Answer all questions. Look at the end for useful information.

1. a) (3 points) Consider the following reaction: $N_2(g) + O_2(g) \rightarrow 2\,NO(g)$
If $\Delta H° = 180.58$ kJ and $\Delta S° = 24.8$ J/K, calculate $\Delta G°$ for this reaction at 300 K.

$$\Delta G° = \Delta H° - T\Delta S°$$
$$= 180.58\,kJ - (300\,K)(24.8\times10^{-3}\,kJ/K)$$
$$\therefore \Delta G° = 173\,kJ$$

b) (1 point) The reaction in part a) is an example of one that switches from being spontaneous to being non-spontaneous at the crossover temperature. Answer the following question by deleting the incorrect choice.

The reaction is spontaneous at ~~low~~/high *(delete one)* temperatures, and non-spontaneous at low/~~high~~ *(delete one)* temperatures.

c) (4 points) Calculate the temperature at which the reaction switches from being spontaneous to being non-spontaneous.

Switch-over temperature when $\Delta G° = 0$.
$$\therefore \Delta G° = \Delta H° - T\Delta S° = 0$$
$$\therefore \Delta H° = T\Delta S°$$
$$\therefore T = \frac{\Delta H°}{\Delta S°} = \frac{180.58\,kJ}{24.8\times10^{-3}\,kJ/K} = 7281\,K$$

2. Consider the two reactions below at 300 K.

$CH_3OH(g) \rightarrow CO(g) + 2\,H_2(g)$ $\Delta G° = 24.4$ kJ
$H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$ $\Delta G° = -190.6$ kJ

a) (3 points) What is the resulting reaction obtained from coupling the above reactions together? ($H_2(g)$ should not appear in the reaction.)

multiply second one by 2, and add
$$CH_3OH(g) + \cancel{2H_2(g)} + 2Cl_2(g) \rightarrow CO(g) + \cancel{2H_2(g)} + 4HCl(g)$$
$$\therefore CH_3OH(g) + 2Cl_2(g) \rightarrow CO(g) + 4HCl(g)$$

b) (3 points) What is $\Delta G°$ for the reaction in part a)?

When equations are added together, their $\Delta G°$'s ~~are~~ also added together (remember to multiply second one by 2)
$$\Delta G° = 24.4\,kJ + 2(-190.6\,kJ) = -356.8\,kJ$$

c) (4 points) What is the equilibrium constant of this reaction in part a) at 300 K?

$$\Delta G° = -RT\ln K$$
$$\therefore \ln K = \frac{-\Delta G°}{RT} = \frac{356.8\,kJ}{(8.314\times10^{-3})kJ/K(300\,K)} = 143.05 \quad \therefore K = 1.4\times10^{62}$$

Consider the following reaction at 298 K

3. Consider the following reaction at 298 K: $2\,NO(g) + Cl_2(g) \rightarrow 2\,NOCl(g)$
$\Delta G°_f$ values: NO (g) = 86.60 kJ/mol; Cl_2 (g) = 0.000 kJ/mol; NOCl (g) = 66.07 kJ/mol.

a) (4 points) Calculate $\Delta G°$ for this reaction.

$$\Delta G° = 2(66.07\,kJ) - 2(86.60\,kJ) - 0.00\,kJ$$
$$= -41.06\,kJ$$

b) (7 points) What is ΔG for this reaction at 298 K when the partial pressures (p) of each species are as follows: p(NO) = 1.2 atmospheres, p(Cl_2) = 0.22 atmospheres, p(NOCl) = 2.1 atmospheres?

$$\Delta G = \Delta G° + RT\ln Q$$
$$= -41.06\,kJ + (8.314\times10^{-3}\,kJ/K)(298\,K)\ln\frac{(2.1)^2}{(1.2)^2(0.22)}$$
$$\therefore \Delta G = -41.06\,kJ + 6.52\,kJ$$
$$\therefore \Delta G = -34.54\,kJ$$

4. (8 points) Balance the following redox reaction in basic solution. You <u>must</u> show the steps involved in reaching the final answer.

$$VO_4^{3-}\ (aq) + N_2O\ (g) \rightarrow V^{3+}\ (aq) + NO_3^-\ (aq)$$

$VO_4^{3-} \rightarrow V^{3+}$ $(V^{5+} \rightarrow V^{3+})$ $\quad N_2O \rightarrow NO_3^-$ $(N^{+} \rightarrow N^{5+})$

balance V/N $\left[\begin{array}{l} VO_4^{3-} \rightarrow V^{3+} \\ N_2O \rightarrow 2NO_3^- \end{array}\right]$ Add e⁻ $\left[\begin{array}{l} VO_4^{3-} + 2e^- \rightarrow V^{3+} \\ N_2O \rightarrow 2NO_3^- + 8e^- \end{array}\right]$

equalize e⁻ $4VO_4^{3-} + 8e^- \rightarrow 4V^{3+}$ and $N_2O \rightarrow 2NO_3^- + 8e^-$

Add $4VO_4^{3-} + N_2O \rightarrow 4V^{3+} + 2NO_3^-$

balance O $4VO_4^{3-} + N_2O \rightarrow 4V^{3+} + 2NO_3^- + 11H_2O$

balance H $4VO_4^{3-} + N_2O + 22H^+ \rightarrow 4V^{3+} + 2NO_3^- + 11H_2O$

Add OH⁻ $4VO_4^{3-} + N_2O + 22H_2O \rightarrow 4V^{3+} + 2NO_3^- + 11H_2O + 22OH^-$

cancel H_2O $\boxed{4VO_4^{3-} + N_2O + 11H_2O \rightarrow 4V^{3+} + 2NO_3^- + 22OH^-}$

5. For the redox reaction

$$Fe\ (s) + 2\ Ag^+\ (aq) \rightarrow Ag\ (s) + Fe^{2+}\ (aq)$$

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

(i) (1 point) The oxidizing agent is (Ag^+).

(ii) (1 point) The substance reduced is (Ag^+).

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

(iii) (1 point) The substance oxidized is (Ag (s)).

(iv) (1 point) The oxidizing agent is (Fe^{2+}).

6. Consider the voltaic cell whose notation is

$$Pb\ (s)\ |\ Pb^{2+}\ (aq)\ \|\ Sn^{4+}\ (aq),\ Sn^{2+}\ (aq)\ |\ Pt(s)$$

(i) (3 points) The anode half-reaction is:

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

(ii) (3 points) The cathode half-reaction is :

$$Sn^{4+}\ (aq) + 2e^- \rightarrow Sn^{2+}\ (aq)$$

(iii) (3 points) The overall spontaneous reaction occurring in this cell is:

$$Pb\ (s) + Sn^{4+}\ (aq) \rightarrow Pb^{2+}\ (aq) + Sn^{2+}\ (aq)$$

$\Delta G = \Delta H - T\Delta S$	$pK_a = -\log K_a$	$\Delta G° = -nFE°_{cell}$
$pH = -\log[H^+]$	$\Delta G = \Delta G° + RT\ln Q$	$R = 8.314$ J/mol.K.
$\Delta S_{rxn} + \Delta S_{surr} = \Delta S_{univ}$	$\Delta G° = -RT\ln K$	$pH + pOH = 14.00$