Suppose one begins with 50.0-mL a 0.100 M solution of nitrous acid. The Ka = 4.5 x 10
What is the pH after (a) 0.0 mL of 0.125 M NaOH is added? (b) After 10.0 mL? (c) After 17.0 mL? (d) After 20.0 mL? (e) After 39.0 mL? (f) After 40.0 mL? (g) After 45.0 mL?

(a) Adding 0.0 mL of 0.125 M NaOH: Calculate the pH using the ICE solution.

	HN02(aq) + H20(I) e	NO/(aq)	1-130 (aq) Ka = 4.5×10^4
Initial conc.	0.100 M	0.000 M	0.000 M
A conc.	—x		
Eq. conc.	0.100-x	x	x

Ka = $\frac{[NO2-][H3O+]}{[HN02J]}$ ➔4.5 X 104 $= \frac{(x)(x)}{(0.100-x)}$ ➔ $X = \sqrt{(4.5x10-4)(0.100-x)}$

$X_1 = 6.71 \times 10^{-3}$ M
$X_2 = 6.48 \times 10^{-3}$ M
$X_3 = 6.49 \times 10^{-3}$ M

X = 0-130+1 = 6.5 x 10-3 M pH = -log [H30+] = -log 6.5 x 10-3 M = 2 19

(b) After initial calculation and before the equivalence point, the system has a buffer.
Initial number of moles of HN02.• 0.100 M (0.050 L) = 0.0050 moles acid
Adding 10.0 of 0.125 M NaOH: 0.125 M (0.010 L) = 0.00125 mol base

pH = pKa + log $\frac{(c.base)}{(acid)}$ = 3.35 + log $\frac{(0+0.00125)}{(0.0050-0.00125)}$ = 3.35 + log $\frac{• 0.00125}{0.00375}$ = 3.35 + (-0.48) = 2.87

(c) After adding 17.0 rnL of 0.125 M NaOH: 0.125 M (0.017 L) = 0.002125 moles base

pH = 3.35 + 16g $\frac{(0.002125)}{(0.0050-0.002125)}$ = 3.35 + log $\frac{0.002125}{0.002875}$ = 3.35 + 0.13) = 3.22

(d) After adding 20.0 rnL of 0.125 M NaOH: 0.125 M (0.020 L) = 0.0025 moles base

$$pH = 3.35 + \log\frac{(0.0025)}{(0.0050-0.0025)} = 3.35 + \log\frac{0.0025}{0.0025} = 3.35 - 0.00 = \underline{3.35}$$

We are halfway to the equivalence point, so pH = pKa

(e) After adding 39.0 of 0.125 M NaOH: 0.125 M (0.0390 L) = 0.004875 moles base

$$pH = 3.35 + \log\frac{(0.004875)}{(0.0050-0.004875)} = 3.35 + \log\frac{0.004875}{0.000125} = 3.35 + 1.59 = \underline{4.94}$$

(f) After adding 40.0 of 0.125 M NaOH: 0.125 M (0.0400 L) = 0.0050 moles base Moles of acid = moles of NaOH added. We have reached the equivalence point!

Do an ICE procedure using the Kb of the c. base of NH02. Namely NO-z- .N02 - (aq) + H20(I) e HN02(aq) + $OH^-_{(aq)}$ Kb = $\frac{...}{Ka}$ = 2.2 x 1

Initial conc.	0.0050 mol/0.090L	0.000 M	0.000 M
conc.	—x		
Eq. conc.	0.0556-x	x	x

$$Kb = \frac{[HNO2][OH-]}{[NO2-]} \rightarrow 2.2 \times 10^{-11} = \frac{(x)(x)}{(0.0556-x)} \rightarrow X = \sqrt{(2.2x10-11)(0.0556-x)}$$

$X_1 = 1.11 \times 10^{-6}$ M

$X_2 = 1.11 \times 10^{-6}$ M

$X = [OH^-] = 1.1 \times 10^{-6}$ M **pOH = -log [OH⁻] = 5.96** **pH = 14.00-5.96 = 8.04**

(g) With the addition of 45.0 mL, we have passed the equivalence point and the only species of any importance is the strong base, NaOH. Remember, at the eq. point, no NaOH exists. It was neutralized into water by the hydronium ion. Soooooo.....

0.125 M (0.005 L) = 0.000625 moles of OW / 0.095 0.00658 M = [Off]

Excess NaOH after eq. pt. total volume

POH = 2.18, pH = 11.82 <u>YOU ARE FINISHED!!!</u>