Student directions for Online Electrochemistry Activity

Video's of lab procedures for Part I and II:.
- The video for Part I shows the set-up and observations for redox reactions between metals and aqueous metal nitrate solutions. These results are used to determine an activity series for these metals: https://drive.google.com/file/d/1lt1q_mzgNMy5L7KVs32WZfeppixt4fHk/view?ts=5e8b0179
- The second video shows the procedure for measuring the cell potential for the Zn-Cu Galvanic cell: https://drive.google.com/file/d/1PTbdDyr8WJh-QQT8ISpa0vPobi0UYBQX/view?ts=5e8b89e6

Part I: Electrochemical Activity from Chemical Reactions

Redox reactions involve the transfer of electrons from the substance oxidized to the substance reduced. An oxidation-reduction reaction occurs when a piece of aluminum metal is placed into a solution of zinc nitrate. The aluminum atoms lose three electrons and are oxidized to Al^{3+} ions, whereas the Zn^{2+} ions gain two electrons and are reduced to zinc atoms. Hence, a coating of zinc metal begins to form at the surface of the aluminum metal due to this electron transfer process.

In Part I, you will look for similar evidence of reaction after drops of metal nitrate solutions are placed on top of different metals. These results will allow you to arrange five metals in order of decreasing activity.

Part II. Electrochemical activity from cell potentials

The redox reaction that occurs between aluminum metal and zinc nitrate is a single replacement reaction. The balanced molecular equation for this reaction can be written as follows:

$$2\ Al(s)\ +\ 3\ Zn(NO_3)_2(aq)\ \rightarrow 2\ Al(NO_3)_3(aq)\ +\ 3\ Zn(s)$$
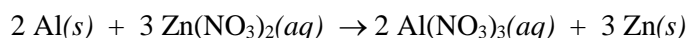
The net ionic equation for this reaction is: $2\ Al(s)\ +\ 3\ Zn^{2+}(aq)\ \rightarrow\ 2\ Al^{3+}(aq)\ +3\ Zn(s)$

This reaction can also be broken down into an oxidation half-reaction and a reduction half-reaction as follows:

$$Al(s) \rightarrow Al^{3+}(aq)\ +\ 3\ e^- \qquad \text{oxidation half-reaction}$$

$$Zn^{2+}(aq)\ +\ 2\ e^- \rightarrow Zn(s) \qquad \text{reduction half-reaction}$$
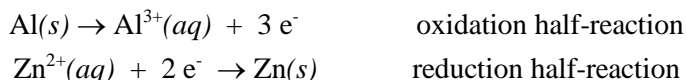
A voltaic cell consisting of two half-cells separates these two half reactions as shown below.

Anode: $Al(s)\ \rightarrow\ Al^{3+}\ +\ 3e\text{-}$ Cathode: $Zn^{2+}\ +\ 2e\text{-} \rightarrow Zn(s)$
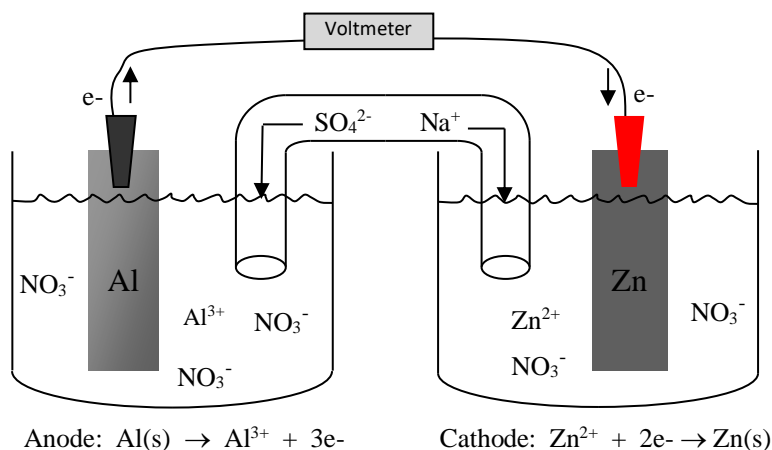
Each half-cell consists of a metal electrode immersed in a solution containing its corresponding cation. The electrode at which oxidation occurs is called the anode. For a voltaic cell, the more active metal will be the anode because it has a greater tendency to be oxidized and lose electrons. Reduction occurs at the electrode called the cathode. The electrodes are connected by a wire through which electrons will flow from the anode to cathode providing an electrical current. The half-cells are also connected by a salt bridge that permits the limited movement of ions into the half-cells to complete the circuit and to maintain charge neutrality in each half-cell. For the voltaic cell shown above, negative ions from the salt bridge move into the anode half-cell as Al (s) atoms are oxidized to form Al^{3+} ions. Positive ions from the salt bridge move into the cathode compartment as Zn^{2+} ions are reduced to Zn(s).

The transfer of electrons from the anode to cathode takes place through the external wire because there is difference in electrical potential between the two electrodes that can be measured in volts. This difference in electrical potential

is referred to as the cell potential. The standard cell potential, $E°_{cell}$, corresponds to 1 M solutions, gases at 1 atm pressure, and a temperature of 298 K.

A standard cell potential can be considered to be the sum of the standard reduction potential for the cathode half reaction and the standard oxidation potential for the anode half reaction:

$$\mathbf{E°_{cell} \ = \ E°_{red} \ + \ E°_{ox}}$$

In this equation, the oxidation potential for the anode reaction, $E°_{ox}$, is the negative of the reduction potential for the reverse reaction.

For example, when Zn(s) is the cathode, the reduction reaction and standard reduction potential would be:
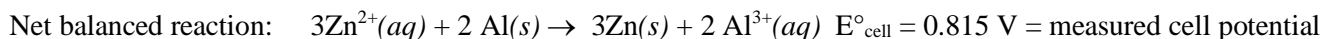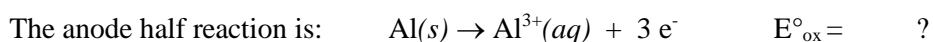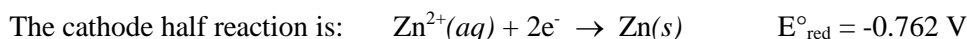
$$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \quad E°_{red} = \text{-0.762 V}$$
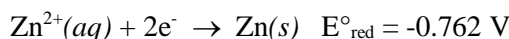
However, if Zn(s) is the anode, we must reverse the reaction and change the sign for the oxidation potential:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \quad E°_{ox} = \text{+0.762 V}$$
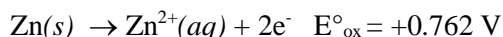
Thus, depending on whether Zn is the cathode or anode, we will use its respective reduction or oxidation potential and the measured cell potential, $E°_{cell}$, to find the half-cell potential for the other metal.

Example for writing the half reactions, balanced net reaction and determining cell potentials:

If the measured cell potential is 0.815 V for the Al-Zn cell described above, the half reactions and net balanced reaction would be:

The cathode half reaction is: $\quad Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \qquad E°_{red} = \text{-0.762 V}$

The anode half reaction is: $\quad \underline{Al(s) \rightarrow Al^{3+}(aq) \ + \ 3 e^- \qquad E°_{ox} = \qquad ?}$

Net balanced reaction: $\quad 3Zn^{2+}(aq) + 2 \ Al(s) \rightarrow \ 3Zn(s) + 2 \ Al^{3+}(aq) \ \ E°_{cell} = 0.815 \ V = $ measured cell potential

Note: We had to multiply the Zn half reaction by 3 and the Al half reaction by 2 to make the number of electrons equal for our two half-reactions. (The 6 electrons transferred cancel out, so they are not shown in the net reaction.)

To find the oxidation potential for Al(s), we use $E°_{cell} \ = \ E°_{red} \ + \ E°_{ox}$

$$0.815 \ V = \text{-0.762 V} + \underline{E°_{ox}}$$

Thus, for Al(s), $\quad \underline{\mathbf{E°_{ox} \ = 1.577 \ V \ \text{(3 decimal places, addition)}}}$

Example calculation for ΔG°$_{rxn,}$

The equation relating the standard free energy change to the standard cell potential is

$$\mathbf{\Delta G°_{rxn} = - nFE°} \qquad \text{where n = moles e}^- \text{ transferred and F = 96,500 } \frac{J}{V \cdot \text{mol } e^-}$$

For the Al-Zn cell described above, n = 6 moles of e⁻ transferred and $E°_{cell} = 0.815$ V

$$\Delta G°_{rxn} = - (6 \text{ mol e-})(96,500 \ \tfrac{J}{V \cdot \text{mol } e^-})(0.815 \text{ V}) = \text{-}4.71885 \times 10^5 \text{ J} \left(\tfrac{1 \ kJ}{1000 J}\right) = \mathbf{\text{-472 kJ}} \text{ (3 sig figs due to Faraday's}$$
$$\text{constant and } E° \text{ having 3 s.f., 6 moles is an exact number so it doesn't count)}$$

Example calculation for K

The standard cell potential is also related to the equilibrium constant by this equation:

$$\mathbf{E°_{cell} = \frac{RT}{nF} ln \ K} \quad \text{where R = 8.314 } \frac{J}{mol \cdot K} \text{ and T } = 298 \text{ K under standard conditions}$$

For the the Al-Zn cell, n = 6 mole e⁻, $E°_{cell} = 0.815$ V, T = 298 K

$$0.815 \ V = \left(\frac{\left(8.314 \ \frac{J}{mol \cdot K}\right)(298 \ K)}{(6 \ mol \ e^-)\left(96500 \ \frac{J}{V \cdot mol}\right)}\right) ln \ K$$

$$0.815 \ V = 0.00427905 \ V \times ln \ K$$

$$ln \ K = \ \frac{0.815 \ V}{0.00427905 \ V} = 190.46$$

$$K = \ e^{(190.46)} = \mathbf{5.21 \times 10^{82}} \quad \text{(3 sig figs, multiplication- least sig figs)}$$

Electrochemistry Lab Report

Part I Data:

Watch the video for part I that shows the reactions of solid metals with metal nitrate solutions.

https://drive.google.com/file/d/1lt1q_mzgNMy5L7KVs32WZfeppixt4fHk/view?ts=5e8b0179

An image from the video is linked below so you can see the appearance of the metals several minutes after drops of the metal nitrate solutions were placed on them. The image was captured near the end of the video (at time 3:33).

Image metals and metal nitrate solutions

Compare the reaction for Mg(s) + AgNO₃(aq) to Ag(s) + Mg(NO₃)₂(aq).

- Hopefully you noticed that only one of these 2 reactions occurs. (What type of reaction is this?)
- Thus, half of the combinations in Table 1 have positive reactions and the other half do not react.
- Additonally, each metal should have a different number of positive reactions.

1) **Look at the image and information above to complete Table 1. Write R if a reaction occurs or NR if no reaction takes place.**

Note: Some of the results have been provided because they may be difficult to interpret from the video image.

The columns in Table 1 represent the metal ions present in the aqueous metal nitrate solutions.

Data Table 1. Reactions of Metals with Metal Nitrate solutions

	Zn^{2+}(aq)	Cu^{2+}(aq)	Pb^{2+}(aq)	Mg^{2+}(aq)	Ag^{+}(aq)	# of reactions
Zn(s)	/////		R	NR		
Cu(s)		/////	NR			
Pb(s)	NR		/////	NR	R	
Mg(s)	R			/////		
Ag(s)		NR			/////	/////

2. Place the **metals** Zn(s), Cu(s), Pb(s), Mg(s), and Ag(s) in order of **decreasing** electrochemical activity based on results for Data Table 1.

_____ > _____ > _____ > _____ > _____

most active   least active

Part II Data. Measured Cell Potentials

Watch the video to observe the procedure for setting up the Zn- Cu cell:

https://drive.google.com/file/d/1PTbdDyr8WJh-QQT8ISpa0vPobi0UYBQX/view?ts=5e8b89e6

1) **Note the colors of the leads attached to copper and zinc at the end of the video. Record this in Table II.**

The color of the electrode leads attached to each metal have been provided for the other 3 voltaic cells. Measured cell potentials for all 4 cells are also provided in Table 2.

2) **Use the electrode lead color and the activity series from Part I to help you identify the anode and cathode metals in Table 2.**

Data Table 2. Voltaic Cell data for Part II.

		Black electrode	Red electrode	Anode Metal	Cathode Metal	Measured Cell Potential, $E°_{cell}$ (V)
#1	Zn-Ag	Zn	Ag			1.410
#2	Zn-Mg	Mg	Zn			0.664
#3	Zn-Pb	Zn	Pb			0.525
#4	Zn-Cu	?	?			0.989

Reactions, Calculations:

Note: you should write the reactions and calculations on scratch paper if you don't print this out BEFORE doing the online canvas assignment.

For each voltaic cell:

- Write each half reaction in the appropriate direction.

- Write the balanced net reaction for the cell.

- Calculate the half reaction potential of the other metal in the cell using the standard reduction potential (E°_{red} = -0.762 V) for $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ and the measured cell potential (E°_{cell}). Use 3 d.p. for E° values.

Voltaic Cell #1: Zn and Ag

A. half reaction in the anode: _____ E°_{ox} = _____

B. half reaction in the cathode: _____ E°_{red} = _____

C. Balanced cell reaction: _____ E°_{cell} = _____

Voltaic Cell #2: Zn and Mg

A. half reaction in the anode: _____ E°_{ox} = _____

B. half reaction in the cathode: _____ E°_{red} = _____

C. Balanced cell reaction: _____ E°_{cell} = _____

Voltaic Cell #3: Zn and Pb

A. half reaction in the anode: _____ E°_{ox} = _____

B. half reaction in the cathode: _____ E°_{red} = _____

C. Balanced cell reaction: _____ E°_{cell} = _____

Voltaic Cell #4: Zn and Cu

A. half reaction in the anode: _____ E°_{ox} = _____

B. half reaction in the cathode: _____ E°_{red} = _____

C. Balanced cell reaction: _____ E°_{cell} = _____

Calculate ΔG^{o} **in kJ** for all 4 cells and write down the calculated ΔG^{o} values (use 3 s.f.).
 (F = 96,500 J/V·mole e-)

Calculate the equilibrium constant, K, for all 4 cells (use 3 s.f.) and write down the calculated K values (3 sig figs). Assume T = 298 K and R = 8.314 J/mol·K