Using the Henderson-Hasselbalch Equation

1. A 1.00L buffer solution is prepared that contains 0.150M nitrous acid and 0.200M sodium nitrite. What is its pH?

$K_a = 7.2 \times 10^{-4}$

Ice Table Method

nitrite

$$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$$

	HNO_2		H_3O^+	NO_2^-
I	0.150M		0	0.200M
C	$-x$		$+x$	$+x$
E	$0.150 - \cancel{x}$ ignore		x	$0.200 + \cancel{x}$ ignore

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{[x][0.200]}{[0.150]} = 7.2 \times 10^{-4}$$

$$\frac{(0.200)(x)}{0.200} = \frac{1.08 \times 10^{-4}}{0.200}$$

$$x = 5.4 \times 10^{-4} = [H_3O^+]$$

$$pH = -\log[H_3O^+]$$

$$= -\log[5.4 \times 10^{-4}]$$

$$= \boxed{3.27}$$

pH = 3.27

Using the Henderson-Hasselbalch Equation

1. A 1.00L buffer solution is prepared that contains 0.150M nitrous acid and 0.200M <u>sodium nitrite</u>. What is its pH?
$K_a = 7.2 \times 10^{-4}$

H-H equation method:

conj base

$$pH = pka + \log \frac{[A^-]}{[HA]}$$ ← acid

A^- = conj base = $NaNO_2$ = 0.200 M

HA = acid = nitrous acid = 0.150 M

pka = $-\log$ Ka = $-\log(7.2 \times 10^{-4})$ = 3.14267

$$pH = 3.14267 + \log \frac{[0.200]}{[0.150]} = \boxed{3.27}$$

pH = 3.27

Using the Henderson-Hasselbalch Equation

2. How many grams of sodium lactate ($CH_3CH(OH)COONa$) should be added to 1.0L of a 0.150M lactic acid ($CH_3CH(OH)COOH$) to form a buffer solution with pH=4.00? K_a = 1.4×10^{-4}; molar mass of sodium lactate = 112.1g/mol

$pH = pka + \log \frac{[A^-]}{[HA]}$

$pka = -\log(ka) = -\log(1.4 \times 10^{-4}) = 3.85387$

HA = lactic acid = 0.150 M

pH = 4.00

A^- = sodium lactate [?]

$4.00 = 3.85387 + \log \frac{[A^-]}{[0.150]}$

$-3.85387 \quad -3.85387$

$0.14613 = \log \frac{[A^-]}{[0.150]}$

pH = 3

$[H_3O^+] = 10^{-3}$

$10^{0.14613} = \frac{[A^-]}{[0.150]}$

$1.400 = \frac{[A^-]}{[0.150]}$

$[A^-] = 0.2100 M$

$0.2100 \frac{mol}{L} \times 1L = 0.2100 \cancel{mol} \times \frac{112.1g}{\cancel{mol}}$

$= 23.54g$

$\rightarrow \boxed{24g}$

A: 24g

Buffer calculations

$H = 1.00794\ g/mol$
$Br = \underline{79.904\ g/mol}$
$80.91194\ g/mol$

1. A 1.0 L buffer solution contains 0.150 M nitrous acid and 0.200 M sodium nitrite. $K_a = 7.2 \times 10^{-4}$ (a) What is the pH of the buffer? (b) What is the pH after adding 1.00 g HBr?

$MM = 80.91194\ g/mol$

$pH = pka + \log \frac{[A^-]}{[HA]}$ $\quad pka = \cancel{-\log}\ -\log(7.2 \times 10^{-4}) = 3.1427$

(a) $pH = 3.1427 + \log \frac{[0.200]}{[0.150]} = \boxed{3.27}$

$1.00\ \cancel{g}\ HBr \left(\frac{1\ mol}{80.91194\ \cancel{g}}\right) = 0.01236\ mol\ HBr$

(b) HBr = strong acid → react with base A^-

$HBr + NO_2^- \longrightarrow HNO_2 + Br^-$

NO_2^-: HBr neutralize NO_2^- $\quad 0.01236\ mol\ HBr \left(\frac{1\ mol\ NO_2^-}{1\ mol\ HBr}\right) = 0.1236\ mol\ NO_2^-$ neutralized

$0.200 \frac{mol\ NO_2}{L} \times 1\ L = 0.200\ mol\ NO_2^- - 0.1236\ mol\ NO_2 = \frac{0.18764\ mol\ NO_2\ remaining}{1.0\ L} = 0.18764 \frac{mol}{L}\ NO_2^-$

HNO_2: formed in reaction $\quad 0.01236\ mol\ HBr \left(\frac{1\ mol\ HNO_2}{1\ mol\ HBr}\right)$

$= 0.01236\ mol\ HNO_2$ formed

$0.150 \frac{mol\ HNO_2}{L} \times 1\ L = 0.150\ mol\ HNO_2 + 0.01236\ mol\ HNO_2\ formed = \frac{0.16236\ mol\ HNO_2}{1.0\ L}$

$= 0.16236\ M$

$pH = 3.1427 + \log \frac{[0.18764]}{[0.16236]}$

$= 3.1427 + 0.062846 = 3.2055 \rightarrow \boxed{3.21}$

A: (a) 3.27
(b) 3.21

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8x10^{-5}$.

(a) What is the pH of the buffer? A: 4.74

(b) Calculate the pH after 0.040 mol HCl is added. A: 4.69

(c) Calculate the pH after 0.040 mol NaOH is added. A: 4.80

$$pH = pka + \log \frac{[A^-] \leftarrow CH_3COO^{\ominus}}{[HA] \leftarrow CH_3COOH}$$

$$pka = -\log(1.8 \times 10^{-5}) = 4.7447$$

$$[A^-] = \frac{0.600\,mol}{2.00L} = 0.300\,M$$

$$[HA] = \frac{0.600\,mol}{2.00L} = 0.300\,M$$

$$pH = 4.7447 + \log \frac{[0.300]}{[0.300]}$$

$$\log 1 = 0$$

$$pH = 4.7447 + \log(1)$$

$$pH = 4.7447 + 0 = 4.7447 \rightarrow \boxed{4.74}$$

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8x10^{-5}$. = 4.7447 = pka

(a) What is the pH of the buffer? A: 4.74

(b) Calculate the pH after 0.040 mol HCl is added. A: 4.69

$$\underset{\text{}}{HCl} + \underbrace{CH_3COO^-}_{\text{decrease}} \rightarrow \underbrace{CH_3COOH}_{\text{increase}} + Cl^-$$

$$pH = pka + \log \frac{[A^-]}{[HA]}$$

0.040 mol HCl:

A^- CH_3COO^- $\quad 0.040 \text{ mol HCl} \left(\frac{1 \text{ mol } CH_3COO^-}{1 \text{ mol HCl}} \right) = 0.040 \text{ mol } CH_3COO^-$

$0.600 \text{ mol initial} - 0.040 \text{ mol neutralized} = \frac{0.560 \text{ mol}}{2.00L} CH_3COO^- \text{ remains}$

$= 0.280 \text{ M } CH_3COO^-$

HA CH_3COOH $\quad 0.040 \text{ mol HCl} \left(\frac{1 \text{ mol } CH_3COOH}{1 \text{ mol HCl}} \right) = 0.040 \text{ mol } CH_3COOH$

$0.600 \text{ mol initial} + 0.040 \text{ mol formed} = \frac{0.640 \text{ mol}}{2.00L} CH_3COOH = 0.320 \text{ M } CH_3COOH$

$$pH = 4.7447 + \log \frac{[0.280]}{[0.320]} = 4.7447 + (-0.05799) = 4.6867 \rightarrow \boxed{4.69}$$

Buffer calculations

2. A buffer is made by adding 0.600 mol CH_3COOH and 0.600 mol CH_3COONa to enough water to make 2.00L of solution. $K_a = 1.8 \times 10^{-5}$. pka = 4.7447

(a) What is the pH of the buffer? A: 4.74

(c) Calculate the pH after 0.040 mol NaOH is added. A: 4.80

$$NaOH + \underset{\text{decrease}}{CH_3COOH} \rightarrow \underset{\text{increase}}{CH_3COO^-} + H_2O$$

[HA] CH_3COOH $\quad 0.040\text{mol NaOH}\left(\frac{1\text{mol } CH_3COOH}{1\text{mol NaOH}}\right) = 0.040\text{ mol } CH_3COOH$ neutralized

$$0.600\text{mol} - 0.040\text{mol} = \frac{0.560\text{ mol}}{2.00L}\ CH_3COOH \text{ remains} = 0.280M\ CH_3COOH$$

[A⁻] CH_3COO^- $\quad 0.040\text{mol NaOH}\left(\frac{1\text{mol } CH_3COO^-}{1\text{mol NaOH}}\right) = 0.040\text{mol } CH_3COO^-$

$$0.600\text{ mol} + 0.040\text{ mol} = \frac{0.640\text{ mol}}{2.00L} = 0.320M\ CH_3COO^-$$

$$pH = 4.7447 + \log\frac{[0.320]}{[0.280]}$$

$$= 4.7447 + 0.05799 = 4.803 \rightarrow 4.80$$

Strong Acid/Strong Base Calculations

In the titration of 25.0mL of 0.100M KOH with 0.100M HNO_3, determine the pH:

(a) At the start of the titration (no acid added) A:13.0

(b) When 24.9mL acid has been added A: 10.3

(c) When 25.1mL acid has been added A: 3.7

(a) $KOH \xrightarrow{H_2O} K^+ + OH^-$ $\quad 0.100\,M\ KOH\left(\frac{1\,OH^-}{1\,KOH}\right) = 0.100\,M\ OH^-$

$pOH = -\log(0.100\,M) = 1.000 \qquad pH = 14 - 1.000 = \boxed{13.000}$

(b) $KOH + HNO_3 \longrightarrow H_2O + KNO_3$

$0.100\,\frac{mol\ HNO_3}{\cancel{L}} \times 0.0249\cancel{L} = 0.00249\ mol\ HNO_3\left(\frac{1\,mol\ KOH}{1\,mol\ HNO_3}\right) = 0.00249\ mol\ KOH$ neutralized

$0.100\,\frac{mol\ KOH}{L} \times 0.0250\,L = \underline{0.00250}\ mol\ KOH$ initial $- 0.00249\ mol = 0.00001\ mol\ OH^-$ remaining

$0.0250\,L + 0.0249\,L = 0.0499\,L \qquad \frac{0.00001\ mol\ OH^-}{0.0499\,L} = 2.004\times10^{-4}\,M\ OH^-$

$pOH = -\log(2.004\times10^{-4}) = 3.6981 \qquad pH = 14 - 3.6981 = \boxed{10.302}$

(c) $0.100\,\frac{mol}{L}\,HNO_3 \times 0.0251\,L = \underline{0.00251}\ mol\ HNO_3$

$0.00250\ mol\ KOH\left(\frac{1\,mol\ HNO_3}{1\,mol\ KOH}\right) = 0.00250\ mol\ HNO_3$

$0.00251\ mol\ HNO_3$ added $- 0.00250\ mol\ HNO_3$ used $= 1.00\times10^{-5}\ mol\ HNO_3$ left

$0.0250\,L + 0.0251\,L = 0.0501\,L \qquad \frac{1.00\times10^{-5}\ mol\ HNO_3}{0.0501\,L} = 1.996\times10^{-4} = [HNO_3]$

$[H_3O^+] = 1.996\times10^{-4} \qquad pH = -\log(1.996\times10^{-4}) = 3.6998 \rightarrow \boxed{3.700}$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH (K_a = $1.8x10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

a.) At the start of the titration A: 2.78

b.) When 20.0mL of 0.150M NaOH has been added A:4.87

c.) At the equivalence point A: 8.81

d.) When 50.0mL of 0.150M NaOH has been added A: 12.42

(a) No base added - only have acetic acid equilibrium

$$CH_3COOH_{(aq)} + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

	CH_3COOH	H_3O^+	CH_3COO^-
I	0.150M	0	0
C	$-x$	$+x$	$+x$
E	$0.150 - \cancel{x}$ ignore	x	x

$$Ka = \frac{[x][x]}{0.150} = 1.8 \times 10^{-5}$$

$$x^2 = 2.7 \times 10^{-6}$$

$$x = 1.643 \times 10^{-3} = [H_3O^+]$$

$$pH = -\log(1.643 \times 10^{-3})$$

$$= \boxed{2.784}$$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8x10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

b.) When 20.0mL of 0.150M NaOH has been added A:4.87

$CH_3COOH + NaOH \rightarrow H_2O + CH_3COO^- + Na^+$

Initial moles acid: $\frac{0.150 mol}{\cancel{L}} \times 0.0350\cancel{L} = 0.00525$ mol CH_3COOH

Amount NaOH added: $\frac{0.150 mol}{L} \times 0.0200L = 0.00300$ mol NaOH

$0.00300 \text{ mol NaOH} \left(\frac{1 \text{mol } CH_3COOH}{1 \text{mol NaOH}}\right) = 0.00300$ mol CH_3COOH neutralized

CH_3COOH remaining: 0.00525 mol initial − 0.00300 mol neut = $\frac{0.00225 \text{ mol}}{0.0550L}$ CH_3COOH remaining

$0.0350L + 0.0200L = 0.0550L$

$= 4.09 \times 10^{-2} M = [CH_3COOH]$

Amount CH_3COO^-:

$0.00300 \text{ mol NaOH} \left(\frac{1 \text{mol } CH_3COO^-}{1 \text{mol NaOH}}\right) = \frac{0.00300 \text{ mol}}{0.0550L}$ $CH_3COO^- = 5.45 \times 10^{-2} M = [CH_3COO^-]$

	CH_3COOH + H_2O ⇌	H_3O^+ +	CH_3COO^-
I	4.09×10^{-2}	0	5.45×10^{-2}
C	$-x$	$+x$	$+x$
E	$4.09 \times 10^{-2} \cancel{-x}$ ignore	x	$5.45 \times 10^{-2} \cancel{+x}$ ignore

$K_a = \frac{[x][5.45 \times 10^{-2}]}{[4.09 \times 10^{-2}]} = 1.8 \times 10^{-5}$

$[5.45 \times 10^{-2}][x] = 7.362 \times 10^{-7}$

$[x] = 1.35 \times 10^{-5} = [H_3O^+]$

$-\log(1.35 \times 10^{-5}) = \boxed{4.87}$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) was titrated with 0.150M NaOH. Determine the pH:

c.) At the equivalence point A: 8.81

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

All original acid & added strong base have been used up in neutralization

pH depends on CH_3COO^- equilibrium $\quad CH_3COO^-_{(aq)} + H_2O \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

Initially $0.00525 \text{ mol } CH_3COOH \left(\frac{1 \text{ mol } CH_3COO^-}{1 \text{ mol } CH_3COOH}\right) = 0.00525 \text{ mol } CH_3COO^-$

New volume: needed $0.00525 \cancel{\text{mol}} \text{ NaOH} \left(\frac{1L}{0.150 \cancel{\text{mol}}}\right) = 0.035L$

$0.0350L + 0.0350L = 0.070L$

$$[CH_3COO^-] = \frac{0.00525 \text{ mol}}{0.070L} = 0.075M$$

	CH_3COO^- + H_2O	$\rightleftharpoons$ CH_3COOH	+ OH^-
I	0.075	0	0
C	$-x$	$+x$	$+x$
E	$0.075 - \cancel{x}$ ignore	x	x

$$K_b = \frac{x^2}{0.075} = 5.56 \times 10^{-10}$$

$x^2 = 4.17 \times 10^{-11}$

$x = 6.458 \times 10^{-6} = [OH^-]$

$pOH = -\log(6.458 \times 10^{-6}) = 5.19$

$pH = 14 - 5.19 = \boxed{8.81}$

Weak Acid/Strong Base Calculations

35.0mL of 0.150M CH_3COOH (K_a = 1.8x10^{-5}) was titrated with 0.150M NaOH. Determine the pH:

d.) When 50.0mL of 0.150M NaOH has been added A: 12.42

pH depends on excess base

$CH_3COOH + NaOH \rightarrow H_2O + Na^+ + CH_3COO^-$

Base added: $\frac{0.150 \text{ mol}}{\cancel{L}} \times 0.0500\cancel{L} = 0.00750 \text{ mol NaOH}$

neutralized 0.00525 mol acid $\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol acid}}\right) = 0.00525$ mol NaOH used

0.00750 mol added − 0.00525 moles used = 0.00225 moles NaOH remaining

New volume: 0.0350L + 0.0500L = 0.085L

$[NaOH] = \frac{0.00225 \text{ mol}}{0.085L} = 0.02647 \text{ M}$ $\quad NaOH \rightarrow Na^+ + OH^-$

0.02647M NaOH = 0.02647M OH^-

pOH = −log(0.02647) = 1.577

pH = 14 − 1.577 = $\boxed{12.42}$