

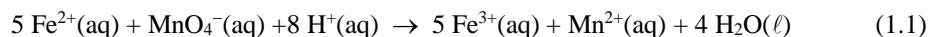
\* While I prefer you turn in a hard copy of the worksheet, I will accept scanned copies sent to my email address, joensj@fiu.edu

Section: (circle one)      M,W,F                      Tu,Tr

Exam 3 is Friday, November 18 (for the M,W,F class) and Thursday, November 17 (for the T,R class) in class. It will cover material from Chapters 17 and 18 of Burdge.

For problems involving calculations you must show your work for credit. Unless otherwise stated, you may assume  $T = 25.0\text{ }^{\circ}\text{C}$ .

1) The concentration of a stock solution of  $\text{Fe}^{2+}$  ion can be found by titrating a sample of the solution with a permanganate ion ( $\text{MnO}_4^-$ ) solution of known concentration. The reaction that occurs is



A 25.00 mL sample of a stock solution of  $\text{Fe}^{2+}$  ion is titrated with a 0.0317 M solution of permanganate ion. After the addition of 11.86 mL of the permanganate ion solution the equivalence point for the titration is reached. Based on this information find  $[\text{Fe}^{2+}]$ , the concentration of  $\text{Fe}^{2+}$  ion in the stock solution.

$$\text{moles MnO}_4^- = 0.01186 \text{ L soln} \frac{0.0317 \text{ mol}}{\text{L soln}} = 3.760 \times 10^{-4} \text{ mol MnO}_4^-$$

$$\text{mol Fe}^{2+} = 3.760 \times 10^{-4} \text{ mol MnO}_4^- \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 1.880 \times 10^{-3} \text{ mol Fe}^{2+}$$

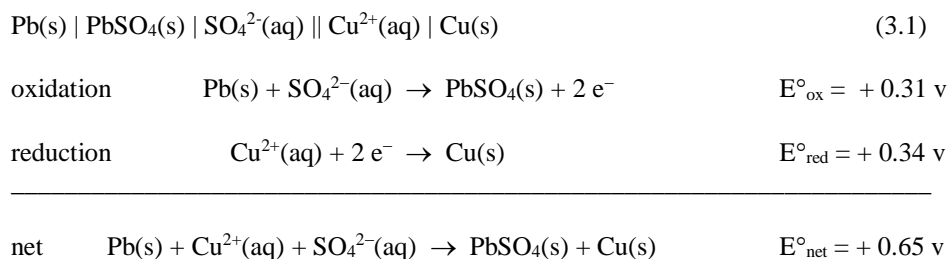
$$[\text{Fe}^{2+}] = \frac{1.880 \times 10^{-3} \text{ mol Fe}^{2+}}{0.02500 \text{ L soln}} = 0.0752 \text{ M Fe}^{2+}$$

2) The half-cell potential for the process  $\text{Sn}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Sn}(\text{s})$  is  $E^\circ = -0.14 \text{ v}$ . The half-cell potential for the process  $3 \text{Sn}(\text{s}) \rightarrow 3 \text{Sn}^{2+}(\text{aq}) + 6 \text{e}^-$  is

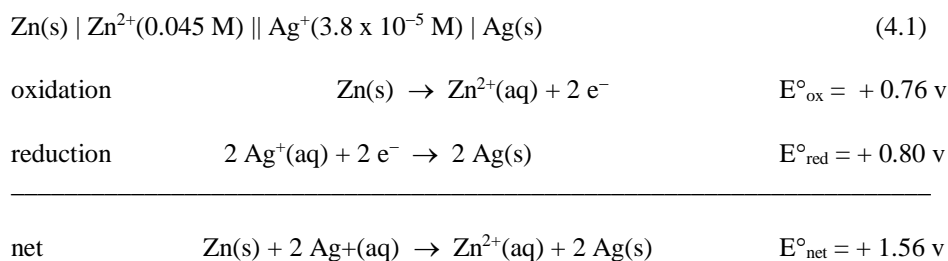
- a)  $E^\circ = -0.42 \text{ v}$   
 b)  $E^\circ = -0.14 \text{ v}$   
**C** c)  $E^\circ = +0.14 \text{ v}$   
 d)  $E^\circ = +0.42 \text{ v}$   
 e) Cannot tell from the information given

Changing the direction of the reaction changes the sign of the standard cell potential. Multiplying the reaction by a positive constant has no effect on the standard cell potential.

3) Find the half-cell oxidation reaction, the half-cell reduction reaction, the net cell reaction, and the cell potential for standard conditions for the galvanic cell given below. Values for half-cell reduction potentials for standard conditions are given on slide 36 of the Chapter 18 powerpoint.



4) Find the values for  $E^\circ_{\text{cell}}$ ,  $E_{\text{cell}}$ ,  $\Delta G^\circ_{\text{rxn}}$ , and  $\Delta G_{\text{rxn}}$  for the galvanic cell



From Nernst  $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$

$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]^2} = \frac{(0.045)}{(3.8 \times 10^{-5})^2} = 3.1 \times 10^7 \quad n = 2$$

$$E_{\text{cell}} = 1.56 \text{ v} - \frac{(8.314 \text{ J/mol}\cdot\text{K})(298. \text{ K})}{(2)(96485. \text{ C/mol})} \ln (3.1 \times 10^7)$$

$$= 1.56 \text{ v} - (0.22 \text{ v}) = 1.34 \text{ v}$$

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

$$\Delta G^\circ_{\text{rxn}} = - (2) (96485 \text{ C/mol}) (+1.56 \text{ v}) = - 3.01 \times 10^5 \text{ J/mol} = - 301. \text{ kJ/mol}$$

$$\Delta G_{\text{rxn}} = - (2) (96485 \text{ C/mol}) (+1.34 \text{ v}) = - 2.59 \times 10^5 \text{ J/mol} = - 259. \text{ kJ/mol}$$

5) Electrolysis is used to cover rings made of inexpensive metals with a coating of a more valuable metal. For example, gold can be deposited onto the outside of a ring using the reaction



In a particular system electrolysis is carried out using a current  $i = 0.46 \text{ amps} = 0.46 \text{ C/s}$ . For how many minutes must electrolysis be carried out to coat a ring with 25.0 mg of gold metal?

$$\text{MW}(\text{Au}) = 196.97 \text{ g/mol}$$

$$\text{moles Au} = 25.0 \times 10^{-3} \text{ g Au} \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} = 1.27 \times 10^{-4} \text{ mol Au}$$

$$\text{mol e}^{-} = 1.27 \times 10^{-4} \text{ mol Au} \frac{3 \text{ mol e}^{-}}{1 \text{ mol Au}} = 3.81 \times 10^{-4} \text{ e}^{-}$$

$$\text{Coulombs e}^{-} = 3.81 \times 10^{-4} \text{ mol e}^{-} \frac{96485 \text{ C}}{\text{mol e}^{-}} = 36.7 \text{ C e}^{-}$$

The current is  $i = 0.46 \text{ amp} = 0.46 \text{ C/s}$ , so the time it takes to deliver this number of Coulombs of electrons is

$$\text{time} = 36.7 \text{ C e}^{-} \frac{1 \text{ s}}{0.46 \text{ C}} = 79.9 \text{ s} \frac{1 \text{ min}}{60 \text{ s}} = 1.33 \text{ minutes}$$

I will not ask a question like this on exam 3, but I could ask such a question on the final exam.