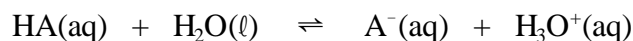


## USEFUL CONCEPTS AND FORMULAS

(1) 
$$\text{pH} = -\log [\text{H}_3\text{O}^+] \quad \text{and} \quad \text{pOH} = -\log [\text{OH}^-] \quad \text{and} \quad \text{pK}_a = -\log K_a$$

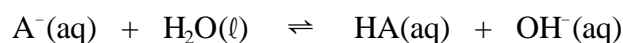
- (2) There are *only six* strong acids: HCl, HBr, HI, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub> and HClO<sub>4</sub>.  
Any other *protonated* acids are weak acids (HA) with K<sub>a</sub> value less than 1.

- (3) **Weak Acid (HA) and Weak Base (A<sup>-</sup>)**



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

! K<sub>c</sub> is usually a small value because HA(aq) is a **weak acid**, which *only partially dissociates* in aqueous solution.



$$K_c = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} = K_b(\text{A}^-)$$

! K<sub>c</sub> is usually a small value because A<sup>-</sup>(aq) is a **weak base**, which *only partially ionizes* in aqueous solution.

(4) 
$$K_w(\text{H}_2\text{O}) = K_a(\text{HA}) \times K_b(\text{A}^-) = 1.0 \times 10^{-14} \text{ M}^2$$

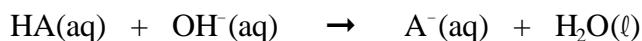
- (5) **Henderson-Hasselbalch equation:**

$$\text{pH} = \text{pK}_a(\text{HA}) + \log \frac{[\text{Weak Base, A}^-]}{[\text{Weak Acid, HA}]}$$

! **Weak Acid** (HA): more positive (∴ one *more* H<sup>+</sup>)

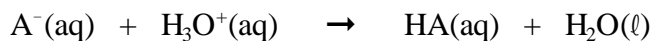
! **Weak Base** (A<sup>-</sup>): more negative (∴ one *less* H<sup>+</sup>)

- (6) **Titration of Weak Acid (HA) with Strong Base (OH<sup>-</sup>)**



$$K_c = \frac{1}{K_b(\text{A}^-)} = \frac{K_a(\text{HA})}{K_w}$$

! K<sub>c</sub> is usually a LARGE value because OH<sup>-</sup>(aq) is a *strong base*, which drives the reaction to completion.

(7) **Titration of Weak Base ( $A^-$ ) with Strong Acid ( $H_3O^+$ )**

$$K_c = \frac{1}{K_a(HA)} = \frac{K_b(A^-)}{K_w}$$

!  $K_c$  is usually a LARGE value because  $H_3O^+(aq)$  is a *strong acid*, which drives the reaction to completion.

(8) Best **BUFFER REGION** should be *within*  $pK_a \pm 1$ .

(9) **Equivalence point:**  $n(HA)$  available =  $n(OH^-)$  used for complete neutralization or  
 $n(A^-)$  available =  $n(H_3O^+)$  used for complete neutralization

**Endpoint:** when indicator changes colour visually

[does not imply  $n(HA) = n(OH^-)$  nor  $n(A^-) = n(H_3O^+)$ ]

[but  $n(HA) \cong n(OH^-)$  or  $n(A^-) \cong n(H_3O^+)$ ]

(10) Best **INDICATOR** chosen for the titration should have:  $pK_{in} \cong$  **pH at equivalence point**

### Acid-Base Properties of Some Common Ions in Aqueous Solution

| ACIDIC ions                 |             | NEUTRAL ions |             | BASIC ions |               |
|-----------------------------|-------------|--------------|-------------|------------|---------------|
| $H^+$                       | $HSO_4^-$   | $Li^+$       | $Cl^-$      | $F^-$      | $HCOO^-$      |
| $NH_4^+$                    | $H_2PO_4^-$ | $Na^+$       | $Br^-$      | $CN^-$     | $C_2H_3O_2^-$ |
| $CH_3NH_3^+$                |             | $K^+$        | $I^-$       | $HS^-$     | $S^{2-}$      |
| $C_5H_5NH^+$                |             | $Mg^{2+}$    | $NO_3^-$    | $NO_2^-$   | $CO_3^{2-}$   |
| $Cu^{2+}, Zn^{2+}, Al^{3+}$ |             | $Ca^{2+}$    | $ClO_4^-$   | $ClO^-$    | $HPO_4^{2-}$  |
| (transition metal ions)     |             | $Ba^{2+}$    | $SO_4^{2-}$ | $HCO_3^-$  | $PO_4^{3-}$   |

|               | <u>cations</u> |           | <u>anion</u> |           | <u>RESULTANT SOLUTION</u> |
|---------------|----------------|-----------|--------------|-----------|---------------------------|
| $NH_4Cl$      | $NH_4^+$       | (acidic)  | $Cl^-$       | (neutral) | <b>acidic</b>             |
| $NaHSO_4$     | $Na^+$         | (neutral) | $HSO_4^-$    | (acidic)  | <b>acidic</b>             |
| $Zn(ClO_4)_2$ | $Zn^{2+}$      | (acidic)  | $ClO_4^-$    | (neutral) | <b>acidic</b>             |
| $MgSO_4$      | $Mg^{2+}$      | (neutral) | $SO_4^{2-}$  | (neutral) | <b>neutral</b>            |
| $Ba(NO_3)_2$  | $Ba^{2+}$      | (neutral) | $NO_3^-$     | (neutral) | <b>neutral</b>            |
| $LiCN$        | $Li^+$         | (neutral) | $CN^-$       | (basic)   | <b>basic</b>              |
| $Na_2S$       | $Na^+$         | (neutral) | $S^{2-}$     | (basic)   | <b>basic</b>              |
| $K_3PO_4$     | $K^+$          | (neutral) | $PO_4^{3-}$  | (basic)   | <b>basic</b>              |

**PART A**  
**QUESTION 1**

| Solutions  | <u>Cation</u>  | <u>Anion</u>   | <u>Resultant Solution</u> | <u>NIE</u><br><u><math>K_c</math> expression</u>   |
|--|--|--|---------------------------|--|
| (i) NaCN   | Na <sup>+</sup> (aq)<br>(neutral)  | CN <sup>-</sup> (aq)<br>(basic)                              | <b>basic</b>              | $\text{CN}^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HCN}(\text{aq}) + \text{OH}^-(\text{aq})$ $K_c = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} = K_b(\text{CN}^-) = \frac{K_w}{K_a(\text{HCN})}$   |
| (ii) LiBr  | Li <sup>+</sup> (aq)<br>(neutral)  | Br <sup>-</sup> (aq)<br>(neutral)                            | <b>neutral</b>            | <b>none</b>  |
| (iii) KH <sub>2</sub> PO <sub>4</sub>  | K <sup>+</sup> (aq)<br>(neutral)   | H <sub>2</sub> PO <sub>4</sub> <sup>-</sup> (aq)<br>(acidic) | <b>acidic</b>             | $\text{H}_2\text{PO}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HPO}_4^{2-}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ $K_c = \frac{[\text{HPO}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^-]} = K_a(\text{H}_2\text{PO}_4^-)$                                       |
| (iv) C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub><br>(organic<br>amine<br>compound) | C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> (aq)<br>a weak base<br>(basic) |  | <b>basic</b>              | $\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$ $K_c = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} = K_b(\text{C}_2\text{H}_5\text{NH}_2)$ |

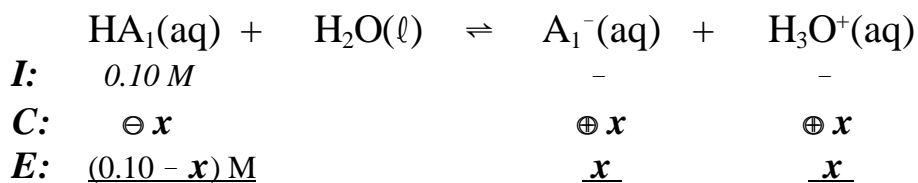
|   |  |                                       |               |   |
|---|--|---------------------------------------|---------------|---|
| (v) $(\text{CH}_3)_3\text{NHCl}$<br>(organic ammonium salt) | $(\text{CH}_3)_3\text{NH}^+$<br>(acidic) | $\text{Cl}^-(\text{aq})$<br>(neutral) | <b>acidic</b> | $(\text{CH}_3)_3\text{NH}^+(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons (\text{CH}_3)_3\text{N}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$<br>$K_c = \frac{[(\text{CH}_3)_3\text{N}] [\text{H}_3\text{O}^+]}{[(\text{CH}_3)_3\text{NH}^+]}$<br>$K_c = K_a [(\text{CH}_3)_3\text{NH}^+] = \frac{K_w}{K_b [(\text{CH}_3)_3\text{N}]}$ |
|---|--|---------------------------------------|---------------|---|

**PART A  
QUESTION 2**

Given: 0.10 M solutions of two weak acids  
 $\text{HA}_1: K_a = 1.5 \times 10^{-5} \text{ M}$   
 $\text{HA}_2: K_a = 4.6 \times 10^{-3} \text{ M}$

To find: pH

(a) 0.10 M  $\text{HA}_1$  ( $K_a = 1.5 \times 10^{-5} \text{ M}$ )



(where  $x = [\text{H}_3\text{O}^+] \text{ produced} = [\text{A}_1^-] \text{ produced} = [\text{HA}_1] \text{ ionized}$ )

$$K_c = \frac{[\text{A}_1^-] [\text{H}_3\text{O}^+]}{[\text{HA}_1]} = K_a (\text{HA}_1)$$

$$\frac{(x)(x)}{0.10 - x} = 1.5 \times 10^{-5}$$

Assuming  $x \ll 1.5 \times 10^{-5}$  (very small)

1<sup>st</sup> approximation:  $1.225 \times 10^{-3} \text{ M}$

2<sup>nd</sup> approximation:  $1.217 \times 10^{-3} \text{ M}$

3<sup>rd</sup> approximation:  $1.217 \times 10^{-3} \text{ M}$

$$\% \text{ dissociation} = \frac{1.217 \times 10^{-3} \text{ M}}{0.10 \text{ M}} \times 100\% = 1.22\% \quad (\because \text{assumption valid})$$

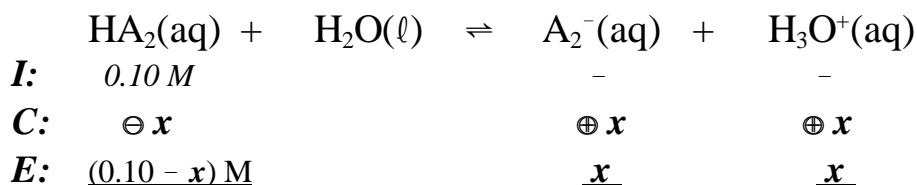
$$x = \sqrt{(1.5 \times 10^{-5})(0.10)} \approx 1.22 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

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

$$\therefore \text{pH} = -\log (1.22 \times 10^{-3}) = 2.914 = \underline{\underline{2.91}}$$

**PART A**  
**QUESTION 2**

(b) **0.10 M HA<sub>2</sub> (K<sub>a</sub> = 4.6 × 10<sup>-3</sup> M)**



(where  $x = [\text{H}_3\text{O}^+] \text{ produced} = [\text{A}_2^-] \text{ produced} = [\text{HA}_2] \text{ dissociated}$ )

$$K_c = \frac{[\text{A}_2^-][\text{H}_3\text{O}^+]}{[\text{HA}_2]} = K_a(\text{HA}_2)$$

$$\frac{(x)(x)}{0.10 - x} = 4.6 \times 10^{-3}$$

To solve for  $x$ , use

either ① “consecutive approximation” or ② “quadratic equation”.

- ①
- |                                |  |
|--------------------------------|--|
| 1 <sup>st</sup> approximation: | $2.145 \times 10^{-2} \text{ M}$                   |
| 2 <sup>nd</sup> approximation: | $1.901 \times 10^{-2} \text{ M}$                   |
| 3 <sup>rd</sup> approximation: | $1.930 \times 10^{-2} \text{ M}$                   |
| 4 <sup>th</sup> approximation: | $1.927 \times 10^{-2} \text{ M}$                   |
| 5 <sup>th</sup> approximation: | <b><math>1.927 \times 10^{-2} \text{ M}</math></b> |

$$\textcircled{2} \quad x^2 + (4.6 \times 10^{-3})x - (4.6 \times 10^{-4}) = 0 \quad (\text{quadratic equation})$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$\begin{aligned} a &= 1 \\ b &= 4.6 \times 10^{-3} \\ c &= -4.6 \times 10^{-4} \end{aligned}$$

$$x = \frac{-(4.6 \times 10^{-3}) \pm \sqrt{(4.6 \times 10^{-3})^2 - 4(1)(-4.6 \times 10^{-4})}}{2(1)}$$

$$x = 1.927 \times 10^{-2} \text{ M} \quad \text{OR} \quad -2.387 \times 10^{-2} \text{ M} \quad (\text{neglected})$$

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

$$\therefore \text{pH} = -\log (1.927 \times 10^{-2}) = 1.715 = \underline{1.72}$$

