

## Electrochemical Conventions

(modified from Dr. Fred Wood)  
UC Davis

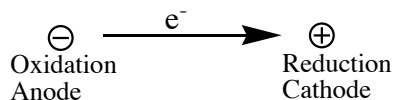
### I. General Oxidation-Reduction (Redox)

- a) Oxidation = loss of electrons. (The charge (oxid. state) is increased by losing electrons.)
- b) Reduction = gain of electrons. (The charge (oxid. state) is reduced by gaining electrons.)
- c) Oxidizing Agent:  
causes an oxidation  
takes electrons  
is itself reduced
- d) Reducing Agent:  
causes a reduction  
gives electrons  
is itself oxidized

### II. Electrochemistry

- a) Electrodes = Where the oxidation/reduction (redox) reactions occur  
Anode = Where oxidation half reaction occurs. (begin with vowels).  
Cathode = Where reduction half reaction occurs. (begin with consonants).
- b) In the external circuit (*ie*, the wire) the electrons always flow from the oxidation site to the reduction site. Therefore, electrons in the external circuit always flow from the anode to the cathode.
- c) Voltaic (Galvanic) Cells- an electrochemical cell in which a spontaneous reaction produces electricity. The voltage is always positive. (If you got a problem where the voltage was negative, it means the cell is spontaneous in the reverse direction.)

- Electrons move in the external circuit from the negative electrode to the positive electrode.

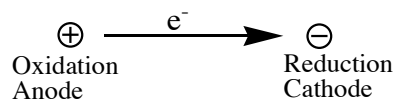


- Thus, in the voltaic cell, anode = negative ( $\ominus$ ) and cathode = positive ( $\oplus$ )

d) Electrolytic Cells- an electrochemical cell in which a non-spontaneous reaction is carried out by electrolysis. (Electrolysis: the decomposition of a substance (in a molten state or in an electrolytic solution) by an electrical current.) The voltage is always negative and a metal is plated out or a gas is evolved.

- Electrons are forced to move in the external circuit from the positive electrode to the negative electrode.

- Thus, in the electrolytic cell, anode = positive( $\oplus$ ) and cathode = negative( $\ominus$ )



e) Drawing cells (both voltaic and electrolytic)

The negative electrode (not necessarily the anode) is shown on the LEFT

Thus, the left-hand electrode is the:

- anode if voltaic cell
- cathode if electrolytic cell

f) Line Diagrams (Cell Diagrams)

- the left hand electrode is the anode (where oxidation occurs)
- The sign on the left hand electrode is negative
- a boundary between different phases is represented by a single line (|).
- reactants are at the left of the double line (||) which represents the boundary between the half cells (usually a salt bridge); products are at the right of ||.
- Different species in the same solution half cell compartment are separated by commas.  $\text{Pt}_{(s)} | \text{Cl}_{2(g)} | \text{Cl}^{-}_{(aq)} || \text{Pb}^{2+}_{(aq)}, \text{H}^{+}_{(aq)} | \text{PbO}_{2(s)}$
- (some professors do not care about the order of the ions presented in a line diagram, while others say that, ions can be ordered in each half cell so that increasingly more positive ions are nearest the double lines. Example for a voltaic cell:



g) Standard Reduction Potentials (or voltages)

- voltages of a half cell in which all gases are 1 atm and all solutions are 1M (25°C)
- half cells are listed as reduction processes.
- the more positive the voltage, the more easily reduced.
- thus, the strongest oxidizing agents are found at the top left (as  $\text{F}_{2(g)}$ ) and the strongest reducing agents are found at the bottom right (as  $\text{Li}_{(s)}$ )
- a useful memory aid is “upper left reacts with lower right.”