Worksheet 15 KEY: Redox Reactions 1

1. C > A > B > E > D. C(s) is most active, as its electrons are removed by ions of A, B, E, and D. There are no visible reactions between E(s) and the ions of C, A, and B, so E(s) is less active, but more active than H2 (there is a reaction between E and HCl).

2. This would be difficult, as Li reacts with water as does the product, Ca(s). Adding Li(s) to water will result in a reaction even without the presence of Ca^{2+}.

3. Separate into half reactions:
   (1) Al(s) → Al(OH)4^−
   (2) H2O(l) → H2(g)

   Balance according to plan on the following worksheet:
   \[
   8\text{H}_2\text{O} + 2\text{Al} \rightarrow 2\text{Al(OH)}_4^- + 8\text{H}^+ + 6\text{e}^-
   \]
   \[
   6\text{e}^- + 6\text{H}^+ + 3\text{H}_2\text{O} \rightarrow 3\text{H}_2 + 3\text{H}_2\text{O}
   \]
   add together and cancel:
   \[
   8\text{H}_2\text{O} + 2\text{Al} + 2\text{OH}^- \rightarrow 2\text{Al(OH)}_4^- + 3\text{H}_2
   \]
   reaction is in basic solution, add 2OH^−
   \[
   6\text{H}_2\text{O} + 2\text{Al} + 2\text{OH}^- \rightarrow 2\text{Al(OH)}_4^- + 3\text{H}_2
   \]
   check to see if atoms and charges balance.
Worksheet 16 key  Redox Reactions

1. From the solid, use $0.500 \text{ L} \times 0.01255 \text{ M} \text{ CaCl}_2 \times 111.0 \text{ g/mol} = 0.697 \text{ g}$ of $\text{CaCl}_2$ dissolve this in water and bring the volume up to 0.500 L

From the solution, use $0.500 \text{ L} \times 0.01255 \text{ mol CaCl}_2 \times 1\text{L/0.425 mol} = 14.8 \text{ mL}$, add this volume of the solution to water and bring the volume up to 0.500 L

2. Oxidation is loss of $e^-$, reduction is gain of $e^-$.

3. (a) $\text{Na}$ is the reducing agent, $\text{Cl}_2$ is the oxidizing agent, and $\text{Na}$ is oxidized.

(b) $\text{MnO}_4^-$ is the oxidizing agent, $\text{H}_2\text{C}_2\text{O}_4$ is the reducing agent, and $\text{C}$ is oxidized

4. a. $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2e^- \quad \text{and} \quad 2\text{Ag}^{+}(\text{aq}) + 2e^- \rightarrow 2\text{Ag(s)}$ so $\text{Cu(s)} + 2\text{Ag}^{+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag(s)}$

b. $2\text{Cr}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 2e^- \quad \text{and} \quad \text{I}_2(\text{aq}) + 2e^- \rightarrow 2\text{I}^-(\text{aq})$ so $2\text{Cr}^{2+}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 2\text{I}^-$

5. 1st half reaction:  
$\text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{ e}^-$

2nd half reaction:  
$\text{VO}_2^+ \rightarrow \text{VO}^{2+}$  
(separate into ½ reactions)

$\text{VO}_2^+ \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$  
(add $\text{H}_2\text{O}$ to balance $\text{O}$ atoms)

$2 \text{H}^+ + \text{VO}_2^+ \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$  
(add $\text{H}^+$ to balance $\text{H}$ atoms)

$\text{e}^- + 2 \text{H}^+ + \text{VO}_2^+ \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$  
(add $\text{e}^-$ to balance charge)

$2\text{e}^- + 4\text{H}^+ + 2\text{VO}_2^+ \rightarrow 2\text{VO}^{2+} + 2\text{H}_2\text{O}$  
(multiply by 2 so $\text{e}^-$ are equal)

$2\text{e}^- + 4\text{H}^+ + 2\text{VO}_2^++ \text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{ e}^- + 2\text{VO}^{2+} + 2\text{H}_2\text{O}$  
(add ½ rxns together)

$4\text{H}^+ + 2\text{VO}_2^+ + \text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{VO}^{2+} + 2\text{H}_2\text{O}$  
(cancel $\text{e}^-$)

6. $\text{Al}$ must lose $\text{e}^-$ more easily, as it is oxidized by the $\text{Sn}$ ions

7. Just like the prelab, $\text{C}$ is most active, then $\text{D}$, then $\text{B}$, $\text{A}$ is least active
Worksheet 17 Key: Electrochemistry

1. The shorthand notation presents the electrochemical cell components from anode to cathode, left to right. The electrodes are on the outer edges (anode to left, cathode to right). Single vertical lines show phase boundaries, and are used between (s), (l) or (aq), and (g). If more than one reactant or product is present in the same phase, commas are used for separation. A double vertical line is for the half-cell boundary. The half-cell reactions are written with reacting substances on the left, and products on the right. Sometimes inactive electrodes are present, like the Pt(s) in the SHE (read the text!). An example cell that operates with a SHE coupled to a copper/copper (II) half-cell would be written as: Pt(s) | H₂(g) | H⁺(aq) || Cu²⁺(aq) | Cu(s) to show that at a platinum electrode (anode), H₂ is oxidized to H⁺ and that Cu²⁺ is reduced to Cu(s) at the cathode.

2. Mg(s) | Mg²⁺(aq) || Pb²⁺(aq) | Pb(s)

3. $E^{\circ}_{cell} = E_{red(Pb)} - E_{red(Mg)} = 0.36 \text{ V} - (-2.71 \text{ V}) = 3.07 \text{ V}$ (great voltage, but Mg reacts with water and air)

4. In a galvanic cell, the negative electrode is the anode (oxidation), so

\[
E^{\circ}_{cell} = E^{\circ}_{red(SHE)} - E^{\circ}_{red(Co)} = 0 - E^{\circ}_{red(Co)} \quad \text{so} \quad 0.28 \text{ V} = 0 \text{ V} - E^{\circ}_{red(Co)}
\]

\[
E^{\circ}_{red(Co)} = -0.28 \text{ V}
\]

5. Mg | Mg²⁺ || Cr³⁺ | Cr, reaction is: 3Mg(s) + 2Cr³⁺ → 3Mg²⁺ + 2 Cr(s)

\[
E^{\circ}_{cell} = -0.74 \text{ V} - (-2.37 \text{ V}) = 1.63 \text{ V}, \text{ the Mg is the anode}
\]

6. $5.00 \text{ A} \times 600 \text{ sec} = 3000 \text{ A} \times \text{s} = 3000 \text{ C} \quad 3000 \text{ C} \times 1 \text{ mol e}^-/96500 \text{ C} = 0.0311 \text{ mole e}^-

0.0311 \text{ mol e}^- \times 1 \text{ mol Ni/2 mole e}^- \times 58.69 \text{ g Ni/1 mol} = 0.912 \text{ g Ni metal}

7. Cu → Cu²⁺ + 2 e⁻

1 mol Cu = 2 mol e⁻

1.50 g Cu (1 mol Cu/63.55 g Cu)(2 mol e⁻/1 mol Cu)(9.65 × 10⁻⁴ C/1 mol e⁻)(A s/C)(1/30 min)(1 min/60 sec) = 2.53 A

8. Ag will get reduced, Zn will get oxidized (reversed)

\[
2 \text{Ag}^+(aq) + 2 \text{e}^- \rightarrow \text{2Ag(s)}
\]

\[
\text{Zn(s)} \rightarrow \text{Zn}^2+(aq) + 2 \text{e}^-
\]

2 Ag⁺(aq) + Zn(s) → 2 Ag(s) + Zn²⁺(aq)

\[
E^{\circ}_{cell} = E^{\circ}_{red} - E^{\circ}_{ox} = E^{\circ}_{Ag^+} - E^{\circ}_{Zn^{2+}} = +0.80 \text{ V} - (-0.76 \text{ V}) = 1.56 \text{ V}
\]

\[
E_{cell} = 1.56 \text{ V} - \frac{RT}{nF} \ln \left[ \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]} \right]
\]

\[
E_{cell} = 1.56 \text{ V} - \frac{(8.314 \text{ J/molK})(298 \text{ K})}{(2 \text{ mol e}^-)(9.65 \times 10^4 \text{ C/mol e}^-)} \ln \left[ \frac{2.8 \times 10^{-3}}{1.5 \times 10^{-2}} \right]^2
\]

= 1.56 V – 0.0324 V = 1.53 V

9. $\Delta G^o = -nFE^{\circ}_{cell} = -(2 \text{ mole e}^-)(9.65 \times 10^4 \text{ C/mol e}^-)(0.60 \text{ J/C})$

= $-1.2 \times 10^7 \text{ J} = -1.2 \times 10^2 \text{ kJ (spontaneous)}$