

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.

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

$$K_a = 10^{-9.21}$$

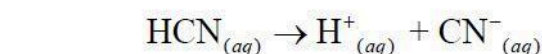
$$K_a = 6.2 \times 10^{-10}$$

$$K_a = \frac{[H^+][CN^-]}{[HCN]} = \frac{(x)(x)}{2.2 - x}$$

$$6.2 \times 10^{-10} = \frac{x^2}{2.2} \text{ assume } 2.2 - x = 2.2$$

$$x^2 = (2.2)(6.2 \times 10^{-10})$$

$$x = 3.7 \times 10^{-5} M$$



I	2.2	0	0
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C	-x	+x	+x
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E	2.2 - x	x	x
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$$[HCN] = 2.2 M$$

$$[H_3O^+] = [CN^-] = 3.7 \times 10^{-5} M$$

- b. Find the pH of the solution.

$$pH = -\log[H_3O^+] = -\log(3.7 \times 10^{-5} M) = 4.43$$

- 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.

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

$$K_a = 10^{-7.46}$$

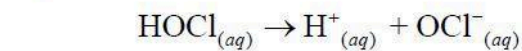
$$K_a = 3.5 \times 10^{-8}$$

$$K_a = \frac{[H^+][OCl^-]}{[HOCl]} = \frac{(x)(x)}{0.50 - x}$$

$$3.5 \times 10^{-8} = \frac{x^2}{0.50} \text{ assume } 0.50 - x = 0.50$$

$$x^2 = (0.50)(3.5 \times 10^{-8})$$

$$x = 1.3 \times 10^{-4} M$$



I	0.50	0	0
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C	-x	+x	+x
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E	0.50 - x	x	x
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$$[HOCl] = 0.5 M$$

$$[H_3O^+] = [OCl^-] = 1.3 \times 10^{-4} M$$

- b. Find the pH of the solution.

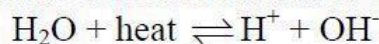
$$\text{pH} = -\log[\text{H}^+] = -\log(1.3 \times 10^{-4} \text{ M}) = 3.89$$

- 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 $[\text{H}_3\text{O}^+] = [\text{OH}^-]$. Explain why the pH drops when the temperature increases.

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



As heat is added, the equilibrium shifts to the right to use up that heat. This causes the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ to increase at the same rate. The pH is lower because $[\text{H}_3\text{O}^+]$ is higher.

- 4) The following questions pertain to a 0.58 M solution of benzoic acid, $\text{HC}_7\text{H}_5\text{O}_2$, at 25°C . $K_a = 6.4 \times 10^{-5}$.
- a. Find the concentration of $\text{H}_3\text{O}^+(\text{aq})$ in the solution.

				$K_a = \frac{[\text{H}^+][\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]} = \frac{(x)(x)}{0.58 - x}$
				assume $0.58 - x = 0.58$
	$\text{HC}_7\text{H}_5\text{O}_2(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{C}_7\text{H}_5\text{O}_2^-(\text{aq})$			
I	0.58	0	0	$6.4 \times 10^{-5} = \frac{x^2}{0.58}$
C	-x	+x	+x	$x^2 = (0.58)(6.4 \times 10^{-5})$
E	$0.58 - x$	x	x	$x = \sqrt{(0.58)(6.4 \times 10^{-5})}$
				$x = [\text{H}_3\text{O}^+] = 6.1 \times 10^{-3} \text{ M}$

- b. Find the pH of the solution.

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

- 5) Find the concentration of $\text{H}_3\text{O}^+(\text{aq})$ in a 1.75 M solution of lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$, at 25°C. $K_a = 1.38 \times 10^{-4}$.

				$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_3\text{H}_5\text{O}_3^-]}{[\text{HC}_3\text{H}_5\text{O}_3]} = \frac{(x)(x)}{1.75 - x}$
				assume $1.75 - x = 1.75\text{M}$
	$\text{HC}_3\text{H}_5\text{O}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{C}_3\text{H}_5\text{O}_3^-(\text{aq})$			
I	1.75	0	0	$1.38 \times 10^{-4} = \frac{x^2}{1.75}$
C	-x	+x	+x	$x^2 = (1.75)(1.38 \times 10^{-4})$
E	1.75 - x	x	x	$x = \sqrt{(1.75)(1.38 \times 10^{-4})}$
				$x = [\text{H}_3\text{O}^+] = 1.55 \times 10^{-2}\text{M}$

- 6) Write the equilibrium expression for the ionization of HOI, and calculate the concentration of $\text{HOI}(\text{aq})$ in solution if $[\text{H}_3\text{O}^+] = 2.3 \times 10^{-5}\text{M}$ and $\text{p}K_a = 10.7$ at 25°C.

$\text{HOI}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OI}^-(\text{aq})$	$K_a = \frac{[\text{H}^+][\text{OI}^-]}{[\text{HOI}]} = \frac{(2.3 \times 10^{-5})(2.3 \times 10^{-5})}{x - 2.3 \times 10^{-5}}$			
$K_a = 10^{-\text{p}K_a}$	assume $x - 2.3 \times 10^{-5} = x$			
$K_a = 10^{-10.7}$				
$K_a = 2 \times 10^{-11}$	$2 \times 10^{-11} = \frac{(2.3 \times 10^{-5})^2}{x}$			
$\text{HOI}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OI}^-(\text{aq})$	$x = \frac{(2.3 \times 10^{-5})^2}{2 \times 10^{-11}}$			
I	x	0	0	$x = 30$
C	x	+ 2.3×10^{-5}	+ 2.3×10^{-5}	
E	$x - 2.3 \times 10^{-5}$	2.3×10^{-5}	2.3×10^{-5}	$x = [\text{HOI}] = 30\text{M}$

- 7) A 0.45 M solution of propanoic acid, $\text{HC}_3\text{H}_5\text{O}_2$, experiences 1.58% ionization.
- a. Find the concentrations of all aqueous species in the solution at equilibrium.

	$\text{HC}_3\text{H}_5\text{O}_{2(aq)} \rightarrow$	$\text{H}^+_{(aq)}$	+	$\text{C}_3\text{H}_5\text{O}_2^-_{(aq)}$
<i>I</i>	0.45	0		0
<i>C</i>	- 0.45(0.0158)	+ 0.45(0.0158)		+ 0.45(0.0158)
<i>E</i>	0.45 - 0.45(0.0158)	0.45(0.0158)		0.45(0.0158)

$$0.45(0.0158) = 7.1 \times 10^{-3} M$$

$$[\text{HC}_3\text{H}_5\text{O}_2] = 0.45 M - 0.0071 M = 0.44 M$$

$$[\text{H}_3\text{O}^+] = [\text{C}_3\text{H}_5\text{O}_2^-] = 7.1 \times 10^{-3} M$$

- b. Find the pH of the solution.

$$\text{pH} = -\log[\text{H}^+] = -\log(7.1 \times 10^{-3} M) = 2.15$$

- c. What concentration of HCl would produce a solution with the same pH as a 0.45 M solution of propanoic acid, $\text{HC}_3\text{H}_5\text{O}_2$? Justify your answer.

A $7.1 \times 10^{-3} M$ 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.

$$\text{pH} = -\log[\text{H}^+] = -\log(7.1 \times 10^{-3} M) = 2.15$$

- 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.

	$\text{HOC}_6\text{H}_{5(aq)} \rightarrow$	$\text{H}^+_{(aq)}$	+	$\text{OC}_6\text{H}_5^-_{(aq)}$
<i>I</i>	0.57	0		0
<i>C</i>	- 0.57(0.000684)	+ 0.57(0.000684)		+ 0.57(0.000684)
<i>E</i>	0.57 - 0.57(0.000684)	0.57(0.000684)		0.57(0.000684)

$$0.57(0.000684) = 3.9 \times 10^{-4} M$$

$$[\text{HOC}_6\text{H}_5] = 0.57 M \text{ and } [\text{H}_3\text{O}^+] = [\text{OC}_6\text{H}_5^-] = 3.9 \times 10^{-4} M$$

- b. Find the pH of the solution.

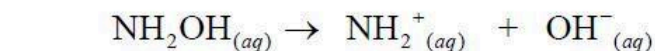
$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(3.9 \times 10^{-4} M) = 3.41$$

- 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 $3.9 \times 10^{-4} M$ 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.

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(3.9 \times 10^{-4} M) = 3.41$$

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



$$I \quad 1.5 \quad 0 \quad 0$$

$$C \quad -x \quad +x \quad +x$$

$$E \quad 1.5 - x \quad x \quad x$$

$$K_b = \frac{[\text{NH}_2^+][\text{OH}^-]}{[\text{NH}_2\text{OH}]} = \frac{(x)(x)}{1.5 - x}$$

$$\text{assume } 1.5 - x = 1.5$$

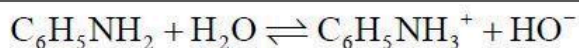
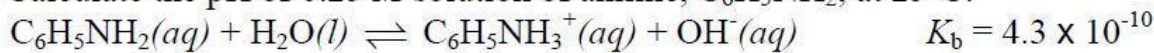
$$9.1 \times 10^{-9} = \frac{(x)^2}{1.5}$$

$$x = \sqrt{1.5 \times 9.1 \times 10^{-9}} = 1.2 \times 10^{-4} = [\text{OH}] = 1.2 \times 10^{-4} M$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.2 \times 10^{-4} M) = 3.92$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.92 = 10.08$$

10) Calculate the pH of 0.25 M solution of aniline, $\text{C}_6\text{H}_5\text{NH}_2$, at 25°C .



<i>I</i>	0.25	0	0
<i>C</i>	$-x$	$+x$	$+x$
<i>E</i>	$0.25 - x$	x	x

$$K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{NH}_2]} = \frac{(x)(x)}{0.25 - x}$$

assume $0.25 - x = 0.25$

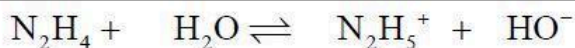
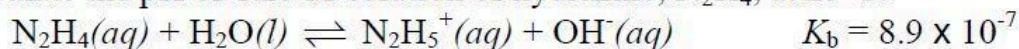
$$4.3 \times 10^{-10} = \frac{(x)^2}{0.25}$$

$$x = \sqrt{0.25 \times 4.3 \times 10^{-10}} = 1.2 \times 10^{-4} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.2 \times 10^{-4} \text{ M}) = 3.92$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.92 = 10.08$$

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



<i>I</i>	1.25	0	0
<i>C</i>	$-x$	$+x$	$+x$
<i>E</i>	$1.25 - x$	x	x

$$K_b = \frac{[\text{N}_2\text{H}_5^+][\text{OH}^-]}{[\text{N}_2\text{H}_4]} = \frac{(x)(x)}{1.25 - x}$$

assume $1.25 - x = 1.25$

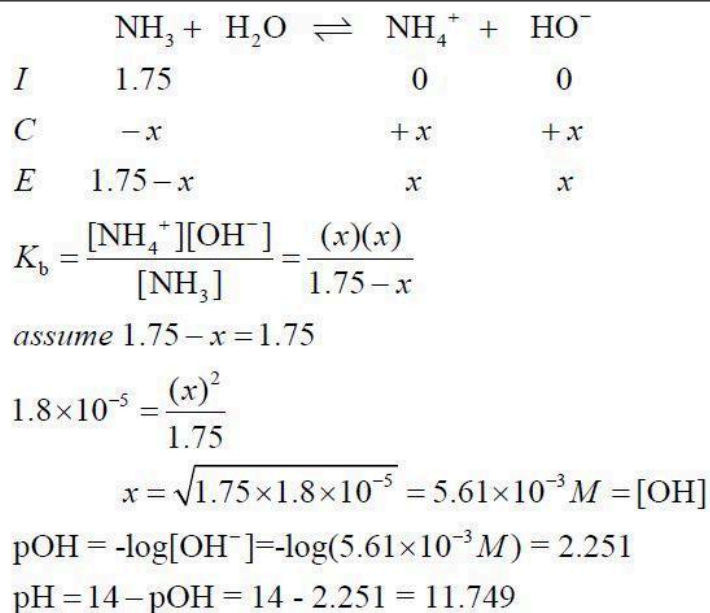
$$8.9 \times 10^{-7} = \frac{(x)^2}{1.25}$$

$$x = \sqrt{1.25 \times 8.9 \times 10^{-7}} = 1.05 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

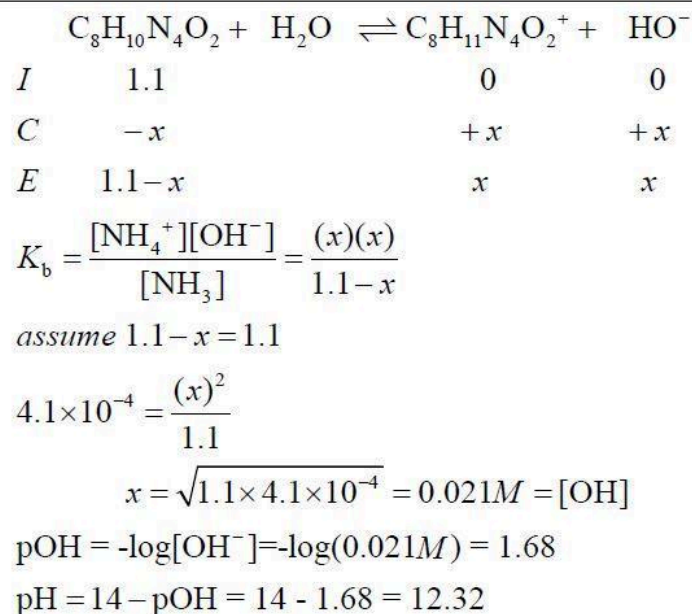
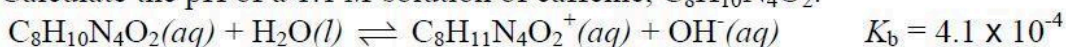
$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.05 \times 10^{-3} \text{ M}) = 2.979$$

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

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



13) Calculate the pH of a 1.1 M solution of caffeine, $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$.



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

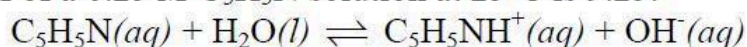
products). For this reason, $[\text{CH}_3\text{COO}^-] > [\text{ClO}^-]$.

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

$$[\text{H}_3\text{O}^+] = 10^{-4.2} = 6.31 \times 10^{-5} M$$

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

16) The pH of a 0.25 M $\text{C}_5\text{H}_5\text{N}$ solution at 25°C is 9.25.



a. Is the solution acidic or basic? Explain.

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

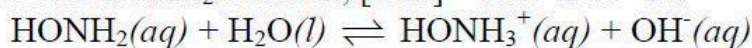
b. Find the concentration of all the aqueous species in the solution.

$$\text{pOH} = 14 - \text{pH} = 14 - 9.25 = 4.75$$

$$[\text{OH}^-] = [\text{C}_5\text{H}_5\text{NH}^+] = 10^{-\text{pOH}} = 10^{-4.75} = 1.8 \times 10^{-5} M$$

$$[\text{C}_5\text{H}_5\text{N}] = 0.25 M$$

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



a. Find $[\text{HONH}_3^+]$.

$$[\text{OH}^-] = [\text{HONH}_3^+] = 5.28 \times 10^{-6} M$$

b. Find K_b .

$$K_b = \frac{[\text{OH}^-][\text{HONH}_3^+]}{[\text{HONH}_2]} = \frac{(5.28 \times 10^{-6})(5.28 \times 10^{-6})}{0.450} = 6.20 \times 10^{-11}$$

c. Find pOH.

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

d. Find pH.

$$\text{pH} = 14 - \text{pOH} = 14 - 5.277 = 8.723$$

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

$$\%ionization = \frac{[\text{HONH}_3^+]_{\text{equilibrium}}}{[\text{HONH}_2]_{\text{initial}}} \times 100 = \frac{5.28 \times 10^{-6}\text{ M}}{0.450\text{ M}} \times 100 = 1.17 \times 10^{-3}\%$$

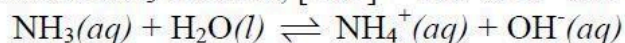
- f. What concentration of NaOH would be required to make a solution with the same pH that was calculated in part d.?

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

- 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, $[\text{OH}^-] = 1.27 \times 10^{-3}\text{ M}$.



- a. Find $[\text{NH}_4^+]$.

$$[\text{OH}^-] = [\text{NH}_4^+] = 1.27 \times 10^{-3}\text{ M}$$

- b. Find K_b .

$$K_b = \frac{[\text{OH}^-][\text{NH}_4^+]}{[\text{NH}_3]} = \frac{(1.27 \times 10^{-3})^2}{0.032 - 0.00127} = 5.25 \times 10^{-5}$$

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.

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

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

$$\%ionization = \frac{[\text{NH}_4^+]_{\text{equilibrium}}}{[\text{NH}_3]_{\text{initial}}} \times 100 = \frac{1.27 \times 10^{-3}\text{ M}}{0.032} \times 100 = 3.9\%$$

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

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

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

100% as KOH is a strong base.

19) Citric acid $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ is a polyprotic acid. $K_{a1} = 8.4 \times 10^{-4}$, and $K_{a2} = 1.8 \times 10^{-5}$ at 25°C .

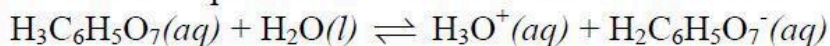
- a. Which of the following species has the lowest concentration in a 1.0 M $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ solution: $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$, or $\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$? Justify your answer.

$[\text{HC}_6\text{H}_5\text{O}_7^{2-}]$ 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 $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ solution: $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$, or $\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$? Justify your answer.

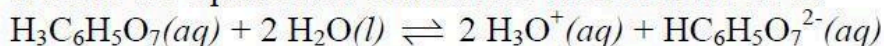
$[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]$ is the largest, because it is a weak acid. The equilibrium lies far to the left (mostly reactants).

- c. What is the equilibrium constant for the reaction below?



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

- d. What is the equilibrium constant for the reaction below?



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

$$\frac{[\text{H}_3\text{O}^+]^2[\text{HC}_6\text{H}_5\text{O}_7^{2-}]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]} = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]} \times \frac{[\text{H}_3\text{O}^+][\text{HC}_6\text{H}_5\text{O}_7^{2-}]}{[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]}$$

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

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

20) $\text{p}K_b = 1.00$ for $\text{CH}_3\text{NH}_2(aq)$ at 25°C . Find $\text{p}K_a$ for $\text{CH}_3\text{NH}_3^+(aq)$ at this temperature.

$$\text{p}K_w = \text{p}K_a + \text{p}K_b$$

$$14 - 1.00 = \text{p}K_a$$

$$\text{p}K_a = 13$$

- 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

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

$$[\text{H}^+] = 10^{\text{pH}} M$$

$$[\text{H}^+] = 1 \times 10^{-1} M$$

$$[\text{H}^+] = \frac{1 \text{ mol}}{10 \text{ L}}$$

when pH = 2

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

$$[\text{H}^+] = 10^{\text{pH}} M$$

$$[\text{H}^+] = 1 \times 10^{-2} M$$

$$[\text{H}^+] = \frac{1 \text{ mol}}{100 \text{ L}}$$

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

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

$$\text{pOH} = 14 - 12$$

$$\text{pOH} = 2$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$[\text{OH}^-] = 10^{\text{pOH}}$$

$$[\text{OH}^-] = 1 \times 10^{-2} M$$

$$[\text{OH}^-] = \frac{1 \text{ mol}}{100 \text{ L}}$$

when pH = 11

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

$$\text{pOH} = 14 - 11$$

$$\text{pOH} = 3$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$[\text{OH}^-] = 10^{\text{pOH}}$$

$$[\text{OH}^-] = 1 \times 10^{-3} M$$

$$[\text{OH}^-] = \frac{1 \text{ mol}}{1000 \text{ L}}$$

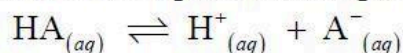
You will add 900 mL of water.

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 $[\text{HBr}]$ less than, equal to, or greater than $[\text{CH}_3\text{COOH}]$? Justify your answer.

$[\text{HBr}] < [\text{CH}_3\text{COOH}]$. 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.)



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

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

Since $0.08 < 1$, $[\text{HA}] > [\text{A}^-]$

A question like this appeared on the 2015 Exam.