CHM 2046 **Quiz 3A Answer Key** **March 28, 2017**

Answer all questions. Show your work. Look at the end for useful information.

1. Consider the two reactions below at 25°C.

$2\ Cu_2O\ (s) \rightarrow 4\ Cu\ (s) + O_2\ (g)$ $\Delta G° = +292$ kJ

$CO_2\ (g) \rightarrow CO\ (g) + \frac{1}{2}O_2\ (g)$ $\Delta G° = +257$ kJ

Flip second eq and x2 $O_2(g) + 2CO\ (g) \rightarrow 2CO_2\ (g)$ $\Delta G° = -514$ kJ

a) (4 pts) What is the overall reaction obtained from coupling together the above reactions to give Cu (s) as a product (no $O_2(g)$ should appear)?

$2\ Cu_2O\ (s) + 2\ CO\ (g) \rightarrow 4\ Cu\ (s) + 2CO_2\ (g)$

b) (3 pts) What is $\Delta G°$ for the reaction in part a) at 25°C?

$\Delta G° = +292$ kJ -514 kJ $= -222$ kJ $\therefore$ **$\Delta G° = -222$ kJ**

c) (3 pts) What is the equilibrium constant K for the reaction in a) at 25°C?

$\Delta G° = -RT\ln K$ $\therefore \ln K = -\Delta G°/RT = -(-222 \times 10^3\ J)/(8.314 J/mol.K)(298K)$

$\therefore \ln K = 89.60$ $\therefore$ **$K = 8.2 \times 10^{38}$**

2. (5 pts) The reaction below at 298 K has $\Delta G° = 191$ kJ.

$2\ HCl\ (g) \rightleftharpoons H_2\ (g) + Cl_2\ (g)$

What is ΔG for this reaction at 298 K when the partial pressures (p) of each species are as follows: $p(Cl_2) = 0.6$ atm, $p(H_2) = 4$ atm, $p(HCl) = 0.2$ atm?

$\Delta G = \Delta G° + RT\ln Q$ $Q = \frac{p(H_2)p(Cl_2)}{p(HCl)^2} = \frac{(4)(0.6)}{(0.2)^2} = 60$

$\Delta G = 191000J + (8.314\ J/mol.K)(298K)\ln(60) = 191000J + 10144\ J = 201144\ J$

$\Delta G = 201$ kJ for the reaction as written

3. a) (5 pts) Calculate the $\Delta G°$ at 298K for the reaction:

$CH_4(g) + 2H_2O(g) \rightleftharpoons CO_2(g) + 4H_2(g)$

Substance:	$CH_4(g)$	$H_2O(g)$	$CO_2(g)$	$H_2(g)$
$\Delta H°_f$ (kJ/mol):	-74.87	-241.8	-393.5	0
$S°$(J/K·mol):	186.1	188.8	213.7	130.7

$\Delta H° = 1mol(-393.5)) + 0 - 1mol(-74.87) - 2mol(-241.8) = 164.97$ kJ/mol

$\Delta S° = 1mol(213.7) + 4mol(130.7) - 1mol(186.1) - 2mol(188.8) = 172.8$ J/mol

$\Delta G° = \Delta H° - T\Delta S° = (164.97\ kJ/mol) - (298K)(172.8 \times 10^{-3}\ kJ) = 113.5$ kJ/mol

b) (3 pts) Above what temperature will the reaction become spontaneous?

$T = \Delta H°/\Delta S° = 164.97\ kJ/mol / 172.8 \times 10^{-3}\ kJ/mol = 954.7$ K

4. For the reaction $Cr^{2+}(aq) + 2\ H^+(aq) + SO_3(g) \rightleftharpoons SO_2(g) + Cr^{3+}(aq) + H_2O(l)$

Complete the following statements (for the reaction from left-to-right).

(i) (1 pt) The oxidizing agent is (**SO_3**).

(ii) (1 pt) The substance oxidized is (**Cr^{2+}**).

(iii) (1 pt) The substance reduced is (**SO_3**).

Complete the following statements (for the reaction from right-to-left).

(iv) (1 pt) The substance oxidized is (**SO_2**).

(v) (1 pt) The reducing agent is (**SO_2**).

(vi) (1 pt) The substance reduced is (Cr^{3+}).

5. (10 points) Balance the following redox reaction in basic solution (you <u>must</u> show the half-reactions and other steps in getting to the final answer.

$$Cr_2O_7^{2-}\ (aq) + I^-\ (aq) \rightarrow IO_4^-\ (aq) + Cr(OH)_3\ (s)$$

$I^- \rightarrow IO_4^- + x\ e^-$ **($I^{1-} \rightarrow I^{7+} + 8\ e^-$ oxidation)**

$Cr_2O_7^{2-} + y\ e^- \rightarrow Cr(OH)_3$ **($Cr^{6+} + 3e^- \rightarrow Cr^{3+}$ reduction)**

Add electrons, balance and make number of e^- the same:

$I^- \rightarrow IO_4^- + 8\ e^-$ **$\therefore\ 3\ I^- \rightarrow 3\ IO_4^- + 24\ e^-$**

$Cr_2O_7^{2-} + 6\ e^- \rightarrow 2\ Cr(OH)_3$ **$\therefore\ 4\ Cr_2O_7^{2-} + 24\ e^- \rightarrow 8\ Cr(OH)_3$**

Now add together: **$4\ Cr_2O_7^{2-} + 3\ I^- \rightarrow 3\ IO_4^- + 8\ Cr(OH)_3$**

Balance O with H_2O: **$4\ Cr_2O_7^{2-} + 3\ I^- + 8\ H_2O \rightarrow 3\ IO_4^- + 8\ Cr(OH)_3$**

Balance H with H^+: **$4\ Cr_2O_7^{2-} + 3\ I^- + 8\ H_2O + 8H^+ \rightarrow 3\ IO_4^- + 8\ Cr(OH)_3$**

Add OH^-: **$4Cr_2O_7^{2-} + 3I^- + 8H_2O + 8H^+ + 8OH^- \rightarrow 3\ IO_4^- + 8\ Cr(OH)_3 + 8OH^-$**

Balanced Equation: $\boxed{4\ Cr_2O_7^{2-} + 3\ I^- + 16\ H_2O \rightarrow 3\ IO_4^- + 8\ Cr(OH)_3 + 8\ OH^-}$

6. **a)** Consider a voltaic cell with half-reactions:

$$MnO_4^-\ (aq) + 8\ H^+\ (aq) + 4\ e^- \rightarrow Mn^{3+}\ (aq) + 4H_2O\ (l)$$

$$Pb\ (s) \rightarrow Pb^{2+}\ (aq) + 2e^-$$

(5 pts) The cell notation for this cell is

$Pb(s)\ |\ Pb^{2+}(aq)\ \|\ MnO_4^-(aq),\ H^+(aq),\ Mn^{3+}(aq)\ |\ Pt(s)$ or graphite(s)

<u>or</u> **$Pb(s)\ |\ Pb^{2+}(aq)\ \|\ H^+\ (aq),\ MnO_4^-(aq),\ Mn^{3+}(aq)\ |\ Pt(s)$ or graphite(s)**

(2 pts) The overall reaction of the cell is

$2\ Pb\ (s) + MnO_4^-\ (aq) + 8\ H^+\ (aq) \rightarrow Mn^{3+}\ (aq) + 2\ Pb^{2+}\ (aq) + 4H_2O(l)$

b) Consider a voltaic cell whose overall reaction is

$$2Cr\ (s) + 3\ I_2(s) \rightarrow 2\ Cr^{3+}(aq) + 6\ I^-(aq)$$

(i) (2 pts) The cathode half-reaction is:

$I_2(s) + 2\ e^- \rightarrow 2\ I^-\ (aq)$

(ii) (2 pts) The anode half-reaction is :

$Cr\ (s) \rightarrow Cr^{3+}(aq) + 3\ e^-$

$R = 8.314$ J/mol.K	$pH + pOH = 14.00$	$\Delta G = \Delta G° + RTlnQ$
$pK_a = -\log K_a$	$\Delta S_{rxn} + \Delta S_{surr} = \Delta S_{univ}$	$\Delta G = -nFE°_{cell}$
$F = 9.65 \times 10^4$ J/V.mol e^-	$\Delta G° = -RTlnK$	$pH = -\log[H^+]$