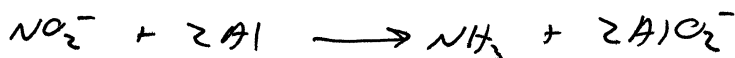
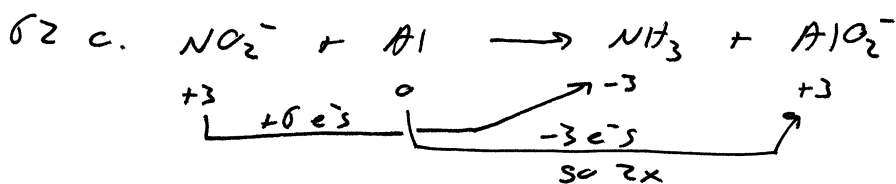
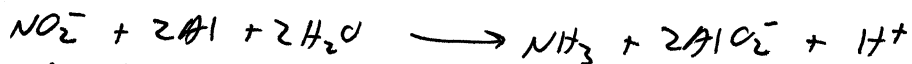


1. CH₄ - 58
 - a. C in CH₄ is oxidized, O₂ is reduced
 - b. Zn is oxidized, H⁺ is reduced
 - c. Cr is in the +6 ox. state in reactant & product - nothing else changes either, so not redox.
 - d. One O from O₃ (ozone) is reduced, N is oxidized
 - e. H₂O₂ is oxidized to O₂ & reduced to H₂O
 - f. Cu is oxidized from +1 to +2, and reduced to Cu metal.
 - g. That's an acid-base rxn! Not redox.
 - h. Not a redox rxn.
 - i. Si is reduced, Mg is oxidized.



+ 2 H₂O to balance O's

+ H⁺ to balance H's

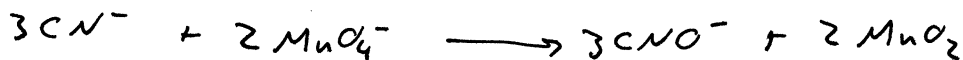
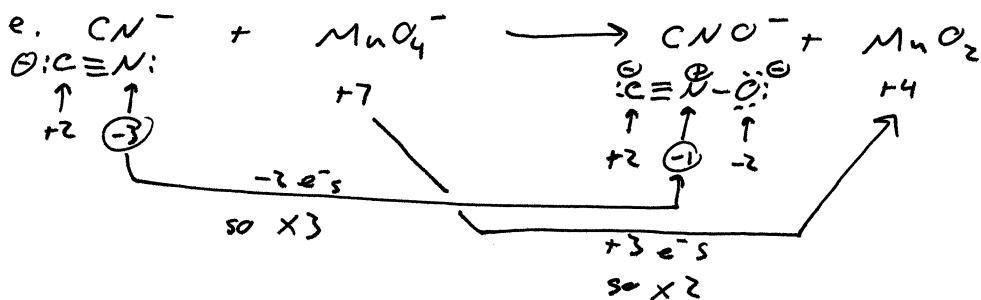
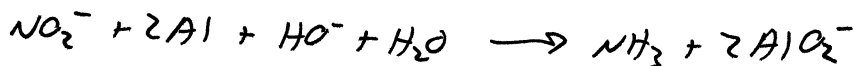


for basic soln rxn, we need to add 1 HO⁻ to each side

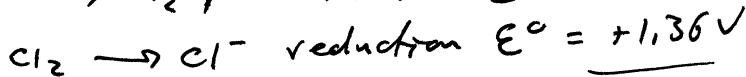
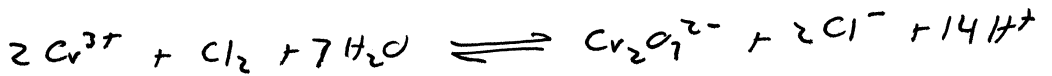
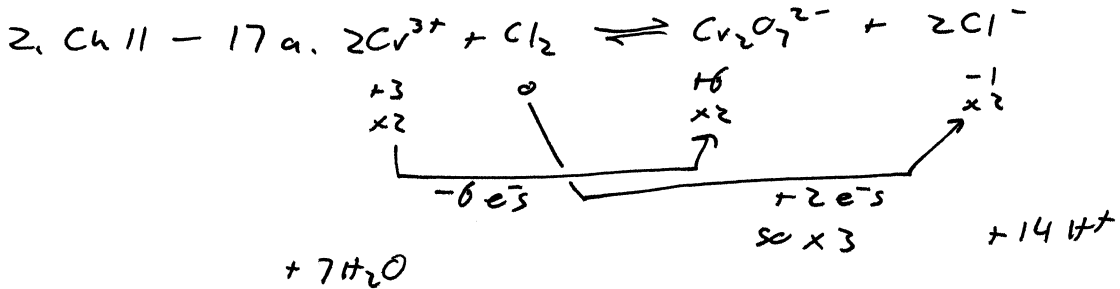
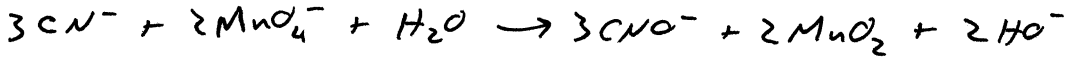
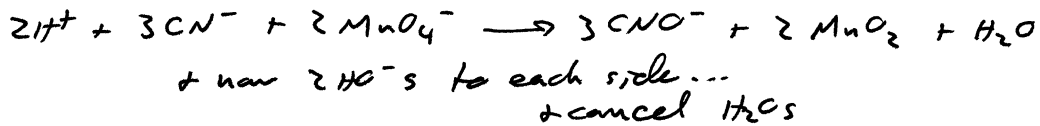
+ HO⁻

+ HO⁻

then cancel an H₂O...



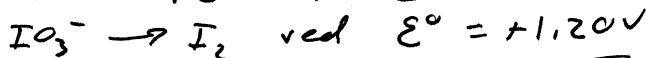
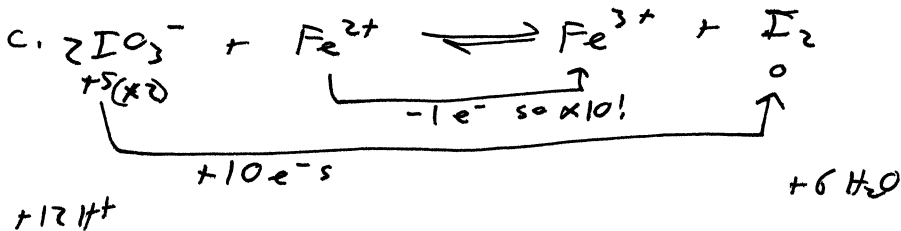
62e (cont'd) now add H₂O + H⁺ -



+0.03V in the direction shown.

Cr oxidation would occur at the (Pt) anode;

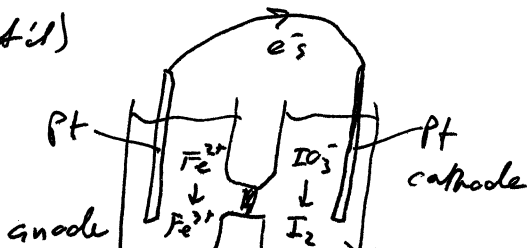
Cl₂ gas would be bubbled in at the cathode, where it would be reduced



+0.43V as written

cell diagram on next page...

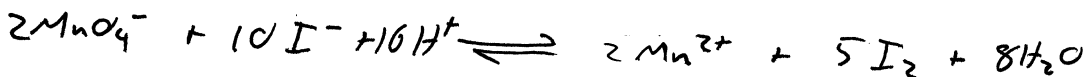
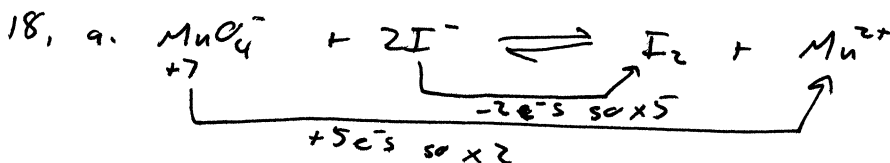
17c (cont'd)



If combination were NO_3^- , the rxn would convert $\text{Fe}(\text{NO}_3)_2$ to $\text{Fe}(\text{NO}_3)_3$, the additional NO_3^- being supplied by the HNO_3 via the salt bridge.

we also need acid here - HNO_3 , perhaps

anions migrate this way (mostly NO_3^- , rather than the reactant, IO_3^- , one would hope!)

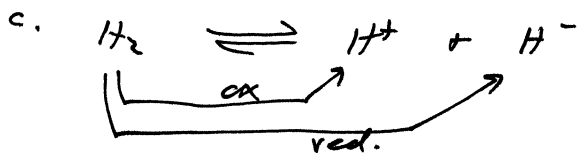


Mn reduction $E^\circ = +1.51\text{V}$
 I^- oxidation $E^\circ = -0.54\text{V}$ $\therefore E_{\text{cell}}^\circ = 0.97\text{V}$
 \therefore goes as written

b. Same thing! F^- to F_2 oxidation

has $E^\circ = -2.87\text{V}$, so $E_{\text{cell}}^\circ = -1.36\text{V}$

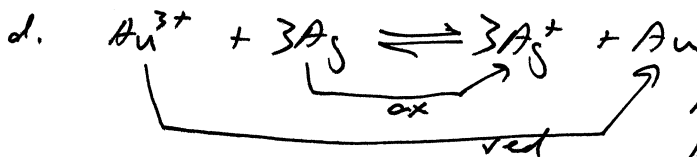
Don't waste time balancing this - you did it in part a! This does not go as written



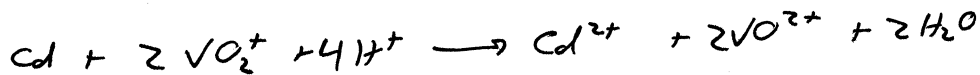
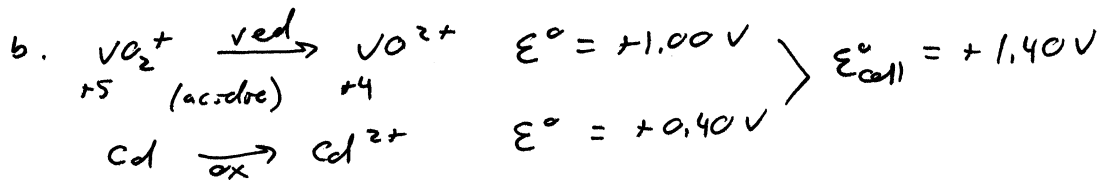
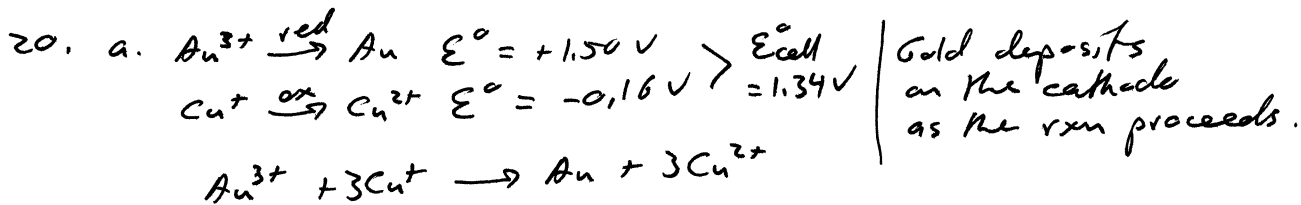
H_2 ox $E^\circ = 0$

H_2 red $E^\circ = -2.23\text{V}$

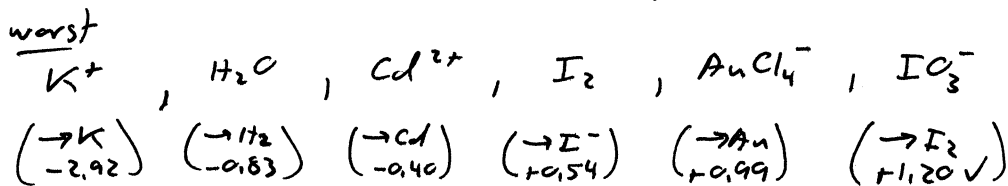
so $E_{\text{cell}}^\circ = -2.23\text{V}$ + this most certainly does not go as written.



Ag ox -0.80V
 Au^{3+} red $+1.50\text{V}$ $\therefore E_{\text{cell}}^\circ = +0.70\text{V}$
 so this does go as written.



24. A good "oxidizing agent" (good at oxidizing something else) has a favorable reduction potential.

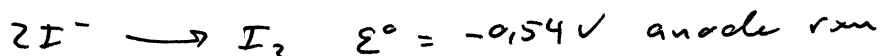
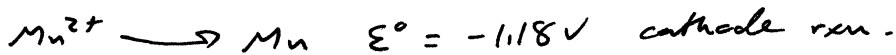


28. a. $\text{H}_2\text{S} (+ \text{aq Cl}^-)$, $\text{As}_2 (+ \text{aq Cl}^-)$, $+ \text{H}_2\text{SO}_3 (\text{aq})$ can all reduce Cu^{2+} to Cu without reducing Cu^{2+} to Cu^+ - write the rxns - the first reduction ($\text{Cu}^{2+} \rightarrow \text{Cu}$) is favorable, the second is not.

b. Reduces Br_2 but not I_2 to their respective halide ions...
 VO^{2+} , $\text{Au} (+ \text{aq Cl}^-)$, NO , ClO_2^- , H_2S^{2+} , As_2 , H_2S , Fe^{2+} , H_2O_2 , MnO_4^{2-}
 assuming standard conditions

61. $1 \text{ hr} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ sec}}{1 \text{ min}} \times \frac{15 \text{ C}}{1 \text{ sec}} \times \frac{1 \text{ mol e}^-}{96,485 \text{ C}} = 0.56 \text{ mol e}^-$
15 Amps

a. $\text{Co}^{2+} \rightarrow \text{Co}$ requires 2 mol e^- s per mol Co , so 0.28 mol Co
 $= 16.5 \text{ g Co}$
 d. $\text{CrO}_3 \rightarrow \text{Cr}$ " " " " " " " " Cr , so 0.0933 mol Cr
 $= 4.85 \text{ g Cr}$

65. c. molten MnF_2 

$$E_{\text{cell}}^\circ = -1.72V$$

f. see #65 on non-graded part ---

reduction of water is easier than reduction of Mn^{2+} , so formation of H_2 will occur at the cathode (ox of water is harder than ox of I^- , so anode rxn will be $I^- \rightarrow I_2$)

As discussed in class, both anode + cathode rxns in charging a lead/ H_2SO_4 car battery are harder than electrolysis of H_2O . If the $PbSO_4$ were in solution, applying the necessary voltage would only produce $H_2 + O_2$. But because the $PbSO_4$ is insoluble in aq H_2SO_4 it sticks to the electrodes + its ox to PbO_2 + red. to Pb can compete with electrolysis of H_2O .

$$74. \quad 15,000 \text{ watt-hrs} \times \frac{3600 \text{ sec}}{1 \text{ hr}} \times \frac{1 \text{ J}}{1 \text{ watt-sec}} = 54,000 \text{ kJ} \\ = 54 \text{ MJ} !$$

Melting Al:

$$\frac{10.7 \text{ kJ}}{1 \text{ mol}} \times 1 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{27 \text{ g}} = 400 \text{ kJ}$$

That's < 1%
of the energy
required for
electrolysis!