AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART F, Day 93:

- Titrations
- pH curve
- Indicators

1)

$$HC_7H_5O_2 + H_2O \rightleftharpoons H_3O^{+1} + C_7H_5O_2^{-1}$$

Benzoic acid, a weak acid, reacts with water according to the reaction represented above.

(a) Write the equilibrium expression, K_a, for the reaction above.

$$K_a = \frac{\left[\ C_7H_5O_2^{-1}\ \right]\left[\ H_3O^{+1}\ \right]}{\left[\ HC_7H_5O_2\ \right]}$$

(b) A sample of benzoic acid is dissolved in water to produce 55.0 mL of a 0.250 *M* solution. The pH of the solution is 2.40. Calculate the equilibrium constant, K_a, for this reaction.

$$pH = 2.40$$

$$[\ H_3O^{+1}\ ] = [\ H^{+1}\ ] = \text{antilog}\ (\ -2.40\ ) = 10^{-2.40} = 3.98 \times 10^{-3}\ M$$

$$[\ H_3O^{+1}\ ] = [\ C_7H_5O_2^{-1}\ ] = 3.98 \times 10^{-3}\ M$$

	$[\ HC_7H_5O_2\ ]$	+	HOH	$\rightleftharpoons$	$[\ C_7H_5O_2^{-1}\ ]$	+	$[\ H_3O^{+1}\ ]$
I	0.250 *M*		-		0		0
C	$0.250 - 3.98 \times 10^{-3}$		-		$+ 3.98 \times 10^{-3}$		$+ 3.98 \times 10^{-3}$
E	0.24602		-		3.98×10^{-3}		3.98×10^{-3}

$$K_a = \frac{\left[\ C_7H_5O_2^{-1}\ \right]\left[\ H_3O^{+1}\ \right]}{\left[\ HC_7H_5O_2\ \right]} = \frac{\left(\ 3.98 \times 10^{-3}\ \right)^2}{0.24602} = 6.44 \times 10^{-5}$$

(c) The solution prepared in part (b) is titrated with 0.10 *M* NaOH. Calculate the pH of the solution when 5.0 mL of the base has been added.

At this moment, we do NOT KNOW if we are at the equivalence point …

INITIAL Number of moles:

Number of moles ($HC_7H_5O_2$): $= \dfrac{55\ mL}{} \times \dfrac{1\ L}{1000\ mL} \times \dfrac{0.25\ mol}{L} \ = \ 0.01375\ mol\ HC_7H_5O_2$

Number of moles (NaOH): $= \dfrac{5\ mL}{} \times \dfrac{1\ L}{1000\ mL} \times \dfrac{0.10\ mol}{L} \ = \ 0.00050\ mol\ NaOH$

Total volume of the titration up to this point.

55 mL + 5 mL = 60 mL formed

$[\ HC_7H_5O_2\ ] \ = \ \dfrac{0.01375\ mol}{0.060\ L} = 0.229\ M$ **;** $[\ OH^{-1}\ ] \ = \ \dfrac{0.00050\ mol}{0.060\ L} = 0.00833\ M$

	$[\ C_7H_5O_2^{-1}\ ]$	+	$[\ H_2O\ ]$	$\rightleftharpoons$	$[\ HC_7H_5O_2\ ]$	+	$[\ OH^{-1}\ ]$
I	~ 0		-		0.229		0.00833
C	+ 0.00833		-		- 0.00833		- 0.00833
E	0.00833		-		0.221		~ 0

$$K_b \ = \ \frac{K_w}{K_a} \ = \ \frac{1.00\times10^{-14}}{6.44\times10^{-5}} \ = \ 1.55\times10^{-10}$$

$$pH \ = \ pK_a \ + \ \log\frac{[\ A^{-1}\ ]}{[\ HA\ ]} \ \ \mathbf{;} \ \ pH \ = \ \left(-\log\ 6.44\times10^{-5}\right) \ + \ \log\frac{[\ 0.00833\ ]}{[\ 0.221\ ]} \ = \ 2.77$$

<u>OR</u>

$$K_b \ = \ 1.55\times10^{-10} \ = \ \frac{[\ HC_7H_5O_2\ ][\ OH^{-1}\ ]}{[\ C_7H_5O_2^{-1}\ ]} \ = \ \frac{(0.221)(\ y\ )}{(0.00833)}$$

$$y \ = \ \frac{(\ 0.00833\ )(\ 1.55\times10^{-10}\ )}{0.221} \ = \ 5.84\times10^{-12} \ = \ [\ OH^{-1}\ ]$$

pOH = - log (6.06 x 10^{-12}) = 11.23

pH = 14 - pOH = 14 - 11.23 = **2.77**

(d) Calculate the pH at the equivalence point of the titration in part (c).

At the equivalence point, MOLES of $HC_7H_5O_2$ (acid) = MOLES of OH^{-1} (base).

$$\frac{0.01375\ mol}{} \times \frac{1\ L}{0.100\ mol} \times \frac{1000\ mL}{L} = 137.5\ mL\ NaOH\ \ needed$$

Need 137.5 mL of 0.10 M NaOH to reach the equivalence point of this titration.

The total volume of the solution is:

Total Volume: 137.5 mL (NaOH) + 55 mL ($HC_7H_5O_2$) = ***192.5 mL***

$$[\ HC_7H_5O_2\ ] = \frac{0.01375\ mol}{0.1925\ L} = 0.0714\ M \quad ; \quad [\ OH^{-1}\ ] = \frac{0.01375\ mol}{0.1925\ L} = 0.0714\ M$$

	[$C_7H_5O_2^{-1}$]	+	HOH	$\rightleftharpoons$	[$HC_7H_5O_2$]	+	[OH^{-1}]
I	~ 0		-		0.0714		0.0714
C	+ 0.0714		-		- 0.0714		- 0.0714
E	0.0714		-		~ 0		~ 0
E	0.0714		-		***y***		***y***

Realize that when one reaches the equivalence point (in the titration), the moles of the acid equal the moles of the base, the number of moles of the acid and base go to zero, and THEN equilibrium is re-established. You will solve for "y" or the concentration of the hydroxide ions (or hydrogen ions depending on the question) at equilibrium.

$$K_b = \frac{\left[\ HC_7H_5O_2\ \right]\left[\ OH^{-1}\ \right]}{\left[\ C_7H_5O_2^{-1}\ \right]} = 1.58\times10^{-10} = \frac{y^2}{0.0714}$$

$$y^2 = (\ 1.58 \text{ x } 10^{-10}\ )\ (\ 0.00714\ ) = 1.13 \text{ x } 10^{-11}$$

$$y = \sqrt{1.13\times10^{-11}} = 3.36\times10^{-6}\ M = \left[\ OH^{-1}\ \right]$$

$$\text{pOH} = -\log\ (\ 3.36 \text{ x } 10^{-6}\ ) = 5.47$$

$$\textbf{pH} = 14 - \text{pOH} = 14 - 5.47 = \mathbf{8.53}$$

(e) The pK_a values for several indicators are given below. Which of the indicators listed is most suitable for this titration? Justify your answer.

Indicator	pK_a
Thymol Blue	1.75
Bromcresol Purple	6.25
m- Nitrophenol	8.50

The pH at the equivalence point is basic. The best indicator is m- Nitrophenol, for which the value of pK_a is closest to the pH at the equivalence point.