(#11)

Oxidation is the process by which electrons are lost and reduction is the process by which electrons are gained. It is impossible for oxidation to occur without reduction and vice-versa. Electrons must be lost and then regained and for this reason oxidation and reduction must take place simultaneously.
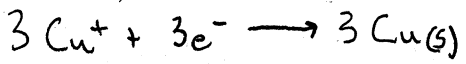
(#12)

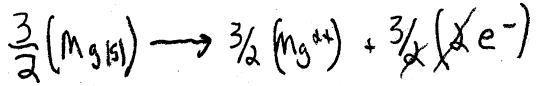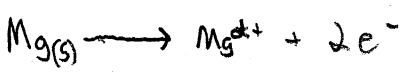For a reaction to spontaneously occur, the copper must oxidize and silver must reduce.

$$E°$$

ox. $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$ -0.337 V

red. $Ag^+(aq) + 1e^- \longrightarrow Ag(s)$ $+0.799$ V
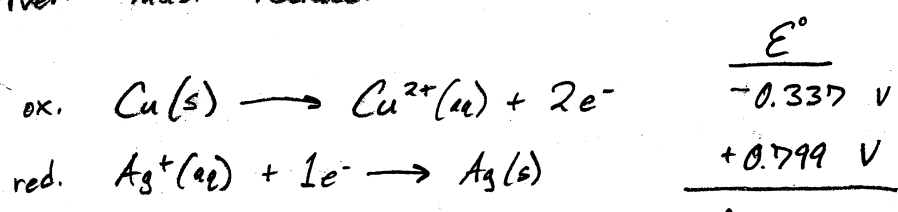
$$E°_{tot} = {}^+0.462 \text{ V} \longrightarrow \text{spontaneous rxn}$$

(#13)

To reduce 3 moles of copper, $3e^-$ are needed.

$$3\,Cu^+ + 3e^- \longrightarrow 3\,Cu(s)$$

To obtain $3e^-$, set the magnesium half reaction to spit out $3e^-$.

$$Mg(s) \longrightarrow Mg^{2+} + 2e^-$$

$$\tfrac{3}{2}(Mg(s)) \longrightarrow \tfrac{3}{2}(Mg^{2+}) + \tfrac{3}{2}(2e^-)$$
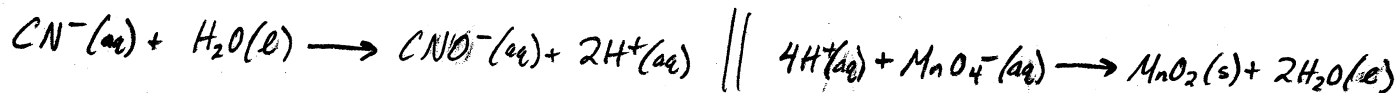
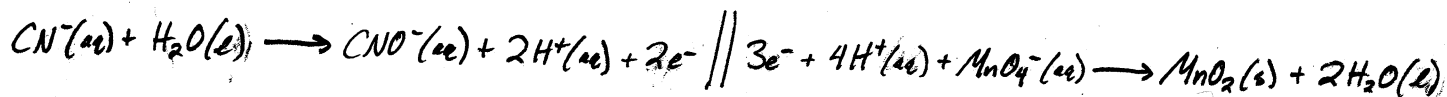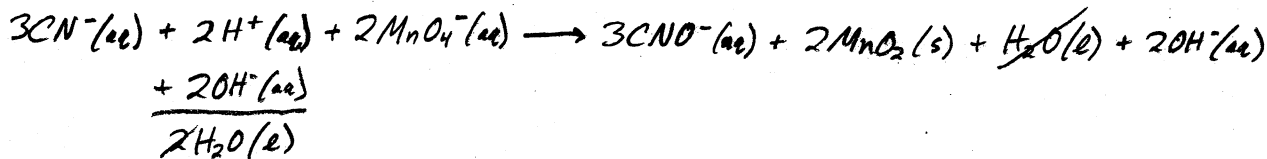$$\tfrac{3}{2}(Mg(s)) \longrightarrow \tfrac{3}{2}(Mg^{2+}) + 3e^-$$

$\tfrac{3}{2}$ moles are needed.
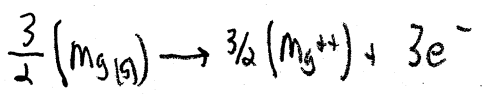
#14

Balance the reduction and oxidation rxns separately.

Step 1
↳ Balance oxygens w/ H_2O

$$CN^-_{(aq)} + H_2O_{(l)} \longrightarrow CNO^-_{(aq)} \quad || \quad MnO_4^-_{(aq)} \longrightarrow MnO_2 (s) + 2H_2O_{(l)}$$
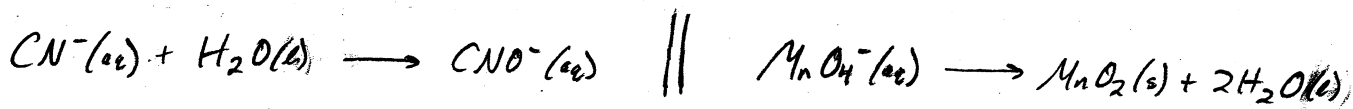
Step 2
↳ Balance H atoms w/ H^+ ions

$$CN^-_{(aq)} + H_2O_{(l)} \longrightarrow CNO^-_{(aq)} + 2H^+_{(aq)} \quad || \quad 4H^+_{(aq)} + MnO_4^-_{(aq)} \longrightarrow MnO_2 (s) + 2H_2O_{(l)}$$

Step 3
↳ Balance charge w/ e^-

$$CN^-_{(aq)} + H_2O_{(l)} \longrightarrow CNO^-_{(aq)} + 2H^+_{(aq)} + 2e^- \quad || \quad 3e^- + 4H^+_{(aq)} + MnO_4^-_{(aq)} \longrightarrow MnO_2 (s) + 2H_2O_{(l)}$$

Step 4
↳ Combine to one equation and add OH^- b/c it is in a <u>basic</u> soln

$$\left(CN^-_{(aq)} + H_2O_{(l)} \longrightarrow CNO^-_{(aq)} + 2H^+_{(aq)} + 2e^- \right) \times 3$$
$$\left(3e^- + 4H^+_{(aq)} + MnO_4^-_{(aq)} \longrightarrow MnO_2 (s) + 2H_2O_{(l)} \right) \times 2$$

NOTE: must multiply by # so that electrons cancel out.

$$3CN^-_{(aq)} + 3H_2O_{(l)} \longrightarrow 3CNO^-_{(aq)} + 6H^+_{(aq)} + 6e^-$$
$$6e^- + 8H^+_{(aq)} + 2MnO_4^-_{(aq)} \longrightarrow 2MnO_2 (s) + 4H_2O_{(l)}$$

$$3CN^-_{(aq)} + 2H^+_{(aq)} + 2MnO_4^-_{(aq)} \longrightarrow 3CNO^-_{(aq)} + 2MnO_2 (s) + H_2O_{(l)} + 2OH^-_{(aq)}$$
$$+ 2OH^-_{(aq)}$$
$$\overline{2H_2O_{(l)}}$$

$$\Downarrow$$

$$3CN^-_{(aq)} + H_2O_{(l)} + 2MnO_4^-_{(aq)} \longrightarrow 3CNO^-_{(aq)} + 2MnO_2 (s) + 2OH^-_{(aq)}$$
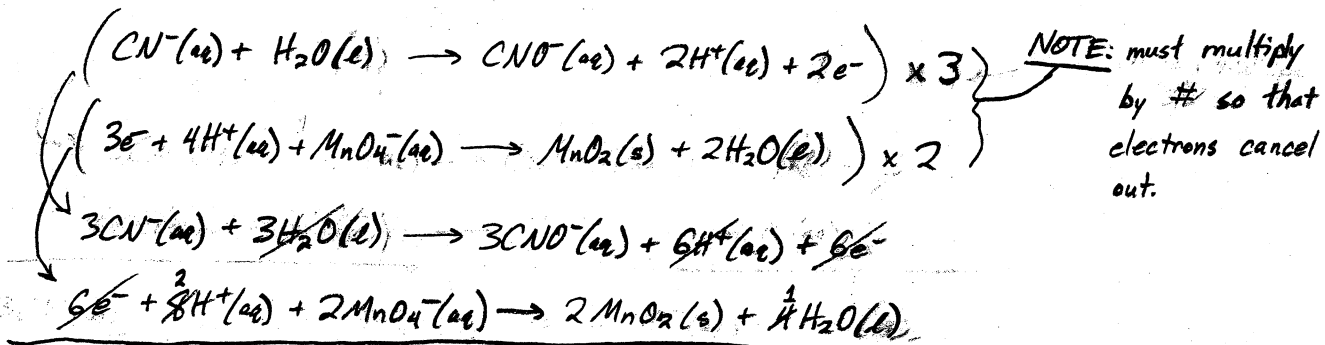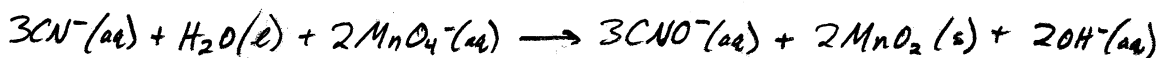
Lithium has a strong affinity to undergo oxidation...

$$Li_{(s)} \longrightarrow Li^+_{(aq)} + e^- \qquad \mathcal{E}^0_{cell} = 3.05 V \text{ (by turning around the } \mathcal{E}^0 \text{ provided)}$$
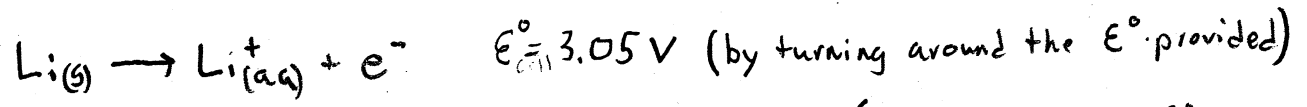
While Zinc also has a natural affinity to be oxidized $(Zn_{(s)} \rightarrow Zn^{2+} + 2e^- \; \mathcal{E}^0 = 0.76 V)$

every oxidation requires a simultaneous reduction for it to occur.

The will of Lithium to be oxidized trumps the will for zinc to be oxidized and

thus zinc is reduced.

$$Li_{(s)} \longrightarrow Li^+_{(aq)} + e^- \qquad \mathcal{E}^0 = 3.05 V$$

$$Zn^{2+} + 2e^- \longrightarrow Zn_{(s)} \qquad \mathcal{E}^0 = -0.76 V$$
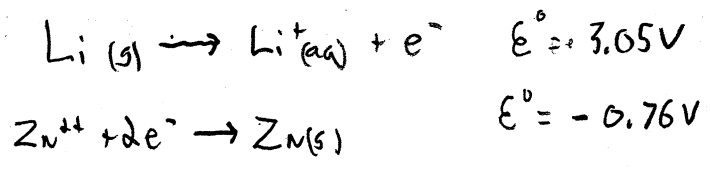
To get a balanced Number of electrons, the lithium equation must be multiplied by 2, but remember this does **Not** effect $\mathcal{E}^0$.

$$2 Li_{(s)} \longrightarrow 2Li^+ + 2e^- \qquad \mathcal{E}^0 = 3.05 V$$

$$Zn^{2+} + 2e^- \longrightarrow Zn_{(s)} \qquad \mathcal{E}^0 = -0.76 V$$

$$\overline{2 Li_{(s)} + Zn^{2+} \longrightarrow Zn_{(s)} + 2Li^+ \qquad \mathcal{E}^0 = 2.29 V}$$

TOTAL RXN: $Mg(s) + MnO_2(s) + 4H^+_{(aq)} \longrightarrow Mn^{2+}_{(aq)} + Mg^{2+}(aq) + 2H_2O$

$\mathcal{E}^0_{cell} = 3.6 V$
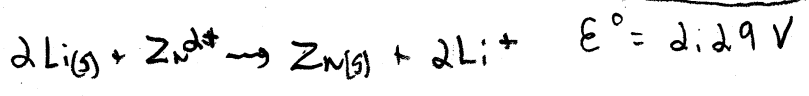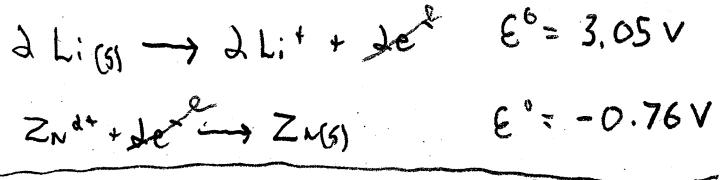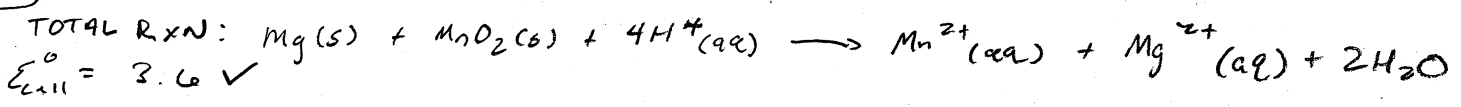
we know the Nernst Eqn Relates $\mathcal{E}^0_{cell}$ to $\mathcal{E}_{cell}$ for different quotient ratios.

For this cell, $Q = \dfrac{[Mn^{2+}][Mg^{2+}]}{[H^+]^4}$

Nernst: $\mathcal{E}_{cell} = \mathcal{E}^0_{cell} - \dfrac{RT}{nF} \ln Q$

At standard conditions, All concentrations are assumed to be 1M,

so $Q^0 = 1$ we need to calculate Q @ the New pH.

$[H^+] = 10^{-pH}$

$[H^+] = 10^{-7}$

$Q = \dfrac{[1M][1M]}{[10^{-7}]^4} = 1 \times 10^{28}$

Q has become much larger, which means $\ln Q$ has grown as

well. The term $\dfrac{RT}{nF} \ln Q$ will be positive, making $\mathcal{E}_{cell} < \mathcal{E}^0_{cell}$.

thus, | **A. Decrease** |

$\hookrightarrow$ Note that $\left(\dfrac{RT}{nF}\right)$ is assumed constant

for the 2 cases.

#17) we'll look at all 3 statements. For $\mathcal{E}_{cell}^\circ = 2.03\,V$,

I) From Nernst Eqn, we know $\mathcal{E}_{cell}^\circ = \dfrac{RT}{nF}\ln K$ $\left(\dfrac{RT}{nF}\right)$ will always be Positive

since $\mathcal{E}_{cell}^\circ > 0$, $K > 1$.

$\boxed{\text{I is true}}$

II) There is a simple relation between ΔG° and $\mathcal{E}_{cell}^\circ$

$$\Delta G^\circ = -nF\mathcal{E}_{cell}^\circ$$

Since $\mathcal{E}_{cell}^\circ > 0$, $\Delta G^\circ < 0$. $\boxed{\text{III is true}}$

II) Finally, we look @ the Gibbs Free Energy Relation,

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ.$$

As you can see, the equation depends on ΔH° and ΔS°, so there could be a scenario where ΔS° is Negative and ΔG° is still negative, therefore it is Not <u>essential</u> for $\Delta S^\circ > 0$. $\boxed{\text{II is Not Necessarily true}}$

Thus, $\boxed{\text{D. I and III only}}$ is <u>correct</u>

#22

For this problem we need the Nernst Equation.

$$E_{cell} = E_{cell}^\circ - \dfrac{RT}{nF}\ln Q$$

We are given E°_{cell}, T

F is a constant $= 96500$ & so is $R = 8.314$

n is the # e- transferred, in this case $n = 2$

We are given the concentrations / partial pressure to solve For Q.

$$Q = \dfrac{[CO^{2+}]^2\, P_{Cl_2}}{[CO^{3+}]^2[Cl^-]^2} = \dfrac{[0.624]^2\,(9.10)}{[0.156]^2[0.723]^2} = 2.785\times10^2$$

Now, we can solve For E_{cell}

$$E_{cell} = 0.71 - \dfrac{(8.314)(298)^*}{(2)(96500)}\ln\left(2.785\times10^2\right)$$

$$E_{cell} = 0.71 - (0.012837)\ln\left(2.785\times10^2\right)$$

$$E_{cell} = 0.71 - 0.0723$$

$$\boxed{E_{cell} = 0.638\,V}$$

* Remember to convert $^\circ C \to K$ $(25\,^\circ C + 273 = 298K)$

#23

$$x \text{ min} = \frac{33.1 \text{ g Cr}^{3+}}{} \times \frac{1 \text{ mol Cr}^{3+}}{52.00 \text{ g Cr}^{3+}} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Cr}^{3+}} \times \frac{96,485 \text{ C}}{1 \text{ mol e}^-} \times \frac{1 \text{ s}}{5 \text{ c}} \times \frac{1 \text{ min}}{60 \text{ s}} = \boxed{614 \text{ min}}$$

#24

$$x \text{ C} = 4.57 \text{ A} \times \frac{7}{2} \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}}$$
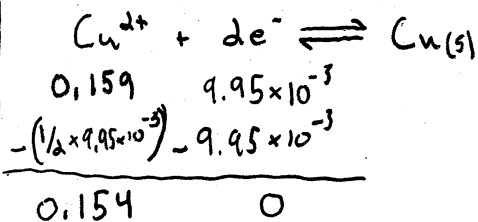
$$= 959.7 \text{ C}$$

$$x \text{ mole}^- = 959.7 \text{ C} \times \frac{1 \text{ mol e}^-}{9.6485 \times 10^4}$$

$$= 9.95 \times 10^{-3} \text{ mol e}^-$$

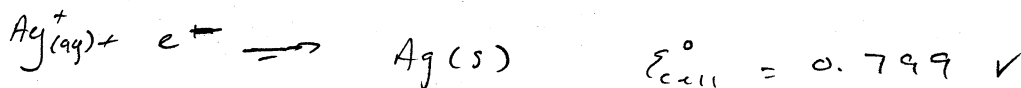$$\text{mol Cu}^{2+} = \frac{0.350 \text{ mol Cu}^{2+}}{1 \text{ L}} \times 455 \times 10^{-3} \text{ L}$$

$$= 0.159 \text{ mol Cu}^{2+}$$

$$Cu^{2+} + 2e^- \rightleftharpoons Cu_{(s)}$$

0.159	9.95×10^{-3}	
$-(\frac{1}{2} \times 9.95 \times 10^{-3})$	-9.95×10^{-3}	
0.154	0	

Final mol $Cu^{2+} = 0.154$ mol e$^-$

$$\frac{0.154 \text{ mol}}{455 \times 10^{-3} \text{ L}} = 0.339 M$$

#25

$$H_2 (g) \longrightarrow 2e^- + 2H^+ (aq) \qquad E^\circ_{cell} = 0.000 \text{ V} \checkmark$$

$$Ag^+_{(aq)} + e^- \longrightarrow Ag (s) \qquad E^\circ_{cell} = 0.799 \text{ V} \checkmark$$

Nernst Eqn

$$0.950 = 0.799 - \frac{RT}{nF} \ln Q \qquad , \qquad Q = \boxed{\frac{[H^+]^2}{P_{H_2} [Ag^+]}}$$

Assumed Equal to <u>1</u>

$$0.151 = -\frac{(8.314)(298)}{(2)(96485)} \ln Q$$

$$11.761 = -\ln [H^+]^2$$

$$[H^+]^2 = 7.8 \times 10^{-6} \text{ M}$$

$$[H^+] = 0.0028$$

$$pH = 2.55 \Rightarrow \boxed{D. 2.6}$$