**Ch 20 Electrochemistry Name:**

*20.1 Lightning and Batteries*

*20.2 Balancing Oxidation-Reduction Equations (AP does not require you to use this technique to balance redox reactions, but these reactions are often used in questions so it is a good skill to be familiar with)*

*Review Example 20.1 & 20.2 (Note once you have the half-reactions written in step 2 then steps 3-5 can be completed on that framework and then can be added together in step 6)*

*Practice 20.1 Practice 20.2*

*Review Example 20.3 (Note that the steps are similar to acidic conditions, but the H+ions must be neutralized – this step can be done last.)*

*Practice 20.3*

*20.3 Voltaic (or Galvanic) Cells: Generating Electricity from Spontaneous Chemical Reactions*

Electrical current is the flow of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ charge either \_\_\_\_\_\_\_\_\_\_\_\_\_\_ flowing through a \_\_\_\_\_\_\_ or

\_\_\_\_\_\_\_ flowing through a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

What is a voltaic/galvanic cell? (How is it different from an electrolytic cell?

*Figure 20.1* shows that Zn(s) will react \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ly react with the Cu2+ ions is solution. Zn(s) is more reactive than Cu(s) – it oxidizes more readily. *(The reverse process Cu(s) + Zn2+ is nonspontaneous.)*

*Figure 20.2* Shows how the 2 half-reactions can be separated in order to be able to harness the transfer of electrons.

When the electrodes in the cells are connected by a \_\_\_\_\_\_\_\_ and the solutions are connected by a \_\_\_\_\_\_ bridge the

Zn anode undergoes \_\_\_\_\_tion the electrons travel through the wire to the Cu \_\_\_\_\_ode which undergoes \_\_\_\_\_\_tion.

The ions in the salt bridge neutralize the buildup of charge in each cell – notice the nitrate anion moves toward the

\_\_\_\_\_\_ode and the potassium cation move towards the \_\_\_\_\_\_\_\_ allowing current to flow. It is important the salt used

is used is a \_\_\_\_\_\_\_\_\_\_ electrolyte which contains ions that do not hydrolyze or react with ions in the cell solutions.

The rate of flow of electrons is measured in \_\_\_\_\_\_\_\_\_\_\_ (A). 1A = 1C/s

(since an electron has a charge of 1.602 x 10-19C, 1A = 1/1.602 x10−19 = 6.242 x 1018 e−/s)

The cell is “driven” by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ difference between the electrodes measured in \_\_\_\_\_\_\_s (V). 1V =1J/C

The larger the difference the stronger the tendency for electron flow which is referred to as \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

force (emf) which is *E*cell. Under standard conditions (\_\_\_\_oC and \_\_\_M reactant solutions, 1atm for reactant gases)

the cell potential is called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ cell potential (\_\_\_\_\_\_\_\_\_\_). The *Eo*cell for the zinc/copper cell in

*Figure 20.2* is +1.10V. What happens if the zinc is replaced with nickel?

In ALL electrochemical cells \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ occurs at the \_\_\_ode and reduction occurs at the \_\_\_\_\_ode. In a

voltaic/galvanic cell the anode is labeled with a (\_\_\_) sign because electrons are flowing out and the cathode is labeled

with a (\_\_\_\_) sign because electrons are flowing in. *(This is what these signs mean on a battery!)*

*CC20.1 \_\_\_\_ 20.2 \_\_\_\_*

*Mnemonic for remembering: ANOX and REDCAT (anode oxidation, reduction cathode)*

*Electrochemical Cell Notation* The notation is not required by AP, but I introduce it in our Electrochemical Cell lab so

you are familiar with it. *Figure 20.4*  shows a voltaic cell with a more complicated reaction occurring at the cathode

and the use of an \_\_\_\_\_\_\_\_\_ platinum electrode which does not take part in the reaction, but supplies a surface for the

reaction to occur and a means to conduct the electrons.

*20.4 Standard Electrode Potentials*

*Eo*cell for a half-cell must be measured vs a “standard”. What is the SHE? Describe the cell.

*Figure 20.7* shows how the *Eo*cell for the Zn/Zn2+ cell is measured. Note the calculation on pg907 to show that the

Zn/Zn2 *Eo*anode = −\_\_\_\_\_ V (this means the Zn electrode is more likely to oxidize). When the Cu/Cu2+ potential is

measured the *Eo*anode = +\_\_\_\_\_ V (this means the Cu electrode is more likely to reduce). By convention, standard

electrode potentials are written for \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ half-reactions. *Table 20.1* shows that \_\_\_(g) has the highest

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ potential (most likely to gain \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_) and is the strongest \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

agent. Lithium ions (\_\_\_+) have the lowest \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ potential because it is more likely that lithium

(solid) will undergo \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (most likely to \_\_\_\_\_\_\_\_\_\_ electrons) and is the strongest \_\_\_\_\_\_\_\_\_\_\_\_\_\_

agent. Using this chart *Eo*\_\_\_\_\_\_\_ = *Eo*cathode − *Eo*anode *(the minus sign is necessary since all potentials are reductions).*

*Eo*cell is \_\_\_\_\_\_\_\_\_\_\_\_\_\_ for \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ reactions and \_\_\_\_\_\_\_\_\_\_\_\_\_\_ for nonspontaneous reactions.

*Example 20.4* The overall reaction has been separated into the oxidation and reduction half-reactions. In the next step the *Eo*cell for the Al ANODE is written as a *reduction potential­* since the half-reaction is an oxidation you can “flip” the sign to + and then just ADD the 2 potentials together. *Try this when you complete* *Practice 20.4.*

*CC20.3 \_\_\_\_*

*Prediction the Spontaneous Direction of a REDOX Reaction* This section can be simplified by remembering that the half-cell with the *highest* (most *positive)* reduction potential will ALWAYS be the cathode.

*Practice 20.5* For “a” there are 2 ways to think about this: 1) Ni2+ has a higher reduction potential than Zn2+ so as written it will reduce by gaining electrons from Zn(s) which will oxidize or 2) Zn2+ has a lower reduction potential so Zn(s) would have a higher oxidation potential and it will act as a reducing agent for the Ni2+. Explain why “b” is nonspontaneous as written. Calculate *Eo*cell to confirm it is positive (nonspontaneous.)

*CC20.4 \_\_\_\_*

What is the general rule for determining if a metal will react with an acid? (I don’t like the use of the word “dissolve” in this section.)

What is an important exception to the rule? (We saw this during the Chemical Rxns of Copper Lab.) *CC20.5 \_\_\_\_*

*20.5 Cell Potential, Free Energy, and the Equilibrium Constant*

Spontaneous: Δ*Go is* \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, *Eo*cell is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, *K \_\_\_\_\_\_\_*

Nonspontaneous: Δ*Go is* \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, *Eo*cell is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, *K \_\_\_\_\_\_\_*

Δ*Go* can be related to *Eo*cell by the equation: Δ*Go* = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ [18.3]

What is *F*? Why is this a more convenient unit than the charge on one electron?

*Practice 20.6 CC 20.6 \_\_\_\_*

*Eo*cell can be related to *K* for a cell by: *Eo*cell = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (This equation is NOT on the AP reference sheet. It is typically easier to use Δ*Go* = *−nFEo*cell and then substitute into Δ*Go = −RTlnK*  to find *K. Try this for:*

*Practice 20.7 CC 20.7 \_\_\_\_*

*20.6 Cell Potential and Concentration*

According to Le Châtlier’s principle why will the zinc/copper cell have a potential greater than +1.10V when the concentration of Cu2+ is increased to 2M and the Zn2+ is reduced to 0.0100M? *Figure 20.11*

The discussion of pg918 shows how the equation for Δ*G* (free energy under \_\_\_\_standard conditions) can be used to

derive a way to calculate *E*cell (cell potential under nonstandard conditions).

This equation is called the \_\_\_\_\_\_\_\_\_\_\_ equation:

This equation is NOT given on the AP reference sheet but it comes in very handy to help with qualitative *Q* vs *K*

explanations. Note under standard conditions *Q* = \_\_\_ and *E*cell \_\_ *Eo*cell.

*Practice 20.8*

When *Q = 1 (standard conditions with 1M concentrations)* log 1= \_\_ and *E*cell \_\_ *Eo*cell.

When *Q < 1* (\_\_\_\_\_\_\_\_\_\_\_s > \_\_\_\_\_\_\_\_\_\_\_\_\_s) the reaction is driven to the \_\_\_\_\_\_\_\_ and *E*cell \_\_ *Eo*cell.

When *Q > 1* (\_\_\_\_\_\_\_\_\_\_\_s > \_\_\_\_\_\_\_\_\_\_\_\_\_s) the reaction is driven to the \_\_\_\_\_\_\_\_ and *E*cell \_\_ *Eo*cell.

When a redox reaction reaches \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ *Q = K* and the reaction has NO tendency to proceed in

either \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and *E*cell = \_\_\_\_\_ (your battery is DEAD). *CC 20.8 \_\_\_\_\_*

Since *E*cell depends not only on the \_\_\_\_\_\_-reactions in the cell but also on \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ we can construct

a voltaic cell in which both half-reactions are the same but which a *difference* in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ drives the

current flow. *Fig 20.12* The flow of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ has the effect of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ the concentration of the

Cu2+ in the \_\_\_\_\_\_\_\_\_ half-cell and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ the concentration in the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ half-cell.

Read *Chemistry and Medicine: Concentration Cells in Human Nerve Cells*

*20.7 Batteries: Using Chemistry to Generate Electricity*

What problem do you see occurring in the dry-cell “a” in *Figure 20.15* if the battery case (Zn) is the anode?

*(What will happen over time?)*

The voltage listed on the *alkaline* battery is \_\_\_\_\_\_\_\_V (*Eo*cell). *Eo*anode for the oxidation of the Zn is +1.245V (as written), what would be the *Eo*cathode for the MnO2?

How many “cells” are in a lead-acid battery (typical car battery)? What is the voltage of each and how must they be wired to produce 12V?

What problem has occurred with rechargeable batteries?

What is the major difference between a battery and a fuel cell? What is the major advantage of the H2-O2 fuel cell?

Read *Chemistry in Your Day: The Fuel-cell Breathalyzer*  Why do you think it would be hard for someone to “beat” a breathalyzer test?

*20.8 Electrolysis: Driving Nonspontaneous Chemical Reactions with Electricity*

How is an electrolytic cell different from a voltaic cell?

*Figure 20.20* shows the classic “electrolysis” reaction of water. *Add the 2 half-cell reactions together to produce the overall reaction for the cell. (Hint: the cathode rxn needs to be multiplied by 2 to balance the exchange of electrons.*

*Figure 20.21* shows how objects can be silver plated. Why doesn’t it matter what metal is used for the cathode?

*Figure 20.22* How can the zinc/copper voltaic cell be reversed to become an electrolytic cell?

Note in an electrolytic cell the power source drives electrons toward the cathode (where \_\_\_\_\_\_\_tion occurs) so the

cathode is now labelled as (\_\_\_) and the anode is labelled as (\_\_\_) but \_\_\_\_\_\_\_\_tion still occurs at the anode.

*Prediction the Products of Electrolysis*

When \_\_\_\_\_\_\_ electrodes are placed in a pure \_\_\_\_\_\_\_\_\_\_\_\_\_ salt the \_\_\_ion is \_\_\_\_\_\_\_\_\_\_\_and the cation is

\_\_\_\_\_\_\_\_\_\_. This is how pure Na and Cl2 can be obtained (a very DANGEROUS process).

When the molten salt is a mixture of NaCl and KCl you need to look at relative potentials (the Standard Red Pot

Table is for *solutions* but the relative order can be used). The table shows that it takes less voltage to reduce \_\_\_

than \_\_\_. *(Note +2.71 V would have to be applied.)* Similarly it takes less voltage to oxidize \_\_\_ than \_\_\_.

*Aqueous Solutions* Read this section as it will apply to a lab we will complete. What is important to note about the concentrations of H+ and OH− in the solution when you complete the electrolysis?

*Practice 20.9 (This question is part of the Electrolysis of Water Lab.) What “species” are in the solution?*

*We will list the possible half-reactions as we prepare for lab..*

*Stoichiometry of Electrolysis*

*Review Example 20.10* and complete *Practice 20.10 (the set-up is basically the reverse of the example)*

*20.9 Corrosion: Undesirable Redox Reactions*

How does the presence of acid (think acid rain) affect the reduction potential of oxygen? (making it a better oxidizing agent)

What’s the advantage of using aluminum over steel for exposed structures?

*Figure 20.27* shows that the corrosion of iron “rusting” is a complex process. The process requires the presence of

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, electrolytes and acid promote the process (think salt on roads in the winter and acid rain).

What is the most obvious way to prevent the rusting of iron?

What is a sacrificial electrode? (same idea for galvanized nails) *CC20.11 \_\_\_\_*

*Exercises* (pgs 939-943) #37c, 41a, 45(a,c), 48, 47, 53(a,b), 57, 63, 65&67a, 71, 73, 79, 87, 89, 99, 107, 108, 113, 118