**Notes - Balancing Redox Equations**

*Rules for Balancing Redox Equations (acidic)*

1. Assign oxidation numbers to all elements

2. Break the reaction into two half reactions

3. Balance everything but H and O

4. Balance H using H+ and O using H2O

5. Balance charge in each half reaction by adding e- to the more positive side

6. Make the # of e- equal in both half reactions

7. Combine the two half reactions together and cancel out everything that will

*Balancing Redox (in basic)*

1. Do all the steps as above..

2. To both sides, add an a # of OH- equal to the H+ in the balanced equation

3. Combine any OH- and H+ on the same side into water, and cancel out anything that will.

For the following examples:

(i) identify what atom is oxidized, what atom is reduced, what compound is the oxidizing agent, and what compound is the reducing agent.

(ii) balance the equation - assume all species are aqueous, and are in an acidic environment.

**EX1** I- + SO42- → I2 + H2S

**EX2 SO2 + MnO4‑ → SO42- + Mn2+**

**EX3** H2C2O4 + MoO42- → CO2 + Mo3+

**EX4** Zn + NO3-  → Zn2+ + NH4+ (BASIC SOLUTION)

Notes: Simple EMF Calculations

EMF – Electromotive Force; measures the push or pull on the electrons in a half reaction, abbreviated E°

Units for E° - E° is measured in volts

Table of Standard Reduction Potentials (Remember – reduction is a gain in electrons) OIL RIG

* All reduction potentials are measured relative to the H2 electrode
* The H2 electrode has been arbitrarily assigned the value 0.00 v
* Half reactions that are more likely to occur as reduction than the H2 half reaction have a positive reduction potential
* Half reactions that are lest likely to occur as reduction than the H2 half reaction has a negative reduction potential

 A Voltaic cell is made up of a spontaneous redox reaction. A redox reaction is spontaneous if it has a positive E°.

Standard Reduction Potential Table

* All half reactions are written as reduction
* All are measured at standard conditions (25°C and 1 atm)
* To get an oxidation potential, flip the half reaction and change the sign for E°

To calculate E°cell for a redox reaction and determine if it is spontaneous or not:

* Break the reaction into half reactions
* Look up the standard reduction potentials for each half reaction (remember about flipping the oxidation)
* Add the two E° values together
* If the E°cell is positive the reaction is spontaneous. If the E°cell is negative the reaction is not spontaneous and will not occur.

Example: Determine if Co(s) + Fe+3(aq) → Co+2(aq) + Fe+2 is spontaneous

 Co(s) → Co+2(aq) + 2e- E° = +0.277 v

 e- + Fe+3(aq) → Fe+2(aq) E° = +0.771 v

 E° = +1.048 v reaction is spontaneous

Example Determine if Ni(s) + Fe+2(aq) → Ni+2 + Fe(s)

 Ni(s) → Ni+2(aq) + 2e- E° = +0.257 v

 2e- + Fe+2(aq) → Fe (s) E° = -0.447 v

 E° = -0.183 v reaction is NOT spontaneous

Notes – Do not multiply E° by any factor in the half reactions even though you may need to multiply to balance electron loss with electron gain to balance the overall reaction.

Determine if the following reactions are spontaneous or not:

A. Co(s) + Fe+3(aq) → Co+2(aq) + Fe+2(aq)

B. H2(g) + Cl2(g) → 2H+1(aq) + 2Cl-1(aq)

C. Ni(s) + Fe+2(aq) → Ni+2(aq) + Fe(s)

D. Al(s) + Fe+2(aq) →Al+3(aq) + Fe(s)

Will a reaction occur is an iron spoon is used to stir a 1M AlCl3 solution? Show all work.

**Notes - Electrolysis**

The process of electrolysis is where an electric current is run through a solution, and by doing so, causes a non-spontaneous reaction to occur. The electric current is usually obtained from a battery, which is a voltaic cell.

What is normally electrolyzed is an ionic compound, and it may or may not be dissolved in water. If the ionic compound is pure, that is, not dissolved in water, then it must be molten. This is because a solid ionic compound will not conduct electricity - the ions are fixed in a given location, and therefore no conduction can occur. A molten solid is a liquefied substance, and in this state the ions are free to move and conduct electricity.

There are some general rules of thumb that can help you predict what the products of electricity will be.

If there is no water present and you have a pure molten ionic compound, then:

* the cation will be reduced (gain electrons/go down in charge) (Bobby, I just ran over a **RED CAT**. What am I gonna do?!)
* the anion will be oxidized (lose electrons/go up in charge) (Marsha look out! It’s **AN OX!!!!**)

 If water is present and you have a solution of an ionic compound, then you have to decide if water will react or if the ion of the salt will react. This is different from what you will read in the book – we will use some simple guidelines to determine which will happen, while the book uses tables of voltages. There are some rules of thumb that can make your decision easier though:

• **no group IA or IIA metal will be reduced in an aqueous solution - water will be reduced instead.**

**• no polyatomic will be oxidized in an aqueous solution - water will be oxidized instead.**

There are some basic differences that exist between a voltaic and electrolytic cell:

i) a voltaic cell is a spontaneous reaction, and an electrolytic cell is a non-spontaneous reaction that is forced to occur.

ii) a voltaic cell is separated into two half cells to generate electricity, where as an electrolytic cell occurs in a single container.

iii) a voltaic cell is a battery since it produces electricity, while an electrolytic cell needs a battery to happen.

1. electrons flow from the negative electrode to the positive electrode in a voltaic cell, while in an electrolytic cell the electrons flow from the positive to the negative electrode.

**EX1** What are the products of the electrolysis of molten sodium chloride? Sketch the cell.

**HALF REACTIONS FOR WATER:**

**H2O →**

**H2O →**

**EX2** What are the products of the electrolysis of aqueous sodium chloride? Sketch the cell.

***Electrolysis calculations***

If you control the current that runs through a solution, you can determine the amount of product that you will form. To do this you need to know two different conversion factors:

Coulomb (C) = a grouping of charge 96485 C = charge of 1 mole of electrons

Amp (A) = a flow of electricity 1 Amp = 1C/sec

**EX3** How many moles of electrons are required to produce 5.00 A for 2.00 hours?

**EX4** What mass of copper metal is produced if 10,000. C are passed through a solution of copper(II) sulfate solution?

**EX5** If it is necessary to deposit 1.50 g of Al on an object, how many minutes must 5.00 A flow through a solution of aluminum nitrate?

**Notes - Nernst Equation**

The voltages we have talked about have been measured for electrochemical cells at standard conditions. Standard conditions are defined to be 1 M solutions of all aqueous substances, 1 atmosphere pressure of all gaseous pressures, and all temperatures at 25oC. If a reaction is carried out at non-standard conditions, the Nernst equation must be used to determine what the voltage would be. The equation is:

You don’t have to memorize this equation, but you do have to know how to use it. The terms are as follows:

E - voltage at non-standard conditions

Eo - voltage at standard conditions, calculated from reduction potential tables

n - number of electrons transferred in the balanced equation

[products]a - concentration of the products raised to the power of their coefficient in the balanced equation.

[reactants]b - concentration of the reactants raised to the power of their coefficient in the balanced equation.

**EX1** What would be the voltage of a voltaic cell made from zinc metal, 1M Zn2+, silver metal and 1M Ag+?

*Shorthand method notation* - this shows the oxidation reaction with each species separated by a single slash, a double slash, and then the reduction reaction with each species separated by a single slash.

**EX2** What would be the shorthand notation for the cell described above in EX1?

**EX3** Diagram the cell above.

**EX4** What would be the voltage of a cell made from zinc metal, 2.00 M Zn2+, silver metal and 0.100 M Ag+?

**EX5** What would be the voltage of a voltaic cell made from solutions of 0.200 M Sn2+, 0.100 M Sn4+, indium metal and 0.500 M In3+ solution?

**EX6** What would be the short and notation of the above voltaic cell?

Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #61**

**Balance the following Redox reaction using the half-reaction method:**

1. S-2 + I2 → HI + S (acidic)

|  |
| --- |
|  |

2. Cl2O-2 + AsO2-1 → Cl-1 + AsO4-3 (basic)

|  |
| --- |
|  |

3. Mn+2 + IO4-1 → MnO4-1 + IO3-1 (acidic)

|  |
| --- |
|  |

4. Mn+2 + ClO-1 → MnO4-1 + Cl-1(basic)

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| --- |
|  |

5.. V2O5 + I-1 → V2O4 + I2(acidic)

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Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #62**

Calculate E° for each of the following reactions and determine if they are spontaneous or not. Show all work.

1. Ag + Cd+2 → Ag+1 + Cd

|  |
| --- |
| E°= |
| Spontaneous? |

2. Cr + Pt+2 → Cr+3 + Pt

|  |
| --- |
| E°= |
| Spontaneous? |

3. Na + Mn+2 → Na+1 + Mn

|  |
| --- |
| E°= |
| Spontaneous? |

4. Pb + Fe+2 → Pb+2 + Fe

|  |
| --- |
| E°= |
| Spontaneous? |

5. Hg + Cl2 → Hg+2 + 2Cl-1

|  |
| --- |
| E°= |
| Spontaneous? |

6. Ir + I2 → Ir+3 + 2I-1

|  |
| --- |
| E°= |
| Spontaneous? |

7. Mg + K+1 → Mg+2 + K

|  |
| --- |
| E°= |
| Spontaneous? |

8. Sn+2 + Pb+2 → Sn+4 + Pb

|  |
| --- |
| E°= |
| Spontaneous? |

9. Cu + Zn+2 → Cu+1 + Zn

|  |
| --- |
| E°= |
| Spontaneous? |

10. Fe+3 + Ti+2 → Fe+2 + Ti+3

|  |
| --- |
| E°= |
| Spontaneous? |

Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #63**

**Give the balanced cell reaction and determine E° for the voltaic cell based on the following half reactions.**

1. Cu+2 + 1e-1 → Cu+1 ; Au+3 + 3e-1 → Au

|  |
| --- |
| Rxn: |
| E° |

2. 2H+1 + 2e-1 → H2 ; Al+3 + 3e-1 → Al

|  |
| --- |
| Rxn: |
| E° |

3. No+3 + 3e-1 → No; Pd+4 +4e-1 → Pd

|  |
| --- |
| Rxn: |
| E° |

4. Li+1 + e-1 → Li; Ti+3 + e-1 → Ti+2

|  |
| --- |
| Rxn: |
| E° |

**Sketch the voltaic cells for each battery below. Show the direction of flow, identify anode and cathode, give the overall reaction and calculate E°.**

5. copper in 1M Cu+2/magnesium in 1M Mg+2

6. zinc in 1M Zn+2/nickel in 1M Ni+2

7. zinc in 1M Zn+2/copper in 1M Cu+2

**Write the shorthand notation for the cells in the problems above.**

8. (Problem #5)

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9. (Problem #6)

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10. (Problem #7)

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Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #64**

**Sketch the voltaic cells for each battery below. Show the direction of flow, identify anode and cathode, give the overall reaction and calculate E°.**

1. calcium in 1M Ca+2/rhodium in 1M Rh+3

2 platinum in 1M Pt+2/bismuth in 1M Bi+3

3. gallium in 1M Ga+3/sodium in 1M Na+1

**Write the shorthand notation for the cells in the problems above.**

4. (Problem #1)

|  |
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|  |

5. (Problem #2)

|  |
| --- |
|  |

6. (Problem #3)

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|  |

**Give the balanced cell reaction and determine E° for the voltaic cell based on the following half reactions.**

7. Ag+1 + e-1 → Ag; Fe+2 + 2e-1 → Fe

|  |
| --- |
| Rxn: |
| E° |

8. Cd+2 + 2e-1 → Cd; Ir+3 + 3e-1 → Ir

|  |
| --- |
| Rxn: |
| E° |

9. Hg+2 + 2e-1 → Hg; Sc+3 + 3e-1 → Sc

|  |
| --- |
| Rxn: |
| E° |

10. K+1 +e-1 → K; Pb+2 + 2e-1 → Pb

|  |
| --- |
| Rxn: |
| E° |

Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #65**

**Calculate E for the following cells (assume standard conditions unless otherwise specified)**

1. Mn⏐Mn+2⏐⏐Fe+3⏐Fe

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|  |

2. Mn⏐Mn+2 (0.5 M)⏐⏐Fe+3 (1 M)⏐Fe

|  |
| --- |
|  |

3. Zn⏐Zn+2⏐⏐Ag+1⏐Ag

|  |
| --- |
|  |

4. Zn⏐Zn+2 (0.01 M)⏐⏐Ag+1 (0.5 M)⏐Ag

|  |
| --- |
|  |

5. Fe⏐Fe+2⏐⏐Cd+2⏐Cd

|  |
| --- |
|  |

6. Fe⏐Fe+2 (0.05 M)⏐⏐Cd+2 (0.1 M)⏐Cd

|  |
| --- |
|  |

7. Pb⏐Pb+2⏐⏐Ag+1⏐Ag

|  |
| --- |
|  |

8. Pb⏐Pb+2 (0.20 M⏐⏐Ag+1 (0.01 M)⏐Ag

|  |
| --- |
|  |

9. Al⏐Al+3⏐⏐Pb+2⏐Pb

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| --- |
|  |

10. Al⏐Al+3 (1 M)⏐⏐Pb+2 (2 M)⏐Pb

|  |
| --- |
|  |

Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period\_\_\_\_\_Date\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**PreAP Chemistry Homework #66**

For each of the following ELECTROLYTIC cells, complete the following:

 a. What reactions are possible at the cathode (in other words, what can be reduced)?

 b. What reactions are possible at the anode (in other words, what can be oxidized)?

 c. If more than one reaction is possible, which reaction occurs at the cathode and which reaction occurs at the anode?

 d. Sketch the cell on the back of this page – label ALL parts as done in class.

1. Molten RbF 2. An aqueous solution of NaCl

3. An aqueous solution of KNO3 4. A solution of PbS

5. How many grams of silver will be deposited by a current of 1 ampere flowing 9650 seconds?

|  |
| --- |
|  |

6. What current must be used to plate 1.75 mol of copper on an electrode in 6.24 minutes?

|  |
| --- |
|  |

7. How many minutes would be necessary to deposit 0.375 g of Ca from a cell with a current of 3.93 amperes?

|  |
| --- |
|  |