

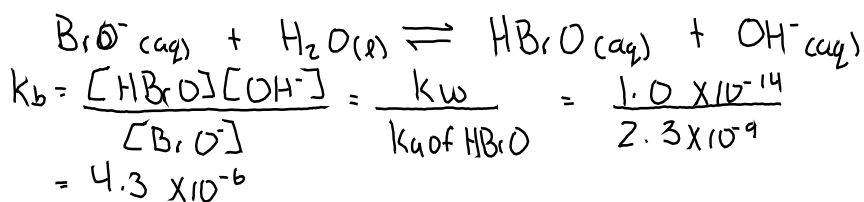
c) at equivalence point \Rightarrow moles acid = moles base

$$[\text{BrO}^-] = \frac{\text{moles BrO}^-}{\text{total volume}} = \frac{\text{moles (HBrO)}_{\text{initial}}}{20 \text{ mL} + \text{Volume used for titration}}$$

- how many mL NaOH needed to reach equivalence?
 $0.2000 \text{ M HBrO} \times 0.0200 \text{ L} = 4.00 \times 10^{-3} \text{ moles}$
 moles acid = moles base = $4.00 \times 10^{-3} \text{ moles NaOH}$
 $\frac{4.00 \times 10^{-3} \text{ moles}}{0.1000 \text{ M}} = 40.00 \text{ mL}$

$$[\text{BrO}^-] = \frac{4.00 \times 10^{-3} \text{ moles}}{(0.0200 \text{ L}) + (0.0400 \text{ L})} = \frac{4.00 \times 10^{-3} \text{ moles}}{0.0600 \text{ L}} = 0.06667 \text{ M}$$

- We need $[\text{BrO}^-]$ here because all HBrO has now been converted to BrO^- so the pH is based on the equilibrium association of BrO^- w/ H_2O to form HBrO + OH^-
- The pH will be basic since we are now in a basic soln.



$X = [\text{HBrO}] = [\text{OH}^-] \Rightarrow x^2$ + solve for $[\text{OH}^-]$, pOH, & then convert to pH
 OR:

convert back to acidic in one step

$$[\text{H}_3\text{O}^+] = \frac{K_w}{\sqrt{K_b \times [\text{BrO}^-]}} = \frac{1.0 \times 10^{-14}}{\sqrt{(4.3 \times 10^{-6})(0.06667)}} = 1.9 \times 10^{-11} \text{ M}$$

$$\text{pH} = 10.72$$