

Chapter 10: Electrochemistry

Elon Musk, an innovator in the field of harnessing renewable sources to generate electric power see a huge potential for electric cars to change the way Americans drive.

Selling an electric sports car creates an opportunity to fundamentally change the way America drives.

- Elon Musk

I've actually made a prediction that within 30 years a majority of new cars made in the United States will be electric. And I don't mean hybrid, I mean fully electric.

- Elon Musk

Given the importance of energy production (and in particular, production from renewable sources) alluded to by Richard Smalley in his address to the United States Congress (see Chapter 1), Elon Musk's vision seems well-aligned with Smalley's priority.

The generation and consumption of electrical energy and how it is harnessed to do work in the universe lends itself very nicely to discussion within the framework of thermodynamics. In this chapter, we will use some of the tools we have developed to relate electrochemical processes to thermodynamic variables, and to frame discussions of a few important topics.

Electricity

Electricity has been known for some time. Ancient Egyptians, for example, referred to electric fish in the Nile River as early as 2750 BC (Moller & Kramer, 1991). In 1600, William Gilbert studied what would later be seen to be electrostatic attraction, by creating static charges rubbing amber (Stewart, 2001). And Benjamin Franklin's famous experiment (although it is actually uncertain if he performed the experiment) of attaching a metal key to a kite string occurred in 1752, and showed that lightening is an electrical phenomenon (Uman, 1987).

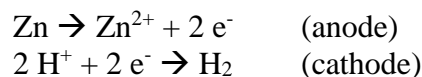
One of the biggest breakthroughs in the study of electricity as a chemical phenomenon was made by Alessandro Volta, who in 1799 showed that electricity could be generated by stacking copper and zinc disks submerged in sulfuric acid (Routledge, 1881). The reactions that Volta produced in his **voltaic**

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Figure 1. Alessandro Volta (1745-1827)

pile included both **oxidation** and **reduction** processes that could be considered as **half-reactions**. The half-reactions can be classified as **oxidation** (the loss of electrons) which happens at the **anode** and **reduction** (the gain of electrons) which occurs at the **cathode**. Those half reactions were



The propensity of zinc to oxidize coupled with that of hydrogen to reduce creates a potential energy difference between the electrodes at which these processes occur. And like any potential energy difference, it can create a force which can be used to do work. In this case, the work is that of pushing electrons through a circuit. The work of such a process can be calculated by integrating

$$dw_e = -E dQ$$

where E is the potential energy difference, and dQ is an infinitesimal amount of charge carried through the circuit. The infinitesimal amount of charge carried through the circuit can be expressed as

$$dQ = e dN$$

where e is the charge carried on one electron (1.6×10^{-19} C) and dN is the infinitesimal change in the number of electrons. Thus, if the potential energy difference is constant

$$w_e = -e E \int_0^N dN = -N e E$$

But since the number of electrons carried through a circuit is an enormous number, it would be far more convenient to express this in terms of the number of moles of electrons carried through the circuit. Noting that the number of moles (n) is given by

$$n = \frac{N}{N_A}$$

and that the charge carried by one mole of electrons is given by

$$F = N_A e = 96484 \text{ C}$$

where F has the magnitude of one Faraday (or the total charge carried by one mole of electrons.) The Faraday is named after Michael Faraday (1791-1867) (Doc, 2014), a British physicist who is credited with inventing the electric motor, among other accomplishments.

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Putting the pieces together, the total electrical work accomplished by pushing n moles of electrons through a circuit with a potential difference E , is

$$w_e = -nFE$$

Example: A current of 2.00 ampere is passed through an electroplating cell containing an initial concentration of silver of $[\text{Ag}^+] = 1.00 \text{ M}$ to electroplate silver metal onto a spoon. If the current is passed through the cell for a time period of 15 min, what mass of silver will be deposited onto the spoon?

Solution: Since $1 \text{ ampere} \times 1 \text{ second} = 1 \text{ coulomb}$, and $96484 \text{ coulomb} =$ the charge carried by 1 mole of electrons

$$(2.00 \text{ amp} \cdot 900 \text{ sec}) \cdot \frac{1 \text{ C}}{1 \text{ amp} \cdot \text{sec}} \cdot \frac{\text{mole } e^-}{96484 \text{ C}} \cdot \frac{1 \text{ mol Ag}}{1 \text{ mole } e^-} \cdot \frac{107.9 \text{ g}}{\text{mol Ag}} = 2.01 \text{ g}$$

The connection to ΔG

Recall that in addition to being used as a criterion for spontaneity, ΔG also indicated the maximum amount of non p-V work a system could produce at constant temperature and pressure. And since w_e is non p-V work, it seems like a natural fit that

$$\Delta G = -nFE$$

If all of the reactants and products in the electrochemical cell are in their standard states, it follows that

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

where E° is the **standard cell potential**. Noting that the molar Gibbs function change can be expressed in terms of the reaction quotient Q by

$$\Delta G = \Delta G^\circ + RT \ln Q$$

it follows that

$$-nFE = -nFE^\circ + RT \ln Q$$

Dividing by $-nF$ yields

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

which is the **Nernst equation**. This relationship allows one to calculate the cell potential of an electrochemical cell as a function of the specific activities of the reactants and products. In the Nernst equation, n is the number of electrons transferred per reaction equivalent. For the specific reaction harnessed by Volta in his original battery, $E^\circ = 0.763 \text{ V}$ (at 25°C) and $n = 2$. So if the Zn^{2+} and H^+ ions are at a concentration that gives them unit activity, and the H_2 gas is at a partial pressure that gives it unit fugacity,

$$E = 0.763 \text{ V} - \frac{RT}{nF} \ln(1) = 0.763 \text{ V}$$

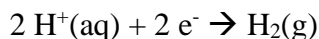
Half Cells and Standard Reduction Potentials



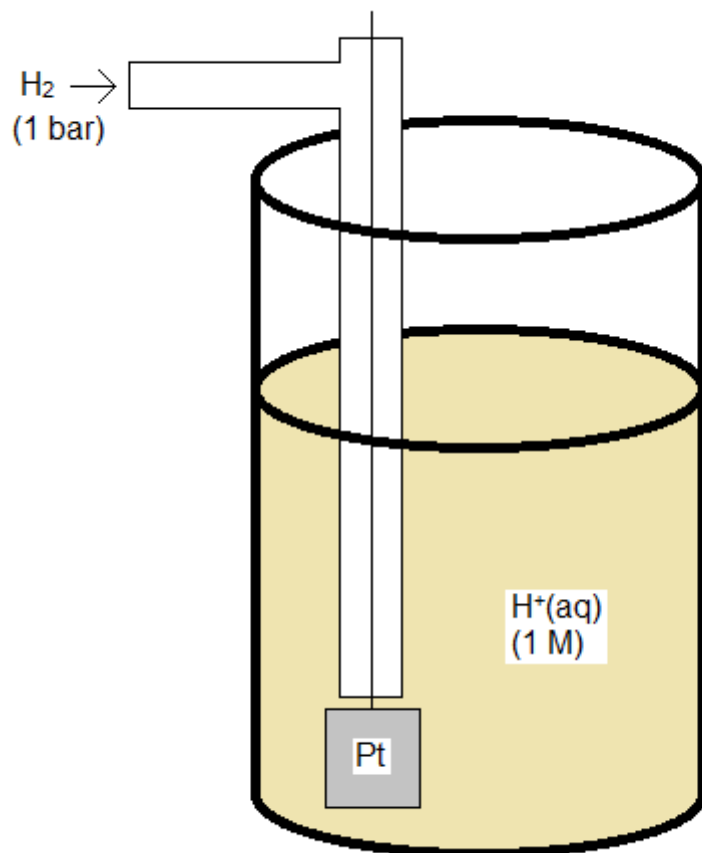
Much like G itself, E can only be measured as a difference, so a convention is used to set a zero to the scale. Toward this end, convention sets the reduction potential of the standard hydrogen electrode (SHE) to 0.00 V .

Standard Hydrogen Electrode

The standard hydrogen electrode is constructed so that H_2 gas flows over an inert electrode made of platinum, and can interact with an acid solution which provides H^+ for the half reaction



Both H^+ and H_2 need to have unit activity (or fugacity), which if the solution and gas behave ideally means a concentration of 1 M and a pressure of 1 bar .



Electrochemical Cells

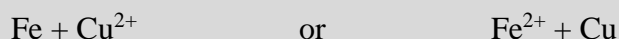
Standard reduction potentials can be measured relative to the convention of setting the reduction potential of the **Standard Hydrogen Electrode** (SHE) to zero. A number of values are shown in the table below. Several more standard reduction potentials are available in (Bratsch, 1989).

Reduction Half Reaction	E° (V)
$\text{F}_2 + 2 \text{e}^- \rightarrow 2 \text{F}^-$	2.87
$\text{MnO}_4^- + 8 \text{H}^+ \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}$	1.51
$\text{Au}^{3+} + 3 \text{e}^- \rightarrow \text{Au}$	1.50
$\text{Cl}_2 + 2 \text{e}^- \rightarrow 2 \text{Cl}^-$	1.360
$\text{Tl}^{3+} + 3 \text{e}^- \rightarrow \text{Tl}$	1.211
$\text{Br}_2 + 2 \text{e}^- \rightarrow 2 \text{Br}^-$	1.065
$\text{Hg}^{2+} + 2 \text{e}^- \rightarrow \text{Hg}$	0.854
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.799
$\text{Hg}_2^{2+} + 2 \text{e}^- \rightarrow 2 \text{Hg}$	0.790
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.771
$\text{Hg}_2\text{SO}_4 + 2 \text{e}^- \rightarrow 2 \text{Hg} + \text{SO}_4^{2-}$	0.615

$\text{I}_2 + 2 \text{e}^- \rightarrow 2 \text{I}^-$	0.536
$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	0.521
$\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu}$	0.337
$\text{HgCl}_2 + 2 \text{e}^- \rightarrow \text{Hg} + 2 \text{Cl}^-$	0.268
$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	0.222
$\text{Sn}^{4+} + 2 \text{e}^- \rightarrow \text{Sn}^{2+}$	0.15
$\text{AgBr} + \text{e}^- \rightarrow \text{Ag} + \text{Br}^-$	0.071
$2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2$	0.000
$\text{Pb}^{2+} + 2 \text{e}^- \rightarrow \text{Pb}$	-0.126
$\text{Sn}^{2+} + 2 \text{e}^- \rightarrow \text{Sn}$	-0.136
$\text{AgI} + \text{e}^- \rightarrow \text{Ag} + \text{I}^-$	-0.152
$\text{CuI} + \text{e}^- \rightarrow \text{Cu} + \text{I}^-$	-0.185
$\text{Ni}^{2+} + 2 \text{e}^- \rightarrow \text{Ni}$	-0.250
$\text{Tl}^+ + \text{e}^- \rightarrow \text{Tl}$	-0.34
$\text{PbSO}_4 + 2 \text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.355
$\text{Fe}^{2+} + 2 \text{e}^- \rightarrow \text{Fe}$	-0.440
$\text{Zn}^{2+} + 2 \text{e}^- \rightarrow \text{Zn}$	-0.763
$\text{Sn(OH)}_6^{2-} + 2 \text{e}^- \rightarrow \text{HSnO}_2 + 3 \text{OH}^- + \text{H}_2\text{O}$	-0.90
$\text{Al}^{3+} + 3 \text{e}^- \rightarrow \text{Al}$	-1.66
$\text{Mg}^{2+} + 2 \text{e}^- \rightarrow \text{Mg}$	-2.37
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.714
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.045

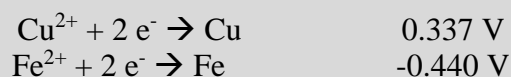
Example:

Which pair of reactants will produce a spontaneous reaction if everything is present in its standard state at 25 °C?

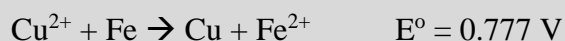
**Solution:**

The species with the higher standard reduction potential will force the other to oxidize.

From the table,

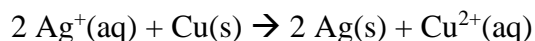


So the iron half-reaction will flip (so that iron is oxidizing) and the spontaneous reaction under standard conditions will be

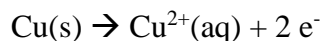


Calculating Cell Potentials

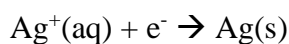
Using values measured relative to the SHE, it is fairly easy to calculate the standard cell potential of a given reaction. For example, consider the reaction



Before calculating the cell potential, we should review a few definitions. The anode half reaction, which is defined by the half-reaction in which oxidation occurs, is



And the cathode half-reaction, defined as the half-reaction in which reduction takes place, is



Using **standard cell notation**, the conditions (such as the concentrations of the ions in solution) can be represented. In the standard cell notation, the anode is on the left-hand side, and the cathode on the right. The two are typically separated by a **salt bridge**, which is designated by a double vertical line. A single vertical line indicates a phase boundary. Hence for the reaction above, if the silver ions are at a concentration of 0.500 M, and the copper (II) ions are at a concentration of 0.100 M, the standard cell notation would be

Example:

Calculate the cell potential at 25 °C for the cell indicated by



Solution:

In order to calculate the cell potential (E), the standard cell potential must first be obtained. The standard cell potential at 25 °C is given by

$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ &= 0.799 \text{ V} - 0.337 \text{ V} \\ &= 0.462 \text{ V} \end{aligned}$$

And for a cell at non-standard conditions, such as those indicated above, the Nernst equation can be used to calculate the cell potential. At 25 °C, The cell potential is given by

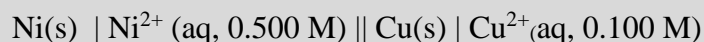
$$\begin{aligned} E &= E^{\circ} - \frac{RT}{nF} \ln \left(\frac{[\text{Cu}^{2+}]}{[\text{Ag}^+]^2} \right) \\ E &= 0.462 \text{ V} - \frac{(8.314 \frac{\text{J}}{\text{mol K}})(298 \text{ K})}{2(96484 \frac{\text{C}}{\text{mol}})} \ln \left(\frac{0.100 \text{ M}}{(0.500 \text{ M})^2} \right) \end{aligned}$$

Noting that $1 \text{ J/C} = 1 \text{ V}$,

$$E = 0.483 \text{ V}$$

Example:

Calculate the cell potential at 25°C for the cell defined by

**Solution:**

We will use the Nernst equation. First, we need to determine E° . Using the table below, it is apparent that



So copper, having the larger reduction potential will be the cathode half-reaction while forcing nickel to oxidize, making it the anode. So E° for the cell will be given by

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

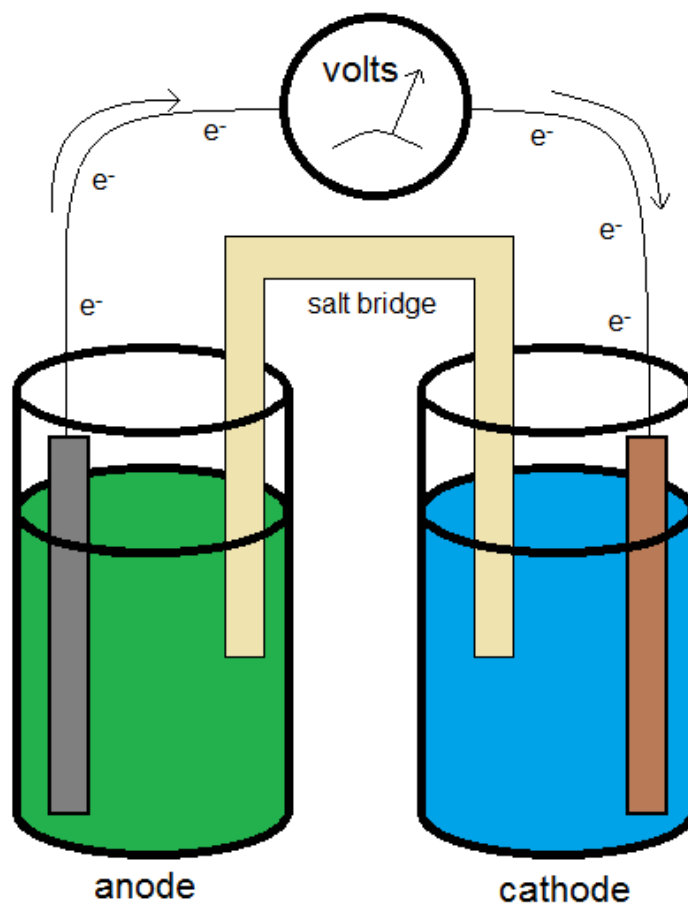
$$E_{\text{cell}}^\circ = 0.337 \text{ V} - (-0.250 \text{ V}) = 0.587 \text{ V}$$

And the cell potential is then given by

$$E_{\text{cell}} = E^\circ - \frac{RT}{nF} \ln Q$$

$$E_{\text{cell}} = 0.587 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol K}}\right)(298 \text{ K})}{2(96484 \text{ C})} \ln \left(\frac{0.500 \text{ M}}{0.100 \text{ M}}\right) = 0.566 \text{ V}$$

Measuring the Voltage



A typical galvanic electrochemical cell can be constructed similar to what is shown in the diagram above. The electrons flow from the anode (the electron source) to the cathode (the electron sink.) The salt bridge allows for the flow of ions to complete the circuit while minimizing the introduction of a junction potential.

Entropy

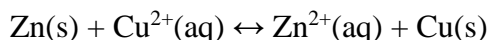
The Gibbs function is related to entropy through its temperature dependence

$$\left(\frac{\partial \Delta G}{\partial T}\right)_p = -\Delta S$$

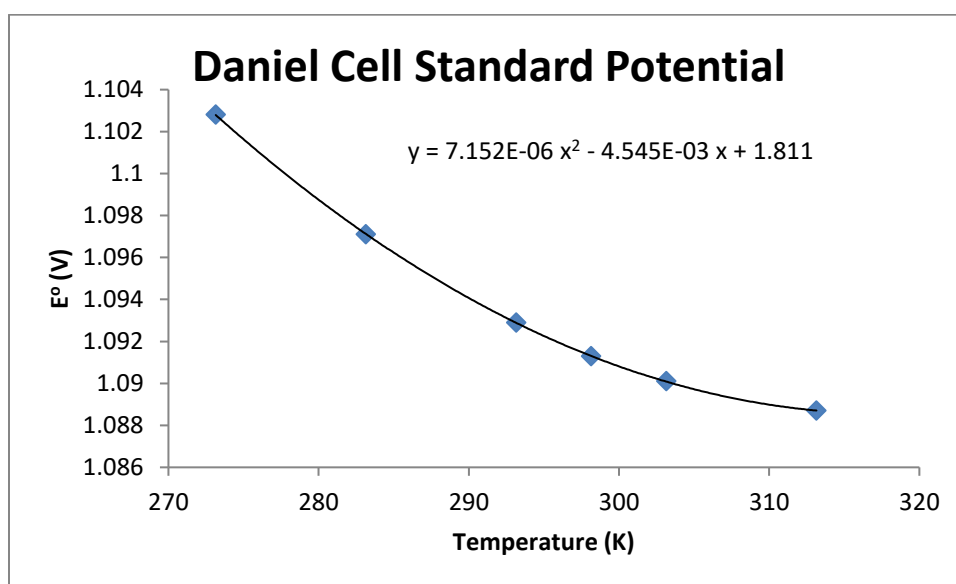
A similar relationship can be derived for the temperature variance of E° .

$$nF\left(\frac{\partial E^\circ}{\partial T}\right)_p = \Delta S$$

Consider the following data for the Daniel Cell (Buckbee, Surdzial, & Metz, 1969) which is defined by the following reaction



T (°C)	E° (V)
0	1.1028
10	1.0971
20	1.0929
25	1.0913
30	1.0901
40	1.0887



From a fit of the data to a quadratic function, the temperature dependence of E° over the measured temperature range is easily established.

$$E^\circ = \left(7.152 \cdot 10^{-6} \frac{\text{V}}{\text{K}^2}\right) T^2 - \left(4.545 \cdot 10^{-3} \frac{\text{V}}{\text{K}}\right) T + 1.811 \text{ V}$$

$$\left(\frac{\partial E^\circ}{\partial T}\right)_p = 3.576 \cdot 10^{-6} \frac{\text{V}}{\text{K}^2}(T) - 4.545 \cdot 10^{-3} \frac{\text{V}}{\text{K}}$$

So, at 298 K,

$$\left(\frac{\partial E^\circ}{\partial T}\right)_p = -3.479 \cdot 10^{-3} \frac{\text{V}}{\text{K}}$$

So the entropy change is calculated by

$$\Delta S^\circ = nF \left(\frac{\partial E^\circ}{\partial T} \right)_p = (2 \text{ mol}) \left(96484 \frac{\text{C}}{\text{mol}} \right) \left(-3.479 \cdot 10^{-3} \frac{\text{V}}{\text{K}} \right) \left(\frac{\text{C} \cdot \text{V}}{\text{J}} \right)$$

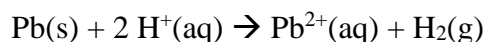
The standard entropy change for the Daniel cell reaction at 25 °C is

$$\Delta S^\circ = -104.5 \frac{\text{J}}{\text{mol K}}.$$

The standard entropies of formation for the species involved in the Daniel cell are

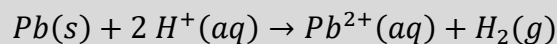
Species	S_f° (J mol ⁻¹ K ⁻¹)
Zn(s)	33.150
Zn ²⁺ (aq)	-112.1
Cu(s)	41.63
Cu ²⁺ (aq)	-99.6

It is the negative entropy change that leads to an increase in standard cell potential at lower temperatures. For a reaction such as

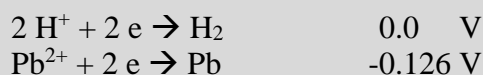


which has a large increase in entropy (due to the production of a gas-phase product), the standard cell potential decreases with decreasing temperature. As this is the reaction used in most car batteries, it explains why it can be difficult to start ones car on a very cold winter morning. The topic of temperature dependence of several standard cell potentials is reported and discussed by Bratsch (Bratsch, 1989).

Example: Consider the redox reaction which is used to generate electricity in most car batteries:



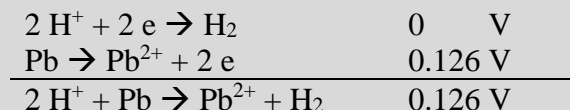
Given the following data



Species	S_f° (J mol ⁻¹ K ⁻¹)
H ⁺ (aq)	0
H ₂ (g)	130.684
Pb ²⁺ (aq)	10.5
Pb(s)	68.81

- Calculate the standard cell potential (E°) at 25 °C,
- the standard reaction entropy ($\Delta S_{\text{rxn}}^\circ$) at 25 °C,
- and the standard cell potential at 0 °C.

Solution: The target reaction can be generated by adding the data reactions (with the lead half-reaction flipped):



The entropy change is given by

$$\begin{aligned} \Delta S_{\text{rxn}}^\circ &= (1 \text{ mol}) \left(10.5 \frac{\text{J}}{\text{mol K}} \right) + (1 \text{ mol}) \left(130.684 \frac{\text{J}}{\text{mol K}} \right) - (2 \text{ mol})(0) \\ &\quad - (1 \text{ mol}) \left(68.81 \frac{\text{J}}{\text{mol K}} \right) \\ \Delta S_{\text{rxn}}^\circ &= 72.374 \frac{\text{J}}{\text{K}} \end{aligned}$$

Since

$$nF \left(\frac{\partial E^\circ}{\partial T} \right)_p = \Delta S^\circ$$

It follows that at constant pressure

$$\int_{T_1}^{T_2} dE^\circ = \frac{1}{nF} \int_{T_1}^{T_2} \Delta S^\circ dT$$

So, if the entropy change is constant over the temperature interval, the change to the standard cell potential is given by

$$\Delta E^\circ = \left(\frac{\Delta S^\circ}{nF} \right) \Delta T$$

So

$$E^\circ(T_2) - E^\circ(T_1) = \left(\frac{\Delta S^\circ}{nF} \right) (T_2 - T_1)$$

And

$$E^{\circ}(273\text{ K}) = 0.126\text{ V} + \left[\frac{72.374 \frac{\text{J}}{\text{K}}}{2 \cdot 96484 \frac{\text{C}}{\text{mol}}} \left(\frac{\text{C} \cdot \text{V}}{\text{J}} \right) \right] (-25\text{ K})$$

$$E^{\circ}(273\text{ K}) = 0.117\text{ V}$$

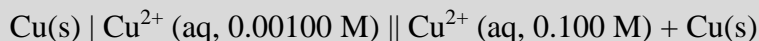
Concentration Cells

The generation of an electrostatic potential difference is dependent on the creation of a difference in chemical potential between two half-cells. One important manner in which this can be created is by creating a concentration difference. Using the Nernst equation, the potential difference for a concentration cell (one in which both half-cells involve the same half-reaction) can be expressed

$$E = -\frac{RT}{nF} \ln \frac{[\text{oxidizing}]}{[\text{reducing}]}$$

Example:

Calculate the cell potential (at 25 °C) for the concentration cell defined by



Solution:

Since the oxidation and reduction half-reactions are the same,

$$E^{\circ} = 0$$

The cell potential at 25 °C is calculated using the Nernst equation:

$$E = E^{\circ} - \frac{RT}{nF} \ln Q$$

Substituting the values from the problem:

$$\begin{aligned} E &= 0 - \frac{(8.314 \frac{\text{J}}{\text{mol K}})(298\text{ K})}{(2)(96484 \frac{\text{C}}{\text{mol}})} \ln \left(\frac{[0.00100\text{ M}]}{[0.100\text{ M}]} \right) \\ &= 0.059\text{ V} \end{aligned}$$

Applications to Equilibrium

Since the standard cell potential, E° , is directly related to the standard Gibbs function change

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

It follows that the standard cell potential will also be related to an equilibrium constant describing a redox reaction. The relationship between ΔG° and K is given by

$$\Delta G^\circ = -RT \ln K$$

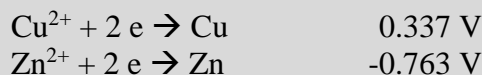
So, it follows that

$$E^\circ = \frac{RT}{nF} \ln K$$

At 298 K, this can be transformed similarly to the Nernst equation

$$E^\circ = \frac{0.0592 \text{ V}}{n} \log K$$

Example: Based in the standard reduction potentials



Calculate the value of the equilibrium constant for the reaction



Solution: The standard cell potential, E° , for the reaction is

$$E^\circ = (0.337 \text{ V}) - (-0.763 \text{ V}) = 1.100 \text{ V}$$

The equilibrium constant can be calculated from the relationship

$$E^\circ = \frac{RT}{nF} \ln K$$

So

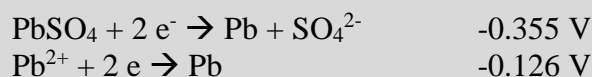
$$1.100 \text{ V} = \left[\frac{\left(8.314 \frac{\text{J}}{\text{mol K}}\right) (298 \text{ K})}{(2 \text{ mol}) \left(96484 \frac{\text{C}}{\text{mol}}\right)} \left(\frac{\text{V} \cdot \text{C}}{\text{J}}\right) \right] \ln K$$

$$K = 2.97$$

Solubility

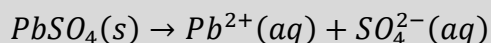
Electrochemical measurements can be used to calculate solubility products (K_{sp}) based on measurements where one of the solid electrodes is the salt of the compound for which the K_{sp} value is desired.

Example: Based on the following data

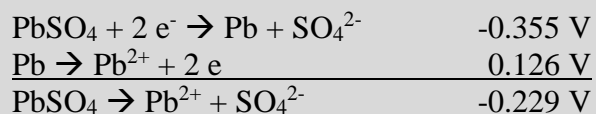


calculate K_{sp} for $\text{PbSO}_4(\text{s})$ at 25°C .

Solution: The target reaction is that for the dissociation of $\text{AgBr}(\text{s})$:



This can be generated from the half reactions by flipping the second half reaction and adding.



And since

$$E^\circ = \frac{RT}{nF} \ln K$$

We see that

$$\begin{aligned} -0.229 \text{ V} &= \left[\left(\frac{(8.314 \frac{\text{J}}{\text{mol K}})(298 \text{ K})}{(2 \text{ mol})(96484 \frac{\text{C}}{\text{mol}})} \right) \left(\frac{\text{V} \cdot \text{C}}{\text{J}} \right) \right] \ln K \\ K &= 1.79 \cdot 10^{-8} \text{ M}^2 \end{aligned}$$

References

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Vocabulary and Concepts

anode	232, 237	reduction	232
cathode	232, 237	salt bridge.....	237
concentration cell	243	standard cell notation	237
half-reactions.....	232	standard cell potential	233
Nernst equation	234	Standard Hydrogen Electrode	235
oxidation	232	voltaic pile.....	232

Learning Objectives

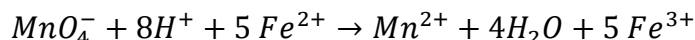
Upon successful completion of this chapter, a student will be able to

1. Articulate and utilize the connection between ΔG° and E° for oxidation-reduction reactions.
2. Define the terms *anode*, *cathode*, *half-cell*, *standard reduction potential*, and other terms found in the “Vocabulary and Concepts” section.
3. Connect the temperature dependence of standard cell potential to entropy changes.
4. Utilize the connections between standard cell potential and equilibrium constants to calculate equilibrium constants and concentrations from electrochemical data, and vice versa.

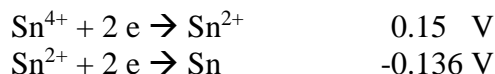
Problems

1. Silver is electroplated onto a spoon by suspending the spoon in a 1.00 M solution of AgNO_3 and passing a 2.00 amp current through it for 15.0 min. What mass of silver is deposited on the spoon?

2. Calculate the standard cell potential for the reaction



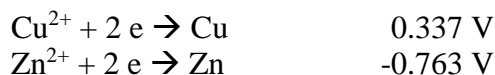
3. The following data can be found in a table of standard reduction potentials at 25 °C:



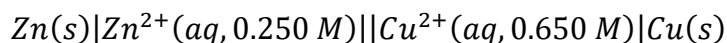
From these data, Calculate the standard reduction potential for the half-cell



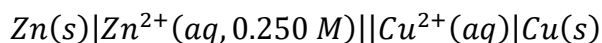
4. From a table of standard reduction potentials at 25 °C, the following is found:



From these data, calculate the cell potential for the cell constructed as

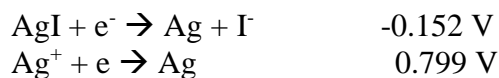


5. A Daniel cell is constructed



and the cell potential is measured to be 1.05 V. Calculate the concentration of Cu^{2+} ions in the cathode half-cell.

6. Given the following standard reduction potentials, calculate the value of K_{sp} for AgI.



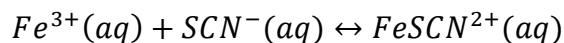
7. K_{sp} for $\text{PbCl}_2(\text{s})$ is $1.8 \times 10^{-4} \text{ M}^3$. Based on this, and the standard reduction potential



calculate the standard reduction potential for



8. At 25 °C, $K_{\text{f}} = 8.9 \times 10^2 \text{ M}^{-1}$ for the complexation reaction



Given that the standard reduction potential at 25 °C



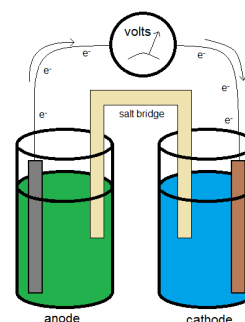
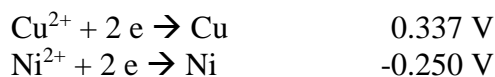
determine the standard reduction potential for the half-reaction



9. Ni^{2+} Ions complex with EDTA by the reaction

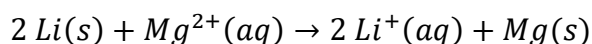


K_f for this complex formation equilibrium is $3.6 \times 10^{18} \text{ M}^{-1}$. A half-cell is prepared to be 0.100 M in Ni^{2+} and 1.00 M in EDTA. Its standard reduction potential is measured relative to a copper half-cell, in which $[Cu^{2+}] = 1.00 \text{ M}$. The standard reduction potentials for the bare-ion reductions are

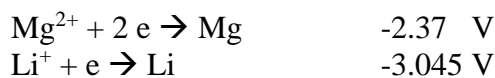


From these data, calculate the cell potential once the complexation reaction of Ni^{2+} with EDTA has reached equilibrium.

10. Consider the redox reaction



Given the following data



Species	$S_r^0 \text{ (J mol}^{-1} \text{ K}^{-1})$
$Mg^{2+}(aq)$	-138.1
$Mg(s)$	32.68
$Li^{+}(aq)$	13.4
$Li(s)$	29.12

- Calculate the standard cell potential (E^0) at 25 °C,
- the standard reaction entropy (ΔS_{rxn}^0) at 25 °C,
- and the standard cell potential at 37 °C.

