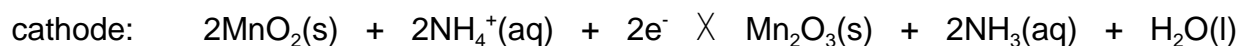
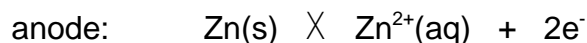


Batteries

(cells with spontaneous redox reactions)

Dry Cell

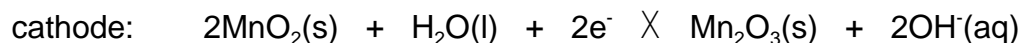
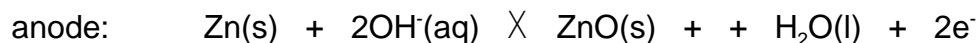


$E_{\text{cell}} = 1.5 \text{ V}$ (Indep. of cell size. Current output depends on cell size. As cell vol. increases, R_{int} decreases and I_{cell} increases; $V=IR$)

Problems: Can't recharge because Zn^{2+} diffuses away.
Acidic NH_4^+ corrodes Zn, leading to poor shelf life.

Alkaline Battery

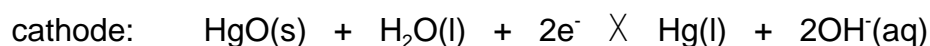
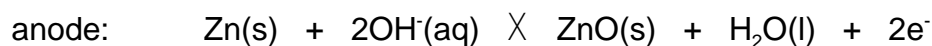
uses KOH instead of NH_4^+ and ZnCl_2



$E_{\text{cell}} = 1.54 \text{ V}$

lasts longer, but slightly more expensive because of the construction

Mercury Battery



$E_{\text{cell}} = 1.34 \text{ V}$

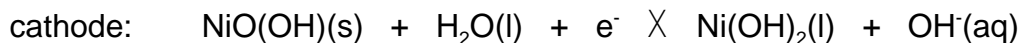
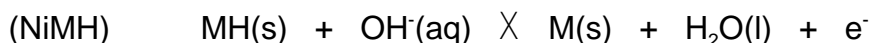
More current for a given size (used for calculators, watches, hearing aids, cameras).

More expensive because of the Hg.

E_{cell} constant during discharge because cell reaction doesn't change cell electrolyte composition.

Not rechargeable. Disposal of Hg is also a problem.

Nickel-Cadmium (nicad) & Nickel-Metal Hydride (NiMH) Batteries



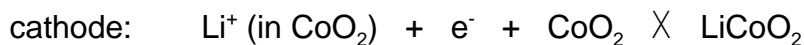
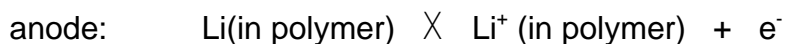
$E_{\text{cell}} = 1.2\text{-}1.4 \text{ V}$ (fairly constant during discharge)

Products adhere to electrodes - reactions can be reversed by charging.

Disposal of cadmium a problem (nicad).

Used in calculators, cordless appliances and portable power tools.

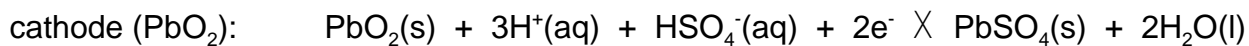
Lithium Ion Battery



Rechargeable, long life, higher voltage ($E_{\text{cell}} = 3.4 \text{ V}$).

Used in portable computers, cellular phones, cameras.

Lead Storage Battery



$E_{\text{cell}} = 2.0 \text{ V}$

Car battery has six cells in series to provide 12 V.

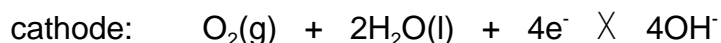
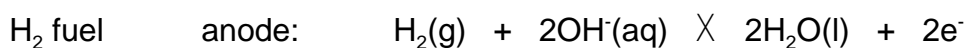
Rechargeable, since PbSO_4 sticks to electrodes.

Fuel Cells

Involve the direct conversion of chemical energy \rightarrow electrical energy (~70% efficiency)
 (otherwise, chemical \rightarrow heat \rightarrow electrical energy is only ~ 30% efficient)

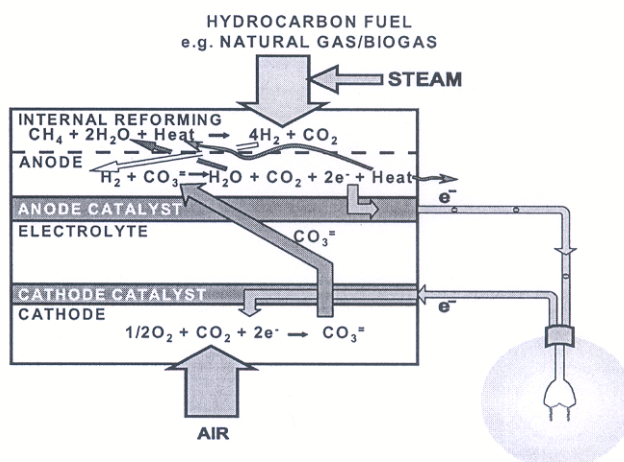
Reactants are continuously supplied.

Example:



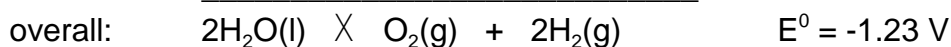
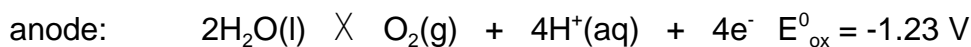
$$E^0_{\text{cell}} = 1.23 \text{ V} \quad (E_{\text{cell}} \sim 0.7 \text{ V})$$

The Direct Fuel Cell installation on campus works in the following way.

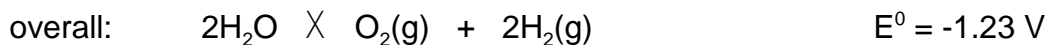
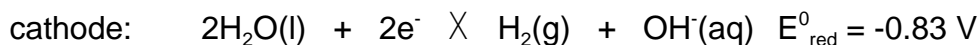
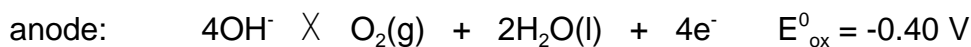


Electrolysis - reactant-favored reactions

H_2O (acidic solution)



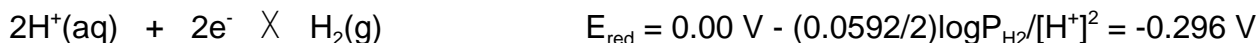
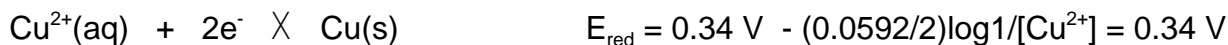
H₂O (basic/neutral solution)



Must have mobile ions in solution so that charge balance can be maintained.

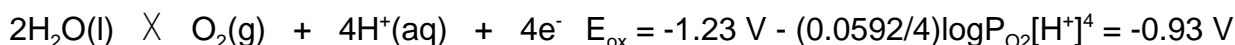
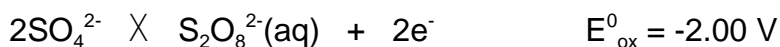
Electrolysis of 1.0 M CuSO₄ at pH = 5.0, P_{H₂} = 1 atm, P_{O₂} = 1 atm

possible cathode reactions



Cu²⁺ is reduced (Cu will plate out on cathode) because it has the more favorable E_{red}.

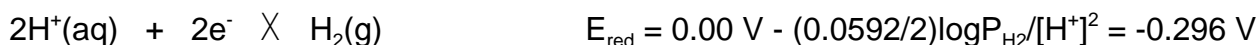
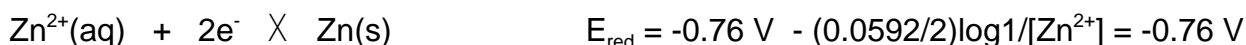
possible anode reactions



O₂ is produced at the anode because it has the more favorable E_{ox}.

Electrolysis of 1.0 M ZnSO₄ at pH = 5.0, P_{H₂} = 1 atm, P_{O₂} = 1 atm

possible cathode reactions



H₂ should form (it has the more favorable E_{red}), but Zn plates out due to **overvoltage**.

For Zn: $E_{\text{red}} = -0.76 \text{ V} - \text{overvoltage}(\text{Zn})$

For H₂: $E_{\text{red}} = -0.296 \text{ V} - \text{overvoltage}(\text{H}_2)$

Overvoltage is a kinetic effect. Electron transfer is inhibited at the electrode. It is usually small for the deposition of metals, but often large (~1 V) for the evolution of gases. This explains why you can plate out Cr³⁺ and Ni²⁺, and produce Cl₂(g) in the electrolysis of NaCl(aq), even though they have unfavorable half-cell potentials.