Non Sibi High School

Andover's Chem 550/580: Advanced Chemistry

Chapter 17, Review Quiz 1 Answers

1

A 65 mL sample of HBr gas, measured at 35°C and 722 mmHg, was dissolved in water to yield 275 mL of solution. Calculate the molarity of hydrogen ion, the molarity of hydroxide ion, pH, and pOH in this solution.

$$\mathrm{n} = \frac{\frac{722}{760}\,\mathrm{atm} \times \frac{65}{1000}\,\mathrm{L}}{0.0821\,\frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}} \times (35+273)\,\mathrm{K}} = 0.00244\,\mathrm{mol\,HBr}$$

$$[\mathrm{HBr}]_\mathrm{i} = \frac{0.00244\,\mathrm{mol}}{\frac{275}{1000}\,\mathrm{L}} = 0.0089\,\mathrm{M}$$

HBr = strong acid:

R)	HBr(aq)	$\longrightarrow$	H^+(aq)	+	Br^-(aq)
I)	0.0089		0		0
C)	-0.0089		+0.0089		+0.0089
E)	0		0.0089		0.0089

$$[\mathrm{H^+}] = 0.0089\,\mathrm{M}$$

$$[\mathrm{OH^-}] = \frac{1.0\times10^{-14}}{0.0089} = 1.1\times10^{-12}\,\mathrm{M}$$

$$\mathrm{pH} = -\log(0.0089) = 2.05$$

$$\mathrm{pOH} = 14.00 - 2.05 = 11.95$$

2

A 0.086 gram sample of strontium hydroxide was dissolved in water to create 58 mL of solution. Calculate the molarity of hydroxide ion, the molarity of hydrogen ion, pOH, and pH in this solution.

$$0.086\,\mathrm{g}\left(\frac{1\,\mathrm{mol}}{121.6\,\mathrm{g}}\right) = 7.07\times10^{-4}\,\mathrm{mol\,Sr(OH)_2}$$

$$[\mathrm{Sr(OH)_2}]_\mathrm{i} = \frac{7.07 \times 10^{-4}\,\mathrm{mol}}{\frac{58}{1000}\,\mathrm{L}} = 0.012\,\mathrm{M}$$

$Sr(OH)_2$ = strong base:

$$\begin{array}{ccccccc}
\mathrm{R)} & \mathrm{Sr(OH)_2(aq)} & \longrightarrow & \mathrm{Sr^{2+}(aq)} & + & \mathrm{2OH^-(aq)} \\
\mathrm{I)} & 0.012 & & 0 & & 0 \\
\mathrm{C)} & -0.012 & & +0.012 & & +2(0.012) \\
\mathrm{E)} & 0 & & 0.012 & & 0.024
\end{array}$$

$$[\mathrm{OH^-}] = 0.024\,\mathrm{M}$$

$$[\mathrm{H^+}] = \frac{1.0 \times 10^{-14}}{0.024} = 4.2 \times 10^{-13}\,\mathrm{M}$$

$$\mathrm{pOH} = -\log(0.024) = 1.62$$

$$\mathrm{pH} = 14.00 - 1.62 = 12.38$$

3

Write the acid ionization equation and calculate the pH and percent ionization of 0.63 M hypobromous acid, HBrO ($K_a = 2.5 \times 10^{-9}$).

$$\begin{array}{cccccc}
\mathrm{R)} & \mathrm{HBrO(aq)} & \rightleftharpoons & \mathrm{H^+(aq)} & + & \mathrm{BrO^-(aq)} \\
\mathrm{I)} & 0.63 & & 0 & & 0 \\
\mathrm{C)} & -\mathrm{x} & & +\mathrm{x} & & +\mathrm{x} \\
\mathrm{E)} & 0.63 - \mathrm{x} & & \mathrm{x} & & \mathrm{x}
\end{array}$$

$$\mathrm{K_a} = 2.5 \times 10^{-9} = \frac{\mathrm{x}^2}{0.63 - \mathrm{x}}$$

$$(0 < \mathrm{x} < 0.63)$$

$$\mathrm{x} = 4.0 \times 10^{-5}\,\mathrm{M} = [\mathrm{H^+}]$$

$$\mathrm{pH} = -\log(4.0 \times 10^{-5}) = 4.40$$

$$\%\,\mathrm{ionization} = \frac{4.0 \times 10^{-5}}{0.63} \times 100\% = 0.0063\%$$

4

A 0.85 M lactic acid solution has a pH of 1.97. Write the acid ionization equation and calculate K_a for lactic acid, $HC_3H_5O_3$.

R)	$HC_3H_5O_3(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_3H_5O_3{}^-(aq)$
I)	0.85		0		0
C)	-x		+x		+x
E)	0.85 - x		x		x

$$[H^+] = 10^{-1.97} = 0.011\,M = x$$

$$K_a = \frac{(0.011)^2}{(0.85 - 0.011)} = 1.4 \times 10^{-4}$$

5

A 0.017 M solution of propanoic acid is 2.7% ionized. Write the acid ionization equation and calculate the pH of the solution and K_a for propanoic acid, $HC_3H_5O_2$.

R)	$HC_3H_5O_2(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_3H_5O_2{}^-(aq)$
I)	0.017		0		0
C)	-x		+x		+x
E)	0.017 - x		x		x

$$\frac{x}{0.017} \times 100\% = 2.7\%$$

$$x = 4.6 \times 10^{-4}\,M = [H^+]$$

$$pH = -\log(4.6 \times 10^{-4}) = 3.34$$

$$K_a = \frac{(4.6 \times 10^{-4})^2}{(0.0.017 - 4.6 \times 10^{-4})} = 1.3 \times 10^{-5}$$

6

A cyanic acid, HCNO, solution has a pH of 2.25. Given that $K_a = 3.5 \times 10^{-4}$ for cyanic acid, write the acid ionization equation and calculate the initial molarity of the cyanic acid solution.

R)	$HCNO(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$CNO^-(aq)$
I)	M_i		0		0
C)	-x		+x		+x
E)	M_i - x		x		x

$$[H^+] = 10^{-2.25} = 0.0056\,M = x$$
$$K_a = 3.5 \times 10^{-4} = \frac{(0.0056)^2}{(M_i - 0.0056)}$$
$$M_i = 0.095\,M$$

7

Write the stepwise acid ionization equations and calculate the pH of 0.23 M selenous acid, H_2SeO_3, which has the following acid ionization constants:

$$Ka_1 = 2.3 \times 10^{-3}$$
$$Ka_2 = 5.3 \times 10^{-9}$$

1st ionization:

R)	$H_2SeO_3(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$HSeO_3{}^-(aq)$
I)	0.23		0		0
C)	-x		+x		+x
E)	0.23 - x		x		x

$$Ka_1 = 2.3 \times 10^{-3} = \frac{x^2}{0.23 - x}$$
$$(0 < x < 0.23)$$
$$x = 0.022\,M = [H^+]$$
$$pH = -\log(0.022) = 1.66$$

2nd ionization:

$$HSeO_3{}^-(aq) \rightleftharpoons H^+(aq) + SeO_3{}^{2-}(aq)$$

8

Identify the Bronsted acids and bases in the forward and reverse directions for the reaction below:

$$H_2C_6H_5O_7{}^-(aq) + HC_3H_2O_4{}^-(aq) \rightleftharpoons HC_6H_5O_7{}^{2-}(aq) + H_2C_3H_2O_4(aq)$$

forward reaction:
Bronsted acid = $H_2C_6H_5O_7{}^-$ (donates proton), Bronsted base = $HC_3H_2O_4{}^-$ (accepts proton)

reverse reaction:
Bronsted acid = $H_2C_3H_2O_4$, Bronsted base = $HC_6H_5O_7{}^{2-}$

9

Write the formula for:

a. the conjugate acid of $HC_2O_4{}^-$
b. the conjugate base of $HAsO_4{}^{2-}$

a. conjugate acid = add H^+ to formula = $H_2C_2O_4$
b. conjugate base = remove H^+ from formula = $AsO_4{}^{3-}$

10

Write the base ionization equation and calculate the pH and percent ionization of 0.44 M dimethylamine, $(CH_3)_2NH$ ($K_b = 5.4 \times 10^{-4}$).

R)	$(CH_3)_2NH(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$(CH_3)_2NH_2{}^+(aq)$	+	$OH^-(aq)$
I)	0.44				0		0
C)	-x				+x		+x
E)	0.44 - x				x		x

$$K_b = 5.4 \times 10^{-4} = \frac{x^2}{0.44 - x}$$
$$(0 < x < 0.44)$$
$$x = 0.015\,M = [OH^-]$$
$$pOH = -\log(0.015) = 1.82$$
$$pH = 14.00 - 1.82 = 12.18$$
$$\%\,\text{ionization} = \frac{0.015}{0.44} \times 100\% = 3.4\%$$

11

A 0.084 M codeine solution has a pH of 10.46. Write the base ionization equation and calculate K_b for codeine, $C_{18}H_{21}O_3N$.

R)	$C_{18}H_{21}O_3N(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$C_{18}H_{21}O_3NH^+(aq)$	+	$OH^-(aq)$
I)	0.084				0		0
C)	-x				+x		+x
E)	0.084 - x				x		x

$$pOH = 14.00 - 10.46 = 3.54$$
$$[OH^-] = 10^{-3.54} = 2.9 \times 10^{-4}\,M = x$$
$$K_b = \frac{(2.9 \times 10^{-4})^2}{(0.084 - 2.9 \times 10^{-4})} = 1.0 \times 10^{-6}$$

12

A 0.077 M solution of piperidine is 12% ionized. Write the base ionization equation and calculate the pH of the solution and K_b for piperidine, $C_5H_{11}N$.

R)	$C_5H_{11}N(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$C_5H_{11}NH^+(aq)$	+	$OH^-(aq)$
I)	0.077				0		0
C)	-x				+x		+x
E)	0.077 - x				x		x

$$\frac{x}{0.077} \times 100\% = 12\%$$

$$x = 0.0092\,M = [OH^-]$$

$$pOH = -\log(0.0092) = 2.04$$

$$pH = 14.00 - 2.04 = 11.96$$

$$K_b = \frac{(0.0092)^2}{(0.077 - 0.0092)} = 1.2 \times 10^{-3}$$

13

A quinoline, C_9H_7N, solution has a pH of 9.00. Given that $K_b = 6.3 \times 10^{-10}$ for quinoline, write the base ionization equation and calculate the initial molarity of the quinoline solution.

R)	$C_9H_7N(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$C_9H_7NH^+(aq)$	+	$OH^-(aq)$
I)	M_i				0		0
C)	-x				+x		+x
E)	M_i - x				x		x

$$pOH = 14.00 - 9.00 = 5.00$$

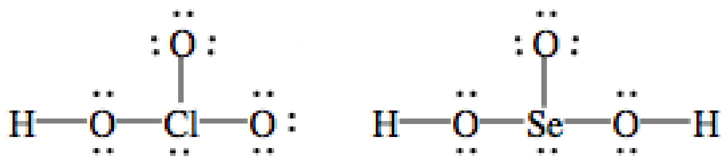
$$[OH^-] = 10^{-5.00} = 1.0 \times 10^{-5}\,M = x$$

$$K_b = 6.3 \times 10^{-10} = \frac{(1.0 \times 10^{-5})^2}{(Mi - 1.0 \times 10^{-5})}$$

$$M_i = 0.16\,M$$

14

Draw Lewis structures for chloric acid, $HClO_3$, and selenous acid, H_2SeO_3. Which is the stronger acid? Give two reasons to justify your answer.

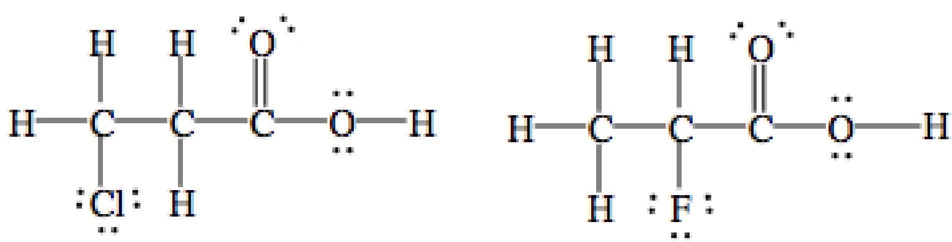


Chloric acid is stronger acid because:

1. $HClO_3$ has more terminal oxygens (3 - 1 = 2) than H_2SeO_3 (3 - 2 = 1).
2. Electronegativity for central atom Cl is higher than for central atom Se.

15

Which of the two acids shown below is the stronger acid? Give two reasons to justify your answer.



The acid on the right is stronger because:

1. Electronegativity of F is higher than Cl.
2. F is closer to ionizable proton than Cl.

16

Predict whether a solution of each compound below will be acidic, basic, or neutral. For solutions that are not neutral, show all relevant hydrolysis reactions that affect the pH and also calculate the equilibrium constant for each reaction you write using information from the following data tables:

Acid	K_a
HCN	4.9×10^{-10}
HIO	2.3×10^{-11}

Base	K_b
$(CH_3)_3N$	6.4×10^{-5}
NH_3	1.8×10^{-5}

a. $(CH_3)_3NHCl$ [composed of $(CH_3)_3NH^+$ and Cl^-]
b. KCN
c. NaI
d. NH_4IO

a. Cl^- = spectator ion

$(CH_3)_3NH^+$ hydrolyzes as weak acid = solution is acidic:

$$(CH_3)_3NH^+(aq) \rightleftharpoons H^+(aq) + (CH_3)_3N(aq)$$
$$K_a = \frac{1.0 \times 10^{-14}}{6.4 \times 10^{-5}} = 1.6 \times 10^{-10}$$

b. K^+ = spectator ion

CN^- hydrolyzes as a weak base = solution is basic:

$$CN^-(aq) + H_2O(l) \rightleftharpoons HCN(aq) + OH^-(aq)$$
$$K_b = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5}$$

c. Na^+ and I^- = spectator ions = solution is neutral

d. Both ions hydrolyze, so we must compare equilibrium constants:

$$NH_4{}^+(aq) \rightleftharpoons H^+(aq) + NH_3(aq)$$
$$K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$IO^-(aq) + H_2O(l) \rightleftharpoons HIO(aq) + OH^-(aq)$$
$$K_b = \frac{1.0 \times 10^{-14}}{2.3 \times 10^{-11}} = 4.3 \times 10^{-4}$$
$$K_a < K_b = \text{solution is basic}$$

17

For each solution below, show any relevant hydrolysis reactions and calculate the pH.

a. 0.31 M $C_6H_5NH_3Br$ [composed of $C_6H_5NH_3{}^+$ and Br^-]
b. 1.2 M $KC_3H_5O_3$

$K_b = 4.3 \times 10^{-10}$ for $C_6H_5NH_2$
$K_a = 1.4 \times 10^{-4}$ for $HC_3H_5O_3$

a. Br^- = spectator ion

$C_6H_5NH_3{}^+$ hydrolyzes as weak acid:

R)	$C_6H_5NH_3{}^+(aq)$	$\rightleftharpoons$	H^+	+	$C_6H_5NH_2(aq)$
I)	0.31		0		0
C)	-x		+x		+x
E)	0.31 - x		x		x

$$K_a = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-10}} = 2.3 \times 10^{-5} = \frac{x^2}{0.31 - x}$$
$$(0 < x < 0.31)$$
$$x = 0.0027\,M = [H^+]$$
$$pH = -\log(0.0027) = 2.57$$

b. K^+ = spectator ion

$C_3H_5O_3{}^-$ hydrolyzes as a weak base:

R)	$C_3H_5O_3{}^-(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$HC_3H_5O_3(aq)$	+	$OH^-(aq)$
I)	1.2				0		0
C)	-x				+x		+x
E)	1.2 - x				x		x

$$K_b = \frac{1.0 \times 10^{-14}}{1.4 \times 10^{-4}} = 7.1 \times 10^{-11} = \frac{x^2}{1.2 - x}$$
$$(0 < x < 1.2)$$
$$x = 9.2 \times 10^{-6}\,M = [OH^-]$$
$$pOH = -\log(9.2 \times 10^{-6}) = 5.04$$
$$pH = 14.00 - 5.04 = 8.96$$

18

Predict whether a solution of sodium hydrogen arsenate, NaH_2AsO_4, will be acidic or basic. Show all relevant reactions that affect the pH and also give the value of the equilibrium constant for each reaction you write. For arsenic acid, H_3AsO_4:

$$Ka_1 = 5.6 \times 10^{-3}$$
$$Ka_2 = 1.0 \times 10^{-7}$$
$$Ka_3 = 3.0 \times 10^{-12}$$

Na^+ = spectator ion

weak acid ionization reaction:

$$H_2AsO_4{}^-(aq) \rightleftharpoons H^+(aq) + HAsO_4{}^{2-}$$

$$K_a \text{ for } H_2AsO_4{}^- = Ka_2 \text{ for } H_3AsO_4 = 1.0 \times 10^{-7}$$

weak base hydrolysis reaction:

$$H_2AsO_4{}^-(aq) + H_2O(l) \rightleftharpoons H_3AsO_4(aq) + OH^-(aq)$$

$$K_b \text{ for } H_2AsO_4{}^- = \frac{1.0 \times 10^{-14}}{Ka_1 \text{ for } H_3AsO_4 = 5.6 \times 10^{-3}} = 1.8 \times 10^{-12}$$

$$K_a > K_b = \text{solution is acidic}$$

19

Will the reaction of P_4O_{10} and water produce H_3PO_3 or H_3PO_4? Write a balanced equation for the reaction.

The reaction of P_4O_{10} (oxidation number of P = +5) and water will produce H_3PO_4 (oxidation number of P = +5) rather than H_3PO_3 (oxidation number of P = +3). The balanced equation will be $P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4$.

20

a. An unknown monoprotic weak acid was found to be 30.60% carbon, 45.16% chlorine, and 3.85% hydrogen by mass, with the remainder being oxygen. Determine the empirical formula of the acid.

b. In a separate experiment, 3.75 grams of the acid was dissolved in 45 mL of water and then titrated with 0.164 M barium hydroxide. The volume of base required to reach the equivalence point was 72.8 mL. Calculate the molar mass and determine the molecular formula of the acid.

a.

$$100\% - 30.60\%\,\mathrm{C} - 45.16\%\,\mathrm{Cl} - 3.85\%\,\mathrm{H} = 20.39\%\,\mathrm{O}$$

Assume one hundred grams of unknown compound:

$$30.60\,\mathrm{g\,C}\left(\frac{1\,\mathrm{mol}}{12.01\,\mathrm{g}}\right) = 2.548\,\mathrm{mol\,C}$$

$$45.16\,\mathrm{g\,I}\left(\frac{1\,\mathrm{mol}}{35.45\,\mathrm{g}}\right) = 1.274\,\mathrm{mol\,Cl}$$

$$3.85\,\mathrm{g\,H}\left(\frac{1\,\mathrm{mol}}{1.008\,\mathrm{g}}\right) = 3.82\,\mathrm{mol\,H}$$

$$20.39\,\mathrm{g\,O}\left(\frac{1\,\mathrm{mol}}{16.00\,\mathrm{g}}\right) = 1.274\,\mathrm{mol\,O}$$

$$\frac{2.548}{1.274}\,\mathrm{mol\,C} : \frac{1.274}{1.274}\,\mathrm{mol\,Cl} : \frac{3.82}{1.274}\,\mathrm{mol\,H} : \frac{1.274}{1.274}\,\mathrm{mol\,O}$$

$$\text{empirical formula} = \mathrm{C_2ClH_3O}$$

b. $0.164\,\mathrm{M\,Ba(OH)_2} \longrightarrow 0.164\,\mathrm{M\,Ba^{2+}}$ and $2 \times 0.164\,\mathrm{M} = 0.328\,\mathrm{M\,OH^-}$

$$\mathrm{HA(aq) + OH^-(aq) \longrightarrow H_2O(l) + A^-(aq)}$$

$$72.8\,\mathrm{mL}\left(\frac{1\,\mathrm{L}}{1000\,\mathrm{mL}}\right)\left(\frac{0.328\,\mathrm{mol\,OH^-}}{1\,\mathrm{L}}\right)\left(\frac{1\,\mathrm{mol\,HA}}{1\,\mathrm{mol\,OH^-}}\right) = 0.02388\,\mathrm{mol\,HA}$$

$$M = \frac{3.75\,\mathrm{g}}{0.02388\,\mathrm{mol}} = 157\,\mathrm{g/mol}$$

$$\frac{M}{EM} = \frac{157}{78.49} = 2$$

$$\text{molecular formula} = \mathrm{C_2ClH_3O} \times 2 = \mathrm{C_4Cl_2H_6O_2}$$

Since the acid is monoprotic, we can rewrite the molecular formula as $\mathrm{HC_4Cl_2H_5O_2}$.

21

Calculate the initial molarity of a sodium fluoride, NaF, solution that has a pH of 8.17 given that $K_a = 6.8 \times 10^{-4}$ for hydrofluoric acid, HF.

Na^+ =spectator ion, F^- hydrolyzes as a weak base:

R)	$F^-(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$HF(aq)$	+	$OH^-(aq)$
I)	M_i				0		0
C)	-x				+x		+x
E)	M_i - x				x		x

$$pOH = 14.00 - 8.17 = 5.83$$

$$[OH^-] = 10^{-5.83} = 1.5 \times 10^{-6}\,M = x$$

$$K_b \text{ for } F^- = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11} = \frac{(1.5 \times 10^{-6})^2}{(Mi - 1.5 \times 10^{-6})}$$

$$M_i = 0.15\,M$$

22

Calculate the initial molarity of a $C_5H_5NHNO_3$ [composed of $C_5H_5NH^+$ and $NO_3{}^-$] solution that has a pH of 2.83 given that $K_b = 1.7 \times 10^{-9}$ for C_5H_5N.

$NO_3{}^-$ =spectator ion, $C_5H_5NH^+$ hydrolyzes as a weak acid:

R)	$C_5H_5NH^+(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_5H_5N(aq)$
I)	M_i		0		0
C)	-x		+x		+x
E)	M_i - x		x		x

$$[H^+] = 10^{-2.83} = 0.0015\,M = x$$

$$K_a \text{ for } C_5H_5NH^+ = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6} = \frac{(0.0015)^2}{(Mi - 0.0015)}$$

$$M_i = 0.38\,M$$

23

Given that a 0.72 M KIO solution has a pH of 12.24, calculate K_a and pK_a for HIO.

K^+ =spectator ion, IO^- hydrolyzes as a weak base:

R)	$IO^-(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$HIO(aq)$	+	$OH^-(aq)$
I)	0.72				0		0
C)	-x				+x		+x
E)	0.72 - x				x		x

$$\mathrm{pOH} = 14.00 - 12.24 = 1.76$$
$$[\mathrm{OH^-}] = 10^{-1.76} = 0.017\,\mathrm{M} = \mathrm{x}$$
$$\mathrm{K_b\;for\;IO^-} = \frac{(0.017)^2}{(0.72 - 0.017)} = 4.1 \times 10^{-4}$$
$$\mathrm{K_a\;for\;HIO} = \frac{1.0 \times 10^{-14}}{4.1 \times 10^{-4}} = 2.4 \times 10^{-11}$$
$$\mathrm{pK_a\;for\;HIO} = -\log(2.4 \times 10^{-11}) = 10.62$$

24

Given that a 0.066 M $C_6H_{15}O_3NHCl$ [composed of $C_6H_{15}O_3NH^+$ and Cl^-] solution has a pH of 4.48, calculate K_b and pK_b for $C_6H_{15}O_3N$.

Cl^- =spectator ion, $C_6H_{15}O_3NH^+$ hydrolyzes as a weak acid:

R)	$C_6H_{15}O_3NH^+(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_6H_{15}O_3N(aq)$
I)	0.066		0		0
C)	-x		+x		+x
E)	0.066 - x		x		x

$$[\mathrm{H^+}] = 10^{-4.48} = 3.3 \times 10^{-5}\,\mathrm{M} = \mathrm{x}$$

$$\mathrm{K_a\ for\ C_6H_{15}O_3NH^+} = \frac{(3.3 \times 10^{-5})^2}{(0.066 - 3.3 \times 10^{-5})} = 1.7 \times 10^{-8}$$

$$\mathrm{K_b\ for\ C_6H_{15}O_3N} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-8}} = 5.9 \times 10^{-7}$$

$$\mathrm{pK_b\ for\ C_6H_{15}O_3N} = -\log(5.9 \times 10^{-7}) = 6.23$$

25

Calculate the pH of a solution containing 4.2 grams of $NaC_7H_5O_2$ in 75 mL of 0.27 M $HC_7H_5O_2$ ($K_a = 6.3 \times 10^{-5}$).

Na^+= spectator ion, $HC_7H_5O_2/C_7H_5O_2{}^-$= buffer solution:

$$4.2\,\mathrm{g\,NaC_7H_5O_2}\left(\frac{1\,\mathrm{mol}}{144.1\,\mathrm{g}}\right) = 0.029\,\mathrm{mol\,NaC_7H_5O_2} \longrightarrow 0.029\,\mathrm{mol\,C_7H_5O_2{}^-}$$

$$75\,\mathrm{mL}\left(\frac{1\,\mathrm{L}}{1000\,\mathrm{mL}}\right)\left(\frac{0.27\,\mathrm{mol\,HC_7H_5O_2}}{1\,\mathrm{L}}\right) = 0.020\,\mathrm{mol\,HC_7H_5O_2}$$

$$\mathrm{pK_a\ for\ HC_7H_5O_2} = -\log(6.3 \times 10^{-5}) = 4.20$$

$$\mathrm{pH} = 4.20 + \log\frac{0.029\,\mathrm{mol\,C_7H_5O_2{}^-}}{0.020\,\mathrm{mol\,HC_7H_5O_2}} = 4.36$$

26

Calculate the pH of a solution containing 2.4 grams of $(CH_3)_2NH_2I$ [composed of $(CH_3)_2NH_2{}^+$ and I^-] in 84 mL of 0.18 M $(CH_3)_2NH$ ($K_b = 5.4 \times 10^{-4}$).

I^-= spectator ion, $(CH_3)_2NH/(CH_3)_2NH_2{}^+$= buffer solution:

$$2.4\,\mathrm{g\,(CH_3)_2NH_2I}\left(\frac{1\,\mathrm{mol}}{173.0\,\mathrm{g}}\right) = 0.014\,\mathrm{mol\,(CH_3)_2NH_2I} \longrightarrow 0.014\,\mathrm{mol\,(CH_3)_2NH_2{}^+}$$

$$84\,\mathrm{mL}\left(\frac{1\,\mathrm{L}}{1000\,\mathrm{mL}}\right)\left(\frac{0.18\,\mathrm{mol\,(CH_3)_2NH}}{1\,\mathrm{L}}\right) = 0.015\,\mathrm{mol\,(CH_3)_2NH}$$

$$\mathrm{pK_b\ for\ (CH_3)_2NH} = -\log(5.4 \times 10^{-4}) = 3.27$$

$$\mathrm{pOH} = 3.27 + \log\frac{0.014\,\mathrm{mol\,(CH_3)_2NH_2{}^+}}{0.015\,\mathrm{mol\,(CH_3)_2NH}} = 3.24$$

$$\mathrm{pH} = 14.00 - 3.24 = 10.76$$

27

Write the net ionic equation for the neutralization reaction that occurs and calculate the pH when:

a. 0.002 mol NaOH is added to the solution in Question 25
b. 0.002 mol HBr is added to the solution in Question 25
c. 0.002 mol NaOH is added to the solution in Question 26
d. 0.002 mol HBr is added to the solution in Question 26

a. Na^+= spectator ion

$$\text{neutralization reaction: } HC_7H_5O_2(aq) + OH^-(aq) \longrightarrow C_7H_5O_2{}^-(aq) + H_2O(l)$$

$$pH = 4.20 + \log \frac{(0.029 + 0.002)\,\text{mol}\, C_7H_5O_2{}^-}{(0.020 - 0.002)\,\text{mol}\, HC_7H_5O_2} = 4.44$$

b. Br^-= spectator ion

$$\text{neutralization reaction: } C_7H_5O_2{}^-(aq) + H^+(aq) \longrightarrow HC_7H_5O_2(aq)$$

$$pH = 4.20 + \log \frac{(0.029 - 0.002)\,\text{mol}\, C_7H_5O_2{}^-}{(0.020 + 0.002)\,\text{mol}\, HC_7H_5O_2} = 4.29$$

c. Na^+= spectator ion

$$\text{neutralization reaction: } (CH_3)_2NH_2{}^+(aq) + OH^-(aq) \longrightarrow (CH_3)_2NH(aq) + H_2O(l)$$

$$pOH = 3.27 + \log \frac{(0.014 - 0.002)\,\text{mol}\,(CH_3)_2NH_2{}^+}{(0.015 + 0.002)\,\text{mol}\,(CH_3)_2NH} = 3.12$$
$$pH = 14.00 - 3.12 = 10.88$$

d. Br^-= spectator ion

$$\text{neutralization reaction: } (CH_3)_2NH(aq) + H^+(aq) \longrightarrow (CH_3)_2NH_2{}^+(aq)$$

$$pOH = 3.27 + \log \frac{(0.014 + 0.002)\,\text{mol}\,(CH_3)_2NH_2{}^+}{(0.015 - 0.002)\,\text{mol}\,(CH_3)_2NH} = 3.36$$
$$pH = 14.00 - 3.36 = 10.64$$

28

Write the net ionic equation for the neutralization reaction that occurs in each aqueous mixture below. Which one of the four reactions creates a buffer solution? For each mixture, describe the method of calculating the pH of the resulting solution after the neutralization reaction is complete.

a. 0.15 mol HNO_3 + 0.15 mol NaCN
b. 0.25 mol HIO + 0.25 mol KOH
c. 0.15 mol NH_4Br + 0.25 mol NaOH
d. 0.15 mol HI + 0.25 mol $C_6H_5NH_2$ (a weak base)

a. ${NO_3}^-$ and Na^+= spectator ions

$$\text{neutralization reaction}: \; CN^-(aq) + H^+(aq) \longrightarrow HCN(aq)$$

$$n_{CN^-} = n_{H^+}$$

HCN solution created, pH calculated using RICE chart for ionization of weak acid HCN

b. K^+= spectator ion

$$\text{neutralization reaction}: \; HIO(aq) + OH^-(aq) \longrightarrow IO^-(aq) + H_2O(l)$$

$$n_{HIO} = n_{OH^-}$$

IO^- solution created, pOH = 14.00 - pH calculated using RICE chart for hydrolysis of weak base IO^-

c. Br^- and Na^+= spectator ions

$$\text{neutralization reaction}: \; {NH_4}^+(aq) + OH^-(aq) \longrightarrow NH_3(aq) + H_2O(l)$$

$$n_{{NH_4}^+} < n_{OH^-}$$

OH^-/NH_3 solution created, pOH = 14.00 - pH = -log $[OH^-]_{excess}$ (ionization of NH_3 negligible)

d. I^-= spectator ion

$$\text{neutralization reaction}: \; C_6H_5NH_2(aq) + H^+(aq) \longrightarrow C_6H_5{NH_3}^+(aq)$$

$$n_{C_6H_5NH_2} > n_{H^+}$$

$$\mathbf{C_6H_5NH_2/C_6H_5{NH_3}^+ \;\text{buffer created}}$$

$$\mathbf{pOH = pK_b + \log\frac{n_{C_6H_5{NH_3}^+}\;\text{produced}}{n_{C_6H_5NH_2}\;\text{excess}} = 14.00 - pH}$$

29

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 22 mL of 0.40 M HNO_3 with 0.20 M NaOH and calculate the volume of base needed to reach the equivalence point.

b. Calculate the following:

i. the initial pH
ii. the pH after 31 mL of base has been added
iii. the pH at the equivalence point
iv. the pH after 49 mL of base has been added

a. $NO_3{}^-$ and Na^+= spectator ions

$$\text{neutralization reaction}: \ H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

$$0.022\,\text{L}\left(\frac{0.40\,\text{mol}\,H^+}{1\,\text{L}}\right) = 0.0088\,\text{mol}\,H^+ = 0.0088\,\text{mol}\,OH^-\left(\frac{1\,\text{L}}{0.20\,\text{mol}\,OH^-}\right)$$
$$= 0.044\,\text{L} = 44\,\text{mL base needed}$$

b.
i.

$$\text{pH} = -\log(0.40) = 0.40$$

ii.

$$0.031\,\text{L}\left(\frac{0.20\,\text{mol}\,OH^-}{1\,\text{L}}\right) = 0.0062\,\text{mol}\,OH^-\ \text{added}$$

$$\text{pH} = -\log\frac{(0.0088 - 0.0062)\,\text{mol}\,H^+\ \text{excess}}{(0.022 + 0.031)\,\text{L total}} = 1.31$$

iii. only H_2O present, pH = 7
iv.

$$0.049\,\text{L}\left(\frac{0.20\,\text{mol}\,OH^-}{1\,\text{L}}\right) = 0.0098\,\text{mol}\,OH^-\ \text{added}$$

$$\text{pOH} = -\log\frac{(0.0098 - 0.0088)\,\text{mol}\,OH^-\ \text{excess}}{(0.022 + 0.049)\,\text{L total}} = 1.85$$
$$\text{pH} = 14.00 - 1.85 = 12.15$$

30

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 24 mL of 0.30 M $HC_3H_5O_2$ ($K_a = 1.3 \times 10^{-5}$) with 0.20 M KOH and calculate the volume of base needed to reach the equivalence point.

b. Calculate the following:

i. the initial pH
ii. the pH after 23 mL of base has been added
iii. the pH at the equivalence point

iv. the pH after 43 mL of base has been added

a. K^+= spectator ion

$$\text{neutralization reaction}: \ HC_3H_5O_2(aq) + OH^-(aq) \longrightarrow C_3H_5O_2{}^-(aq) + H_2O(l)$$

$$0.024\,L\left(\frac{0.30\,mol\,HC_3H_5O_2}{1\,L}\right) = 0.0072\,mol\,HC_3H_5O_2 = 0.0072\,mol\,OH^-\left(\frac{1\,L}{0.20\,mol\,OH^-}\right)$$
$$= 0.036\,L = 36\,mL\,\text{base needed}$$

b.
i.

R)	$HC_3H_5O_2(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_3H_5O_2{}^-(aq)$
I)	0.30		0		0
C)	-x		+x		+x
E)	0.30 - x		x		x

$$K_a \text{ for } HC_3H_5O_2 = 1.3 \times 10^{-5} = \frac{x^2}{0.30 - x}$$
$$x = 0.0020\,M = [H^+]$$
$$pH = -\log(0.0020) = 2.70$$

ii.

$$0.023\,L\left(\frac{0.20\,mol\,OH^-}{1\,L}\right) = 0.0046\,mol\,OH^-\text{ added}$$
$$= 0.0046\,mol\,HC_3H_5O_2\text{ reacted} = 0.0046\,mol\,C_3H_5O_2{}^-\text{ produced}$$
$$0.0072\,mol\,HC_3H_5O_2\text{ initial} - 0.0046\,mol\,HC_3H_5O_2\text{ reacted} = 0.0026\,mol\,HC_3H_5O_2\text{ excess}$$
$$pK_a \text{ for } HC_3H_5O_2 = -\log(1.3 \times 10^{-5}) = 4.89$$
$$pH = 4.89 + \log\frac{0.0046\,mol\,C_3H_5O_2{}^-}{0.0026\,mol\,HC_3H_5O_2} = 5.14$$

iii.

R)	$C_3H_5O_2{}^-(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$HC_3H_5O_2(aq)$	+	$OH^-(aq)$
I)	M_i				0		0
C)	-x				+x		+x
E)	M_i - x				x		x

$$M_i = \frac{n_{C_3H_5O_2{}^-}}{L\,\text{total}} = \frac{n_{OH^-}\text{ added}}{L\,\text{total}} = \frac{0.0072\,mol}{(0.024 + 0.036)\,L} = 0.12\,M$$

$$\mathrm{K_b\ for\ C_3H_5O_2}^- = \frac{\mathrm{K_w}}{\mathrm{K_a\ for\ HC_3H_5O_2}} = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10} = \frac{\mathrm{x}^2}{0.12 - \mathrm{x}}$$

$$\mathrm{x} = 9.6 \times 10^{-6}\,\mathrm{M} = [\mathrm{OH}^-]$$

$$\mathrm{pOH} = -\log(9.6 \times 10^{-6}) = 5.02$$

$$\mathrm{pH} = 14.00 - 5.02 = 8.98$$

iv.

$$0.043\,\mathrm{L}\left(\frac{0.20\,\mathrm{mol\,OH^-}}{1\,\mathrm{L}}\right) = 0.0086\,\mathrm{mol\,OH^-\ added}$$

$$\mathrm{pOH} = -\log\frac{(0.0086 - 0.0072)\,\mathrm{mol\,OH^-\ excess}}{(0.024 + 0.043)\,\mathrm{L\ total}} = 1.68$$

$$\mathrm{pH} = 14.00 - 1.68 = 12.32$$

31

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 48 mL of 0.10 M C_5H_5N ($K_b = 1.7 \times 10^{-9}$) with 0.30 M HI and calculate the volume of acid needed to reach the equivalence point.

b. Calculate the following:

i. the initial pH
ii. the pH after 12 mL of acid has been added
iii. the pH at the equivalence point
iv. the pH after 27 mL of acid has been added

a. I^-= spectator ion

$$\mathrm{neutralization\ reaction:}\ \ \mathrm{C_5H_5N(aq) + H^+(aq) \longrightarrow C_5H_5NH^+(aq)}$$

$$0.048\,\mathrm{L}\left(\frac{0.10\,\mathrm{mol\,C_5H_5N}}{1\,\mathrm{L}}\right) = 0.0048\,\mathrm{mol\,C_5H_5N} = 0.0048\,\mathrm{mol\,H^+}\left(\frac{1\,\mathrm{L}}{0.30\,\mathrm{mol\,H^+}}\right)$$

$$= 0.016\,\mathrm{L} = 16\,\mathrm{mL\ acid\ needed}$$

b.
i.

R)	$C_5H_5N(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$C_5H_5NH^+(aq)$	+	$OH^-(aq)$
I)	0.10				0		0
C)	-x				+x		+x
E)	0.10 - x				x		x

$$\mathrm{K_b\ for\ C_5H_5N} = 1.7 \times 10^{-9} = \frac{\mathrm{x}^2}{0.10 - \mathrm{x}}$$

$$\mathrm{x} = 1.3 \times 10^{-5}\,\mathrm{M} = [\mathrm{OH^-}]$$

$$\mathrm{pOH} = -\log(1.3 \times 10^{-5}) = 4.89$$

$$\mathrm{pH} = 14.00 - 4.89 = 9.11$$

ii.

$$0.012\,\mathrm{L}\left(\frac{0.30\,\mathrm{mol\,H^+}}{1\,\mathrm{L}}\right) = 0.0036\,\mathrm{mol\,H^+\ added}$$

$$= 0.0036\,\mathrm{mol\,C_5H_5N\ reacted} = 0.0036\,\mathrm{mol\,C_5H_5NH^+\ produced}$$

$$0.0048\,\mathrm{mol\,C_5H_5N\ initial} - 0.0036\,\mathrm{mol\,C_5H_5N\ reacted} = 0.0012\,\mathrm{mol\,C_5H_5N\ excess}$$

$$\mathrm{pK_b\ for\ C_5H_5N} = -\log(1.7 \times 10^{-9}) = 8.77$$

$$\mathrm{pOH} = 8.77 + \log\frac{0.0036\,\mathrm{mol\,C_5H_5NH^+}}{0.0012\,\mathrm{mol\,C_5H_5N}} = 9.25$$

$$\mathrm{pH} = 14.00 - 9.25 = 4.75$$

iii.

R)	$C_5H_5NH^+(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_5H_5N(aq)$
I)	M_i		0		0
C)	-x		+x		+x
E)	M_i - x		x		x

$$\mathrm{M}_i = \frac{\mathrm{n_{C_5H_5NH^+}}}{\mathrm{L\ total}} = \frac{\mathrm{n_{H^+}\ added}}{\mathrm{L\ total}} = \frac{0.0048\,\mathrm{mol}}{(0.048 + 0.016)\,\mathrm{L}} = 0.075\,\mathrm{M}$$

$$\mathrm{K_a\ for\ C_5H_5NH^+} = \frac{\mathrm{K_w}}{\mathrm{K_b\ for\ C_5H_5N}} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6} = \frac{\mathrm{x}^2}{0.075 - \mathrm{x}}$$

$$\mathrm{x} = 6.6 \times 10^{-4}\,\mathrm{M} = [\mathrm{H^+}]$$

$$\mathrm{pH} = -\log(6.6 \times 10^{-4}) = 3.18$$

iv.

$$0.027\,\mathrm{L}\left(\frac{0.30\,\mathrm{mol\,H^+}}{1\,\mathrm{L}}\right) = 0.0081\,\mathrm{mol\,H^+\ added}$$

$$\mathrm{pH} = -\log\frac{(0.0081 - 0.0048)\,\mathrm{mol\,H^+\ excess}}{(0.048 + 0.027)\,\mathrm{L\ total}} = 1.36$$

32

Which indicator, bromphenol blue ($K_a = 1 \times 10^{-4}$) or phenolphthalein ($K_a = 5 \times 10^{-10}$), would be the better choice for the titration in Question 30?

pH at equivalence point = 8.98
pK_a for bromphenol blue $= -\log(1 \times 10^{-4}) = 4.0$
pK_a for phenolphthalein $= -\log(5 \times 10^{-10}) = 9.3$
pK_a for phenolphthalein close to pH at equivalence point, so phenolphthalein is better choice

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