

Exam 4A

Chem 1142

Spring 2017

Name: KEY

MULTIPLE CHOICE. [3 pts ea.] Record the best response on the scantron sheet. [45 pts total.]

Assume all solutions are aqueous and at a temperature of 25 °C, unless stated otherwise.

- Q1. Which version of the exam do you have?
a) 4A
b) 4B
- Q2. Which law of thermodynamics states that the entropy of the universe keeps increasing?
a) First law
b) Second law
c) Third law
d) Fourth law
- Q3. Which of the following substances would be expected to have the highest entropy at a given temperature?
a) H₂O(s)
b) Au(s)
c) Hg(l)
d) He(g)
- Q4. Which of the following chemical reactions would likely have a $\Delta S^\circ \approx 0$?
a) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
b) $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
c) $\text{CO}(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CH}_2\text{O}(\text{g})$
d) $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$
- Q5. A chemical reaction loses 25 kJ of heat to its surroundings at a temperature of 25 °C. What will the entropy change of the surrounding be?
a) -1 kJ/°C
b) +1000 J/°C
c) -84 J/K
d) +84 J/K
- Q6. An endothermic chemical reaction has $\Delta S^\circ_{\text{rxn}} < 0$. What can you say about the spontaneity of this reaction?
a) The reaction is always spontaneous
b) The reaction is always non-spontaneous
c) The reaction is spontaneous at low temperatures, but non-spontaneous at high temperatures
d) The reaction is non-spontaneous at low temperatures, but spontaneous at high temperatures

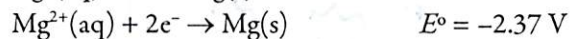
$$\Delta H = +ve$$

$$\Delta S = -ve$$

$$\Delta G = \Delta H - T \Delta S$$

- Q7. Which of the following substances will have a Gibbs free energy of formation of zero?
- $\text{H}_2(\text{l})$
 - $\text{CH}_4(\text{g})$
 - $\text{C}(\text{s, graphite})$
 - $\text{CO}_2(\text{s})$
- Q8. A reaction with a large and negative value of ΔG° will have an equilibrium constant, K , whereby which statement best applies:
- $K \gg 1$
 - $K \ll 1$
 - $K = 1$
 - $K = 0$
- Q9. At equilibrium, what can you say about the value of ΔG ?
- $\Delta G = 1$
 - $\Delta G > 0$
 - $\Delta G = 0$
 - $\Delta G \gg 1$
- Q10. The oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$ is:
- +8
 - +7
 - +6
 - +5
- Q11. Which of the following is **not** a redox reaction?
- $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
 - $\text{Zn}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
 - $\text{PCl}_3(\text{l}) + \text{Cl}_2(\text{g}) \rightarrow \text{PCl}_5(\text{l})$
 - $\text{KCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{KNO}_3(\text{aq})$
- Q12. Oxidation takes place at which part of a galvanic cell?
- Salt bridge
 - Voltmeter
 - Anode
 - Cathode

- Q13. Given the following two standard electrode potentials:



Which of the following species would be the best **oxidizing** agent?

- $\text{Ag}^+(\text{aq})$
- $\text{Ag}(\text{s})$
- $\text{Mg}^{2+}(\text{aq})$
- $\text{Mg}(\text{s})$

-cations loss e⁻ (-=> good @ gaining)

Q14. A spontaneous redox reaction would best be described as having:

- a) $E^\circ_{\text{cell}} > 0, \Delta G^\circ > 0$
- b) $E^\circ_{\text{cell}} > 0, \Delta G^\circ < 0$
- c) $E^\circ_{\text{cell}} < 0, \Delta G^\circ > 0$
- d) $E^\circ_{\text{cell}} < 0, \Delta G^\circ < 0$

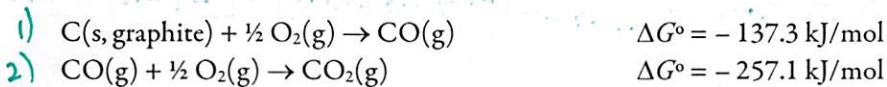
Q15. The charge on 2 moles of electrons is equal to:

- a) -193,000 C
- b) $-2 \times 6.022 \times 10^{23}$
- c) 96,500 C
- d) 2-

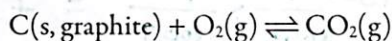
Short Response.

Show ALL work to receive credit.

Q16. [10 pts.] Given the following chemical equations:



Calculate the value of the equilibrium constant, K , at 25 °C for the reaction:



Be sure to show all work and explain clearly your solution.

$$1) + 2) : \text{ C} + \frac{1}{2} \text{ O}_2 + \cancel{\text{CO}} + \frac{1}{2} \text{ O}_2 \rightarrow \cancel{\text{CO}} + \text{CO}_2 ; \Delta G^\circ = -137.3 \frac{\text{kJ}}{\text{mol}} \oplus -257.1 \frac{\text{kJ}}{\text{mol}}$$



$$\Delta G^\circ = -RT \ln K \quad (+2) \quad (2)$$

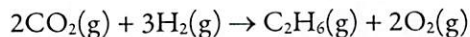
$$\Rightarrow \ln K = \frac{-\Delta G^\circ}{RT} \quad (+2) \quad (3)$$

$$\Rightarrow K = e^{-\Delta G^\circ/RT} \quad (+1) \quad (3)$$

$$= e^{\frac{+394,400 \text{ J/mol}}{8.3145 \text{ J/mol}\cdot\text{K} \times 298 \text{ K}}}$$

$$= e^{159.18} = 1.35 \times 10^{69} \quad (+3) \quad (3)$$

Q17. [15 pts.] (a) Show how to, then calculate, ΔG° at 45 °C for the chemical reaction:



Substance	CO ₂ (g)	H ₂ (g)	C ₂ H ₆ (g)	O ₂ (g)
ΔH_f° (kJ/mol)	-393.5	0	-84.7	0
S° (J/mol·K)	213.6	131.0	229.5	205.0

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta H^\circ = \sum \Delta H_f^\circ(\text{P}) - (\text{R}) = [1 \times \cancel{-84.7} + 2 \times 0] - [2 \times \cancel{-393.5} + 3 \times 0] = +702.3 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta S^\circ = \sum S^\circ(\text{P}) - (\text{R}) = [1 \times 229.5 + 2 \times 205.0] - [2 \times 213.6 + 3 \times 131.0] = -180.7 \frac{\text{J}}{\text{mol}\cdot\text{K}}$$

$$T = 45 + 273 = 318\text{K}$$

$$\Delta G^\circ = +702.3 \frac{\text{kJ}}{\text{mol}} - 318\text{K} \times -180.7 \frac{\text{J}}{\text{mol}\cdot\text{K}} \times \frac{1\text{kJ}}{1000\text{J}} = +759.76 \frac{\text{kJ}}{\text{mol}}$$

$$= 757.95 \frac{\text{kJ}}{\text{mol}}$$

(b) Comment on the value you obtained from part (a), and its meaning.

the ΔG° means Non-spontaneous.

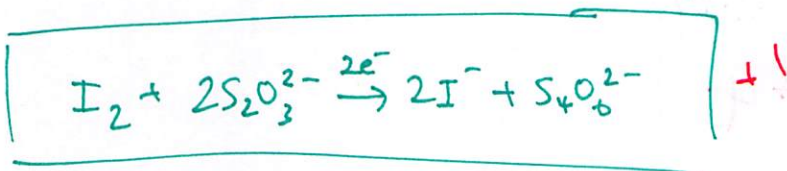
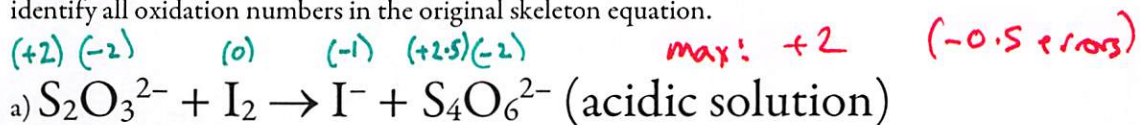
(would need to provide 760 kJ of energy to allow rxn to happen)

(c) Explain what will happen to the reaction as the temperature is increased.

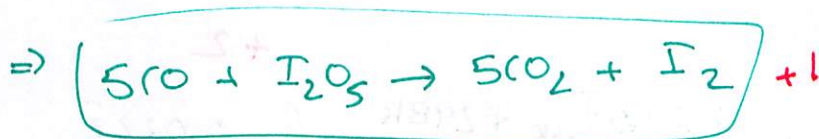
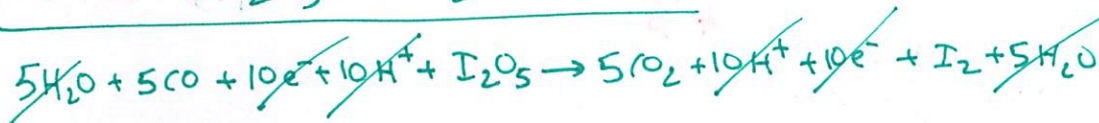
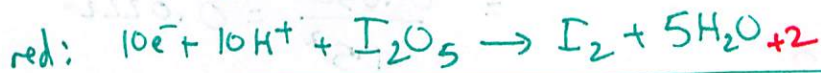
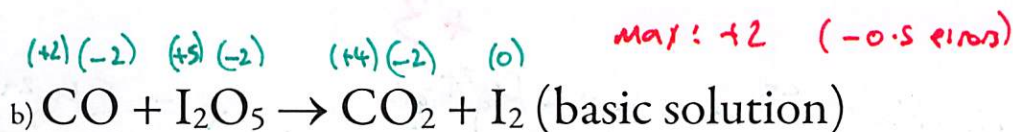
Since both ΔH° is +ve and $-\Delta S^\circ$ is -ve, will always

be have +ve ΔG° no matter what T is!

Q18. [15 pts.] Balance the following two redox equations using the half-reaction method. Be sure to clearly identify all oxidation numbers in the original skeleton equation.

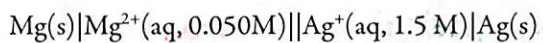


(7)



(8)

Q19. [15 pts.] Calculate E_{cell} for the following cell diagram:

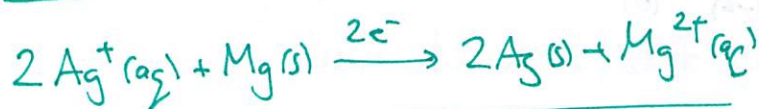


As part of your answer you should calculate E°_{cell} and also write the overall balanced chemical equation. Assume the cell temperature is 25 °C.

$$E^{\circ}_{\text{cell}} = E_{\text{Ag}^{+}/\text{Ag}} - E^{\circ}_{\text{Mg}^{2+}/\text{Mg}}$$

$$= +0.80\text{V} - (-2.37\text{V})$$

$$= +3.17\text{V} \quad +4$$



$$Q = \frac{[\text{Mg}^{2+}]}{[\text{Ag}^{+}]^2} \quad +2$$

$$= \frac{0.050}{1.5^2} = 0.022\bar{2}$$

Nernst:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q \quad +3$$

$$= +3.17\text{V} - \frac{8.3145 \frac{\text{J}}{\text{mol}\cdot\text{K}} \times 298\text{K}}{2 \times 96,500 \frac{\text{C}}{\text{mol}}} \ln 0.022\bar{2} \quad +2$$

$$= +3.17\text{V} - -0.0489\text{V} \quad (\text{IV} = \text{I}^{\circ}/\text{C})$$

$$= +3.17\text{V} + 0.0489\text{V}$$

$$= \boxed{3.22\text{V}} \quad +1$$

Periodic Table of the Elements

IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA																												
1 H 1.008	2 He 4.003											3 B 10.81	4 C 12.01	5 N 14.01	6 O 16.00	7 F 19.00	8 Ne 20.18																												
3 Li 6.941	4 Be 9.012											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95																												
11 Na 22.99	12 Mg 24.31	3 Sc 44.96	4 Ti 47.87	5 V 50.94	6 Cr 52.00	7 Mn 54.94	8 Fe 55.85	9 Co 58.93	10 Ni 58.69	11 Cu 63.55	12 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.90	36 Kr 83.80																												
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.60	53 I 126.9	54 Xe 131.3																												
55 Cs 132.9	56 Ba* 137.3	71 Lu 175.0	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po [210]	85 At [210]	86 Rn [222]																												
87 Fr [223]	88 Ra** [226]	103 Lr [262]	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [265]	109 Mt [269]	110 [269]	111 [272]	112 [277]	113 [285]	114 [285]	115 [289]	116 [289]	117 [293]	118 [293]																												
<table border="1" style="width: 100%; border-collapse: collapse; text-align: center;"> <tr> <td>57 La 138.9</td> <td>58 Ce 140.1</td> <td>59 Pr 140.9</td> <td>60 Nd 144.2</td> <td>61 Pm [145]</td> <td>62 Sm 150.4</td> <td>63 Eu 152.0</td> <td>64 Gd 157.3</td> <td>65 Tb 158.9</td> <td>66 Dy 162.50</td> <td>67 Ho 164.9</td> <td>68 Er 167.3</td> <td>69 Tm 168.9</td> <td>70 Yb 173.0</td> </tr> <tr> <td>89 Ac [227]</td> <td>90 Th 232.0</td> <td>91 Pa 231.0</td> <td>92 U 238.0</td> <td>93 Np [237]</td> <td>94 Pu [244]</td> <td>95 Am [243]</td> <td>96 Cm [247]</td> <td>97 Bk [247]</td> <td>98 Cf [251]</td> <td>99 Es [252]</td> <td>100 Fm [257]</td> <td>101 Md [258]</td> <td>102 No [259]</td> </tr> </table>																		57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm [145]	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.50	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	89 Ac [227]	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]
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$$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$$

$$R = 8.3145 \text{ J/mol}\cdot\text{K} = 0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K}$$

$$\Delta G = -nFE_{\text{cell}}$$

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

$$\Delta G = \Delta H - T\Delta S$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$F = 96,500 \text{ C/mol } e^- \quad 1 \text{ V} = 1 \text{ J/C}$$

$$\Delta S_{\text{surr}} = q_{\text{surr}}/T$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\Delta G^\circ = -RT \ln K$$

Partial list of standard electrode potentials:

