

Electrochemistry Practice

Name hey

Block _____

1. Using the half-reactions given, answer the questions below:

<u>Half - Reaction</u>	<u>E° (V)</u>
1. $Ce^{4+}_{(aq)} + e^- \rightarrow Ce^{3+}_{(aq)}$	+1.61
2. $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$	+ .80
3. $Hg^{2+}_{(aq)} + 2e^- \rightarrow 2Hg_{(l)}$	+ .79
4. $Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}$	- .14
5. $Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$	- .25
6. $Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$	- 1.66

- a. Which metal in the table is most likely to be oxidized in a galvanic cell? Al
- b. Which metal in the table is most likely to be reduced in a galvanic cell? Ce⁴⁺
- c. Will Sn spontaneously reduce Ag⁺ to Ag? Yes
- d. Will Hg spontaneously reduce Sn²⁺ to Sn? No

2. Using the reduction table in #1, answer the following:

a. If reactions (1) and (2) are added together as a redox reaction, which do you think will proceed as a reduction and which as an oxidation? Support with a standard potential calculation.

Ce⁴⁺ would proceed as a reduction and Ag would be oxidized.

$$+1.61 - .80 = .81$$

b. If reactions (4) and (5) are added together as a redox reaction, which do you think will proceed as a reduction and which as an oxidation? Support with a standard potential calculation.

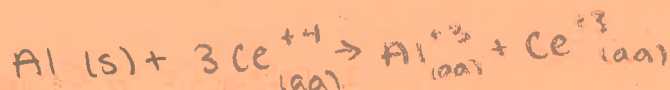
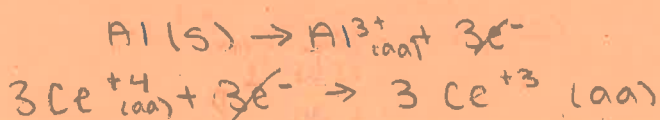
Reduction - Sn²⁺ Oxidation - Ni

$$-.14 - (-.25) = .11$$

c. Explain your reasoning behind your answers to a & b.

You want your E° to be positive, as it is the spontaneous. So the higher E° value will be reduction and the lower E° value will be oxidation.

d. Write a balanced equation for the reaction leading to the maximum positive potential.



3. Assume that you assemble an electrochemical cell based on the half reactions:



Answer the following questions:

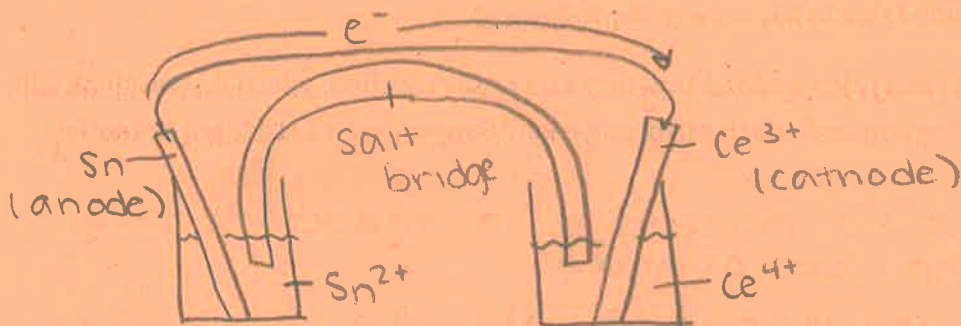
a. Write the balanced equation for the reaction that occurs spontaneously.



b. Which species was oxidized? Sn

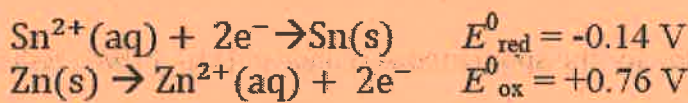
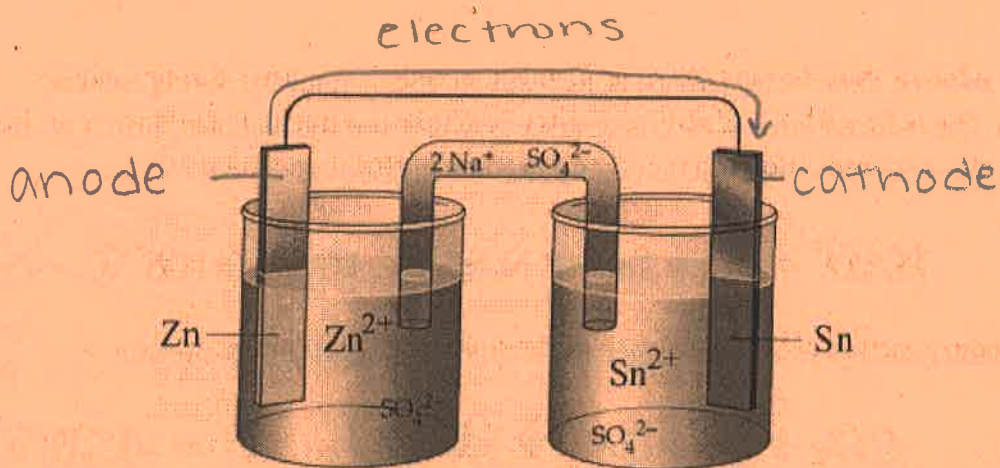
c. Which species was reduced? Ce⁺⁴

d. Diagram the cell. Label all parts.



e. You want to plate tin onto the Sn electrode; explain how you could allow for this to happen using the cell above.

In order to plate Sn on the Sn electrode, Sn must be reduced. In a spontaneous reaction Sn will be oxidized, so you must add energy to do the reverse reaction.



4. Which electrode will lose mass and which one will gain mass. Explain.

- Since Zn is being oxidized / losing e^{-} , it will lose mass
- Since Sn^{2+} is being reduced / gaining e^{-} , it will gain mass

5. Does oxidation or reduction occur at the cathode?

Reduction

6. Which of the following statements is/are correct.

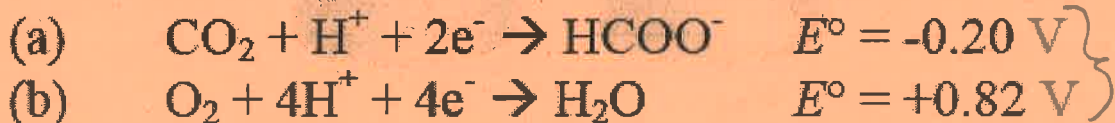
- a) Electrons flow through the wire towards the zinc electrode
* will flow towards Sn
- b) Electrons flow through the wire towards the tin electrode
- c) Na^{+} ions flow through the salt bridge towards the zinc half-cell
- d) Na^{+} ions flow through the salt bridge towards the tin half-cell
- Balances negative charge
- e) The standard potential of the cell is -0.90V
- f) The standard potential of the cell is $+0.62\text{V}$
* Add the 2 E° values together, since one is already labeled reduction and one is oxidation.

7. Nicotine adenine dinucleotide (NAD) is involved in redox chemistry during cellular respiration. The reduced form of NAD is written as NADH and the oxidized form is written as NAD⁺. The standard reduction reaction and potential of NAD is given by:



*Nonspontaneous
but if coupled/
paired with*

The NAD redox reaction occurs with each of the following reactions, as shown:



overall potential is positive & spont.

Write the overall reaction for the spontaneous change of NAD. Is NAD reduced or oxidized in this reaction?

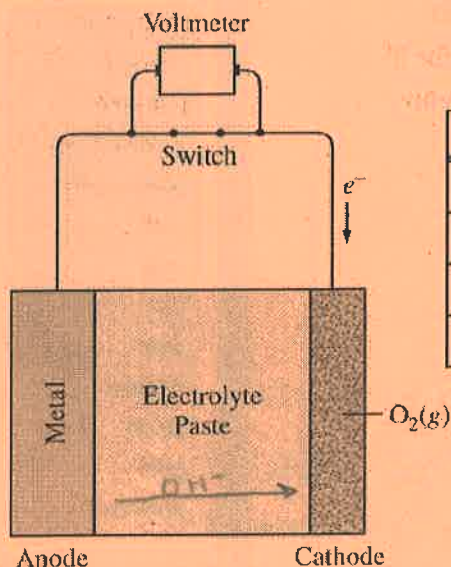
since (a) & (b) combined potential is a +0.62V, if they happen with the NAD reaction, the NAD will oxidize since it has a negative reduction potential (flipped)



** NADH is being oxidized, as it is giving off electrons.*

** the reaction becomes spont.*

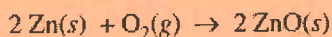
2015 Exam Question



Half Reaction	E at pH 11 and 298 K (V)
$O_2(g) + 2 H_2O(l) + 4 e^- \rightarrow 4 OH^-(aq)$	+0.34
$ZnO(s) + H_2O(l) + 2 e^- \rightarrow Zn(s) + 2 OH^-(aq)$	-1.31
$Na_2O(s) + H_2O(l) + 2 e^- \rightarrow 2 Na(s) + 2 OH^-(aq)$	-1.60
$CaO(s) + H_2O(l) + 2 e^- \rightarrow Ca(s) + 2 OH^-(aq)$	-2.78

Metal-air cells are a relatively new type of portable energy source consisting of a metal anode, an alkaline electrolyte paste that contains water, and a porous cathode membrane that lets in oxygen from the air. A schematic of the cell is shown above. Reduction potentials for the cathode and three possible metal anodes are given in the table.

- (a) Early forms of metal-air cells used zinc as the anode. Zinc oxide is produced as the cell operates according to the overall equation below.



- (i) Using the data in the table above, calculate the cell potential for the zinc-air cell. $+0.34 + 1.31 = 1.65$
- (ii) The electrolyte paste contains OH^- ions. On the diagram of the cell above, draw an arrow to indicate the direction of migration of OH^- ions through the electrolyte as the cell operates.
- (b) A fresh zinc-air cell is weighed on an analytical balance before being placed in a hearing aid for use.
- (i) As the cell operates, does the mass of the cell increase, decrease, or remain the same?
- (ii) Justify your answer to part (b)(i) in terms of the equation for the overall cell reaction.
You end with the same moles of solids you started with
- (c) The zinc-air cell is taken to the top of a mountain where the air pressure is lower.
- (i) Will the cell potential be higher, lower, or the same as the cell potential at the lower elevation?
- (ii) Justify your answer to part (c)(i) based on the equation for the overall cell reaction and the information above. *Less moles of O_2 to convert into OH^-*
- (d) Metal-air cells need to be lightweight for many applications. In order to transfer more electrons with a smaller mass, Na and Ca are investigated as potential anodes. A 1.0 g anode of which of these metals would transfer more electrons, assuming that the anode is totally consumed during the lifetime of a cell? Justify your answer with calculations. *Ca, it has a higher E° when oxidized than Na*
- (e) The only common oxide of zinc has the formula ZnO. Oxidized then Na
- (i) Write the electron configuration for a Zn atom in the ground state. $[Ar] 4s^2 3d^{10}$
- (ii) From which sublevel are electrons removed when a Zn atom in the ground state is oxidized?
4s

9. Electrolysis

- a. Use the half-reactions and the reactivity series provided to write a balanced equation for the reaction that would occur if aluminum is placed into a silver nitrate solution. Explain your reasoning.

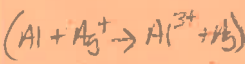
Most reactive

potassium
sodium
calcium
magnesium
aluminium
zinc
iron
tin
lead
copper
silver
gold
platinum

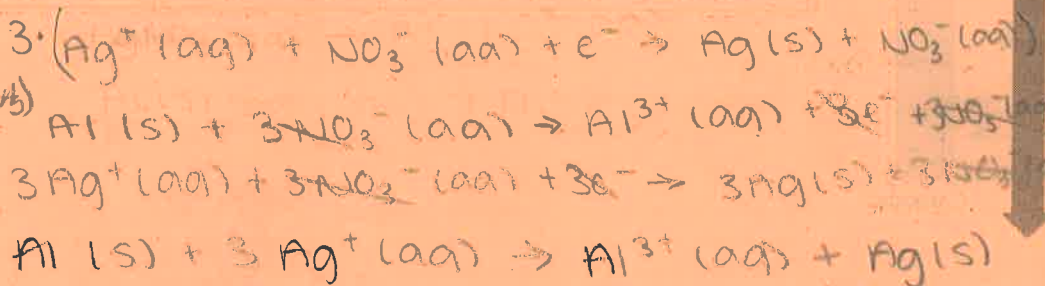
Least reactive

1) Al is more reactive, therefore will kick Ag out of preferred ion state

Half - Reaction	E° (V)
1. $Ag^+(aq) + e^- \rightarrow Ag(s)$	+ .80
2. $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66



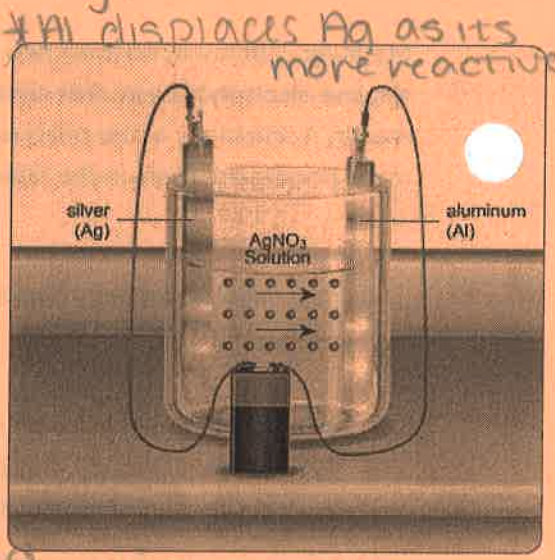
2) Al will oxidize due to lower reduction potential. Oxidation is becoming more positive, which is shown w/ $Al \rightarrow Al^{3+}$



- b. A cell is constructed as shown. If a piece of aluminum is placed into the silver nitrate solution, will it be plated by silver? If yes, explain why. If no, explain what must be done to cause plating to occur in the silver nitrate cell.

Therefore electrolytic cell, for something to plate where the reverse (non-spontaneous) reaction is forced. If silver plating is spontaneous (as seen in (a)) $Ag^+ \rightarrow Ag(s)$, the opposite will happen in this cell.

No, this cell is hooked up to (energy), as seen in the diagram, so in order for it requires energy.



This is due to Al's preferred state being $Al^{3+}(aq)$, not $Al(s)$, $Al(s)$ will plate, not $Ag(s)$.

- c. Calculate the following for the plating of 100.0 g of silver onto the aluminum electrode.

- 1) Coulombs of electricity necessary to deposit the silver from the $AgNO_3$.

$$100 \text{ g Ag} \left(\frac{1 \text{ mol Ag}}{107.86 \text{ g Ag}} \right) = 0.9271 \text{ moles Ag} \left(\frac{1 \text{ mol Ag}^+}{1 \text{ mol Ag}} \right) \left(\frac{1 \text{ mol } e^-}{1 \text{ mol Ag}^+} \right)$$

$$0.9271 \text{ moles } e^- \left(\frac{96,500 \text{ C}}{1 \text{ mol } e^-} \right) = 89,470 \text{ C}$$

- 2) How many minutes will take to plate out the silver using a current of 3.450 amp.

$$I = \frac{q}{t} \quad q = It$$

$$t = 25,930 \text{ s} \left(\frac{1 \text{ min}}{60 \text{ sec}} \right)$$

$$(3.450 \text{ amp}) t = 89,470 \text{ C}$$

$$= 432.1 \text{ min}$$