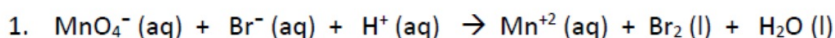
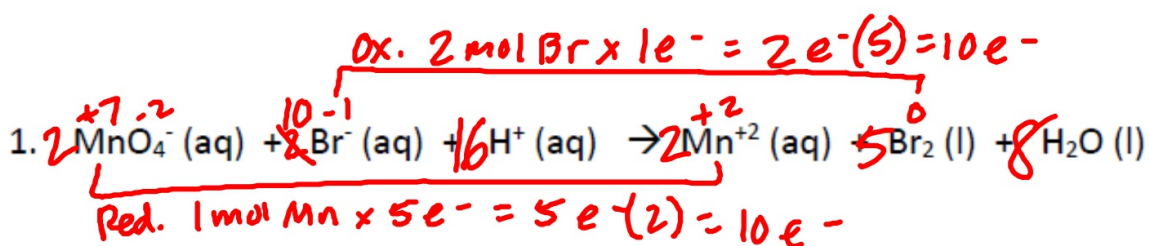


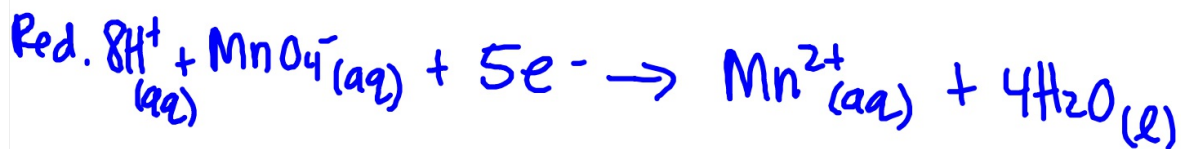
ΔG , K_{eq} , and Nernst



- a. After balancing the above reaction with the $\frac{1}{2}$ box method how many electrons are transferred in the reaction represented by the balanced eq'n above?
A) 2 B) 4 C) 5 D) 8 E) 10
- b. Write the 2 half reaction equations.
- c. Calculate E°_{cell} for the redox reaction above.
- d. Calculate ΔG°_{rxn} for the above reaction.
- e. Calculate K_{eq} .



- a. After balancing the above reaction with the $\frac{1}{2}$ box method how many electrons are transferred in the reaction represented by the balanced eq'n above?
A) 2 B) 4 C) 5 D) 8 E) 10
- b. Write the 2 half reaction equations.



c. Calculate E°_{cell} for the redox reaction above.

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$$

$$E^\circ_{\text{cell}} = 1.51 \text{ V} - 1.065 \text{ V} = \boxed{0.45 \text{ V}}$$

d. Calculate $\Delta G^\circ_{\text{rxn}}$ for the above reaction.

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

$$\Delta G^\circ = -(10 \text{ mol } e^-)(96485 \frac{\text{J}}{\text{mol} \cdot \text{V}})(0.45 \text{ V})$$

$$\boxed{\Delta G^\circ = -4.3 \times 10^5 \text{ J}} \quad \text{spontaneous rxn}$$

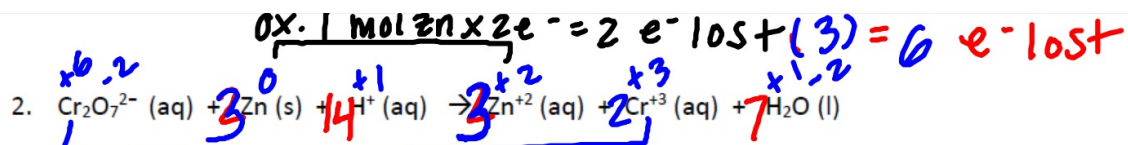
e. Calculate K_{eq} .

$$\Delta G = -RT \ln(K_{\text{eq}})$$

$$-4.3 \times 10^5 \text{ J} = -\left(8.31 \frac{\text{J}}{\text{K}}\right)(298 \text{ K}) \ln(K_{\text{eq}})$$

$$\ln(K_{\text{eq}}) = 173$$

$$K_{\text{eq}} = e^{173} = \boxed{2.6 \times 10^{75}} \quad \text{product favored}$$

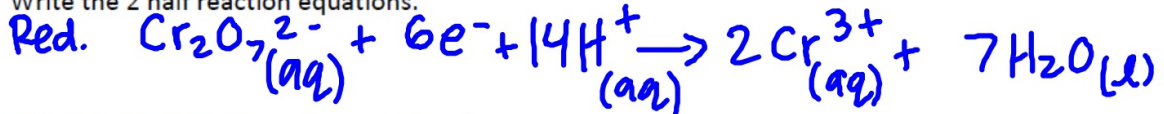


Red. $2 \text{ mol Cr} \times 3e^- = 6e^- \text{ gained}$

- a. After balancing the above rxn with the 1/2 box method, how many electrons are transferred?

$6e^-$

- b. Write the 2 half reaction equations.



- c. Calculate E°_{cell} for the redox reaction above.

Ox. $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2e^-$

$E^\circ_{\text{cell}} = 1.33 \text{ V} + 0.763 \text{ V} = 2.09 \text{ V}$

- d. Calculate $\Delta G^\circ_{\text{rxn}}$ for the above reaction.

$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -(6 \text{ mol } e^-)(96485 \frac{\text{J}}{\text{mol} \cdot \text{V}})(2.09 \text{ V})$

$\Delta G^\circ = -1.21 \times 10^6 \text{ J}$

- e. Calculate K_{eq} .

$\Delta G = -RT \ln K_{\text{eq}}$

$-1.21 \times 10^6 \text{ J} = -(8.31 \frac{\text{J}}{\text{mol} \cdot \text{K}})(298 \text{ K}) \ln K_{\text{eq}}$

$\ln K_{\text{eq}} = 488.6$

$K_{\text{eq}} = e^{488.6} = \text{overflow}!!!$

Huge value

$K_{\text{eq}} = e^{488.6} = \text{greater than } 1$

$K \gg 1$

- f. E°_{cell} is the potential of the reaction as long as all solutions are 1.0 Molar and all gases (if the reaction has gases) are at 1.0 atm. If even one of these values changes then the E°_{cell} changes to E_{cell} .

The equation for finding E_{cell} is $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} (\ln Q)$

where $R = 8.31 \text{ J/mol}\cdot\text{K}$; $T = 298\text{K}$; $n = \text{moles of electrons transferred}$; $F = 96485 \text{ J/V}\cdot\text{mol}$; $Q = (P/R)$ where (aq) = [] and (g) = atm

For example: $\text{Ni (s)} + 2\text{H}^+ \text{ (aq)} \rightarrow \text{Ni}^{2+} \text{ (aq)} + \text{H}_2 \text{ (g)}$

$$Q = \frac{[\text{Ni}^{2+}] (P_{\text{H}_2})}{[\text{H}^+]^2}$$

* Find E_{cell} for the reaction in part A if $[\text{Zn}^{2+}] = 0.100\text{M}$ when everything else is standard.

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} (\ln Q)$$

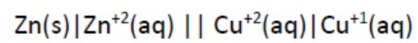
$$E_{\text{cell}} = 2.093 - \frac{(8.31)(298)}{(6)(96485)} (\ln(0.001))$$

$$E_{\text{cell}} = 2.12 \text{ V}$$

$$Q = \frac{[\text{Zn}^{2+}]^3 [\text{Cr}^{3+}]^2}{[\text{Cr}_2\text{O}_7^{2-}]^1 [\text{H}^+]^{14}}$$

$$Q = \frac{(0.1)^3 (1)^2}{(1)^1 (1)^{14}}$$

$$Q = .001$$



3. Write the balanced redox reaction, find the E°_{cell} , draw a picture of this cell with a salt bridge of KNO_3 and then predict how the voltage would change with each of the following:
- the products are less than standard molarity? More than standard molarity?
 - the reactants are less than standard molarity? More than standard molarity?