b) Batteries (V.11)

i) "Eveready" (Zinc-carbon battery)

- carbon cathode and zinc casing acts as anode
- NH_4Cl / MnO_2 paste inside

- carbon cathode is inert: a place for $Mn^{+4} + e^- \rightarrow Mn^{+2}$ to occur
- zinc anode: $Zn \rightarrow Zn^{+2} + 2e^-$

- why do batteries die?

- E°cell is initial cell voltage
- Overtime, the redox reaction reaches equilibrium.
- Equilibrium is when no more reactant is left or electrode gets coated with by-product.
- In a battery, the battery dies when you run out of zinc or electrode get covered with $Zn(NH_4)_4^{+2}$
- No zinc? No electrode access? E°cell = 0 V

ii) "Duracell or Energizer" (Alkaline Battery)

- the same set up as the zinc-carbon battery
- except: NH_4Cl is replaced by KOH base (alkaline!!)

- lasts longer cause no $Zn(NH_4)_4^{+2}$ build up at electrodes.

iii) "Car Battery" (Lead-Acid Battery)

- anode is solid lead: $Pb_{(s)} \rightarrow Pb^{+2} + 2e^-$
- cathode is PbO_2: $Pb^{+4} + 2e^- \rightarrow Pb^{+2}$

- both sit in H_2SO_4 electrolyte

- why can car batteries be *recharged*?

- the battery dies cause of $PbSO_4$ buildup on the electrodes
- "boosting" the car, sends electricity through the battery reversing the reaction!
- $PbSO_4 + H_2O \rightarrow Pb_{(s)} + PbO_{2\,(s)} + 2H_2SO_4$

NOTE: above is a ***"Disproportionation"*** reaction. (the same reactant is both ox. and red.) The Pb^{+2} oxidizes to Pb^{+4} and the Pb^{+2} reduces to $Pb_{(s)}$

- why do car batteries eventually die?

- $PbSO_4$ can fall off the electrodes, so it can't be made back into Pb and PbO_2

iv) Fuel Cells –see page 233 Hebden

c) Corrosion (V.12)

i) What is Corrosion?

- Oxidation of metals
- "Rusting" is oxidation of iron

ii) How Does "Rusting" Occur? (See page 233-234 Hebden)

① The "electrochemical cell" is a drop of water:

Anode is the center of the drop.	Cathode is the edge of the drop.
$Fe_{(s)} \rightarrow Fe^{+2} + 2e-$	$\frac{1}{2}O_2 + H_2O + 2e- \rightarrow 2OH^-$

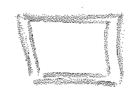
② The Fe^{+2} reacts with the OH^- to make $Fe(OH)_{2\,(s)}$

③ The $Fe(OH)_2$ is then oxidized to $Fe_2O_{3\,(s)}$ Rust!

iii) How Can we Stop Corrosion of Iron?

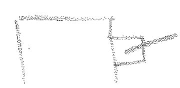
① *Protect the metal surface:*

cover



- paint
- "protective oxide" …coat iron with a thin layer of magnesium or tin
  …the Mg or Sn will oxidize, but cover the iron!
  …"stainless steel" is steel with a thin SnO coat

② *Cathodic Protection*



small piece (not covered)

- attach a sacrificial anode of zinc
  (or a substance more easily oxidized than iron)
- the zinc will oxidize preferentially, leaving the iron alone.
- think what we learned from the section on *"Multiple Electrodes"* above!

③ *Shift Reactions in Reverse*

- add OH^- or remove O_2 from the situation
- the cathode reaction above will shift to the left, decreasing oxidation!

Read through section V.11 and V.12
Do Questions: #49 page 229; #52-53 page 231; #56 page 233; #57-63 page 234 & 236