Unit 4
ELECTROCHEMISTRY
Specific Curriculum Outcomes

Suggested Time: 25 Hours

Electrochemistry

Introduction

Matter is electrical in nature and some of its most important particles—electrons, protons, and ions—carry electric charge. When an electrical potential is applied between electrodes placed in a solution of ions, ions migrate to oppositely charged electrodes and chemical reactions take place. Quantitative aspects of this electrolysis are important in analytical chemistry and the chemical industry.

Focus and Context

This unit builds on concepts dealing with electric forces, matter and energy in chemical change, and quantitative relationships in chemical changes. Energy is involved in electrochemical changes. Problem solving and decision making in this unit will be helpful in creating an interest in the application of technology. Students should investigate, through laboratory work and relevant problems, the ways in which science and technology advanced in relation to each other. The oxidation-reduction reactions that occur in everyday life, the products and processes used in industry, or the relationship of global environmental problems to oxidation-reduction reactions could be investigated.

Science Curriculum Links

Students will have studied the mole and electronegativity in Chemistry 2202. This unit will provide a more indepth look at what is happening in chemical reactions such as single replacements. Solutions, ionization, and chemical equilibrium in Chemistry 3202 should be completed before beginning electrochemistry.

Curriculum Outcomes

STSE

Students will be expected to

Nature of Science and Technology

115-1 distinguish between scientific questions and technological problems

Relationships Between Science and Technology

116-3 identify examples where technologies were developed based on scientific understanding

116-5 describe the functioning of domestic and industrial technologies, using scientific principles

116-6 describe and evaluate the design of technological solutions and the way they function, using scientific principles

116-7 analyse natural and technological systems to interpret and explain their structure and dynamics

Social and Environmental Contexts of Science and Technology

118-1 compare the risks and benefits to society and the environment of applying scientific knowledge or introducing a technology

118-4 evaluate the design of a technology and the way it functions on the basis of a variety of criteria that they have identified themselves

Skills

Students will be expected to

Initiating and Planning

212-1 identify questions to investigate that arise from practical problems and issues

212-2 define and delimit problems to facilitate investigation

212-7 formulate operational definitions of major variables

Performing and Recording

213-2 carry out procedures controlling the major variables and adapting or extending procedures where required

213-8 select and use apparatus and materials safely

Analysing and Interpreting

214-7 compare theoretical and empirical values and account for discrepancies

214-8 evaluate the relevance, reliability, and adequacy of data and data collection methods

214-14 construct and test a prototype of a device or system and troubleshoot problems as they arise

214-16 evaluate a personally designed and constructed device on the basis of criteria they have developed themselves

214-18 identify and evaluate potential applications of findings

Communication and Teamwork

215-7 evaluate individual and group processes used in planning, problem solving and decision making, and completing a task

Knowledge

Students will be expected to

322-1 define oxidation and reduction experimentally and theoretically

322-2 write and balance half reactions and net reactions

322-3 compare oxidation-reduction reactions with other kinds of reactions

322-4 illustrate and label the parts of electrochemical and electrolytic cells and explain how they work

322-5 predict whether oxidation-reduction reactions are spontaneous based on their reduction potentials

322-6 predict the voltage of various electrochemical cells

322-7 compare electrochemical and electrolytic cells in terms of energy efficiency, electron flow/transfer and chemical change

322-8 explain the processes of electrolysis and electroplating

322-9 explain how electrical energy is produced in a hydrogen fuel cell

ACC-9 differentiate between the terms mineral and ore

ACC-10 define and describe the process of *flotation*, as it applies to the mining industry

ACC-11 define the terms flotation, slag, and leaching, as they apply to pyrometallurgy and hydrometallurgy

ACC-12 describe the process of electrowinning pure nickel from a nickel(II) subsulfide matte, used in the industrial production of pure nickel

Oxidation and Reduction

Outcomes

Students will be expected to

- observe oxidation and reduction experimentally (322-1)
- define oxidation and reduction theoretically (322-1)
  - define the terms: oxidation and reduction in terms of loss or gain of electrons
  - identify electron transfer in redox equations
  - use mnemonics such as "LEO the lion says GER", or "OIL RIG" to remember that when a chemical entity loses electrons it is oxidized, when it gains electrons it is reduced
  - identify oxidation and reduction half-reaction equations in an oxidation-reduction (redox) equation

Elaborations—Strategies for Learning and Teaching

Students should understand from the beginning that electrochemistry involves the study of all reactions in which transfer of electrons takes place. Some of these reactions produce electricity for our use (e.g., batteries) while others need electrical energy to occur (e.g., electrolysis of water). However, many electrochemical reactions occur where no electricity is produced or used (e.g., $Na + ½\ Cl_2 \rightarrow NaCl$).

The transfer of electrons may be illustrated using simple Bohr models to show formation of Na^+ from Na and Cl^- from Cl and the resulting electron loss and gain.

One way to begin this topic is to ask students to observe an electrochemical reaction. A simple one is to place a strip of zinc metal in a $CuSO_4$ solution. Copper metal forms on the surface of the zinc, zinc metal disappears, and the blue color of the Cu^{2+} ions fades. The two half-reactions represent electron loss and electron gain at the zinc surface. These can be determined, along with the overall reaction, by doing questions similar to the following:

e.g., What happened to the Cu^{2+} ions? They changed to Cu metal.

What is the equation for this? $Cu^{2+} \rightarrow Cu$

How can this occur? Cu^{2+} gains two electrons.

Is the equation balanced? It is not balanced for charge but it can be balanced using the two electrons lost as follows:

$Cu^{2+} + 2e^- \rightarrow Cu$ This is one half reaction.

Where do the two electrons come from? The Zn as it changes to Zn^{2+} in solution.

What is the second half reaction? $Zn \rightarrow Zn^{2+} + 2e^-$

At this point the definitions of oxidation and reduction can be introduced using the two half-reactions above as examples. The mnemonics OIL (**O**xidation **I**nvolves **L**oss) and RIG (**R**eduction **I**nvolves **G**ain) can be introduced to help students remember the meaning of both terms.

Oxidation and Reduction

Tasks for Instruction and/or Assessment

Journal

- All halogens kill bacteria and other microorganisms. Chlorine is a halogen that is safe enough and readily available for large-scale treatment of public water supplies. Students could determine what happens to the hypochlorous acid formed when $Cl_{2(g)}$ is added to $H_2O_{(l)}$. (212-1, 115-1, 322-1)
- Teachers could ask, "What does a study of electrochemistry involve?" (115-1, 322-1)

Paper and Pencil

- Students could write half-reactions for each of the following:

  $Br_2 + 2Cl^- \rightarrow Cl_2 + 2Br^-$

  $Cu + Cd^{2+} \rightarrow Cu^{2+} + Cd$ (322-1)

Journal

- Now that students have been introduced to redox, they could list examples that might occur in their homes. (322-3)

Resources/Notes

www.gov.nl.ca/edu/science_ref/main.htm

MGH Chemistry, pp. 712-713

MGH Chemistry, pp. 714-715

MGH Chemistry, p. 714

MGH Chemistry, pp. 715-716

Redox and Half Reactions

Outcomes

Students will be expected to

- define oxidation and reduction theoretically (322-1) (**Cont'd**)
  - identify a redox equation as the sum of the oxidation half-reaction and the reduction half-reaction
  - identify the species oxidized, the species reduced, the oxidizing agent and the reducing agent in simple redox equations

Elaborations—Strategies for Learning and Teaching

The previous line of questioning above can be continued.

How can we get the equation for the overall reaction? Add them together:

$$Zn \rightarrow Zn^{2+} + 2e^- \quad \text{(oxidation) [O]}$$

$$Cu^{2+} + 2e^- \rightarrow Cu \quad \text{(reduction) [R]}$$

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

Using this overall reaction, the electrons lost by Zn are the same electrons gained by Cu^{2+} ions. The SO_4^{2-} was a spectator ion, therefore, it does not show up in the overall reaction.

A redox reaction is the combination of a reduction half reaction and an oxidation half reaction.

Using the overall reaction, students must determine the substance oxidized, the substance reduced, the oxidizing agent, and the reducing agent. For example:

Cu^{2+} is being reduced by Zn, therefore, Zn is the reducing agent.

Zn is being oxidized by Cu^{2+}, therefore, Cu^{2+} is the oxidizing agent.

Teachers should note that a common error by students is to assign oxidation, reduction, oxidizing agent and reducing agent with four different substances (reactants and products) involved in a redox reaction. It is important for students to remember that oxidation, reduction, oxidizing agent and reducing agent only involve the reactants.

Outcomes

- compare oxidation-reduction reactions with other kinds of reactions (322-3)
  - define oxidation number
  - use oxidation number rules to find the oxidation numbers of the atoms in molecules or ions
  - identify changes in oxidation number in half-reactions and in redox equations

Elaborations—Strategies for Learning and Teaching

Students should learn the oxidation number rules and be able to apply them to determine the oxidation number of all species in molecules or ions. By applying oxidation number rules to equations students should be able to predict if the reaction is an oxidation-reduction reaction or not. Using the numbers assigned to individual species, they should be able to identify the species oxidized, species reduced, oxidizing agent and reducing agent.

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(aq)}$$

Ox # +1 -1 | +1 +5 -2 | +1 +5 -2 | +1 -1

no change in ox. #, therefore not a redox reaction.

$$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$$

Ox # 0 | +1 +5 -2 | 0 | +2 +5 -2

red.

oxd.

change in ox. #, therefore a redox reaction.

Ag^+ ions (ox.# = +1) have been **reduced** to Ag metal atoms (ox.# = 0) Cu metal atoms (ox.# = 0) have been **oxidized** to Cu^{2+} ions (ox.# = +2). **Note: reduction** involves a **reduction** in oxidation number. In this reaction Cu is the reducing agent and the $AgNO_3$ is the oxidation agent.

Redox and Half Reactions

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could show the species that is oxidized and the species reduced. They should identify the oxidizing agent, reducing agent, oxidation number of each species, and electron transfer in each of the following unbalanced equations:

  $Fe^{2+} + MnO_4^- + H^+ \rightarrow Fe^{3+} + Mn^{2+} + H_2O$

  $Cu + NO_3^- + H^+ \rightarrow Cu^{2+} + NO + H_2O$

  (322-1)

- Students could identify which of the following are redox reactions. If they are redox, identify the oxidizing agent and reducing agent.

  (i) $BaCl_{2(aq)} + 2\ KIO_{3(aq)} \rightarrow Ba(IO_3)_{2(aq)} + 2\ KCl_{(aq)}$
  (ii) $Mg_{(s)} + Br_{2(g)} \rightarrow MgBr_{2(s)}$
  (iii) $NH_4NO_{2(s)} \rightarrow N_{2(g)} + 2\ H_2O_{(l)}$
  (iv) $CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(l)}$ (322-3)

Resources/Notes

MGH Chemistry, pp. 715-716, p. 833

MGH Chemistry, pp. 714-715

MGH Chemistry, p. 721

MGH Chemistry, pp. 721-725

MGH Chemistry, pp. 726-728

Balancing Redox Reactions

Outcomes

Students will be expected to

- compare oxidation-reduction reactions with other kinds of reactions (322-3) (**Cont'd**)
  – given an equation for composition or decomposition of simple ionic compounds identify oxidation and reduction half reactions by balancing the number of electrons lost and gained

- write and balance half-reactions and complete redox reactions (322-2)

Elaborations—Strategies for Learning and Teaching

Once oxidation numbers for all species have been assigned in oxidation-reduction (redox) reactions the half-reactions can be identified. The number of electrons required to balance each half-reaction equals the change in oxidation number for species oxidized and reduced.

$$\underset{0}{Cu} \rightarrow \underset{+2}{Cu^{2+}} + 2e^- \qquad \underset{+1}{Ag^+} + e^- \rightarrow \underset{0}{Ag}$$

- $2\ e^-$ oxidation   + $1\ e^-$ reduction

Teachers could emphasize identification of redox reactions by indicating that any reaction involving an element as a reactant or product is definitely redox. This includes all composition, decomposition, single replacement, and combustion reactions. Double replacement reactions are **never** redox reactions.

Students should be able to determine and balance the half-reactions in the following reaction:

$$2\underset{0}{Na_{(s)}} + \underset{0}{Cl_{2(g)}} \rightarrow 2\underset{+1}{Na}\underset{-1}{Cl_{(s)}}$$

oxidation   reduction

$$2Na \rightarrow 2Na^+ + 2e^- \text{ (Oxd)}$$

$$Cl_{2(g)} + 2e^- \rightarrow 2Cl^- \text{ (Red)}$$

Hint: All half-reactions must be balanced for both atoms and charge.

Teachers could review that redox equations for reactions are combinations of an oxidation half-reaction and a reduction half-reaction. Students should understand that the electrons lost during oxidation equals the electrons gained during reduction. If the numbers of electrons lost in the balanced oxidation half-reaction are not equal to the electrons gained in the balanced reduction half-reaction; one or both equations will have to be multiplied by appropriate integers before adding them to the overall equation.

For example: (1) $Au^{3+} + 3e^- \rightarrow Au$

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

Multiply equation (1) by 2 and equation (2) by 3 before adding to give

$$2Au^{3+} + \cancel{6e^-} \rightarrow 2Au$$

$$3Zn \rightarrow 3Zn^{2+} + \cancel{6e^-}$$

Now that the numbers of electrons lost equals the number of electrons gained the half reactions can be added to give:

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

Students should check each final redox equation to ensure that atoms and charge are conserved.

Balancing Redox Reactions

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could balance the following redox equations using half-reactions:

  (i) $Co^{3+}_{(aq)} + Au_{(s)} \rightarrow Co^{2+}_{(aq)} + Au^{3+}_{(aq)}$
  (ii) $Co^{3+}_{(aq)} + Cd_{(s)} \rightarrow Co^{2+}_{(aq)} + Cd^{2+}_{(aq)}$
  (iii) $Mg_{(s)} + P_{4(s)} \rightarrow Mg_3P_{2(s)}$
  (iv) $Fe_2O_{3(s)} \rightarrow Fe_{(s)} + O_{2(g)}$
  (v) $Mg_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + Cu_{(s)}$
  (vi) $Al_{(s)} + Ni(NO_3)_{2(aq)} \rightarrow Al(NO_3)_{3(aq)} + Ni_{(s)}$
  (vii) $Cu^{2+}_{(aq)} + Al_{(s)} \rightarrow Al^{3+}_{(aq)} + Cu_{(s)}$ (322-2)

Resources/Notes

MGH Chemistry, p. 728

MGH Chemistry, pp. 734-736

Balancing Redox Reactions *(continued)*

Outcomes

Students will be expected to

- write and balance half-reactions and complete redox reactions (322-2) (**Cont'd**)
  - write and balance equations for complex oxidation-reduction reactions occurring in acidic or basic solutions
  - solve redox stoichiometry problems for redox reactions

Elaborations—Strategies for Learning and Teaching

Many redox reactions occur only in acidic or basic solutions. Acidic solutions have available H^+ ions and H_2O molecules available to take part in reaction while basic solutions have available OH^- ions and H_2O molecules available. The H^+ ions or OH^- ions, and H_2O molecules are used to balance the number of oxygen atoms and hydrogen atoms in these acidic or basic redox reactions. Balancing these type of redox equations involves several steps which students must complete in order. To balance under acidic conditions:

1. separate 2 half-reactions (oxd./red.)
2. balance all atoms except O, H
3. balance O using H_2O
4. balance H using H^+
5. balance charges with e^-
6. balance number of e^-
7. add half reactions

To balance redox equations under basic conditions, students follow the same procedure as under acidic conditions and then, as a final step, add the equation: $H^+ + OH^- \rightleftharpoons H_2O$ as many times as necessary to cancel out any H^+ (note: this is an equilibrium reaction, and therefore, could also be used as: $H_2O \rightleftharpoons H^+ + OH^-$).

Teachers may choose to use the oxidation number method for balancing redox equations in acidic and basic conditions.

Redox reactions are reactions that proceed to form products when reactants are mixed. Redox stoichiometry problems are standard stoichiometry problems which were covered in Chemistry 2202 which involve a balanced redox equation. An application example could involve doing a redox titration experiment. This gives students another look at titrations involving a process other than acid-base reactions.

Balancing Redox Reactions *(continued)*

Tasks for Instruction and/or Assessment

Journal

- Students could answer questions such as, "Will corrosion be a greater problem in an acidic or a basic solution?" (116-7, 214-18, 322-2)

Paper and Pencil

- Students could solve redox stoichiometry problems such as:
  - (a) (i) write a balanced equation for the redox equation below: $MnO_4^-{}_{(aq)} + H_2C_2O_{4(aq)} + H^+{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$
    (ii) The titration of the $H_2C_2O_4$ required 35.62 mL of 0.1092 mol/L MnO_4^- solution. Calculate the mass of $H_2C_2O_4$ that reacted with the MnO_4^-.
    (iii) Indicate which substance is the oxidizing agent and which is the reducing agent.
  - (b) How many grams of MnO_2 will form when 0.500 g of Na_2SO_3 is titrated to the endpoint with $KMnO_4$ in the presence of a base? *Hint: The sulfite ions, SO_3^{2-} are oxidized to sulfate ions, SO_4^{2-}.* (322-2)

Resources/Notes

MGH Chemistry, pp. 732-739, 747-750, p. 833

MGH Chemistry, p. 742, p. 833

Electrochemical Cells

Outcomes

Students will be expected to

- illustrate and label the parts of electrochemical cells and explain how they work (322-4)
  - define electrochemical cells
  - draw and label an electrochemical cell. Include:
    (i) anode
    (ii) cathode
    (iii) salt bridge
    (iv) direction of flow of electrons
    (iv) direction of flow of ions

Elaborations—Strategies for Learning and Teaching

Electrochemical cells are generally defined as cells with a spontaneous redox reaction occurring. The terms galvanic cell and voltaic cell are also used interchangeably with electrochemical cell. Students should illustrate, label, define, and identify the parts of an electrochemical cell: anode, cathode, anion, cation, salt bridge/porous cup, and internal and external circuit. Students should identify the flow of electrons and the migration of ions for both electrochemical and electrolytic cells.

The teacher could develop the concept of an electrochemical cell which produces electricity by using the same Zn + Cu^{2+} reaction in questioning and demonstration. For example:

Question: Can the 2 half-reactions occurring at the Zn surface be separated?

Demo: Set up 2 beakeers, one containing a solution of Cu^{2+} ions and a piece of Cu metal, and another containing Zn^{2+} ions and a piece of Zn metal. Use alligator clips to attach the 2 metal strips to a voltmeter. No voltage is observed.

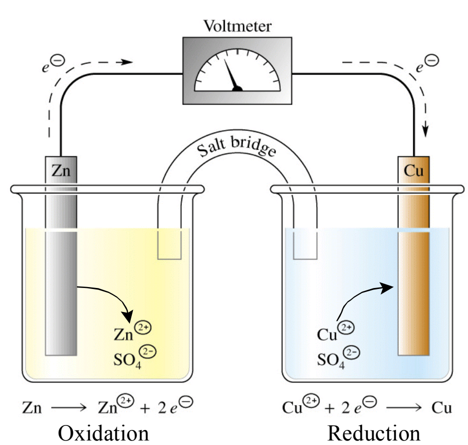
Question: Why is no voltage provided? *Answer:* Not a complete circuit.

Question: How can we create a complete circuit? *Answer:* Connect the the solutions in the 2 beakers.

Students might suggest using a piece of twine or string, which would not work. This allows the teacher to introduce the concept of a salt bridge.

Demo: Soak a piece of paper towel in a salt solution (e.g., $NaCl_{(aq)}$) and place one end in the Cu^{2+} solution and one end in the Zn^{2+} solution. a voltage is observed.

Using the electrochemical cell produced, the direction of flow of electrons can be demonstrated. The anode can be identified as the negative electrode where oxidation occurs. The cathode can be identified as the positive electrode where reduction occurs.



Electrochemical Cells

Tasks for Instruction and/or Assessment

Performance

- Students could determine what is being conserved in a balanced redox reaction. (322-4)

Journal

- Students could explain how the flow of electrons in a flashlight produces light. (116-6, 215-7, 212-2, 322-4)
- Paper clips are sometimes used as electrodes. Students could determine if plastic coated clips could work. (214-16, 322-4)

Paper and Pencil

- Students could report on a personally designed electrochemical device. Include their criteria, procedures, variables, and materials. (322-4, 214-16, 213-2, 213-8)
- Students could answer, "What is an electrochemical cell? (322-4)
- Students could draw a concept map for the following terms: anode, cathode, anion, cation, salt bridge, internal circuit, external circuit, power supply. (322-4, 212-7)

Performance

- Using simple materials like an orange, a potato, or a lemon, students could design an electrochemical cell. (213-8, 322-4)

Presentation

- Students could illustrate and label the parts of an electrochemical cell and explain how it works. (322-4)
- Students could role play the behaviour of a particle in an electrochemical cell. (322-4)
- Students could diagram a cell in which the reaction consists of the displacement of silver from $AgNO_3$ by metallic copper to produce Ag(s). They should write the equation for the half-reaction that takes place in each half-cell, identifying each as oxidation or reduction. Then they should determine the equation for the total cell reaction. (322-4)

Resources/Notes

MGH Chemistry, p. 757, pp. 833-834

MGH Chemistry, pp. 757-760

Electrochemical Cells *(continued)*

Outcomes

Students will be expected to

- illustrate and label the parts of electrochemical cells and explain how they work (322-4) (**Cont'd**)
  - use electrochemical cell notation to represent an electrochemical cell
  - draw and label an electrochemical cell using electrochemical cell notation

Elaborations—Strategies for Learning and Teaching

Electrochemical cell notation is a shorthand method of summarizing the anode and cathode reactions in an electrochemical cell. We will use the convention where the anode reaction appears on the left and the cathode reaction appears on the right. In the notation the half-cells are separated by double vertical lines which represent the salt bridge. Single lines separate anode and cathode from their ion forms in solution. For example:

anode/anodic ion solution//cathodic ion solution/cathode

For the previous electrochemical cell example the cell notation is:

$Zn/Zn^{2+}//Cu^{2+}/Cu$

Students should recognize the use of inert electrodes in electrochemical cells and in electrochemical cell notation.

Electrochemical Cells *(continued)*

Tasks for Instruction and/or Assessment

Performance

- Students could design and test a basic cell of their choice. Compare the results with the table. Comment. (214-8, 322-4, 214-7)

Paper and Pencil

- Students could draw a labelled diagram, indicating what is the anode and cathode and show how they would set up this electrochemical cell: $Zn/Zn^{2+}//Ag^{+}/Ag$. (322-4)
- Students could draw a diagram for an electrochemical cell using $Cu/Cu^{2+}//Ag^{+}/Ag$. On the diagram clearly indicate the anode and cathode, and the directions of electron and ion flow. (322-4)

Resources/Notes

MGH Chemistry, p. 760

MGH Chemistry, p. 761

Redox Reactions with Standard Reduction Potentials

Outcomes

Students will be expected to

- predict cell voltage and whether oxidation-reduction reactions are spontaneous using a table of reduction potentials (322-5, 322-6)
  - identify electrochemical cells as cells which produce electrical energy in spontaneous oxidation-reduction reactions
  - define half-cell voltage, standard half-cell, cell voltage, E°
  - define a spontaneous reaction as one that produces a positive cell potential
  - write and balance equations for oxidation-reduction reactions using half-reaction equations obtained from a standard reduction potential table for simple redox reactions

Elaborations—Strategies for Learning and Teaching

Different electrochemical cells result from different combinations of half-cells which produce different voltages called cell voltages, E°. These cell voltages are combinations of voltages for each half-cell called half-cell voltages. The half-cell voltages are obtained by connecting each half-cell to the standard half-cell, $2H^+ + 2e^- \rightarrow H_2$ (The Hydrogen Half-Cell). The hydrogen half-cell is assigned a voltage of 0.00 volts and is assigned this value as a reference. When connected to the hydrogen cell, the voltage recorded on the voltmeter (the voltage relative to the Hydrogen Half-Cell) is the voltage assigned to each half-cell. For example, when the $Cu^{2+} + 2e^- \rightarrow Cu$ is connected to H_2 hydrogen half-cell the voltage is + 0.34 volts. This is the voltage of the Cu half-cell. When the $Zn^{2+} + 2e^- \rightarrow Zn$ is connected the voltage is -0.76 volts. This is the voltage of the Zn half-cell.

If the voltage of many half-cells are measured they can be listed from most positive voltage to most negative to produce a Table of Standard Half-Cell Potentials.

Many redox reactions are more complex than simple composition, decomposition and single replacement reactions. The half-reactions for such reactions are provided in tables of half-reactions called standard Reduction Potential Tables. Selection of the appropriate half-reactions and following proper procedure will give the correctly balanced redox equation.

Redox Reactions with Standard Reduction Potentials

Tasks for Instruction and/or Assessment

Performance

- Students could construct their own activity series table of strongest to weakest oxidizing agents for metals based on a lab they have designed and completed. They should compare it with the standard electrode potential table. (322-5, 322-6)

Presentation

- Students could sketch a cell that forms iron metal from iron (II) ions while changing chromium metal to chromium (III) ions. They should calculate the voltage, show the electron flow, label the anode and cathode, and balance the overall cell equations. (322-5, 322-6)

Paper and Pencil

- An electric eel can produce a charge of 600 V. It does this by combining the voltages of individual electroplates. If each electroplate produces 150 mV, students could determine how many plates are required to give off the total discharge. (212-2, 215-7, 322-6, 322-7)

Journal

- Students could explain why some reactions are spontaneous and some are not. (322-5, 322-6)

Resources/Notes

MGH Chemistry, p. 761, p. 834

MGH Chemistry, pp. 768-774

MGH Chemistry, p. 761

MGH Chemistry, pp. 772-775

Redox Reactions with Standard Reduction Potentials *(continued)*

Outcomes

Students will be expected to

- predict cell voltage and whether oxidation-reduction reactions are spontaneous using a table of reduction potentials (322-5, 322-6) (**Cont'd**)
  - use a standard reduction potential table to predict cell voltage and to predict if the redox reaction is spontaneous (occurs as is written)
  - develop a table of redox half-reactions from experimental results

Elaborations—Strategies for Learning and Teaching

Half-cell voltages from the Standard Reduction Potential Table can be combined to give cell voltage, E^0, for the overall Redox reaction. A redox reaction is spontaneous if it produces a positive cell voltage. For example, what is the cell voltage for the following reaction? Is the reaction spontaneous? $Cu_{(s)} + Fe^{2+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Fe_{(s)}$

Answer: I. Obtain the half-cell voltages from the Standard Reduction Potential Table.

$Cu^{2+} + 2e^- \rightarrow Cu \quad E^0 = +0.34\ V$

$Fe^{2+} + 2e^- \rightarrow Fe \quad E^0 = -0.45\ V$

II. Identify the two half reactions from the overall reaction. Since the Cu half-reaction is reversed its voltage changes sign to -0.34 volts. While the Fe half-cell voltage remains at -0.45 volts.

$Cu \rightarrow Cu^{2+} + 2e^- \quad E^0 = -0.34\ V$

$Fe^{2+} + 2e^- \rightarrow Fe \quad E^0 = -0.45\ V$

Add: $Cu + Fe^{2+} \rightarrow Cu^{2+} + Fe \quad E^0 = -0.79\ V$

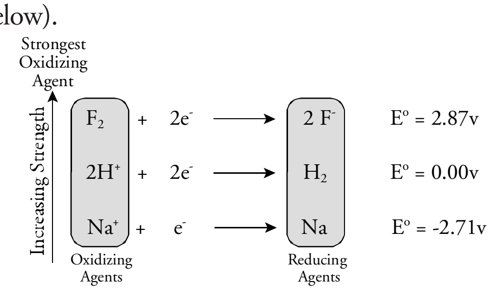
Since E^0 is negative the reaction is not spontaneous and Cu will not react with Fe^{2+} ions. The reverse reaction is spontaneous.

Spontaneity could be demonstrated using the original Zn metal in a Cu^{2+} ion solution, which is spontaneous and produces a precipitate of Cu metal. Then try Cu metal in a Zn^{2+} ion solution and this reaction gives no reaction.

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu \quad E^0 = +1.10\ V \quad$ Spontaneous

$Zn^{2+} + Cu \rightarrow Zn + Cu^{2+} \quad E^0 = -1.10\ V \quad$ Non-spontaneous

Standard Reduction Tables contain two columns of substances, a column of oxidizing agents on the left side and a column of reducing agents on the right side. The strongest oxidizing agent is at the top (F_2 in the mini-table below) and the strongest reducing agent is at the bottom (Na in the mini-table below).



Rule: Spontaneous reactions only occur when the oxidizing agent is above the reducing agent in the Table of Reduction Potentials.

To balance the number of electrons between the oxidizing half-reaction and the reduction half-reaction sometimes one or both must be multiplied by an integer before adding to get the overall redox equation. When this is required **never** multiply the half-cell voltages by the integers used. Greater amounts do not lead to greater voltages.

Redox Reactions with Standard Reduction Potentials *(continued)*

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could predict whether the following reactions are possible:
  - oxidation of iron atoms by silver ions
  - oxidation of bromide ion by chlorine
  - reduction of iodine by fluoride ion (322-5, 322-6)
- Students could write the balanced equation for the reaction of copper with dilute nitric acid. They should determine if this reaction is spontaneous. They should support their answer. (322-5, 322-6)
- Students could determine if the reaction of cadmium metal and copper II ions would be spontaneous? They should support their answer. (322-5, 322-6)
- Students could discuss various methods used to find reduction potential. They should compare theoretical and experimental reduction potential values. (214-7, 214-8, 322-5, 322-6)

Resources/Notes

MGH Chemistry, p. 733, pp. 807-808

MGH Chemistry, p. 774

Electrolytic Cells

Outcomes

Students will be expected to

- evaluate a personally constructed cell by selecting and using apparatus for electrochemistry labs (213-8, 214-16)
  - deduce from lab activities that electrochemical cells operate on the energy of spontaneous oxidation-reduction reactions
  - write half-reaction equations from their lab results
- evaluate processes used in planning, problem-solving and decision-making, and completing a task (215-7)
- evaluate the relevance, reliability, and adequacy of data and data collection methods (214-8)
- construct and test a system and troubleshoot problems as they arise (214-14)
- evaluate a constructed galvanic cell (214-16)
- explain the processes of electrolysis and electroplating (322-8)
  - explain the historical development of electrolysis, e.g. Sir Humphry Davy used electrolysis to separate table salt into sodium and chlorine

Elaborations—Strategies for Learning and Teaching

The Laboratory outcomes 213-8, 214-16, 215-7, 214-8, 214-14, and, in part, 214-16 are addressed by completing *Measuring Cell Potentials of Galvanic Cells,* CORE LAB #7. Note, a good alternative to using glass tubing as a salt bridge is a rolled piece of paper towel soaked in salt solution.

This outcome can be achieved by having students perform a lab activity where they use several different metal strips placed in aqueous solutions of the metal ions to create several half-cells. Then by using a voltmeter and a salt bridge these half-cells can be connected to each other. If the concentrations of the solutions are 1.0M and the temperature is near 25^0C the voltages obtained should match cell voltages calculated using a Standard Reduction Potentials Table. Comparisons can be made by students between the experimental and the technical cell voltages. Teachers should note that a good starting point would be to use the $Zn/Zn^{2+}//Cu^{2+}/Cu$ cell for comparison purposes. Connect these 2 half-cells to give a positive voltage (+1.10 V) on the voltmeter. This indicates electrons are flowing from the anode (Zn) to the cathode (Cu). When students connect other combinations of half-cells they may not get a reading on the voltmeter. They should realize the reaction is non-spontaneous and would require a power supply to react.

Electrolytic Cells

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could illustrate and label the parts of an electrolytic cell. They should explain how it works. (322-4)

Resources/Notes

Core Lab #7: *"Measuring Cell Potentials of Galvanic Cells"*, pp. 762-763

MGH Chemistry, p. 778

Electrolytic Cells *(continued)*

Outcomes

Students will be expected to

- explain the processes of electrolysis and electroplating (322-8) (**Cont'd**)
  - define electrolytic cells as requiring electrical energy to cause non-spontaneous oxidation-reduction reactions to occur
- illustrate and label the parts of electrolytic cells and explain how they work (322-4)
  - draw and label an electrolytic cell. Include:
    - (i) anode
    - (ii) cathode
    - (iii) salt bridge
    - (iv) power supply
    - (v) direction of flow of electrons and ions

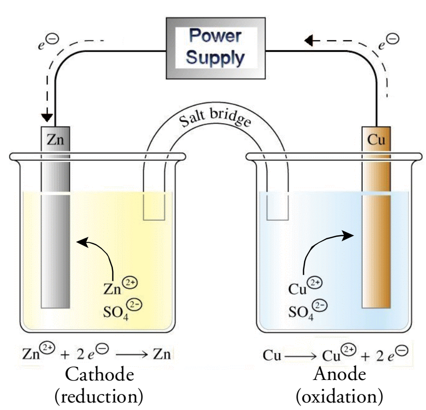
Elaborations—Strategies for Learning and Teaching

Teachers could first ask students if it might be possible to cause a non-spontaneous redox reaction to occur. Students should recognize that many reactions are reversible and this could be possible if electrons were forced to flow in the opposite direction by applying voltage greater than the cell potential , E^o. This leads to the definition of an electrolytic cell and differentiates it from electrochemical (galvanic, voltaic) cells.

At this point the teacher should briefly outline the historical changes in chemistry brought about by the discovery of the electric cell (battery) and the use of the voltage produced to cause non-spontaneous chemical reactions to occur. This leads to the early work of Sir Humphrey Davy and how electrolysis led to his discovery of several elements. This also led to the important industrial processes of electroplating and metal production purification.

Teachers could demonstrate examples of operating electrolytic cells. Simple ones would be electrolysis of water or reversing the reaction in the $Zn/Zn^{2+}//Cu^{2+}/Cu$ electrochemical cell previously demonstrated. Before the demonstration, students could calculate the required voltage to cause the redox reaction. To make the demo a practical example and show electroplating as well, an Fe nail could be substituted for the Zn electrode. The product would be a common galvanized (Zn coated), corrosion resistant nail.

Many other electrolysis examples may be used but caution should be exercised because of the toxicity of products. For example, electrolyzing $NaCl_{(aq)}$ produces poisonous $Cl_{2(g)}$. Identification of the relevant parts of the electrolyte cell should be covered at this point. Students could be required to produce a labelled diagram for the demonstration set up.



Students could perform a lab at this point which illustrates electrolysis or electroplating.

e.g. 19C - McGraw Hill - Electroplating (p. 794-795)

Electrolytic Cells *(continued)*

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could ask, "Should you use zinc, copper or neither to 'plate out' nickel metal from a nickel (II) nitrate solution?" Students should support their answers. (116-7, 322-8, 214-18)
- A sunken ship is to be lifted from the ocean bottom. Plastic bags, containing seaweed and equipped with an arrangement of inert, internal and external electrodes, are attached to the ship. Electrolysis current is applied to the electrodes, to fill the bags with hydrogen gas. Students could determine if the internal electrode is the anode or cathode. They should explain their answer. As well, students should determine what are the products at the other electrode. (322-8, 116-7)

Resources/Notes

MGH Chemistry, pp. 776-777

MGH Chemistry, p. 780

Electrolytic Cells *(continued)*

Outcomes

Students will be expected to

- explain the processes of electrolysis and electroplating (322-8) (**Cont'd**)
  - predict and write balanced half-reactions for reactions at the cathode and the anode of electrolytic cells
  - write equations for electrolytic cells from half-reactions
  - perform stoichiometry calculations related to electroplating using Q=It and Q=nF

Elaborations—Strategies for Learning and Teaching

Students should be able to predict the most likely half-reactions to occur at the anode and the cathode of an electrolytic cell. For example, the electrolysis of aqueous sodium bromide:

Anode (ox):

$$2Br \rightarrow Br_2 + 2e^- \qquad \varepsilon^o = -1.07V$$

$$2H_2O \rightarrow O_2 + 4H^+ + 4e^- \qquad \varepsilon^o = -1.23V$$

Cathode (red):

$$Na + e^- \rightarrow Na \qquad \varepsilon^o = -2.71V$$

$$2H_2O + 2e^- \rightarrow H_2 + 2OH^- \qquad \varepsilon^o = -0.83V$$

The half-reaction with the more positive value (ie: the least negative) at each electrode is most likely to occur. From the example, the overall reaction would be

$$2Br^- \rightarrow Br_2 + 2e^- \qquad \varepsilon^o = -1.07V$$

$$2H_2O + 2e^- \rightarrow H_2 + 2OH^- \qquad \varepsilon^o = -0.83V$$

$$2H_2O + 2Br^- \rightarrow H_2 + 2OH^- + Br_2 \qquad \varepsilon^o = -1.90V$$

Note: ε^o values for half-reactions are taken from the Standard Reduction Potential Table produced by the Department of Education for the public exams (values may differ from text).

Using Faraday's Law and stoichiometry, the amount of products produced by an electrolytic cell can be determined. Faraday's Law determines the number of moles of electrons which pass through a cell when a particular current is applied over a period of time. Current (in amperes) multiplied by time (in seconds) gives the total charge, in coulombs, carried by the electrons. Faraday's Constant (F=96 500 coulombs/mole) gives the charge on 1 mole of electrons, therefore by dividing, the moles of electrons are obtained.

$$n = moles\ e^- = \frac{current \times time}{Faraday's\ Constant} = \frac{I \times t}{F}$$

Using stoichiometry and the appropriate half reaction occurring, the amount of product produced can be determined. For example, calculate the moles of Mg produced when a current of 60.0 A is passed through a magnesium chloride solution for 4.00 hours.

Electrolytic Cells *(continued)*

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could research industrial applications of Faraday's Law, e.g., painting cars, galvanizing nails, silver plating and chrome plating. (212-1, 322-8)
- Students could calculate the charge on the iridium ion if a current of 0.200 A is passed through a solution of iridium bromide for 1.00 h, resulting in the deposition of 0.478 g of iridium at the cathode. (322-8)

Resources/Notes

MGH Chemistry, pp. 778-779, 781-783

MGH Chemistry, pp. 790-793, p. 834

Applications of Electrochemistry

Outcomes

Students will be expected to

- identify questions to investigate that arise from practical problems and issues (212-1)
- define and delimit problems to facilitate investigation (212-2)
- carry out procedures controlling the major variables and adapting or extending procedures where required (213-2)
- select and use apparatus and materials safely (213-8)
- compare theoretical and empirical values and account for discrepancies (214-7)
- evaluate the relevance, reliability, and adequacy of data and data collection methods (214-8)

Elaborations—Strategies for Learning and Teaching

The laboratory outcomes 212-1, 212-2, 213-2, 213-8, 214-7, 214-8 and, in part, 322-4 are addressed by completing *Electroplating,* CORE LAB #8.

Applications of Electrochemistry

Tasks for Instruction and/or Assessment

Resources/Notes

Core Lab #8: *"Electroplating"*, pp. 794-795

Applications of Electrochemistry *(continued)*

Outcomes

Students will be expected to

- define metallurgy, pyrometallurgy and hydrometallurgy using scientific principles (116-5)
- differentiate between the terms mineral and ore (ACC-9)
- define and describe the process of *flotation*, as it applies to the mining industry (ACC-10)
- define the terms flotation, slag, and leaching, as they apply to pyrometallurgy and hydrometallurgy (ACC-11)
- identify and describe, using chemical reactions and chemical equations, the purification of copper metal using an electrolytic cell

  (116-3)
- identify the risks and benefits to society and the environment of pyrometallurgy (118-1)
- describe the process of electrowinning pure nickel from a nickel(II) subsulfide matte, used in the industrial production of pure nickel (ACC-12)
- evaluate the processes of pyrometallurgy versus hydrometallurgy for the production of copper and nickel from two representative minerals, and determine which industrial process would be best (118-4)

Elaborations—Strategies for Learning and Teaching

The CORE STSE component of this unit incorporates a broad range of Chemistry 3202 outcomes. More specifically, it targets (in whole or in part) 116-5, ACC-9, ACC-10, ACC-11, 116-3, 118-1, ACC-12, and 118-4. The STSE component, *From Mineral to Metal: Metallurgy and Electrolytic Refining,* can be found in Appendix A.

Applications of Electrochemistry *(continued)*

Tasks for Instruction and/or Assessment

Resources/Notes

Core STSE #4: *"From Mineral to Metal: Metallurgy and Electrolytic Refining"*, **Appendix A**

Energy Production and Electrochemical Cells

Outcomes

Students will be expected to

- compare electrochemical and electrolytic cells in terms of energy efficiency, electron flow/ transfer, and chemical change (322-7)

Elaborations—Strategies for Learning and Teaching

Electrochemical cells (galvanic or voltaic cells) produce a particular amount of electrical energy spontaneously. All of this energy can then be used to produce heat or do work. For example, the $Zn/Zn^{2+}//Cu^{2+}/Cu$ cell produces 1.10V. However, to cause a non-spontaneous electrolytic reaction to occur requires electrical energy in excess of the energy produced in the spontaneous reverse reaction. For example, in the $Zn/Zn^{2+}//Cu^{2+}/Cu$ cell a voltage >1.10V must be applied to make the reverse non-spontaneous process proceed. If a voltage of 2.00 V is applied only 0.90V are available to create heat or do work. This makes a spontaneous redox reaction (i.e., an electrochemical cell) more efficient than non-spontaneous redox reactions (i.e., an electrolytic cells). For example, to recharge a 12V car battery, a voltage >12V must be applied by the charger.

When a spontaneous redox reaction, occurring in an electrochemical cell, is compared to its reverse non-spontaneous redox reaction, occurring in an electrolytic cell, there are differences in the reactions occurring at the electrodes - the anode and cathode are reversed.

E.g., Spontaneous (electrochemical) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Electrode	Reaction
Oxidation: Zn (anode)	$Zn \rightarrow Zn^{2+} + 2e^-$
Reduction: Cu (cathode)	$Cu^{2+} + 2e^- \rightarrow Cu$

E.g., Non-spontaneous (electrolytic) $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$

Electrode	Reaction
Oxidation: Cu (anode)	$Cu \rightarrow Cu^{2+} + 2e^-$
Reduction: Zn (cathode)	$Zn^{2+} + 2e^- \rightarrow Zn$

This is the result of electron flow being reversed by the applied voltage in the electrolyte cell.

- define primary battery, secondary battery and identify common examples of each. Include:
  (i) dry cell
  (ii) alkaline
  (iii) button
  (iv) lead storage
  (v) Ni-Cd

Batteries are a common source of energy used to power many personal devices. Some are disposable (primary batteries - not rechargeable), for example, dry cells, while others can be used again (secondary batteries - rechargeable). Teachers and/or students could easily produce examples of these in any classroom. e.g., calculator batteries, watch batteries, etc.

Students are not required to memorize the particular half-cell reactions for each battery.

Energy Production and Electrochemical Cells

Tasks for Instruction and/or Assessment

Performance

- For a week/day, students could keep a record of everything they use that is powered by batteries. They should record the device used and the number and the type of batteries it contains. (116-7, 214-18, 118-4, 322-7)
- Students could research and illustrate the familiar flashlight battery, either the acidic version with the central carbon rod or the alkaline version. (116-6, 322-4, 322-7)
- Students could write a newspaper article offering ways to reduce the amount of waste produced by batteries. (118-4, 214-18, 322-7)
- Students could write a short essay about technology that was not and could not have been available before the development of the nickel-cadmium battery. (118-4, 116-7, 214-18, 322-7)

Presentation

- In small groups, students could research the environmental effects of different types of batteries. They should analyze both the production and waste costs and share their research process and findings with their class. (322-7, 118-4, 116-7, 214-18)

Resources/Notes

MGH Chemistry, pp. 764-766

MGH Chemistry, pp. 747-766, 787-788

Energy Production and Electrochemical Cells *(continued)*

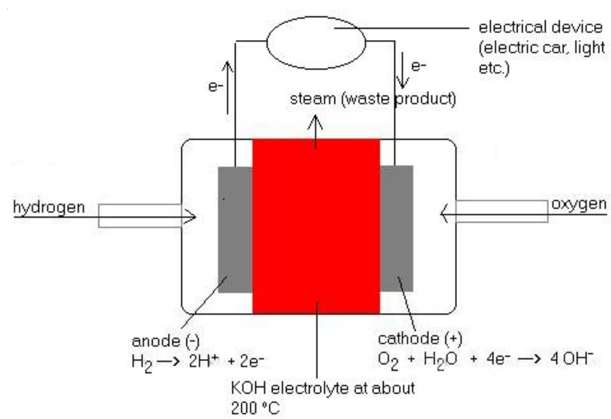
Outcomes

Students will be expected to

- explain the production of energy in a hydrogen fuel cell (322-9)
  - compare the efficiency of burning hydrogen in a fuel cell with burning it in a combustion reaction

Elaborations—Strategies for Learning and Teaching

Research on fuel cell technology has made major progress in recent years and prototype cars driven by fuel cell technology are being tested. An understanding of how fuel cells work is an important future application of electrochemistry. Students should understand the electrochemical reactions occurring in the hydrogen fuel cell only as described in the textbook. A comparison of the energy efficiency of using the H_2 in a fuel cell ($2H_2 + O_2 \rightarrow 2H_2O$ + electr.) compared to burning the H_2 in a combustion reaction ($2H_2 + O_2 \rightarrow 2H_2O$ + heat) will illustrate the high efficiency of fuel cells.



Energy Production and Electrochemical Cells *(continued)*

Tasks for Instruction and/or Assessment

Paper and Pencil

- Students could describe how electrical energy is produced in a hydrogen fuel cell. (322-9)
- Students could investigate the benefits and drawbacks associated with hydrogen fuel cells. (322-9)

Resources/Notes

MGH Chemistry, pp. 802-803