CHEM 1423
Chapters 21
Homework Answers

TEXTBOOK HOMEWORK

21.5 (a) Cl^- is oxidized
(b) MnO_4^- is reduced
(c) MnO_4^- is the oxidizing agent
(d) Cl^- is the reducing agent
(e) Electrons flow from Cl^- which is losing electrons, to MnO_4^- which is gaining electrons
(f) $8\ H_2SO_4(aq) + 2\ KMnO_4(aq) + 10\ KCl(aq) \rightarrow$
$2\ MnSO_4 + 5\ Cl_2(g) + 8\ H_2O(l) + 6\ K_2SO_4(aq)$

21.9 (a) $4\,NO_3^- + 4\,H^+ + 4\,Sb \rightarrow 4NO + 2\,H_2O + Sb_4O_6$
(b) $5\,BiO_3^- + 14\,H^+ + 2\,Mn^{2+} \rightarrow 5\,Bi^{3+} + 7\,H_2O + 2\,MnO_4^-$
(c) $Pb(OH)_3^- + 2\,Fe(OH)_2 \rightarrow Pb + 2Fe(OH)_3 + OH^-$

21.27 (a)
$$\text{Red}:\ 2\,SO_3^{2-} + 3\,H_2O + 4e^- \rightarrow S_2O_3^{2-} + 6\,OH^-$$
$$Oxid : Se^{2-} \rightarrow Se + 2e^-$$
(b) $E^o_{red}[Se] = -E^o_{ox}[Se^{2-}] = -0.92V$

21.33 (a)
$$2\,Ag + Cu^{2+} \rightarrow 2\,Ag^+ + Cu$$
$$E^o_{cell} = E^o_{red}(Cu^{2+}) + E^o_{oxid}(Ag) = E^o_{red}(Cu^{2+}) - E^o_{red}(Ag^+)$$
$$= +0.34 - 0.80 = -0.46V$$
This reaction is NOT spontaneous.

(b) $Cr_2O_7^{2-} + 14\,H^+ + 3\,Cd \rightarrow 2\,Cr^{3+} + 7\,H_2O + 3Cd^{2+}$
$$E^o_{cell} = E^o_{red}(Cr_2O_7^{2-}) + E^o_{oxid}(Cd) = E^o_{red}(Cr_2O_7^{2-}) - E^o_{red}(Cd^{2+})$$
$$= +1.33 - (-0.40) = +1.73V$$
This reaction IS Spontaneous

(c)
$$Ni^{2+} + Pb \rightarrow Ni + Pb^{2+}$$
$$E^o_{cell} = E^o_{red}(Ni^{2+}) + E^o_{oxid}(Pb) = E^o_{red}(Ni^{2+}) - E^o_{red}(Pb^{2+})$$
$$= -0.25 - (-0.13) = -0.12V$$
This reaction is NOT spontaneous.

21.36 (a) N_2O is reduced, Au is oxidized.

$$N_2O + 2H^+ + 2Au \rightarrow 2N_2 + 2H_2O + 2Au^+$$

$$E^o_{cell} = E^o_{red}(N_2O) + E^o_{oxid}(Au) = E^o_{red}(N_2O) - E^o_{red}(Au^+)$$

$$= +1.77 - (+1.69) = +0.08V$$

(b) N_2O is reduced, Cr is oxidized

$$3N_2O + 6H^+ + 2Cr \rightarrow 3N_2 + 3H_2O + 2Cr^{2+}$$

$$E^o_{cell} = E^o_{red}(N_2O) + E^o_{oxid}(Cr) = E^o_{red}(N_2O) - E^o_{red}(Cr^{2+})$$

$$= +1.77 - (-0.74) = +2.51V$$

(c) Au^+ is reduced, Cr is oxidized

$$3Au^+ + Cr \rightarrow 3Au + Cr^{2+}$$

$$E^o_{cell} = E^o_{red}(Au^+) + E^o_{oxid}(Cr) = E^o_{red}(N_2O) - E^o_{red}(Cr^{2+})$$

$$= +1.69 - (-0.74) = +2.43V$$

Oxidizing Agents: $N_2O > Au^+ > Cr^{3+}$

Reducing Agents: $Cr > Au > N_2$

21.45 (a) $K = 3.0 \times 10^{35}$
(b) $K = 4 \times 10^{-31}$

21.50 $E^o_{cell} = -0.035$ V
$\Delta G^o = +6.8$ kJ

21.52 $[Pb^{2+}] = 3.5 \times 10^{-21}$ M

21.76 (a) $n(e^-) = 3.75$ mol e^-
(b) $q = 3.62 \times 10^5$ C
(c) $I = 28.7$ A

21.80 $t = 9.2 \times 10^3$ s

21.84 62.5 g Zn

SUPPLEMENTARY HOMEWORK

S1. B

S2. B

S3. D

S4. $\Delta G^o = -86.3$ kJ , $K = 1.3\times10^{15}$

S5. $E^o_{cell} = 1.22$ V

S6. $E^o_{cell} = 1.45$ V

S7. $[Zn^{2+}] = 30.3$ mg/L

S8. $K_{sp} = 3.4\times10^{-13}$

S9. $K_{sp} = 8.7\times10^{-17}$

S10. 71 g Cu

S11.

(a) Molten AlF_3(liq)

Cathode (-) Electrode: $Al^{3+} + 3\,e^- \rightarrow Al$ $E^o_{red} = -1.66$ V

Anode (+) Electrode: $2\,F^- \rightarrow F_2 + 2\,e^-$ $E^o_{ox} = -2.87$ V

$E_{cell}^o = -1.66 - 2.87 = -4.53$ V - Therefore, Voltage > +4.53 V required for electrolysis

Net Reaction: $2\,Al^{3+} + 6\,F^- \rightarrow 2\,Al + 3\,F_2$

(b) Molten NaI(liq) + Molten $ZnBr_2$(liq)

Cathode (-) Electrode: $Zn^{2+} + 2\,e^- \rightarrow Zn$ $E^o_{red} = -0.76$ V

Zn^{2+} reduced because it has a smaller negative reduction potential than Na^+

Anode (+) Electrode: $2\,I^- \rightarrow I_2 + 2\,e^-$ $E^o_{ox} = -0.54$ V

I^- oxidized because it has a smaller negative oxidation potential than Br^-

$E_{cell}^o = -0.76 - 0.54 = -1.30$ V - Therefore Voltage > +1.30 V required for electrolysis

Net Reaction: $Zn^{2+} + 2\,I^- \rightarrow Zn + I_2$

(c) An aqueous solution of $MnF_2(aq)$

Cathode (-) Electrode: $2\ H_2O + 2\ e^- \rightarrow H_2 + 2\ OH^-$ $E_{red}^{o} = -0.83$ V

H_2O is reduced because it has a smaller negative reduction potential than Mn^{2+}

Anode (+) Electrode: $2\ H_2O \rightarrow O_2 + 4\ H^+ + 4\ e^-$ $E_{ox}^{o} = -1.23$ V

H_2O is oxidized because it has a smaller negative oxidation potential than F^-

$E_{cell}^{o} = -0.83 - 1.23 = -2.06$ V - Therefore Voltage > 2.06 V required for electrolysis

Net Reaction: $2\ H_2O \rightarrow 2\ H_2 + O_2$

Note: Get cancellation of some waters when combine the H^+ and OH^- products to H_2O

(d) An aqueous solution of $FeI_2(aq)$

Cathode (-) Electrode: $Fe^{2+} + 2\ e^- \rightarrow Fe$ $E_{red}^{o} = -0.44$ V

Fe^{2+} is reduced because it has a smaller negative reduction potential than H_2O

Anode (+) Electrode: $2\ I^- \rightarrow I_2 + 2\ e^-$ $E_{ox}^{o} = -0.54$ V

I^- is oxidized because it has a smaller negative oxidation potential than H_2O

$E_{cell}^{o} = -0.44 - 0.54 = -0.98$ V - Therefore Voltage > +0.98 V required for electrolysis

Net Reaction: $Fe^{2+} + 2\ I^- \rightarrow Fe + I_2$

(e) An aqueous solution of $NaI(aq)$

Cathode (-) Electrode: $2\ H_2O + 2\ e^- \rightarrow H_2 + 2\ OH^-$ $E_{red}^{o} = -0.83$ V

H_2O is reduced because it has a smaller negative reduction potential than Na^+

Anode (+) Electrode: $2\ I^- \rightarrow I_2 + 2\ e^-$ $E_{ox}^{o} = -0.54$ V

I^- is oxidized because it has a smaller negative oxidation potential than H_2O

$E_{cell}^{o} = -0.83 - 0.54$ V $= -1.37$ V - Therefore Voltage > 1.37 V is required for electrolysis

Net Reaction: $2\ H_2O + 2\ I^- \rightarrow H_2 + 2\ OH^- + I_2$

(f) An aqueous solution of $ZnF_2(aq)$

Cathode (-) Electrode: $Zn^{2+} + 2\ e^- \rightarrow Zn$ $E_{red}^{o} = -0.76$ V

Zn is reduced because it has a smaller negative reduction potential than H_2O

Anode (+) Electrode: $2\ H_2O \rightarrow O_2 + 4\ H^+ + 4\ e^-$ $E_{ox}^{o} = -1.23$ V

H_2O is oxidized because it has a smaller negative oxidation potential than F^-

$E_{cell}^{o} = -0.76 - 1.23 = -1.99$ V - Therefore Voltage > 1.99 V is required for electrolysis

Net Reaction: $2\ Zn^{2+} + 2\ H_2O \rightarrow 2\ Zn + O_2 + 4\ H^+$