Electrochemical Cells II – Supplemental Worksheet  KEY

All the electrochemical cells on this worksheet are the same ones on the first Electrochemical Cells worksheet. To make the work on this worksheet easier, refer to the work you did on the previous Electrochemical Cells worksheet.

Experimental Observations of Electrochemical Cells

1. Consider the voltaic cell that contains standard Co^{2+}/Co and Au^{3+}/Au electrodes. The following experimental observations were noted: (1) Metallic gold plates out on one electrode and the gold ion concentration decreases around that electrode, and (2) The mass of the cobalt electrode decreases and the cobalt (II) ion concentration increases around that electrode.

   a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

   ![Diagram](image)

   **Anode**
   Co → Co^{2+} + 2e-
   (oxidation)

   **Cathode**
   Au^{3+} + 3e^- → Au
   (reduction)

   Overall Redox Reaction: 2Au^{3+} + 3Co → 3Co^{2+} + 2Au

   <Co(s)|Co^{2+}(aq)||Au^{3+}(aq)|Au(s)>
   
   \[ E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 1.50 \text{ V} - (-0.28 \text{ V}) = 1.78 \text{ V} \]

   b. What is the standard reaction free energy for this cell?

   \[ \Delta G^\circ = -nF E^\circ = -\left(6 \text{ mol e}^- \right) \left(96,485 \text{ C/mol} \right) \left(1.78 \text{ J/C} \right) \left(1 \text{ mol e}^- \right) = -1030459.8 \frac{\text{J}}{\text{mol reactant}} \]

   \[ = -1030.46 \text{ kJ/mol reactant} \]
c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.

Because we are looking for \( K \) at 25°C, we can use the equation 
\[
\log K = \frac{n \times E^\circ}{0.0591}
\]

So, 
\[
K = 10^{\frac{n \times E^\circ}{0.0591}} = 10^{\frac{6 \times 1.78}{0.0591}} = 5.14 \times 10^{180}
\]

This is a very large \( K \), this reflects the spontaneous nature of this reaction.

d. Calculate the emf at 25°C of this cell in which the concentration of \( \text{Co}^{2+} \) ions is 0.30 mol/L and that of the \( \text{Au}^{3+} \) ions is 0.0010 mol/L.

Here we use the Nerst equation. Since we are finding the emf at 25°C we can use this version of the Nerst equation:
\[
E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log Q = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log \left[ \frac{[\text{Co}^{2+}]^3}{[\text{Au}^{3+}]^2} \right] = 1.78 \text{ V} - \frac{0.0591}{6} \log \left[ \frac{0.30}{0.0010} \right]^3
\]
\[
= 1.74 \text{ V}
\]

e. Assume you use this cell as a battery to power one of your electric devices that draws 1 mA of current. If you ran the battery at a constant 1 mA for 3000 hours, how many grams of Co would be consumed?

To figure this problem out, we will do a series of conversions: current and time \( \rightarrow \) charge \( \rightarrow \) moles of e\(^-\) \( \rightarrow \) moles of Co \( \rightarrow \) grams of Co

current and time \( \rightarrow \) charge: charge = current * time = 0.001 C/s * (3000 hrs * 60 min/hr * 60 s/min) = 10800 C

charge \( \rightarrow \) moles of e\(^-\): 10800 C * \( \frac{1 \text{ mol e}^-}{96,485 \text{ C}} \) = 0.1119 mol e\(^-\)

moles of e\(^-\) \( \rightarrow \) moles of Co: 0.1119 mol e\(^-\) * \( \frac{1 \text{ mol Co}}{2 \text{ mol e}^-} \) = 0.0560 mol Co

moles of Co \( \rightarrow \) grams of Co: 0.0560 mol Co * \( \frac{58.9 \text{ g Co}}{1 \text{ mol Co}} \) = 3.30 g Co

3.30 g Co would be consumed
2. Consider the electrolysis of molten calcium chloride with inert electrodes. The following experimental observations were noted: (1) Bubbles of pale green chlorine gas are produced at one electrode, and (2) Silvery white molten metallic calcium is produced at the other electrode.

   a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

   \[
   <\text{Pt(s)}|\text{Cl}^{-}(\text{aq})|\text{Cl}_2(\text{g})||\text{Ca}^{2+}(\text{aq})|\text{Ca(s)}|\text{Pt(s)}>  
   \]

   \[E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -2.76 \text{ V} - (1.36 \text{ V}) = -4.12 \text{ V}\]

   b. What is the standard reaction free energy for this cell?

   \[\Delta G^\circ = -nF E^\circ = - \left( \frac{2 \text{ mol e}^-}{1 \text{ mol rxn}} \right) \left( \frac{96485 \text{ C}}{1 \text{ mol e}^-} \right) \left( -4.12 \frac{1}{\text{C}} \right) = 795.04 \text{ kJ mol}^{-1}\text{rxn}^{-1}\]

   c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.

   Because we are looking for K at 25°C, we can use the equation \[\log K = \frac{n E^\circ}{0.0591}\]

   So, \[K = 10^{\frac{-4.12}{0.0591}} = 3.76 \times 10^{-140}\]

   This is a very small K, this reflects the non-spontaneous nature of this reaction.
d. How many hours are required to plate 12.00 g of metallic calcium from 1.00 M CaCl$_2$ (aq) by using a current of 3.00 A?

To figure this problem out, we will do a series of conversions: grams of Ca → moles of Ca → moles of e− → charge → charge and current → time

grams of Ca → moles of Ca: 12 g Ca * $\frac{1 \text{ mol Ca}}{40.1 \text{ g Ca}} = 0.299 \text{ mol Ca}$

moles of Ca → moles of e−: 0.299 mol Ca * $\frac{2 \text{ mol e−}}{1 \text{ mol Ca}} = 0.599 \text{ mol e−}$

moles of e− → charge: 0.599 mol e− * $\frac{96,485 \text{ C}}{1 \text{ mol e−}} = 57746.0 \text{ C}$

charge and current → time: time = $\frac{\text{charge}}{\text{current}} = \frac{57746.0 \text{ C}}{3 \text{ C/s}} = 19249 \text{ s} * $ $\frac{1 \text{ hr}}{3600 \text{ s}} = 5.4 \text{ hr}$

It would require 5.4 hours to plate 12.00 g of metallic calcium from 1.00 M CaCl$_2$ (aq) by using a current of 3.00 A.

e. Determine the volume (in liters, at STP) of chlorine gas that can be produced in this cell by using a current of 7.30 mA for 2.11 hours.

To figure this problem out, we will do a series of conversions: current and time → charge → moles of e− → moles of Cl$_2$ → liters of Cl$_2$

current and time → charge: charge = current * time = 0.0073 C/s * (2.11 hrs * 60 min/hr * 60 s/min) = 55.45 C

charge → moles of e−: 55.45 C * $\frac{1 \text{ mol e−}}{96,485 \text{ C}} = 0.000575 \text{ mol e−}$

moles of e− → moles of Cl$_2$: 0.000575 mol e− * $\frac{1 \text{ mol Cl}_2}{2 \text{ mol e−}} = 0.000287 \text{ mol Cl}_2$

moles of Cl$_2$ → liters of Cl$_2$: 0.000287 mol Cl$_2$ * $\frac{22.4 \text{ L Cl}_2 \text{ at STP}}{1 \text{ mol Cl}_2 \text{ at STP}} = 0.00644 \text{ L Cl}_2 = 6.44 \text{ mL Cl}_2$

6.44 mL Cl$_2$ would be produced
Short Hand Notation of Electrochemical Cells

3. Consider the following cell <Ni(s)|Ni²⁺(aq)||Ag⁺(aq)|Ag(s)>

a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

\[ E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.80 \text{ V} - (-0.23 \text{ V}) = 1.03 \text{ V} \]

b. What is the standard reaction free energy for this cell?

\[ \Delta G^{\circ}_{r} = -nF E^{\circ} = - \left( 2 \text{ mol e}^- \right) \left( \frac{96,485 \text{ C}}{1 \text{ mol e}^-} \right) \left( \frac{1.03}{1} \right) = -198.76 \frac{\text{kJ}}{\text{mol rxn}} \]

c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.

Because we are looking for K at 25°C, we can use the equation \( \log K = \frac{n \cdot E^{\circ}}{0.0591} \).

So, \( K = 10^{\frac{2 \cdot 1.03}{0.0591}} = 10^{23.18} = 7.18 \times 10^{34} \)

This is a very large K, this reflects the spontaneous nature of this reaction.
d. What is the concentration of Ni\(^{2+}\) ions if the emf at 25°C of this cell is +1.68 V and the concentration of Ag\(^+\) ions is 0.0280 mol/L?

Here we use the Nerst Equation. Since we are finding the emf at 25°C we can use this version of the Nerst Equation:

\[ E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log Q \Rightarrow E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log \left( \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \right) \]

\[ \Rightarrow \frac{n}{0.0591} (E^\circ_{\text{cell}} - E_{\text{cell}}) = \log \left( \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \right) \Rightarrow 10^{\frac{n}{0.0591}(E^\circ_{\text{cell}} - E_{\text{cell}})} = \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \]

\[ \Rightarrow [\text{Ni}^{2+}] = [\text{Ag}^+]^2 \times 10^{\frac{n}{0.0591}(E^\circ_{\text{cell}} - E_{\text{cell}})} = [0.0280 \text{ mol/L}]^2 \times 10^{\frac{1.68 V - 1.1 V}{0.0591}} = 1.2 \times 10^{-4} \text{ M} \]

e. How many hours would it take for this galvanic cell to plate 30.00 g of metallic silver from 1.00 M solutions of Ni\(^{2+}\)(aq) and Ag\(^+\)(aq) assuming it produces a constant current of 2.25 A?

To figure this problem out, we will do a series of conversions: grams of Ag → moles of Ag → moles of e- → charge → charge and current → time.

grams of Ag → moles of Ag: 30.0 g Ag * \( \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} = 0.278 \text{ mol Ag} \)

moles of Ag → moles of e-: 0.278 mol Ag * \( \frac{1 \text{ mol e}^-}{1 \text{ mol Ca}} = 0.278 \text{ mol e}^- \)

moles of e- → charge: 0.278 mol e- * \( \frac{96,485 \text{ C}}{1 \text{ mol e}^-} = 26825.9 \text{ C} \)

charge and current → time: time = \( \frac{\text{charge}}{\text{current}} = \frac{26825.9 \text{ C}}{2.25 \text{ C/s}} = 11923 \text{ s} \approx 3.3 \text{ hr} \)

It would require 3.3 hours to plate 30.00 g of metallic silver from 1.00 M solutions of Ni\(^{2+}\)(aq) and Ag\(^+\)(aq) assuming it produces a constant current of 2.25 A.
4. Consider the following cell \(<\text{Pt(s)}|\text{Ce}^{3+}(aq)|\text{Ce}^{4+}(aq)|\text{Cu}^{2+}(aq) |\text{Cu(s)}>\)
   
a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

\[
E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 \text{ V} - (1.70 \text{ V}) = -1.36 \text{ V}
\]

b. What is the standard reaction free energy for this cell?
\[
\Delta G_r^{\circ} = -nF E^{\circ} = -\left(\frac{2 \text{ mol e}^-}{1 \text{ mol rxn}}\right)\left(\frac{96,485 \text{ C}}{1 \text{ mol e}^-}\right)\left(-1.36 \frac{1}{\text{C}}\right) = 262.44 \frac{\text{kJ}}{\text{mol rxn}}
\]

c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.
Because we are looking for K at 25°C, we can use the equation 
\[
\log K = \frac{nE^{\circ}}{0.0591}
\]
So, 
\[
K = 10^{\frac{nE^{\circ}}{0.0591}} = K = 10^{\frac{2(-1.36)}{0.0591}} = 9.47 \times 10^{-47}
\]
This is a very small K, this reflects the non-spontaneous nature of this reaction.
d. What is the concentration of Ce$^{4+}$ ions if the emf at 25°C of this cell is –1.20 V, the concentration of Cu$^{2+}$ ions is 0.60 mol/L and the concentration of Ce$^{3+}$ ions is 0.30 mol/L?

Here we use the Nerst Equation. Since we are finding the emf at 25°C we can use this version of the Nerst Equation:

\[ E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log Q \]

\[ \Rightarrow \frac{n}{0.0591} (E^\circ_{\text{cell}} - E_{\text{cell}}) = \log \frac{[\text{Ce}^{4+}]^2}{[\text{Cu}^{2+}] \cdot [\text{Ce}^{3+}]^2} \]

\[ \Rightarrow [\text{Ce}^{4+}]^2 = [\text{Cu}^{2+}] \cdot [\text{Ce}^{3+}]^2 \cdot 10^{\frac{n}{0.0591}(E^\circ_{\text{cell}} - E_{\text{cell}})} \]

\[ \Rightarrow [\text{Ce}^{4+}] = \sqrt{[\text{Cu}^{2+}] \cdot [\text{Ce}^{3+}]^2 \cdot 10^{\frac{n}{0.0591}(E^\circ_{\text{cell}} - E_{\text{cell}})}} = \sqrt{[0.60 \text{ mol/L}] \cdot [0.30 \text{ mol/L}]^2 \cdot 10^{\frac{2}{0.0591}(-1.36 \text{ V} - (-1.20 \text{ V})}} = 4.56 \times 10^{-4} \text{ M} \]

e. Determine the mass (in grams) of metal copper that can be produced in this cell by using a current of 5.0 A for 2.7 days.

383.8 g Cu would be produced (see problem 1e for a similar explanation for this problem)
Electrochemical Cell Diagrams
5. Consider the following cell:

![Electrochemical Cell Diagram](image)

a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

\[ \text{Anode} \quad \text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^- \quad \text{(oxidation)} \]

\[ \text{Cathode} \quad \text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn} \quad \text{(reduction)} \]

Overall Redox Reaction: \[ \text{Zn}^{2+} + \text{Cu} \rightarrow \text{Cu}^{2+} + \text{Zn} \]

\[ \langle \text{Cu(s)}|\text{Cu}^{2+}(aq)||\text{Zn}^{2+}(aq)|\text{Zn(s)} \rangle \]

\[ E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.76 \text{ V} - (0.34 \text{ V}) = -1.10 \text{ V} \]

b. What is the standard reaction free energy for this cell?

\[ \Delta G^\circ = -nF E^\circ = 212.27 \text{ kJ mol}^{-1} \text{ rxn} \]
c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.

\[ K = 5.96 \times 10^{-38} \]

This is a very small \( K \), this reflects the non-spontaneous nature of this reaction.

d. What is the potential for this cell if the \( \text{Cu}^{2+} \) concentration is \( 5.8 \times 10^{-3} \) M and the \( \text{Zn}^{2+} \) concentration is \( 1.3 \times 10^{-1} \) M?

Here we use the Nerst Equation to find that

\[
E = E^\circ - 0.0591 \frac{n}{n} \log Q = 1.10 \quad 0.0591 \log \left[ \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \right] = 1.10 \quad 0.0591 \log \left( \frac{0.13}{0.0058} \right) = 1.14 \text{V}
\]

e. Determine the mass (in grams) of zinc metal that can be produced in this cell by using a current of 6.0 A for 1.5 days.

263.5 g Zn would be produced (see problem 1e for a similar explanation for this problem)
6. Consider the following cell:

![Diagram of a voltaic cell with a salt bridge, copper, and zinc electrodes.]

a. Recall the diagram, overall balanced redox reaction, and the standard cell potential for this cell?

\[ <\text{Zn}(s)|\text{Zn}^{2+}(aq)||\text{Cu}^{2+}(aq)|\text{Cu}(s)> \]

\[ E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V} \]

b. What is the standard reaction free energy for this cell?

\[ \Delta G^{\circ} = -nF E^{\circ} = -212.27 \text{ kJ/mol rxn} \]
c. Calculate the equilibrium constant for the overall redox reaction of this cell at 25°C.
   \[ K = 1.1 \times 10^{37} \]
   This is a very large \( K \), this reflects the spontaneous nature of this reaction.

d. Calculate the emf at 25°C of this cell in which the concentration of Cu\(^{2+}\) ions is 0.1250 mol/L and that of the Zn\(^{2+}\) ions is 0.0020 mol/L.
   Here we use the Nerst Equation. Since we are finding the emf at 25°C we can use this version of the Nerst Equation:
   \[
   E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log Q = E^\circ_{cell} - \frac{0.0591}{n} \log \left[ \frac{[Zn^{2+}]}{[Cu^{2+}]} \right] = 1.10 \text{ V} - \frac{0.0591}{2} \log \left[ \frac{0.0020}{0.1250} \right]
   \]
   \[= 1.15 \text{ V} \]

e. How many hours would it take for this galvanic cell to plate 15.00 g of metallic copper from 1.00 M solutions of Cu\(^{2+}\)(aq) and Zn\(^{2+}\)(aq) assuming it produces a constant current of 5.0 A?
   It would require 2.53 hours to plate 15.00 g of metallic copper from 1.00 M solutions of Cu\(^{2+}\)(aq) and Zn\(^{2+}\)(aq) assuming it produces a constant current of 5.0 A.