Chemistry 112 Spring 2007 Prof. Metz Exam 3

Each question is worth 5 points, unless otherwise indicated.

1. HBrO reacts with water via

$$HBrO + H_2O \rightleftharpoons BrO^- + H_3O^+$$

In this acid-base reaction, _____ and _____ act as Bronsted acids; while _____ and _____ are bases:

(A) HBrO, H_2O; BrO^-, H_3O^+
(B) HBrO, H_3O^+; H_2O, BrO^-
(C) HBrO, BrO^-; H_2O, H_3O^+
(D) H_2O, H_3O^+; HBrO, BrO^-
(E) BrO^-, H_3O^+; HBrO, H_2O

HBrO is an acid, it donates a proton to H_2O (base), producing BrO^- (conjugate base) and H_3O^+ (acid)

2. The pH of a 0.150 M solution of formic acid, HCOOH is (K_a (formic acid) = 1.8×10^{-4}).

(A) 5.38 (C) 3.74 (E) 0.82
(B) 4.57 (D) 2.28

	$HCOOH + H_2O$ <->	H_3O^+	$+ COOH^-$	$K_a = 1.8 \times 10^{-4}$
Initial	*0.150*	*0*	*0*	
Change	*-x*	*x*	*x*	
Equilibrium	*0.150-x*	*x*	*x*	

So, $K_a = [H_3O^+][COOH^-]/[HCOOH] = (x)(x)/(0.150-x) = 1.8 \times 10^{-4}$
Assume x is small, so 0.15-x = 0.15, then
$(x^2)/0.15 = 1.8 \times 10^{-4}$
$x^2 = 2.7 \times 10^{-5}$
$x = [H_3O^+] = 5.2 \times 10^{-3}$ (and note that x is much less than 0.15)
$pH = -\log [H_3O^+] = 2.28$

3. An aqueous solution has a pH of 4.62. The $[OH^-]$ in the solution is

(A) 4.2×10^{-10} (C) 1.0×10^{-7} (E) 2.2×10^{-1}
(B) 8.3×10^{-8} (D) 2.4×10^{-5}

pH = $-\log[H_3O^+]$, so
$[H_3O+] = 10\text{-}pH = 10^{-4.62} = 2.40 \times 10^{-5}$
$[OH-] = 10^{-14}/[H_3O+] = 10^{-14}/(2.40 \times 10^{-5}) = 4.16 \times 10^{-10}$

4. The pH of a 0.05 M solution of nitric acid, HNO_3 is

(A) 2.00 (C) 0.70 (E) 0.05
(B) 1.30 (D) 0.10

HNO_3 is a strong acid, so it ionizes completely to form $H^+ + NO_3^-$. The H^+ combines with water to form H_3O^+. So, the H_3O^+ concentration is 0.05 M
pH = $-\log[H_3O^+] = -\log(0.05) = 1.30$

Questions 5 and 6 refer to a solution made by dissolving 0.020 mol of hydrazoic acid (HN_3) and 0.030 moles of sodium azide (NaN_3) in enough water to make 1.00 L of solution. pK_a for HN_3 is 4.72.

5. What is the approximate pH of this solution?

(A) 4.90 (C) 4.54 (E) 4.18
(B) 4.72 (D) 4.30

pH=pKa + log([conjugate base]/[acid])
pH=4.72 + log(0.030/0.020)
pH = 4.72 + 0.18 = 4.90

6. What is the approximate pH after the addition of 0.010 liters of 1M HCl to the solution?

(A) 4.72 (C) 4.30 (E) 2.00
(B) 4.54 (D) 4.18

Start with 0.020 moles hydrazoic acid (HN_3) and 0.030 moles azide (N_3^-) (Initial)

The 0.010 moles of HCl ionize completely to Cl^- and H^+, which combines with water to form H_3O^+, then H_3O^+ reacts with the base: (Change)

$H_3O^+ + N_3^- \rightarrow HN_3 + H_2O$

	HN_3	N_3^-
Initial	*0.020 moles*	*0.030 moles*
Change	*+0.010 moles*	*-0.010 moles*
Final	*0.030 moles*	*0.020 moles*

Concentrations: $[HN_3]$ = 0.03 moles/1.01L = 0.0297 M

$[N_3^-]$ = 0.020 moles/1.01 L = 0.0198 M
pH=pKa + log([conjugate base]/[acid])
pH=4.72 + log(0.0198/0.0297)
pH = 4.72 - 0.18 = 4.54

7. Calculate the $[CH_3COO^-]/[CH_3COOH]$ ratio necessary to give a buffer with a pH of 5.00.
K_a for CH_3COOH (acetic acid) is 1.8×10^{-5}
(A) 1.82 (B) 1.07 (C) 1.00 (D) 0.55 (E) 0.26

Ka = 1.8×10^{-5}, so pKa = -log(Ka) = 4.74
pH =pKa + log([conjugate base]/[acid])
5.00 = 4.74 + log([conjugate base]/[acid])
log([conjugate base]/[acid]) = 0.26
[conjugate base]/[acid] = $10^{0.26}$ = 1.82
(We did this as an example in class)

8. What is the $[Ca^{2+}]$ in a saturated $CaCO_3$ solution?
K_{sp} ($CaCO_3$) = 8.7×10^{-9}

(A) 4.4×10^{-9} M (D) 9.3×10^{-5} M
(B) 8.7×10^{-9} M (E) 2.1×10^{-3} M
(C) 6.2×10^{-7} M

	$CaCO_{3(s)}$ <->	$Ca^{2+}_{(aq)}$ +	$CO_3^{2-}{}_{(aq)}$
Initial		0	0
Change		x	x
Equilibrium		x	x

$K_{sp} = [Ca^{2+}][CO_3^{2-}]^2 = (x)\ (x) = 8.7 \times 10^{-9}$

$x^2 = 8.7 \times 10^{-9}$

$x = [Ca^{2+}] = 9.33 \times 10^{-5}\ M$

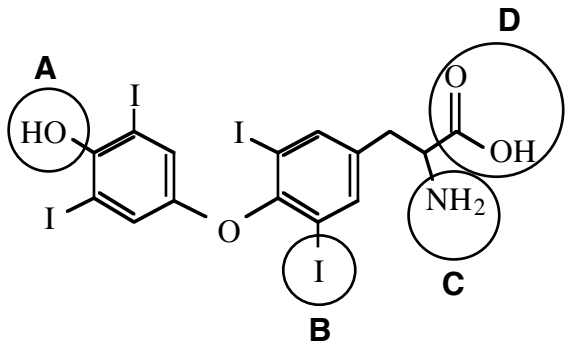
9. The solubility of calcium fluoride is 0.018 grams per liter of water and the molecular weight of CaF_2 is 78.0 g/mol. The solubility product constant (K_{sp}) for CaF_2 is _____ .

(A) 2.3×10^{-5} (C) 5.3×10^{-8} (E) 4.9×10^{-11}
(B) 5.8×10^{-6} (D) 5.8×10^{-9}

Concentrations:
*0.018 grams/L * 1 mol/78.0 grams = 2.31×10^{-4} M CaF_2, which dissolves to form 2.31×10^{-4} M Ca^{2+} and $2 \times 2.31 \times 10^{-4} = 4.62 \times 10^{-4}$ M F^-*

$$K_{sp} = [Ca^{2+}][F^-]^2 = (2.31 \times 10^{-4})(4.62 \times 10^{-4})^2 = 4.9 \times 10^{-11}$$

10. Thyroxine (shown below) is a metabolism regulator produced in the thyroid gland. It is a complex molecule with many functional groups.



Which of the groups circled is the most basic ?
(A) A (B) B (C) C (D) D

Group A is an alcohol (not acidic or basic), B is not acidic or basic, D is a carboxylic acid, C is an amine, which is basic.

11. The molar solubility of gold chloride ($AuCl_3$) in a 0.030 M NaCl solution is _____ M.
For $AuCl_3$ $K_{sp} = 3.2 \times 10^{-25}$
Note: Ksp is so small that little of the gold chloride dissolves

(A) 3.2×10^{-25} (C) 1.1×10^{-23} (E) 1.2×10^{-20}
(B) 5.7×10^{-13} (D) 3.3×10^{-7}

$AuCl_{3(s)}$ <->	$Au^+_{(aq)}$	+ 3 $Cl^-_{(aq)}$
Initial	0	0.030
Change	x	x
Equilibrium	x	0.030+x

$K_{sp} = [Au^+][Cl^-]^3 = (x)(0.030+x)^3 = 3.2 \times 10^{-25}$
Can assume that x is much smaller than 0.030, so 0.030 + x is approximately 0.030
$x(0.030)^3 = 3.2 \times 10^{-25}$
$x(2.7 \times 10^{-5}) = 3.2 \times 10^{-25}$
x = number of moles of $AuCl_3$ that dissolve = 1.19×10^{-20} M

12. If a solution of Na_2SO_4 is added slowly to an aqueous solution which is 0.010 M in Ba^{2+} and 0.50 M in Sr^{2+}, which salt precipitates first, and at what concentration of SO_4^{2-}? Assume no volume change.

$BaSO_4$ $K_{sp} = 1.1 \times 10^{-10}$; $SrSO_4$ $K_{sp} = 2.8 \times 10^{-7}$

(A) $SrSO_4$, at $[SO_4^{2-}] = 2.2 \times 10^{-10}$ M
(B) $SrSO_4$, at $[SO_4^{2-}] = 1.1 \times 10^{-8}$ M
(C) $SrSO_4$, at $[SO_4^{2-}] = 5.6 \times 10^{-7}$ M
(D) $BaSO_4$, at $[SO_4^{2-}] = 1.1 \times 10^{-8}$ M
(E) $BaSO_4$, at $[SO_4^{2-}] = 5.6 \times 10^{-7}$ M

Minimum $[SO_4^{2-}]$ to start precipitation of $BaSO_4$ is when
$[Ba^{2+}][SO_4^{2-}] = 1.1 \times 10^{-10}$
$(0.01\ M)(x) = 1.1 \times 10^{-10}$
$x = 1.1 \times 10^{-8}$ M

Minimum $[SO_4^{2-}]$ to start precipitation of $SrSO_4$ is when
$[Sr^{2+}][SO_4^{2-}] = 2.8 \times 10^{-7}$
$(0.50\ M)(x) = 2.8 \times 10^{-7}$
$x = 5.6 \times 10^{-7}$ M

So, Ba^{2+} will precipitate first, at the lower concentration $[SO_4]=1.1 \times 10^{-8}$ M
3 pts. Partial credit for (C)

Questions 13-16 are worth 2 points each
For the following compounds, state whether a 0.1 M solution will be
(A) acidic
(B) basic
(C) pH neutral

13. NH_4Cl
(A) NH_4Cl is formed by reaction of NH_3 (weak base) and HCl (strong acid), so it's an acidic salt

14. NaF
(B) NaF is formed by reaction of NaOH (strong base) with HF (weak acid), so it's a basic salt

15. $NaNO_3$
(C) $NaNO_3$ is formed by reaction of NaOH (strong base) with HNO3 (strong acid), so it's pH neutral

16. KOH
(B) KOH is a strong base

17. What is the concentration of CO_2 in a soft drink that is bottled with a CO_2 partial pressure of 4 atmospheres at 25 °C ? The Henry's Law constant for CO_2 in water is 4.08×10^{-5} M/(mm Hg)
(A) 0.124 M (B) 0.031 M (C) 1.6×10^{-4} M (D) 4.1×10^{-5} M (E) 2.1×10^{-7} M

$S_g = k_H P_g$
P_g is the pressure of CO_2.
P_g = 4 atm (760 mm Hg/atm) = 3040 mm Hg
$S_g = 4.08 \times 10^{-5}$ M/(mm Hg) x 3040 mm Hg = 0.124 M

18. A chemist dissolves 5.60 grams of $MgCl_2$ (Molec. wt. = 95.2 g/mol) in 300 grams of water (Molec. wt. = 18.0 g/mol) to give a solution with a volume of 320 mL.

What is the molality of $MgCl_2$ in the solution?

(A) 0.0187 (B) 0.0163 (C) 0.0175 (D) 0.196 (E) 0.184

molality = moles solute/kg of solvent
moles $MgCl_2$ (solute) = 5.60 g/(95.2 g/mol) = 0.0588 moles
molality = 0.0588 moles $MgCl_2$/0.300 kg water = 0.196 m
(Note: you were asked to calculate the molality, not the molarity, so you don't use the volume information given)

19. The vapor pressure of pure water at 25°C is 23.8 mm Hg. What is the vapor pressure of a solution that contains 50 grams water (MW=18.0 g/mol) and 50 grams ethylene glycol (MW = 62.1 g/mol). Ethylene glycol is a nonelectrolyte and assume that its vapor pressure is negligible.

(A) 3.2 mm Hg (B) 6.7 mm Hg (C) 11.9 mm Hg (D) 18.5 mm Hg (E) 23.8 mm Hg

$P_{solvent} = X_{solvent} P^0_{solvent}$, so need to calculate mole fraction of solvent:
Moles solvent = moles water = 50 g water/(18.0 g/mol) = 2.78 moles water
Moles solute = moles ethylene glycol = 50 g glycol/(62.1 g/mol) = 0.805 moles ethylene glycol
$X_{solvent}$ = moles solvent/(moles solvent + moles solute)
$X_{solvent}$ = 2.78 moles water/(2.78 moles water + 0.805 moles ethylene glycol)
$X_{solvent}$ = 0.775 (note: don't round off until later !)
$P_{solvent} = X_{solvent} P^0_{solvent}$ = 0.775 (23.80 mm Hg) = 18.46 mm Hg

20. 1.00 g of chloroform ($CHCl_3$, MW = 119.5 g/mol), a non-electrolyte, is dissolved in 50.0 g of carbon tetrachloride (CCl_4, MW = 154 g/mol) to give a solution with a volume of 40.0 mL. Assume that the solution is ideal. For CCl_4, the normal freezing point is -22.3 °C, the normal boiling point is 76.8 °C, K_{fp} = -29.8 °C/m and K_{bp} = 5.02 °C/m.
The freezing point of this solution is:
(A) –28.5 °C (B) –27.3 °C (C) –24.2 °C (D) –22.3 °C (E) –19.3 °C

$\Delta T_{fp} = K_{fp}\, m_{solute}\, i$, and K_{fp} = -29.8 °C/m, and i=1 (nonelectrolyte)
Moles solute = 1.00 g $CHCl_3$/119.5 g/mol = 8.368 x 10^{-3} moles $CHCl_3$
Molality solute = (8.368 x 10^{-3} moles $CHCl_3$)/(0.050 kg solvent) = 0.167 m
$\Delta T_{fp} = K_{fp}\, m_{solute}\, i$, and K_{fp} = -29.8 °C/m, and i=1 (nonelectrolyte)
ΔT_{fp} = (-29.8 °C/m) (0.167 M) (1) = -4.99 °C
Freezing point of solution = normal freezing point of solvent + ΔT_{fp}
Freezing point of solution = -22.3 °C – 4.99 °C = -27.3 °C

21. (2 pts) Which of the following is a strong acid (pick one) ?
(A) apple juice
(B) sulfuric acid
(C) gasoline
(D) water
(E) milk
This question is a double-check of which version of the exam you had…

22. Which of the following aqueous solutions would have the lowest freezing point?

(A) 0.10 m NaCl
(B) 0.10 m $Ba(OH)_2$
(C) 0.10 m HCl
(D) 0.10 m CH_3COOH (acetic acid)
(E) 0.10 m sugar

$\Delta T_{fp} = K_{fp}\, m_{solute}\, i$, and for water, K_{fp} = -1.86 °C/m, so lower freezing point means higher value of $m_{solute}\, i$ (or you can calculate ΔT_{fp} for each solution):
0.10 m NaCl: electrolyte, with i = 2, ΔT_{fp} = (-1.86 °C/m) 0.10 m (2) = -0.372 °C
0.10 m acetic acid, so i lies between 1 (doesn't ionize) and 2 (ionizes completely), so ΔT_{fp} lies between -0.186 and -0.372 °C
0.10 m HCl: strong acid, ionizes completely, so i = 2, ΔT_{fp} = (-1.86 °C/m) 0.10 m (2) = -0.372 °C
0.10 m $Ba(OH)_2$: strong base, with i = 3 (make Ba^{2+} and two OH^-), ΔT_{fp} = (-1.86 °C/m) 0.10 m (3) = -0.558 °C
0.10 m sugar: non-electrolyte, so i = 1, ΔT_{fp} = (-1.86 °C/m) 0.10 m (1) = -0.186 °C

23. A container has a 1.0 m aqueous solution of potassium chloride in one compartment and a 0.10 m aqueous solution of potassium chloride in the other compartment. The two compartments are separated by a semi-permeable membrane. As the system approaches equilibrium , _____ molecules will move through the membrane from the ____ m solution to the other solution.

(A) water, 1.0
(B) water, 0.10
(C) potassium chloride, 1.0
(D) potassium chloride, 0.10
(E) no molecules move through the membrane

Ion concentration is higher in the 1.0 m compartment, so the water concentration is lower in the 1.0 m compartment
This is a semipermeable membrane, so water goes through it, but ions don't.
Water moves from region with higher water concentration (the 0.1 m compartment) to region with lower water concentration (the 1.0 m compartment).

Extra Credit (5 pts.)
0.16 liters of 1.00 M NaOH are added to 1 liter of 0.1 M H_3PO_4, making the total volume 1.16 liters. What is the pH of the resulting solution ?
For H_3PO_4, $K_{a1} = 7.5 \times 10^{-3}$, $K_{a2} = 6.2 \times 10^{-8}$ and $K_{a3} = 3.6 \times 10^{-13}$.

Please answer ON THIS SHEET. Partial credit may be given, but only for work shown on this sheet !

Start with 0.1 moles of H_3PO_4 …it reacts with 0.1 moles of OH^-
$H_3PO_4 + OH^- \rightarrow H_2PO_4^- + H_2O$
End up with 0.1 moles of $H_2PO_4^-$ and 0.16 – 0.10 = 0.06 moles of OH^- remain.
As $H_2PO_4^-$ is a weak acid, it reacts with the 0.06 moles of OH^- that remain:
$H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_2O$
End up with 0.10 – 0.06 = 0.04 moles of $H_2PO_4^-$ and 0.06 moles of HPO_4^{2-}
This is a buffer (weak acid + conjugate base)
pH = pKa + log([conjugate base]/[acid])
[conjugate base] = $[HPO_4^{2-}]$ = 0.06 moles/1.16L = 0.0517 M
[acid] = $[H_2PO_4^-]$ = 0.04 moles/1.16L = 0.0345 M
K_{a1} refers to loss of first proton ($H_3PO_4 \rightarrow H_2PO_4^-$)
K_{a2} refers to loss of second proton ($H_2PO_4^- \rightarrow HPO_4^{2-}$)
so, need to use pKa2 = -log(K_{a2}) = 7.21
pH = 7.21 + log(0.0517/0.0345) = 7.21 + 0.18 = 7.39

(Note: Get the same answer if you don't divide by 1.16L, as the pH of a buffer doesn't change when you dilute it, as it depends on the ratio of concentrations)