PART A QUESTION 9	

Given: 10 mL of 0.10 M NaA (a *weak base*)
0.10 M HCl (a *strong acid*)
$K_a = 1.0 \times 10^{-5}$ ($pK_a = 5.00$)

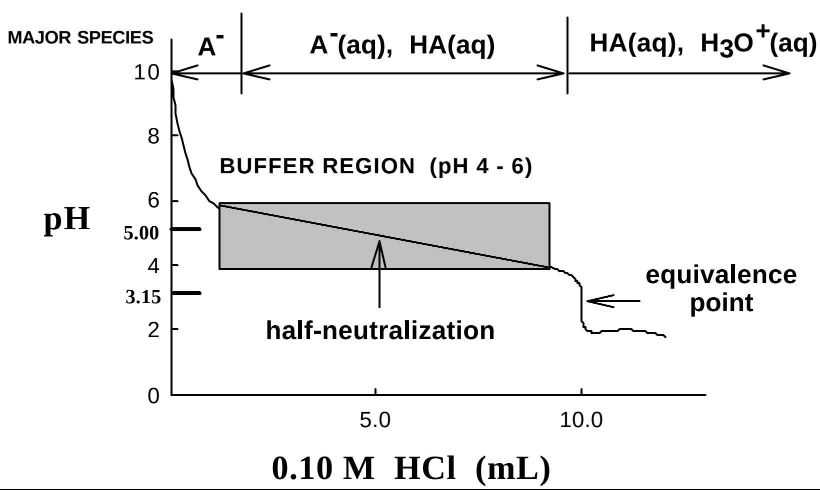
To sketch: a titration curve of NaA with HCl (**WB + SA**)

(a) addition of 5.0 mL HCl

NIE:	$A^-(aq)$ +	$H_3O^+(aq)$ →	$HA(aq) + H_2O(\ell)$	(**WB + SA**)
I:	*10 mL, 0.10 M*	***5.0 mL**, 0.10 M*	–	
	1.0×10^{-3} *mol*	5.0×10^{-4} *mol*	–	
C:	⊖ 5.0×10^{-4} mol	⊖ 5.0×10^{-4} mol	⊕ 5.0×10^{-4} mol	
E:	**5.0×10^{-4} mol**	–	**5.0×10^{-4} mol**	

(b) addition of 10 mL HCcl

NIE:	$A^-(aq)$ +	$H_3O^+(aq)$ →	$HA(aq) + H_2O(\ell)$	(**WB + SA**)
I:	*10 mL, 0.10 M*	***10 mL**, 0.10 M*	–	
	1.0×10^{-3} *mol*	1.0×10^{-3} *mol*	–	
C:	⊖ 1.0×10^{-4} mol	⊖ 1.0×10^{-3} mol	⊕ 1.0×10^{-3} mol	
E:	–	–	**1.0×10^{-3} mol**	



At half-neutralization

(1) $n(A^-)$ leftover = $n(HA)$ produced
= 5.0×10^{-4} mol
(*i.e.* $[A^-] = [HA]$ or **WB = WA**)
(2) 1:1 buffer
(3) **$pH = pK_a$** = 5.00
(4) major species: $A^-(aq)$, $HA(aq)$, $Na^+(aq)$, $Cl^-(aq)$

At complete neutralization (*i.e.* at **equivalence point**)

(1) $n(A^-)$ consumed = $n(H_3O^+)$ reacted
= 1.0×10^{-3} mol (*i.e.* **WB = SA**)
(2) $n(HA)$ produced = 1.0×10^{-3} mol
(*i.e. all* $A^-(aq)$ is *theoretically* converted to $HA(aq)$)
(3) [HA] = 0.10 M ÷ 2 = 0.050 M (∵ volume double)
(4) pH at equivalence point = 3.15
(5) major species: 100% HA, $Na^+(aq)$, $Cl^-(aq)$

PART A QUESTION 10

Given:
① 250 mL of 0.10 M $CH_3COOH(aq)$ (a *weak acid*)
② 25 mL of NaOH (a *strong base*)
③ pH = 4.90
④ K_a (CH_3COOH) = 1.8×10^{-5} M
∴ pK_a (CH_3COOH) = − log K_a = − log (1.8×10^{-5}) = 4.745

To find: c(NaOH)

WA + SB Titration

	$CH_3COOH(aq)$	+	$OH^-(aq)$	⇌	$CH_3COO^-(aq)$	+	$H_2O(\ell)$
I:	*250 mL*		*25 mL*		–		
	0.10 M		*? M*		–		
	0.025 mol		***x***		–		
C:	⊖ ***x***		⊖ ***x***		⊕ ***x***		
E:	**(0.025 − *x*) mol**		–		***x***		

(*where* ***x*** = n(OH^-) needed to neutralize with CH_3COOH to reach pH 4.90)
x = n(CH_3COO^-) produced = n(CH_3COOH) ionized

$$pH = pK_a + \log \frac{[WB]}{[WA]}$$

$$pH = pK_a\ (CH_3COOH) + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$$

$$4.90 = 4.745 + \log \frac{x}{0.025 - x}$$

x = 1.47×10^{-2} mol = n(OH^-) required

∴ c(NaOH) = n(NaOH) ÷ V(NaOH)
= (1.47×10^{-2} mol) ÷ (25×10^{-3} L)
= **0.59 mol L^{-1}**

PART A
QUESTION 11

Given: ① $[C_2H_5COOH] = 0.100\ M$ (a *weak acid*)

In a total of 1.00 L, $n(C_2H_5COOH) = 0.100\ mol$

② $[C_2H_5COONa] = [C_2H_5COO^-] = 0.150\ M$ (its *conjugate weak base*)

In a total of 1.00 L, $n(C_2H_5COO^-) = 0.150\ mol$

③ n(HCl) added = 0.020 mol

④ $K_a\ (C_2H_5COOH) = 1.34 \times 10^{-5}\ M$

$\therefore\ pK_a\ (C_2H_5COOH) = -\log K_a = -\log(1.34 \times 10^{-5}) = 4.\underline{873}$

To find: the change in pH (ΔpH)

BEFORE addition of H_3O^+

$$pH = pK_a + \log \frac{[WB]}{[WA]}$$

$$pH = pK_a\ (C_2H_5COOH) + \log \frac{[C_2H_5COO^-]}{[C_2H_5COOH]}$$

$$pH = 4.873 + \log \frac{0.150\ M\ /\ 1\ L}{0.100\ M\ /\ 1\ L} = 4.873 + (0.176) = 5.049$$

AFTER addition of 0.020 mol H_3O^+

∴ SA + WB Titration

	$H_3O^+(aq)$ +	$C_2H_5COO^-(aq)$ ⇌	$C_2H_5COOH(aq)$ +	$H_2O(\ell)$
I:	*0.020 mol*	*0.150 mol*	*0.100 mol*	
C:	⊖ 0.020 mol	⊖ 0.020 mol	⊕ 0.020 mol	
E:	–	**0.130 mol**	**0.120 mol**	

$$pH = 4.873 + \log \frac{0.130\ mol\ /\ \sim 1.00\ L}{0.120\ mol\ /\ \sim 1.00\ L} = 4.873 + (0.03476) = 4.9078$$

$\therefore\ \Delta pH = \text{final pH} - \text{initial pH} = 4.\underline{907}8 - 5.\underline{049} = -0.\underline{141}2$

$= \underline{\underline{\mathbf{-0.141}}}$ (***decreased*** by 0.141 pH units)

PART A QUESTION 12

Given: [sodium butyrate] = $[Bu^-]$ = 0.25 M (a *weak base*)
[butyric acid] = [HBu] = ? (its *conjugate weak acid*)
K_a (HBu) = 1.5×10^{-5} M
∴ pK_a (HBu) = $-\log K_a$ = $-\log(1.5 \times 10^{-5})$ = 4.824

(a) pH of a 0.25 M unbuffered solution of sodium butyrate

	$Bu^-(aq)$ +	$H_2O(\ell)$ ⇌	$HBu(aq)$ +	$OH^-(aq)$
I:	*0.25 M*		–	–
C:	⊖ ***x***		⊕ ***x***	⊕ ***x***
E:	$(0.25 - x)$ M		***x***	***x***

(where ***x*** = $[OH^-]$ produced = [HBu] produced = $[Bu^-]$ ionized)

$$K_c = \frac{[HBu]\,[OH^-]}{[Bu^-]} = K_b\,(Bu^-) = \frac{K_w}{K_a\,(HBu)}$$

$$\frac{(x)\,(x)}{0.25 - x} = \frac{1.0 \times 10^{-14}\ M^2}{1.5 \times 10^{-5}\ M} = 6.67 \times 10^{-10}$$

1st approximation: 1.291×10^{-5} M
2nd approximation: **1.291×10^{-5} M**

x = 1.291×10^{-5} M = $[OH^-]$

pOH = − log [OH⁻]

pOH = $-\log(1.291 \times 10^{-5})$ = 4.889

∴ pH = 14.00 − pOH = 14.00 − 4.889 = 9.111 = **9.11**

(b) net ionic equation and the equilibrium constant for the reaction

! HCl(aq) is a ***strong acid***, and must react with the ***conjugate base*** buffer component (*i.e.* Bu^-).

$$H_3O^+(aq) + Bu^-(aq) \rightleftharpoons HBu(aq) + H_2O(\ell)$$

$$K_c = \frac{[HBu]}{[H_3O^+]\,[Bu^-]} = \frac{1}{K_a\,(HBu)}$$

$$= \frac{1}{1.5 \times 10^{-5}\ M} = 6.7 \times 10^{4}\ M^{-1}$$

(c) resultant pH after addition of 20.0 mL of 0.100 M HCl

	$H_3O^+(aq)$	+	$Bu^-(aq)$	⇌	$HBu(aq)$ +	$H_2O(\ell)$
I:	*20.0 mL*		*100 mL*		–	
	1.00 M		*0.25 M*		–	
	0.0200 mol		*0.025 mol*		–	
C:	⊖ 0.0200 mol		⊖ 0.0200 mol		⊕ 0.0200 mol	
E:	–		**0.005 mol**		**0.0200 mol**	

$$pH = pK_a + \log \frac{[WB]}{[WA]}$$

$$pH = pK_a\ (HBu) + \log \frac{[Bu^-]}{[HBu]}$$

① K_a (HBu) = 1.5×10^{-5} M

∴ pK_a (HBu) = − log K_a = − log (1.5×10^{-5}) = 4.824

② *total volume = 100 mL $Bu^-(aq)$ + 20.0 mL HCl = 120 mL*

$$pH = 4.824 + \log \frac{0.005\ mol\ /\ 120\ mL}{0.0200\ mol\ /\ 120\ mL}$$

∴ pH = 4.824 + (−0.602) = 4.222 = **4.22**

(d) Is the solutions prepared in Part (b) a buffer solution? Explain.

Yes, it is a buffer. This buffer is prepared by mixing a *strong acid* (HCl) with a *weak base* (sodium butyrate, Bu^-), where HCl(aq) is a **limiting reactant**. Some of $Bu^-(aq)$ is converted to HBu(aq) by the addition of *strong acid*, Hcl(aq).

PART A
QUESTION 13

<u>Given</u>: V(water sample) = 250 mL
alkalinity of water = 4.31 mmol H^+ per litre
= 4.31×10^{-3} mol H^+ per L
= acidity consumed per L
V(HCl) = 20.48 mL to reach the methyl orange endpoint

<u>To find</u>: c(HCl)

$$\begin{aligned}
\mathbf{n(H^+)} &= \textbf{(acidity consumed per L)} \times \textbf{V(water sample)} \\
&= \text{(alkalinity of water sample)} \times \text{V(water sample)} \\
&= (4.31 \times 10^{-3}\ \text{mol H}^+\ \text{per L}) \times (250 \times 10^{-3}\ \text{L}) \\
&= \underline{1.07}75 \times 10^{-3}\ \text{mol H}^+
\end{aligned}$$

$$\begin{aligned}
\therefore \quad \mathbf{c(HCl)} &= \mathbf{n(H^+) \div V(HCl)} \\
&= (\underline{1.07}75 \times 10^{-3}\ \text{mol H}^+) \div (20.48 \times 10^{-3}\ \text{L}) \\
&= \underline{5.26}12 \times 10^{-2}\ \text{mol L}^{-1} \\
&= \underline{\underline{\mathbf{5.26 \times 10^{-2}\ M}}}
\end{aligned}$$