AP CHEMISTRY

TOPIC 11: ELECTROCHEMISTRY, EXAMPLE AP QUESTION

Day 127:

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS

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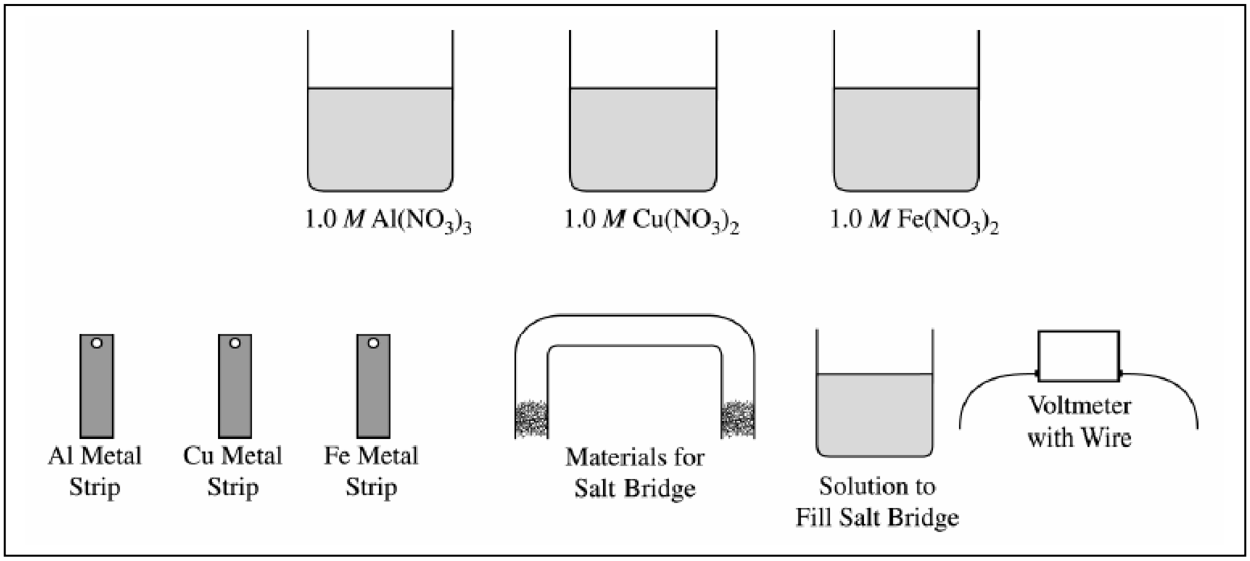
1. a) Several different electrochemical cells can be constructed using the materials shown below. Write the balanced net-ionic equation for the reaction that occurs in the cell that would have the greatest positive value of E_{cell}.



$Al^{+3} + 3\ e^{-1} \rightarrow Al$ $E^0 = -1.66$ V (reverse reaction to oxidize)

$Cu^{+2} + 2\ e^{-1} \rightarrow Cu$ $E^0 = +0.34$ V (best reduction reaction)

~~$Fe^{+2} + 2\ e^{-1} \rightarrow Fe$ $E^0 = -0.44$ V~~ (can only have one reduction)

$2\ (Al \rightarrow Al^{+3} + 3\ e^{-1})$

$3\ (Cu^{+2} + 2\ e^{-1} \rightarrow Cu)$

$$2\ Al + 3\ Cu^{+2} \rightarrow 2\ Al^{+3} + 3\ Cu$$

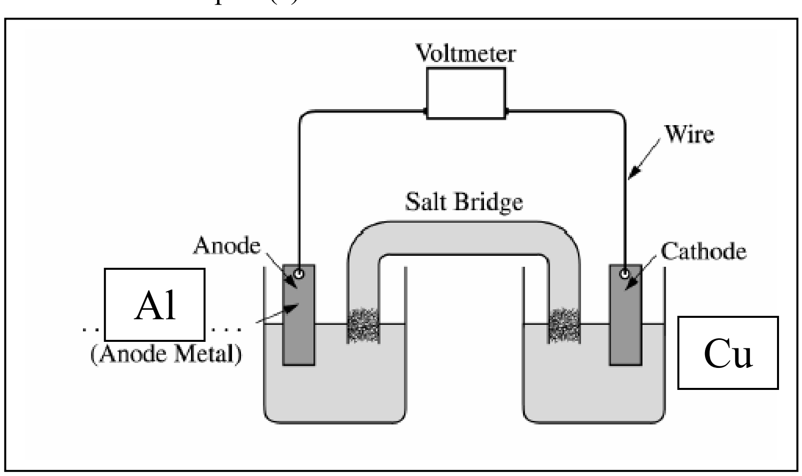
b) Calculate the standard cell potential, E_{cell}, for the reaction written in part (a).

$Al \rightarrow Al^{+3} + 3\ e^{-1}$ $E^0 = +1.66$ V

$Cu^{+2} + 2\ e^{-1} \rightarrow Cu$ $E^0 = +0.34$ V

$E^0 = +2.00$ V

c) A cell is constructed based on the reaction in part (a) above. Label the metal used for the anode on the cell shown in the figure below:



d) of the compounds, NaOH, CuS, and $NaNO_3$, which one is appropriate to use in a salt bridge? Briefly explain your answer, and for each of the other compounds, include a reason why it is not appropriate.

NaOH dissociates nicely in water. However, OH^{-1} ions would move toward the Anode (to balance the charges in solution), as the OH^{-1} meets the Al^{+3} ions (newly created from the oxidation of aluminum metal) a precipitate would form. Bad.

CuS does not dissociate in water at all. This is bad for the system because now we have NO migration of ions to either chambers and this STOPS the flow of negative charges and the flow of positive charges. Salt bridges need to be soluble and NOT FORM precipitates !!! Bad.

$NaNO_3$ is the winner! $NaNO_3$ dissociates in water nicely and does not form a precipitate with either metal. $NaNO_3$ allows for the flow of negative (and positive) charges. Good !!!

e) Another standard cell is based on the following reaction.

$$Zn + Pb^{+2} \rightarrow Zn^{+2} + Pb$$

If the concentration of Zn^{+2} is decreased from 1.0 *M* to 0.25 *M*, what effect does this have on the cell potential? Justify your answer.

Since Zn^{+2} is a product in the oxidation of Zn (solid) a slight INCREASE in the cell potential would occur. Because the decrease in the concentration would change the equilibrium quotient, Q, to a lower value since the oxidation reaction PRODUCES electrons. The concentration of Pb^{+2} stays at 1 M (standard cell).

$$E_{cell} = E^0 - \frac{0.0592}{n} \log\left(\frac{[prod.]}{[react.]}\right)$$

$$E_{cell} = 0.63\ V - \frac{0.0592}{2}\log\left(\frac{(0.25)^1}{(1)^1}\right) = 0.63\ V - (-0.01782\ V) = 0.6478\ V$$

$$E_{cell} = 0.65\ V$$

2. The reaction between silver ion and solid zinc is represented by the following equation.

$$2\ Ag^{+1}\,(aq) + Zn\,(s) \rightarrow Zn^{+2}\,(aq) + 2\ Ag\,(s)$$

a) A 1.50 g sample of Zn is combined with 250 mL of 0.110 *M* $AgNO_3$ at 25^0C.

(i) Identify the limiting reactant. Show calculations to support your answer.

$$\frac{250\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.110\ mol\ Ag^{+1}}{1\ L} \times \frac{1\ mol\ Zn^{+2}}{2\ mol\ Ag^{+1}} \times \frac{65.39\ g}{1\ mol\ Zn^{+2}} = 0.89911\ g\ Zn^{+2}$$

Ag^{+1} = L.R.

If one reacted ALL 250 mL of the 0.110 M $AgNO_3$ with 1.50 grams of Zn, ONLY 0.89911 grams of Zn metal would be consumed in the reaction. We started with 1.50 grams of Zinc – there is EXTRA zinc remaining after the reaction. Therefore, $AgNO_3$ or Ag^{+1} are the Limiting Reactant.

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of $[Zn^{+2}]$ after the reaction is complete. Assume the volume change is negligible

From part 2 a (i), we calculated that 0.89911 grams of Zinc would be consumed in the reaction. The question is asking about the concentration AFTER the reaction is complete. We started with 1.50 grams of Zinc and USED 0.89911 grams, therefore, we will have:

1.50 grams Zn – 0.89911 grams Zn = 0.60089 grams Zn REMAINING

$$\frac{0.60089\ g\ Zn}{} \times \frac{1\ mole\ Zn}{65.39\ g} = 0.0091893\ mol\ Zn$$

$$\left[Zn^{+2}\right] = \frac{0.0091893\ mol\ Zn}{250\ mL} \times \frac{1000\ mL}{1\ L} = 0.0367572\ M$$

$$[\ Zn^{+2}\ ] = 3.68 \times 10^{-2}\ M$$

b) Determine the value of the **standard potential**, E^0, for a galvanic cell based on the reaction between $AgNO_3$ (*aq*) and solid Zn **at 25^{0}C**.

$$2\ (\ Ag^{+1} + e^{-1} \rightarrow Ag\ ) \qquad E^0 = +0.80\ V$$

$$Zn \rightarrow Zn^{+2} + 2\ e^{-1} \qquad E^0 = +0.76\ V$$

$$\boxed{E^0 = +1.56\ V}$$