

Write very clearly and **show all of your work** for partial credit. A list of equations and constants as well as a periodic table are on the last two pages of your exam.

1. (20 points) Place a correct response in each blank.

(a.) What is  $K$  if  $\Delta G^\circ = 1.2 \times 10^{-2}$ ?

1.0

(b.) If  $\Delta H_{rxn}^\circ = -84.9 \text{ kJ}$ , what is  $\Delta S_{surr}^\circ$  in  $\frac{\text{J}}{\text{K}}$ ?

285  $\frac{\text{J}}{\text{K}}$

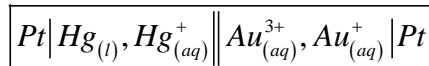
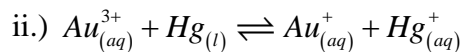
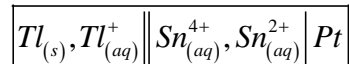
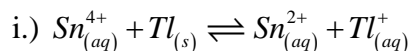
(c.) Which law is related to  $\Delta E$ ,  $q$ , and  $w$ ?

1<sup>st</sup>

(d.) Is the  $S_f^0$  for graphite zero (yes or no)?

no

(e.) Write the line/short notation for the redox reactions given below.



(f.) What conditions for Gibbs and cell potential will never give a spontaneous process?

$\Delta G > 0, E < 0$

(g.) What do we call the cell described in f.)?

electrolytic

(h.) What do we call the process in f.)?

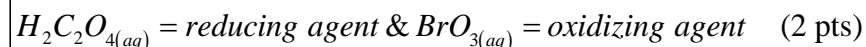
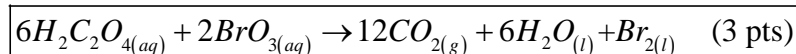
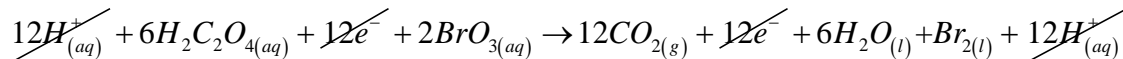
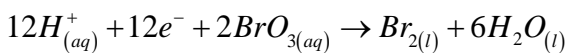
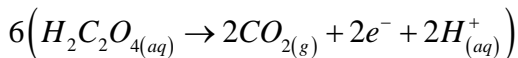
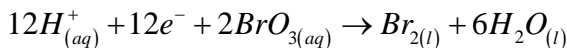
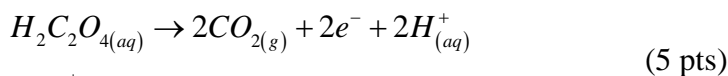
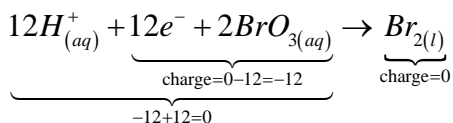
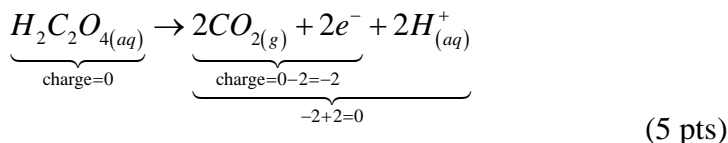
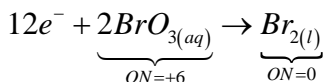
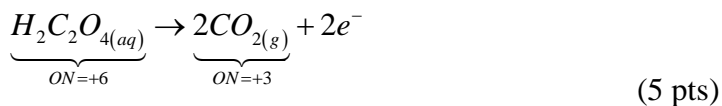
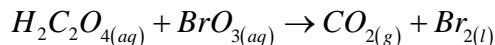
endergonic/nonspontaneous

(i.) What is the  $n$  for the cell given in (e.) ii.)?

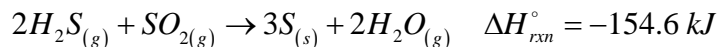
2

(j.) For electrons to flow nonspontaneously is the potential positive or negative?

2. (20 points) Balance the following redox reaction under acidic conditions. Identify which reagent is the oxidizing agent and which is the reducing agent.



3. (25 points) Consider the following reaction and corresponding thermodynamic data



$S^\circ \left( \frac{J}{mol \cdot K} \right)$	20.5	264.1	64.8	89.9
$\Delta_f H^\circ \left( \frac{kJ}{mol} \right)$		-256.8	0	-214.8

(a) Determine the heat of formation for  $H_2S_{(g)}$ . (5pts)

$$\Delta H_{rxn}^\circ = 3\Delta H_f^\circ(S) + 2\Delta H_f^\circ(H_2O) - [2\Delta H_f^\circ(H_2S) + \Delta H_f^\circ(SO_2)] \quad (2 \text{ pts})$$

$$-154.6 \text{ kJ} = 3\Delta H_f^\circ(0) + 2\Delta H_f^\circ(-214.8) - [2\Delta H_f^\circ(H_2S) + \Delta H_f^\circ(-256.8)]$$

$$\Delta H_f^\circ(H_2S) = \boxed{-9.1 \text{ kJ}} \quad (3 \text{ pts})$$

(b) Determine the entropy change for the above reaction. (5pts)

$$\Delta S_{rxn}^{\circ} = 3S_f^{\circ}(S) + 2S_f^{\circ}(H_2O) - [2S_f^{\circ}(H_2S) + S_f^{\circ}(SO_2)] \quad (2pts)$$

$$\Delta S_{rxn}^{\circ} = 3 \cdot 64.8 \frac{J}{K} + 2 \cdot 89.9 \frac{J}{K} - [2 \cdot 20.5 \frac{J}{K} + 264.1 \frac{J}{K}] = \boxed{69.1 \frac{J}{K}} \quad (3pts)$$

(c) Determine  $\Delta G^{\circ}$  for the above reaction. (5pts)

$$\Delta G_{rxn}^{\circ} = \Delta H_{rxn}^{\circ} - T\Delta S_{rxn}^{\circ} \quad (2pts)$$

$$\Delta G_{rxn}^{\circ} = -154.6 kJ - 298.15K \cdot (69.1 \frac{J}{K}) \times \frac{1kJ}{1000J} = \boxed{-175 kJ} \quad (3pts)$$

(d) Determine the equilibrium constant for the above reaction. (5pts)

$$\{R = 8.314 J/K \cdot mol \text{ and } F = 96485 C/mol\}$$

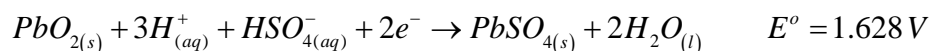
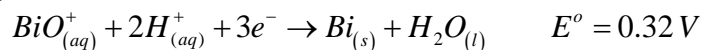
$$K = e^{-\Delta G_{rxn}^{\circ}/RT} \quad (2pts)$$

$$K = e^{-(-175 kJ)/8.3145 \frac{J}{K} \times \frac{1kJ}{1000J} \times 298.15K} = \boxed{4.94 \times 10^{30}} \quad (3pts)$$

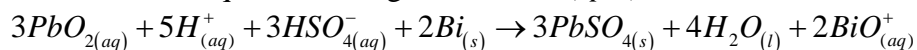
(e) Is the reaction spontaneous? Why or why not? (5pts)

Yes, Gibbs is less than zero.

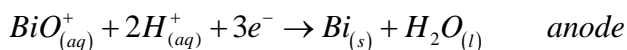
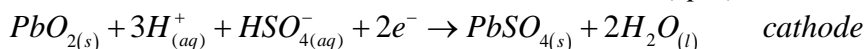
4. (35 points) Given the data below:



a.) Write the balanced equation for a galvanic cell. (5pts)



b.) State which half reaction is the cathode and the anode. (5pts)



c.) State the species which is oxidized and reduced. (5pts)

$Bi_{(s)}$  is oxidized &  $PbO_{2(s)}$  is reduced

d.) State the species which is the oxidizing and reducing agent. (5pts)

$PbO_{2(s)}$  is the oxidizing &  $Bi_{(s)}$  is reducing agent

e.) Determine the standard emf,  $E_{cell}^{\circ}$ . (5pts)

$$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} \quad (2pts)$$

$$E_{cell}^{\circ} = 1.628 - 0.32 = \boxed{1.308V} \quad (3pts)$$

f.) Is the reaction spontaneous? Why? (5pts)

yes,  $E_{cell}^{\circ} > 0$

g.) Determine the emf,  $E_{cell}$ , given:  $[HSO_4^{-}] = 0.500M$ ,  $[BiO^{+}] = 0.250M$ ,  $pH = 5$  (5pts)

$$E_{cell} = E_{cell}^{\circ} - \frac{0.0592}{n} \log \frac{[Bi^{3+}]^2}{[H^{+}]^5 [HSO_4^{-}]^3} \quad (2pts)$$

$$E_{cell} = 1.308V - \frac{0.0592}{6} \log \frac{[0.500]^2}{[10^{-5.00}]^5 [0.250]^3} = \boxed{1.085V} \quad (3pts)$$