$$HC_3H_5O_3(aq) \rightleftarrows H^+(aq) + C_3H_5O_3^-(aq)$$

1. Lactic acid, $HC_3H_5O_3$, is a monoprotic acid that dissociates in aqueous solution, as represented by the equation above. Lactic acid is 1.66 percent dissociated in 0.50 *M* $HC_3H_5O_3(aq)$ at 298 K. For parts (a) through (d) below, assume the temperature remains at 298 K.

   (a) Write the expression for the acid-dissociation constant, K_a, for lactic acid and calculate its value.

   (b) Calculate the pH of 0.50 *M* $HC_3H_5O_3$.

   (c) Calculate the pH of a solution formed by dissolving 0.045 mole of solid sodium lactate, $NaC_3H_5O_3$, in 250. mL of 0.50 *M* $HC_3H_5O_3$. Assume that volume change is negligible.

   (d) A 100. mL sample of 0.10 *M* HCl is added to 100. mL of 0.50 *M* $HC_3H_5O_3$. Calculate the molar concentration of lactate ion, $C_3H_5O_3^-$, in the resulting solution.

(a) $K_a = \frac{[H^+][C_3H_5O_3^-]}{[HC_3H_5O_3]}$

	$HC_3H_5O_3$	$\rightleftarrows$	H^+	$+ C_3H_5O_3^-$
I	0.50		0	0
C	−0.00833		+0.00833	+0.00833
E	0.49167		0.00833	0.00833

$K_a = \frac{(0.00833)(0.00833)}{0.49167}$

$= 1.41 \times 10^{-4}$

(b) $[H^+] = 0.00833\ M$ pH = 2.08

(c) $(0.250\ L)(0.50\ M) = 0.125$ mole $HC_3H_5O_3$

$pH = pK_a + \log \frac{C_3H_5O_3^-}{HC_3H_5O_3}$

$= -\log(1.41\times10^{-4}) + \log\frac{0.045}{0.125}$ $(3.85 + -0.44)$

$= 3.41$

(d) Hmm – not a buffer. ∵ H^+ does NOT neutralize (from HCl) $HC_3H_5O_3$ to form $C_3H_5O_3^-$

So Use behaviour equation & ICE table

	$HC_3H_5O_3$	$\rightleftarrows$ H^+	$+ C_3H_5O_3^-$
I	0.50 M	0.05 M	0
C	−x	+x	+x
E	0.50−x	0.05+x	x

↑ Ignore x in both of these ↑

Can you Find $[H^+]$?

plug into K_a

$[C_3H_5O_3^-] = 7.0\times10^{-4}\ M$

1. Aniline, $C_6H_5NH_2$, is a neutral weak base similar to ammonia.

(a) Write an equation to show the weak base behavior of aniline.

(b) Write the equilibrium constant expression K_b for the equation shown in (a).

(c) A 0.10 M solution of aniline is prepared. The pH of this solution is measured to be 8.82. Calculate the value of K_b.

(d) A 50.00 mL sample of 0.10 M aniline is titrated with 0.125 M HCl.
(i) Write a net ionic equation to show the reaction that takes place between aniline and HCl.
(ii) Calculate the volume of HCl needed to reach the equivalence point.
(iii) Calculate the pH of the titration mixture after 10.00 mL of HCl has been added.
(iv) Calculate the pH of the titration mixture at the equivalence point.

(a) $C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$

(b) $K_b = \dfrac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]}$

(c) $pH = 8.82$ so $[OH^-] = 6.61\times10^{-6}$ M (this is x)

$$K_b = \frac{(6.61\times10^{-6})(6.61\times10^{-6})}{0.10 - 6.61\times10^{-6}} \approx 0.10$$ (You could ignore x)

$K_b = 4.4\times10^{-10}$

(d) (i) $H^+ + C_6H_5NH_2 \rightarrow C_6H_5NH_3^+$ (OR $H_3O^+ + C_6H_5NH_2 \rightarrow C_6H_5NH_3^+ + H_2O$)

(ii) $(0.05000\text{L})(0.10\text{M}) = 0.0050$ moles $C_6H_5NH_2 \times \frac{1\ H^+}{1\ C_6H_5NH_2} = 0.0050$ moles H^+ needed

$0.125\text{M} = \frac{0.0050\text{ mol}}{\text{vol}}$ vol = 40 mL

$K_a = 2.3\times10^{-5}$

(iii) $H^+ + C_6H_5NH_2 \rightarrow C_6H_5NH_3^+$

	H^+	$C_6H_5NH_2$	$C_6H_5NH_3^+$
10 mL of HCl →	0.00125	0.0050	0
	−0.00125	−0.00125	+0.00125
	0	0.00375 (A^-)	0.00125 (HA) ← BUFFER

OR $pH = pK_a + \log\frac{A^-}{HA}$ OR $pOH = pK_b + \log\frac{HA}{A^-}$

pH = 5.12

(iv) At end point, solution is 0.0050 moles $C_6H_5NH_3^+$ in 90 mL

Weak acid $[C_6H_5NH_3^+] = 0.0556$ M $K_a = 2.3\times10^{-5}$ $x = [H^+] = 2.7\times10^{-4}$ M

pH = 3.6