

Now that you're grounded in fundamentals, we extend the discussion into other aspects of potential importance.

63.1 Two halves make a whole.

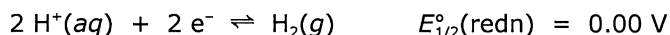
We will be working quite a bit with Appendix C and its list of half-reactions and half-cell potentials. Look it over and be ready to follow along with it as we go. Let me make five Points right off the bat.

First Point: By convention, such tables list each half-reaction as a reduction; as such, the electrons are on the left side of each equation. Thus, the potentials are specifically reduction potentials. A whole reaction needs a reduction half-reaction and an oxidation half-reaction. The oxidation half-reaction will be the reverse of a reduction; in reversing the reaction, the signs of $\Delta G_{1/2}^\circ$ and $E_{1/2}^\circ$ will negate. I'll illustrate this in a moment.

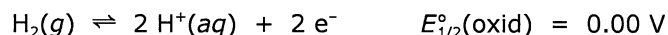
Second Point: The equations are written using a double half-arrow notation, \rightleftharpoons , instead of a single arrow, \rightarrow . Although we have in the past associated the double half-arrow notation with an equilibrium system, many (not all) redox reactions will go to completion, entirely to the right. Nevertheless, the double half-arrow notation emphasizes that these half-reactions can operate forward or backward, as a reduction or as an oxidation, depending on the reaction and the conditions. For our purposes here, we will not make a large distinction between the use of \rightleftharpoons versus \rightarrow .

Third Point: The equations for some half-reactions include H^+ and some include OH^- , corresponding to acidic or basic conditions. The presence of H^+ or of OH^- follows Scenario A or B for balancing half-reactions in Chapter 61. Thus, if the conditions of a half-cell are stated to be acidic or if no condition is specified, then H^+ can be present in the half-reaction. If the conditions are stated to be basic, then OH^- can be present in the half-reaction. Does that matter? It can, and in a big way. For example, compare half-reactions #6 and #20 in Appendix C: both are for the ClO_3^-/Cl^- couple but #6 includes H^+ while #20 includes OH^- ; their half-cell potentials are 1.45 V and 0.62 V, which are very different. For any half-reaction which involves H^+ or OH^- , then the potential will change if the acid/base conditions change.

Fourth Point: SHE is #30. Note the zero V.



The corresponding oxidation

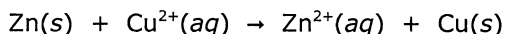


is also a zero.

Fifth Point: The half-cell potentials in Appendix C are given to two decimal places, although many are known beyond that. Two decimals will suffice for most of our purposes, although I may occasionally throw in more depending on the type of calculation. Besides, these calculations will again run into ideal versus real considerations, as with all aqueous solutions and especially so with ionic solutes.

This ends the Points.

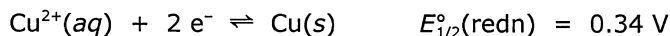
OK, consider the $Zn + Cu^{2+}$ reaction again.



We'll start with the reduction half-reaction.

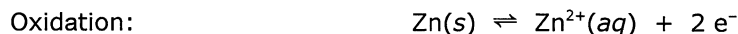


Go to Appendix C, look for this reaction. Be careful with this! When you search Appendix C, be aware that several half-reactions can look alike if you're not careful; you have to be sure to get the correct one. In Appendix C, #25 and #28 both involve the reduction of $Cu^{2+}(aq)$; #25 is the correct equation for the present usage with $Cu(s)$ as product. The value for $E_{1/2}^\circ$ is 0.34 V.

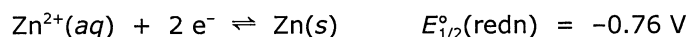


This compares very well to the value as calculated from $\Delta G_{1/2}^\circ$ at the tail end of the last Chapter.

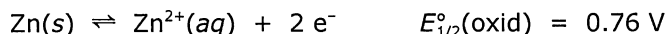
Now turn to the oxidation half-reaction.



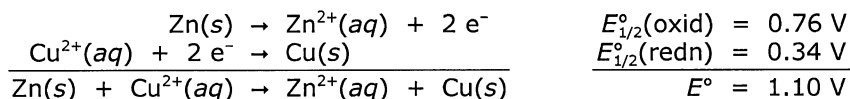
If you look this up in Appendix C, you will not find it in this form but you will find it written as a reduction half-reaction, #41.



Reverse that for oxidation and negate the sign.

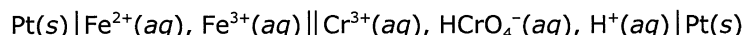


This compares very well to the value as also calculated from $\Delta G_{1/2}^{\circ}$ at the tail end of the last Chapter. By the rationale demonstrated at that time, the whole cell potential is the sum of the half-cell potentials.

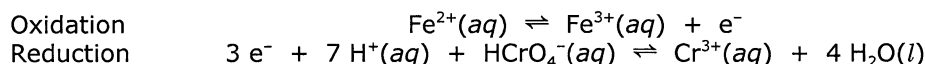


This illustrates the typical approach for determining a whole cell potential from its halves.

As another illustration, consider the cell notation from Example 4 in Chapter 62

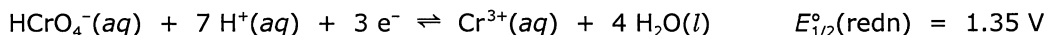


for which you derived the following half-reactions.

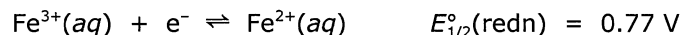


Now, find E° for the whole cell.

Look up these reactions in Appendix C. One will be written as the reduction shown above, while the other will be reversed. The reduction is #9 in Appendix C



while the oxidation is the reverse of #18. Here it is as it appears in Appendix C.



Go ahead and reverse and negate.



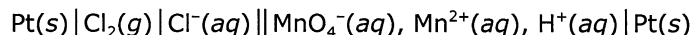
The potential of the whole cell is then

$$E^{\circ} = E_{1/2}^{\circ}(\text{redn}) + E_{1/2}^{\circ}(\text{oxid}) = 1.35 \text{ V} + (-0.77 \text{ V}) = 0.58 \text{ V}$$

and that's all there is to this.

Let's start in on a number of new Examples. We'll first start from cell notation.

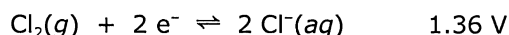
Example 1. Consider the following cell.



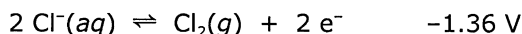
Derive the balanced equation for the whole reaction and find E° of the cell.

Although we covered the full process for balancing equations in Chapter 61, there is now a big shortcut available: Appendix C. When you work with half-cell potentials and tables of them, then you are given balanced equations. If your half-reactions are in the table, then you're in luck because the equation is already balanced. If your half-reaction is not in the table, then you still have to do the balance.

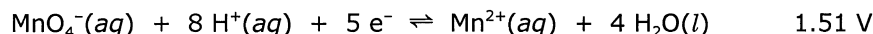
To derive the balanced equation for this problem, find the half-reactions in Appendix C. We'll start with the half-cell on the left in the notation, which means it is the anode. The cell notation says that $\text{Cl}_2(\text{g})$ and $\text{Cl}^{-}(\text{aq})$ are involved; look for the half-reaction with those two reagents on opposite sides. You will find this as #8, written as a reduction.



In the present cell, this is for the anode so we need it written as an oxidation. Flip it around.

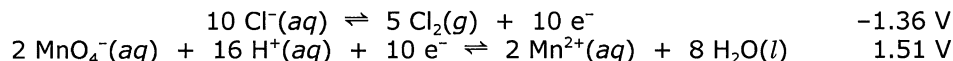


For the cathode, the cell notation says that $\text{MnO}_4^-(aq)$, $\text{Mn}^{2+}(aq)$ and $\text{H}^+(aq)$ are involved. Find the half-reaction with those reagents on either side of the equation. You will find this as #4.



Notice that the half-reaction includes $\text{H}_2\text{O}(l)$ but the cell notation does not. Remember: H_2O does not show in cell notation although it may be present in the equation for the half-reaction.

We have the balanced half-reactions but we need to get loss equal to gain before we can add these for the whole reaction. The oxidation loses two electrons and the reduction gains five electrons. Get those equal; we can set them both to ten by multipliers. Multiply the Cl^-/Cl_2 half-reaction by five and $\text{MnO}_4^-/\text{Mn}^{2+}$ by two.

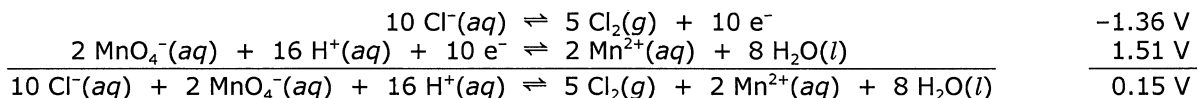


Note! Although each half-reaction was multiplied, the potentials were not. Remember that $\Delta G^\circ = -nFE^\circ$ and remember what I said in Section 62.3.

“ By the way, like ΔG° itself, the value of n depends on the coefficients for the reaction as written. If you multiply an equation, you multiply ΔG° and you multiply n ; this point will be important later. ”

Later is now. Keep in mind that E° represents the push/pull on the charges. When you multiply an equation, you multiply the number of moles of everything in the equation; for charges, that means you multiply n . The push/pull per charge stays the same, so E° stays the same.

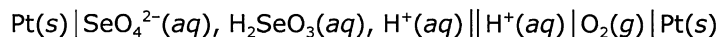
Now, add the half-reactions.



That's the balanced equation and that's the E° for the cell.

Your turn.

Example 2. Consider the following cell.



Derive the balanced equation for the whole reaction and find E° of the cell.

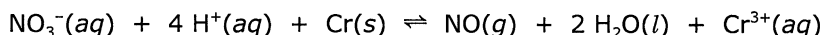
Find the half-reactions in Appendix C. Clue: be careful picking out the cathode reaction. There are three entries in Appendix C which involve the reduction of O_2 but only one fits the given information for this Example. To decide which of the three is correct, watch your reagents; only $\text{H}^+(aq)$ and $\text{O}_2(g)$ are given in the cell notation, but don't forget that water can be involved. Once you've identified your two half-reactions, write them out here, but flip the half-reaction for the anode.

Set loss equal to gain, then add the equations. Cancel whatever needs to be canceled.

(As a clue, all H_2O will cancel out in the final equation.) Finally, what is E° ? You should get 0.08 V.

Next, we'll start with a balanced equation.

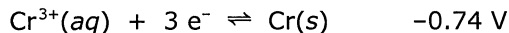
Example 3. Consider the following reaction.



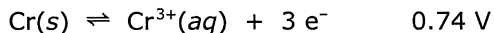
Derive this equation from half-cell equations and find E° of the cell.

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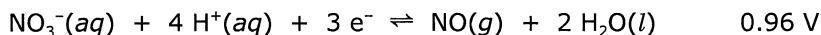
Go to Appendix C and find the half-reactions. The prior Examples started with the cell notation which spells out the ingredients in each half-cell. Now we start with a whole equation; when starting from there, the ingredients in each half-cell may not be so obvious. This particular reaction is not too bad, however. Hopefully you can readily see Cr/Cr^{3+} as being one couple. Find the half-reaction in Appendix C which involves those two.



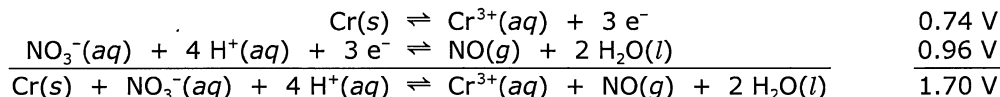
In the whole equation as given in the problem, $\text{Cr}(\text{s})$ is on the left and $\text{Cr}^{3+}(\text{aq})$ is on the right, which is the opposite of the half-reaction from Appendix C. That means this is the oxidation portion for the cell. Reverse the half-reaction from Appendix C.



The other half-reaction will involve the other ingredients in the equation, namely NO_3^- , H^+ , NO and H_2O . Find that half-reaction in the Appendix.



Both half-reactions involve three electrons, so we won't need any multipliers here. Go ahead and add.

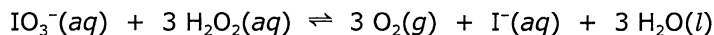


This gives the equation as presented in the problem, along with $E^\circ = 1.70 \text{ V}$.

Your turn.

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Example 4. Consider the following reaction.



Derive this equation from half-cell equations and find E° of the cell.

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Find the half-reactions in Appendix C. Reverse one of them. Clue: both half-reactions also have $\text{H}^+(\text{aq})$ although that is not shown in the equation as given.

Set loss equal to gain, then add.

Notice that all H^+ cancel out in the final equation. Did you arrive at the balanced equation as given for the problem? Did you get $E^\circ = 0.38 \text{ V}$?

Pause a moment.

In all of the Examples so far, we arrived at a positive E° and that designates a galvanic cell and an exergonic reaction (at standard conditions). There was a subtle characteristic to all of these Examples which now needs to be pointed out. In all cases, the half-reaction used for reduction was higher on the list in Appendix C than the half-reaction which was used for oxidation (after reversing). This aspect relates to how the half-reactions are listed in the Appendix: the list goes from the most positive $E_{1/2}^\circ$ to the most negative $E_{1/2}^\circ$. This numerical listing is very common in tables for half-cell potentials, although some tables just do an alphabetical list based on the couple. For tables which use a numerical list based on the values of $E_{1/2}^\circ$, then there are a few useful Consequences to note.

The most positive $E_{1/2}^\circ(\text{redn})$ values are at the top of the list, and those correspond to the most negative $\Delta G_{1/2}^\circ(\text{redn})$ values. Thus, the half-reactions at the top have the most driving force to occur, in

a whole cell application. Since they are written for reduction, then these are the reductions with the greatest driving force relative to any half-reaction lower on the list. The topmost half-reaction in Appendix C is the reduction of ozone.

Most driving force for reduction: $\text{O}_3(g) + 2 \text{H}^+(aq) + 2 e^- \rightleftharpoons \text{O}_2(g) + \text{H}_2\text{O}(l)$ 2.08 V

This half-reaction has the greatest driving force for reduction out of all half-reactions listed. As you go down from there, the driving force for reduction lessens. That brings us to Consequence 1.

- Consequence 1. For any two half-reactions on the list, the higher one will have the greater driving force for reduction.

Recall that it is the oxidizing agent which undergoes reduction. A stronger oxidizing agent will have a stronger driving force for reduction, which means the stronger oxidants will be higher in the list. That is Consequence 2.

- Consequence 2. The strongest oxidants are at the top of the list and the strength of the oxidant decreases going down.

$\text{O}_3(g)$ is the strongest oxidizing agent of all in Appendix C.

Now get to the bottom of the list: there we find the most negative $E_{1/2}^\circ(\text{redn})$ values. Those correspond to the most positive $\Delta G_{1/2}^\circ(\text{redn})$ values. Thus, the half-reactions at the bottom have the least driving force to occur as written, which means they have the least driving force for reduction. The last and the least on the list is the reduction of $\text{Mg}^{2+}(aq)$.

Least driving force for reduction: $\text{Mg}^{2+}(aq) + 2 e^- \rightleftharpoons \text{Mg}(s)$ -2.37 V

But now consider the reverse, which means consider oxidation instead. For the reverse (oxidation), the signs in Appendix C are negated. This means the most positive $E_{1/2}^\circ(\text{oxid})$ values would fall at the bottom of the list and this corresponds to the most negative values for $\Delta G_{1/2}^\circ(\text{oxid})$. Thus, the half-reactions at the bottom of the list have the most driving force for the reverse reaction, oxidation.

Most driving force for oxidation: $\text{Mg}(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2 e^-$ +2.37 V

That brings us to Consequence 3.

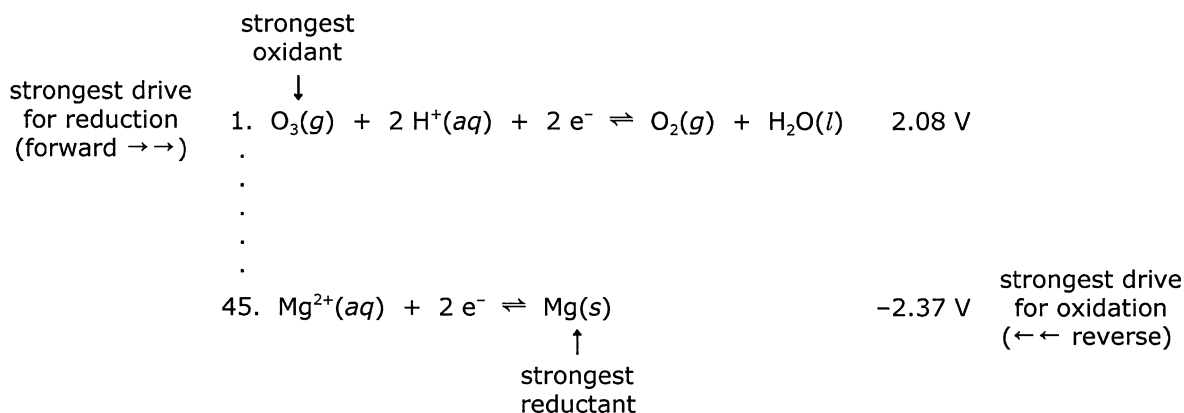
- Consequence 3. For any two half-reactions on the list, the lower one will have the greater driving force for the reverse direction, oxidation.

Recall that it is the reducing agent which undergoes oxidation. A stronger reducing agent has a stronger driving force for oxidation, which means the stronger reductants will be lower in the list. That is Consequence 4.

- Consequence 4. The strongest reductants are at the bottom of the list and the strength of the reductant decreases going up.

Since this oxidation aspect is the reverse of the half-reactions as written, then the reducing agents are on the right side of each equation in Appendix C. Overall, the strongest reducing agents are at the bottom in Appendix C and on the right side of the equation. $\text{Mg}(s)$ is the strongest reducing agent in that list.

Here's a schematic of the four Consequences so far, using the first and last equations in Appendix C for illustration.



Now we come to the final Consequence, and this relates to the exergonic aspect for the whole cells in all Examples so far. Consider any whole reaction composed of its two half-reactions.

- Consequence 5. A whole reaction will have a positive E° and it will be exergonic at standard conditions when the reduction half-reaction is higher on the list than the oxidation half-reaction.

From the prior Consequences, for any two half-reactions on the list, the higher one is better at reduction and has the more positive $E_{1/2}^\circ(\text{redn})$; the reverse of the lower half-reaction is better at oxidation and has the more positive $E_{1/2}^\circ(\text{oxid})$. When you add these two, you will always get a positive E° and an exergonic reaction. If you go back and look at the relative ranking of the half-reactions in Examples 1 - 4, you will see that the reduction half-reaction was indeed higher on the list than the oxidation, and the outcome was a positive E° in each case. On the other hand, you can force the upper half-reaction to do oxidation while the lower half-reaction does reduction, but that opposes the driving force and that is endergonic at standard conditions. We will see those cases in Section 65.1.

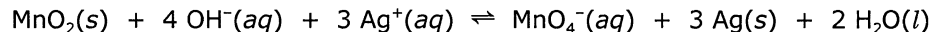
As another result of this setup, you can get a quick assessment for the exergonic direction for some desired reaction. Let's say you want to oxidize $\text{I}^-(\text{aq})$ to $\text{I}_2(\text{s})$. What oxidizing agent can you use? Look at Appendix C. The oxidation of $\text{I}^-(\text{aq})$ to $\text{I}_2(\text{s})$ would be the reverse of #22. You could use any oxidizing agent higher on the list, and the reaction would be exergonic at standard conditions. Let's now say you want to reduce $\text{Cu}^{2+}(\text{aq})$ to $\text{Cu}(\text{s})$. What reducing agent can you use? The reduction of $\text{Cu}^{2+}(\text{aq})$ to $\text{Cu}(\text{s})$ is #25. You could use any reducing agent from the right side of any reaction lower on the list. We were using $\text{Zn}(\text{s})$ previously for this task, and it's at #41.

OK, time to move on. We now return to $\Delta G^\circ = -nFE^\circ$ and we bring in a new application for half-reactions.

63.2 Two halves make another half.

We'll first do a calculation of $\Delta G^\circ = -nFE^\circ$ as a variation on the coverage so far and because we'll need it here. We had calculated values for E° from ΔG° in the last Chapter, and now we will go the other direction using E° derived from half-cell potentials.

Example 5. Consider the following reaction.



Find E° and ΔG° .

This Example does not specifically ask for a bunch of balanced equations, but we will need to at least identify the relevant half-reactions in order to get their $E_{1/2}^\circ$ values. The Ag^+/Ag couple is evident in the given equation and it is written for reduction; this half-reaction is #17 in Appendix C. That means the other half-reaction is oxidation, and it involves the $\text{MnO}_2/\text{MnO}_4^-$ couple along with some H_2O and OH^- . That half-reaction is #21 in Appendix C. #21 would be reversed and its listed $E_{1/2}^\circ$ value would be negated. We then add these potentials together.

$$E^\circ = 0.80 \text{ V} + (-0.60 \text{ V}) = 0.20 \text{ V}$$

Now find ΔG° .

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

We've got the constant F and we've found E° , but what is the value for n ? How many moles of electrons are transferred for the equation as written? There are three approaches for determining this.

First Way. Half-reaction #21 has one MnO_2 and one MnO_4^- showing in the equation, just like the equation for the whole reaction as given here. Half-reaction #21 involves three electrons, so $n = 3 \text{ mol}$ will also apply to the whole reaction.

Second Way. Half-reaction #17 has one silver species on each side and it involves one mol of electrons. The equation for the whole reaction here has three silvers, so it derives from multiplying #17 by three. That gives you $n = 3 \text{ mol}$ for the whole reaction.

Third Way. Find the total change in oxnos for the whole equation as written; that is equal to the value of n . The total change in oxnos was introduced in Section 14.2. You can do this for either of the couples ($\text{MnO}_2/\text{MnO}_4^-$ or Ag^+/Ag) and it will give the same answer: $n = 3 \text{ mol}$.

Now plug n into ΔG° .

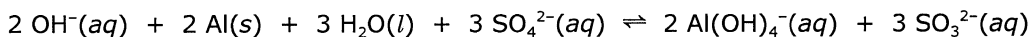
$$\Delta G^\circ = -nFE^\circ = -3 \text{ mol} \times 96,485 \text{ C/mol} \times 0.20 \text{ V}$$

$$= -58,000 \text{ C} \cdot \text{V} = -58,000 \text{ J} = -58 \text{ kJ}$$

Done.

Remember that n is the number of moles of electrons which are transferred for a reaction as written. That can apply for a half-reaction or for a whole reaction. The Third Way connects n to the total change in oxnos for the whole reaction. This is yet another application for oxidation numbers in analyzing redox processes.

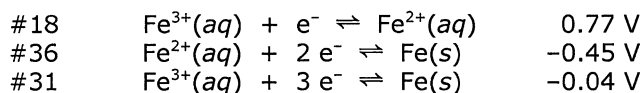
If you'd like a little practice, take a shot at the following equation. Find n , using any of the above three Ways.



This equation is from Example 3 in Chapter 61. The value for n can be found in Step 4 or 5 for that Example. Try it first here before looking back.

We now consider a new application of the half-cell potentials in Appendix C. These tables allow you to determine a potential for an unlisted half-reaction, as long as you can derive that half-reaction from a combination of the ones in the table. This is yet another result of the additivity of reactions.

Consider three equations from Appendix C which relate $\text{Fe}(s)$, $\text{Fe}^{2+}(aq)$ and $\text{Fe}^{3+}(aq)$.



You can arrive at the third equation by adding the first two, but you do not obtain the correct $E_{1/2}^\circ$ by simple addition. Why? You need to account for each n .

Recall that $E_{1/2}^\circ$'s are additive when n 's are the same. End of Chapter 62:

“ As long as n 's are the same for both half-reactions and for the total reaction, then you can add $E_{1/2}^\circ$'s. Fortunately, this will always be true when deriving a whole reaction from two half-reactions; in this application, the n 's represent the number of electrons lost and gained, and the numbers must be equal for balancing. The additivity of E° 's in this kind of application is an important result which we will utilize heavily.

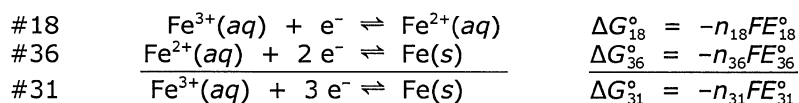
Note the underlines. There are other applications in which the n 's can differ, and in those cases you cannot add E° 's directly. We will see such an application later in Chapter 63. ”

Later is now. This new application is different from what we were doing and I want to spell out the distinction for clarity.

In Sections 62.3 and 63.1, we used the additivity of $E_{1/2}^\circ$'s to determine the potential of a whole cell from two half-cells; that was possible because the n 's were the same for the two half-reactions and for the whole reaction, as required for balancing. In these cases, the electrons cancel out in the end.

Now we will derive a third half-cell potential from two given half-cell potentials. In these applications, the n 's cannot all be the same, and the electrons do not cancel out in the end. Thus, we can no longer do simple addition of $E_{1/2}^\circ$'s. But ΔG° still rules, and the $\Delta G_{1/2}^\circ$'s are still additive.

Let's see how this works out for the iron equations. Watch the labels in the subscripts since these will now show which half-reaction they belong to.



Since $\Delta G_{1/2}^\circ$'s are additive, we get the following.

$$-n_{18}FE_{18}^\circ + (-n_{36}FE_{36}^\circ) = -n_{31}FE_{31}^\circ$$

F is a constant; factor it out along with the negatives.

$$n_{18}E_{18}^\circ + n_{36}E_{36}^\circ = n_{31}E_{31}^\circ$$

For illustration purposes, let's solve for E_{31}° . The n 's are simply the number of electrons showing in each equation. Notice that the three half-reactions have different values for their n 's.

$$(1 \text{ mol})E_{18}^\circ + (2 \text{ mol})E_{36}^\circ = (3 \text{ mol})E_{31}^\circ$$

The mol units factor out. Re-arrange.

$$[(1)E_{18}^{\circ} + (2)E_{36}^{\circ}] \times 1/3 = E_{31}^{\circ}$$

Plug in some volts.

$$E_{31}^{\circ} = [(1)(0.77 \text{ V}) + (2)(-0.45 \text{ V})] \times 1/3 = (-0.13 \text{ V}) \times 1/3 = -0.043 \text{ V}$$

There you have it: using the additivity of ΔG° 's, we have derived -0.043 V for #31, while the value from Appendix C is -0.04 V .

As a point to note, watch your sigfigs on these types of problems. They are a combination of +/- and \times/\div , and can be a bit tricky.

This method provides additional half-cell potentials beyond what is available in a given listing. The process breaks down very nicely as follows. Note the one equation above

$$n_{18}E_{18}^{\circ} + n_{36}E_{36}^{\circ} = n_{31}E_{31}^{\circ}$$

which we will now generalize: when adding half-reaction 'a' to half-reaction 'b' to derive half-reaction 'c', then their n 's and $E_{1/2}^{\circ}$'s will be related as follows.

$$n_a E_a^{\circ} + n_b E_b^{\circ} = n_c E_c^{\circ}$$

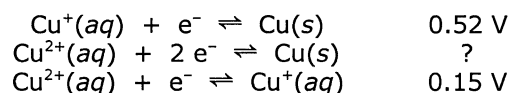
Thus, the $n_c E_c^{\circ}$ term for the resulting half-reaction equals the sum of the nE° terms for the parts. By the way, when all n 's are equal, they also factor out. This gives the direct additivity of $E_{1/2}^{\circ}$'s as we had been doing for whole cell potentials.

$$E_a^{\circ} + E_b^{\circ} = E_c^{\circ} \quad (\text{all } n\text{'s equal})$$

But that won't work when one or more n 's are different, as in the present application.

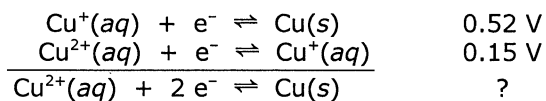
Time for another Example.

Example 6. The following half-reactions relate $\text{Cu}(s)$, $\text{Cu}^+(aq)$ and $\text{Cu}^{2+}(aq)$.



Determine the value (in V) for ?.

OK, first you want to set these up as a summation of two half-reactions to give a third half-reaction. This setup will not necessarily be in the same sequence as in the provided information. The clue for you to follow is that the sum half-reaction has the largest value of n . Here, the Cu^{2+}/Cu half-reaction has $n = 2 \text{ mol}$, while the others have $n = 1 \text{ mol}$; we need the Cu^{2+}/Cu half-reaction to be the final sum equation. Set it up as such.



Again, we cannot sum the half-cell potentials directly, but we can sum the $\Delta G^{\circ} = -nFE^{\circ}$ terms for each half-reaction and that, as seen above, simplifies to the general nE° equation.

$$n_a E_a^{\circ} + n_b E_b^{\circ} = n_c E_c^{\circ}$$

Start plugging things in.

$$(\text{___ mol})(\text{___ V}) + (\text{___ mol})(\text{___ V}) = (\text{___ mol})(?)$$

Mole units factor out again. Re-arrange and solve.

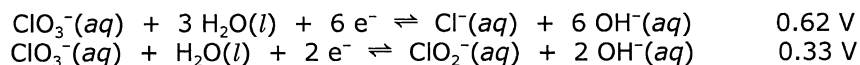
All of these half-cells are in Appendix C. You can check your answer there.

This method is fairly general and can even involve more complicated half-reactions. There is one situation where you cannot do this, however, at least within our coverage: you cannot do this when mixing

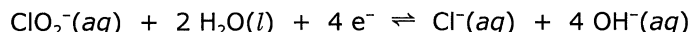
a half-cell reaction with H^+ and a half-cell reaction with OH^- . That can involve a neutralization component and that would modify the approach; we will not get into those cases here.

Let's take a another look with a bit more complication involved.

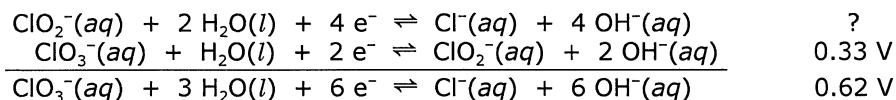
Example 7. From the following half-reactions and potentials,



find $E_{1/2}^\circ$ (in V) for the following.



The three half-reactions for this Example are much busier than those in Example 6 or in the prior iron series, but the methodology is the same. Set up a summation. Note that the $\text{ClO}_3^-/\text{Cl}^-$ half-cell has the largest n value (6 mol), so that will be the sum half-reaction. Set everything up.



The general nE equation

$$n_a E_a^\circ + n_b E_b^\circ = n_c E_c^\circ$$

leads to

$$(4 \text{ mol})(?) + (2 \text{ mol})(0.33 \text{ V}) = (6 \text{ mol})(0.62 \text{ V})$$

which becomes

$$4(?) = 3.7 \text{ V} - 0.66 \text{ V} = 3.0 \text{ V}$$

which finally gives 0.75 V for ?.

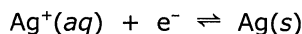
Wrap up!

We've now completed the methods for finding a whole cell potential from two half-cell potentials and for finding a half-cell potential from two others. The general process requires correctly identifying which half-reactions are involved, re-arranging the equations as needed, and summing the resulting $E_{1/2}^\circ$ or $nE_{1/2}^\circ$ terms. A very large number of potentials can thus be obtained.

But everything so far is for standard conditions, and life's not always a standard condition.

Problems

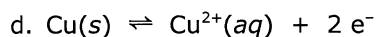
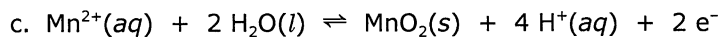
- Find E° (in V) for each of the following balanced equations.
 - $\text{O}_2(g) + 4 \text{H}^+(aq) + 4 \text{Ag}(s) \rightleftharpoons 2 \text{H}_2\text{O}(l) + 4 \text{Ag}^+(aq)$
 - $\text{MnO}_4^-(aq) + 2 \text{H}_2\text{O}(l) + 3 \text{Cu}^+(aq) \rightleftharpoons \text{MnO}_2(s) + 4 \text{OH}^-(aq) + 3 \text{Cu}^{2+}(aq)$
 - $3 \text{Sn}^{2+}(aq) + 2 \text{NO}_3^-(aq) + 8 \text{H}^+(aq) \rightleftharpoons 2 \text{NO}(g) + 3 \text{Sn}^{4+}(aq) + 4 \text{H}_2\text{O}(l)$
- Find E° (in V) for each cell as given by the following notations.
 - $\text{Pt}(s) | \text{H}_3\text{PO}_3(aq), \text{H}_3\text{PO}_2(aq), \text{H}^+(aq) || \text{Ni}^{2+}(aq) | \text{Ni}(s)$
 - $\text{Al}(s) | \text{OH}^-(aq), \text{Al}(\text{OH})_4^-(aq) || \text{OH}^-(aq) | \text{O}_2(g) | \text{Pt}(s)$
 - $\text{Pt}(s) | \text{IO}_3^-(aq), \text{H}^+(aq), \text{I}^-(aq) || \text{HCrO}_4^-(aq), \text{H}^+(aq), \text{Cr}^{3+}(aq) | \text{Pt}(s)$
- A galvanic cell at standard conditions is to be composed of the following reduction half-reaction.



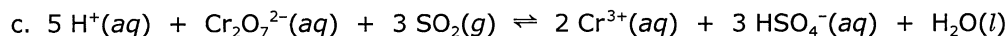
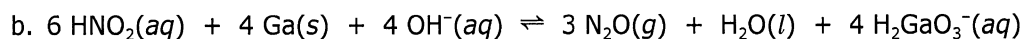
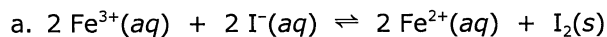
Which of the following would be suitable for use as an oxidation half-reaction?

- $2 \text{Br}^-(aq) \rightleftharpoons \text{Br}_2(l) + 2 e^-$

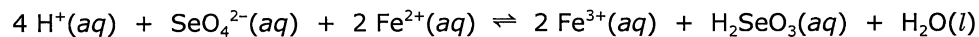
b. SHE



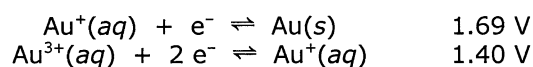
4. Consider each of the following balanced equations. For each equation, what is the value of n (in mol)?



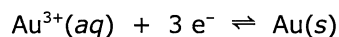
5. Find E° (in V) and ΔG° (in kJ) for the following reaction.



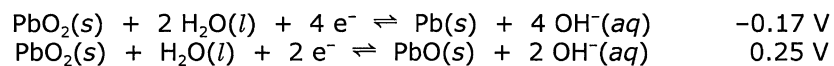
6. From the following half-reactions and potentials,



find $E_{1/2}^\circ$ (in V) for the following.



7. From the following half-reactions and potentials,



find $E_{1/2}^\circ$ (in V) for the following.

