~~80~~ 36

FR Practice #2 (2012 #1, shortened, 7 points)

2. A 1.22 g sample of a pure monoprotic acid, HA, was dissolved in distilled water. The HA solution was then titrated with 0.250 *M* NaOH. The pH was measured throughout the titration, and the equivalence point was reached when 40.0 mL of the NaOH solution had been added. The data from the titration are recorded in the table below.

Volume of 0.250 *M* NaOH Added (mL)	pH of Titrated Solution
0.00	?
10.0	3.72
20.0	4.20
30.0	?
40.0	8.62
50.0	12.40

$\frac{1}{2}$ eq. pt → 20.0

eq. pt → 40.0

a. Explain how the data in the table above provide evidence that HA is a weak acid rather than a strong acid. (1 point)

b. Write the balanced net-ionic equation for the reaction that occurs when the solution of NaOH is added to the solution of HA. (1 point)

c. Calculate the number of moles of HA that were titrated. (1 point)

d. Calculate the molar mass of HA. (1 point)

The equation for the dissociation reaction of HA in water is shown below.

$$HA(aq) + H_2O(l) \leftrightarrow A^-(aq) + H_3O^+(aq)$$

e. Assume that the initial concentration of the HA solution (before any NaOH solution was added) is 0.200 *M*. Determine the pH of the initial HA solution. (3 points)

a.) The pH @ eq. pt. > 7, so HA must be a weak acid (if HA were a strong acid being titrated w/ NaOH, a strong base, the pH @ eq. pt. would be 7).

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

c.) moles HA = moles base @ eq. pt = $M_b V_b$ = 0.250 M × 0.0400 L = **0.0100 mol HA**

d.) MM (HA) = $\frac{1.22 g}{0.0100 mol}$ = **122 g/mol** e.) $K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{x^2}{0.200-x} \approx \frac{x^2}{0.200}$

HA is a weak acid,
$\Rightarrow K_a \ll 1$, x negligible

@ ½ eq. pt, pH = 4.20 = pK_a (of HA)
$\Rightarrow K_a = 10^{-4.20} = 6.3 \times 10^{-5}$

$\frac{x^2}{0.200} = 6.3E{-5} \Rightarrow x = \sqrt{(0.200)(6.3E{-5})} = 3.5E{-3}\,M = [H_3O^+]$

$pH = -\log[H_3O^+] = -\log(3.5E{-3}) = \boxed{2.45}$