Week 3 Worksheet

Chem 11300-2: Section 33

April 12, 2022

***Problem* 1:** Use the following information to answer the questions below:

$$2\,H_2O + 2\,e^- \rightleftharpoons H_2 + 2\,OH^- \qquad E^\circ_{red} = -0.83V$$
$$O_2 + 4\,H^+ + 4\,e^- \rightleftharpoons 2\,H_2O \qquad E^\circ_{red} = 1.229V$$

a) In the electrolysis of water, a disproportionation reaction occurs in which water is simultaneously oxidized and reduced into two new species. What are the respective oxidation states of hydrogen and oxygen in water before electrolysis? During electrolysis, what atom is likely reduced? What atom is likely oxidized?

b) Use the half-cell reactions listed above to write a net equation for the electrolysis reaction (*Hint: you will produce hydronium at one electrode and hydroxide at the other.*)

c) What is the overall potential (E°) for this reaction? What is the minimum potential needed for electrolysis to occur?

d) If you were to collect the gases produced during electrolysis, what is the ratio of volumes (O_2 vs H_2 gas) that you would expect to see?

***Problem* 2:** Faraday's law simply states: the amount of substance consumed or produced at one of the electrodes in an electrolytic cell is directly proportional to the amount of electricity that passes through the cell.

It is useful to be able to relate the amount of charge (measured in coulombs) that passes through the cell for a given time. 1 Coulomb (C) of charge is transferred when a 1 ampere (A) current is passed through the cell for 1 second:

$$A = \frac{C}{sec} \tag{1}$$

The amount of charge in C can be directly related to the number of electrons by using Faraday's constant

$$F = 96485\frac{C}{\text{mol e}^-} \tag{2}$$

a) A 9V battery in this setup generates about 140 mA current. If we assume that the current remained steady for the 10.0 minutes that we electrolyzed water, calculate how many moles of O_2 could be produced (*Hint: use results from question 1*).

b) What volume of hydrogen and oxygen gas would be collected when water is electrolyzed for 10.0 minutes with a 140 mA current? ($R = 0.08206$ atm/mol K $= 8.134$ J/mol K)

c) Do these volumes match your prediction from problem 1?

***Problem* 3:** Given the following half reactions, answer the following questions.

$$\begin{aligned} Zn^{2+} + 2\,e^- &\longrightarrow Zn \quad E^\circ_{red} = -0.76V \\ Cu^{2+} + 2\,e^- &\longrightarrow Cu \quad E^\circ_{red} = 0.34V \end{aligned}$$

a) You want to plate zinc metal onto a copper penny. What is the spontaneous reaction that would occur between Zn/Zn^{2+} and Cu/Cu^{2+}?

b) Electroplating allows us to force the non-spontaneous redox proces to occur by using a voltage source to run the redox reaction in the opposite direction. Draw a diagram showing what processes will occur at the anode and cathode if we force Zn (s) to plate onto a copper penny.

c) How much potential is required to make this reaction happen?

d) What terminal of your voltage source should be connected to the penny during electroplating?

***Problem* 4:** The following will be useful:

$$\begin{aligned} Ag^+\,(aq) + e^- &\rightleftharpoons Ag\,(s) \quad E^\circ_{Ag} = 0.7996V \\ Cu^{2+}\,(aq) + 2\,e^- &\rightleftharpoons Cu\,(s) \quad E^\circ_{Cu} = 0.3402V \end{aligned}$$

a) Consider a voltaic cell constructed from silver and copper. Calculate the cell potential at the moment this battery starts to run, assuming we are using 1M solutions for all salts.

b) When a cell is under standard conditions, what is true of E_{cell} and E°?

c) Suppose that we now leave this cell running for several hours and then come back and measure the cell potential. If we find that the cell potential is now +0.20 volts, what must be the concentration of silver and copper ions in each solution? Assume equal volumes of solution at the anode and cathode.

***Problem* 5:** You construct a voltaic cell that uses the 2 half reactions below:

$$\begin{aligned} Ni^{2+}\,(aq) + 2\,e^- &\rightleftharpoons Ni\,(s) \quad E^\circ = -0.23V \\ Mn^{2+}\,(aq) + 2\,e^- &\rightleftharpoons Mn\,(s) \quad E^\circ = -1.18V \end{aligned}$$

How long will this voltaic cell be able to supply a potential of at least 0.94 V if, at 298 K, you start the reaction using 1.5 M solution of Ni^{2+} and 0.50 M solution of Mn^{2+}? Assume the voltaic cell delivers a constant current of 150 mA and the volume of each cell is 1.00 L.

***Problem* 6:** Balance the following reactions in an **acidic** environment.

$$MnO_4{}^-\,(aq) + C_2O_4{}^{2-} \longrightarrow Mn^{2+}(aq) + CO_2\,(g)$$

***Problem* 7:** Balance the following reaction in both an acidic and basic environment:

$$NO_2{}^-\,(aq) + Al(s) \longrightarrow NH_3\,(g) + AlO_2{}^-\,(aq)$$

The following problems are written by Professor Mcleod or Head TA Miah Turke. They may mimic homework problems closely, but will be highly beneficial for the midterms and final.

***Problem* 8:** Do acidic or basic conditions favor the solubility of lithium nitrite ($LiNO_2$)? Please explain.

***Problem* 9:** The following balanced chemical equation occurs in a galvanic cell.

$$18\,Br^-\,(aq) + 16\,H_3O^+\,(aq) + 2\,MnO_4{}^-\,(aq) \longrightarrow 2\,Mn^{2+}\,(aq) + 5\,Br_2\,(l) + 24\,H_2O\,(l)$$

Calculate $[Mn^{2+}]$ given that $\Delta E^\circ_{cell} = 0.420V$, $\Delta E_{cell} = 0.412V$, pH$=$ 0, $[Br^-]=$ 0.8M, and $[MnO_4{}^-]=$ 0.3M.

***Problem* 10:**

a) What is the definition of an average reaction rate? An instantaneous reaction rate? For the following reaction, what are all definitions of both the average and the instantaneous reaction rate in terms of the concentrations of products and reactants?

$$2\,NO\,(g) + 2\,H_2\,(g) \longleftrightarrow N_2\,(g) + 2\,H_2O\,(g)$$

b) The rate of a reaction depends upon the concentrations of reactants and products. If the reaction in (a) is mth orger in NO (g) and nth order in H_2 (g) with a reaction rate constant given by k, what is the rate law? What is the overall order of the reaction?

c) Nitrogen oxide reacts with hydrogen at 1280°C. A set of 3 experiments-described below-were run in order to determine the rate constant of the reaction. From this information, determine the initial rate law, rate constant, and order of the reaction.

Run	$[NO]_0$	$[H_2]_0$	Initial Rate [M/min]
1	0.0100 M	0.0100 M	0.00600
2	0.0200 M	0.0300 M	0.14400
3	0.0100 M	0.0200 M	0.01200