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Basketball – Acid and Base Practice Problems

Strong Acid/Base Titrations

1. It takes 40.00 mL of 0.151 M NaOH to neutralize 25.0 mL of and H_2SO_4 solution, what is the concentration of the H_2SO_4?

$0.04000\,L \left(\frac{0.151\,mol\ NaOH}{1\,L}\right)\left(\frac{1\ mol\ H_2SO_4}{2\ mol\ NaOH}\right)\left(\frac{1}{0.0250}\right) = 0.120\,M$

2. It takes 154.0 mL of 0.831 M HBr to neutralize 258.9 mL of and KOH solution, what is the concentration of the KOH?

$0.1540\,L \left(\frac{0.831\,mol}{L}\right)\left(\frac{1\ KOH}{1\ HBr}\right)\left(\frac{1}{0.25892}\right) = 0.494\,M$

Weak Acid/Base Equilibrium

3. Calculate (a) the pH and (b) the percent ionization of a 0.250 M $HC_2H_3O_2$ solution. $K_a(HC_2H_3O_2) = 1.8 \times 10^{-5}$. (The formula for acetic acid may also be written as CH_3COOH.)

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

	0.250	0	0
	$-x$	$+x$	$+x$
	$0.250 - x$	x	x

$1.8 \times 10^{-5} = \frac{x^2}{0.25 - x}$

$\sqrt{4.5 \times 10^{-6}} = \sqrt{x^2}$

$2.12 \times 10^{-3}\,M = [H^+] = x$

$pH = -\log(2.12 \times 10^{-3}) = 2.67$

$\% = \frac{2.12 \times 10^{-3}}{0.25} \times 100\%$

$\% = 0.00848\%$

4. Calculate the pH of a 0.600 M solution of methylamine CH_3NH_2. $K_b = 4.4 \times 10^{-4}$.

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

	0.6	0	0
	$-x$	$+x$	$+x$
	$0.6 - x$	x	x

$4.4 \times 10^{-4} = \frac{x^2}{0.6 - x}$

$\sqrt{2.64 \times 10^{-4}} = \sqrt{x^2}$

$1.62 \times 10^{-2}\,M = x = [OH^-]$

$pOH = -\log[OH^-] = 1.79$

$pH = 14 - 1.79 = 12.21$

5. The pH of a 0.22 M solution of a weak acid is 1.67. What is the K_a of the acid? Look at Appendix D of your textbook. Predict which weak acid you are working with.

$HA \rightleftharpoons H^+ + A^-$

	0.22	0	0
	$-x$	$+x$	$+x$
	$0.22 - x$	x	x

Use pH!

$-\log[H^+] = 1.67$

$10^{-\log[H^+]} = 10^{-1.67}$

$x = [H^+] = 0.0214\,M$

$K_a = \frac{(2.14 \times 10^{-2})^2}{0.22 - (2.14 \times 10^{-2})}$

$K_a = 2.3 \times 10^{-3}$

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6. What is the pH of a 0.423 M solution of hypochlorous acid, HClO? The K_a for hypochlorous acid is 3.0 x 10^{-8}. What is the percent ionization of this solution?

$HClO \rightleftharpoons H^+ + ClO^-$

0.423	0	0
$-x$	$+x$	$+x$
$0.423-x$	x	x

$3\times10^{-8} = \frac{x^2}{0.423}$

$1.1\times10^{-4} M = x = [H^+]$

$pH = -\log(1.1\times10^{-4}) = 3.95$

$\% = \frac{1.1\times10^{-4}}{0.423}\times100\% = 2.6\times10^{-2}\%$

Buffers

7. To study the effect of a weekly acidic medium on the rate of corrosion on a metal, a chemist prepares a buffer solution by mixing 20 mL of 0.11 M $NaC_2H_5O_2$ and 23 mL of 0.090 M $HC_2H_3O_2$. What is the pH of the solution? NOTE – the pK_a of $HC_2H_3O_2$ is 4.74.

$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$

$pH = 4.74 + \log\left(\frac{0.11}{0.09}\right) = 4.83$

8. A buffer is made that is 500 mL of 0.1 M HCO_3^- and 520 mL of 0.11 M CO_3^{2-}. What is the pH of this solution? What is the pH of the solution after 5 mL of 4 M NaOH is added? Note – The K_a is 5.6×10^{-11}.

$pH = -\log(5.6\times10^{-11}) + \log\left(\frac{0.11}{0.1}\right) = 10.29$

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

B	0.05 mol	0.02 mol	0.0572
C	-0.02	-0.02	+0.02
A	0.03	0	0.0772

$[HCO_3^-] = \frac{0.03}{1.025} = 0.029\ M$

$[CO_3^{2-}] = \frac{0.0772}{1.025} = 0.0753\ M$

$pH = 10.29 + \log\left(\frac{0.0753}{0.029}\right) = 10.66$

9. What is the pH of a solution made by adding 0.40 g of NaOH to 100 mL of 0.4 M CH_3COOH? Is it a buffering solution? NOTE – The pK_a of CH_3COOH is 4.74.

$0.40\,g\left(\frac{1\ mol}{40\,g}\right) = 0.01\ mol\ OH^-$

$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

0.04 mol	0.01	0
-0.01	-0.01	+0.01
0.03 mol	0	0.01

$[HA] = \frac{0.03}{0.1} = 0.3\ M$ $[A^-] = \frac{0.01}{0.1} = 0.1\ M$

$pH = 4.74 + \log\left(\frac{0.1}{0.3}\right)$

$pH = 4.26$

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10. What is the pH of the solution formed from mixing 50 mL of a 0.250 M benzoic acid (HA, $K_a = 6.5 \times 10^{-5}$) and 40.0 mL of 0.200 M solution of NaOH? Is this solution a buffer?

$HA + OH^- \rightarrow A^- + H_2O$

	0.0125	0.008	0	
-0.008	-0.008	+0.008	
	0.0045	0	0.008	

$[HA] = \frac{0.0045}{0.09} = 0.05\ M$

$[A^-] = \frac{0.008}{0.09} = 0.089\ M$

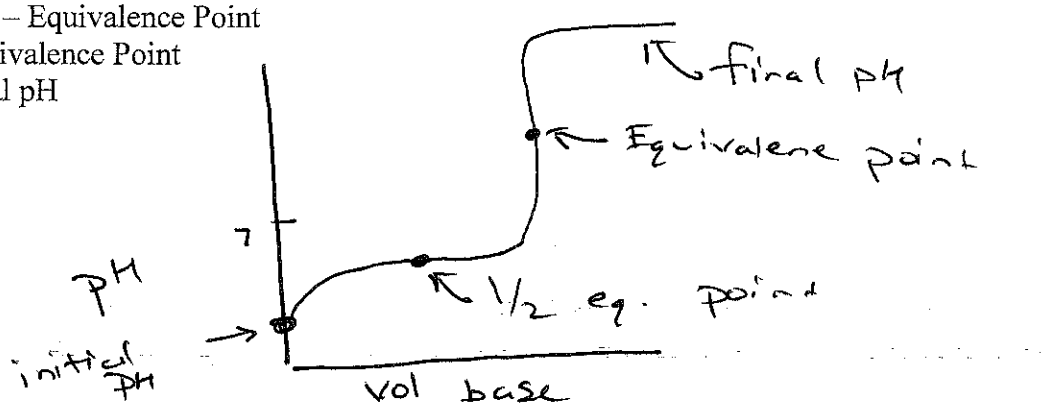
$pH = 4.19 + \log\left(\frac{0.089}{0.05}\right)$

$pH = 4.44$ ±1 pK_a so it is a buffer

Weak Acid/Base Titrations

11. You generate a titration curve for the titration of 50.0 mL of benzoic acid (HA, $K_a = 6.5 \times 10^{-5}$) with 0.200 M solution of NaOH. 0.20 M

a. Draw a titration curve for this titration and indicate the following regions:
(i) Initial pH
(ii) Half – Equivalence Point
(iii) Equivalence Point
(iv) Final pH



b. Give the primary species in solution and write the equilibrium reaction that determines the pH at the three designated points given below:

(i) Initial pH: Primary species HA Equilibrium: $HA \rightleftharpoons H^+ + A^-$

(ii) Half – Equivalence Point: Primary Species: HA + A^- Equilibrium: $HA \rightleftharpoons A^- + H^+$

(iii) Equivalence Point: Primary Species: A^- Equilibrium: $A^- + H_2O \rightleftharpoons HA + OH^-$

c. Calculate the pH after the addition of 40.0 mL of NaOH solution. Be sure to check any simplifying approximations made!

$HA + OH^- \rightarrow H_2O + A^-$

	0.01	0.008	0	
-0.008	-0.008	+0.008	
	0.002	0	0.008	

$[HA] = \frac{0.002}{0.09} = 0.022\ M$

$[A^-] = \frac{0.008}{0.09} = 0.089\ M$

$pH = 4.19 + \log\left(\frac{0.089}{0.022}\right) = 4.15$

d. Calculate the pH after the addition of 70.0 mL of NaOH. Be sure to check any simplifying approximations made.

Past the equivalence point ∴

$pOH = -\log(0.2\ M) = 0.70$

$pH = 14 - 0.7 = 13.30$