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

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

- how many mL NaOH needed to reach equivalence?

$0.2000\ M\ HBrO \times 0.0200L = 4.00 \times 10^{-3}$ moles

moles acid = moles base = $4.00 \times 10^{-3}$ moles NaOH

$$\frac{4.00 \times 10^{-3}\text{ moles}}{0.1000\ M} = 40.00\text{ mL}$$

$$[BrO^-] = \frac{4.00 \times 10^{-3}\text{ moles}}{(0.02000L)(0.04000L)} = \frac{4.00 \times 10^{-3}\text{ moles}}{0.06000\ L} = 0.06667\ M$$

- We need $[BrO^-]$ 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 <u>basic</u> since we are now in a basic soln.

$$BrO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HBrO_{(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{[HBrO][OH^-]}{[BrO^-]} = \frac{K_w}{K_a \text{ of } HBrO} = \frac{1.0 \times 10^{-14}}{2.3 \times 10^{-9}}$$

$= 4.3 \times 10^{-6}$

$X = [HBrO] = [OH^-] \Rightarrow x^2$ & solve for $[OH^-]$, pOH, & then convert to pH

OR:

convert back to acidic in one step

$$[H_3O^+] = \frac{K_w}{\sqrt{K_b \times [BrO^-]}} = \frac{1.0 \times 10^{-14}}{\sqrt{(4.3 \times 10^{-6})(0.06667)}} = 1.9 \times 10^{-11} M$$

pH = 10.72