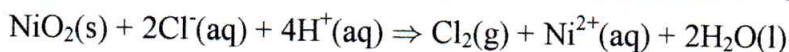


Name  
**electrochemistry: Equilibrium Calculations**

$\log K = nE^\circ/0.0592$   
 (R = 8.31 T = 298K)

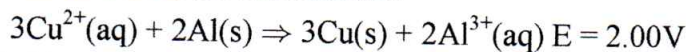
$\Delta G^\circ = -nFE^\circ \quad \Delta G^\circ = -RT \ln K$

1. Calculate  $\Delta G^\circ$  for the following reaction, given that its standard cell potential is 0.320V at 25C.



$\Delta G^\circ = -nFE$   
 $-2 \cdot 96500 \cdot 0.32 = -617$

2. Calculate Kc for this reaction:



2 options

$\Delta G^\circ = -nFE$

$-6 \cdot 96500 \cdot 2$   
 $= -1,158,000 J$

$\Delta G^\circ = -RT \ln K$

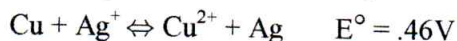
$-1158 = -8.31(298) \ln K$

solve

$\log K = \frac{nE}{0.0592}$

$\log K = \frac{6 \cdot 2}{0.0592}$

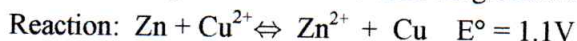
4. As this reaction proceeds over time explain the following:



- a) As this reaction proceed the  $[Cu^{2+}]$  (increases or decreases)  
 b) As this reaction proceed the  $[Ag^+]$  (increases or decreases)  
 c) If you were to increase the concentration of the  $[Ag^+]$  to 2M the V will (increase or decrease)  
 d) If you were to decrease the concentration of the  $[Cu^{2+}]$  to .01M the V will (increase or decrease)  
 e) By the  $E^\circ = .46V$ , this reaction will proceed in which direction to get to products?

$E$  is + so  $\rightarrow$

5. In the drawing below label the following items for the given reaction:



\*\*\* Use arrows on your drawing and label the tops of the arrows to indicate location of item.

Label:

$E^\circ$

$E$  when conditions are  $[Cu^{2+}] = .1M \quad [Zn^{2+}] = 1M$

$\Delta G^\circ$

K

$\Delta G^\circ$  at equilibrium

