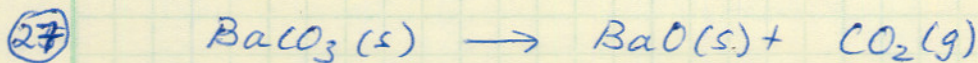


INTRODUCTION TO NATURAL SCIENCE

CHEMISTRY HOMEWORK - SPRING - WEEK 8

Chapter 19



$$\Delta G_{\text{rn}}^{\ominus} = \Delta G_{\text{f}}^{\ominus}[\text{BaO}(s)] + \Delta G_{\text{f}}^{\ominus}[\text{CO}_2(g)] - \Delta G_{\text{f}}^{\ominus}[\text{BaCO}_3(s)]$$

$$219.7 \text{ kJ} = -520.38 \text{ kJ} + (-394.359) \text{ kJ} - \Delta G_{\text{f}}^{\ominus}[\text{BaCO}_3(s)]$$

$$\begin{aligned} \Delta G_{\text{f}}^{\ominus}[\text{BaCO}_3(s)] &= (-520.38 - 394.359 - 219.7) \text{ kJ} \\ &= \underline{\underline{-1134.44 \text{ kJ}}} \end{aligned}$$

(29) (a) disfavored  $\Delta S^{\ominus} > 0$  at high T, this can lead to  $\Delta G_{\text{rn}}^{\ominus} < 0$

(b) favored  $\Delta S^{\ominus} < 0$  at high T, can lead to  $\Delta G_{\text{rn}}^{\ominus} > 0$

(c) disfavored  $\Delta S^{\ominus} > 0$  at high T, can lead to  $\Delta G_{\text{rn}}^{\ominus} < 0$

(d) favored  $\Delta S^{\ominus} < 0$  at high T, can lead to  $\Delta G_{\text{rn}}^{\ominus} > 0$

(33)  $\Delta G_{\text{rn}}^{\ominus} = -RT \ln K$

$$86.58 \text{ kJ} = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298.15 \text{ K}) \ln K$$

$$\ln K = \frac{-86.58 \times 10^3 \text{ J}}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298.15 \text{ K})} = -34.93$$



$$K = \underline{\underline{6.78 \times 10^{-16}}}$$

$\Delta G_{\text{rxn}}^{\ominus} > 0$  ( $r^{\ominus}$  is not favored as written)

Note  $K$  is very small (not favored as written)

$$(35) \quad \Delta G_{\text{rxn}}^{\ominus} = \Delta G_{\text{f}}^{\ominus} [\text{C}_2\text{H}_6(\text{g})] - \Delta G_{\text{f}}^{\ominus} [\text{C}_2\text{H}_4(\text{g})] - \Delta G_{\text{f}}^{\ominus} [\text{H}_2(\text{g})]$$

$$\Delta G_{\text{rxn}}^{\ominus} = [-31.89 - (68.35)] \text{ kJ} = \underline{\underline{-100.24 \text{ kJ}}}$$

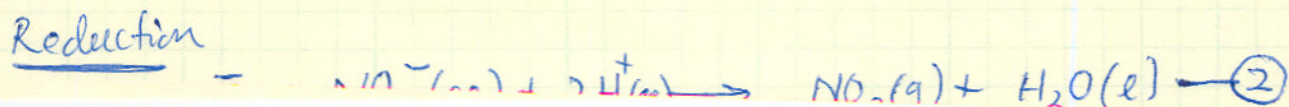
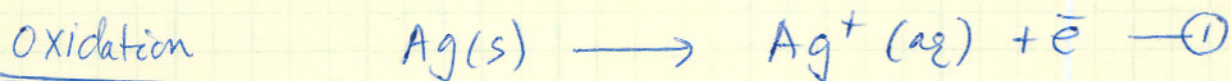
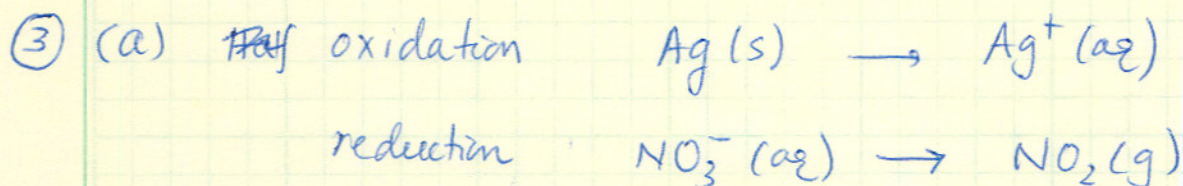
$$\Delta G_{\text{rxn}}^{\ominus} = -RT \ln K$$

$$\ln K = \frac{-100.24 \times 10^3 \text{ J}}{- (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) (298.15 \text{ K})} = 40.44$$

$$K = \underline{\underline{3.65 \times 10^{17}}}$$

$\Delta G_{\text{rxn}}^{\ominus} < 0$  spontaneous as written.  $K$  is very large.

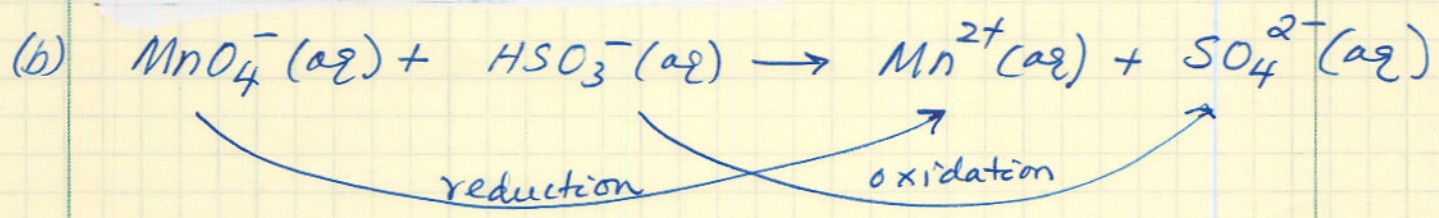
## Chapter 20



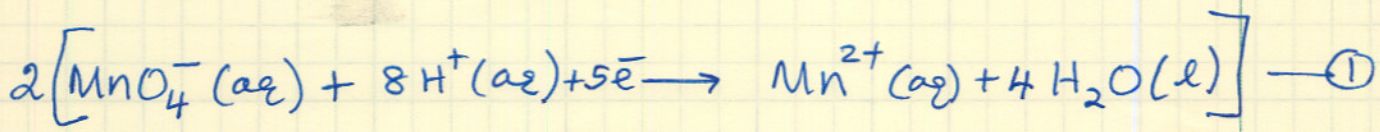


a

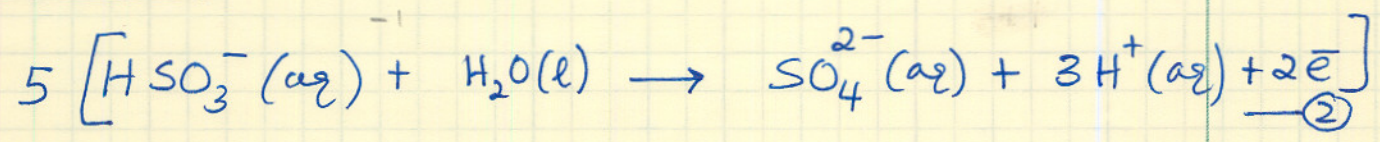
Redox reaction



Reduction

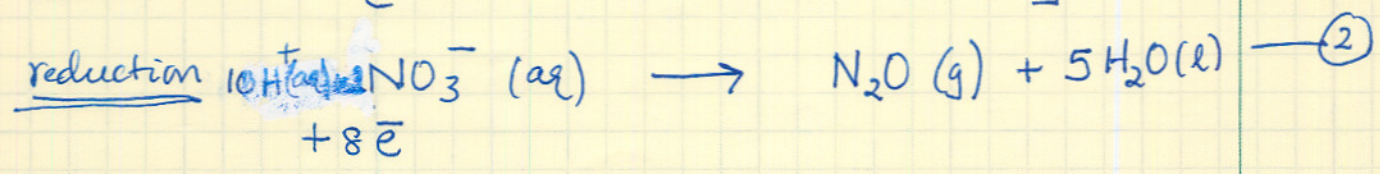
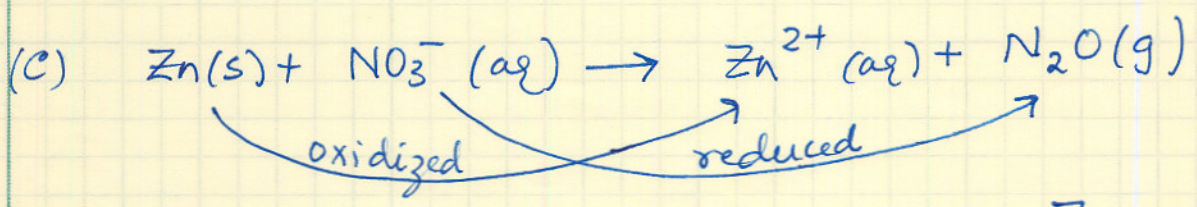
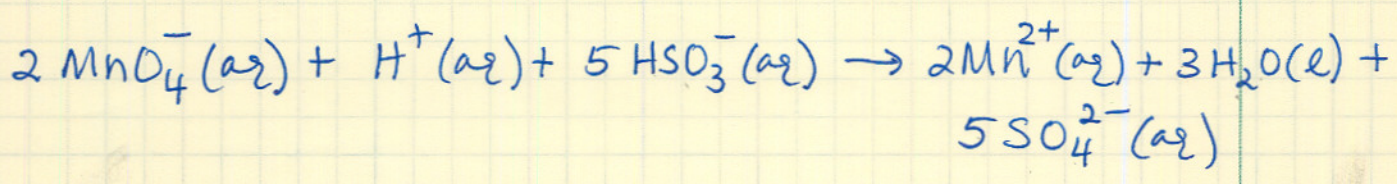
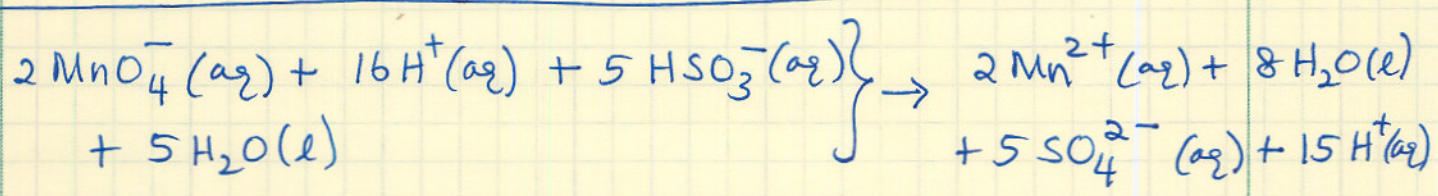


oxidation



+

↓

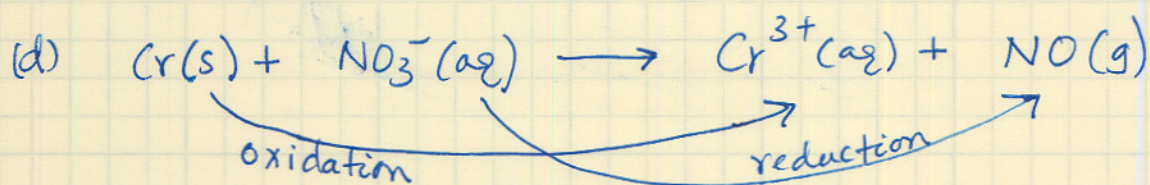
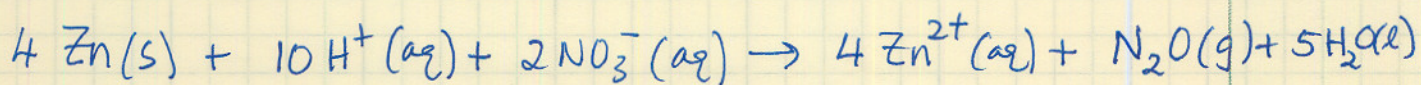
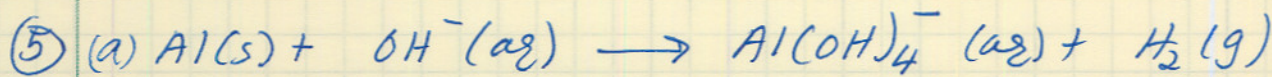
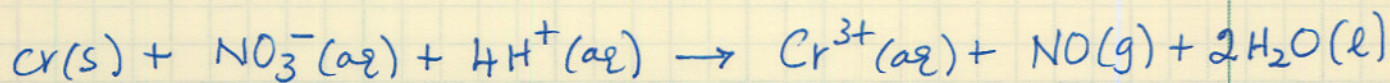
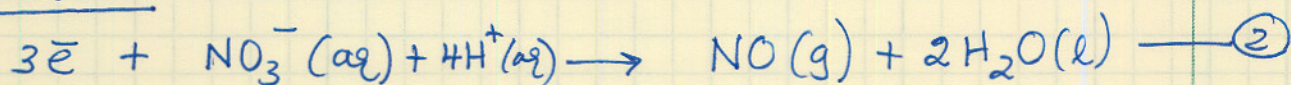


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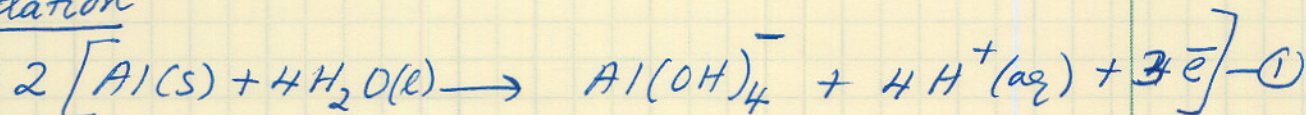
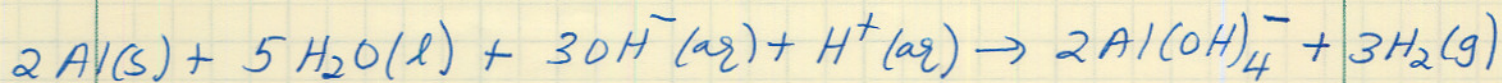
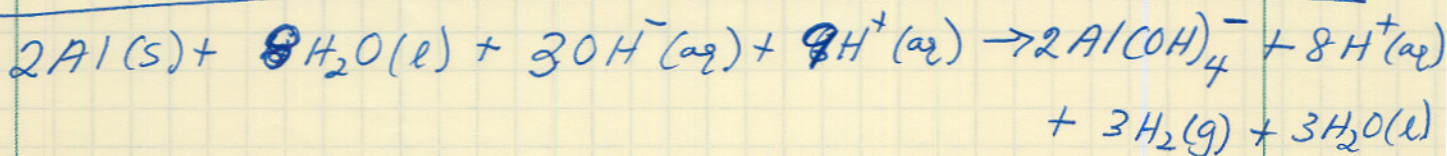
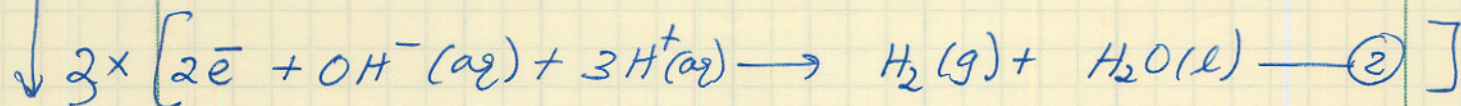
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b

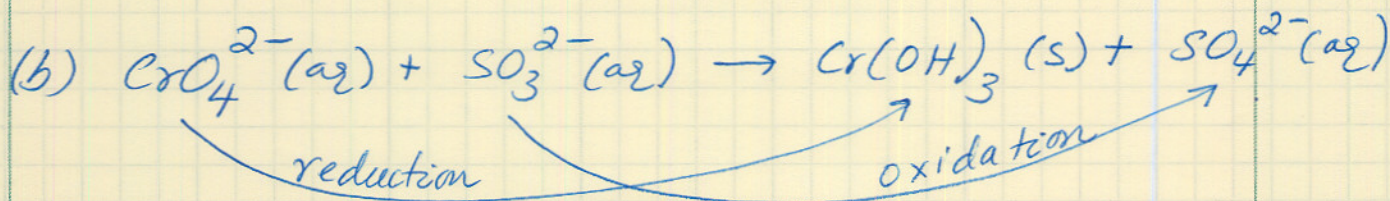
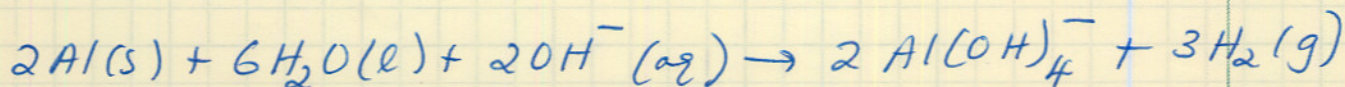
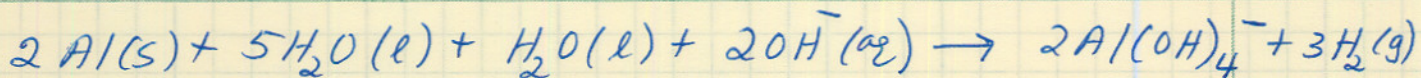
oxidationreduction

oxidation
reduction

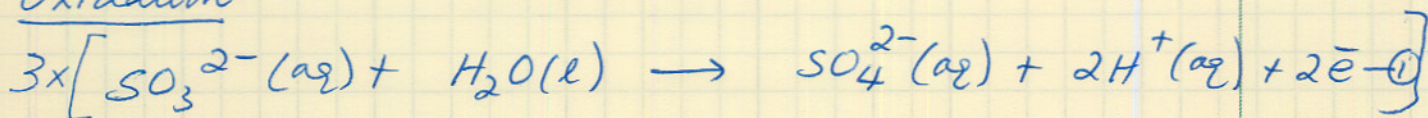
oxidationreduction



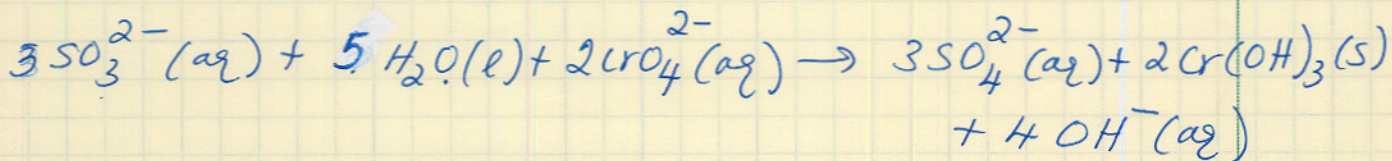
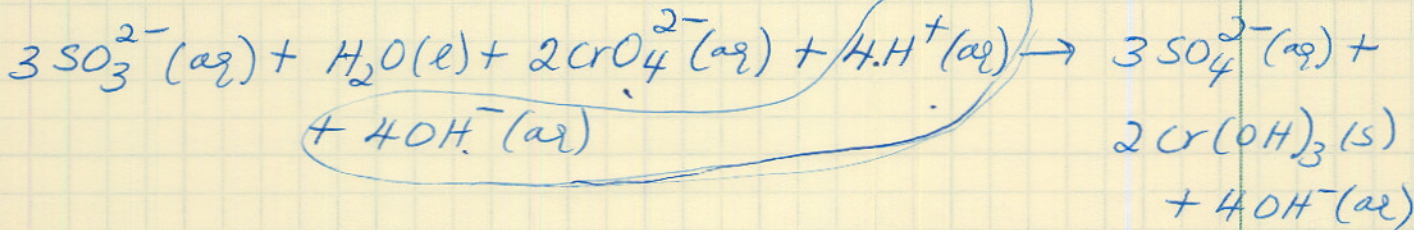
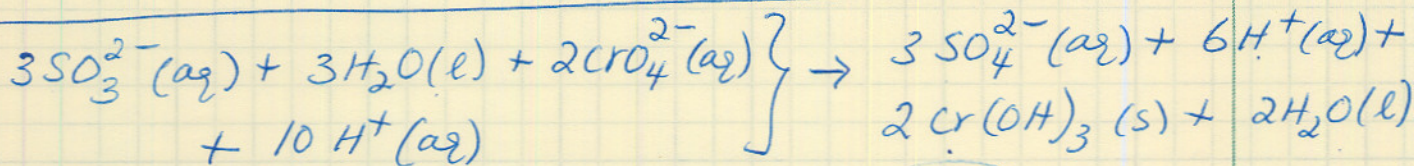
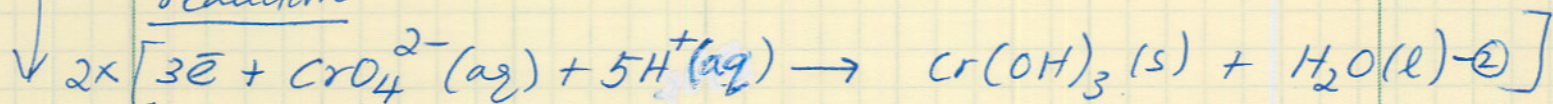
c



oxidation

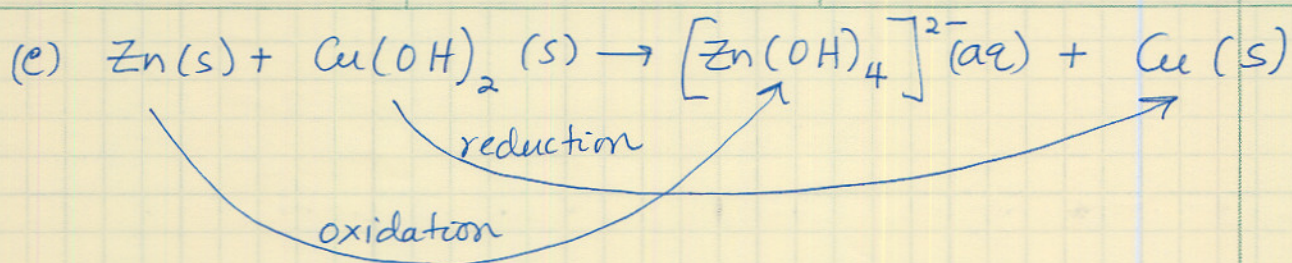
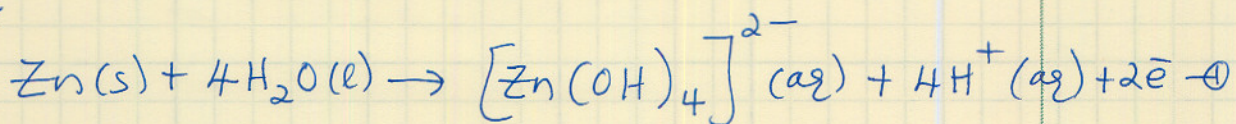
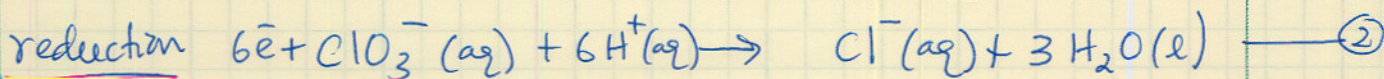
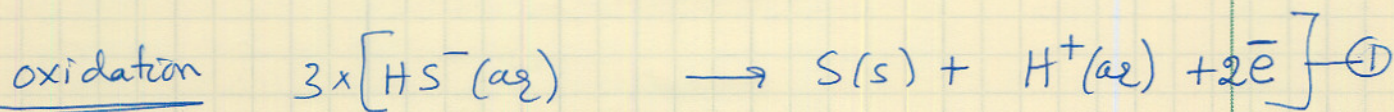
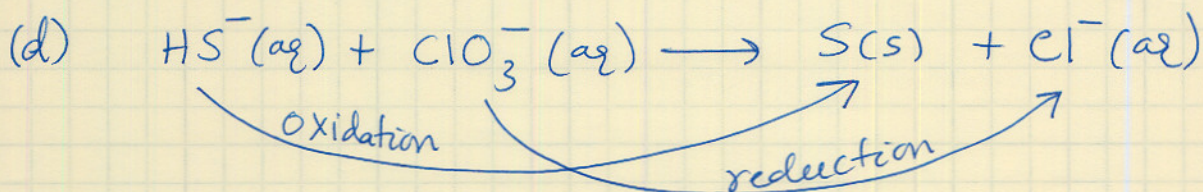
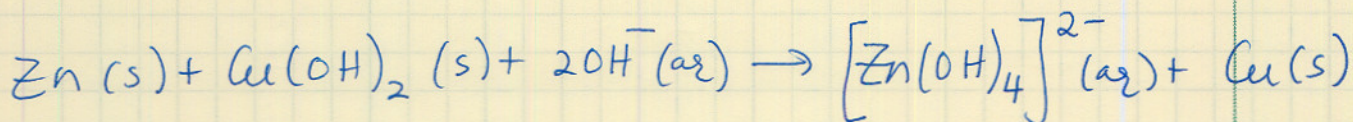
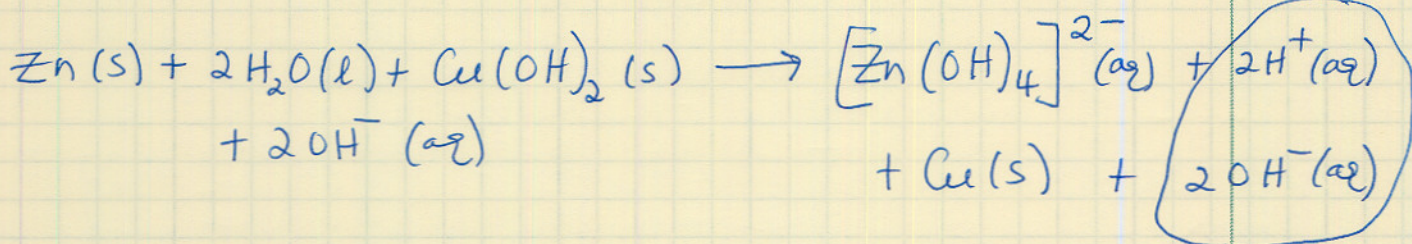
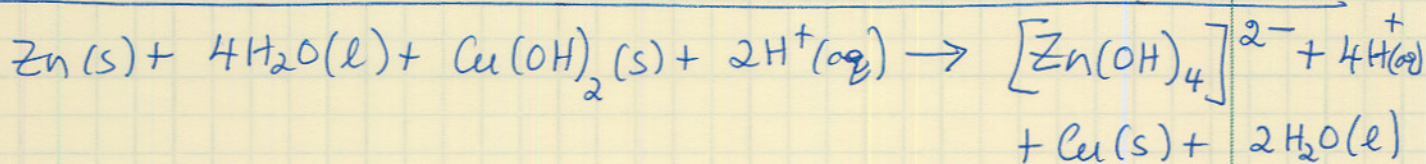
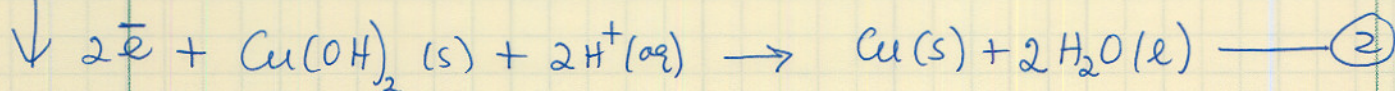


reduction



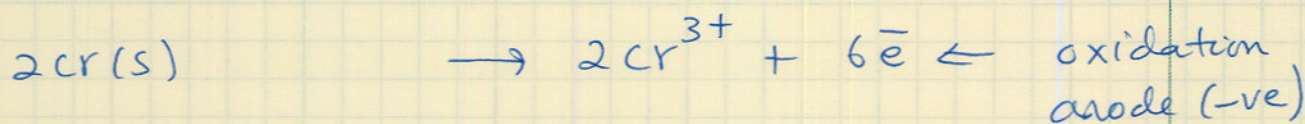
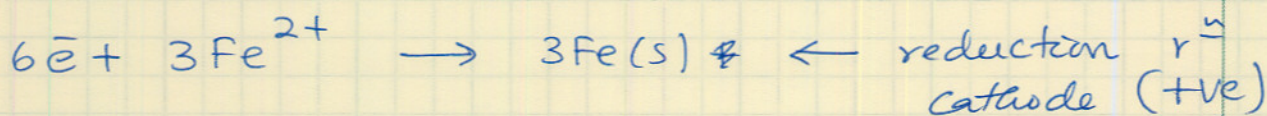
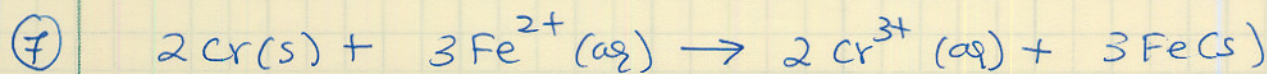
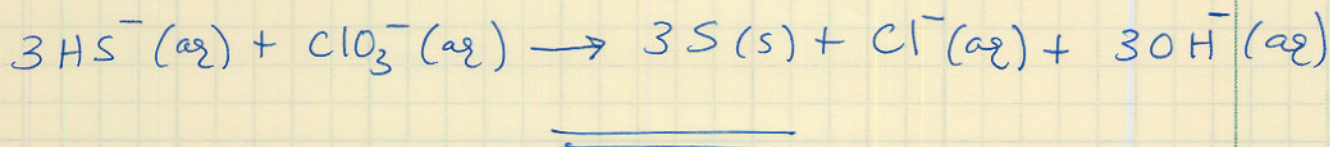
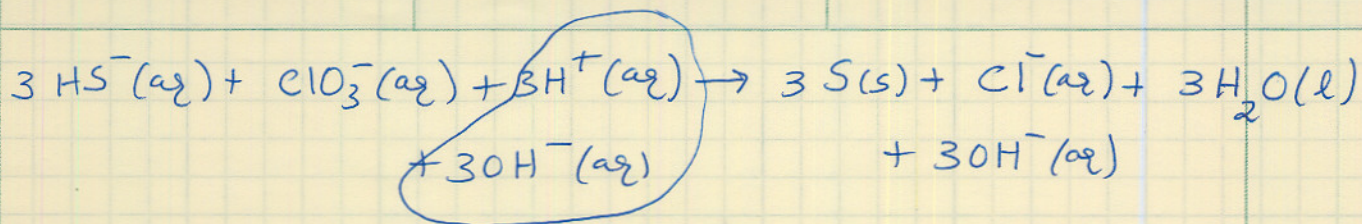


d

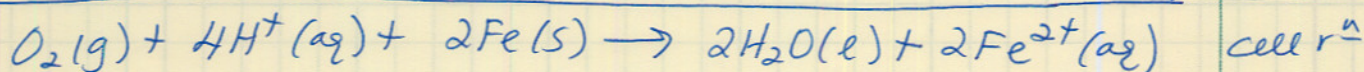
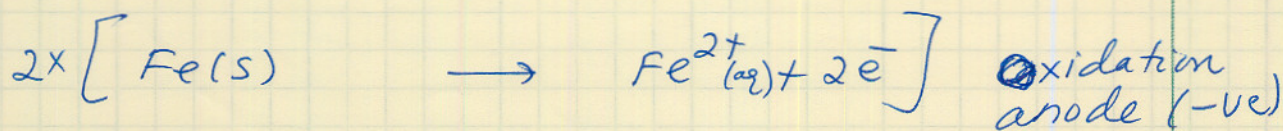
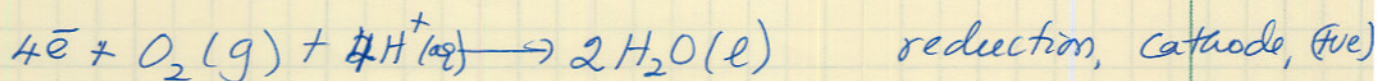
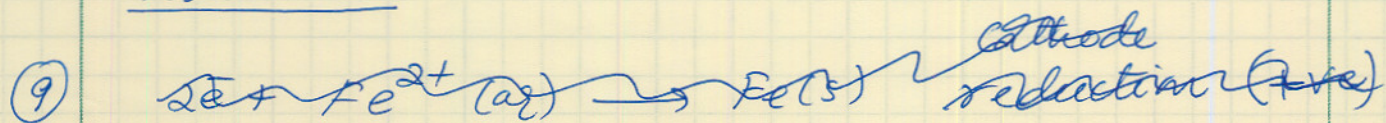
oxidationreduction



e



Electrons in the external circuit flow from the Cr anode (-) to the Fe cathode (+) electrode. Negative ions move in the salt bridge from the cathode (Fe/Fe<sup>2+</sup>) half-cell to the anode (Cr/Cr<sup>3+</sup>) half-cell. The half-reaction at the anode is oxidation and that at the cathode is reduction.





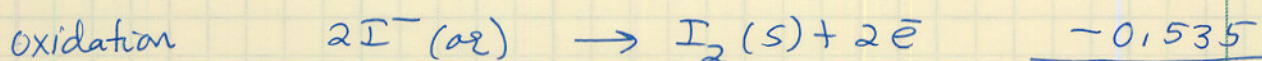
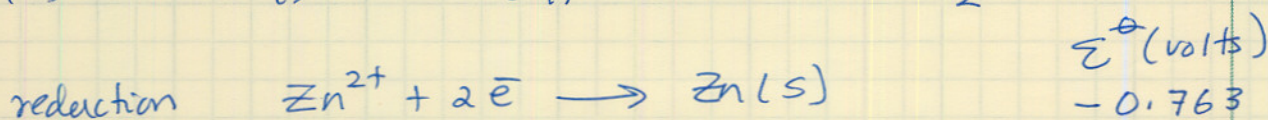
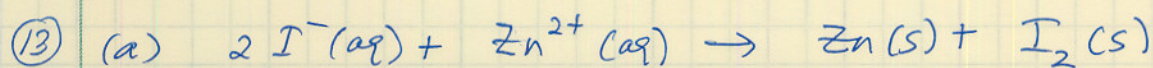
f

(b) anode: Fe(s) is oxidized

cathode O<sub>2</sub>(g) is reduced.

(c) Electrons in the external circuit flow from the (Fe) anode (-) electrode to the cathode (+)

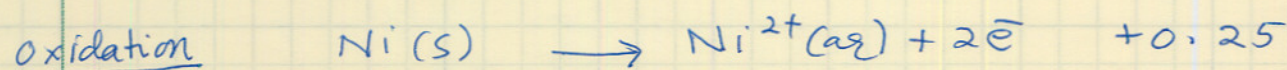
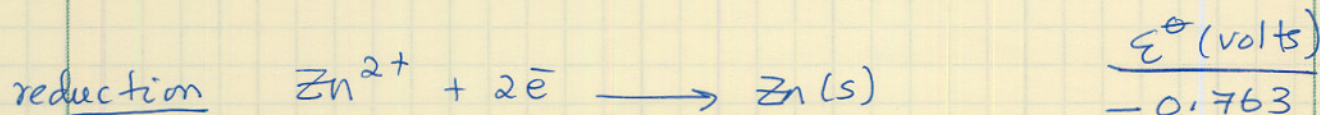
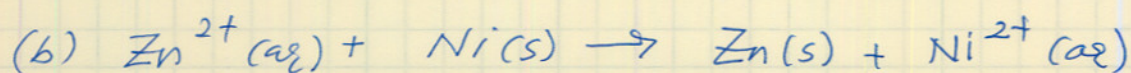
(O<sub>2</sub>) electrode. Negative ions move in the salt bridge from the cathode half-cell to the anode half-cell.



$$E_{\text{cell}}^{\ominus} = -0.763 - 0.535$$

$$= \underline{\underline{-1.298 V}}$$

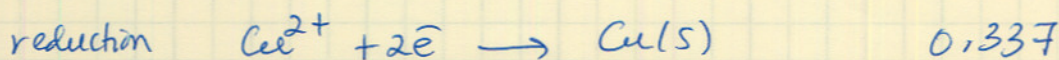
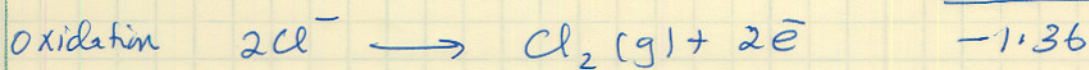
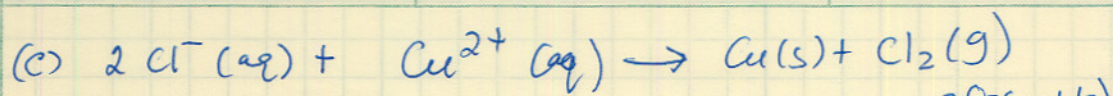
Reaction is not product favored.



$$E_{\text{cell}}^{\ominus} = (-0.763 + 0.25) V = \underline{\underline{-0.513 \text{ volts}}}$$

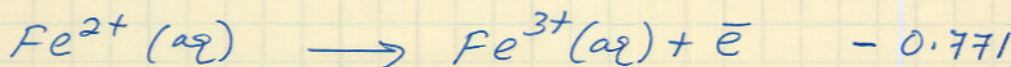
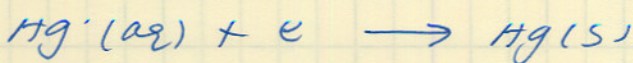
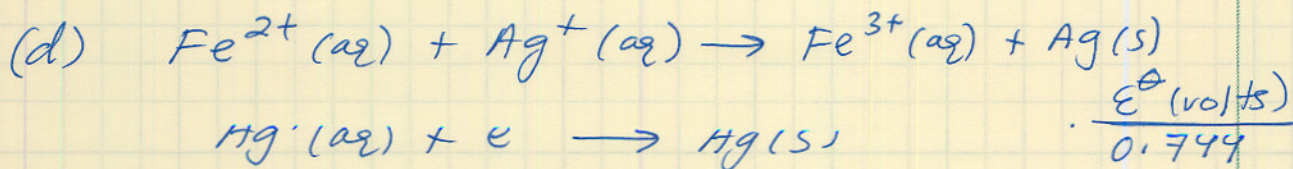
Reaction is not product favored.





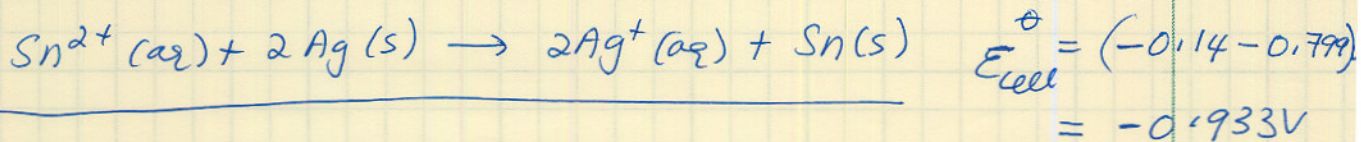
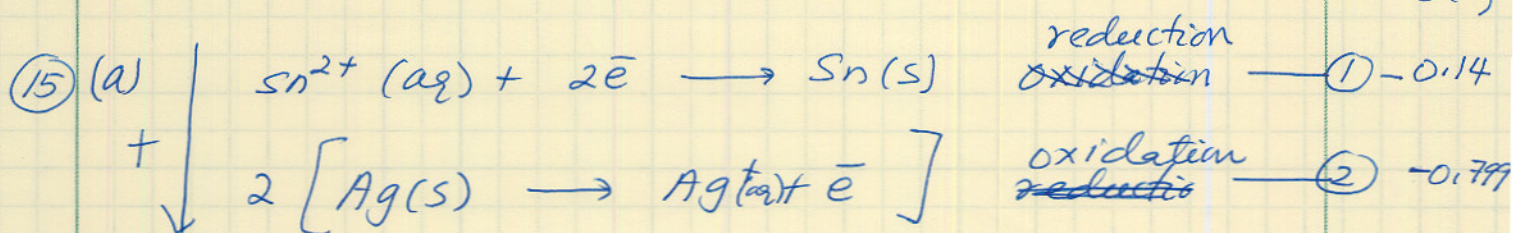
$$E_{\text{cell}}^{\ominus} = -1.36 + 0.337 = \underline{\underline{-1.023 \text{ volts}}}$$

This reaction is not product favored.



$$E_{\text{cell}}^{\ominus} = (0.799 - 0.771) \text{ volts} = \underline{\underline{0.028 \text{ volts}}}$$

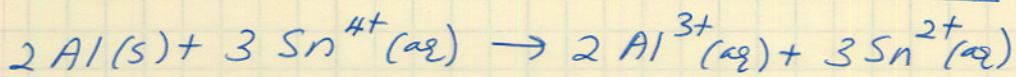
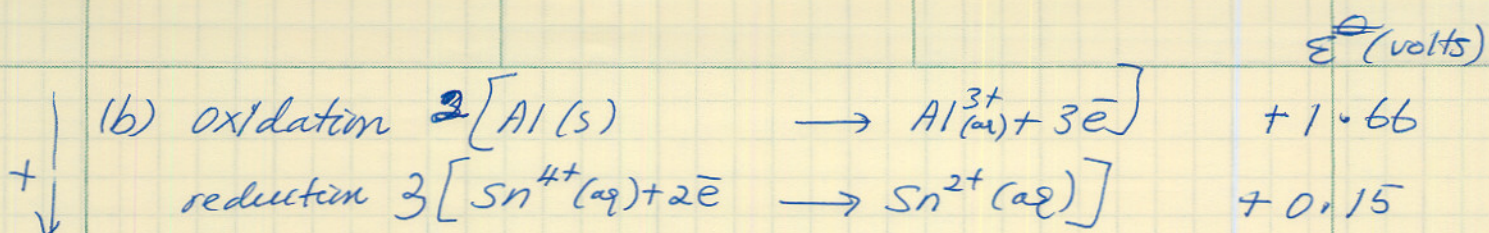
Reaction is product favored.



Not product favored.

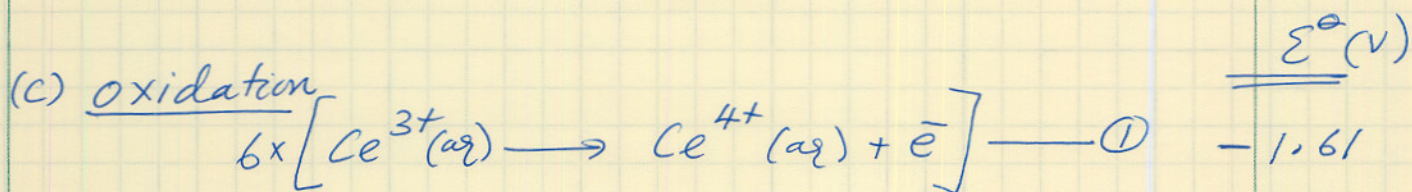


h

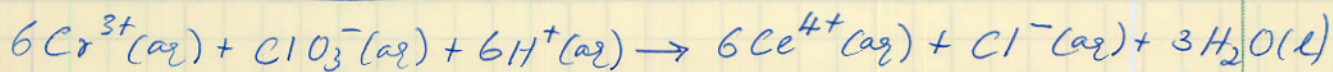
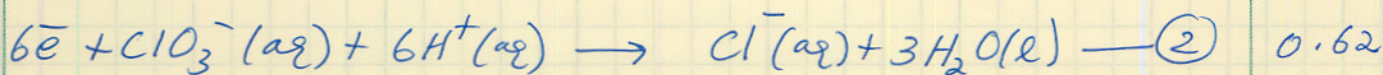


$$E_{cell}^\ominus = (1.66 + 0.15)V = \underline{\underline{1.81 \text{ volts}}}$$

Reaction is product favored.

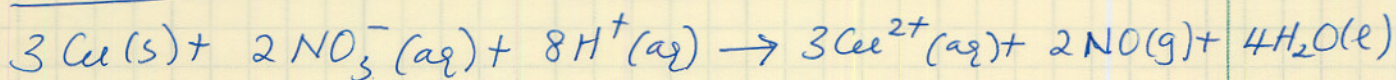
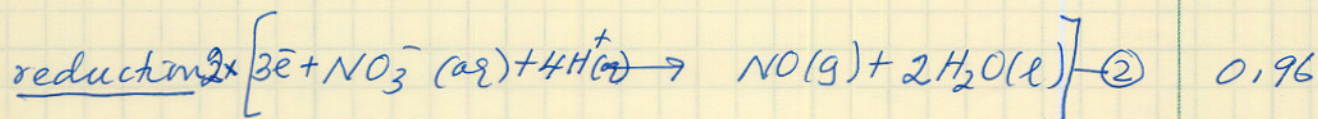
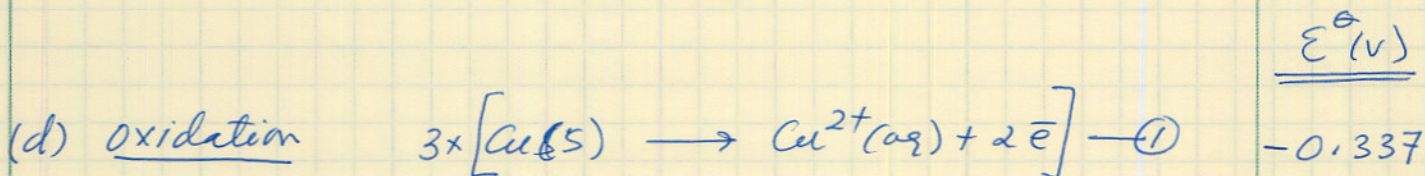


reduction



$$E_{cell}^\ominus = (-1.61 + 0.62)V = \underline{\underline{-0.99V}}$$

not product favored.

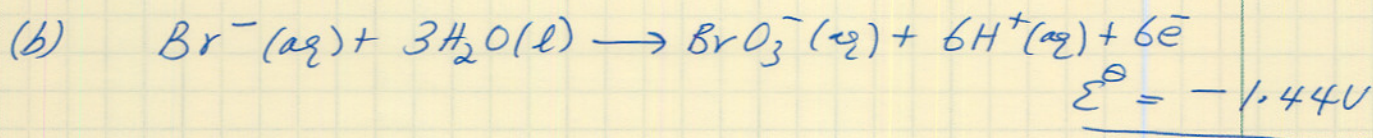


$$E_{cell}^\ominus = (-0.337 + 0.96)V = \underline{\underline{0.623V}}$$

Product favored reaction

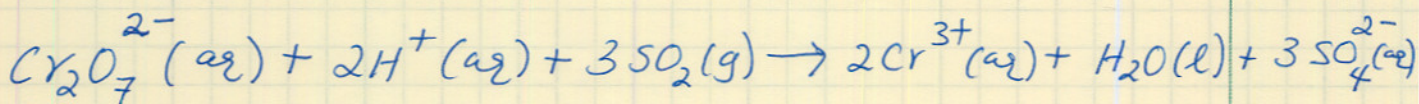
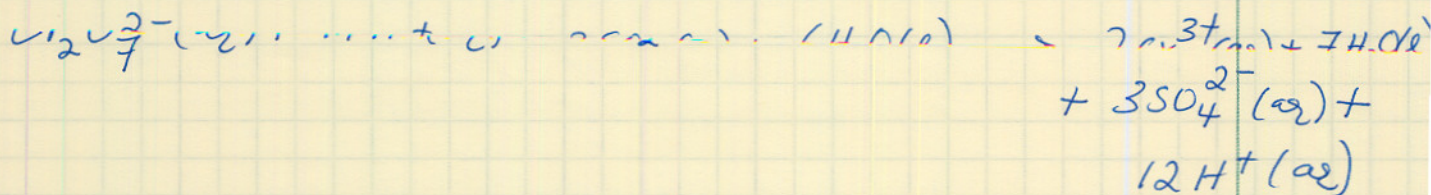
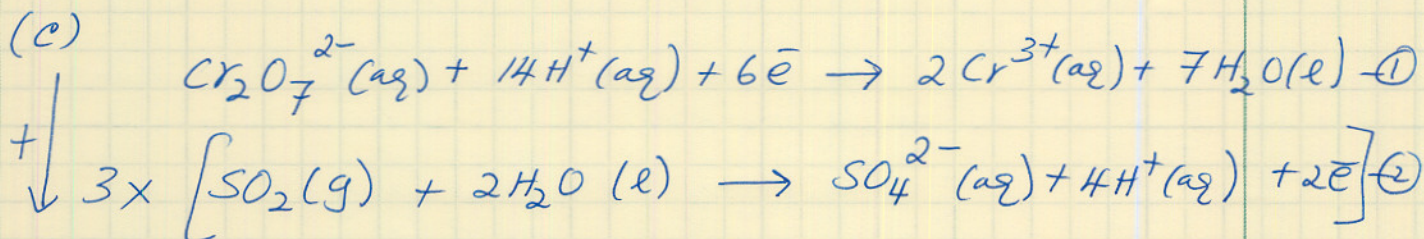


- (18) (a) Strongest oxidizing agent =  $\text{MnO}_4^- (\text{aq})$   
 weakest oxidizing agent =  $\text{SO}_4^{2-} (\text{aq})$



In order for ~~the~~ <sup>an</sup> oxidizing agent to promote the above reaction, the half-cell voltage of that reaction must be greater than 1.44V.

$\therefore \text{MnO}_4^- (\text{aq})$  ~~and~~ is the only one

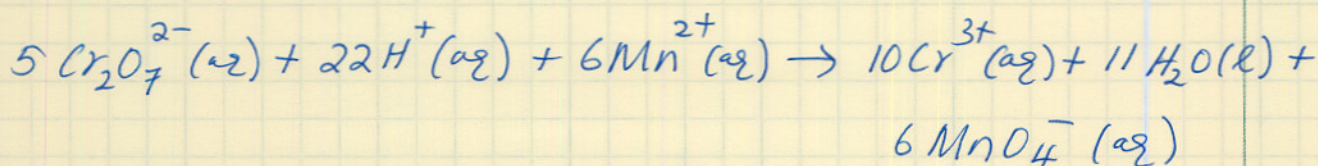
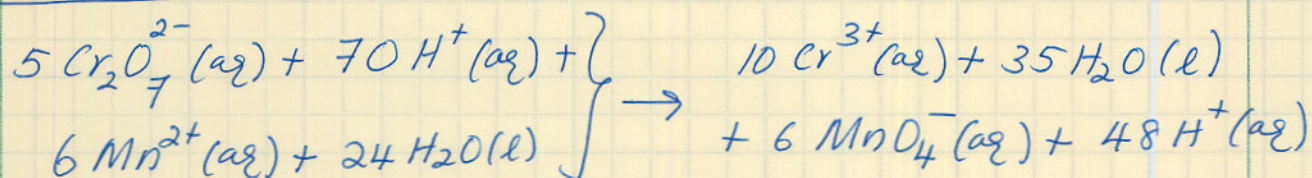
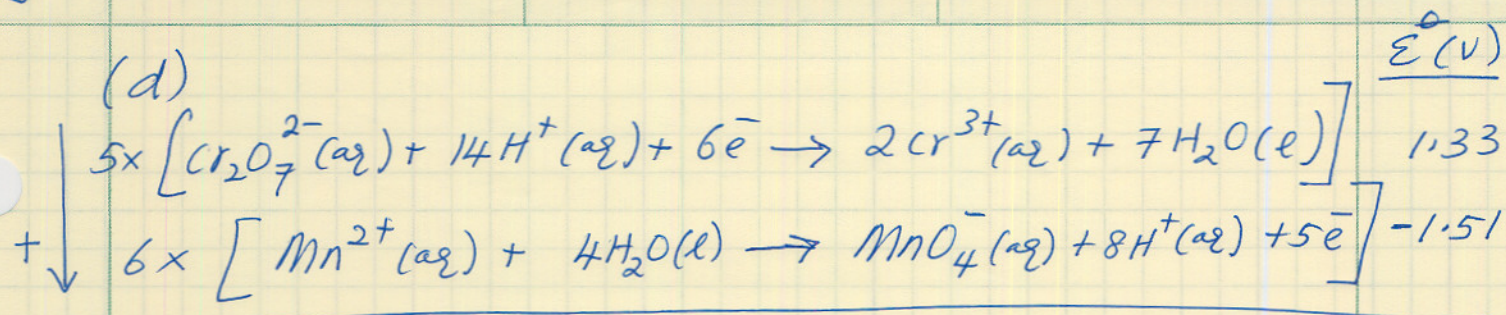


$\underline{\underline{\mathcal{E}_{\text{cell}}^\ominus = \mathcal{E}_1^\ominus + \mathcal{E}_2^\ominus = [1.33 + (-0.20)]\text{V} = 1.13\text{V}}}$

Reaction is product favored

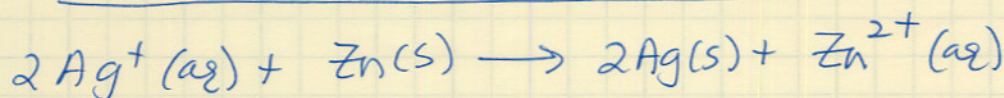
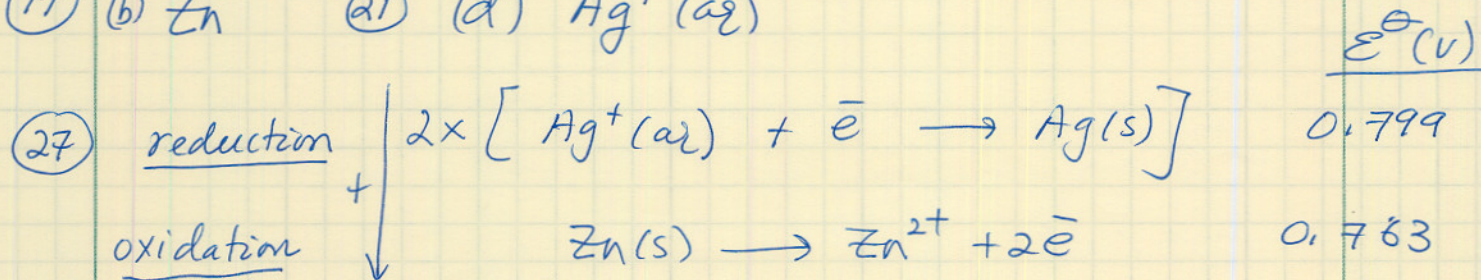
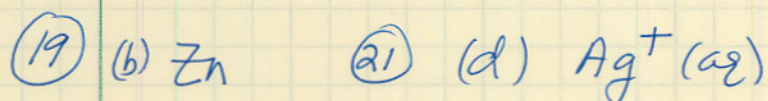


j



$$E_{cell}^{\ominus} = (1.33 - 1.51)V = \underline{\underline{-0.18V}}$$

Reactant favored.



$$E_{cell}^{\ominus} = (0.799 + 0.763)V = \underline{\underline{1.562V}}$$

$$E_{cell} = E_{cell}^{\ominus} - \frac{0.0257}{n} \ln Q \quad \text{assume } T = \underline{\underline{298K}}$$



k

$$E_{\text{cell}} = 1.562 \text{ V} - \frac{0.0257}{2} \ln Q$$

$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]^2} = \frac{0.010}{(0.25)^2} = 0.16$$

$$E_{\text{cell}} = 1.562 \text{ V} - \frac{0.0257}{2} \ln(0.16) \\ = \underline{\underline{1.586 \text{ V}}}$$

(29) cell reaction is the same as above.

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0257}{2} \ln Q$$

$$1.48 \text{ V} = 1.562 \text{ V} - \frac{0.0257}{2} \ln Q$$

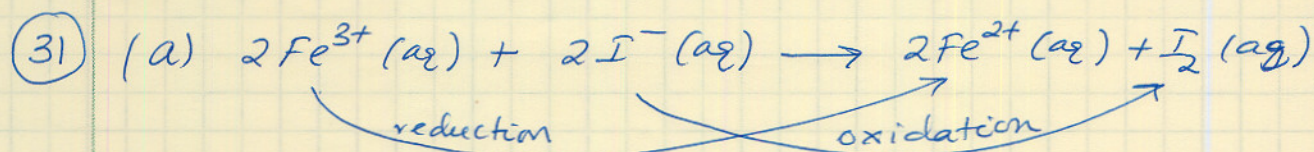
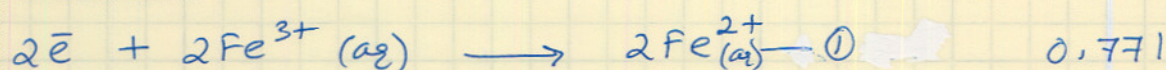
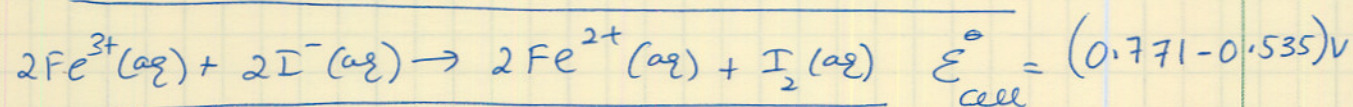
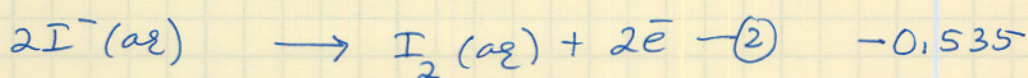
$$\text{cross } \ln Q = \frac{(1.562 - 1.48) \cdot 2}{0.0257} = 6.381$$

$$Q = 590.71 = \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]^2} = \frac{1}{[\text{Ag}^+]^2}$$

$$[\text{Ag}^+]^2 = \frac{1}{590.71} = 1.693 \times 10^{-3}$$

$$[\text{Ag}^+] = \underline{\underline{4.11 \times 10^{-2} \text{ M}}}$$



lreductionoxidation

$$\varepsilon_{\text{cell}}^{\ominus} = 0.236\text{V}$$

$$\Delta G_{\text{rn}}^{\ominus} = -nF\varepsilon_{\text{cell}}^{\ominus} = -(2\text{mol})\left(\frac{96,500\text{C}}{\text{mol}}\right)(0.236\text{V})$$

$$= -45,548\text{J}$$

$$\underline{\underline{\Delta G_{\text{rn}}^{\ominus} = -45.548\text{kJ}}}$$

$$\Delta G_{\text{rn}}^{\ominus} = -RT \ln K \Rightarrow \ln K = -\frac{\Delta G_{\text{rn}}^{\ominus}}{RT}$$

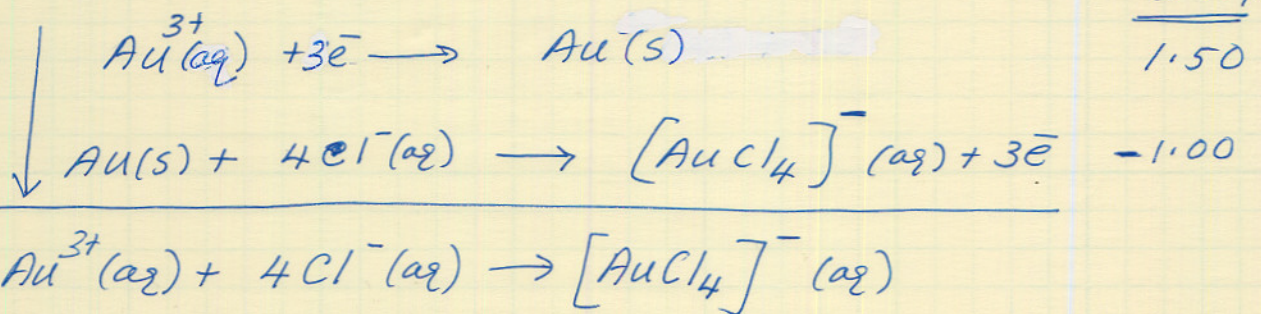
$$\ln K = \frac{-45.548 \times 10^3 \text{J}}{(8.314 \text{J K}^{-1} \text{mol}^{-1})(298\text{K})} = 18.38$$

$$\underline{\underline{K = 9.64 \times 10^7}}$$



m

(35)



$$E_{cell}^\ominus = (1.50 - 1.00)V = \underline{0.50V}$$

$$\begin{aligned} \Delta G_m^\ominus &= \Delta G_{formation}^\ominus = -nF E_{cell}^\ominus = -(3 \text{ mol}) \left( \frac{96,500 \text{ C}}{\text{mol}} \right) (0.50 \text{ V}) \\ &= -144,750 \text{ J} \\ &= -144.750 \text{ kJ} \end{aligned}$$

$$\Delta G_{formation}^\ominus = -RT \ln K$$

$$\ln K = \frac{-\Delta G_{formation}^\ominus}{RT} = \frac{144.750 \times 10^3 \text{ J}}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}$$

$$\ln K = 58.424$$

$$K = \underline{\underline{2.36 \times 10^{25}}}$$