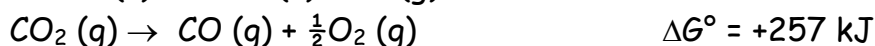
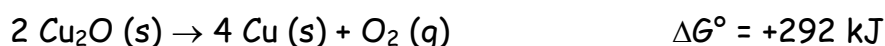
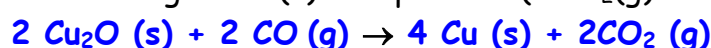


Answer all questions. Show your work. Look at the end for useful information.

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



- a) (4 pts) What is the overall reaction obtained from coupling together the above reactions to give Cu (s) as a product (no O₂(g) should appear)?



- b) (3 pts) What is ΔG° for the reaction in part a) at 25°C?

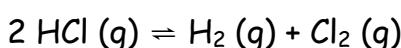
$$\Delta G^\circ = +292 \text{ kJ} - 514 \text{ kJ} = -222 \text{ kJ} \quad \therefore \Delta G^\circ = -222 \text{ kJ}$$

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

$$\Delta G^\circ = -RT \ln K \quad \therefore \ln K = -\Delta G^\circ / RT = -(-222 \times 10^3 \text{ J}) / (8.314 \text{ J/mol.K})(298 \text{ K})$$

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

2. (5 pts) The reaction below at 298 K has $\Delta G^\circ = 191 \text{ kJ}$.



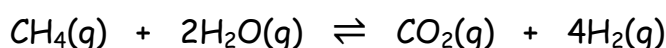
What is ΔG for this reaction at 298 K when the partial pressures (p) of each species are as follows: p(Cl₂) = 0.6 atm, p(H₂) = 4 atm, p(HCl) = 0.2 atm?

$$\Delta G = \Delta G^\circ + RT \ln Q \quad Q = \frac{p(\text{H}_2)p(\text{Cl}_2)}{p(\text{HCl})^2} = \frac{(4)(0.6)}{(0.2)^2} = 60$$

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

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

3. a) (5 pts) Calculate the ΔG° at 298K for the reaction:



Substance:	$\text{CH}_4(\text{g})$	$\text{H}_2\text{O}(\text{g})$	$\text{CO}_2(\text{g})$	$\text{H}_2(\text{g})$
ΔH_f° (kJ/mol):	-74.87	-241.8	-393.5	0
S° (J/K·mol):	186.1	188.8	213.7	130.7

$$\Delta H^\circ = 1\text{mol}(-393.5) + 0 - 1\text{mol}(-74.87) - 2\text{mol}(-241.8) = 164.97 \text{ kJ/mol}$$

$$\Delta S^\circ = 1\text{mol}(213.7) + 4\text{mol}(130.7) - 1\text{mol}(186.1) - 2\text{mol}(188.8) = 172.8 \text{ J/mol}$$

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

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

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

4. For the reaction $\text{Cr}^{2+}(\text{aq}) + 2 \text{H}^+(\text{aq}) + \text{SO}_3(\text{g}) \rightleftharpoons \text{SO}_2(\text{g}) + \text{Cr}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

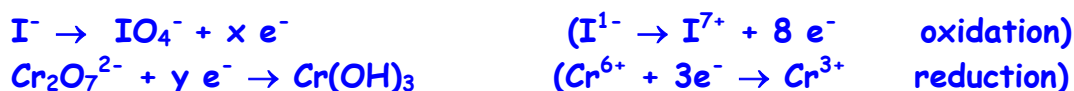
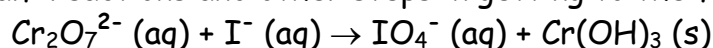
Complete the following statements (for the reaction from left-to-right).

- (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).

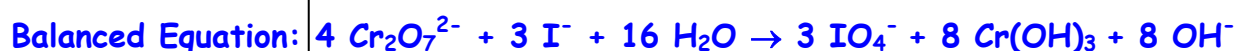
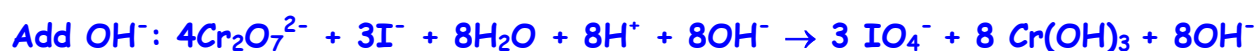
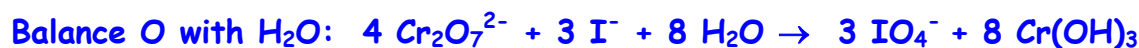
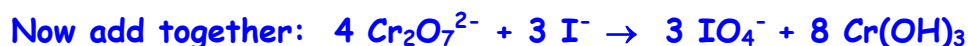
Complete the following statements (for the reaction from right-to-left).

- (iv) (1 pt) The substance oxidized is (SO_2).
 (v) (1 pt) The reducing agent is (SO_2).

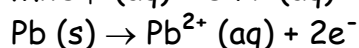
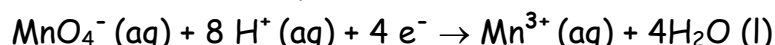
- (vi) (1 pt) The substance reduced is (Cr^{3+}).
5. (10 points) Balance the following redox reaction in basic solution (you must show the half-reactions and other steps in getting to the final answer.



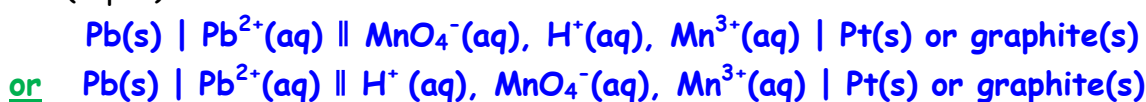
Add electrons, balance and make number of e^- the same:



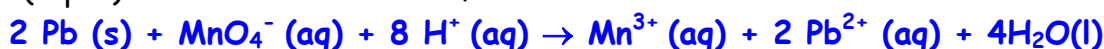
6. a) Consider a voltaic cell with half-reactions:



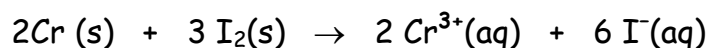
(5 pts) The cell notation for this cell is



(2 pts) The overall reaction of the cell is



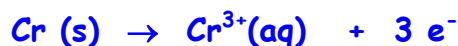
- b) Consider a voltaic cell whose overall reaction is



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



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



$$R = 8.314 \text{ J/mol.K}$$

$$\text{pK}_a = -\log K_a$$

$$F = 9.65 \times 10^4 \text{ J/V.mol } e^-$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\Delta S_{\text{rxn}} + \Delta S_{\text{surr}} = \Delta S_{\text{univ}}$$

$$\Delta G^\circ = -RT \ln K$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\Delta G = -nFE^\circ_{\text{cell}}$$

$$\text{pH} = -\log[\text{H}^+]$$