1. Consider the two reduction processes and their standard reduction potentials ($E°$).

$$Cu^{+}\,(aq) + 1\,e^{-} \rightarrow Cu\,(s) \qquad E° = 0.521\text{ V}$$

$$Ag^{+}\,(aq) + 1\,e^{-} \rightarrow Ag\,(s) \qquad E° = 0.800\text{ V}$$

A) Write the net ionic equation for a Galvanic/voltaic cell based on these reactions.

$$Ag^{+}\,(aq) + Cu\,(s) \rightleftharpoons Cu^{+}\,(aq) + Ag\,(s)$$

B) Determine the value of the $E°_{cell}$.

$$E°_{\text{cell}} = E°_{\text{cathode}} - E°_{\text{anode}} = 0.800\text{ V} - 0.521\text{ V} = 0.279\text{ V}$$

C) Determine the value of the standard free energy change of the cell ($\Delta G°_{cell}$).

$$\Delta G°_{\text{cell}} = -nFE°_{cell} = -(1\text{ mol } e^{-})\left(96500\frac{\text{C}}{\text{mol } e^{-}}\right)(0.279\text{ V}) = -2.69 \times 10^{4}\text{ J}$$

D) Determine the equilibrium constant (K) for the reaction.

$$E_{\text{cell}} = E°_{\text{cell}} - \frac{RT}{nF}\ln Q$$

$$\ln K = \frac{nF}{RT}E°_{\text{cell}}$$

$$K = \exp\left\{\frac{(1\text{ mol } e^{-})\left(96500\frac{\text{C}}{\text{mol } e^{-}}\right)}{\left(8.314\frac{\text{J}}{\text{mol}\cdot\text{K}}\right)(298.15\text{ K})} \times 0.279\text{ V}\right\} = 5.21 \times 10^{4}$$

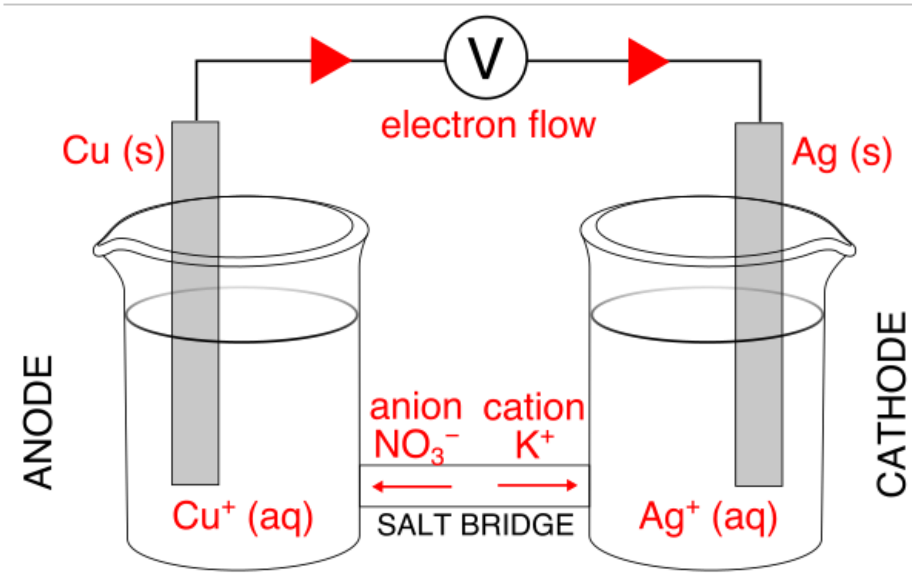
E) Given below is an unlabeled diagram. Label the following components in the diagram:

i. The solid electrodes on the anode and cathode sides.

ii. The ions in solutions on the anode and cathode sides.

iii. The direction of the flow of electrons through the voltmeter and wire.

iv. The direction of the flow of cations and anions in a salt bridge made of KNO_3 (aq).



F) Write the cell diagram for this electrochemical cell.

Cu (s) I Cu+ (1.0 M) II Ag+ (1.0 M) I Ag (s)

2. You have constructed a Galvanic cell with the following reaction under standard conditions.

$$Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s) \qquad E°_{cell} = +1.104\ V$$

What will the potential of the cell be when 0.50 M of Cu^{2+} (aq) has reacted?

Assume that volume and temperature do not change.

	Zn (s)	+	Cu^{2+} (aq)	→	Zn^{2+} (aq)	+	Cu (s)
I	n/a		1.00 M		1.00 M		n/a
C	n/a		– 0.50		+ 0.50		n/a
"E"	n/a		0.50		1.50		n/a

Now use the Nernst equation to find the new cell potential:

$$E_{cell} = E°_{cell} - \frac{RT}{nF} \ln Q$$

$$= 1.104\ V - \frac{\left(8.314 \frac{J}{mol \cdot K}\right)(298.15\ K)}{(2\ mol\ e^-)\left(96500 \frac{C}{mol\ e^-}\right)} \cdot \ln\left(\frac{1.50}{0.50}\right)$$

$$E_{cell} = 1.090\ V$$

3. Consider an electrochemical cell with the following cell diagram at 298.15 K.

$$Ni\ (s)\ |\ Ni^{2+}\ (1.25\ M)\ ||\ Cu^{2+}\ (0.225\ M)\ |\ Ni\ (s)$$

Given the following $E°$ values, determine whether each statement is true or false.

$$Cu^{2+}\ (aq) + 2\ e^- \rightarrow Cu\ (s) \qquad E° = 0.342\ V$$

$$Ni^{2+}\ (aq) + 2\ e^- \rightarrow Ni\ (s) \qquad E° = -0.257\ V$$

$$Ni\ (s) + Cu^{2+}\ (aq) \rightarrow Cu\ (s) + Ni^{2+}\ (aq) \qquad E°_{cell} = 0.599\ V$$

A) E_{cell} is a smaller value than $E°_{cell}$.
   True, $E_{cell} = 0.578\ V < E°_{cell}$.
B) The oxidation reaction takes place at the anode.
   True, oxidation always takes place at the anode.
C) Adding 1.0 L of water to both the anodic and cathodic solutions will increase the cell potential.
   False, this will have no effect on the E_{cell} because the reaction quotient Q does not change.
D) Decreasing the concentration of Ni^{2+} will increase the cell potential.
   True, this will shift reaction to the right, thereby increasing the E_{cell} relative to 0.578 V.
E) Increasing the concentration of Cu^{2+} will increase the cell potential.
   True, this will shift reaction to the right, thereby increasing the E_{cell} relative to 0.578 V.
F) Using a Pt electrode in place of the Ni electrode will not change the cell potential.
   False, this will eliminate the concentration of Ni^{2+} over time.
G) The mass of the Cu electrode will decrease over time.
   False, the mass of the Cu electrode (a product) increases over time.