Department of Chemistry
CHM 1220/1225
Exam V – B
April 22, 2016

Directions

1. Skim the entire exam before you begin so that you have a sense of the whole: what parts you can do quickly and what parts will require more time. The points for each problem are shown in parenthesis in the left margin. Try to use your time in proportion to the points assigned for each question.

2. You must show all the work necessary to arrive at your answer. **No credit will be given for numerical answers unless your work is shown.** (We want to be able to follow your thought process in order to be able to help make corrections and allot partial credit.)

3. Be sure to include the correct number of significant figures and the appropriate unit when reporting your answers.

Academic Integrity Pledge

During the exam I will

- turn off my cell phone and put it away (out of sight and not on my person)
- close all books, notebooks, etc. and put them under the seat in which I sit
- use only a permitted calculator
- keep my eyes down and focused on my own paper
- write only in ink
- keep my answers covered
- sit in the area assigned to my section
- stop writing when the end of the exam is announced

During the exam I will not

- have any papers other than those provided
- have any writing on my clothing or person or desk
- talk to anyone other than a TA or the instructor

I understand that the *minimum consequence* of any behavior contrary to this pledge is that I will receive a *zero on this exam* that will not be replaced by the percent earned on my final exam.

Scoring

1	15	15	5	15 /	15	MC 20 /	30
2	15 /	15	6	3 /	15		
3	6 /	15	7	15 /	15	Total 11 /	150
4	15 /	15	8	07 /	15		

Multiple Choice Answers

1. C
2. A
3. D
4. C
5. C
6. D

1. A soluble salt, MX_2, is added to water in a beaker. The equation for this dissolution is:

$$MX_2(s) \rightarrow M^{2+}(aq) + 2X^-(aq); \Delta H > 0 \text{ surroundings} \quad \text{endothermic}$$

a. Immediately after the salt dissolve, the solution is (circle one) WARMER ⟨COLDER⟩.

b. While the salt dissolves, the flow of heat is (circle one) IN ⊘(OUT) of the beaker.

c. Predict the sign of $\Delta S°$ for this reaction. Explain.

If the enthalpy is ~~above~~ >0 then
(s→l is ~~spo~~ endothermic)   it is a spontaneous process, ΔS is $+$
(more disorder from
s→aq)

d. Predict the sign of $\Delta G°$. Explain.

$\Delta S = \Delta \&H$
sign
reverse of ΔG

(negi)

⊘ $\underline{\Delta G°}$ is $-$ because it is a spontaneous
process

→soluble
Discuss your answer to part d. in light of the description of MX_2 as soluble.

When you go from a (S) to (aq), it is getting more disordered
and ~~has~~ is spontaneous, so ΔG is negative

2. Ammonia burns in the presence of a copper catalyst to form nitrogen gas.

$$NH_3(g) + O_2(g) \rightarrow N_2(g) + 6H_2O(g); \Delta H = -1.40 \text{ kJ}$$
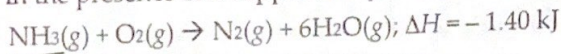
What is the enthalpy change when 48.1 g NH_3 burn?

$\Delta H = -1.40 \text{ kJ}$

$\Delta H = \dfrac{q}{n} \, \text{kJ}$

$q = (\Delta H)(n)$

$= (-1.40 \text{ kJ})(2.82434 \frac{1241)}{\text{mol}}$

$\boxed{= -3.95 \text{ kJ}}$ ✓ 15

Molar mass $NH_3 = 17.03052 \frac{g}{mol}$

$m_{NH_3} = 48.1 \text{ g}$

$n = \dfrac{m}{M} = 2.82434$

3. When 23.6 g $CaCl_2$ is dissolved in water in a constant–pressure calorimeter, the temperature changed from 25.00°C to 21.56°C. The heat capacity of the solution and the calorimeter is 1258 J/°C.

   a. What is the sign of q for the calorimeter? Explain.

$\Delta T = -3.44$

endothermic

$q = C\Delta T \rightarrow q$ is negative since the ΔT is negative ~~while the calorimeter~~ (in the calorimeter) surroundings

   b. What is the sign of q for the reaction? Explain.

q of the reaction is positive because → because this is ~~the~~ the system itself so if

$q_{calorimeter} = -q_{~~surround~~ reaction}$ the calorimeter had a negative temp. change and lost heat, the reaction gained it
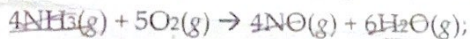
   c. What is ΔH for the reaction?

ΔH is positive for the reaction since it is endothermic and q is positive

$q = C\Delta T$
$= 1258(-3.44)$
$= -4.33 \times 10^3$
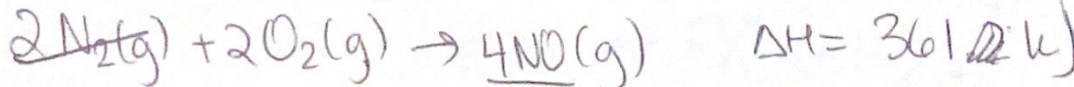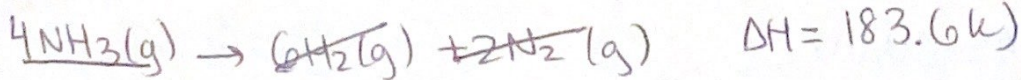$\Delta H = q/n = (-4.33\times10^3 / 0.213) = 20.4 kJ$

$\Delta H = (23.6g) \cdot (1258 J/°C)(-3.44°C)$
$\quad (110.978)$
$\quad = -20,349.97 J$
$\quad = -20.3 kJ$

4. Ammonia will burn in the presence of a platinum catalyst to form nitric oxide gas.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g);$$

Calculate the enthalpy change using the following data:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l); \Delta H° = -483.7 \text{ kJ}$$
$$2N_2(g) + 3H_2(g) \rightarrow 2NH_3(g); \Delta H° = -91.8 \text{ kJ} \quad \times 2$$
$$2N_2(g) + 2O_2(g) \rightarrow 2NO(g); \Delta H° = 180.6 \text{ kJ} \quad \times 2$$

$4NH_3(g) \rightarrow 6H_2(g) + 2N_2(g) \quad\quad \Delta H = 183.6 kJ$

$2N_2(g) + 2O_2(g) \rightarrow 4NO(g) \quad\quad \Delta H = 361.2 kJ$

$6H_2(g) + 3O_2(g) \rightarrow 6H_2O(g) \quad\quad \Delta H = -483.7 kJ \times 3$
$\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad = -1451.1 kJ$

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$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g) \quad\quad \Delta H = -906.3 kJ$
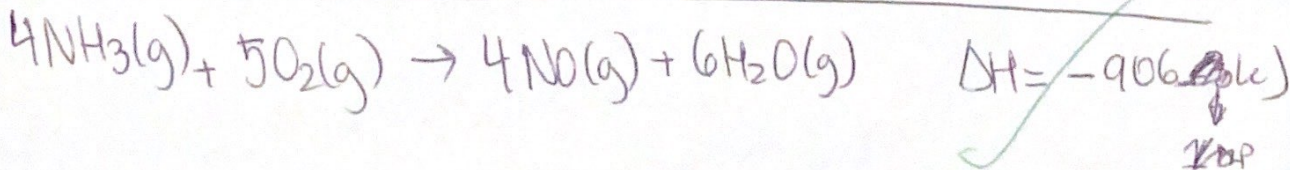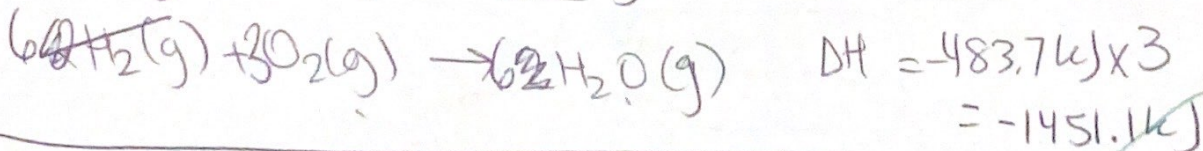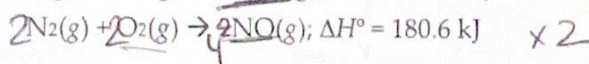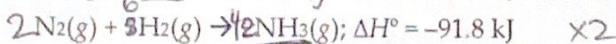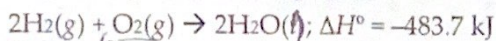
5. Given the balanced reaction :

$$2A(g) + B(g) \rightarrow 3C(g)$$
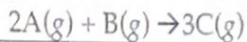
	ΔH_f° (kJ/mol)	S_f° (J/K)
A(g)	191	244
B(g)	70.8	300.
C(g)	-197	164

a. Find ΔH° for the reaction.

$$\Delta H^\circ_{rxn} = (3(C)) - [B + 2A] = (3 \cdot -197) - [70.8 + 2(191)]$$
$$= -591 - [452.8] = -1044. \frac{kJ}{mol}$$

b. Find ΔS° for the reaction

$$\Delta S^\circ = [3 \cdot C] - [B + 2 \cdot A] = [3 \cdot 164] - [300 + 2 \cdot 244]$$
$$= 492 - 788 = -296 \, J/k$$

c. Using your answer to parts a and b, find ΔG°

$$\Delta G = \Delta H - T\Delta S$$
$$\Delta G = (-1043.8) - (298)(-296)$$
$$\frac{1000}{}$$
$$= -10438 - [-88.208] = \cancel{932.0 kJ} \Rightarrow -955.6 \, kJ)$$

d. What does the value of ΔG° mean?

It means the free energy. the reaction is spontaneous since it's above < 0

e. Will the sign of ΔG change from low to high temperature?

~~No, if you have ΔS takes precedence at high temp.~~

Yes it does change

6. Given that ΔG_f° for HBr(g) = -53.50 kJ/mol, find the thermodynamic equilibrium constant for the reaction: H₂(g) + Cl₂(l) → 2HBr(g)

why divide by 1000? what if this was entropy?

$$\Delta G = \frac{-RT\ln k}{1000}$$

$$-53.50 = \frac{-(8.31451)(298)(\ln k)}{1000}$$
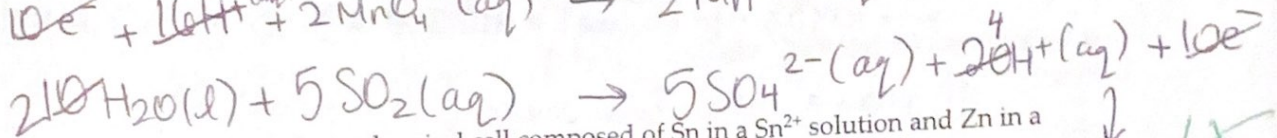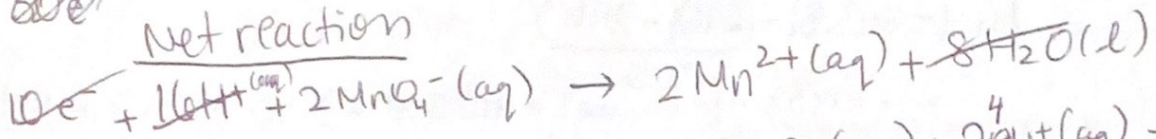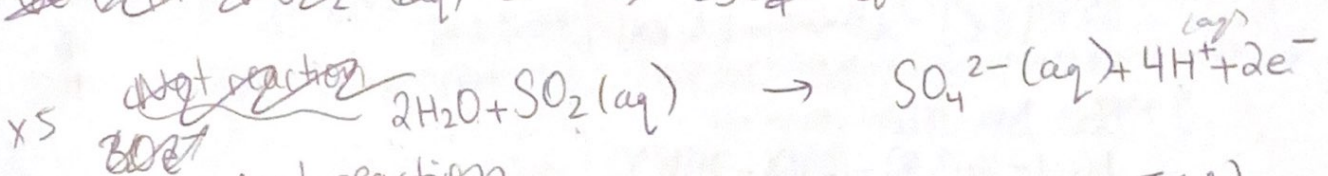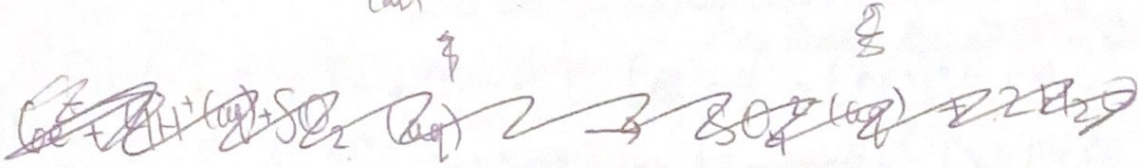
should be 2(-53.50) because formation is /mol
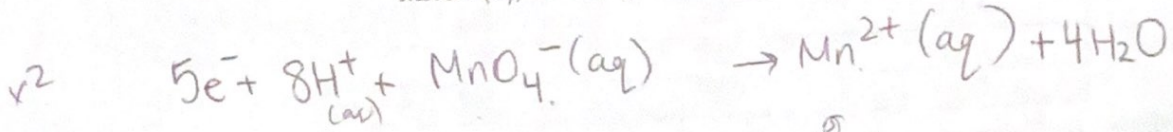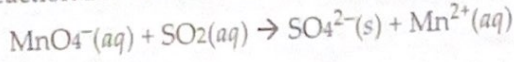
$$\Rightarrow 4SD \Rightarrow 2.385 \times 10^9$$

$$K_{th} = \underline{\hspace{3cm}}$$

4SD

CHM 1220/1225 - Exam V   B

7. Balance the following reaction in acidic conditions:

$$MnO_4^-(aq) + SO_2(aq) \rightarrow SO_4^{2-}(s) + Mn^{2+}(aq)$$

×2 $5e^- + 8H^+_{(aq)} + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O$

~~$6e^- + 8H^+(aq) + SO_2(aq) \rightarrow SO_4^{2-}(aq) + 22H_2O$~~

×5   ~~Net reaction~~ $2H_2O + SO_2(aq) \rightarrow SO_4^{2-}(aq) + 4H^+ + 2e^-$

Net reaction

$10e^- + 16H^+(aq) + 2MnO_4^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

$2\cancel{10}H_2O(l) + 5SO_2(aq) \rightarrow 5SO_4^{2-}(aq) + 2\cancel{0}H^+(aq) + 10e^-$

8. You are given an electrochemical cell composed of Sn in a Sn²⁺ solution and Zn in a Zn²⁺ solution.

Standard reduction potentials:

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E_{red}^\circ = -0.76$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $E_{red}^\circ = -0.14$ V

$2MnO_4^-(aq) + 5SO_2(aq) + 2H_2O(l) \rightarrow$
$2Mn^{2+}(aq) + 5SO_4^{2-}(aq) + 4H^+(aq)$
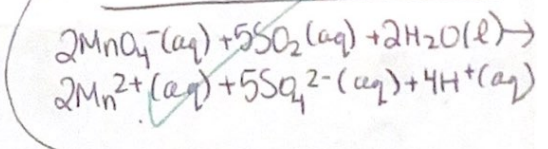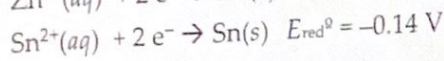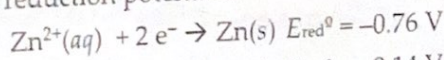
a. What is E_{cell}° for the spontaneous reaction.

$0.76V \rightarrow$ oxidation
$-0.14 V \rightarrow$ reduction $\Big\} \rightarrow 0.76 + (-0.14) = \boxed{0.62V}$

b. Find the thermodynamic equilibrium constant.

① $\Delta G = -nFe$

② K_{Th}? $\Delta G = -RT\ln K$

Spontaneous, so ΔG is <0

$= (-10)(96485)(0.62)$
2 for this q.

$= -598207 \dfrac{JV}{mol}$

$-598.207 = \dfrac{-(8.3145)(298)}{1000}(\ln k)$

$= -598.207 \dfrac{kJ·V}{mol}$

$\ne 0$

$= \dfrac{JV}{mol}$

c. Discuss the relationship between your values of E and K_{Th}.

$K>1$, $E^\circ cell >0$, $G<0$

$K<1$, $E^\circ cell <0$, $G>0$

E_{cell} is a spontaneous rxn, as it's +, K_{Th} favors products ($k>1$) causing a large K_{Th}

$K = 9.3\times10^{20}$

$G = -nFE = -1.2\times10^5 J$

$E_{cell}>0$, K should be $>1 \rightarrow$ favors products