Review Question 7

	$NH_{3\,(aq)}$	$+ H_2O_{(l)}$	<->	$NH_4^+{}_{(aq)}$	$+ OH^-{}_{(aq)}$
Initial	0.25			0.25	
Change	- x			+x	+x
Eq'm	0.25 – x			0.25 + x	x

At this point you can solve this one of two ways: like a normal equilibrium questions, using method of approximation or using a modification of the Henderson-Hasselback equation.

Method 1	Method 2
$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$ assumption $1.8 \times 10^{-5} = \frac{(0.25)\,[OH^-]}{(0.25)}$ $[OH^-] = 1.8 \times 10^{-5}$ $pOH = -\log [OH^-] = -\log (1.8 \times 10^{-5}) = 4.745$ $pH = 14 - pOH = 9.255$	If we modify the eq`m expression: $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$ get $K_b \times [NH_3] / [NH_4^+] = [OH^-]$ then take – log of both sides $-\log\ [OH^-] = -\log K_b \times [NH_3] / [NH_4^+]$ Then apply two log rules: $pOH = -\log K_b \times \log [NH_4^+] / [NH_3]$ now sub in $pOH = -\log(1.8 \times 10^{-5}) + \log (0.25)/(0.25)$ $pOH = 4.745$ $pH = 14 - pOH = 9.255$

Part b: add add 0.10 mol HCl

You know that the addition of 0.10 mol HCl will react with the OH- and remove 0.10 mol of OH^- from solution. This will cause the eq`m to shift left and will change the amounts of NH_3 and NH_4^+.

	$NH_{3\,(aq)}$	$+ H_2O_{(l)}$	<->	$NH_4^+{}_{(aq)}$	$+ OH^-{}_{(aq)}$	
Initial []	0.25			0.25		
Initial moles (1 L)	0.25			0.25		
Change in moles	-0.1			+0.1		due to the reaction with 0.10 mol HCl
New initial moles	0.15			0.35		
New initial []	0.15			0.35		now this can be done exactly like part a using either method

$pOH = pK_b + \log ([NH_4^+] / [NH_3])$

$pOH = -\log (1.8 \times 10^{-5}) + \log (0.35 / 0.15)$

$pOH = 4.745 + 0.368$

$pOH = 5.113$

$pH = 14 - pOH = 8.887$