PLI Chem 104 Review
Worksheet 1

This is the first practice exam of 2. This exam may not reflect the level of difficulty of the actual Chem 104 final exam or its length.There are many ways to utilize this sheet i.e. only doing specific problems, using it as a diagnostic to guide studying moving forward, etc.

1) A 35.0-g sample of ethylene glycol, $HOCH_2CH_2OH$, is dissolved in 500.0 g of water. The vapor pressure of water at 32 °C is 35.7 mm Hg. What is the vapor pressure of the water–ethylene glycol solution at 32 °C? (Ethylene glycol is nonvolatile.) (1 atm = 1.01325 bar)

2) An aqueous solution contains 3.00% phenylalanine ($C_9H_{11}NO_2$) by mass. (Phenylalanine is nonionic and nonvolatile.) Find the following:
   (a) the freezing point of the solution (K_{fp} = -1.86 °C/m)

   (b) the boiling point of the solution (K_{bp} = 0.5121 °C/m)

   (c) the osmotic pressure of the solution at 25 °C

3) An aqueous solution contains 0.180 g of an unknown, nonionic solute in 50.0 g of water. The solution freezes at −0.040 °C. What is the molar mass of the solute? (K_{fp} = -1.86 °C/m)

4) List the following aqueous solutions in order of increasing melting point.
(a) 0.1 *m* sugar (b) 0.1 *m* NaCl
(c) 0.08 *m* $CaCl_2$ (d) 0.04 *m* Na_2SO_4

5) Data for the reaction NO(g) + ½ O_2(g) n NO_2(g) are given (for a particular temperature) in the table.

Concentration (mol/L)			
Experiment	[NO]	[O_2]	Initial Rate (mol NO / (L*h))
1	3.6×10^{-4}	5.2×10^{-3}	3.4×10^{-8}
2	3.6×10^{-4}	1.04×10^{-2}	6.8×10^{-8}
3	1.8×10^{-4}	1.04×10^{-2}	1.7×10^{-8}
4	1.8×10^{-4}	5.2×10^{-3}	?

(a) What is the rate law for this reaction?

(b) What is the rate constant for the reaction?

(c) What is the initial rate of the reaction in experiment 4?

6) Data for the decomposition of dinitrogen monoxide:

$$N_2O(g) \rightarrow N_2(g) + \tfrac{1}{2} O_2(g)$$

on a gold surface at 900 °C are given below. Determine the order of the reaction and the rate constant. Using the rate law and value of *k,* determine the decomposition rate at 900 °C when $[N_2O]$ = 0.035 mol/L.

Time (min)	$[N_2O]$ (mol/L)
15.0	0.0835
30.0	0.0680
80.0	0.0350
120.0	0.0220

7) Ozone, O_3, in the Earth's upper atmosphere decomposes according to the equation:

$$2\ O_3(g) \rightarrow 3O_2\ (g)$$

The mechanism of the reaction is thought to proceed through an initial fast, reversible step, follow by a slow, second step.

Step 1: Fast reversible $O_3(g) \leftrightarrow O_2(g) + O(g)$
Step 2: Slow $O_3(g) + O(g) \rightarrow 2O_2(g)$

(a) Which of the steps is rate-determining?

(b) Write the rate equation for the rate-determining step.

8) Carbonyl bromide decomposes to carbon monoxide and bromine:

$$COBr_2(g) \leftrightarrow CO(g) + Br_2(g)$$

K_c is 0.190 at 73 °C. If you place 0.0500 mol of $COBr_2$ in a 2.00-L flask and heat it to 73 °C, what are the equilibrium concentrations of $COBr_2$, CO, and Br_2? What percentage of the original $COBr_2$ decomposed at this temperature?

9) Calculate K for the reaction

$$SnO_2(s) + 2CO(g) \leftrightarrow Sn(s) + 2CO_2(g)$$

given the following information:

1. $SnO_2(s) + 2H_2(g) \leftrightarrow Sn(s) + 2H_2O(g)$ $K_1 = 8.12$
2. $H_2(g) + CO_2(g) \leftrightarrow H_2O(g) + CO(g)$ $K_2 = 0.771$

10) The reaction of hydrogen and iodine to give hydrogen iodide has an equilibrium constant, K_c, of 56 at 435 °C.

$$H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

(a) What is the value of K_p?

(b) Suppose you mix 0.045 mol of H_2 and 0.045 mol of I_2 in a 10.0-L flask at 425 °C. What is the total pressure of the mixture before and after equilibrium is achieved?

(c) What is the partial pressure of each gas at equilibrium?

11) A 0.015M solution of hydrogen cyanate, HOCN, has a pH of 2.67.

(a) What is the hydronium ion concentration in the solution?

(b) What is the ionization constant, K_a, for the acid?

12) What are the equilibrium concentrations of H_3O^+, CN^-, and HCN in a 0.025 M solution of HCN? What is the pH of the solution? (HCN $K_a = 4.0 \times 10^{-10}$)

13) Sodium cyanide is the salt of the weak acid HCN. Calculate the concentrations of H_3O^+, OH^-, HCN, and Na^+ in a solution prepared by dissolving 10.8 g of NaCN in enough water to make 5.00×10^2 mL of solution at 25 °C. (HCN $K_a = 4.0 \times 10^{-10}$)

14) What mass of sodium acetate, $NaCH_3CO_2$, must be added to 1.00 L of 0.10 M acetic acid to give a solution with a pH of 4.50? ($HC_2H_3O_2$ $K_a = 1.8 \times 10^{-5}$)

15) You titrate 25.0 mL of 0.10 M NH_3 with 0.10 M HCl. (NH_3 $K_b = 1.8 \times 10^{-5}$)

(a) What is the pH of the NH_3 solution before the titration begins?

(b) What is the pH at the equivalence point?

(c) What is the pH at the half equivalence point of the titration?

(d) Calculate the pH of the solution after adding 5.00, 15.0, 20.0, 22.0, and 30.0 mL of the acid.

16) Calculate the molar solubility of silver thiocyanate, AgSCN, in pure water and in water containing 0.010 M NaSCN. (AgSCN $K_{sp} = 1.0 \times 10^{-12}$)

17) A solution contains Ca^{2+} and Pb^{2+} ions, both at a concentration of 0.010 M. You wish to separate the two ions from each other as completely as possible by precipitating one but not the other using aqueous Na_2SO_4 as the precipitating agent. ($CaSO_4$ $K_{sp} = 4.9 \times 10^{-5}$, $PbSO_4$ $K_{sp} = 2.5 \times 10^{-8}$)

(a) Which will precipitate first as sodium sulfate is added, $CaSO_4$ or $PbSO_4$?

(b) What will be the concentration of the first ion that precipitates (Ca^{2+} or Pb^{2+}) when the second, more soluble salt begins to precipitate?

Chapter 18: Thermodynamics

Question 18:
Fill in the following table. Assume no two rows are the same

ΔH	ΔS	ΔG	Low Temp	High Temp
-		-		
		Temp. dependent		Non-spontaneous
	+	Temp. dependent		

Question 19 (18.25):
Heating some metal carbonates, among them magnesium carbonate, leads to their decomposition.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$

$S_f°$ MgO = 26.85 J/mol K
$S_f°$ CO_2 = 213.74 J/mol K
$S_f°$ $MgCO_3$ = 65.84 J/mol K

$H_f°$ MgO = -601.85 kJ/mol K
$H_f°$ CO_2 = -393.51 kJ/mol K
$H_f°$ $MgCO_3$ = -1111.69 kJ/mol K

(a) Calculate $\Delta_r G°$ for the reaction.

(b) Over what temperature range is the reaction predicted to be product-favored at equilibrium?

(c) What is the equilibrium constant of the above reaction at T = 500 K?

Question 20 (18.31):
For the synthesis of ammonia from its elements at 25 °C, $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$
$\Delta_f G°\ N_2(g) = \Delta_f G°\ H_2(g) = 0$ kJ/mol $\Delta_f G°\ NH_3(g) = -16.37$ kJ/mol

(a) Calculate $\Delta_r G$. Is the reaction product-favored at equilibrium?

(b) Calculate $\Delta_r G$ when the reactants and product are each present at a partial pressure of 0.10 atm. Is the reaction spontaneous under these conditions?

Question 21 (18.33):
For each set of compounds below, circle the one of highest entropy. Assume that all are at the same temperature. Briefly explain your reasoning.

(a) HF(g) HCl(g) HBr(g)

(b) $NH_4Cl(s)$ $NH_4Cl(aq)$

(c) $C_2H_4(g)$ $N_2(g)$ (same molar mass)

(d) NaCl(s) NaCl(g)

Chapter 19: Electrochemistry

Question 22 (19.21):

Consider the following half reactions

Half Reaction	E^{o}(V)
$Cu^{2+}(aq) + 2\ e^{-}$ -> Cu(s)	+ 0.34 V
$Sn^{2+}(aq) + 2\ e^{-}$ -> Sn(s)	- 0.14 V
$Fe^{2+}(aq) + 2\ e^{-}$ -> Fe(s)	- 0.44 V
$Zn^{2+}(aq) + 2\ e^{-}$ -> Zn(s)	- 0.76 V
$Al^{3+}(aq) + 3\ e^{-}$ -> Al(s)	- 1.66 V

(a) Based on *E*° values, which metal is the most easily oxidized?

(b) Which metals on this list are capable of reducing $Fe^{2+}(aq)$ to Fe(s)?

(c) Write a balanced chemical equation for the reaction of $Fe^{2+}(aq)$ with Sn(s). Is this reaction product- favored or reactant-favored?

(d) Write a balanced chemical equation for the reaction of $Zn^{2+}(aq)$ with Sn(s). Is this reaction product- favored or reactant-favored?

Question 23 (19.33):
One half-cell in a voltaic cell is constructed from a silver wire electrode in a $AgNO_3$ solution of unknown concentration. The other half-cell consists of a zinc electrode in a 1.0 M solution of $Zn(NO_3)_2$. A potential of 1.48 V is measured for this cell. Use this information to calculate the concentration of $Ag^+(aq)$.

Zn^{2+} (*aq)* + $2e^-$ -> Zn (s) E^o = -0.76 V
Ag^+ (*aq)* + e^- -> Ag (s) E^o = 0.80 V

Question 24 (19.37):
Use standard reduction potentials below to calculate the value of K_{sp} for AgBr (s)

- $AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$ 0.0713 V
- $Ag^+(aq) + e^- \rightarrow Ag(s)$ 0.7994 V

Question 25 (19.49):
Electrolysis of a solution of $CuSO_4$ (aq) to give metal Cu (s) is carried out using a current of 0.66 A. How long should electrolysis continue to produce 0.50 g of copper?

Question 26 (19.3 & 19.5):
Balance the following redox equations. (a) is in acidic solution (b) is in basic solution

(a) $MnO_4^-(aq) + HSO_3^-(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$

(b) $Fe(OH)_2(s) + CrO_4^{2-}(aq) \rightarrow Fe(OH)_3(s) + [Cr(OH)_4]^-(aq)$

Chapter 25: Nuclear Chemistry **Note: NOT covered in class 8/4/21 but will be on the final exam.**

Question 27:
Give an example of each of the four types of nuclear reactions (α emission, β emission, positron emission, electron capture) for $^{235}_{92}U$.

Question 28 (25.15):
The uranium-235 radioactive decay series, beginning with $^{235}_{92}U$ and ending with $^{207}_{82}Pb$, occurs in the following sequence: α, β, α, β, α, α, α, α, β, β, α. Write an equation for each step in this series.

Question 29 (25.19): Predict the probable mode of decay for each of the following radioactive isotopes, write an equation to show the products of decay, and briefly explain your reasoning.

(a) bromine-80

(b) californium-240

(c) cobalt-61

(d) carbon-11

Question 30: Given ${}_{1}^{1}H = 1.00783\ amu$ ${}_{0}^{1}n = 1.00867\ amu$, calculate the binding energy per mole of nucleons for the following.

(a) ^{11}B (11.00931 g/mol)

(b) ^{40}Ca (39.96259 g/mol)

(c) ^{16}O (15.99492 g/mol)

Question 31: Gallium-67 ($t_{1/2}$ = 78.25 hours) is used in the medical diagnosis of certain kinds of tumors. If you ingest a compound containing 0.015 mg of this isotope, what mass (in milligrams) remains in your body after 13 days? (Assume none is excreted.)