1.) pH of the 0.02M solution of potassium hypoiodate (KOI) is 11.44. Calculate K_a of hypoiodic acid (HOI). (12 pts)

weak base Spectator

$$KOI + H_2O_{(l)} \rightleftharpoons HOI + OH^- + K^+$$

	KOI		HOI	OH^-
i	0.02		0	0
C	$-x$		$+x$	$+x$
e	$0.02-x$		x	x

$$K_b = \frac{(x)(x)}{0.02-x}$$

$$K_b = \frac{(2.75\times10^{-3})^2}{(0.02-2.75\times10^{-3})}$$

$$K_b = 4.40\times10^{-4}$$

$pH = 11.44$

$14 = pH + pOH$

$2.56 = pOH$

$2.56 = -\log[OH^-]$

$10^{-2.56} = [OH^-]$

$2.75\times10^{-3} = [OH^-] = x$

$KOI \leftrightarrow HOI$

weak base conjugate acid

so

$$K_a \cdot K_b = K_w$$

$$K_a \cdot 4.4\times10^{-4} = 1\times10^{-14}$$

$$K_a = 2.27\times10^{-11}$$

2.) How many mL of a 0.2 M HCl must be added to 50 mL of 0.3M $Ba(OH)_2$ to get a pH of 7.00?) (12 points)

Start w/ reaction equation

$HCl + Ba(OH)_2 \rightarrow H_2O + Cl^- + Ba^{2+}$

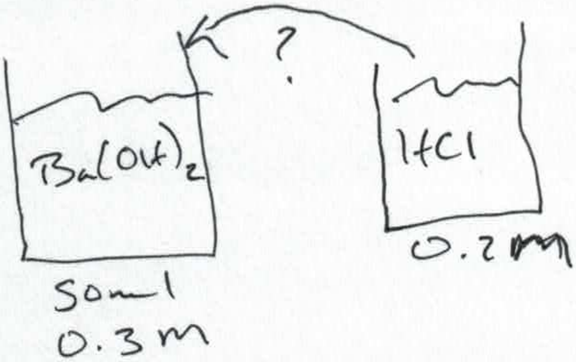
Spectators

Balance equation

$2HCl + Ba(OH)_2 \rightarrow 2H_2O + 2Cl^- + Ba^{2+}$

Strong acid + strong base makes pH=7 water

need 2 HCl molecules to neutralize one $Ba(OH)_2$



How many mmols of $Ba(OH)_2$ in beaker?

mmol $Ba(OH)_2$ = 0.3M · 50ml = 15 mmols

need 2 HCl for every one $Ba(OH)_2$ ~~Sor~~ So
need 30 mmol HCl

$$\text{ml HCl} = \cancel{\frac{0.2M}{30mmol}}\ \frac{30\ \text{mmol}}{0.2M} = \underline{\underline{150\ \text{ml}}}$$

3.) Calculate a change in pH when 10 mL of 0.2 M LiOH is added to the 100 mL of buffer containing 0.1 M hydrobromous acid (HOBr) and 0.2 M of sodium hydrobromide (NaOBr). K_a of HOBr = 2.0×10^{-9}. (15 points)

First realize LiOH is a strong base and will react ~~s~~ with the acid HOBr

spectator

Next write chemical equation

$$HOBr + LiOH \rightarrow OBr^- + H_2O + Li^+$$

So for every one mmol of LiOH added the buffer loses one mmol of HOBr and gains one mmol of OBr^-

initial mmol HOBr = 0.1M · 100ml = 10 mmol

initial mmol OBr^- = 0.2M · 100ml = 20 mmol

mmol LiOH added = 0.2M · 10 ml = 2 mmol

final mmol HOBr = 10 − 2 = 8 mmol

final mmol OBr^- = 20 + 2 = 22 mmol

Final pH:

$$pH = pKa + \log \frac{\text{base}}{\text{acid}}$$

$$pH = -\log(2.0 \times 10^{-9}) + \log \frac{22}{8}$$

$$pH = 9.14$$

4.) Titration: This is a multistep problem.

Calculate the pH for the titration of 100 mL of 0.10 M NH_3, $K_b = 1.8 \times 10^{-5}$, with 0.25 M HBr.

a) (10 points) Before any acid is added.

b) (12 points) After 30 mL of acid is added.

c) (15 points) After 40 mL of acid is added

d) (12 points) After 60 mL of acid is added

a.) pH = 11.13

b.) pH = 8.78

c.) pH = 5.20

d.) pH = 1.51

5.) 5. Multiple choice (4 questions, 12 pts)

A.) In the unusual acid-base reaction $HClO_4 + H_2SO_4 \rightleftharpoons ClO_4^- + H_3SO_4^+$, one of the conjugate acid-base pairs is:

a. $HClO_4$, H_2SO_4
b. H_2SO4, ClO_4^-
c. ClO_4-, $H_3SO_4^+$
d. $H_3SO_4^+$, H_2SO4
e. $HClO_4$, $H_3SO_4^+$

B.) Consider the following Bronsted-Lowry acid-base reaction:
$C_6H_5NH_2 + HCOOH \rightleftharpoons C_6H_5NH_3^+ + HCOO^-$
Which of the following is a conjugate acid-base pair?

a. $C_6H_5NH_2$ and HCO_2H
b. $HCOOH$ and $C_6H_5NH_3^+$
c. $HCOO^-$ and $C_6H_5NH_2$
d. $C_6H_5NH_3^+$ and $C_6H_5NH_2$
e. $C_6H_5NH_3^+$ and $HCOO^-$

C.) Which aqueous solution(s) below is(are) considered *basic*?
I. A solution with a pH = 6
II. A solution with $[OH^-] = 1x10^{-7}$
III. A solution with $[H_3O^+] = 1x10^{-9}$

a. I only
b. II only
c. III only
d. I and II
e. I and III

D.) After completing the following *unfinished* Base Ionization table, what is the order of *increasing base strength*?

Base	K_b	pK_b
Bicarbonate ion, HCO_3^-	2.2×10^{-8}	
Pyridine, C_5H_5N		8.77
Acetate ion, CH_3COO^-	5.7×10^{-10}	

a. (weakest) Bicarbonate ion < Pyridine < Acetate ion (strongest)
b. (weakest) Bicarbonate ion < Acetate ion < Pyridine (strongest)
c. (weakest) Pyridine < Bicarbonate ion < Acetate ion (strongest)
d. (weakest) Acetate ion < Pyridine < Bicarbonate ion (strongest)
e. (weakest) Acetate ion < Bicarbonate ion < Pyridine (strongest