1. Consider the following half reactions and standard reduction potentials.

Au^{3+} (*aq*) + 3 e⁻ → Au (*s*)	$E°$ = +1.400 V
Cl_2 (*g*) + 2 e⁻ → 2 Cl^- (*aq*)	$E°$ = +1.358 V
Cu^{2+} (*aq*) + 2 e⁻ → Cu (*s*)	$E°$ = +0.342V
2 H^+ (*aq*) + 2 e⁻ → H_2 (*g*)	$E°$ = 0 V
Cd^{2+} (*aq*) + 2 e⁻ → Cd (*s*)	$E°$ = −0.403 V
Zn^{2+} (*aq*) + 2 e⁻ → Zn (*s*)	$E°$ = −0.762 V

A) Circle (○) the strongest oxidizing agent and box (□) the strongest reducing agent.

B) Write the balanced chemical equation and cell diagram for the Galvanic cell with the largest standard cell potential (E^o_{cell}).

The largest standard cell potentials would arise from: 2 Au^{3+} (*aq*) + 3 Zn (*s*) → 2 Au (*s*) + 3 Zn^{2+} (*aq*)

$$E^o_{cell} = E°(Au^{3+}/Au) - E°(Zn^{2+}/Zn) = 1.400\ V - (-0.762\ V) = +2.162\ V$$

The cell (line) diagram would be: Zn (*s*) I Zn^{2+} (1.00 M) II Au^{3+} (1.00 M) I Au (*s*)

2. How long will it take to plate 0.0625 g of solid copper from an aqueous solution of copper(II) sulfate with a current of 0.200 A? Note: 1 A = 1 C/s and F = 96,500 C/mol e^-

$$0.0625\ g\ Cu \times \frac{1\ mol\ Cu}{63.55\ g\ Cu} \times \frac{2\ mol\ e^-}{1\ mol\ Cu} \times \frac{96500\ C}{1\ mol\ e^-} \times \frac{1\ s}{0.200\ C} = 949\ s$$

-or-

$$0.200\frac{C}{s} \times \frac{1\ mol\ e^-}{96500\ C} \times \frac{1\ mol\ Cu}{2\ mol\ e^-} \times \frac{63.55\ g\ Cu}{1\ mol\ Cu} \times \frac{1}{0.0625\ g\ Cu} = 0.00105_4\ s^{-1} \rightarrow 949\ s$$

3. A Galvanic cell is constructed using two Au electrodes and two Au^{3+} solutions: one is 0.123 M and the other is 0.449 M. Fill in the concentrations in the cell diagram below for this cell?

Au (*s*) | Au^{3+} (0.123 M) || Au^{3+} (0.449 M) | Au (*s*)

4. You construct the following Galvanic/voltaic cell at 298.15 K.

Pt (*s*) | Cr^{2+} (0.30 M), Cr^{3+} (2.00 M) || Co^{2+} (0.20 M) | Co (*s*) E^o_{cell}= +0.220 V

What will the potential of the cell be after 0.10 M of Cr^{2+} is consumed?

Assume that volume and temperature do not change.

	2 Cr^{2+} (aq)	+	Co^{2+} (aq)	→	2 Cr^{3+} (aq)	+	Co (s)
I	0.30 M		0.20 M		2.00 M		n/a
C	– 0.10		– ½ × 0.10		+ 0.10		n/a
F	0.20		0.15		2.10		n/a

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

$$E_{cell} = E°_{cell} - \frac{RT}{nF}\ln\left(\frac{[Cr^{3+}]^2}{[Cr^{2+}]^2[Co^{2+}]}\right)$$

$$= 0.220\ V - \frac{\left(8.314\frac{J}{mol \cdot K}\right)(298.15\ K)}{(2\ mol\ e^-)\left(96500\frac{C}{mol\ e^-}\right)} \cdot \ln\left(\frac{[2.10]^2}{[0.20]^2 \times [0.15]}\right)$$

$$E_{cell} = 0.135\ V$$

5. You construct the following Galvanic cell at 298.15 K.

$$Fe\ (s)\ |\ Fe^{2+}\ (0.10\ M)\ ||\ Cd^{2+}\ (0.95\ M)\ |\ Cd\ (s) \qquad E^{o}_{cell} = +0.044\ V$$

The initial mass of the Fe electrode is 100.0 g and the volumes of the solutions are 1.00 L each. What will the cell potential be when the mass of the Fe electrode is 62.0 g? {Fe = 55.85 g/mol}

Assume that temperature does not change.

The net ionic equation is: $Fe\ (s) + Cd^{2+}\ (aq) \rightarrow Fe^{2+}\ (aq) + Cd\ (s)$

$$n_{Fe} = (100.0 - 62.0)\ g\ Fe \times \frac{1\ mol\ Fe}{55.85\ g\ Fe} = 0.680_4\ mol\ Fe\ consumed$$

$$n_{Cd^{2+}} = 0.95\ mol - \left(0.680_4\ mol\ Fe \times \frac{1\ mol\ Cd^{2+}}{1\ mol\ Fe}\right) = 0.26_{96}\ mol\ Cd^{2+}\ leftover$$

$$n_{Fe^{2+}} = 0.10\ mol + \left(0.680_4\ mol\ Fe \times \frac{1\ mol\ Fe^{2+}}{1\ mol\ Fe}\right) = 0.78_{04}\ M\ Fe^{2+}\ total$$

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

$$E_{cell} = 0.044\ V - \frac{\left(8.314 \frac{J}{mol \cdot K}\right)(298.15\ K)}{(2\ mol\ e^-)\left(96500 \frac{C}{mol\ e^-}\right)} \cdot \ln\left(\frac{0.78_{04}\ mol\ Fe^{2+}/1.00\ L}{0.26_{96}\ mol\ Cd^{2+}/1.00\ L}\right)$$

$$E_{cell} = 0.030\ V$$

6. You construct a voltaic cell using the two reactions given.

$$Sn^{2+}\ (aq) + 2\ e^- \rightarrow Sn\ (s) \qquad E^\circ = -0.143\ V$$

$$Pb^{2+}\ (aq) + 2\ e^- \rightarrow Pb\ (s) \qquad E^\circ = -0.131\ V$$

If the cell starts with $[Sn^{2+}]$ = 1.35 M and $[Pb^{2+}]$ = 2.11 M at 298.15 K, what will be the concentration of Pb^{2+} when the cell is "dead"?

A "dead" cell is one that has reached equilibrium: $\Delta G = 0$ and $E_{cell} = 0\ V$.

	Sn (s)	+	Pb^{2+} (aq)	→	Sn^{2+} (aq)	+	Pb (s)
I	n/a		2.11 M		1.35 M		n/a
C	n/a		− x		+ x		n/a
F	n/a		2.11 − x		1.35 + x		n/a

$$0\ V = 0.012\ V - \frac{\left(8.314 \frac{J}{mol \cdot K}\right)(281\ K)}{(2\ mol\ e^-)\left(96500 \frac{C}{mol\ e^-}\right)} \cdot \ln\left(\frac{1.35 + x}{2.11 - x}\right)$$

$$x = 1.17_4\ M$$

Now solve for the concentration of Pb^{2+}

$$[Pb^{2+}] = 2.11 - 1.17_4 = 0.94\ M$$

7. Consider a voltaic cell with the following cell diagram at 298.15 K.

$$Pb\ (s)\ |\ Pb^{2+}\ (1.00\ M)\ ||\ Cu^{+}\ (1.00\ M)\ |\ Cu\ (s)$$

A) What will happen to the cell potential if $[Pb^{2+}]$ is doubled?

Increases   <u>Decreases</u>   Stays the same

B) What will happen to the cell potential if $[Cu^{+}]$ is doubled?

<u>Increases</u>   Decreases   Stays the same

C) What will happen to the cell potential if we added enough water to double the volumes of both the anodic and cathodic solutions?

Increases   <u>Decreases</u>   Stays the same