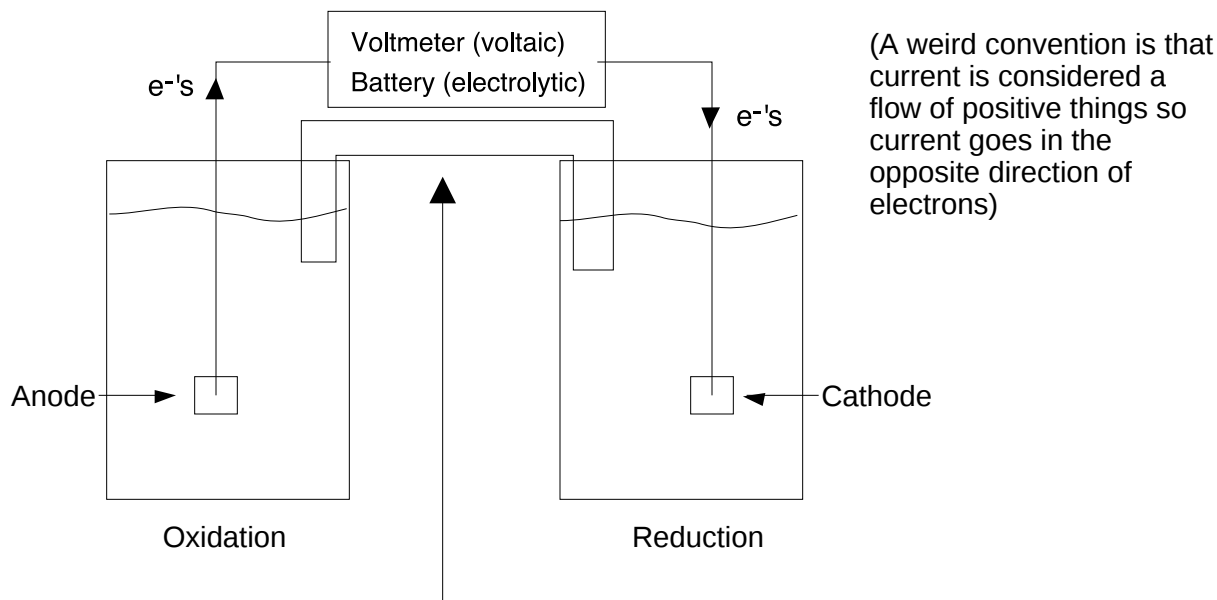


Student Packet: 8.2: Electrochemical Cells A: Student Handout AP[®] Chemistry

Electrochemical cells either use spontaneous redox reactions to generate a voltage (voltaic cells or batteries) or they use a voltage source to drive a nonspontaneous redox reaction (electrolytic cells). Basically you just separate the reactants physically, connect them with a wire and either drive something with the spontaneous flow of e⁻'s or drive the flow of e⁻'s for a nonspontaneous rxn.

Basically it looks like this:



This is called a salt bridge. It's just an ionic solid that provides a conductive return path for charge carriers so that the circuit can be complete. If a return path wasn't there, the cell would build up a charge separation and you'd get lightning (well maybe not lightning but it wouldn't work so well.)

The conventions for these cells can be very confusing because they change whether it's a voltaic (also called galvanic) versus electrolytic cell.

Here they are:

	<u>CATHODE</u>	<u>ANODE</u>
ions attracted	cations(+)	anions(-)
how electrons move	into the cell/out of the cathode	into the anode out of the cell
half-reaction	Reduction	Oxidation
electrode Sign		
Voltaic Cell	positive	negative
Electrolytic Cell	negative	positive

There are also some units you must know

Coulomb = Metric unit of charge: 1 mole of e⁻'s has a total charge of 96,485C. This is also called Faraday's constant (the number 96,485C).

Ampere = Unit of current: amperes = coulombs/second or more usefully coulombs = amps x se.conds
This measures how many e⁻'s were transferred and thus how much reaction took place.

Student Packet: 8.2: Electrochemical Cells A: Student Handout AP® Chemistry

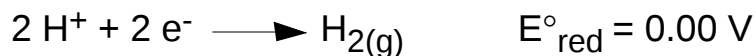
In electrochemistry you can take redox reactions or cell configurations and determine whether the reaction will occur spontaneously or if not, how much voltage is required to make it go. Just like for most of the problems you've been doing for the past two months there's tabulated data that you can use to analyze the system. These are called standard reduction potentials, E°_{red} and there are a couple of things you have to know.

1. All of the values are for **reduction half-reactions** so when you write out your two half reactions use the value off the table for the reduction reaction and **switch the sign of the one for the oxidation reaction.**

$$E^\circ_{\text{ox}} = -E^\circ_{\text{red}}$$

So the table really has all E°_{ox} and E°_{red} values; you just have to switch signs for E°_{ox} .

2. The table is set up with an arbitrary zero point for the reduction of hydrogen ion.



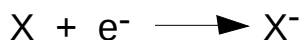
3. For the whole Redox reaction you just add the oxidation potential and the reduction potential.

$$E^\circ_{\text{TOT}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$$

From the value of E°_{TOT} for the reaction, you can tell whether it's spontaneous or not.

$E^\circ_{\text{TOT}} > 0$	Reaction is spontaneous (voltaic cell)
$E^\circ_{\text{TOT}} < 0$	Reaction is nonspontaneous (electrolytic cell) (You would have to supply a voltage of at least E°_{TOT} to make it go. The reverse reaction will be spontaneous.)
$E^\circ_{\text{TOT}} = 0$	rxn is at equilibrium

- The **more positive E°_{red} , the stronger the oxidizing agent** is.



- The **more positive E°_{ox} , the stronger the reducing agent** is.

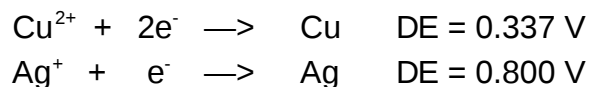
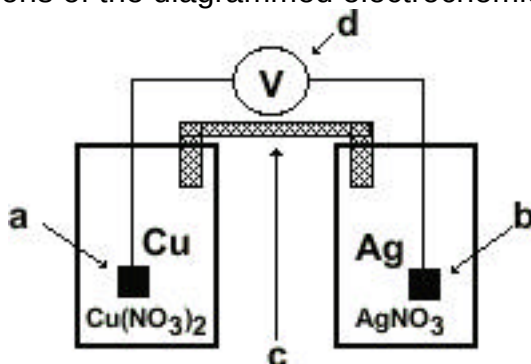


Student Packet: 8.2: Electrochemical Cells A: AssessMe questions

- 1) Select the set of coefficients which balances the chemical equation:
$$\text{B}_3\text{N}_3\text{H}_{12} + \text{O}_2 \longrightarrow \text{N}_2\text{O}_5 + \text{B}_2\text{O}_3 + \text{H}_2\text{O}$$
- A) 1,5,1,1,2
B) 1,15,3,3,6
C) 2,5,3,3,3
D) 2,18,3,3,12
E) I'm clueless
- 2) Which of the following ions can be oxidized by appropriate chemical means and reduced by a different chemical reaction?
- A) Fe^{2+}
B) F^-
C) CO_3^{2-}
D) NO_3^-
E) I'm clueless
- 3) For the reaction: $2\text{Fe} + 3\text{CdCl}_2 \rightleftharpoons 2\text{FeCl}_3 + 3\text{Cd}$, which statement is true?
- A) Fe is the oxidizing agent
B) Cd undergoes oxidation
C) Cd is the reducing agent
D) Fe undergoes oxidation
E) I'm clueless
- 4) In the reaction: $\text{SO}_2 + 2\text{H}_2\text{S} \longrightarrow 3\text{S} + 2\text{H}_2\text{O}$
- A) S is oxidized and H is reduced.
B) S is reduced and there is no oxidation.
C) S is reduced and H is oxidized.
D) S is both reduced and oxidized.
E) I'm clueless
- 5) When considering thermodynamics, it is often necessary to establish a "standard state" in temperature and pressure, (STP). The normal pressure condition employed is:
- A) 1 atm
B) 10^6 Pa
C) 25 torr
D) 1000 mm of Hg
E) I'm clueless
- 6) Which statement is true?
- A) At equilibrium no chemical reaction occurs.
B) At equilibrium the forward rate of reaction and the backward rate of reaction are equal.
C) Chemical reactions are never exactly in equilibrium since reaction rates slow as one approaches equilibrium.
D) All of the above statements are true.
E) I'm clueless

Student Packet: 8.2: Electrochemical Cells A: ShowMe questions

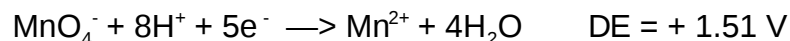
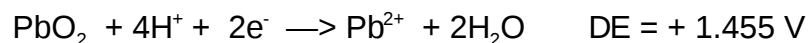
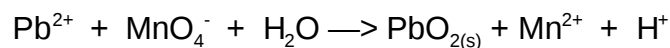
1) Identify the portions of the diagrammed electrochemical voltaic cell:



- A) 1 : salt bridge
 B) 2 : potentiometer
 C) 3 : anode
 D) 4 : cathode

2) Reduction of Ag^+ is accomplished at a Pt electrode through electrolysis. 1.3 g of Ag is deposited in 626 seconds. How much charge was passed? What was the magnitude of the current?

3) Calculate the equilibrium constant for:



4) What is the emf for a cell constructed from a cadmium electrode and a silver electrode?
 ($\text{Cd}/\text{Cd}^{2+}||\text{Ag}^+/\text{Ag}$)



Student Packet: 8.2: Electrochemical Cells A: TestMe questions

- 1) An aqueous solution of CuSO_4 is electrolyzed to deposit copper metal. If 0.1A is passed for 5 hr, how much Cu metal is deposited?

A) 0.296 g B) 0.593 g C) 1.18 g D) 1.32 g E) 0.484 g

- 2) What is the equilibrium constant for this cell?



A) 1×10^{38} B) 1×10^{72} C) 1×10^{92} D) 1×10^{108} E) 272

- 3) What is the net reaction for this cell:

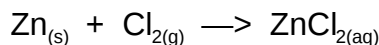


- A) $\text{Br}_2 + \text{H}_2 \longrightarrow 2\text{H}^+ + 2\text{Br}^-$
 B) $2\text{Br}_2 + \text{H}_2 \longrightarrow 2\text{H}^+ + 4\text{Br}^-$
 C) $\text{Br}_2 + 2\text{H}^+ \longrightarrow 2\text{Br}^- + \text{H}_2$
 D) $\text{H}_2\text{O} + \text{Pt} + 2\text{H}^+ \longrightarrow \text{Pt}^{4+} + 2\text{OH}^-$
 E) $2\text{H}^+ + 2\text{Br}^- \longrightarrow \text{Br}_2 + \text{H}_2$

- 4) If a given cell has an emf of 1.6 V, and each half reaction involves two electrons, what is the standard free energy change for this cell?

A) -309 J B) 309 J C) 154.5 kJ D) -154.5 kJ E) -309 kJ

- 5) What is the potential of a voltaic cell using the reaction:



if the free energy charge is calculated to be -409.7 kJ?

A) 1.080 V B) 1.349 V C) 2.005 V D) 2.123 V E) 2.364 V