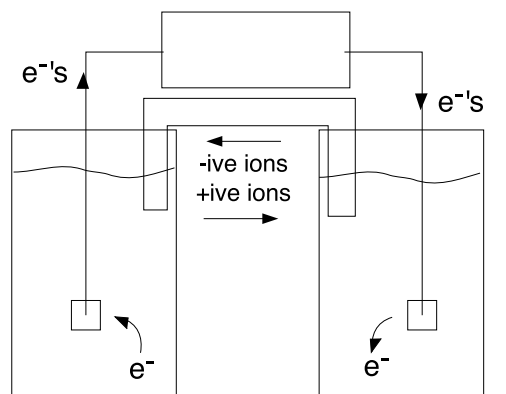


Student Packet: 8.3: Electrochemical Cells B: Student Handout AP[®] Chemistry

Since voltaic and electrolytic cells are hard to get straight, here's a picture of each.

Voltaic Cell

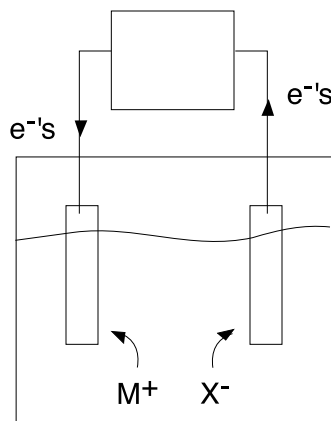


Anode (-)
Oxidation

Cathode (+)
Reduction

The desire of the electrons to go from the oxidized species to the reduced species drives the circuit.

Electrolytic Cell

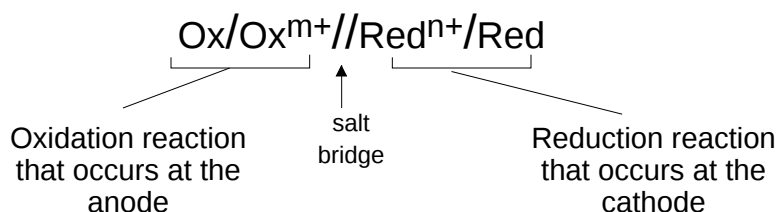


Cathode (-)
Reduction

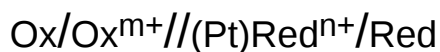
Anode (+)
Oxidation

The battery pumps electrons from the anode to the cathode. It sucks them in at the anode and pushes them out the cathode.

- Cell abbreviations are used instead of having to draw a complete cell picture every time you do an electrochemistry problem. The convention is as follows



If an inert electrode (platinum) is used to deliver or accept electrons you just put (Pt) in for the electrode you are talking about. Like let's say the cathode is made of Pt then



You do all this stuff so that you can use tabulated data of half-reactions and determine if a reaction will be spontaneous or not.

Student Packet: 8.3: Electrochemical Cells B: Student Handout AP[®] Chemistry

The last thing to look at in electrochemistry is how values of E°_{red} change with concentration (standard reduction potentials are taken at 1M concentration).

This can be calculated from the **Nernst Equation**

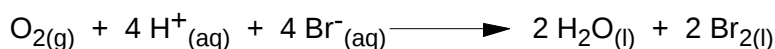
$$E = E^\circ_{\text{TOT}} - \frac{0.0591}{n} \log Q$$

E°_{TOT} = the standard reaction potential

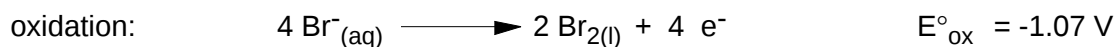
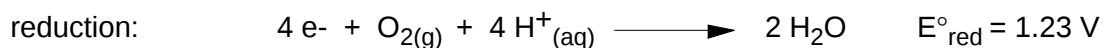
Q = reaction quotient (just the K_c expression at whatever concentrations are given. Remember, you have to use molarity for aqueous solutions and partial pressures for gases and that pure liquids and solids appear as 1.)

n = The number of electrons transferred for the given equation.

For example: Find the potential for the following reaction when $[H^+] = [Br^-] = .1M$ and $P_{O_2} = 1\text{atm}$.



A) first find E°_{TOT}



$$E^\circ_{\text{TOT}} = 0.16 \text{ V}$$

B) Calculate E at $[H^+] = [Br^-] = .1M$ and $P_{O_2} = 1\text{atm}$

$$E = E^\circ_{\text{TOT}} - \frac{0.0591}{n} \log Q = E^\circ_{\text{TOT}} - \frac{0.0591}{4} \log \frac{1}{(P_{O_2})[H^+]^4[Br^-]^4}$$

$$E = .16 \text{ V} - \frac{0.0591}{4} \log \frac{1}{(1)[.1]^4[.1]^4}$$

$$E = .16 \text{ V} - 0.1182 \text{ V} = 0.042 \text{ V}$$

NOTE: You can get K_c from this expression because at equilibrium $E = 0$.

$$0 = E^\circ_{\text{TOT}} - \frac{0.0591}{n} \log K_c$$

Student Packet: 8.3: Electrochemical Cells B: AssessMe questions

- 1) What is the potential in volts for the spontaneous reaction between the Ag/Ag⁺ and Zn/Zn²⁺ half-cells?



- A) -2.361 V B) -1.562 V C) 1.562 V D) 2.361 V E) I'm clueless

- 2) What is the [Cu²⁺] in the cell Zn / Zn²⁺ (0.05M) // Cu²⁺ (? M) / Cu if the cell voltage is 1.10 V?

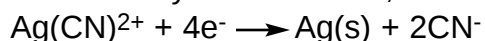


- A) 0.12 M B) 0.0002 M C) 0.05 M D) 0.0035 M E) I'm clueless

- 3) 10 amps are passed through molten aluminum chloride for 5.5 hours. How many grams of aluminum metal could be produced by this electrolysis?

- A) 18.5 g B) 55.4 g C) 91.2 g D) 273 g E) I'm clueless

- 4) In the electroplating of silver from cyanide solution, the cathode reaction is:



How many grams of silver should be deposited by a current of 4.5 A in 28.0 minutes?

- A) 0.141 g B) 4.23 g C) 2.11 g D) 12.53 g E) I'm clueless

- 5) Consider a solution in which the pOH is measured to be 6. What is the hydrogen ion concentration?

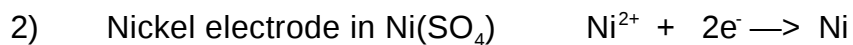
- (A) 1×10^{-6} (B) 6 (C) 8 (D) 1×10^{-8} (E) I'm clueless

- 6) The autoionization pK value for D₂O is 15 at room temperature. Which is a stronger acid, water or deuterium oxide?

- A) Water
 B) Deuterium oxide
 C) They are equivalent acids.
 D) Neither of these species is an acid.
 E) I'm clueless

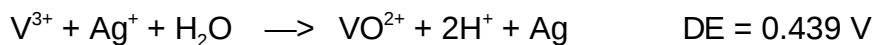
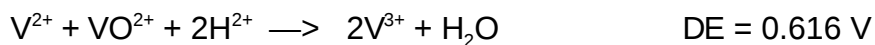
Student Packet: 8.3: Electrochemical Cells B: ShowMe questions

1) A voltaic cell is constructed from two half cells:



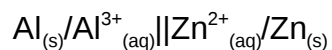
Which electrode is the anode, which is the cathode?

2) Given:



Calculate DE for $\text{V}^{3+} + \text{e}^- \longrightarrow \text{V}^{2+}$

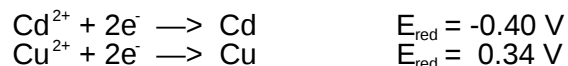
3) What is the net reaction for this electrochemical cell?



4) An aqueous solution of KBr is electrolyzed. What are the expected reactions at each electrode, and what is the overall reaction?

Student Packet: 8.3: Electrochemical Cells B: TestMe questions

1) What is the emf produced by a voltaic cell that uses Cd and Cu electrodes?



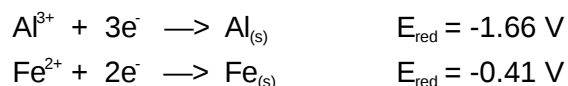
- A) -0.06 V B) 0.06 V C) 0.34 V D) 0.40 V E) 0.74 V

2) What is the cell diagram for this cell?

(get picture)

- A) $\text{H}^+ || \text{Ag}$
 B) $\text{H}_2/\text{H}^+ || \text{Ag}/\text{Ag}^+$
 C) $\text{Pt}/\text{H}_2/\text{H}^+ || \text{Ag}^+/\text{Ag}$
 D) $\text{Pt}/\text{H}^+ || \text{Ag}^+/\text{Ag}$
 E) $\text{Ag}/\text{Ag}^+ || \text{H}^+/\text{H}_2/\text{Pt}$

3) What is the cell potential for:

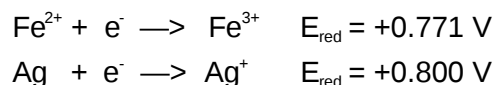
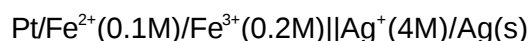


- A) -2.07 V B) -1.25 V C) 2.51 V D) 1.25 V E) 2.07 V

4) An aqueous solution of CuCl_2 is electrolyzed; another aqueous solution of CuSO_4 is also electrolyzed. What is the difference between these two experiments?

- A) Cu^{2+} is not reduced in CuSO_4 solution.
 B) Cu^{2+} is not reduced in CuCl_2 solution.
 C) SO_4^{2-} is oxidized, but Cl^- is not.
 D) Cl^- is oxidized, but SO_4^{2-} is not.
 E) Nothing: both cells will perform the same chemistry.

5) What is the emf for the following electrochemical cell?



- A) -0.029 V B) 0.0 V C) 0.029 V D) 0.047 V E) 0.063 V