10-2. Given 100.0 mL of 0.100 M NaOH titrated with 1.00 M HBr. ⇒ strong acid - strong base titration

Calc. pH at V_a = 5.00, 9.90 and 12.0 mL

Work:

Calc. V_e first

$$V_e = \frac{M_b V_b}{M_a} = \frac{(0.100\ M)(100.0\ mL)}{1.00\ M} = 10.0\ mL$$

Thus, V_a = 5.00 and 9.90 mL < V_e (before the equiv. pt.)

(a) V_a = 5.00 mL } mol H^+ = $(1.00\ mol/L)(5.00 \times 10^{-3}\ L) = 5.00 \times 10^{-3}$ mol

	H^+	+ OH^-	⟶ H_2O
mol Initial	0.005	0.0100	Liquid
Change	−0.005	−0.005	
Final	0	0.005	

$$[OH^-]_{unreacted} = \frac{0.005\ mol}{(0.105\ L)} \leftarrow \begin{matrix} 100\ mL + 5\ mL \\ OH^- \quad H^+ \end{matrix}$$

$$= 0.04762\ M$$

pOH = 1.32 (pH = 12.68)

(b) V_a = 9.90 mL ⇒ mol H^+ = 9.90×10^{-3}

mol unreacted OH^- = $0.01 - 9.90 \times 10^{-3} = 1 \times 10^{-4}$

$$[OH^-] = \frac{1 \times 10^{-4}\ mol}{0.1099\ L} = 9.099 \times 10^{-4}\ M$$

pOH = 3.04 ; (pH = 10.96)

(c) V_a = 12.0 mL ⇒ beyond the equiv. pt., so excess mol H^+ determines pH ; mol H^+ added = 1.2×10^{-2}

mol H^+ in excess = $1.2 \times 10^{-2} - 0.010 = 2 \times 10^{-3}$

$$[H^+]_{\text{in excess}} = \frac{2 \times 10^{-3}\ mol}{0.112\ L} = 0.01786 \Rightarrow (pH = 1.75)$$