



Acid – Base Equilibrium Cheat Sheet

Relationships

Equilibrium Expression

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

$$K_a = \frac{[\text{x}][\text{x}]}{[\text{M}_0]} \text{ where } \text{x} = [\text{H}_3\text{O}^+]$$

$$K_b = \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]}$$

$$K_b = \frac{[\text{x}][\text{x}]}{[\text{M}_0]} \text{ where } \text{x} = [\text{OH}^-]$$

$$\% = \frac{[\text{x}]}{[\text{M}_0]} \times 100 \quad \text{where } \text{x} = [\text{H}_3\text{O}^+] \text{ or } [\text{OH}^-]$$

$$K_a \times K_b = 1 \times 10^{-14}$$

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} + \text{pOH} = 14$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} @ 25^\circ\text{C}$$

$$\text{Acidic } \text{pH} < 7 \quad [\text{H}_3\text{O}^+] > [\text{OH}^-]$$

$$\text{Neutral } \text{pH} = 7 \quad [\text{H}_3\text{O}^+] = [\text{OH}^-]$$

$$\text{Basic } \text{pH} > 7 \quad [\text{H}_3\text{O}^+] < [\text{OH}^-]$$

Conjugate acids and bases

$$K_{sp} = [\text{M}^+]^x[\text{A}^-]^y \quad \text{MA(s)} \rightleftharpoons x\text{M}^+(\text{aq}) + y\text{A}^-(\text{aq})$$

Connections

Equilibrium

Buffers and Titrations

Precipitation and Qualitative Analysis

Bonding and Lewis Structures – justify oxyacid strengths

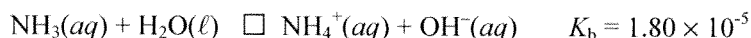
Potential Pitfalls

K_a or K_b with salt pH

Weak acids and bases – be sure you know what you are using, acid or base – K_a or K_b ; solving a K_b problem gives $[\text{OH}^-]$ thus you are finding the pOH!

Weak is about IONIZATION not CONCENTRATION

NMSI SUPER PROBLEM



1. Ammonia reacts with water as indicated in the reaction above.

(a) Write the equilibrium constant expression for the reaction represented above.

$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$	1 point is earned for the correct expression
--	---

(b) Calculate the pH of a 0.150 M solution of NH_3

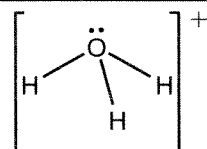
$K_b = \frac{x^2}{M}$ where $x = [\text{OH}^-]$ $1.80 \times 10^{-5} = \frac{x^2}{0.150}$ $1.64 \times 10^{-3} = x$ $-\log x = pOH = 2.784$ $pH = 11.216$	1 point is earned for the correct set up and for calculating the concentration of hydroxide ions 1 point is earned for the correct pH
---	--

(c) Determine the percent ionization of the weak base NH_3 .

$\% = \frac{[x]}{[M_o]} \times 100$ $\% = \frac{1.64 \times 10^{-3}}{[0.150]} \times 100 = 1.09\%$	1 point is earned for the correct percent ionization
---	---

(d) Calculate the hydronium ion, H_3O^+ , concentration in the above solution. Be sure to include units with your answer.

$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}$ $[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14}}{1.64 \times 10^{-3}} = 6.10 \times 10^{-12} M$ OR $[\text{H}_3\text{O}^+] = 10^{-pH}$ $[\text{H}_3\text{O}^+] = 10^{-11.216} = 6.08 \times 10^{-12} M$	1 point is earned for the concentration of hydronium ions with units of mol L^{-1} or M
---	---

N

2 points possible

(i) molarity (M) of the ammonium ions in this solution

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.827} = 1.49 \times 10^{-5}$$

$$K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.80 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$5.56 \times 10^{-10} = \frac{[1.49 \times 10^{-5}][1.49 \times 10^{-5}]}{[M]}$$

$M = 0.400$

1 point is earned for calculating the concentration of ions in the solution.

(ii) number of moles ammonium ions in 250 mL of the above solution.

 $M = 0.400$

$$0.400\text{ M} = \frac{x\text{ mol}}{0.250\text{ L}}$$

$$x = 0.100 \text{ mol}$$

1 point is earned for calculating the molar concentration and the number of moles of ammonium ions

AP[®] CHEMISTRY
2005 SCORING GUIDELINES (Form B)

5 Points

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OCl}^-]}{[\text{HOCl}]} = 3.2 \times 10^{-8}$$

Hypochlorous acid, HOCl, is a weak acid in water. The K_a expression for HOCl is shown above.

(a) Write a chemical equation showing how HOCl behaves as an acid in water.

$\text{HOCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{OCl}^-(aq) + \text{H}_3\text{O}^+(aq)$	One point is earned for the correct chemical equation.
---	--

(b) Calculate the pH of a 0.175 M solution of HOCl.

	$\text{HOCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{OCl}^-(aq) + \text{H}_3\text{O}^+(aq)$			
I	0.175	—	0	~ 0
C	$-x$	—	$+x$	$+x$
E	$0.175 - x$	—	$+x$	$+x$
$K_a = \frac{[\text{H}_3\text{O}^+][\text{OCl}^-]}{[\text{HOCl}]} = \frac{(x)(x)}{(0.175)}$				
$3.2 \times 10^{-8} = \frac{x^2}{0.175}$				
$x^2 = (3.2 \times 10^{-8})(0.175) = 5.6 \times 10^{-9}$				
$x = [\text{H}_3\text{O}^+] = 7.5 \times 10^{-5} M$				
$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (7.5 \times 10^{-5}) = 4.13$				

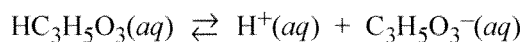
One point is earned for calculating the value of $[\text{H}_3\text{O}^+]$.

One point is earned for calculating the pH.

(c) Write the net ionic equation for the reaction between the weak acid HOCl(aq) and the strong base NaOH(aq).

$\text{HOCl}(aq) + \text{OH}^-(aq) \rightarrow \text{OCl}^-(aq) + \text{H}_2\text{O}(l)$	<p>One point is earned for both of the correct reactants.</p> <p>One point is earned for both of the correct products.</p>
--	--

AP[®] CHEMISTRY
SCORING GUIDELINES - Modified
Short Free Response
5 Points



1. Lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$, is a monoprotic acid that dissociates in aqueous solution, as represented by the equation above. Lactic acid is 1.66 percent dissociated in 0.50 M $\text{HC}_3\text{H}_5\text{O}_3(aq)$ at 298 K. For parts (a) through (c) below, assume the temperature remains at 298 K.

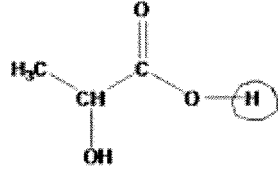
(a) Write the expression for the acid-dissociation constant, K_a , for lactic acid and calculate its value.

$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_3^-]}{[\text{HC}_3\text{H}_5\text{O}_3]}$	1 point earned for equilibrium expression												
$0.50 \text{ M} \times 0.0166 = 0.0083 \text{ M} = x$	1 point earned for amount of $\text{HC}_3\text{H}_5\text{O}_3$ dissociating												
$\text{HC}_3\text{H}_5\text{O}_3(aq) \rightarrow \text{H}^+(aq) + \text{C}_3\text{H}_5\text{O}_3^-(aq)$													
<table><tr><td>I</td><td>0.50</td><td>~0</td><td>0</td></tr><tr><td>C</td><td>-x</td><td>+x</td><td>+x</td></tr><tr><td>E</td><td>0.50 - x</td><td>+x</td><td>+x</td></tr></table>	I	0.50	~0	0	C	-x	+x	+x	E	0.50 - x	+x	+x	
I	0.50	~0	0										
C	-x	+x	+x										
E	0.50 - x	+x	+x										
$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_3^-]}{[\text{HC}_3\text{H}_5\text{O}_3]} = \frac{[0.0083][0.0083]}{[0.50]}$	1 point earned for $[\text{H}^+] = [\text{C}_3\text{H}_5\text{O}_3^-]$ set up and solution												
$K_a = 1.4 \times 10^{-4}$													

(b) Calculate the pH of 0.50 M $\text{HC}_3\text{H}_5\text{O}_3$.

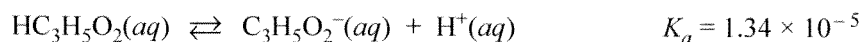
<p>From part (a):</p> <p>$[\text{H}^+] = 0.0083 \text{ M}$</p> <p>$\text{pH} = -\log [\text{H}^+] = -\log (0.0083) = 2.08$</p>	1 point earned for correctly calculating pH
--	---

(c) The structure of lactic acid is shown below. Circle the H^+ that is ionized in $\text{HC}_3\text{H}_5\text{O}_3$.

 <p>The circled hydrogen is a part of the carboxylic acid group (COOH).</p>	1 point for identifying the correct hydrogen
--	--

AP[®] CHEMISTRY
2005 SCORING GUIDELINES

Question 1



Propanoic acid, $\text{HC}_3\text{H}_5\text{O}_2$, ionizes in water according to the equation above.

(a) Write the equilibrium-constant expression for the reaction.

$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]}$ <p><u>Notes:</u> Correct expression without K_a earns 1 point. Entering the value of K_a is acceptable. Charges must be correct to earn 1 point.</p>	<p>One point is earned for the correct equilibrium expression.</p>
---	--

(b) Calculate the pH of a 0.265 *M* solution of propanoic acid.

$\text{HC}_3\text{H}_5\text{O}_2(aq) \rightleftharpoons \text{C}_3\text{H}_5\text{O}_2^-(aq) + \text{H}^+(aq)$			
I	0.265	0	~0
C	-x	+x	+x
E	0.265 - x	+x	+x
$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]} = \frac{(x)(x)}{(0.265 - x)}$			
Assume that $0.265 - x \approx 0.265$,			
then $1.34 \times 10^{-5} = \frac{x^2}{0.265}$			
$(1.34 \times 10^{-5})(0.265) = x^2$			
$3.55 \times 10^{-6} = x^2$			
$x = [\text{H}^+] = 1.88 \times 10^{-3} M$			
$\text{pH} = -\log [\text{H}^+] = -\log (1.88 \times 10^{-3})$			
$\text{pH} = 2.725$			

One point is earned for recognizing that $[\text{H}^+]$ and $[\text{C}_3\text{H}_5\text{O}_2^-]$ have the same value in the equilibrium expression.
One point is earned for calculating $[\text{H}^+]$.
One point is earned for calculating the correct pH.

AP[®] CHEMISTRY
2005 SCORING GUIDELINES

Question 1 (continued)

- (c) A 0.496 g sample of sodium propanoate, $\text{NaC}_3\text{H}_5\text{O}_2$, is added to a 50.0 mL sample of a 0.265 M solution of propanoic acid. Assuming that no change in the volume of the solution occurs, calculate each of the following.

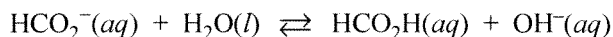
- (i) The concentration of the propanoate ion, $\text{C}_3\text{H}_5\text{O}_2^-(aq)$ in the solution

$\text{mol NaC}_3\text{H}_5\text{O}_2 = 0.496 \text{ g NaC}_3\text{H}_5\text{O}_2 \times \frac{1 \text{ mol NaC}_3\text{H}_5\text{O}_2}{96.0 \text{ g NaC}_3\text{H}_5\text{O}_2}$ $\text{mol NaC}_3\text{H}_5\text{O}_2 = 5.17 \times 10^{-3} \text{ mol NaC}_3\text{H}_5\text{O}_2 = \text{mol C}_3\text{H}_5\text{O}_2^-$ $[\text{C}_3\text{H}_5\text{O}_2^-] = \frac{\text{mol C}_3\text{H}_5\text{O}_2^-}{\text{volume of solution}} = \frac{5.17 \times 10^{-3} \text{ mol C}_3\text{H}_5\text{O}_2^-}{0.050 \text{ L}} = 0.103 \text{ M}$	<p>One point is earned for calculating the number of moles of $\text{NaC}_3\text{H}_5\text{O}_2$.</p> <p>One point is earned for the molarity of the solution.</p>
--	---

- (ii) The concentration of the $\text{H}^+(aq)$ ion in the solution

$\text{HC}_3\text{H}_5\text{O}_2(aq) \rightleftharpoons \text{C}_3\text{H}_5\text{O}_2^-(aq) + \text{H}^+(aq)$				<p>One point is earned for calculating the value of $[\text{H}^+]$.</p>
I	0.265	0.103	~ 0	
C	$-x$	$+x$	$+x$	
E	$0.265 - x$	$0.103 + x$	$+x$	
$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]} = \frac{(x)(0.103 + x)}{(0.265 - x)}$				
<p>Assume that $0.103 + x \approx 0.103$ and $0.265 - x \approx 0.265$</p>				
$K_a = 1.34 \times 10^{-5} = \frac{(x)(0.103)}{0.265}$				
$x = [\text{H}^+] = (1.34 \times 10^{-5}) \times \frac{0.265}{0.103} = 3.45 \times 10^{-5} \text{ M}$				

The methanoate ion, $\text{HCO}_2^-(aq)$, reacts with water to form methanoic acid and hydroxide ion, as shown in the following equation.



- (d) Given that $[\text{OH}^-]$ is $4.18 \times 10^{-6} \text{ M}$ in a 0.309 M solution of sodium methanoate, calculate each of the following.

**AP[®] CHEMISTRY
2005 SCORING GUIDELINES**

Question 1 (continued)

(i) The value of K_b for the methanoate ion, $\text{HCO}_2^-(aq)$

$\text{HCO}_2^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_2\text{H} + \text{OH}^-(aq)$					<p>One point is earned for substituting 4.18×10^{-6} for both $[\text{OH}^-]$ and $[\text{HCO}_2\text{H}]$, and for calculating the value of K_b.</p>
I	0.309	-	0	~ 0	
C	$-x$	-	$+x$	$+x$	
E	$0.309 - x$	-	$+x$	$+x$	
$x = [\text{OH}^-] = 4.18 \times 10^{-6} M$ $K_b = \frac{[\text{OH}^-][\text{HCO}_2\text{H}]}{[\text{HCO}_2^-]} = \frac{(x)(x)}{(0.309 - x)} = \frac{(4.18 \times 10^{-6})^2}{(0.309 - x)}$ x is very small ($4.18 \times 10^{-6} M$), therefore $0.309 - x \approx 0.309$ $K_b = \frac{(4.18 \times 10^{-6})^2}{0.309} = 5.65 \times 10^{-11}$					

(ii) The value of K_a for methanoic acid, HCO_2H

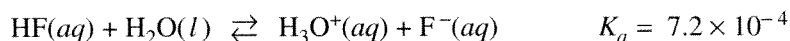
$K_w = K_a \times K_b$ $K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{5.65 \times 10^{-11}}$ $K_a = 1.77 \times 10^{-4}$	<p>One point is earned for calculating a value of K_a from the value of K_b determined in part (d)(i).</p>
--	--

(e) Which acid is stronger, propanoic acid or methanoic acid? Justify your answer.

<p>K_a for propanoic acid is 1.34×10^{-5}, and K_a for methanoic acid is 1.77×10^{-4}. For acids, the larger the value of K_a, the greater the strength; therefore methanoic acid is the stronger acid because $1.77 \times 10^{-4} > 1.34 \times 10^{-5}$.</p>	<p>One point is earned for the correct choice and explanation based on the K_a calculated for methanoic acid in part (d)(ii).</p>
--	--

AP[®] CHEMISTRY
2007 SCORING GUIDELINES

Question 1



Hydrofluoric acid, $\text{HF}(aq)$, dissociates in water as represented by the equation above.

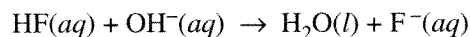
(a) Write the equilibrium-constant expression for the dissociation of $\text{HF}(aq)$ in water.

$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$	One point is earned for the correct expression.
--	---

(b) Calculate the molar concentration of H_3O^+ in a 0.40 M $\text{HF}(aq)$ solution.

$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(x)}{0.40 - x} = 7.2 \times 10^{-4}$ <p>Assume $x \ll 0.40$, then $x^2 = (0.40)(7.2 \times 10^{-4})$</p> $x = [\text{H}_3\text{O}^+] = 0.017 \text{ M}$	<p>One point is earned for the correct setup (or the setup consistent with part (a)).</p> <p>One point is earned for the correct concentration.</p>
--	---

$\text{HF}(aq)$ reacts with $\text{NaOH}(aq)$ according to the reaction represented below.



A volume of 15 mL of 0.40 M $\text{NaOH}(aq)$ is added to 25 mL of 0.40 M $\text{HF}(aq)$ solution. Assume that volumes are additive.

(c) Calculate the number of moles of $\text{HF}(aq)$ remaining in the solution.

$\begin{aligned} \text{mol HF}(aq) &= \text{initial mol HF}(aq) - \text{mol NaOH}(aq) \text{ added} \\ &= (0.025 \text{ L})(0.40 \text{ mol L}^{-1}) - (0.015 \text{ L})(0.40 \text{ mol L}^{-1}) \\ &= 0.010 \text{ mol} - 0.0060 \text{ mol} = 0.004 \text{ mol} \end{aligned}$	<p>One point is earned for determining the initial number of moles of HF and OH^-.</p> <p>One point is earned for setting up and doing correct subtraction.</p>
---	---

(d) Calculate the molar concentration of $\text{F}^-(aq)$ in the solution.

$\begin{aligned} \text{mol F}^-(aq) \text{ formed} &= \text{mol NaOH}(aq) \text{ added} = 0.0060 \text{ mol F}^-(aq) \\ \frac{0.0060 \text{ mol F}^-(aq)}{(0.015 + 0.025) \text{ L of solution}} &= 0.15 \text{ M F}^-(aq) \end{aligned}$	<p>One point is earned for determining the number of moles of $\text{F}^-(aq)$.</p> <p>One point is earned for dividing the number of moles of $\text{F}^-(aq)$ by the correct total volume.</p>
---	--

AP[®] CHEMISTRY
2007 SCORING GUIDELINES

Question 1 (continued)

(e) Calculate the pH of the solution.

$[\text{HF}] = \frac{0.004 \text{ mol HF}}{0.040 \text{ L}} = 0.10 \text{ M HF}$ $K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \Rightarrow \frac{[\text{HF}] \times K_a}{[\text{F}^-]} = [\text{H}_3\text{O}^+]$ $\Rightarrow \frac{0.10 \text{ M} (7.2 \times 10^{-4})}{0.15 \text{ M}} = 4.8 \times 10^{-4}$ $\Rightarrow \text{pH} = -\log (4.8 \times 10^{-4}) = 3.32$ <p style="text-align: center;">OR</p> $\text{pH} = \text{p}K_a + \log \frac{[\text{F}^-]}{[\text{HF}]}$ $= -\log (7.2 \times 10^{-4}) + \log \frac{0.15 \text{ M}}{0.10 \text{ M}}$ $= 3.14 + 0.18$ $= 3.32$	<p>One point is earned for indicating that the resulting solution is a buffer (e.g., by showing a ratio of $[\text{F}^-]$ to $[\text{HF}]$ or moles of F^- to HF).</p> <p>One point is earned for the correct calculation of pH.</p>
--	--

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

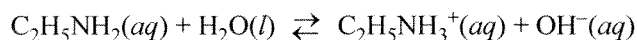
Question 1 (10 points)

A pure 14.85 g sample of the weak base ethylamine, $\text{C}_2\text{H}_5\text{NH}_2$, is dissolved in enough distilled water to make 500. mL of solution.

- (a) Calculate the molar concentration of the $\text{C}_2\text{H}_5\text{NH}_2$ in the solution.

$n_{\text{C}_2\text{H}_5\text{NH}_2} = 14.85 \text{ g C}_2\text{H}_5\text{NH}_2 \times \frac{1 \text{ mol C}_2\text{H}_5\text{NH}_2}{45.09 \text{ g C}_2\text{H}_5\text{NH}_2}$ $= 0.3293 \text{ mol C}_2\text{H}_5\text{NH}_2$ $M_{\text{C}_2\text{H}_5\text{NH}_2} = \frac{0.3293 \text{ mol C}_2\text{H}_5\text{NH}_2}{0.500 \text{ L}} = \mathbf{0.659 \text{ M}}$	<p>One point is earned for the correct number of moles.</p> <p>One point is earned for the correct concentration.</p>
--	---

The aqueous ethylamine reacts with water according to the equation below.



- (b) Write the equilibrium-constant expression for the reaction between $\text{C}_2\text{H}_5\text{NH}_2(aq)$ and water.

$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$	One point is earned for the correct expression.
--	---

- (c) Of $\text{C}_2\text{H}_5\text{NH}_2(aq)$ and $\text{C}_2\text{H}_5\text{NH}_3^+(aq)$, which is present in the solution at the higher concentration at equilibrium? Justify your answer.

$\text{C}_2\text{H}_5\text{NH}_2$ is present in the solution at the higher concentration at equilibrium. Ethylamine is a weak base, and thus it has a small K_b value. Therefore only partial dissociation of $\text{C}_2\text{H}_5\text{NH}_2$ occurs in water, and $[\text{C}_2\text{H}_5\text{NH}_3^+]$ is thus less than $[\text{C}_2\text{H}_5\text{NH}_2]$.	One point is earned for the correct answer with justification.
--	--

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

Question 1 (continued)

- (d) A different solution is made by mixing 500. mL of 0.500 *M* C₂H₅NH₂ with 500. mL of 0.200 *M* HCl. Assume that volumes are additive. The pH of the resulting solution is found to be 10.93.

- (i) Calculate the concentration of OH[−](*aq*) in the solution.

$\text{pH} = -\log[\text{H}^+]$ $[\text{H}^+] = 10^{-10.93} = 1.17 \times 10^{-11}$ $[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1.00 \times 10^{-14}}{1.17 \times 10^{-11}} = \mathbf{8.5 \times 10^{-4} \text{ M}}$ <p>OR</p> $\text{pOH} = 14 - \text{pH} = 14 - 10.93 = 3.07$ $\text{pOH} = -\log[\text{OH}^-]$ $[\text{OH}^-] = 10^{-3.07} = \mathbf{8.5 \times 10^{-4} \text{ M}}$	<p>One point is earned for the correct concentration.</p>
--	---

- (ii) Write the net-ionic equation that represents the reaction that occurs when the C₂H₅NH₂ solution is mixed with the HCl solution.

$\text{C}_2\text{H}_5\text{NH}_2 + \text{H}_3\text{O}^+ \rightarrow \text{C}_2\text{H}_5\text{NH}_3^+ + \text{H}_2\text{O}$	<p>One point is earned for the correct equation.</p>
---	--

- (iii) Calculate the molar concentration of the C₂H₅NH₃⁺(*aq*) that is formed in the reaction.

moles of $\text{C}_2\text{H}_5\text{NH}_2 = 0.500 \text{ L} \times \frac{0.500 \text{ mol}}{1.00 \text{ L}} = \mathbf{0.250 \text{ mol}}$

moles of $\text{H}_3\text{O}^+ = 0.500 \text{ L} \times \frac{0.200 \text{ mol}}{1.00 \text{ L}} = \mathbf{0.100 \text{ mol}}$

	$[\text{C}_2\text{H}_5\text{NH}_2]$	$[\text{H}_3\text{O}^+]$	$[\text{C}_2\text{H}_5\text{NH}_3^+]$
initial value	0.250	0.100	~ 0
change	-0.100	-0.100	+0.100
final value	0.150	~ 0	0.100

$[\text{C}_2\text{H}_5\text{NH}_3^+] = \frac{0.100 \text{ mol C}_2\text{H}_5\text{NH}_3^+}{1.00 \text{ L}} = \mathbf{0.100 \text{ M}}$

One point is earned for the correct number of moles of $\text{C}_2\text{H}_5\text{NH}_2$ and H_3O^+ .

<

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

Question 1 (continued)

(iv) Calculate the value of K_b for $\text{C}_2\text{H}_5\text{NH}_2$.

$[\text{C}_2\text{H}_5\text{NH}_2] = \frac{0.150 \text{ mol C}_2\text{H}_5\text{NH}_2}{1.00 \text{ L}} = \mathbf{0.150 \text{ M}}$ $K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} = \frac{(0.100)(8.5 \times 10^{-4})}{0.150} = \mathbf{5.67 \times 10^{-4}}$	<p>One point is earned for the correct calculation of the molarity of $\text{C}_2\text{H}_5\text{NH}_2$ after neutralization.</p> <p>One point is earned for the correct value.</p>
--	--

