

### Mathematics of Redox Reactions

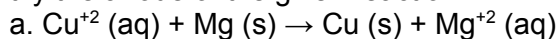
This worksheet will cover the equations and mathematics of reduction-oxidation reactions. It will discuss standard and nonstandard conditions of an electrolytic cell,  $E^\circ_{\text{cell}}$ , and the Faraday constant. As you progress through the worksheet, you will develop the skills necessary to calculate  $E^\circ_{\text{cell}}$ , determine the spontaneity of a cell, find the  $\Delta G^\circ$  (Gibbs free energy), and use the Nernst equation for nonstandard conditions.

Useful equations:

$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, cathode}} - E^\circ_{\text{reduction, anode}}$
$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} (\ln Q)$
$\Delta G^\circ = -nFE^\circ$

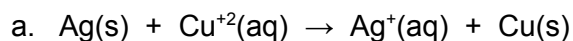
Practice Questions:

1. Identify the anode of the given reaction.

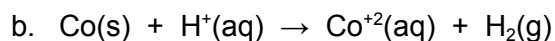


- i. Anode:
- ii. Cathode:

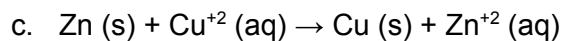
2. Calculate  $E^\circ_{\text{cell}}$  using [standard reduction potentials](#) for the reactions below to determine which are spontaneous.



$\text{Cu}^{+2}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.34 V
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80 V



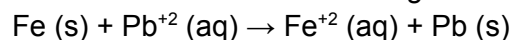
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00 V
$\text{Co}^{+2}(\text{aq}) + 2\text{e}^- \rightarrow \text{Co}(\text{s})$	-0.28 V



$\text{Cu}^{2+} \text{(aq)} + 2\text{e}^{-} \rightarrow \text{Cu (s)}$	+0.34 V
$\text{Zn}^{2+} \text{(aq)} + 2\text{e}^{-} \rightarrow \text{Zn (s)}$	-0.76 V

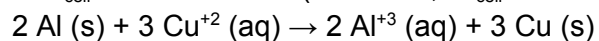
3. If a cell is at the conditions of 237K, 1.5 atm, and 2.0M for its solutions, what equation should be used to determine the  $E_{\text{cell}}$ ?

4. Determine the value of  $\Delta G^{\circ}$  for the following reaction. (Hint: find  $E^{\circ}_{\text{cell}}$  first.)



$\text{Fe}^{2+} \text{(aq)} + 2\text{e}^{-} \rightarrow \text{Fe (s)}$	-0.44 V
$\text{Pb}^{2+} \text{(aq)} + 2\text{e}^{-} \rightarrow \text{Pb (s)}$	-0.13 V

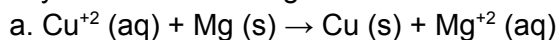
5. Determine the  $E_{\text{cell}}$  of the reaction. ( $T = 298\text{K}$ ,  $E^{\circ}_{\text{cell}} = 2.00 \text{ V}$ ,  $[\text{Al}^{3+}] = 0.1\text{M}$ ,  $[\text{Cu}^{2+}] = 2.5\text{M}$ )



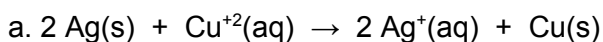
**ANSWER KEY**  
**Mathematics of Redox Reactions**

Practice Questions:

1. Identify the anode of the given reaction.

ii. Anode: **Mg**iii. Cathode: **Cu**

Oxidation occurs at the anode (electrons are lost), whereas reduction occurs at the cathode (electrons are gained). Thus, the anode element will lose electrons in the forward reaction, and the cathode element will gain electrons.

2. Calculate  $E^\circ_{\text{cell}}$  using [standard reduction potentials](#) for the reactions below to determine which are spontaneous.

First, determine which elements are being reduced and oxidized. Since Ag is becoming more positive as a cation, it is losing electrons and is therefore oxidized, thus, Ag would be the anode.  $\text{Cu}^{+2}$  is reduced to become Cu, thus Cu would be the cathode.

The standard reduction potentials needed are:

$\text{Cu}^{+2}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.34 V
$\text{Ag}^{+}(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80 V

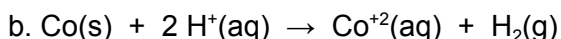
Now, we may use this equation:

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, cathode}} - E^\circ_{\text{reduction, anode}}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, Cu}} - E^\circ_{\text{reduction, Ag}}$$

$$E^\circ_{\text{cell}} = 0.34 \text{ V} - 0.80 \text{ V}$$

$$E^\circ_{\text{cell}} = \mathbf{-0.46 \text{ V}}$$

**The reaction is not spontaneous because  $E^\circ_{\text{cell}}$  is negative.**

Determine which elements are being reduced and oxidized. Since Co is becoming more positive as a cation, it is losing electrons and is therefore oxidized; Co would be the anode.  $\text{H}^{+}$  is reduced to become  $\text{H}_2$ ;  $\text{H}_2$  would be the cathode.

Unit 16: Electrochemistry



The standard reduction potentials needed are:

$2 \text{H}^+ (\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2 (\text{g})$	0.00 V
$\text{Co}^{+2} (\text{aq}) + 2\text{e}^- \rightarrow \text{Co} (\text{s})$	-0.28 V

Now, we may use this equation:

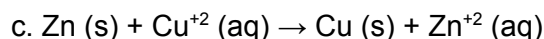
$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, cathode}} - E^\circ_{\text{reduction, anode}}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, H}_2} - E^\circ_{\text{reduction, Co}}$$

$$E^\circ_{\text{cell}} = 0.00 \text{ V} - (-0.28 \text{ V})$$

$$E^\circ_{\text{cell}} = \mathbf{+0.28 \text{ V}}$$

The reaction is spontaneous because  $E^\circ_{\text{cell}}$  is positive.



First, determine which elements are being reduced and oxidized. Since Zn is becoming more positive as a cation, it is losing electrons and is therefore oxidized; Zn would be the anode.  $\text{Cu}^{+2}$  is reduced to become Cu; Cu would be the cathode.

The standard reduction potentials needed are:

$\text{Cu}^{+2} (\text{aq}) + 2\text{e}^- \rightarrow \text{Cu (s)}$	+0.34 V
$\text{Zn}^{+2} (\text{aq}) + 2\text{e}^- \rightarrow \text{Zn (s)}$	-0.76 V

Now, we may use this equation:

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, cathode}} - E^\circ_{\text{reduction, anode}}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction, Cu}} - E^\circ_{\text{reduction, Zn}}$$

$$E^\circ_{\text{cell}} = 0.34 \text{ V} - (-0.76 \text{ V})$$

$$E^\circ_{\text{cell}} = \mathbf{1.10 \text{ V}}$$

The reaction is spontaneous because  $E^\circ_{\text{cell}}$  is positive.

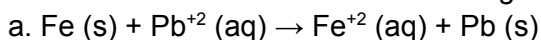
3. If a cell is at the conditions of 237K, 1.5 atm, and 2.0M for its solutions, what equation should be used to determine the  $E_{\text{cell}}$ ?

Since these conditions are not standard conditions (298K, 1.0 atm, 1.0 M solutions), the Nernst equation should be used. This equation is  $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} (\ln Q)$  (you may also use

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592 \text{ V}}{n} (\log Q) \text{ if the temperature is at } 25^{\circ}\text{C}$$

$Q$  refers to the reaction quotient,  $F$  refers to the Faraday constant,  $T$  refers to temperature in Kelvin,  $n$  refers to the number of electrons transferred in the reaction, and  $E^{\circ}$  refers to the cell potential under standard conditions.

4. Determine the value of  $\Delta G^{\circ}$  for the following reaction. (Hint: find  $E^{\circ}_{\text{cell}}$  first.)



First, let's find  $E^{\circ}_{\text{cell}}$  of this reaction. Since Fe is becoming more positive as a cation, it is losing electrons and is therefore oxidized; Fe would be the anode.  $\text{Pb}^{2+}$  is reduced to become Pb; Pb would be the cathode.

The standard reduction potentials needed are:

$\text{Fe}^{2+} \text{ (aq)} + 2\text{e}^{-} \rightarrow \text{Fe (s)}$	-0.44 V
$\text{Pb}^{2+} \text{ (aq)} + 2\text{e}^{-} \rightarrow \text{Pb (s)}$	-0.13 V

Now, we may use this equation:

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{reduction, cathode}} - E^{\circ}_{\text{reduction, anode}}$$

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{reduction, Pb}} - E^{\circ}_{\text{reduction, Fe}}$$

$$E^{\circ}_{\text{cell}} = -0.13 \text{ V} - (-0.44 \text{ V})$$

$$E^{\circ}_{\text{cell}} = 0.31 \text{ V}$$

Now, we may use the equation to find Gibbs free energy.  $n$  will be 2, because there are 2 electrons transferred Fe and  $\text{Pb}^{2+}$ .

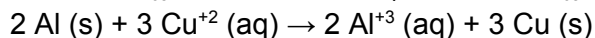
$$\Delta G^{\circ} = -nFE^{\circ}$$

$$\Delta G^{\circ} = -(2 \text{ mole e}^{-})(96485 \text{ C/mol e}^{-})(0.31 \text{ V})$$

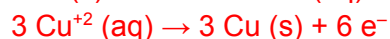
$$\Delta G^{\circ} = -59820.6 \text{ J}$$

$$\Delta G^{\circ} = -6.0 \times 10^4 \text{ J}$$

5. Determine the  $E_{\text{cell}}$  of the reaction. ( $T = 298\text{K}$ ,  $E^{\circ}_{\text{cell}} = 2.00 \text{ V}$ ,  $[\text{Al}^{3+}] = 0.1\text{M}$ ,  $[\text{Cu}^{2+}] = 2.5\text{M}$ )



Since some conditions of this reaction are not at standard conditions, we must use the Nernst equation. First, we must identify how many electrons are being transferred ( $n$ ) by splitting the reaction into half-reactions and balancing if needed.



## Unit 16: Electrochemistry



There are six electrons being transferred in the balanced equation. Now, we must find Q, the reaction quotient.

$$Q = (\text{Product ion A})^{\text{coefficient of A}} / (\text{Reactant ion B})^{\text{coefficient of B}}$$

$$Q = [\text{Al}^{+3}]^2 / [\text{Cu}^{+2}]^3$$

$$Q = (0.1\text{M})^2 / (2.5\text{M})^3$$

$$Q = 0.000064$$

Now, we can plug everything into the modified Nernst equation for 298K, since that is one of our conditions.

$$E = E^{\circ} - \frac{0.0592 \text{ V}}{n} (\log Q)$$

$$E = (2.00 \text{ V}) - \frac{0.0592 \text{ V}}{6} (\log 0.000064)$$

$$\mathbf{E = 2.0413 \text{ V}}$$