Review Problems

Problem #1

a. Calculate the pH of a solution which is 1.00 M HNO_2 and 1.00 M $NaNO_2$.
b. Calculate the pH after 0.10 mole of NaOH is added to 1.00 L of the above solution.
c. Calculate the pH after 0.20 mol of HCl is added to 1.00 L of the solution.

Problem #2
Salts of the following weak acids are available in lab. The Ka values for the weak acids of these salts are listed below. Describe how you would make 1.00 L of a buffer solution with a pH of 11.00.

Ka of HSO_4^- = 1.2×10^{-2}
Ka of HSO_3^- = 1.0×10^{-7}
Ka of $H_2PO_4^-$ = 6.2×10^{-8}
Ka of HCO_3^- = 5.6×10^{-11}
Ka of HPO_4^{2-} = 4.8×10^{-13}

Problem #3
Consider the titration of 100. mL of 0.10 M HCl with 0.20 M NaOH.

Calculate the pH of the solution at various selected points in the course of the titration. (The volumes are added to the original solution)

POINT A:	No NaOH has been added to the acid
POINT B:	50.0 mL of NaOH has been added
POINT C:	100. mL of LiOH has been added

Problem #4
Consider the titration of 50.0 mL of 0.200 M NH_3 ($K_b = 1.8 \times 10^{-5}$) by 0.100 M HNO_3.

Calculate the pH of the solution at various selected points in the course of the titration. (The volumes are added to the original solution)

POINT A:	No acid is added
POINT B:	10.0 mL of acid are added
POINT C:	50.0 mL of acid aré added
POINT D:	150.0 mL of acid are added

Problem #5
Calculate the molarity of the NaOH when titrated against a standardized HCl solution:

Molarity of the HCl	0.100 M
Volume of HCl	50.00 mL
Original volume of NaOH in buret	12.25 mL
Final volume of NaOH in buret	23.31 mL

Calculate the molar mass of an unknown diprotic acid that is titrated with the above NaOH solution

Mass of unknown acid	0.325 g
Original volume of NaOH in buret	23.31 mL
Final volume of NaOH in buret	39.30 mL

*The acid is either sulfuric acid or oxalic acid. Which one, and why?

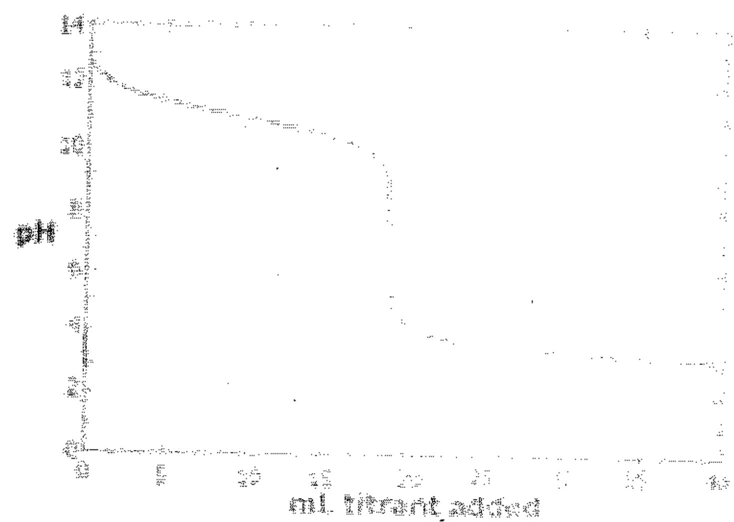
Problem 7:

You titrate 100.00 ml of 1.00 M NaCN with 1.00M HCl. The Ka of HCN is 6.2×10^{-10}

Point A: What is the pH after you add 0.00 ml of acid?
Point B: What is the pH after you add 25.00 ml of acid?
Point C: What is the pH after you add 50.00 ml of acid?
Point D: What is the pH after you add 100.00 ml of acid?
Point E: What is the pH after you add 150.00 ml of acid?

Problem 8:

25.00 mL of NaOH is titrated with 0.500 M acetic acid and the data is graphed below.



Calculations:

1. Determine the equivalence point for the titration.
2. Calculate the moles of acetic acid required to reach the equivalence point.
3. Calculate the moles of NaOH in the original solution.
4. Calculate the molarity of the NaOH.
5. Calculate the theoretical pH at the equivalence point.
6. Compare the theoretical pH at the equivalence point with the actual pH from the graph.

7) a) H_2O Na^+ CN^-

$$CN^- + H_2O \rightleftharpoons HCN + OH^-$$

	CN^-		HCN	OH^-
I	1.00		0	0
C	$-x$		$+x$	$+x$
E	$1.00-x$		x	x

$$K_b = \frac{1.0\times10^{-14}}{6.2\times10^{-10}} = 1.61\times10^{-5} = \frac{[HCN][OH^-]}{[CN^-]}$$

$$1.61\times10^{-5} = \frac{x^2}{1.00-x}$$

$$x = 4.01\times10^{-3}\,M = [OH^-]$$

5% $\frac{4.01\times10^{-3}}{1.00}\times100 = 0.40\%$ $pOH = -\log(4.01\times10^{-3})$

2.40

$$\boxed{pH = 14 - 2.40 = 11.60}$$

b) H_2O ~~Na^+~~ CN^- $\underline{\underline{H^+}}$ ~~Cl^-~~

$$H^+ + CN^- \rightarrow HCN$$

mmole H^+ = (1.00 M)(25.00 mL) = 25.00 mmol

mmole CN^- = (1.00 M)(100.00 mL) = 100 mmol

		H^+	CN^-	HCN
m	I	25.00	100.	0
m	C	-25.00	-25.00	$+25.00$
l	E	0	75.00	25.00

$$pH = pKa + \log\frac{B}{A}$$

$$= -\log(6.2\times10^{-10}) + \log\frac{75.00}{25.00}$$

$$\boxed{pH = 9.68}$$

c)

		H^+	+	CN^-	$\rightarrow$	HCN
m	I	50.00		100.		0
m	C	-50.00		-50.00		+50.00
o	E	0		50.00		50.00
l						

mmol H^+ = (1.00 M)(50.00 mL) = 50.00 mmol

$\frac{1}{2}$ equivalence
pH = pKa = $-\log(6.2\times10^{-10})$ = $\boxed{9.21}$

d)

		H^+	+	CN^-	$\rightarrow$	HCN
m	I	100.		100.		0
m	C	-100.		-100.		+100
o	E	0		0		100.
l						

mmol H^+ = (1.00 M)(100.00 mL) = 100. mmol

$[HCN] = \frac{100.\ mmol}{100. + 100\ mL} = 0.500\ M$

	HCN	$\rightleftharpoons$	H^+	+	CN^-
I	0.500		0		0
C	$-x$		$+x$		$+x$
E	$0.500 - x$		x		x

$K_a = 6.2\times10^{-10} = \frac{[H^+][CN^-]}{[HCN]}$

$6.2\times10^{-10} = \frac{x^2}{0.500 - x}$

$x = 1.76\times10^{-5} M = [H^+]$

5% $\frac{1.76\times10^{-5}}{0.500}\times100 = 0.0035\%$ ✓

pH = $-\log(1.76\times10^{-5})$: 4.75

e)

	H^+	+	CN^-	$\longrightarrow$	HCN
m I	150.		100.		0
m C	-100		$-100.$		$+100$
o E	50.		0		100.

$$\text{mmol } H^+ = (1.00\ M)(150.00\ mL) = 150.\ \text{mmol}$$

$$[H^+] = \frac{50.00\ \text{mmol}}{100.00 + 150.00\ mL} = 0.200\ M$$

$$pH = -\log(0.200) = \boxed{0.70}$$

8) 1.) 18.0 mL $HC_2H_3O_2$

2) mole = (0.500 M)(18.00 mL) = 9.00 mmole

3) 9.00 mmole $HC_2H_3O_2$ $\frac{1 \text{ mole NaOH}}{1 \text{ mole } HC_2H_3O_2}$ = 9.00 mmole NaOH

4) $M = \frac{9.00 \text{ mmol}}{25.00 \text{ mL}} = 0.36 \text{ M NaOH}$

5) $OH^- + HC_2H_3O_2 \rightarrow H_2O + C_2H_3O_2^-$

		OH^-	$HC_2H_3O_2$	$C_2H_3O_2^-$
m	I	9.00	9.00	0
m	C	-9.00	-9.00	+9.00
o	E	0	0	9.00
l				

$[C_2H_3O_2^-] = \frac{9.00 \text{ mmol}}{25.00 + 18.00 \text{ mL}} = 0.209 \text{ M}$

$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$

	$C_2H_3O_2^-$	$HC_2H_3O_2$	OH^-
I	0.209	0	0
C	$-x$	$+x$	$+x$
E	$0.209-x$	x	x

$K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$

$5.56 \times 10^{-10} = \frac{x^2}{0.209 - x}$

$x = 1.08 \times 10^{-5} \text{ M} = [OH^-]$

5% $\frac{1.08 \times 10^{-5}}{0.209} \times 100 = 0.005\%$ ✓

$pOH = -\log(1.08 \times 10^{-5}) = 4.97$

$pH = 14 - 4.97 = 9.03$