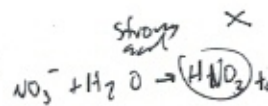
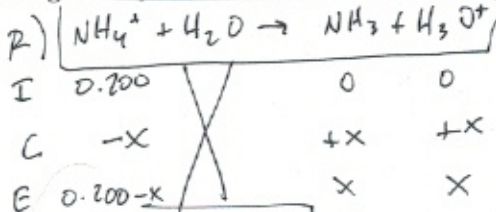
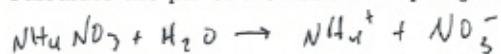


Salts, Buffers, and Titrations Practice

1. Calculate the pH of a 0.200 M NH_4NO_3 solution given $K_b = 1.8 \times 10^{-5}$.



$$\frac{K_w}{K_b} = K_a = \frac{x^2}{0.200 - x}$$

$$\frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{x^2}{0.200 - x}$$

$$x = 1.05 \times 10^{-5}, \quad \text{pH} = 4.98$$

2. For the solution created using 20 mL of 0.10 M HF and 0.050 g of NaF, calculate the pH assuming the K_a of HF is 3.5×10^{-4} .

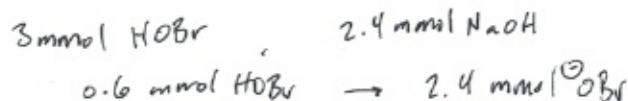
$$\text{MW}_{\text{NaF}} = 41.988173 \text{ g/mol} \quad 0.050 \text{ g NaF} \left| \frac{1 \text{ mol NaF}}{41.988173 \text{ g}} \right| = 0.00119 \text{ mol NaF} = 1.19 \text{ mmol NaF}$$

$$20 \text{ mL } (0.10 \text{ M}) = 2 \text{ mmol HF}$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{1.19}{2}\right) = 3.23$$

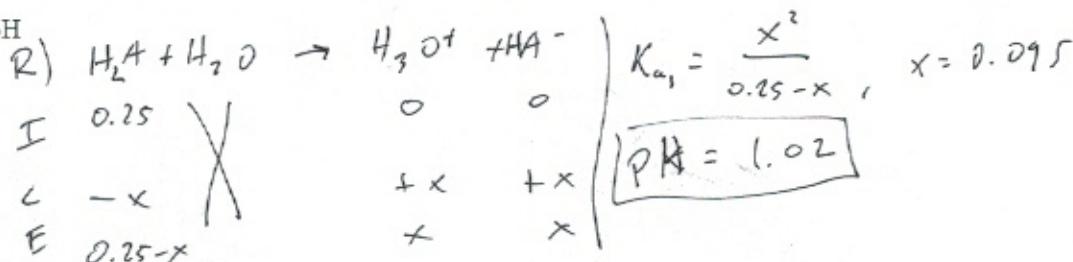
3. Calculate the pH of the solution resulting from titrating 30 mL of 0.10 M HOBr with 24 mL of 0.10 M NaOH assuming the K_a of HOBr is 2.0×10^{-9} .

$$\text{pH} = \text{p}K_a + \log\left(\frac{2.4}{2.6}\right) = 9.30$$



4. Sketch a titration curve for the titration of 30 mL of the 0.25 M diprotic acid HA with 0.25 M NaOH by calculating the following ($K_{a1} = 5.9 \times 10^{-2}$ and $K_{a2} = 6.4 \times 10^{-5}$):

- (a) The initial pH



- (b) The pH of both equivalence points

$$\text{p}K_{a1} = 1.23, \quad \text{p}K_{a2} = 4.19$$

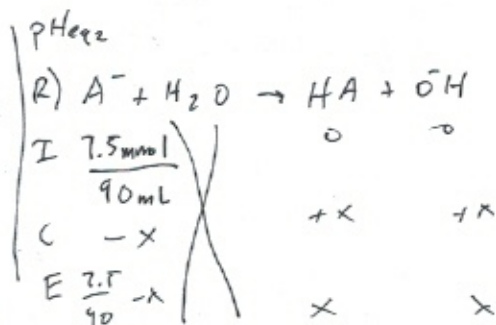
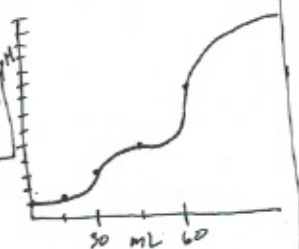
$$\text{pH}_{\text{eq}} = \frac{\text{p}K_{a1} + \text{p}K_{a2}}{2} = \frac{1.23 + 4.19}{2} = 2.71$$

$$\text{pH} = 2.71$$

- (c) The pH of both half-equivalence point

$$\text{half-equivalence 1} = 1.23$$

$$\text{half-equivalence 2} = 4.19$$



$$K_{a2} = \frac{x^2}{7.5 - x}, \quad x = 3.61 \times 10^{-6}$$

$$\text{pOH} = 5.44$$

$$\text{pH} = 14 - 5.44 = 8.56$$

$$\text{pH}_2 = 8.56$$