{or} \quad \text{pH} = \text{pK}_a + \log\left(\frac{n_b}{n_a}\right)$$

-  $C_a$  and  $C_b$  are the analytical concentrations of the acid and the conjugate base, respectively

-  $n_a$  and  $n_b$  are the number of moles of the acid and the conjugate base, respectively

$$C_a = n_a/V \quad \text{and} \quad C_b = n_b/V \quad (\text{concentration ratio equals mole ratio})$$

# PREPARING BUFFERS

- Measure the amount of weak acid (HA) to be used  
[or weak base (B)]
- Calculate the amount of strong base ( $\text{OH}^-$ ) to be added  
[or strong acid ( $\text{H}^+$ )]
- This makes a mixture of HA and  $\text{A}^-$  which is a buffer  
[or B and  $\text{BH}^+$ ]

Or

- Add the correct proportions of the acid and its conjugate base,  
then check the pH

# Question 1

A buffer solution is made by adding 0.3 mol of acetic acid and 0.3 mol of sodium acetate to enough water to make 1L of solution. Determine the pH of the buffer solution. The  $K_a$  for acetic acid is  $1.8 \times 10^{-5}$

$$\text{pH} = \text{p}K_a + \log [A^-] / [HA] \quad A^- = \text{sodium acetate}$$
$$\text{HA} = \text{acetic acid}$$

$$\text{pH} = -\log (1.8 \times 10^{-5}) + \log (0.3) / (0.3)$$

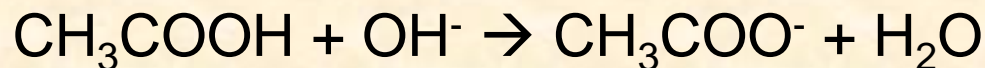
$$\text{pH} = 4.74$$

## Question 2

Calculate the pH of the buffer solution after 0.02 moles of sodium hydroxide is added?

1 litre of buffer contains 0.3 moles of sodium acetate and 0.3 moles of acetic acid. Sodium hydroxide is a strong base. The acetic acid will react with the base added to try to maintain the pH. What the acid loses in concentration, the salt (sodium acetate) will gain.





	$\text{CH}_3\text{COOH}$	$\text{OH}^-$	$\text{CH}_3\text{COO}^-$
Initial conc	0.3	0	0.3
Change	$0.3 - 0.02$	0.02	$0.3 + 0.02$
Equil. Conc.	0.28	0.02	0.32

$$\text{pH} = \text{pK}_a + \log [\text{A}^-] / [\text{HA}]$$

$$\text{pH} = -\log (1.8 \times 10^{-5}) + \log (0.32) / (0.28)$$

$$\text{pH} = 4.74 + 0.058$$

$$\text{pH} = 4.8$$



## Calculating the pH of an acidic buffer solution

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



# Calculating the pH of an acidic buffer solution

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

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

