

Standard Redox Potentials, E°

Ox. agent	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
	$2H^+ + 2e^- \rightarrow H_2$	0.00
	$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76 Red. agent

Any substance on the right will reduce any substance higher than it on the left.

Northwest-southeast rule: product-favored reactions occur between

- reducing agent at southeast corner
- oxidizing agent at northwest corner

Using Standard Potentials, E°

Table 20.1

- In which direction do the following reactions go?
- $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
 - Goes right as written
- $2Fe^{2+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Sn(s)$
 - Goes LEFT opposite to direction written
- What is E°_{net} for the overall reaction?

Standard Redox Potentials, E°

CATHODE	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
	$2H^+ + 2e^- \rightarrow H_2$	0.00
	$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76 ANODE

Northwest-southeast rule:

- reducing agent at southeast corner = ANODE
- oxidizing agent at northwest corner = CATHODE

Standard Redox Potentials, E°

E°_{net} = "distance" from "top" half-reaction (cathode) to "bottom" half-reaction (anode)

$$E^\circ_{net} = E^\circ_{cathode} - E^\circ_{anode}$$

"Distance" from $E^\circ_{cathode}$ to E°_{anode} is $0.80V - 0.337V = 0.46V$.

$E^\circ_{cathode} = +0.80V$	$Ag^+(aq) + e^- \rightarrow Ag(s)$
$E^\circ_{anode} = +0.337V$	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Cu is "southeast" of Ag^+

E°_{net} for Cu/Ag^+ reaction = +0.46 V

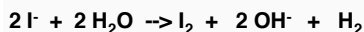
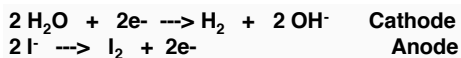
Echem Cell Notation

- Cells are given in an abbreviated notation:

Oxidation (anode) || Reduction (cathode)
- Where the || means two phase boundaries, such one at each end of a salt bridge.
- $Cu | CuSO_4 (1.0 M) || AgNO_3 (1.0 M) | Ag(s)$
- $E^\circ_{cell} = E^\circ_{right} - E^\circ_{left}$

More About Calculating Cell Voltage

Assume I⁻ ion can reduce water.



Assuming reaction occurs as written,

$$E^\circ_{\text{net}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

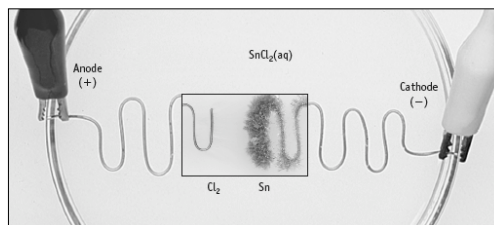
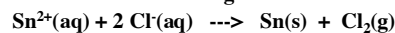
$$= (-0.828 \text{ V}) - (+0.535 \text{ V}) = -1.363 \text{ V}$$

Minus E^o means rxn. occurs in opposite direction

Electrolysis is required to force the I⁻ + H₂O reaction.

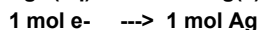
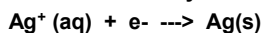
Electrolysis

Using electrical energy to produce chemical change.



Quantitative Aspects of Electrochemistry

Consider electrolysis of aqueous silver ion.



If we could measure the moles of e⁻, we could know the quantity of Ag formed.

But how to measure moles of e⁻?

$$\text{Current} = \frac{\text{charge passing}}{\text{time}}$$

$$I \text{ (amps)} = \frac{\text{coulombs}}{\text{seconds}}$$

Quantitative Aspects of Electrochemistry

$$I \text{ (amps)} = \frac{\text{coulombs}}{\text{seconds}} \quad \text{Current} = \frac{\text{charge passing}}{\text{time}}$$

But how is charge related to moles of electrons?

$$= 96,500 \text{ C/mol e}^-$$

$$= 1 \text{ Faraday}$$



Michael Faraday
1791-1867

Quantitative Aspects of Electrochemistry

$$I \text{ (amps)} = \frac{\text{coulombs}}{\text{seconds}}$$

1.50 amps flow thru a Ag⁺(aq) solution for 15.0 min. What mass of Ag metal is deposited?

Solution

(a) Calc. charge

$$\text{Charge (C)} = \text{current (A)} \times \text{time (t)}$$

$$= (1.5 \text{ amps})(15.0 \text{ min})(60 \text{ s/min}) = 1350 \text{ C}$$

Quantitative Aspects of Electrochemistry

$$I \text{ (amps)} = \frac{\text{coulombs}}{\text{seconds}}$$

1.50 amps flow thru a Ag⁺(aq) solution for 15.0 min. What mass of Ag metal is deposited?

Solution

(a) Charge = 1350 C

(b) Calculate moles of e⁻ used

$$1350 \text{ C} \cdot \frac{1 \text{ mol e}^-}{96,500 \text{ C}} = 0.0140 \text{ mol e}^-$$

(c) Calc. quantity of Ag

$$0.0140 \text{ mol e}^- \cdot \frac{1 \text{ mol Ag}}{1 \text{ mol e}^-} = 0.0140 \text{ mol Ag or } 1.51 \text{ g Ag}$$

Quantitative Aspects of Electrochemistry

The anode reaction in a lead storage battery is
 $\text{Pb(s)} + \text{HSO}_4^-(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2\text{e}^-$
If a battery delivers 1.50 amp, and you have 454 g of
Pb, how long will the battery last?

Solution

a) $454 \text{ g Pb} = 2.19 \text{ mol Pb}$

b) Calculate moles of e^-

$$\frac{2.19 \text{ mol Pb} \cdot 2 \text{ mol e}^-}{1 \text{ mol Pb}} = 4.38 \text{ mol e}^-$$

c) Calculate charge

$$4.38 \text{ mol e}^- \cdot 96,500 \text{ C/mol e}^- = 423,000 \text{ C}$$

Quantitative Aspects of Electrochemistry

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If a battery delivers 1.50 amp, and you have 454 g of
Pb, how long will the battery last?

Solution

a) $454 \text{ g Pb} = 2.19 \text{ mol Pb}$

b) Mol of $\text{e}^- = 4.38 \text{ mol}$

c) Charge = 423,000 C

d) Calculate time $\text{Time (s)} = \frac{\text{Charge (C)}}{I (\text{amps})}$

$$\text{Time (s)} = \frac{423,000 \text{ C}}{1.50 \text{ amp}} = 282,000 \text{ s} \quad \text{About 78 hours}$$