1. Assign oxidation numbers to each atom in:

\[ \text{NiCl}_2 \]
- **Nickel ion charge would be +2, so oxidation number is +2**
- **Chloride ion charge would be –1, so each chlorine has an ox # of -1**

\[ \text{Mg}_2\text{TiO}_4 \]
- **Magnesium ion charge would be +2, so oxidation number is +2**
- **Oxygen ox # (from the rules) is -2**
- **Titanium ox # of +4, so that the whole thing adds up to zero charge**

\[ \text{K}_2\text{Cr}_2\text{O}_7 \]
- **Oxygen ox # (from the rules) is -2**
- **Potassium ion charge would be +1, so oxidation number is +1**
- **Chromium ox # of +6, so that the whole thing adds up to zero charge**

\[ \text{HPO}_4^{2-} \]
- **Oxygen ox # (from the rules) is -2**
- **Hydrogen ox # (from the rules) is +1**
- **Phosphorus ox # of +6, so that the whole thing adds up to –1 charge**

2. Identify whether each reaction is a redox reaction. If so, identify the oxidizing agent, the reducing agent, the substance being oxidized, and the substance being reduced.

a. \[ \text{Zn (s) + 2HCl (aq)} \rightarrow \text{ZnCl}_2 \text{ (aq) + H}_2 \text{(g)} \]
- **Reactant ox #** Zn is 0, H is +1, Cl is -1
- **Product ox #** Zn is +2, H is 0, Cl is –1
- Since there is a change in ox #, this is a redox reaction. Since Zn’s ox # increased, it was oxidized and is the reducing agent. Since hydrogen’s ox # decreased, it was reduced and is the oxidizing agent

b. \[ \text{Cr}_2\text{O}_7^{2-} \text{(aq) + 2OH- (aq)} \rightarrow \text{CrO}_4^{2-} \text{(aq) + H}_2\text{O (l)} \]
- **Reactant ox #** Cr is +6, O is -2, H is +1
- **Product ox #** Cr is +6, O is -2, H is +1
- Since there is no change in ox #s, this is not a redox reaction

3. Identify the substances oxidized and reduced in the reaction of calcium and chlorine to form calcium chloride. What are the oxidizing and reducing agents?

**Balanced equation:** \[ \text{Ca} \ + \ \text{Cl}_2 \rightarrow \text{CaCl}_2 \]
- **Reactant ox #** Ca is 0, Cl is 0
- **Product ox #** Ca is +2, Cl is –1
- So oxidizing agent is Cl (which is reduced) and the reducing agent is Ca (which is oxidized)
1. Balance the following equations:

CH$_3$OH (aq) + Cr$_2$O$_7^{2-}$ (aq) $\rightarrow$ CH$_2$O (aq) + Cr$^{3+}$ (aq) in acidic solution

Write half reactions:

CH$_3$OH $\rightarrow$ CH$_2$O

Balance (not H or O)

CH$_3$OH $\rightarrow$ CH$_2$O

Balance O using H$_2$O

CH$_3$OH $\rightarrow$ CH$_2$O

Balance H using H$^+$

CH$_3$OH $\rightarrow$ CH$_2$O + 2H$^+$

Balance charge using e-

3CH$_3$OH $\rightarrow$ 3CH$_2$O + 6H$^+$ + 6e-

Add and cancel

3CH$_3$OH + 8 H$^+$ + Cr$_2$O$_7^{2-}$ $\rightarrow$ 3CH$_2$O + 2Cr$^{3+}$ + 7 H$_2$O

Cl$_2$ (g) $\rightarrow$ Cl$^-$ (aq) + ClO$^-$ (aq) in basic solution

Write half reactions:

Cl$_2$ $\rightarrow$ Cl$^-$

Balance (not H or O)

Cl$_2$ $\rightarrow$ 2Cl$^-$

Balance O using H$_2$O

Cl$_2$ $\rightarrow$ 2Cl$^-$

Balance H using H$^+$

Cl$_2$ $\rightarrow$ 2Cl$^-$

Balance charge using e-

Cl$_2$ + 2e$^-\rightarrow$ 2Cl$^-$

Add and cancel

2Cl$_2$ + 2H$_2$O $\rightarrow$ 2Cl$^-$ + 2ClO$^-$ + 4H$^+$

Add 4 OH$^-$ to each side

4 OH$^-\rightarrow$ 2Cl$_2$ + 2H$_2$O $\rightarrow$ 2Cl$^-$ + 2ClO$^-$ + 4H$^+$ + 4 OH$^-$

Make H$_2$O from H$^+$ and OH$^-$

4 OH$^-\rightarrow$ 2Cl$_2$ + 2H$_2$O $\rightarrow$ 2Cl$^-$ + 2ClO$^-$ + 4H$_2$O

Cancel H$_2$O

4 OH$^-\rightarrow$ 2Cl$_2$ $\rightarrow$ 2Cl$^-$ + 2ClO$^-$ + 2H$_2$O

Reduce coefficients

2 OH$^-\rightarrow$ Cl$^- +$ ClO$^- +$ H$_2$O
1. Given the following cell notation:

\[
\text{Cr (s)} \quad || \quad \text{Cr}^{3+} (aq) \quad || \quad \text{Cu}^{2+} (aq) \quad || \quad \text{Cu (s)}
\]

Write a balanced equation for the cell reaction and give a brief description of the cell.

- **The anode is first:** \(\text{Cr (s)} \rightarrow \text{Cr}^{3+} (aq) + 3e^-\)
- **The cathode is second:** \(\text{Cu}^{2+} (aq) + 2e^- \rightarrow \text{Cu (s)}\)

Make the number of electrons the same and add the reactions to get:

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

The cell is a chromium anode dipped into a solution of \(\text{Cr}^{3+}\) ions and a copper cathode dipped into a solution of \(\text{Cu}^{2+}\) ions. The two solutions are connected by a salt bridge.

2. Write the shorthand notation for a galvanic cell that uses the reaction:

\[
\text{Mg (s)} + \text{Fe}^{2+} (aq) \rightarrow \text{Mg}^{2+} (aq) + \text{Fe (s)}
\]

Splitting into half reactions:

- \(\text{Mg (s)} \rightarrow \text{Mg}^{2+} (aq) + 2e^-\)
- \(\text{Fe}^{2+} (aq) + 2e^- \rightarrow \text{Fe (s)}\)

So Mg is oxidized (this is the anode) and the Fe is reduced (this is the cathode).

\[
\text{Mg (s)} || \text{Mg}^{2+} (aq) || \text{Fe}^{2+} (aq) || \text{Fe (s)}
\]
1. A galvanic cell consists of a Mg electrode in a 1.0 M Mg(NO$_3$)$_2$ solution and a Ag electrode in a 1.0 M AgNO$_3$ solution. Calculate the E° for this cell.

The reduction potentials are:
E° (Mg$^{2+}$/Mg) $= -2.37$ V
E° (Ag$^+$/Ag) $= 0.80$ V

Silver has the more positive reduction potential, so it will be reduced.
So E°$_{cell}$ $= E°_{Ag^+/Ag} - E°_{Mg^{2+}/Mg} = 0.80 - (-2.37) = 3.17$ V

2. Determine whether the following reactions are spontaneous as written:

a. 2Al$^{3+}$ + 6I$^-$ $\rightarrow$ 3I$_2$ + 2Al

Al is reduced, I- is oxidized.
So E°$_{cell}$ $= E°_{Al^{3+}/Al} - E°_{I_2/I^-} = (-1.66) - (+0.54) = -2.20$ V
Since E° is negative the reaction is NOT spontaneous

b. 3Fe$^{2+}$ + 2NO + 4H$_2$O $\rightarrow$ 3Fe + 2NO$_3^-$ + 8H$^+$

Fe$^{2+}$ is reduced, NO is oxidized.
So E°$_{cell}$ $= E°_{Fe^{2+}/Fe} - E°_{NO_3^-/NO} = (-0.45) - (+0.96) = -1.41$ V
Since E° is negative the reaction is NOT spontaneous

c. 3Ca + 2Cr$^{3+}$ $\rightarrow$ 2Cr + 3Ca$^{2+}$

Cr$^{3+}$ is reduced, Ca is oxidized.
So E°$_{cell}$ $= E°_{Cr^{3+}/Cr} - E°_{Ca^{2+}/Ca} = (-0.74) - (-2.87) = +2.13$ V
Since E° is positive the reaction is spontaneous
1. A current of 0.452 A is passed through an electrolytic cell containing molten CaCl$_2$ for 1.5 hours. Write the electrode reactions and calculate the quantity of products (in grams) formed at the electrode.

**Cathode:**  $\text{Ca}^{2+} (l) + 2\text{e}^- \rightarrow \text{Ca} (l)$  
**Anode:**  $2\text{Cl}^- (l) \rightarrow \text{Cl}_2 (g) + 2\text{e}^-$

$$1.50 \text{ hours} \times (60 \text{ min/hr}) \times (60 \text{ sec/min}) = 5.40 \times 10^3 \text{ seconds}$$

Charge = current x time  
Charge = 0.452 A x (5.40 x 10$^3$ seconds) = 2.44 x 10$^3$ C  
$(2.44 \times 10^3 \text{ C}) \times (1 \text{ mole e}^- / 96485 \text{ C}) = 0.0253 \text{ mole e}^-$

Using the half reactions:  
**Cathode:**  0.0253 mole e$^- \times (1 \text{ mole Ca}/2 \text{ mole e}^-) \times (40.08 \text{ g/mol}) = 0.507 \text{ g Ca}$  
**Anode:**  0.0253 mole e$^- \times (1 \text{ mole Cl}_2/2 \text{ mole e}^-) \times (70.906 \text{ g/mol}) = 0.897 \text{ g Cl}_2$

2. A constant current is passed through an electrolytic cell containing molten MgCl$_2$ for 18 hours. If 4.8 x 10$^5$ grams of Cl$_2$ are obtained, what was the current (in amperes)?

**Anode:**  $2\text{Cl}^- (l) \rightarrow \text{Cl}_2 (g) + 2\text{e}^-$  
4.8 x 10$^5$ grams x (1 mole Cl$_2$/70.906 g) x (2 mole e$^-$/1 mole Cl$_2$) x (96485 C/mol e$^-$) = 1.31 x 10$^9$ C = 1.31 x 10$^9$ As  

18 hrs x (3600 sec/hr) = 6.48 x 10$^4$ sec  
$1.31 \times 10^9 \text{ As} / 6.48 \times 10^4 \text{ sec} = 2.02 \times 10^4 \text{ A}$

3. How many hours are required to produce 1000 kg of sodium by electrolysis of molten NaCl with a constant current of 30,000 A? How many liters of Cl$_2$ at STP will be obtained as a byproduct?

**Cathode:**  $\text{Na}^+ (l) + \text{e}^- \rightarrow \text{Na} (l)$  
**Anode:**  $2\text{Cl}^- (l) \rightarrow \text{Cl}_2 (g) + 2\text{e}^-$

1000 kg Na x (1000 g / 1 kg) x (1 mol Na/22.99 g) x (1 mol e$^-$/1 mole Na) = 4.350 x 10$^4$ mole e$^-$

$4.350 \times 10^4 \text{ mole e}^- \times (96485 \text{ C} / 1 \text{ mole e}^-) = 4.197 \times 10^9 \text{ C} = 4.197 \times 10^9 \text{ As}$  

4.197 x 10$^9$ As / 30,000 A = 1.399 x 10$^5$ sec = 38.86 hrs

$4.350 \times 10^5 \text{ mole e}^- \times (1 \text{ mol Cl}_2 / 2 \text{ mole e}^-) = 2.175 \times 10^4 \text{ mole Cl}_2$

PV = nRT  
STP:  1 atm and 273 K  
V = $(2.175 \times 10^4 \text{ mole}) \times (0.082057 \text{ L} \cdot \text{atm/mol} \cdot \text{K}) \times (273 \text{ K}) / 1 \text{ atm} = 4.872 \times 10^5 \text{ L}$
1. Calculate the standard free energy change and \( K_c \) for the following reaction at 25 \( ^\circ \text{C} \).

\[
\text{Sn (s) + 2Cu}^{2+} (\text{aq}) \rightarrow \text{Sn}^{2+} (\text{aq}) + 2\text{Cu}^+ (\text{aq})
\]

**Half reactions:**

**Oxidation:** \( \text{Sn} \rightarrow \text{Sn}^{2+} + 2e^- \)

**Reduction:** \( 2\text{Cu}^{2+} + 2e^- \rightarrow 2\text{Cu}^+ \)

So \( E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cu}^{2+/Cu^+}} - E^{\circ}_{\text{Sn}^{2+/Sn}} = (0.16) - (-0.14) = +0.29 \text{ V} \)

\( \Delta G^\circ = -nF E^{\circ}_{\text{cell}} \)

\( \Delta G^\circ = -(2 \text{ mole e}) (96485 \text{ C/mole e}) (0.29 \text{ V}) \)

\( \Delta G^\circ = -58 \text{ kJ} \) (recall that \( C = J/V \))

\( E^{\circ}_{\text{cell}} = \frac{RT}{nF} \ln K \)

\( \ln K = \frac{nE^{\circ}_{\text{cell}}}{0.0257 \text{ V}} = \frac{(2) (0.29 \text{ V})}{0.0257 \text{ V}} = 22.6 \)

\( K = 6.4 \times 10^9 \)

Note 0.0257 \text{ V} is the result of RT/F for 25 \( ^\circ \text{C} \)

2. Predict whether the following reaction would proceed spontaneously as written at 298 K:

\[
\text{Co (s) + Fe}^{2+} (\text{aq}) \rightarrow \text{Co}^{2+} (\text{aq}) + \text{Fe (s)}
\]

Given that \([\text{Co}^{2+}] = 0.15 \text{ M} \) and \([\text{Fe}^{2+}] = 0.68 \text{ M} \)

**The half reactions:**

**Oxidation:** \( \text{Co} \rightarrow \text{Co}^{2+} + 2e^- \)

**Reduction:** \( \text{Fe}^{2+} + 2e^- \rightarrow \text{Fe} \)

\( E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Fe}^{2+/Fe}} - E^{\circ}_{\text{Co}^{2+/Co}} = (-0.45) - (-0.28) = -0.17 \text{ V} \)

\( E = E^{\circ}_{\text{cell}} - 0.0257 \text{ V} / n \ln Q \)

From the overall reaction (recalling that pure solids do not appear in Q expression)

\( Q = [\text{Co}^{2+}] / [\text{Fe}^{2+}] = 0.15 / 0.68 = 0.22 \)

\( E = (-0.17) - (0.0257 \text{ V} / 2) (\ln 0.22) = -0.15 \text{ V} \)

Since \( E \) is negative, the reaction is NOT spontaneous in the direction written.

3. A galvanic cell was set up using the following reaction:

\[
\text{Mg (s) + Cd}^{2+} (\text{aq}) \rightarrow \text{Mg}^{2+} (\text{aq}) + \text{Cd (s)}
\]

The magnesium electrode was dipped into a 1.0 M solution of MgSO\(_4\) and the cadmium electrode was dipped into a solution of unknown Cd\(^{2+}\) concentration. The potential of this cell was measured to be 1.54 V. What is the Cd\(^{2+}\) concentration?

**The half reactions:**

**Oxidation:** \( \text{Mg} \rightarrow \text{Mg}^{2+} + 2e^- \)

**Reduction:** \( \text{Cd}^{2+} + 2e^- \rightarrow \text{Cd} \)

\( E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cd}^{2+/Cd}} - E^{\circ}_{\text{Mg}^{2+/Mg}} = (-0.40) - (-2.37) = +1.97 \text{ V} \)

\( E = E^{\circ}_{\text{cell}} - (0.0257 \text{ V} / n) \ln Q \)

From the overall reaction (recalling that pure solids do not appear in Q expression)

\( Q = [\text{Mg}^{2+}] / [\text{Cd}^{2+}] = 1.0 / [\text{Cd}^{2+}] \)