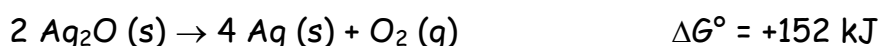
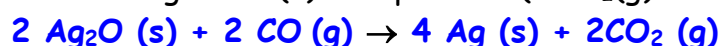


Answer all questions. Show your work. Look at the end for useful information.

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 O<sub>2</sub>(g) should appear)?



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

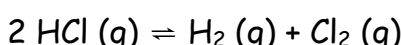
$$\Delta G^\circ = +152 \text{ kJ} - 514 \text{ kJ} = -222 \text{ kJ} \quad \therefore \Delta G^\circ = -362 \text{ kJ}$$

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

$$\Delta G^\circ = -RT \ln K \quad \therefore \ln K = -\Delta G^\circ / RT = -(-362 \times 10^3 \text{ J}) / (8.314 \text{ J/mol.K})(298 \text{ K})$$

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

2. (5 pts) The reaction below at 320 K has  $\Delta G^\circ = 191 \text{ kJ}$ .



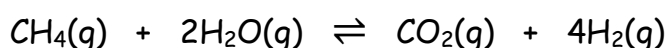
What is  $\Delta G$  for this reaction at 320 K when the partial pressures (p) of each species are as follows: p(Cl<sub>2</sub>) = 0.3 atm, p(H<sub>2</sub>) = 6 atm, p(HCl) = 0.3 atm?

$$\Delta G = \Delta G^\circ + RT \ln Q \quad Q = \frac{p(\text{H}_2)p(\text{Cl}_2)}{p(\text{HCl})^2} = \frac{[(6) \times (0.3)]}{(0.6)^2} = 5$$

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

$$\Delta G = 195 \text{ kJ for the reaction as written}$$

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



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

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

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

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

- 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  $\text{Cr}^{2+}(aq) + 2 \text{H}^+(aq) + \text{SO}_3(g) \rightleftharpoons \text{SO}_2(g) + \text{Cr}^{3+}(aq) + \text{H}_2\text{O}(l)$

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

(i) (1 pt) The reducing agent is (  $\text{Cr}^{2+}$  ).

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

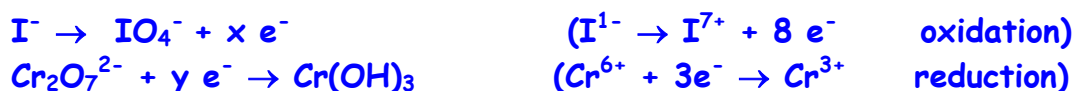
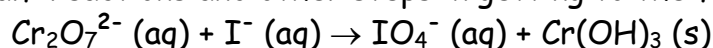
(iii) (1 pt) The substance reduced is (  $\text{SO}_3$  ).

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

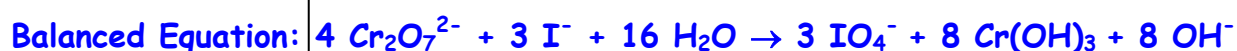
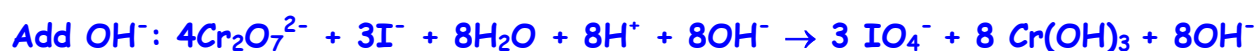
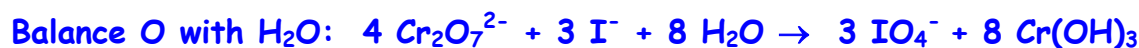
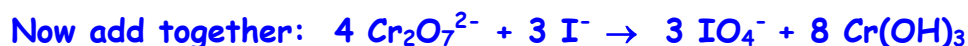
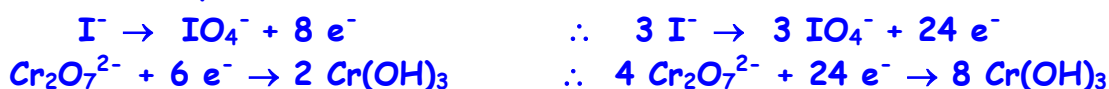
(iv) (1 pt) The substance oxidized is (  $\text{SO}_2$  ).

(v) (1 pt) The oxidizing agent is (  $\text{Cr}^{3+}$  ).

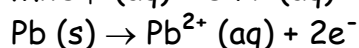
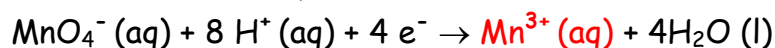
- (vi) (1 pt) The substance reduced is (  $\text{Cr}^{3+}$  ).
5. (10 points) Balance the following redox reaction in basic solution (you must show the half-reactions and other steps in getting to the final answer.



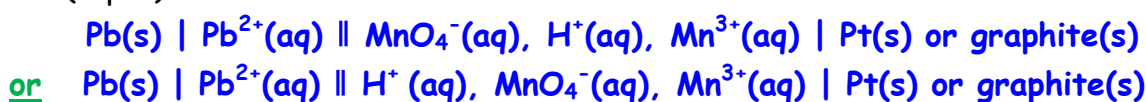
Add electrons, balance and make number of  $\text{e}^-$  the same:



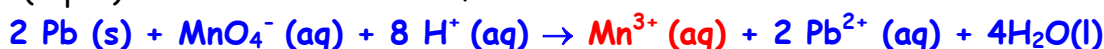
6. a) Consider a voltaic cell with half-reactions:



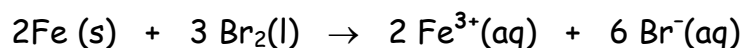
(5 pts) The cell notation for this cell is



(2 pts) The overall reaction of the cell is



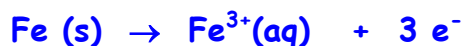
- b) Consider a voltaic cell whose overall reaction is



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



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



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

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

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

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

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

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

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

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

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