1. Using the copper/zinc "General Chemistry Cell" discussed in lecture as a model, construct a lead/cobalt cell assembled with a piece of lead, PbSO₄ (aq), a piece of cobalt, and CoSO₄ (aq) [the electrons flow to the lead compartment]. Identify all cell components, half-reactions, and details—which species is being reduced, oxidized, the oxidizing agent, the reducing agent, gaining e⁻, and losing e⁻.

The question states that the electrons flow to the lead compartment. This is in agreement with Table 18.1; the "lead" half-reaction appears above the "cobalt" half reaction:

\[
Pb^{2+} (aq) + 2 e^- \rightarrow Pb (s) \quad \text{E}^\circ = -0.125 \text{ V} \quad | \quad \text{Difference} = 0.152 \text{ V}
\]

\[
Co^{2+} (aq) + 2 e^- \rightarrow Co (s) \quad \text{E}^\circ = -0.277 \text{ V} \quad |
\]

Pb²⁺ (aq) desires electrons more than Co²⁺ (aq); in other words, Pb²⁺ (aq) is a better oxidizing agent than Co²⁺ (aq). Electrons will flow from the "cobalt compartment" to the "lead compartment."

The half-reactions are:

\[
Pb^{2+} (aq) + 2 e^- \rightarrow Pb (s) \quad [\text{This is the half-reaction shown in Table 18.1 because } \text{"lead" is gaining electrons}].
\]

and

\[
Co (s) \rightarrow Co^{2+} (aq) + 2 e^- \quad [\text{This is flipped from the half reaction shown in Table 18.1 because } \text{"cobalt" is not gaining electrons, it is having \text{electrons taken away}].
\]

Pb²⁺ (aq) gained two electrons to become Pb (s).
Co (s) lost two electrons to become Co²⁺ (aq).
Pb²⁺ (aq) is being reduced to Pb (s).
Co (s) is being oxidized to Co²⁺ (aq).
Pb²⁺ (aq) is the oxidizing agent—it is oxidizing (taking electrons away) from Co (s).
Co (s) is the reducing agent—it is reducing (giving electrons to) Pb²⁺ (aq).

Anions (negatively charged ions) flow through the electrolyte to the "cobalt" compartment. This flow is in the opposite direction than the electron flow. This makes for a "complete circuit" and maintains neutrally charged beakers.
Can you write the cell components, electron flow, anion flow, and half-reactions on this figure?

2. (a) Determine the oxidation state of manganese in Mn(NO₃)₂.
(b) Determine the oxidation state of calcium in Ca(NO₃)₂.
(c) Determine the oxidation state of tin in Sn(NO₃)₄.
(d) Determine the oxidation state of iron in Fe(NO₃)₂.
(e) Determine the oxidation state of iron in Fe(NO₃)₃.
(f) Determine the oxidation state of all species in KMnO₄.
(g) Determine the oxidation state of all species in Pb₃O₄ (this is not a typo!).

(a) The oxidation number of Mn is 2+ because each nitrate ion is 1-.

\[
\begin{array}{c}
\text{NO}_3^- \\
\text{Mn}^{2+} \\
\text{NO}_3^-
\end{array}
\]

(b) The oxidation number of Ca is 2+ because each nitrate ion is 1-.

\[
\begin{array}{c}
\text{NO}_3^- \\
\text{Ca}^{2+} \\
\text{NO}_3^-
\end{array}
\]

Additionally, Ca is 2+ because Ca is in Group 2 of the Periodic Table of the Elements.

(C) The oxidation number of Sn is 4+ because each nitrate ion is 1-.

\[
\begin{array}{c}
\text{NO}_3^- \\
\text{Sn}^{4+} \\
\text{NO}_3^-
\end{array}
\]
The oxidation number of Fe is 2+ because each nitrate ion is 1-.

\[
\begin{array}{c}
\text{NO}_3^- \\
\text{Fe}^{2+} \\
\text{NO}_3^-
\end{array}
\]

The oxidation number of Fe is 3+ because each nitrate ion is 1-.

\[
\begin{array}{c}
\text{NO}_3^- \\
\text{Fe}^{3+} \\
\text{NO}_3^-
\end{array}
\]

The oxidation number of Mn is 7+ because each O is 2- and the K is 1+.

\[
\begin{array}{c}
\text{O}^2- \\
\text{K}^+ \\
\text{Mn}^{7+} \\
\text{O}^2-
\end{array}
\]

Each oxygen is 2- so the total negative charge is 8-. The total positive charge must be 8+. Therefore, each Pb has an oxidation number of +8/3.

3. List a primary cell. What can it be used in?

Primary cells are cells that have one life—use once and then recycle. Today's AAA, AA, C, D, and 9-volt primary batteries are usually referred to as "alkaline batteries." Eveready, Duracell, and Rayovac are several manufacturers of these batteries. These batteries are commonly used in flashlights, remote controls...

4. List four secondary cells. What are their uses?

Pb-ACID, Ni-Cd, Ni-MH, Li-ION are four examples of secondary cells. These cells are rechargeable (they can be cycled hundreds and even thousands of times). Used in automobiles (starting, lighting, and ignition), portable consumer electronic devices (laptop computers, cellular telephones, pagers, CD players...)

5. We observed zinc metal in CuSO\(_4\) (aq). Write the two half-reactions for this process.

The half-reactions are:

\[\text{Cu}^{2+} (\text{aq}) + 2 \text{e}^- \rightarrow \text{Cu} (s)\]  
[This is the half-reaction shown in Table 18.1 because "copper" is gaining electrons].
and

\[ \text{Zn (s)} \rightarrow \text{Zn}^{2+} \text{(aq)} + 2 \text{e}^- \]  
 [This is flipped from the half reaction shown in Table 18.1 because "zinc" is not gaining electrons, it is having electrons taken away].

Why is sulfate ion not shown in the half-reactions? It is a spectator ion.

6. **Consider the placement of aluminum metal in Cu(NO}_3 \text{)}_2 \text{(aq) to produce Al(NO}_3 \text{)}_3 \text{(aq) and copper metal. Write the two half-reactions for this process. Which species is gaining electrons? Which species is losing electrons? ID the reducing and oxidizing agents.**

The half-reactions are:

\[ \text{Cu}^{2+} \text{(aq)} + 2 \text{e}^- \rightarrow \text{Cu (s)} \]  
 [This is the half-reaction shown in Table 18.1 because "copper" is gaining electrons].

and

\[ \text{Al (s)} \rightarrow \text{Al}^{3+} \text{(aq)} + 3 \text{e}^- \]  
 [This is flipped from the half reaction shown in Table 18.1 because "aluminum" is not gaining electrons, it is having electrons taken away].

Why is nitrate ion not shown in the half-reactions? It is a spectator ion.

\[ \text{Cu}^{2+} \text{(aq)} \text{ has gained two electrons. It is reduced to Cu (s). It is the oxidizing agent because it is oxidizing Al (s).} \]

\[ \text{Al (s)} \text{ has lost three electrons. It is oxidized to Al}^{3+} \text{(aq). It is the reducing agent because it is reducing Cu}^{2+} \text{(aq).} \]

Note that the two half reactions involve a different number of transferred electrons (the "copper" half-reaction involves the gaining of two electrons while the "aluminum" half-reaction involves the loss of three electrons). Let's balance these half-reactions and add them together to determine the overall reaction:

\[ \text{Cu}^{2+} \text{(aq)} + 2 \text{e}^- \rightarrow \text{Cu (s)} \]
\[ \text{Al (s)} \rightarrow \text{Al}^{3+} \text{(aq)} + 3 \text{e}^- \]

If we multiply the "copper" half-reaction by three we see six electrons transferred:

\[ 3 \left[ \text{Cu}^{2+} \text{(aq)} + 2 \text{e}^- \rightarrow \text{Cu (s)} \right] = 3 \text{Cu}^{2+} \text{(aq)} + 6 \text{e}^- \rightarrow 3 \text{Cu (s)} \]
If we multiply the "aluminum" half-reaction by two we see six electrons transferred:

$$2 \ [\text{Al (s) } \rightarrow \text{Al}^{3+} (aq) + 3 \text{ e}^-] = 2 \ \text{Al (s) } \rightarrow 2 \ \text{Al}^{3+} (aq) + 6 \ \text{e}^-$$

Adding the two half-reactions with the transfer of six electrons:

$$3 \ \text{Cu}^{2+} (aq) + 6 \ \text{e}^- \rightarrow 3 \ \text{Cu} (s)$$

$$+ \quad 2 \ \text{Al (s) } \rightarrow 2 \ \text{Al}^{3+} (aq) + 6 \ \text{e}^-$$

$$\underline{3 \ \text{Cu}^{2+} (aq) + 6 \ \text{e}^- + 2 \ \text{Al (s)} \rightarrow 3 \ \text{Cu} (s) + 2 \ \text{Al}^{3+} (aq) + 6 \ \text{e}^-}$$

The electrons cancel on each side of the reaction arrow (six electrons leave the anode and six electrons arrive at the cathode):

$$3 \ \text{Cu}^{2+} (aq) + 6 \ \text{e}^- + 2 \ \text{Al (s)} \rightarrow 3 \ \text{Cu} (s) + 2 \ \text{Al}^{3+} (aq) + 6 \ \text{e}^-$$

Are we mass balanced? Yes, there are three copper atoms on the left side of the reaction arrow and three on the right side. There are two aluminum atoms on the left of the reaction arrow and two on the right side.

Are we charge balanced? Yes, the charge on each side of the reaction arrow is 6+. How did we get 6+ on the left? How did we get 6+ on the right?

Note that the overall reaction:

$$3 \ \text{Cu}^{2+} (aq) + 2 \ \text{Al (s)} \rightarrow 3 \ \text{Cu} (s) + 2 \ \text{Al}^{3+} (aq)$$

appears deceptively simple! It does not represent all the work we did to mass balance, charge balance, and describe the redox details. One thing is clear from the overall reaction, redox occurred—electrons were transferred. We know this because the oxidation numbers of species changed. In this case the oxidation number of copper changed from 2+ to zero and the oxidation number of aluminum changed from zero to 3+.

7. Consider the following balanced reaction:

$$\text{SnCl}_2 (aq) + 2 \ \text{FeCl}_3 (aq) \rightarrow \text{SnCl}_4 (aq) + 2 \ \text{FeCl}_2 (aq)$$

Write the two half-reactions. Identify the oxidation numbers of each species. Identify the reducing agent, the oxidizing agent...
I will break down the overall reaction into two half-reactions:

$\text{Fe}^{3+} (\text{aq}) \rightarrow \text{Fe}^{2+} (\text{aq})$

$\text{Sn}^{2+} (\text{aq}) \rightarrow \text{Sn}^{4+} (\text{aq})$

How did I determine the oxidation numbers? Why did I leave chloride ion out of the half-reactions?

$\text{Fe}^{3+} (\text{aq})$ is being reduced to $\text{Fe}^{2+} (\text{aq})$. Recognizing the oxidation number decreased from 3+ to 2+ shows reduction:

\[ \text{Fe}^{3+} (\text{aq}) + e^- \rightarrow \text{Fe}^{2+} (\text{aq}) \]

$\text{Sn}^{2+} (\text{aq})$ is being oxidized to $\text{Sn}^{4+} (\text{aq})$. Recognizing the oxidation number increased from 2+ to 4+ shows oxidation:

\[ \text{Sn}^{2+} (\text{aq}) \rightarrow \text{Sn}^{4+} (\text{aq}) + 2 \text{e}^- \]

Summary:
Fe$^{3+}$ (aq) has gained an electron.
Fe$^{3+}$ (aq) is reduced.
Fe$^{3+}$ (aq) is the oxidizing agent.

Sn$^{2+}$ (aq) has lost two electrons.
Sn$^{2+}$ (aq) is oxidized.
Sn$^{2+}$ (aq) is the reducing agent.

8. Balance Ca (s) + VO$_4^{3-}$ (aq) → V$^{2+}$ (aq) + Ca$^{2+}$ (aq) in acid. Identify details (species reduced, oxidized...)

Steps 1 and 2: Write the half-reactions and identify which is reduction and which is oxidation

Ca (s) → Ca$^{2+}$ (aq)
(Oxidation—the oxidation number of calcium is increasing from 0 to 2+)

VO$_4^{3-}$ (aq) → V$^{2+}$ (g)
( Reduction—the oxidation number of vanadium is decreasing from 5+ to 2+)

Step 3: Balance atoms (except H and O)

This step is not necessary—calcium and vanadium are balanced.

Step 4: Balance oxygen by adding water

This step is not necessary for the calcium half-reaction.

VO$_4^{3-}$ (aq) → V$^{2+}$ (g) + 4 H$_2$O (l)

Step 5: Balance hydrogen by adding protons

This step is not necessary for the copper half-reaction.

8 H$^+$ (aq) + VO$_4^{3-}$ (aq) → V$^{2+}$ (g) + 4 H$_2$O (l)

Step 6: Charge balance by adding electrons

Consider: Ca (s) → Ca$^{2+}$ (aq)

(Because the total charge of the reactants is 0 and the total charge of the products is 2+, two electrons are required on the products side. This makes the total charge on each side equal to 0. This is consistent with the identification of this half-reaction as being oxidation—see Steps 1 and 2 above)

Ca (s) → Ca$^{2+}$ (aq) + 2 e$^-$

Consider: 8 H$^+$ (aq) + VO$_4^{3-}$ (aq) → V$^{2+}$ (g) + 4 H$_2$O (l)
(Because the total charge of the reactants is +5 and the total charge of the products is +2, three electrons are required on the reactants side. This makes the total charge on each side equal to +2. This is consistent with the identification of this half-reaction as being reduction—see Steps 1 and 2 above)

\[3 \text{e}^- + 8 \text{H}^+ (\text{aq}) + \text{VO}_4^{3-} (\text{aq}) \rightarrow \text{V}^{2+} (\text{g}) + 4 \text{H}_2\text{O} (\text{l})\]

**Step 7:** Use "multipliers" to make the number of electrons gained equal to the number of electrons lost

\[3 \times [\text{Ca} (s) \rightarrow \text{Ca}^{2+} (\text{aq}) + 2 \text{e}^-]\]

\[2 \times [3 \text{e}^- + 8 \text{H}^+ (\text{aq}) + \text{VO}_4^{3-} (\text{aq}) \rightarrow \text{V}^{2+} (\text{g}) + 4 \text{H}_2\text{O} (\text{l})]\]

The oxidation half-reaction (the one on the top) was multiplied by three so each half-reaction shows six electrons being transferred.

The reduction half-reaction (the one on the bottom) was multiplied by two so each half-reaction shows six electrons being transferred.

**Step 8:** Add the half-reactions for the Overall Reaction

\[3 \text{Ca} (s) \rightarrow 3 \text{Ca}^{2+} (\text{aq}) + 6 \text{e}^-\]

\[+ \quad 6 \text{e}^- + 16 \text{H}^+ (\text{aq}) + 2 \text{VO}_4^{3-} (\text{aq}) \rightarrow 2 \text{V}^{2+} (\text{g}) + 8 \text{H}_2\text{O} (\text{l})\]

\[\underline{3 \text{Ca} (s) + 6 \text{e}^- + 16 \text{H}^+ (\text{aq}) + 2 \text{VO}_4^{3-} (\text{aq}) \rightarrow 2 \text{V}^{2+} (\text{g}) + 8 \text{H}_2\text{O} (\text{l}) + 3 \text{Ca}^{2+} (\text{aq}) + 6 \text{e}^-}\]

**Step 9:** Cancel

\[3 \text{Ca} (s) + 16 \text{H}^+ (\text{aq}) + 2 \text{VO}_4^{3-} (\text{aq}) \rightarrow 2 \text{V}^{2+} (\text{g}) + 8 \text{H}_2\text{O} (\text{l}) + 3 \text{Ca}^{2+} (\text{aq})\]

**Step 10:** Check mass and charge balancing

Same atoms on each side of the reaction arrow
Same charge (10+) on each side of the reaction arrow

**9.** What is *current*? What are the units of *current*? Give an analogy for *current*. What are the units for *charge*? What charge does a mole of electrons have? What charge does an electron have?

Current is the rate of electron flow (the quantity of charge per second). The units of current are Amps (1 Amp = 1 coulomb per second). An analogy for current might be cars...
on a freeway: the current is high when a large number of cars are traveling on the freeway (many cars per second flow); the current is low when few cars are traveling down a residential street (few cars per second flow). The units for charge are coulombs. One mole of electrons \((6.02 \times 10^{23} \text{ electrons})\) has a charge of 96,500 coulombs; this is known as Faraday's Constant. A single electron has a charge of 96,500 coulombs/6.02 x \(10^{23}\) = 1.60 x \(10^{-19}\) coulombs.

10. A student provides a current of 6.00 amps through a solution of AgNO\(_3\) (aq) for 4.00 hours. The voltage is such that silver metal is deposited at the cathode. Determine the mass of silver deposited.

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

4.00 hours (60 min/1 hour)(60 s/1 min) = 14400 s

14400 s (6.00 Coulombs/1 s) = 86400 Coulombs

\(\text{1 Amp} = 1 \text{ Coulomb/s}\)

86400 Coulombs(1 mole e\(^-\)/96500 Coulombs) = 0.895 moles e\(^-\)

0.895 moles e\(^-\) (1 mol Ag/1 mole e\(^-\)) = 0.895 moles Ag

0.895 moles Ag (107.9 g/mol) = 96.6 grams Ag