**CHE1031 module 5 HW: Electrochemistry KEY**

**5.1: Balancing oxidation-reduction reactions**

**1.** If a 2.5 A current is run through a circuit for 35 minutes, how many coulombs of charge moved through the circuit?  
current A = C/s 🡪 C = (A)(s) = (2.5 A)(35 min)(60 s/min) = 5250 C

**2.** For the scenario in the previous question, how many electrons moved through the circuit?

5250 C 6.25 E18 e- = 3.28 E22 electrons

1 C

**3.** Given the following pairs of balanced half-reactions, determine the balanced reaction for each pair of half-reactions in an acidic solution. Be sure you are able to identify each half-reaction as either oxidation or reduction.

(a) Ca ⟶ Ca+2 + 2e−, F2 + 2e− ⟶ 2F−1

(b) Li ⟶ Li+1 + e−, Cl2 + 2e− ⟶ 2Cl−1

(c) Fe ⟶ Fe+3 + 3e−, Br2 + 2e− ⟶ 2Br−1

(d) Ag ⟶ Ag+1 + e−, MnO4−1 + 4H+1 + 3e− ⟶ MnO2 + 2H2O

(a) Ca + F2 ⟶ CaF2 Ca ox; F2 reduced

(b) 2Li + Cl2 ⟶ 2LiCl Li ox; Cl2 reduced

(c) 2Fe + 3Br2 ⟶ 2FeBr3 Fe ox; Br2 reduced

(d) 3Ag + MnO4−1 + 4H+1 + ~~3e−~~ ⟶ 3Ag+1 + ~~3e−~~ + MnO2 + 2H2O Ag ox; MnO4-1 reduced

**4.** Balance the following in acidic solution:

(a) H2O2 + Sn+2 ⟶ H2O + Sn+4

(b) PbO2 + Hg ⟶ Hg2+2 + Pb+2

(c) Al + Cr2O7-2 ⟶ Al+3 + Cr+3

(a) H2O2 + 2H+1 + 2e-⟶ 2H2O

Sn+2 ⟶ Sn+4 + 2e-

H2O2 + 2H+1 + 2e- + Sn+2 ⟶ 2H2O + Sn+4 + 2e-

(b) PbO2 + 4H+1 + 2e- ⟶ Pb+2 + 2H2O

2Hg ⟶ Hg2+2 + 2e-

PbO2 + 4H+1 + ~~2e-~~ + 2Hg ⟶ Pb+2 + 2H2O + Hg2+2 + ~~2e-~~

(c) 2(Al ⟶ Al+3 + 3e-)

Cr2O7-2 + 14H+1 + 6e- ⟶ 2Cr+3 + 7H2O

2Al + Cr2O7-2 + 14H+1 + ~~6e-~~ ⟶ 2Cr+3 + 7H2O + 2Al+3 + ~~6e-~~

**5.** Identify the species that undergoes oxidation, the species that undergoes reduction, the oxidizing agent, and the reducing agent in each of the reactions of the previous problem.

(a) H2O2 reduced / oxidizing agent

Sn+2 oxidized / reducing agent

(b) PbO2 reduced / oxidizing agent

2Hg oxidized / reducing agent

(c) Al oxidized / reducing agent

Cr2O7-2 reduced / oxidizing agent

**6.** Why must the charge be balanced in oxidation-reduction reactions?   
To indicate where the electrons are acting and how many are involved. And because the number of electrons exchanged in oxidation and reduction are equal.

**5.2: Galvanic cells**

**7.** 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)

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

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

(c) Mn(s) | Mn(NO3)2(aq) || Sn(NO3)2(aq) | Sn(s)

(d) Pt | Cu(NO3)(aq) | Cu(NO3)2(aq) || Au(NO3)3(aq) | Au(s)

**8.** 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)

(a) Mg ox/red agent; Cu+2 red/ ox agent

(b) Ni ox/red agent; Ag+1 red/ox agent

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

(a) Mg(s) + Cu+2(aq) 🡪 Mg+2(aq) + Cu(s)

(b) Ni(s) + 2Ag+1(aq) 🡪 Ni+2(aq) + 2Ag(s)

**10.** 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)

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

Al(s) | Al+3(aq)  || Zr+4(aq) | Zr(s)

(b) 3Ag+1(aq) + NO(g) + 2H2O ⟶ 3Ag(s) + NO3−1(aq) + 4H+1 (acidic solution)

Pt(s) | NO(g), NO3−1(aq) || Ag+1 | Ag(s)

**11.** 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.

(a) Cu | Cu(NO3)2 (1 M) || Pt(NO3)2 (1 M) | Pt

(b) Mg | Mg(NO3)2 (0.75 M) || Au(NO3)3 (0.55 M) | Au

(c) Cu | Cu(NO3)2 (1 M) || Ag(NO3) (1 M) | Ag

**12.** Why is a salt bridge necessary in galvanic cells?  
The salt bridge allows ions to move between half-cells. Anions to the anode and cations move to the cathode. This allows half-cells to remain electrically neutral so that the flow of electrons from the anode to the cathode continues.

**13.** 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.  
Cathode: metal ions are reduced at the cathode and added to it.

**14.** 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.

Anode: metal atoms are oxidized, becoming soluble cations and leaving the anode.

**15.** 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).  
A is and active cathode.  
B is an inert electrode.  
C is an active anode.

**5.3: Standard reduction potentials**

**16.** 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) Mg(s) + Ni+2(aq) ⟶ Mg+2(aq) + Ni(s) = (- 0.23) - (-2.37) = **+2.14 V**

(b) 2Ag+1(aq) + Cu(s) ⟶ Cu+2(aq) + 2Ag(s) = (+0.80) – (+0.34) = **+0.46 V**

(c) Mn(s) + Sn(NO3)2(aq) ⟶ Mn(NO3)2(aq) + Sn(s) = (-0.14) – (-1.18) = **+1.04 V**

(d) 3Fe(NO3)2(aq) + Au(NO3)3(aq) ⟶ 3Fe(NO3)3(aq) + Au(s) = (+1.5) – (+0.77) = **+0.73 V**All are spontaneous as their Ecell values are positive.

**17.** 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) Mn(s) + Ni+2(aq) ⟶ Mn+2(aq) + Ni(s) = (-0.23) – (-1.18) = **+0.95 V**

(b) 3Cu+2(aq) + 2Al(s) ⟶ 2Al+3(aq) + 3Cu(s) = (+0.34) – (-1.66) = **+2.00 V**

(c) Na(s) + LiNO3(aq) ⟶ NaNO3(aq) + Li(s) = (-3.05) – (-2.71) = **-0.34 V (not)**

(d) Ca(NO3 )2(aq) + Ba(s) ⟶ Ba(NO3 )2(aq) + Ca(s) = (-2.76) – (-2.90) = **+0.14 V**

**18.** 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)

= (+1.50) – (+0.34) = +1.16 V

**19.** 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?

Zn | Zn(NO3)2 (1 M) || Ag(NO3) (1 M) | Ag = (+0.80) – (-0.76) = +1.56 V; spont

**20.** 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?  
Cd | Cd+2 (1 M) || Al(NO3)3 (1 M) | Al = (-1.66) – (-0.40) = - 1.26 V; not

**21.** 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) - (0) = +1.09 V; spontaneous

**5.5: Batteries and fuel cells**

**22.** What are the desirable qualities of an electric battery?  
A positive and high Ecell, concentrated reactants, rechargeable, minimally toxic.

**23.** 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?

(a) Cu is the cathode: Cu+2 + 2e- 🡪 Cu  
Pb is the anode: Pb 🡪 Pb+2 + 2e-

(b) E°cell = (+0.34) – (-0.13) = +0.47 V

(c) If you connected 2-3 cells in serial.

(d) Increase because the SRP for Pb(SO4) is more negative than for Pb+2

**24.** 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?

H2 is highly flammable, so no.

**25.** Why do batteries go dead, but fuel cells do not?  
Fuel cells are open systems that can accept a constant stream of new fuel.

**5.6: Corrosion**

**26.** 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

**27.** 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

**28.** 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.

**29.** 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

**30.** 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.

**31.** 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.

**5.7: Electrolysis**

**32.** Identify the reaction at the anode, reaction at the cathode, the overall reaction, and the approximate potential required for the electrolysis of the following molten salts. Assume standard states and that the standard reduction potentials in Appendix L are the same as those at each of the melting points. Assume the efficiency is 100%.

(a) CaCl2

(b) LiH

(c) AlCl3

(d) CrBr3

(a) CaCl2 anode: 2Cl-1 🡪 Cl2 + 2e-; cathode: Ca+2 + 2e- 🡪 Ca  
 overall: Ca+2 + ~~2e-~~ + 2Cl-1 🡪 Ca + Cl2 + ~~2e-~~ E°cell = (-2.76) – (1.36) = **- 4.12 V**

(b) LiH anode: 2H-1 🡪 H2 + 2e-; cathode: Li+1 + e- 🡪 Li  
 overall: 2H-1 + 2Li+1 + ~~2e-~~ 🡪 H2 + ~~2e-~~ + 2Li E°cell = (-3.05) – (-2.23) = **- 0.82 V**

(c) AlCl3 anode: 3Cl-1 🡪 1.5Cl2 + 3e-; cathode: Al+3 + 3e- 🡪 Al

overall: 3Cl-1 + Al+3 + ~~3e-~~ 🡪 1.5Cl2 + ~~3e-~~ + Al E°cell = (-1.66) – (1.36) = **- 3.02 V**

(d) CrBr3 anode: 3Br-1 🡪 1.5Br2 + 3e-; cathode: Cr+3 + 3e- 🡪 Cr

overall: 3Br-1 + Cr+3 + ~~3e~~- 🡪 1.5Br2 + ~~3e-~~ + Cr E°cell = (-0.73)–(+1.09) = **- 1.82 V**

**33.** What mass of each product is produced in each of the electrolytic cells of the previous problem if a total charge of 3.33 × 105 C passes through each cell? Assume the voltage is sufficient to perform the reduction.

(a) n = Q = 3.33 E5 C = 3.45 mole e- 1 mole Ca 40.08 g = 69.1 g Ca

F 96,485 C/mole e- 2 mole e- 1 mole Ca

= 3.45 mole e- 1 mole Cl2 70.90 g = 122 g Cl2

2 mole e- 1 mole Cl2

(b) n = Q = 3.33 E5 C = 3.45 mole e- 2 mole Li 6.94 g = 23.9 g Li

F 96,485 C/mole e- 2 mole e- 1 mole Li

= 3.45 mole e- 1 mole H2 2.02 g = 3.48 g H2

2 mole e- 1 mole H2

(c) n = Q = 3.33 E5 C = 3.45 mole e- 2 mole Al 26.98 g = 31.0 g Al

F 96,485 C/mole e- 6 mole e- 1 mole Al

= 3.45 mole e- 3 mole Cl2 70.90 g = 122 g Cl2

6 mole e- 1 mole Cl2

(d) n = Q = 3.33 E5 C = 3.45 mole e- 1 mole Cr 51.99 g = 59.6 g Cr

F 96,485 C/mole e- 3 mole e- 1 mole Cr

= 3.45 mole e- 3 mole Br 159.8 g = 551 g Br2

3 mole e- 1 mole Br2

**34.** How long would it take to reduce 1 mole of each of the following ions using the current indicated? Assume the voltage is sufficient to perform the reduction.

(a) Al+3, 1.234 A

(b) Ca+2, 22.2 A

(c) Cr+5, 37.45 A

(d) Au+3, 3.57 A

(a) For every 1 mole of Al+3 ions, 3 moles of electrons are needed.  
 Q = nF = (3 moles)(96,485 C/mole e-) = 2.89 E5 C

Q = It 🡪 t = Q = 2.86 E5 C = 2.35 E5 seconds

I 1.234 A

(b) For every 1 mole of Ca+2 ions, 2 moles of electrons are needed.  
 Q = nF = (2 moles)(96,485 C/mole e-) = 1.92 E5 C

Q = It 🡪 t = Q = 1.91 E6 C = 8.69 E3 seconds

I 22.2 A

(c) For every 1 mole of Cr+5 ions, 5 moles of electrons are needed.  
 Q = nF = (5 moles)(96,485 C/mole e-) = 4.82 E5 C

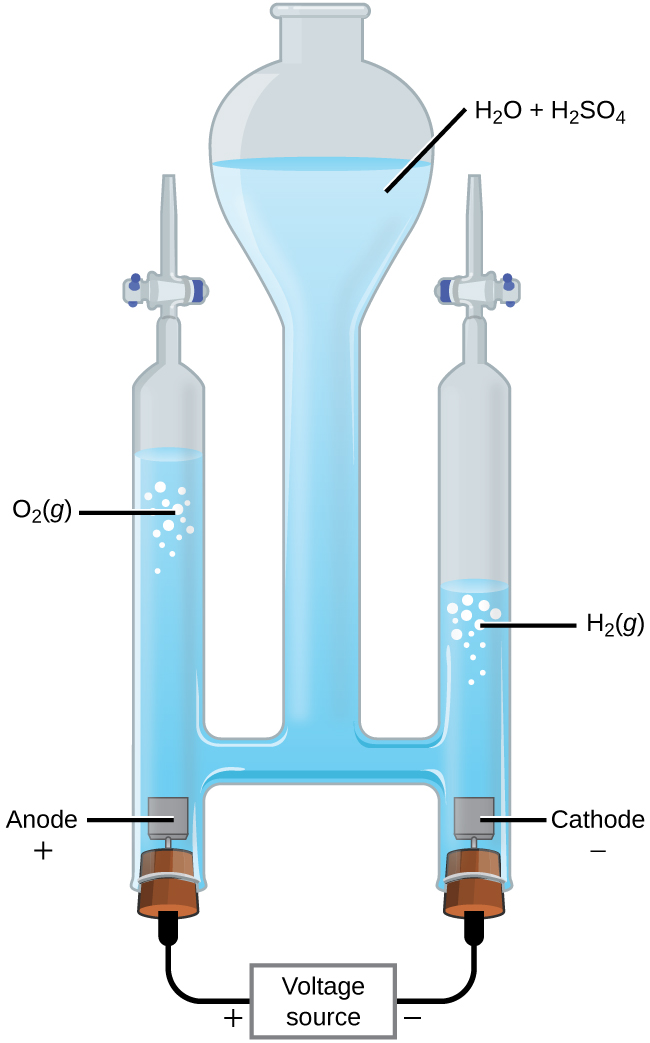
Q = It 🡪 t = Q = 4.77 E6 C = 1.29 E4 seconds

I 37.45 A

(d) For every 1 mole of Au+3 ions, 3 moles of electrons are needed.  
 Q = nF = (3 moles)(96,485 C/mole e-) = 2.89 E5 C

Q = It 🡪 t = Q = 2.86 E5 C = 8.11 E4 seconds

I 3.57 A

**35.** A current of 2.345 A passes through the cell shown here for 45 minutes. What is the volume of the hydrogen collected at room temperature if the pressure is exactly 1 atm? Assume the voltage is sufficient to perform the reduction. (Hint: Is hydrogen the only gas present above the water?)  
Q = It = (2.345 A)(45 min)(60 s/min) = 6.33 E3 C

Q = nF 🡪 n = Q = 6.33 E3 C = 6.56 E-2 mole e-

F 96,485 C/mole e-  
6.56 E-2 mole e- 1 mole H2 22.4 L = 0.735 L H2 gas.

2 mole e- 1 mole H2

**36.** An irregularly shaped metal part made from a particular alloy was galvanized with zinc using a Zn(NO3)2 solution. When a current of 2.599 A was used, it took exactly 1 hour to deposit a 0.01123-mm layer of zinc on the part. What was the total surface area of the part? The density of zinc is 7.140 g/cm3. Assume the efficiency is 100%.  
Q = It = (2.599 A)(3600 s) = 9.36 E3 C

Q = nF 🡪 n = Q = 9.36 E3 C = 9.70 E-2 mole e- 1 mole Zn = 4.85 E-2 mole Zn

F 96,485 C/mole e- 2 mole e-  
4.85 E-2 mole Zn 65.38 g 1 cm3 = 0.444 cm3

1 mole Zn 7.140 g

SA = volume/thickness

0.01123 mm 1 cm = 1.123 E-3 cm SA = 0.444 cm3 = 395 cm2

10 mm 1.123 E-3 cm