84.122- Section 204 Quiz 8 Name KEY
Dr. N. De Luca March 31, 2016

Assume all solutions are aqueous at 25°C. $K_w = 1.0 \times 10^{-14}$

1. A 0.10M solution of H_2SO_4 will have a pH (circle your choice): (5 pts)

   a) equal to 1.00
   b) equal to .70
   c) slightly greater than 1.00
   [✓ d) slightly less than 1.00]

   diprotic K_a, very large
   K_{a_2} } weak acid $\sim 10^{-2}$

2. Consider aqueous solutions of each of the following salts. Indicate if they will be acidic, basic or neutral. For each non-neutral solution, write the net-ionic chemical reaction that affects the pH. Be sure to include physical states (l, aq, etc.), any charges on ions, and double or single arrows as needed. (15 pts)

	Acidic, basic or neutral	Chemical reaction, if non-neutral
a) NH4Cl	acidic	$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftarrows NH_{3(aq)} + H_3O^+_{(aq)}$
b) NaF	basic	$F^-_{(aq)} + H_2O_{(l)} \rightleftarrows HF_{(aq)} + OH^-_{(aq)}$
c) LiI	neutral	

3. Calculate the pH of a 1.60 M KBrO solution. K_a for hypobromous acid, HBrO, is 2.0×10^{-9}. Be sure to include a chemical reaction, ICE table, equilibrium constant expression and your method to receive credit. (10 pts)

	$BrO^-_{(aq)}$ +	H_2O ⇌	$HBrO_{(aq)}$ +	$OH^-_{(aq)}$
I	1.60		0	0
C	$-x$		$+x$	$+x$
E	$1.60-x$		x	x

$$K_b = \frac{[HBrO][OH^-]}{[BrO^-]}$$

$$= \frac{K_w}{K_a \text{ for } HBrO}$$

$$= \frac{1.0\times10^{-14}}{2.0\times10^{-9}} = 5.0\times10^{-6}$$

$$\frac{(x)(x)}{1.60-\cancel{x}} = 5.0\times10^{-6}$$

assume x is small bc K is small

$$x^2 = 5.0\times10^{-6}(1.60) = 8.0\times10^{-6}$$

$$x = 2.8\times10^{-3}$$

$1.60 - 2.8\times10^{-3} = 1.60$ ✓

$[OH^-] = 2.8\times10^{-3}$

$pOH = 2.55$; $pH = 14.00 - 2.55$ Ans. pH = 11.45

4. Calculate the pH of a solution that contains 0.058 M HF and 0.070 M LiF. The Ka for HF is 3.5×10^{-4}. Clearly show your method, (including chemical reactions, an ICE table, etc. or a clear method) for partial credit. (10 pts)

buffer solution

	$HF_{(aq)}$ +	$H_2O_{(l)}$ ⇌	$H_3O^+_{(aq)}$ +	$F^-_{(aq)}$
I	.058		0	.070
C	$-x$		$+x$	$+x$
E	$.058-x$		x	$.070+x$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$3.5\times10^{-4} = \cancel{.058}\ \frac{(x)(.070+x)}{.058-x}$$

assume x will be small, so $.070+x \cong .070$, $.058-x \cong .058$

$.070 + .00029 = .070$ ✓

$.058 - 2.9\times10^{-4} = .058$ ✓

$$x = \frac{3.5\times10^{-4}(.058)}{.070} = 2.9\times10^{-4}$$

$x = [H_3O^+] = 2.9\times10^{-4}$, $pH = -\log[H_3O^+]$

$= 3.54$

Henderson-Hasselbalch alternate method

$$pH = pK_a + \log\frac{[F^-]}{[HF]}$$

$$= -\log 3.5\times10^{-4} + \log\frac{.070}{.058}$$

$$= 3.456 + .0817 = 3.538 = 3.54$$

Ans. pH = 3.54