SCH 4UI **9.1 Characterizing Oxidation and Reduction** - **Intro to Electrochemistry**

Read section 9.1 and complete the questions below.

Historical Reference, Where did the term “reduction” originate?

Using the equation below, explain the terms “**oxidation**” and “**reduction**”

$Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$ Hint: What are the charges on each element?

Define the term “**Redox” Reactions**:

Using the equation below, identify which elements are being oxidized and which are being reduced;

$\mathbf{Zn_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + ZnSO_{4(aq)}}$

Define the terms “**oxidizing agent**” and “**reducing agent**”. In the equation above, which is which agent?

Explain how you can predict if a redox reaction is spontaneous or not. **Write the steps** and include a reaction to illustrate the process. See page 588.

Complete Review Questions, **page 589, #’s 5, 6, 7, 8, 11 and 12**

Section 9.3 **Redox Rxtns Involving Molecular Compounds**,

Read pages 603-611, Complete PP26-34

Copy Table 9.3 Rules for Assigning Oxidation Numbers

Rules	Examples
1.	
2.	
3.	
4.	
5.	
6.	
7.	

Key Information:

-A redox rxtn is a reaction in which the oxidation numbers of atleast 2 atoms change

-Oxidation is an increase in oxidation number – IO

-Reduction is a decrease in oxidation number – DR **Therefore DR. IO**

Use the following as an example and identify which elements are being reduced and which are being oxidized by assigning oxidation numbers

a) $2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(l)}$ b) $3\ HNO_{2(aq)} \rightarrow HNO_{3(aq)} + 2\ NO_{(g)} + H_2O_{(l)}$

SCH 4UI 9.1 **Predicting the Spontaneity of Redox Reaction**

Table 9.2 **Relative Strengths of Oxidizing Agents and Reducing Agents**
-Complete the second half of the table and answer the questions below

Strongest Oxidizing Agent	**Weakest Reducing Agent**
Au+	
Pt^{2+}	
Ag^{+}	
Hg^{2+}	
Cu^{2+}	
Sn^{2+}	
Ni^{2+}	
Co^{2+}	
Tl^{+}	
Cd^{2+}	
Fe^{2+}	
Cr^{3+}	
Zn^{2+}	
Al^{3+}	
Mg^{2+}	
Ca^{2+}	
Ba^{2+}	
Weakest Oxidizing Agent	**Strongest Reducing Agent**

1. Which of the following reactions will proceed spontaneously?
   a) solid aluminum and aqueous copper (II) sulphate
   b) aqueous calcium nitrate and solid nickel
   c) solid chromium and aqueous silver nitrate
   d) aqueous barium sulphate and solid tin
   e) solid copper and aqueous magnesium chloride

2. List 3 solid metals that, when placed in an aqueous solution of cadmium chloride will react spontaneously.

3. If you want to use solid cobalt to demonstrate a spontaneous reaction between a solid and an ionic solution, which of the following aqueous solutions would you use? Explain why?

   a) silver nitrate

   b) zinc sulphate

4. Write the net ionic equations for the reactions that you listed in question 3.

SCH 4UI Sect 9.2 **Half Reaction for Balancing Equations** Hk: Read pages 590-602
Review Q's 1,4,6,10,13

Balancing half reactions in neutral conditions:

*For each half reaction... 1) Write correct formulas for Reactants and Products
2) Balance number of atoms
3) Add correct number of electrons to balance charges

Balancing half reactions in Acidic Solutions aka under Acidic Conditions:

*For each half reaction 1) Write unbalanced half reaction including R and P
2) Balance any atoms other than O and H
3) Balance O by adding H_2O molecules to Other Side
4) Balance H by adding H^+ to O/S
5) Balance charges by adding electrons to correct side

Example: Balance ½ reaction of $MnO_4^- \rightarrow Mn^{+2}$ in acidic conditions

Example; Balance ½ reaction for $ClO_3 \rightarrow Cl^-$ in acidic conditions

Balancing Half reactions in Basic Conditions aka Under Basic Conditions:

*For each ½ reaction

1) write unbalanced half reaction
2) Balance any atoms but O and H
3) Balance O and H by adding H_2O and H^+
4) Adjust for basic conditions by adding to both sides same number of OH^- as number of H^+
5) Simplify by combining H^+ and OH^- to make H_2O
6) Adjust reaction by cancelling number of H_2O molecules on both sides
7) Balance charges by adding electrons

Example: Balance the following under Basic Conditions:

$$S_2O_3 \rightarrow SO_3^{2-}$$

Half Reaction Method for Balancing Redox Reactions:

*When balancing redox reactions it is important to note that no electrons can be created or destroyed in a redox reaction. Ie **The number of electrons must balance and therefore cancel out when the two half reactions** are added together.

Steps:

1) Write ½ reactions
2) Balance reduction and oxidation reactions independently
3) Manipulate equations so that the number of electrons equal and therefore cancel out
4) Add the two half reactions together

Example: Balance the following in Acidic Conditions; $MnO_4^- + Ag \rightarrow Mn^{+2} + Ag^+$

SCH 4UI 10.1 con't - Cell Potential…pages 642-648 PP #11-20

*The difference of electric potential between an anode and a cathode is electric potential, E
*Electric potential is measured in volts (V) with the aid of a voltmeter. Electric potential is also known as cell voltage and/or cell potential.

Tables titled "Standard half-cell potentials" include half reactions of reduction reactions with the electric potential values. **These tables can be used to determine the potential for a cell/battery under standard conditions.** Standard conditions include;

*ions and molecules in their standard states
*aqueous ions and molecules with a standard concentration of 1 mol/L
*gases at standard pressure 101.3 kPa or 1 atm
*standard temperature 25°C or 298K

*E° - indicates the reduction potential for standard conditions

NOTE: half cell reduction potential is relative to the reduction potential of the standard hydrogen electrode which has a value of 0 volts

There are 2 methods to determining standard cell potential

Use the following cell for the 2 methods below

$$\mathbf{Al_{(s)} | Al^{3+}_{(aq)} || Br^{2+}_{(aq)} | Br_{(l)}}$$

Note: Al is the **anode** and therefore is being **oxidized**, Br is the **cathode** and is therefore being **reduced**

Method 1

$\mathbf{E°_{cell} = E°_{cathode} - E°_{anode}}$

$E°_{cell} = 1.066 – (-1.662)$

$\mathbf{= 2.728\ V}$

Method 2

$\mathbf{E°_{cell} = E°_{red} + E°_{ox}}$

*Since all values in table are for the reduct reaction, in order to obtain the oxidation reaction and potential values, the equation must be flipped and therefore the sign must switch ie

$Al^{3+} + 3e^- \rightarrow Al \quad E° = -1.662$

$Al \rightarrow Al^{3+} + 3e^- \quad E° = +1.662$

Therefore…

$E°_{cell} = 1.066 + 1.662$

$\mathbf{= 2.728\ V}$

Note: All galvanic cells have a positive cell potential
*Standard cell potential depends only on the identities of the anode and cathode. You do not need to take into consideration the amounts, however standard cell potential only applies to those under standard conditions
*The **most easily reduced** substances have high E° values and are at the top of list. These are the **best oxidizing agents.**
*Vice versa is also true, the **least easily reduced** substances are the worst oxidizing agents ie. at the bottom of the list, but they make the **best reducing agents**.

SCH 4UI **Sect 10.1 Galvanic Cells** Hk: Read pages 634-648, LC #1-6

*An application of redox reactions is converting chemical energy to mechanical energy, via the flow of electrons
*An examples is the battery… stores chemical energy

Electrochemistry – studies the process involved in converting chemical energy to electrical energy and vice versa

*1 type of battery, a galvanic cell aka, voltaic cell operates in the following manner…

***Galvanic Cell Notation**: a short hand notation used to illustrate the components of a cell

Application

$$\mathbf{Pb_{(s)} + 2\ FeCl_{3(aq)} \rightarrow 2\ FeCl_{2(aq)} + PbCl_{2(aq)}}$$

*Which element is being oxidized and which is being reduced?
*Which is the anode and which is the cathode
*Draw a galvanic cell
*Write the galvanic cell notation for the reaction

SCH 4UI **Sect 10.3 Electrolytic Cells** Hk: Read pages 660-664, LC #19-24

*An electrolytic cell converts electrical energy to chemical energy, the reverse of a galvanic cell
*In this process electrons move…

*This is a _____________________________ reaction and therefore requires…

*The process that takes place in an electrolytic cell is called ____________________

Just like in an Galvanic Cell

*Reduction occurs at the …

*Oxidation occurs at the …

However due to the fact that an external source of power is required to make the electrolytic cell operate, the electrodes have different charges, ie…

*Electrons still flow from the …

Example:

In Galvanic and Electrolytic cells, the anode eventually disappears and there is a build up on the cathode. In an electrolytic cell this build up is referred to as Electroplating…See Faradays Law

Faradays Law… Read pages 538-541, pp#21-24… OLD Textbook

… Ties electrochemistry to stoichiometry

Faraday's Law states that ….

BASICS:

Electric Current: flow of electrons through a circuit

Charge or Quantity of Electricity … is the product of

Example: Find the quantity of charge if 3.5A flows for 10 minutes

Charge on 1 mole of electrons = Charge on 1 electron X # of electrons in 1 mol

= ____

Apply info from above to a question…

*Calculate the mass of aluminum produced by the electrolysis of molten aluminum chloride if the current is 5A and runs for 2 hours.
NOTE: This reaction can only occur at the cathode… why?

Faraday's Practice Problems #21-24 … OLD Text

21. Calculate the mass of zinc plated onto the cathode of an electrolytic cell by a current of 750 mA in 3.25 hours.

22. How many minutes does it take to plate 0.925 g of silver onto the cathode of an electrolytic cell using a current of 1.55A?

23. the nickel anode in an electrolytic cell decreases in mass by 1.20g in 35.5 minutes. The oxidation half reaction converts nickel atoms to nickel (II) ions. What is the constant current?

24. The following two half reactions take place in an electrolytic cell with an iron anode and a chromium cathode.
    Oxidation: $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$
    Reduction: $Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$
    During the process, the mass of the iron anode decreases by 1.75g.
    a) Find the change in mass of the chromium cathode.
    b) Explain why you do not need to know the electric current or the time to complete part a.