**Redox and Electrochem Practice Test Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

Determination the oxidation number for the underlined element.

1. Na2SO3 \_\_\_\_\_\_\_ 2. HNO3 \_\_\_\_\_\_\_\_\_\_\_
2. H2Se \_\_\_\_\_\_\_\_ 4. NiSO4 \_\_\_\_\_\_\_\_\_\_\_
3. HClO \_\_\_\_\_\_\_\_\_ 6. AuCl3 \_\_\_\_\_\_\_\_\_\_\_

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

1. C + H2SO4 🡪 CO2 + SO2 + H2O

*Red = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

*Oxid = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

1. HNO3 + HI 🡪 NO + I2 + H2O

*Red = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

*Oxid = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

1. Mg + 2HCl 🡪 MgCl2 + H2

*Red = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

*Oxid = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

1. 2Na + 2H2O 🡪 2NaOH + H2

*Red = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

*Oxid = \_\_\_\_\_\_\_\_ Start Charge =\_\_\_\_\_\_\_\_\_\_ End Charge = \_\_\_\_\_\_\_\_*

Balance the following redox equations in acidic solution.

1. H2S + NO3- 🡪 NO + S
2. MnO4- + Cl - 🡪 Cl2 + Mn2+

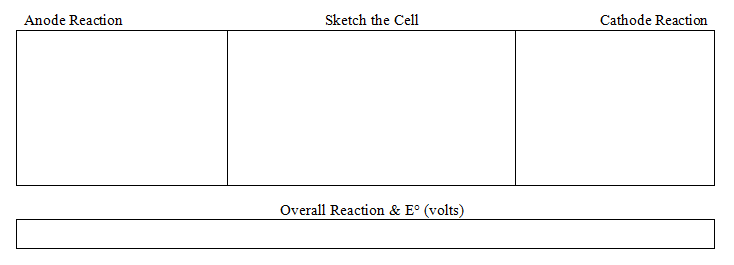
Balance the following redox reactions in basic solution.

1. Br2 🡪 BrO3- + Br -
2. KIO3 + H2SO3 🡪 KI + H2SO4

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. Co2+(aq) + 2e- → Co(s)

Ag+(aq) + e- → Ag(s)



b. Ni2+(aq) + 2e- → Ni(s)

Cu2+(aq) + 2e- → Cu(s)

