A weak base, methylamine (CH_3NH_2), is titrated with hydrochloric acid. When a 0.250 M HCl solution is titrated with 10.0 mL of methylamine solution, it requires 15.0 mL of the HCl to reach the equivalence point. (K_b of methylamine = 4.38×10^{-4})

a) What is the concentration of the methylamine solution?

b) Sketch a titration curve for this reaction. The x-axis should be labeled "mL of HCl added" and the y-axis should be labeled "pH". (THINK about this before you draw your sketch)

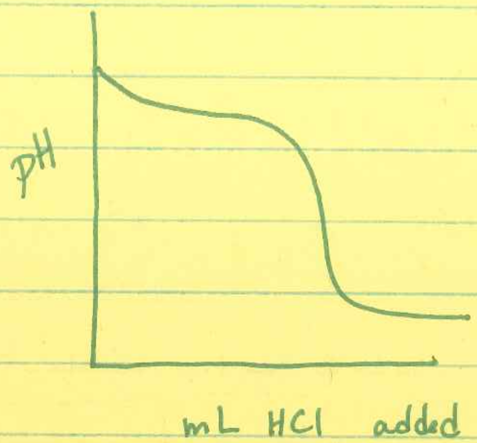
c) What is the pH of the solution when the following volumes of HCl solution are added to the 10.0 mL of methylamine solution?

   i. 0 mL HCl
   ii. 5.00 mL HCl
   iii. 7.50 mL HCl
   iv. 10.0 mL HCl
   v. 15.0 mL HCl
   vi. 17.5 mL HCl

a) $M_A V_A = M_B V_B$ $\quad (0.250\,M)(15.0\,mL) = x\,(10.0\,mL)$

$x = 0.375\,M$ CH_3NH_2

b)



c) (i) 0 mL

soln contains 0.375 M CH_3NH_2 only
(need equation for hydrolysis of CH_3NH_2 + To do an ICE calc)

Molarity

	$H_2O + CH_3NH_2$	$\rightleftharpoons$ $CH_3NH_3^+$	$+ OH^-$
I	0.375 M	0	0
C	$-x$	$+x$	$+x$
E	$0.375-x$	x	x

$$K_b = \frac{x^2}{0.375} = 4.38 \times 10^{-4}$$

$$x = [OH^-] = 0.0128\,M$$

$$pOH = -\log(0.0128\,M) = \cancel{1.466} = 1.893$$

$$pH = 14 - pOH = \cancel{12.534} = 12.107$$

ii) 5.00 mL HCl added to 10.0 mL CH_3NH_2

(need neutralization equation between HCl + CH_3NH_2 + To Do an ICF calc)

$$HCl + CH_3NH_2 \longrightarrow CH_3NH_3^+ + Cl^-$$

b/c strong acid, ionizes completely. Can simplify equation to net ionic below

must be in moles.

	H^+	$+ CH_3NH_2$	$\longrightarrow CH_3NH_3^+$
I	0.00125 mol	0.00375 mol	0
C	−0.00125	−0.00125	+ 0.00125
F	0	0.0025 mol	0.00125 mol

⟸ Look to see what you have for a system.

Recognize this as a buffer!

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

$$pH = -\log\left(\frac{1\times10^{-14}}{4.38\times10^{-4}}\right) + \log \frac{\left(\frac{0.0025\,mol}{0.0150\,L}\right)}{\left(\frac{0.00125\,mol}{0.0150\,L}\right)}$$

$$pH = 10.942$$

iii) 7.50 mL HCl added to 10 mL CH_3NH_2

Did you recognize this as the "halfway point"?

If so, $pH = pK_a$ (If not, you went through same process as part ii)

$$= -\log\left(\frac{K_w}{K_b}\right) = -\log\left(2.28\times10^{-11}\right)$$

$$pH = 10.641$$

iv) 10.0 mL HCl added to 10.0 mL CH_3NH_2.

(Need neutralization equation + To Do a ICF calc)

moles

	H^+ +	CH_3NH_2 $\longrightarrow$	$CH_3NH_3^+$
I	.0025 mol	0.00375 mol	0
C	-.0025	-.0025	+.0025
F	0	0.00125 mol	0.0025 mol

Recognize a BUFFER!

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

$$= 10.641 + \log \left(\frac{0.00125/.0020L}{0.0025/.0020L} \right)$$

$pH = 10.340$

v) 15.0 mL HCl added to 10.0 mL CH_3NH_2

(need neutralization eqn + To Do ICF calc) $\Rightarrow$ OR (recognize that it is the equivalence pt + all moles will be converted to product.)

moles

	H^+ +	CH_3NH_2 $\longrightarrow$	$CH_3NH_3^+$
I	.00375 mol	.00375 mol	0
C	-.00375	-.00375	+.00375
F	0	0	.00375 mol

ONLY $CH_3NH_3^+$ remains...
Do ICE calc for hydrolysis of $CH_3NH_3^+$

Molarity

	$CH_3NH_3^+$ $\rightleftharpoons$	CH_3NH_2 +	H^+
I	0.150 M	0	0
C	-x	+x	+x
E	0.150-x	x	x

$$K_a = \frac{x^2}{0.150} = 2.28 \times 10^{-11}$$

$$x = 1.85 \times 10^{-6} = [H^+]$$

$pH = 5.733$

vi) 17.5 mL HCl added to 10.0 mL CH_3NH_2

need neutralization eqn & To Do ICF calc

moles

	H^+ +	CH_3NH_2 →	$CH_3NH_3^+$
I	0.004375 mol	0.00375 mol	0
C	−.00375	−.00375	+.00375
F	6.25×10^{-4} mol	0	.00375 mol

← Recognize Excess H^+. Use this to determine pH.

$pH = -\log[H^+]$

$pH = -\log\left(\frac{6.25\times10^{-4}\text{ mol}}{0.0275\text{ L}}\right)$

$pH = 1.643$