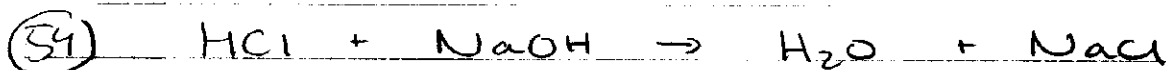


Chapter 1a



flask buret
50.00 mL
0.1000 M 0.1000 M

The first thing you need to do is to calculate how many mL of NaOH you need to add to reach the equivalence point of the titration.
@ equivalence moles HCl = moles NaOH

$$C_{\text{HCl}} V_{\text{HCl}} = C_{\text{NaOH}} V_{\text{NaOH}}$$

$$V_{\text{NaOH}} = \frac{C_{\text{HCl}} V_{\text{HCl}}}{C_{\text{NaOH}}} = \frac{(0.1000 \text{ M})(50.00 \text{ mL})}{0.1000 \text{ M}} = \boxed{50.0 \text{ mL}}$$

So when you've added $< 50.0 \text{ mL}$ of NaOH, there will still be HCl in the flask - and you will therefore have a soln of a strong acid.

When you've added $> 50.0 \text{ mL}$ of NaOH, all the HCl will be gone and you'll have excess NaOH. At that point you will have a soln of a strong base.

(a) 0 mL of NaOH added

What's in the flask? HCl + water

Since HCl is a strong acid $[\text{H}_3\text{O}^+] = [\text{HCl}] = 0.1000 \text{ M}$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.1000) = \boxed{1.0000}$$

(b) 25.0 mL of NaOH added

What's in the flask? Since the NaOH completely reacts with the HCl, it will all be gone & you will have the unreacted HCl, NaCl + H₂O.

Na⁺ and Cl⁻ ions are neutral so they don't affect the pH of soln.

$$[\text{H}_3\text{O}^+] = [\text{HCl}]_{\text{left}} = \frac{\text{moles HCl left}}{\text{tot V}}$$

(54) (b) cont.

From the stoichiometry, each mole of NaOH added reacts w/ 1 mole of HCl.

$$\begin{aligned} \frac{\text{moles HCl left}}{\text{total V}} &= \frac{\text{moles HCl}_{\text{orig}} - \text{moles NaOH added}}{\text{total V}} \\ &= \frac{C_{\text{HCl, orig}} V_{\text{HCl, orig}} - C_{\text{NaOH}} V_{\text{NaOH added}}}{\text{total V}} \\ &= \frac{(0.1000 \text{ M})(50.00 \text{ mL}) - (0.1000 \text{ M})(25.00 \text{ mL})}{75.00 \text{ mL}} \end{aligned}$$

total volume = orig 50.00 mL + 25.00 mL added

$$= 0.03333 \text{ M} = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log(0.03333) = \boxed{1.4771}$$

(c) 49.00 mL of NaOH added

Whats in the flask? Still HCl, NaCl + H₂O

$$\begin{aligned} [\text{H}_3\text{O}^+] = [\text{HCl}]_{\text{left}} &= \frac{C_{\text{HCl, orig}} V_{\text{HCl, orig}} - C_{\text{NaOH}} V_{\text{NaOH added}}}{\text{tot V}} \\ &= \frac{(0.1000 \text{ M})(50.00 \text{ mL}) - (0.1000 \text{ M})(49.00 \text{ mL})}{99.00 \text{ mL}} \end{aligned}$$

$$= 1.010 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log(1.010 \times 10^{-3}) = \boxed{2.9956}$$

(d) 49.90 mL NaOH added

Whats in the flask? Still below the equiv. point so still HCl, NaCl + H₂O

$$[\text{H}_3\text{O}^+] = [\text{HCl}]_{\text{left}} = \frac{(0.1000 \text{ M})(50.00 \text{ mL}) - (0.1000 \text{ M})(49.90 \text{ mL})}{99.90 \text{ mL}}$$

$$= 1.001 \times 10^{-4} \text{ M} = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log(1.001 \times 10^{-4}) = \boxed{4.0000}$$

(54) (e) + 50.00 mL NaOH

Whats in the flask? Reached the equivalence point - so we've added just enough NaOH to react with all the HCl and we've left with NaCl + H₂O - Since Na⁺ + Cl⁻ are neutral ions, pH = 7.0000

(f) + 50.10 mL NaOH

Whats in the flask? Now have gone beyond the equiv. pt. There is no more HCl left for the NaOH to react with so it just stays in the flask & we have a soln of NaOH, NaCl + H₂O. Since NaOH is a strong base it will now determine the pH -

$$\begin{aligned} [\text{OH}^-] &= [\text{NaOH}]_{\text{excess}} = \frac{\text{moles NaOH excess}}{\text{tot } V} \\ &= \frac{C_{\text{NaOH}} V_{\text{NaOH in excess}}}{\text{tot } V} \quad \leftarrow \text{this will be the volume added past the equiv pt.} \\ &= \frac{(0.1000 \text{ M})(0.10 \text{ mL})}{100.1 \text{ mL}} = 9.990 \times 10^{-5} \text{ M} = [\text{OH}^-] \end{aligned}$$

need to convert [OH⁻] to [H₃O⁺] to calculate pH

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{9.990 \times 10^{-5}} = 1.009 \times 10^{-10} \text{ M}$$

$$\text{pH} = -\log(1.009 \times 10^{-10}) = \boxed{10.00}$$

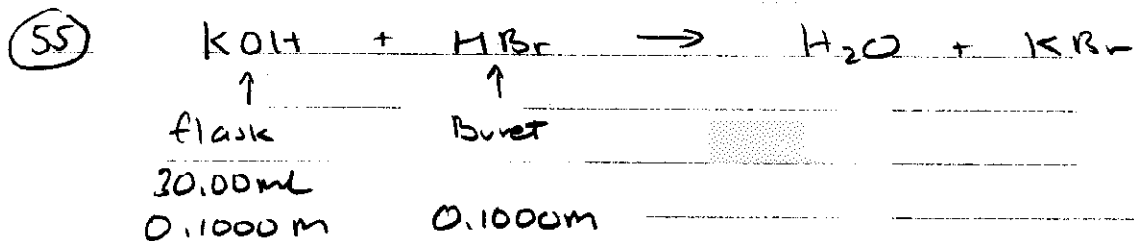
(g) + 60.00 mL NaOH

Whats in the flask? Still NaOH, NaCl + H₂O

$$[\text{OH}^-] = [\text{NaOH}]_{\text{excess}} = \frac{(0.1000 \text{ M})(10.00 \text{ mL})}{110 \text{ mL}} = 9.091 \times 10^{-3} \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{9.091 \times 10^{-3}} = 1.109 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log(1.109 \times 10^{-12}) = \boxed{11.9552}$$



As with previous problem, first need to calculate the equivalence pt.

@ equivalence moles KOH = moles HBr

$$C_{\text{KOH}} V_{\text{KOH}} = C_{\text{HBr}} V_{\text{HBr}}$$

$$V_{\text{HBr}} = \frac{C_{\text{KOH}} V_{\text{KOH}}}{C_{\text{HBr}}} = \frac{(0.1000 \text{ M})(30.00 \text{ mL})}{(0.1000 \text{ M})}$$

$$= \boxed{30.00 \text{ mL}}$$

So when you've added < 30.00 mL of HBr there will still be KOH left and you will have a soln of a strong base

When you've added > 30.00 mL of HBr, all the KOH will be gone and you will have excess HBr. This will make a soln of a strong acid.

(a) 0 mL of HBr added

Whats in the flask? $\text{KOH} = \text{H}_2\text{O}$

Since KOH is a strong base $[\text{OH}^-] = [\text{KOH}] = 0.1000 \text{ M}$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{0.1000 \text{ M}} = 1.008 \times 10^{-13}$$

$$\text{pH} = -\log(1.008 \times 10^{-13})$$

$$= \boxed{12.9965}$$

(b) 15.00 mL of HBr added

Whats in the flask? Since HBr completely reacts with KOH, its all gone. So just have KOH remaining, $\text{KBr} = \text{H}_2\text{O}$

$$[\text{OH}^-] = [\text{KOH}]_{\text{left}} = \frac{\text{moles KOH left}}{\text{tot } V}$$

(55) (b) cont.

each mole of HBr added reacts w/ 1 mole of KOH

$$\begin{aligned} \frac{\text{moles KOH left}}{\text{total V}} &= \frac{\text{moles KOH orig} - \text{moles HBr added}}{\text{total V}} \\ &= \frac{C_{\text{KOH orig}} V_{\text{KOH, orig}} - C_{\text{HBr}} V_{\text{HBr added}}}{\text{total V}} \\ &= \frac{(0.1000 \text{ M})(30.00 \text{ mL}) - (0.1000 \text{ M})(15.00 \text{ mL})}{45.00 \text{ mL}} \\ &= 0.03333 \text{ M} = [\text{OH}^-] \quad [\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{0.03333 \text{ M}} = 3.0243 \times 10^{-13} \\ \text{pH} &= -\log [\text{H}_3\text{O}^+] = -\log (3.0243 \times 10^{-13}) = \boxed{12.5194} \end{aligned}$$

(c) 29.00 mL HBr added

What's in flask? Still KOH, KBr, H₂O

$$\begin{aligned} [\text{OH}^-] &= [\text{KOH}]_{\text{left}} = \frac{C_{\text{KOH, orig}} V_{\text{KOH, orig}} - C_{\text{HBr}} V_{\text{HBr added}}}{\text{total V}} \\ &= \frac{(0.1000 \text{ M})(30.00 \text{ mL}) - (0.1000 \text{ M})(29.00 \text{ mL})}{59.00 \text{ mL}} \\ &= 1.695 \times 10^{-3} \text{ M} = [\text{OH}^-] \quad [\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{1.695 \times 10^{-3}} = 5.947 \times 10^{-12} \\ \text{pH} &= -\log (5.947 \times 10^{-12}) = \boxed{11.2257} \end{aligned}$$

(d) 29.90 mL HBr added

What's in the flask? Still KOH, KBr, H₂O

$$\begin{aligned} [\text{OH}^-] &= [\text{KOH}]_{\text{left}} = \\ &= \frac{(0.1000 \text{ M})(30.00 \text{ mL}) - (0.1000 \text{ M})(29.90 \text{ mL})}{59.90 \text{ mL}} = 1.669 \times 10^{-4} \text{ M} = [\text{OH}^-] \\ [\text{H}_3\text{O}^+] &= \frac{1.008 \times 10^{-14}}{1.669 \times 10^{-4}} = 2.228 \times 10^{-11} \text{ M} \\ \text{pH} &= -\log (2.228 \times 10^{-11}) = \boxed{10.2190} \end{aligned}$$

(e) 30.00 mL HBr added

Whats in the flask? This is the equiv. pt
So have added just enough HBr to react with
all the KOH & now have a soln of
KBr in H₂O. Since K⁺ & Br⁻ are neutral
ions, pH = 7.0000

(f) 30.10 mL HBr added

Whats in the flask? Now have gone
beyond the equiv pt so will have excess
HBr along with KBr & H₂O.
Since HBr is a strong acid

$$\begin{aligned} [\text{H}_3\text{O}^+] &= [\text{HBr}]_{\text{excess}} = \frac{\text{mols HBr excess}}{\text{tot } V} \\ &= \frac{C_{\text{HBr}} V_{\text{HBr excess}}}{\text{tot } V} = \frac{(0.1000 \text{ M})(0.10 \text{ mL})}{60.10 \text{ mL}} \\ &= 1.664 \times 10^{-4} = [\text{H}_3\text{O}^+] \end{aligned}$$

$$\text{pH} = -\log(1.664 \times 10^{-4}) = \boxed{3.78}$$

(g) 40.00 mL HBr added

Whats in the flask? HBr, KBr & H₂O

$$\begin{aligned} [\text{H}_3\text{O}^+] &= [\text{HBr}]_{\text{excess}} = \frac{\text{mols HBr excess}}{\text{tot } V} \\ &= \frac{C_{\text{HBr}} V_{\text{HBr excess}}}{\text{tot } V} = \frac{(0.1000 \text{ M})(10.00 \text{ mL})}{(70.00 \text{ mL})} \\ &= 1.429 \times 10^{-2} \text{ M} = [\text{H}_3\text{O}^+] \end{aligned}$$

$$\text{pH} = -\log(1.429 \times 10^{-2}) = \boxed{1.8451}$$

56, cont

Not making the approximation -

$$\frac{x^2}{0.1000-x} = 1.54 \times 10^{-5} \quad x^2 = (1.54 \times 10^{-5})(0.1000-x)$$

$$x^2 = 1.54 \times 10^{-6} - 1.54 \times 10^{-5}x$$

$$x^2 + 1.54 \times 10^{-5}x - 1.54 \times 10^{-6}$$

a b c

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-1.54 \times 10^{-5} \pm \sqrt{(1.54 \times 10^{-5})^2 - 4(1)(-1.54 \times 10^{-6})}}{2}$$

$$= 1.23 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(1.23 \times 10^{-3}) = \boxed{2.910}$$

In this case it probably wasn't worth the extra effort to use the quadratic formula!

(b) + 10.00 mL NaOH

What's in the flask? Before the equivalence point, so some, but not all, of the $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}$ has been converted to the conjugate base.

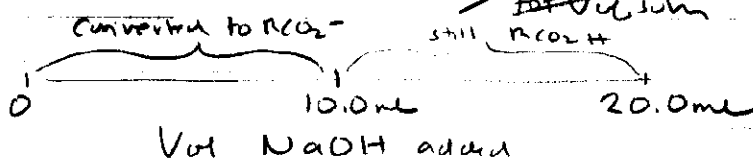
Flask contains $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}$ and $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2^-$ weak acid and its conjugate base - BUFFER SOLN - use Henderson-Hasselbalch equation to calculate pH

$$\text{pH} = \text{pK}_a + \log \frac{[\text{RCO}_2^-]}{[\text{RCO}_2\text{H}]}$$

"CH₃CH₂CH₂CO₂⁻" "CH₃CH₂CH₂CO₂H"

$$\text{pK}_a = -\log K_a = -\log(1.54 \times 10^{-5}) = 4.812$$

$$\frac{[\text{RCO}_2^-]}{[\text{RCO}_2\text{H}]} = \frac{\text{moles RCO}_2^- / \text{tot vol soln}}{\text{moles RCO}_2\text{H} / \text{tot vol soln}} = \frac{\text{moles RCO}_2^-}{\text{moles RCO}_2\text{H}}$$



Since we're 1/2 way to the equivalence pt, 1/2 of the RCO_2H has been converted to RCO_2^- and 1/2 remains as RCO_2H

$$\frac{\text{moles RCO}_2^-}{\text{moles RCO}_2\text{H}} = \frac{1/2 \text{ orig. mol of RCO}_2\text{H}}{1/2 \text{ orig. mol of RCO}_2\text{H}} = 1$$

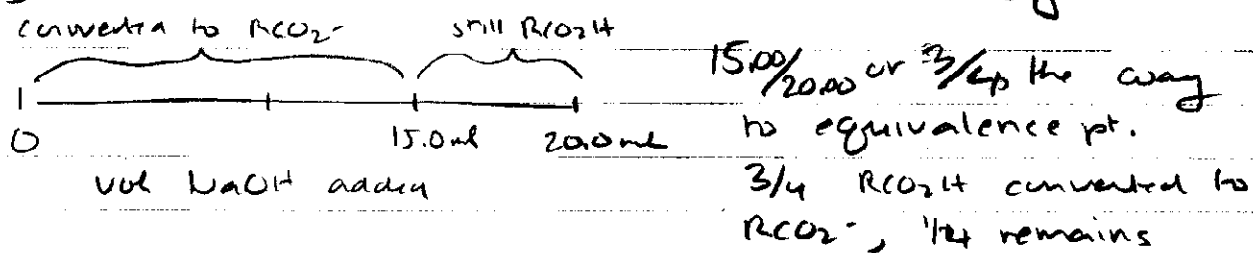
56 cont.

(b) cont.

$$\text{pH} = \text{pK}_a + \log(1) = \text{pK}_a = \boxed{4.812}$$

(c) + 15.00 mL NaOH

What's in the flask? Still before the equiv. pt so still a mix of HCO_2H and HCO_2^- , a buffer soln - so use Henderson-Hasselbalch eqn.



$$\frac{\text{moles } \text{HCO}_2^-}{\text{moles } \text{HCO}_2\text{H}} = \frac{\cancel{3/4} \text{ original } \text{HCO}_2\text{H}}{\cancel{1/4} \text{ original } \text{HCO}_2\text{H}} = 3.00$$

$$\text{pH} = \text{pK}_a + \log \frac{\text{moles } \text{HCO}_2^-}{\text{moles } \text{HCO}_2\text{H}} = 4.812 + \log 3 = \boxed{5.289}$$

(d) + 19.00 mL NaOH

What in the flask? Still before equivalence pt, so still HCO_2H + HCO_2^-

19/20 the way to equivalence, so 19/20 HCO_2H converted to HCO_2^- , and 1/20 remains as HCO_2H

$$\frac{\text{moles } \text{HCO}_2^-}{\text{moles } \text{HCO}_2\text{H}} = \frac{\cancel{19.00/20.00} \text{ original } \text{HCO}_2\text{H}}{\cancel{1/20.00} \text{ original } \text{HCO}_2\text{H}} = 19.00$$

$$\text{pH} = \text{pK}_a + \log \frac{\text{moles } \text{HCO}_2^-}{\text{moles } \text{HCO}_2\text{H}} = 4.812 + \log 19.00 = \boxed{6.091}$$

(e) + 19.95 mL NaOH

What's in the flask? Still before equiv. pt so still RCO_2H and RCO_2^-

$\frac{19.95}{20.00}$ the way to equivalence, so $\frac{19.95}{20.00} \text{RCO}_2\text{H}$ converted to RCO_2^- , and $\frac{0.05}{20.00}$ remains as RCO_2H .

$$\frac{\text{moles } \text{RCO}_2^-}{\text{moles } \text{RCO}_2\text{H}} = \frac{\frac{19.95}{20.00} \text{ original } \text{RCO}_2\text{H}}{\frac{0.05}{20.00} \text{ original } \text{RCO}_2\text{H}} = \frac{19.95}{0.05} = 399$$

$$\text{pH} = \text{pKa} + \log \frac{\text{moles } \text{RCO}_2^-}{\text{moles } \text{RCO}_2\text{H}} = 4.812 + \log 399 = \boxed{7.41}$$

(f) + 20.00 mL NaOH

What's in the flask? Equivalence pt! Have added just enough NaOH to convert all the RCO_2H to RCO_2^- - so have just a soln of RCO_2^- (or $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2^-$). Since this is the conj. base of a weak acid it will be basic. Need to use K_b to calculate pH

$$K_b \text{ of } \text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2^- = \frac{K_w}{K_a \text{ of } \text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}} \\ = \frac{1.008 \times 10^{-14}}{1.54 \times 10^{-5}} = 6.55 \times 10^{-10}$$

Also need to know the conc. of base to calculate the pH.

$$[\text{RCO}_2^-] = \frac{\text{moles } \text{RCO}_2^-}{\text{tot. V}} = \frac{\text{orig moles of } \text{RCO}_2\text{H}}{\text{tot. V}} \leftarrow \text{because all of the } \text{RCO}_2\text{H} \text{ has been converted to } \text{RCO}_2^- \text{ at this point in the titration}$$
$$= \frac{C_{\text{RCO}_2\text{H, orig}} V_{\text{RCO}_2\text{H, orig}}}{\text{tot. V}} = \frac{(0.1000\text{M})(20.00\text{mL})}{(40.00\text{mL})} \\ = 0.05000\text{M}$$

$$K_b = \frac{x^2}{0.05000 - x} \approx \frac{x^2}{0.05000} = 6.55 \times 10^{-10}$$

$$x = [\text{OH}^-] \text{ (Since this is } K_b \text{ and not } K_a)$$

56, cont.

(f) cont.

$$x = \sqrt{(6.55 \times 10^{-9})(0.05000)} = 5.721 \times 10^{-6} = [\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{5.721 \times 10^{-6}} = 1.762 \times 10^{-9} \text{ M}$$

$$\text{pH} = -\log(1.762 \times 10^{-9}) = \boxed{8.754}$$

(g) + 20.05 mL NaOH

Whats in the flask? Past the equivalence pt, so will have excess NaOH along with the BCO_2^- . Since NaOH is a strong base and BCO_2^- is a weak base, the NaOH will control the pH and you can just ignore the BCO_2^- .

$$[\text{OH}^-] = [\text{NaOH}]_{\text{excess}} = \frac{\text{moles NaOH excess}}{\text{tot. V}}$$
$$= \frac{C_{\text{NaOH}} V_{\text{NaOH, excess}}}{\text{tot. V}} = \frac{(0.1000)(0.05 \text{ mL})}{40.05 \text{ mL}}$$

$$= 1.25 \times 10^{-4} = [\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{1.25 \times 10^{-4}} = 8.06 \times 10^{-11}$$

$$\text{pH} = -\log(8.06 \times 10^{-11}) = \boxed{10.1}$$

(h) + 25.00 mL NaOH

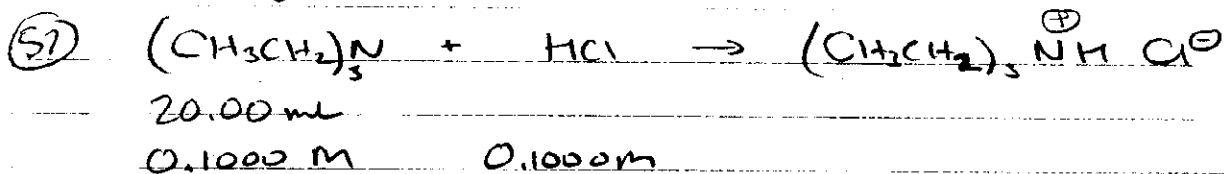
Whats in the flask? Still have NaOH + BCO_2^- , NaOH controls the pH

$$[\text{OH}^-] = [\text{NaOH}]_{\text{excess}} = \frac{\text{moles NaOH excess}}{\text{tot. V}}$$
$$= \frac{C_{\text{NaOH}} V_{\text{NaOH, excess}}}{\text{tot. V}} = \frac{(0.1000 \text{ M})(5.00 \text{ mL})}{45.00 \text{ mL}} = 1.11 \times 10^{-2} \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{1.11 \times 10^{-2}} = 9.07 \times 10^{-13}$$

$$\text{pH} = -\log(9.07 \times 10^{-13}) = \boxed{12.042}$$

Weak base!



c equivalence mols $(\text{CH}_3\text{CH}_2)_3\text{N} = \text{mols HCl}$

$$C_{(\text{CH}_3\text{CH}_2)_3\text{N}} V_{(\text{CH}_3\text{CH}_2)_3\text{N}} = C_{\text{HCl}} V_{\text{HCl}}$$

$$V_{\text{HCl}} = \frac{C_{(\text{CH}_3\text{CH}_2)_3\text{N}} V_{(\text{CH}_3\text{CH}_2)_3\text{N}}}{C_{\text{HCl}}}$$

$$= \frac{(0.1000\text{M})(20.00\text{ml})}{(0.1000\text{M})}$$

$$= \boxed{20.00\text{ml}} \leftarrow \text{Vol. of HCl needed to reach equiv.}$$

(a) 0 ml HCl added

What's in the flask? $(\text{CH}_3\text{CH}_2)_3\text{N}$ - weak base
Use K_b to calculate pH

$$K_b \text{ of } (\text{CH}_3\text{CH}_2)_3\text{N} = 5.2 \times 10^{-4}$$

$$K_b = \frac{x^2}{0.1000 - x} \sim \frac{x^2}{0.1000} = 5.2 \times 10^{-4} \quad x = [\text{OH}^-] \text{ since } K_b$$

$$x = \sqrt{(5.2 \times 10^{-4})(0.1000)} = 7.21 \times 10^{-3} = [\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{7.21 \times 10^{-3}} = 1.398 \times 10^{-12}$$

$$\text{pH} = -\log(1.398 \times 10^{-12}) = \boxed{11.85}$$

(b) 10.00 ml HCl added

What's in the flask? Before the equivalence pt, so both $(\text{CH}_3\text{CH}_2)_3\text{N}$ and $(\text{CH}_3\text{CH}_2)_3\text{NH}^+$ are present - weak base and its conj. acid so use Henderson Hasselbalch eqn to calculate the pH.

(b) cont.

shorthand for
(CH₃CH₂)₃N

$$pH = pK_a + \log \frac{[Et_3N]}{[Et_3NH^+]} = pK_a + \log \frac{\text{mole } Et_3N}{\text{mole } Et_3NH^+}$$

$$pK_a = -\log(K_a \text{ of } Et_3NH^+)$$

$$K_a \text{ of } Et_3NH^+ = \frac{K_w}{K_b \text{ of } Et_3N} = \frac{1.008 \times 10^{-14}}{5.2 \times 10^{-4}} = 1.938 \times 10^{-11}$$

$$pK_a = -\log(1.938 \times 10^{-11}) = 10.71$$

Since were 10.00/20.00 the way towards equiv.,
10.00/20.00 = 1/2 the original moles of Et₃N have
been converted to Et₃NH⁺ and 1/2 remains

$$\frac{\text{moles } Et_3N}{\text{moles } Et_3NH^+} = \frac{1/2 \text{ original moles of } Et_3N}{1/2 \text{ original moles of } Et_3N} = 1$$

$$pH = pK_a + \log 1 = pK_a = \boxed{10.71}$$

(c), (d) and (e) are also before equivalence &
the pH is calculated in a similar manner

$$(c) \quad pH = 10.71 + \log \frac{5/20}{15/20} = 10.71 + \log(1/3) = \boxed{10.23}$$

$$(d) \quad pH = 10.71 + \log \frac{1/20}{19/20} = 10.71 + \log(1/19) = \boxed{9.43}$$

$$(e) \quad pH = 10.71 + \log \frac{0.05/20}{19.95/20} = 10.71 + \log\left(\frac{0.05}{19.95}\right) = \boxed{8.11}$$

(f) + 20.00 mL HCl

What is in the flask? Equivalence pt so have added just enough HCl to react with all the $(\text{CH}_3\text{CH}_2)_3\text{N}$ - will just have a soln of $(\text{CH}_3\text{CH}_2)_3\text{NH}^{\oplus}$ since this is the conj. acid of a weak base, it will be a weak acid, use K_a to calculate the pH.

Also need to know the actual conc. of $(\text{CH}_3\text{CH}_2)_3\text{NH}^{\oplus}$

$$[\text{Et}_3\text{NH}^{\oplus}] = \frac{\text{moles Et}_3\text{NH}^{\oplus}}{\text{tot. V}} = \frac{\text{orig. moles of Et}_3\text{N}}{\text{tot. V}}$$

$$= \frac{C_{\text{Et}_3\text{N, orig}} V_{\text{Et}_3\text{N, orig}}}{\text{tot. V}} = \frac{(0.1000 \text{ M})(20.00 \text{ mL})}{(40.00 \text{ mL})} = 0.05000 \text{ M}$$

$$K_a = \frac{x^2}{0.05000 \text{ M} - x} \sim \frac{x^2}{0.05000} = 1.9 \times 10^{-11} \quad x = [\text{H}_3\text{O}^{\oplus}]$$

since now dealing w/ K_a

$$x = \sqrt{(1.9 \times 10^{-11})(0.05000)} = 9.75 \times 10^{-7} = [\text{H}_3\text{O}^{\oplus}]$$

$$\text{pH} = -\log(9.75 \times 10^{-7}) = \boxed{6.01}$$

(g) + 20.05 mL HCl

What's in the flask? Past the equivalence pt so will have excess HCl along with the $\text{Et}_3\text{NH}^{\oplus}$. The pH will be controlled by the HCl since it's a strong acid.

$$[\text{H}_3\text{O}^{\oplus}] = [\text{HCl}]_{\text{excess}} = \frac{\text{moles HCl excess}}{\text{tot. V}}$$
$$= \frac{C_{\text{HCl}} V_{\text{HCl excess}}}{\text{tot. V}} = \frac{(0.1000 \text{ M})(0.05 \text{ mL})}{40.05 \text{ mL}} = 1.25 \times 10^{-4}$$

$$\text{pH} = -\log(1.25 \times 10^{-4}) = \boxed{3.9}$$

(h) is calculated in the same manner

$$[\text{H}_3\text{O}^{\oplus}] = \frac{(0.1000 \text{ M})(5.00 \text{ mL})}{45.00 \text{ mL}} = 1.11 \times 10^{-2} \quad \text{pH} = -\log(1.11 \times 10^{-2}) = \boxed{1.954}$$