Analytical Chemistry
Lecture III
by/ Dr. Ekhlas Q. J.

BUFFER SOLUTIONS

Buffer solutions

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

a solution that contains a weak acid/conjugate base mixture or a weak base/conjugate acid mixture

- **able to neutralize acids and bases without allowing the pH of the solution to change greatly**

Acidic Buffer (pH < 7) made from a weak acid (ethanoic acid) + its sodium or potassium salt (sodium ethanoate)

Alkaline Buffer (pH > 7) made from a weak base (ammonia) + its chloride salt (ammonium chloride)
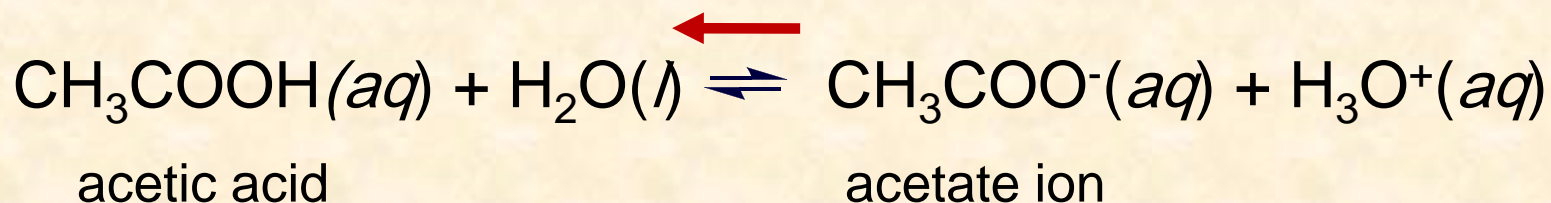
Uses
- ❖ **Standardising pH meters**
- ❖ **Buffering biological systems (eg in blood)**
- ❖ **Maintaining the pH of shampoos**

Buffers and the Common-ion Effect

A buffer works through the ***common-ion effect***.

Acetic acid in water dissociates slightly to produce some acetate ion:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

acetic acid   acetate ion

If CH_3COONa is added, it provides a source of CH_3COO^- ion, and the equilibrium shifts to the left. CH_3COO^- is ***common*** to both solutions.
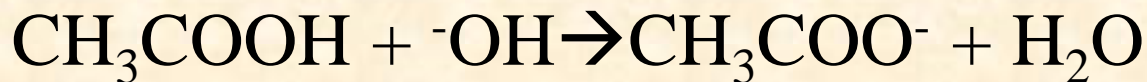
The addition of CH_3COO^- reduces the % dissociation of the acid.

Buffers

- A buffer consisting of a mixture of the weak acid(CH_3COOH) and its salt (CH_3COONa^+), and

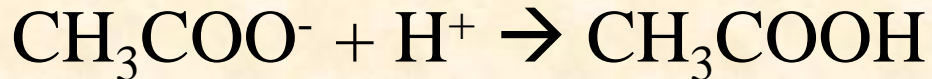 Will undergo the following changes on the addition of acid or base:

- If the base is added (^-OH) to the solution it will be buffered by the following reaction with acetic acid:

 $CH_3COOH + {}^-OH \rightarrow CH_3COO^- + H_2O$

 So the pH will not change significantly

Buffers

- If acid (H^+) is added, it will be buffered by another reaction, this time using the salt (CH_3COO^-):

  $CH_3COO^- + H^+ \rightarrow CH_3COOH$

  The pH will not alter significantly because the CH_3COOH formed is a weak acid

- Addition of more base increases A^- and decreases (HA) and this doesn't alter the pH much until $[A^-] >>>> [HA]$

The buffering power is greatest when pH=pKa , i.e. when the acid and the salt are at the same concentration

How a Buffer Works

The buffer components (HA and A⁻) are able to consume **small** amounts of added OH⁻ or H_3O^+ **by a shift in equilibrium position**.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

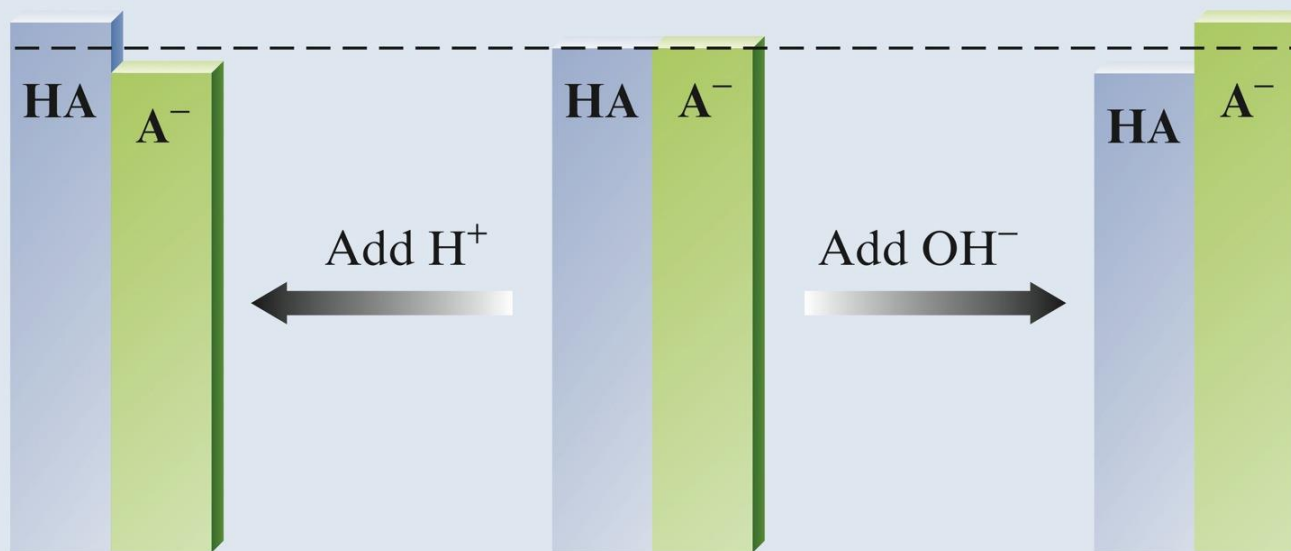Added OH⁻ reacts with CH_3COOH, causing a shift to the right.

Added H_3O^+ reacts with CH_3COO^-, causing a shift to the left.

The shift in equilibrium position absorbs the change in $[H_3O^+]$ or $[OH^-]$, and the pH changes only slightly.

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

Ratio [HA]/[A$^-$] does not increase very much.

Ratio [HA]/[A$^-$] does not decrease very much.

HA A$^-$

HA A$^-$

HA A$^-$

Add H$^+$

Add OH$^-$

H$^+$ reacts with A$^-$

OH$^-$ reacts with HA

Effect of added H$^+$ or OH$^-$ on a buffered system.

in buffer solution with HA & A⁻:

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

- addition of H⁺
  neutralization:
    $$H^+ + A^- \rightarrow HA$$
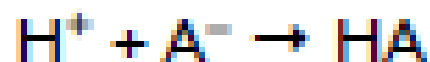
- [HA] increases slightly
  [A⁻] decreases slightly
  [HA]/[A⁻] increases
  [H⁺] increases
   pH decreases

- addition of OH⁻
  neutralization:
    $$OH^- + HA \rightarrow A^- + H_2O$$

- [HA] decreases slightly
  [A⁻] increases slightly
  [HA]/[A⁻] decreases
  [H⁺] decreases
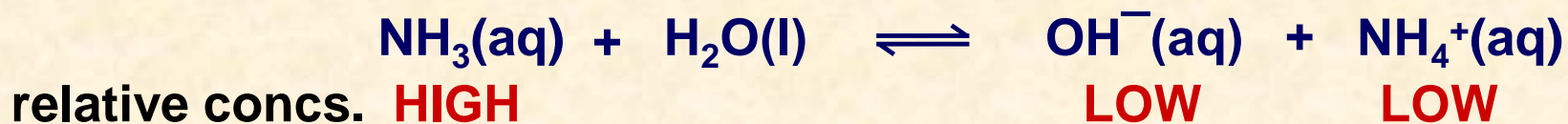   pH increases

Buffer Action

Buffer action occurs as

- the weak acid in a buffer neutralizes base.

- the conjugate base in the buffer neutralizes acid.

- the pH of the solution is maintained.

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.
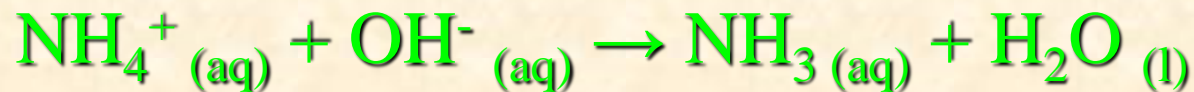
Use **AMMONIA** (a weak base) + **AMMONIUM CHLORIDE** (one of its salts)

Buffers

- Example: Ammonia is combined with its salt, NH_4Cl, in sol'n:

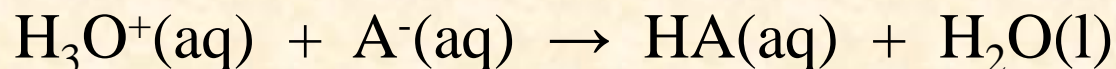- If acid is added to this solution, ammonia reacts with the H^+ :

$$NH_{3\ (aq)} + H^+_{\ (aq)} \rightarrow NH_4^+_{\ (aq)}$$
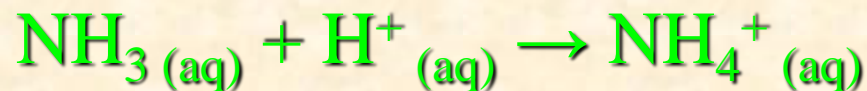
- If a base is added to this solution, the NH_4^+ from the dissolved salt will react with the OH^- :

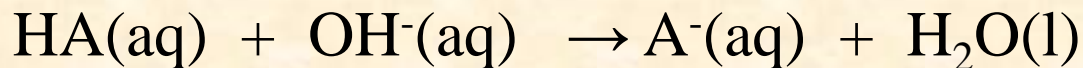$$NH_4^+_{\ (aq)} + OH^-_{\ (aq)} \rightarrow NH_{3\ (aq)} + H_2O_{\ (l)}$$

BUFFER SOLUTION

- When an acid is added, the conjugate base converts the excess H_3O^+ ion into its acid (conjugate base removes excess H_3O^+)
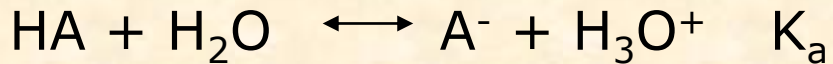
$$H_3O^+(aq) \; + \; A^-(aq) \; \rightarrow \; HA(aq) \; + \; H_2O(l)$$

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

$$HA(aq) \; + \; OH^-(aq) \; \rightarrow A^-(aq) \; + \; H_2O(l)$$

- These reactions go to completion (large equilibrium constants)

pH calculations of buffer solutions

Buffer consisting of a <u>weak acid</u> and <u>its salt with a strong base</u>
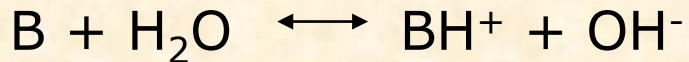
$$HA + H_2O \longleftrightarrow A^- + H_3O^+ \quad K_a$$

Henderson – Hasselbalch equation

pH = pKa + log[A-]/[HA]   HA – weak acid
                                  A^- – conjugated base

Buffer consisting of a <u>weak base</u> and <u>its salt with a strong acid</u>

$$B + H_2O \longleftrightarrow BH^+ + OH^-$$

pOH = pK$_b$ + log[BH$^+$]/[B]   B – weak base
                                     BH^+ - conjugated acid

The Henderson-Hasselbalch Equation

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \qquad [H_3O^+] = K_a \times \frac{[HA]}{[A^-]}$$

$$-\log[H_3O^+] = -\log K_a - \log\left(\frac{[HA]}{[A^-]}\right)$$

$$\boxed{pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)}$$

HENDERSON-HASSELBALCH EQUATION

In general

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

- 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

$C_a = n_a/V$ and $C_b = n_b/V$ (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 (OH^-) to be added
[or strong acid (H^+)]

- This makes a mixture of HA and A^- which is a buffer
[or B and 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 1.8×10^{-5}

$$pH = pK_a + \log [A^-] / [HA]$$

A^- = sodium acetate
HA = acetic acid
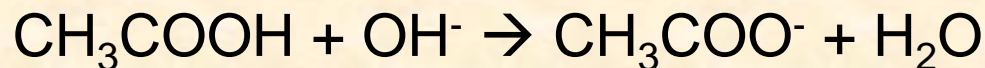
$$pH = -\log (1.8 \times 10^{-5}) + \log (0.3) / (0.3)$$
$$pH = 4.74$$

Question2

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.

$$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$$

	CH_3COOH	OH^-	CH_3COO^-
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

$pH = pK_a + \log [A^-] / [HA]$

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

$pH = 4.74 + 0.058$

$pH = 4.8$

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}

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 x 10^{-4} mol dm^{-3}

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

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}.

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

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

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}.

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

re-arrange $\qquad [H^+(aq)] = \dfrac{[HA(aq)] \times K_a}{[A^-(aq)]}$

from information given $\qquad$ [A¯] = 0.1 mol dm^{-3}

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

◄ $\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad$ ►

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}.

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

re-arrange

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

from information given

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

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

If the K_a of the weak acid HA is 2 x 10^{-4} mol dm^{-3}.

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

Calculating the pH of an acidic buffer solution

Calculate the pH of a buffer whose [HA] is 0.1 mol dm^{-3} and [A$^-$] of 0.1 mol dm^{-3}.

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

re-arrange

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

from information given

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

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

If the K_a of the weak acid HA is 2 x 10^{-4} mol dm^{-3}.

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

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

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 x 10^{-5} mol dm^{-3}.

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

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 x 10^{-5} mol dm^{-3}.

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

re-arrange $\qquad [H^+(aq)] \quad = \quad \dfrac{[HX(aq)] \; K_a}{[X^-(aq)]}$

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 x 10^{-5} mol dm^{-3}.

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

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

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}

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 x 10^{-5} mol dm^{-3}.

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

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

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}

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

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

Buffer Solutions

Problem: What is the pH of a buffer that has [HOAc] = 0.700 M and [OAc$^-$] = 0.600 M?

$$HOAc + H_2O \rightleftharpoons OAc^- + H_3O^+$$

$$K_a = 1.8 \times 10^{-5}$$

	[HOAc]	[OAc$^-$]	[H$_3$O$^+$]
equilib	0.700 - x	0.600 + x	x
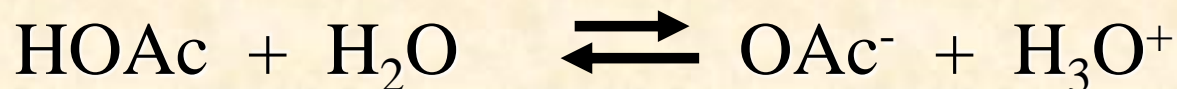
Assuming that x << 0.700 and 0.600, we have

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+](0.600)}{0.700}$$

[H$_3$O$^+$] = 2.1 x 10^{-5} and pH = 4.68

$$pH = pKa + \log \frac{[A^-]}{[\text{HA}]}$$

$$pH = 4.74 + \log \frac{[0.600]}{[0.700]}$$

$$pH = 4.68$$

Ex: Calculate pH change when 1.0 mL of 1.0 M HCl is added to 1.0 L acetic acid/sodium acetate buffer with $[CH_3COOH] = 0.700$ M and $[CH_3COO^-] = 0.600$ M

$$CH_3COOH_{(aq)} \leftrightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$$

0.701 M   0.599 M
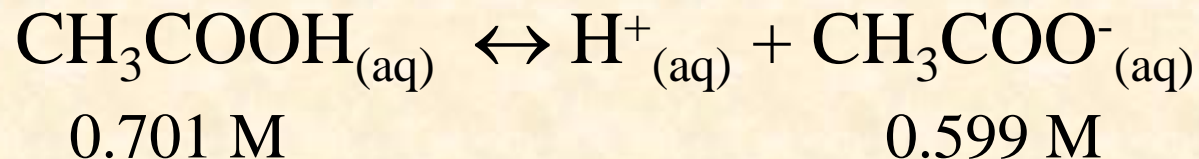
$$[H_3O^+] = ka\frac{[CH_3COOH]}{[CH_3COO^-]}$$

$$[H_3O^+] = (1.8x10^{-5})\frac{[0.701]}{[0.599]}$$

$$[H_3O^+] = 2.1x10^{-5}M$$

$$pH = -\log[H_3O^+]$$

$$pH = 4.67$$

The Henderson Hasselbalch Equation (a little short cut)

- Ex: Calculate pH change when 1.0 mL of 1.0 M HCl is added to 1.0 L acetic acid/sodium acetate buffer with $[CH_3COOH]$ = 0.700 M and $[CH_3COO^-]$ = 0.600 M
  - As before, after the buffer has neutralized the added HCl, the $[CH3COOH]$ in the buffer will be 0.701 M and the $[CH3COO^-]$ in the buffer will be 0.599 M.
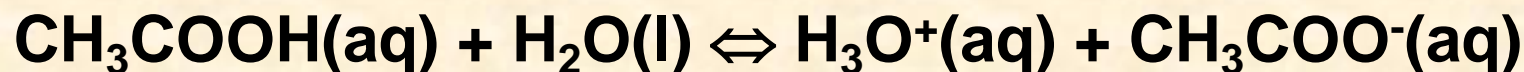
$$pH = pKa + \log\frac{[A^-]}{[HA]}$$

$$pH = 4.74 + \log\frac{[0.599]}{[0.701]}$$

$$pH = 4.67$$

Calculate the pH of a buffer solution that is 0.040 M CH_3COONa (aq) and 0.080 M CH_3COOH (aq).

$pK_a(CH_3COOH(aq)) = 4.75$

$CH_3COOH(aq) + H_2O(l) \Leftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

	$CH_3COOH(aq)$	$H_3O^+(aq)$	$CH_3COO^-(aq)$
Initial	0.080 M	0	0.040M
Change	-x	x	0.040 + x
Equilibrium	0.080 - x	x	0.040 + x

K_a = (0.040 + x) x / (0.080 - x)

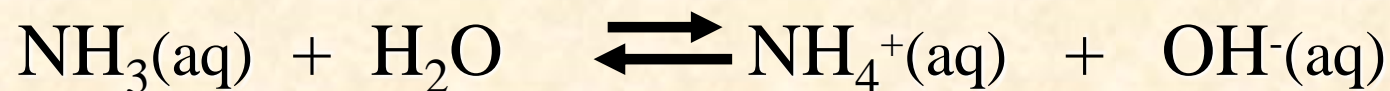**Assume x << 0.040**

x = 3.6 x 10^{-5}

pH = 4.44

$$pH = pKa + \log\frac{[A^-]}{[HA]}$$

$$pH = 4.75 + \log\frac{[0.040]}{[0.080]}$$

$$pH = 4.44$$

The Common Ion Effect

QUESTION: What is the effect on the pH of adding NH_4Cl to 0.25 M NH_3(aq)?

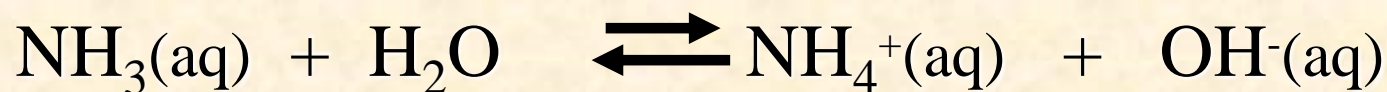$$NH_3(aq) \; + \; H_2O \; \rightleftharpoons \; NH_4^+(aq) \; + \; OH^-(aq)$$

Let us first calculate the pH of a 0.25 M NH_3 solution.

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
initial	0.25	0	0
change	-x	+x	+x
equilib	0.25 - x	x	x

The Common Ion Effect

QUESTION: What is the effect on the pH of adding NH_4Cl to 0.25 M $NH_3(aq)$?

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

$$K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{x^2}{0.25 - x}$$

Assuming x is $<<$ 0.25, we have

$[OH^-] = x = [K_b(0.25)]^{1/2} = 0.0021$ M

This gives pOH = 2.67 and so - - -

pH = 14.00 - 2.67 = 11.33 for 0.25 M NH_3

What is the pH of a buffer that is 0.12 M in lactic acid, $CH_3CH(OH)COOH$, and 0.10 M in sodium lactate? K_a for lactic acid is 1.4×10^{-4}.

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

$$pH = -\log (1.4 \times 10^{-4}) + \log \frac{(0.10)}{(0.12)}$$

$$pH = 3.85 + (-0.08)$$

$$pH = 3.77$$

Suppose you want to make a pH 5.00 buffer---using acetic acid (HA)and sodium acetate (A^-)

pK_a of acetic acid = 4.76;

What ratio of HA and A should be used?

$5.00 = 4.76 + \log x$;   where $x = ([A^-]/[HA])$

$0.24 = \log x$

$\qquad 10^{0.24} = x = 1.74 =$ ratio of moles conjugate base to acid in solution

Example :- Calculate the pH of a buffer prepared by adding 10mL of 0.1M acetic acid to 20mL of 0.1M sodium acetate ?

$$pH = pK_a + \log\frac{[salt]}{[acid]}$$

$$pH = -\log 1.75\times10^{-5} + \log\frac{\frac{0.1\times20}{30}}{\frac{0.1\times10}{30}}$$

$$pH = 4.76 + \log 2.0 = 5.06$$

We can use millimoles of acids and salt in place of molarity.

Calculate the volume of concentrated ammonia and the weight of ammonium chloride you would have take to prepare 100mL of a buffer at pH 10.0 if the final concentration of salt is to be 0.2M ? Where, density of ammonia = 0.9g/cm^3 and percentage = 28%.

pOH=pKb+log[salt]/[base]   pH+POH=14

10+POH=14   POH=14-10=4

4=4.76+log0.2[NH3]   −0.76=log0.2/[NH3]

4=4.76+log0.2/[NH3]

−0.76=log0.2/[NH3]

$$10^{-0.76} = \frac{0.2}{[NH_3]}$$

$$[NH_3] = \frac{0.2}{10^{-0.76}} = \frac{0.2}{0.173} = 1.15M$$

$$M_1V_1 = M_2V_2$$

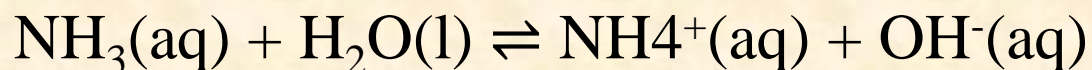$$(\frac{0.28 \times 0.9 \times 1000}{17}) \times V_1 = 1.15 \times 100$$

$$14.8 \times V_1 = 1.15 \times 100 \quad , \quad V_1 = 7.8 \text{ mL}$$

$$M_{NH_4Cl} = \frac{wt}{M.\,wt} \times \frac{1000}{V\,mL}$$

$$0.2 = \frac{wt}{53.5} \times \frac{1000}{100} \qquad wt = 1.07 \text{ gm } NH_4Cl$$

Example: A buffer is prepared from 60.0 mL of 0.100 M NH_3 and 40.0 mL of 0.100 M NH_4Cl. What is the pH of the buffer solution? $K_b = 1.8 \times 10^{-5}$ for NH_3.

Answer: The buffer involves the NH_3/NH_4^+ conjugate acid/base pair.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH4^+(aq) + OH^-(aq)$$

Conc's: Since two solutions are mixed, both have been diluted.

Use this: $V_f = 60.0 \text{ mL} + 40.0 \text{ mL} = 100.0 \text{ mL}$

$[NH_3]_f = V_iM_i/V_f = (60.0 \text{ mL})(0.100 \text{ M})/100.0 \text{ mL} = 0.0600 \text{ M}$

$[NH_4^+]_f = V_iM_i/V_f = (40.0 \text{ mL})(0.100 \text{ M})/100.0 \text{ mL} = 0.0400 \text{ M}$

Now, we can calculate pH:

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

	NH_3		NH_4^+	OH^-
[init]	0.0600		0.0400	0
[change]	-x		+x	+x
[equil]	(0.0600-x)		(0.0400+x)	x

$$K_b = 1.8 \times 10^{-5} = \frac{x(0.0400+x)}{(0.0600-x)} \approx \frac{0.0400x}{0.0600} \qquad \therefore x = = 2.7 \times 10^{-5} \text{ M}$$

$$\therefore pOH = -\log(2.7 \times 10^{-5}) = 4.57 \qquad \therefore \boxed{pH = 9.43}$$

If the K_a of an acid is 1.0×10^{-6}, pH = 3.00, and the concentration of the unionized acid is 1.0×10^{-2} M, what is the concentration of the salt?

Solution

Given: pH = 3.00, $K_a = 1.0 \times 10^{-6}$ M, and [HA] = 1.0×10^{-2} M

solve for $[H_3O^+] = 10^{-pH} = 10^{-3} = 1.0 \times 10^{-3}$ M

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

rearrange equation to solve for the salt [A$^-$]

$$[A^-] = K_a \frac{[HA]}{[H_3O^+]} = 1.0 \times 10^{-6} \text{ M} \frac{[1.0 \times 10^{-2} \text{ M}]}{[1.0 \times 10^{-3} \text{ M}]}$$

$$[A^-] = 1.0 \times 10^{-5} \text{ M}$$

Q1 (a) The pK_a of HF is 3.17. Calculate the pH of a 1.00-L solution that is 1.00 M HF and 1.50 M NaF.
(b) What is the pH of this solution after addition of 50.0 mLof 10.0 M HCl?

Answer: (a) pH = 3.35
(b) pH = 2.99

Q2;Calculate the pH of a solution prepared by adding 25mL of 0.1M sodium hydroxide to 30mL of 0.2M acetic acid?

• A 1.0 L buffer solution contains 0.50 M NH3 (pKb = 4.75) and 0.20 M NH4Cl. Calculate the pH upon addition of 0.010 mol of solid NaOH.