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) $H_2O(s)$
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) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
b) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
c) $CO(g) + H_2(g) \rightarrow CH_2O(g)$
d) $H_2O(g) \rightarrow H_2O(l)$

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

ΔH = +ve ΔS = –ve $\Delta G = \Delta H - T\Delta S$

Q6. An endothermic chemical reaction has $\Delta S^\circ_{rxn} < 0$. 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

Q7. Which of the following substances will have a Gibbs free energy of formation of zero?
a) $H_2(l)$
b) $CH_4(g)$
c) C(s, graphite)
d) $CO_2(s)$

Q8. A reaction with a large and negative value of $\Delta G°$ will have an equilibrium constant, K, whereby which statement best applies:
a) $K \gg 1$
b) $K \ll 1$
c) $K = 1$
d) $K = 0$

Q9. At equilibrium, what can you say about the value of ΔG?
a) $\Delta G = 1$
b) $\Delta G > 0$
c) $\Delta G = 0$
d) $\Delta G \gg 1$

Q10. The oxidation number of chromium in $Cr_2O_7^{2-}$ is:
a) +8
b) +7
c) +6
d) +5

Q11. Which of the following is **not** a redox reaction?
a) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
b) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
c) $PCl_3(l) + Cl_2(g) \rightarrow PCl_5(l)$
d) $KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$

Q12. Oxidation takes place at which part of a galvanic cell?
a) Salt bridge
b) Voltmeter
c) Anode
d) Cathode

Q13. Given the following two standard electrode potentials:

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E° = +0.80$ V
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$ $E° = -2.37$ V

– causes loss e^- (⇒ good @ gaining)

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

a) $Ag^+(aq)$
b) $Ag(s)$
c) $Mg^{2+}(aq)$
d) $Mg(s)$

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

a) $E^o_{cell} > 0, \Delta G^o > 0$
b) $E^o_{cell} > 0, \Delta G^o < 0$
c) $E^o_{cell} < 0, \Delta G^o > 0$
d) $E^o_{cell} < 0, \Delta G^o < 0$

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:

1) $C(s, graphite) + \frac{1}{2} O_2(g) \rightarrow CO(g)$ $\Delta G^o = -137.3$ kJ/mol
2) $CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$ $\Delta G^o = -257.1$ kJ/mol

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

$$C(s, graphite) + O_2(g) \rightleftharpoons CO_2(g)$$

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

1) + 2) : $C + \frac{1}{2}O_2 + \cancel{CO} + \frac{1}{2}O_2 \rightarrow \cancel{CO} + CO_2$; $\Delta G^\circ = -137.3 \frac{kJ}{mol} \oplus -257.1 \frac{kJ}{mol}$

or $C(s, graph) + O_2(g) \rightarrow CO_2(g)$; $\Delta G^\circ = -394.4$ kJ/mol (+4) 4

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

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

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

$= e^{\frac{+394{,}400 \, J/mol}{8.3145 \, J/mol \cdot K \times 298K}}$ 3

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

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

$$2CO_2(g) + 3H_2(g) \rightarrow C_2H_6(g) + 2O_2(g)$$

Substance	$CO_2(g)$	$H_2(g)$	$C_2H_6(g)$	$O_2(g)$
ΔH_f° (kJ/mol)	–393.5	0	–84.7	0
S° (J/mol·K)	213.6	131.0	229.5	205.0

3

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

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

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

$T = 45 + 273 = 318K$

$\Delta G^\circ = +702.3 \frac{kJ}{mol} - 318K \times -180.7 \frac{J}{mol\cdot K} \times \frac{1kJ}{1000J} = +759.76 \, kJ/mol$ 3

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

(b) Comment on the value you obtained from part (a), and its meaning. @ 35°C 3

+ve ΔG° means Non-spontaneous.
(Would need to provide 760kJ 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 ΔS° (is –ve), will always have +ve ΔG° no matter what T is! 3

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.

(+2) (−2) (0) (−1) (+2.5)(−2) max: +2 (−0.5 error)

a) $S_2O_3^{2-} + I_2 \rightarrow I^- + S_4O_6^{2-}$ (acidic solution)

red: $2e^- + I_2 \rightarrow 2I^-$ +2

ox: $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$ +2 −0.5 error

$\boxed{I_2 + 2S_2O_3^{2-} \xrightarrow{2e^-} 2I^- + S_4O_6^{2-}}$ +1

7

(+2) (−2) (+5) (−2) (+4) (−2) (0) max: +2 (−0.5 error)

b) $CO + I_2O_5 \rightarrow CO_2 + I_2$ (basic solution)

ox: $(H_2O + CO \rightarrow CO_2 + 2H^+ + 2e^-)5$ +2 +1

red: $10e^- + 10H^+ + I_2O_5 \rightarrow I_2 + 5H_2O$ +2

$\cancel{5H_2O} + 5CO + \cancel{10e^-} + \cancel{10H^+} + I_2O_5 \rightarrow 5CO_2 + \cancel{10H^+} + \cancel{10e^-} + I_2 + \cancel{5H_2O}$

$\Rightarrow \boxed{5CO + I_2O_5 \rightarrow 5CO_2 + I_2}$ +1

8

/15

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

$Mg(s)|Mg^{2+}(aq, 0.050M)||Ag^{+}(aq, 1.5\ M)|Ag(s)$

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

$$E^{o}_{cell} = E_{Ag^+/Ag} - E^{o}_{Mg^{2+}/Mg}$$

$$= +0.80V - -2.37V$$

$$= +3.17V \quad +4$$

$$(e^- + Ag^+ \rightarrow Ag) \times 2$$

$$Mg \rightarrow Mg^{2+} + 2e^-$$

+3 +2

$$2Ag^+(aq) + Mg(s) \xrightarrow{2e^-} 2Ag(s) + Mg^{2+}(aq) \quad , \quad Q = \frac{[Mg^{2+})}{[Ag^+]_o^2}$$

Nernst:

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

+3

$$E_{cell} = E^{o}_{cell} - \frac{RT}{nF} \ln Q$$

+2

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

$$= +3.17V - -0.0489\ J/C \qquad (1V = 1\ J/C)$$

$$= +3.17V + 0.0489V$$

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

Useful Information

Periodic Table of the Elements

1 (IA)	2 (IIA)	3	4	5	6	7	8	9	10	11	12	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **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** [268]	110 [269]	111 [272]	112 [277]	113	114 [285]	115	116 [289]	117	118 [293]

*	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]

$N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$ $\quad R = 8.3145\ \text{J/mol·K} = 0.08206\ \text{L·atm/mol·K}$

$\Delta G = -nFE_{cell}$ $\quad \Delta G° = -nFE°_{cell}$ $\quad \Delta G = \Delta H - T\Delta S$

$$E_{cell} = E°_{cell} - \frac{RT}{nF}\ln Q$$

$E°_{cell} = E°_{cathode} - E°_{anode}$ $\quad F = 96{,}500\ \text{C/mol e}^-$ $\quad 1\ \text{V} = 1\ \text{J/C}$

$\Delta S_{surr} = q_{surr}/T$ $\quad \Delta G = \Delta G° + RT\ln Q$ $\quad \Delta G° = -RT\ln K$

Partial list of standard electrode potentials:

$Au^{3+} + 3e^- \rightarrow Au \quad E° = +1.50\ \text{V}$

$Ag^+ + e^- \rightarrow Ag \quad E° = +0.80\ \text{V}$

$Cu^{2+} + 2e^- \rightarrow Cu \quad E° = +0.34\ \text{V}$

$Pb^{2+} + 2e^- \rightarrow Pb \quad E° = -0.13\ \text{V}$

$Mg^{2+} + 2e^- \rightarrow Mg \quad E° = -2.37\ \text{V}$

$Li^+ + e^- \rightarrow Li \quad E° = -3.05\ \text{V}$