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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.

\[ \text{HOBr}(aq) \rightleftharpoons \text{H}^+(aq) + \text{OBr}^-(aq) \quad K_a = 2.3 \times 10^{-9} \]

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 \( 1.8 \times 10^{-5} \text{M} \).

(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} \text{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.
c. \[ \text{HOBr} + \text{OH}^- \rightarrow \text{OBr}^- + \text{H}_2\text{O} \]

\[
0.05\text{mol}\text{Ba(OH)}_2 \times 1.4 \text{mol}\text{HOBr} \times 1 \text{mol}\text{OH}^- \times 1 \text{mol}\text{Ba(OH)}_2 \times \frac{1 \text{L}}{1 \text{L}} = 0.07 \text{mol}\text{Ba(OH)}_2
\]

\[
0.0413 \text{L} \text{Ba(OH)}_2
\]

ii. When an acid is mixed with a base, the result is a neutralization reaction.

Since the molar ratio equals 1:1, the reaction will go to completion.

When all the HOBr is reacted in the acid-base reaction, the OBr\(^-\) reacts in the solution as follows:

\[
\text{HOBr} + \text{H}_2\text{O} \rightarrow \text{OBr}^- + \text{H}_2\text{O}
\]

Since OBr\(^-\) is a weak base, some hydroxide ion will be present in the solution, making it basic and having a pH above 7.

d. \[ \text{HOBr} + \text{H}_2\text{O} \rightarrow \text{OBr}^- + \text{H}_2\text{O} \]

\[
1.25 \text{L} \times 1.00 \text{M} \times 0.00 \times 10^{-9} = 2.3 \times 10^{-9}
\]

\[
K_a = \left(1.00 \times 10^{-9}\right) = 2.3 \times 10^{-9}
\]

e. \[ \text{HOBr} \quad \text{H}_2\text{O} \xrightarrow{\text{OBr}^-} \text{Br}^- + \text{H}^+ \]

[\( x = [\text{OBr}^-] = 1.92 \times 10^{-9} \) M]

Both substances are weak acids. An acid will be in

**Substances** since the Br \(-\) atom is more electrophilic.

Since both are Br \(-\) ions, the electrophilic \(-\) factor is the same. These extra O atoms in H\(_2\)O will pull the electron cloud to the O\(^-\) more towards O\(^-\). The O\(^-\) will be more negative and easier to remove, thus the \(-\) will attach \(-\) more and more easily and with a slight dissociation.

GO ON TO THE NEXT PAGE.
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(e) HOBr is a weaker acid than HBrO\(_3\). Account for this fact in terms of molecular structure.

\[
\begin{align*}
\text{(a)} & \quad \rho H = 4.75 = -\log [H^+] \\
\quad [H^+] &= 1.1 \times 10^{-5} \text{ M} \\

\text{(b)} & \quad K_a = \frac{[OBr^-][H^+]}{[HOBr]} \\
\quad [H^+] &= [OBr^-] = 1.8 \times 10^{-5} \text{ M} \\
\quad 2.3 \times 10^{-7} &= \left( \frac{1.8 \times 10^{-5} \text{ M}}{\text{HOBr}} \right)^2 \\
\quad [\text{HOBr}] &= 7.1 \text{ M}
\end{align*}
\]
c) i) \( (0.146 \text{ M})(0.0650 \text{ L}) = 9.49 \times 10^{-3} \text{ mol} \) \( \text{H}^+ \)

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

ii) The pH is greater than 7 since \( \text{Ba(OH)}_2 \)

is a strong base. A strong base and a weak acid give a pH greater than 7 at the equivalence point. The \( \text{OH}^- \) left after reaction will act as a weak base to make a basic solution.

d)

\[ x = \text{concentration of Na}_2\text{OBr} \]

\[
\begin{array}{ccc}
\text{H}_2\text{OBr} & \text{H}^+ & \text{OBr}^- \\
I & 0.160 \text{ M} & 0 & x \\
C & 5.00 \times 10^{-9} \text{ M} & 5.00 \times 10^{-9} \text{ M} & 5.00 \times 10^{-9} \text{ M} \\
E & 0.160 \text{ M} & 5.00 \times 10^{-9} \text{ M} & 5.00 \times 10^{-9} \text{ M} + x \\
\end{array}
\]

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

\[ x = 0.074 \text{ M} \]

\[ \text{OBr}^- = 0.074 \text{ M Na}_2\text{OBr} \]

\[ (0.074 \text{ M})(0.125 \text{ L}) = 9.3 \times 10^{-3} \text{ mol of Na}_2\text{OBr} \]

2) \[
\begin{array}{c}
\text{H}^+ \quad \text{O}^- \\
\text{O} \quad \text{Br}^- \\
\text{H} \quad \text{O} \\
\end{array}
\]

Stronger acids dissociate more than weaker acids.

In \( \text{H}_2\text{OBr} \) there is a \( \text{Br}^- \) atom that is negatively charged by the polar bond with \( \text{O} \). This attracts the positive \( \text{H}^+ \) atom, thus keeping the molecule from dissociating.

In \( \text{H}_3\text{O}_2 \) for the same reason, there are three \( \text{O} \) atoms that are positively charged and repel the positive \( \text{H}^+ \) atom. This makes \( \text{H}_3\text{O}_2 \) dissociate more than \( \text{H}_2\text{OBr} \).
CHEMISTRY
Section II
(Total time—90 minutes)

Part A
Time—40 minutes

YOU MAY USE YOUR CALCULATOR FOR PART A.

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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.

\[
\begin{align*}
\text{a) } & \quad \text{pH} = -\log [\text{H}^+] \\
& \frac{[\text{H}^+]}{[\text{H}^+] + [\text{OH}^-]} = 10^{-\text{pH}} \\
& \text{[H}\(^+\)] = 1.12 \times 10^{-5} \text{ M} \\
\text{b) } & \quad K_a = \frac{[\text{OBr}^-][\text{H}^+]}{[\text{HOBr}]} \\
& K_a = 2.3 \times 10^{-9} = \frac{[\text{OBr}^-][\text{H}^+]}{[\text{HOBr}]} \\
& [\text{HOBr}] = 0.14 \text{ M}
\end{align*}
\]

GO ON TO THE NEXT PAGE.
\[ i - 0.115 \text{ M} \times V = 0.146 \text{ M} \times \left( \frac{0.05 \text{ L}}{1000 \text{ L}} \right) \]
\[ V = 0.08245 \text{ L} \]
\[ V = 82.5 \text{ mL} \]

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

d) \[ K_a = \frac{\text{[OH}^-\text{][H}^+]}{\text{[HOBr]}} \]
\[ \frac{2.3 \times 10^{-9} \text{ [HOBr]}}{[\text{H}^+]}} = [\text{OH}^-] \]
\[ [\text{OH}^-] = \frac{2.3 \times 10^{-9} (0.125)}{5.0 \times 10^{-9}} \]
\[ [\text{OH}^-] = 0.0736 \text{ M} \]

\[ \text{moles} = M(V) \]
\[ \text{moles OBr}^- = 0.0736(0.125 \text{ L}) \]
\[ \text{moles OBr}^- = 9.2 \times 10^{-3} \]

e) \( \text{HBrO}_3 \) is a stronger acid than \( \text{HOBr} \) because the bonds in \( \text{HBrO}_3 \) are covalent while those in \( \text{HOBr} \) are ionic.