

AP Chemistry Homework Answers

Chapter 20 Principles of Reactivity: Electron Transfer Reactions (Kotz)

Problems: 3ab, 5ab, 9, 13, 17, 23, 25, 29, 31, 34, 41, 45, 49, 59

- 20.3 (a)  $\text{Ag(s)} \rightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$   
 $\text{NO}_3^-(\text{aq}) + 2 \text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$   


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 $\text{Ag(s)} + \text{NO}_3^-(\text{aq}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ag}^+(\text{aq}) + \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$
- (b)  $2[\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 \text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\ell)]$   
 $5[\text{HSO}_3^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 3 \text{H}^+(\text{aq}) + 2 \text{e}^-]$   


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 $2 \text{MnO}_4^-(\text{aq}) + \text{H}^+(\text{aq}) + 5 \text{HSO}_3^-(\text{aq}) \rightarrow 2 \text{Mn}^{2+}(\text{aq}) + 3 \text{H}_2\text{O}(\ell) + 5 \text{SO}_4^{2-}(\text{aq})$
- 20.5 (a)  $2[\text{Al(s)} + 4 \text{OH}^-(\text{aq}) \rightarrow 3 \text{e}^- + \text{Al(OH)}_4^-(\text{aq})]$   
 $3[2 \text{H}_2\text{O}(\ell) + 2 \text{e}^- \rightarrow 2 \text{OH}^-(\text{aq}) + \text{H}_2(\text{g})]$   

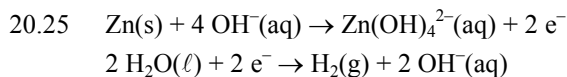

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 $2 \text{Al(s)} + 6 \text{H}_2\text{O}(\ell) + 2 \text{OH}^-(\text{aq}) \rightarrow 2 \text{Al(OH)}_4^-(\text{aq}) + 3 \text{H}_2(\text{g})$
- (b)  $2[\text{CrO}_4^-(\text{aq}) + 4 \text{H}_2\text{O}(\ell) + 3 \text{e}^- \rightarrow \text{Cr(OH)}_3(\text{s}) + 5 \text{OH}^-(\text{aq})]$   
 $3[\text{SO}_3^{2-}(\text{aq}) + 2 \text{OH}^-(\text{aq}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell) + 2 \text{e}^-]$   


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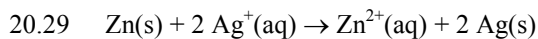
 $2 \text{CrO}_4^-(\text{aq}) + 5 \text{H}_2\text{O}(\ell) + 3 \text{SO}_3^{2-}(\text{aq}) \rightarrow 2 \text{Cr(OH)}_3(\text{s}) + 4 \text{OH}^-(\text{aq}) + 3 \text{SO}_4^{2-}(\text{aq})$
- 20.9 (a) Oxidation:  $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2 \text{e}^-$   
Reduction:  $\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O}(\ell)$   
Overall:  $2 \text{Fe(s)} + \text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\ell) + 2 \text{Fe}^{2+}(\text{aq})$
- (b) Oxidation occurs in the anode compartment and reduction occurs in the cathode compartment.
- (c) Electrons in the external circuit flow from the Fe electrode to the positive (site of  $\text{O}_2$  reduction) electrode.  
Negative ions move in the salt bridge from the  $\text{O}_2/\text{H}_2\text{O}$  half-cell to the  $\text{Fe}/\text{Fe}^{2+}$  half-cell.
- 20.13 (a)  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.763 \text{ V}) - (+0.535 \text{ V}) = -1.298 \text{ V}$  not product-favored
- (b)  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.763 \text{ V}) - (-0.25 \text{ V}) = -0.51 \text{ V}$  not product-favored
- (c)  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (+0.337 \text{ V}) - (+1.360 \text{ V}) = -1.023 \text{ V}$  not product-favored
- (d)  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (+0.80 \text{ V}) - (+0.771 \text{ V}) = +0.03 \text{ V}$  product-favored
- 20.17 (a) Al(s)
- (b) Zn(s) and Al(s)
- (c)  $\text{Fe}^{2+}(\text{aq}) + \text{Sn(s)} \rightarrow \text{Fe(s)} + \text{Sn}^{2+}(\text{aq})$   
 $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.44 \text{ V}) - (-0.14 \text{ V}) = -0.30 \text{ V}$  reactant-favored
- (d)  $\text{Zn}^{2+}(\text{aq}) + \text{Sn(s)} \rightarrow \text{Zn(s)} + \text{Sn}^{2+}(\text{aq})$   
 $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.76 \text{ V}) - (-0.14 \text{ V}) = -0.62 \text{ V}$  reactant-favored

- 20.23 (a)  $F_2$   
 (b)  $Cl_2$  and  $F_2$



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.8277 \text{ V}) - (-1.22 \text{ V}) = +0.39 \text{ V}$$

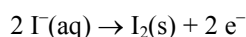
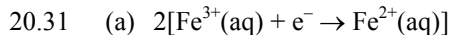
$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0257}{n} \ln \frac{[Zn(OH)_4^{2-}]P_{H_2}}{[OH^-]^2} = +0.39 \text{ V} - \frac{0.0257}{2} \ln \frac{(0.025)(1.0)}{(0.025)^2} = 0.34 \text{ V}$$



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (0.80 \text{ V}) - (-0.763 \text{ V}) = +1.56 \text{ V}$$

$$E_{\text{cell}} = 1.48 \text{ V} = 1.56 \text{ V} - \frac{0.0257}{2} \ln \frac{1.0}{[Ag^+]^2}$$

$$[Ag^+] = 0.040 \text{ M}$$

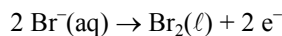
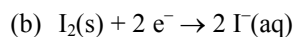


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (0.771 \text{ V}) - (0.535 \text{ V}) = +0.236 \text{ V}$$

$$\Delta G^\circ = -nFE^\circ = -(2 \text{ mol } e^-)(96,500 \text{ C/mol } e^-)(0.236 \text{ V})(1 \text{ J/1 C}\cdot\text{V})(1 \text{ kJ}/10^3 \text{ J}) = -45.5 \text{ kJ}$$

$$\ln K = \frac{nE^\circ}{0.0257 \text{ V}} = \frac{(2)(0.236 \text{ V})}{0.0257 \text{ V}} = 18.4$$

$$K = 9 \times 10^7$$

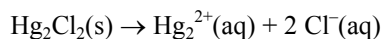
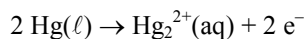
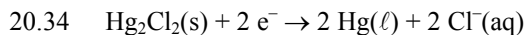


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (0.535 \text{ V}) - (1.08 \text{ V}) = -0.55 \text{ V}$$

$$\Delta G^\circ = -nFE^\circ = -(2 \text{ mol } e^-)(96,500 \text{ C/mol } e^-)(-0.55 \text{ V})(1 \text{ J/1 C}\cdot\text{V})(1 \text{ kJ}/10^3 \text{ J}) = 110 \text{ kJ}$$

$$\ln K = \frac{nE^\circ}{0.0257 \text{ V}} = \frac{(2)(-0.55 \text{ V})}{0.0257 \text{ V}} = -42$$

$$K = 4 \times 10^{-19}$$



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (0.27 \text{ V}) - (0.789 \text{ V}) = -0.52 \text{ V}$$

$$\ln K = \frac{nE^\circ}{0.0257 \text{ V}} = \frac{(2)(-0.52 \text{ V})}{0.0257 \text{ V}} = -40.$$

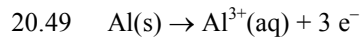
$$K = 3 \times 10^{-18}$$

- 20.41  $F^-$  is much more difficult to oxidize than water, so  $O_2$  is more likely to be formed at the anode.

20.45 Charge = current  $\times$  time = (0.150 A)(12.2 min)(60.0 s/min) = 110. C

$$\text{mol e}^- = (110. \text{ C}) \left( \frac{1 \text{ mol e}^-}{96,500 \text{ C}} \right) = 0.00114 \text{ mol e}^-$$

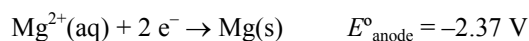
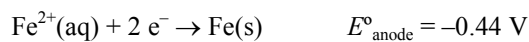
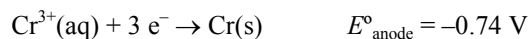
$$\text{mass of Ni} = (0.00114 \text{ mol e}^-) \left( \frac{1 \text{ mol Ni}}{2 \text{ mol e}^-} \right) \left( \frac{58.69 \text{ g}}{1 \text{ mol Ni}} \right) = 0.0334 \text{ g Ni}$$



$$(84 \text{ g Al}) \left( \frac{1 \text{ mol Al}}{27.0 \text{ g}} \right) \left( \frac{3 \text{ mol e}^-}{1 \text{ mol Al}} \right) \left( \frac{96,500 \text{ C}}{1 \text{ mol e}^-} \right) = 9.0 \times 10^5 \text{ C}$$

$$(9.0 \times 10^5 \text{ C}/1.0 \text{ A})(1 \text{ min}/60.0 \text{ s})(1 \text{ h}/60.0 \text{ min}) = 250 \text{ h}$$

20.59 (a) The combination of a negative  $E^\circ_{\text{anode}}$  half-cell with a SHE cathode (0 V) will result in a positive  $E^\circ_{\text{cell}}$



(b) Highest voltage (Mg):  $E^\circ_{\text{cell}} = (0.00 \text{ V}) - (-2.37 \text{ V}) = 2.37 \text{ V}$

Lowest voltage (Cu):  $E^\circ_{\text{cell}} = (0.337 \text{ V}) - (0.00 \text{ V}) = 0.337 \text{ V}$