Acid-Base #2
(K calculations - Summary)

Write response in the space provided. Express your answer in correct units & sig figs where appropriate.

---

1. (5 marks)

Calculate the initial concentration of a KF $\boxed{\text{salt}}$ solution that has a pH = 8.65 .
Begin by writing the equation for the predominant equilibrium reaction.

$KF \rightarrow \cancel{K^+} + F^-$

$F^- + H_2O \rightleftharpoons HF + OH^-$

$K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{1.00 \times 10^{-14}}{3.5 \times 10^{-4}}$

$K_b = 2.857... \times 10^{-11}$

	F^-	$\rightleftharpoons$	HF	+	OH^-
I	x		0		0
C	$-4.46... \times 10^{-6}$ M		$+4.46.. \times 10^{-6}$ M		$+4.46.. \times 10^{-6}$ M
E	$x - 4.46.. \times 10^{-6}$ M				$4.46... \times 10^{-6}$ M $\rightarrow 4.4668359 \times 10^{-6}$

pH = 8.65
pOH = 14 − 8.65
= 5.35
$[OH^-] = 10^{-5.35}$

$$K_b = \frac{(4.46... \times 10^{-6} M)(4.46.. \times 10^{-6} M)}{(x - 4.46... \times 10^{-6} M)}$$

$(4.46... \times 10^{-6} M)^2 = 2.857... \times 10^{-11}(x - 4.46... \times 10^{-6} M)$

$(4.46... \times 10^{-6} M)^2 + (4.46... \times 10^{-6})(2.857... \times 10^{-11}) = 2.857... \times 10^{-11} x$

$x = 0.6983...$

$x = [KF]_{initial} =$ ~~0.6~~ 0.70 M

5

2. A 0.0200 M solution of methylamine, CH_3NH_2, has a pH = 11.40.
Calculate the K_b for methylamine. (4 marks)

base

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$$

$$Kb = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = ?$$

	CH_3NH_2	$\rightleftharpoons CH_3NH_3^+$	$+ OH^-$
I	0.0200M		
C	$-2.51.... \times 10^{-3}M$	$+2.51... \times 10^{-3}M$	$+2.51... \times 10^{-3}M$
E	0.0175M	$2.51... \times 10^{-3}M$	$2.51... \times 10^{-3}M$

pH = 11.40
pOH = 14 − 11.40
= 2.60
$[OH^-] = 10^{-2.60}$

$$Kb = \frac{(2.51.. \times 10^{-3}M)^2}{0.0175M}$$

$Kb = 3.6 \times 10^{-4}$ (2 S.F. b/c pH = 11.40 is 2 S.F.)

4

3. Using calculations, show why the electrical conductivity of 1.0 M H_2CO_3 will be less than that for 0.10 M HCl. (4 marks)

= higher concentration of ions

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$

0.10M → 0.10M, 0.10M

$[ions] = 0.10M + 0.10M$
$= 0.20M$

$$H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^-$$

	H_2CO_3	H_3O^+	HCO_3^-
I	0.10M	0	0
C	$-x$	$+x$	$+x$
E	$0.10-x$	? x	? x

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} = 4.3\times10^{-7}$$

$$= \frac{x^2}{0.10-x} = 4.3\times10^{-7}$$

$$x^2 + 4.3\times10^{-7}x - 4.3\times10^{-8} = 0$$

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

$$x = 2.07\ldots\times10^{-4}$$

$[H_3O^+] = x = 2.07\ldots\times10^{-4}M$
$[HCO_3^-] = x = 2.07\ldots\times10^{-4}M$ + $[ions] = 4.14\ldots\times10^{-4}M$

[ions] in HCl = 0.20M
[ions] in H_2CO_3 = $4.1\times10^{-4}M$

* see back page in case you're wondering.....

4

4. **(5 marks)**

Calculate the pH of a 0.35 M solution of the salt ammonium bromide.
Begin by writing the equation for the predominant equilibrium.

$$\underset{0.35M}{NH_4Br} \rightarrow \underset{0.35M}{NH_4^+} + \cancel{Br^-}$$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$Ka = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.6 \times 10^{-10}$$

	NH_4^+	$\rightleftharpoons$	NH_3	+	H_3O^+
I	0.35M		0		0
C	$-x$		$+x$		$+x$
E	$0.35-x$		x		x

$$\frac{(x)(x)}{(0.35-x)} = 5.6 \times 10^{-10}$$

Assume:

$0.35 - x \simeq 0.35$

$$\frac{x^2}{0.35} = 5.6 \times 10^{-10}$$

$$x = \sqrt{(5.6 \times 10^{-10})(0.35M)}$$
$$= 1.4 \times 10^{-5}M$$

justification:

$$\% \text{ dissociation} = \frac{[H_3O^+]}{[NH_4^+]} \times 100$$

$$= \frac{1.4 \times 10^{-5}M}{0.35M} \times 100$$

$= 0.004\%$

approximation is valid.

$x = [H_3O^+]$

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

$= -\log(1.4 \times 10^{-5}M)$

$pH = 4.85$

$\overline{\ \ \ }$
5

9. A 0.20 M solution of a weak acid, HA, has a pH = 1.32.
Use calculations and the table of "Relative Strengths of Brønsted-Lowry Acids and Bases" from the *Data Booklet* to determine the identity of the acid. **(5 marks)**

you can identify an acid with its $\underline{\underline{K_a}}$

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

	HA	$\rightleftharpoons$ H^+	$+$ A^-
I	0.20M	0	0
C	-0.0478...	+0.0478...	+0.0478...
E	0.15M	0.0478... M	0.0478... M

$\uparrow$
pH = 1.32
$[H^+] = 10^{-1.32}$

$$K_a = \frac{(0.0478...M)^2}{0.15M}$$

$K_a = 1.5 \times 10^{-3}$

$\therefore$ the acid must be H_2SO_3

5

Acid-Base #2
(K calculations - Summary)

Write response in the space provided. Express your answer in correct units & sig figs where appropriate.

---

1. **(5 marks)**

Calculate the initial concentration of a KF $\boxed{\text{salt}}$ solution that has a pH = 8.65 . Begin by writing the equation for the predominant equilibrium reaction.

$$KF \rightarrow \cancel{K^+} + F^-$$

$$F^- + H_2O \rightleftharpoons HF + OH^- \qquad K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{1.00 \times 10^{-14}}{3.5 \times 10^{-4}}$$

$$K_b = 2.857... \times 10^{-11}$$

	F^-	$\rightleftharpoons$	HF	+	OH^-
I	x		0		0
C	$-4.46... \times 10^{-6}$ M		$+4.46.. \times 10^{-6}$ M		$+4.46.. \times 10^{-6}$ M
E	$x - 4.46.. \times 10^{-6}$ M				$4.46... \times 10^{-6}$ M $\rightarrow 4.4668359 \times 10^{-6}$

pH = 8.65
pOH = 14 − 8.65
= 5.35
$[OH^-] = 10^{-5.35}$

$$K_b = \frac{(4.46... \times 10^{-6} M)(4.46.. \times 10^{-6} M)}{(x - 4.46... \times 10^{-6} M)}$$

$$(4.46.. \times 10^{-6} M)^2 = 2.857... \times 10^{-11}(x - 4.46... \times 10^{-6} M)$$

$$(4.46... \times 10^{-6} M)^2 + (4.46... \times 10^{-6})(2.857... \times 10^{-11}) = 2.857... \times 10^{-11} x$$

$x = 0.6983...$

$x = [KF]_{initial} =$ ~~0.46~~ 0.70M

$\overline{\quad}$ 5