1. What is the pH of a 1 L solution made with 3.26 g of trichloroacetic acid and 37.0 g of sodium trichloroacetate? The K_b of sodium trichloroacetate is 10^{-12}. Molecular weights of tichloroacetic acid and sodium trichloroacetate are 163 g/mol and 185 g/mol respectively.

a. 4
b. 3.5
c. 3
d. 2

mols of acid = 3.26/163 = .02
mols of base = 37.0/185 = .2
$pK_a = pK_w - pK_b = 14 - 12 = 2$
$K_a = 10^{-2}$
$[H^+] = 10^{-2} (.02/.2) = 10^{-2} * 10^{-1} = 10^{-3}$
$pH = -\log(10^{-3}) = 3$

2. If 500 ml of each of the following solutions is combined, which will not produce a buffer?

a. .1 M trichloroacetic acid + .1 M sodium trichloroacetate
b. .3 M trichloroacetic acid + .15 M barium hydroxide
c. .3 M trichloroacetic acid + .2 M sodium hydroxide
d. .2 M sodium trichloroacetate + .18 M hydrochloric acid

Option B neutralizes to produce a weak base solution.

3. Consider the following table of acids and their pK_a values. Rank these acids in order of increasing strength of their conjugate bases.

Acid	pK_a
CHOOH	3.77
CH_3COOH	4.75
HF	3.20

a. HF < CH3COOH < CHOOH
b. CHOOH < CH3COOH < HF
c. CH3COOH < CHOOH < HF
d. HF < CHOOH < CH3COOH

The stronger an acid, the weaker its conjugate base. The lower the pK_a, the stronger the acid. So, rank the acids in order of increasing pKas.

4. What volume of 1 M calcium hydroxide could be added to 500 mL of a buffer that is 0.4 M ammonia and 0.6 M ammonium chloride before exhausting the buffer?

a. 600 mL
b. 300 mL
c. 150 mL
d. 400 mL
e. 200 mL

When a strong base such as $Ca(OH)_2$ is added to a buffer, it will react with the acidic species, in this case ammonium. The buffer contains 0.3 moles of ammonium, and this will be exhausted by 150 mL of calcium hydroxide.

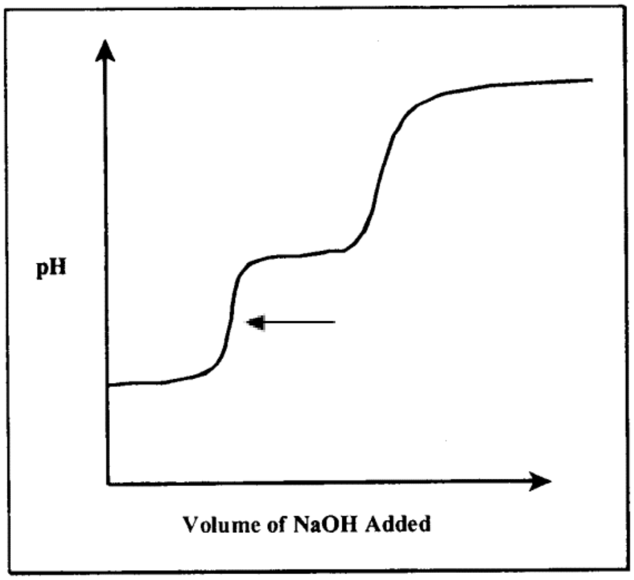
5. What is the pH of a 100 mL of a buffer solution containing .25 M formic acid and .30 M sodium formate if you add .01 mols of $Ba(OH)_2$? The pKa of formic acid is 4.

a. 3.7
b. 4
c. 4.4
d. 5

R HA + OH^- --> A^-

I .025 .02 .03
C -.02 -.02 +.02
E .005 0 .05
$[H^+] = 10^{-4} (.005/.05) = 10^{-5}$
pH = 5

6. Consider the following titration curve of an acid, H_2A:



At the point indicated by the arrow, which species is predominant?

a. H_2A
b. HA^-
c. A^{2-}
d. No species is predominant, $[H_2A] = [HA^-]$

At the first equivalence point, all of H2A has been converted to HA^-.

7. If 25 mL of a .17 M solution of barium hydroxide is titrated with 75 mL of a .1 M nitric acid solution, what is the final pH?

a. 12
b. 1.5
c. 11
d. 7

R H^+ + OH^- --> H_2O
I .0075 .0085 ~
C -.0075 -.0075 ~
E 0 .001 ~
$[OH^-] = .001/.1 = 10^{-2}$
pOH = 2
pH = 12

8. A 100 mL solution of .25 M $(CH_3)_3NHCl$ is titrated with 80 mL of a .25 M solution of NaOH. What is the pH of the final solution? The pK_a of $(CH_3)_3NHCl$ is $4 * 10^{-4}$.

a. 8
b. 3.5
c. 4

d. 2

R	BH^+	+	OH^-	-->	B
I	.025		.020		0
C	-.020		-.020		+.020
E	.005		0		.020

$[H^+] = 4*10^{-4} (.005/.020) = 4*10^{-4}(1/4) = 1 * 10^{-4}$
pH = 4

9. A .36 M solution of ethylamine ($C_2H_5NH_2$) is titrated with a .12 M perchloric acid solution to the stoichiometric point. What is left in solution? What is the molarity? What is the pH range? The pKb of ethylamine is 3.
   a. equal amounts of $C_2H_5NH_2$ and $C_2H_5NH_3^+$, .3 M, 3
   b. equal amounts of $C_2H_5NH_2$ and $C_2H_5NH_3^+$, .18 M, 3
   c. only $C_2H_5NH_3^+$, .18 M, 11-12
   d. only $C_2H_5NH_3^+$, .09 M, 9-10

To reach the stoichiometric (equivalence) point, 3 equivalents of the strong acid solution must be used. At the equivalence point, all that is left in solution is the conjugate acid of ethylamine, and the molarity is 1/4th of what it was originally. The pH should be in the weak base range.

10. Consider the table of salts and their corresponding K_{sp} values.

Salt	K_{sp}
$BaCO_3$	10^{-20}
Cu_3PO_4	10^{-20}
$Ba_3(PO_3)_2$	10^{-20}

Rank the salts in order of increasing solubility in water.
   a. $BaCO_3 < Cu_3PO_4 < Ba_3(PO_3)_2$
   b. $Ba_3(PO_3)_2 < Cu_3PO_4 < BaCO_3$
   c. $Cu_3PO_4 < Ba_3(PO_3)_2 < BaCO_3$
   d. $BaCO_3 < Ba_3(PO_3)_2 < Cu_3PO_4$

Approximate the molar solubility by taking the nth root of the K_{sp}, where n is the number of ions in the salt. The approximate molar solubilities for $BaCO_3$, Cu_3PO_4, and $Ba_3(PO_3)_2$ are 10^{-10}, 10^{-5}, and 10^{-4}, respectively.

11. Which of the following compounds has the highest molar solubility in water?
   a. $Cu_3(PO_4)_2$ ($K_{sp} = 1.9 \times 10^{-37}$)
   b. Bi_2S_3 ($K_{sp} = 1.8 \times 10^{-99}$)
   c. Fe_2S_3 ($K_{sp} = 1.4 \times 10^{-88}$)
   d. PbI_2 ($K_{sp} = 7.1 \times 10^{-9}$)

$Cu_3(PO_4)_2$: $1.9 \times 10^{-37}) = (3x)^3(2x)^2 = 108x^5 = 1.8x10^{-8}$ mol/L
Bi_2S_3: $1.8 \times 10^{-99} = (2x)^2(3x)^3$ $x = 7x10^{-21}$ mol/L
Fe_2S_3: 1.4×10^{-88} $x = 1.1x10^{-18}$ mol/L
PbI_2: $7.1 \times 10^{-9} = 4x^3$ $x = 0.0012$ mol/L

12. Which equation will you use to calculate the molar solubility, x, of $HgCl_2$ in water? ($K_{sp} = 1.1x10^{-18}$)
   a. $1.1x10^{-18} = 4x^3$
   b. $x^3 = 1.1x10^{-18}$
   c. $(1.1x10^{-18})^3 = x$
   d. $1.1x10^{-18} / 3 = x^4$

$HgCl_2 \rightleftharpoons Hg^{+2} + 2Cl^{-1}$

$Ksp = [Hg^{2+}][Cl^-]^2 = [x][2x]^2 = 4x^3 = 1.1x10^{-18}$

13. How many grams of $CaCrO_4$ ($K_{sp} = 2.3x10^{-2}$) could you dissolve in a 100 mL of 2.3 M $Ca(OH)_2$?
   a. 10^{-2}
   b. 10^{-1}
   c. 156
   d. 1.56
   e. 10^{-3}
   f. 0.156

$K_{sp} = [Ca^{2+}][CrO_4^{2-}]$
$2.3x10^{-2} = x(2.3)$
$x = 10^{-2}$ moles/L
mass of $CaCrO_4 = 10^{-2}$ moles/L (0.1 L)(156 g/mole) = 0.156 g of $CaCrO_4$

14. Each of the following answers are the solubility constants for a metal ions. Which one will be the most difficult to separate by precipitation?
   a. 10^{-9}, 10^{-15}
   b. 10^{-5}, 10^{-33}
   c. 10^{-2}, 10^{-3}
   d. 10^{-3}, 10^{-5}

Closest set of Ksp makes them difficult to separate by precipitation

15. When can you use the weak acid approximation?
   a. Always
   b. When K_a is near K_w
   c. When the concentration of acid is smaller than [H+] and Kw is negligible
   d. When K_a and K_w are far apart

16. In an attempt to prepare a large quantity of a solution with a pH of 9, a student dissolves 100.5 micrograms of perchloric acid in 1000 L of water. Roughly, what will be the actual pH of the solution and why?
   a. 7, because the leveling effect will prevent the dissociation of the acid.
   b. 9, because the final concentration of the strong acid will 10^{-9} M.
   c. 9, because the conjugate base of the strong acid will predominate.
   d. 3, because the final concentration of the strong acid will 10^{-3} M.
   e. 7, because the strong acid too dilute to contribute significantly to the $[H^+]$.

The strong acid will dissociate completely, but tis will only result in a small amount of protons relative to that contributed by water itself.

17. How many equations would be need to completely describe a solution of $(NH_4)_3AsO_3$ and NaCl?
   1. 6
   2. 8
   3. 4
   4. 7
   5. 5
   6. 9

The species with unknown concentrations are: $[H^+]$, $[OH^-]$, $[NH_4^+]$, $[NH_3]$, $[AsO_3^{3-}]$, $[HAsO_3^{2-}]$, $[H_2AsO_3^-]$, $[H_3AsO_3]$, $[Na^+]$, $[Cl^-]$.

18. Which of the following equations is correct in describing a solution of $(NH_4)_2CO_3$?
   a. $[NH_4^+] + [H^+] = [OH^-] + [CO_3^{2-}] + [HCO_3^-]$
   b. $2\ C_{(NH_4)_2CO_3} = [NH_4^+] + [NH_3]$

c. $C_{(NH_4)_2CO_3} = 2[NH_4^+] + 2[NH_3]$
d. $[NH_4^+] = 2[CO_3^{2-}] + [HCO_3^-]$

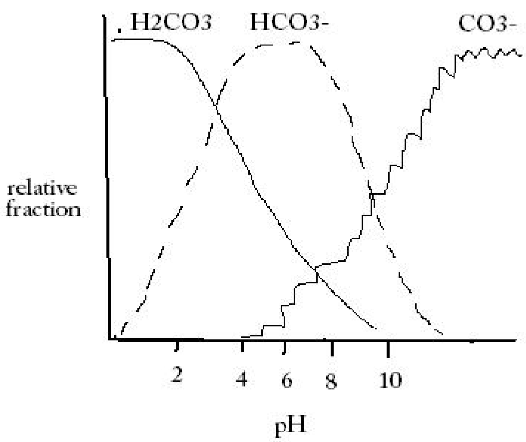
19. Which of the following is a likely pOH for a solution of 10^{-9} M NaOH?
a. 7.02
b. 6.98
c. 7.00
d. 9.00
e. 5.00

The NaOH will contribute a small amount of OH^-, which will make the $[OH^-]$ slightly greater than 10^{-7} and the pOH slightly less than 7.

20. Consider the chemical EDTA, which has 6 acid/base groups on it. If all 6 are initially protonated and we titrate until all 6 are deprotonated, how many buffer regions and amphoteric species will occur?
a. 5, 5
b. 6, 5
c. 6, 6
d. 5, 6

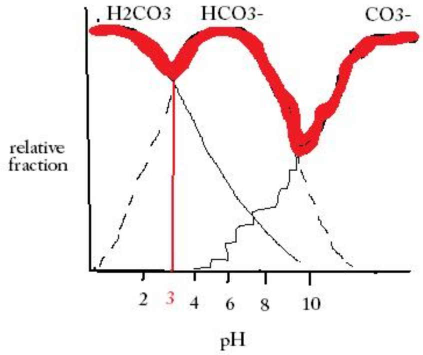
All 6 K_a values can be used to describe a buffer. The product of the first 5 deprotonations (but not the sixth) will be amphoteric.

21. Given the following alpha diagram. What is(are) the predominant specie(s) at pH 3?



a. H_2CO_3
b. H_2CO_3, HCO_3^-
c. HCO_3^-
d. HCO_3^-, CO_3^{2-}

the red bold line shows the predominant species at the different pH: those are the lines with the greatest Y value at that point as you move across the x axis. At pH 3, the greatest y value is where H2CO3 and HCO3- lines intersect. Thus, they are both the dominant species. At pH below 3, H2CO3 is the predominant specie; at pH above 10, CO3--is the dominant species...so on and so forth.



22. The pK_a values for H_3PO_4 are as following: pK_{a1}=2.1, pK_{a2}=7.2, pK_{a3}=12.8. What is the pOH pf a 0.001M HPO_4^{-2}?

a. 10
b. 9
c. 5
d. 4

Since HPO_4^{-2} is amphiprotic, to determine its pOH, we use the amphiprotic equation.
With the following equilibrium equations:

Ka1 H_3PO_4 --> H^+ + $H_2PO_4^-$
Ka2 $H_2PO_4^-$ --> H^+ + **HPO_4^{2-}**
Ka3 **HPO_4^{2-}** --> H^+ + PO_4^{3-}

Since HPO_4^{2-} is found in K_{a2} and K_{a3}, those are the values we plug into the equation.
pH = 1/2 (pK_{a2}+pK_{a3}) = (7.2+12.8)/2 = 10
pOH = 14 -pH = 14-10 = 4

23. What is the pH of a 0.1 M solution of Na_2CO_3? pK_{a1}= 4, pK_{a2}=11.

a. 12
b. 7.5
c. 6.5
d. 2

CO_3^{-2} is basic and not amphiprotic, therefore we use the weak base equation to calculate the pH of the CO_3^{-2} solution. K_{a2} is the equilibrium equation that involves CO_3^{-2}.

K_{a1} H_2CO_3 --> H^+ + HCO_3^-
K_{a2} HCO_3- --> H^+ + CO_3^{-2}

to determine the K_b, we use the equation $K_a*K_b = K_w = 10^{-14}$
$K_b = K_w/K_{a2} = (10^{-14})/(10^{-11}) = 10^{-3}$

$[OH^-] = (C_bK_b)^{0.5} = (0.1*10^{-3})^{0.5} = (10^{-1} * 10^{-3})^{0.5} = 10^{-2}$
pOH =2
pH= 14-pOH = 14-2 =12

24. When calculating the PH of a 1 M solution of a diprotic acid (H_2A), for which of the following K_a values would it be safe to ignore K_{a2} and simply treat the acid as a monoprotic species?

a. $K_{a1} = 1.14x10^{-5}$, $K_{a2} = 2.38x10^{-6}$
b. $K_{a1} = 4.08x10^{-3}$, $K_{a2} = 5.60x10^{-7}$
c. $K_{a1} = 3.26x10^{-2}$, $K_{a2} = 9.98x10^{-7}$
d. $K_{a1} = 4.55x10^{-1}$, $K_{a2} = 8.19x10^{-3}$

The farther apart the K_a values, the safer it is to treat the deprotonation events independently.

25. When calculating the pH of a 0.3M H_2SO_4 solution, which equation do we use for calculations?

a. strong acid
b. weak acid
c. amphiprotic
d. buffer
e. a full quadratic solution is necessary

H_2SO_4 is a strong acid, which means the first dissociation goes to completion. Thus, there is no longer H_2SO_4 in solution, and we are only left with the dissociated products HSO_4^- and H^+.

Ka1	R	H_2SO_4	--> H^+	+ HSO_4^-
	I	0.3	0	0
	C	-0.3	+0.3	+0.3
	E	0	0.3	0.3

Ka2	R	$H_SO_4^-$	--> H^+	+ SO_4^{2-}
	I	0.3	0.3	0
	C	-X	+X	+x
	E	0.3-x	0.3+X	x

$K_{a2} = (0.3+X)*x/(0.3-2)$
This will requires us to solve a quadratic.