Redox and Electrochem Practice Test **Name** Key

Determination the oxidation number for the underlined element.

1. $Na_2\underline{S}O_3$ 4
2. $H\underline{N}O_3$ 5
3. $H_2\underline{Se}$ -2
4. $Ni\underline{S}O_4$ 6
5. $H\underline{Cl}O$ 1
6. $\underline{Au}Cl_3$ +3

For each of the following, determine the species being oxidized, the species being reduced, the starting and ending charge.

7. $\overset{0}{C} + \overset{1}{H_2}\overset{6}{S}\overset{-2}{O_4} \rightarrow \overset{4}{C}\overset{-2}{O_2} + \overset{4}{S}\overset{-2}{O_2} + \overset{1}{H_2}\overset{-2}{O}$

   Red = S *Start Charge =* 6 *End Charge =* 4

   Oxid = C *Start Charge =* 0 *End Charge =* 4

8. $\overset{+5}{H}\overset{-2}{N}O_3 + \overset{+-}{H}I \rightarrow \overset{2}{N}\overset{-2}{O} + \overset{0}{I_2} + \overset{+}{H_2}\overset{-2}{O}$

   Red = N *Start Charge =* 5 *End Charge =* 2

   Oxid = I *Start Charge =* -1 *End Charge =* 0

9. $\overset{0}{Mg} + \overset{+-}{2HCl} \rightarrow \overset{+2}{Mg}\overset{-}{Cl_2} + \overset{0}{H_2}$

   Red = H *Start Charge =* 1 *End Charge =* 0

   Oxid = Mg *Start Charge =* 0 *End Charge =* 2

10. $2\overset{0}{Na} + 2\overset{+}{H_2}O \rightarrow 2\overset{1,2,1}{NaOH} + \overset{0}{H_2}$

    Red = H *Start Charge =* 1 *End Charge =* 0

    Oxid = Na *Start Charge =* 0 *End Charge =* 1

Balance the following redox equations in acidic solution.

11. $H_2S + NO_3^- \rightarrow NO + S$

$3(H_2S \rightarrow S + 2H^+ + 2e^-) \xrightarrow{\text{Rewrite}} 3H_2S \rightarrow 3S + \cancel{6H^+} + \cancel{6e^-}$

$2(3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O) \rightarrow \cancel{6e^-} + \cancel{8H^+} + 2NO_3^- \rightarrow 2NO + 4H_2O$

$\boxed{3H_2S + 2H^+ + 2NO_3^- \rightarrow 3S + 2NO + 4H_2O}$

12. $MnO_4^- + Cl^- \rightarrow Cl_2 + Mn^{2+}$

$2(5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O) \xrightarrow{\text{Rewrite}} \cancel{10e^-} + 16H^+ + 2MnO_4^- \rightarrow 2Mn^{2+} + 8H_2O$

$5(2Cl^- \rightarrow Cl_2 + 2e^-)$ $\quad 10Cl^- \rightarrow 5Cl_2 + \cancel{10e^-}$

$\boxed{16H^+ + 2MnO_4^- + 10Cl^- \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2}$

Balance the following redox reactions in basic solution.

13. $Br_2 \rightarrow BrO_3^- + Br^-$

$12\,OH^- + 6H_2O + Br_2 \rightarrow 2BrO_3^- + 12\,H^+ + 12\,OH^-$

$12\,H_2O$

Rewrite From Above

$12\,OH^- + \cancel{6H_2O} + Br_2 \rightarrow 2BrO_3^- + \cancel{12\,H_2O}$

$12\,OH^- + Br_2 \rightarrow 2BrO_3^- + 6H_2O + 10e^-$

$10e^- + 5Br_2 \rightarrow 10\,Br^-$

$Br_2 \rightarrow 2Br^-$

Rewrite

$5(2e^- + Br_2 \rightarrow 2Br^-)$

13. Final Answer:

$12\,OH^- + 6Br_2 \rightarrow 2BrO_3^- + 6H_2O + 10Br^-$

14. $KIO_3 + H_2SO_3 \rightarrow KI + H_2SO_4$

$6OH^- + 6H^+ + KIO_3 \rightarrow KI + 3H_2O + 6\,OH^-$

$6\,H_2O + KIO_3 \rightarrow KI + 3H_2O + 6OH^-$

$3H_2O + KIO_3 \rightarrow KI + 6OH^-$

$2\,OH^- + H_2SO_3 \rightarrow H_2SO_4 + H_2O$

$H_2O + H_2SO_3 \rightarrow H_2SO_4 + 2H^+ + 2OH^-$

$2OH^- + H_2SO_3 \rightarrow H_2SO_4 + H_2O$

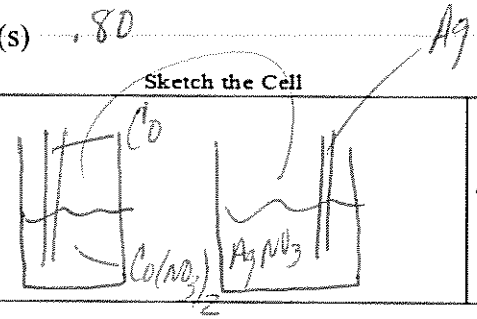
Final Answer

$2H_2O + KIO_3 + H_2SO_3 \rightarrow KI + 4OH^- + H_2SO_4$

15. Consider the following pairs of half-reactions, decide which of the two half-reactions will occur at the anode and which will occur at the cathode, draw diagrams for the cells, and calculate the standard cell potentials:

a. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ $-.28$ (flip)

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $.80$

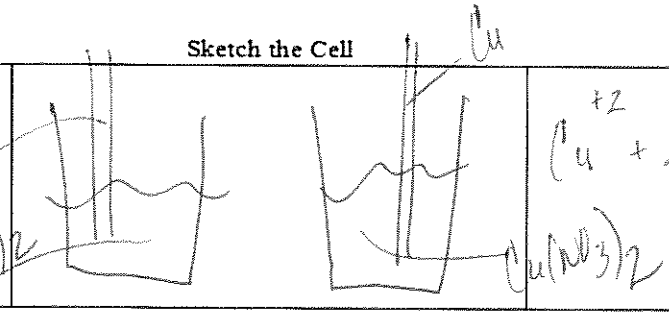
Anode Reaction	Sketch the Cell	Cathode Reaction
$Co \rightarrow Co^{2+} + 2e^-$		$Ag^+ + e^- \rightarrow Ag$

Overall Reaction & E° (volts)

$2Ag^+ + Co \rightarrow Co^{2+} + 2Ag$ $E = .80 + .28 = 1.08$ ✓

b. $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $-.25$ (flip) $E = .34 + .25 = .59$ ✓

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $.34$

Anode Reaction	Sketch the Cell	Cathode Reaction
$Ni \rightarrow Ni^{2+} + 2e^-$		$Cu^{+2} + 2e^- \rightarrow Cu$

Overall Reaction & E° (volts)

$Ni + Cu^{+2} \rightarrow Ni^{+2} + Cu$ $E_{cell} = .59$ ✓

4. $6OH^- + 6H^+ + KIO_3 \longrightarrow KI + 3H_2O + 6OH^-$

$6e^- + \cancel{6H_2O} + KIO_3 \longrightarrow KI + \cancel{3H_2O} + 6OH^-$

$6e^- + 3H_2O + KIO_3 \longrightarrow KI + 6OH^-$

<u>Second Half-Reaction</u>

$2OH^- + H_2O + H_2SO_3 \longrightarrow H_2SO_4 + 2H^+ + 2OH^- + 2e^-$

$2OH^- + \cancel{H_2O} + H_2SO_3 \longrightarrow H_2SO_4 + \cancel{2H_2O} + 2e^-$

$3(2OH^- + H_2SO_3 \longrightarrow H_2SO_4 + H_2O + 2e^-)$

$\cancel{6OH^-} + 3H_2SO_3 \longrightarrow 3H_2SO_4 + \cancel{3H_2O} + \cancel{6e^-}$

From Above: $\cancel{6e^-} + \cancel{3H_2O} + KIO_3 \longrightarrow KI + \cancel{6OH^-}$

$$\boxed{3H_2SO_3 + KIO_3 \longrightarrow 3H_2SO_4 + KI}$$