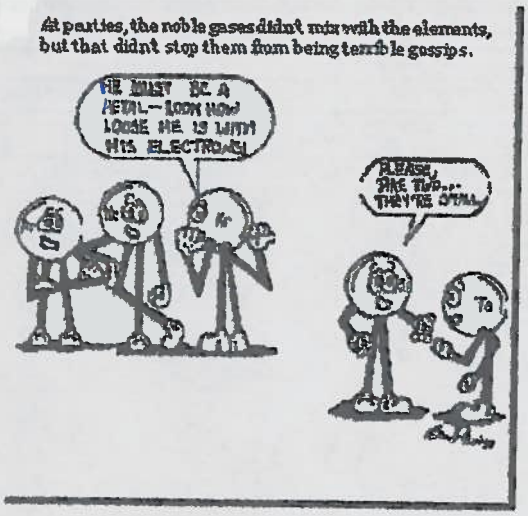
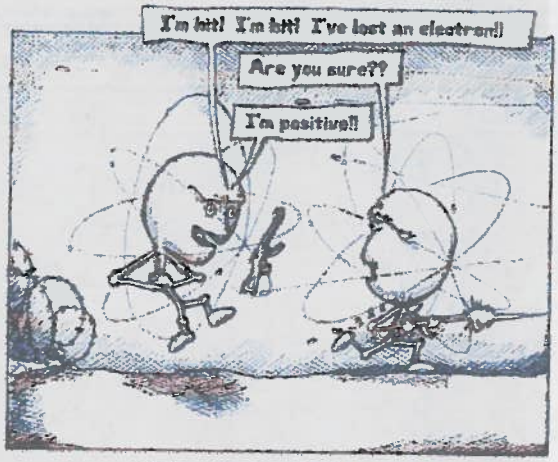
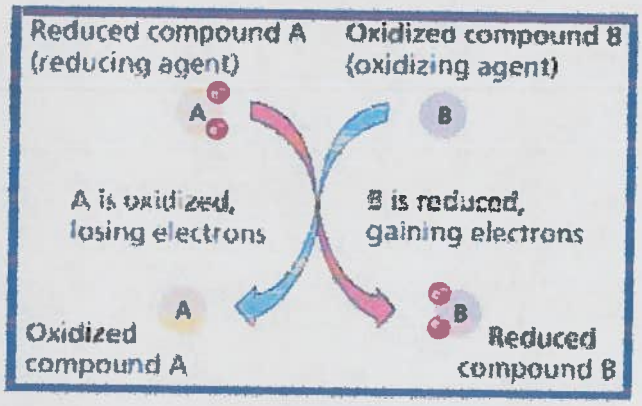
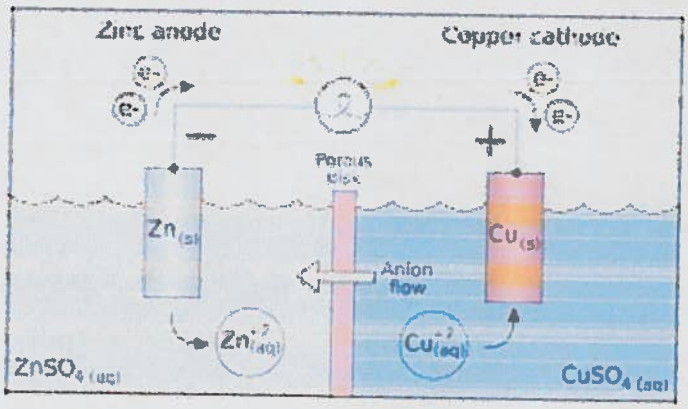




Another casualty in the War of the Atoms

Unit 12: Redox





Name_____ Per____

Mendoza Chemistry

2014-15

Rm 323/326

www.wbamendoza.weebly.com

April 2015

Monday	Tuesday	Wednesday	Thursday	Friday	Saturday	Sunday
		1	2	3 No School	4	5
6 No School	7 No School	8 No School	9 No School	10	11	12
13	14	15	16	17	18	19
20	21	22	23	24	25	26
27	28	29	30 Family STEM Night 7p-8:30pm			

Chapter 12: Oxidation & Reduction

Learning Objectives

1. An oxidation-reduction (redox) reaction involves the transfer of electrons (e^-). (3.2d)
   - Student should be able to determine a missing reactant or product in a balanced equation 3.2iii)
2. Reduction is the gain of electrons. (3.2e)
3. A half-reaction can be written to represent reduction. (3.2f)
   - Student should be able to write and balance half-reactions for oxidation and reduction of free elements and their monatomic ions (3.2vi)
4. Oxidation is the loss of electrons. (3.2g)
5. A half-reaction can be written to represent oxidation. (3.2h)
6. In a redox reaction the number of electrons lost is equal to the number of electrons gained. (3.3b)
7. Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred. (3.2i)
8. An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode. (3.2j)
   - Student should be able to compare and contrast voltaic and electrolytic cells (3.2ix)
9. A voltaic cell spontaneously converts chemical energy to electrical energy. (3.2k)
   - Student should be able to identify and label the parts of a voltaic cell (cathode, anode, salt bridge) and direction of electron flow, given the reaction equation (3.2vii)
   - Student should be able to use an activity series to deter- mine whether a redox reaction is spontaneous (3.2x)
10. An electrolytic cell requires electrical energy to produce chemical change. This process is known as electrolysis. (3.2l)

Unit Requirements

- Minimum of 2 POGILs. Additional may be done for extra credit.
- One formal lab report, as per the instructions in your lab book, typed and attached to the lab
  - Lab Report should be written in PARAGRAPH form, not as a bulleted list.
- Labs or any assignments with ** next to its name.
- One current event article
- Each Quiz must be 'mastered' or additional Learning Opportunities will be assigned before you may retake a quiz.
- The Test must be mastered or corrections must be completed in order to move on.

Section Learning Objectives	Date Completed	Learning Opportunities Minimum Suggested: 3-4 per section	Section Quiz Grade Mastery Level: 70 75 80	Activities, Labs & POGILs ** = Required Activity	Date Completed
12.1 Oxidation Numbers Obj # 7	_____ _____ _____ _____ _____ _____ _____ _____ _____ _____	1. Video Notes 12.1 pgs 2-4 2. 12.1 Video/Animations-on website 3. TB pgs 639-643 4. RevBook2013 pg 158-160 5. Vocab pg 2-Quizlet 12.1 6. Prac. WS pgs 5-8 7. Castle Learning 12.1 8. Poster or Concept Map 9. Review Sheet pg 10-12 10. Discuss with peers 11. Mini Lecture-Teacher	Grade 1_____ If mastery not achieved, complete _____ more Learning Opps Grade 2_____ Avg. Quiz Grade:	Activity Oxidation Connect 4- pg 9 (2-4 students per game) -Get cards from teacher	_____
12.2 Oxidation & Reduction Obj # 1,2,4,6	_____ _____ _____ _____ _____ _____ _____ _____ _____ _____	1. Video Notes 12.2 pgs 13-15 2. 12.2 Video/Animations-website 3. TB pgs 631-638 4. RevBook2013 pg 161-162 5. Vocab-Quizlet 12.2 6. Prac. WS pgs 18-19 7. Castle Learning 12.2 8. Poster or Concept Map 9. Review Sheet pg 20-21 10. Discuss with peers 11. Mini Lecture-Teacher	Grade 1_____ If mastery not achieved, complete _____ more Learning Opps Grade 2_____ Avg. Quiz Grade:	**CRT Practice Table J-pgs 16-17 **POGIL Oxidation & Reduction Half Reactions pgs 163-166 ** Lab Oxidation-Reduction Reactions -handout	_____ _____
12.3 Half Reactions Obj # 3,5	_____ _____ _____ _____ _____ _____ _____ _____ _____ _____	1. Video Notes 12.3 pgs 22-23 2. 12.3 Video/Animation-on website 3. TB pgs 645-653 4. RevBook2013 pg 163-164 5. Vocab-Quizlet 12.3 6. Prac. WS pgs 24-30 7. Castle Learning 12.3 8. Poster or Concept Map 9. Review Sheet pg 31-35 10. Discuss with peers 11. Mini Lecture-Teacher	Grade 1_____ If mastery not achieved, complete _____ more Learning Opps Grade 2_____ Avg. Quiz Grade:	** Labs #45: pgs 153-154 #46: pgs 155-156	_____ _____ _____

Section Learning Objectives	Date Completed	Learning Opportunities Minimum Suggested: 3-4 per section	Section Quiz Grade Mastery Level: 70 75 80	Activities, Labs & POGILs ** = Required Activity	Date Completed
12.4 Electro-chemical Cells Obj # 8,9,10	_____ _____ _____ _____ _____ _____ _____ _____ _____ _____ _____	1. Video Notes 12.4 pgs 36-39 2. Suppl.Video/Animation 3. TB pgs 663-683 4. RevBook2013 pg 165-168 5. Vocab-Quizlet 12.4 6. Prac. WS pgs 40-49 7. Castle Learning 12.4 8. Poster or Concept Map 9. Review Sheet pgs 52-61 10. Discuss with peers 11. Mini Lecture-Teacher	Grade 1_____ If mastery not achieved, complete _____ more Learning Opps Grade 2_____ Avg. Quiz Grade:	** Lab Handout: Voltaic Cell Computer Lab POGILs ** Choose 1: Voltaic Cells pgs 167-170 Batteries pgs 171-175 ** Webquest Power- pgs 50-51	_____ _____
Choose 1-2: Review & Test	_____ _____ _____ _____ _____ _____ _____ _____ _____ _____	1. Review Sheet pgs 62 2. Review Your Section Quizzes 3. Review Unit Study Guide pg 4. Create Unit Poster/Concept map 5. Castle Learning Review 6. Practice Exam pgs 63-71 7. www.regentsprep.org Unit Rev 8. Discuss Questions with Teacher In groups of 2-4: 9. Whiteboard Review 10. Discuss with peers	COMPLETE LAST: Unit Test Grade	** Current Event Article** Flaking Away pgs 72-75 Corrections must be completed if test grade is below mastery level	_____

Grade Sheet Attached:

For a 5-pt test bonus, get grade sheet signed by_____

Date:_____

_____

Parent/Guardian Signature Date

Checkpoints will occur at Progress Reports and Quarter Ends. Average will be reduced by a full grade if unit is not completed within a reasonable time of the posted deadline, based upon effort & participation during class time and attendance at extra help.

Unit 12: Oxidation-Reduction Vocabulary

1. Anode

2. Cathode

3. Electrochemical cell

4. Electrode

5. Electrolysis

6. Electrolytic cell

7. Half-reaction

8. Oxidation

9. Oxidation number (state)

10. Redox

11. Reduction

12. Salt bridge

13. Voltaic cell

14. LEO says GER

15. RED CAT and AN OX

Topic 12: Redox

12.1: Oxidation Numbers

AIM: _____

- ***Oxidation Numbers***
  - Keeps track of _____
    Oxidation numbers (states) can be _____
    - Identify how many electrons are either gained or lost by an atom or ion
      - 0 = neutral = _____ lost or gained
      - +1 = positive = one electron _____ : Group 1
      - -1 = negative = one electron _____ : Group 17
  - Some oxidation numbers _____
  - Others _____
- ***Rules for Assigning Oxidation Numbers***

1. An atom that is not combined with anything has an oxidation number of _____
   - $2Na + Cl_2$ Sodium and chlorine have oxidation numbers of _____
   - O_2, Au, Fe, Al Each has an oxidation number of _____
2. Monatomic ions have an oxidation number = to the _____ (Use CRT)
   - NaCl
     - Sodium oxidation number = _____; charge 1+
     - Chlorine oxidation number = _____; charge 1-
   - $Mg(Cl)_2$ MgO H_2O

3. Metals in group 1: oxidation number is _____
   Metals in group 2: oxidation number is _____
   - NaCl KBr MgO BaF_2

4. Fluorine is always _____. The other halogens are also _____ when they are the _____
   - NaF OF_2 $CaBr_2$

5. Hydrogen is _____ unless its combined with a metal when its _____

- HCl hydrogen = _____
- LiH hydrogen = _____

6. Oxygen is usually _____

- When combined with fluorine it becomes _____
  - H_2O oxygen = _____
  - OF_2 oxygen = _____
  - O_2^{2-} oxygen = _____

7. The sum of the oxidation numbers in all compounds = _____

– NaCl H_2O $MgSO_4$

8. The sum of the oxidation numbers in polyatomic ions must be

_____

– H_3O^+ CO_3^{-2} OH^-

- ***Assigning Oxidation Numbers***
  - What are the oxidation numbers of the atoms in HNO_3 ?
  - What is the oxidation number of chromium in the dichromate ion ($Cr_2O_7^{2-}$)?
- ***Regents Questions***
  - The oxidation number of an uncombined Group 2 metal is:

1) +1

2) +2

3) -2

4) 0

- Hydrogen has an oxidation number of:

1) 0 only
2) +1 only
3) -1 only
4) 0, +1 or -1

- What is the oxidation number of carbon in $NaHCO_3$?

1) -2
2) +2
3) -4
4) +4

- Chlorine has an oxidation state of +3 in the compound?

1) HClO
2) $HClO_2$
3) $HClO_3$
4) $HClO_4$

- What are the two oxidation states of nitrogen in the compound NH_4NO_3?

1) -3 and +5
2) -3 and -5
3) +3 and -5
4) +3 and +5

Chemistry: _____ Name _____

REDOX AND ELECTROCHEMISTRY Date _____ Period _____

Finding the Oxidation State

Study the rules for assigning oxidation numbers and examine the sample problem below. Then determine the unknown oxidation state in each example.

Rules for Assigning Oxidation Numbers

1. Oxidation numbers for atoms that are free elements are always zero
2. The oxidation numbers of ions are the same as the charge on the ion
3. Some elements have only one oxidation state
   a. group 1 metals always form 1+ ions and always have a +1 oxidation state
   b. group 2 metals always form 2+ ions and always have a +2 oxidation state
4. Some elements usually have a particular oxidation state
   a. oxygen has a –2 oxidation state except in peroxides where it is –1 and in compounds with fluorine (OF_2) where it is +2
   b. hydrogen has a +1 oxidation state except in hydrides with group 1 and group 2 metals
5. the sum of the oxidation numbers
   a. in a compound it is always zero
   b. in a polyatomic ion it is equal to the charge on the ion

Sample Problem

Find the oxidation state of the elements in $K_2Cr_2O_7$.

Element	K	Cr	O	
Subscript	2	2	7	TOTAL
Oxidation state	+1	?	–2	
Sum of oxidation states	+2	??	–14	0

[a] potassium is a group one metal; its oxidation state is always +1

[b] oxygen usually has an oxidation state of –2

[c] the sum of oxidation states of each element is the product of the subscript and the oxidation state

[d] find the sum of the oxidation states of chromium (??) by setting the sum of all the oxidation states to zero

$$(+2) + ?? + (-14) = 0$$

$$?? = +12$$

[f] find the oxidation state of chromium (?) by dividing the sum (+12) by the subscript (2)

$$+12 \div 2 = +6$$

1. Chlorine in $KClO_4$ 1. _____
2. Nitrogen in $Ba(NO_3)_2$ 2. _____
3. Phosphorus in $Ca_3(PO_4)_2$ 3. _____
4. Manganese in $LiMnO_4$ 4. _____
5. Sulfur in Na_2SO_3 5. _____
6. Chromium in $CaCrO_4$ 6. _____
7. Sulfur in MgS_2O_3 7. _____
8. Nitrogen in $Zn(NO_2)_2$ 8. _____
9. Chlorine in $HClO_3$ 9. _____
10. Carbon in CaC_2O_4 10. _____
11. Sulfur in $KHSO_4$ 11. _____

Determining Oxidation Numbers

Example: NO_3^{-1}

1) Sum must = -1
2) Oxygen = -2 (from Rule)
3) $-2 \times 3 = -6$ (there are 3 Os each worth -2)
4) $X + -6 = -1$ $= +5$ $N = +5$

Find the oxidation # of each element:

. $KClO_4$

. CuO_2

. Fe_2S_3

. K_2SO_4

. $LiOH$

$Mg(ClO_3)_2$

H_2O_2

. NO_2

. $NaMnO_4$

. $Al_2(SO_3)_3$

Name _____ Date _____ Class _____

REVIEW ACTIVITY

Oxidation Numbers

The *oxidation number* of an atom is the apparent charge assigned to it in a particular molecule or ion. Certain rules are followed in assigning oxidation numbers.

Use the rules at the right to assign oxidation numbers to each element in each of the given formulas.

Example

H_2O	+1, −2
N_2	0

Rules

The oxidation number of:

- an element in the uncombined state is 0.
- a monatomic ion equals the charge on the ion.
- hydrogen is generally +1; in hydrides, −1.
- oxygen is generally −2; in peroxides, −1.
- the more electronegative element in a binary covalent compound is negative, while that of the other element is positive.
- elements other than oxygen and hydrogen in a neutral compound is such that the sum of the oxidation numbers for all atoms in the compound is 0.
- elements other than oxygen and hydrogen in a polyatomic ion is such that the sum of the oxidation numbers for all atoms in the ion equals the charge on the ion.

1. Cl_2	1. _____
2. Cl^-	2. _____
3. Na	3. _____
4. Na^+	4. _____
5. KCl	5. _____
6. H_2S	6. _____
7. CaO	7. _____
8. H_2SO_4	8. _____
9. NO_3^-	9. _____
10. $Cr_2O_7^{2-}$	10. _____
11. NH_4Cl	11. _____
12. NH_3	12. _____
13. NO_2	13. _____
14. CaH_2 (calcium hydride)	14. _____
15. Na_2O_2 (sodium peroxide)	15. _____

(7)

ASSIGNING OXIDATION NUMBERS

Name _____

Assign oxidation numbers to all of the elements in each of the compounds or ions below.

1. HCl	11. H_2SO_3
2. KNO_3	12. H_2SO_4
3. OH^-	13. BaO_2
4. Mg_3N_2	14. $KMnO_4$
5. $KClO_3$	15. LiH
6. $Al(NO_3)_3$	16. MnO_2
7. S_8	17. OF_2
8. H_2O_2	18. SO_3
9. PbO_2	19. NH_3
10. $NaHSO_4$	20. Na

Connect 4

In pairs take it in turns to turn over a card, find the answer in the grid below and colour in the circle with your coloured crayon. If you do not know the answer, place your card at the bottom of the pile and miss your turn. The first person to get 4 in a row wins. Don't forget to check your opponent's answer! Each answer appears in more than one place, you choose which one would be best to colour in.

VII	II	-1	+6	+5	+4	+4
+4	+6	-3	VII	0	+4	-1
+4	+6	I	+5	-1	V	II
II	+5	+5	V	VII	+6	+5
+5	+6	0	-3	I	+4	+5
0	-3	VI	+5	0	+5	I
VI	-3	+6	-3	VI	-3	0

Oxidation Numbers

Rules:

1) Each uncombined element has an oxidation number of zero.
2) Monatomic ions have an oxidation number equal to the ionic charge.
3) The metals of Group 1 always have an oxidation number of +1 in compounds, and the metals of Group 2 always have an oxidation number of +2 in compounds.
4) Fluorine is always –1 in compounds. The other halogens are also –1 when they are the most electronegative element in the compound.
5) Hydrogen is +1 in compounds unless it is combined with a metal, in which case it is –1.
6) Oxygen is usually –2 in compounds. When it is combined with fluorine, which is more electronegative, it is +2. In the peroxide ion (O_2^{2-}), oxygen is –1.
7) The sum of the oxidation numbers in all compounds must be zero.
8) The sum of the oxidation numbers in polyatomic ions must be equal to the charge on the ion.

Definitions:

Electronegativity – a measure of the attraction of a nucleus for a bonded electron.

Oxidation – the loss of electrons by an atom or ion and a gain in oxidation number.

Reduction – the gain of electrons by an atom or ion and a loss in oxidation number.

Redox – term used to describe reactions where oxidation and reduction are occurring at the same time. IMPORTANT: Neither reduction nor oxidation can ever occur alone. They both must happen at the same time.

Oxidation Numbers (states) – identify how many electrons are either gained or lost by an atom or ion.

LEO says GER
LEO = Loss of Electrons in Oxidation
GER = Gain of Electrons in Reduction

How to write oxidation numbers: Mg^{2+} is the charge on a magnesium ion while Mg^{+2} is the oxidation number of Magnesium.

Practice:

What are the oxidation numbers of the atoms in HNO_3?

What is the oxidation number of chromium in the dichromate ion ($Cr_2O_7^{2-}$)?

What is the oxidation number of sulfur in SO_2?

What is the oxidation number of Oxygen in $Ca(OH)_2$?

1. What is the sum of the oxidation numbers of the atoms in the compound CO_2?
(1) 0 (3) −4
(2) −2 (4) +4

2. The oxidation numbers of all the atoms in H_2SO_4 must add up to
(1) 0 (3) +9
(2) +5 (4) +16

3. The oxidation number of nitrogen in N_2 is
(1) +1 (3) +3
(2) 0 (4) −3

4. If element X forms the oxides XO and X_2O_3, the oxidation numbers of element X are
(1) +1 and +2 (3) +1 and +3
(2) +2 and +3 (4) +2 and +4

5. In which compound does chlorine have the highest oxidation number?
(1) KClO (3) $KClO_3$
(2) $KClO_2$ (4) $KClO_4$

6. In which compound does chlorine have an oxidation number of +5?
(1) HClO (3) $HClO_3$
(2) $HClO_2$ (4) $HClO_4$

7. In which compound does hydrogen have a negative oxidation number?
(1) CaH_2 (3) NaOH
(2) H_3PO_4 (4) NH_3

8. In which substance does sulfur have an oxidation state of +4?
(1) H_2S (3) SO_3
(2) S (4) SO_2

9. What is the oxidation number of carbon in $NaHCO_3$?
(1) +6 (3) −4
(2) +2 (4) +4

10. In which species does hydrogen have an oxidation number of −1?
(1) H_2O (3) NaH
(2) H_2 (4) NaOH

11. What is the oxidation number of nitrogen in HNO_2?
(1) −1 (3) +3
(2) −2 (4) +4

12. Oxygen has an oxidation number of −2 in
(1) O_2 (3) Na_2O_2
(2) NO_2 (4) OF_2

13. What is the oxidation number of oxygen in hydrogen sulfate ion, HSO_4^-?
(1) +1 (3) +6
(2) −2 (4) −4

14. What is the oxidation number of oxygen in OF_2?
(1) +1 (3) −1
(2) +2 (4) −2

15. Oxygen has an oxidation number of −1 in
(1) H_3O^+ (3) H_2O
(2) OH^- (4) H_2O_2

16. What are the two oxidation states of nitrogen in the compound NH_4NO_3?
(1) −3 and −5 (3) +3 and −5
(2) −3 and +5 (4) +3 and +5

17. In which substance does phosphorus have a +3 oxidation state?
(1) P_4O_{10} (3) $Ca_3(PO_4)_2$
(2) PCl_5 (4) KH_2PO_3

(12)

Unit 12: Redox

12.2: Oxidation and Reduction

AIM: _____

- ***Oxidation and Reduction***
  - Oxidation = _____
  - Reduction = _____
  - They _____ occur together
  - Today we see it as a _____
  - Magnesium burned in oxygen
  - $Mg + O \rightarrow Mg^{2+} + O^{2-}$
    - _____
      - Magnesium _____
      - Oxygen _____
  - Magnesium reacting with chlorine
  - $Mg + Cl \rightarrow Mg^{2+} + 2\ Cl^{-}$

    Cl
    - _____
    - Magnesium _____
    - Each chlorine _____
  - Oxidation = _____ by an atom or ion

    $Mg + O \rightarrow Mg^{2+} + O^{2-}$
  - Reduction = _____ by an atom or ion

    $Mg + O \rightarrow Mg^{2+} + O^{2-}$
  - When at atom loses an electron (undergoes _____)

    another atom must gain the electron (undergoes _____)

- Because they always occur together we usually call them

  _____ (oxidation and reduction reactions)

- ***Oxidation***
  - Oxidation = _____
    - As something loses an electron it loses a _____
    - there is a _____ (becomes more _____)

$$Fe^0(s) \rightarrow Fe^{3+}(aq) + 3e^-$$

oxidation number gets more _____=electrons are _____

- ***Reduction***
  - Reduction = _____
    - As something gains electrons it _____
    - there is a _____ (becomes more _____)

$$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$$

oxidation number gets more _____=electrons are _____

- ***Oxidation – Reduction***
  - _____
    - LEO = _____
    - GER = _____
- ***Redox Reactions***
  - Not all reactions are _____
  - To determine if a reaction _____
    - Assign oxidation numbers to each atom in the reactants and products
    - If there is a change in oxidation number for an atom the reaction is redox
      - There is a _____
  - $MnO_2 + 4HCl \rightarrow MnCl_2 + 2Cl + 2H_2O$
    - _____ is oxidized (_____)
    - _____ is reduced (_____)

Regents Questions

- During a reduction reaction there is:

1) Loss of electrons and a loss of oxidation number
2) Loss of electrons and a gain of oxidation number
3) Gain of electrons and a loss of oxidation number
4) Gain of electrons and a gain of oxidation number

- As an atom is oxidized, the number of protons in the nucleus

1) Decreases
2) Increases
3) Remains the same
4) Depends on the atom

- In a redox reaction , the species reduced

1) Gains electrons
2) Gains oxidation number
3) Loses electrons and is the oxidation agent
4) Loses electrons and is the reducing agent

- A redox reaction always involves

1) A change in oxidation number
2) A change of phase
3) A transfer of protons
4) The formation of ions

- Which equation represents a redox reaction?

1) $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$
2) $HCl + KOH \rightarrow H_2O + KCl$
3) $2KClO_3 \rightarrow 2KCl + 3O_2$
4) $H_2CO_3 \rightarrow H_2O + CO_2$

- In the reaction $Cl_2 + H_2O \rightarrow HClO + HCl$, hydrogen is

1) Oxidized only
2) Reduced only
3) Both oxidized and reduced
4) Neither oxidized or reduced

- In the reaction $2KCl(l) \rightarrow 2K(s) + Cl_2(g)$, the K^+ ions are

1) Reduced by losing electrons
2) Reduced by gaining electrons
3) Oxidized by losing electrons
4) Oxidized by gaining electrons

Activity Series - Table J

Aim

- compare the activity of metals to hydrogen

Notes

- ☆ More active metals can replace less active metals
- ☆ Metals that are more active than hydrogen can replace hydrogen
- ☆ Hydrogen is used as a standard for comparing the activity of metals
  - ☆ Lithium (*MOST ACTIVE*)
  - ☆ Rubidium
  - ☆ Potassium
  - ☆ Cesium
  - ☆ Barium
  - ☆ Strontium
  - ☆ Calcium
  - ☆ Sodium
  - ☆ Magnesium
  - ☆ Aluminum
  - ☆ Titanium
  - ☆ Manganese
  - ☆ Zinc
  - ☆ Chromium
  - ☆ Iron
  - ☆ Nickel
  - ☆ Tin
  - ☆ Lead
  - ☆ HYDROGEN
  - ☆ Copper
  - ☆ Mercury
  - ☆ Silver
  - ☆ Platinum
  - ☆ Gold (*LEAST ACTIVE*)
- ☆ Acids release hydrogen when they react with active metals
- ☆ Active metals corrode easily
  - ☆ Definition: CORROSION — loss of metallic properties due to action of air, water, and chemicals
  - ☆ Examples
    - ☆ Rust: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
    - ☆ Action of Acids: $2Fe + 6HCl \rightarrow 2FeCl_3 + 3H_2$
- ☆ Spontaneous reactions - replacement of a less active metal by a more active metal occurs spontaneously

Answer the questions below by circling the number of the correct response

1. Which element is used as a standard for comparing the activity of metals? (1) gold (2) iron (3) francium (4) hydrogen

2. In which of the following pairs of metals is the more active metal listed first? (1) iron/sodium (2) copper/tin (3) lithium/platinum (4) zinc/magnesium

3. Based on the activity series, which of the following reactions is likely to occur? (1) $2Fe + 6HCl \rightarrow 2FeCl_3 + 3H_2$ (2) $MgSO_4 + Zn \rightarrow ZnSO_4 + Mg$ (3) $3BaCl_2 + 2Al \rightarrow 2AlCl_3 + 3Ba$ (4) $H_2 + 2LiOH \rightarrow 2Li + 2H_2O$

4. Based on the activity series, which of the following metals could ***NOT*** replace any of the others? (1) Calcium (2) Sodium (3) Magnesium (4) Aluminum

5. The standard on which the activity series is based is (1) fluorine, (2) lithium, (3) hydrogen, (4) oxygen.

6. Of the following, which is ***NOT*** a way to prevent corrosion? (1) painting (2) galvanizing (3) electroplating (4) coating with acid.

7. During a single displacement reaction, which of the following is true? (1) The more active metal steals electrons from the less active metal. (2) The more active metal is oxidized. (3) The more active metal is reduced. (4) The less active metal loses electrons.

8. During the reaction
$AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$,
which ion was reduced? (1) Ag^{+1} (2) Na^{+1} (3) Cl^{-1} (4) none of these

9. Which metal will react with 1.0 M $Pb^{+2}(aq)$ but not with 1.0 M Mg^{+2}? 1. Ba 2. Al 3. Cu 4. Ag

10. If the reaction $X + Zn^{+2} \rightarrow X^{+2} + Zn$ is spontaneous, then X may be 1. Mg 2. Pb 3. Cu 4. Sn

11. Which metal can reduce Pb^{+2}? (1) Cu (2) Hg (3) Fe (4) Ag

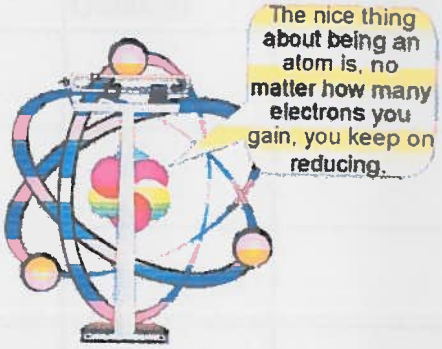
12. Which ion can be most easily reduced?
1. Cu^{+2} 2. Zn^{+2} 3. Ni^{+2} 4. Ca^{+2}

Chemistry: Name _____

REDOX AND ELECTROCHEMISTRY Date _____ Period _____

Analyzing Oxidation-Reduction Reactions

When chemical bonds form, electrons are either lost, gained or shared. Metals lose electrons. This is what happens when iron rusts. When the iron, a metal, combines with oxygen, a non metal, to form rust, it loses electrons. This process is called oxidation even when the nonmetal is not oxygen. Nonmetals gain electrons causing their oxidation states to go down. This is called reduction. It is possible to tell what was oxidized and what was reduced in a chemical reaction by checking the oxidation states of the elements before and after the reaction. The element that has an increase in oxidation state was oxidized while the one that has a decrease in oxidation state was reduced.



Example

$2FeCl_2 + Cl_2 \rightarrow 2FeCl_3$

$Fe^{+2} \rightarrow Fe^{+3}$ Iron was oxidized

$Cl^0 \rightarrow Cl^{-1}$ Chlorine was reduced

For each of the examples below, determine the oxidation states of the elements on both sides of the equation. Then determine which element was oxidized and which was reduced. Write your answer in the space provided.

Reaction	Element:	
	Oxidized	Reduced
Example: $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$ $\overset{0}{Cu} + 2\overset{+1\ +5\ -2}{AgNO_3} \rightarrow \overset{+2\ +5\ -2}{Cu(NO_3)_2} + 2\overset{0}{Ag}$	Cu	Ag
1. $2Mg + O_2 \rightarrow 2MgO$		
2. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$		
3. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$		
4. $2K_2Cr_2O_7 + 2H_2O + 3S \rightarrow 4KOH + 2Cr_2O_3 + 3SO_2$		

(☞ Go on to the next page.)

Oxidation-Reduction Computer Activity

Go to the following website and complete the activity

http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7604

In oxidation and reduction reactions, sometimes called _____ reactions, one or more _____ are transferred between chemicals.

An example of an oxidation-reduction reaction is the following:

$$3CuCl_2(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

In this reaction the Cu^{2+} ions have been converted into copper atoms. This requires that each copper(II) ion gains two electrons.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

The electrons for this transformation come from the aluminum atoms: they are converted into Al^{3+} ions by losing 3 electrons.

$$Al(s) \rightarrow Al^{3+} + 3e^-$$

The net result is that Cu^{2+} ions come out of solution as Cu metal and Al metal goes into solution as Al^{3+} ion.

$$3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$$

Oxidation is defined as the _____ of electrons from a substance. In this reaction the _____ _____ are oxidized because they _____ electrons.

$$Al(s) \rightarrow Al^{3+} + 3e^-$$

Reduction is defined as the _____ of electrons by a substance. In this reaction _____ are reduced because they _____ electrons.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

Each of these reactions is part of the overall oxidation-reduction reactions and is called a _____.

Which of the following best describes an oxidation-reduction reaction?

(20)

1. All redox reactions involve
(1) the gain of electrons, only
(2) the loss of electrons, only
(3) both the gain and the loss of electrons
(4) neither the gain nor the loss of electrons

2. A redox reaction always involves
(1) a change in oxidation number
(2) a change of phase
(3) the transfer of protons
(4) the formation of ions

3. For a redox reaction to occur, there must be a transfer of
(1) protons
(2) neutrons
(3) electrons
(4) ions

4. Oxidation-reduction reactions occur because of the competition between particles for
(1) neutrons
(2) electrons
(3) protons
(4) positrons

5. As a sodium atom is oxidized, the number of protons in its nucleus
(1) decreases
(2) increases
(3) remains the same

6. Which change in oxidation number represents reduction?
(1) −1 to +1
(2) −1 to −2
(3) −1 to +2
(4) −1 to 0

7. Which change in oxidation number represents reduction?
(1) −3 to 0
(2) −2 to −3
(3) 0 to +1
(4) +1 to +2

8. Which is true when an Sn^{2+} ion is reduced?
(1) Its oxidation number increases.
(2) It gains electrons.
(3) Its mass decreases.
(4) It acts as a reducing agent.

9. Which change occurs when an Sn^{2+} ion is oxidized?
(1) Two electrons are lost.
(2) Two electrons are gained.
(3) Two protons are lost.
(4) Two protons are gained.

10. Which statement correctly describes a redox reaction?
(1) The oxidation half-reaction and the reduction half-reaction occur simultaneously.
(2) The oxidation half-reaction occurs before the reduction half-reaction.
(3) The oxidation half-reaction occurs after the reduction half-reaction.
(4) The oxidation half-reaction occurs spontaneously but the reduction half-reaction does not.

11. When 1 mole of Ni^{3+} changes to Ni^{2+}, the Ni^{3+} undergoes
(1) oxidation by losing electrons
(2) oxidation by gaining electrons
(3) reduction by losing electrons
(4) reduction by gaining electrons

12. Given the reaction:
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$
This reaction can be classified as
(1) a reduction reaction, because there is a decrease in oxidation number
(2) a reduction reaction, because there is an increase in oxidation number
(3) an oxidation reaction, because there is a decrease in oxidation number
(4) an oxidation reaction, because there is an increase in oxidation number

(21)

Topic 12: Redox
12.3: Half Reactions

AIM: _____

- ***Half Reactions***
  - Show either the _____
    - Include the _____
  - Must have _____
    - same number of atoms on both sides
  - Must have _____
  - same _____
  - the charge does not _____
- ***Reduction Half Reactions***
  - shows atom or ion _____
    - decreasing _____
  -
    - Conservation of mass: _____
    - Conservation of charge: _____
- ***Oxidation Half Reactions***
  - Shows atoms or ions _____
    - Shows oxidation number _____
  -
    - Conservation of mass : _____
    - Conservation of charge : _____
- ***Half Reactions***
  - Write the half reactions from the equation and then balance the equation

$$Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$$

- ***Regents Question***

- In a half-reaction

1) Only mass if conserved
2) Charge only is conserved
3) Both mass and charge are conserved
4) Neither mass nor charge are conserved

- In the reaction $Mg + Cl_2 \rightarrow MgCl_2$, the correct half-reaction for the oxidation that occurs is

1) $Mg + 2e^- \rightarrow Mg^{2+}$
2) (2) $Mg \rightarrow Mg^{2+} + 2e^-$
3) (3) $Cl_2 + 2e^- \rightarrow 2Cl^-$
4) (4) $Cl_2 \rightarrow 2Cl^- + 2e^-$

- Given the reaction:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Write the balanced oxidation half-reaction for this oxidation-reduction reaction

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Competition for Electrons

Aim

- write equations for oxidation and reduction half reactions

Notes

Atoms compete for each other's electrons

- ☆ When chemical bonds form, electrons are either lost, gained or shared
- ☆ Oxidation-Reduction reactions (Redox reactions)
  - ☆ Metals
    - ☆ lose electrons (OXIDATION)[*NOTE: as when metals combine with oxygen*]
    - ☆ are oxidized
    - ☆ are reducing agents
  - ☆ Nonmetals
    - ☆ gain electrons reducing their oxidation states (REDUCTION)
    - ☆ are reduced
    - ☆ are oxidizing agents

Oxidation Is Loss

Reduction Is Gain

- ☆ Example 1 - $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Mg	O_2
★ loses electrons	★ gains electrons
★ gets oxidized to Mg^{2+}	★ gets reduced to O^{2-}
★ is the reducing agent for O_2	★ is the oxidizing agent for Mg

  - ☆ Half reactions — reaction showing either a gain or loss of electrons
    - ☆ $2Mg^0 \rightarrow 2Mg^{2+} + 4e^-$
    - ☆ $O_2^0 + 4e^- \rightarrow 2O^{2-}$
  - ☆ Net equation (REDOX REACTION)— combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- ☆ Example 2 - More active metals replace less active metals in compounds by transferring electrons to them
  - ☆ Sample Reaction:

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

  - ☆ Half reactions — reaction showing either a gain or loss of electrons
    - ☆ $Zn^0 \rightarrow Zn^{2+} + 2e^-$
    - ☆ $Cu^{2+} + 2e^- \rightarrow Cu^0$
  - ☆ Net equation — combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$Cu^{2+} + Zn^0 \rightarrow Zn^{2+} + Cu^0$$

  - ☆ Spectator ions — ions that are present during a reaction but do not participate in the reaction: $2NO_3^-$

Oxidation number (Oxidation state) - number assigned to keep track of electrons based on the arbitrary assumption that shared electrons belong to the more electronegative element

- ☆ Rules for assigning oxidation numbers
  - ☆ Oxidation numbers for atoms that are free elements are always zero
  - ☆ The oxidation numbers of ions are the same as the charge on the ion
  - ☆ Some elements have only one oxidation state
    - ☆ group 1 metals always form 1+ ions and always have a +1 oxidation state
    - ☆ group 2 metals always form 2+ ions and always have a +2 oxidation state
  - ☆ Some elements usually have a particular oxidation state
    - ☆ oxygen has a -2 oxidation state except in peroxides where it is -1 and in compounds with fluorine (OF_2) where it is +2
    - ☆ hydrogen has a +1 oxidation state except in hydrides with group 1 and group 2 metals
  - ☆ the sum of the oxidation numbers
    - ☆ in a compound it is always zero
    - ☆ in a polyatomic ion it is equal to the charge on the ion
- ☆ Finding oxidation numbers
  - ☆ apply the rules
  - ☆ construct a table if necessary

Sample Problem

Find the oxidation state of the elements in $K_2Cr_2O_7$.

Element	K	Cr	O	TOTAL
Subscript	2	2	7	
Oxidation state	+1	?	-2	
Sum of oxidation states	+2	??	-14	0

[a] potassium is a group one metal; its oxidation state is always +1

[b] oxygen usually has an oxidation state of -2

[c] the sum of oxidation states of each element is the product of the subscript and the oxidation state

[d] find the –sum of the oxidation states of chromium (??) by setting the sum of all the oxidation states to zero

$(+2) + ?? + (-14) = 0$

$?? = +12$

[f] find the oxidation state of chromium (?) by dividing the sum (+12) by the subscript (2)

$+12 \div 2 = +6$

Answer the questions below by circling the number of the correct response

1. In this reaction, the oxidation number (oxidation state) of C changes from: $2CO_2 \rightarrow 2CO + O_2$
(1) 0 to +4 (2) +2 to +4 (3) +3 to 0 (4) +4 to +2

2. In the reaction:
$2KMnO_4 + 3H_2SO_4 + 5H_2S \rightarrow 5S + 2MnSO_4 + K_2SO_4 + 8H_2O$
the oxidation number of sulfur changes from
(1) +5 to -5 (2) -5 to +5 (3) 0 to -2 (4) -2 to 0

3. What is the oxidation number of Cr in Na_2CrO_4?
(1) +1 (2) +2 (3) +3 (4) +6

4. What is the oxidation state of the chromium in $K_2Cr_2O_7$?
(1) +5 (2) +6 (3) +3 (4) +12

5. In the reaction $Pb + 2Ag^+ \rightarrow Pb^{+2} + 2Ag$, the reducing agent is
(1) Ag (2) Ag^+ (3) Pb (4) Pb^{+2}

6. Which is not an oxidation-reduction reaction?
(1) $4Na + O_2 \rightarrow 2Na_2O$
(2) $Fe + 2HCl \rightarrow FeCl_2 + H_2$
(3) $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$
(4) $2H_2O \rightarrow 2H_2 + O_2$

7. Given: $2Al + 3Zn^{+2} \rightarrow 2Al^{+3} + 3Zn$. In this reaction, the oxidizing agent is (1) Al (2) Al^{+3} (3) Zn (4) Zn^{+2}

8. Given: $2Al + 3Zn^{+2} \rightarrow 2Al^{+3} + 3Zn$ In this reaction, electrons are transferred from (1) Al to Al^{+3} (2) Zn^{+2} to Zn (3) Al to Zn^{+2} (4) Zn^{+2} to Al

9. What is the oxidation number of nitrogen in N_2O_3? (1) +1 (2) +2 (3) +3 (4) +6

10. In the reaction $3CO + Fe_2O_3 \rightarrow 3CO_2 + 2Fe$, the oxidation number of the iron changes from (1) +2 to 0 (2) +2 to +3 (3) +3 to +2 (4) +3 to 0

11. What is the oxidation number of Br in BrO_3^{-2}?
(1) +1 (2) +6 (3) +5 (4) +4

12. Which is the reducing agent in the following reaction?
$Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$
(1) Cl_2 (2) H_2O (3) K^+ (4) Br^-

13. What is the oxidation number of carbon in $C_2O_4^{-2}$?
(1) +1 (2) +2 (3) +3 (4) +4

14. Which is an oxidation-reduction reaction?
(1) $CaCO_3 \rightarrow CaO + CO_2$
(2) $KOH + HBr \rightarrow KBr + H_2O$
(3) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
(4) $Mg + Cl_2 \rightarrow MgCl_2$

15. $MnSO_4$ is a product in a reaction that contained $KMnO_4$ as a reactant. The oxidation number of the manganese changed from
(1) -2 to +5 (2) +7 to +2 (3) +5 to -2 (4) -7 to +2

16. Given the balanced equation:
$2HNO_3 + 3H_2S \rightarrow 4H_2O + 2NO + 3S$
Which is reduced? (1) S (2) S-2 (3) N+2 (4) N+5

17. During the reaction $Ca + H_2 \rightarrow CaH_2$, the oxidation number of the hydrogen changes from
(1) 0 to +1 (2) +1 to 0 (3) 0 to -1 (4) -1 to 0

18. In the reaction $Sn^{+4} + H_2(g) \rightarrow Sn^{+2} + 2H^+$, the reducing agent is
(1) Sn^{+4} (2) H_2 (3) Sn^{+2} (4) H^+

19. Given: $3Ag + 4HNO_3 \rightarrow NO + 3AgNO_3 + 2H_2O$. The reducing agent in this reaction is
(1) Ag (2) Ag^{+1} (3) H^{+1} (4) N^{+2}

20. The reaction $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ is an example of
(1) an oxidation reaction, only
(2) a reduction reaction, only
(3) both an oxidation and a reduction reaction
(4) neither an oxidation nor a reduction reaction

21. The oxidation number of manganese in $KMnO_4$ is
(1) +1 (2) +7 (3) +3 (4) +4

22. In the reaction $Sn^{+2} + 2Fe^{+3} \rightarrow Sn^{+4} + 2Fe^{+2}$, the reducing agent is
(1) Fe^{+2} (2) Fe^{+3} (3) Sn^{+2} (4) Sn

23. An oxidizing agent will always
(1) lose electrons (3) be reduced
(2) increase in oxidation number (4) increase in mass

Chemistry: [redacted]

Name _____

REDOX AND ELECTROCHEMISTRY

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Writing Half Reactions

During a redox reaction electrons are both lost and gained. The metal loses and the non metal gains. An equation showing either the gain or the loss of electrons but not both is called a half reaction. Consider the reaction below:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Magnesium loses electrons while oxygen gains. The reaction can be split into two half reactions showing each. The oxidation half reaction shows the loss of electrons. Electrons are shown on the product side of the equation. The reduction half reaction shows the electron gain. Electrons are shown on the reactant side of the equation.

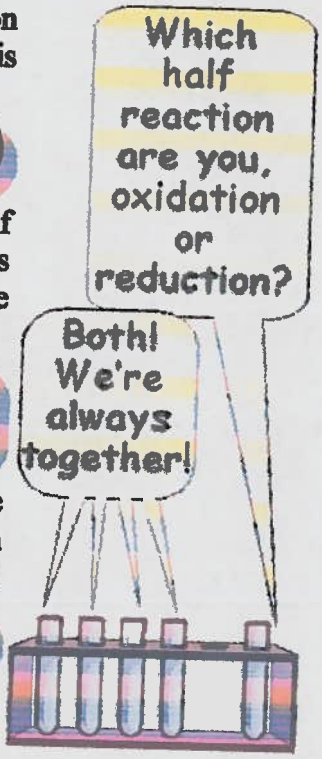
Oxidation Half: $2[Mg^0 \rightarrow Mg^{2+} + 2e^-]$

Reduction Half: $2[O^0 + 2e^- \rightarrow O^{2-}]$

The net equation, the redox reaction, is a combination of the half reactions such that the number of electrons lost equals the number of electrons gained. The electrons are not shown in the net equation because the electrons that were lost are the same ones that were gained.

Net Reaction: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

To write the half reactions, it is first necessary to determine the oxidation states of the elements on both sides of the equation so you know what was oxidized and what was reduced. Then write the oxidation and reduction halves as shown above, making sure the equation is balanced so the number of electrons lost equals the number gained.



Write the half reactions for each of the redox reactions below:

1. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$

2. $CdS + I_2 + HCl \rightarrow CdCl_2 + HI + S$

3. $NaClO + H_2S \rightarrow NaCl + H_2SO_4$

☞ Go on to the next page.

4. $Sn + HNO_3 + H_2O \rightarrow H_2SnO_3 + NO$

5. $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + H_2O + Cl_2$

6. $Fe(OH)_2 + H_2O_2 \rightarrow Fe(OH)_3$

7. $Na + H_2O \rightarrow NaOH + H_2$

8. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$

9. $H_2O_2 \rightarrow H_2O + O_2$

10. $K_2Cr_2O_7 + H_2O + S \rightarrow SO_2 + KOH + Cr_2O_3$

____1. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-}?

1) +8 3) +6
2) +2 4) +3

____2. Given the reaction that occurs in an electrochemical cell:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

During this reaction, the oxidation number of Zn changes from

1) +2 to 0 3) 0 to −2
2) 0 to +2 4) −2 to 0

____3. In a redox reaction, the total number of electrons lost is

1) greater than the total number of electrons gained
2) equal to the total number of protons gained
3) less than the total number of electrons gained
4) equal to the total number of electrons gained

____4. Which changes occur when Pt^{2+} is reduced?

1) The Pt^{2+} loses electrons and its oxidation number decreases.
2) The Pt^{2+} loses electrons and its oxidation number increases.
3) The Pt^{2+} gains electrons and its oxidation number increases.
4) The Pt^{2+} gains electrons and its oxidation number decreases.

____5. Half-reactions can be written to represent all

1) neutralization reactions
2) double-replacement reactions
3) oxidation and reduction reactions
4) fission and fusion reactions

____6. Which particles are gained and lost during a redox reaction?

1) electrons 3) neutrons
2) protons 4) positrons

____7. In a redox reaction, there is a conservation of

1) both mass and charge 3) mass, only
2) neither mass nor charge 4) charge, only

____8. Which half-reaction correctly represents oxidation?

1) $Fe^{3+}(aq) \rightarrow Fe(s) + 2e^-$
2) $Fe(s) + 2e^- \rightarrow Fe^{2-}(aq)$
3) $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$
4) $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

____9. An oxidation half-reaction always involves the

1) gain of electrons and a decrease in the oxidation number
2) gain of electrons and an increase in the oxidation number
3) loss of electrons and a decrease in the oxidation number
4) loss of electrons and an increase in the oxidation number

10. Base your answer to the following question on the information below.

During a laboratory activity, a student reacted a piece of zinc with 0.1 *M* HCl(aq).

Based on Reference Table *J*, identify *one* metal that does *not* react spontaneously with HCl(aq).

11. What is the oxidation number of nitrogen in NO(g)?

Base your answers to questions 12 and 13 on the following redox reaction, which occurs spontaneously.

$$Zn + Cr^{3+} \rightarrow Zn^{2+} + Cr$$

12. Write the half-reaction for the reduction that occurs.

13. Write the half-reaction for the oxidation that occurs.

1. Which half-reaction correctly represents oxidation?
   (1) $Sn^{2+} + 2e^- \rightarrow Sn^0$
   (2) $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$
   (3) $Sn^{2+} \rightarrow Sn^0 + 2e^-$
   (4) $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

2. Which half-reaction correctly represents oxidation?
   (1) $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
   (2) $Fe^{2+}(aq) \rightarrow Fe(s) + 2e^-$
   (3) $Fe(s) + 2e^- \rightarrow Fe^{2+}(aq)$
   (4) $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

3. Which half-reaction correctly represents oxidation?
   (1) $F_2 \rightarrow 2F^- + 2e^-$
   (2) $F_2 + 2e^- \rightarrow 2F^-$
   (3) $H_2 \rightarrow 2H^+ + 2e^-$
   (4) $H_2 + 2e^- \rightarrow 2H^+$

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

5. A redox reaction always involves
   (1) a change in oxidation number
   (2) a change of phase
   (3) the transfer of protons
   (4) the formation of ions

6. As a sodium atom is oxidized, the number of protons in its nucleus
   (1) decreases
   (2) increases
   (3) remains the same

7. In the half-reaction
   $Pb^0 \rightarrow Pb^{2+} + 2e^-$, the Pb^0
   (1) gains protons
   (2) loses protons
   (3) is oxidized
   (4) is reduced

8. When a redox reaction occurs, there must be a transfer of
   (1) electrons
   (2) neutrons
   (3) protons
   (4) ions

9. An oxidation half-reaction always involves the
   (1) gain of electrons and a decrease in the oxidation number
   (2) gain of electrons and an increase in the oxidation number
   (3) loss of electrons and a decrease in the oxidation number
   (4) loss of electrons and an increase in the oxidation number

10. What occurs when a substance in an oxidation-reduction reaction is reduced?
    (1) It loses electrons, and its oxidation number decreases.
    (2) It loses electrons, and its oxidation number increases.
    (3) It gains electrons, and its oxidation number decreases.
    (4) It gains electrons, and its oxidation number increases.

11. When a neutral atom undergoes oxidation, the atom's oxidation state
    (1) decreases as it gains electrons
    (2) decreases as it loses electrons
    (3) increases as it gains electrons
    (4) increases as it loses electrons

12. Which type of reaction occurs when nonmetal atoms become negative nonmetal ions?
    (1) oxidation
    (2) reduction
    (3) substitution
    (4) condensation

13. The transfer of which particle is required for a redox reaction to occur?
    (1) electron
    (2) ion
    (3) neutron
    (4) proton

14. As an atom of nitrogen gains electrons, its oxidation number
    (1) decreases
    (2) increases
    (3) remains the same

15. In a redox reaction, how does the total number of electrons lost by the oxidized substance compare to the total number of electrons gained by the reduced substance?
    (1) The number lost is always greater than the number gained.
    (2) The number lost is always equal to the number gained.
    (3) The number lost is sometimes equal to the number gained.
    (4) The number lost is sometimes less than the number gained.

Redox

Just to remind you...

Oxidation
Oxidation reactions involve the loss of electrons.

Example:
$Fe \rightarrow Fe^{2+} + 2e^-$

Remember "OIL RIG" (Oxidation Is Loss, Reduction Is Gain)

Reduction
Reduction is the reverse of oxidation — reduction is the gain of electrons.

Example:
$Cl_2 + 2e^- \rightarrow 2Cl^-$

Redox Reactions

Redox is short for Reduction-Oxidation — a redox reaction is when oxidation and reduction happen at the same time. A typical redox reaction occurs when a metal reacts with oxygen to form an oxide — for example, magnesium and oxygen react together to form magnesium oxide.
The magnesium is oxidized and loses electrons, and the oxygen is reduced — it gains electrons from the magnesium.
All redox reactions involve electron transfers like this.

Example:
$2Mg + O_2 \rightarrow 2MgO$

In redox reactions, each substance has an oxidation number

The oxidation number of a simple ion, such as Na^+, is the same as its charge. So the oxidation number for an atom increases by 1 for each electron lost and decreases by 1 for each electron gained.
For example, sodium and chlorine react together to produce sodium chloride.
In this redox reaction, sodium loses an electron and its oxidation number changes from 0 to +1.
Chlorine gains an electron and its oxidation number changes from 0 to –1:

Redox reaction electron transfer: sodium loses e^- ... chlorine gains e^-

$Na + \frac{1}{2}Cl_2 \rightarrow Na^+ Cl^-$

Oxidation Number: 0 0 +1 –1

A change in the oxidation number shows that oxidation or reduction has occurred.

The Half-Equations — make sure the electrons balance

You can write a half-equation for each part of a redox reaction — one for the reduction and one for the oxidation. The main thing is to make sure the number of electrons is the same for both half-equations.

Below are the half-equations for this reaction: $2Mg + O_2 \rightarrow 2MgO$

Oxidation: $Mg \rightarrow Mg^{2+} + 2e^-$

Reduction: $O + 2e^- \rightarrow O^{2-}$

The equations aren't right because oxygen is found as the gas O_2. They need rewriting like this:

Oxidation: $2Mg \rightarrow 2Mg^{2+} + 4e^-$

Reduction: $O_2 + 4e^- \rightarrow 2O^{2-}$

If you combine the half-equations and cancel out the electrons, you get the overall equation.

Note that there are two electrons in both half-equations, which means they're nice and balanced.

Oxidation and Reduction — they make a lovely couple...

Redox reactions aren't as hard as they may at first appear. My advice would be to read over this page again until you absolutely and totally understand every word of it. Then read over it one more time to be sure...

Redox

Q1 For each of these ionic equations, state whether it involves oxidation or reduction.

a) $Na \rightarrow Na^+ + e^-$ b) $Br_2 + 2e^- \rightarrow 2Br^-$ c) $2H^+ + 2e^- \rightarrow H_2$

Q2 What is a redox reaction?

Please Note:
Due to printing restrictions, a "Red Ox" could not be shown on this page.
Please be amused by this blue goat instead.

Q3 What does "OIL RIG" stand for in relation to redox reactions?

Q4 When molten aluminum oxide is electrolyzed, there is a redox reaction.

Here are the two half-equations for the reaction:

$4Al^{3+} + 12e^- \rightarrow 4Al$ $6O^{2-} \rightarrow 3O_2 + 12e^-$

a) State whether the aluminum ions and the oxygen ions have been oxidized or reduced.

b) What is the oxidation number of the Al^{3+} ions? What is the oxidation number of the Al atoms?

c) What is the oxidation number of the O^{2-} ions? What is the oxidation number of the O atoms in the O_2 molecules?

d) Copy and complete these sentences:

Each time an atom _____ an electron its oxidation number increases by one.
Each time an atom _____ an electron its oxidation number decreases by one.

Q5 When light hits silver bromide, it splits it into silver and bromine. This is a redox reaction.

Here are the two half-equations for the reaction:

$2Ag^+ + 2e^- \rightarrow 2Ag$ $2Br^- \rightarrow Br_2 + 2e^-$

a) Explain why this is a redox reaction.

b) Write down the oxidation numbers of the following:

i) Ag^+ ii) Ag iii) Br^- iv) Br

Q6 Balance each of these pairs of half-equations.

a) $Na \rightarrow Na^+ + e^-$ b) $K \rightarrow K^+ + e^-$ c) $Al \rightarrow Al^{3+} + 3e^-$ d) $Al^{3+} + 3e^- \rightarrow Al$

$Br_2 + 2e^- \rightarrow 2Br^-$ $O_2 + 4e^- \rightarrow 2O^{2-}$ $Cl_2 + 2e^- \rightarrow 2Cl^-$ $2O^{2-} \rightarrow O_2 + 4e^-$

REDOX reactions are just REDuction/OXidation reactions...

A redox reaction is any reaction where one thing gives electrons to another — reduction and oxidation both happen. You can't really have one without the other, but you can have equations that only show one (half-equations). If you have a pair of these (one for the oxidation and one for the reduction), you need to make sure there's the same number of electrons in each before you can say they're balanced.

1. Which half-reaction correctly represents reduction?
(1) $Ca^{2+} \rightarrow Ca + 2e^-$ (3) $2F^- + 2e^- \rightarrow F_2$
(2) $Ca^{2+} + 2e^- \rightarrow Ca$ (4) $2F^- \rightarrow F_2 + 2e^-$

2. Which half-cell reaction correctly represents oxidation?
(1) $Pb^{2+} + 2e^- \rightarrow Pb$ (3) $Pb^{2+} \rightarrow Pb + 2e^-$
(2) $Pb + 2e^- \rightarrow Pb^{2+}$ (4) $Pb \rightarrow Pb^{2+} + 2e^-$

3. Which half-reaction represents reduction?
(1) $Ca^0 \rightarrow Ca^{2+} + 2e^-$ (3) $Ca^{2+} + 2e \rightarrow Ca^0$
(2) $Cl_2^0 - 2e \rightarrow 2Cl^+$ (4) $2Cl^- \rightarrow Cl_2^0 + 2e^-$

4. Which half-reaction correctly represents a reduction reaction?
(1) $Sn^0 + 2e^- \rightarrow Sn^{2+}$ (3) $Li^0 + e^- \rightarrow Li^+$
(2) $Na^0 + e^- \rightarrow Na^+$ (4) $Br_2^0 + 2e^- \rightarrow 2Br^-$

5. Which half-reaction correctly represents reduction?
(1) $Cr^{3+} + 3e^- \rightarrow Cr(s)$ (3) $Cr(s) \rightarrow Cr^{3+} + 3e^-$
(2) $Cr^{3+} \rightarrow Cr(s) + 3e^-$ (4) $Cr(s) + 3e^- \rightarrow Cr^{3-}$

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

7. Which half-reaction correctly represents oxidation?
(1) $Sn^{2+} + 2e^- \rightarrow Sn^0$ (3) $Sn^{2+} \rightarrow Sn^0 + 2e^-$
(2) $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$ (4) $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

8. Which species is produced when a hydrogen atom is oxidized?
(1) H:H (3) H·
(2) $H:^-$ (4) H^+

9. Which half-reaction correctly represents oxidation?
(1) $F_2 \rightarrow 2F^- + 2e^-$ (3) $H_2 \rightarrow 2H^+ + 2e^-$
(2) $F_2 + 2e^- \rightarrow 2F^-$ (4) $H_2 + 2e^- \rightarrow 2H^+$

10. In the reaction
$2CO + O_2 \rightarrow CO_2$,
the oxidation number of carbon changes from
(1) +2 to +1 (3) +4 to +1
(2) +2 to +4 (4) +4 to +2

11. In the half-reaction
$Pb^0 \rightarrow Pb^{2+} + 2e^-$, the Pb^0
(1) gains protons (3) is oxidized
(2) loses protons (4) is reduced

12. In the reaction
$Sn^{2+}(aq) + 2Ag^+(aq) \rightarrow Sn^{4+}(aq) + 2Ag(s)$,
the species reduced is
(1) Sn^{2+} (3) Sn^{4+}
(2) Ag^+ (4) Ag

13. In the reaction:
$Pb + 2Ag^+ \rightarrow Pb^{2+} + 2Ag$,
the Ag^+ is
(1) reduced, and the oxidation number changes from +1 to 0
(2) reduced, and the oxidation number changes from +2 to 0
(3) oxidized, and the oxidation number changes from 0 to +1
(4) oxidized, and the oxidation number changes from +1 to 0

14. In the reaction
$2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe^0(s)$,
the species oxidized is
(1) Al(s) (3) Fe(s)
(2) $Al^{3+}(aq)$ (4) $Fe^{2+}(aq)$

15. In the reaction
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$,
the species oxidized is
(1) Zn(s) (3) $Cu^{2+}(aq)$
(2) Cu(s) (4) $Zn^{2+}(aq)$

16. In the half-cell reaction, $Ba^0 \rightarrow Ba^{2+} + 2e^-$, which is true of the barium atom?
(1) It gains protons. (3) It gains electrons.
(2) It loses protons. (4) It loses electrons.

1. Given the reaction:

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Which half-cell reaction represents the reduction that occurs?

(1) $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

(2) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

(3) $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

(4) $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

2. Given the reaction:

$Ca(s) + Cu^{2+}(aq) \rightarrow Ca^{2+}(aq) + Cu(s)$

What is the correct reduction half-reaction?

(1) $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

(2) $Cu^{2+}(aq) \rightarrow Cu(s) + 2e^-$

(3) $Cu(s) + 2e^- \rightarrow Cu^{2+}(aq)$

(4) $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

3. In the reaction

$Mg + Cl_2 \rightarrow MgCl_2$,

the correct half-reaction for the oxidation that occurs is

(1) $Mg + 2e^- \rightarrow Mg^{2+}$

(2) $Cl_2 + 2e^- \rightarrow 2Cl^-$

(3) $Mg \rightarrow Mg^{2+} + 2e^-$

(4) $Cl_2 \rightarrow 2Cl^- + 2e^-$

4. Given the reaction:

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Which equation represents the correct oxidation half-reaction?

(1) $Zn(s) \rightarrow Zn^{2+} + 2e^-$

(2) $2H^+ + 2e^- \rightarrow H_2(g)$

(3) $Zn^{2+} + 2e \rightarrow Zn(s)$

(4) $2Cl^- \rightarrow Cl_2(g) + 2e^-$

5. Given the reaction:

$3Sn^{4+}(aq) + 2Cr(s) \rightarrow 3Sn^{2+}(aq) + 2Cr^{3+}(aq)$

Which half-reaction correctly represents the reduction that occurs?

(1) $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$

(2) $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$

(3) $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$

(4) $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$

6. Given the cell reaction:

$Sn(s) + Pb^{2+}(aq) \rightarrow Sn^{2+}(aq) + Pb(s)$

The reduction half-reaction for this cell is

(1) $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

(2) $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$

(3) $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

(4) $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

7. Given the reaction:

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Which half-reaction correctly shows the oxidation that occurs?

(1) $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

(2) $Fe(s) + 2e^- \rightarrow Fe^{2+}(aq)$

(3) $Cu^{2+}(aq) \rightarrow Cu(s) + 2e^-$

(4) $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

8. Given the reaction:

$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Which equation represents the oxidation that takes place?

(1) $Mg^{2+} + 2e^- \rightarrow Mg$

(2) $Mg \rightarrow Mg^{2+} + 2e^-$

(3) $Cu^{2+} + 2e^- \rightarrow Cu$

(4) $Cu \rightarrow Cu^{2+} + 2e^-$

9. Given the reaction:

$2NaCl + 2H_2O \rightarrow 2NaOH + H_2 + Cl_2$

Which electronic equation correctly represents the oxidation that occurs in this reaction?

(1) $2Na^0 \rightarrow 2Na^+ + 2e^-$

(2) $2Cl^- \rightarrow Cl_2^0 + 2e^-$

(3) $2H^+ + 2e^- \rightarrow H_2^0$

(4) $O_2^0 + 2e^- \rightarrow 2O^{-2}$

10. Given the reaction:

$2Al^0(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni^0(s)$

What is the total number of moles of electrons lost by 2 moles of $Al^0(s)$?

(1) 6

(2) 2

(3) 3

(4) 8

11. How many moles of electrons would be required to completely reduce 1.5 moles of Al^{3+} to Al?

(1) 0.50

(2) 1.5

(3) 3.0

(4) 4.5

Redox Practice Name:

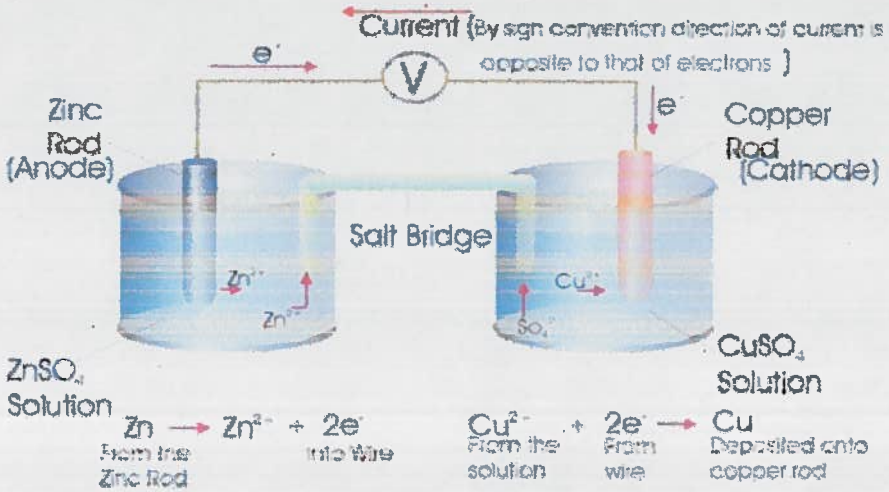
- Put electrons in the half reaction to balance it.
- State if the reaction is oxidation or reduction.

1. $Mn^{+6} \rightarrow Mn^{+4}$
2. $Cl^{-1} \rightarrow Cl^{+5}$
3. $Fe \rightarrow Fe^{3+}$
4. $Fe^{3+} \rightarrow Fe^{2+}$
5. $Cl_2 \rightarrow 2Cl^-$
6. $H_2 \rightarrow 2H^+$
7. $2Br^{+3} \rightarrow Br_2$
8. $Al \rightarrow Al^{3+}$
9. $N^{+5} \rightarrow N^{-3}$
10. $S^{-2} \rightarrow S^{+4}$
11. $C \rightarrow C^{+4}$
12. $N_2 \rightarrow 2N^{+3}$

Unit 12: Redox
12.4: Electrochemical Cells

AIM: _____

- ***Electrochemical Cells***
  - Redox
    - Chemical reaction
    - Exchange of _____ between what is oxidized and what is reduced
  - _____ = involves a chemical reaction and a flow of electrons
  - Two types
    - Voltaic cells: _____
      _____
    - Electrolytic cell: _____
      _____
  - Electrodes: _____
    - Site_____
  - ***How do the parts of an electrochemical cell work?***
    - ***Anode- The site where _____ occurs and electrons are _____.***
    - ***Cathode- The site where _____ occurs and electrons are _____.***
    - ***Salt bridge-Allows _____ from one cell to another.***
    - ***Wire- Acts as an external conductor by connecting two half cells***
      (_____).
    - _____- ***Place where half reactions occur.***
    - _____- ***Measures the current potential of a cell.***
    - ***As the current _____***

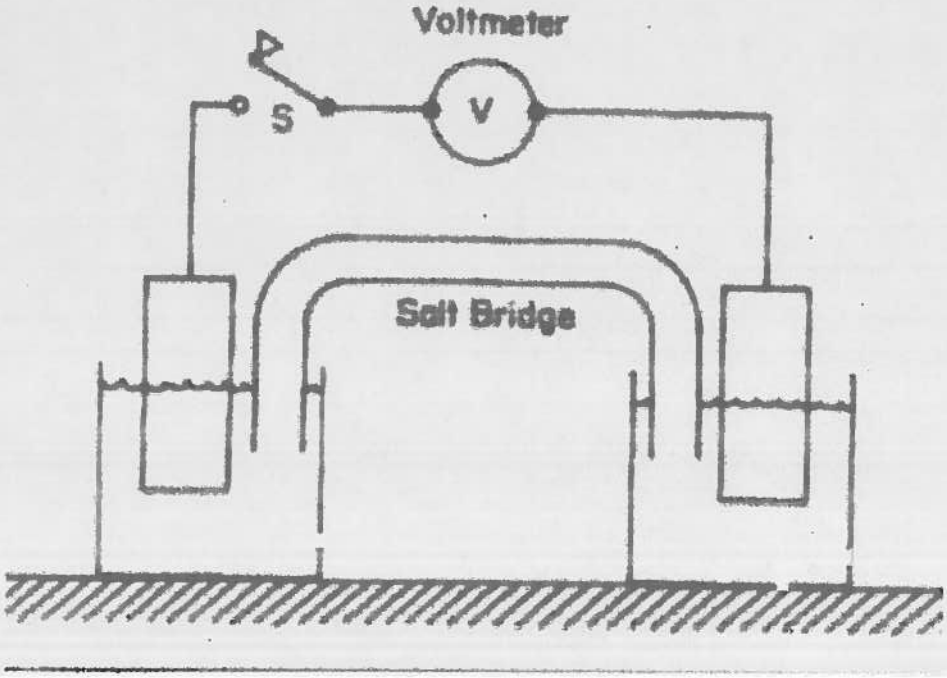


- Electrons lost during _____ travel through wire to what is being _____
  - Oxidized = _____ electrons
  - Reduced = _____ electrons

- ***RED CAT and AN OX***
  - _____ occurs at the _____
  - _____ is the site of _____

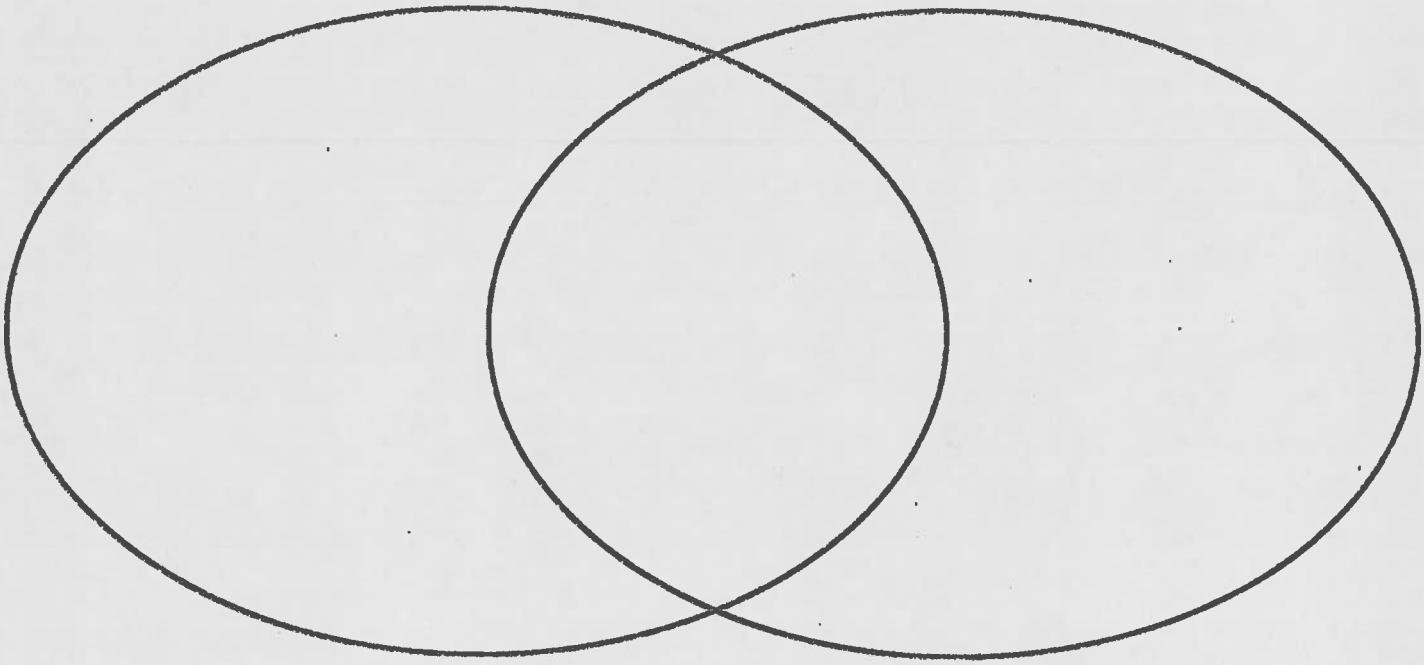
- ***TABLE J***
  - Identify cathode and anode
  - the metal that is _____
  - The metal that is _____
    - Ions are _____
    - Electrons flow from _____
  - $Zn^0 (s) + Pb^{2+} (aq) \rightarrow Zn^{2+} (aq) + Pb^0 (s)$
    - Identify the anode, cathode and the flow of electrons

  - $Zn^0 (s) + Pb^{2+} (aq) \rightarrow Zn^{2+} (aq) + Pb^0 (s)$
    - Label the picture correctly



_____

- ***Electrolytic Cell***
  - _____ reactions
  - Needs an _____ placed into the circuit to force the electrons to flow from the _____ (does not usually occur)
  - Electrolysis: _____ is used to force a chemical reaction to occur
  - What is the purpose of electrolysis?
    - Is used to separate _____
  - How do the parts of an electrolytic reaction work?
    - An electrolytic cell uses a _____



Regents Questions

- The function of a salt bridge in a voltaic cell is to

1) Allow the flow of electrons
2) Allow the flow of protons
3) Allow the flow of ions
4) Provide a site for electron transfer

- Which reaction occurs at the anode in voltaic and electrolytic cells?

1) Reduction only
2) Oxidation only
3) Both reduction and oxidation
4) Neither reduction nor oxidation

- The overall reaction in an electrochemical cell is $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}$. As the reaction in this cell takes place,

1) Oxidation occurs at the cathode
2) The Cu^{2+} is oxidized
3) The concentration of Zn^{2+} increases
4) The concentration of Cu^{2+} increases

- Consider the following equation

$$2Cr(s) + 3Cu^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cu(s)$$

Which reaction occurs at the cathode in this voltaic cell?

(1) Reduction of $Cu^{2+}(aq)$
(2) Reduction of $Cu(s)$
(3) Oxidation of $Cr^{3+}(aq)$
(4) Oxidation of $Cr(s)$

- Which of the following occurs in an electrolytic cell?

1) A chemical reaction produces an electric current
2) An electric current produces a chemical reaction
3) An oxidation reactions takes place at the cathode
4) A reduction reaction takes place at the anode

- Voltaic cells differ from electrolytic cells because in a voltaic cell

1) The cathode is positively charged
2) The anode is positively charged
3) Electrons flow from anode to cathode
4) Electrons flow from cathode to anode

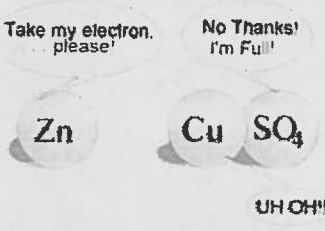
Chemistry: Form WS10.2.1A

Name _____

REDOX AND ELECTROCHEMISTRY

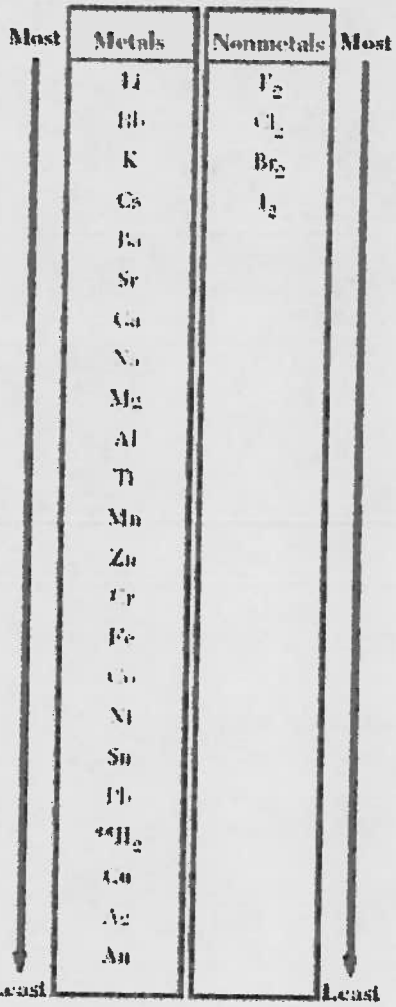
Date _____ Period _____

Applying the Activity Series



During a single replacement reaction, one element takes the place of another in a compound. Many compounds, such as the copper II sulfate, consist of two parts, a metal (copper) and a nonmetal (sulfate). When a metal such as zinc is dropped into a solution containing copper II sulfate, its natural tendency is to combine with the sulfate by giving electrons to it. The sulfate's outer shell is already full, however, because it has already gained electrons from the copper. As a result, however, the copper has room for zinc's electrons. If zinc can force copper to take its electrons, zinc can become a cation and take copper's place in the compound. Whether or not the zinc can take the copper's place depends upon which metal has the greater tendency to lose electrons. Scientists have determined by experimentation which metals can replace each other in aqueous solution. This resulted in the development of the *Activity Series* as shown in Chart J to the right. The most active metals and nonmetals are shown toward the top of the chart. Elements at the top of the activity series can replace those below them.

Table J
Activity Series**



**Activity Series based on hydrogen standard

For each example below, if a reaction will occur based on the elements' positions in the *Activity Series*, complete the equation and balance it. If there is no reaction, write no reaction. [*NOTE:* for metals, the format for single replacement reactions is AB + C → CB + A; for nonmetals the format is AB + D → AD + B]

1. $Mg(s) + HCl(aq) \rightarrow$ _____
2. $Ag(s) + Cu(NO_3)_2(aq) \rightarrow$ _____
3. $Zn(s) + Mn(CH_3COO)_2(aq) \rightarrow$ _____
4. $Al(s) + HCl(aq) \rightarrow$ _____
5. $Cu(s) + HBr(aq) \rightarrow$ _____
6. $Cu(s) + AgCH_3COO(aq) \rightarrow$ _____
7. $Sn(s) + H_2SO_4(aq) \rightarrow$ _____
8. $Mg(s) + Pb(NO_3)_2(aq) \rightarrow$ _____
9. $Pb(s) + AuCl(aq) \rightarrow$ _____
10. $Au(s) + LiCl(aq) \rightarrow$ _____

Answer the questions below by circling the number of the correct response

1. Which reaction will take place in a 1.0 molar aqueous solution?
   1. $Cu + Ag^{+} \rightarrow$
   2. $Ag + Mn^{+2} \rightarrow$
   3. $Co + Zn^{+2} \rightarrow$
   4. $Sn + Fe^{+2} \rightarrow$

2. Which reaction occurs at the positive electrode during the electrolysis of molten sodium chloride?
   1. chloride ions are reduced
   2. sodium ions are reduced
   3. chloride ions are oxidized
   4. sodium ions are oxidized

3. Strips of zinc are placed in solutions of the salts listed below. In which solution will a redox reaction take place?
   1. $Ca(NO_3)_2$
   2. $Mg(NO_3)_2$
   3. $Ni(NO_3)_2$
   4. $Sr(NO_3)_2$

4. When the reaction of a chemical cell reaches equilibrium, the potential difference of the cell
   1. decreases
   2. increases
   3. remains the same

5. When electroplating with silver, the mass of the positive electrode
   (1) decreases (2) increases (3) remains the same

6. When electroplating with silver, the mass of the negative electrode
   (1) decreases (2) increases (3) remains the same

7. Which of the following half cells is used as the standard?
   1. $F_2 + 2e^{-} = 2F^{-}$
   2. $Li^{+} + e^{-} = Li(s)$
   3. $2H^{+} + 2e^{-} = H_2$
   4. $Ag^{+} + e^{-} = Ag$

8. Oxygen and copper are produced during the electrolysis of a $CuSO_4$ solution. Which reaction occurs at the negative electrode?
   1. the copper atom is oxidized
   2. the copper ion is reduced
   3. the oxygen atom is oxidized
   4. the oxygen ion is reduced

9. Oxidation will occur in the Ni, Ni^{2+}(1 M) half-cell when it forms a cell with
   1. Al, Al^{+3} (1 M)
   2. Au, Au^{+3} (1 M)
   3. Sr, Sr^{+2} (1 M)
   4. Zn, Zn^{+2} (1 M)

10. In the electrolysis of fused $CaCl_2$, the species that reacts at the negative electrode is (1) Ca (2) Ca^{+2} (3) Cl_2 (4) Cl^{-}

Name _____ Date _____ Class _____

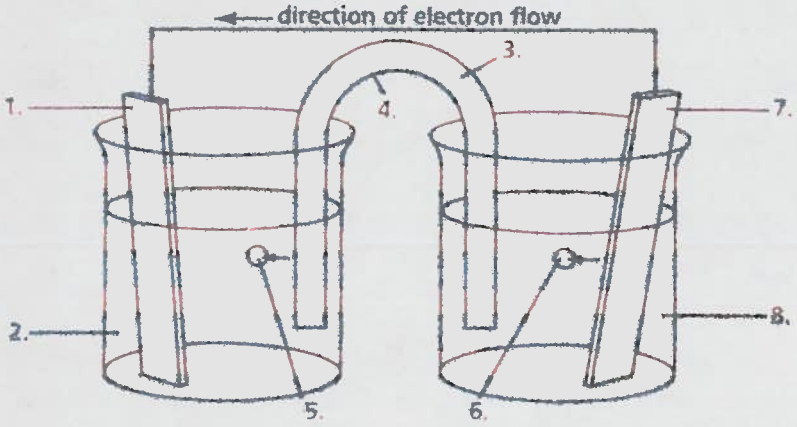
REVIEW ACTIVITY

Labelling a Galvanic Cell

The diagram below represents a Daniell galvanic cell involving zinc and copper.

Match each numbered item in the diagram with the letter of the correct label from the list below. Write your answers in the space provided.

A. Zn^{2+} solution
B. salt bridge
C. cathode
D. zinc cation
E. anode
F. U-tube
G. copper cation
H. Cu^{2+} solution



1. _____
2. _____
3. _____
4. _____
5. _____
6. _____
7. _____
8. _____

9. a. Write the equation for the half-reaction that occurs at the cathode.

9. a. _____

b. Is this an example of oxidation or of reduction?

b. _____

10. a. Write the equation for the half-reaction that occurs at the anode.

10. a. _____

b. Is this an example of oxidation or of reduction?

b. _____

11. Write the overall ionic equation.

11. _____

12. Write the overall molecular equation.

12. _____

Name _____ Date _____ Class _____

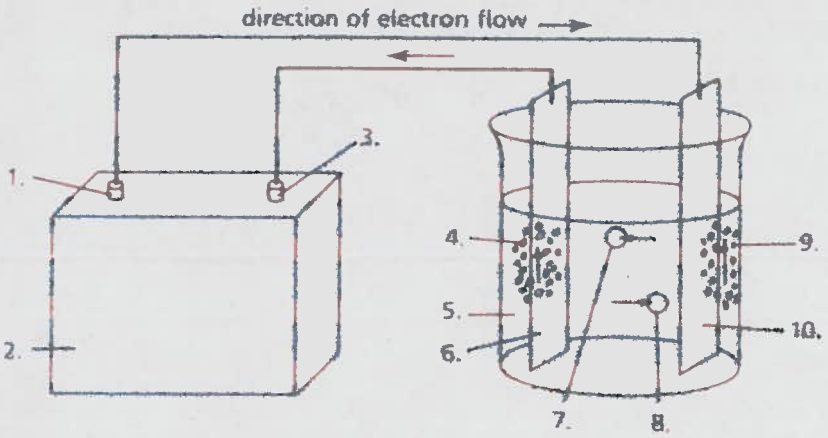
REVIEW ACTIVITY

Labelling an Electrolytic Cell

The diagram below represents the electrolysis of a sodium chloride solution (brine).

Match each numbered item in the diagram with the letter of the correct label from the list below.

A. sodium ion
B. cathode
C. electrolyte
D. hydrogen gas
E. chloride ion
F. positive terminal
G. battery
H. chlorine gas
I. anode
J. negative terminal



1. _____
2. _____
3. _____
4. _____
5. _____
6. _____
7. _____
8. _____
9. _____
10. _____

11. a. Write the equation for the half-reaction that occurs at the cathode. 11. a. _____

b. Is this an example of oxidation or of reduction? b. _____

12. a. Write the equation for the half-reaction that occurs at the anode. 12. a. _____

b. Is this an example of oxidation or of reduction? b. _____

13. Write the overall ionic equation. 13. _____

14. Write the overall molecular equation. 14. _____

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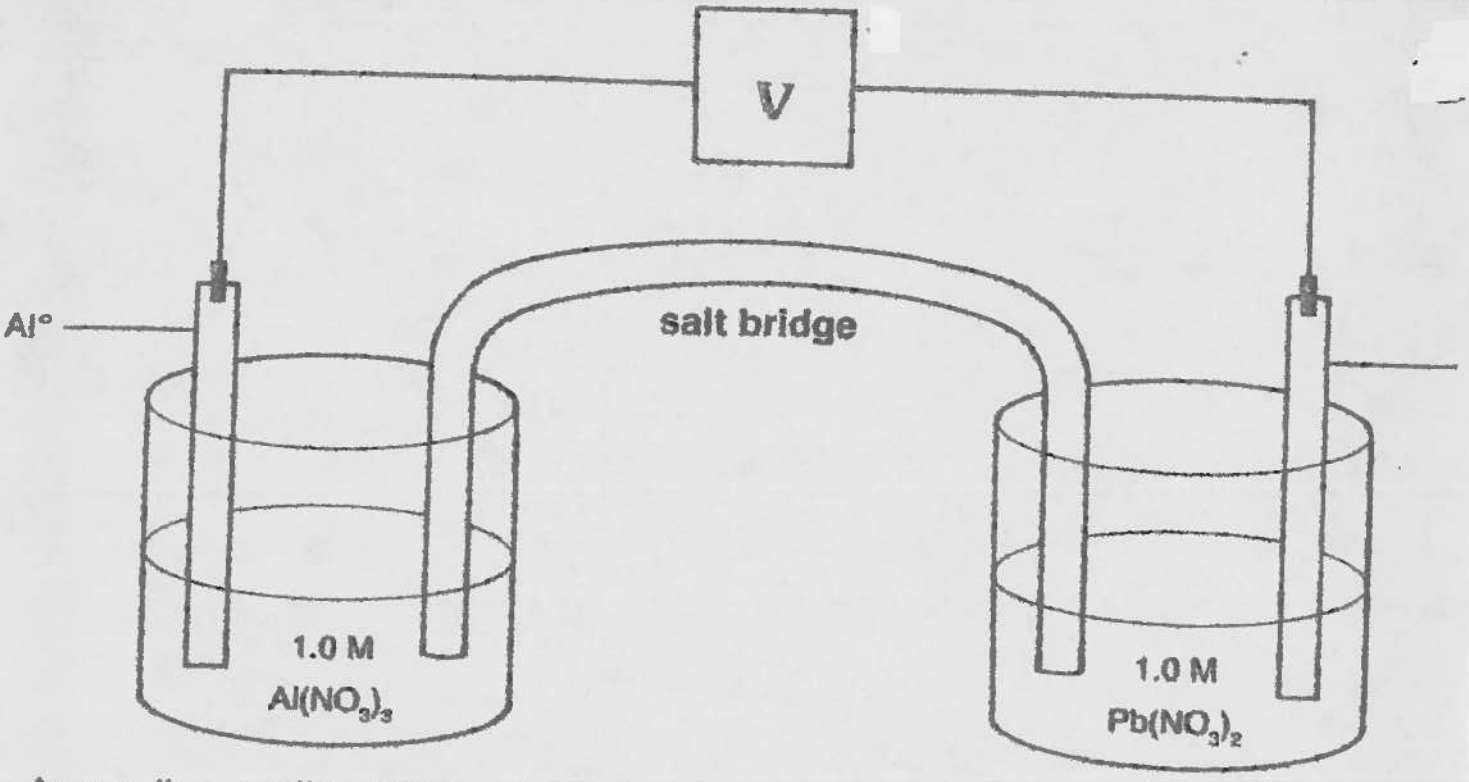
THE ELECTROCHEMICAL CELL

Name _____



Answer the questions below referring to the above diagram and a Table of Standard Electrode Potentials.

1. Which is more easily oxidized, metal, aluminum or lead? _____
2. What is the balanced equation showing the spontaneous reaction that occurs?

   _____
3. What is the maximum voltage that the above cell can produce? _____
4. What is the direction of electron flow in the wire? _____
5. What is the direction of positive ion flow in the salt bridge? _____
6. Which electrode is decreasing in size? _____
7. Which electrode is increasing in size? _____
8. What is happening to the concentration of aluminum ions? _____
9. What is happening to the concentration of lead ions? _____
10. What is the voltage in this cell when the reaction reaches equilibrium? _____
11. Which is the anode? _____
12. Which is the cathode? _____
13. What is the positive electrode? _____
14. What is the negative electrode? _____

Name _____

REDOX AND ELECTROCHEMISTRY Date _____ Period _____

Activity and Electricity

Aim

- describe an electrochemical cell
- describe voltaic cells and electrolytic cells

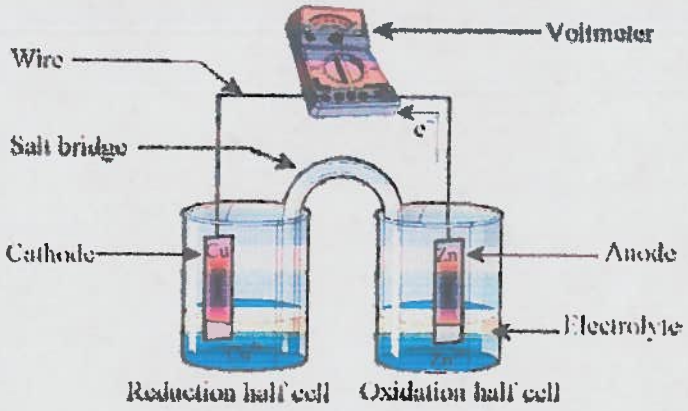
Notes

Electrochemical cells

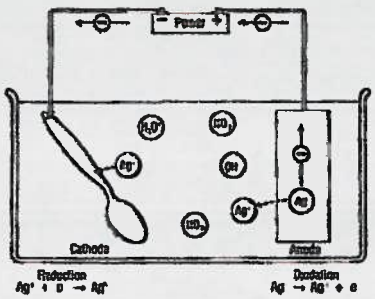
- ☆ Functioning of the electrochemical cell
  - ☆ During a single replacement reaction, more active metals transfer electrons to less active metals
    - ☆ the more active metal is oxidized
    - ☆ the less active metal is reduced
  - ☆ If the oxidation and reduction half reactions are physically separated and attached by a wire, electrons will flow through the wire during the reaction
- ☆ Parts of an electrochemical cell
  - ☆ electrodes
    - ☆ anode — place where oxidation occurs
    - ☆ cathode — place where reduction occurs
  - ☆ half cells — separate containers in which oxidation and reduction half reactions occur

The Electrode Zoo

AN OX – ANode = OXidation
RED CAT – CAThode = REDuction



  - ☆ U-tube or salt bridge — lets ions travel between half cells to complete the circuit
- ☆ Examples of electrochemical cells
  - ☆ Voltaic Cells (Spontaneous Reactions)
    - ☆ Definition — a system that uses a chemical reaction to produce electricity
    - ☆ Examples
      - ☆ lead acid storage battery (automobile battery)
      - ☆ dry cell (zinc container anode, carbon center post cathode)
  - ☆ Electrolytic cells (Nonspontaneous Reactions)
    - ☆ Definition — a system that uses electricity to cause a chemical reaction
    - ☆ Examples
      - ☆ recharging a car battery:
        $2PbSO_4 + 2H_2O \rightarrow PbO_2 + Pb + 2H_2SO_4$
      - ☆ electrolysis of molten sodium chloride
        $2NaCl \rightarrow 2Na^0 + Cl_2^0$
      - ☆ electroplating



45

Chemistry: _____

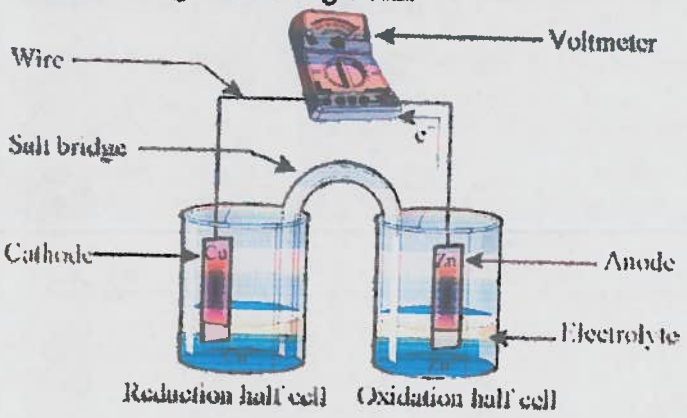
Name _____

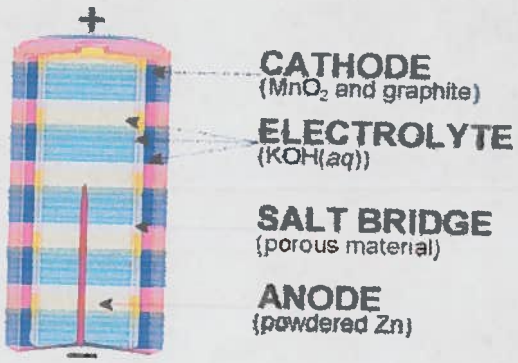
REDOX AND ELECTROCHEMISTRY

Date _____ Period _____

A Salt and Battery

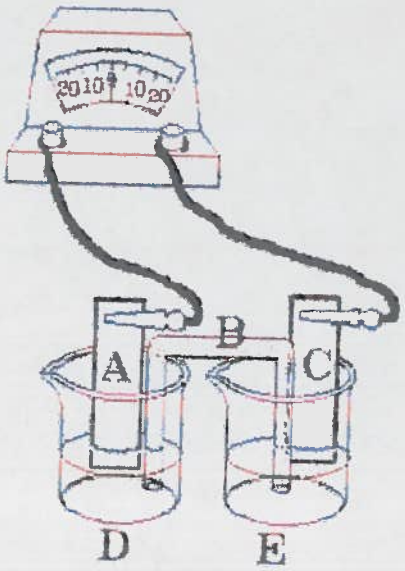
Portable electronic devices run on batteries. The electricity generated by a battery comes from a chemical reaction known as an oxidation-reduction reaction. During an a single replacement, a type of oxidation-reduction reaction, more active metals transfer electrons to less active metals. As a result, the more active metal is oxidized, and the less active metal is reduced. If the oxidation and reduction half reactions are physically separated and attached by a wire, electrons will flow through the wire during the reaction and can be used to power our portable electronics. This is done by putting electrolytes, usually aqueous acids, bases, or salts, into separate containers. The separate containers are called half cells because the half reactions are isolated in them. They are connected by a salt bridge which lets ions travel between half cells. Electrodes are immersed into the electrolytes. The electrodes are merely metals with differing activity. Completing the circuit by connecting the electrodes enables electrons to flow from the more active metal to the less active metal, reducing it. The electrode where reduction occurs is called the **cathode**. The electrode where oxidation occurs is called the **anode**. The device that produces electric current from a chemical reaction is called a **voltaic cell**. Several voltaic cells attached together form a battery of cells. A **battery**, produces a higher voltage than a single cell.





Answer the questions below based on your reading above and on your knowledge of chemistry.

Answer questions1-4 by referring to the diagram to the right showing an electrochemical cell. The metal at electrode A is silver. The metal at electrode C is lead. The electrolytes at locations B, D, and E are potassium nitrate, silver nitrate, and lead nitrate respectively.



1. **In what direction do electrons flow in the electrochemical cell pictured to the right (A to C *or* C to A)?** _____
2. **What type of chemical change is taking place in the half-cell contained in the beaker at location *E*?** _____
3. **At which location are electrons being gained?** _____
4. **Which metal is being replaced during the reaction in this electrochemical cell?**

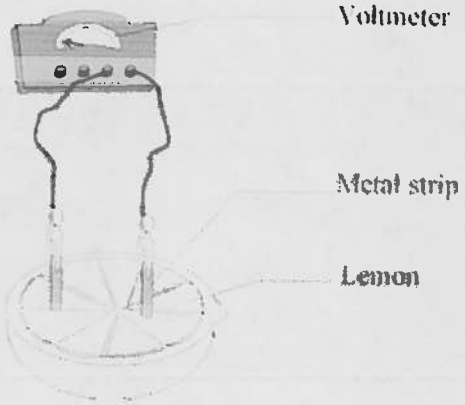
_____

Continue ☞

Answer questions 5-16 by referring to Table J. For each of the electrode pairs, which would be the anode in an electrochemical cell?

5. Cu/Zn ..... _____
6. Pb/Sn ..... _____
7. K/Al ..... _____
8. Ba/Li ..... _____
9. Au/Pb ..... _____
10. Mn/Zn ..... _____
11. Fe/Zn ..... _____
12. Co/Ca ..... _____
13. Co/Ni ..... _____
14. H_2/Ag ..... _____
15. Cu/Mg ..... _____
16. Zn/Al ..... _____

Answer questions17-19 by referring to t he setup shown to the right using a lemon and metal strips. It actually produces measurable electricity.



17. Explain how the lemon battery works? _____

_____

_____

_____

_____

_____

18. What parts of a typical voltaic cell are missing in the lemon battery? What effect does this have on how well it functions? Explain. _____

_____

_____

19. If the metal strip on the right is iron and the metal strip on the left is aluminum, in what direction will electricity flow?

_____

20. What happens at the anode of an electrochemical cell? _____

_____

21. There are two voltaic cells pictured on the previous page. The one on the left is called a wet cell, while the one at the left is called a dry cell. The one at the right is also called an alkaline cell. What is the difference between these cells that accounts for the difference in the way they are named? _____

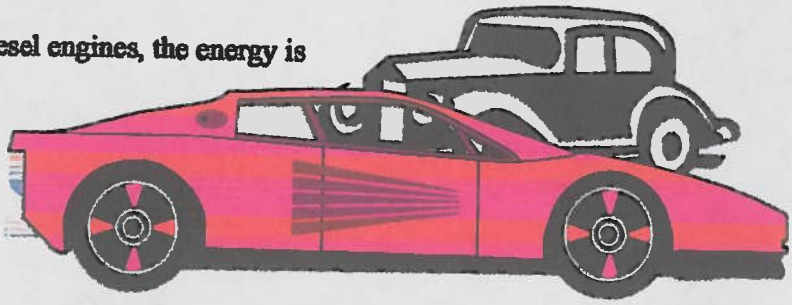
_____

_____

Chemistry: Name _____

REDOX AND ELECTROCHEMISTRY Date _____ Period _____

Forcing Electrons to Move



The energy to run most cars from gasoline. Except in diesel engines, the energy is released from the gasoline by exploding it with a tiny spark from a spark plug. The energy to make the spark comes from the car's battery. The battery in the car is called a "wet cell." It contains sulfuric acid [$H_2SO_4(aq)$], a liquid electrolyte. The electricity is generated by the following chemical reaction:

$$PbO_2 + Pb + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$$

Car batteries can last for several years. This is because they get recharged. As the engine spins, a moving magnet in the alternator pushes electrons in a direction opposite to the way they normally flow from the battery. These electrons reverse the chemical reaction that generated electricity in the battery. A cell that uses electricity to produce a chemical reaction in this way is called an electrolytic cell. When the car battery is generating electricity it is an electrochemical cell. When it is being recharged, it is an electrolytic cell.

Answer the questions below based on the reading above and on your knowledge of chemistry.

1. Write the chemical reaction that occurs when a car battery generates electricity. _____

   a. Write the half reactions: _____

   b. What is oxidized, and what is reduced? _____

2. Write the chemical reaction that occurs when a car battery is recharged. _____

   a. Write the half reactions: _____

   b. What is oxidized, and what is reduced? _____

3. Aluminum is found in the mineral bauxite (Al_2O_3). To get pure aluminum, the aluminum needs to be separated from oxygen.

   a. Imagine bauxite forms by the following reaction: $4Al + 3O_2 \rightarrow 2Al_2O_3$. Write the half reactions. _____

   b. During the formation of bauxite from its elements, what is oxidized, and what is reduced? Does this make sense considering that aluminum is a metal? Explain. _____

   c. Write the reaction for the purification of aluminum from bauxite. _____

☞ Continue on the next page.

d. Write the half reactions for the purification of aluminum. During the purification, what is oxidized, and what is reduced? _____

e. Considering that aluminum is a metal, suggest a method to purify it. Explain. _____

4. Iron is often protected from rusting by a process called galvanizing. When a metal is galvanized, it is coated with zinc. One way to coat iron with zinc is through a single replacement reaction: $Fe + Zn(NO_3)_2 \rightarrow Fe(NO_3)_3 + Zn$. Since the reaction occurs at the surface of the iron, the iron becomes plated with zinc.

a. Write the half reactions for this reaction. What was oxidized, and what was reduced? _____

b. Consult the activity series on Chart J. How likely is this reaction to occur? Explain. _____

c. Suggest a method to plate iron with zinc. _____

5. What is an electrolytic cell? What are some of its functions? _____

6. What type of cell is represented by the following reaction: $Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$? Write the half reactions associated with it. Identify the oxidation and reduction half reactions. _____

Power Webquest

Directions: Go to this website. Click on the links to find the answers to each of the questions that are listed below and on the website. Work on your own. You may turn in work by

a. Writing answers on this sheet or
b. Making a word document and emailing it to me at amendoza@wbschools.org.
c. Put your name and document title (Power Webquest) in the Subject of the email.
d. Do NOT cut and paste. Responses MUST BE IN YOUR OWN WORDS to receive credit.
e. Be sure to answer ALL parts of each question.
f. If a site does not work use www.google.com.

http://wayback.archive-it.org/2118/20100924225911/http://64.251.202.97/EducationalServices/webquest/power.htm

1. How do batteries work? Who invented the modern battery? How do the batteries made by this inventor differ from the ones used today?

2. What are semiconductors, and how are they used in harnessing solar power?

3. Thomas Midgley and Eugene Houdry were both chemists who found ways to make more powerful gasoline. What were their two methods? How did the methods differ? Which method was safer environmentally?

4. What is the greenhouse effect? What energy-producing chemical reactions might contribute to it? What is the carbon cycle and what does it have to do with energy and the greenhouse effect?

50

5. What is a fuel cell and how does it work? What kinds of fuels are normally used in fuel cells? How are these fuels produced? Which production methods would you say are the safest environmentally?

6. What are some alternative fuels besides gasoline? What chemical reactions do they undergo when burned? Compare the combustion products of these fuels with those of gasoline in terms of environmental safety.

7. What are nuclear fission and nuclear fusion? How can they provide energy? What are the advantages and drawbacks of each?

8. What are nuclear fission and nuclear fusion? How can they provide energy? What are the advantages and drawbacks of each?

Voltaic Cells

Some redox reactions are spontaneous

Some redox reactions just happen when you mix the reactants. For example, if you place a piece of tin into a solution containing gold ions (Au^{3+}), gold metal forms on the tin.

$$3Sn(s) + 2Au^{3+}(aq) \rightarrow 3Sn^{2+}(aq) + 2Au(s)$$

This is because the gold ions have been reduced, and the tin has been oxidized.
This happens spontaneously because tin oxidizes (loses electrons) more easily than gold.

Voltaic cells use spontaneous redox reactions to make electricity

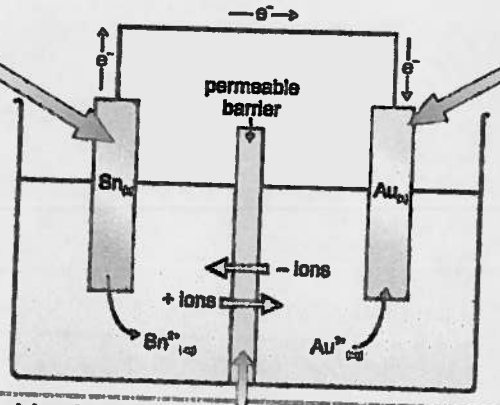
1) Redox reactions involve the transfer of electrons.
2) A voltaic cell makes these electrons flow through a wire — it converts stored chemical energy to electrical energy.
3) The two half-reactions are separated, so that they occur in two half-cells linked together.
4) Each half-cell consists of a solid electrode submerged in a solution.

Oxidation at the Anode

At this electrode (the anode), tin is oxidized:

$Sn \rightarrow Sn^{2+} + 2e^-$

The Sn^{2+} ions go into the solution and the electrons travel through the wire to the cathode.



Reduction at the Cathode

At this electrode (the cathode), gold is reduced:

$Au^{3+} + 3e^- \rightarrow Au$

The electrons from the wire attach to the Au^{3+} ions in the solution. Solid gold is deposited on the cathode.

The reaction releases positive ions on one side, and removes them from the other. This would make the solution at the anode positively charged, and the solution at the cathode negatively charged, causing the reaction to stop.
The permeable barrier allows ions to move between the half-cells to even out the charges.

You can identify the anode and cathode using the activity series

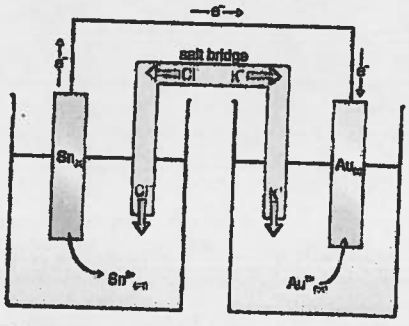
Table J of the Reference Tables.

Here's how to figure out which metal is the cathode and which metal is the anode in a voltaic cell: First find where the two metals involved in the cell are in the activity series. The more reactive metal will be oxidized — so it's the anode. The other metal will be the cathode.

So, for a nickel and zinc cell, zinc is more reactive and so will be oxidized at the anode. This means nickel must be the cathode.

A salt bridge can be used instead of a barrier

1) The half-cells can be separated completely, and a salt bridge used to allow ions to flow between them.
2) The salt bridge is made of an electrolyte (see p.85), like KCl, dispersed through a gel. Ions can flow through it.
3) The diagram shows how the Cl^- ions flow out of the bridge to balance the positive ions being produced on the left side. K^+ ions flow into the other cell to balance the charges there.

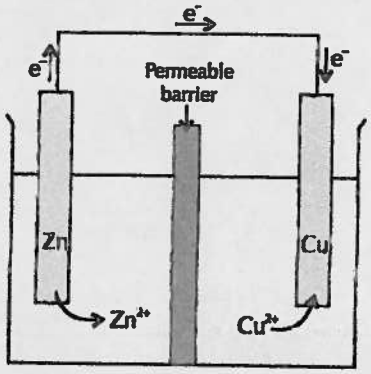


Voltaic cells — a battery by any other name doth smell as sweet...

Another complicated page — just keep your head, and read it carefully. It will make sense in the end, I promise. Then try covering the page and scribbling down what you've learned.

Voltaic Cells

Q1 The diagram below shows a voltaic cell.



a) Is zinc being oxidized or reduced? Explain your answer.

b) Write down the half-equation for Zn becoming Zn^{2+}.

c) Write down the half-equation for Cu^{2+} becoming Cu.

d) Why is a permeable barrier used instead of an impermeable barrier?

e) Which substance loses electrons more easily — copper or zinc?

Q2 The diagram below shows two half-cells connected by a salt bridge.

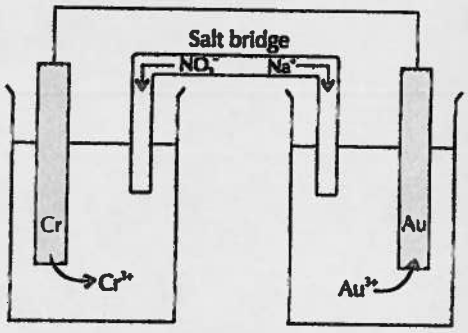
a) What is the purpose of the salt bridge?

b) What electrolyte is being used in the salt bridge in the diagram?

c) On a copy of the diagram, show the direction of the current through the wire.

d) On your diagram, label the anode and the cathode.

e) Why will the chromium electrode need replacing regularly?



Q3 The equation below shows the reaction in a voltaic cell.

$$2Ag^{+}(aq) + Mg(s) \rightarrow 2Ag(s) + Mg^{2+}(aq)$$

a) Is the anode made from silver or magnesium?

b) Is magnesium being oxidized or reduced in this reaction?

c) Write down a pair of half-equations for the reaction.

d) In which direction will the electrons flow through the wire?

Q4 Identify the anode and the cathode when the following pairs of electrodes are used.

a) lead and zinc

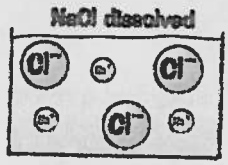
b) copper and chromium

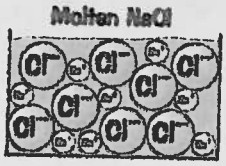
c) nickel and cobalt

Electrolysis and the Half-Equations

Electrolysis means "Splitting Up with Electricity"

1) Electrolysis is the breaking down (decomposition) of a substance using electricity.
2) It requires an electrolyte, a liquid that will conduct electricity. Electrolytes are usually free ions dissolved in water, such as dilute acids like HCl, or dissolved salts, like NaCl solution.
3) Electrolytes can also be molten ionic substances, but this involves higher temperatures. In either case, it's the free ions that conduct the electricity, and allow the whole thing to work.
4) The electrical supply acts like an electron pump, taking electrons away from the positive anode and supplying them to the negative cathode. Ions gain or lose electrons at the electrodes and neutral atoms and molecules are formed.



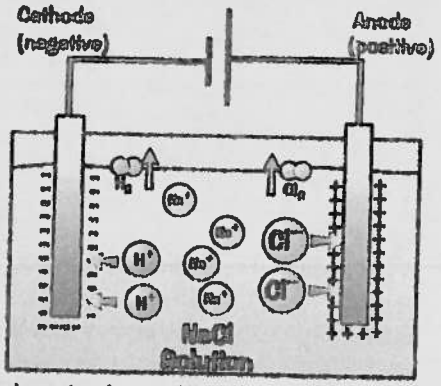


1 Metals will always be produced at the cathode (where reduction takes place) because metals form positive ions.

Positive ions are called CATIONS because they're attracted to the negative cathode.

Hydrogen is also produced at the negative cathode.

EXAMPLE: Electrolysis of NaCl



One hydrogen ion plus one electron becomes one hydrogen atom.

One chloride ion becomes one chlorine atom and an electron.

In this solution, hydrogen gas forms (from H⁺ ions in the water), rather than sodium metal — because sodium is too reactive to form.

2 ALL nonmetals (except hydrogen) have negative ions and so they'll be produced at the positive anode (where oxidation happens).

Negative ions are called ANIONS because they're attracted to the anode.

Electrolytic and voltaic cells are different types of electrochemical cells...
Both voltaic cells and electrolytic cells involve redox reactions, but the reactions in electrolysis are not spontaneous. That's why they need electrical energy to make them go. Electrolytic cells consume electricity, while voltaic cells produce it.

Half-Equations are used to show what happens at each electrode

The reaction at each electrode can be represented by a half-equation.
Remember that the number of electrons in the two equations must balance.

Cathode: $H^+(aq) + e^- \rightarrow H$

Anode: $Cl^-(aq) \rightarrow Cl + e^-$

These equations aren't finished because both the hydrogen and the chlorine come off as diatomic gases. They must be rewritten as H_2 and Cl_2:

Cathode: $2H^+ + 2e^- \rightarrow H_2(g)$

Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

There are two electrons in each equation, so it all balances.

You can combine the two balanced half-equations to give the full equation for the electrolysis:

$$2H^+(aq) + 2Cl^-(aq) \rightarrow H_2(g) + Cl_2(g)$$

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Electrolysis and the Half-Equations

Q1 Answer these questions on electrolysis:

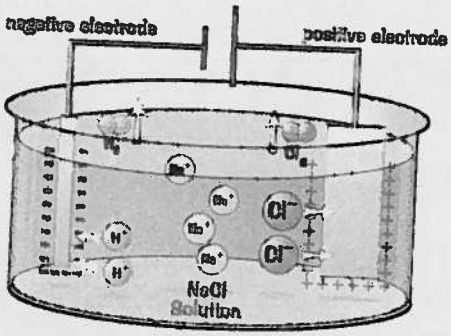
a) What is another name for the positive electrode?

b) What is another name for the negative electrode?

c) Which electrode are cations attracted to?

d) Which electrode are anions attracted to?

Q2 The diagram below shows electrolysis in action.



a) What is the name of the electrolyte in this electrolysis?

b) Explain why solid sodium chloride would not be suitable as an electrolyte.

c) Why does sodium metal not form at the cathode?

Q3 Do electrons flow through the wire from the anode to the cathode or from the cathode to the anode?

Q4 Write down the headings "Voltaic cells" and "Electrolytic cells." Put each of the following phrases under the correct heading. Some phrases may go under both headings.

redox reactions | reduction occurs at the cathode | oxidation occurs at the anode | needs electricity

spontaneous reactions | not spontaneous reactions | produces electricity

Q5 In the electrolysis of copper(II) chloride solution, a layer of copper appears on one electrode and chlorine gas is released at the other.

a) On which electrode will the copper form?

b) Write two balanced half-equations to show what happens at each electrode. (Copper(II) chloride solution contains Cu^{2+} ions.)

Q6 Which is the only nonmetallic element that is produced at the cathode?

55

Voltaic cells and electrolytic cells are not the same...

Confusing these two types of cells is a pitfall that it's all too easy to fall head first into. If you try to power your radio using an electrolytic cell you're going to hit a problem — it's voltaic cells that produce electricity. Before tackling a question on cells in the Regents exam, check what type of cell it's about.

Electrochemical (Voltaic) Cell Questions

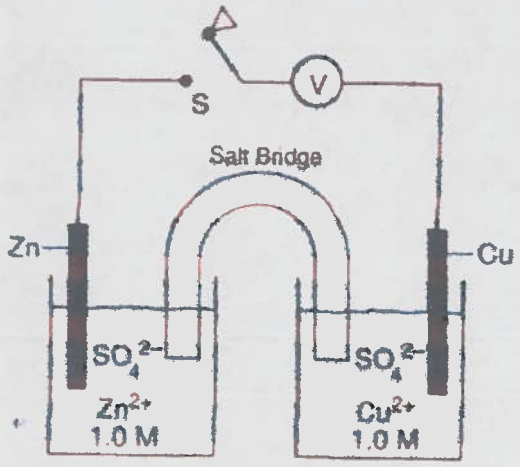
____1. In a chemical cell composed of two half-cells, ions are allowed to flow from one half-cell to another by means of
1) electrodes
2) a voltmeter
3) an external conductor
4) a salt bridge

____2. Given the overall cell reaction:

$$Zn(s) + 2\,Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2\,Ag(s)$$

Which will occur as the cell operates?
1) The amount of Zn(s) will increase.
2) The amount of Ag(s) will decrease.
3) The concentration of $Zn^{+2}(aq)$ will increase.
4) The concentration of $Ag^{+}(aq)$ will increase.

Base your answers to questions 3 and 4 on the diagram below which represents a chemical cell at 298 K and 1 atmosphere.



____3. Which species represents the cathode?
1) Zn
2) Zn^{2+}
3) Cu
4) Cu^{2+}

____4. When switch S is closed, electrons in the external circuit will flow from
1) Zn to Zn^{2+}
2) Zn to Cu
3) Cu to Zn^{2+}
4) Cu to Zn

____5. The reaction that takes place in a chemical cell is best classified as
1) fusion
2) redox
3) transmutation
4) cracking

____6. An electrochemical cell that generates electricity contains half-cells that produce
1) oxidation half-reactions, only
2) reduction half-reactions, only
3) spontaneous redox reactions
4) non-spontaneous redox reactions

____7. In a chemical cell, electrical energy will be produced when
1) only oxidation occurs
2) only reduction occurs
3) both oxidation and reduction occur
4) neither oxidation nor reduction occurs

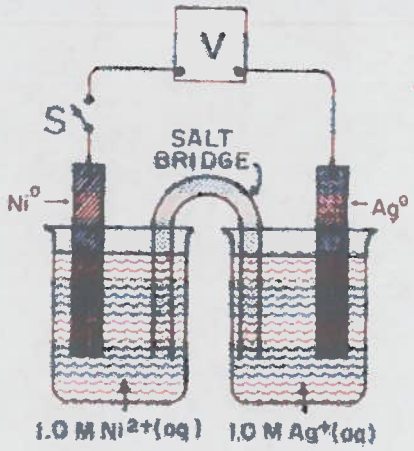
____8. A chemical cell differs from an electrolytic cell because in a chemical cell there is
1) a positive and negative electrode
2) an anode and a cathode
3) a redox reaction that produces an electric current
4) an electric current that causes a redox reaction

____9. Given the nickel oxide–cadmium reaction:

$$2\,NiO(OH)_3 + Cd + 2\,H_2O \rightarrow 2\,Ni(OH)_2 + Cd(OH)$$

During discharge, the Cd electrode
1) is oxidized
2) is reduced
3) gains electrons
4) gains mass

____10. Base your answer to the following question on the diagram of the chemical cell at 298 K and on the equation below.

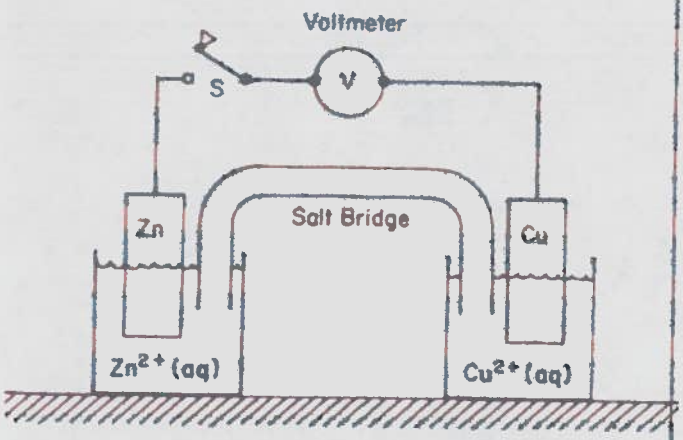


$$Ni^0(s) + 2Ag^{+}(aq) \rightarrow Ni^{2+}(aq) + 2Ag^0(s)$$

In the given reaction, the Ag^{+} ions
1) gain electrons
2) lose electrons
3) gain protons
4) lose protons

____11. In a chemical cell, the function of the salt bridge is to provide a path for the migration of
1) electrons
2) neutrons
3) molecules
4) ions

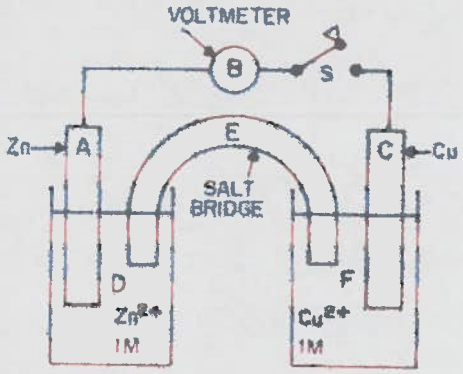
___12. The diagram below represents an electrochemical cell.



When switch *S* is closed, which particles undergo reduction?

1) Zn^{2+} ions
2) Zn atoms
3) Cu^{2+} ions
4) Cu atoms

___13. On the diagram of the chemical cell below which is at 298 K and 1 atmosphere.



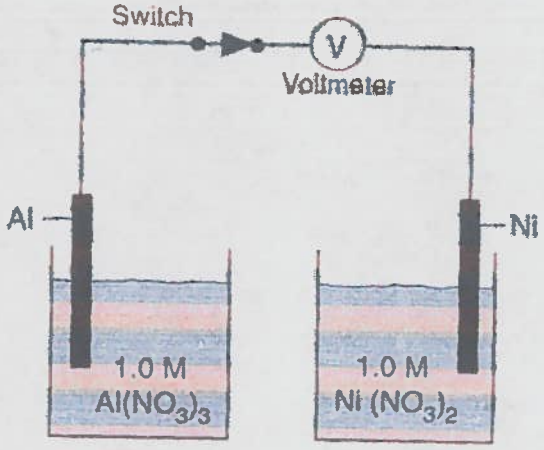
When switch S is closed, which series of letters show the path and direction of the Zn^{2+} ion flow?

1) *ABC*
2) *CBA*
3) *DEF*
4) *FED*

___14. What is the electron flow in a wire connecting the Zn and Cu electrodes of a zinc-copper chemical cell at standard conditions?

1) from negative Zn to positive Cu
2) from positive Zn to negative Cu
3) from negative Cu to positive Zn
4) from positive Cu to negative Zn

___15. The diagram below represents a chemical cell.

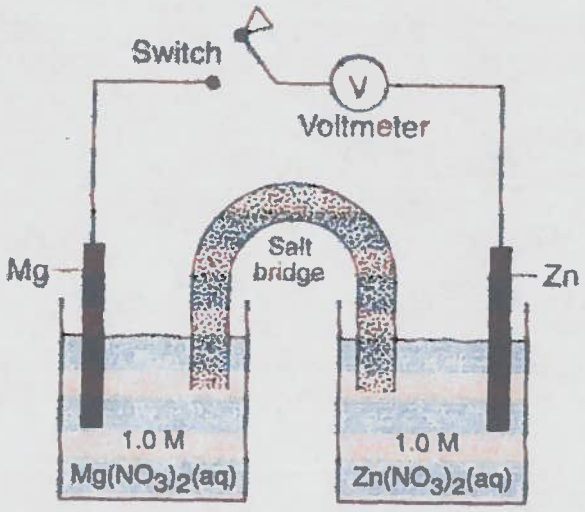


In order for the cell to operate, it should be provided with

1) a cathode
2) an anode
3) a salt bridge
4) an external path for electrons

16. Base your answers to the following questions on the information below.

A student constructed an electrochemical cell shown below. One cell contains a strip of magnesium and the other a strip of zinc. The concentration of both solutions is 1.0 mol/L.



a Write the half-reaction that will occur in the cell with the magnesium strip.

b Draw above the voltmeter an arrow indicating the direction that the electrons will flow when the switch is closed.

c As the reaction proceeds the voltage will eventually drop to zero. Explain.

d What is the purpose of the salt bridge joining each half-cell?

____1. An electrolytic cell is different from an electrochemical cell because in an electrolytic cell

1) a redox reaction occurs
2) a spontaneous reaction occurs
3) an electric current is produced
4) an electric current causes a chemical reaction

____2. In an electrolytic cell, to which electrode will a positive ion migrate and undergo reduction?

1) the anode, which is negatively charged
2) the anode, which is positively charged
3) the cathode, which is negatively charged
4) the cathode, which is positively charged

____3. Which statement best describes the reaction represented by the equation below?

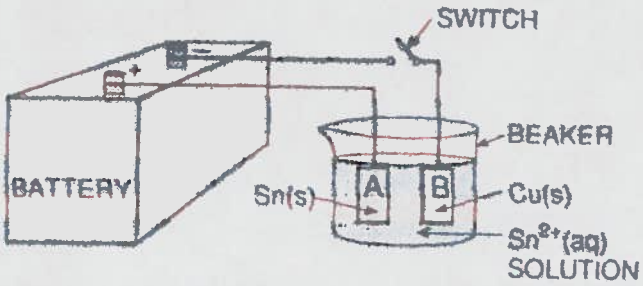
$$2\,NaCl + 2\,H_2O + \text{electricity} \rightarrow Cl_2 + H_2 + 2\,NaOH$$

1) The reaction occurs in a chemical cell and releases energy.
2) The reaction occurs in a chemical cell and absorbs energy.
3) The reaction occurs in an electrolytic cell and releases energy.
4) The reaction occurs in an electrolytic cell and absorbs energy.

____4. In an electrolytic cell, which ion would migrate through the solution to the positive electrode?

1) a hydrogen ion
2) a chloride ion
3) an ammonium ion
4) a hydronium ion

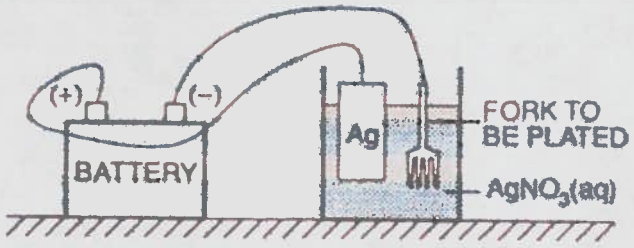
____5. Base your answer to the following question on the diagram below of an electrolytic cell in which the electrodes are tin and copper.



When the switch is closed, what will happen to the two electrodes?

1) *B* will dissolve and *A* will become coated with tin.
2) *A* will dissolve and *B* will become coated with tin.
3) *B* will dissolve and *A* will become coated with copper.
4) *A* will dissolve and *B* will become coated with copper.

Base your answers to questions 6 and 7 on the diagram below which represents the electroplating of a metal fork with Ag(s).



____6. Which part of the electroplating system is provided by the fork?

1) the anode, which is the negative electrode
2) the cathode, which is the negative electrode
3) the anode, which is the positive electrode
4) the cathode, which is the positive electrode

____7. Which equation represents the half-reaction that takes place at the fork?

1) $Ag^+ + NO_3^- \rightarrow AgNO_3$
2) $AgNO_3 \rightarrow Ag^+ + NO_3^-$
3) $Ag^+ + e^- \rightarrow Ag(s)$
4) $Ag(s) \rightarrow Ag^+ + e^-$

____8. In an electrolytic cell, oxidation takes place at the

1) anode, which is positive
2) anode, which is negative
3) cathode, which is positive
4) cathode, which is negative

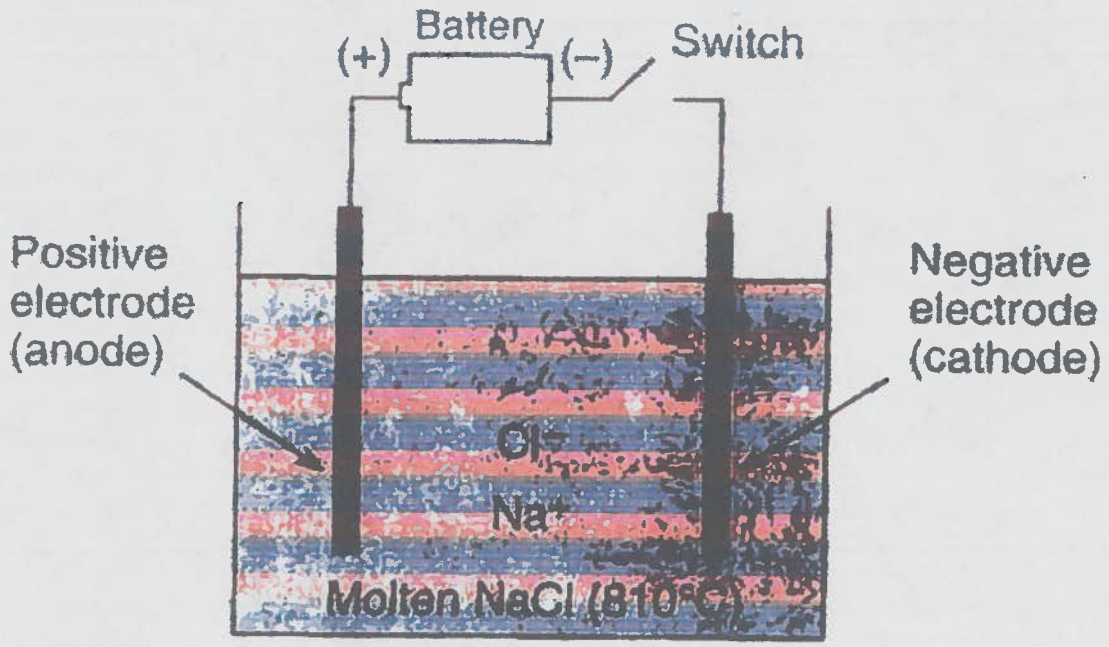
____9. In an electrolytic cell, the negative electrode is called the

1) anode, at which oxidation occurs
2) anode, at which reduction occurs
3) cathode, at which oxidation occurs
4) cathode, at which reduction occurs

____10. In an electrolytic cell, a negative ion will migrate to and undergo oxidation at the

1) anode, which is negatively charged
2) anode, which is positively charged
3) cathode, which is negatively charged
4) cathode, which is positively charged

Base your answers to questions 11 through 13 on the diagram and balanced equation below, which represent the electrolysis of molten NaCl.



$$2NaCl \longrightarrow Cl_2 + 2Na$$

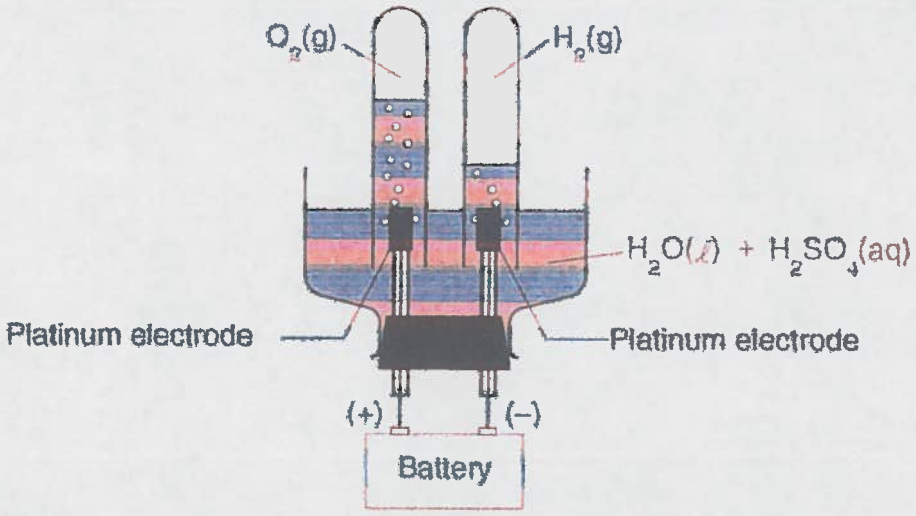
11. When the switch is closed, which electrode will attract the sodium ions?

12. What is the purpose of the battery in this electrolytic cell?

13. Write the balanced half-reaction for the reduction that occurs in this electrolytic cell.

14. Base your answer to the following question on the information and diagram below.

The apparatus shown in the diagram consists of two inert platinum electrodes immersed in water. A small amount of an electrolyte, H_2SO_4, must be added to the water for the reaction to take place. The electrodes are connected to a source that supplies electricity.



What type of electrochemical cell is shown?

Exam Question Review

1 In order for a substance to be reduced, what is required?

1 the loss of an electron
2 the gain of a proton
3 the loss of a proton
4 the gain of an electron

Note that question 2 has only 3 choices.

2 For each electron gained, the oxidation number

1 stays the same
2 decreases by 1
3 increases by 1

3 The following equation shows the reaction between iron and copper(II) sulfate, making iron(II) sulfate and pure copper: $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

a Write the balanced oxidation half-equation for this oxidation-reduction reaction. [1]

b What is the oxidation number of iron in $FeSO_4$? [1]

Base your answers to questions 4 to 7 on the following redox reaction, which occurs spontaneously in an electrochemical cell.

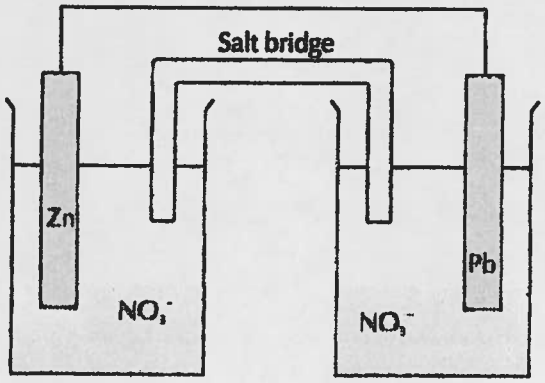
$Zn(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Zn^{2+}(aq)$

4 In this reaction, which substance gains electrons and which substance loses electrons? [1]

5 Which half-reaction happens at the anode — oxidation or reduction? [1]

6 State what happens to the number of protons in a Zn atom when it changes to Zn^{2+} as the redox reaction occurs. [1]

7 A voltaic cell can be set up as shown in the diagram below.



What is the purpose of the salt bridge? [1]

8 During the electrolysis of hydrochloric acid (HCl), hydrogen gas (H_2) and chlorine gas (Cl_2) are produced.

a Complete the following unfinished half-equations for this reaction. [2]

$H^+(aq) + e^- \rightarrow H$

$Cl^-(aq) \rightarrow Cl + e^-$

b This reaction does not happen spontaneously. How can this reaction be forced to happen? [1]

1. What is the oxidation number assigned to manganese in $KMnO_4$?

1) +7 2) +2
3) +3 4) +4

2. Given the reaction that occurs in an electrochemical cell:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

During this reaction, the oxidation number of Zn changes from

1) +2 to 0 2) 0 to -2
3) -2 to 0 4) 0 to +2

3. In which substance does chlorine have an oxidation number of +1?

1) HClO 2) HCl
3) Cl_2 4) $HClO_2$

4. What is the oxidation number of chromium in $K_2Cr_2O_7$?

1) +6 2) +3
3) +12 4) +2

5. In which substance does hydrogen have an oxidation number of zero?

1) H_2S 2) LiH
3) H_2 4) H_2O

6. If element X forms the oxides XO and X_2O_3, the oxidation numbers of element X are

1) +1 and +2 2) +2 and +3
3) +1 and +3 4) +2 and +4

7. In which substance does sulfur have a negative oxidation number?

1) $CaSO_4$ 2) Na_2S
3) SO_2 4) S

8. What is the oxidation number of oxygen in OF_2?

1) +1 2) +2
3) -1 4) -2

9. Given the balanced ionic equation:

$$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$$

Compared to the total charge of the reactants, the total charge of the products is

1) less 2) greater
3) the same

10. Which type of reaction occurs when nonmetal atoms become negative nonmetal ions?

1) condensation 2) substitution
3) reduction 4) oxidation

11. Given the reaction:

$$Mg(s) + 2\,H^+(aq) + 2\,Cl^-(aq) \rightarrow Mg^{2+}(aq) + 2\,Cl^-(aq) + H_2(g)$$

Which species undergoes oxidation?

1) $Mg(s)$
2) $Cl^-(aq)$
3) $H^+(aq)$
4) $H_2(g)$

12. Given the electrochemical cell reaction:

$$Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$$

Which species is the reducing agent?

1) Ni^{2+} 2) Zn
3) Ni 4) Zn^{2+}

13. Given the reaction:

$$F_2(g) + 2\,Br^-(aq) \rightarrow Br_2(\ell) + 2\,F^-(aq)$$

Which species is reduced?

1) $F^-(aq)$
2) $Br^-(aq)$
3) $F_2(g)$
4) $Br_2(\ell)$

14. Which oxidation number change could occur during an oxidation of an element?

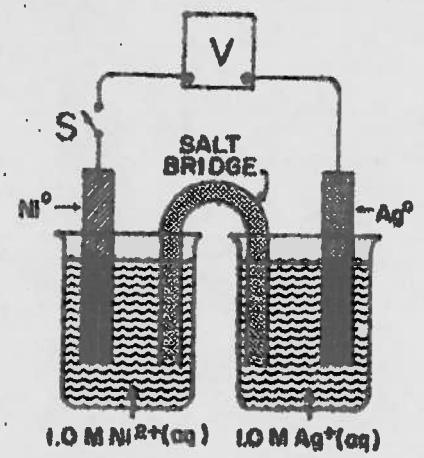
1) +3 to +1 2) -2 to -3
3) +1 to -1 4) +2 to +3

Practice Test

15. Which half-reaction correctly represents oxidation?

1) $F_2 \rightarrow 2\,F^- + 2e^-$
2) $H_2 \rightarrow 2\,H^+ + 2e^-$
3) $F_2 + 2e^- \rightarrow 2\,F^-$
4) $H_2 + 2e^- \rightarrow 2\,H^+$

16. Base your answer to the following question on the diagram of the chemical cell at 298 K and on the equation below.



$$Ni^0(s) + 2Ag^+(aq) \rightarrow Ni^{2+}(aq) + 2Ag^0(s)$$

As the reaction in this cell takes place, the concentration of Ni^{2+} ions

1) decreases and the concentration of Ag^+ ions decreases
2) increases and the concentration of Ag^+ ions decreases
3) decreases and the concentration of Ag^+ ions increases
4) increases and the concentration of Ag^+ ions increases

17. Which half-reaction correctly represents reduction?

1) $S^{2-} \rightarrow S^0 + 2e^-$
2) $Mn^{7+} \rightarrow Mn^{4+} + 3e^-$
3) $Mn^{7+} + 3e^- \rightarrow Mn^{4+}$
4) $S^{2-} + 2e^- \rightarrow S^0$

18. Which balanced equation represents a redox reaction?

1) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
2) $CuO + CO \rightarrow Cu + CO_2$
3) $BaCl_2 + K_2CO_3 \rightarrow BaCO_3 + 2KCl$
4) $HCl + KOH \rightarrow KCl + H_2O$

19. Which statement correctly describes a redox reaction?

1) The oxidation half-reaction occurs after the reduction half-reaction.
2) The oxidation half-reaction and the reduction half-reaction occur simultaneously.
3) The oxidation half-reaction occurs spontaneously but the reduction half-reaction does not.
4) The oxidation half-reaction occurs before the reduction half-reaction.

20. Given the reaction:

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Which statement correctly describes what occurs when this reaction takes place in a closed system?

1) There is a net gain of mass.
2) Atoms of Zn(s) lose electrons and are oxidized.
3) Atoms of Zn(s) gain electrons and are reduced.
4) There is a net loss of mass.

21. Which is a redox reaction?

1) $Fe + 2\,HCl \rightarrow FeCl_2 + H_2$
2) $H^+ + Cl^- \rightarrow HCl$
3) $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
4) $NaOH + HCl \rightarrow NaCl + H_2O$

22. Which is a redox reaction?

1) $2\,HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$
2) $4\,HCl + MnO_2 \rightarrow MnCl_2 + 2\,H_2O + Cl_2$
3) $2\,HCl + FeS \rightarrow FeCl_2 + H_2S$
4) $HCl + KOH \rightarrow KCl + H_2O$

23. The oxidation number of a reducing agent can change from

1) -1 to -3
2) -2 to -1
3) 3 to -1
4) 4 to -3

24. In a redox reaction, the species reduced

1) gains electrons and is the oxidizing agent
2) loses electrons and is the oxidizing agent
3) loses electrons and is the reducing agent
4) gains electrons and is the reducing agent

25. Which of the following elements is the poorest reducing agent?

1) Al
2) Zn
3) H_2
4) Ba

26. As the elements in Period 3 of the Periodic Table are considered in order of increasing atomic number, the ability of each successive element to act as a reducing agent

1) decreases
2) increases
3) remains the same

27. Which metal reacts spontaneously with a solution containing zinc ions?

1) magnesium
2) silver
3) nickel
4) copper

28. According to Reference Table *J*, which of these metals will react most readily with 1.0 M HCl to produce $H_2(g)$?

1) Ca
2) K
3) Zn
4) Mg

29. According to Reference Table J, which of these ions is most easily reduced?

1) Cu^+
2) Cr^{3+}
3) Ca^{2+}
4) Ag^+

30. Referring to Reference Table J, which reaction will not occur under standard conditions?

1) $Sn(s) + 2\ HCl(aq) \rightarrow SnCl_2(aq) + H_2(g)$
2) $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
3) $Cu(s) + 2\ HCl(aq) \rightarrow CuCl_2(aq) + H_2(g)$
4) $Ba(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$

31. Based on Reference Table J, which of the following elements will replace Pb from $Pb(NO_3)_2(aq)$?

1) Mg(s)
2) Cu(s)
3) Ag(s)
4) Au(s)

32. According to Reference Table J, which species is the strongest oxidizing agent?

1) Li^+
2) $F_2(g)$
3) Li(s)
4) F^-

33. Which metal is used as a coating on steel to limit corrosion?

1) Zn
2) Na
3) Ca
4) K

34. Given the nickel-cadmium battery reaction:

$$2NiOOH + Cd + 2H_2O \underset{\text{charge}}{\overset{\text{discharge}}{\rightleftharpoons}} 2Ni(OH)_2 + Cd(OH)_2$$

During the discharge of the battery, Ni^{3+} ions are
1) oxidized, and cadmium metal is oxidized
2) reduced, and cadmium metal is oxidized
3) reduced, and cadmium metal is reduced
4) oxidized, and cadmium metal is reduced

35. Given the reaction:

$$2\ Al^0(s) + 3\ Ni^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Ni^0(s)$$

What is the total number of moles of electrons lost by 2 moles of $Al^0(s)$?
1) 6
2) 2
3) 3
4) 8

36. Given the unbalanced equation:

$$\_\_\_Ag_2S + 8HNO_3 \rightarrow \_\_\_AgNO_3 + 2NO + \_\_\_S + \_\_\_H_2O$$

What is the coefficient of Ag_2S when the equation is completely balanced using the smallest whole numbers?
1) 6
2) 2
3) 3
4) 4

37. Which simple oxidation-reduction reaction is *not* correctly balanced?
1) $Ni(s) + Sn^{2+}(aq) \rightarrow Sn(s) + Ni^{2+}(aq)$
2) $2\ I^-(aq) + Fe^{3+}(aq) \rightarrow Fe^{2+}(aq) + I_2(s)$
3) $Sn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Sn^{2+}(aq)$
4) $2\ I^-(aq) + Hg^{2+}(aq) \rightarrow Hg(\ell) + I_2(s)$

38. Iron corrodes more easily than aluminum and zinc because aluminum and zinc both
1) form oxides
2) are oxidizing agents
3) are reduced
4) form oxides that are self-protective

39. Iron corrodes according to the equation

$$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3.$$

This redox process occurs because
1) iron loses electrons and is oxidized
2) oxygen loses electrons and is reduced
3) iron gains electrons and is reduced
4) oxygen gains electrons and is oxidized

40. Which two metals resist corrosion by forming self-protective coatings?
1) Al and Fe
2) Fe and Zn
3) Al and Zn
4) Fe and Na

41. The chemical reaction that causes corrosion of metals in contact with water and oxygen is
1) a substitution reaction
2) an addition reaction
3) a reduction and oxidation reaction
4) a neutralization and ionization reaction

42. Which of the following metals forms a self-protective coating when exposed to air and moisture?
1) zinc
2) iron
3) sodium
4) calcium

43. Which conversion of energy always occurs in a voltaic cell?
1) electrical energy to chemical energy
2) chemical energy to light energy
3) light energy to chemical energy
4) chemical energy to electrical energy

44. A voltaic cell spontaneously converts
1) chemical energy to electrical energy
2) nuclear energy to electrical energy
3) electrical energy to chemical energy
4) electrical energy to nuclear energy

45. A voltaic cell differs from an electrolytic cell in that in a voltaic cell
1) neither oxidation nor reduction occurs
2) energy is required for the reaction to occur
3) energy is produced when the reaction occurs
4) both oxidation and reduction occur

46. What is the voltage for a chemical cell that has reached equilibrium?
1) 1.00 V
2) greater than 1.00 V
3) greater than 0.00 V and less than 1.00 V
4) 0.00 V

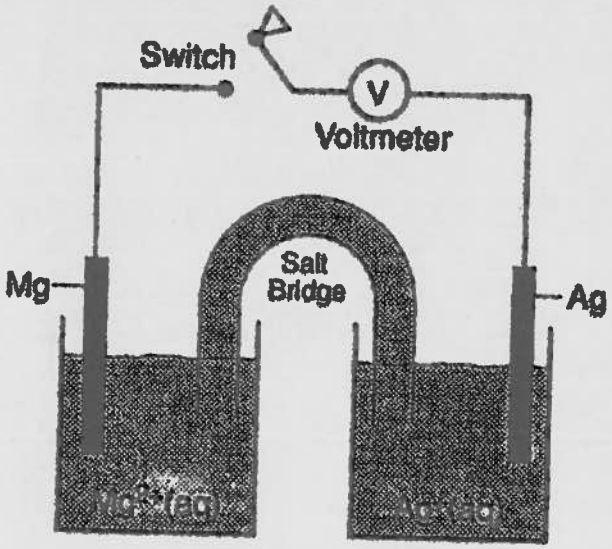
47. Which substance functions as the electrolyte in an automobile battery?
1) H_2O
2) PbO_2
3) H_2SO_4
4) $PbSO_4$

48. The type of reaction in an electrochemical cell is best described as a
1) non-spontaneous oxidation-reduction reaction
2) spontaneous oxidation reaction, only
3) spontaneous oxidation-reduction reaction
4) non-spontaneous oxidation reaction, only

49. Which statement is true about oxidation and reduction in an electrochemical cell?
1) Both occur at the anode.
2) Both occur at the cathode.
3) Oxidation occurs at the anode and reduction occurs at the cathode.
4) Oxidation occurs at the cathode and reduction occurs at the anode.

50. Which component of an electrochemical cell is correctly paired with its function?
1) external conductor – allows the solutions to mix
2) external conductor – permits the migration of ions
3) salt bridge – permits the migration of ions
4) salt bridge – allows the solutions to mix

51. Base your answer to the following question on the equation and diagram below represent an electrochemical cell at 298 K and 1 atmosphere.



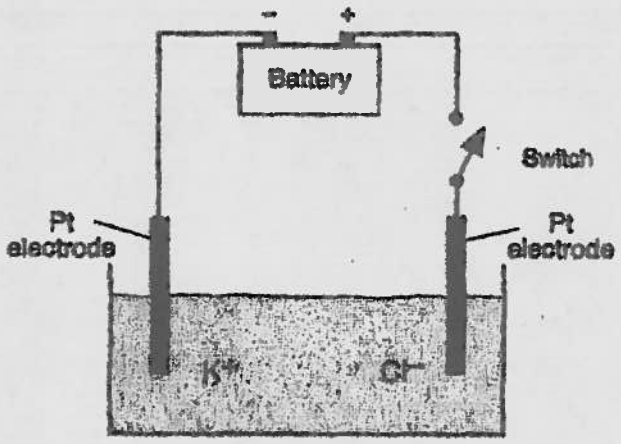
$Mg(s) + 2Ag^+(aq) \longrightarrow Mg^{2+}(aq) + 2Ag(s)$

Which species is oxidized when the switch is closed?
1) $Ag(s)$
2) $Mg^{2+}(aq)$
3) $Mg(s)$
4) $Ag^+(aq)$

52. A redox reaction is set up so that both half reactions take place in separate beakers that are connected by a salt bridge and an external conductor. A path for the transfer of ions is provided by the
1) external conductor
2) salt bridge
3) cathode
4) anode

53. In an electrolytic cell, the negative electrode is called the
1) anode, at which oxidation occurs
2) anode, at which reduction occurs
3) cathode, at which oxidation occurs
4) cathode, at which reduction occurs

54. The diagram below shows the electrolysis of fused KCl.

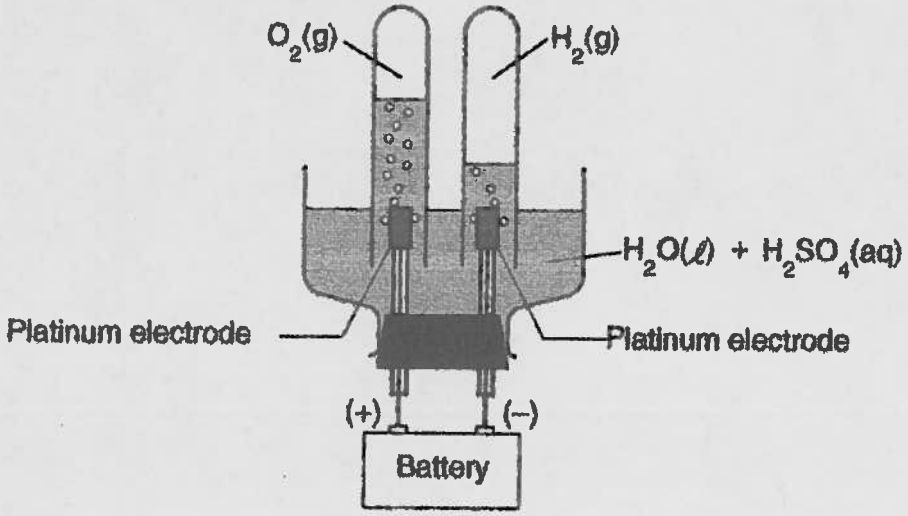


What occurs when the switch is closed?

1) Positive ions migrate toward the cathode, where they gain electrons.
2) Positive ions migrate toward the anode, where they gain electrons.
3) Positive ions migrate toward the anode, where they lose electrons.
4) Positive ions migrate toward the cathode, where they lose electrons.

Base your answers to questions 55 and 56 on the information and diagram below.

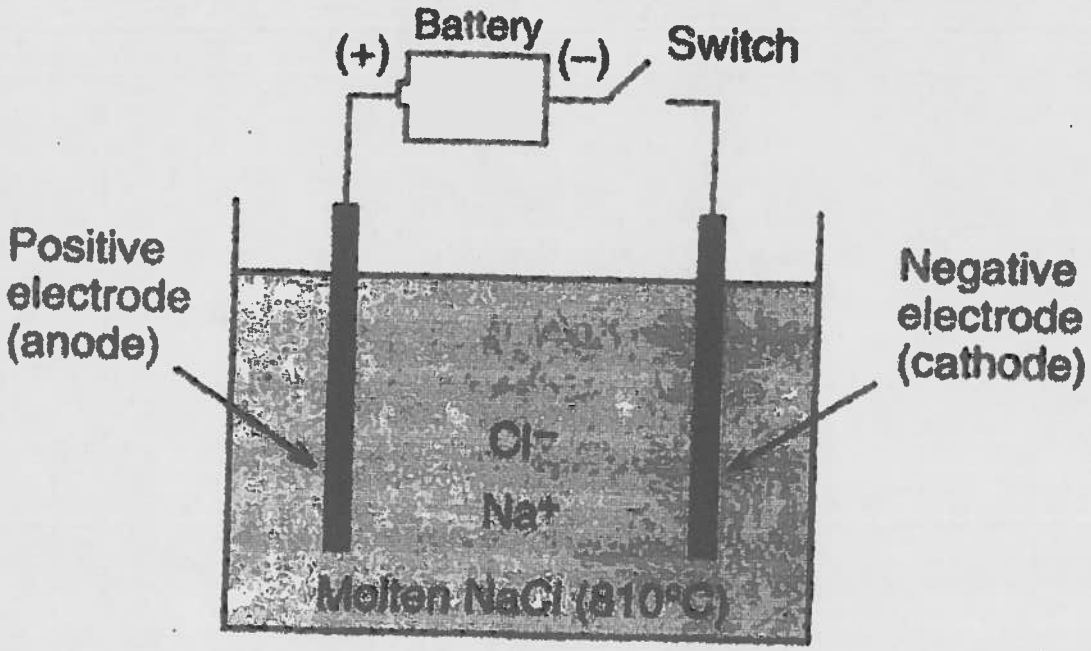
The apparatus shown in the diagram consists of two inert platinum electrodes immersed in water. A small amount of an electrolyte, H_2SO_4, must be added to the water for the reaction to take place. The electrodes are connected to a source that supplies electricity.



55. What type of electrochemical cell is shown?

56. What particles are provided by the electrolyte that allow an electric current to flow?

Base your answers to questions 57 through 59 on the diagram and balanced equation below, which represent the electrolysis of molten NaCl.



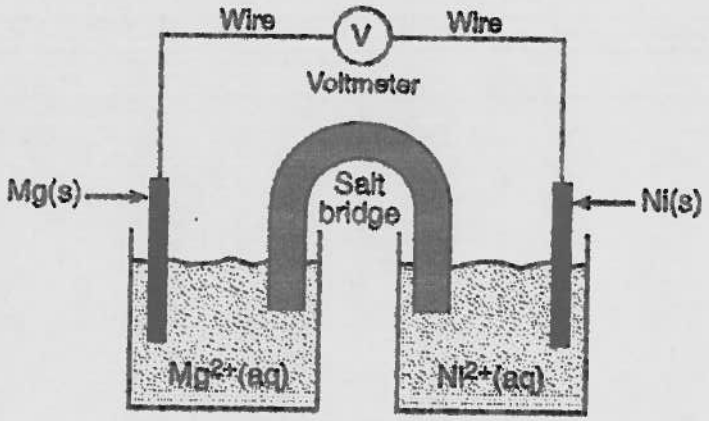
$$2NaCl \longrightarrow Cl_2 + 2Na$$

57. When the switch is closed, which electrode will attract the sodium ions?

58. What is the purpose of the battery in this electrolytic cell?

59. Write the balanced half-reaction for the reduction that occurs in this electrolytic cell.

Base your answers to questions 59 and 60 on the diagram of a voltaic cell and the balanced ionic equation below.



$$Mg(s) + Ni^{2+}(aq) \rightarrow Mg^{2+}(aq) + Ni(s)$$

60. What is the total number of moles of electrons needed to completely reduce 6.0 moles of $Ni^{2+}(aq)$ ions?

61. Explain the function of the salt bridge in the voltaic cell.

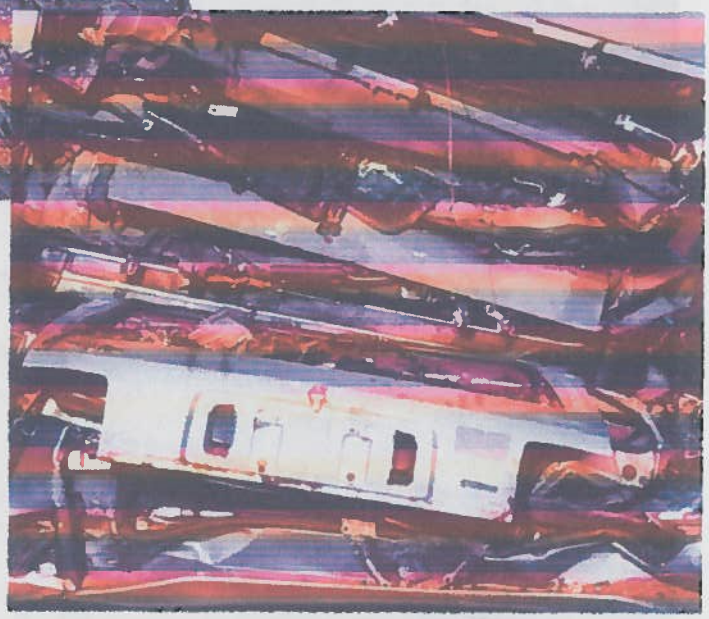
Base your answers to questions 62 and 63 on the unbalanced redox reaction below.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

62. Write the reduction half-reaction.

63. Balance the redox equation using the smallest whole-number coefficients.

EATING AWAY



From Ferraris to Ford Pintos, almost every car is fighting a losing battle to rust.

By Christen Brownlee

The old-style Volkswagen Beetle: Is it a classic car, or an endangered species? The answer depends on where you live. Although there are few classic cars hanging around the northern states or on the coastlines, plenty of vintage automobiles still exist in the mild southern climates and in America's interior states.

The reason that Volkswagen Bugs and other older cars are dropping like flies isn't the typical habitat loss or human encroachment that's plaguing other endangered species. Classic car fleets are constantly shrinking due to a chemical reaction that you're no doubt already familiar with: rusting.

But why does rust unequally strike cars in the snowy states and coastal towns but leave vehicles elsewhere virtually untouched? And more importantly, how can you keep your beloved grocery-getter safe, no matter what parking place you call home? Read on to get the lowdown on how rust works and what measures you can take to stop corrosion in its tracks.

Electron swap meet

Like all types of corrosion, rust is actually a chemical bargain, with two reactions in one: reduction, in which some atoms gain electrons, and oxidation, in which other atoms lose electrons. With all those electrons flowing from one place to another, rust-making is also considered an electrochemical reaction. According to John Scully, a

corrosion expert at the University of Virginia in Charlottesville, the redox reaction that forms rust needs just three ingredients to take place: an anode, or metal that readily gives up electrons; a cathode, or substance (in this case, oxygen) that easily accepts electrons; and an electrolyte solution, which shuttles ions between cathode and anode.

Most cars are made mostly of steel, a tough mixture of iron, carbon, and small amounts of other ingredients like manganese, silicon, phosphorus, and sulfur. It's the iron part of steel that corrodes to make rust.

Iron doesn't hold onto its electrons very tightly, says Scully, making it the perfect anode for an electrochemical reaction to take place. Other metal atoms in the steel mixture, or even another point on the piece of iron, make excellent cathodes. Steel has a nonuniform surface because the chemical composition is not completely homogeneous. Also, physical strains leave stress points in the metal. These defects create anodic regions where the iron is more easily oxidized than it is at others (cathodic regions). However, without a bridge to connect potential anodes and cathodes, rusting would be such a time-consuming process that cars would virtually last forever.

The water on the steel surface is a solvent for ions produced when the iron metal at the anodic region loses electrons (is oxidized to form ferrous ions) and the electrons are conducted through the metal to the cathodic region where they react with water and oxygen from the air to form hydroxide ions, as shown in these equations:

$$2Fe \rightarrow 2Fe^{2+} + 4e^-$$ (oxidation at anodic sites)

$$4e^- + O_2 + 2H_2O \rightarrow 4OH^-$$ (reduction at cathodic sites)

The ions in this electrolyte solution can migrate together and react to form ferrous hydroxide, which reacts further with oxygen from the air to oxidize the ferrous ion and form insoluble ferric oxide, the chemical name for rust, as shown in these equations:

$$Fe^{2+} + 2OH^- \rightarrow Fe(OH)_2$$ (formation of ferrous hydroxide)

$$4Fe(OH)_2 + O_2 \rightarrow 2Fe_2O_3 + 4H_2O$$ (oxidation to ferric oxide or red 'rust')

The movement of ions through the electrolyte solution completes the electric circuit that allows the electrons from iron to move from the anode to the cathode.

A redox reaction on wheels. The bumper of this vehicle wears the product of the reaction between iron and oxygen—rust!

But all this still doesn't explain why colder climates and coastal areas get an unfair share of rust. The magic ingredient that both areas share—which is missing in the basic recipe for rust—is a high abundance of salt. Coastal areas have plenty of salt sailing through the air from ocean spray, and with each cold, snowy winter, northern states smear tons of rock salt on roads to lower the freezing point of water and help keep roads free or ice and snow. [See "Salting Roads: The Solution for Winter Driving" in this issue]

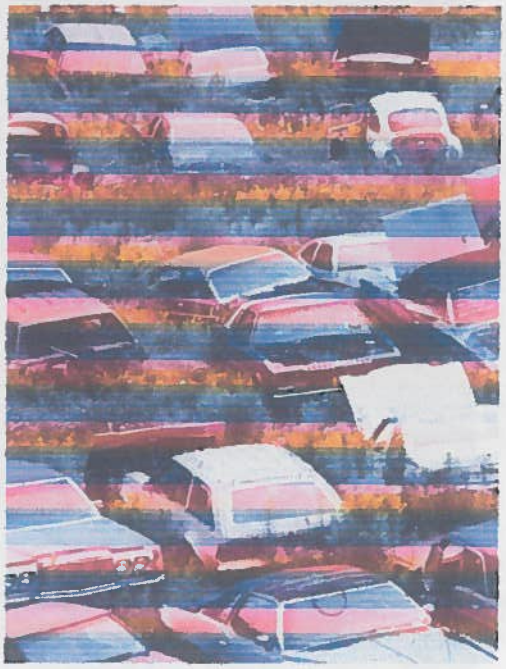
Salt speeds rust's redox reaction along by making water a better conductor. "Salt allows the anode and cathode to be in touch even better," says Scully, making corrosion happen even quicker. Also, chloride ions form very stable complex ions with Fe^{3+}, which helps dissolve iron and accelerate corrosion.

Costly corrosion

Scully points out that iron can't help rusting—existing as an oxide in its thermodynamically favored state. In fact, the metal rarely exists in a pure state in nature. Before it becomes a side panel in your car, engineers must convert rusty iron ore into a pure metal.

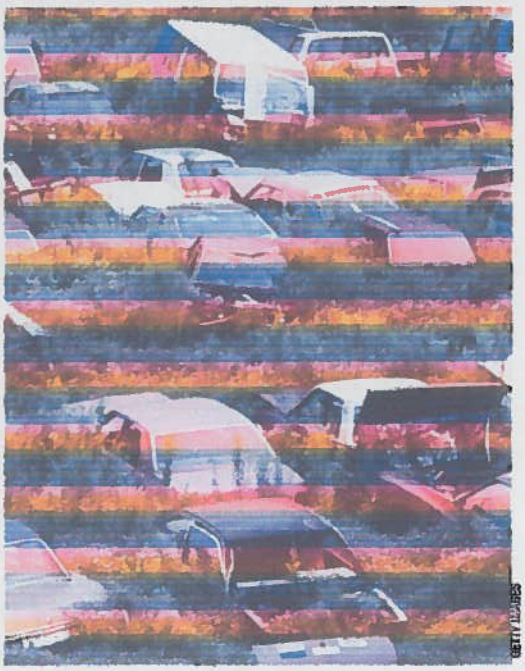
With rust being iron's favored state, it is of little use trying to fix rust after it has already happened. By putting energy into rust, it's possible to plate metal back onto a car, says Scully. However, it's also incredibly costly and impractical. Plus, by the time most car owners notice a rust spot, a significant amount of iron has already sloughed off of the car, lost to wind and rain.

One solution to stopping iron's thermodynamic conversion, says Scully, would be to make cars out of a metal that doesn't corrode, such as gold or silver. "Converting to an oxide isn't thermodynamically favorable



It's not only cars that are rusting away. Corrosion costs the United States a whopping $276 billion per year!

This student investigates the role of salt on the rate of rusting by putting nails in various salt water solutions.



reaction for these metals, so they won't corrode spontaneously," he says. "Archaeologists can dig up gold coins that have survived through the ages without corroding."

But while driving a gold car might make you feel like a million bucks, making such a vehicle would cost substantially more. Not to mention the fact that these soft metals would be unable to support the car's weight and would squash like putty in a collision!

Even using noncorroding metals that are cheaper than gold or silver, such as stainless steel that Deloreans are made of, is still more expensive than using the plain steel that most cars are made of today.

The best way to prevent corrosion is still the cheapest. A good coating of paint removes the connection between anode and cathode by preventing water from making contact with steel. Without water, rusting slows down to a snail's pace.

Today's high tech paints have evolved far from being just a simple barrier, says Scully. Researchers are currently working on paints that release rust inhibitors on demand when paint's seal on steel is breached, for example, when the paint on a car is scraped or scratched. Other "smart paints" that ooze together to close gaps whenever a car's panels get scratched are also in the works.

Although rust is a big deal for car owners, it's an even bigger deal for industries that rely on machines with metal parts, ranging from farm tools to factory equipment to fighter jets. According to Scully, corrosion costs the United States a whopping annual toll of about $276 billion in lost goods and services. With the exorbitant cost of new military equipment, the Department of Defense (DOD) is one of the largest investors in antirust research, says Scully. Scientists at the DOD hope they can keep the aging Blackhawk helicopters and B52 bombers that are currently in use running smoothly for decades to come—a cost savings of millions of dollars per machine.

But one military asset sometimes harmed by corrosion is extremely difficult to put a price on, says Scully. "If a soldier goes to war and the rifle he's using to protect himself doesn't fire when he needs it to, how do you estimate the cost of corrosion then?"

Christen Brownlee is a contributing editor to *ChemMatters*. Her article "Super Fibers" also appears in this issue.

(14)

Flaking Away Article

- Include Paragraph # in your answer and highlight or underline the evidence in the article.

1. Where in the United States are more vintage cars found? Why?

2. What is reduction?

3. What is oxidation?

4. What is an electrochemical reaction?

5. What are the three ingredients needed for the redox reaction to take place that forms rust?

   What is an anode?

   What is a cathode?

   What is an electrolyte solution?

6. What is the part of steel that corrodes to make rust?

7. What does water allow to happen when a car rusts?

8. Why do colder climates and coastal areas have an increased rate of rusting? Explain.

9. How does salt affect the rusting of cars?

10. What is one solution to stop the rusting of cars? What is the problem with this solution?

11. What areas, other than cars, are affected by rusting?

(75)