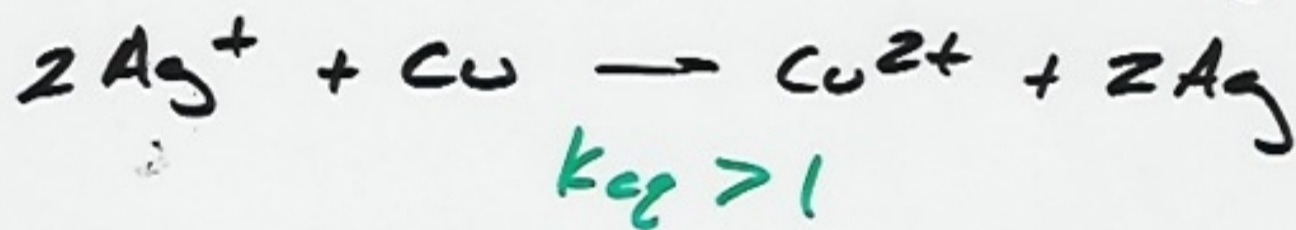
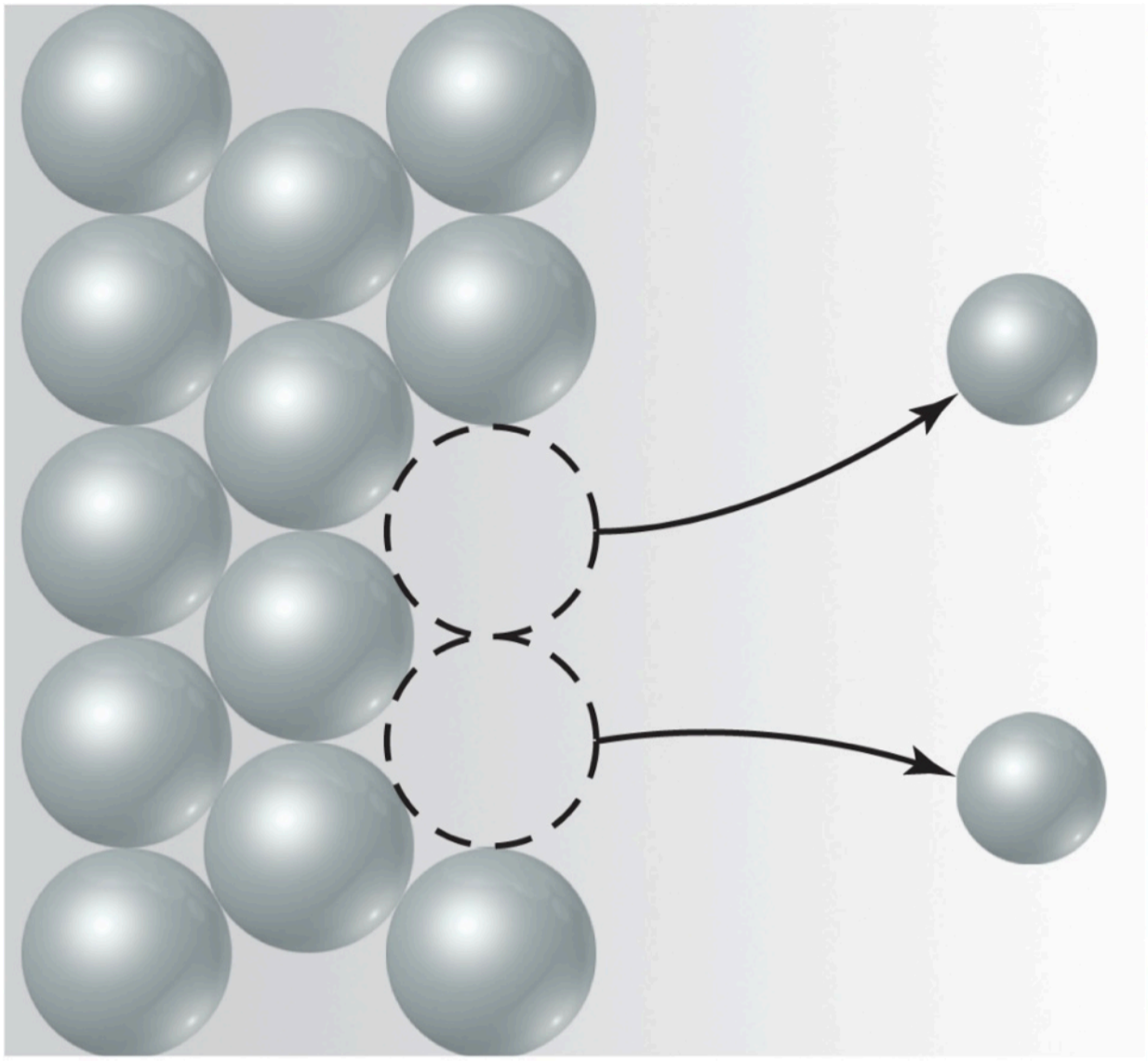


1] Acid/Base reactions are H^+ transfer reactions

Red/Ox reactions are e^- transfer reactions

equilibrium lies to the side of the weaker oxidizing agent.

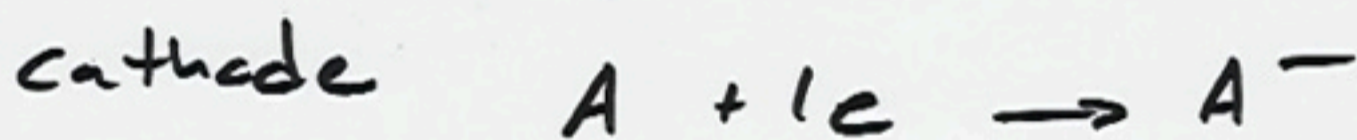
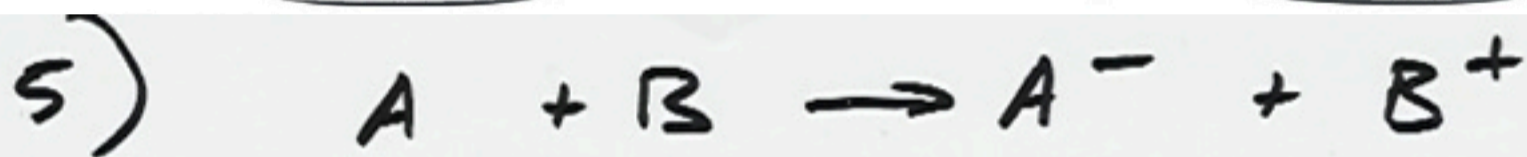
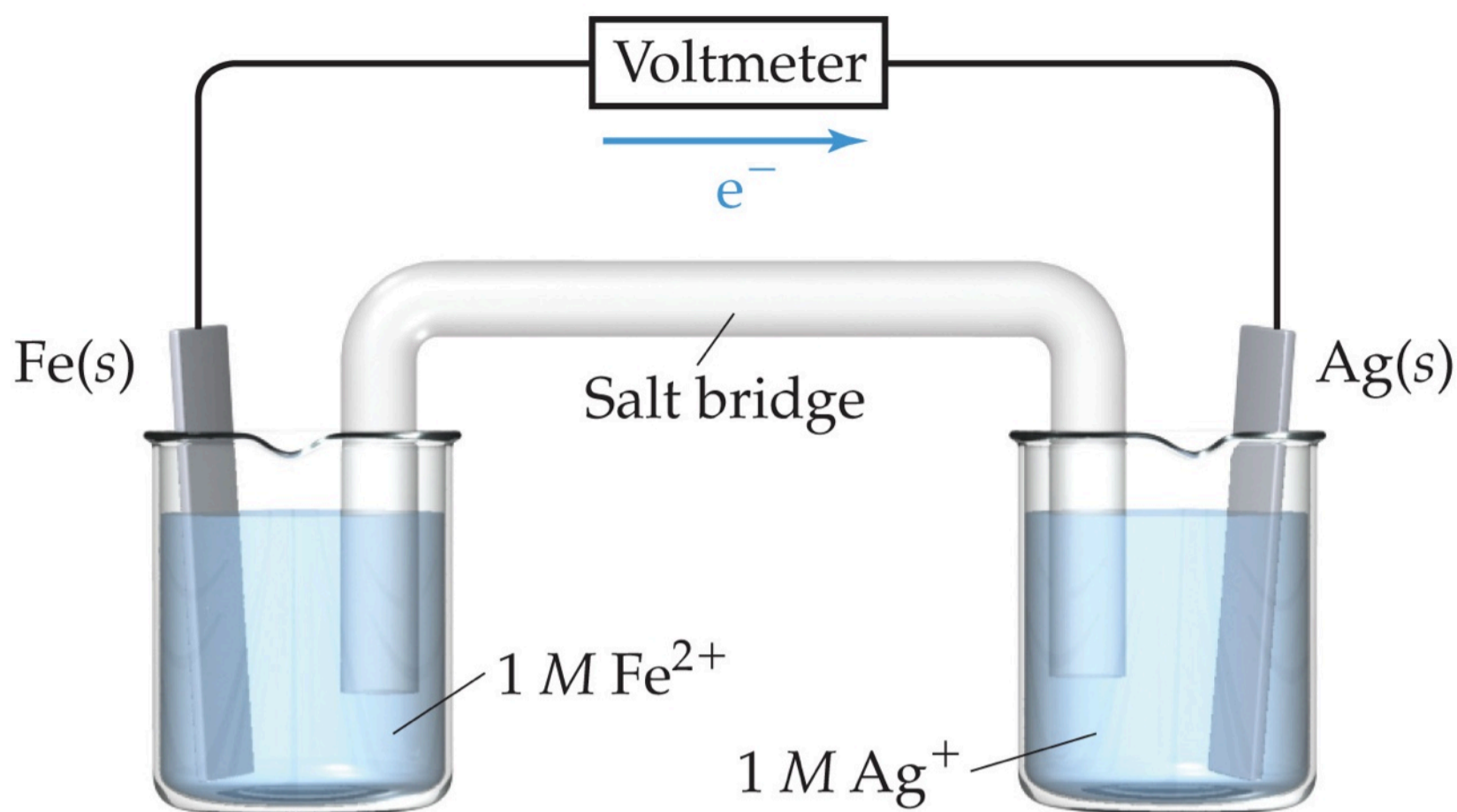




3] The diagram represents oxidation occurring at the Anode



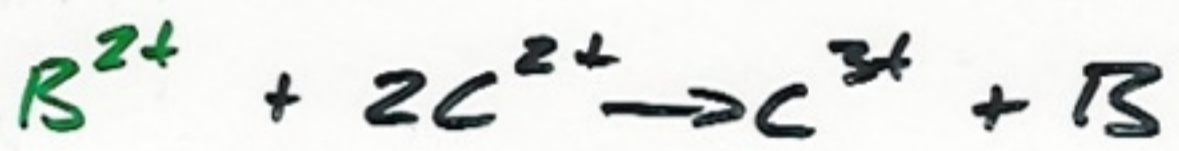
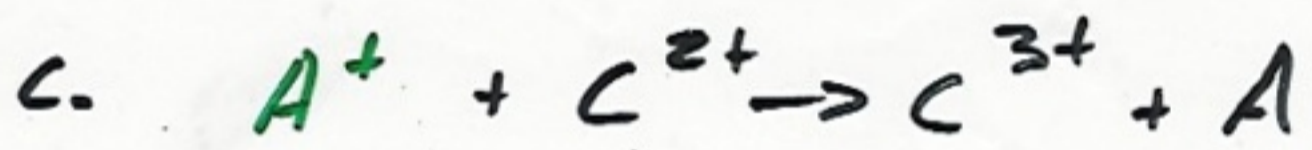
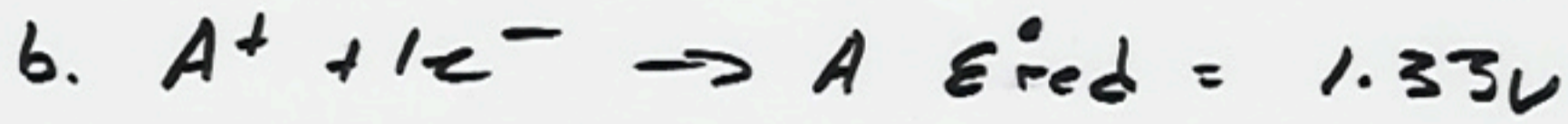
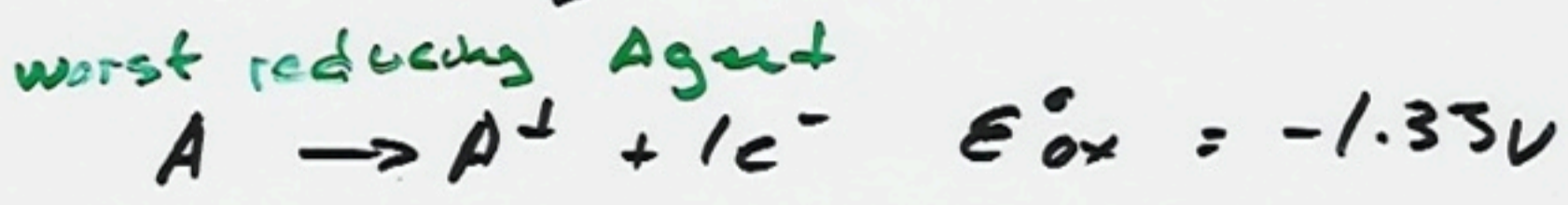
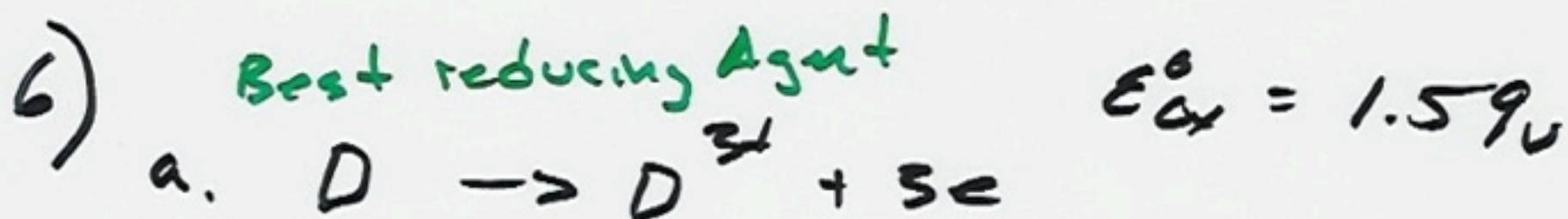
metal atoms being oxidized to form cations.



b. the anode would be at a higher potential energy



c. E°_{cell} is (+) in a spontaneous reaction



7 $E^{\circ}_{\text{cell}} = (-)$ $\Delta G^{\circ} = -nFE^{\circ}$ $(+)$ $(-)$

a. $\Delta G^{\circ} = (+)$

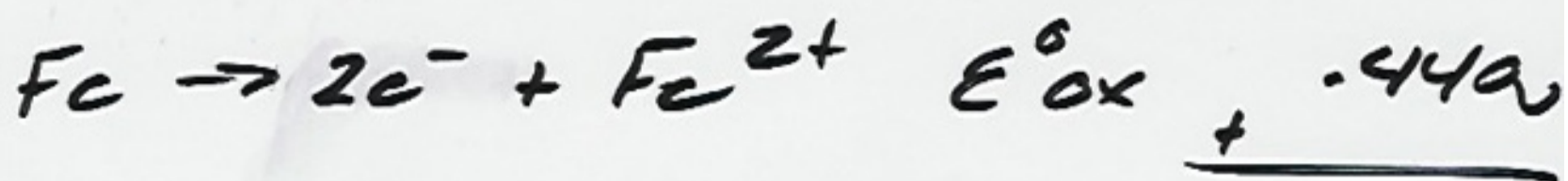
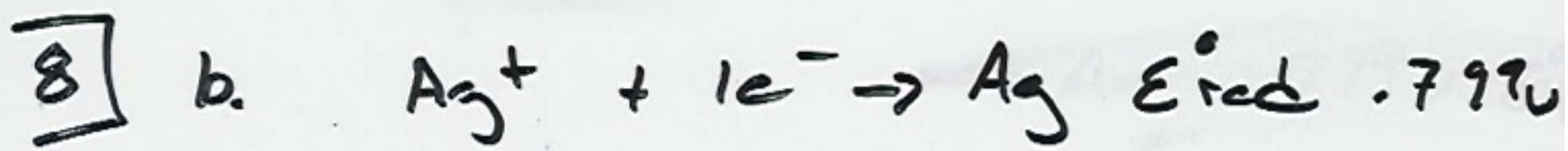
b. $E^{\circ} = \frac{-0.0592}{n} \log K$ $K < 1$ $(-)$

c. ? depends on Q

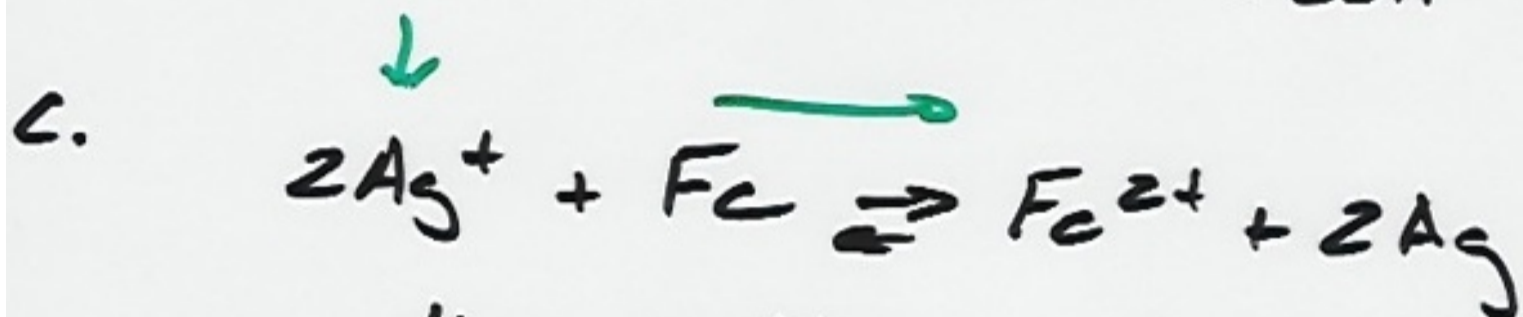
$$E = E^{\circ} - \frac{0.0592}{n} \log Q$$

8 a. by convention the cathode is on the right



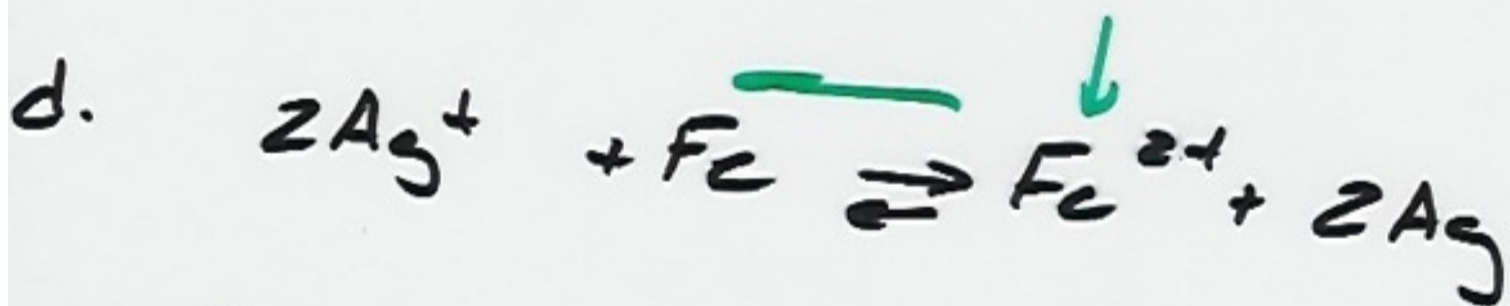


$E^{\circ}_{\text{cell}} \quad 1.24\text{V}$



the voltage will increase

$$E = 1.24 - \frac{.0592}{2} \log\left(\frac{1}{(10)^2}\right)$$
$$= 1.30\text{V}$$



$$E = 1.24 - \frac{.0592}{2} \log\left(\frac{10}{(1)^2}\right)$$
$$= 1.21\text{V}$$

13 a. Oxidation - losing electrons

b. electrons appear on the product side of the $\frac{1}{2}$ reaction



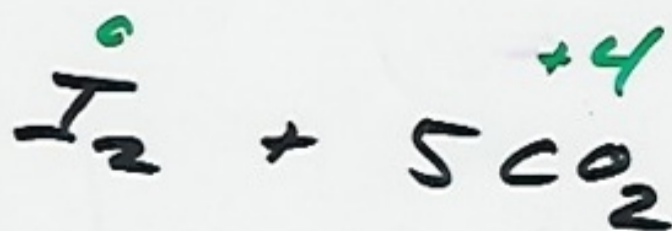
c. Oxidizing agent is the electron taker. (it is reduced)



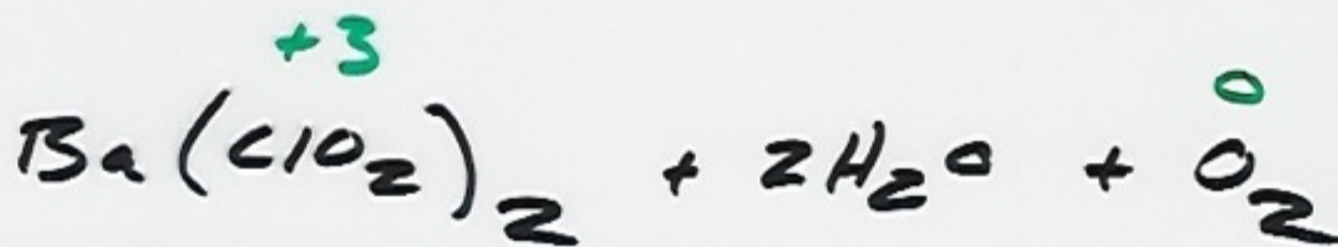
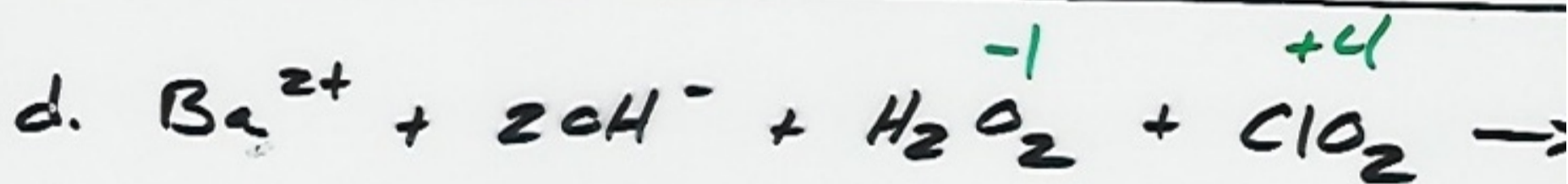
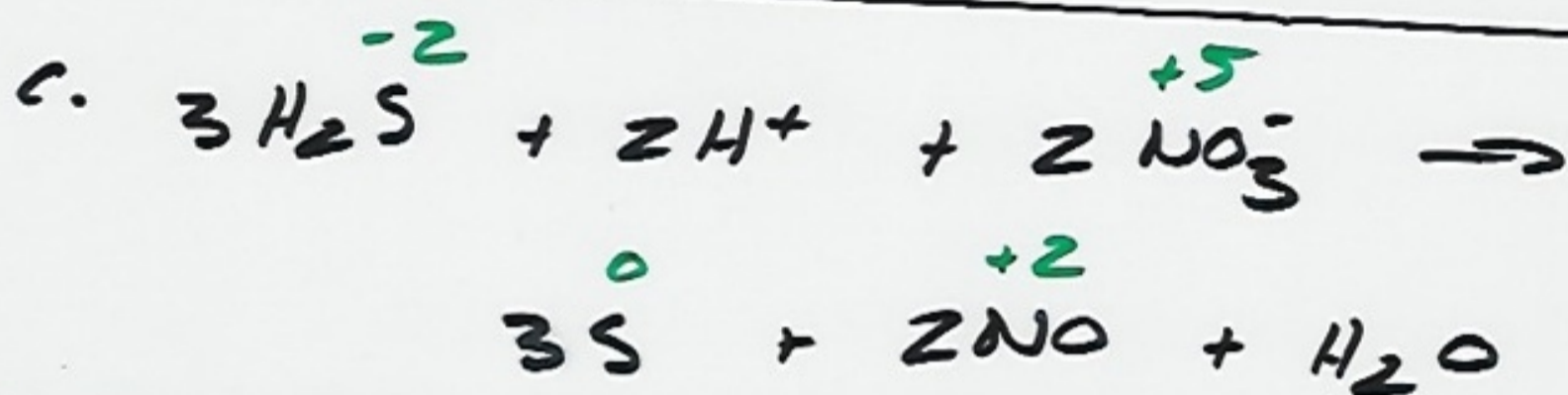
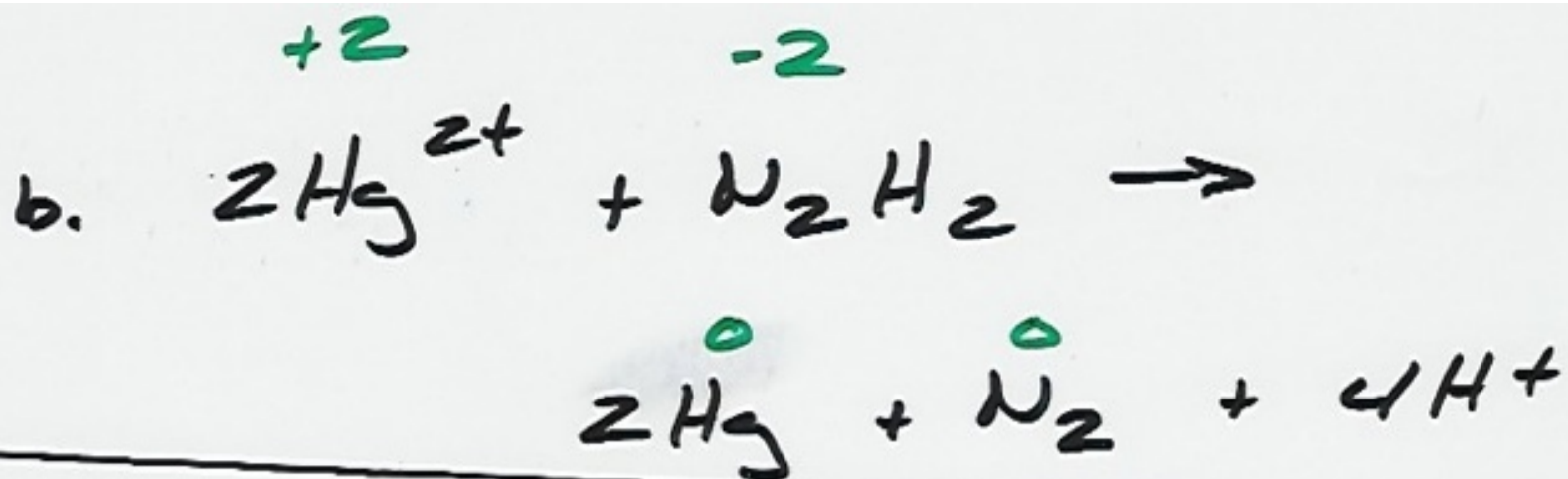
15 a. True

b. False: $\text{Fe}^{3+} + \text{Co}^{2+} \rightarrow \text{Co}^{3+} + \text{Fe}^{2+}$
oxidizing agent

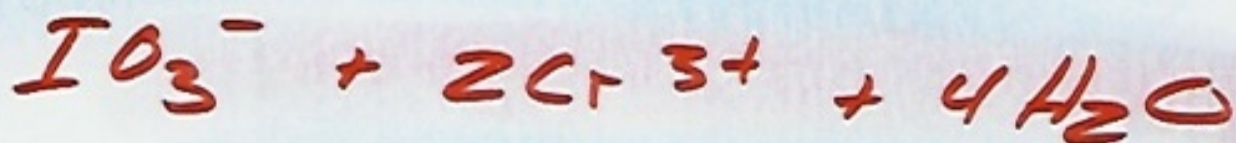
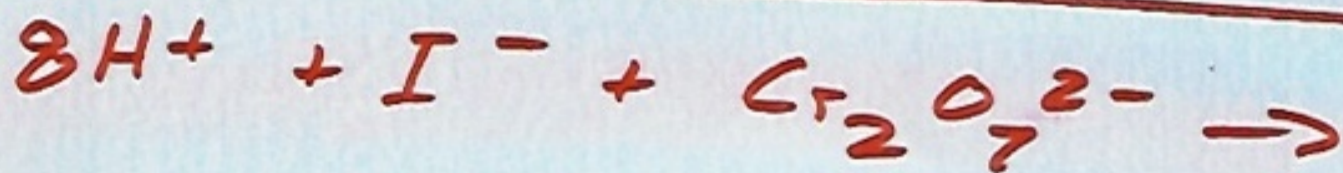
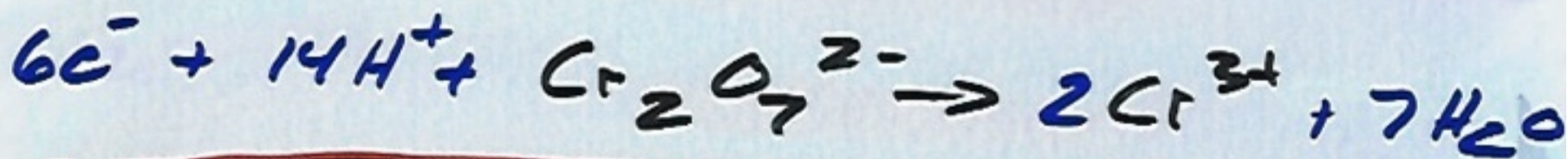
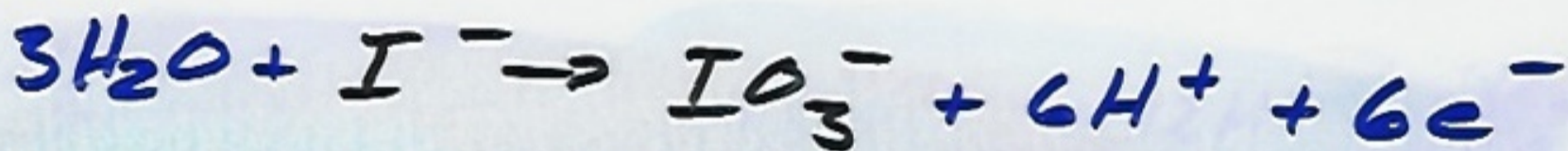
c. True



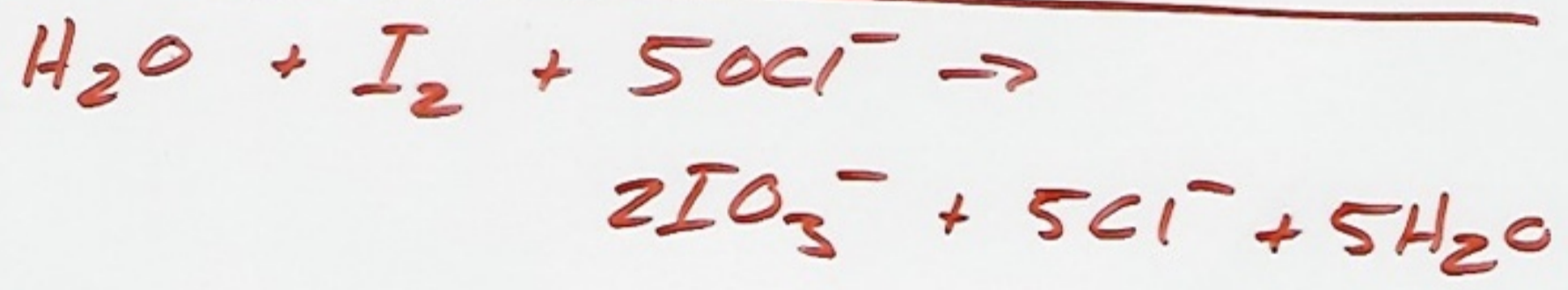
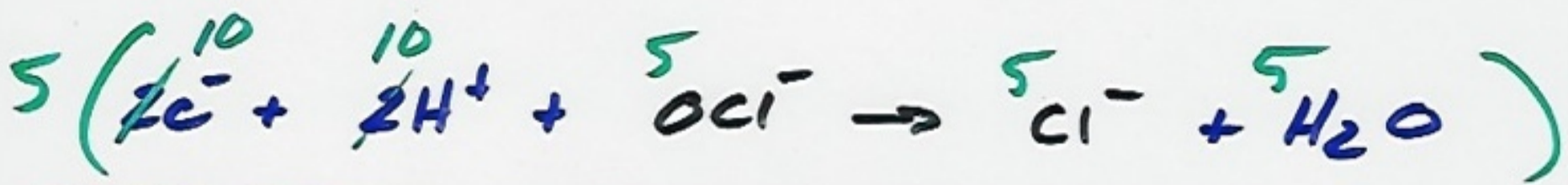
17



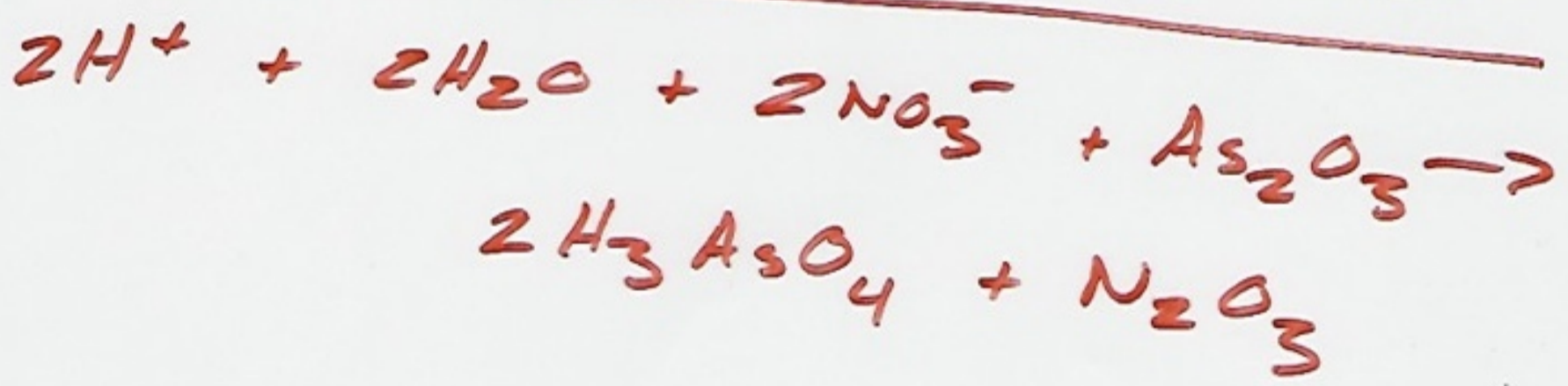
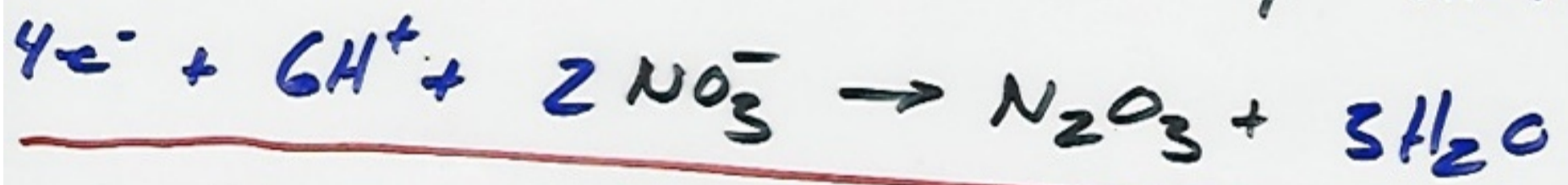
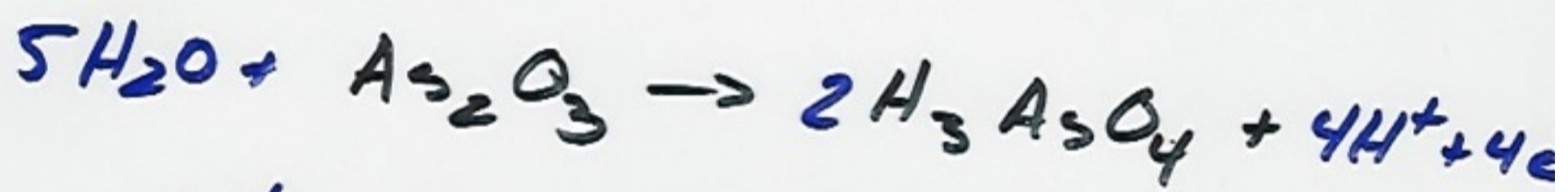
23



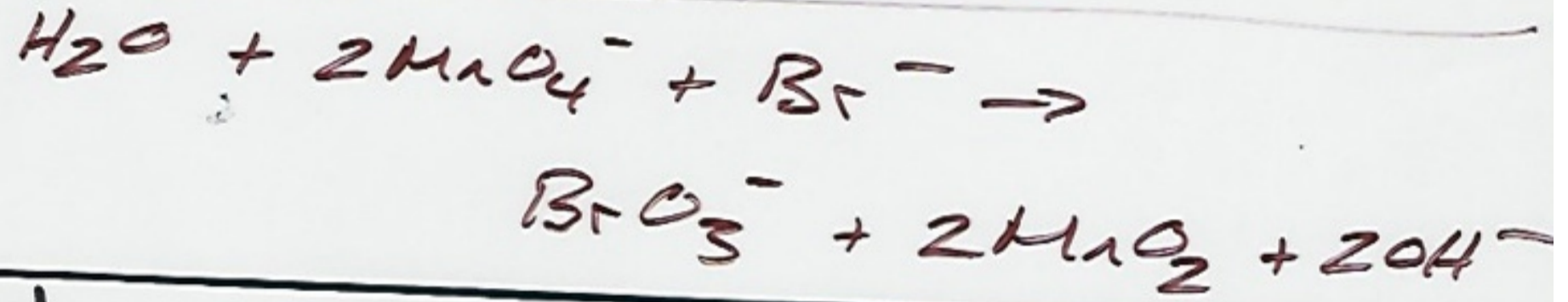
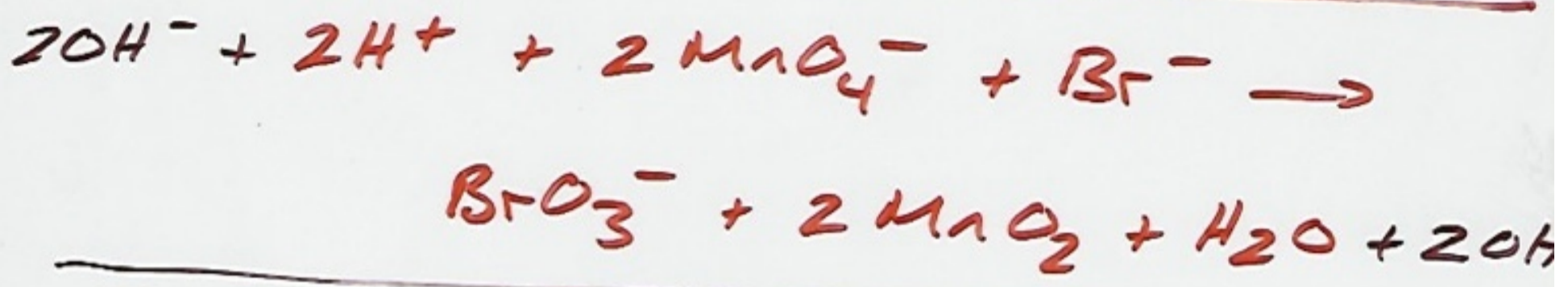
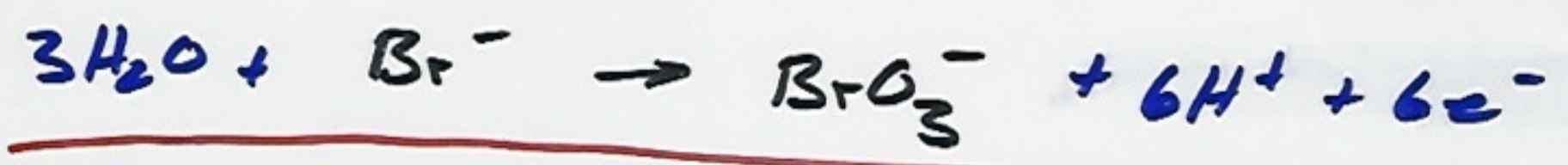
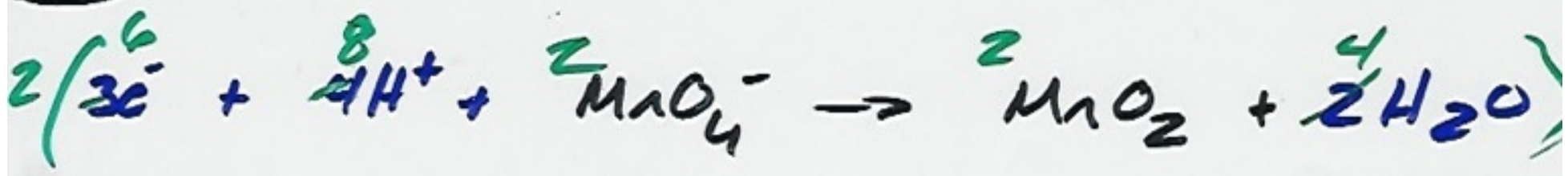
23 c.



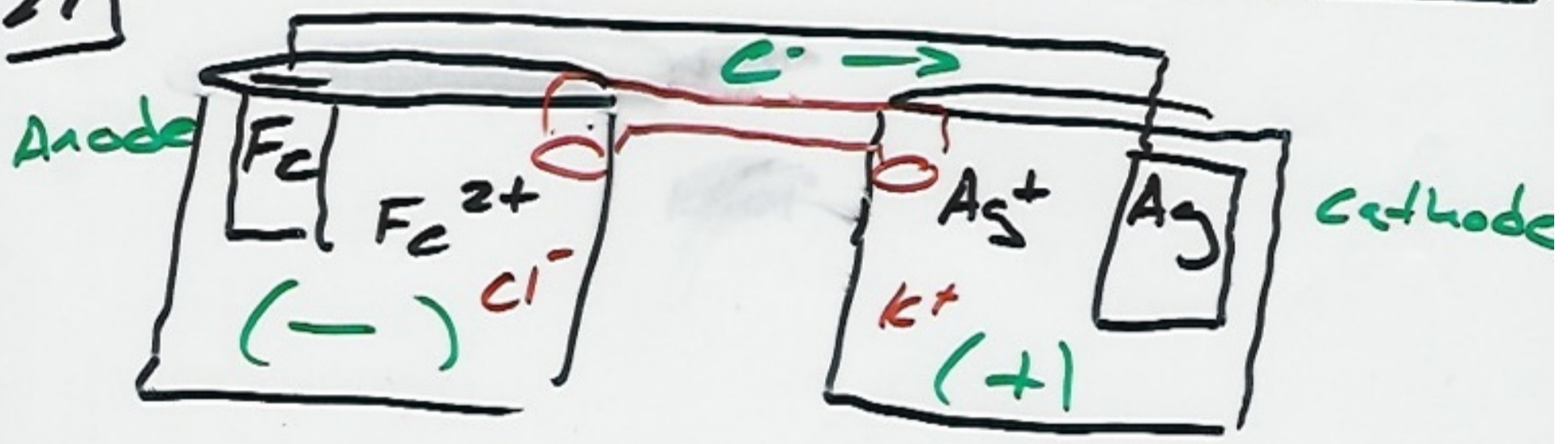
d.



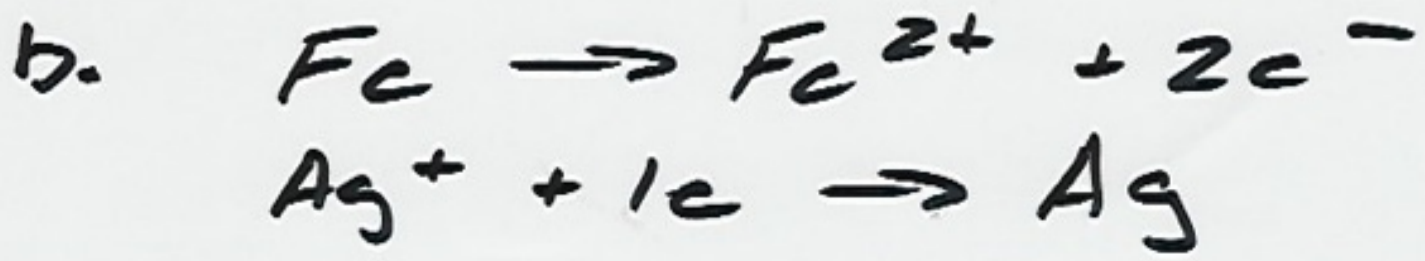
23) c.



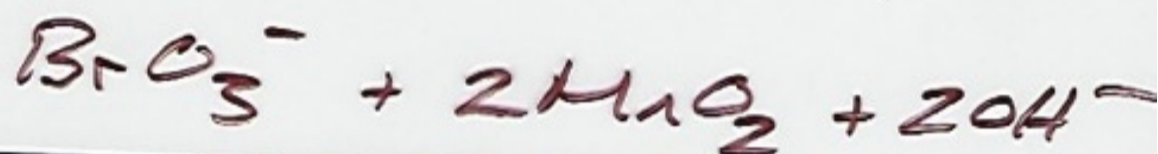
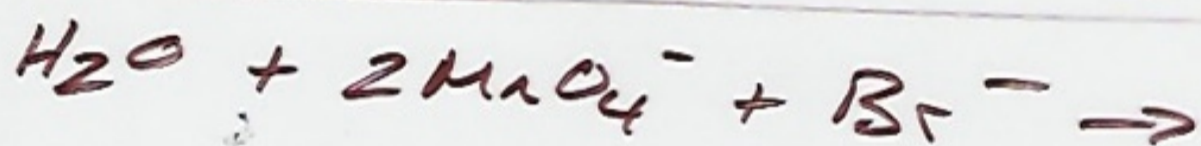
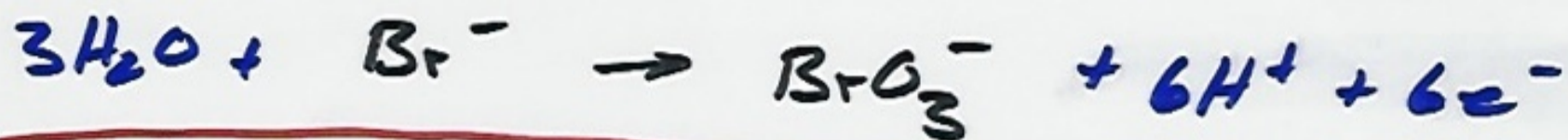
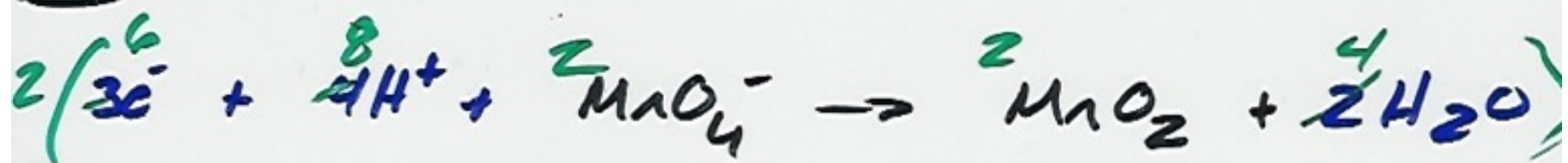
27



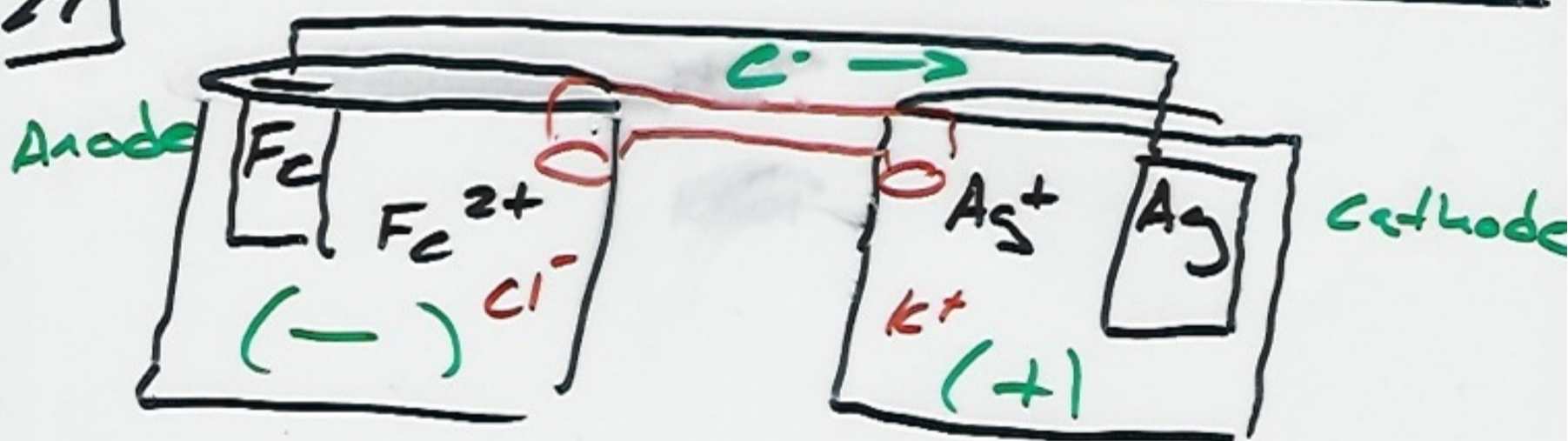
- a. Fe is oxidized
- Ag⁺ is reduced



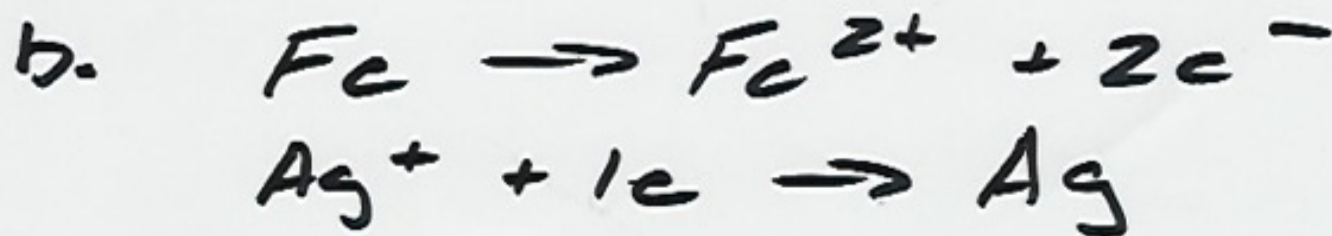
23) c.



27



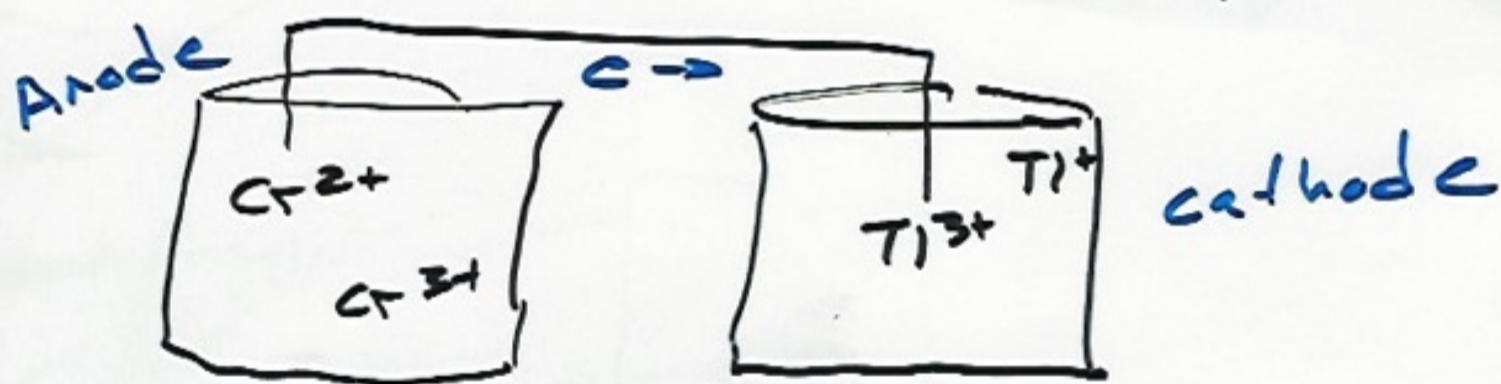
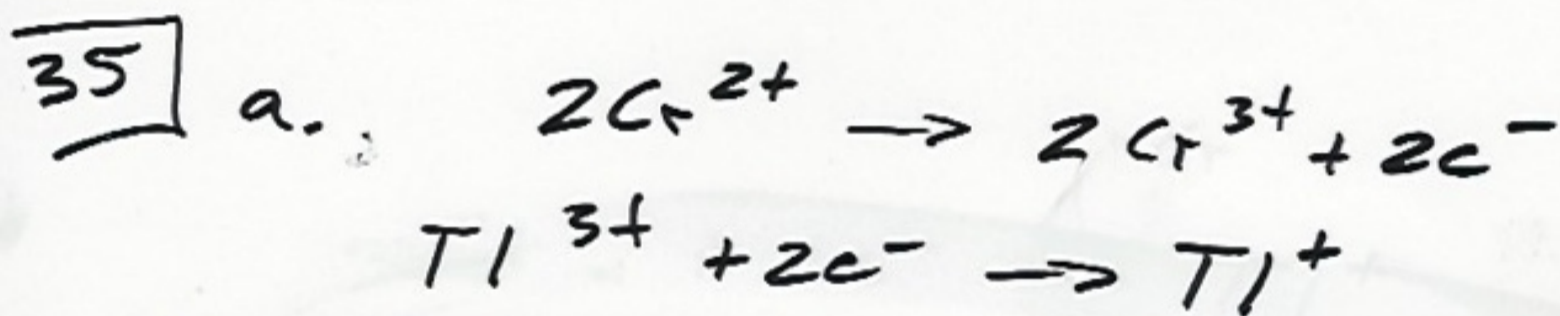
- a. Fe is oxidized
Ag⁺ is reduced



29 a. EMF - is the push (volts) on the electrons going thru the wire.

b. $V = \frac{J}{C}$

c. E_{cell} is the voltage of the cell.

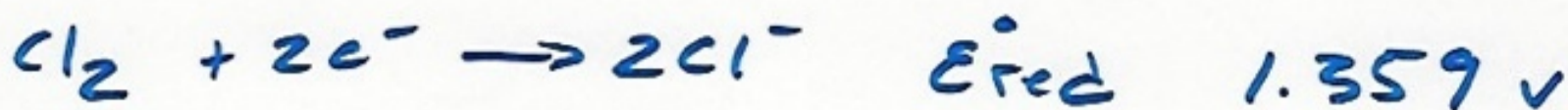
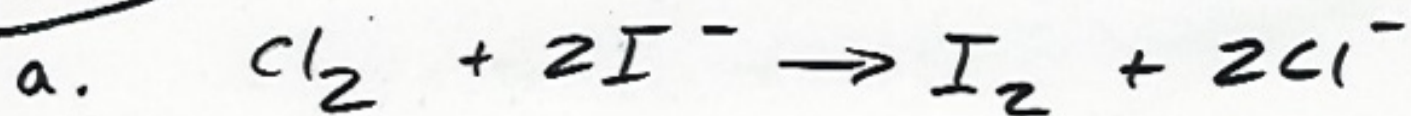


$$E_{cell}^{\circ} = E_{red}^{\circ} + E_{ox}^{\circ}$$

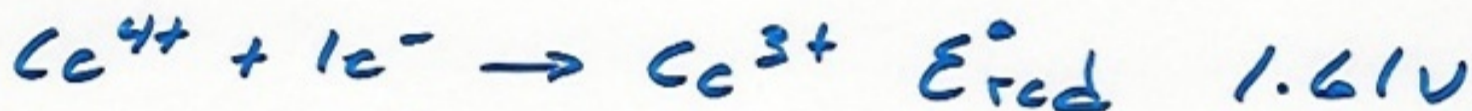
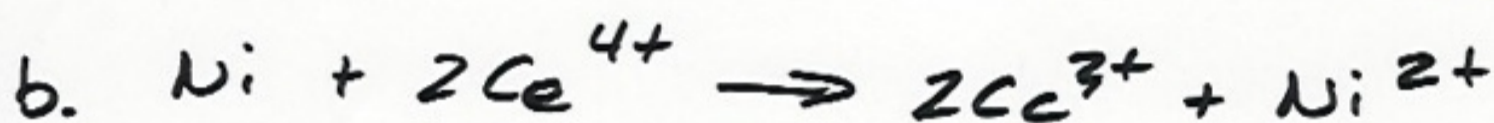
$$1.19V = E_{red}^{\circ} + .41V$$

$$E_{red}^{\circ} = .78V$$

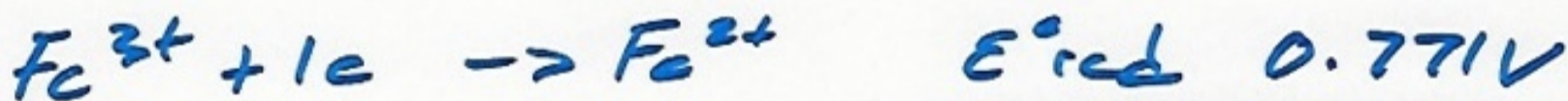
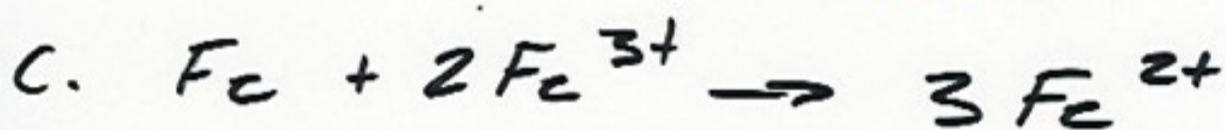
37



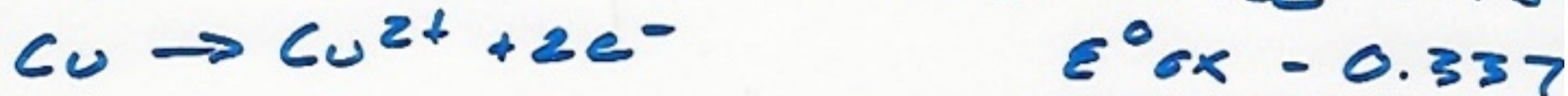
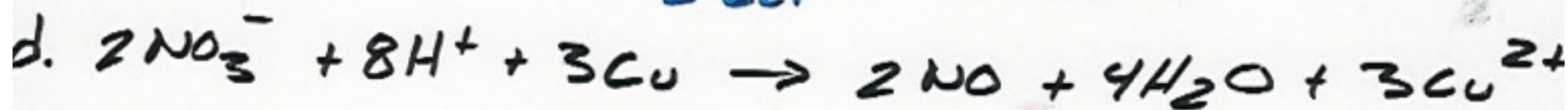
$$E_{\text{cell}}^{\circ} \quad 0.823 \text{ V}$$



$$E_{\text{cell}}^{\circ} \quad 1.89 \text{ V}$$



$$E_{\text{cell}}^{\circ} \quad 1.211 \text{ V}$$



$$E_{\text{cell}}^{\circ} \quad 0.62 \text{ V}$$

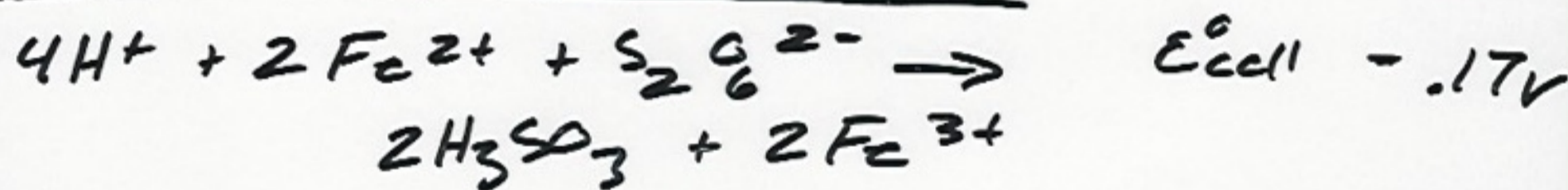
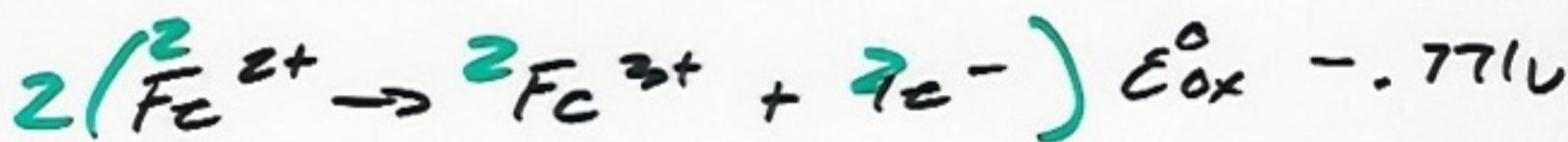
44

- a. Cl_2 c. ClO_3^-
 b. Cd^{2+} d. O_3

45

- a. Cl_2 oxidizing agent
 b. MnO_4^- oxidizing agent
 c. Ba reducing agent
 d. Zn reducing agent

51

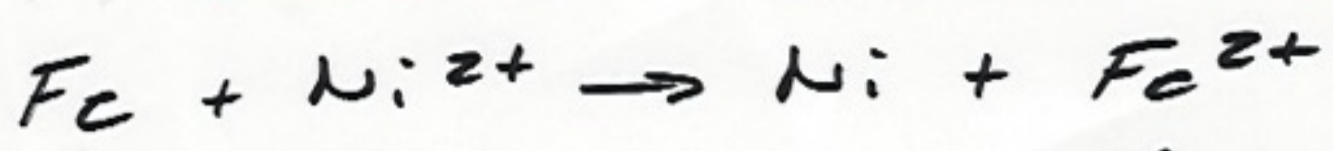


$$\Delta G^{\circ} = -nFE^{\circ}$$

$$= (-2\text{mol e}^-) \left(\frac{96485\text{C}}{1\text{mol e}^-} \right) \left(\frac{-.17\text{J}}{\text{C}} \right)$$

$$= 32.8\text{ kJ}$$

55 a.
$$E_{\text{cell}}^{\circ} = \frac{.0592 \text{ V}}{n} \log K$$



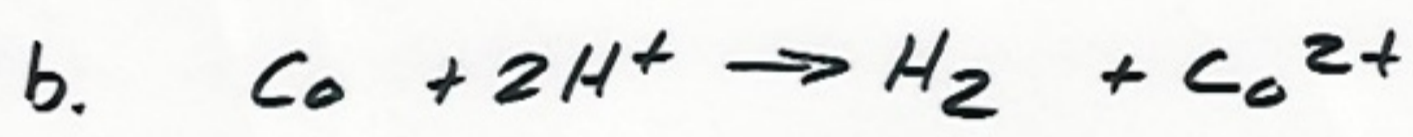
$$E_{\text{ox}}^{\circ} = .440 \text{ V}$$

$$E_{\text{red}}^{\circ} = -.28 \text{ V}$$

$$E_{\text{cell}}^{\circ} \quad .16 \text{ V}$$

$$0.16 \text{ V} = \frac{.0592}{2} \log K$$

$$K = 2.5 \times 10^5$$



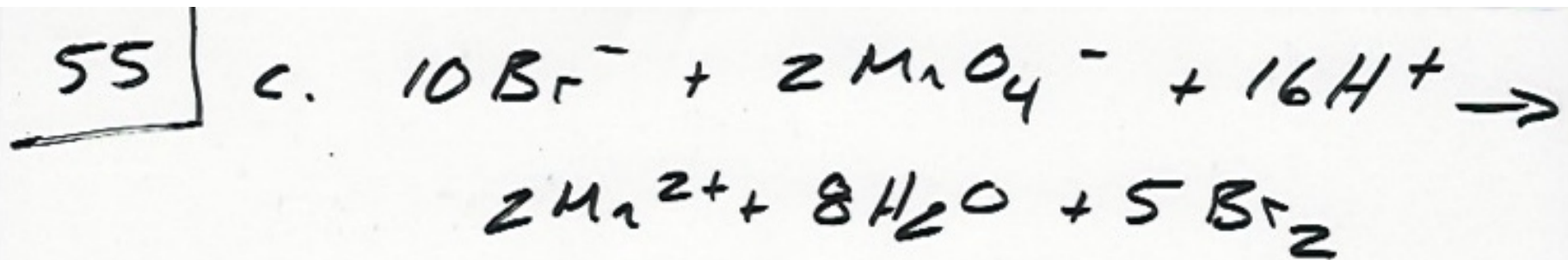
$$E_{\text{ox}}^{\circ} \quad .277$$

$$E_{\text{red}}^{\circ} \quad 0$$

$$0.277$$

$$0.277 = \frac{.0592}{2} \log K$$

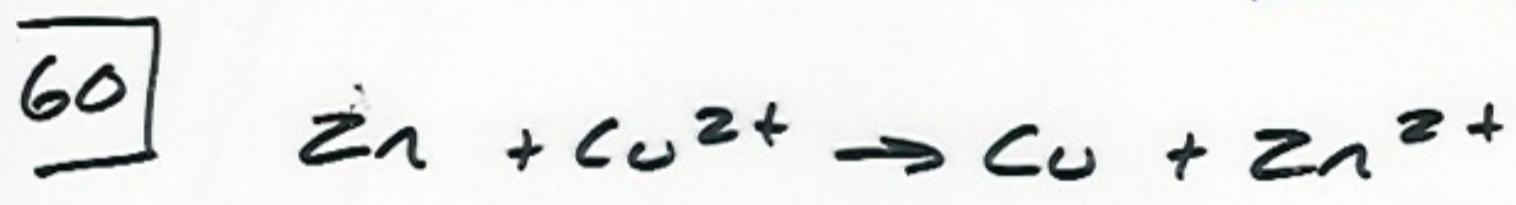
$$K = 2.3 \times 10^9$$



E°_{red}	1.51
E°_{ox}	-1.065
E°_{cell}	<u>0.445</u>

$$0.445\text{V} = \frac{-0.0592}{10} \log K$$

$$K = 1.5 \times 10^{75}$$



$$\Delta G^\circ = -nFE^\circ$$

E°_{red}	0.337
E°_{ox}	<u>0.763</u>
E°_{cell}	1.10

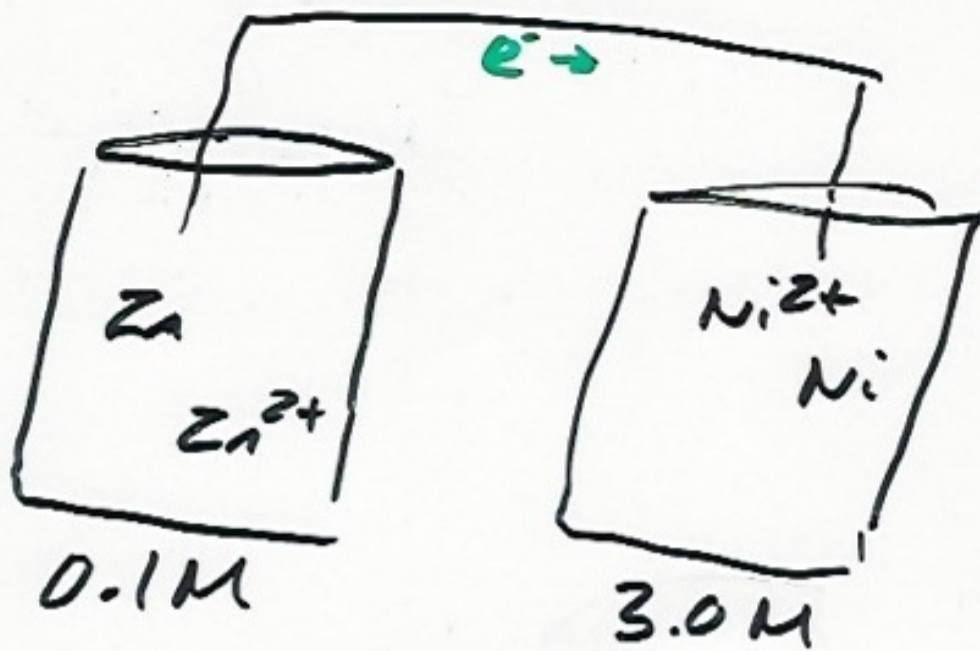
$$= 2\text{mol } e^- \times \frac{96485\text{C}}{1\text{mol } e^-} \times \frac{1.10\text{J}}{\text{C}}$$

$$\Delta G^\circ = \frac{212.3\text{kJ}}{1\text{mol Cu}}$$

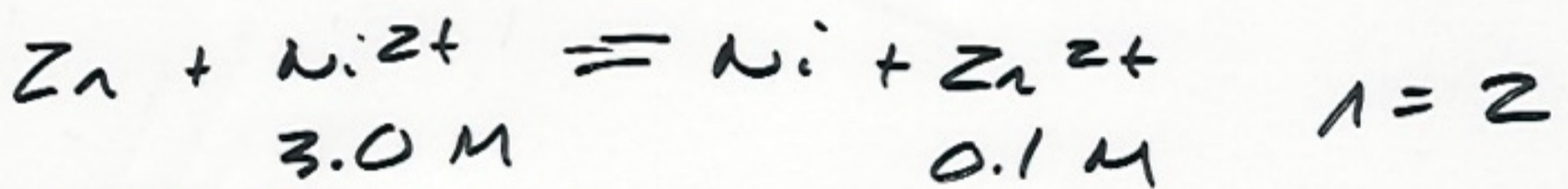
$$50\text{g Cu} \times \frac{1\text{mol Cu}}{63.55\text{g}} \times \frac{212.3\text{kJ}}{1\text{mol Cu}} = -167\text{kJ work required}$$

65

a.



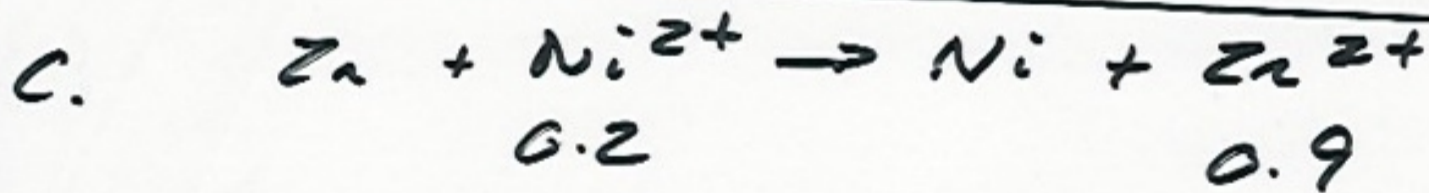
$$\begin{array}{r}
 E_{\text{ox}}^{\circ} \quad .763 \text{ V} \\
 E_{\text{red}}^{\circ} \quad -.28 \text{ V} \\
 \hline
 E_{\text{cell}}^{\circ} \quad .483 \text{ V}
 \end{array}$$



$$E = E^{\circ} - \frac{.0592}{n} \log Q$$

$$E = 0.483 - \frac{.0592}{2} \log \frac{(0.1)}{(3)}$$

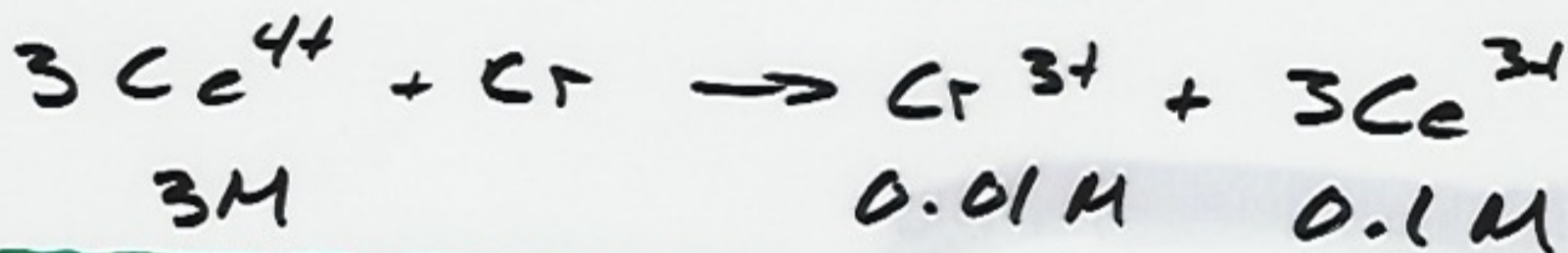
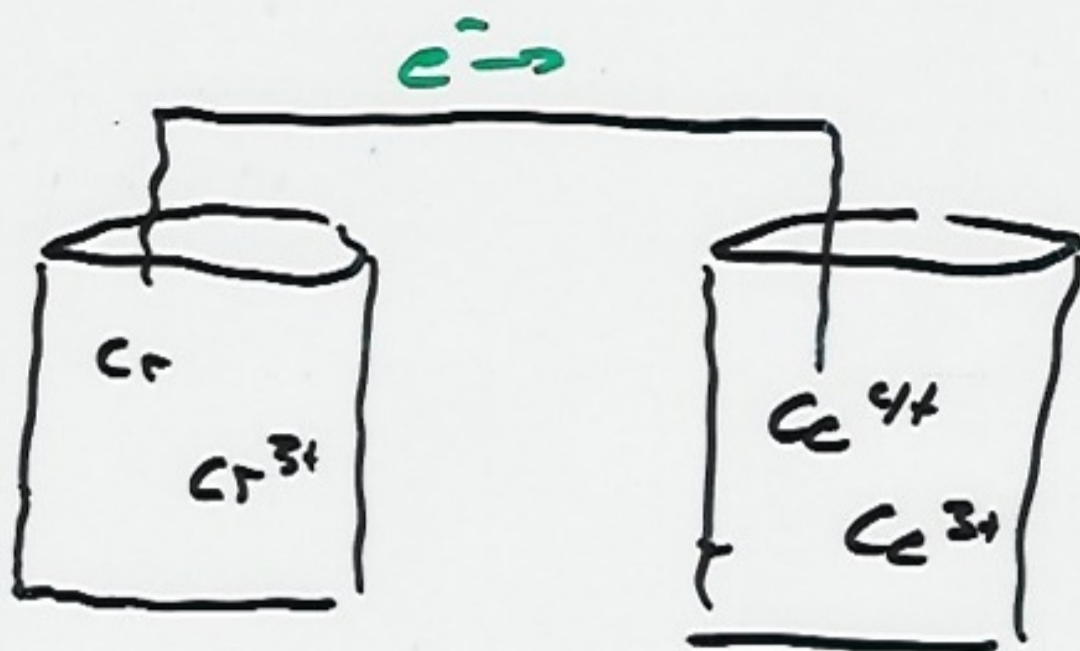
$$E = 0.53 \text{ V}$$



$$E = 0.483 - \frac{.0592}{2} \log \frac{0.9}{0.2}$$

$$E = 0.46 \text{ V}$$

66



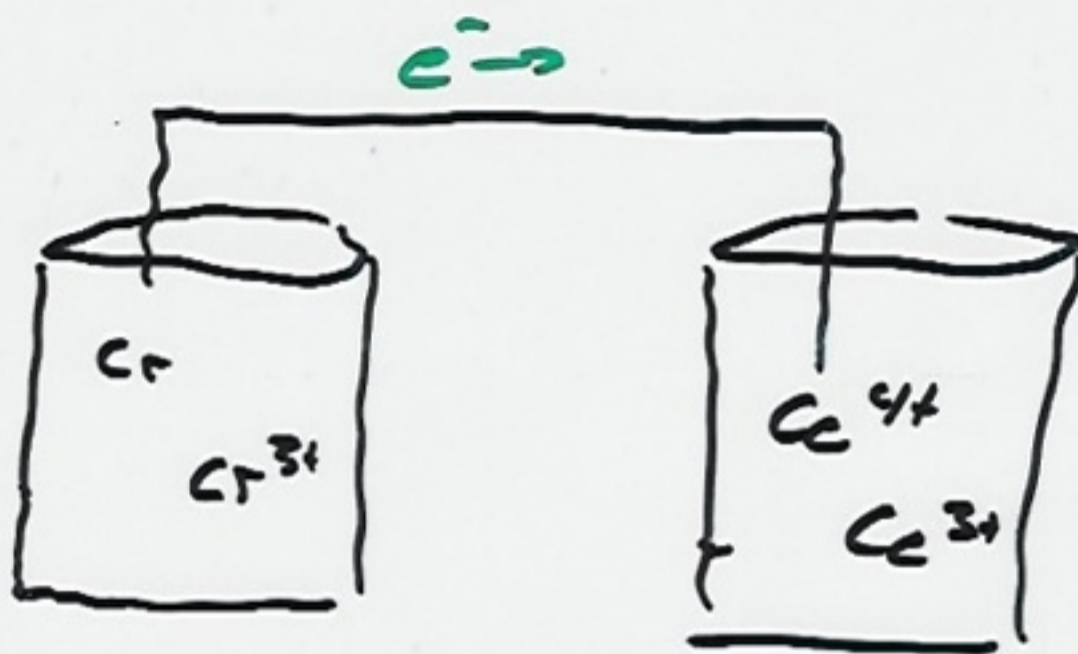
$$E = E^{\circ} - \frac{.0592}{n} \log Q$$

$$n = 3$$

E°_{red}	1.61V
E°_{ox}	.71V
<hr/>	
E°_{cell}	2.35V

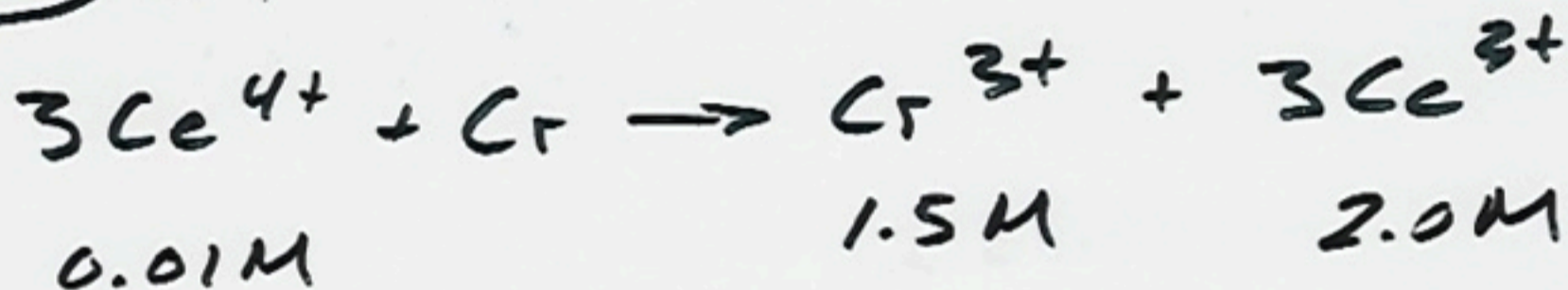
$$E = 2.35V - \frac{.0592}{3} \log \left(\frac{(0.1)^3 (0.01)}{(3)^3} \right)$$

$$E = 2.48V$$



66

c.

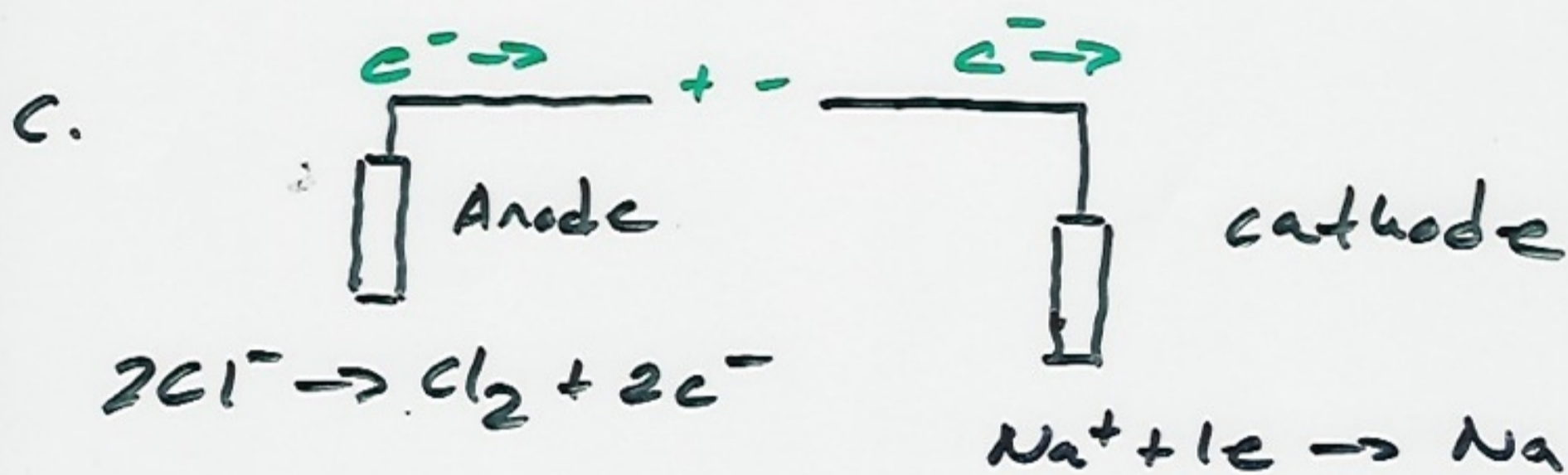


$$E = 2.35\text{V} - \frac{0.0592}{3} \log \left(\frac{(2.0)^3 (1.5)}{(0.01)^3} \right)$$

$$E = 2.21\text{V}$$

89) a. Electrolysis uses electrical energy (from the surroundings) to drive nonspontaneous chemical reactions (the system) forward.

b. nonspontaneous



d. The reduction of Na^+ can not occur in an aqueous environment because the electrical energy would go toward the electrolysis of H_2O .

$$\begin{aligned}
 & \boxed{91} \quad 2 \text{ days} \times \frac{24 \text{ hr}}{1 \text{ day}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} \\
 & \times \frac{7.6 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mole } e^-}{96485 \text{ C}} \times \frac{\text{Cr}^{3+}}{3 \text{ mole } e^-} \times \frac{52 \text{ g}}{\text{mol Cr}} \\
 & = 235.9 \text{ g Cr collected}
 \end{aligned}$$

$$\begin{aligned}
 \text{b. } & 0.25 \text{ mol Cr} \times \frac{3 \text{ mole } e^-}{1 \text{ mol Cr}^{3+}} \times \frac{96485 \text{ C}}{1 \text{ mole } e^-} \\
 & \times \frac{1}{8 \text{ hr}} \times \frac{1 \text{ hr}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} = \frac{2.51 \text{ C}}{\text{s}} \\
 & \text{or } 2.51 \text{ A}
 \end{aligned}$$

