Oxidation-Reduction Reactions

LESSONS

RESOURCES

NAME: ______________________________

REDOX LESSON LEARNING GOALS

http://redoxanswers.weebly.com

Lesson 1: Introduction to Redox

- ☐ Relate to examples of oxidation-reduction reactions in the real-world.
- ☐ Understand oxidation and reduction as transfer of oxygen atoms.
- ☐ Understand oxidation and reduction as transfer of electrons.

Lesson 2: Assigning Oxidation Numbers

- ☐ Learn to apply the rules of assigning oxidation numbers.
- ☐ Be able to assign oxidation numbers to atoms in polyatomic ions and compounds.

Lesson 3: Balancing Redox Reactions in Acidic Solution

- ☐ Learn to apply the steps for balancing redox reactions in acidic solution using the half-reaction method.
- ☐ Learn the definition of oxidizing and reducing agents and be able to identify them in redox reactions.

Lesson 4: Balancing Redox Reactions in Basic Solution

- ☐ Learn to apply the steps for balancing redox reactions in basic solution using the half-reaction method.
- ☐ Learn the definition of oxidizing and reducing agents and be able to identify them in redox reactions.

Lesson 5: Electrochemical Cells

- ☐ Understand the basic function of electrochemical cells.
- ☐ Be able to identify the various components of electrochemical cells.
- ☐ Be able to explain the role and function of each of the components of electrochemical cells.
- ☐ Learn how to use the activity series of metals to determine how metals can be used as electrodes in electrochemical cells.
- ☐ Develop an understanding of how the components of electrochemical cells work together to generate electric current.

Lesson 6: Calculating Cell Potentials

- ☐ Learn how to read a table of standard reduction potentials.
- ☐ Learn how to calculate the theoretical voltage of electrochemical chemical cells based on the metals used in the electrodes.

Lesson 7: Electrochemical Cell Pre-Lab

- ☐ Understand the purpose and theoretical background for the lab demonstration on electrochemical cells.
- ☐ Complete the Pre-Demonstration exercise.

Redox Reactions and Electrochemistry: Learning Goals

Learning Goals Scale			
1	2	3	4
Beginning	Developing	Applying	Mastery

#	Learning Goal		1	2	3	4
1	Students are able to relate to examples of oxidation-reduction reactions in the real-world.					
2	Students are able to define oxidation and reduction in terms of transfer of oxygen.					
3	Students are able to define oxidation and reduction in terms of transfer of electrons.					
4	Students are able apply the rules for assigning oxidation numbers to atoms of elements, polyatomic ions, and compounds.					
5	Students are able balance redox reactions in acidic/neutral solution using the half-reaction method for balancing redox reactions.					
6	Students are able balance redox reactions in basic solution using the half-reaction method for balancing redox reactions.					
7	Students are able to define the terms *oxidizing agent* and *reducing agent*.					
8	Students are able to identify oxidizing/reducing agents in redox reactions.					
9	Students are able explain the basic function of electrochemical cells.					
10	Students are able to identify the various components of electrochemical cells.					
11	Students are able explain the role and function of each of the components of electrochemical cells.					
12	Students are able to use the activity series of metals to determine how metals can be used as electrodes in electrochemical cells.					
13	Students are able to explain how the components of electrochemical cells work together to generate electric current.					
14	Students are able to read and obtain values from a Standard Reduction Potential table.					
15	Students are able to calculate theoretical voltage for electrochemical cells.					
16	Students are able to apply their knowledge of electrochemical cells to enhance their understanding of how conventional batteries work.					
17	Students are able distinguish similarities and differences between electrochemical and electrolytic cells.					

Oxidation-Reduction Reactions and Electrochemistry

1. Examples of Oxidation

- Rusting (slow oxidation)
- Burning (rapid oxidation)
- Tarnishing of silver
- Production of energy from glucose (long step)

2. Defining Oxidation and reduction

a) Former Definition- Transfer of Oxygen

Term	Definition	Example
Oxidation	gain of oxygen	$2\ Fe + O_2 \rightarrow FeO$
Reduction	loss of oxygen	$FeO \rightarrow 2\ Fe + O_2$

b) Current Definition- Transfer of electrons

Term	Definition	Example
Oxidation	loss of electron(s) in a chemical reaction	$Mg \rightarrow Mg^{2+} + 2e^{-}$
Reduction	gain of electron(s) in a chemical reaction	$\frac{1}{2}\ O_2 + 2e^{-} \rightarrow O^{2-}$

3. Definition of Oxidation Number

The oxidation number of an atom in a molecule or ion is defined as the electric charge an atom has, or appears to have, as determined by some guidelines for assigning oxidation numbers.

4. Rules for Assigning Oxidation Numbers

i) **Each atom in a pure element has an oxidation number of zero.** The oxidation number of Cu in metallic copper, as well as for I_2 or S_8, is zero.

ii) **For ions consisting of a single atom, the oxidation number is equal to the charge on the ion.** Elements of groups IA-IIIA form monatomic ions with a positive charge and an oxidation number equal to the group number. Magnesium therefore forms Mg^{2+}, and its oxidation number is +2.

iii) **Fluorine is always -1 in compounds with other elements.**

iv) **Cl, Br, and I are -1 in compounds except when combined with oxygen and fluorine.** This means that Cl has an oxidation number of -1 in NaCl (in which Na is +1). In the ion ClO^{-}, however, the Cl atom has an oxidation number of +1 (and O has an oxidation number of -2)

v) **The oxidation number of H is +1 and of O is -2 in most compounds.** Some exceptions occur:

- When H forms a binary compound with a metal, the metal forms a positive ion and H becomes a hydride ion,H^-.
- Oxygen can have an oxidation number of -1 in a class of compounds called peroxides, compounds based on the O_2^{2-} ion. For example, in H_2O_2, hydrogen peroxide, H is assigned its usual oxidation number of +1, and oxygen is assigned an oxidation number of -1.

vi) **The algebraic sum of the oxidation numbers in a neutral compound must be zero; in a polyatomic ion, the sum must be equal to the charge on the ion.**

5. Defining Oxidation and Reduction According to Oxidation Numbers

Term	Definition
Oxidation	increase in oxidation number
Reduction	decrease in oxidation number

Example: $2Ag^+{}_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)}\ Cu^{2+}{}_{(aq)}$

On the reactant side: Ag^+ has an oxidation number of +1
Cu has an oxidation number of 0

On the products side: Ag has an oxidation number of 0 (Ag^+ is reduced to Ag)
Cu^{2+} has an oxidation number of +2 (Cu is oxidized to Cu^{2+})

6. Redox Reactions

If a substance is oxidized in a chemical reaction, then another substance in the same reaction must be reduced.

- Reducing Agent: the substance that brings about reduction **(the reducing agent itself is oxidized)**→ In the example under section 5, Ag^+ is reduced, and therefore, is considered to be the oxidizing agent.

- Oxidizing Agent: the substance that brings about oxidation **(the oxidizing agent is itself reduced)**→ In the example under section 5, Cu is oxidized, and therefore, is considered to be the reducing agent.

Balancing Redox Reactions Using the Half-Reaction Method

In Acidic or Neutral Solutions:

1. Assign oxidation numbers to all of the elements present in the reaction.

2. Identify the element being oxidized and the element being reduced.
   a. Element oxidized loses electrons ⇨ the oxidation number increases.
   b. Element reduced gains electrons ⇨ the oxidation number decreases.

3. Write separate oxidation and reduction half-reactions and balance them independently according to the following steps:

4. Balance for charge (# of electrons) and for elements other than oxygen and hydrogen.

5. Balance for oxygen by adding H_2O to the side of the equation that is short of oxygen. (add one H_2O for each O needed)

6. Balance for hydrogen by adding H^+ ions to the side of the equation that is short of hydrogen. (add one H^+ for each H needed)

7. Multiply each half-reaction by a whole number so that the number of electrons lost in the oxidation half equals the number of electrons gained by the reduction half.

8. Add the two half-reactions and cancel common terms on either side of the equation.

9. Check that your final balanced reaction is:
   a. Balanced for numbers of atoms
   b. Balanced for charge
   c. Includes the phases of all substances

In Basic Solutions:

Apply Steps 1–10 above with the following changes to steps 5 and 6:

5. Balance for oxygen by adding OH^- ions to the side of the equation that is short of oxygen. (add two OH^- ions for every O needed)

6. Balance for hydrogen by adding the appropriate number of H_2O molecules to the side of the equation that is short of hydrogen.

ELECTROCHEMICAL CELLS

Electrochemical Cell
- Uses a spontaneous redox reaction to generate electric current.
- Also known as voltaic cells

Two Major Components
- External Circuit
- Internal Circuit

External Circuit (flow of electrons)
- Two metal electrodes
- Positive (cathode, red)
- Negative (anode, black)
- Wire connects the anode and cathode
- Electrons flow from negative to positive electrodes
- Current flows opposite electron flow

Internal Circuit (flow of ions)
- Half-cells (separate contains for oxidation and reduction)
- Electrolytic solutions
- Salt-bridge

Oxidation and Reduction
- Oxidation is a loss of electrons and occurs at the Anode
- Reduction is a gain of electrons and occurs at the Cathode

Activity Series
- Used to predict single replacement reactions
- Used to determine whether a metal is used as the anode or cathode in an electrochemical cell
- More active = higher on the activity series (anode)
- Less active = lower on the activity series (cathode)

Solutions in the Half-Cells and Salt Bridge
- Solutions that conduct electricity
- Contains the same metal cation that corresponds to the metal of the electrode in each half-cell
- Salt Bridge solution is a strong electrolyte (usually sodium nitrate, $NaNO_3$)

Electrochemical Cell Short-hand Notation

Calculating Cell Potential (Voltage)
- Look up the reduction potentials for the two electrodes
- Identify the anode and cathode as stated or using the activity series if not stated
- Substitute into the equation: $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
- $E^\circ_{cell} > 0$; corresponds to a spontaneous redox reaction

STANDARD REDUCTION POTENTIALS OF HALF-CELLS

Ionic concentrations are at 1M in water at 25°C.

STRONG ↑ STRENGTH OF OXIDIZING AGENT ↑ WEAK

WEAK ↓ STRENGTH OF REDUCING AGENT ↓ STRONG

Oxidizing Agents		Reducing Agents	E° (Volts)
$F_{2(g)} + 2e^-$	$\rightleftarrows$	$2F^-$	+2.87
$S_2O_8{}^{2-} + 2e^-$	$\rightleftarrows$	$2SO_4{}^{2-}$	+2.01
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftarrows$	$2H_2O$	+1.78
$MnO_4{}^- + 8H^+ + 5e^-$	$\rightleftarrows$	$Mn^{2+} + 4H_2O$	+1.51
$Au^{3+} + 3e^-$	$\rightleftarrows$	$Au_{(s)}$	+1.50
$BrO_3{}^- + 6H^+ + 5e^-$	$\rightleftarrows$	$\frac{1}{2}Br_{2(\ell)} + 3H_2O$	+1.48
$ClO_4{}^- + 8H^+ + 8e^-$	$\rightleftarrows$	$Cl^- + 4H_2O$	+1.39
$Cl_{2(g)} + 2e^-$	$\rightleftarrows$	$2Cl^-$	+1.36
$Cr_2O_7{}^{2-} + 14H^+ + 6e^-$	$\rightleftarrows$	$2Cr^{3+} + 7H_2O$	+1.23
$\frac{1}{2}O_{2(g)} + 2H^+ + 2e^-$	$\rightleftarrows$	H_2O	+1.23
$MnO_{2(s)} + 4H^+ + 2e^-$	$\rightleftarrows$	$Mn^{2+} + 2H_2O$	+1.22
$IO_3{}^- + 6H^+ + 5e^-$	$\rightleftarrows$	$\frac{1}{2}I_{2(s)} + 3H_2O$	+1.20
$Br_{2(\ell)} + 2e^-$	$\rightleftarrows$	$2Br^-$	+1.09
$AuCl_4{}^- + 3e^-$	$\rightleftarrows$	$Au_{(s)} + 4Cl^-$	+1.00
$NO_3{}^- + 4H^+ + 3e^-$	$\rightleftarrows$	$NO_{(g)} + 2H_2O$	+0.96
$Hg^{2+} + 2e^-$	$\rightleftarrows$	$Hg_{(\ell)}$	+0.85
$\frac{1}{2}O_{2(g)} + 2H^+(10^{-7}\text{ M}) + 2e^-$	$\rightleftarrows$	H_2O	+0.82
$2NO_3{}^- + 4H^+ + 2e^-$	$\rightleftarrows$	$N_2O_4 + 2H_2O$	+0.80
$Ag^+ + e^-$	$\rightleftarrows$	$Ag_{(s)}$	+0.80
$\frac{1}{2}Hg_2{}^{2+} + e^-$	$\rightleftarrows$	$Hg_{(\ell)}$	+0.80
$Fe^{3+} + e^-$	$\rightleftarrows$	Fe^{2+}	+0.77
$O_{2(g)} + 2H^+ + 2e^-$	$\rightleftarrows$	H_2O_2	+0.70
$MnO_4{}^- + 2H_2O + 3e^-$	$\rightleftarrows$	$MnO_{2(s)} + 4OH^-$	+0.60
$I_{2(s)} + 2e^-$	$\rightleftarrows$	$2I^-$	+0.54
$Cu^+ + e^-$	$\rightleftarrows$	$Cu_{(s)}$	+0.52
$H_2SO_3 + 4H^+ + 4e^-$	$\rightleftarrows$	$S_{(s)} + 3H_2O$	+0.45
$Cu^{2+} + 2e^-$	$\rightleftarrows$	$Cu_{(s)}$	+0.34
$SO_4{}^{2-} + 4H^+ + 2e^-$	$\rightleftarrows$	$H_2SO_3 + H_2O$	+0.17
$Cu^{2+} + e^-$	$\rightleftarrows$	Cu^+	+0.15
$Sn^{4+} + 2e^-$	$\rightleftarrows$	Sn^{2+}	+0.15
$S_{(s)} + 2H^+ + 2e^-$	$\rightleftarrows$	$H_2S_{(g)}$	+0.14
$2H^+ + 2e^-$	$\rightleftarrows$	$H_{2(g)}$	+0.00
$Pb^{2+} + 2e^-$	$\rightleftarrows$	$Pb_{(s)}$	−0.13
$Sn^{2+} + 2e^-$	$\rightleftarrows$	$Sn_{(s)}$	−0.14
$Ni^{2+} + 2e^-$	$\rightleftarrows$	$Ni_{(s)}$	−0.26
$H_3PO_4 + 2H^+ + 2e^-$	$\rightleftarrows$	$H_3PO_3 + H_2O$	−0.28
$Co^{2+} + 2e^-$	$\rightleftarrows$	$Co_{(s)}$	−0.28
$Se_{(s)} + 2H^+ + 2e^-$	$\rightleftarrows$	H_2Se	−0.40
$Cr^{3+} + e^-$	$\rightleftarrows$	Cr^{2+}	−0.41
$2H_2O + 2e^-$	$\rightleftarrows$	$H_2 + 2OH^-(10^{-7}\text{ M})$	−0.41
$Fe^{2+} + 2e^-$	$\rightleftarrows$	$Fe_{(s)}$	−0.45
$Ag_2S_{(s)} + 2e^-$	$\rightleftarrows$	$2Ag_{(s)} + S^{2-}$	−0.69
$Cr^{3+} + 3e^-$	$\rightleftarrows$	$Cr_{(s)}$	−0.74
$Zn^{2+} + 2e^-$	$\rightleftarrows$	$Zn_{(s)}$	−0.76
$Te_{(s)} + 2H^+ + 2e^-$	$\rightleftarrows$	H_2Te	−0.79
$2H_2O + 2e^-$	$\rightleftarrows$	$H_{2(g)} + 2OH^-$	−0.83
$Mn^{2+} + 2e^-$	$\rightleftarrows$	$Mn_{(s)}$	−1.19
$Al^{3+} + 3e^-$	$\rightleftarrows$	$Al_{(s)}$	−1.66
$Mg^{2+} + 2e^-$	$\rightleftarrows$	$Mg_{(s)}$	−2.37
$Na^+ + e^-$	$\rightleftarrows$	$Na_{(s)}$	−2.71
$Ca^{2+} + 2e^-$	$\rightleftarrows$	$Ca_{(s)}$	−2.87
$Sr^{2+} + 2e^-$	$\rightleftarrows$	$Sr_{(s)}$	−2.89
$Ba^{2+} + 2e^-$	$\rightleftarrows$	$Ba_{(s)}$	−2.91
$K^+ + e^-$	$\rightleftarrows$	$K_{(s)}$	−2.93
$Rb^+ + e^-$	$\rightleftarrows$	$Rb_{(s)}$	−2.98
$Cs^+ + e^-$	$\rightleftarrows$	$Cs_{(s)}$	−3.03
$Li^+ + e^-$	$\rightleftarrows$	$Li_{(s)}$	−3.04

Activity Series Metals

Most Active ↓ Least Active

- lithium
- rubidium
- potassium
- calcium
- sodium
- magnesium
- aluminum
- manganese
- zinc
- iron
- nickel
- tin
- lead
- copper
- silver
- platinum
- gold

Practice: Assigning Oxidation Numbers

Assign oxidation numbers to each of the elements in the given compounds.

a)	$Mg(HSO_4)_2$	Mg	H	S	O
b)	$(NH_4)_2C_2O_4$	N	H	C	O
c)	$Sn(CrO_4)_2$	Sn	Cr	O	
d)	$AlPO_4$	Al	P	O	
e)	IF_7	I	F		
f)	$Pb(HPO_4)_2$	Pb	H	P	O

REDOX REVIEW

Indicate whether the following statements are TRUE or FALSE.

___________ 1. Oxidation can occur without reduction taking place.

___________ 2. In most compounds, the oxidation number of hydrogen is +1.

___________ 3. A substance has been oxidized when it gains electrons.

___________ 4. The oxidation number of a free element, such as O_2 or Fe, is zero.

___________ 5. Oxygen atoms must be present if oxidation is to occur.

___________ 6. A substance has been reduced when it gains electrons, which results in a reduction of its oxidation number.

___________ 7. In most compounds, the oxidation number of oxygen is +2.

___________ 8. A half-reaction is an equation that represents either the oxidation or reduction in a reaction.

___________ 9. An oxidizing agent is a substance in a reaction that loses electrons.

___________ 10. The oxidation numbers of Group IA (1) and IIA(2) are numerically equal to the group number.

Indicate the oxidation number for the following.

___________ 11. S in H_2SO_3

___________ 12. Cl in $HC1O_4$

___________ 13. I in IO_3^-

___________ 14. Mn in MnO_2

___________ 15. N in HNO_3

___________ 16. As in H_3AsO_4

Calculate the number of electrons gained or lost and whether the substance is oxidized or reduced in each of the following cases.

___________ 17. S in $H_2SO_4 \rightarrow SO_4^{2-}$

___________ 18. N_2 in $N_2O \rightarrow 2NO_2^-$

___________ 19. Br_2 in $Br_2 \rightarrow 2BrO_3^-$

___________ 20. Cl in $C1O_3^- \rightarrow Cl^-$

___________ 21. Sb in $SbO_2^- \rightarrow Sb$

___________ 22. Cr in $CrO_2^- \rightarrow CrO_4^{2-}$

1. Write half-reactions for each of the following unbalanced reactions. Indicate which is the oxidation and reduction half-reaction,

a. $Zn + Ag^{+} \rightarrow Zn^{2+} + Ag$

b. $Fe^{3+} + Cr^{3+} \rightarrow Fe^{2+} + Cr_2O_7^{2-}$

c. $Mn^{2+} + Cl_2 \rightarrow Cl^{-} + MnO_4^{-}$

d. $Cu^{2+} + NO \rightarrow NO_3^{-} + Cu$

2. Balance the following reactions by the redox half-reaction method,

a. $H_5IO_6 + NO \rightarrow NO_3^{-} + IO_3^{-}$

b. $H_3PO_3 + NO_3^{-} \rightarrow PO_4^{3-} + N_2O_4$

c. $Sb_2O_5 + I^{-} \rightarrow Sb^{3+} + I_2$

d. $H_6TeO_6 + Br_2 \rightarrow TeO_2 + BrO_3^{-}$

Practice: Balancing Redox Reactions

Use the half-reaction method to balance the oxidation-reduction reactions below.
Identify the oxidizing and reducing agents for each redox reaction.

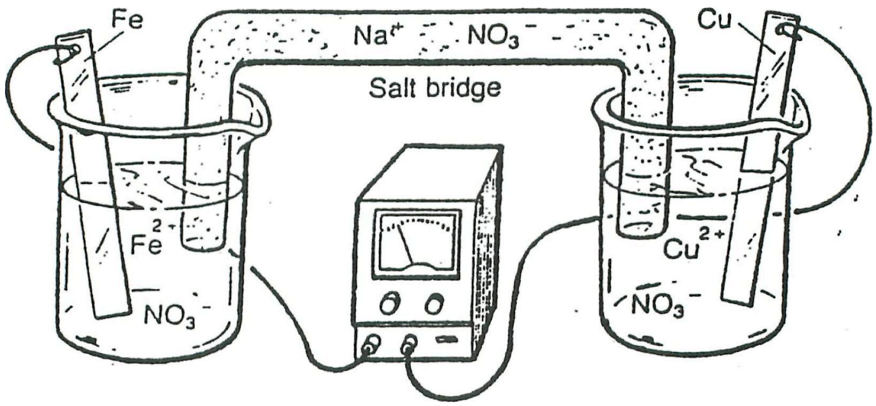
#	Unbalanced Redox Reaction	Solution Type
1	$VO_2^{+}{}_{(aq)} + Zn_{(s)} \rightarrow VO^{2+}{}_{(aq)} + Zn^{2+}{}_{(aq)}$	Acid
2	$Ag_{(s)} + NO_3^{-}{}_{(aq)} \rightarrow NO_{2\,(g)} + Ag^{+}{}_{(aq)}$	Acid
3	$MnO_4^{-}{}_{(aq)} + HSO_3^{-}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$	Acid
4	$Zn_{(s)} + NO_3^{-}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + N_2O_{(g)}$	Acid
5	$Cr_{(s)} + NO_3^{-}{}_{(aq)} \rightarrow Cr^{3+}{}_{(aq)} + NO_{(g)}$	Acid
6	$Al_{(s)} + OH^{-}{}_{(aq)} \rightarrow Al(OH)_4^{-}{}_{(aq)} + H_{2\,(g)}$	Base
7	$CrO_4^{2-}{}_{(aq)} + SO_3^{2-}{}_{(aq)} \rightarrow Cr(OH)_{3\,(s)} + SO_4^{2-}{}_{(aq)}$	Base
8	$Zn_{(s)} + Cu(OH)_{2\,(s)} \rightarrow Zn(OH)_4^{2-}{}_{(aq)} + Cu_{(s)}$	Base
9	$HS^{-}{}_{(aq)} + ClO_3^{-}{}_{(aq)} \rightarrow S_{(s)} + Cl^{-}{}_{(aq)}$	Base
10	$NiO_{2\,(s)} + Zn_{(s)} \rightarrow Ni(OH)_{2\,(s)} + Zn(OH)_{2\,(s)}$	Base
11	$MnO_4^{-} + H_2SO_3 + H^{+} \rightarrow Mn^{2+} + HSO_4^{-} + H_2O$	Acid
12	$Cr_2O_7^{2-} + H^{+} + I^{-} \rightarrow Cr^{3+} + I_2 + H_2O$	Acid
13	$As_2O_3 + H^{+} + NO_3^{-} + H_2O \rightarrow H_3AsO_4 + NO$	Acid

Practice: Balancing Redox Reactions

#	Unbalanced Redox Reaction	Solution Type
14	$I_2 + H_2SO_3 + H_2O \rightarrow I^- + HSO_4^- + H^+$	Acid
15	$H_3AsO_4 + Zn \rightarrow AsH_3 + Zn^{2+}$	Acid
16	$MnO_4^{2-} + H^+ \rightarrow MnO_4^- + MnO_2$	Acid
17	$MnO_4^- + SO_2 \rightarrow Mn^{2+} + SO_4^{2-} + H^+$	Acid
18	$HgS + Cl^- + NO_3^- \rightarrow HgCl_4^{2-} + S + NO$	Acid
19	$NH_3 + O_2 \rightarrow NO + H_2O$	Base
20	$NO_2 + OH^- \rightarrow NO_2^- + NO_3^-$	Base
21	$H_2O_2 + Co(OH)_2 \rightarrow Co_2O_3 + H_2O$	Base
22	$Cu(NH3)_4^{2+} + CN^- \rightarrow CNO^- + Cu(CN)_3^{2-} + NH_3$	Base
23	$ClO^- + Mn(OH)_2 \rightarrow MnO_2 + Cl^-$	Base

Pencast solutions for this sheet are published at http://redoxanswers.weebly.com

ELECTROCHEMICAL CELL



____________________ 1. At which electrode does oxidation occur?

____________________ 2. At which electrode does reduction occur?

____________________ 3. Which electrode is the anode?

____________________ 4. Which electrode is the cathode?

____________________ 5. Electrons will flow from which electrode?

____________________ 6. Which electrode will gain in mass?

____________________ 7. Which electrode will lose in mass?

____________________ 8. To which electrode will the anions migrate?

____________________ 9. List all of the anions in the above cell.

____________________ 10. List all of the cations in the above cell.

____________________ 11. To which electrode will the cations migrate?

____________________ 12. What is the theoretical voltage on the meter?

____________________ 13. What is the oxidizing agent?

____________________ 14. What is the reducing agent?

15. What is the purpose of the salt bridge?

____

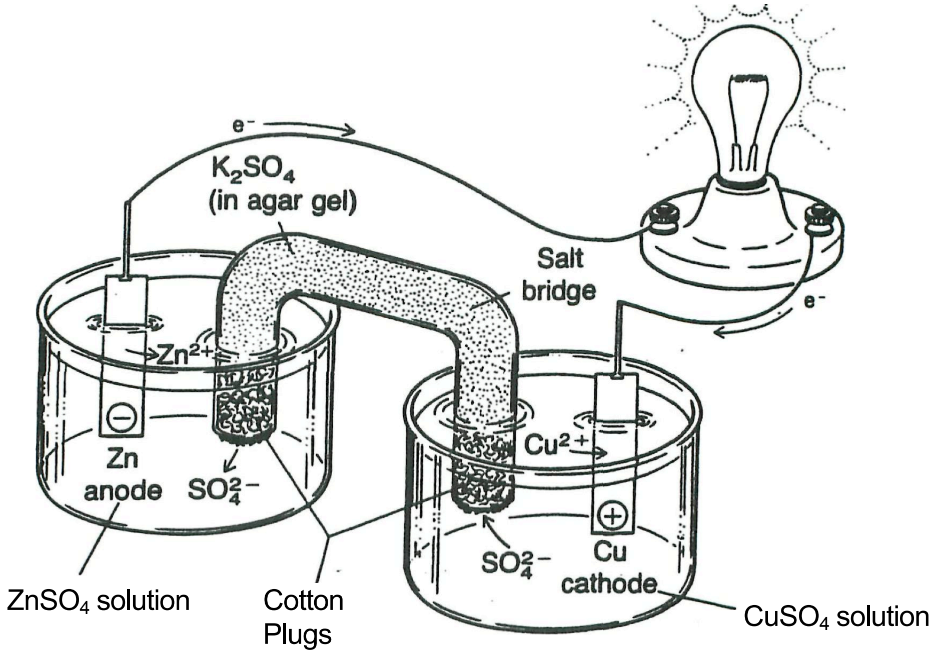
16. Write the half-reaction for the left half-cell. ____

17. Write the half-reaction for the right half-cell. ____

18. Write the overall redox reaction for the electrochemical cell.

____

Electrochemical Cell



Determine the location where each process occurs in this electrochemical cell. Write the letter for that location in the blank provided.

Locations:

a. in the wire and light bulb

b. at the cathode

c. in the salt bridge

d. at the anode

Processes:

_____ Electrons pass through the external circuit to the copper strip.

_____ Positive and negative ions move through the aqueous solution to maintain electrical neutrality.

_____ Electrons are passed to copper ions, and reduction takes place.

_____ Electrons are produced by oxidation.

Write the half-reactions and overall cell reaction for this electrochemical cell.

Oxidation half-reaction: ____

Reduction half-reaction: ____

Overall cell reaction: ____

Lab Demo: Reduction Potentials & Micro-Voltaic Cells

Purpose

The purpose of this laboratory demonstration is to establish reduction potentials of four metals relative to copper. This will be accomplished by measuring the voltage, or potential difference, between various pairs of half-cells.

Theory

A voltaic cell uses a spontaneous redox reaction to produce electrical energy. Placing a piece of metal into a solution containing the cation of the metal normally produces half-cells. In this demonstration, the standard half-cell will be composed of copper, Cu, and a solution of copper II sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$.

Label	Metal Electrode	Metal Cation	Aqueous Solution
M_1	Cu	Cu^{2+}	$CuSO_4 \cdot 5H_2O$
M_2	Zn	Zn^{2+}	$ZnSO_4 \cdot 7H_2O$
M_3	Pb	Pb^{2+}	$Pb(NO_3)_2$
M_4	Ag	Ag^{+}	$AgNO_3$
M_5	Fe	Fe^{2+}	$FeSO_4 \cdot 7H_2O$

Salt Bridge: $NaNO_3$ Red Lead: (+) positive Black Lead (–) negative

Pre-Demonstration Exercise

Write out reduction half reactions for each of the 5 metals in the table above and look up their reduction potentials.

Metal	Reduction Half Reaction	Reduction Potential (V)
Cu		
Zn		
Pb		
Ag		
Fe		

Using Activity Series to arrange the five metals from most active to least active.

Most Active	
1	
2	
3	
4	
5	
Least Active	

How is the activity level of a metal related to its ability to be oxidized?

____

____

Why is the sodium nitrate, NaNO3, important in this demonstration?

____

____

Data Table I: Measured Potential Differences Between Metals

Voltaic Cell (metals used)	Metal being Oxidized Black (−)	Metal Being Reduced Red (+)	Measured Potential (V)
M1/M2 (Cu/Zn)			
M1/M3 (Cu/Pb)			
M1/M4 (Cu/Ag)			
M1/M5 (Cu/Fe)			

Data Table II: Measured Reduction Potentials, E° (V)

Metal	Lowest (-) Reduction Potential
	Highest (+) Reduction Potential

Data Table III: Predicting Potential Difference for Various Voltaic Cells

Instructions for Predicting Potential Difference: Calculate potential for each cell below using the equation: $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$.

Voltaic Cell (metals used)	Predicted Potential	Measured Potential	Percent Difference
M2/M3 (Zn/Pb)			
M2/M4 (Zn/Ag)			
M2/M5 (Zn/Fe)			
M3/M4 (Pb/Ag)			
M3/M5 (Pb/Fe)			
M4/M5 (Ag/Fe)			

Post-Demonstration Exercise

Calculate the %difference for each of the voltaic cells described in Data Table III.

$$\%difference = \frac{|measured - predicted|}{measured} \times 100$$

REDOX ANSWER KEY

Assigning Oxidation Numbers

a)	$Mg(HSO_4)_2$	Mg	H	S	O
		+2	+1	+6	−2
b)	$(NH_4)_2C_2O_4$	N	H	C	O
		−3	+1	+3	−2
c)	$Sn(CrO_4)_2$	Sn	Cr	O	
		+4	+6	−2	
d)	$AlPO_4$	Al	P	O	
		+3	+5	−2	
e)	IF_7	I	F		
		+7	−1		
f)	$Pb(HPO_4)_2$	Pb	H	P	O
		+4	+1	+5	−2

Redox Review

1	F	12	+7
2	T	13	+5
3	F	14	+4
4	T	15	+5
5	F	16	+5
6	T	17	not redox
7	F	18	$4e^-$ lost; oxidation
8	T	19	$10e^-$ lost; oxidation
9	F	20	$6e^-$ gained; reduction
10	T	21	$3e^-$ gained; reduction
11	+4	22	$3e^-$ lost; oxidation

1. Write half-reactions for each of the following unbalanced reactions. Indicate which is the oxidation and reduction half-reaction.

   a) Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$ Reduction: $Ag^+ + 1e^- \rightarrow Ag$

   b) Oxidation: $2Cr \rightarrow Cr_2O_7^{2-} + 6e^-$ Reduction: $Fe^{3+} + 1e^- \rightarrow Fe^{2+}$

   c) Oxidation: $Mn^{2+} \rightarrow MnO_4^- + 5e^-$ Reduction: $Cl_2 + 2e^- \rightarrow 2Cl^-$

   d) Oxidation: $NO \rightarrow NO_3^- + 3e^-$ Reduction: $Cu^{2+} + 2e^- \rightarrow Cu$

2. Balance the following reactions by the redox half-reaction method.

   a) $3\ H_5IO_6 + 2\ NO \rightarrow 2NO_3^- + 3\ IO_3^- + 5\ H_2O + 5\ H^+$

   b) $H_3PO_3 + 2\ NO_3^- \rightarrow PO_4^{3-} + N_2O_4 + H_2O + H^+$

   c) $Sb_2O_5 + 4\ I^- + 10H^+ \rightarrow 2\ Sb^{3+} + 2\ I_2 + 5\ H_2O$

   d) $5\ H_6TeO_6 + Br_2 \rightarrow 5\ TeO_2 + 2\ BrO_3^- + 14\ H_2O + 2\ H^+$

REDOX ANSWER KEY

Electrochemical Cell 1

1	iron (anode)
2	copper (cathode)
3	Fe
4	Cu
5	Fe
6	Cu
7	Fe
8	Fe
9	NO_3^-
10	Na^+, Fe^{2+}, Cu^{2+}
11	Cu
12	0.78V
13	$Cu^{2+}{}_{(aq)}$
14	$Fe_{(s)}$
15	completes the circuit; maintains electric neutrality
16	$Fe \rightarrow Fe^{2+} + 2e^-$
17	$Cu^{2+} + 2e^- \rightarrow Cu$
18	$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$

Electrochemical Cell 2

A	Electrons pass through the external circuit to the copper strip.
C	Positive and negative ions move through the aqueous solution to maintain electrical neutrality.
B	Electrons are passed to copper ions and reduction takes place.
D	Electrons are produced by oxidation.
Oxidation Half	$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$
Reduction Half	$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$
Overall	$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$

Practice: Balancing Redox Reactions

Pencast solutions for this sheet are published at http://redoxanswers.weebly.com

#	Balanced Redox Reaction	Oxidizing Agent	Reducing Agent
1	$2VO_2^+{}_{(aq)} + Zn_{(s)} + 4H^+{}_{(aq)} \rightarrow 2VO^{2+}{}_{(aq)} + Zn^{2+}{}_{(aq)} + 2H_2O_{(l)}$	VO_2^+	Zn
2	$Ag_{(s)} + NO_3^-{}_{(aq)} + 2H^+{}_{(aq)} \rightarrow NO_{2\,(g)} + Ag^+{}_{(aq)} + H_2O_{(l)}$	NO_3^-	Ag
3	$2MnO_4^-{}_{(aq)} + H^+{}_{(aq)} + 5HSO_3^-{}_{(aq)} \rightarrow 2Mn^{2+}{}_{(aq)} + 5SO_4^{2-}{}_{(aq)} + 3H_2O_{(l)}$	HSO_3^-	MnO_4^-
4	$4Zn_{(s)} + 2NO_3^-{}_{(aq)} + 10H^+{}_{(aq)} \rightarrow 4Zn^{2+}{}_{(aq)} + N_2O_{(g)} + 5H_2O_{(l)}$	NO_3^-	Zn
5	$Cr_{(s)} + NO_3^-{}_{(aq)} + 4H^+{}_{(aq)} \rightarrow Cr^{3+}{}_{(aq)} + NO_{(g)} + 2H_2O_{(l)}$	NO_3^-	Cr
6	$2Al_{(s)} + 2OH^-{}_{(aq)} + 6H_2O_{(l)} \rightarrow 2Al(OH)_4^-{}_{(aq)} + 3H_{2\,(g)}$	H_2O	Al
7	$2CrO_4^{2-}{}_{(aq)} + 3SO_3^{2-}{}_{(aq)} + 5H_2O_{(l)} \rightarrow 2Cr(OH)_{3(s)} + 3SO_4^{2-}{}_{(aq)} + 4OH^-{}_{(aq)}$	SO_3^{2-}	CrO_4^{2-}
8	$Zn_{(s)} + 2OH^-{}_{(aq)} + Cu(OH)_{2\,(s)} \rightarrow Zn(OH)_4^{2-}{}_{(aq)} + Cu_{(s)}$	Cu	Zn
9	$3HS^-{}_{(aq)} + ClO_3^-{}_{(aq)} \rightarrow 3S_{(s)} + Cl^-{}_{(aq)} + 3OH^-{}_{(aq)}$	HS^-	ClO_3^-
10	$NiO_{2\,(s)} + Zn_{(s)} + 2H_2O_{(l)} \rightarrow Ni(OH)_{2\,(s)} + Zn(OH)_{2\,(s)}$	NiO_2	Zn
11	$2MnO_4^- + 5H_2SO_3 + H^+ \rightarrow 2Mn^{2+} + 5HSO_4^- + 3H_2O$	MnO_4^-	H_2SO_3
12	$Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O$	$Cr_2O_7^{2-}$	I^-
13	$3As_2O_3 + 4H^+ + 4NO_3^- + 7\,H_2O \rightarrow 6H_3AsO_4 + 4NO$	NO_3^-	As_2O_3

Practice: Balancing Redox Reactions

#	Balanced Redox Reaction	Oxidizing Agent	Reducing Agent
14	$I_2 + H_2SO_3 + H_2O \rightarrow 2I^- + HSO_4^- + 3H^+$	I_2	H_2SO_3
15	$H_3AsO_4 + 8H^+ + 4Zn \rightarrow AsH_3 + 4Zn^{2+} + 4H_2O$	H_3AsO_4	Zn
16	$3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$	MnO_4^{2-}	MnO_4^{2-}
17	$2MnO_4^- + 5SO_2 + 2H_2O \rightarrow 2Mn^{2+} + 5SO_4^{2-} + 4H^+$	MnO_4^-	SO_2
18	$3HgS + 12Cl^- + 2NO_3^- + 8H^+ \rightarrow 3HgCl_4^{2-} + 3S + 2NO + 4H_2O$	NO_3^-	HgS
19	$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$	O_2	NH_3
20	$2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^-$	NO_2	NO_2
21	$H_2O_2 + 2Co(OH)_2 \rightarrow Co_2O_3 + 3H_2O$	H_2O_2	$Co(OH)_2$
22	$2Cu(NH_3)_4^{2+} + 7CN^- + 2OH^- \rightarrow CNO^- + 2Cu(CN)_3^{2-} + 8NH_3 + H_2O$	$Cu(NH_3)_4^{2+}$	CN^-
23	$ClO^- + Mn(OH)_2 \rightarrow MnO_2 + Cl^- + H_2O$	ClO^-	$Mn(OH)_2$