CHEMISTRY 1B – Fall, 2015
FINAL EXAM 1 – VERSION A – KEY

Use Scantron Form SC982-E and select the letter corresponding to the correct answer. Make sure to put **your full name, lab section number, and exam version** (under test no.) on the Scantron Form.

Reference pages, including a periodic table and a page of equations and constants, followed by two pages of scratch paper are provided at the end of the test.

Part I. Multiple Choice Section. All Questions have only one correct answer. Each Question is worth 5 points.

1. Which of the following concentration based equilibrium equations correctly corresponds to the following chemical equation: $CaCO_3(s) + H_2CO_3(aq) \rightleftharpoons Ca^{2+}(aq) + 2HCO_3^-(aq)$?

a) $K_C = \frac{[Ca^{2+}][HCO_3^-]}{[CaCO_3][H_2CO_3]}$

b) $K_C = \frac{[Ca^{2+}][HCO_3^-]}{[H_2CO_3]}$

c) $K_C = \frac{[Ca^{2+}][HCO_3^-]^2}{[H_2CO_3]}$

d) $K_C = \frac{[H_2CO_3]}{[Ca^{2+}][HCO_3^-]}$

2. Given that the K_{sp} (sp = solubility constant) for AgCl is 1.8×10^{-10} and the K_f for formation of $Ag(NH_3)_2^+$ from Ag^+ and NH_3 is 1.7×10^7, K for: $AgCl(s) + 2NH_3(aq) \leftrightarrow Ag(NH_3)_2^+(aq) + Cl^-(aq)$ is:

a) 1.8×10^{-10} b) 3.1×10^{-3} c) 6.1×10^{-3} d) 1.0 e) 1.7×10^7

$K = K_{sp} \cdot K_f$ (this is equivalent to adding the K_{sp} reaction with the K_f reaction to get the listed reaction)

3. For the reaction $3H_2(g) + CO(g) \leftrightarrow CH_4(g) + H_2O(g)$, If the initial partial pressure of H_2, CO, CH_4 and H_2O are 2.00, 1.00, 0.00 and 0.00 atm, respectively, and the equilibrium partial pressure of CH_4 is 0.62 atm, K_P =

a) 0.019 b) 0.27 c) 7.2 d) 52 e) 370

ICE Table:	$3H_2(g)$	$+ CO(g) \leftrightarrow$	$CH_4(g)$	$+ H_2O(g)$
I	*2.00*	*1.00*	*0*	*0*
C	*-3x*	*-x*	*+x*	*+x*
E	*2.00 – 3x*	*1.00 – x*	*x*	*x*

where x = 0.62 atm (given above) so $K_P = P_{CH4}P_{H2O}/P_{H2}^3P_{CO}$

or $K_P = (0.62)^2/[(2.0 - 3 \cdot 0.62)^3(1.0 - 0.62)] = 369$ *(but only 2 sig fig)*

4. Which of the following changes will shift the reaction in problem 3 to the products if you are given that ΔH for that reaction is negative?

a) increase the flask volume
b) increase the temperature
c) add a catalyst
d) remove some water by condensation
e) all of the above

d) removes product causing shift to more products

5. If K_C for the reaction $2Na_2O_2(s) + 2CO_2(g) \leftrightarrow 2Na_2CO_3(s) + O_2(g)$ is 6.3×10^4, what is the equilibrium concentration of CO_2 (in M) if the equilibrium concentration of O_2 is 0.0100 M?

a) 1.6×10^{-7} b) 4.0×10^{-4} c) 2.0×10^{-4} d) 0.020 e) 1.3×10^4

$K_C = [O_2]/[CO_2]^2$ or $6.3 \times 10^4 = 0.010\ M/[CO_2]^2$ or $[CO_2] = (0.010/6.3 \times 10^4)^{0.5}$ =

6. A solution has a pH of 4.89. $[H^+]$ is:

a) 7.8×10^{-10} b) 1.3×10^{-5} c) 7.5×10^{-3} d) 0.69 e) 1.3

$[H^+] = 10^{-4.89} = 1.3 \times 10^{-5}$

7. An acid is considered a strong acid if it:

a) completely dissociates in water
b) can protonate acetic acid
c) is known as a Lewis acid
d) releases multiple protons per compound
e) reacts completely with a strong base

8. What is the pH of a 0.0050 M $Ba(OH)_2$ solution? (Note: K_w given in reference material)

a) 0.01 b) 2.00 c) 11.70 d) 12.00 e) 12.30

$[OH^-] = 2[Ba(OH)_2] = 0.01M$ $pOH = -\log[OH^-] = 2$ and $pH = 14 - pOH$ =

9. What is the pH of a 0.20 M $HC_2H_3O_2$ acid ($K_a = 1.8 \times 10^{-5}$)?

a) 0.70 b) 2.72 c) 4.74 d) 7.00 e) 9.26

ICE Table:	$HC_2H_3O_2(aq)$ $\leftrightarrow$	$C_2H_3O_2^-(aq)$ +	$H^+(aq)$
I	*0.20*	*0*	*0*
C	*-x*	*+x*	*+x*
E	*0.20 – x*	*x*	*x*

$1.8 \times 10^{-5} = [C_2H_3O_2^-][H^+]/[HC_2H_3O_2] = x^2/(0.20 - x) = x^2/0.20$ (assuming $x << 0.20$)
$x = [(1.8 \times 10^{-5})(0.20)] = 0.00190$ (so assumption is valid) $pH = -\log[H^+]$

10. In chlorine containing oxyacids, oxygen is more electronegative than chlorine and stabilizes conjugate anions. Which acid is the strongest?

a) HClO b) $HClO_2$ c) $HClO_3$ d) $HClO_4$

11. = BONUS #1. Given K_a values below, determine which salt solutions will be acidic?

a) $NaHCO_3$ b) $Na_2C_2O_4$ c) $NaHSO_3$ d) Na_2HPO_4 e) Na_2SO_4

(Note: K_w given in reference material)

Acid	K_{a1}	K_{a2}	K_{a3}
H_2CO_3	4.5×10^{-6}	4.7×10^{-11}	NA
$H_2C_2O_4$	6.0×10^{-2}	6.1×10^{-5}	NA
H_2SO_3	1.6×10^{-2}	6.4×10^{-8}	NA
H_2SO_4	large	1.2×10^{-2}	NA
H_3PO_4	7.5×10^{-3}	6.2×10^{-8}	4.2×10^{-13}

b and e are only bases, while a, c, and d can act as acids or bases. For a, c, and d we can compare K_a of acid with K_b (K_w/K_a) of base. Only HSO_3^- has a larger $K_{a2(3)}$ than K_{b2}.

12. What is the pH of a solution that contains 0.080 M NH_3 and 0.12 M NH_4Cl?
$K_b(NH_3) = 1.76 \times 10^{-5}$

a) 4.93 b) 9.07 c) 9.25 d) 9.42 e) 9.91

This is a buffer so we determine pH as $pH = pK_a + \log\{[NH_3]/[NH_4^+]\}$
$= (14 - pK_b) + \log(0.08/0.12) =$

13. A chemist has 100.0 mL of a 0.100 M $NaC_2H_3O_2$ solution. 100.0 mL of what solution can be added to the $NaC_2H_3O_2$ solution to make a buffer with pH = $pK_a \pm 1$? (with pK_a for $HC_2H_3O_2$)

a) 0.050 M HCl b) 0.15 M HCl c) 0.050 M $KC_2H_3O_2$ d) 0.050 M NaOH

A buffer requires a weak acid and its conjugate base. We are starting with the conjugate base, but $HC_2H_3O_2$*, the weak acid, is not a choice. Alternatively, a strong acid can be added to convert some of* $C_2H_3O_2^-$*, to* $HC_2H_3O_2$*. a) is the only choice since we only want to convert some of the conjugate base to acid.*

14. A chemist creates a new indicator which he calls vomit green that has a pK_a value of 5.44. It is yellow at pH values less than 5.44 and blue at pH values above 5.44 (with a vomit green shade near 5.44). This best use of this indicator would be for:

a) titrations of strong bases with strong acids
b) titrations of strong acids with strong bases
c) titrations of weak acids with strong bases
d) titrations of weak bases with strong acids
e) titrations of greenish colored solutions

The indicator is best when the equivalence point pH is around 5.5. This will occur for a weak base/strong acid titration.

Use the following information to answer questions 15 and 16. A buret is filled with 0.100 M NaOH and it is used to titrate 25.0 mL of HClO, a weak acid of unknown concentration. The titration requires 44.0 mL of NaOH to reach the equivalence point. The pK_a of HClO is 7.53

15. The original concentration of HClO in the 25.0 mL was:

a) 0.176 M b) 0.0568 M c) 0.114 M d) 0.0880 M e) 7.53 M

$n(OH^-) = n(HClO)$ or $(44.0\ mL)(0.100\ M) = (25.0\ mL)[HClO]$ or $[HClO] = 4.40/25.0 =$

16. The pH after adding 22.0 mL of NaOH will be:

a) 0.75 b) 6.77 c) 7.53 d) 8.27 e) 13.00

This is half of the equivalence point. At this point, half of HClO has been converted to ClO^-*, and* $pH = pK_a$*.*

17. Solid calcium fluoride (CaF_2) is added to water to make a saturated solution. Which of the following additions will **decrease** the equilibrium concentration of Ca^{2+}?

a) NaF b) HNO_3 c) EDTA (forms complex with Ca^{2+}) d) more $CaF_2(s)$

Adding more solid has no effect (since solids aren't in equilibrium equations). More NaF decreases solubility (common ion effect). b) and d) remove product increasing solubility

18. Calcium fluoride (CaF_2) has a $K_{sp} = 1.46 \times 10^{-10}$. Its molar solubility in water is:
a) 1.21×10^{-5} M   b) 3.32×10^{-4} M   c) 6.63×10^{-4} M   d) 1.00×10^{-3} M   e) 4.67×10^{-2} M
ICE table and equilibrium gives $K_{sp} = 1.46 \times 10^{-10} = [Ca^{2+}][F^-]^2 = x(2x)^2$ *or* $x = (K_{sp}/4)^{1/3}$

19. Which of the following reactions leads to a decrease in entropy for the system?
a) $H_2O(s) \leftrightarrow H_2O(l)$
b) $2H_2O(g) + O_2(g) \leftrightarrow 2H_2O_2(g)$
c) $Ag_2CrO_4(s) \leftrightarrow 2Ag^+(aq) + CrO_4^{2-}(aq)$
d) $H_2O(l) \leftrightarrow H_2O(g)$
e) $Cl_2(g) \leftrightarrow 2Cl(g)$
b) has more moles of reactants than products

20. Hydrazine (N_2H_4) is used to make rocket fuel and other products but it has positive ΔG_f over all temperatures. A strategy to make it would be to:
a) make from N_2 and H_2, but using higher temperatures
b) make from N_2 and H_2, but use catalysts
c) use a more stable reactant than H_2 (such as CH_4)
d) have the reaction produce another very stable co-product (such as H_2O)
e) have the reaction also produce another unstable product (such as N_3)

21. Under what temperature regimes will the reaction: $NH_4NO_3(s) \rightleftharpoons NH_4^+(aq) + NO_3^-(aq)$ ($\Delta H° = +25.7$ kJ/mol) be spontaneous?
a) never   b) high temperature   c) low temperature   d) always
$\Delta S > 0$ (more moles of products and in solution vs. solid) $\Delta G < 0$ only at high T

22. The reaction $H_2(g) + I_2(s) \leftrightarrow 2HI(g)$ has a $\Delta G° = +3.4$ kJ mol^{-1}. If a system starts with $P_{H2} = 0.85$ atm and $P_{HI} = 0.010$ atm and T = 298K, ΔG =
a) -19 kJ mol^{-1} b) -9.5 kJ mol^{-1} c) -7.6 kJ mol^{-1} d) +3.4 kJ mol^{-1} e) +19 kJ mol^{-1}
$\Delta G = \Delta G° + RT\ln Q$ where $Q = P_{HI}^2/P_{H2} = 1.18 \times 10^{-4}$
or $\Delta G = +3.4 + (8.314\ J\ mol^{-1}\ K^{-1})(298\ K)(-9.05)(1\ kJ/1000\ J) = 3.4 - 22.4 = -19$ kJ mol^{-1}

23. The $\Delta G_f°$ values are listed below for molecules in the reaction $4H_2(g) + 2CO(g) \leftrightarrow C_2H_4(g) + 2H_2O(g)$. (see reference page for constants)

Compound	$H_2(g)$	$CO(g)$	$C_2H_4(g)$	$H_2O(g)$
$\Delta G_f°$ (kJ mol^{-1})	0	-137.2	68.4	-228.6

The equilibrium constant at 298K is:
a) 1.3×10^{-25} b) 0.131   c) 1.0   d) 1.1×10^{20} e) 7.5×10^{24}
$\Delta G° = -RT\ln K$ or $K = e^{-\Delta G°(rxn)/RT}$ and $\Delta G°(rxn) = \Delta G_f°(C_2H_4) + 2\Delta G_f°(H_2O) - 2\Delta G_f°(CO)$
$\Delta G°(rxn) = 68.4 + 2(-228.6) - 2(-137.2) = -114.4$ kJ mol^{-1}
$K = e^{-(-114.4\ kJ/mol)(1000\ J/kJ)/[(8.314\ J/molK)(298K)]} = e^{46.2} =$

24. The following reaction is an UNBALANCED reaction showing reactants and products of a redox reaction:
$HClO(aq) + Cr(s) \leftrightarrow Cl_2(g) + Cr^{3+}(aq)$
When balanced (can assume acidic conditions), the coefficients in front of HClO and Cr (with no fractional coefficients anywhere in the equation) are:
a) 1 and 1 b) 2 and 1 c) 3 and 1 d) 3 and 2 e) 6 and 2
, respectively.
Cr half reaction: $Cr(s) \leftrightarrow Cr^{3+}(aq) + 3e^-$
HClO half reaction: $2HClO(aq) + 2H^+(aq) + 2e^- \leftrightarrow Cl_2(g) + 2H_2O(l)$
(Cr rxn)·2 + (HClO rxn)·3 gives net rxn with $6e^-$ *each*

25. The oxidation state of C in CH_3OH is:
a) +3 b) +2 c) +1 d) 0 e) -2
H is +1 and O is -2 so $x + 4\cdot(+1) - 2 = 0$ *or* $x = 2-4 = -2$

For questions 26 to 30, consult the following table of standard reduction potentials as needed.

Reaction	E°(V)
$Cl_2(g) + 2e^- \leftrightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \leftrightarrow 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \leftrightarrow 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \leftrightarrow Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^- \leftrightarrow Cu(s)$	+0.34
$2H^+(aq) + 2e^- \leftrightarrow H_2(g)$	0.00
$Ni^{2+}(aq) + 2e^- \leftrightarrow Ni(s)$	-0.23
$Cd^{2+}(aq) + 2e^- \leftrightarrow Cd(s)$	-0.40
$Fe^{2+}(aq) + 2e^- \leftrightarrow Fe(s)$	-0.45
$2H_2O(l) + 2e^- \leftrightarrow H_2(g) + 2OH^-(aq)$	-0.83

26. Which of the following reactants can oxidize copper metal (Cu(s)) under standard conditions?
a) $Ag^+(aq)$ b) $H^+(aq)$ c) $Ni^{2+}(aq)$ d) $Fe^{2+}(aq)$
e) all of the above
only oxidized compounds above Cu on the list can oxidize Cu as they are being reduced

27. A voltaic cell is made **under standard conditions** by Fe(s) and $FeCl_2(aq)$ in one half cell and $AgNO_3(aq)$ and Ag(s) in another half cell (with the two half cells connected by a salt bridge). The voltage measured from the silver (+) to the iron electrode (-) will be:
a) -0.10 V b) +0.10 V c) +0.35 V d) +1.25 V e) +2.05 V
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.80 - (-0.45) =$

28. For the reaction in the voltaic cell made in problem 27, if we want to increase the cell voltage, we can do so by:
a) increasing the mass of Fe
b) increasing the mass of Ag
c) increasing the concentration of $FeCl_2$
d) decreasing the concentration of $FeCl_2$
e) using a larger cell with twice the volume (at same concentration) of $FeCl_2$ and $AgNO_3$
We can determine this by looking at the Nernst Equation: $E_{Cell} = E°_{Cell} - \frac{0.0592}{n} logQ$
Anything decreasing Q will increase E_{cell} *(due to the minus sign in front of the logQ term)*
In problem 27, Ag^+ *is being reduced and Fe(s) is being oxidized. Thus* $Q = [Fe^{2+}]/[Ag^+]^2$
The only change that will decrease Q is d)

29. Which of the metals (in oxidation state of zero) in the list above are oxidized by H^+?
a) Ni and Cd b) Ni only c) Cd only d) Cu and Ag e) Ni, Cd, and Fe
All of the metals with negative reduction potentials [at least under standard conditions]

30. BONUS #2 In the following cell, the measured voltage is 0.99 V. Determine x (the concentration of $AgNO_3$(aq) in half of the cell.
cell: Ni(s)|$NiCl_2$(aq, 1.0 M)||$AgNO_3$(aq, x M)|Ag(s)
a) 0.21 M b) 0.045 M c) 0.459 M d) 2.2 M e) 4.8 M
Cell reaction: $2Ag^+(aq) + Ni(s) \leftrightarrow 2Ag(s) + Ni^{2+}(aq)$
$E_{Cell} = E°_{Cell} - \frac{0.0592}{n} logQ$ *where* $E°_{cell} = E°_{cathode} - E°_{anode} = 0.80 - (-0.23) = 1.03$ V
0.99 V = 1.03 V – 0.0592/2log{$[Ni^{2+}]/[Ag^+]^2$*} or -0.04 V = -0.0592/2log{*$1/x^2$*}*
or 1.35 = log{$1/x^2$*} or* $x^2 = 1/10^{1.35} = 0.0445$ *or x = 0.21*

31. What is a difference between voltaic cells and electrolytic cells?
a) Voltaic cells have both oxidation and reduction occurring, while electrolytic cells only have reduction occurring
b) In voltaic cells, oxidation occurs at the anode, while in electrolytic cells, oxidation occurs at the cathode
c) In voltaic cells, the cells provide electrical energy, while in electrolytic cells, electrical energy is consumed to drive the reaction
d) In voltaic cells, inert electrodes are used, while in electrolytic cells, metals being oxidized are needed for electrodes
e) Electrolytic cells can be used for batteries, while voltaic cells can't be used as batteries

32. A redox reaction has a negative ΔG° value. This tells us
a) a negative cell potential is expected under all conditions
b) a positive cell potential is expected under standard conditions
c) a K value of less than 1 is expected
d) this cell would only be useful if it were an electrolytic cell

Questions 33 and 34 involve $Na[Cr(H_2O)_2(ox)_2]$. Note that the Latin root for Cr is the same as in English and "ox" stands for $C_2O_4^{2-}$ or oxalate, a bidentate ligand.

33. The name for this complex is:
a) Sodium diaquadioxalatochromate(III) b) Sodium biaquabioxochromium(III)
c) Diaquadioxalatochromium(III) sodiate d) Sodium diaquadioxalatocrobaltate(III)
e) Sodium bis(dihydromonoxide)dioxalatochromate(III)

34. The oxidation state of the Cr atom is:
a) +5 b) +3 c) +2 d) +1 e) 0
ox = -2 charge, Na = +1 charge so x + 1 +2(-2) = 0 or x = +3

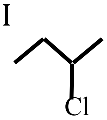
35. In coordination complexes, electrons in bonds between ligands and metals almost always come from:
a) metal s shells b) metal d shells c) ligand lone pair electrons
d) ligand inner shell electrons e) ligand sigma bonds

36. The coordination complex $[Ni(en)_3]^{2+}$ absorbs yellow light (at 560 nm). Calculate Δ in kJ/mol. (See reference page for constants)
a) 214 kJ/mol b) 0.214 kJ/mol c) 21 kJ/mol d) 83,113 kJ/mol
$E = hc/\lambda = (6.63 \times 10^{-34}\ J\cdot s)(3.00 \times 10^{8}\ m/s)/(560 \times 10^{-9}\ m) = 3.55 \times 10^{-19}\ J/complex$
$E = (3.55 \times 10^{-19}\ J/complex)(6.02 \times 10^{23} complex/mol)(1\ kJ/1000\ J) = 214\ kJ/mol$

37. The complex $[Co(NO_2)_6]^{4-}$ will have how many unpaired d electrons:
a) 0 b) 1 c) 2 d) 3 e) 4
Co^{2+} is a d^7 metal and NO is a strong ligand causing a low spin state - 6 lowest electrons fill first leaving one unpaired electron

38. Which type of hydrocarbons will have at least two carbons with sp hybridization?
a) alkanes b) alkenes c) alkynes d) aromatics
sp hybridization occurs with triple bonds

Using the structures shown below, answer the following 2 questions:

I	II	III	IV	V
				

39. Which structure is 2,2-dimethylpropane
a) I b) II c) III d) IV e) V
only IV and V are propanes and V is di-substituted (2 methyl groups)

40. Which two structures are isomers (have same molecular formula)?
a) I and II b) I and IV c) IV and V d) II and IV e) I and V
I and IV both have 4 carbons and no double bonds or other substituents

41. What is the product of CH_2=$CHCH_2CH_3$ and Cl_2?

a) $ClCH_2CH_2CH_2CH_3$

b) $CH_3CClHCH_2CH_3$

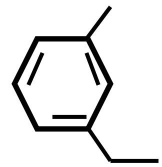
c) $ClCH_2CClHCH_2CH_3$

d) CH_2=$CHCClHCH_3$

e) CH_2=$CHCH_2CH_2Cl$

C_2 add across double bond

42. Give the name of the compound to the right.



a) 1-methyl-3-ethylbenzene

b) methylphenylethene

c) 1-ethyl-5-methylbenzene

d) 1-ethyl-3-methylbenzene

ethyl first (e before m in alphabet), so methyl is at 3 position (always use lower number)

Constants and Equations

Constants: 0°C = 273K

Universal Gas constant = 8.314 J mol^{-1} K^{-1}

K_w = 1.0 x 10^{-14} (self hydrolysis constant for water)

h = Planck's constant = 6.63 x 10^{-34} J·s and c = speed of light = 3.00 x 10^{8} m/s

Avagadro's number = 6.02 x 10^{23}.

F = Faraday's constant = 96,485 C/mol e

1 A (amp) = 1 C/s

Equations:

Quadratic Equation – not needed for this test

for $ax^2 + bx + c = 0$ is $x = \frac{-b\pm\sqrt{b^2-4ac}}{2a}$

Thermodynamic Equations

$\Delta G_{rxn} = \Delta G_{rxn}° + RT\ln Q$ $\Delta G° = \Delta H° - T\Delta S°$

Nernst + related Equations:

$E_{Cell} = E°_{Cell} - \frac{0.0592}{n} \log Q$

$\Delta G = -nFE°$

Strong ligands: $CN^- > NO_2^- >$ en $> NH_3$ **Weak ligands:** $H_2O > OH^- > F^- > Cl^- > Br^-$