USEFUL CONCEPTS AND FORMULAS

(1) $pH = -\log [H_3O^+]$ and $pOH = -\log [OH^-]$ and $pK_a = -\log K_a$

(2) There are *only* **six** strong acids: HCl, HBr, HI, H_2SO_4, HNO_3 and $HClO_4$.
Any other ***protonated*** acids are weak acids (**H**A) with K_a value less than 1.

(3) **Weak Acid (HA) and Weak Base (A^-)**

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

$$K_c = \frac{[A^-]\,[H_3O^+]}{[HA]} = K_a\,(HA)$$

! K_c is usually a small value because HA(aq) is a **weak acid**, which *only partially dissociates* in aqueous solution.

$$A^-(aq) + H_2O(\ell) \rightleftharpoons HA(aq) + OH^-(aq)$$

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

! K_c is usually a small value because $A^-(aq)$ is a **weak base**, which *only partially ionizes* in aqueous solution.

(4) $K_w\,(H_2O) = K_a\,(HA) \times K_b\,(A^-) = 1.0 \times 10^{-14}\ M^2$

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

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

! **Weak Acid** (HA): more positive (∵ one *more* H^+)
! **Weak Base** (A^-): more negative (∵ one *less* H^+)

(6) **Titration of Weak Acid (HA) with Strong Base (OH^-)**

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

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

! K_c is usually a LARGE value because $OH^-(aq)$ is a *strong base*, which drives the reaction to completion.

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

$$A^-(aq) + H_3O^+(aq) \rightarrow HA(aq) + H_2O(\ell)$$

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

! 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*** $\mathbf{pK_a \pm 1}$.

(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) ≅ n(OH^-) or n(A^-) ≅ n(H_3O^+)]

(10) Best **INDICATOR** chosen for the titration should have: $\mathbf{pK_{In}}$ ≅ **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-}

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

PART A
QUESTION 1

Solutions	Cation	Anion	Resultant Solution	NIE K_c expression
(i) **NaCN**	$Na^+(aq)$ (*neutral*)	$CN^-(aq)$ (*basic*)	**basic**	$CN^-(aq) + H_2O(\ell) \rightleftharpoons HCN(aq) + \mathbf{OH^-(aq)}$ $K_c = \frac{[HCN]\,[OH^-]}{[CN^-]} = K_b\,(CN^-) = \frac{K_w}{K_a\,(HCN)}$
(ii) **LiBr**	$Li^+(aq)$ (*neutral*)	$Br^-(aq)$ (*neutral*)	**neutral**	**none**
(iii) **KH_2PO_4**	$K^+(aq)$ (*neutral*)	$H_2PO_4^-(aq)$ (*acidic*)	**acidic**	$H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + \mathbf{H_3O^+(aq)}$ $K_c = \frac{[HPO_4^{2-}]\,[H_3O^+]}{[H_2PO_4^-]} = K_a\,(H_2PO_4^-)$
(iv) **$C_2H_5NH_2$** (organic **amine** compound)	$C_2H_5NH_2(aq)$ a weak base (*basic*)		**basic**	$C_2H_5NH_2(aq) + H_2O(\ell) \rightleftharpoons C_2H_5NH_3^+(aq) +$ **$OH^-(aq)$** $K_c = \frac{[C_2H_5NH_3^+]\,[OH^-]}{[C_2H_5NH_2]} = K_b\,(C_2H_5NH_2)$

(v) **$(CH_3)_3NHCl$** (oraganic ammonium salt)	$(CH_3)_3NH^+$ (*acidic*)	$Cl^-(aq)$ (*neutral*)	**acidic**	$(CH_3)_3NH^+(aq) + H_2O(\ell) \rightleftharpoons (CH_3)_3N(aq) + \mathbf{H_3O^+(aq)}$ $K_c = \dfrac{[(CH_3)_3N]\ [H_3O^+]}{[(CH_3)_3NH^+]}$ $K_c = K_a\ [(CH_3)_3NH^+] = \dfrac{K_w}{K_b\ [(CH_3)_3N]}$

PART A QUESTION 2

Given: 0.10 M solutions of two weak acids
HA_1: $K_a = 1.5 \times 10^{-5}$ M
HA_2: $K_a = 4.6 \times 10^{-3}$ M

To find: pH

(a) 0.10 M HA_1 ($K_a = 1.5 \times 10^{-5}$ M)

	$HA_1(aq)$ +	$H_2O(\ell)$	$\rightleftharpoons$	$A_1^-(aq)$ +	$H_3O^+(aq)$
I:	*0.10 M*			–	–
C:	$\ominus\ \boldsymbol{x}$			$\oplus\ \boldsymbol{x}$	$\oplus\ \boldsymbol{x}$
E:	$(0.10 - \boldsymbol{x})$ M			$\boldsymbol{x}$	$\boldsymbol{x}$

(where $\boldsymbol{x} = [H_3O^+]$ produced $= [A_1^-]$ produced $= [HA_1]$ ionized

$$K_c = \frac{[A_1^-]\ [H_3O^+]}{[HA_1]} = K_a\ (HA_1)$$

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

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

1st approximation:	1.225×10^{-3} M
2nd approximation:	1.217×10^{-3} M
3rd approximation:	1.217×10^{-3} M

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

$$x = \sqrt{(1.5 \times 10^{-5})(0.10)} \approx 1.22 \times 10^{-3} = [H_3O^+]$$

pH = − log [H⁺]

$\therefore$ pH = − log (1.22×10^{-3}) = 2.914 = **2.91**

PART A QUESTION 2

(b) 0.10 M HA_2 ($K_a = 4.6 \times 10^{-3}$ M)

	$HA_2(aq)$ +	$H_2O(\ell)$	⇌	$A_2^-(aq)$ +	$H_3O^+(aq)$
I:	*0.10 M*			–	–
C:	⊖ x			⊕ x	⊕ x
E:	$(0.10 - x)$ M			x	x

(where $x = [H_3O^+]$ produced $= [A_2^-]$ produced $= [HA_2]$ dissociated

$$K_c = \frac{[A_2^-]\,[H_3O^+]}{[HA_2]} = K_a\,(HA_2)$$

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

To solve for x, use

either ① "*consecutive approximation*" or ② "*quadratic equation*".

① 1st approximation: 2.145×10^{-2} M

2nd approximation:	1.901×10^{-2} M
3rd approximation:	1.930×10^{-2} M
4th approximation:	1.927×10^{-2} M
5th approximation:	**1.927×10^{-2} M**

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

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

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

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

$x = \underline{1.9}27 \times 10^{-2}$ M **OR** -2.387×10^{-2} M (*neglected*)

pH = − log [H⁺]

$\therefore$ pH = − log ($\underline{1.9}27 \times 10^{-2}$) = $1.\underline{71}5$ = **1.72**