

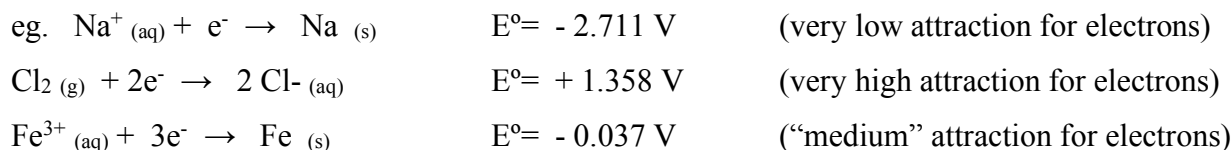
Unit 8, Lesson 06: Using Reduction Potentials to Calculate the Voltage of an Electrochemical Cell

Tyler DeWitt does not yet have a video on calculating cell potentials. Instead, please watch the video by ChemistNATE called Finding Ecell for a Reaction. (Please note that in his example, he has the anode on the right and the cathode on the left...this is not the order used in 'galvanic cell notation', but it is fine for the example that he is doing.)

An activity series is used to compare how easily an element loses electrons relative to other elements. The most reactive metals lose their electrons most easily and are at the top of the metal activity series.

Alternatively, we can measure and compare the level of attraction that an element has for new electrons. This is called the element's "reduction potential". In this series, the elements with the strongest attraction for electrons (the highest reduction potentials) are at the top of the list.

Standard reduction potentials (E°) are measured for elements in their standard states and at standard conditions, ie. in 1.0 M solution, at 25.0°C and 1.00 atm pressure. They are measured against the hydrogen electrode, $2 \text{H}^+ (\text{aq}) + 2 \text{e}^- \rightarrow \text{H}_2 (\text{g})$, which has been assigned a reduction potential of exactly 0.00 V. Standard reduction potentials (E°) are reported on page 598 of your text. This chart is photocopied on the last page of your exam resource package:

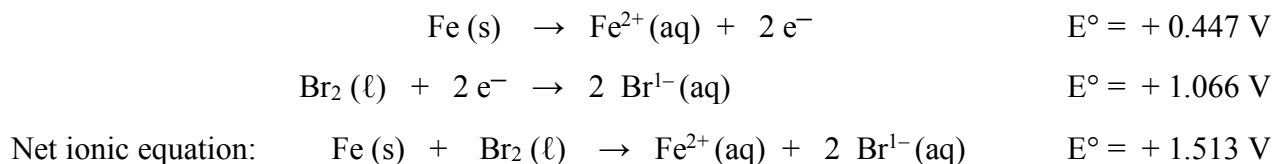


If you are given two half reactions, you can calculate the voltage that will be produced when these reactions are combined in an electrochemical cell:

1. Determine which half reaction will be the anode (lose electrons). The half reaction with the lowest (most negative) E° will be the anode. (**the **lower** half reaction **loses** electrons)
2. Flip this half reaction to write it as an oxidation and flip (change) the sign on the E° value.
3. Write the other half reaction as a reduction, keeping the E° as reported.
4. Balance the number of electrons so that the number of electrons lost at the anode equals the number of electrons gained at the cathode. *** Do **NOT** multiply the E° values by the molar coefficient (the **strength** of attraction for electrons does not increase as the number of moles goes up).
5. Add the E° values for the half reactions. The larger and more positive the E° for an electrochemical cell, the higher the voltage it will produce.

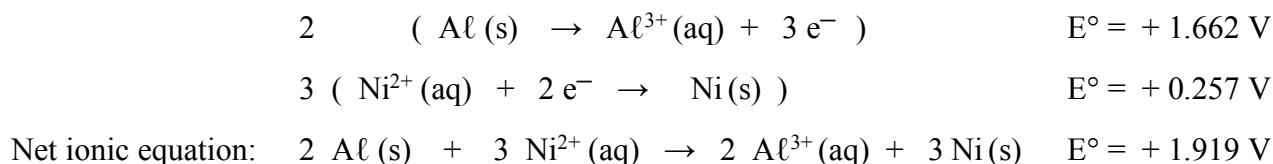
eg. Calculate the voltage produced by an electrochemical cell made of $\text{Br}_2 | \text{Br}^-$ and $\text{Fe} | \text{Fe}^{2+}$

Fe is lower on the list of standard reduction potentials, so it will lose electrons (lower loses). This is the anode. Write the iron half reaction and change the sign of its voltage (reduction potential). Write the bromine half reaction as a reduction (gaining electrons) and keep its voltage (reduction potential) as is:



eg. Calculate the voltage produced by an electrochemical cell made of $\text{Ni} | \text{Ni}^{2+}$ and $\text{Al} | \text{Al}^{3+}$

The aluminum half reaction has a lower reduction potential, so it will be oxidized and is the anode.



There is a mathematical 'shortcut' that we can use to calculate the voltage produced by any electrochemical cell (E° cell):

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

Note: Redox reactions are reversible, so they are equilibrium reactions. An electrochemical cell eventually stops producing electricity (ie. a battery 'dies') when the rate of the forward and reverse reactions are equal and the reaction achieves equilibrium.

FYI: The Ultimate Big Picture in Chemistry

There are three values that indicate the *extent* of an equilibrium reaction (how far it will proceed to the right, or how much product will be produced compared to the amount of reactant that is present):

$$1. K_{eq} = \frac{[\text{products}]}{[\text{reactants}]} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

The larger the value of K_{eq} , the more the reaction favours the products.

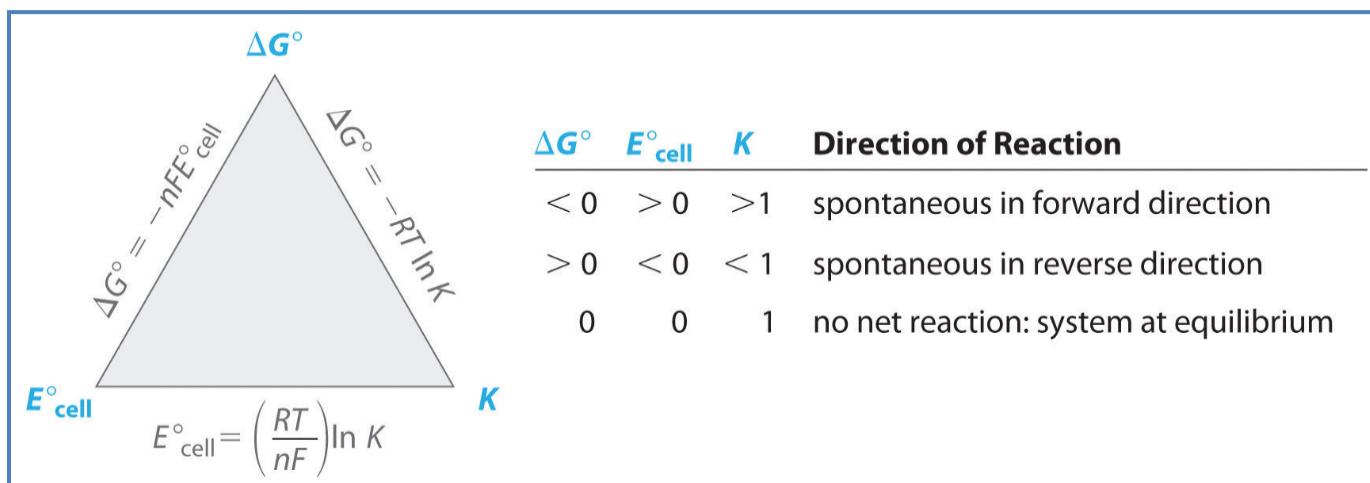
$$2. \Delta G^\circ = \Delta H - T\Delta S$$

The larger and more negative the value of ΔG , the more the reaction favours the products.

$$3. E^\circ \text{ cell} = E^\circ \text{ cathode} - E^\circ \text{ anode}$$

The larger and more positive the value of E° cell, the more the redox reaction favours the products.

Because all of these values indicate how far a reaction will proceed to the right, they are connected by other equations that you will see next year:



* R is the Ideal gas law constant, 8.31 kPa·L/mol·K (I told you that you would see 'R' again:)

* T is the temperature, in Kelvins

* K is the equilibrium constant K_{eq}

* n is the number of moles

* F is Faraday's constant, the amount of charge carried by one mole of electrons = 96485.3 Coulombs/mol

* ln is the natural logarithm. You will meet ln in your Functions class toward the end of the semester (it is on your calculator)