$$Zn^{2+} + 2e^- \rightarrow Zn \quad E^\circ_{red} -0.763V$$
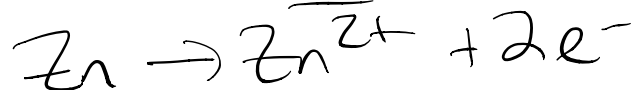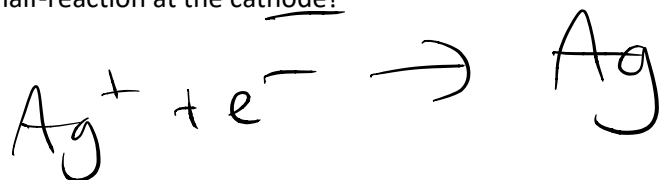$$Ag^+ + 1e^- \rightarrow Ag \quad E_{red} = 0.7994$$

1. One half cell of a voltaic cell uses a silver electrode and 1.0 M silver nitrate solution. Another half cell uses a zinc electrode and 1.0 M zinc nitrate. Answer the following questions

a. What is the half-reaction at the anode? $\rightarrow$ oxidation

$$Zn \rightarrow Zn^{2+} + 2e^-$$

b. What is the half-reaction at the cathode? $\rightarrow$ reduction

$$Ag^+ + e^- \rightarrow Ag$$

c. What is the balanced redox reaction?
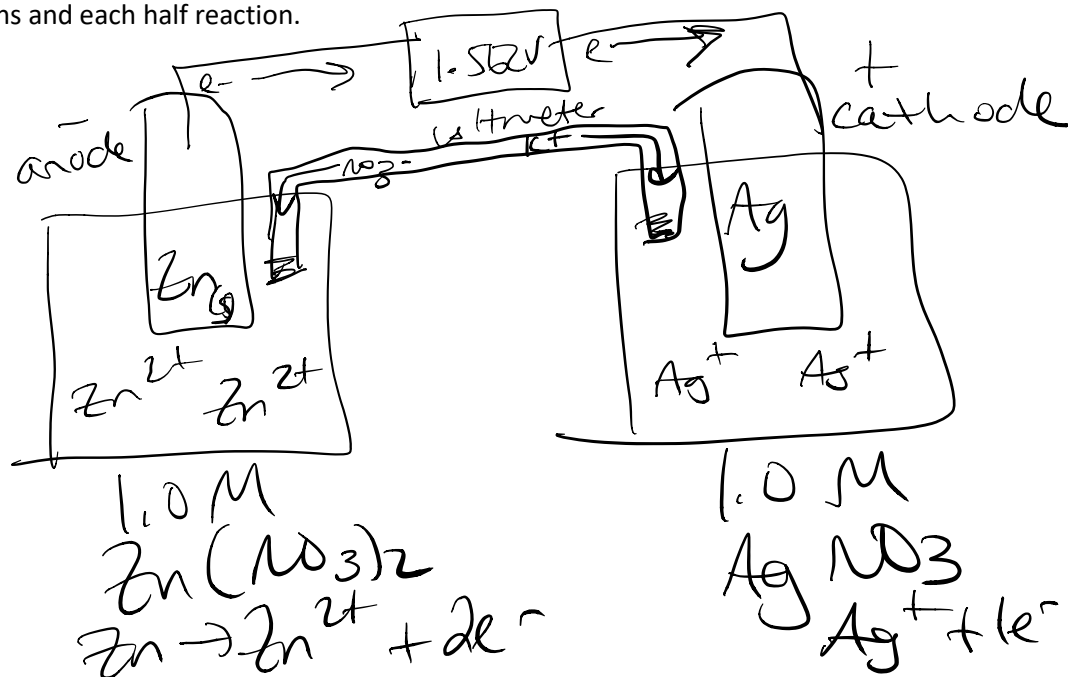
$$Zn_{(s)} + 2Ag^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)}$$

d. At 25.0 °C, what is the standard cell potential (E°cell?)

$$E^\circ cell = E^\circ_{cathode} - E^\circ_{anode}$$
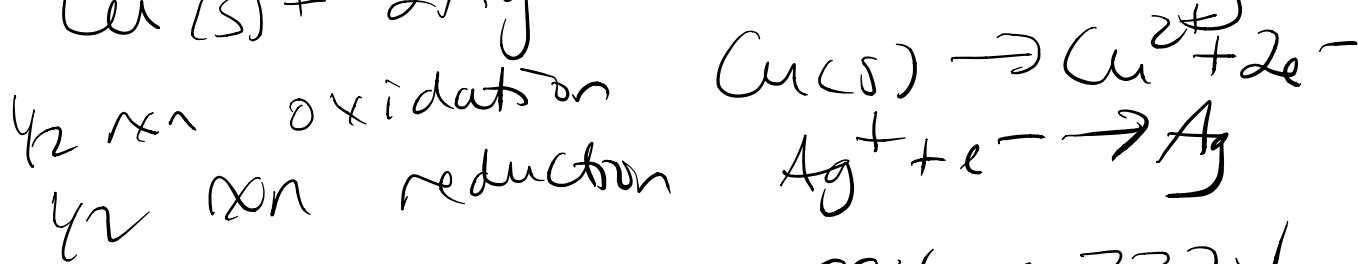$$E^\circ cell = \frac{0.7994V - (-0.763V)}{}$$
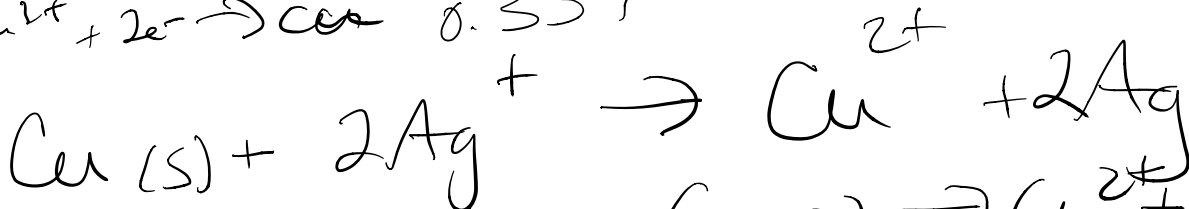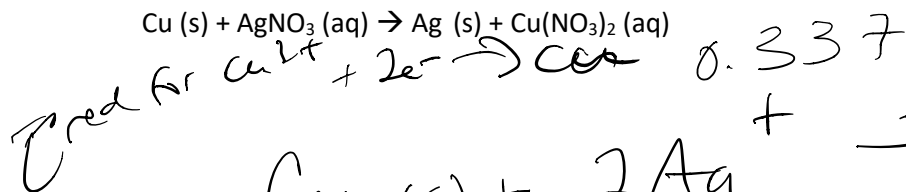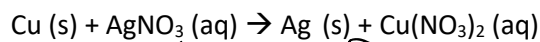$$\boxed{E^\circ cell = 1.562V}$$

e. Make sketch of the voltaic cell label anode, cathode, salt bridge ions, voltage, movement of electrons and each half reaction.

NO_3^- spectator
anode
cathode
$1.562V$
saltmeter
NO_3^-
Zn
Ag
Zn^{2+} Zn^{2+}
Ag^+ Ag^+
1.0 M
$Zn(NO_3)_2$
$Zn \rightarrow Zn^{2+} + 2e^-$
1.0 M
$Ag NO_3$
$Ag^+ + 1e^- \rightarrow Ag$

2. Is the following reaction spontaneous at 25.0 °C. Start by writing a net ionic equation and balancing the redox reaction.

$E°$ cell > 0 for spon. rxn.
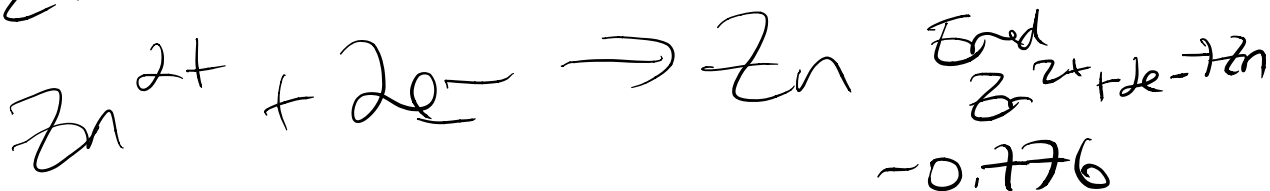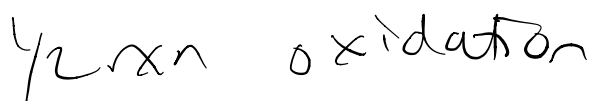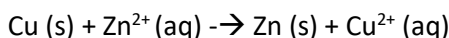
Cu (s) + AgNO$_3$ (aq) → Ag (s) + Cu(NO$_3$)$_2$ (aq)

E_{red} for $Cu^{2+} + 2e^- → Cu$ 0.337

$$Cu (s) + 2Ag^+ → Cu^{2+} + 2Ag$$

½ rxn oxidation $Cu(s) → Cu^{2+} + 2e^-$

½ rxn reduction $Ag^+ + e^- → Ag$

$E_{red} = 0.7994$

$$E° cell = 0.7994 - 0.337 V =$$

$$\boxed{0.4574}$$

yes!
$E°$ cell
> 0

3. What is the E°cell for the following reaction, is the reaction spontaneous or non-spontaneous at 25.0°C.

Cu (s) + Zn^{2+} (aq) -→ Zn (s) + Cu^{2+} (aq)

½ rxn oxidation

$$Cu(s) → Cu^{2+} + 2e^-$$

E_{red}
$Cu^{2+} + 2e^- → Cu$
0.337v

½ rxn reduction

$$Zn^{2+} + 2e^- → Zn$$

E_{red}
$Zn^{2+} + 2e^- → Zn$
~0.776

$$E° cell = -0.776 - 0.337 V =$$

$$\boxed{-1.113 V}$$

non- spontaneous !

$\overline{E}_{cell} < 0$