

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

1. a) (3 points) Consider the following reaction: $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$
If $\Delta H^\circ = 180.58 \text{ kJ}$ and $\Delta S^\circ = 24.8 \text{ J/K}$, calculate ΔG° for this reaction at 300 K.

$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= 180.58 \text{ kJ} - (300 \text{ K})(24.8 \times 10^{-3} \text{ kJ/K}) \\ \therefore \Delta G^\circ &= 173 \text{ kJ}\end{aligned}$$

- 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.

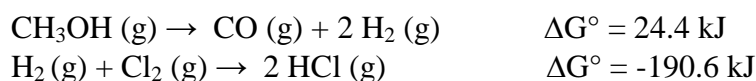
The reaction is spontaneous at **low**/high (delete one) temperatures, and non-spontaneous at low/**high** (delete one) temperatures.

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

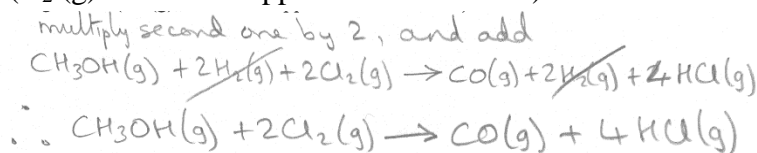
Switch-over temperature when $\Delta G^\circ = 0$.

$$\begin{aligned}\therefore \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ = 0 \\ \therefore \Delta H^\circ &= T\Delta S^\circ \\ \therefore T &= \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{180.58 \text{ kJ}}{24.8 \times 10^{-3} \text{ kJ/K}} = 7281 \text{ K}\end{aligned}$$

2. Consider the two reactions below at 300 K.



- a) (3 points) What is the resulting reaction obtained from coupling the above reactions together? ($\text{H}_2(\text{g})$ should not appear in the reaction.)



- b) (3 points) What is ΔG° for the reaction in part a)?

When equations are added together, their ΔG° 's are also added together (remember to multiply second one by 2)

$$\Delta G^\circ = 24.4 \text{ kJ} + 2(-190.6 \text{ kJ}) = -356.8 \text{ kJ}$$

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

$$\begin{aligned}\Delta G^\circ &= -RT \ln K \\ \therefore \ln K &= \frac{-\Delta G^\circ}{RT} = \frac{356.8 \text{ kJ}}{(8.314 \times 10^{-3} \text{ kJ/K})(300 \text{ K})} = 143.05 \\ \therefore K &= 1.4 \times 10^{62}\end{aligned}$$

Consider the following reaction at 298 K

3. Consider the following reaction at 298 K: $2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NOCl}(\text{g})$
 ΔG°_f values: $\text{NO}(\text{g}) = 86.60 \text{ kJ/mol}$; $\text{Cl}_2(\text{g}) = 0.000 \text{ kJ/mol}$; $\text{NOCl}(\text{g}) = 66.07 \text{ kJ/mol}$.

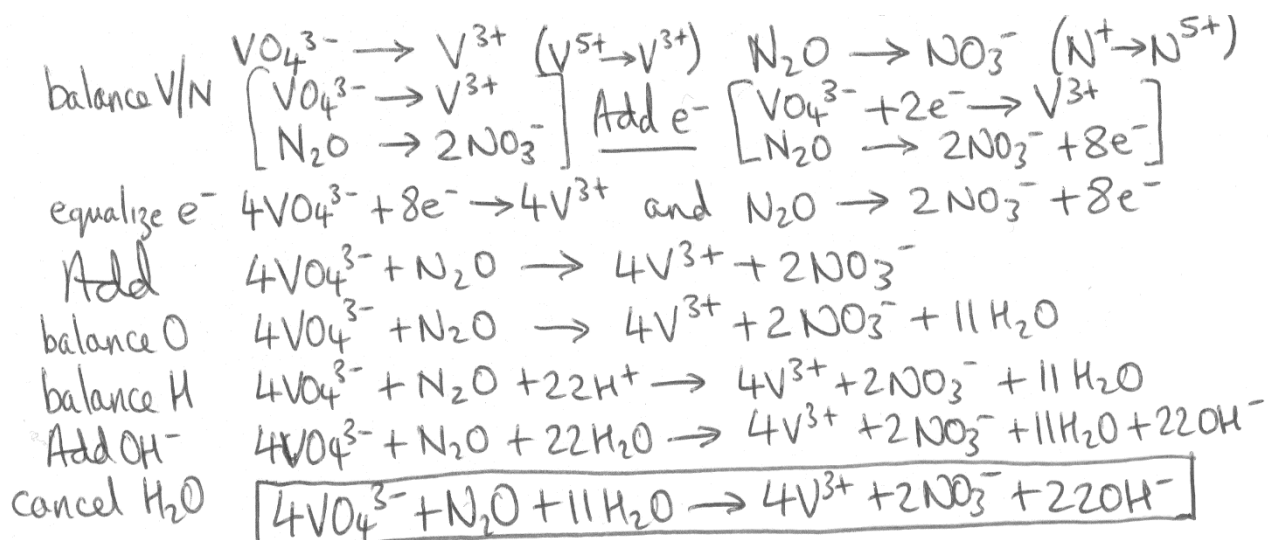
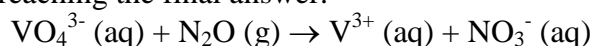
- a) (4 points) Calculate ΔG° for this reaction.

$$\begin{aligned}\Delta G^\circ &= 2(66.07 \text{ kJ}) - 2(86.60 \text{ kJ}) - 0.000 \text{ kJ} \\ &= -41.06 \text{ kJ}\end{aligned}$$

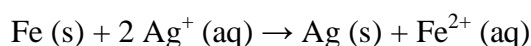
- b) (7 points) What is ΔG for this reaction at 298 K when the partial pressures (p) of each species are as follows: $p(\text{NO}) = 1.2$ atmospheres, $p(\text{Cl}_2) = 0.22$ atmospheres, $p(\text{NOCl}) = 2.1$ atmospheres?

$$\begin{aligned}\Delta G &= \Delta G^\circ + RT \ln Q \\ &= -41.06 \text{ kJ} + (8.314 \times 10^{-3} \text{ kJ/K})(298 \text{ K}) \ln \frac{(2.1)^2}{(1.2)^2(0.22)} \\ \therefore \Delta G &= -41.06 \text{ kJ} + 6.52 \text{ kJ} \\ \therefore \Delta G &= -34.54 \text{ kJ}\end{aligned}$$

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



5. For the redox reaction



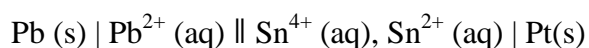
Complete the following statements (referring to the reaction from left-to-right).

- (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 (**Fe²⁺**).

6. Consider the voltaic cell whose notation is



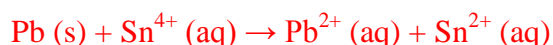
- (i) (3 points) The anode half-reaction is:



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



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



$$\begin{aligned} \Delta G &= \Delta H - T\Delta S \\ \text{pH} &= -\log[\text{H}^+] \\ \Delta S_{\text{rxn}} + \Delta S_{\text{surr}} &= \Delta S_{\text{univ}} \end{aligned}$$

$$\begin{aligned} \text{pK}_a &= -\log K_a \\ \Delta G &= \Delta G^\circ + RT \ln Q \\ \Delta G^\circ &= -RT \ln K \end{aligned}$$

$$\begin{aligned} \Delta G^\circ &= -nFE^\circ_{\text{cell}} \\ R &= 8.314 \text{ J/mol}\cdot\text{K} \\ \text{pH} + \text{pOH} &= 14.00 \end{aligned}$$