

XII) Standard Potentials

April 30, 2018 8:35 PM

2. Draw an electrochemical cell with one half-cell having an iron electrode and iron III nitrate solution and the other half-cell having an inert platinum electrode with $\text{Cl}_2(g)$ being bubbled in. The electrolyte in that half-cell is aqueous sodium chloride. Use KNO_3 in the salt bridge.

- Identify the cathode.
- Write the half-reaction occurring in the iron half-cell.
- What happens to the $[\text{NO}_3^-]$ in the iron half-cell?
- Write the half-reaction occurring in the Cl_2 half-cell.
- What happens to the extra Cl^- ions being produced at the cathode?
- What is the function of the platinum electrode at the cathode?
- Write the net redox reaction.

3. Read Hebden p.215-217 and do #34-35

The tendency for e^- to flow in an electrochemical cell is called **VOLTAGE**, or **Electrical Potential** "...to do work"

XII) Standard Potentials

The tendency of electron to flow in an electrochemical cell is called **voltage** or E_{cell}° to do work.

Cell potential, measured in volts (V), is the 'pull' on electrons. Potential, or voltage, acts on electrons just like 'pressure' acts on water, or like gravity acts on objects. In an **electrochemical cell, the substance that wants to reduce pulls the electrons (cathode)** through the wire, and this **movement of electrons is what is used to do work such as power a motor**. One volt is defined as 1 joule of work per coulomb of charge transferred, so voltage gives an indication of how much work the electrochemical cell can do. The higher the voltage, the stronger the battery.

work done / e^- transferred.

The **magnitude of the voltage in an electrochemical cell depends on its half-reactions**. If a substance that is **very good at giving up electrons oxidizes** (low on the right side of the table), and a substance that is **very good at reducing gains the electrons** (high on the left side), the cell will have a large voltage (large gap between half-reactions on table). The smaller the gap between the reduction and oxidation half-reactions, the smaller the voltage.

strong oxidising agent

strong reducing Agent (metals)

↑ voltage ($\Delta E_{\text{cell}}^\circ$) reduction + oxidation rxns to be far apart. Large Δ in E° (voltage) $\frac{1}{2}$ rxns.

Quiz 2 Tues 29th May

VII → XII
Assign #6 - #11

Unit Test June 8 Friday

"XIII) Independent study"

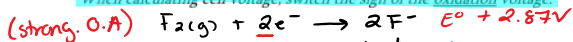
* June 11 - 15 > Re-Test Week *
1 test only ↓
(June 1st) ↓ Thurs.

oxidize / reduce

Table shows E° for the FWDS rxn (reduction only)

Find the cell voltage of a lithium/fluorine cell:

When calculating cell voltage, switch the sign of the oxidation voltage:



$E^\circ_{cell} + 5.91V$

← Flip sign largest cell potential

*Coefficients don't affect voltage, as it's joule per coulomb of charge

The E° values on the table are at 'standard state', meaning 25°C, 1 atm pressure for gases (room pressure), 1M concentrations for all solutions. If any of these conditions are altered, the voltage is just an E value as it's no longer at standard state.

Notice that each half-reaction is assigned a certain voltage on the redox table. It is impossible to set up just a half-cell (without another half-cell) and get a voltage reading. You can only get a voltage from a complete cell with oxidation and reduction half-reactions occurring simultaneously. So where do these values on the table come from?

must have an oxidation + reduction

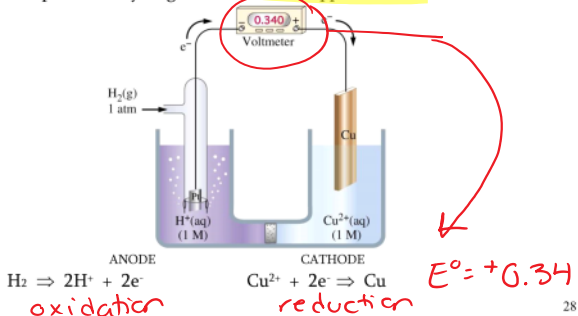
The hydrogen half-cell at standard state was used as a reference half-cell, and arbitrarily assigned a voltage of 0.00V. $2H^+ + 2e^- \rightarrow H_2 \quad E^\circ = 0.00V$

every other metal/non-metal was compared to the H-1/2 cell to calculate E°

(voltage) The potential of an electrochemical cell is just the difference between the half-cell voltages for the reduction and the oxidation reaction.

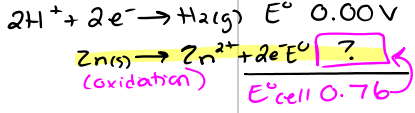
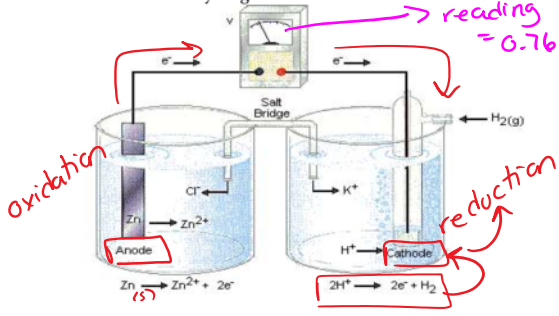
Every other half-cell was hooked up to the hydrogen half-cell and the voltage was recorded, and became the half-cell voltage for that respective half reaction.

Here is an example of the hydrogen half-cell and copper half-cell:



* higher than H on Table.

What about a hydrogen half-cell with a zinc half-cell?



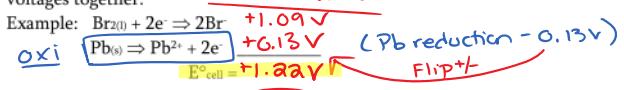
∴ The reduction 1/2 rxn for Zn
 $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}(\text{s}) \quad E^\circ = -0.76$
 (in Table)

anytime you rvs the rxn you flip the sign.

*Any other half-reaction could have been chosen as the reference half-cell and been assigned 0.00V, but the hydrogen cell was used.

The voltages given on the table are for reduction half-reactions.
 To get oxidation half-reaction voltages, reverse the sign on the voltage.
 Reduction of Cu^{2+} : $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} \quad E^\circ = 0.34\text{V}$ (in table)
 Oxidation of Cu: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^- \quad E^\circ = -0.34\text{V}$ (reading in reverse)

To predict the cell voltage of two half-cells, simply add the two half-reaction voltages together:



* Setup cell to be spontaneous

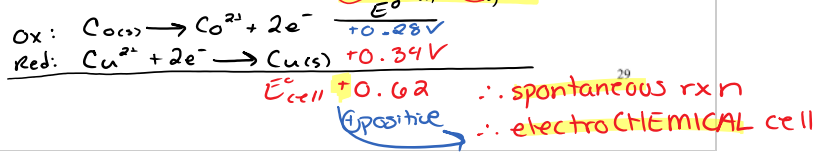
A spontaneous reaction has a positive E°_{cell} and it's an electrochemical cell.

A non-spontaneous reaction has a negative E°_{cell} or zero E°_{cell} and is an electrolytic cell (it won't occur unless supplied with voltage).

more on this later.

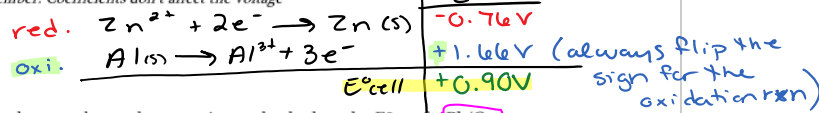
Practice Questions

1. Calculate the standard potential of the cell: $\text{Cu}^{2+} + \text{Co}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \text{Co}^{2+}$

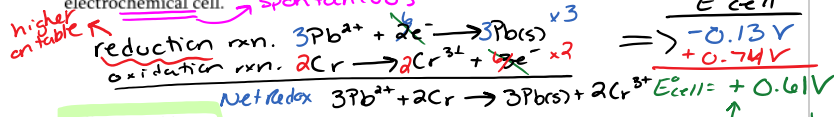


2. Calculate the potential of the cell: $3\text{Zn}^{2+} + 2\text{Al} \Rightarrow 3\text{Zn} + 2\text{Al}^{3+}$

*Remember: Coefficients don't affect the voltage



3. Write the complete redox equation and calculate the E°_{cell} of a Pb/Cr electrochemical cell.



Cells at Equilibrium

→ redox \neq equilibrium either FWD or REV. rxn. ↑ spontaneous.

The E°_{cell} of a cell is the voltage at standard state, so solution concentrations are 1.0M. As the cell operates, reactants are being used up, so their concentrations are continually decreasing, thus the voltage of the cell starts to decrease. When all reactants are used up, the cell will cease to operate (the battery is dead), and the cell voltage will be zero. At this point, the cell is said to be at equilibrium.

Therefore, the voltage of a cell at equilibrium is 0 (a dead battery). ← of a complete cell

Assignment 11: 1) Hebden p.224-225 #36abcdf, 37, p.226 #46

(36f is a disproportionation reaction)

XIII) Cell Potential Practice Questions

1. Draw and label an electrochemical cell using a copper anode and having an E°_{cell} greater than 1.0V (The product of the oxidation is Cu^{2+}).