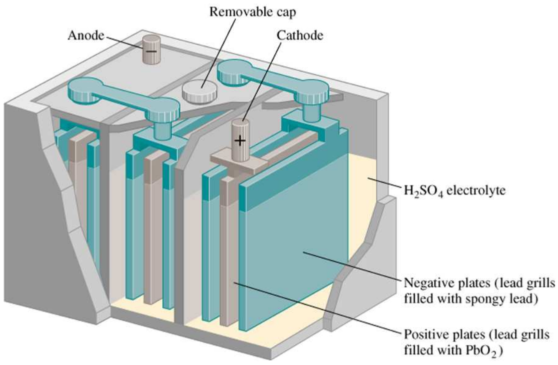
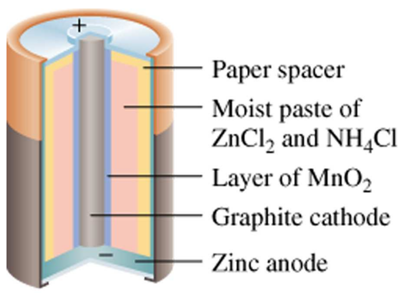
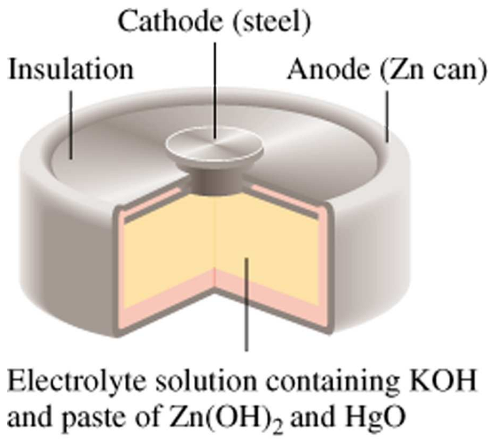
Notes on Chapter 17

Commercial Batteries

A Battery is a galvanic cell or several galvanic cells in series. The following table summarizes information on some of the well known batteries.

Type	Reaction	Comments
Lead storage	$Pb + HSO_4^- \rightarrow PbSO_4 + H^+ + 2e^-$ $PbO_2 + HSO_4^- + 3H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$ 	$\varepsilon_{cell} \approx 2$ V. Six cells in series yield about 12 V. Checked using H_2SO_4 density. Rechargeable. Figure 17.13
Dry Cell (Leclanche)	$Zn \rightarrow Zn^{2+} + 2e^-$ $2NH_4^+ + 2MnO_2 + 2e^- \rightarrow Mn_2O_3 + 2NH_3 + H_2O$ 	$\varepsilon_{cell} \approx 1.5$ V. Zn corrodes fast in NH4Cl (source of $NH4^+$). Figure 17.14
Dry Cell (alkaline)	$Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$ $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$	KOH (or NaOH) replaces NH4Cl. Last longer than Leclanche as Zn corrodes much more slowly

Ag or Hg Cells	Both use Zn anode. An Ag cell uses Ag_2O as an oxidizing agent in a basic medium. An Hg cell uses HgO in a basic medium 	Often used in calculators. Details in Figure 17.15
Ni –Cd (Nicad)	$Cd + 2OH^- \rightarrow Cd(OH)_2 + 2e^-$ $NiO2 + 2H_2O + 2e^- \rightarrow Ni(OH)_2 + 2OH^-$	Theoretically indefinitely rechargeable

<u>Aluminium Production (see figure 17.22)</u>

The reactions involved leading to the final product are believed to occur as follows.

CATHODE: (1) $AlF_6^{3-} + 3e^- \rightarrow Al + 6F^-$

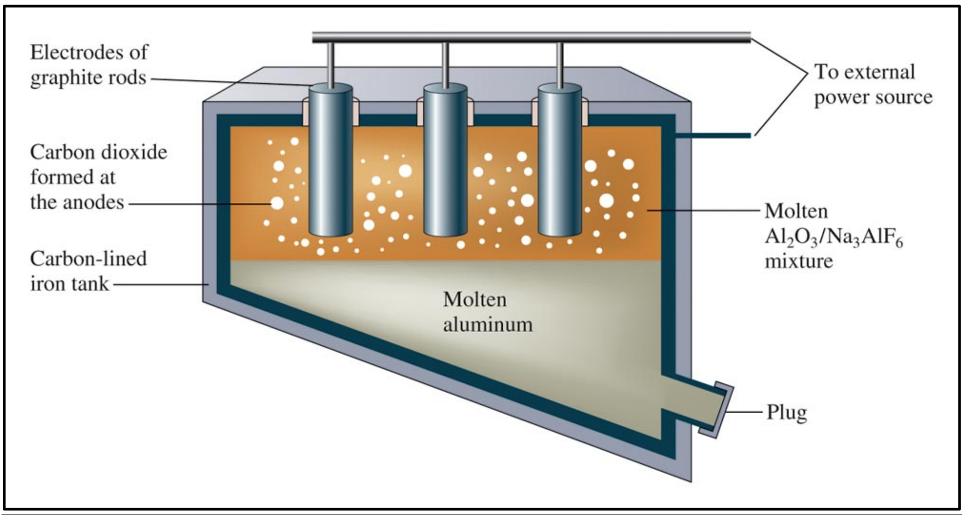
ANODE: (2) $AlOF_6^{2-} + 12\,F^- + C \rightarrow 4\,AlF_6^{3-} + CO_2 + 4\,e^-$

One of the products, AlF_6^{3} , is believed to react with Al_2O_3 as follows

Reaction (3) $Al_2O_3 + 4\,AlF_6^{3-} \rightarrow 3\,AlOF_6^{2-} + 6F^-$

Multiplying (1) by 4 and (2) by 3 to balance the electrons and adding them to (3) multiplied by yields the following overall reaction.

$$2\,Al_2O_3 + 3\,C \rightarrow 4\,Al + 3\,CO_2$$



Electro-refining

After Cu is obtained from its ore its still contains impurities like Fe, Zn , and other metals M. This impure Cu is used as anode and a very pure copper is used as a cathode.

Anode half reaction $Cu \rightarrow Cu^{2+} + 2\,e^-$

Cathode half reaction $Cu^{2+} + 2\,e^- \rightarrow Cu$

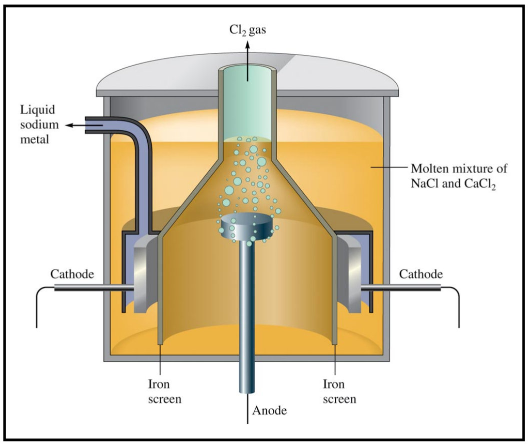
Thus basically the Cu is transferred from the impure anode and collected pure on the cathode. The other impurities like Fe, Zn etc usually leave the anode as Fe^{2+} and Zn^{2+} and so on but remain as sludge and are not deposited on the cathode because their reduction potential is lower than that of Cu^{2+}.

Metal Plating (see figure 17.24 b)

The metal (Ag in the case of figure 17.24 b) leaves the anode as Ag^+ and Ag^+ from the solution is deposited as Ag on the item to be plated (a silver spoon in the case of figure 17.24) which acts as a cathode.

Production of Na and Cl_2 in a Downs Cell (figure 17.25)

A molten NaCl / $CaCl_2$ mixture is electrolyzed.



ANODE $2\,Cl^-(aq) + 2\,e^- \rightarrow Cl_2(g)$

CATHODE $Na^+(aq) + e^- \rightarrow Na(s)$

Electrolysis of NaCl (aq) does not yield Na because H_2O which has a higher reduction potential than Na^+ is preferentially reduced.

(1) $Na^+(aq) + e^- \rightarrow Na(s) \qquad \varepsilon^o = -2.71\ V$

(2) $2\,H_2O + 2\,e^- \rightarrow H_2 + 2\,OH^- \qquad \varepsilon^o = -0.83\ V$

In the Hg cell (figure 17.26) a Hg electrode is used on which half-cell reaction (1) occurs because the production of H_2 on Hg requires a very high overvoltage. The NaHg amalgam forms then the Na is removed from it by the following reaction with water.

$$2\,Na(s) + 2\,H_2O(l) \rightarrow 2\,Na^+(aq) + 2\,OH^-(aq) + H_2(g)$$

Using electrolysis of NaCl (aq) to get Cl_2 and Na is possible (done in Japan) by surrounding the cathode with a membrane that only passes the Na^+ ions to it.