**CHE1031 module 5 HW: Electrochemistry**

**5.1: Galvanic cells**

**1.** Write the following balanced reactions using cell notation. Use platinum as an inert electrode, if needed.

(a) Mg(s) + Ni+2(aq) ⟶ Mg+2(aq) + Ni(s)

(b) 2Ag+1(aq) + Cu(s) ⟶ Cu+2(aq) + 2Ag(s)

(c) Mn(s) + Sn(NO3)2(aq) ⟶ Mn(NO3)2(aq) + Sn(s)

(d) 3Cu(NO3)(aq) + Au(NO3)3(aq) ⟶ 3Cu(NO3)2(aq) + Au(s)

**2.** Given the following cell notations, determine the species oxidized, species reduced, and the oxidizing agent and reducing agent, without writing the balanced reactions.

(a) Mg(s) │ Mg+2(aq) ║ Cu+2(aq) │ Cu(s)

(b) Ni(s) │ Ni+2(aq) ║ Ag+1(aq) │ Ag(s)

**3.** For the cell notations in the previous problem, write the corresponding balanced reactions.

**4.** Balance the following reactions and also write the reactions using cell notation. Ignore any inert electrodes, as they are never part of the half-reactions.

(a) Al(s) + Zr+4(aq) ⟶ Al+3(aq) + Zr(s)

(b) Ag+1(aq) + NO(g) ⟶ Ag(s) + NO3−1(aq) (acidic solution)

**5.** From the information provided, use cell notation to describe the following systems:

(a) In one half-cell, a solution of Pt(NO3)2 forms Pt metal, while in the other half-cell, Cu metal goes into a Cu(NO3)2 solution with all solute concentrations 1 M.

(b) The cathode consists of a gold electrode in a 0.55 M Au(NO3)3 solution and the anode is a magnesium electrode in 0.75 M Mg(NO3)2 solution.

(c) One half-cell consists of a silver electrode in a 1 M Ag(NO3) solution, and in the other half-cell, a copper electrode in 1 M Cu(NO3)2 is oxidized.

**6.** Why is a salt bridge necessary in galvanic cells?

**7.** An active (metal) electrode was found to gain mass as the oxidation-reduction reaction was allowed to proceed. Was the electrode part of the anode or cathode? Explain.

**8.** An active (metal) electrode was found to lose mass as the oxidation-reduction reaction was allowed to proceed. Was the electrode part of the anode or cathode? Explain.

**9.** The mass of three different metal electrodes, each from a different galvanic cell, were determined before and after the current generated by the oxidation-reduction reaction in each cell was allowed to flow for a few minutes. The first metal electrode, given the label A, was found to have increased in mass; the second metal electrode, given the label B, did not change in mass; and the third metal electrode, given the label C, was found to have lost mass. Make an educated guess as to which electrodes were active and which were inert electrodes, and which were anode(s) and which were the cathode(s).

**5.2: Standard reduction potentials**

**10.** For each reaction listed, determine its standard cell potential at 25 °C and whether the reaction is spontaneous at standard conditions.

(a) Mg(s) + Ni+2(aq) ⟶ Mg+2(aq) + Ni(s)

(b) 2Ag+1(aq) + Cu(s) ⟶ Cu+2(aq) + 2Ag(s)

(c) Mn(s) + Sn(NO3)2(aq) ⟶ Mn(NO3)2(aq) + Sn(s)

(d) 3Fe(NO3)2(aq) + Au(NO3)3(aq) ⟶ 3Fe(NO3)3(aq) + Au(s)

*(a) +2.14 V*

*(b) +0.46 V*

*(c) +1.04 V*

*(d) +0.73 V*

**11.** For each reaction listed, determine its standard cell potential at 25 °C and whether the reaction is spontaneous at standard conditions.

(a) Mn(s) + Ni+2(aq) ⟶ Mn+2(aq) + Ni(s)

(b) 3Cu+2(aq) + 2Al(s) ⟶ 2Al+3(aq) + 3Cu(s)

(c) Na(s) + LiNO3(aq) ⟶ NaNO3(aq) + Li(s)

(d) Ca(NO3 )2(aq) + Ba(s) ⟶ Ba(NO3 )2(aq) + Ca(s)

*(a) +0.95 V*

*(b) +2.00 V*

*(c) +0.34 V*

*(d) +0.14 V*

**12.** Determine the overall reaction and its standard cell potential at 25 °C for this reaction. Is the reaction spontaneous at standard conditions?

Cu(s) │ Cu+2(aq) ║ Au+3(aq) │ Au(s)

*+1.16 V*

**13.** Determine the overall reaction and its standard cell potential at 25 °C for the reaction involving the galvanic cell made from a half-cell consisting of a silver electrode in 1 M silver nitrate solution and a half-cell consisting of a zinc electrode in 1 M zinc nitrate. Is the reaction spontaneous at standard conditions?

 *+1.56 V*

**14.** Determine the overall reaction and its standard cell potential at 25 °C for the reaction involving the galvanic cell in which cadmium metal is oxidized to 1 M cadmium (II) ion and a half-cell consisting of an aluminum electrode in 1 M aluminum nitrate solution. Is the reaction spontaneous at standard conditions?
*+ 1.26 V*

**15.** Determine the overall reaction and its standard cell potential at 25 °C for these reactions. Is the reaction spontaneous at standard conditions? Assume the standard reduction for Br2(l) is the same as for Br2(aq).

Pt(s) │ H2(g) │ H+1(aq) ║ Br2(aq), Br−1(aq) │ Pt(s)
*+1.09 V*

**5.3: Batteries and fuel cells**

**16.** What are the desirable qualities of an electric battery?

**17.** Consider a battery made from one half-cell that consists of a copper electrode in 1 M Cu(SO4) solution and another half-cell that consists of a lead electrode in 1 M Pb(NO3)2 solution.

(a) What are the reactions at the anode, cathode, and the overall reaction?

(b) What is the standard cell potential for the battery?

(c) Most devices designed to use dry-cell batteries can operate between 1.0 and 1.5 V. Could this cell be used to make a battery that could replace a dry-cell battery? Why or why not.

(d) Suppose sulfuric acid is added to the half-cell with the lead electrode and some Pb(SO4)(s) forms. Would the cell potential increase, decrease, or remain the same?

*(b) +0.47 V*

**18.** An inventor proposes using a SHE (standard hydrogen electrode) in a new battery for smartphones that also removes toxic carbon monoxide from the air:

Anode: CO(g) + H2O(l) ⟶ CO2(g) + 2H+1(aq) + 2e− E°anode = −0.53 V

Cathode: 2H+1(aq) + 2e− ⟶ H2(g) E°cathode = 0 V

Overall: CO(g) + H2O(l) ⟶ CO2(g) + H2(g) E°cell = +0.53 V

Would this make a good battery for smartphones? Why or why not?

**19.** Why do batteries go dead, but fuel cells do not?

**5.4: Corrosion**

**20.** Which member of each pair of metals is more likely to corrode (oxidize)?

(a) Mg or Ca

(b) Au or Hg

(c) Fe or Zn

(d) Ag or Pt

(a) Ca: more negative SRP

(b) Hg

(c) Zn

(d) Ag

**21.** Consider the following metals: Ag, Au, Mg, Ni, and Zn. Which of these metals could be used as a sacrificial anode in the cathodic protection of an underground steel storage tank? Steel is mostly iron, so use −0.447 V as the standard reduction potential for steel.
Zn & Mg have more negative SRPs, -0.76 V & -2.37 V respectively

**22.** Aluminum (EAl+3 /Al ° = −2.07 V) is more easily oxidized than iron (EFe+3 /Fe ° = −0.477 V), and yet when both are exposed to the environment, untreated aluminum has very good corrosion resistance while the corrosion resistance of untreated iron is poor. Explain this observation.
Al is a self-oxidizing metal. Exposed Al surfaces self-oxidize to an aluminum oxide surface that is impenetrable to air and water and protects the unoxidized Al below it.

**23.** If a sample of iron and a sample of zinc come into contact, the zinc corrodes but the iron does not. If a sample of iron comes into contact with a sample of copper, the iron corrodes but the copper does not. Explain this phenomenon.
Zn is more easily oxidized than iron and iron is more easily oxidized than copper. SRPs: Zn = -2.37; Fe = -0.44; Cu = +0.34 V

**24.** Suppose you have three different metals, A, B, and C. When metals A and B come into contact, B corrodes and A does not corrode. When metals A and C come into contact, A corrodes and C does not corrode. Based on this information, which metal corrodes and which metal does not corrode when B and C come into contact?
Ease of corrosion = ease of oxidation = B > A > C. So, B will be corroded and C will not.

**25.** Why would a sacrificial anode made of lithium metal be a bad choice despite its E°Li+ /Li = −3.04 V, which appears to be able to protect all the other metals listed in the standard reduction potential table?
Li oxidizes so quickly that the protection it affords would soon be gone.