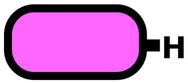
TITRATION CALCULATIONS

In a titration 20.0 mL of 0.300 M $HC_2H_3O_{2(aq)}$ with 0.30 M $NaOH_{(aq)}$, what is the pH of the solution after the following volumes of $NaOH_{(aq)}$ have been added?

a) 0.00 mL
b) 10.0 mL
c) equivalence pt
d) 30.0 mL



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

	$HA_{(aq)}$	$H^+_{(aq)}$	$A^-_{(aq)}$
I	0.300M	0	0
C	-x	+x	+x
E	0.300-x	x	x

$$K_a = \frac{[H^+_{(aq)}][C_2H_3O_2^-{}_{(aq)}]}{[HC_2H_3O_{2(aq)}]}$$

$$1.8\text{x}10^{-5} = \frac{[x][x]}{[0.300-x]}$$ *← Assumption used*

$$2.32379\text{x}10^{-3} = x$$

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

$pH = -\log [2.32\text{x}10^{-3}]$

$pH = 2.63$

.: pH = 2.63

STRONG-WEAK TITRATIONS

b)

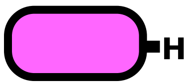
$n_{HC_2H_3O_2}$=(0.300M)(0.020L)

$n_{HC_2H_3O_2}$=0.006mol

n_{NaOH}=CV

n_{NaOH}=(0.3M)(0.010L)

n_{NaOH}= 0.003 mol



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

0.003 mol remaining

0.003 mol formed

Solution is a buffer (HA and A^- present) so we can use this formula:

$$[H_3O^+] = K_a \times \frac{n_{HA}}{n_{A-}}$$

$$= 1.8 \times 10^{-5}\ \frac{(0.003)}{(0.003)}$$

$$= 1.8 \times 10^{-5}\ M$$

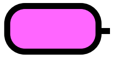
pH = -log [H^+]

pH = -log [1.8×10^{-5}]

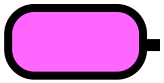
pH = 4.74

.: pH = 4.74

c)

Since  is all that remains,

1) convert to concentration
2) change the equation to start with the conjugate base
3) set-up a new ICE table



$$A^-_{(aq)} + H_2O_{(l)} \Longleftrightarrow HA_{(aq)} + OH^-_{(aq)}$$

$$V_{total} = V_{acid} + V_{base}$$

$$V_{total} = 0.020L + 0.020L$$

$$V_{total} = 0.040L$$

$$C = \frac{n}{V_{total}}$$

$$C = \frac{0.006mol}{0.040L}$$

$$C = 0.15\ M$$

$$V_{base} = \frac{n_{NaOH}}{C}$$

$$V_{base} = \frac{0.006mol}{0.3M}$$

$$V_{base} = 0.02L$$

c) $A^-_{(aq)} + H_2O_{(l)} \Longleftrightarrow HA_{(aq)} + OH^-_{(aq)}$

	$A^-_{(aq)}$	$HA_{(aq)}$	$OH^-_{(aq)}$
I	0.15	0	0
C	-x	+x	+x
E	0.15-x	9.1287×10^{-6}	9.1287×10^{-6}

Since the equation is for the conjugate base, you must use k_b instead of k_a

$$K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{1.0\times10^{-14}}{1.8\times10^{-5}}$$

$$K_b = 5.\overline{5}\times10^{-10}$$

$$K_b = \frac{[OH^-_{(aq)}][HA_{(aq)}]}{[A^-_{(aq)}]}$$

$$K_b = \frac{(x)(x)}{(0.15-x)}$$

$$5.\overline{5}\times10^{-10} = \frac{(x)(x)}{(0.15)}$$ ← *Assumption used*

$$9.1287\times10^{-6}\ M = x$$

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

$$pOH = -\log[9.1287\times10^{-6}]$$

$$pOH = 5.03959$$

$$pH = 14-5.03959$$

$$pH = 8.96$$

.: pH = 8.96

d) Solve for pH from remaining concentration of base (no ICE table needed)

$n_{HA} = 0.006mol$

$n_{NaOH}=CV$

$n_{NaOH}=(0.3M)(0.030L)$

$n_{NaOH}=0.009mol$

$n_{NaOH} - n_{HA} = 0.009mol - 0.006mol$

$= 0.003mol$ NaOH remaining

$C = \frac{n_{NaOH}}{V_{total}}$

$C = \frac{0.003mol}{0.050L}$

$C = 0.06M$

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

$pOH = -\log [0.06]$

$pOH = 1.2218$

$pH = 14-1.2218$

$pH = 12.78$

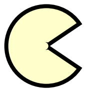
.: pH = 12.8

STRONG-WEAK TITRATIONS

In a titration 20.0 mL of 0.1 M $NH_{3(aq)}$ is titrated with 0.1 M $HCl_{(aq)}$. What is the pH of the resulting solution after the following volumes of $HCl_{(aq)}$ have been added?

a) 0.00 mL b) 10.0 mL c) equivalence pt

d) 30.0 mL



a) $$NH_{3(aq)} + H_2O_{(l)} \Longleftrightarrow OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

	$NH_{3(aq)}$	$OH^-_{(aq)}$	$NH_4^+{}_{(aq)}$
I	0.1M	0	0
C	-x	+x	+x
E	0.1-x	x	x

$$K_b = \frac{[OH^-_{(aq)}][NH_4^+{}_{(aq)}]}{[NH_{3(aq)}]}$$

$$1.8\text{x}10^{-5} = \frac{[x][x]}{[0.1-x]}$$ ← *Use assumption*

$$1.34164\text{x}10^{-3} = x$$

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

$$pOH = -\log [1.34\text{x}10^{-3}]$$

$$pOH = 2.87$$

$$pH = 14 - 2.87$$

$$pH = 11.13$$

.: pH = 11.13

b)

$n_{NH_3} = CV$

$n_{NH_3} = (0.1M)(0.020L)$

$n_{NH_3} = 0.002\ mol$

$n_{HCl} = CV$

$n_{HCl} = (0.1M)(0.010L)$

$n_{HCl} = 0.001mol$

H+

$$NH_{3(aq)} + H_2O_{(l)} \Longleftrightarrow OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

0.001 mol remaining — 0.001 mol formed

Solution is a buffer (B and HB^+ present) so we can use this formula:

$$[OH^1] = K_b \times \frac{n_B}{n_{HB+}}$$

$$= 1.8 \times 10^{-5} \frac{(0.001)}{(0.001)}$$

$$= 1.8 \times 10^{-5}\ M$$

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

$pOH = -\log [1.8x10^{-5}]$

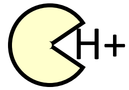
$pOH = 4.74$

$pH = 14 - 4.74$

$pH = 9.26$

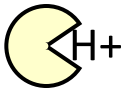
.: pH = 9.26

c)

Since 

is all that remains,

1) convert to concentration
2) change the equation to start with the conjugate acid
3) set-up a new ICE table



$$NH_4^+{}_{(aq)} <===> NH_{3(aq)} + H^+_{(aq)}$$

$$V_{total} = V_{base} + V_{acid}$$

$$V_{total} = 0.020L + 0.020L$$

$$V_{total} = 0.040L$$

$$C = \frac{n}{V_{total}}$$

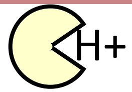
$$C = \frac{0.002mol}{0.040L}$$

$$C = 0.05\ M$$

$$V_{acid} = \frac{n_{HCl}}{C}$$

$$V_{acid} = \frac{0.002mol}{0.1M}$$

$$V_{acid} = 0.02L$$



c)

$$NH_4^+{}_{(aq)} \Longleftrightarrow NH_{3(aq)} + H^+{}_{(aq)}$$

	NH_4^+	NH_3	H^+
I	0.05	0	0
C	-x	+x	+x
E	0.05-x	x	x

Since the equation is flipped, you must use k_a instead of k_b

$$5.555\times10^{-10} = \frac{[x][x]}{[0.05-x]}$$ ← *Use assumption*

$$5.555\times10^{-10} = \frac{[x][x]}{[0.05]}$$

$$5.27\times10^{-6} = x$$

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

$$pH = -\log [5.27\times10^{-6}]$$

$$pH = 5.28$$

.: pH = 5.28

d) Solve for pH from remaining concentration of acid (no ICE table needed)

$n_{NH_3} = CV$

$n_{NH_3} = (0.1M)(0.020L)$

$n_{NH_3} = 0.002\ mol$

$n_{HCl} = CV$

$n_{HCl} = (0.1M)(0.030L)$

$n_{HCl} = 0.003mol$

$$n_{HCl} - n_{NH_3} = 0.003mol - 0.002mol$$
$$= 0.001mol\ HCl\ remaining$$

$$C = \frac{n_{HCl}}{V_{total}}$$

$$C = \frac{0.001mol}{0.050L}$$

$$C = 0.02M$$

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

$pH = -\log [0.02]$

$pH = 1.69897$

$.: pH = 1.70$