7.4 Electrochemistry Worksheet

1. Calculate the voltage produced for each of the electrochemical cells containing the substances below. On each diagram, label:

   i. A cathode and the substance it is made of

   ii. An anode and the substance it is made of

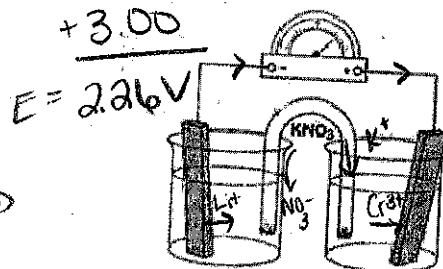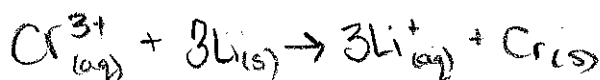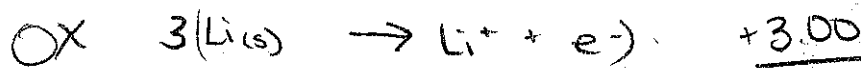   iii. The correct half reaction under each beaker

   iv. The ions coming off or attaching to each electrode

   v. The direction of the ions in the salt bridge

   vi. The direction of the flow of electrons

-0.41 -0.13

a) $(Fe \mid Fe^{2+})$ and $(Pb \mid Pb^{2+})$

ox   red

anode   cathode

red. $Pb^{2+} + 2e \rightarrow Pb$ -0.13

ox $Fe_{(s)} \rightarrow Fe^{2+} + 2e$ $+0.44$

$Pb^{2+} + Fe \rightarrow Pb + Fe^{2+}$ $E = 0.31V$
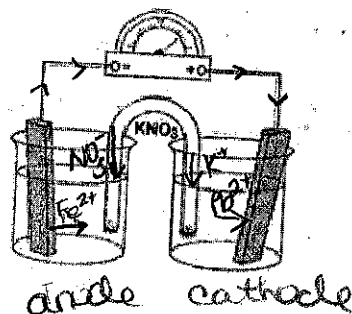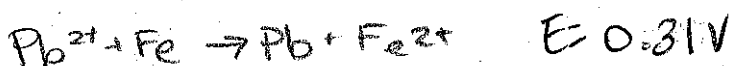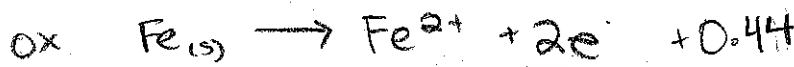
anode   cathode

red   ox

b) $(Cr \mid Cr^{3+})$ and $(Li \mid Li^{+})$

cath.   ano.

RED $Cr^{3+}_{(aq)} + 3e^{-} \rightarrow Cr_{(s)}$ -0.74

OX $3(Li_{(s)} \rightarrow Li^{+} + e^{-})$ $+3.00$

$E = 2.26V$

$Cr^{3+}_{(aq)} + 3Li_{(s)} \rightarrow 3Li^{+}_{(aq)} + Cr_{(s)}$

$3Li \rightarrow 3Li^{+} + 3e^{-}$

ANODE $Cr^{3+} + 3e^{-} \rightarrow Cr$

CATHODE

2. Determine the E° for a (Ag+ | Ag) and (Zn|Zn²⁺) electrochemical cell

RED $(Ag^+ + 1e^- \rightarrow Ag)2$ $+0.80$

OX $\underline{Zn \rightarrow Zn^{2+} + 2e^- \quad +0.76}$

$2Ag^+ + Zn \rightarrow 2Ag^+ Zn^{2+}$ $E = 1.56V$
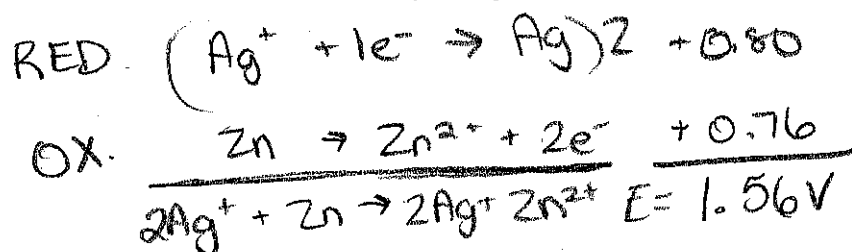
3. Determine E° for a $Fe_{(s)} | Fe^{2+}_{(aq)} | | Cu^{2+}_{(aq)} | Cu_{(s)}$

OX $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^- \quad +0.44$

RED $\underline{Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \quad +0.34}$

$Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$ $E = 0.78V$
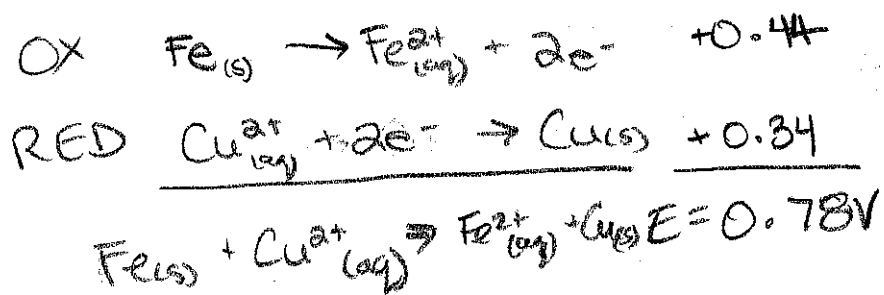
4. a) An electrochemical cell is created using <u>gold</u> and <u>magnesium</u> half-cells. Write the redox reaction. Determine which half-cell will undergo oxidation and which will undergo reduction, identify which substance is the anode and cathode, and calculate the voltage for the cell. You do not need to draw a diagram of the cell.

RED $2(Au^{3+} + 3e^- \rightarrow Au_{(s)})$ $+1.50$ CATHODE

OX $\underline{(Mg_{(s)} \rightarrow Mg^{2+} + 2e^-)3 \quad +2.37}$ ANODE

$2Au^{3+} + 3Mg_{(s)} \rightarrow 3Mg^{2+} + 2Au$ $E = 3.87V$
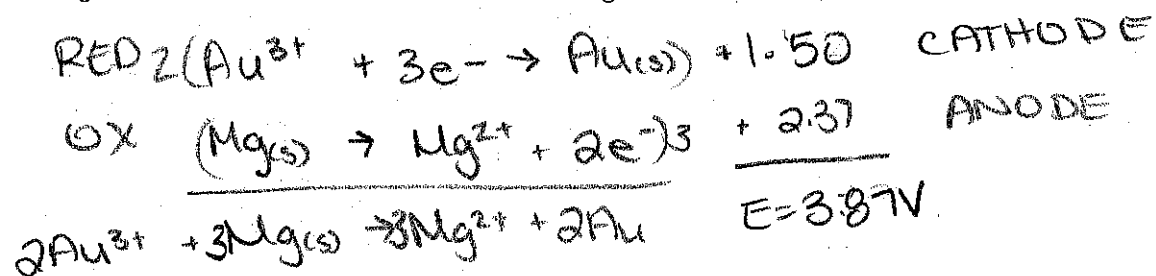
b. If the mass of the magnesium electrode ~~changes~~ *decreased* by 5.0 g, what will be the change in mass of the gold electrode, and will its mass increase or decrease? (hint – use mass to mass stoichiometry using the redox equation found in part a)

↳mass will increase

Mg: $n = \dfrac{wt}{mm} = \dfrac{5.0g}{24.3g/mol} = 0.206$ mol Mg

$\dfrac{3mol\ Mg}{0.206\ mol} : \dfrac{2mol\ Au}{x}$

$x = 0.137\ mol\ Au$

$wt = n \cdot mm = (0.137\ mol)(196.97g/mol) = 27.0g\ Au$