

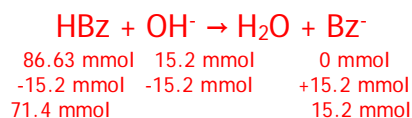
Practice Exam for Exam 2

1. Calculate the **pH of a solution** when 0.365 g of lithium hydroxide are added to 100.0 mL of 0.8663 M benzoic acid.

Do neutralization step first. Find limiting reactant between OH^- and benzoic acid.

$$n_{\text{HBz}} = 100.0 \text{ mL} \times \frac{0.8663 \text{ mol HBz}}{\text{L}} = 86.63 \text{ mmol HBz}$$

$$n_{\text{OH}^-} = 0.365 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.948 \text{ g LiOH}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol LiOH}} \times \frac{1 \text{ mmol OH}^-}{10^{-3} \text{ mol LiOH}} = 15.2 \text{ mmol OH}^-$$



This is a buffer. We can use the Henderson-Hasselbalch equation:

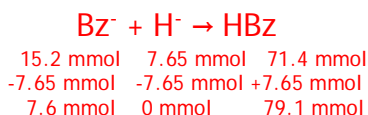
$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{\text{base}}{\text{acid}} = -\log(6.5 \times 10^{-5}) + \log \frac{15.2}{71.4} \\ &= 4.19 + (-0.671) = 3.52 \end{aligned}$$

The pH of the solution is 3.52.

Calculate the **pH of the solution** when 50.0 mL of 0.153 M hydrochloric acid is added to the mixture from above.

Again, because HCl is a strong acid we need to do the neutralization first. The acid will react with the conjugate base Bz^- .

$$n_{\text{H}^+} = 50.0 \text{ mL} \times \frac{0.153 \text{ mol HCl}}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HCl}} = 7.65 \text{ mmol H}^+$$



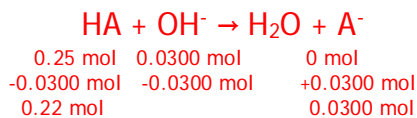
This is right at the limit for a buffer (10:1 or 1:10 ratio).

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{\text{base}}{\text{acid}} = -\log(6.5 \times 10^{-5}) + \log \frac{7.6}{79.1} \\ &= 4.19 + (-1.02) = 3.17 \end{aligned}$$

2. A 0.25 mol sample of a weak acid with an unknown pK_a was combined with 10.0 mL of 3.00 M KOH, and the resulting solution was diluted to 1.500 L. The measured pH of the solution was 3.85. What is the ***pKa of the weak acid?***

Do neutralization step first. Find limiting reactant between OH^- and acid.

$$n_{\text{OH}^-} = 10.0 \text{ mL KOH} \times \frac{3.00 \text{ mol KOH}}{1000 \text{ mL KOH}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol KOH}} = 0.0300 \text{ mol OH}^-$$



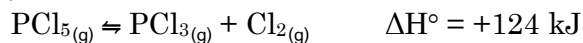
This is a buffer. We can use the Henderson-Hasselbalch equation:

$$pH = pK_a + \log \frac{\text{base}}{\text{acid}} = pK_a + \log \frac{0.0300}{0.22}$$

$$3.85 = pK_a + (-0.865)$$

$$pK_a = 3.85 - (-0.865) = 4.72$$

3. Consider the reaction:



a. A reaction mixture at 453 K initially contains 522 bar of PCl_5 . At equilibrium, the total pressure in the reaction mixture is 748 bar. Calculate K_p at this temperature.

	PCl_5	PCl_3	Cl_2
I	522	0	0
C	-x	+x	+x
E	522-x	x	x

$$P_{\text{total}} = 543 \text{ bar} = P_{\text{PCl}_5} + P_{\text{PCl}_3} + P_{\text{Cl}_2} = (522 - x) + x + x = 522 + x$$

$$x = 543 \text{ bar} - 522 \text{ bar} = 21 \text{ bar}$$

$$K_p = \frac{P_{\text{PCl}_3} P_{\text{Cl}_2}}{P_{\text{PCl}_5}} = \frac{(21)(21)}{(522 - 21)} = 0.88$$

b. A second reaction mixture at 175 K initially contains 255 bar of PCl_3 and 168 bar Cl_2 . What is the equilibrium **partial pressure of PCl_5** in this mixture?

	PCl_5	PCl_3	Cl_2
I	0	255	168
C	+x	-x	-x
E	x	255-x	168-x

$$K_p = \frac{P_{\text{PCl}_5} P_{\text{Cl}_2}}{P_{\text{PCl}_3}} = \frac{(x)(168 - x)}{(255 - x)} = 0.88$$

$$42840 - 423x + x^2 = 0.88x$$

$$42840 - 423.88x + x^2 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} : a = 1; b = -423.88; c = 42840$$

$$x = 166$$

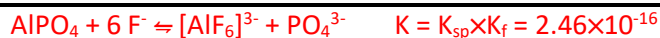
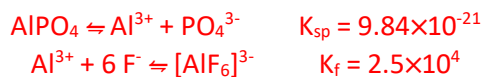
$$P_{\text{PCl}_5} = 166 \text{ bar}$$

c. How will the **equilibrium pressure of PCl_5 change** if the temperature is decreased?

Because this reaction is exothermic, if the temperature is decreased then K will also decrease which means that the reaction will favor the reactants more so the equilibrium pressure of PCl_5 will increase.

4. A solution is prepared so that the initial concentration of aluminum nitrate is 0.10 M, the initial concentration of sodium fluoride is 1.50 M, and the initial concentration of sodium phosphate is 2.00 M. Assume the complex ion formed contains the highest number of ligands to the central metal ion.

- a. Does a precipitate form from this solution?



Assume all the Al^{3+} is initially in the complex ion. This means that the F^- concentration becomes 0.90 M.

$$Q = \frac{[\text{AlF}_6]^{3-}][\text{PO}_4^{3-}]}{[\text{F}^-]^6} = \frac{(0.10)(2.00)}{(0.90)^6} = 0.376 > K \quad \text{so a precipitate does form}$$

- b. What is the equilibrium concentration of aluminum ions in the solution?

Because Q is so much larger than K we will go to completion first in the reverse direction (because a precipitate forms) then go back to equilibrium in the forward direction.

	F^-	$[\text{AlF}_6]^{3-}$	PO_4^{3-}
I	0.90	0.10	2.00
completion	+0.60	-0.10	-0.10
C	-6x	+x	+x
E	1.50-6x	x	1.90-x

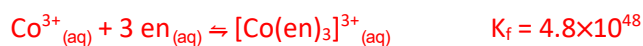
$$K = 2.46 \times 10^{-16} = \frac{[\text{AlF}_6]^{3-}][\text{PO}_4^{3-}]}{[\text{F}^-]^6} = \frac{(x)(1.90+x)}{(1.50-6x)^6} \approx \frac{(x)(1.90)}{(1.50)^6}$$

$$x = \frac{(2.46 \times 10^{-16})(1.50)^6}{(1.90)} = 1.4748 \times 10^{-15} \approx [\text{AlF}_6]^{3-}_e$$

$$K_f = \frac{[\text{AlF}_6]^{3-}}{[\text{Al}^{3+}][\text{F}^-]^6} \Rightarrow [\text{Al}^{3+}] = \frac{[\text{AlF}_6]^{3-}}{K_f [\text{F}^-]^6} = \frac{1.4748 \times 10^{-15}}{(2.5 \times 10^4)(1.50)^6} = 5.2 \times 10^{-21} \text{ M}$$

$$\text{or } K_{sp} = [\text{Al}^{3+}][\text{PO}_4^{3-}] \Rightarrow [\text{Al}^{3+}] = \frac{K_{sp}}{[\text{PO}_4^{3-}]} = \frac{9.84 \times 10^{-21}}{1.90} = 5.17 \times 10^{-21} \text{ M}$$

5. What is the **equilibrium concentration of cobalt(III) ion** in a solution that is 0.20 M ethylenediamine (en) and 0.0150M cobalt(III) nitrate.



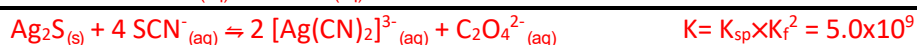
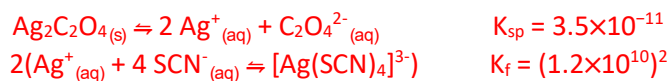
	Co^{3+}	en	$[\text{Co}(\text{en})_3]^{3+}$
I	0.0150	0.20	0.00
Comp	-0.0150	-0.0450	+0.0150
C	+x	+3x	-x
E	x	0.15+3x	0.0150-x

$$K_f = \frac{[\text{Co}(\text{en})_3]^{3+}}{[\text{Co}^{3+}][\text{en}]^3} = \frac{(0.0150 - x)}{(x)(0.155 + 3x)^3} = 4.8 \times 10^{48} \approx \frac{(0.0150)}{(x)(0.155)^3}$$

$$x = \frac{(0.0150)}{(4.8 \times 10^{48})(0.155)^3} = 8.4 \times 10^{-49} \quad \text{approximation is valid}$$

$$x \approx [\text{Co}^{3+}] = 8.4 \times 10^{-49} \text{ M}$$

6. Calculate the **molar solubility of silver oxalate** in 1.00 M sodium thiocyanate.



	SCN^-	$[\text{Ag}(\text{SCN})_2]^{3-}$	$\text{C}_2\text{O}_4^{2-}$
I	1.00	0	0
Comp	-1.00	+0.50	+0.25
C	+4x	-2x	-x
E	4x	0.50-2x	0.25-x

$$K = \frac{[\text{Ag}(\text{SCN})_2]^{3-}]^2 [\text{C}_2\text{O}_4^{2-}]}{[\text{SCN}^-]^4} = \frac{[0.50 - 2x]^2 [0.25 - x]}{[4x]^4} \approx \frac{[0.50]^2 [0.25]}{[4x]^4} = 5.0 \times 10^9$$

$$x = \sqrt[4]{\frac{(0.50)^2 (0.25)}{5.0 \times 10^9 (4^4)}} = 4.7 \times 10^{-4}$$

The molar solubility is the value of $[\text{C}_2\text{O}_4^{2-}]$ which is $0.25 - x$ which is essentially the same as 0.25 mol/L.