From IS 14 (12/5/06)

1) $14\ H_3O^+_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} + 6\ Cl^-_{(aq)} \rightarrow 2\ Cr^{+3}_{(aq)} + 21\ H_2O_{(l)} + 3\ Cl_{2(g)}$
2) $Cu_{(s)} + 4\ H_3O^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow Cu^{+2}_{(aq)} + 2NO_{2(g)} + 6H_2O_{(l)}$
3) $3CN^-_{(aq)} + H_2O_{(l)} + 2MnO_4^-{}_{(aq)} \rightarrow 3CNO^-_{(aq)} + 2MnO_{2(s)} + 2OH^-_{(aq)}$
4) $3NO_2^-{}_{(aq)} + 8H_3O^+_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} \rightarrow 3NO_3^-{}_{(aq)} + 2Cr^{+3}_{(aq)} + 12H_2O_{(l)}$
5) $2Cr(OH)_{3(s)} + 6ClO^-_{(aq)} \rightarrow 2\ CrO_4^{-2}{}_{(aq)} + 3Cl_{2(g)} + 2H_2O_{(l)} + 2OH^-_{(aq)}$
6) $4H_2O_{2(aq)} + 2OH^-_{(aq)} + Cl_2O_{7(aq)} \rightarrow 4O_{2(g)} + 2ClO_2^-{}_{(aq)} + 5H_2O_{(l)}$
7) $2H_3O^+_{(aq)} + As_2O_{3(s)} + 2NO_3^-{}_{(aq)} \rightarrow 2H_3AsO_{4(aq)} + N_2O_{3(aq)}$
8) $4NH_2OH_{(aq)} + Tl_2O_{3(s)} \rightarrow 2N_{2(g)} + 5H_2O_{(l)} + 2TlOH_{(s)}$
9) $2H_3O^+_{(aq)} + 3H_2S_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow 3S_{(s)} + 2NO_{(g)} + 6H_2O_{(l)}$

From Chapter 19

11a) $8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{+2} + 4\ H_2O$
b) $6e^- + 14H^+ + Cr_2O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$
c) $2\ H_2O + H_{2(g)} \rightarrow 2H_3O^+ + 2e^-$
d) $2H^+ + H_2O_2 + 2\ e^- \rightarrow 2\ H_2O$

15 a) $2\ FeO_{(s)} + 1/3\ O_{3(g)} \rightarrow Fe_2O_{3(s)}$ $O^0 \rightarrow O^{-2}$ reduced, O_3 oxidizing agent
$6\ FeO_{(s)} + O_{3(g)} \rightarrow 3Fe_2O_{3(s)}$ $Fe^{+2} \rightarrow Fe^{+3}$ oxidized, FeO reducing agent

b) $P_{4(s)} + 10\ Br_{2(l)} \rightarrow 4\ PBr_{5(l)}$ $P^0 \rightarrow P^{+5}$ oxidized, P_4 reducing agent
$Br^0 \rightarrow Br^{-1}$ reduced, Br_2 oxidizing agent

c) $H_2O_{2(aq)} + Co^{+2}_{(aq)} \rightarrow H_2O_{(l)} + Co^{+3}_{(aq)}$
$2H^+_{(aq)} + 2e^- + H_2O_{2(aq)} \rightarrow 2H_2O_{(l)}$ $O^{-1} \rightarrow O^{-2}$ reduced, H_2O_2 oxidizing agent
$2\ (Co^{+2}_{(aq)} \rightarrow Co^{+3}_{(aq)} + e^-)$ $Co^{+2} \rightarrow Co^{+3}$ oxidized, Co^{+2} reducing agent
$2H^+_{(aq)} + H_2O_{2(aq)} + 2Co^{+2}_{(aq)} \rightarrow 2H_2O_{(l)} + 2\ Co^{+3}_{(aq)}$
$2H_3O^+_{(aq)} + H_2O_{2(aq)} + 2Co^{+2}_{(aq)} \rightarrow 4H_2O_{(l)} + 2\ Co^{+3}_{(aq)}$

e) $Cl^-_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} \rightarrow Cl_{2(g)} + Cr^{+3}_{(aq)}$
$6e^- + 14H^+_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} \rightarrow 2Cr^{+3}_{(aq)} + 7H_2O_{(l)}$
$Cr^{+7} \rightarrow Cr^{+3}$ reduced $Cr_2O_7^{-2}$ oxidizing agent
$3(2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-)$
$Cl^- \rightarrow Cl^0$ oxidation, Cl^- reducing agent
$14H^+_{(aq)} + 6Cl^-_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} \rightarrow 3Cl_{2(g)} + 2Cr^{+3}_{(aq)} + 7H_2O_{(l)}$
$14H_3O^+_{(aq)} + 6Cl^-_{(aq)} + Cr_2O_7^{-2}{}_{(aq)} \rightarrow 3Cl_{2(g)} + 2Cr^{+3}_{(aq)} + 21H_2O_{(l)}$

f) $H_2CO_{(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
$H_2O_{(l)} + H_2CO_{(g)} \rightarrow CO_{2(g)} + 4H^+_{(aq)} + 4e^-$ $C^0 \rightarrow C^{+4}$ oxidation, H_2CO reducing agent
$4H^+ + 4e^- + O_{2(g)} \rightarrow 2H_2O$ $O^0 \rightarrow O^{-2}$ reduction, O_2 oxidizing agent
$H_2CO_{(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

g) $C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

$6H_2O_{(l)} + C_3H_{8(g)} \rightarrow 3CO_{2(g)} + 20\,H^+_{(aq)} + 20e^-$

$C^{-8/3} \rightarrow C^{+4}$ oxidation, C_3H_8 reducing agent

$5\,(4H^+ + 4e^- + O_{2(g)} \rightarrow 2H_2O)$ $O^0 \rightarrow O^{-2}$ reduction, O_2 oxidizing agent

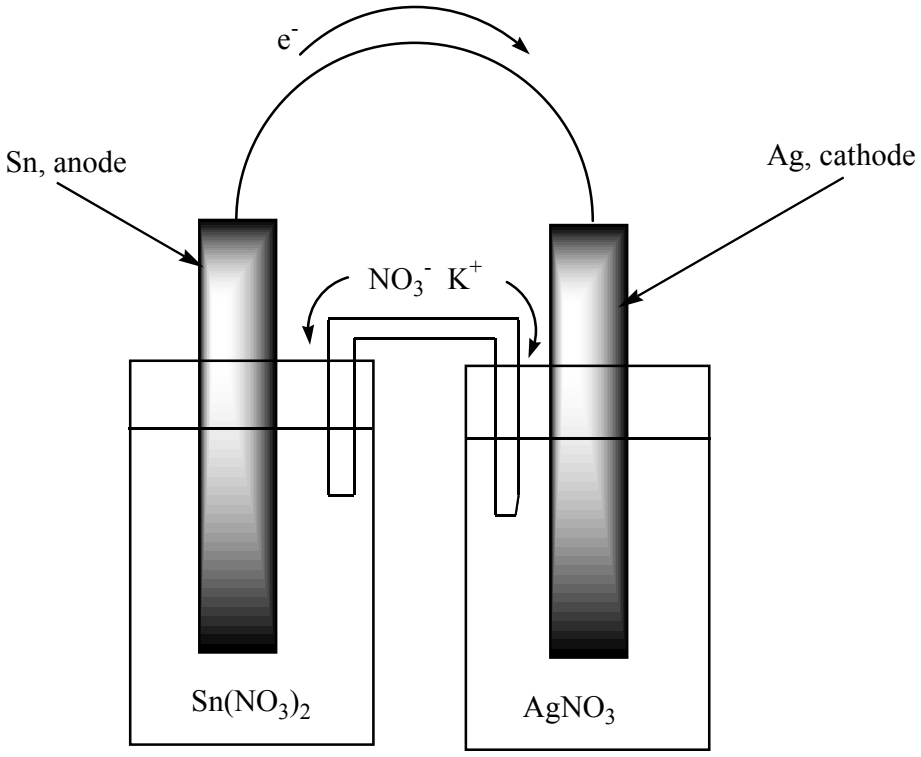
$C_3H_{8(g)} + 5\,O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

19) False. The value for an electrode potential is the same when the half reaction is multiplied by a factor. E^o for $Li^+ + e^- \rightarrow Li$ is the same as for $2Li^+ + e^- \rightarrow 2Li$.

21) a/b) $2(Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)})$ cathode

$Sn_{(s)} \rightarrow Sn^{+2}_{(aq)} + 2e^-$ anode

$2Ag^+_{(aq)} + Sn_{(s)} \rightarrow 2Ag_{(s)} + Sn^{+2}_{(aq)}$



27) $Na > Fe > H_2 > Ag > Cl_2$

31 a) Au^{+3}; b) Mn^{+2}; c) Mn; d) Au; e) yes, $E^o_{cell} = 1.48$; f) no, $E^o_{cell} = -1.48$; g) Au^{+3}, Pt^{+2}

43) $Zn_{(s)} + Cd^{+2}_{(aq)} \rightarrow Zn^{+2}_{(aq)} + Cd_{(s)}$

a) $Zn \rightarrow Zn^{+2} + 2e^-$ 0.763 V

$Cd^{+2} + 2\,e^- \rightarrow Cd_{(s)}$ -0.40 V

0.363 V

b) $E_{cell} = 0.363V - \frac{0.0592\ V}{n} \log Q$ $\quad Q = [Zn^{+2}]/[Cd^{+2}]$

$0.390 = 0.363V - \frac{0.0592\ V}{2} \log ([x]/[2.00\ M])$

$-0.912 = \log(x/2)$

$0.245\ M = x = [Zn^{+2}]$

c) $E_{cell} = 0.363V - \frac{0.0592\ V}{2} \log ([1.00]/[0.068\ M])$

$= 0.33$ V

50) $Cd_{(s)} \rightarrow Cd^{+2} + 2e^-$ 0.40V

$Ni^{+2} + 2e^- \rightarrow Ni_{(s)}$ -0.25

$Cd_{(s)} + Ni^{+2} \rightarrow Cd^{+2} + Ni_{(s)}$ 0.15 V

b) Cd is oxidized; reducing agent
Ni^{+2} is reduced; oxidizing agent

c) Cd anode; Ni cathode

d) 0.15 V

e) from Cd to Ni

f) NO_3^- toward Cd

61)

1.0 mA	1A	$= 1.0x10^{-3}$ C
	1000mA	sec

$= 1.0x10^{-3}$ C	14.5 min	60 sec	1 mol e^-	1 mol Ag	107.87 g Ag
sec		1 min	$9.65x10^4$ C	1 mol e^-	1 mol Ag

$= 9.7\ x10^{-4}$ g

65) $V^{+3} + e^- \rightarrow V^{+2}$

0.125 L	0.015 mol V^{+3}	1 mol e^-	$9.65\ x10^4$ C
	1 L	1 mol V^{+3}	1 mol e^-

$= 1.809\ x10^2$ C

0.268 C/sec * (x sec) $= 1.809\ x10^2$C

x = 675 sec = 11 min

81) Increasing ability to be reducing agents (get oxidized, lose e^-)

B < D < A< C

82)
I worked under the assumption that the metals were losing 2 electrons to achieve a M^{+2} oxidation state. Nowhere is this indicated in the problem. However it is a very common oxidation state for many metals

$B^{+2} + 2\ e^- \rightarrow B_{(s)}$ 0.68 V
$C^{+2} + 2\ e^- \rightarrow C_{(s)}$ 0.37 V
$E^{+2} + 2\ e^- \rightarrow E_{(s)}$ 0.00 V
$A^{+2} + 2\ e^- \rightarrow A_{(s)}$ -0.21 V
$D^{+2} + 2\ e^- \rightarrow D_{(s)}$ -0.56 V