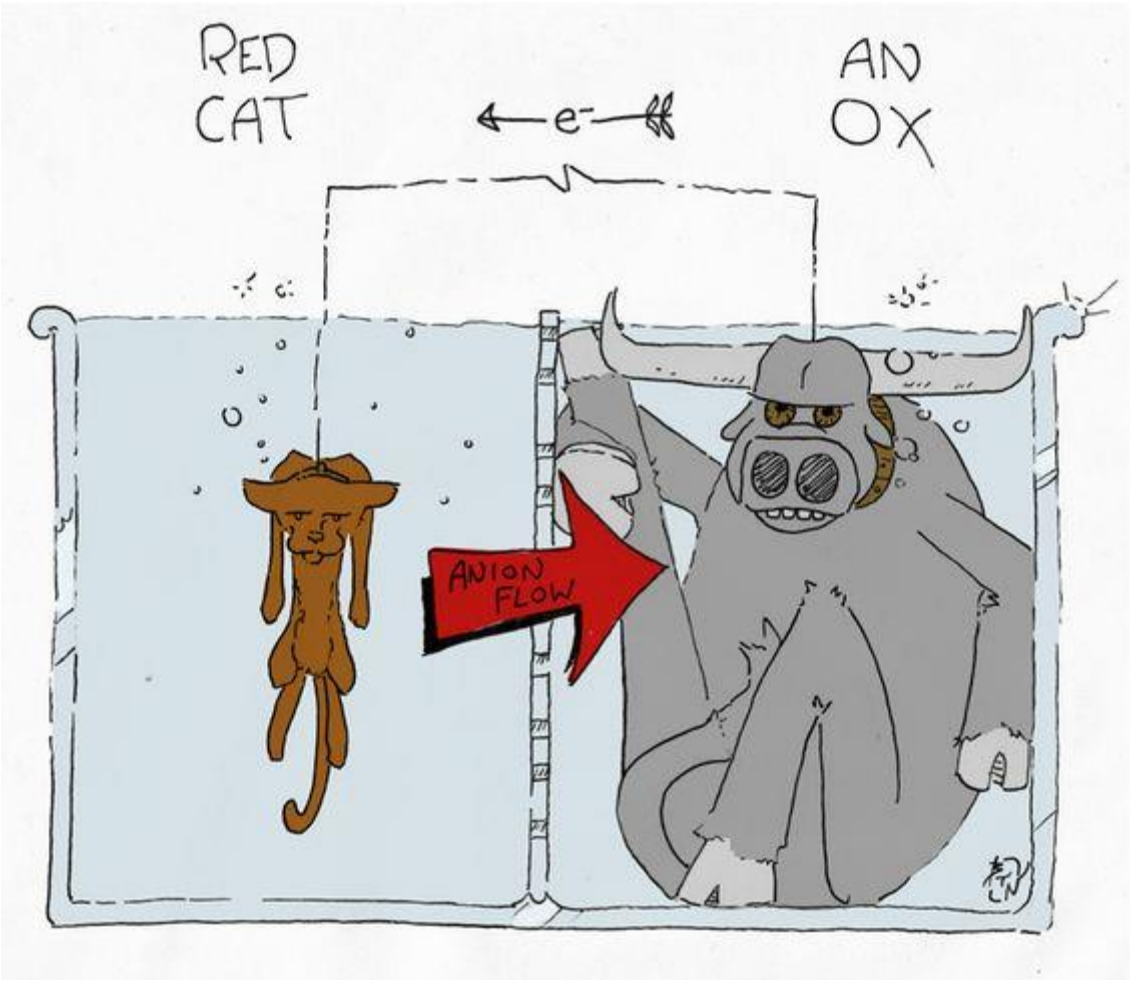
Unit 11: Redox Reactions and Electrochemistry

Regents Chemistry

Ms. Monaghan



Name: _____

Class: _____

Table of Contents

May 2016

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
1	2 F	3 A	4 B	5 C	6 D	7
8	9 E	10 F	11 A	12 B	13 C	14
15	16 D	17 E	18 F	19 A	20 B	21
22	23 C	24 D	25 E	26 F	27 A	28

Unit Learning Targets~ Unit 11: Properties of Solutions

Students will understand that by manipulating redox reactions we can create useful devices and processes in which electricity can flow due to a presence of movable charges

Topic 1: Oxidation-Reduction (Redox) Reactions

1. Assign oxidation numbers to species within a reaction
2. Identify whether or not a reaction is an oxidation- reduction (redox) reaction.
3. Identify, within a redox reaction, which species was oxidized and which was reduced.
4. Write the oxidation and reduction half-reactions for a redox reaction.
5. Balance ionic equations to show conservation of mass and charge.
6. Explain how redox reactions show conservation of charge.

Topic 2: Spontaneity of Redox Reactions

1. Use the Metal Reactivity Series (Table J) to determine whether or not a redox reaction will occur spontaneously.

Topic 3: Electrochemistry

1. Compare and contrast the two types of electrochemical cells: Voltaic and Electrolytic.
2. Identify the parts of an electrochemical cell
3. Determine the direction of electron flow, given the reaction equation.

Vocabulary

- ☐ Redox Reactions
  - Oxidation Numbers
  - Conservation of Charge
  - Oxidation (LEO)
  - Reduction (GER)
  - ½ Reactions
  - Activity of Metals or Nonmetals (Table J)
- ☐ Electrochemistry
  - Energy Conversion
    - Chemical to Electrical
    - Electrical to Chemical
- ☐ Electrochemical Cells
  - Voltaic Cells (Batteries)
    - Anode (AN – OX)
    - Cathode (RED – CAT)
    - Salt bridge
    - Electrolytes
    - Spontaneous
  - Electrolytic Cells
    - Electrolysis
    - Refining Metals
    - Electroplating
    - Recharging Batteries

BIG PICTURE!!!

How does Electricity Flow?

→ Movable Charges from an ELECTRICAL SOURCE (BATTERY)

→ Movable Charges from a CHEMICAL SOURCE (SOLUTION)

The Fruit and Vegetable Clock- HOW DOES IT WORK!?

First Viewing:

Observations:

Proposed Explanation: How does this potato clock work?	**Lingering Questions:** What questions/ wonderings do you have, that if answered could improve your understanding and explanation. *Give me at least 2!!*

Redox Reaction Reminders:

Complete the following questions on your own with no notes or help from friends.
After each part, check your answers. If you score lower than an 80% (if you miss more than 1 question) on any part, see Ms. Monaghan for a mini-lesson to rehab that topic.

Part 1-Assigning Oxidation Numbers

1. What is the formula of titanium (II) oxide?
   a. TiO b. Ti_2O c. TiO_2 d. Ti_2O_3

2. What is the oxidation state of chlorine in NaCl?
   a. -1 b. +1 c. +3 d. +5

3. What is the oxidation state of nitrogen in $NaNO_2$?
   a. +1 b. +3 c. +2 d. +4

4. What is the oxidation number of chromium in the chromate ion, CrO_4^{-2}?
   a. +6 b. +3 c. +2 d. +8

5. Given the following reaction: $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$ The oxidation number of Zn changes from:
   a. 0 to +2 b. +2 to 0 c. 0 to –2 d. –2 to 0

Part 2-Identifying Redox Reactions

1. All chemical reactions have a conservation of
   a. mass, only
   b. charge and energy, only
   c. mass and charge, only
   d. mass, charge, and energy

2. Which equation shows conservation of both mass and charge?
   a. $Cl_2 + Br^- \longrightarrow Cl^- + Br_2$
   b. $Zn + Cr^{3+} \longrightarrow Zn^{2+} + Cr$
   c. $Cu + 2\ Ag^+ \longrightarrow Cu^{2+} + Ag$
   d. $Ni + Pb^{2+} \longrightarrow Ni^{2+} + Pb$

3. Given the balanced ionic equation: $2Al(s) + 3Cu^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Cu(s)$
   Compared to the total charge of the reactants, the total charge of the products is
   a. Less b. Greater c. The same

4. During which process does an atom gain one or more electrons?
   a. Transmutation
   b. Oxidation
   c. Reduction
   d. Neutralization

5. Which reaction is an example of an oxidation reduction reaction?
   a. $AgNO_3 + KI \longrightarrow AgI + KNO_3$
   b. $Cu + 2\ AgNO_3 \longrightarrow Cu(NO_3)_2 + 2\ Ag$
   c. $2\ KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2\ H_2O$
   d. $Ba(OH)_2 + 2\ HCl \longrightarrow BaCl_2 + 2\ H_2O$

Part 3- Identifying Species Oxidized and Reduced

1. When a neutral atom undergoes oxidation, the atom's oxidation state
   a. decreases as it gains electrons
   b. increases as it gains electrons
   c. decreases as it loses electrons
   d. increases as it loses electrons

2. Which change in oxidation number indicates oxidation?
   a. –1 to +2 b. –1 to –2 c. +2 to –3 d. +3 to +2

3. Which changes occur when Pt^{2+} is reduced?
   a. The Pt^{2+} gains electrons and its oxidation number increases.
   b. The Pt^{2+} gains electrons and its oxidation number decreases.
   c. The Pt^{2+} loses electrons and its oxidation number increases.
   d. The Pt^{2+} loses electrons and its oxidation number decreases.

4. Given the equation: $C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$ Which species undergoes reduction?
   a. C(s) b. C^{2+} c. H^+ d. $H_2(g)$

5. Given the equation: $3Au(s) + 2Fe^{+3}(aq) \rightarrow 3Au^{+2}(aq) + 2Fe(s)$
   a. Gold is reduced as it loses electrons
   b. Iron is reduced as it gains electrons
   c. Gold is oxidized as it loses electrons
   d. Iron is Oxidized as it gains electrons

Part 4- Putting it All Together

Base your answers to questions 1 through 4 on the information below.

In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.

Cell with iron and copper electrodes: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$

Cell with zinc and iron electrodes: $Fe^{2+}(aq) + Zn(s) \rightarrow Fe(s) + Zn^{2+}(aq)$

1. State evidence from the balanced equation for the cell with iron and copper electrodes that indicates the reaction in the cell is an oxidation-reduction reaction.

2. Identify the particles transferred between Fe^{2+} and Zn during the reaction in the cell with zinc and iron electrodes.

3. In terms of oxidation state, explain how you know zinc is oxidized.

4. In terms of electrons lost or gained, explain how you know zinc is oxidized.

NOTES

Redox Reminders

Half-Reactions

Redox Half-Reactions Practice

Name: _____

Identify the following half reactions as oxidation or reduction. THEN, complete the reaction showing electrons in the right place. The first one has been done for you.

Reduction or Oxidation?

1. Oxidation $Li \rightarrow Li^{1+} + 1\ e^{-1}$

2. _____ $P \rightarrow P^{3-}$

3. _____ $Ca^{2+} \rightarrow Ca$

4. _____ $2\ Br^{1-} \rightarrow Br_2$

5. _____ $Fe^{2+} \rightarrow Fe^{3+}$

6. _____ $Mn^{4+} \rightarrow Mn^{7+}$

7. _____ $Fe^{2+} \rightarrow Fe$

8. _____ $H_2 \rightarrow 2\ H^{1+}$

NOTES

Balancing Half-Reactions

NOTES

Balancing Net Ionic Equations

Balancing Half-Reactions and Net Ionic Equations Practice Name:_____

Part 1:

- ☐ Balance the half reactions in each pairing using the correct number of electrons.
- ☐ Identify which is the oxidation and which is the reduction half reaction.
- ☐ Combine the half reactions in order to produce the **balanced** net ionic equation.
- ☐ **Explain or show how charge is conserved**

1. Fe^{+3} + ___ e^{-1} → Fe^{+2} and Zn → Zn^{+2} + ___ e^{-1}

2. Ca → Ca^{+2} + ___ e^{-1} and Al^{+3} + ____ e^{-1} → Al

3. Sr^{+2} → Sr and Rb → Rb^{+1}

Part 2:

- ☐ Write balanced oxidation and reduction half reactions for each of the following (indicating which is which)
- ☐ Combine the half reactions in order to produce the balance the given net ionic equation.

4. ____ Fe^{+3} + ____ Mg → ____ Mg^{+2} + ____ Fe^{+2}

5. ____ Ba + ____ Al^{+3} → ____ Ba^{+2} + ____ Al

6. ____ Al + ____ Ti^{+4} → ____ Al^{+3} + ____ Ti

7. ____ H^{+1} + ____ Zn → ____ Zn^{+2} + ____ H_2

Redox and Half Reactions Regents Practice Name: _____

1. Which balanced equation represents an oxidation-reduction reaction?
   (1) $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$
   (2) $C + H_2O \rightarrow CO + H_2$
   (3) $CaCO_3 \rightarrow CaO + CO_2$
   (4) $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

2. Given the balanced ionic equation representing a reaction:
   $2Al^{3+}(aq) + 3Mg(s) \rightarrow 3Mg^{2+}(aq) + 2Al(s)$
   In this reaction, electrons are transferred from
   (1) Al to Mg^{2+}
   (2) Al^{3+} to Mg
   (3) Mg to Al^{3+}
   (4) Mg^{2+} to Al

3. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-}?
   (1) +6
   (2) +2
   (3) +3
   (4) +8

4. Given the balanced equation representing a reaction:
   $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
   The oxidation state of chlorine in this reaction changes from
   (1) –1 to +1
   (2) –1 to +5
   (3) +1 to –1
   (4) +5 to –1

5. Given the balanced equation representing a reaction:
   $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$
   During this reaction, the oxidation number of Fe changes from
   (1) +2 to 0 as electrons are transferred
   (2) +2 to 0 as protons are transferred
   (3) +3 to 0 as electrons are transferred
   (4) +3 to 0 as protons are transferred

6. Which balanced equation represents a redox reaction?
   (1) $PCl_5 \rightarrow PCl_3 + Cl_2$
   (2) $KOH + HCl \rightarrow KCl + H_2O$
   (3) $LiBr \rightarrow Li^+ + Br^-$
   (4) $Ca^{2+} + SO_4^{2-} \rightarrow CaSO_4$

7. When lithium reacts with bromine to form the compound LiBr, each lithium atom
   (1) gains one electron and becomes a negatively charged ion
   (2) gains three electrons and becomes a negatively charged ion
   (3) loses one electron and becomes a positively charged ion
   (4) loses three electrons and becomes a positively charged ion

Base your answers to questions 15 and 16 on the information below.

The unbalanced equation below represents the decomposition of potassium chlorate.

$$KClO_3(s) \rightarrow KCl(s) + O_2(g)$$

8. Balance the equation *below*, using the smallest whole-number coefficients.

_____ $KClO_3(s)$ → _____ $KCl(s)$ + _____ $O_2(g)$

9. Determine the oxidation number of chlorine in the reactant.

10. In an oxidation-reduction reaction, the number of electrons lost is
    (1) equal to the number of electrons gained
    (2) equal to the number of protons gained
    (3) less than the number of electrons gained
    (4) less than the number of protons gained

11. Given the balanced equation representing a reaction: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + 55.8\ kJ$

    In this reaction there is conservation of
    (1) mass, only
    (2) mass and charge, only
    (3) mass and energy, only
    (4) mass, charge, and energy

12. Given the unbalanced ionic equation: $3Mg + \_\_\_\_ Fe^{3+} \rightarrow 3Mg^{2+} + \_\_\_\_ Fe$

    When this equation is balanced, both Fe^{3+} and Fe have a coefficient of
    (1) 1, because a total of 6 electrons is transferred
    (2) 2, because a total of 6 electrons is transferred
    (3) 1, because a total of 3 electrons is transferred
    (4) 2, because a total of 3 electrons is transferred

13. Given the balanced equation representing a reaction: $Mg(s) + Ni^{2+}(aq) \rightarrow Mg^{2+}(aq) + Ni(s)$

    What is the total number of moles of electrons lost by Mg(s) when 2.0 moles of electrons are gained by $Ni^{2+}(aq)$?
    (1) 1.0 mol
    (2) 2.0 mol
    (3) 3.0 mol
    (4) 4.0 mol

14. Which half-reaction correctly represents reduction?
    (1) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
    (2) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
    (3) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
    (4) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

15. Which equation shows conservation of mass and charge?
    (1) $NH_4Br \rightarrow NH_3 + Br_2$
    (2) $2Mg + Fe^{3+} \rightarrow Mg^{2+} + 3Fe$
    (3) $H_2SO_4 + LiOH \rightarrow Li_2SO_4 + H_2O$
    (4) $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

16. Which half-reaction equation represents the reduction of a potassium ion?
    (1) $K^+ + e^- \rightarrow K$
    (2) $K + e^- \rightarrow K^+$
    (3) $K^+ \rightarrow K + e^-$
    (4) $K \rightarrow K^+ + e^-$

17. Given the balanced equation representing a reaction: $2Fe + 3Cu^{2+} \rightarrow 2Fe^{3+} + 3Cu$
    When the iron atoms lose six moles of electrons, how many moles of electrons are gained by the copper ions?
    (1) 12 moles
    (2) 2 moles
    (3) 3 moles
    (4) 6 moles

18. Which half-reaction equation represents the reduction of an iron(II) ion?
    (1) $Fe^{2+} \rightarrow Fe^{3+} + e^-$
    (2) $Fe^{2+} + 2e^- \rightarrow Fe$
    (3) $Fe^{3+} + e^- \rightarrow Fe^{2+}$
    (4) $Fe \rightarrow Fe^{2+} + 2e^-$

NOTES

Spontaneity of Redox Reactions: HOW DO YOU KNOW IF IT WILL HAPPEN?

Table J
Activity Series**

Most Active	Metals	Nonmetals	Most Active
↓	Li	F_2	↓
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	H_2		
	Cu		
	Ag		
Least Active	Au		Least Active

**Activity Series is based on the hydrogen standard. H_2 is *not* a metal.

Spontaneity of Reactions- Table J Practice Name: _____

1. Based on Table J, which of the following metals is most reactive?
   a. Ag b. Au c. Ca d. Cu

2. Based on Table J, which of the following metals is most likely to be oxidized?
   a. Li b. Na c. K d. Cs

3. Based on Table J, circle the reaction below that is going to be "spontaneous." Explain your choice:

   $Ca + 2\ Na^{+1} \rightarrow Ca^{+2} + 2\ Na$ OR $2\ Na + Ca^{+2} \rightarrow 2\ Na^{+1} + Ca$

4. Circle all of the pairs for which a spontaneous reaction will occur:

   a. Al, $CuCl_2$ c. Sn, Pb e. Mg^{2+}, Co g. Li, CaO
   b. Cu, HCl d. Au, Ag^{+} f. Ni^{2+}, Sn^{2+} h. H^{+}, Sn

5. Which metal is more active than Ni and *less* active than Zn?
   a. Cu b. Mg c. Cr d. Pb

6. Which metal is more active than H_2?
   a. Ag b. Cu c. Au d. Pb

Base your answers to questions 7 through 9 on the information below.
In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.

$$Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

7. State, in terms of the relative activity of elements, why this reaction is spontaneous.

8. Balance the equation *below,* using the smallest whole-number coefficients.

   _____ $Mg(s)$ + _____ $HCl(aq) \rightarrow$ _____ $H_2(g)$ + _____ $MgCl_2(aq)$

9. Write a balanced half-reaction equation for the oxidation that occurs.

10. On the non-metals side of Chart J, explain why it makes sense that F_2 is most reactive and I_2 is least. Explain using the definition of electronegativity, as well as the electronegativity values.

The Fruit and Vegetable Clock- HOW DOES IT WORK!?

Second Viewing:

Observations:

Proposed Explanation: How does this potato clock work?	**Lingering Questions:** What questions/ wonderings do you have, that if answered could improve your understanding and explanation. *Give me at least 2!!*