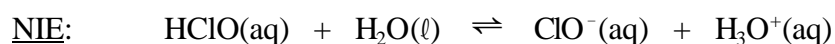


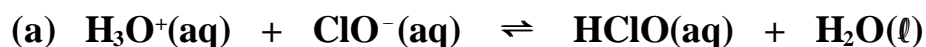
PART A
QUESTION 3

Given: hypochlorous acid (HClO, a *weak acid*)

$$K_a(\text{HClO}) = 3.0 \times 10^{-8} \text{ M (at } 25^\circ\text{C)}$$

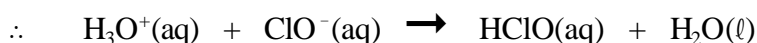


$$K_c = \frac{[\text{ClO}^{\text{-}}][\text{H}_3\text{O}^{\text{+}}]}{[\text{HClO}]} = K_a(\text{HClO}) = 3.0 \times 10^{-8} \text{ M}$$

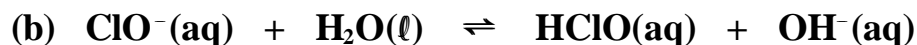


$$K_c = \frac{[\text{HClO}]}{[\text{H}_3\text{O}^{\text{+}}][\text{ClO}^{\text{-}}]} = \frac{1}{K_a(\text{HClO})}$$

$$K_c = \frac{1}{K_a(\text{HClO})} = \frac{1}{3.0 \times 10^{-8} \text{ M}} = 3.3 \times 10^7 \text{ M}^{-1} \quad (\text{LARGE})$$

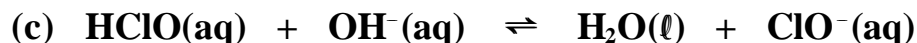


($\text{H}_3\text{O}^{\text{+}}(\text{aq})$, being a *strong acid*, drives the reaction to completion.)



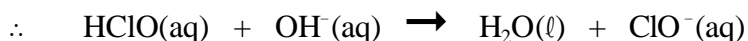
$$K_c = \frac{[\text{HClO}][\text{OH}^{\text{-}}]}{[\text{ClO}^{\text{-}}]} = K_b(\text{ClO}^{\text{-}}) = \frac{K_w}{K_a(\text{HClO})}$$

$$K_c = K_b(\text{ClO}^{\text{-}}) = \frac{1.0 \times 10^{-14} \text{ M}^2}{3.0 \times 10^{-8} \text{ M}} = 3.3 \times 10^{-7} \text{ M}$$



$$K_c = \frac{[\text{ClO}^{\text{-}}]}{[\text{HClO}][\text{OH}^{\text{-}}]} = \frac{1}{K_b(\text{ClO}^{\text{-}})} = \frac{K_a(\text{HClO})}{K_w}$$

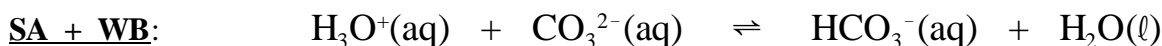
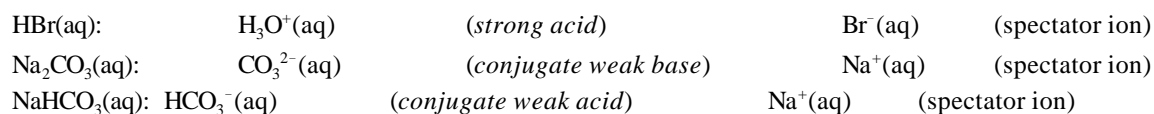
$$K_c = \frac{1}{K_b(\text{ClO}^{\text{-}})} = \frac{3.0 \times 10^{-8} \text{ M}}{1.0 \times 10^{-14} \text{ M}^2} = 3.0 \times 10^6 \text{ M}^{-1} \quad (\text{LARGE})$$



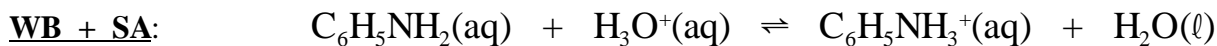
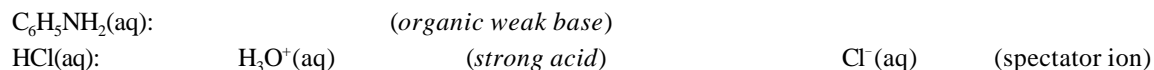
($\text{OH}^{\text{-}}(\text{aq})$, being a *strong base*, drives the reaction to completion.)

PART A
QUESTION 4

To write: a balanced net ionic equation
To calculate: equilibrium constant (K_c)

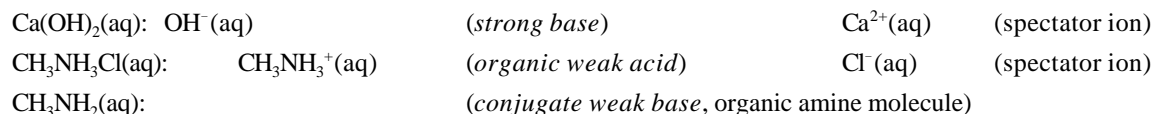
(a) HBr(aq) + Na₂CO₃-NaHCO₃ buffer [$K_a(\text{HCO}_3^-)$: 5.6×10^{-11}]


$$K_c = \frac{[\text{HCO}_3^-]}{[\text{H}_3\text{O}^+][\text{CO}_3^{2-}]} = \frac{1}{K_a(\text{HCO}_3^-)} = \frac{1}{5.6 \times 10^{-11}} = 1.8 \times 10^{10}$$

(b) C₆H₅NH₂(aq) + HCl(aq) [$K_b(\text{C}_6\text{H}_5\text{NH}_2)$: 3.8×10^{-10}]


$$K_c = \frac{[\text{C}_6\text{H}_5\text{NH}_3^+]}{[\text{C}_6\text{H}_5\text{NH}_2][\text{H}_3\text{O}^+]} = \frac{1}{K_a(\text{C}_6\text{H}_5\text{NH}_3^+)} = \frac{K_b(\text{C}_6\text{H}_5\text{NH}_2)}{K_w}$$

$$K_c = \frac{K_b(\text{C}_6\text{H}_5\text{NH}_2)}{K_w} = \frac{3.8 \times 10^{-10}}{1.0 \times 10^{-14}} = 3.8 \times 10^4$$

(c) Ca(OH)₂(aq) + CH₃NH₃Cl-CH₃NH₂ buffer [$K_b(\text{CH}_3\text{NH}_2)$: 4.2×10^{-4}]


$$K_c = \frac{[\text{CH}_3\text{NH}_2]}{[\text{OH}^-][\text{CH}_3\text{NH}_3^+]} = \frac{1}{K_b(\text{CH}_3\text{NH}_2)} = \frac{1}{4.2 \times 10^{-4}} = 2.4 \times 10^3$$

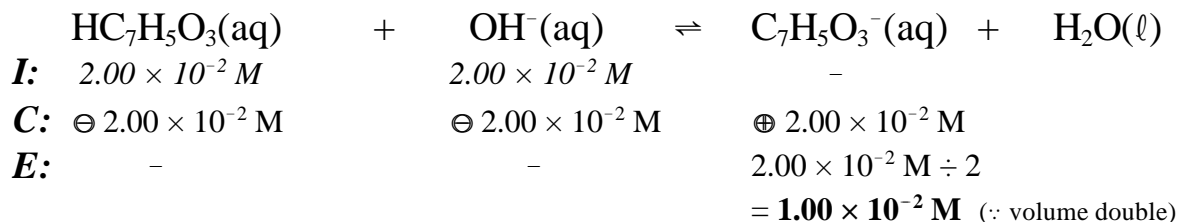
! In all cases, the **large** K_c values indicate the completion of the reaction due to the presence of either strong base, (Ca(OH)₂(aq)) or strong acids (HBr(aq) or HCl(aq)).

! **ACID must react with BASE**, neither SA reacts with WA nor SB reacts with WB.

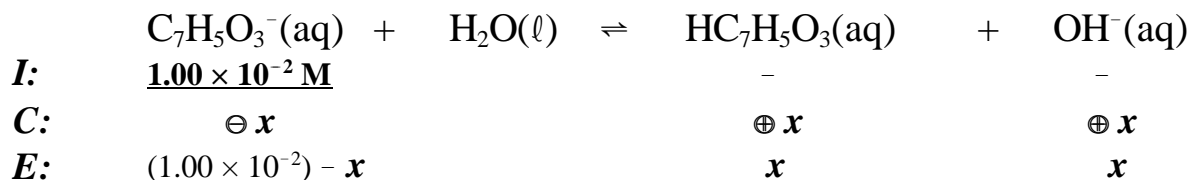
PART A
QUESTION 5

- Given:
- ① $[\text{HC}_7\text{H}_5\text{O}_3] = 2.00 \times 10^{-2} \text{ M}$ (salicylic acid, a *weak acid*)
 - ② $K_a [\text{HC}_7\text{H}_5\text{O}_3] = 1.49 \times 10^{-3} \text{ M}$ (at 25°C)
 - ③ $[\text{NaOH}] = 2.00 \times 10^{-2} \text{ M}$ (a *strong base*)

To find: pH at the equivalence point


At equivalence point,

- (1) All $\text{HC}_7\text{H}_5\text{O}_3(\text{aq})$ is consumed and converted into $\text{C}_7\text{H}_5\text{O}_3^-(\text{aq})$.
(i.e. Theoretically **no more $\text{HC}_7\text{H}_5\text{O}_3(\text{aq})$ LEFT**, but *only* $\text{C}_7\text{H}_5\text{O}_3^-(\text{aq})$ is present.)
- (2) Since $\text{C}_7\text{H}_5\text{O}_3^-(\text{aq})$ is a *weak base*, which only partially ionizes in water until the equilibrium is reached.



where $x = [\text{OH}^-] \text{ produced} = [\text{HC}_7\text{H}_5\text{O}_3] \text{ produced} = [\text{C}_7\text{H}_5\text{O}_3^-] \text{ ionized}$

$$K_c = \frac{[\text{HC}_7\text{H}_5\text{O}_3][\text{OH}^-]}{[\text{C}_7\text{H}_5\text{O}_3^-]} = K_b(\text{C}_7\text{H}_5\text{O}_3^-) = \frac{K_w}{K_a(\text{HC}_7\text{H}_5\text{O}_3)}$$

$$\frac{(x)(x)}{(1.00 \times 10^{-2}) - x} = \frac{1.00 \times 10^{-14} \text{ M}^2}{1.49 \times 10^{-3} \text{ M}}$$

$$x = \underline{2.591 \times 10^{-7} \text{ M}} = [\text{OH}^-]$$

$$\boxed{\text{pOH} = -\log [\text{OH}^-]}$$

$$\text{pOH} = -\log(2.591 \times 10^{-7}) = \underline{6.5866}$$

$$\therefore \text{pH at the equivalence point} = 14.000 - \text{pOH} = 14.000 - 6.5866 = \underline{7.413}$$

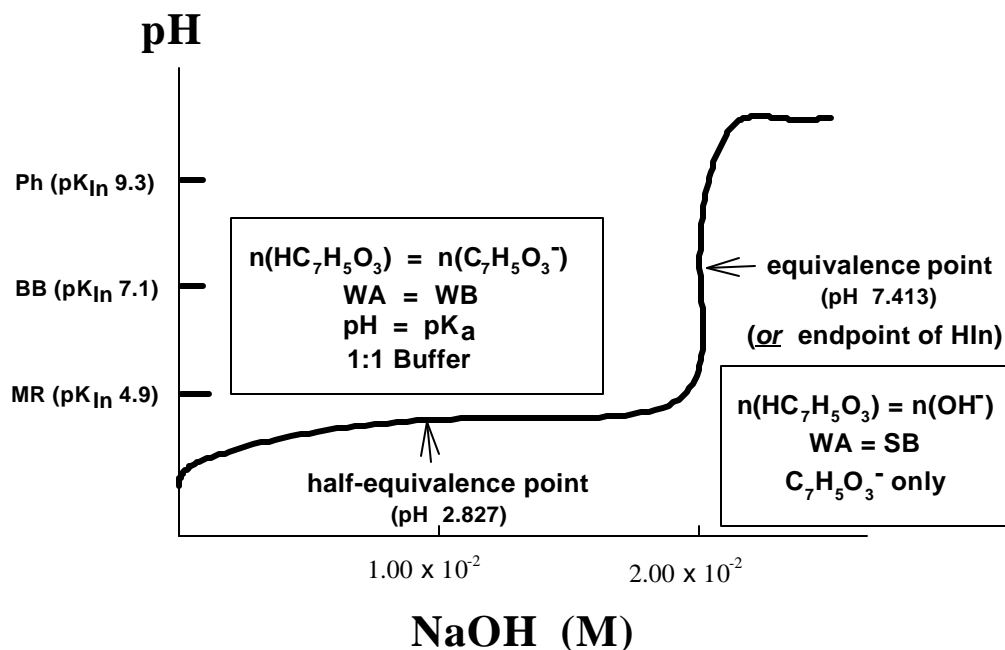
PART A
QUESTION 5

Salicylic acid (WA) + NaOH (SB)

Best **INDICATOR** chosen for the titration should have: $pK_{In} \cong \text{pH at equivalence point}$

Choose the indicator whose pK_{In} is **close to 7.413** (*i.e.* pH at the equivalence point)

1. methyl red (MR, pK_{In} 4.9)
→ not suitable (color change *far before* equivalence point)
2. bromothymol blue (BB, pK_{In} 7.1)
→ *best* (very close to pH at equivalence point)
3. phenolphthalein (Ph, pK_{In} 9.3)
→ O.K. (colour change *after* equivalence point)

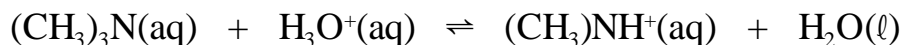


In general,

- ① **WA + SB:** bromothymol blue (pK_{In} 7.1); phenolphthalein (pK_{In} 9.3)
- ② **SA + SB:** methyl red (pK_{In} 4.9);
bromothymol blue (pK_{In} 7.1); phenolphthalein (pK_{In} 9.3)
- ③ **WB + SA:** methyl red (pK_{In} 4.9); methyl orange (pK_{In} 3.4)

PART A
QUESTION 6

- Given:
- ① $(\text{CH}_3)_3\text{N}(\text{aq}) = 0.100 \text{ M}$ (a weak base)
 - ② $K_b(\text{CH}_3)_3\text{N}(\text{aq}) = 6.1 \times 10^{-5} \text{ M}$
 - ③ $[\text{HClO}_4(\text{aq})] = 0.100 \text{ M} = [\text{H}_3\text{O}^+]$ (a strong acid)

WB + SA


$$K_c = \frac{[(\text{CH}_3)_3\text{NH}^+]}{[(\text{CH}_3)_3\text{N}][\text{H}_3\text{O}^+]} = \frac{1}{K_a [(\text{CH}_3)_3\text{NH}^+]}$$

(a) pH at half-equivalence point (or half-neutralization point)

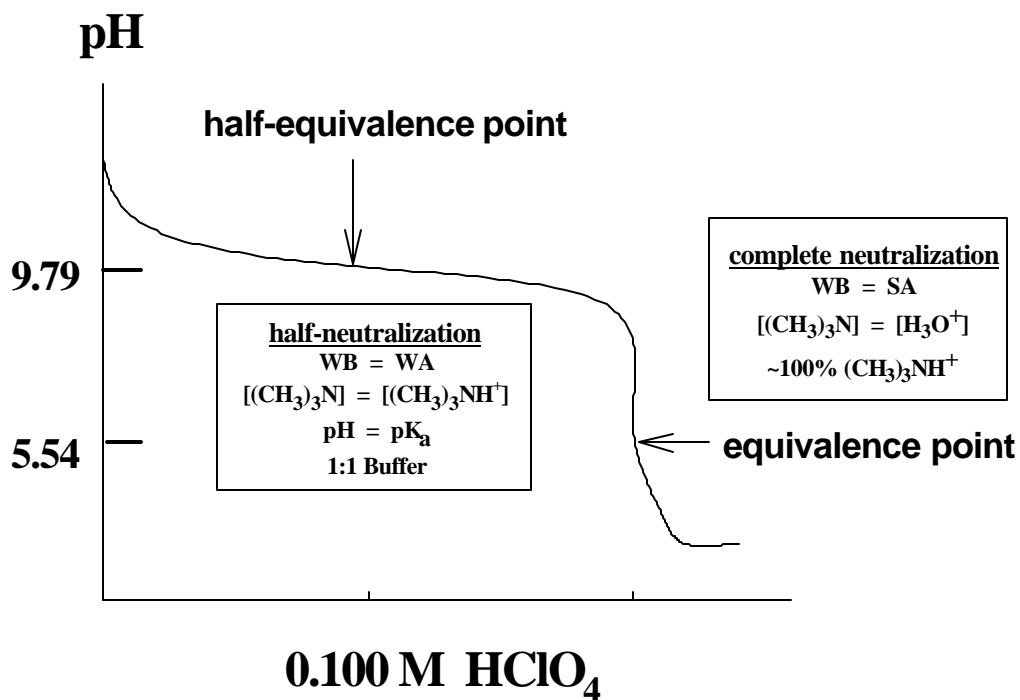
At half-equivalence point, $\text{pH} = \text{p}K_a [(\text{CH}_3)_3\text{NH}^+]$

$$K_a [(\text{CH}_3)_3\text{NH}^+] = \frac{K_w}{K_b [(\text{CH}_3)_3\text{N}]} = \frac{1.0 \times 10^{-14} \text{ M}^2}{6.1 \times 10^{-5} \text{ M}} = 1.64 \times 10^{-10} \text{ M}$$

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log K_a = -\log (1.64 \times 10^{-10}) = \underline{\underline{9.79}}$$

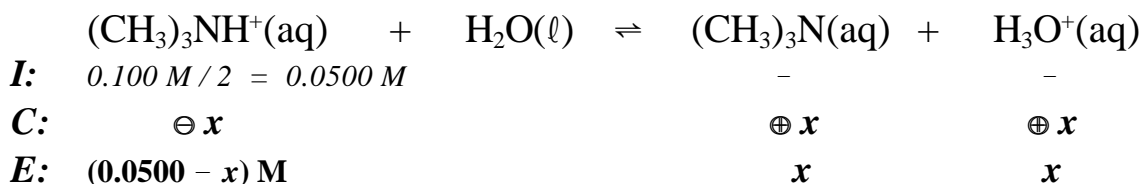
$$\therefore \text{pH at half-equivalence point} = \text{p}K_a [(\text{CH}_3)_3\text{NH}^+] = \underline{\underline{9.79}}$$



PART A
QUESTION 6

(b) pH at equivalence point

- (1) All Weak Base $(\text{CH}_3)_3\text{N}(\text{aq})$ are *theoretically* converted into Weak Acid $(\text{CH}_3)_3\text{NH}^+(\text{aq})$.
- (2) Since $(\text{CH}_3)_3\text{NH}^+(\text{aq})$ is a *weak acid*, which only partially dissociates in water until the equilibrium is reached.
- (3) Since the volume of the solution is doubled, the concentration of $(\text{CH}_3)_3\text{NH}^+(\text{aq})$ is reduced by half to 0.0500 M.



where $x = [\text{H}_3\text{O}^+] \text{ produced} = [(\text{CH}_3)_3\text{N}] \text{ produced} = [(\text{CH}_3)_3\text{NH}^+] \text{ dissociated}$

$$K_c = \frac{[(\text{CH}_3)_3\text{N}] [\text{H}_3\text{O}^+]}{[(\text{CH}_3)_3\text{NH}^+]} = K_a [(\text{CH}_3)_3\text{NH}^+] = \frac{K_w}{K_b [(\text{CH}_3)_3\text{N}]}$$

$$\frac{(x)(x)}{(0.0500 - x) \text{ M}} = \frac{1.0 \times 10^{-14} \text{ M}^2}{6.1 \times 10^{-5} \text{ M}}$$

$$x = \underline{2.86 \times 10^{-6} \text{ M}} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

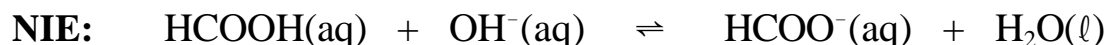
$$\therefore \text{pH at equivalence point} = -\log (2.86 \times 10^{-6}) = 5.543 = \underline{\underline{5.54}}$$

PART A
QUESTION 7

There are *only three* different methods to prepare buffer:
 (a) **WA** (e.g. HA) + **SB** (e.g. NaOH)
 (b) **WB** (e.g. A⁻) + **SA** (e.g. HCl)
 (c) **WA** (e.g. HA) + **WB** (e.g. A⁻)

<u>Given:</u>	0.10 M aqueous solutions available			
formic acid	(HCOOH)	HCOOH(aq)	<i>weak acid</i>	(WA)
sodium formate	(HCOONa)	HCOO ⁻ (aq)	<i>weak base</i>	(WB)
perchloric acid	(HClO ₄)	H ₃ O ⁺ (aq)	<i>strong acid</i>	(SA)
sodium hydroxide	(NaOH)	OH ⁻ (aq)	<i>strong base</i>	(SB)

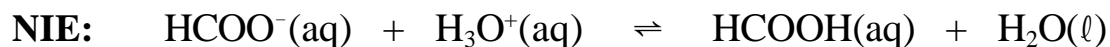
(a) **WA + SB**: **formic acid + sodium hydroxide**



$$K_c = \frac{[\text{HCOO}^{-}]}{[\text{HCOOH}][\text{OH}^{-}]} = \frac{1}{K_b(\text{HCOO}^{-})} = \frac{K_a(\text{HCOOH})}{K_w}$$

! A *strong base*, OH⁻(aq) (being a **limiting reagent**), converts some *weak acid*, HCOOH(aq), into its *conjugate base buffer component*, HCOO⁻(aq).

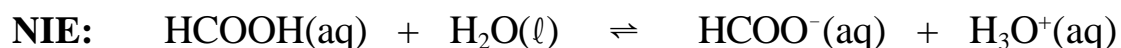
(b) **WB + SA**: **sodium formate + perchloric acid**



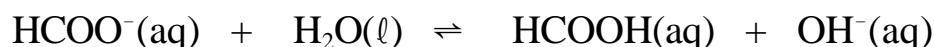
$$K_c = \frac{[\text{HCOOH}]}{[\text{HCOO}^{-}][\text{H}_3\text{O}^{+}]} = \frac{1}{K_a(\text{HCOOH})}$$

! A *strong acid*, H₃O⁺(aq) (being a **limiting reagent**), converts some *weak base*, HCOO⁻(aq), into its *conjugate acid buffer component*, HCOOH(aq).

(c) **WA + WB**: **formic acid + sodium formate**



OR



$$K_c = K_a(\text{HCOOH}) \quad \text{OR} \quad K_c = K_b(\text{HCOO}^{-})$$

PART A
QUESTION 8

Given: $[\text{HNO}_2] = 0.10 \text{ M}$ (HNO_2 is a *weak acid*)
 $[\text{NaNO}_2] = 0.10 \text{ M}$ (NO_2^- is its *conjugate base*)

$$K_a(\text{HNO}_2) = 4.6 \times 10^{-4}$$
$$\therefore \text{p}K_a(\text{HNO}_2) = -\log K_a = -\log(4.6 \times 10^{-4}) = 3.337$$

To find: ratio of NO_2^- : HNO_2 at $\text{pH} = 3.0$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{WB}]}{[\text{WA}]}$$

$$\text{pH} = \text{p}K_a(\text{HNO}_2) + \log \frac{[\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$3.0 = 3.337 + \log \frac{[\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$-0.337 = \log \frac{[\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$\therefore \frac{[\text{NO}_2^-]}{[\text{HNO}_2]} = 0.46 = \frac{0.46}{1} = \frac{46}{100}$$

$$\% \text{NO}_2^- = \frac{46}{46 + 100} \times 100\% = 31.5\%$$

$$\% \text{HNO}_2 = \frac{100}{46 + 100} \times 100\% = 68.5\%$$

$$\therefore \text{ratio of } \text{NO}_2^- : \text{HNO}_2 = \underline{\underline{0.46 : 1.0}}$$