

ADVANCED PLACEMENT CHEMISTRY STUDY GUIDE/CHAPTER OUTLINE

17

Electrochemistry

I. The Voltaic Cell

A. How does it work?

It produces electrical energy by carrying out a spontaneous redox reaction.

B. Components

1. Cathode

a. Where the ion or molecule undergoes reduction (gain of e^-)

b. The ions that move toward the cathode are called cations

2. Anode

a. Where the ion or molecule undergoes oxidation (loss of e^-)

b. The ions that move toward the anode are called anions

c. Electrons are formed at the anode and move toward the cathode

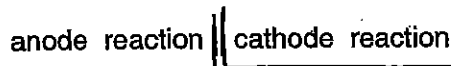
3. Salt bridge

a. Function: allows electrons to flow from cathode to anode w/out mixing solutions

b. Brief description: it keeps the net charge of components ZERO!

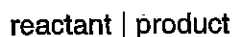
C. Notation

1. A voltaic cell can be abbreviated in the following manner:

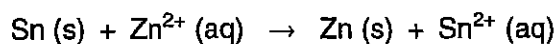


2. The salt bridge is indicated by the symbol \parallel

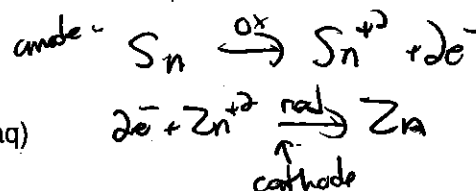
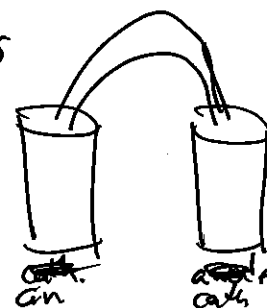
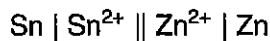
3. The anode and cathode reactions are further abbreviated



4. The voltaic cell



is thus abbreviated



5. If no metal is involved in either anode or cathode reaction, or another metal is used as an electrode, the electrode is separated from the reactant/product entry. If the electrode is used as the anode then the symbol for the metal is written before the reactant for the anode and separated by |. If the electrode is used as the cathode then the symbol for the metal is written after the product for the cathode and separated by |.

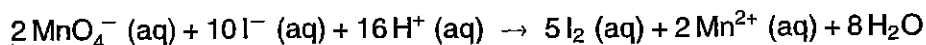
D. Drawing a diagram of the cell

Key points to remember:

1. There are always two separate half-cells connected by a wire and a salt bridge.
2. One of the compartments is the anode and the other is the cathode.
All species (reactants and products of a half-reaction) except water, are shown in the appropriate compartment.
3. If a metal precipitates in a cell half-reaction (either as product or reactant), it is ordinarily chosen as the electrode for the compartment where that half-reaction occurs.
If no metal is involved in the half-reaction, an electrically conducting non-reactive solid like platinum (Pt) or graphite (C) should be used.
4. Electron flow from anode to cathode is denoted by an arrow in the external circuit.
5. The flow of ions through the salt bridge is also indicated by an arrow. Cations move to the cathode, anions move to the anode.

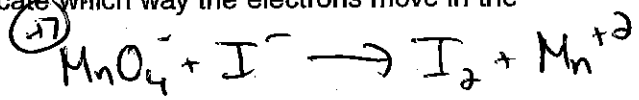
E. Exercises

1. When potassium permanganate and sodium iodide are mixed, a spontaneous reaction occurs:



This reaction can serve as a source of energy in a voltaic cell.

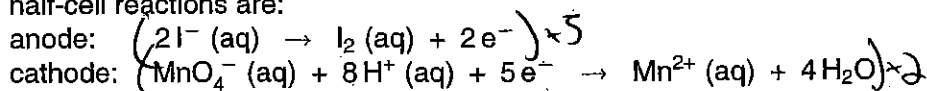
- Write the anode and cathode reactions.
- Draw a diagram of the cell. Indicate which way the electrons move in the external circuit.
- Write the cell notation.



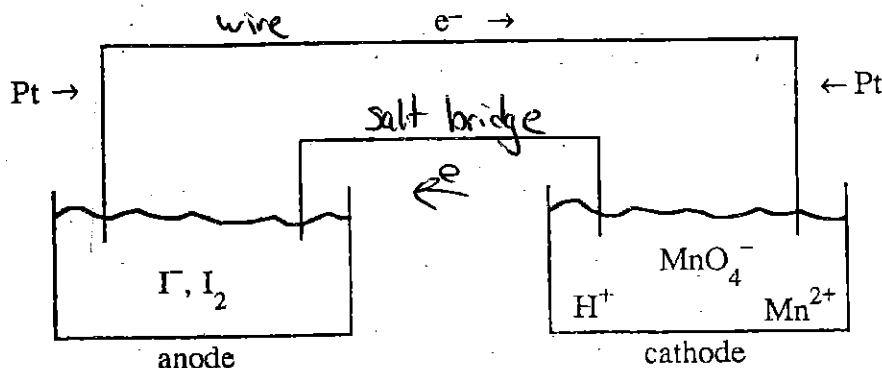
Solution:

- a. Since Mn goes from an oxidation number of 7 to 2, it is being reduced. I goes from an oxidation number of -1 to 0, so it is oxidized. Thus the half-cell reactions are:

CH 4 { BALANCING IN ACID SOLUTION }

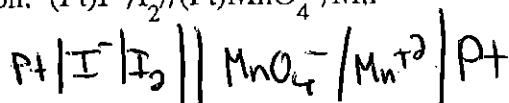


- b. The diagram of the cell is



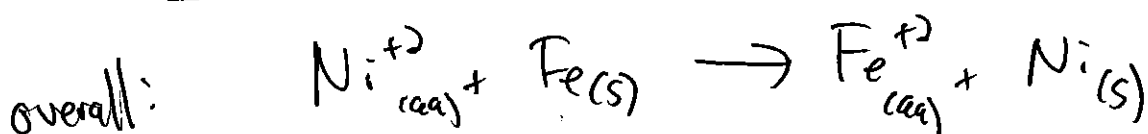
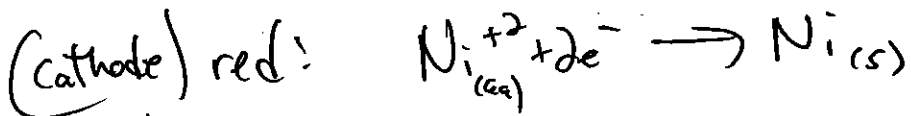
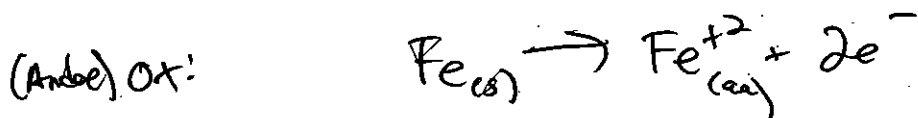
within the cell: anions move to anode; cations move to cathode

- c. Cell notation: $(\text{Pt})\text{I}^-/\text{I}_2//(\text{Pt})\text{MnO}_4^-/\text{Mn}^{2+}$

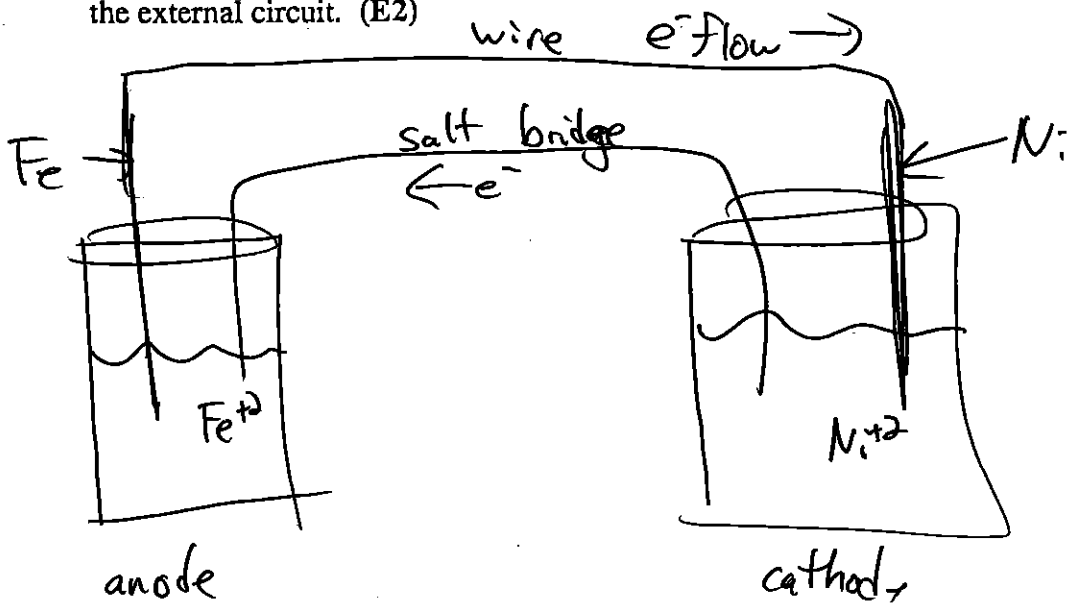


2. Given the following electrochemical cell notation:
 $\overset{\text{Anode}}{\text{Fe}}/\text{Fe}^{+2} \parallel \text{Ni}^{+2}/\overset{\text{Cathode}}{\text{Ni}}$

- a. Write the reaction that takes place at the anode; at the cathode.
 Write the overall reaction. (E1)



- b. Draw a diagram of the cell, indicating the anode, the cathode, all species in each half-cell and the direction of electron flow in the external circuit. (E2)



$$1V = 1 \text{ Joule} / \text{Coulomb}$$

II. Standard Voltages

A. Qualitative description

1. Definition: the reduction potential of a redox reaction

2. Symbol for

a. the standard voltage of a redox reaction: E_{tot}° E_{total}°

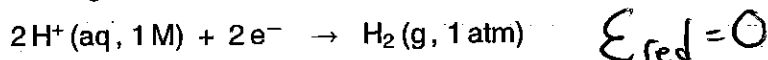
b. the standard voltage of an oxidation half-reaction: E_{ox}°

c. the standard voltage of a reduction half-reaction: E_{red}°

B. Quantitative aspects

1. $E^{\circ} = E_{\text{ox}}^{\circ} + E_{\text{red}}^{\circ}$ ✓

2. E_{red}° for the following reaction



is zero.

C. Standard potentials

1. The standard potentials listed in Table 18.1 of the text give you the standard voltages for reduction half-reactions.

Remember: These are half-reactions in which e^- are on the reactant side and there is a reduction in oxidation number.

2. Standard voltages for oxidation half-reactions are obtained by changing the *sign* of the reduction half-reaction. You can do this only if the oxidation half-reaction is the reverse of the reduction half-reaction as written in Table 18.1.

D. The strength of oxidizing and reducing agents

1. The larger E_{red}° is, the stronger the oxidizing agent is. Remember the oxidizing agent is the *reactant* in a reduction half-reaction.

2. The larger E_{ox}° is, the stronger the reducing agent is.

3. In order to compare the strength of reducing agents, you must:

a. Find the species on the right side of the standard potential table. If the ion or molecule is on the left side, it is not acting as a reducing agent.

b. Reverse the reactions, making sure to change E_{red}° to E_{ox}° and reverse the sign.

c. Compare the magnitude of E_{ox}° values.

4. In order to compare the strength of oxidizing agents, you must:

a. Find the species on the left-hand side of the standard potential table. If the ion or molecule is on the right-hand side, it is not acting as an oxidizing agent.

b. Compare the magnitude of E_{ox}° values.

5. The product formed by a strong oxidizing agent is itself a weak reducing agent.

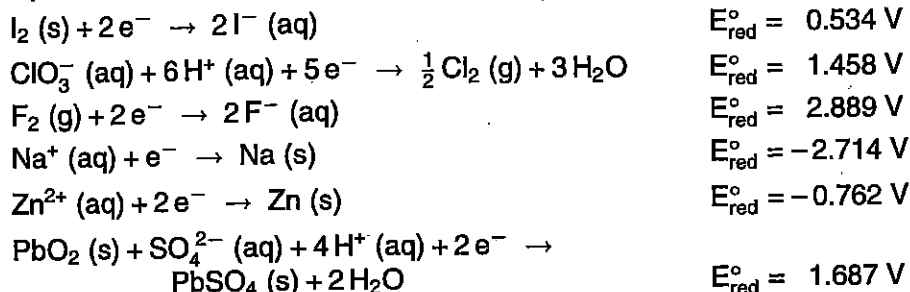
6. The product formed by a strong reducing agent is itself a weak oxidizing agent.

7. Exercises

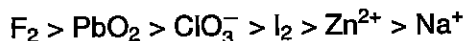
- a. Arrange, in decreasing order of strength, the following oxidizing agents in acidic solution.



Solution: Since we are ranking oxidizing agents, these are the species that are reduced. Hence, we should find them on the left side of the equations listed in Table 18.1. The equations are



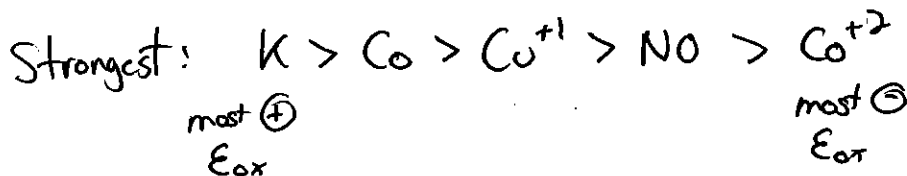
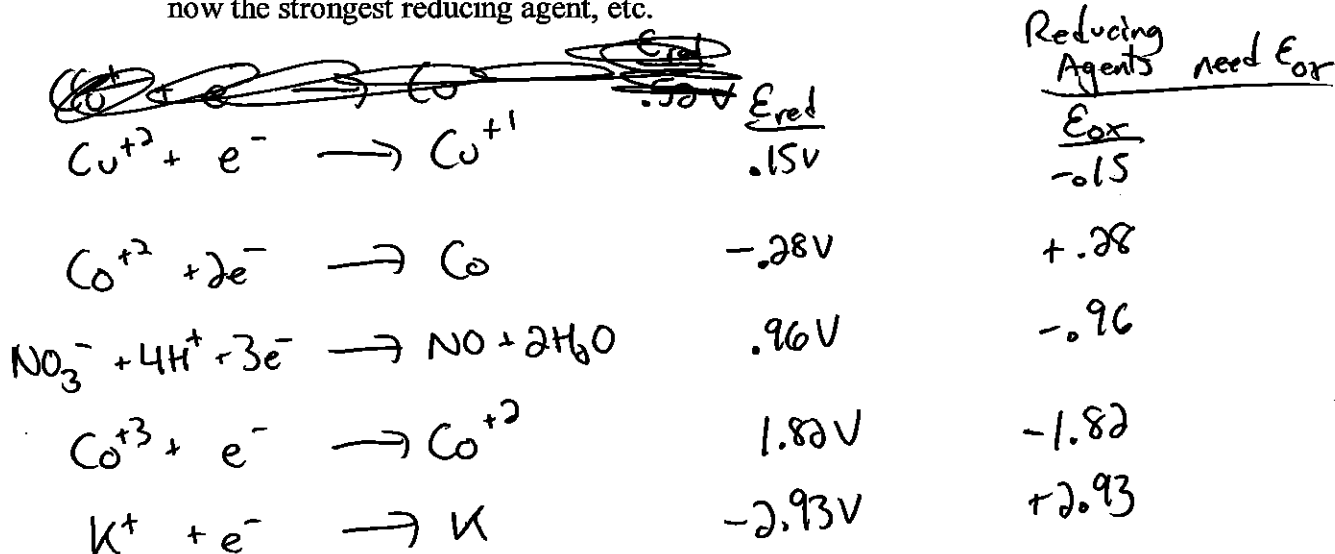
From this table we see that the largest E_{red}° is 2.889, which belongs to the reduction half-reaction of F_2 . Thus, F_2 is the strongest oxidizing agent. We go through the same reasoning for all other oxidizing agents and get the order



- b. Arrange in increasing order of strength the following reducing agents in acidic solution. (E3)



Solution: Since we are ranking reducing agents, these are the species that are oxidized. Hence, we should find them on the right side of table 18.1. The equations must be written in reverse and the signs of the E_{red}° must also be reversed. The largest E_{ox}° is now the strongest reducing agent, etc.

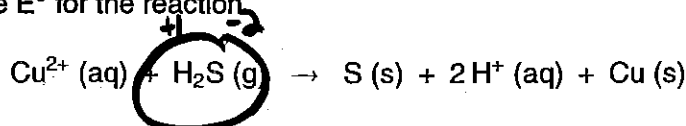


E. Calculation of E° from E°_{red} and E°_{ox}

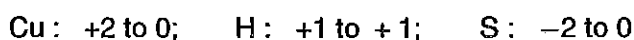
1. $E^\circ = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$
2. When given a redox equation, you have to break it up into two half-equations.
 - a. The reduction half-reaction has voltage E°_{red} . You get this half-reaction "as is" from Table 18.1.
 - b. The oxidation half-reaction has voltage E°_{ox} . You have to look for the reverse of that reaction in Table 18.1. When you have found it, write your oxidation half-reaction, and take the voltage of the reverse reaction. Do not forget to reverse the sign.

3. Exercises

- a. Calculate E° for the reaction



Solution: In order to decide how to split the reaction, we have to determine oxidation numbers.

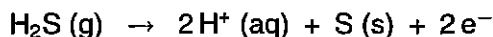


Thus, Cu is reduced and H_2S is oxidized.

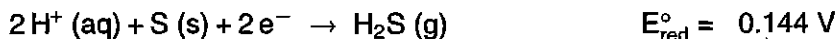
Reduction half-reaction:



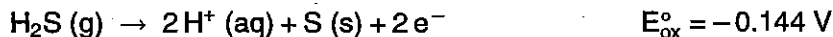
Oxidation half-reaction:



Looking this up in the table, we find the following equation



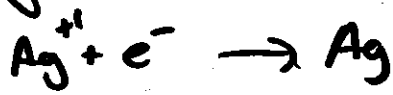
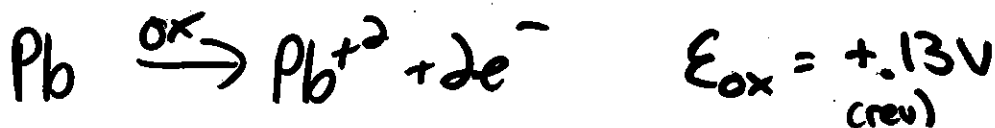
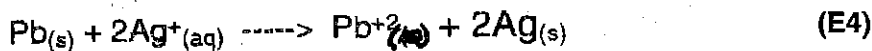
Since we want E°_{ox} for the reverse reaction, we write



Thus,

$$E^\circ = 0.339 + (-0.144) = 0.195 \text{ V}$$

- b. Calculate E° for the cell



#e = 2

$$\Delta G = -nFE$$

$$E_{\text{tot}} = E_{\text{red}} + E_{\text{ox}} = .93 \text{ V}$$

$$\Delta G = -(2)(96,485)(.93)$$

$$.13 + .8 = .93 \text{ V}$$

$$\Delta G = -179,462 \text{ kJ/mole} = -179 \text{ kJ/mole}$$

F. Spontaneity of redox reactions

1. If $E^\circ > 0$, the reaction is spontaneous.
2. If $E^\circ < 0$, the reaction is nonspontaneous; the reverse reaction will tend to occur.
3. If $E^\circ = 0$, the reaction is at equilibrium.

III: Relations between E° , ΔG° , and K

A. Indicators of spontaneity

1. E°
 - a. If $E^\circ > 0$, the reaction is spontaneous.
 - b. If $E^\circ < 0$, the reaction is nonspontaneous.
 - c. If $E^\circ = 0$, the reaction is at equilibrium.
2. ΔG°
 - a. If $\Delta G^\circ > 0$, the reaction is nonspontaneous.
 - b. If $\Delta G^\circ < 0$, the reaction is spontaneous.
 - c. If $\Delta G^\circ = 0$, the reaction is at equilibrium.
3. K
 - a. If $K > 1$, the reaction occurs spontaneously from left to right at standard conditions.
 - b. If $K < 1$, the reverse reaction occurs spontaneously.

$K=1$ equilibrium

B. Relation between E° and ΔG°

1. Mathematical relationship

$\Delta G = -nFE$

$\frac{\text{kJ}}{\text{mole}} = - \frac{\text{J}}{\text{mole}}$

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

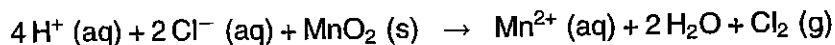
where n is the number of electrons transferred. F is the Faraday constant $-96,485$. Thus, in kJ

$$\Delta G^\circ = -96.5 n E^\circ$$

F

2. Exercises

- a. Determine ΔG° for the equation

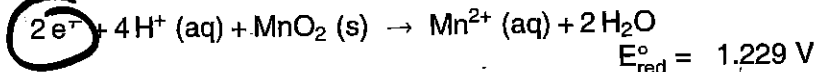


Solution:

- (1) We first determine E° by writing the two half-reactions and their standard potentials. The oxidation half-reaction is



The reduction half-reaction is



Calculating E° , we get

$$E^\circ = -1.360 + 1.229 = -0.131 \text{ V}$$

Since E° is negative, we know that the reaction is nonspontaneous and expect ΔG° to be positive.

- (2) We note that two electrons are transferred, thus $n = 2$.
- (3) To calculate ΔG° , we substitute into the equation

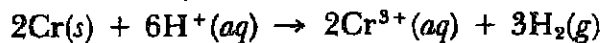
$$\Delta G^\circ = -96.5 n E^\circ$$

and get

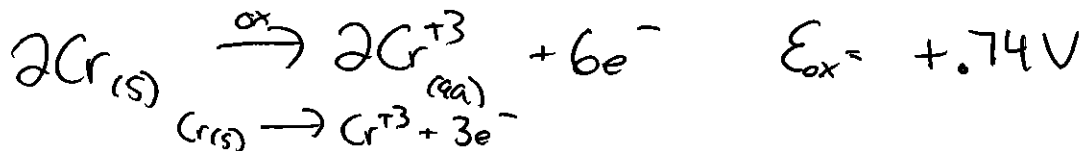
$$\Delta G^\circ = -96.5 (2) (-0.131) = 25.3 \text{ kJ/mole}$$

$\Delta G = (+)$ NON-SPONT.

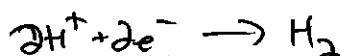
b. Consider a voltaic cell in which the following reaction takes place:



Calculate E_{tot}^0 , ΔG^0 (E5)



$$E_{\text{red}} = 0$$



$$E_{\text{tot}} = +.74\text{V}$$

$$\Delta G = -nFE^0$$

$$= -(6) \left(\frac{96,500 \text{ J/mole}}{\text{mole}} \right) (.74 \text{ J/K}) = -428,000 \text{ J}$$

$$\underline{\underline{-428 \text{ kJ}}}$$

C. Relation between E^0 and K

1. Mathematical relationship at 25°C

$$\ln K = \frac{nE^0}{0.0257}$$

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

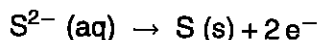
$$\frac{nFE}{RT} = \ln K$$

$$n \cdot E \left(\frac{F}{RT} \right)$$

2. Exercises

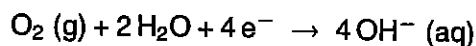
a. Write a balanced redox equation for the reaction between sulfide ion and oxygen gas to produce solid sulfur and hydroxide ion. Calculate K for the reaction at 25°C .

Solution: The oxidation half-reaction is



$$E_{\text{ox}}^0 = 0.445\text{V}$$

The reduction half-reaction is



$$E_{\text{red}}^0 = 0.401\text{V}$$

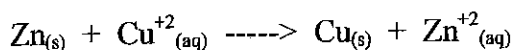
E^0 therefore is 0.846V . We multiply the oxidation half-reaction by 2 to get the same number of electrons lost and gained. Thus, $n = 4$. We next calculate K.

$$\ln K = \frac{4 \times 0.846}{0.0257} = 132$$

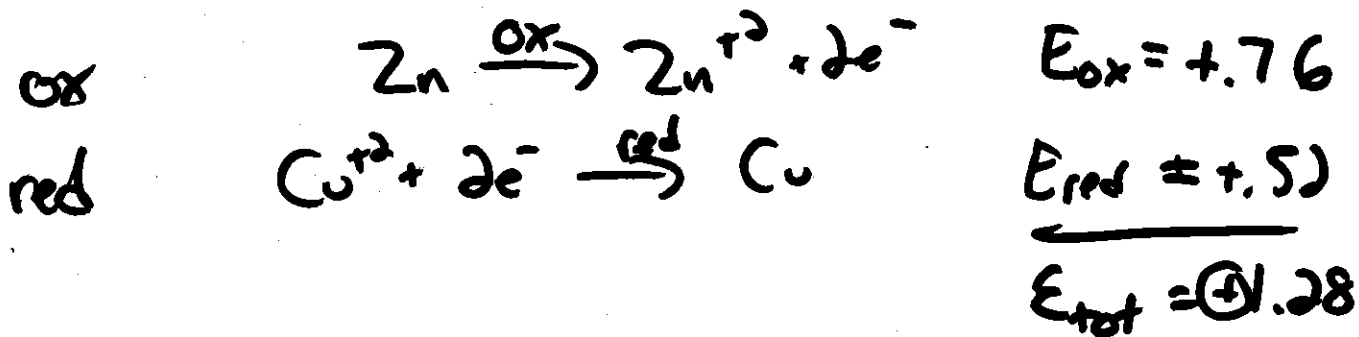
Taking the inverse ln of both sides, we obtain

$$K = \text{inv ln } 132 = 2 \times 10^{57}$$

b. Given the following reaction found in the Daniel type voltaic cell:



Compute: E°_{tot} and K (E6)



$$\ln K = \frac{nE}{.0257} = \frac{2(1.28)}{.0257} = 99.6$$

$$K = 1.82 \times 10^{43}$$

LARGE K
SPONTANEOUS

IV. Effect of Concentration Upon Voltage

A. Qualitative aspects of concentration change

1. A reaction becomes more spontaneous (voltage increases) if

- the concentration of the reactant increase
- the concentration of the product decreases

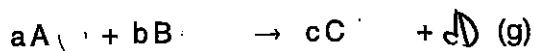
2. A reaction becomes less spontaneous (voltage decreases) if

- the concentration of the reactant decrease
- the concentration of the product increase

B. Nernst Equation

- Note that E_{red}° , E_{ox}° , and E° apply only at standard conditions, that is, 1 M for species in solution, and 1 atm for gases.
- Mathematical relationship between concentration, E° , and E , the voltage at a given concentration.

For the reaction



the voltage E , at given conditions, is calculated using the equation

$$E = E_{\text{tot}}^{\circ} - \frac{0.0257}{n} \ln \frac{(\text{conc. C})^c (\text{conc. D})^d}{(\text{conc. A})^a (\text{conc. B})^b}$$

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

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

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

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

where

E = voltage at given concentration

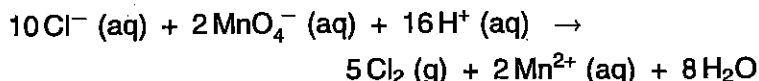
E° = standard voltage

n = number of moles of e^{-} transferred

- The condition for species in solution is expressed in molarity, while the condition for gases is expressed in partial pressure (in atm).
- Pure liquids and solids do not appear.

5. Exercises calculating E

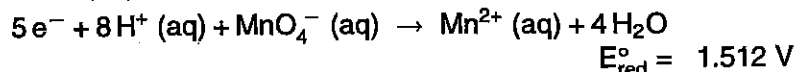
- Calculate the voltage of the cell at 25°C for the reaction



The partial pressure of chlorine gas is 1.00 atm, the pH is 0.50, and the concentrations of all the ionic species except H^{+} are 0.10 M.

Solution: We need to use the Nernst equation to calculate E , since the conditions are no longer the standard conditions of 1 M for ionic species and 1 atm for gases.

- First, we must calculate E° . We do that by writing the two half-reactions.

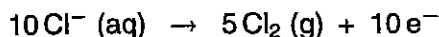


Note: To calculate E_{ox}° and E_{red}° , we do not pay any attention to the coefficients used to balance the species in the final reaction. Each half-reaction is balanced independently and looked up.

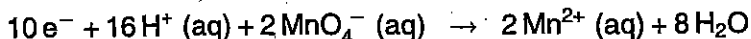
$$E^{\circ} = E_{\text{ox}}^{\circ} + E_{\text{red}}^{\circ}$$

$$= -1.360 + 1.512 = 0.152 \text{ V}$$

- Next, we determine n . To do so, we multiply both equations by integers to make the number of electrons gained and lost the same. In this case, we multiply the oxidation half-reaction by 5.



We multiply the reduction half-reaction by 2.



Notice that the coefficients of the species are now just like those given for the cell equation. Since there are 10 electrons gained as well as lost, $n = 10$.

It is important that you first find E° before you determine n , so that you will not be tempted to multiply E_{ox}° and E_{red}° by the integers you use to

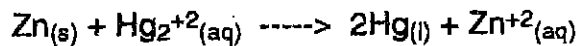
(10)

(3) Lastly, we determine the conditions for the ionic species in molar concentrations, and those for gases in atm pressure. All the ionic species except H^+ have a concentration of 0.10 M, that is, $[Cl^-] = [MnO_4^-] = [Mn^{2+}] = 0.10$ M. $[H^+]$ is expressed in terms of pH. Since the pH is 0.50, $[H^+] = 0.32$ M. Chlorine gas has pressure of 1.0 atm.

(4) Substituting, we get

$$\begin{aligned} E &= 0.152 - \frac{0.0257}{10} \ln \frac{(P_{Cl_2})^5 [Mn^{2+}]^2}{[Cl^-]^{10} [MnO_4^-]^2 [H^+]^{16}} \\ &= 0.152 - 0.00257 \ln \frac{(1.0)^5 (0.10)^2}{(0.10)^{10} (0.10)^2 (0.32)^{16}} \\ &= 0.152 - 0.00257 \ln (8.3 \times 10^{17}) \\ &= 0.046 \text{ V} \end{aligned}$$

b. Consider the following cell reaction:

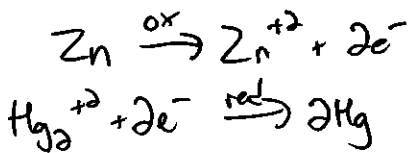


Given the following data for the above reaction

$$\begin{aligned} E^{\circ}_{tot} &= 1.56 \text{ V} \\ [Hg_2^{+2}] &= 0.750 \text{ M} \\ [Zn^{+2}] &= 0.380 \text{ M} \end{aligned}$$

Using the Nernst equation, compute the potential of the cell under these concentration conditions at 25°C.

(E7)



$$\begin{aligned} E^{\circ}_{tot} &= 1.56 \text{ V} \\ n &= 2 \end{aligned}$$

$$Q = \frac{[Zn^{+2}]}{[Hg_2^{+2}]}$$

Nernst

$$E = E^{\circ} - \left(\frac{0.0257}{n} \ln Q \right) = E^{\circ} - \left(\frac{0.0257}{n} \ln \left(\frac{[Zn^{+2}]}{[Hg_2^{+2}]} \right) \right)$$

$$E = 1.56 \text{ V} - \left(\left(\frac{0.0257}{2} \right) \ln \left(\frac{0.380 \text{ M}}{0.750 \text{ M}} \right) \right)$$

$$E = 1.56 \text{ V} - (-0.00874) = \boxed{1.57 \text{ V}}$$

V. Electrolytic Cells

p. 847 in Zumdahl

A. Qualitative aspects

1. The reaction in this cell is nonspontaneous.
2. A battery (source of electrical energy) acts as an electron pump, pushing electrons into the cathode and removing them from the anode.
3. An oxidation-reduction carried out in an electrolytic cell is called electrolysis.

B. Cell Reactions in Aqueous Solutions

1. Possible reactions at the cathode

a. The cation gets reduced to metal (neutral charge) $Cu^{2+} + 2e^- \rightarrow Cu$

H_2O b. Water gets reduced to Hydrogen Gas (H_2)

Half-reaction: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E_{red} = -0.828V$

2. Possible reactions at the anode

a. The anion gets oxidized to gas (neutral charge) $2Cl^- \rightarrow Cl_2 + 2e^-$

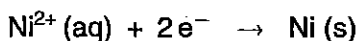
b. Water gets oxidized to oxygen gas (O_2).

Half-reaction: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $E_{red} = +1.229V$

$E_{ox} = -1.229V$

C. Quantitative aspects of electrolysis

1. When considering a half-reaction as in



you can write the following equivalences:

$$1 \text{ mol } Ni^{2+} = 2 \text{ mol } e^- = 1 \text{ mol } Ni$$

2. A mole of electrons is equal to 96,485 coulombs. This is known as the Faraday constant.

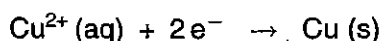
3. One coulomb is equal to 1 amp-s.

$$1A = 1 \frac{C}{s}$$

4. Exercises

a. How many grams of copper metal can be obtained by passing 1.20 amp of electric current for half an hour through an aqueous solution of copper(II) ions?

Solution: First, we write the half-reaction



$$30 \text{ min} \left(\frac{60 \text{ sec}}{1 \text{ min}} \right) \left(\frac{1.2 \text{ C}}{\text{sec}} \right) \left(\frac{1 \text{ mole } e^-}{96,485 \text{ C}} \right)$$

Next, we write the equivalences

$$2 \text{ mol } e^- \cong 1 \text{ mol } Cu^{2+} \cong 1 \text{ mol } Cu$$

Now, since we are given the number of amperes and how long the current goes through, we can calculate how many coulombs are involved. Remember that coulombs are ampere-seconds, hence time has to be expressed in seconds.

$$0.500 \text{ hr} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} \times 1.20 \text{ amp} = 2.16 \times 10^3 \text{ amp-s}$$

Since 2160 amp-s is equal to 2160 coul, we can now convert coulombs to moles of copper and then to grams of copper.

$$2.16 \times 10^3 \text{ coul} \times \frac{1 \text{ mol } e^-}{96485 \text{ C}} \times \frac{1 \text{ mol } Cu}{2 \text{ mol } e^-} \times \frac{63.55 \text{ g}}{1 \text{ mol } Cu} = 0.711 \text{ g } Cu$$

$$1.20 \frac{C}{s}$$

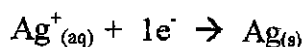
$$\left(\frac{1 \text{ mole } Cu}{2 \text{ mole } e^-} \right)$$

$$\left(\frac{63.6 \text{ g}}{1 \text{ mole } Cu} \right)$$

$$= 0.712 \text{ g } Cu$$

(2)

- b. A spoon made of copper is to be plated with silver. The spoon ~~is~~ acts as the cathode immersed in a $\text{Ag}(\text{CN})_2^-$ solution. The reduction reaction is:



$$I = \frac{Q}{t}$$

- a) What mass of silver would be deposited by a current of 0.500 amperes flowing for two hours? (E8)

? g Ag

$$I = .500 \text{ A}$$

$$t = 2 \text{ hr}$$

$$2 \text{ hr} \left(\frac{3600 \text{ s}}{1 \text{ hr}} \right) \left(\frac{.500 \cancel{\text{ A}}}{\text{sec}} \right) \left(\frac{1 \cancel{\text{ mole e}^-}}{96,485 \cancel{\text{ C}}} \right) \left(\frac{1 \cancel{\text{ mole Ag}}}{1 \cancel{\text{ mole e}^-}} \right) \left(\frac{108 \text{ g}}{1 \cancel{\text{ mole Ag}}} \right) =$$

4.03 g Ag formed

- b) If the density of silver is 10.5 g/cm^3 , what volume of silver would be deposited on the spoon?

$$4.03 \text{ g Ag} \left(\frac{1 \text{ cm}^3}{10.5 \text{ g}} \right) = .384 \text{ cm}^3$$

- c) If Area of the spoon being plated was 0.853 cm^2 , how thick would the silver plating be?

$$V = l \times w \times h = A \times h$$

$$A = l \times w$$

$$h = \frac{V}{A} = \frac{.384 \text{ cm}^3}{.853 \text{ cm}^2} = \underline{.450 \text{ cm}}$$

TABLE 8.1 STANDARD POTENTIALS IN WATER SOLUTION AT 25°C

OXIDIZING AGENT	REDUCING AGENT	E_{red}° (V)
$\text{Li}^+(\text{aq}) + e^-$	$\rightarrow \text{Li}(\text{s})$	-3.05
$\text{K}^+(\text{aq}) + e^-$	$\rightarrow \text{K}(\text{s})$	-2.93
$\text{Ba}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Ba}(\text{s})$	-2.90
$\text{Ca}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Ca}(\text{s})$	-2.87
$\text{Na}^+(\text{aq}) + e^-$	$\rightarrow \text{Na}(\text{s})$	-2.71
$\text{Mg}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Mg}(\text{s})$	-2.37
$\text{Al}^{3+}(\text{aq}) + 3 e^-$	$\rightarrow \text{Al}(\text{s})$	-1.66
$\text{Mn}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Mn}(\text{s})$	-1.18
$\text{Zn}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Zn}(\text{s})$	-0.76
$\text{Cr}^{3+}(\text{aq}) + 3 e^-$	$\rightarrow \text{Cr}(\text{s})$	-0.74
$\text{Fe}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Fe}(\text{s})$	-0.44
$\text{Cr}^{3+}(\text{aq}) + e^-$	$\rightarrow \text{Cr}^{2+}(\text{aq})$	-0.41
$\text{Cd}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Cd}(\text{s})$	-0.40
$\text{PbSO}_4(\text{s}) + 2 e^-$	$\rightarrow \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq})$	-0.36
$\text{Tl}^+(\text{aq}) + e^-$	$\rightarrow \text{Tl}(\text{s})$	-0.34
$\text{Co}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Co}(\text{s})$	-0.28
$\text{Ni}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Ni}(\text{s})$	-0.25
$\text{AgI}(\text{s}) + e^-$	$\rightarrow \text{Ag}(\text{s}) + \text{I}^-(\text{aq})$	-0.15
$\text{Sn}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Sn}(\text{s})$	-0.14
$\text{Pb}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Pb}(\text{s})$	-0.13
$2 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow \text{H}_2(\text{g})$	0.00
$\text{AgBr}(\text{s}) + e^-$	$\rightarrow \text{Ag}(\text{s}) + \text{Br}^-(\text{aq})$	0.07
$\text{S}(\text{s}) + 2 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow \text{H}_2\text{S}(\text{aq})$	0.14
$\text{Sn}^{4+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Sn}^{2+}(\text{aq})$	0.15
$\text{Cu}^{2+}(\text{aq}) + e^-$	$\rightarrow \text{Cu}^+(\text{aq})$	0.15
$\text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow \text{SO}_2(\text{g}) + 2 \text{H}_2\text{O}$	0.20
$\text{Cu}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Cu}(\text{s})$	0.34
$\text{Cu}^+(\text{aq}) + e^-$	$\rightarrow \text{Cu}(\text{s})$	0.52
$\text{I}_2(\text{s}) + 2 e^-$	$\rightarrow 2 \text{I}^-(\text{aq})$	0.53
$\text{Fe}^{3+}(\text{aq}) + e^-$	$\rightarrow \text{Fe}^{2+}(\text{aq})$	0.77
$\text{Hg}_2^{2+}(\text{aq}) + 2 e^-$	$\rightarrow 2 \text{Hg}(\text{l})$	0.79
$\text{Ag}^+(\text{aq}) + e^-$	$\rightarrow \text{Ag}(\text{s})$	0.80
$2 \text{Hg}^{2+}(\text{aq}) + 2 e^-$	$\rightarrow \text{Hg}_2^{2+}(\text{aq})$	0.92
$\text{NO}_3^-(\text{aq}) + 4 \text{H}^+(\text{aq}) + 3 e^-$	$\rightarrow \text{NO}(\text{g}) + 2 \text{H}_2\text{O}$	0.96
$\text{AuCl}_4^-(\text{aq}) + 3 e^-$	$\rightarrow \text{Au}(\text{s}) + 4 \text{Cl}^-(\text{aq})$	1.00
$\text{Br}_2(\text{l}) + 2 e^-$	$\rightarrow 2 \text{Br}^-(\text{aq})$	1.07
$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 e^-$	$\rightarrow 2 \text{H}_2\text{O}$	1.23
$\text{MnO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow \text{Mn}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}$	1.23
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 e^-$	$\rightarrow 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}$	1.33
$\text{Cl}_2(\text{g}) + 2 e^-$	$\rightarrow 2 \text{Cl}^-(\text{aq})$	1.36
$\text{ClO}_3^-(\text{aq}) + 6 \text{H}^+(\text{aq}) + 5 e^-$	$\rightarrow \frac{1}{2} \text{Cl}_2(\text{g}) + 3 \text{H}_2\text{O}$	1.47
$\text{Au}^{3+}(\text{aq}) + 3 e^-$	$\rightarrow \text{Au}(\text{s})$	1.50
$\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 e^-$	$\rightarrow \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}$	1.52
$\text{PbO}_2(\text{s}) + \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}$	1.68
$\text{H}_2\text{O}_2(\text{aq}) + 2 \text{H}^+(\text{aq}) + 2 e^-$	$\rightarrow 2 \text{H}_2\text{O}$	1.77
$\text{Co}^{3+}(\text{aq}) + e^-$	$\rightarrow \text{Co}^{2+}(\text{aq})$	1.82
$\text{F}_2(\text{g}) + 2 e^-$	$\rightarrow 2 \text{F}^-(\text{aq})$	2.87