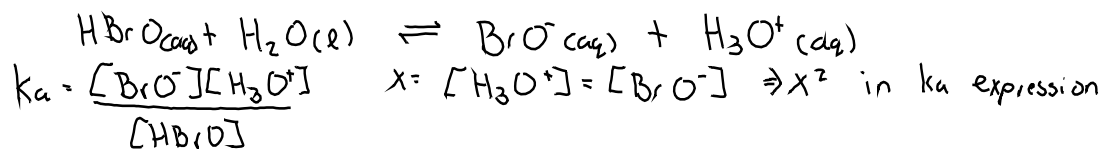


Weak acid - Strong Base

20.00 mL of 0.2000 M HBrO is titrated with 0.1000 M NaOH.
What is pH:

- before any base is added
 - when $[HBrO] = [BrO^-]$
 - at the equivalence point
 - when moles of OH^- added is 2x the moles of HBrO initially present
- K_a of HBrO = 2.3×10^{-9}
- Also, sketch the titration curve

a) Before base:



$$K_a = \frac{x^2}{[HBrO]} \quad \therefore [H_3O^+] = \sqrt{[HBrO] K_a} = \sqrt{(0.2000)(2.3 \times 10^{-9})}$$

$$= 2.1 \times 10^{-5} \text{ M}$$

$$pH = -\log [H_3O^+] = -\log (2.1 \times 10^{-5}) = 4.68$$

$$pH = 4.68$$

b) $[HBrO] = [BrO^-]$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$\Rightarrow [H_3O^+] = K_a \times \frac{[HBrO]}{[BrO^-]}$$

\swarrow HA
 \nwarrow A⁻

$$[H_3O^+] = 2.3 \times 10^{-9} \times \frac{[HBrO]}{[BrO^-]} = 2.3 \times 10^{-9} \times 1$$

$$= 2.3 \times 10^{-9} \text{ M}$$

$$pH = -\log (2.3 \times 10^{-9}) = 8.64$$

$$pH = 8.64$$

Since $[HBrO] = [BrO^-]$
the ratio is 1 regardless
of the actual numbers
for concentration