Answers

Acid/Base Equilibrium Practice

1. Calculate the percent dissociation of the acid in each of the following solutions:
   a. 0.50 M acetic acid (Ka = 1.8 x 10^{-5})

$$HC_2H_3O_2 \rightleftarrows H^+ + C_2H_3O_2^-$$

	$HC_2H_3O_2$	H^+	$C_2H_3O_2^-$
	0.50M	~0	0
	$-x$	$+x$	$+x$
	0.50M$-x$	x	x

$$K_a = 1.8\times10^{-5} = \frac{x^2}{0.50 - \cancel{x}}$$

$x = 0.0030$ Assump. is good.

$$\%\text{ diss} = \frac{[H^+]}{[\text{acid}]} = \frac{0.0030\,M}{0.50\,M}\times100 = \boxed{0.60\%}$$

   b. 0.050 M acetic acid

$$K_a = \frac{x^2}{0.050 - \cancel{x}}$$

$x = 9.5\times10^{-4}$

Assump. is good

$$\%\text{ diss.} = \frac{9.5\times10^{-4}}{0.050}\times100 = 1.9\%$$

   c. Use Le Chatelier's principle to explain why percent dissociation increases as the concentration of a ~~weak acid~~ [H+] decreases.

↓[H+], shift right, ↑ dissociation

2. The pH of a 0.010 M solution of cyanic acid (HOCN) is 2.77 at 25°C. Calculate the Ka for HOCN from this result.

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

$10^{-2.77} = 1.7\times10^{-3}M = [H^+] = x$

$$HOCN \rightleftarrows H^+ + OCN^-$$

	HOCN	H^+	OCN^-
	0.010	~0	0
	$-x$	$+x$	$+x$
	0.010$-x$	x	x

$$K_a = \frac{(1.7\times10^{-3}M)(1.7\times10^{-3}M)}{(0.010-(1.7\times10^{-3}))}$$

$$\boxed{K_a = 3.5\times10^{-4}}$$

3. Calculate the percent ionization of 0.10 M NH_3 (K_b = 1.8 x 10^{-5}).

$$NH_3 + \cancel{H_2O} \rightleftarrows NH_4^+ + OH^-$$

	NH_3	NH_4^+	OH^-
	0.10M	0	~0
	$-x$	$+x$	$+x$
	0.10M$-x$	x	x

$$K_b = 1.8\times10^{-5} = \frac{x^2}{0.10M - \cancel{x}}$$

$x = 1.3\times10^{-3}$

Assump. is good.

$$\%\text{ diss} = \frac{1.3\times10^{-3}M}{0.10M}\times100 = \boxed{1.3\%}$$

4. A solution of formic acid (HCOOH, K_a = 1.8 x 10^{-4}) has a pH of 2.70. Calculate the initial concentration of formic acid in this solution.

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

$10^{-2.70} = 2.0\times10^{-3}M$

$$HCOOH \rightleftarrows H^+ + COOH^-$$

	HCOOH	H^+	$COOH^-$
	C	~0	0
	$-x$	$+x$	$+x$
	C$-x$	x	x

$x = [H^+] = 2.0\times10^{-3}M$

$$K_a = 1.8\times10^{-4} = \frac{x^2}{C-x} = \frac{(2.0\times10^{-3}M)^2}{C-(2.0\times10^{-3}M)}$$

$$\boxed{C = 2.4\times10^{-2}M}$$

$X = \text{acetate}$ $\quad HX \rightleftarrows X^- + H^+$

5. A solution with a total volume of 250.0 mL is prepared by diluting 20.0 mL of glacial acetic acid with water. Calculate the [H+] and the pH of this solution. Assume that glacial acetic acid is pure ~~liquid~~ acetic acid with a density of 1.05 g/cm^3. ($K_a = 1.8 \times 10^{-5}$).

$$\frac{1.05\,g}{mL}\cdot 20.0\,mL \cdot \frac{1\,mol}{60\,g} = 0.35\,mol$$

$$\frac{0.35\,mol}{0.25\,L} = 1.4\,M$$

CH_3COOH	$\rightleftarrows$	H^+	+	CH_3COO^-
1.4 M		~0		0
$-x$		$+x$		$+x$
$1.4-x$		x		x

$$1.8 \times 10^{-5} = \frac{x^2}{1.4-\cancel{x}}$$

$x = 5.0 \times 10^{-3}\,M = [H+]$

Assump. is good.

$-\log[H+] = pH$

$\boxed{pH = 2.30}$

6. The pH of a 1.00×10^{-3} M solution of a base, pyrrolidine, is 10.82. Calculate the K_b.

P	+ $\cancel{H_2O}$	$\rightleftarrows$	HP	+ OH^-
1.00×10^{-3} M			0	~0
$-x$			$+x$	$+x$
$(1.00 \times 10^{-3}) - x$			x	x

$pH = 10.82$

$pOH = 14 - 10.82 = 3.18$

$10^{-pOH} = [OH^-] = 6.6 \times 10^{-4} = x$

$$K_b = \frac{(6.6 \times 10^{-4})^2}{(1.00 \times 10^{-3}) - (6.6 \times 10^{-4})}$$

$\boxed{K_b = 1.3 \times 10^{-3}}$

7. Calculate the pH of a 0.0050 M solution of H_2SO_4 ($K_{a2} = 1.2 \times 10^{-2}$).

H^+, HSO_4^-, H_2O

H^+: 0.0050 M; HSO_4^-: 0.0050 M

HSO_4^-	$\rightleftarrows$	H^+	+	SO_4^{2-}
0.0050		0.0050		0
$-x$		$+x$		$+x$
$0.0050 - x$		$0.0050 + x$		x

$$K_{a_2} = \frac{(0.0050 + \cancel{x})(x)}{(0.0050 - \cancel{x})} = 1.2 \times 10^{-2}$$

Assump. Not good.

$(6 \times 10^{-5}) - (1.2 \times 10^{-2})x = 0.0050x + x^2$

$x^2 + 0.017x - (6 \times 10^{-5}) = 0$

$$\frac{(-0.017) \pm \sqrt{(0.017)^2 - 4(1)(-6 \times 10^{-5})}}{2(1)}$$

$x = 3.0 \times 10^{-3}\,M$

$[H^+] = (0.0050) + (3.0 \times 10^{-3}\,M)$

$[H^+] = 8.0 \times 10^{-3}\,M$

$\boxed{pH = 2.10}$

8. Arrange the following 0.10 M solutions in order of most acidic to most basic:

$CaBr_2$, K_2CO_3, $HClO_4$, FeI_3 $\Rightarrow Fe^{+3} + H_2O \rightleftarrows Fe(OH)_3 + H^+$

$CaBr_2$: SB, SA — Neutral

K_2CO_3: SB, WA — Basic

$CO_3^{2-} + H_2O \rightleftarrows HCO_3^- + OH^-$

$HClO_4$: Strong acid

FeI_3: WB, SA — Acidic

$HClO_4 > FeI_3 > CaBr_2 > K_2CO_3$

most acidic

9. Given that the K_a value for acetic acid is 1.8×10^{-5} and the K_a value for hypochlorous acid is 3.5×10^{-8}, which is the stronger base, OCl^- or $C_2H_3O_2^-$?

Calculate K_b for each compound & compare. $\uparrow K_b = \uparrow$ base

OR

Bigger Ka for acetic acid means stronger acid. Stronger acid means weaker conj. base

Answers: 1.a. 0.60%

b. 1.9%

c. decrease [H$^+$] = shift right = more dissociation

2. $K_a = 3.5 \times 10^{-4}$

3. 1.3%

4. [HCOOH] = 2.4×10^{-2} M

5. pH = 2.30

6. $K_b = 1.3 \times 10^{-3}$

7. pH = 2.20

8. $HClO_4 > FeI_3 > CaBr_2 > K_2CO_3$

9. OCl^-