**CHM 2046** **Quiz 3B 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.

$2\ Ag_2O\ (s) \rightarrow 4\ Ag\ (s) + O_2\ (g)$ $\Delta G° = +152$ 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)?

**$2\ Ag_2O\ (s) + 2\ CO\ (g) \rightarrow 4\ Ag\ (s) + 2CO_2\ (g)$**

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

$\Delta G° = +152$ kJ $-514$ kJ $= -222$ kJ $\therefore$ **$\Delta G° = -362$ kJ**

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

$\Delta G° = -RT\ln K$ $\therefore \ln K = -\Delta G°/RT = -(-362 \times 10^3\ J)/(8.314 J/mol.K)(298K)$

$\therefore \ln K = 146.11$ $\therefore$ **$K = 2.8 \times 10^{63}$**

**2.** (5 pts) The reaction below at 320 K has $\Delta G° = 191$ kJ.

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

What is $\Delta G$ for this reaction at 320 K when the partial pressures (p) of each species are as follows: $p(Cl_2) = 0.3$ atm, $p(H_2) = 6$ atm, $p(HCl) = 0.3$ atm?

$\Delta G = \Delta G° + RT\ln Q$ $Q = \dfrac{p(H_2)p(Cl_2)}{p(HCl)^2} = [(6)X(0.3)]/(0.6)^2 = 5$

$\Delta G = 191000J + (8.314\ J/mol.K)(298K)\ln(60) = 191000J + 4282\ J = 195282\ J$

**$\Delta G = 195$ 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

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

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 reducing agent is ( **$Cr^{2+}$** ).

(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 oxidizing agent is ( **$Cr^{3+}$** ).

(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

$$2Fe(s) + 3\,Br_2(l) \rightarrow 2\,Fe^{3+}(aq) + 6\,Br^-(aq)$$

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

**$Br_2(l) + 2\,e^- \rightarrow 2\,Br^-(aq)$**

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

**$Fe(s) \rightarrow Fe^{3+}(aq) + 3\,e^-$**

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| | | |
|---|---|---|
| $R = 8.314$ J/mol.K | $pH + pOH = 14.00$ | $\Delta G = \Delta G^\circ + RT\ln Q$ |
| $pK_a = -\log K_a$ | $\Delta S_{rxn} + \Delta S_{surr} = \Delta S_{univ}$ | $\Delta G = -nFE^\circ_{cell}$ |
| $F = 9.65 \times 10^4$ J/V.mol $e^-$ | $\Delta G^\circ = -RT\ln K$ | $pH = -\log[H^+]$ |