A. $Ch\ 4$ —
   55, 57, 61, 63 - see text answers

B. Ch 11 — 19 a - $Br_2 \rightarrow Br^-$ must be reversed - so $\mathcal{E}^{o}_{ox} = -1.09 V$

$Cl_2 + 2Br^- \longrightarrow 2Cl^- + Br_2$
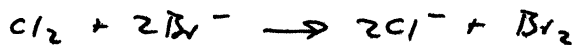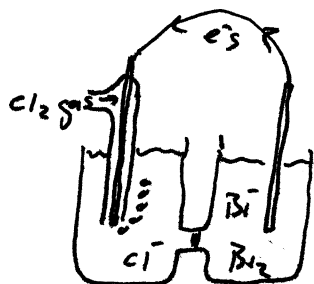
$\mathcal{E}^o = 1.36 - 1.09 = 0.27 V$

Pt electrodes
- both sides

Cl_2 gas reduced to Cl^- — cathode

Br^- oxidized to Br_2 — anode

Cl^- would migrate $L \rightarrow R$; pos counterion of Br^- (e.g. Na^+ or K^+) would migrate $R \rightarrow L$ via salt bridge.

b - see text answer

c - skip this one - there are too many puzzling aspects to this.

d. - see text answer - note that we can use Mn & Fe metals as electrodes here.

23, 25 - see text ans.

26 - a. strongest "oxidizing agent" is the species that is most easily reduced - Ag^+

b. similarly, we're looking for the species that is easiest to oxidize - Zn (that's Zn metal, not Zn^{2+}!)

c. SO_4^{2-} reduction potential is $+0.20 V$ - we're looking for species with oxidation potentials $> -0.20 V$ - only Zn & Pb qualify

d. Al metal $\xrightarrow{ox} Al^{3+}$ $+1.66 V$ - we're looking for things with reduction potentials $> -1.66 V$ - Ag^+, Zn^{2+}

27, 29 - see text answer.

60 - $M^{+n} \xrightarrow{+ne^-} M$
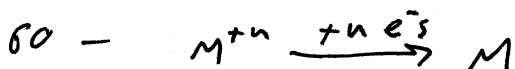
mass of $M(g) \times \dfrac{1\ mol\ M}{AW\ g\ M} \times \dfrac{n\ mol\ e^-}{1\ mol\ M} \times \dfrac{96,485\ C}{1\ mol\ e^-} \times \dfrac{1\ sec}{100\ C}$
(atomic wt.)

1 Amp $= 1\ C/sec$

60 (cont'd) a. $1000 \, g$, $n = 3$, Al $26.98 \, g/mol$

$$107{,}300 \text{ sec} \times \frac{1 \text{ min}}{60 \text{ sec}} \times \frac{1 \text{ hr}}{60 \text{ min}} = 29.8 \text{ hrs} \, !$$
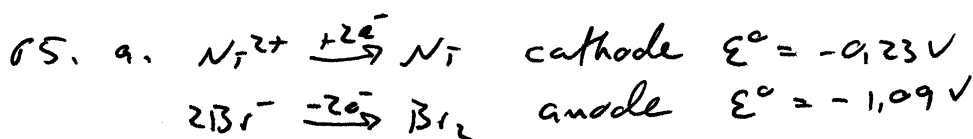
or, only 2 sig figs---
30 hrs !

b. $1 \, g$, $n = 2$, Ni $58.69 \, g/mol$

33 sec.
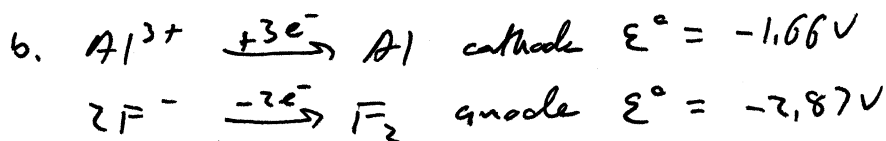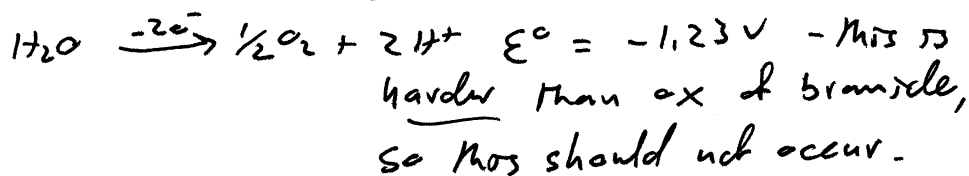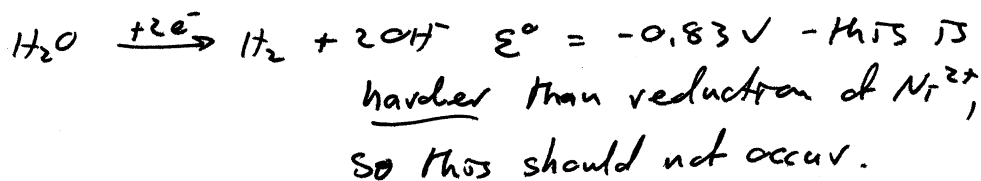
c. $5.0 \text{ mol Ag} \times \frac{1 \text{ mol e}^-}{1 \text{ mol Ag}} \times \frac{96{,}485 C}{1 \text{ mol e}^-} \times \frac{1 \text{ sec}}{100 C} = 4{,}824 \text{ sec}$
$(540 \, g)$

1.3 hrs
(80 min)

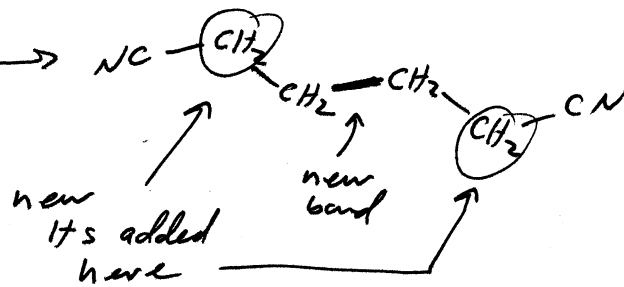64. with molten salts, hydrolysis of H_2O is not a problem

65. a. $Ni^{2+} \xrightarrow{+2e^-} Ni$ cathode $\varepsilon° = -0.23 \, V$

$2Br^- \xrightarrow{-2e^-} Br_2$ anode $\varepsilon° = -1.09 \, V$

d. 1 M aq $NiBr_2$ — can electrolysis of water occur at either electrode?

$H_2O \xrightarrow{+2e^-} H_2 + 2OH^-$ $\varepsilon° = -0.83 \, V$ –this is harder than reduction of Ni^{2+}, so this should not occur.

$H_2O \xrightarrow{-2e^-} \frac{1}{2}O_2 + 2H^+$ $\varepsilon° = -1.23 \, V$ –this is harder than ox of bromide, so this should not occur.

b. $Al^{3+} \xrightarrow{+3e^-} Al$ cathode $\varepsilon° = -1.66 \, V$

$2F^- \xrightarrow{-2e^-} F_2$ anode $\varepsilon° = -2.87 \, V$

e. 1 M aq AlF_3 — as above, consider oxidation + reduction of H_2O — these rxns are both easier (less negative potentials) than reduction of Al^{3+} + oxidation of F^-, so only electrolysis of water would occur !

72.

$$2 \ H_2C = \overset{H}{\underset{C \equiv N}{C}} \ + \ 2 H^+ + 2e^- \longrightarrow NC - CH_2 \overset{\displaystyle \uparrow}{\underset{CH_2}{}} - CH_2 \overset{CH_2}{} - CN$$

new
H's added
here

new
bond

$\downarrow H_2$ catalyst

$$H_2N - (CH_2)_6 - NH_2$$

$$\downarrow$$

$$\{ \overset{}{\underset{H}{N}} - (CH_2)_6 - \overset{H}{N} - \overset{}{\underset{O}{C}} - (CH_2)_x - \overset{O}{C} - N - \}$$

Nylon
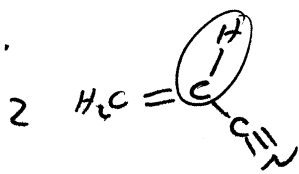
$$150 \ kg \ of \ this \times \frac{1000 \ g}{1 \ kg}$$

$$= 150,000 \ g \times \frac{1 \ mol}{108 \ g}$$

$$= 1.39 \ kmol$$

$$(1.39 \times 10^3 \ mol)$$

This requires 2.78 kmol of e⁻s

$$\times \frac{96,485 \times 10^3 \ C}{1 \ kmol}$$

$$= 2.68 \times 10^8 \ C$$

This amount of charge per hr (per 3600 s)

$$\Rightarrow 74,400 \ A \ !$$