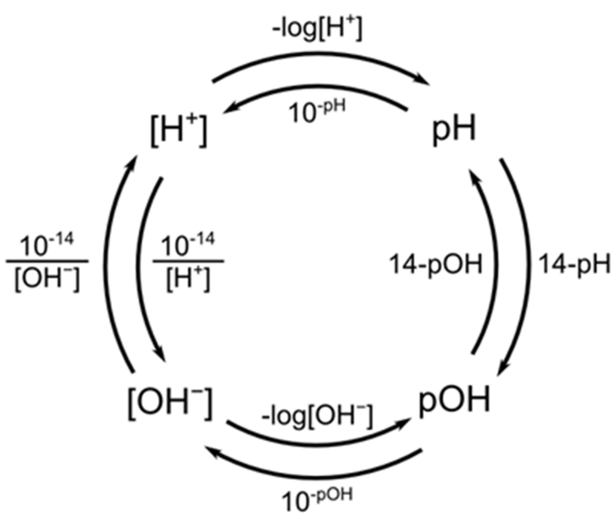
Chem 101B Study Questions Name: ____________________________
Chapters 14,15,16
Review Tuesday 3/26/2019
Due on Exam Thursday 3/28/2019 (Exam 3 Date)
This is a homework assignment. Please show your work for full credit. If you do work on separate paper, attach the work to these.

Exam 2 Sections Covered:

14.6, 14.8, 14.9, 14.10, 14.11, 14.12
15.1 – 15.5
16.1

Useful Info to be provided on exam:

$\%\text{diss} = \dfrac{[\text{H}^+]_{eq}}{[\text{HA}]_o} \times 100\%$ $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$K_a K_b = K_w$ $[\text{H}^+] = K_a \dfrac{[\text{HA}]}{[\text{A}^-]}$ $\quad$ $\text{pH} = \text{p}K_a + \log\dfrac{[\text{A}^-]}{[\text{HA}]}$ $[\text{OH}^-] = K_b \dfrac{[\text{B}]}{[\text{BH}^+]}$ $\quad$ $\text{pOH} = \text{p}K_b + \log\dfrac{[\text{BH}^+]}{[\text{B}]}$



In questions 1-7, determine the acidity of an aqueous solution made from each of the following salts:

1. solid calcium hydroxide, $Ca(OH)_2$
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

2. solid sodium nitrate, $NaNO_3$
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

3. solid ammonium bromide, NH_4Br
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

4. solid aluminum chloride, $AlCl_3$
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

5. solid sodium hydrogen carbonate $NaHCO_3$ (K_a HCO3 = 4.7×10^{-11} , Kb HCO3 = 2.2×10^{-8})
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

6. solid sodium carbonate, Na_2CO_3
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

7. solid ammonium acetate ($NH_4C_2H_3O_2$). For NH_4^+, $K_a = 5.6 \times 10^{-10}$; for $C_2H_3O_2^-$, $K_b = 5.6 \times 10^{-10}$.
   A) acidic
   B) basic
   C) neutral
   D) cannot tell
   E) none of these (A-D)

8. Calculate the pH of a 1.0 *M* sodium acetate solution. ($HC_2H_3O_2$ $K_a = 1.8\times10^{-5}$)
   A) 7.0
   B) 9.37
   C) 5.46
   D) 8.54
   E) none of these (A-D)

9. Calculate the pH of the following aqueous solution:
   0.69 *M* NH_4Cl (K_b for $NH_3 = 1.8\times10^{-5}$)
   A) 9.29
   B) 4.71
   C) 9.42
   D) 4.58
   E) none of these

10. What is the pH of a 0.44 *M* KCl solution?
   A) 0.36
   B) 7.00
   C) 13.64
   D) 2.20
   E) 9.20

11. Which is the strongest acid of the following (hint: consider the structures)?
    A) $HClO_2$
    B) $HClO$
    C) $HBrO$
    D) HIO
    E) $HOAt$

12. Which of the following species <u>cannot</u> act as a Lewis base?
    A) N^{3-}
    B) NH^{2-}
    C) NH_2^-
    D) NH_3
    E) NH_4^+

13. Which of the following species <u>cannot</u> act as a Lewis acid?
    A) CH_4
    B) H^+
    C) BF_3
    D) $BeCl_2$
    E) Ag^+

14. Explain why 0.1 *M* NaCN is basic while 0.1 *M* $NaNO_3$ is neutral.

Use the following to answer questions 15-17:

Determine whether the following oxides produce an acidic, basic, or neutral solution when dissolved in water:

15. K_2O

16. Cl_2O

17. SO_2

18. What is the percent dissociation of HNO_2 when 0.057 g of sodium nitrite is added to 115.0 mL of a 0.064 *M* HNO_2 solution? K_a for HNO_2 is 4.0×10^{-4}.
   A) 14%
   B) 0.36%
   C) 4.0%
   D) 0.081%
   E) 7.9%

19. Which of the following **will not** produce a buffered solution? *hint: Do the (before/after) stoichiometry first*
   A) 100 mL of 0.1 *M* Na_2CO_3 and 50 mL of 0.1 *M* HCl
   B) 100 mL of 0.1 *M* $NaHCO_3$ and 25 mL of 0.2 *M* HCl
   C) 100 mL of 0.1 *M* H_2CO_3 and 25 mL of 0.2 *M* NaOH
   D) 50 mL of 0.2 *M* Na_3PO_4 and 5 mL of 1.0 *M* HCl
   E) 100 mL of 0.1 *M* Na_2CO_3 and 50 mL of 0.1 *M* NaOH

20. Suppose a buffer solution is made from formic and ($HCHO_2$) and sodium formate ($NaCHO_2$). What is the **net ionic** equation for the reaction that occurs when a small amount of hydrochloric acid is added to the buffer?
   A) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
   B) $H^+(aq) + HCHO_2(aq) \rightarrow H_2O(l) + H_2CHO_2^+(aq)$
   C) $HCl(aq) + OH^-(aq) \rightarrow H_2O(l) + Cl^-(aq)$
   D) $HCl(aq) + CHO_2^-(aq) \rightarrow HCHO_2(aq) + Cl^-(aq)$
   E) $H^+(aq) + CHO_2^-(aq) \rightarrow HCHO_2(aq)$

21. A 100. mL sample of 0.10 *M* HCl is mixed with 50. mL of 0.14 *M* NH_3. What is the resulting pH? (K_b for $NH_3 = 1.8 \times 10^{-5}$)
   A) 3.04
   B) 10.96
   C) 12.30
   D) 1.52
   E) 1.70

22. The following question refers to a **2.0-liter** buffered solution created from 0.38 *M* NH_3 ($K_b = 1.8 \times 10^{-5}$) and 0.26 *M* NH_4Cl. When 0.10 mol of H^+ ions is added to the solution what is the pH?
   A) 4.72
   B) 4.77
   C) 10.43
   D) 9.28
   E) 7.98

23. You have a 250.-mL sample of 1.93 *M* acetic acid ($K_a = 1.8 \times 10^{-5}$). Calculate the pH of the **best buffer** using this weak acid..
A) 7.00
B) 4.74
C) 4.42
D) 9.26
E) none of these

24. A solution contains 0.250 *M* HA ($K_a = 1.0 \times 10^{-6}$) and 0.45 *M* NaA (buffer). What is the pH after 0.24 mole of HCl is added to 1.00 L of this solution?
A) 0.62
B) 7.63
C) 5.63
D) 2.04
E) 8.37

25. The following question refers to the following system: A 1.0-liter solution contains 0.25 *M* HF and 0.36 *M* NaF (K_a for HF is 7.2×10^{-4}).
If one adds 0.30 liters of 0.020 *M* KOH to the solution, what will be the change in pH?
hint: this is kind of like a titration!
A) 0.02
B) 3.32
C) 0.18
D) –0.11
E) –0.27

26. How many moles of solid NaF would have to be added to 1.0 L of 2.48 *M* HF solution to achieve a buffer of pH 3.35? Assume there is no volume change. (K_a for HF = 7.2×10^{-4})
A) 4.0
B) 0.51
C) 0.65
D) 1.0
E) 1.6

27. A solution contains 0.500 M HA ($K_a = 1.0 \times 10^{-8}$) and 0.482 M NaA. What is the $[H^+]$ after 0.10 mole of HCl is added to 1.00 L of this solution?
    A) 1.0×10^{-8} M
    B) 2.6×10^{-8} M
    C) 6.4×10^{21} M
    D) 1.6×10^{-8} M
    E) none of these

28. Which of the following solutions will be the best buffer for a pH of 9.26? (K_a for $HC_2H_3O_2$ is 1.8×10^{-5}, K_b for NH_3 is 1.8×10^{-5}).
    A) 0.10 M $HC_2H_3O_2$ and 0.10 M Na $C_2H_3O_2$
    B) 5.0 M $HC_2H_3O_2$ and 5.0 M Na $C_2H_3O_2$
    C) 0.10 M NH_3 and 0.10 M NH_4Cl
    D) 5.0 M NH_3 and 5.0 M NH_4Cl
    E) 5.0 M $HC_2H_3O_2$ and 5.0 M NH_3

Use the following to answer questions 29-30:

You have two buffered solutions. Buffered solution 1 consists of 5.0 M $HC_2H_3O_2$ and 5.0 M $NaC_2H_3O_2$; Buffered solution 2 is made of 0.050 M $HC_2H_3O_2$ and 0.050 M $NaC_2H_3O_2$.

29. How does the pH of each buffered solution compare?
    A) The pH of buffered solution 1 is greater than that of buffered solution 2.
    B) The pH of buffered solution 2 is greater than that of buffered solution 1.
    C) The pHs are the same.
    D) Cannot be determined without the K_a values.
    E) None of these (A-D).

30. Which solution has the greater buffering capacity?
    A) Solution 1
    B) Solution 2
    C) Both solutions have the same buffering capacity.

31. A 50.00-mL sample of 0.100 *M* KOH is titrated with 0.293 *M* HNO_3. Calculate the pH of the solution after 52.00 mL of HNO_3 is added.
A) 13.00
B) 0.83
C) 1.00
D) 13.17
E) none of these

32. Consider the titration of 300.0 mL of 0.700 *M* NH_3 ($K_b = 1.8 \times 10^{-5}$) with 0.550 *M* HNO_3. How many milliliters of 0.550 *M* HNO_3 are required to reach the equivalence point of the reaction?
A) 4.32×10^2 mL
B) 4.82×10^2 mL
C) 7.00×10^2 mL
D) 3.82×10^2 mL
E) none of these

33. A 50.0-mL sample of 0.10 *M* HNO_2 ($K_a = 4.0 \times \square 10^{-4}$) is titrated with 0.13 M NaOH. The pH after 25.0 mL of NaOH have been added is
A) 10.33
B) 7.00
C) 6.67
D) 3.67
E) none of these

34. The pH at the equivalence point of a titration of a weak acid with a strong base will be
A) less than 7.00
B) equal to 7.00
C) greater than 7.00
D) equal to the pK_a of the acid
E) more data needed to answer this question

35. A 75.0-mL sample of 0.0500 *M* HCN ($K_a = 6.2 \times 10^{-10}$) is titrated with 0.348 *M* NaOH. What is the $[H^+]$ in the solution after 3.0 mL of 0.348 *M* NaOH have been added?
A) 6.2×10^{-6} *M*
B) 1.0×10^{-7} *M*
C) 2.6 *M*
D) 1.6×10^{-9} *M*
E) none of these

36. You have 75.0 mL of 0.12 *M* HA. After adding 30.0 mL of 0.10 *M* NaOH, the pH is 5.50. What is the K_a value of HA? *hint: use buffer equation*
A) 3.2×10^{-6}
B) 1.6×10^{-6}
C) 0.50
D) 1.1×10^{-6}
E) none of these

37. You have 100.0 mL of 0.100 *M* aqueous solutions of each of the following acids: HCN, HF, HCl, and $HC_2H_3O_2$. You titrate each with 0.100 *M* NaOH (*aq*). Rank the pHs of each of the solutions at the equivalence point, from highest to lowest pH.

K_a for HCN = 6.2×10^{-10}
K_a for HF = 7.2×10^{-4}
K_a for $HC_2H_3O_2$ = 1.8×10^{-5}

A) HCN, $HC_2H_3O_2$, HF, HCl
B) HCl, HF, HCN, $HC_2H_3O_2$
C) HF, HCN, $HC_2H_3O_2$, HCl
D) $HC_2H_3O_2$, HCl, HCN, HF
E) none of these

38. Consider the following indicators and their pH ranges:

Methyl orange	3.2-4.4
Methyl red	4.8-6.0
Bromothymol blue	6.0-7.6
Phenolphthalein	8.2-10.0
Alizarin yellow	10.1-12.0

For which of the following titrations would **methyl red** be a good indicator?
A) 0.100 *M* HNO_3 + 0.100 *M* KOH
B) 0.100 *M* aniline ($K_b = 3.8 \times 10^{-10}$) + 0.100 *M* HCl
C) 0.100 *M* NH_3 ($K_b = 1.8 \times 10^{-5}$) + 0.100 *M* HCl
D) 0.100 *M* HF ($K_a = 7.2 \times 10^{-4}$) + 0.100 *M* NaOH
E) 0.100 *M* acetic acid ($K_a = 1.8 \times 10^{-5}$) + 0.100 *M* NaOH

39. A certain indicator HIn has a pK_a of 9.00 and a color change becomes visible when 7.00% of it is In^-. At what pH is this color change visible? *hint: try the Hendersson formula!*
A) 10.2
B) 3.85
C) 6.15
D) 7.88
E) none of these

40. The solubility of $CaSO_4$ in pure water at 0ºC is 1.17 gram(s) per liter. Calculate K_{sp}.
A) 8.59×10^{-3}
B) 1.17×10^{-3}
C) 9.27×10^{-2}
D) 7.38×10^{-5}
E) none of these

41. The solubility in mol/L of Ag_2CrO_4 is 1.4×10^{-4} *M*. Calculate the K_{sp} for this compound.
A) 3.9×10^{-8}
B) 1.4×10^{-4}
C) 1.1×10^{-11}
D) 2.7×10^{-12}
E) 2.8×10^{-4}

42. The solubility of silver phosphate, Ag_3PO_4, at 25°C is 1.64×10^{-5} mol/L. What is the K_{sp} for the silver phosphate at 25°C?
A) 1.19×10^{-13}
B) 1.95×10^{-18}
C) 8.07×10^{-10}
D) 7.23×10^{-20}
E) none of these

43. Find the solubility (in mol/L) of lead(II) chloride, $PbCl_2$, at 25°C. $K_{sp} = 1.58 \times 10^{-5}$.
A) 1.58×10^{-2}
B) 2.51×10^{-2}
C) 6.16×10^{-17}
D) 1.99×10^{-3}
E) 1.99×10^{-2}

44. Calculate the concentration of the silver ion in a saturated solution of silver chloride, AgCl ($K_{sp} = 1.62 \times 10^{-10}$).
A) 1.62×10^{-10}
B) 1.27×10^{-5}
C) 2.62×10^{-20}
D) 3.24×10^{-10}
E) none of these

45. The molar solubility of $BaCO_3$ ($K_{sp} = 1.6 \times 10^{-9}$) in 0.10 *M* $BaCl_2$ solution is:
A) 1.6×10^{-10}
B) 4.0×10^{-5}
C) 7.4×10^{-4}
D) 0.10
E) none of these

46. The solubility of $Mg(OH)_2$ ($K_{sp} = 8.9 \times 10^{-12}$) in 1.0 L of a solution buffered (with large capacity) at pH 9.47 is:
A) 7.8×10^{7} moles
B) 1.0×10^{-2} moles
C) 3.0×10^{-7} moles
D) 3.0×10^{-5} moles
E) none of these

47. The solubility in mol/L of $M(OH)_2$ in 0.030 *M* KOH is 1.0×10^{-5} mol/L. What is the K_{sp} for $M(OH)_2$?
A) 9.0×10^{-9}
B) 3.0×10^{-7}
C) 9.0×10^{-4}
D) 4.0×10^{-15}
E) 1.7×10^{-6}

48. Which of the following solid salts is more soluble in 1.0 M H^+ than in pure water?
A) NaCl
B) $CaCO_3$
C) KCl
D) AgCl
E) KNO_3

49. Without performing calculations, which of the following is the most soluble?

Salt	***K_{sp}***
$Pb(OH)_2$	1.4×10^{-20}
$Mn(OH)_2$	2.0×10^{-13}
$Zn(OH)_2$	2.1×10^{-16}

A) $Pb(OH)_2$
B) $Mn(OH)_2$
C) $Zn(OH)_2$
D) All have the same solubility.

Answer Key

1.	B
	Chapter/Section: 14.8
2.	C
	Chapter/Section: 14.8
3.	A
	Chapter/Section: 14.8
4.	A
	Chapter/Section: 14.8
5.	B
	Chapter/Section: 14.8
6.	B
	Chapter/Section: 14.8
7.	C
	Chapter/Section: 14.8
8.	B
	Chapter/Section: 14.8
9.	B
	Chapter/Section: 14.8
10.	B
	Chapter/Section: 14.8
11.	A
	Chapter/Section: 14.9
12.	E
	Chapter/Section: 14.11
13.	A
	Chapter/Section: 14.11
14.	When NaCN dissolves in water, it produces Na^+ and CN^- ions. The Na^+ ion is the cation of a strong base, and so does not have any effect on the $[H^+]$ or $[OH^-]$ in water. The CN^- ion, however, is the anion of a weak acid. It will react with water to produce OH^- and the conjugate acid, HCN. Since $[OH^-]$ increases by this reaction, the solution is basic. When $NaNO_3$ dissolves in water, the solvated ions are Na^+ and NO_3^-. Again, Na^+ does not affect $[H^+]$ or $[OH^-]$. Neither does NO_3^- since it is the anion of a strong acid, and so it does not act as a base, and does not affect $[H^+]$ or $[OH^-]$. See Sec. 14.8 of Zumdahl, *Chemistry*.
	Chapter/Section: 14.8
15.	basic $K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$; see Sec 14.10, Zumdahl *Chemistry*.
	Chapter/Section: 14.10

16.	acidic $Cl_2O(g) + H_2O(l) \rightarrow 2HClO(aq)$; see Sec 14.10, Zumdahl *Chemistry*.
	Chapter/Section: 14.10
17.	acidic $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$; see Sec 14.10, Zumdahl *Chemistry*.
	Chapter/Section: 14.10
18.	C
	Chapter/Section: 15.1
19.	E
	Chapter/Section: 15.2
20.	E
	Chapter/Section: 15.2
21.	E
	Chapter/Section: 15.2
22.	D
	Chapter/Section: 15.2
23.	B
	Chapter/Section: 15.2
24.	C
	Chapter/Section: 15.2
25.	A
	Chapter/Section: 15.2
26.	A
	Chapter/Section: 15.2
27.	D
	Chapter/Section: 15.2
28.	D
	Chapter/Section: 15.3
29.	C
	Chapter/Section: 15.3
30.	A
	Chapter/Section: 15.3
31.	C
	Chapter/Section: 15.4
32.	D
	Chapter/Section: 15.4
33.	D
	Chapter/Section: 15.4
34.	C
	Chapter/Section: 15.4
35.	D
	Chapter/Section: 15.4
36.	B
	Chapter/Section: 15.4

37.	A
	Chapter/Section: 15.4
38.	A, C
	Chapter/Section: 15.5
39.	D
	Chapter/Section: 15.5
40.	D
	Chapter/Section: 16.1
41.	C
	Chapter/Section: 16.1
42.	B
	Chapter/Section: 16.1
43.	A
	Chapter/Section: 16.1
44.	B
	Chapter/Section: 16.1
45.	E
	Chapter/Section: 16.1
46.	B
	Chapter/Section: 16.1
47.	A
	Chapter/Section: 16.1
48.	B
	Chapter/Section: 16.1
49.	B
	Chapter/Section: 16.2