Acids and Bases
8.3 Weak Acid and Base Equilibria
Worksheet Key

1) The following questions pertain to a 2.2 M solution of hydrocyanic acid, HCN, at 25°C. pK_a = 9.21 at 25°C.
   a. Find the concentrations of all species present in the solution at equilibrium.

$$K_a = 10^{-pK_a}$$
$$K_a = 10^{-9.21}$$
$$K_a = 6.2\times10^{-10}$$

	$HCN_{(aq)} \rightarrow H^+_{(aq)}$		$+ CN^-_{(aq)}$
I	2.2	0	0
C	$-x$	$+x$	$+x$
E	$2.2 - x$	x	x

$$K_a = \frac{[H^+][CN^-]}{[HCN]} = \frac{(x)(x)}{2.2 - x}$$
$$6.2\times10^{-10} = \frac{x^2}{2.2}\ \textit{assume}\ 2.2 - x = 2.2$$
$$x^2 = (2.2)(6.2\times10^{-10})$$
$$x = 3.7\times10^{-5}\,M$$
$$[HCN] = 2.2\,M$$
$$[H_3O^+] = [CN^-] = 3.7\times10^{-5}\,M$$

   b. Find the pH of the solution.

$$pH = -\log[H_3O^+] = -\log(3.7\times10^{-5}\,M) = 4.43$$

   c. Identify the strongest base in this system.

CN^- is the strongest base in this system. H_2O and CN^- compete for protons and CN^- wins most of the time. We know this because the equilibrium lies far to the left.

2) The following questions pertain to a 0.50 M solution of HOCl at 25°C. pK_a = 7.46 at 25°C.
   a. Find the concentrations of all species present in the solution at equilibrium.

$$K_a = 10^{-pK_a}$$
$$K_a = 10^{-7.46}$$
$$K_a = 3.5\times10^{-8}$$

	$HOCl_{(aq)} \rightarrow H^+_{(aq)}$		$+ OCl^-_{(aq)}$
I	0.50	0	0
C	$-x$	$+x$	$+x$
E	$0.50 - x$	x	x

$$K_a = \frac{[H^+][OCl^-]}{[HOCl]} = \frac{(x)(x)}{0.50 - x}$$
$$3.5\times10^{-8} = \frac{x^2}{0.50}\ \textit{assume}\ 0.50 - x = 0.50$$
$$x^2 = (0.50)(3.5\times10^{-8})$$
$$x = 1.3\times10^{-4}\,M$$
$$[HOCl] = 0.5M$$
$$[H_3O^+] = [OCl^-] = 1.3\times10^{-4}\,M$$

b. Find the pH of the solution.

$$pH = -\log[H^+] = -\log(1.3 \times 10^{-4} M) = 3.89$$

c. Identify the strongest base in this system.

OCl^- is the strongest base in this system. H_2O and OCl^- compete for protons and OCl^- wins most of the time. We know this because the equilibrium lies to the left.

3) The pH of distilled water at 25°C is 7.0. When the temperature is increased to 37°C the pH drops to 6.8. At both temperatures the water is considered to be neutral as $[H_3O^+] = [OH^-]$. Explain why the pH drops when the temperature increases.

K_w is temperature dependent and the autodissociation of water is an endothermic process.

$$H_2O + \text{heat} \rightleftharpoons H^+ + OH^-$$

As heat is added, the equilibrium shifts to the right to use up that heat. This causes the $[H_3O^+]$ and $[OH^-]$ to increase at the same rate. The pH is lower because $[H_3O^+]$ is higher.

4) The following questions pertain to a 0.58 *M* solution of benzoic acid, $HC_7H_5O_2$, at 25°C. $K_a = 6.4 \times 10^{-5}$.

a. Find the concentration of $H_3O^+(aq)$ in the solution.

	$HC_7H_5O_{2(aq)}$	$\rightarrow H^+_{(aq)}$	$+ C_7H_5O_2^-{}_{(aq)}$
I	0.58	0	0
C	$-x$	$+x$	$+x$
E	$0.58 - x$	x	x

$$K_a = \frac{[H^+][C_7H_5O_2^-]}{[HC_7H_5O_2]} = \frac{(x)(x)}{0.58 - x}$$

$$\textit{assume } 0.58 - x = 0.58$$

$$6.4 \times 10^{-5} = \frac{x^2}{0.58}$$

$$x^2 = (0.58)(6.4 \times 10^{-5})$$

$$x = \sqrt{(0.58)(6.4 \times 10^{-5})}$$

$$x = [H_3O^+] = 6.1 \times 10^{-3} M$$

b. Find the pH of the solution.

$$pH = -\log[H^+] = -\log(6.1 \times 10^{-3} M) = 2.21$$

5) Find the concentration of $H_3O^+(aq)$ in a 1.75 *M* solution of lactic acid, $HC_3H_5O_3$, at 25°C. $K_a = 1.38 \times 10^{-4}$.

$$HC_3H_5O_{3(aq)} \rightarrow H^+_{(aq)} + C_3H_5O_3^-{}_{(aq)}$$

	$HC_3H_5O_3$	H^+	$C_3H_5O_3^-$
I	1.75	0	0
C	$-x$	$+x$	$+x$
E	$1.75 - x$	x	x

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]} = \frac{(x)(x)}{1.75 - x}$$

$$\textit{assume } 1.75 - x = 1.75M$$

$$1.38\times10^{-4} = \frac{x^2}{1.75}$$

$$x^2 = (1.75)(1.38\times10^{-4})$$

$$x = \sqrt{(1.75)(1.38\times10^{-4})}$$

$$x = [H_3O^+] = 1.55\times10^{-2}M$$

6) Write the equilibrium expression for the ionization of HOI, and calculate the concentration of $HOI(aq)$ in solution if $[H_3O^+] = 2.3 \times 10^{-5}$ *M* and $pK_a = 10.7$ at 25°C.

$$HOI_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OI^-_{(aq)}$$

$$K_a = 10^{-pK_a}$$

$$K_a = 10^{-10.7}$$

$$K_a = 2\times10^{-11}$$

$$HOI_{(aq)} \rightleftharpoons H^+_{(aq)} + OI^-_{(aq)}$$

	HOI	H^+	OI^-
I	x	0	0
C	x	$+2.3\times10^{-5}$	$+2.3\times10^{-5}$
E	$x - 2.3\times10^{-5}$	2.3×10^{-5}	2.3×10^{-5}

$$K_a = \frac{[H^+][OI^-]}{[HOI]} = \frac{(2.3\times10^{-5})(2.3\times10^{-5})}{x - 2.3\times10^{-5}}$$

$$\textit{assume } x - 2.3\times10^{-5} = x$$

$$2\times10^{-11} = \frac{(2.3\times10^{-5})^2}{x}$$

$$x = \frac{(2.3\times10^{-5})^2}{2\times10^{-11}}$$

$$x = 30$$

$$x = [HOI] = 30M$$

7) A 0.45 *M* solution of propanoic acid, $HC_3H_5O_2$, experiences 1.58% ionization.
   a. Find the concentrations of all aqueous species in the solution at equilibrium.

	$HC_3H_5O_{2(aq)} \rightarrow$	$H^+_{(aq)}$ +	$C_3H_5O_2^-{}_{(aq)}$
I	0.45	0	0
C	- 0.45(0.0158)	+0.45(0.0158)	+0.45(0.0158)
E	0.45 - 0.45(0.0158)	0.45(0.0158)	0.45(0.0158)

$0.45(0.0158) = 7.1\times10^{-3}\,M$

$[HC_3H_5O_2] = 0.45M - 0.0071M = 0.44M$

$[H_3O^+] = [C_3H_5O_2^-] = 7.1\times10^{-3}\,M$

   b. Find the pH of the solution.

$$pH = -\log[H^+] = -\log(7.1\times10^{-3}\,M) = 2.15$$

   c. What concentration of HCl would produce a solution with the same pH as a 0.45 *M* solution of propanoic acid, $HC_3H_5O_2$? Justify your answer.

A $7.1\times10^{-3}\,M$ solution of HCl would have the same pH as a 0.45 *M* solution of propanoic acid. We know this because HCl experiences 100% ionization.

$$pH = -\log[H^+] = -\log(7.1\times10^{-3}\,M) = 2.15$$

8) In a 0.57 *M* solution of propanoic acid, HOC_6H_5, 0.0684% of the acid has ionized.
   a. Find the concentrations of all aqueous species in the solution at equilibrium.

	$HOC_6H_{5(aq)} \rightarrow$	$H^+_{(aq)}$ +	$OC_6H_5^-{}_{(aq)}$
I	0.57	0	0
C	- 0.57(0.000684)	+0.57(0.000684)	+0.57(0.000684)
E	0.57 - 0.57(0.000684)	0.57(0.000684)	0.57(0.000684)

$0.57(0.000684) = 3.9\times10^{-4}\,M$

$[HOC_6H_5] = 0.57M$ and $[H_3O^+] = [OC_6H_5^-] = 3.9\times10^{-4}\,M$

b. Find the pH of the solution.

$$pH = -\log[H_3O^+] = -\log(3.9\times10^{-4}M) = 3.41$$

c. What concentration of HBr would produce a solution with the same pH as a 0.57 *M* solution of propanoic acid, HOC_6H_5? Justify your answer.

A $3.9\times10^{-4}M$ solution of HBr would have the same pH as a 0.57 *M* solution of propanoic acid. We know this because HBr experiences 100% ionization.

$$pH = -\log[H_3O^+] = -\log(3.9\times10^{-4}M) = 3.41$$

9) Calculate the pH of 1.5 *M* solution of hydroxylamine, NH_2OH, at 25°C.
$K_b = 9.1 \times 10^{-9}$

	$NH_2OH_{(aq)}$ →	$NH_2^+{}_{(aq)}$	+ $OH^-{}_{(aq)}$
I	1.5	0	0
C	$-x$	$+x$	$+x$
E	$1.5-x$	x	x

$$K_b = \frac{[NH_2^+][OH^-]}{[NH_2OH]} = \frac{(x)(x)}{1.5-x}$$

assume $1.5-x=1.5$

$$9.1\times10^{-9} = \frac{(x)^2}{1.5}$$

$$x = \sqrt{1.5\times9.1\times10^{-9}} = 1.2\times10^{-4} = [OH] = 1.2\times10^{-4}M$$

$$pOH = -\log[OH^-] = -\log(1.2\times10^{-4}M) = 3.92$$

$$pH = 14 - pOH = 14 - 3.92 = 10.08$$

10) Calculate the pH of 0.25 *M* solution of aniline, $C_6H_5NH_2$, at 25°C.

$C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$ $K_b = 4.3 \times 10^{-10}$

$$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + HO^-$$

I	0.25	0	0
C	$-x$	$+x$	$+x$
E	$0.25-x$	x	x

$$K_b = \frac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]} = \frac{(x)(x)}{0.25-x}$$

assume $0.25 - x = 0.25$

$$4.3\times10^{-10} = \frac{(x)^2}{0.25}$$

$$x = \sqrt{0.25\times4.3\times10^{-10}} = 1.2\times10^{-4}M = [OH^-]$$

$$\text{pOH} = \text{-log}[OH^-] = \text{-log}(1.0\times10^{-5}M) = 5.00$$

$$\text{pH} = 14 - \text{pOH} = 14 - 5.00 = 9.00$$

11) Calculate the pH of 1.25 *M* solution of hydrazine, N_2H_4, at 25°C.

$N_2H_4(aq) + H_2O(l) \rightleftharpoons N_2H_5^+(aq) + OH^-(aq)$ $K_b = 8.9 \times 10^{-7}$

$$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + HO^-$$

I	1.25	0	0
C	$-x$	$+x$	$+x$
E	$1.25-x$	x	x

$$K_b = \frac{[N_2H_5^+][OH^-]}{[N_2H_4]} = \frac{(x)(x)}{1.25-x}$$

assume $1.25 - x = 1.25$

$$8.9\times10^{-7} = \frac{(x)^2}{1.25}$$

$$x = \sqrt{1.25\times8.9\times10^{-7}} = 1.05\times10^{-3}M = [OH]$$

$$\text{pOH} = \text{-log}[OH^-] = \text{-log}(1.05\times10^{-3}M) = 2.979$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2.979 = 11.021$$

12) Calculate the pH of 1.75 M solution of ammonia, NH_3, at 25°C. $K_b = 1.8 \times 10^{-5}$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + HO^-$$

I	1.75	0	0
C	$-x$	$+x$	$+x$
E	$1.75-x$	x	x

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{1.75-x}$$

$assume\ 1.75-x = 1.75$

$$1.8\times10^{-5} = \frac{(x)^2}{1.75}$$

$$x = \sqrt{1.75\times1.8\times10^{-5}} = 5.61\times10^{-3}M = [OH]$$

$$pOH = \text{-log}[OH^-] = \text{-log}(5.61\times10^{-3}M) = 2.251$$

$$pH = 14 - pOH = 14 - 2.251 = 11.749$$

13) Calculate the pH of a 1.1 M solution of caffeine, $C_8H_{10}N_4O_2$.

$C_8H_{10}N_4O_2(aq) + H_2O(l) \rightleftharpoons C_8H_{11}N_4O_2^+(aq) + OH^-(aq)$ $K_b = 4.1 \times 10^{-4}$

$$C_8H_{10}N_4O_2 + H_2O \rightleftharpoons C_8H_{11}N_4O_2^+ + HO^-$$

I	1.1	0	0
C	$-x$	$+x$	$+x$
E	$1.1-x$	x	x

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{1.1-x}$$

$assume\ 1.1-x = 1.1$

$$4.1\times10^{-4} = \frac{(x)^2}{1.1}$$

$$x = \sqrt{1.1\times4.1\times10^{-4}} = 0.021M = [OH]$$

$$pOH = \text{-log}[OH^-] = \text{-log}(0.021M) = 1.68$$

$$pH = 14 - pOH = 14 - 1.68 = 12.32$$

14) K_a for acetic acid is 1.8×10^{-5}, and K_a for hypochlorous acid is 3.5×10^{-8} at 25°C. If 500.0 mL of 1.0 M acetic acid was mixed with 500.0 mL 1.0 M hypochlorous acid, which conjugate base would have the highest concentration? Justify your answer.

Acetic acid is stronger, as it has a larger K_a value. Acid strength increases as K_a increases. The larger the K_a value, the further to the right the equilibrium position (more products). For this reason, $[CH_3COO^-] > [ClO^-]$.

15) The pH is 4.2 in a 0.50 *M* solution of an unknown acid. Find K_a for the acid.

$$[H_3O^+] = 10^{-4.2} = 6.31 \times 10^{-5} M$$
$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(6.31 \times 10^{-5})^2}{0.50} = 7.96 \times 10^{-9}$$

16) The pH of a 0.25 *M* C_5H_5N solution at 25°C is 9.25.

$$C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$$

a. Is the solution acidic or basic? Explain.

The solution is basic, as the pH is greater than 7.

b. Find the concentration of all the aqueous species in the solution.

pOH = 14 - pH = 14 - 9.25 = 4.75

$$[OH^-] = [C_5H_5NH^+] = 10^{-pOH} = 10^{-4.75} = 1.8 \times 10^{-5} M$$
$$[C_5H_5N] = 0.25 M$$

17) In a 0.450 *M* $HONH_2$ solution, $[OH^-] = 5.28 \times 10^{-6}$ *M*.

$$HONH_2(aq) + H_2O(l) \rightleftharpoons HONH_3^+(aq) + OH^-(aq)$$

a. Find $[HONH_3^+]$.

$$[OH^-] = [HONH_3^+] = 5.28 \times 10^{-6} M$$

b. Find K_b.

$$K_b = \frac{[OH^-][HONH_3^+]}{[HONH_2]} = \frac{(5.28 \times 10^{-6})(5.28 \times 10^{-6})}{0.450} = 6.20 \times 10^{-11}$$

c. Find pOH.

$$pOH = -\log[OH^-] = -\log(5.28 \times 10^{-6} M) = 5.277$$

d. Find pH.

$$pH = 14 - pOH = 14 - 5.277 = 8.723$$

e. Find the percent ionization of $HONH_2$ in a 0.450 *M* $HONH_2$ solution.

$$\%ionization = \frac{[HONH_3^+]_{equilibrium}}{[HONH_2]_{initial}} \times 100 = \frac{5.28 \times 10^{-6} M}{0.450M} \times 100 = 1.17 \times 10^{-3}\%$$

f. What concentration of NaOH would be required to make a solution with the same pH that was calculated in part d.?

$$5.28 \times 10^{-6} M$$

g. Find the percent ionization of NaOH in the above solution.

100% as NaOH is a strong base.

18) In a 0.032 *M* NH_3 solution, $[OH^-] = 1.27 \times 10^{-3}$ *M*.

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

a. Find $[NH_4^+]$.

$[OH^-] = [NH_4^+] = 1.27 \times 10^{-3}$ *M*

b. Find K_b.

$$K_b = \frac{[OH^-][NH_4^+]}{[NH_3]} = \frac{(1.27 \times 10^{-3})^2}{0.032 - 0.00127} = 5.25 \times 10^{-5}$$

The subtraction (0.032 – 0.00127) was made in the denominator, because it affects the value of the denominator when significant digits are taken into consideration.

c. Find pH.

$$pOH = -\log[OH^-] = -\log(1.27 \times 10^{-3}\ M) = 2.896$$
$$pH = 14 - pOH = 14 - 2.896 = 11.104$$

d. Find the percent ionization of NH_3 in a 0.032 *M* NH_3 solution.

$$\%ionization = \frac{[NH_4^+]_{equilibrium}}{[NH_3]_{initial}} \times 100 = \frac{1.27 \times 10^{-3} M}{0.032} \times 100 = 3.9\%$$

e. What concentration of KOH would be required to make a solution with the same pH that was calculated in part c.?

$$1.27 \times 10^{-3}\ M$$

f. Find the percent ionization of KOH in the above solution.

100% as KOH is a strong base.

19) Citric acid $H_3C_6H_5O_7$ is a polyprotic acid. $K_{a1} = 8.4 \times 10^{-4}$, and $K_{a2} = 1.8 \times 10^{-5}$ at 25°C.

a. Which of the following species has the lowest concentration in a 1.0 *M* $H_3C_6H_5O_7$ solution: $H_3C_6H_5O_7(aq)$, $H_2C_6H_5O_7^-(aq)$, or $\mathbf{HC_6H_5O_7^{2-}(aq)}$? Justify your answer.

$[HC_6H_5O_7^{2-}]$ is the lowest, because K_2 is smaller than K_1.

b. Which of the following possesses the highest concentration in a 1.0 *M* $H_3C_6H_5O_7$ solution: $\mathbf{H_3C_6H_5O_7(aq)}$, $H_2C_6H_5O_7^-(aq)$, or $HC_6H_5O_7^{2-}(aq)$? Justify your answer.

$[H_3C_6H_5O_7]$ is the largest, because it is a weak acid. The equilibrium lies far to the left (mostly reactants).

c. What is the equilibrium constant for the reaction below?
$H_3C_6H_5O_7(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2C_6H_5O_7^-(aq)$

$K_{a1} = 8.4 \times 10^{-4}$

d. What is the equilibrium constant for the reaction below?
$H_3C_6H_5O_7(aq) + 2\ H_2O(l) \rightleftharpoons 2\ H_3O^+(aq) + HC_6H_5O_7^{2-}(aq)$

$$K^I = K_1 \times K_2$$

$$\frac{[H_3O^+]^2[HC_6H_5O_7^{2-}]}{[H_3C_6H_5O_7]} = \frac{[H_3O^+][H_2C_6H_5O_7^-]}{[H_3C_6H_5O_7]} \times \frac{[H_3O^+][HC_6H_5O_7^{2-}]}{[H_2C_6H_5O_7^-]}$$

$$K^I = (8.4 \times 10^{-4})(1.8 \times 10^{-5})$$

$$K^I = 1.5 \times 10^{-8}$$

20) $pK_b = 1.00$ for $CH_3NH_2(aq)$ at 25°C. Find pK_a for $CH_3NH_3^+(aq)$ at this temperature.

$$pK_w = pK_a + pK_b$$

$$14 - 1.00 = pK_a$$

$$pK_a = 13$$

21) You are given 10 mL of a hydrochloric acid, HCl, solution with a pH of 1.0. You are required to change the pH to 2.0 by adding water. How much water do you add?

when pH = 1	*when* pH = 2
$pH = -\log[H^+]$	$pH = -\log[H^+]$
$[H^+] = 10^{-pH} M$	$[H^+] = 10^{-pH} M$
$[H^+] = 1\times10^{-1} M$	$[H^+] = 1\times10^{-2} M$
$[H^+] = \frac{1\text{ mol}}{10\text{ L}}$	$[H^+] = \frac{1\text{ mol}}{100\text{ L}}$

You will add 90 mL of water.
HCl is a strong acid, thus the number of moles of H^+ will not change. To change the pH we must change the concentration of H^+. Reducing the concentration of H^+ by a factor of ten will cause the pH to increase by one. If the volume is increased by a factor of ten, the concentration is reduced by a factor of ten. Thus, adding 90 mL of water will raise the pH from 1.0 to 2.0.

22) You are given 100 mL of a solution of potassium hydroxide with a pH of 12.0. You are required to change the pH to 11.0 by adding water. How much water do you add?

when pH = 12	*when* pH = 11
pOH = 14 - pH	pOH = 14 - pH
pOH = 14 - 12	pOH = 14 - 11
pOH = 2	pOH = 3
$pOH = -\log[OH^-]$	$pOH = -\log[OH^-]$
$[OH^-] = 10^{-pOH}$	$[OH^-] = 10^{-pOH}$
$[OH^-] = 1\times10^{-2} M$	$[OH^-] = 1\times10^{-3} M$
$[OH^-] = \frac{1\text{ mol}}{100\text{ L}}$	$[OH^-] = \frac{1\text{ mol}}{1000\text{ L}}$

You will add 900 mL of water.
KOH is a strong base, so the number of moles of OH^- will not change. To change the pH we must change the concentration of OH^-. Reducing the concentration of OH^- by a factor of ten will cause the pOH to increase by one, and the pH to drop by one. If the volume is increased by a factor of ten, the concentration is reduced by a factor of ten. Thus, adding 900 mL of water will reduce the pH from 12.0 to 11.0.

23) If the pH of a HBr solution is the same as the pH of a CH_3COOH solution, is [HBr] less than, equal to, or greater than [CH_3COOH]? Justify your answer.

[HBr] < [CH_3COOH]. HBr is a strong acid and experiences 100% ionization. CH_3COOH is a weak acid and experiences much less than 100% ionization. For this reason, fewer moles of HBr are required to produce the same molar concentration of H_3O^+ in the solution.

24) At 25°C, the following weak acid solution has a pH of 3.6 and a K_a of 1.9×10^{-5}. Which species has the highest concentration – A^- or HA? (Hint: This problem can be solved with a manipulation of the equilibrium expression.)

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

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

$$\frac{[A^-]}{[HA]} = \frac{K_a}{[H^+]} = \frac{1.9 \times 10^{-5}}{10^{-3.6}} = 0.08$$

Since 0.08 < 1, [HA] > [A^-]

A question like this appeared on the 2015 Exam.