

CHM 152 Group Work 5 (over Ch. 15) Crib

- Calculate the pH of a 2.00 L solution containing 0.885 moles of hypochlorous acid (HClO) and 0.905 moles of NaClO. Given K_a for HClO is 3.0×10^{-8} .

What is in the beaker? A weak acid HClO, and its conjugate base, ClO^- ions from NaClO. (Na^+ ions are spectators)

So we have a buffer and can use the buffer equation. So we need the concentrations of these in the beaker.

$$\text{HClO } .885 \text{ moles} / 2.00\text{L} = 0.4425\text{M HClO (acid)}$$

$$\text{NaClO } .905 \text{ moles} / 2.00\text{L} = 0.4525\text{M NaClO that dissociates 100\% so} = 0.4525\text{M } \text{ClO}^- \text{ (c. base)}$$

$$\text{pH} = 7.5229 + \log(0.4525 / 0.4425) = 7.53 \text{ (need 2 decimal places since } K_a \text{ had two sig dig)}$$

- Give the formulas for two chemicals that would make a buffer solution in water. **Like HF and KF**
- Calculate the pH when 25.0 mL of 0.100M HBr is added to 15.0 mL of 0.100M LiOH.

Strong acid and strong base. Reacts one way.



Need moles of each. acid: $0.0250\text{L} (0.100 \text{ mol} / \text{L}) = 0.00250 \text{ moles acid}$

base: $0.0150\text{L} (0.100\text{mol} / \text{L}) = 0.00150 \text{ moles base}$

Set up initial final table

HBr	+ LiOH	$\rightarrow \text{H}_2\text{O (l)}$	+ LiBr
0.00250 moles	0.00150 moles	---	0
- 0.00150, not all gone = excess	all reacts, limiting - 0.00150	---	+ 0.00150
.00100 mol	0	---	0.00150 moles

NOT a buffer by the way!!! LiBr is a neutral salt, not a conjugate base.

Note the new volume is 40.0 mL.

pH will depend on the strong acid left over not the neutral salt. So HBr dissociates 100%.

Thus $[\text{H}^+] = 0.00100 \text{ moles} / 0.0400\text{L} = 0.0250\text{M}$

$$\text{pH} = 1.602 \text{ (final answer needs 3 decimal places since everything had three sig dig)}$$

- How many milliliters of 0.95M sodium hydroxide must be added to 35.0 mL of 0.85M acetic acid to reach the equivalence point? What is the pH at the equivalence point? Given: K_a for acetic acid is 1.8×10^{-5}

Find the volume of NaOH by stoichiometry:

$$0.0350\text{L} (0.85 \text{ mol/L}) (1 \text{ NaOH} / 1 \text{ acid}) (1000\text{mL} / 0.95\text{M}) = 31.316 \text{ mL} = V_b = 31.3 \text{ mL}$$

First they react together one way since NaOH is strong. Set up an initial final table. Calculate moles of each. Note they are equal since we are at the equivalence point cause nothing is in excess at the equivalence point, only product salt exists in the beaker.

NaOH +	CH_3COOH	$\rightarrow \text{H}_2\text{O (l)}$	+ NaCH_3COO
0.02975 moles	0.02975 moles	---	0

all reacts	all reacts	---	+0.02975
0	0	---	0.02975 moles

Now what happens? No acid left, no base left = equivalence point!!! We have only product. But this salt is not neutral - it contains the conjugate base acetate ion. Basic ions react in water just like any base. We need the molarity of acetate ion. Note the new volume of 66.316 mL.

The basic salt NaCH_3COO will dissolve completely leaving 0.02975 moles sodium ion and 0.02975 moles acetate ion. Acetate ion is basic and will react further. Sodium ions are neutral and will not react further. We must put concentrations in ICE tables, so we need the molarity of the acetate ion.

M CH_3COO^- is 0.02975 moles / 0.066316 L = 0.4486 M Set up an ICE table for the C. base reacting with water.

$\text{H}_2\text{O (l) +}$	CH_3COO^-	$\rightleftharpoons \text{OH}^-$	$+ \text{CH}_3\text{COOH}$
---	0.4486 M	0	0
---	-x	+x	+x
---	0.4486 - x	x	x

This is a base reaction, need K_b . Get it from K_w / K_a . $K_b = 5.556 \times 10^{-10} = x^2 / 0.4486$
 $x = 1.5787 \times 10^{-5} \text{ M}$ (note I'm not rounding anything till the final answer)
 pOH = 4.80 so **pH = 9.20** (two decimal places since the M given have two sig dig)

5. Solid calcium fluoride is added to 1.00 liter of pure water. After several hours of stirring, some of the solid remains undissolved. If the concentration of the calcium ions is $7.2 \times 10^{-5} \text{ M}$, calculate the solubility product (K_{sp}).

	$\text{CaF}_2(\text{s})$	\rightleftharpoons	$\text{Ca}^{2+}(\text{aq})$	$+ 2 \text{F}^-(\text{aq})$	
I	--		0	0	
C	-x		x	2x	
E	---		x	2x	

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = x(2x)^2 = 4x^3$$

$$x = 7.2 \times 10^{-5}$$

So substitute and **$K_{sp} = 1.5 \times 10^{-12}$**

6. Will the solubility of barium carbonate increase, decrease, or remain the same if solid barium nitrate is added to a saturated solution of barium carbonate? Explain and support with appropriate chemical reactions.

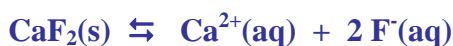
$\text{BaCO}_3(\text{s}) \rightleftharpoons \text{Ba}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$ When $\text{Ba}(\text{NO}_3)_2$ added it is soluble so it completely dissolves. Adding barium ions shifts the equilibrium rxn left toward the solid. So the solubility and the amount of dissolved BaCO_3 **decreases**.

7. Why is magnesium hydroxide considered a weak base? **It does not ionize 100%**

8. In lab Sally adds 0.0244 grams of solid calcium fluoride to 1.50 liters of pure water and stirs vigorously. For calcium fluoride $K_{sp} = 4.0 \times 10^{-11}$. Will the resulting solution be unsaturated, just saturated, or will solid be present? Show your work for full credit.

$0.0244 \text{ g CaF}_2 (\text{mol} / 78.08 \text{ g}) = 3.125 \times 10^{-4} \text{ mol}$ then divide by 1.50L = $2.083 \times 10^{-4} \text{ M CaF}_2$ if it all actually dissolves.

If all of the solid dissolves there will be $[\text{Ca}^{2+}] = 2.083 \times 10^{-4} \text{ M}$ and twice that for $[\text{F}^-] = 4.167 \times 10^{-4} \text{ M}$



So $Q = [\text{Ca}^{2+}][\text{F}^{-}]^2 = 3.62 \times 10^{-11}$ which is less than K_{sp} so **unsaturated**

9. If you mix 225.0 mL of 0.015 M aqueous lead(II) nitrate with 125.0 mL of 0.045 M aluminum bromide, does a precipitate form? K_{sp} for $\text{PbBr}_2 = 6.9 \times 10^{-6}$. Must show your work mathematically by calculating no guessing.



We do not know if enough PbBr_2 was made to pass the saturation point, so we will calculate Q based on the Pb and Br ion concentrations.

Since soluble the $[\text{Pb}^{2+}] = .015 \text{M Pb}(\text{NO}_3)_2 (1 \text{ Pb}^{2+} / 1 \text{ Pb}(\text{NO}_3)_2) = 0.015 \text{ M Pb}^{2+}$ originally

Since soluble $[\text{Br}^{-}] = \frac{0.045 \text{ mol AlBr}_3}{1 \text{ L}} \times \frac{3 \text{ mol Br}^{-}}{1 \text{ mol AlBr}_3} = 0.135 \text{ M Br}^{-}$ originally

Now these two solutions were added together, thus diluted, with a final volume of 350.0 mL

In the final mixture: $[\text{Pb}^{2+}] = \left(\frac{0.015 \text{ M Pb}^{2+} \times 225.0 \text{ ml}}{350.0 \text{ ml}} \right) = 9.643 \times 10^{-3} \text{ M Pb}^{2+}$

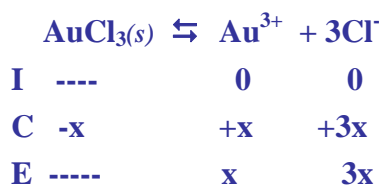
In the final mixture: $[\text{Br}^{-}] = \left(\frac{0.135 \text{ M Br}^{-} \times 125.0 \text{ ml}}{350.0 \text{ ml}} \right) = 4.821 \times 10^{-2} \text{ M Br}^{-}$



$$Q = [\text{Pb}^{2+}][\text{Br}^{-}]^2 = (9.643 \times 10^{-3})(4.821 \times 10^{-2})^2$$

$$Q = 2.2 \times 10^{-5} \quad Q > K_{\text{sp}}, \text{ so } \text{PbBr}_2 \text{ does precipitate}$$

10. Calculate the solubility of $\text{AuCl}_3(\text{s})$ in pure water. K_{sp} for $\text{AuCl}_3 = 3.2 \times 10^{-25}$.



$$K_{\text{sp}} = [\text{Au}^{3+}][\text{Cl}^{-}]^3 = (x)(3x)^3$$

$$3.2 \times 10^{-25} = 27x^4$$

$$x^4 = 1.185 \times 10^{-26}$$

$$x = 3.299 \times 10^{-7} \text{ M} = \text{molar solubility of } \text{AuCl}_3 \text{ in water}$$

$$3.299 \times 10^{-7} \text{ mol/L} (303.32 \text{ g/mol}) = 1.0 \times 10^{-4} \text{ g/L}$$

11. Calculate the molar solubility of $\text{AuCl}_3(\text{s})$ in a solution containing 0.075 M CaCl_2 . Explain the difference between the solubility of AuCl_3 in question 10 and 11.



We are interested in the common ion so $[\text{Cl}^{-}] = \frac{0.075 \text{ mol CaCl}_2}{1 \text{ L}} \times \frac{2 \text{ mol Cl}^{-}}{1 \text{ mol CaCl}_2} = 0.150 \text{ M Cl}^{-}$



$$K_{\text{sp}} = [\text{Au}^{3+}][\text{Cl}^{-}]^3 \text{ thus } 3.2 \times 10^{-25} = (x)(0.150)^3$$

$$x = 3.2 \times 10^{-25} / .003375 = 9.481 \times 10^{-23}$$

$$x = \text{M} = \text{molar solubility}$$

$$9.481 \times 10^{-23} \text{ mol/L} (303.32 \text{ g/mol}) = 2.9 \times 10^{-20} \text{ g/L}$$

Solubility ↓ due to the common ion Cl^{-} .