

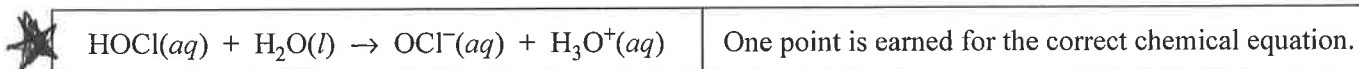
**AP<sup>®</sup> CHEMISTRY**  
**2005 SCORING GUIDELINES (Form B)**

**Question 1**

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

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

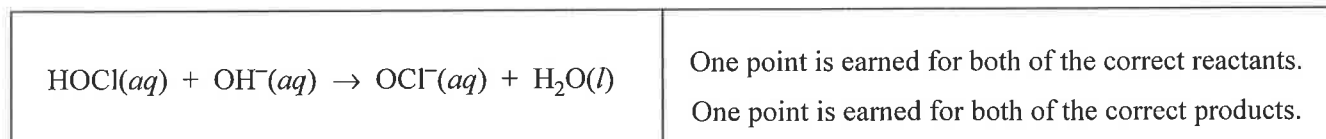
(a) Write a chemical equation showing how HOCl behaves as an acid in water.



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

$\text{HOCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{OCl}^-(aq) + \text{H}_3\text{O}^+(aq)$ <table style="margin-left: 20px;"> <tr> <td style="padding-right: 10px;">I</td> <td style="padding-right: 10px;">0.175</td> <td style="padding-right: 10px;">-</td> <td style="padding-right: 10px;">0</td> <td style="padding-right: 10px;">~ 0</td> </tr> <tr> <td>C</td> <td>-x</td> <td>-</td> <td>+x</td> <td>+x</td> </tr> <tr> <td>E</td> <td>0.175 - x</td> <td>-</td> <td>+x</td> <td>+x</td> </tr> </table> $K_a = \frac{[\text{H}_3\text{O}^+][\text{OCl}^-]}{[\text{HOCl}]} = \frac{(x)(x)}{(0.175 - x)}$ <p>Assume that <math>0.175 - x \approx 0.175</math></p> $3.2 \times 10^{-8} = \frac{x^2}{0.175}$ $x^2 = (3.2 \times 10^{-8})(0.175) = 5.6 \times 10^{-9}$ $x = [\text{H}_3\text{O}^+] = 7.5 \times 10^{-5} M$ $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (7.5 \times 10^{-5}) = 4.13$	I	0.175	-	0	~ 0	C	-x	-	+x	+x	E	0.175 - x	-	+x	+x	<p>One point is earned for calculating the value of <math>[\text{H}_3\text{O}^+]</math>.</p> <p>One point is earned for calculating the pH.</p>
I	0.175	-	0	~ 0												
C	-x	-	+x	+x												
E	0.175 - x	-	+x	+x												

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



(d) In an experiment, 20.00 mL of 0.175 M HOCl(aq) is placed in a flask and titrated with 6.55 mL of 0.435 M NaOH(aq).

(i) Calculate the number of moles of NaOH(aq) added.

$\text{mol}_{\text{NaOH}} = 6.55 \text{ mL} \times \frac{1 \text{ L}}{1,000 \text{ mL}} \times \frac{0.435 \text{ mol NaOH}}{1 \text{ L}}$ $\text{mol}_{\text{NaOH}} = 2.85 \times 10^{-3} \text{ mol NaOH}$	<p style="text-align: right; margin-right: 20px;"><i>or: (Molarity)(Volume) = moles</i></p> <p>One point is earned for the correct number of moles of NaOH.</p>
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**Question 1 (continued)**

(ii) Calculate  $[H_3O^+]$  in the flask after the  $NaOH(aq)$  has been added.

$$\text{mol}_{\text{HOCl}} = 20.00 \text{ mL} \times \frac{1 \text{ L}}{1,000 \text{ mL}} \times \frac{0.175 \text{ mol NaOH}}{1 \text{ L}} = 3.50 \times 10^{-3} \text{ mol}$$

$OH^-(aq)$  is the limiting reactant, therefore all of it reacts

	$HOCl(aq)$	$OH^-(aq)$	$\rightarrow$	$OCl^-(aq)$	$+$	$H_2O(l)$	
I	0.00350	0.00285		0		-	
C	-0.00285	-0.00285		+0.00285		-	
E	0.00065	0		0.00285		-	

★  $M_{\text{HOCl}} = \frac{0.00065 \text{ mol}}{0.02655 \text{ L}} = 0.0245 \text{ M}$

★  $M_{\text{OCl}^-} = \frac{0.00285 \text{ mol}}{0.02655 \text{ L}} = 0.107 \text{ M}$

	$HOCl(aq)$	$+ H_2O(l)$	$\rightarrow$	$H_3O^+(aq)$	$+$	$OCl^-(aq)$
I	0.0245	-		~0		0.107
C	-x	-		+x		+x
E	0.0245 - x	-		+x		0.107 + x

$$K_a = \frac{[H_3O^+][OCl^-]}{[HOCl]} = \frac{(x)(0.107 + x)}{(0.0245 - x)}$$

Assume that  $0.107 + x \approx 0.107$  and that  $0.0245 - x \approx 0.0245$

$$3.2 \times 10^{-8} = \frac{(x)(0.107)}{(0.0245)}$$

★  $x = [H_3O^+] = 7.3 \times 10^{-9} \text{ M}$

One point is earned for calculating the initial number of moles of HOCl.

This is the behavior equation for HOCl + OCl<sup>-</sup> to set up their equilibrium. It is NOT needed to work the problem.

One point is earned for the concentration or number of moles of HOCl and OCl<sup>-</sup> after the neutralization reaction.

One point is earned for the correct  $[H_3O^+]$ .

(iii) Calculate  $[OH^-]$  in the flask after the  $NaOH(aq)$  has been added.

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14} = K_w$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{7.3 \times 10^{-9}} = 1.4 \times 10^{-6} \text{ M}$$

One point is earned for the correct concentration of  $OH^-$ .

This is NOT the  $OH^-$  that came from the buret. (That's all gone)  
It is the tiny amount of  $OH^-$  that's present in any acid solution ~~that~~ with a water solvent.

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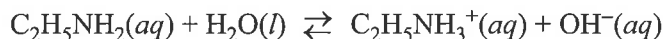
**Question 1 (10 points)**

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>
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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.

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

<p><math>\text{C}_2\text{H}_5\text{NH}_2</math> is present in the solution at the higher concentration at equilibrium. Ethylamine is a weak base, and thus it has a small <math>K_b</math> value. Therefore only partial dissociation of <math>\text{C}_2\text{H}_5\text{NH}_2</math> occurs in water, and <math>[\text{C}_2\text{H}_5\text{NH}_3^+]</math> is thus less than <math>[\text{C}_2\text{H}_5\text{NH}_2]</math>.</p>	<p>One point is earned for the correct answer with justification.</p>
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**Question 1 (continued)**

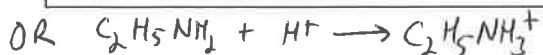
(d) A different solution is made by mixing 500. mL of 0.500 M C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub> with 500. mL of 0.200 M HCl. Assume that volumes are additive. The pH of the resulting solution is found to be 10.93.

(i) Calculate the concentration of OH<sup>-</sup>(aq) in the solution.

<p>pH = -log[H<sup>+</sup>]  [H<sup>+</sup>] = 10<sup>-10.93</sup> = 1.17 × 10<sup>-11</sup>  [OH<sup>-</sup>] = <math>\frac{K_w}{[H^+]}</math> = <math>\frac{1.00 \times 10^{-14}}{1.17 \times 10^{-11}}</math> = <b>8.5 × 10<sup>-4</sup> M</b>  <b>OR</b>  pOH = 14 - pH = 14 - 10.93 = 3.07  pOH = -log[OH<sup>-</sup>]  [OH<sup>-</sup>] = 10<sup>-3.07</sup> = <b>8.5 × 10<sup>-4</sup> M</b></p>	<p>One point is earned for the correct concentration.</p>
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(ii) Write the net-ionic equation that represents the reaction that occurs when the C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub> solution is mixed with the HCl solution. *weak base/strong acid*

$C_2H_5NH_2 + H_3O^+ \rightarrow C_2H_5NH_3^+ + H_2O$	<p>One point is earned for the correct equation.</p>
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(iii) Calculate the molar concentration of the C<sub>2</sub>H<sub>5</sub>NH<sub>3</sub><sup>+</sup>(aq) that is formed in the reaction.

<p>moles of C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub> = 0.500 L × <math>\frac{0.500 \text{ mol}}{1.00 \text{ L}}</math> = <b>0.250 mol</b>  moles of H<sub>3</sub>O<sup>+</sup> = 0.500 L × <math>\frac{0.200 \text{ mol}}{1.00 \text{ L}}</math> = <b>0.100 mol</b></p> <table border="1" style="margin-left: auto; margin-right: auto; border-collapse: collapse; text-align: center;"> <thead> <tr> <th></th> <th>[C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub>]</th> <th>[H<sub>3</sub>O<sup>+</sup>]</th> <th>[C<sub>2</sub>H<sub>5</sub>NH<sub>3</sub><sup>+</sup>]</th> </tr> </thead> <tbody> <tr> <td>initial value</td> <td>0.250</td> <td>0.100</td> <td>~ 0</td> </tr> <tr> <td>change</td> <td>-0.100</td> <td>-0.100</td> <td>+0.100</td> </tr> <tr> <td>final value</td> <td>0.150</td> <td>~ 0</td> <td>0.100</td> </tr> </tbody> </table> <p><math>[C_2H_5NH_3^+] = \frac{0.100 \text{ mol } C_2H_5NH_3^+}{1.00 \text{ L}} = \mathbf{0.100 \text{ M}}</math></p>		[C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> ]	[H <sub>3</sub> O <sup>+</sup> ]	[C <sub>2</sub> H <sub>5</sub> NH <sub>3</sub> <sup>+</sup> ]	initial value	0.250	0.100	~ 0	change	-0.100	-0.100	+0.100	final value	0.150	~ 0	0.100	<p>One point is earned for the correct number of moles of C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub> and H<sub>3</sub>O<sup>+</sup>.</p> <p style="text-align: center;">One point is earned for the correct concentration.</p>
	[C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> ]	[H <sub>3</sub> O <sup>+</sup> ]	[C <sub>2</sub> H <sub>5</sub> NH <sub>3</sub> <sup>+</sup> ]														
initial value	0.250	0.100	~ 0														
change	-0.100	-0.100	+0.100														
final value	0.150	~ 0	0.100														

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Question 1 (continued)

(iv) Calculate the value of  $K_b$  for  $C_2H_5NH_2$ .

Use  
into  
from  
ICE  
table

$$\rightarrow [C_2H_5NH_2] = \frac{0.150 \text{ mol } C_2H_5NH_2}{1.00 \text{ L}} = 0.150 \text{ M}$$

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{(0.100)(8.5 \times 10^{-4})}{0.150} = 5.67 \times 10^{-4}$$

$[OH^-]$  from (i)

One point is earned for the correct calculation of the molarity of  $C_2H_5NH_2$  after neutralization.

One point is earned for the correct value.