AP® Chemistry
2002 Sample Student Responses

1A, 1

CHEMISTRY

Section II

(Total time—90 minutes)

Part A

Time—40 minutes

YOU MAY USE YOUR CALCULATOR FOR PART A.

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

Be sure to write all your answers to the questions on the lined pages following each question in the booklet with the pink cover. Do NOT write your answers on the green insert.

Answer Question 1 below. The Section II score weighting for this question is 20 percent.

$$HOBr(aq) \rightleftarrows H^+(aq) + OBr^-(aq) \qquad K_a = 2.3 \times 10^{-9}$$

1. Hypobromous acid, HOBr , is a weak acid that dissociates in water, as represented by the equation above.

   (a) Calculate the value of $[H^+]$ in an HOBr solution that has a pH of 4.95.

   (b) Write the equilibrium constant expression for the ionization of HOBr in water, then calculate the concentration of HOBr(*aq*) in an HOBr solution that has $[H^+]$ equal to 1.8×10^{-5} *M*.

   (c) A solution of $Ba(OH)_2$ is titrated into a solution of HOBr.

   (i) Calculate the volume of 0.115 *M* $Ba(OH)_2(aq)$ needed to reach the equivalence point when titrated into a 65.0 mL sample of 0.146 *M* HOBr(*aq*).

   (ii) Indicate whether the pH at the equivalence point is less than 7, equal to 7, or greater than 7. Explain.

   (d) Calculate the number of moles of NaOBr(*s*) that would have to be added to 125 mL of 0.160 *M* HOBr to produce a buffer solution with $[H^+] = 5.00 \times 10^{-9}$ *M*. Assume that volume change is negligible.

   (e) HOBr is a weaker acid than $HBrO_3$. Account for this fact in terms of molecular structure.

1. a $[H^+] = 10^{-pH}$

$[H^+] = 1.1 \times 10^{-5} M$

b.

$K_a = \dfrac{[H^+][OBr^-]}{[HOBr]} = 2.3 \times 10^{-9} = \dfrac{(1.8 \times 10^{-5})(1.8 \times 10^{-5})}{x}$

$x = [HOBr] = .14\ M$

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1A2

c. $HOBr + OH^- \rightarrow OBr^- + H_2O$

i.

$$.0650L \times \frac{.146 mol\ HOBr}{1L} \times \frac{1 mol\ OH^-}{1 mol\ HOBr} \times \frac{1 mol\ Ba(OH)_2}{2 mol\ OH^-} \times \frac{1L}{.115 mol\ Ba(OH)_2}$$

$.0413\ L\ Ba(OH)_2$

ii.

When at the eq pt the moles of weak acid (HOBr) used are equal to the moles of strong base (OH^-) used. Since this neutralization rxn goes virtually all the way no HOBr or OH^- is left in soln. It has all been converted into OBr^- & H_2O the OBr^- reacts in the H_2O as follows:

$OBr^- + H_2O \rightleftharpoons HOBr + OH^-$

Since OBr^- is a weak base some hydroxide will be present ∴ the soln will be basic and have a pH above 7

d. $HOBr + H_2O \rightleftharpoons OBr^- + H_3O^+$

.125L × .160m   .160m   x   0

+ ~~?~~ 5.00×10^{-9}

.160m   x

$$K_a = \frac{(x)(5.00\times10^{-9})}{.160M} = 2.3\times10^{-9}$$

$x = [OBr^-] = .0736M \cdot .125L = \boxed{.0920 mol\ OBr^-}$

e. HOBr $HBrO_3$

Br-O-H   H-O-Br(O)(O)

Both substances are oxy acids. An oxy acid is stronger in substances whose central atoms are more EN since both are Br this is not the case, the next factor is # of O atoms, These extra O atoms in $HBrO_3$ will pull the electron cloud btw O & H more towards O ∴ the H will be more δ^+ and easier to remove, Thus H_2O will "attack it" more and more easily and the acid will dissociate more.

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1 B 1

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1. Hypobromous acid, $HOBr$, is a weak acid that dissociates in water, as represented by the equation above.

   (a) Calculate the value of $[H^+]$ in an $HOBr$ solution that has a pH of 4.95.

   (b) Write the equilibrium constant expression for the ionization of $HOBr$ in water, then calculate the concentration of $HOBr(aq)$ in an $HOBr$ solution that has $[H^+]$ equal to 1.8×10^{-5} *M*.

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      (i) Calculate the volume of 0.115 *M* $Ba(OH)_2(aq)$ needed to reach the equivalence point when titrated into a 65.0 mL sample of 0.146 *M* $HOBr(aq)$.

      (ii) Indicate whether the pH at the equivalence point is less than 7, equal to 7, or greater than 7. Explain.

   (d) Calculate the number of moles of $NaOBr(s)$ that would have to be added to 125 mL of 0.160 *M* $HOBr$ to produce a buffer solution with $[H^+] = 5.00 \times 10^{-9}$ *M*. Assume that volume change is negligible.

   (e) $HOBr$ is a weaker acid than $HBrO_3$. Account for this fact in terms of molecular structure.

a) $pH = 4.95 = -\log[H^+]$

$[H^+] = 1.1 \times 10^{-5}$ M

b) $K_a = \frac{[OBr^-][H^+]}{[HOBr]}$ $\qquad [H^+] = [OBr^-] = 1.8 \times 10^{-5}$ M

$2.3 \times 10^{-9} = \frac{(1.8 \times 10^{-5} M)^2}{[HOBr]}$

$[HOBr] = 7.1$ M

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ADDITIONAL PAGE FOR ANSWERING QUESTION 1. 1 B 2

c) i) $(0.146\ M)(0.0650\ L) = 9.49 \times 10^{-3}\ mol\ H^+$

$\frac{9.49 \times 10^{-3}\ mol\ OH^-}{0.230\ M\ OH^-} = 0.0413\ L = 41.3\ mL$

ii) The pH is greater than 7 since $Ba(OH)_2$ is a strong base. A strong base and a weak acid give a pH greater than 7 at the equivalence point. The OBr^- left after reaction will act as a weak base to make a basic solution.

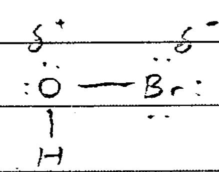
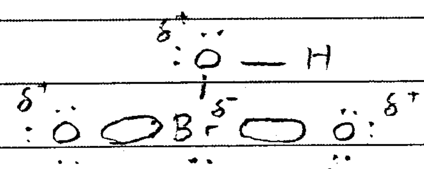
d) x = concentration of NaOBr

	HOBr	H^+	OBr^-
I	0.160 M	0	x
C	-5.00×10^{-9} M	5.00×10^{-9} M	5.00×10^{-9} M
E	0.160 M	5.00×10^{-9} M	$5.00 \times 10^{-9} + x$

$$2.3 \times 10^{-9} = \frac{(5.00 \times 10^{-9})(5.00 \times 10^{-9} + x)}{0.160\ M}$$

$x = 0.074\ M\ OBr^- = 0.074\ M\ NaOBr$

$(0.074\ M)(0.125\ L) = 9.3 \times 10^{-3}$ mol of NaOBr

e)  

Stronger acids dissociate more than weaker acids. In HOBr there is a Br atom that is negatively charged by the polar bond with O. This attracts the positive H atom, thus keeping the molecule from dissociating. In $HBrO_3$ for the same reason, there are three O atoms that are positively charged and repel the positive H atom. This makes $HBrO_3$ dissociate more than HOBr.

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1C1

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(i) Calculate the volume of 0.115 M $Ba(OH)_2(aq)$ needed to reach the equivalence point when titrated into a 65.0 mL sample of 0.146 M $HOBr(aq)$.

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(e) $HOBr$ is a weaker acid than $HBrO_3$. Account for this fact in terms of molecular structure.

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

$[H^+] = 10^{-pH}$

$\boxed{[H^+] = 1.12\times10^{-5} M}$

b) $K_c = \dfrac{[OBr^-][H^+]}{[HOBr]}$

$K_a = \dfrac{[H^+]^2}{[HOBr]}$

$2.3\times10^{-9} = \dfrac{[1.8\times10^{-5}]^2}{[HOBr]}$

$\boxed{[HOBr] = 0.14 M}$

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ADDITIONAL PAGE FOR ANSWERING QUESTION 1.

c) i- $0.115M(V) = 0.146M\left(\frac{65.0}{1000}L\right)$

$V = 0.0825\,L$

$\boxed{V = 82.5\,mL}$

ii- The titrated solution would be greater than 7 because the strong base added will release more OH^- ions to attract the H^+ ions, therefore creating more of the conjugate base.

d) $K_a = \frac{[OBr^-][H^+]}{[HOBr]}$

$\frac{2.3\times10^{-9}\,[HOBr]}{[H^+]} = [OBr^-]$

$[OBr^-] = \frac{2.3\times10^{-9}\,(0.160)}{5.00\times10^{-9}}$

$[OBr^-] = 0.0736\,M$

moles = $M(V)$

moles $OBr^- = 0.0736(0.125\,L)$

$\boxed{\text{moles } OBr^- = 9.20\times10^{-3}}$

e) $HBrO_3$ is a stronger acid than HOBr because the bonds in $HBrO_3$ are covalent while those in HOBr are ionic.

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