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Identifying, attributing, and dispelling student misconceptions in electrochemistry

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Identifying, attributing, and dispelling student misconceptions in electrochemistry

by

Michael James Sanger

A dissertation submitted to the graduate faculty
in partial fulfillment of the requirements for the degree of
DOCTOR OF PHILOSOPHY

Departments: Chemistry; Curriculum and Instruction
Co-majors: Chemistry; Education
Major Professors: Thomas J. Greenbowe and Lynn W. Glass

Iowa State University
Ames, Iowa
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For the Graduate College

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GENERAL INTRODUCTION

Dissertation Organization

This dissertation consists of six chapters: the first chapter is a review of the chemical education literature concerning oxidation-reduction reactions and electrochemistry, the second through fifth chapters contain the results of my research on student misconceptions in electrochemistry as they were submitted for journal publication, and the sixth chapter contains an opinion paper concerning the possible mismatch between Science-Technology-Society/ChemCom-based high school chemistry courses and traditional introductory college-level chemistry courses. Literature citations, tables, and figures are numbered independently in the literature review and in each paper. Following the papers is a general summary of this work and suggestions for future research.

Summary of Research

The four research studies that comprise this dissertation are primarily concerned with student misconceptions in electrochemistry. The first study was a replication, with additions, of a clinical interview study performed by Garnett and Treagust (*1*). Garnett and Treagust interviewed 32 high school students in Western Australia using a set of semi-structured interview questions to probe student understanding of electrochemical and electrolytic cells and reported a list of misconceptions demonstrated by these students. The replication study used these semi-structured interview questions with modifications and included new interview questions on concentration cells.

The second study described in greater detail and elaborated on one of the student misconceptions reported in the first study. This study briefly discussed student misconceptions concerning current flow in electrolyte solutions and the salt bridge, identified chemistry textbooks as possible sources of these misconceptions, and discussed the use of computer animations and a confrontational teaching method as an effective method of preventing and dispelling these misconceptions. There was some criticism of this study from my colleagues in the science education seminar that the comparison of my subjects to those of Ogude and Bradley (*2*) was not statistically rigorous and that the differences in the performance of these two groups could be explained by other confounding variables.

The third study investigated introductory college-level chemistry textbooks as a possible source of student misconceptions in electrochemistry. The oxidation-reduction and electrochemistry chapters of ten chemistry textbooks were analyzed for vague or misleading statements and illustrations that could lead to student misconceptions reported in the first study and the results were reported in this study. As a result of this analysis, several suggestions for textbooks authors were reported.

The fourth study investigated the effects of computer animations and conceptual change instruction on students' conceptions concerning the flow of current in electrolyte solutions. Computer animations had little effect on students' conceptions, but conceptual change instruction significantly decreased the proportion of student responses suggesting that electrons flow in aqueous solutions.

Also included in this dissertation is a provocative opinion concerning the possible mismatch between high school Science-Technology-Society (STS) or ChemCom chemistry courses and traditional introductory college-level chemistry courses, which originated as a preliminary examination question concerning the advantages and disadvantages of the STS movement in science education.

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1. Garnett, P. J.; Treagust, D. F. *J. Res. Sci. Teach.*, **1992**, *29*, 1079-1099.

2. Ogude, A. N.; Bradley, J. D. *J. Chem. Educ.*, **1994**, *71*, 29-34.

CHEMICAL EDUCATION LITERATURE CONCERNING OXIDATION-REDUCTION REACTIONS AND ELECTROCHEMISTRY

The chemical education literature concerning oxidation-reduction reactions and electrochemistry can be divided into three categories: Suggestions or personal opinions promoting a particular method of teaching oxidation-reduction reactions or electrochemistry, descriptions of real-world examples of oxidation-reduction reactions, and empirical data aimed at identifying student difficulties or misconceptions related to oxidation-reduction reactions or electrochemistry.

Opinions Promoting Particular Teaching Methods

A majority of the articles in the chemical education literature concerning oxidation-reduction reactions and electrochemistry topics are suggestions about particularly effective teaching methods that are based largely on personal opinions or anecdotal evidence on the part of the authors and are not based on empirical evidence. Suggestions concerning effective teaching methods for oxidation-reduction reactions fall into three categories: Definitions of oxidation and reduction processes, assignment of oxidation numbers, and methods for balancing oxidation-reduction reactions. Suggestions regarding effective teaching methods for electrochemistry concepts can be organized into the following categories: Descriptions of electrode charges, methods for calculating cell potentials, advanced electrochemistry topics, and suggestions based on empirical data.

Definitions of Oxidation and Reduction Processes. Several authors have advocated different methods for identifying oxidation and reduction reactions. Antoine Lavoisier first used the term "oxidation" to mean "reaction with oxygen" and used the term "reduction" to mean "the removal of oxygen" (*1*). After the discovery of the electron and the introduction of Bohr's model of the atom, chemical reactions were gradually reinterpreted in terms of electron activity and the term "oxidation" came to mean "a loss of electrons" while the term "reduction" came to mean "a gain of electrons". With the introduction of oxidation numbers as a way to account for electrons in oxidation-reduction reactions, the term "oxidation" has also been defined as "an increase in oxidation number" and the term "reduction" has been defined as "a decrease in oxidation number".

Herron (*2*) and Sisler and VanderWerf (*3*) recommended the use of oxidation numbers for determining which elements have been oxidized or reduced. Herron (*2*) argued that defining oxidation as an increase in oxidation number and reduction as a decrease in oxidation number should appear more logical to students since it makes use of the traditional meaning of the word "reduction". Defining oxidation and reduction in terms of electron transfer may suggest to students that all oxidation-reduction reactions involve an exchange of electrons (*3*). Sisler and VanderWerf (*3*) proposed several alternative definitions for identifying oxidation-reduction reactions, but these definitions would include simple precipitation, acid-base, and Lewis acid-base reactions. The authors concluded that oxidation and reduction processes are best defined in terms of changes in oxidation states.

Several authors responded to Herron's discussion about the definition of oxidation and reduction processes (*2*). Schug (*4*) pointed out that the rules for assigning oxidation states are completely arbitrary and frequently bear no relation to the actual charge distribution with a molecule. Koellner (*5*) stated that oxidation is most properly defined as the *apparent* loss of electrons and that oxidation states are the means for making observations about oxidation-reduction reaction, not the end in themselves. Loveridge (*6*) suggested that both the oxidation number and the electron transfer definitions of oxidation and reduction processes should be taught to students. Janke (*7*) stated that defining oxidation as an increase in oxidation number and reduction as a decrease in oxidation number is based upon artificial and arbitrary assignments of oxidation numbers. It is important to distinguish between the concept of oxidation numbers (which is an artificial concept) and the concept of oxidation and reduction as processes that are always related to the chemistry of the reaction under study.

Goodstein (*8*) argued against the use of oxidation states because the rules for assigning oxidation states are largely arbitrary. Instead, she proposed a new definition of an oxidation-reduction reaction. Goodstein argued that only reactions which are driven by the electronegativity difference of the atoms undergoing oxidation state changes should be considered oxidation-reduction reactions. Reactions which are not driven by this electronegativity difference should not be considered oxidation-reduction reactions. As examples, the hydrolysis of chlorine ($Cl_2 + H_2O \rightarrow HCl + HOCl$) and the bromination of ethylene ($C_2H_4 + Br_2 \rightarrow CH_2BrCH_2Br$) would not considered oxidation-reduction reactions by this author. Sisler and VanderWerf (*3*) argued that defining oxidation states using electronic conditions within the molecules or ions is just as futile as defining oxidation and reduction in terms of electron transfer.

In the course of interviewing students to determine common misconceptions concerning oxidation-reduction reactions, Garnett and Treagust (*9*) discovered that many students had difficulty identifying oxidation-reduction reactions. Students who had little difficulty identifying these reactions consistently used oxidation state changes. However, students experiencing difficulty tended to use several methods to identify oxidation-reduction reactions, including oxidation states, gain/loss of oxygen, loss/gain of hydrogen, and loss/gain of electrons. These students often appeared uncertain as to which definition to apply and tended to use the definition that could be most easily applied to a particular equation. Garnett, Garnett, and Treagust (*10*) suggested that if multiple definitions are to be used, then the limitations for each definition should be clearly pointed out to the students.

Assignment of Oxidation Numbers. Kolb (*1*) provided a list of hierarchical rules for determining oxidation states for each element in a compound. Oxidation numbers are positive or negative numbers assigned to individual atoms for purposes of electron "bookkeeping". Although the assignment of oxidation numbers is arbitrary, it provides a consistent method for determining which elements have been oxidized or reduced and how many electrons have been transferred. Holleran and Jespersen (*11*) provided a simpler list of rules for assigning oxidation states that addresses all situations that would be encountered in general chemistry. Neman and Logan (*12*) argued that a distinction should be made between oxidation number, which refers to the "degree of oxidation" of the atom, and valence, which describes the number of atoms bound to the atom.

Although oxidation states are arbitrarily assigned, there is some relationship between oxidation numbers and chemical reactivities of elements. As an example, Stonestreet (*13*) pointed out that although fluorine is always assigned an oxidation state of -1 in its compounds, the reactivity of hypofluorous acid (HOF) suggests that it is as strong an oxidizing agent as F_2 ($O.N._F = 0$) and XeF_2

(O.N.$_F$ = -1) and therefore the F atom should be considered to have an oxidation state of +1 instead of -1. Several authors voiced an objection to this line of reasoning. Geanangel (*14*) pointed out that oxidation numbers are arbitrary and that neither assignment is likely to be an accurate description of this compound. He also pointed out that other strong oxidizing agents containing fluorine (such as ClF_3 and AgF_2) have fluorine with oxidation numbers of -1. Woolf (*15*) argued that the calculated electronegativity for OH of 2.89 suggests that the HOF is likely to contain a net negative charge on F, consistent with a -1 oxidation state for fluorine. Smith (*16*) argued that oxidation numbers do not determine relative oxidizing abilities and cited examples of compounds with the same oxidation states but different oxidizing abilities.

Traditional rules used to calculate oxidation numbers often result in fractional oxidation numbers (most commonly in covalent and organic compounds), which make no sense according to modern atomic theory. Kauffman (*17*) proposed a method of calculating oxidation numbers that he called the "exploded structure method" (esm). A Lewis dot structure is drawn for the compound and then the atoms are drawn further apart, with each pair of bonding electrons associated with the more electronegative atom (the bond is split equally if the atoms are the same). Oxidation numbers are determined by subtracting the number of electrons around each atom from the number of valence electrons associated with the free atom. This method determines oxidation states for each atom; the traditional method only determines "average" oxidation numbers for each atom. The esm method results in two oxidation numbers for the carbon atoms in CH_3COOH: –3 for the methyl carbon and +3 for the carboxylic acid carbon. The traditional method results in an average oxidation number of zero for both carbon atoms. Goodstein (*18*) proposed a similar method for determining oxidation numbers of carbon atoms in organic molecules. A Lewis structure is drawn for the molecule and oxidation numbers are assigned for each carbon atom as follows: Each carbon atoms starts out at zero; for every bond to a more electronegative atom (N, O, Cl, etc.), the oxidation number is decreased by one; for every bond to a less electronegative (H, Li, etc.), the oxidation number is increased by one. This method also has the advantage of calculating oxidation numbers for each carbon atom instead of calculating an average value for all of the carbon atoms in the molecule.

Woolf (*19*) identified several other limitations of calculating oxidation numbers using simple electronegativity arguments. Using the esm method with compounds having more than one resonance structure can result in different oxidation numbers for the same atom in the different resonance forms. Woolf also argued that using atomic electronegativities to assign oxidation numbers is less sophisticated than calculating "mean" electronegativities. Traditional and the esm methods of assigning oxidation numbers result in the same oxidation numbers for the I atoms in methyl iodide (CH_3I) and trifluoromethyl iodide (CF_3I) and the C atoms in cyanogen iodide (CNI) and cyanogen chloride (CNCl). Calculating mean electronegativities (which yields values of 3.49 for CF_3, 3.00 for Cl, 2.69 for CN, 2.56 for I, and 2.35 for CH_3), however, correctly explains the fact that the atoms mentioned above have a net negative charge in their former compounds and a net positive charge in their latter compounds.

Kjonaas (*20*) proposed a method of recognizing and quantifying oxidation-reduction reactions in organic chemistry called the Number of Oxidations Relative to Methylene (NORM). The NORM (*N*) is defined as the net number of two-electron oxidations required to synthesize the compound from methylene groups. The NORM of non-cyclic alkanes is -1, which is interpreted as suggesting that a two-electron reduction of the methylene groups (i.e., the addition of H_2) is required to synthesize alkanes. If no oxidation-reduction process occurs, the NORM of the reactant and the product are the same. If ΔN is positive, the reactant has been oxidized; if ΔN is negative, the reactant has been reduced. If more than one reactant or product is involved in the reaction, the sum of NORM's for the reactants and products are compared. The change in NORM's can be useful in proposing mechanisms, determining whether an added reagent acts as a catalyst or as an oxidizing/reducing agent, and predicting unidentified products.

Methods for Balancing Oxidation-Reduction Reactions. Kolb (*1*) briefly discussed several methods for balancing oxidation-reduction reactions including the algebraic method, simple inspection, the oxidation number method, the ion electron method, and the matrix method. In balancing oxidation-reduction reactions using the algebraic method, each reactant is assigned an algebraic coefficient. Mathematical formulas relating these coefficients are created and solved by requiring mass balance for each atom in the equation and charge balance (*1, 21, 22*). Advantages of this method include the fact that it is not necessary to determine which element is oxidized or reduced (*1*), reactions can be balanced without determining oxidation number (*22*), and reactions with more than one oxidant or reductant can be solved easily (*22*).

Kolb (*23*) reported a simple method of balancing oxidation-reduction reactions by inspection proposed by Jason Ling. This method has three steps: (1) Elements that appear only once on each side of the equation and in equal numbers of both sides are marked with arrows; (2) Elements that appear only once on each side of the equation but have unequal numbers of atoms are balanced first; (3) Balance elements that must be present in a constant ratio on each side of the equation. If these three steps do not work, algebraic coefficients for the unbalanced elements are assigned and these coefficients are solved for algebraically. This method can also be used for balancing ionic equations, using fictitious cations (Q^+) or anions (X^-) that are removed from the final equation.

Using the ion electron method, each reaction is separated into two half-reactions that are balanced separately (*1*). In each half-reaction, O atoms are balanced by adding H_2O, H atoms are balanced by adding H^+ ions, and charge is balanced by adding electrons (e^-). After both half-reactions are balanced, they are multiplied by the appropriate coefficients to cancel the electrons from each half-reaction and are added together. Reactants that appear on both sides of the equation are canceled. Reactions that occur in basic solutions are balanced as above, and enough OH^- ions to completely react the H^+ ions present in the reaction are added to both sides of the equation. The H^+ and OH^- ions combine to produce H_2O and any H_2O molecules that appear on both sides of the equation are canceled. Garcia (*24*) proposed a slightly modified version of the ion electron method that uses O atoms to balance oxidation-reduction reactions. In neutral solutions, H atoms are balanced in each half-reaction using H_2O molecules and O atoms are balanced using O atoms. The two half-reactions are added after multiplying by appropriate coefficients to cancel the O atoms. In reactions containing ions, charge is balanced using either H^+ or OH^- ions.

Balancing oxidation-reduction reactions using the oxidation number method requires determining the oxidation states of the elements that are oxidized and reduced (*1*). Once the number of electrons lost by the reductant and the number of electrons gained by the oxidant are determined, these compound are multiplied by coefficients to make the number of electrons lost and gained the same. After the oxidation and reduction processes are balanced, the other atoms are balanced by inspection. Davis (*25*) proposed a method of balancing oxidation-reduction reactions that is a combination of the ion electron and the oxidation number methods. In this method, oxidation numbers are assigned and the elements that are oxidized and reduced are identified. The reaction is then split into two half-reactions that are balanced separately. First, the number of electrons lost or gained are added to the half-reaction as e^-. Charge is then balanced using H^+ or OH^- ions and O atoms are balanced using H_2O molecules. The two-half reactions are multiplied by appropriate coefficients to cancel the electrons and are added together, canceling any reactants that appear on both sides of the equation. This procedure works equally well for acidic and basic solutions and students found this method easier to use than the ion electron or the oxidation number methods.

Blakley (*26*) discussed the use of a simple computerized program to solve oxidation-reduction reactions using the matrix method. The matrix method has the advantage that it can determine relatively quickly whether an oxidation-reduction reaction, as written, cannot be balanced, whether it has only one distinct solution, or whether it has a class of solutions. It can also be used to balance incredibly complex oxidation-reduction reactions with 20 or more reactant and products containing 20 or more elements. Swinehart (*27*) pointed out that although a computerized program using the matrix method can solve complex oxidation-reduction reactions, these reactions can also be balanced by inspection.

Kolb (*1*) listed several problems that students experience with oxidation-reduction reactions including superscripts, dual function reagents, ionic equations, incomplete equations, disproportionation, and more than two oxidants and reductants. Blakley (*26*), Kolb (*28*), and Carrano (*29*) pointed out two particularly difficult problems in balancing oxidation-reduction reactions. Some reactions can be balanced as a linear combination of two or more distinct balanced equations (*26, 28*). The other difficulty occurs when the same element (in four distinct oxidation states) serves as both the oxidant and the reductant (*28, 29*).

Descriptions of Electrode Charges. The electrical charges associated with electrodes in galvanic and electrolytic cells have been the source of great debate in the chemical education literature (*30-33*). Moran and Gileadi (*30*) described the generally-accepted polarities of electrodes in galvanic and electrolytic cells: In galvanic cells, electrons flow from the negative electrode (anode) to the positive electrode (cathode) and in electrolytic cells, electrons flow from the positive electrode (anode) to the negative electrode (cathode). This change in polarity can be confusing to students (*31*) and as a result, several authors have suggested the use of alternate sign conventions for electrodes.

Al-Soudi (*31*) proposed that cathodes always be referred to as electron rich, the source of electrons, and the negative electrode. Al-Soudi suggested that this assignment will prevent the confusion student experience when trying to explain why negatively-charged electrons flow from the positive to the negative electrode in electrolytic cells. However, in her own description of this new method, she assigns the zinc electrode of a Daniel cell as the source of electrons and as electron rich, thus implying that the zinc electrode is the cathode.

MacDonald (*32*) argued that students have difficulty determining electrode charges because instructors fail to distinguish between a terminal (a part of the voltmeter) and an electrode (a conductor where electrons enter or leave a cell) and do not recognize that the terminal of an electrode always has a sign opposite to that of the electrode. MacDonald defined the cathode as the negative electrode and the anode as the positive electrode. In describing a galvanic zinc-magnesium cell, he pointed out that before the electrodes are connected, the magnesium electrode undergoes more dissociation than the zinc electrode and therefore has a larger negative charge (cf. *34-35*). However, when the electrodes are connected, some of the electrons on the magnesium electrode flow onto the zinc electrode, making it more negative and making the magnesium electrode more positive, consistent with his definitions. MacDonald (*33*) rejected the argument that the cathode in a galvanic cell must be positive because electrons flow from the negative anode toward the positive cathode by pointing out that anions would not be attracted to a negatively-charged anode and that cations would not be attracted to a positively-charged cathode.

Methods for Calculating Cell Potentials. Most of the chemical education articles concerning the calculation of cell potentials advocate the use of the difference method for calculating cell potentials. Birss and Truax (*35*) argued that cell potentials should be calculated as differences (i.e., $E_{cell} = E_{red}(\text{cathode}) - E_{red}(\text{anode})$), instead of changing the sign of the reduction potential for the oxidation half-reaction and adding it to the reduction potential for the reduction half-reaction ($E_{cell} = E_{red} + E_{ox}$). The latter method causes student confusion with respect to the relative oxidizing-reducing capabilities of half-reactions, while the former method retains the sense of the "potential difference" between two half-cells. Stevenson (*36*) pointed out an additional difficulty with the additive method. In convincing students that the sign of a half-cell potential changes when the equation is reversed, students are led to believe that half-cell potentials obey Hess' Law. This explains why some students want to multiply half-cell potentials by coefficients when they multiply the equation by this coefficient. It also explains why students have difficulty adding two half-cell potential to derive another half-cell potential. West (*37*) described an obvious difficulty in using the equation $E_{cell} = E_{right} - E_{left}$ to calculate cell potentials in lecture: The right-hand half-cell for the teacher is the left-hand cell for the students and vice versa. The left-right convention also assumes that it is already known which half-cell is the anode and the cathode before the cell is set up.

Haight (*38*) pointed out that, using gas-phase calculations concerning the hydration of electrons, absolute reduction potentials can be calculated. The standard hydrogen electrode (SHE), which is now arbitrarily set equal to 0.00 V, would have an absolute half-cell potential of 2.67 V. Redefining the half-cell potentials would make most reduction potentials positive—only the most active metals such as Na or Li would still have negative half-cell potentials, which suggests that these cations would not be strong enough oxidizing agents to oxidize solvated electrons. Moran and Gileadi (*30*) proposed adding a constant value of 3.00 V to each reduction potential. They argued that this would make calculating cell potentials easier and would make the differences in oxidizing or reducing powers more obvious.

Several authors (*39-41*) have proposed the use of horizontal or vertical number lines to plot potential differences of cell reactions. Vella (*39*) described the use of an electropotential axis to determine the relative activities of metals. The electropotential axis is a horizontal number line in which cell potentials are plotted with the negative electrode at the left and the positive electrode at the right. The axis makes it clear that relative potentials can be measured, but that absolute potentials cannot be determined unless an arbitrary zero point (like the SHE) is assigned. The axis can be used to determine the cell potential, the direction of electron flow, and the cell reaction of a system in question.

Borrell and Dixon (*40*) provided a description of the electrode potential diagrams that are typically used in the Hill-Bendall or Z-scheme concerning the overall process of photosynthesis. In electrode potential diagrams, half-cell potentials are plotted on a vertical number line with negative potentials on top and positive potentials on bottom. Although the scale may seem backwards for cell potentials, it is correct for the free energy (ΔG), whose sign is opposite of that for the cell potential. In these drawings, electrons flow spontaneously from top to bottom, but can flow from bottom to top under electrolytic conditions. Runo and Peters (*41*) described the use of a potential ladder to determine the spontaneity of cell reactions. Their number line differs from the electrode potential diagram proposed by Borrell and Dixon (*40*) in that positive half-cell potentials are plotted at the top and negative potentials are plotted at the bottom of the potential ladder. In this case, electrons flow spontaneously from bottom to top, but can be forced to flow from top to bottom in electrolytic cells. The relative placement of the half-cell potentials in the methods outlined by Borrell and Dixon (*40*) and Runo and Peters (*41*) are slightly affected by changing the concentrations of the species present. Moran and Gileadi (*30*) and Runo and Peters (*41*) also demonstrated how overpotential can be explained using potential ladder diagrams.

Advanced Electrochemistry Topics. Some of the teaching methods proposed in the chemical education literature are too complex or are concerned with topics that are not typically in freshman-level chemistry courses. Chambers (*34*) and Birss and Truax (*35*) advocated the use of current-potential curves to explain to students what happens to the cell potential when current is allowed to flow in the electrochemical cell. Current-potential curves plot half-cell potentials on the horizontal axis (negative potentials on the left, positive potentials on the right) and current on the vertical axis (anodic or oxidizing currents on top, cathodic or reducing currents on bottom). Chambers (*34*) also discussed the topic of membrane potentials, which can be used to explain ion selective electrodes, including glass electrodes used to measure pH.

In his advanced description of electrochemistry, Faulkner (*42*) came up with five simple statements to help students appreciate the heterogeneity of electrochemical systems: (1) Electrochemical systems are not homogeneous, (2) Many reactions can happen at once, (3) Current is an expression of rate, (4) Potential is an expression of electron energy, and (5) One cannot control both current and potential simultaneously. Bockris (*43*) adapted his discussion to junior-level college students enrolled in a course entitled "Kinetics at Charged Interfaces". Bockris discussed several advanced topics: Measuring potential difference changes at the interphase, analyzing absolute potential differences, metal-metal potential differences and the electron overlap potential difference, thermodynamics of electrified interphases, and the interphase structure.

Suggestions Based on Empirical Data. Few of the opinions concerning teaching methods mentioned up to this point have been based on empirical data. However, several authors (*9-10, 44-45*) have provided a list of proposed teaching methods based on the analysis of student responses to interview protocols and conceptually-based multiple-choice questions. The first implication reported by these authors was that chemistry and physics should be treated as related disciplines (*9-10, 45*). Physics and chemistry courses teach a different model of current flow (physics teaches the flow of holes as current; chemistry teaches the flow of electrons as current) and Garnett and Treagust (*9*) suggested that physics courses should teach the electronic model of current flow. Ogude and Bradley (*45*) pointed out that physics courses tend to teach the flow of current as the migration of electrons, even in electrolyte solutions. These authors also cautioned about the careless or inappropriate use of language (*10, 44-45*). Each of these articles listed examples of words or phrases concerning electrochemical cells that are vague or subject to misinterpretation. The use of multiple definitions or models can be confusing to students (*9-10*). These include the multiple definitions of oxidation and reduction reactions (discussed above) and the conflicting chemical and physical models of current flow. Garnett and Treagust (*44*) cautioned against the inappropriate application of unqualified general statements, such as "an electrolytic cell works on the reverse principle of a galvanic cell" and "like charges repel and unlike charges attract".

Real-World Examples of Oxidation-Reduction Reactions

Although most introductory college-level chemistry textbooks contain some information regarding real-world examples or applications of oxidation-reduction reactions, articles in chemical education journals also provide the learner with examples of chemistry used in everyday situations. This review discusses real-world examples of oxidation-reduction reactions regarding corrosion reactions, combustion reactions, photography, batteries and fuel cells, and metallurgy and electrolysis.

Corrosion Reactions. A corrosion reaction is typically defined as a chemical reactions in which a metallic species is oxidized. Slabaugh (*46*) used standard oxidation potentials to explain the corrosion reactions of iron. Passive iron (created by placing iron in concentrated nitric acid) appears to be coated with a thin oxide layer (represented as $Fe{\cdot}O{\cdot}O_2$) which protects the iron and raises the reduction potential of iron. Metal alloys of iron (including those with Ni or Cr additives) also have a higher reduction potential than iron alone, and this is generally attributed to the formation of protective metal oxide surface films which slow the rate of dissolution of iron (*47*). It is generally agreed that the corrosion of iron occurs by the following electrochemical processes (*46-47*): Electrons flow from anode to cathode as iron is oxidized to Fe^{2+} at the anode. Oxygen is reduced at the anode to produce hydroxide ions. The Fe^{2+} ions are rapidly oxidized to Fe^{3+} ions in an oxygen-rich atmosphere, and the result is hydrated $Fe(OH)_3$, or rust. Slabaugh also discussed the corrosion of aluminum, which is largely prevented due to the formation of a stable (and adherent) Al_2O_3 oxide layer. Chloride ions can disrupt the stable oxide coatings that typically appear on aluminum and therefore aluminum metal appears to corrode in the presence of chloride ions.

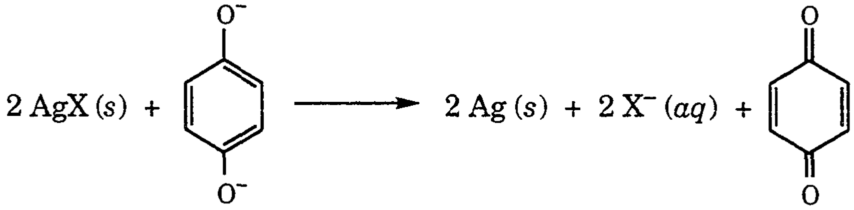
DeLorenzo (*48*) included a list of several real-world examples of corrosion reactions that could be explained using chemical formulas. These include the corrosion of the iron armature bars in the Statue of Liberty by its copper skin

($2\ Fe + 3\ Cu^{2+} \rightarrow 2\ Fe^{3+} + 3\ Cu$), the formation of a gaseous hydrogen bubble (*49*) inside the core of the reactor at Three Mile Island ($Zr + 2\ H_2O \rightarrow ZrO_2 + 2\ H_2$), the disintegration of aluminum water pumps that were attached to cast iron engine blocks in American automobiles in the 1960's ($Al + Fe^{3+} \rightarrow Al^{3+} + Fe$), to name a only a few.

Combustion Reactions. In this review, a combustion reaction is defined as a reaction in which a non-metallic species is oxidized. The carbon cycle (*50*) describes how carbon is stored and released (as CO_2 in the air, $CaCO_3$ in the ocean, and in the form of plants and animals) as the result of oxidation and reduction reactions. A similar cycle also exists for nitrogen and demonstrates the importance of these oxidation-reduction reactions in the biosphere. The formation of sugars from the photosynthesis of carbon dioxide and water, and the subsequent respiration of these sugars into CO_2, water, and energy represents the food cycle. The combustion of fossil fuels (e.g., $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$) is another example of oxidation-reduction reactions. Explosives are single chemical substances that contain strong oxidizing agents and strong reducing agents in the same molecule or are mixtures of a strong oxidizing agent and a strong reducing agent (*50*). Nitroglycerine, $NO_2CH_2CH(NO_2)CH_2NO_2$, is an example of an explosive and contains a strong reducing agent (the C_3H_5 framework) and a strong oxidizing agent (the NO_2 fragments).

Photography. The chemical processes that occur in converting light into visual images span several disciplines, including solid state chemistry, photochemistry, electrochemistry, coordination chemistry, kinetics and catalysis, and organic chemistry (*51*). The silver halide emulsion typically used in photography consists of sub-micron sized crystals, which tend to be non-stoichiometric silver chlorobromides and bromoiodides. The general steps of image formation in the silver halide emulsion include forming the latent image by irradiating the silver halide emulsion, making the latent image permanent (using a reducing agent), and removing the unreacted emulsion (*51*). Since the first two steps involve oxidation-reduction reactions, these reactions will be discussed in greater detail.

Silver halides exhibit *n*-type photoconductivity (*51*); the absorption of a quantum of radiation by a silver halide crystal occurs with the transfer of an electron from the valence to the conduction band and creation of a positive hole in the valence band (this corresponds to a "free" halogen atom). An atom of silver formed serves as an electron trap or a hole trap, which results in a two-atom aggregate and finally a four-atom aggregate, which is the smallest effective latent image size (*51*). The latent image is transformed into a permanent silver image using a mild reducing agent such as hydroquinone, *p*-aminophenol, or *p*-phenylenediamine. A balanced equation for the reaction of silver halide and hydroquinone is shown here. Recent results (*51*) suggest that two methods of



silver reduction at the latent image site are important in forming a permanent image. In the early autocatalytic phase of image development, the triple-phase theory appears to be important, in which reduction occurs at the boundary of the latent image (silver), the silver halide, and the solution. As the reaction proceeds and the latent image reaches a certain size, the electrode theory becomes more prominent. The electrode theory treats the developing grain as a short-circuited electrochemical cell, with the latent image as both the cathode (silver-silver halide interface) and the anode (silver-solution interface). Guida and Raber (*52*) and Simon (*53*) describe the chemistry of color photography, which differs from traditional black and white photography in its use of organic dyes. Both articles include detailed descriptions (and chemical formulas) of the particular organic dyes used in color photography.

Batteries and Fuel Cells. Alkire (*47*) described in detail the reactions occurring in the lead storage battery. At the negative pole, or anode, porous lead is oxidized to lead sulfate ($PbSO_4 + 2\ H^+ + 2\ e^- \rightarrow Pb + H_2SO_4$, $E° = -0.356$ V), while lead dioxide, an *n*-type semiconductor, is reduced to lead sulfate at the positive pole, or cathode, ($PbO_2 + 2\ H^+ + H_2SO_4 + 2\ e^- \rightarrow PbSO_4 + 2\ H_2O$, $E° = 1.685$ V), resulting in the following reaction: $Pb + PbO_2 + 2\ H_2SO_4 \rightarrow 2\ PbSO_4 + 2\ H_2O$, $E°_{cell} = 2.041$ V. Highly porous surfaces are needed to promote high reaction rates from small volumes and additives are added to maintain the porous structure during cycling so that it stays in place. Sulfuric acid is consumed during discharge and the diffusion of H_2SO_4 within the porous electrodes plays an important role on discharge rate. After a few hard cranks from a weak battery on cold winter days, it appears to die; however, if it is allowed to set for about ten minutes, the H_2SO_4 that was depleted from the electrode pores will have had time to diffuse back in the electrodes and the battery appears to recover (*47*).

Sammells (*54*) discussed several batteries that are presently being investigated as energy storage devices. There are several major technical features that are sought in storage batteries. These features include the abilities to (1) deliver high energy densities, (2) deliver high power densities for practical sustainable periods, (3) achieve high electrochemical energy efficiencies, and (4) sustain a long cycle life (i.e., be able to be electrically recharged indefinitely and be able to supply a large amount of energy per cycle). Sammells (*54*) discussed several promising electrochemical batteries in detail, including nickel-iron batteries, nickel-zinc batteries, zinc-chlorine and zinc-bromine batteries, metal air batteries (Zn and Al), and lithium alloy-metal sulfide and sodium-sulfur batteries. In batteries containing zinc, the most obvious problem with these batteries is the formation of dendritic zinc upon electrical charging which results in a loss of cell capacity, intercell electrical shorting, and electrochemical cell asymmetry (*54*).

There are three types of fuel cells that are typically used to convert hydrogen and oxygen gases into water (*54*). Although the kinetics for the reduction of oxygen at a cathode are more rapid in alkaline than acid electrolytes, the presence of CO_2 in atmospheric O_2 leads to a build-up of carbonate salts in alkaline electrolytes and the electrode pores. Therefore, the emphasis in fuel cell research has been aimed at fuels cells with acidic electrolytes. The phosphoric acid fuel cell operates at 180°C and uses highly dispersed platinum on graphitized carbon as electrodes. Research is presently being directed toward minimizing the amount of platinum required to catalyzed the spontaneous electrochemical reaction. The platinum electrodes are highly susceptible to carbon monoxide poisoning. The molten carbonate fuel cell operates at 650°C and consists of a porous nickel anode ($H_2 + CO_3^{2-} \rightarrow H_2O + CO_2 + 2\ e^-$), a porous nickel oxide cathode ($O_2 + 2\ CO_2 + 4\ e^- \rightarrow 2\ CO_3^{2-}$), and an electrolyte tile, which contains an inert matrix of $LiAlO_2$ into which a variety of alkali carbonate mixtures can dissolve at the operating temperatures. The CO_2 produced at the anode is shuttled to the cathode, where it is allowed to react. The solid oxide fuel cell is based on the solid electrolytes of yttrium- or calcium-stabilized zirconia which can act as highly ionic conducting materials for oxygen ions at temperature of 900-1000°C. The fuel is oxidized at a nickel-coated anode within the center of a tube made of the solid electrolyte and the oxidant (oxygen) is reduced at a tin-doped indium oxide catalyst placed on the outside wall of the solid electrolyte tube (*54*).

In discussing NASA and the space program to teach college students chemistry, Kelter, Snyder, and Buchar (*55*) listed several reasons why hydrogen-oxygen fuel cells are exclusively used in the space program. These reasons include the fact that the half-cells involving these gases have been extensively researched and improved, hydrogen and hydroxide ions have higher mobilities than other ions (which results in larger current densities), and the product, water, can be used by the astronauts.

Metallurgy and Electrolysis. Several examples have been published regarding the formation of metals from their oxide ores (metallurgy) and the formation of new chemicals from non-spontaneous reactions through the use of electricity (electrolysis). The metallurgy of iron (*56*) the formation of aluminum metal (*47, 57*) and the formation of hydrogen, chlorine, and sodium hydroxide from brine (*58*) will be discussed in greater detail here.

In his discussion of the chemical processes that occur when iron oxide ore is transformed into steel, Sellers (*56*) pointed out that the conversion of iron ore to steel is essentially a purification process relying on two systems of oxidation-reduction reactions. In the first stage, the iron oxide ore must be reduced to its metallic state. This reaction occurs in a blast furnace, with carbon monoxide as the main reductant (e.g., $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$). The reduced product is called "pig iron". The second step involves the removal of impurities from the pig iron (including carbon, phosphorus, silicon, aluminum, calcium, magnesium, and sulfur) by oxidation using pure oxygen or air. Carbon is removed as gaseous CO or CO_2. Aluminum (as Al_2O_3), silicon (as SiO_2), and sulfur (as CaS) dissolve in the calcium oxide slag, which is physically separated from the blast furnace.

Nearly all aluminum is produced by the electrolysis of alumina (Al_2O_3) in a molten cryolite electrolyte, which is referred to as the Hall-Heroult process (*57*). The alumina used in this process has been purified by the Bayer process to remove iron impurities. The electrolyte consists of molten cryolite, which can be written as Na_3AlF_6, or $3NaF{\cdot}AlF_3$, with calcium fluoride or lithium fluoride impurities. These impurities lower the melting point of the electrolyte but also decrease the solubility of Al_2O_3 in the electrolyte, so they are generally limited to 10% by weight of the electrolyte. In the electrolyte, aluminum ions exist as AlF_6^{3-}, AlF_4^-, and as dimeric $Al_2OF_6^{2-}$ and $Al_2O_2F_4^{2-}$ ions. The cathode consists of iron electrodes and research performed on the reduction reaction occurring at the cathode has eliminated the possibilities of reducing sodium ions with subsequent aluminum reaction (i.e., $3\ Na + Al^{3+} \rightarrow 3\ Na^+ + Al$) and of reducing free Al^{3+} ions (i.e., $Al^{3+} + 3\ e^- \rightarrow Al$). The anode is made of pre-baked carbon electrodes, which are consumed during the electrolysis reaction. Although the primary reaction at the anode can be written as the oxidation of oxide ions (i.e., $C + 2\ O^{2-} \rightarrow CO_2 + 4\ e^-$), the oxide ions are complexed in the electrolyte and therefore this reaction must involve complex ions (e.g., $Al_2O_2F_4^{2-} + 8\ F^- + C \rightarrow 2\ AlF_6^{3-} + CO_2 + 4\ e^-$). Carbon dioxide is formed exclusively, even though it is thermodynamically unstable with respect to CO in the presence of graphite electrodes. This is explained by the fact that the desorption of CO from the electrode is kinetically hindered, resulting in CO_2 formation which desorbs rapidly.

The electrolysis of aqueous sodium chloride (the Chlor-Alkali process) to produce chlorine, hydrogen, and sodium hydroxide is the second largest user of electricity (next to the Hall-Heroult process for producing aluminum) among electrolytic industries (*58*). The balanced equation for this reaction is: $2\ NaCl + 2\ H_2O \rightarrow H_2 + Cl_2 + 2\ NaOH$. Chlorine is used to make polyvinylchloride and as a bleaching agent; sodium hydroxide has wide industrial applications in mineral processing, the paper industry, and textile and glass manufacturing. Although this reaction only requires 2.23 V to decompose the aqueous sodium chloride, larger voltages (usually around 3.5 V) are used to overcome kinetic barriers to these reactions. There are three types of cells used in the Chlor-Alkali process: Diaphragm cells, membrane cells, and mercury cells. These cells differ primarily in how they separate the NaOH and Cl_2 products and prevent them from reacting (e.g., $2\ NaOH + Cl_2 \rightarrow NaOCl + NaCl + H_2O$). diaphragm cells use an asbestos diaphragm or a polymer-modified asbestos composite to prevent mixing of NaOH and Cl_2. The liquid at the cathode tends to be 12% NaOH and 15% NaCl. Membrane cells use an ion-exchange membrane that allows the exclusive transfer of Na^+ ions from the anode to the cathode. The liquid at the cathode in membrane cells tends to be 10-35% NaOH with little NaCl impurities. Mercury cells physically separate the anode and cathode compartments using the mercury cathode itself. Chlorine is generated at the anode and sodium ins are reduced and enter the mercury cathode as an amalgam (i.e., $Na^+ + e^- + Hg \rightarrow Na{\cdot}Hg$). The amalgam passes into another cell where it reacts with water to produce

NaOH and H_2 (i.e., 2 Na·Hg + 2 $H_2O \rightarrow H_2$ + 2 NaOH + 2 Hg) and the regenerated mercury returns to the anode compartment. Due to the damaging environmental effects of mercury, these cells are gradually being phased out of use.

Blatt (*59*) and Doeltz, Tharaud, and Sheehan (*60*) discussed the process of anodizing aluminum, in which the protective aluminum oxide coating (which is typically about 20 Å thick) is increase to about 10^5 Å (*60*). Anodizing aluminum results in a thin porous film formed on the surface of a thin, non-porous, passive film. Anodized aluminum surfaces are relatively maintenance-free with excellent resistance to weathering and they provide a surface which can be colored with organic dyes and mineral pigments (*59*). The extraction of metals from oxide ores using carbon, carbon monoxide, and hydrogen gas were also briefly discussed (*50*). Wagenknecht (*61*) also discussed the use of electrolysis in the synthesis of several organic compounds including adiponitrile, tetramethyl lead, perfluorinated organic compounds, and several other organic products.

Empirical Data Concerning Student Misconceptions

Little of the chemical education research concerning oxidation-reduction reactions and electrochemistry contains empirical data that identifies actual student difficulties or misconceptions. Student misconceptions regarding oxidation-reduction reactions and electrochemistry are usually identified from student responses to conceptually-based multiple-choice questions or semi-structured interview questions.

Misconceptions Regarding Oxidation-Reduction Reactions. Andersson (*62*) summarized the results of several studies concerning student explanations of oxidation-reduction reactions and created a general category ranking the chemical knowledge demonstrated by these responses. Andersson's five levels of explaining oxidation-reduction reactions are: (1) It is just like that (which is really no explanation at all); (2) Displacement (the new compound was really inside the old compound or the air all along but is now visible); (3) Modification (the new compound is really the old compound in another form—a phase change has occurred); (4) Transmutation (atoms are transformed into new types of atoms or into energy); (5) Chemical Interaction (atoms are conserved, but they react with each other). In the first three levels, students still do not differentiate between the macroscopic properties of solids, liquids, and gases and the microscopic properties of atoms and molecules. In the fourth level, students recognize the differences between macroscopic and microscopic properties of objects, but still try to apply macroscopic rules to microscopic objects. In the fifth level, students fully differentiate between macroscopic and microscopic properties.

Barral, Fernändez, and Otero (*63*) reported student explanations of a simple electrochemical reaction. The students observed that when zinc is added to dilute hydrochloric acid (HCl), bubbles form on the surface and the teacher explained that zinc is oxidized by hydrogen ions, which are reduced to make hydrogen gas (i.e., $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$). The students then added copper to dilute HCl and observed no bubbles. Finally, a piece of copper wire was attached to a piece of zinc and was dropped in dilute HCl; the students then observed bubbles appearing on both surfaces and were asked to explain this reaction. The authors reported three types of student responses: (1) Purely descriptive replies (in which students merely described the reaction interpreting their observations); (2) Alternative representations (in which students suggested than zinc shared heat or hydrogen atoms with copper to produce hydrogen gas); (3) Electronic interpretations (in which students correctly interpreted that zinc shared electrons with copper so that hydrogen ions could now react at the copper surface). The authors suggested that it is the instructor's duty to confront student misconceptions through class discussions, counterexamples, and clear explanations of the accepted theories that describe these phenomena.

Allsop and George (*64*) analyzed and interpreted student responses to multiple-choice questions concerning oxidation-reduction reactions. These responses suggested that students have difficulty properly identifying oxidation-reduction reactions and could not consistently use the oxidation number method. In particular, 42% of the students incorrectly suggested that the decomposition of H_2SO_4 into SO_3 and H_2O was an oxidation-reduction reaction. From the analysis of student responses to additional multiple-choice questions with carefully written distractors, it was clear that students have difficulty balancing charge in chemical reactions. In fact, several students were satisfied writing equations that are not charge balanced (e.g., $H_2 + 2\ Cl^- \rightarrow 2\ H^+ + Cl_2$). Students also had trouble determining the proper reaction stoichiometry, given two balanced half-reactions. The students didn't seem to realize that half-reactions are balanced to conserve overall charge and many students arbitrarily chose a 1:1 ratio.

Garnett and Treagust (*9*) reported student misconceptions related to electrical circuits and oxidation-reduction reactions from interviews with 11-12 grade students in Western Australia. During these interviews, students were asked to determine which of these reactions ($2\ Mg + O_2 \rightarrow 2\ MgO$; $Mg + 2\ HCl \rightarrow MgCl_2 + H_2$; $H^+ + OH^- \rightarrow H_2O$; $2H^+ + CO_3{}^{2-} \rightarrow H_2O + CO_2$) represent oxidation-reduction reactions. From these interviews, the authors proposed four general misconceptions experienced by the students when trying to identify oxidation-reduction reactions: (1) The oxidation state of the free element is the same as the charge of its monatomic ion; (2) Oxidation numbers can be assigned to polyatomic molecules that equal its charge, and changes in the total charge represent oxidation-reduction reactions; (3) The addition of oxygen is always oxidation and the removal of oxygen is always reduction; and (4) Oxidation and reduction reactions can occur independently. The authors' conclusion of this study is that students are confused by the fact that there can be several different and conflicting definitions of oxidation and reduction reactions (e.g., oxidation has been defined in the past as the loss of hydrogen, the gain of oxygen, the loss of electrons, and an increase in the oxidation number).

De Jong, Acampo, and Verdonk (*65*) analyzed how two high school teachers taught the subject of oxidation-reduction reactions to determine why students have difficulty with this topic. The authors listed eight problems that may have lead to student difficulties: (1) The teachers introduced the concept of oxidation numbers without explaining why the concept is necessary; (2) The teachers introduced superfluous explanations to explain the concept of electron transfer; (3) The teachers told students what they should observe in the laboratory instead of allowing them to make their own observations; (4) The teachers introduced imprecise terminology which may have confused students; (5) The teachers introduced the concept of oxidation numbers without explaining that they do not represent actual atomic charges; (6) The teachers did not discuss incorrect responses with the class as a way to facilitate student conceptual change; (7) The teachers tended to ignore important applications of oxidation-reduction reactions in the field of technology and society; and (8) The teachers overemphasized the use of algorithmic procedures to solve simple problems. The teachers were unaware of many of these problems, and cited a lack of time or teaching ability as reasons why some of these problems arose. The authors also suggested that the teachers' high level of subject matter knowledge may make it difficult for them to teach oxidation-reduction reactions at a level that students can understand.

Misconceptions Regarding Electrochemistry. Hillman, Hudson, and McLean (*66*) compared student responses to electrochemistry questions containing technical words to questions that were rephrased to minimize the use of terminology. When these questions were rephrased, several of them had higher difficulty indices (i.e., more students could answer them) as well as higher discrimination indices (i.e., these questions are better at distinguishing between low- and high-ability students). The authors pointed out that the use of extraneous phrases can make a problem more difficult to interpret or understand, but that technical words or phrases in algorithmic problems did not affect students' abilities to solve these problems.

Allsop and George (*64*) also probed student misconceptions concerning electrochemical concepts. In particular, they found that students had difficulty using standard reduction potentials to predict the direction of chemical reactions. Only 20% of their students were able to produce an acceptable diagram of an electrochemical cell, 11% stated that a salt bridge provides a pathway for the flow of electrons, 36% suggested that reduction occurs at the anode, and 30% thought that oxidation occurs at the cathode. These questions also showed that students have a very poor appreciation for the effects of electrolyte concentration on the cell potential.

Garnett and Treagust (*9*) probed student misconceptions regarding electrical circuits and oxidation-reduction reactions in simple electrochemical cells using student interviews. As the result of the protocol analysis of these interviews, the authors reported several student misconceptions. These misconceptions include the notions that: (1) Electrons can flow in the electrolyte solutions and the salt bridge to complete the electrical circuit; (2) Conventional current (as taught in many physics classes) consists of protons (H^+) flowing through the wire and is different from "chemical" current, which is flowing electrons; (3) Ions flowing in aqueous solutions do not constitute a current; (4) EMF differences in an electrochemical cell are a result of differences in electron concentrations (the anode has a higher electron concentration and therefore electrons flow through the wire to the cathode, which has a lower electron concentration).

In a follow-up study, Garnett and Treagust (*44*) interviewed students using a similar interview protocol to determine student misconceptions concerning simple electrochemical and electrolytic cells. The authors reported the following student misconceptions as a result of their interview analysis: (1) Standard reduction potentials list metals in the order of decreasing metal reactivities; (2) The fact that the EMF of the standard hydrogen electrode equals 0.00 V is not arbitrary, and there is no need for some standard half-cell; (3) Electrons flow in electrolyte solutions and the salt bridge to complete the electrical circuit; (4) The anode is either positively charged (because it has lost electrons) or negatively charged (because the electrons originated at the anode); (5) In electrolytic cells, the battery does not affect the direction of electron flow; (6) No reaction will occur at inert electrodes; (7) In electrolytic cells, water does not react under any conditions; and (8) In electrolytic cells, the EMF can be positive.

Garnett, Garnett, and Treagust (*10*) discussed some probable origins of the student misconceptions identified in their previous studies (*9, 44*) and the implications of these misconceptions on improving the chemistry curriculum. The origins of these misconceptions include: (1) Compartmentalization of subjects (e.g., treating chemistry and physics as distinct and independent subject areas by using different terminologies to describe the same phenomena); (2) Inadequate prerequisite knowledge (instructors need to establish student knowledge levels and plan instruction accordingly); (3) Misuse of everyday language in chemical situations (instructors and textbook authors must be exceedingly careful in their wording so that students are less likely to misinterpret their comments); (4) Use of multiple definitions and models (when multiple definitions or models are used, the uses and limitations of these models must be clearly explained; if the models contradict each other, it is advisable to limit the number of models used or to clearly explain when each method is useful); and (5) Rote application of algorithms (students should be taught in ways that encourage them to understand the concepts instead of memorizing information and accepting the use of algorithms without question).

Ogude and Bradley (*45*) noted that although many students could solve quantitative electrochemical problems that appear on most chemistry exams, very few students were able to answer qualitative questions that required a deeper conceptual knowledge. The authors administered a 20-item multiple-choice, true-false, and assertion-reason exam to high school and college students to probe students' understanding of the microscopic processes involved in electrochemical cells. The exam was analyzed to determine how widespread each misconception was and to determine possible causes for each misconception. Several students believed that electrons can flow through the salt bridge and into the electrolyte solutions. In this exam, 20% of the students consistently replied that electrons cannot flow in the electrolyte, 35% consistently replied that electrons can flow in the electrolyte, and 45% were inconsistent in their responses. The results of the 25th National Youth Science Olympiad in South Africa in 1989 (N = 6900, referenced in *45*) showed similar results in which 30% of the students suggested that ions flow to complete the circuit in electrolyte solution, while 61% suggested that electrons flow in the electrolyte.

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COMMON STUDENT MISCONCEPTIONS IN ELECTROCHEMISTRY: GALVANIC, ELECTROLYTIC, AND CONCENTRATION CELLS

A paper submitted for publication to the *Journal of Research in Science Teaching*

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Abstract

This study replicates, with additions, research done by Garnett and Treagust (*1*). Garnett and Treagust's interview questions for galvanic and electrolytic cells were used with modifications; concentration cell questions were constructed by the method used in Garnett and Treagust's article. These questions were administered to 16 introductory college chemistry students after electrochemistry instruction. Student misconceptions most commonly encountered involved the notions that electrons flow through the salt bridge and electrolyte solutions to complete the circuit, plus and minus signs assigned to the electrodes represent net electronic charges, and water is unreactive in the electrolysis of aqueous solutions. New misconceptions identified include the notions that half-cell potentials are absolute and can be used to predict the spontaneity of individual half-cells and electrochemical cell potentials are independent of ion concentrations. Most students demonstrating misconceptions were still able to correctly calculate cell potentials, which is consistent with research (*2*) suggesting that students capable of solving quantitative examination problems often lack an understanding of the underlying concepts. Probable origins of these student misconceptions were attributed to the fact that students are ignorant about the

relative nature of electrochemical potentials and that some chemistry textbooks make misleading and incorrect statements. A minor technical flaw in the Garnett and Treagust study is also addressed.

Introduction

Research in the field of chemical education in the past twenty years has focused on problem-solving skills and the identification and investigation of student misconceptions experienced in solving chemical problems (*3*). Some of the topics investigated include balancing chemical equations, gas laws, chemical equilibrium, the concept of a mole, heat, and the conceptions of matter. This research confirms that student beliefs about problem complexity affect student performance and learning. Although Herron did not specifically mention studies concerning electrochemistry, student and teacher surveys suggest that students find this topic difficult (*4, 5*).

George Bodner (*6*), on the constructivist approach in chemical education, suggested that knowledge cannot simply be handed down from instructor to students—it is something that a student must actively construct from new information and his or her existing experiences and knowledge. A student uses his or her existing knowledge base to evaluate new information—if the new information is consistent with his or her existing knowledge base, it can be assimilated; however, if the new information contradicts the student's existing knowledge base, this knowledge base must be changed to accommodate the new information. Because knowledge is constructed by the student (*7, 8*), any erroneous information that is part of the student's knowledge base may adversely affect subsequent learning.

Although the constructivist model of learning suggests that the student constructs knowledge from his or her existing experiences and knowledge, it recognizes that this knowledge must be consistent with the circumstances that led to its formation—"the only thing that matters is whether the knowledge we construct from this information functions satisfactorily in the context in which it arises." (*6*, p. 874). Research on student understanding of scientific phenomena indicates that student explanations are often inconsistent with, inferior to, and incapable of explaining observable phenomena when compared to the scientifically-accepted descriptions (*9-11*).

In this paper, we define the term *misconception* as student conceptual and propositional knowledge that is inconsistent with or different from the commonly-accepted scientific consensus and is unable to adequately explain observable scientific phenomena (*6, 12*). It is important to note that some student misconceptions are capable of adequately explaining the student's experiences and observations, appear quite logical to the student, and are consistent with his or her understanding of the world. In these instances, student misconceptions are very resistant to change (*3*).

Several researchers have documented student misconceptions concerning electrochemistry. Allsop and George (*13*) reported that students had difficulty using standard reduction potentials to predict chemical reactions and were unable to produce an acceptable diagram of an electrochemical cell. Garnett and Treagust (*1, 14*) probed student misconceptions on oxidation-reduction reactions and electrochemical and electrolytic cells through student interviews. As a result of the protocol analysis of these interviews, the authors reported several common student misconceptions. A paraphrased list of these common student misconceptions appears in Table 1. In a subsequent article, Garnett, Garnett, and Treagust (*15*) discussed some probable origins of these student misconceptions and their implications on improving the chemistry curriculum. Ogude and Bradley (*16*) noted that although many students can solve quantitative electrochemical problems that appear on most chemistry exams, few were able to

Table 1. Paraphrased List of Common Student Misconceptions Reported by Garnett and Treagust (*1*).

Galvanic cells: Questions 14-21

8a. In an ordered table of reduction potentials, the species with the most positive E° value is the anode.
8b. Standard reduction potentials list metals by decreasing reactivity.
9a. The fact that the E° for H_2(*1 atm*)/H^+(*1 M*) is zero is somehow based on the chemistry of H^+ and H_2.
9b. There is no need for a standard half-cell.
10a. Electrons enter the solution from the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit.
10b. Anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode.
10c. Cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode.
10d. Cations and anions move until their concentrations are uniform.
11a. The anode is negatively charged and releases electrons; the cathode is positively charged and attracts electrons.
11b. The anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons.

Electrolytic cells: Questions 22-28

12a. In electrolytic cells, the direction of the applied voltage has no effect on the reaction or the site of the anode and cathode.
12b. No reaction will occur if inert electrodes are used.
12c. In electrolytic cells, oxidation now occurs at the cathode and reduction occurs at the anode.
13a. In electrolytic cells, water is unreactive towards oxidation and reduction.
13b. When predicting an electrolytic reaction, the half-cell reactions are reversed prior to combining them.
13c. The calculated cell potentials in electrolytic cells can be positive.
13d. There is no relationship between the calculated cell potential and the magnitude of the applied voltage.

answer qualitative questions that required a deeper conceptual knowledge of electrochemistry.

The purpose of this investigation is to report the identification of student misconceptions concerning electrochemistry related to galvanic (electrochemical), electrolytic, and concentration (Nernst) cells. This study replicates the results of Garnett and Treagust's interviews (*1*) on electrochemical and electrolytic cells and extends this study by addressing student misconceptions on concentration cells. Replication of the Garnett and Treagust study using college chemistry students from a midwestern American university is warranted for several reasons: (1) Teaching styles and content for high school chemistry courses are very different in Europe and Australia than in the United States and the former study contains material that is considered to be advanced for American high school <u>and</u> college students; (2) Several student misconceptions reported by Garnett and Treagust have not been observed with American students; (3) Several of the questions used by Garnett and Treagust have minor technical content flaws that appear to affect the interpretation of student responses. Garnett and Treagust's questions with flaws were rewritten to make the identification of student misconceptions explicit and to ensure correct content presentation; and (4) Good (*17*) pointed out that very few replication studies have been published in science education. We believe that this replication study does add to the knowledge of teaching and learning in electrochemistry.

Three questions provide a focus for this replication study: (1) What propositional knowledge and concepts are necessary for students to understand galvanic, electrolytic, and concentration cells in electrochemistry? (2) What misconceptions do students acquire about galvanic, electrolytic, and concentration cells in electrochemistry? (3) What are the implications of this research for classroom practice and science curriculum development?

Method

This study replicates the investigation of Garnett and Treagust (*1*) concerning common student misconceptions on galvanic and electrolytic cells. It also investigates common student misconceptions on the related topic of concentration (Nernst) cells.

Sample Size and Selection. The sample consisted of sixteen student volunteers (nine male and seven female students) from three introductory college chemistry courses taught at an American midwestern university. The three chemistry courses contained students with differing backgrounds and interests in chemistry (the courses are roughly divided into students majoring in the liberal arts, engineering, and the physical sciences). No attempt was made to ensure that all students received equal instruction; in fact, the engineering students were not exposed to concentration cells during instruction and therefore were not interviewed on this subject. Each student was individually interviewed for 40-50 minutes by the authors after receiving regular classroom instruction on electrochemistry using a set of semi-structured interview questions.

Identification and Validation of Conceptual and Propositional Knowledge Statements. The propositional statements derived by Garnett and Treagust (*1*) for galvanic and electrolytic cells were used without modification. Consistent with the techniques used by Garnett and Treagust, propositional knowledge statements necessary to fully understand concentration cells were derived by the researchers after reviewing several introductory college chemistry textbooks for relevant conceptual and propositional knowledge. The propositional knowledge statements for concentration cells appear in Table 2.

Table 2. Conceptual and Propositional Knowledge Statements about Concentration (Nernst) Cells

10. In a concentration (Nernst) cell:
    (a) there is a spontaneous chemical reaction which converts stored chemical energy into electrical energy,
    (b) the oxidation-reduction reaction which takes place is controlled and the oxidation and reduction half-reactions usually occur in separate compartments called half-cells,
    (c) the chemical contents of the two half-cells are the same and the oxidation-reduction reaction is determined by the relative concentrations of the reactants in the half-cells, and
    (d) the cell e.m.f. generated depends on the relative concentrations of the reactants in the two half-cells.
11. Half-cells:
    (a) are compartments in which separate oxidation and reduction half-reactions occur,
    (b) consist of an electrode immersed in an electrolyte,
    (c) are linked by a salt bridge which allows the transfer of ions in the internal circuit, and
    (d) enable the transfer of electrons from one reactant to another to take place through an external circuit or metallic conductor which links the electrodes.
12. Electrodes:
    (a) are electrical conductors that are placed in an electrolyte to provide a surface for oxidation or reduction half-reactions,
    (b) and the electrolyte determine the oxidation and reduction reactions that will occur,
    (c) that are considered to be inert, such as graphite and platinum, are made from substances which conduct electricity and are not chemically altered in cell reactions, and
    (d) are labeled as the anode or cathode depending on the site of the oxidation and reduction half-reactions; the electrode at which oxidation occurs is called the anode, which is labeled as (–), while the electrode at which reduction occurs is called the cathode, which is labeled as (+).
13. Transfer of charge:
    (a) In a concentration cell in which the electrodes react, the direction of the reaction is determined by the relative concentrations of the

Table 2 (continued)

electrolyte solutions in the half-cells; electrons move directly from the electrode in the less concentrated half-cell (anode) to the electrode in the more concentrated half-cell (cathode) through the external circuit and positive ions are released into the less concentrated solution as the anode dissolves. At the cathode, the positive ions in solution accept electrons and are plated out onto the cathode.

- (b) If the electrodes are inert, electrons are transferred directly from the oxidized substance to the anode and then through the external circuit to the cathode. At the cathode the substance being reduced accepts electrons.
- (c) An electrolyte conducts electricity within a cell by the movement of dissolved positively- and negatively-charged ions. The movement of ions completes the circuit and maintains electrical neutrality.
- (d) Negative ions are called anions and positive ions are called cations.
- (e) Anions move through the electrolyte to the anode and cations move to the cathode.
- (f) A salt bridge contains ions in solution and provides a continuous path for the movement of ions between separate half-cells.

14. Predicting cell potentials:
    - (a) The potential of a concentration cell depends on the temperature of the cell and the relative concentrations of the reactants in the half cells and can be expressed by the Nernst equation:

$$E = E^\circ - \frac{RT}{nF} \times \ln Q = E^\circ - \frac{0.0591\ \mathrm{V}}{n} \times \log Q$$

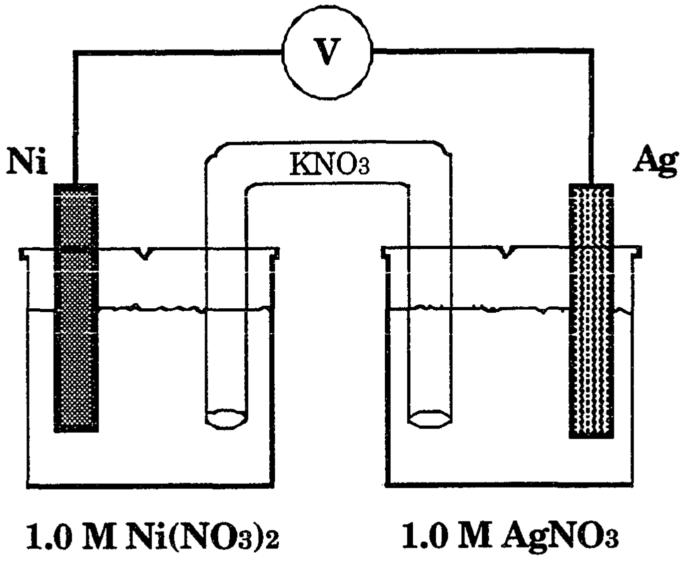
where E° is the standard reduction potential, R the gas constant, T is the temperature, n is the number of electrons transferred determined from the half-reactions, F is the Faraday constant, and Q is the reaction quotient.

- (b) The standard reduction potentials assume 1.0 mol / L concentration, 1 atm pressure, and 25°C temperature. In a concentration cell, the standard reduction potential refers to an electrochemical cell that has identical half-cells (both at standard conditions) and therefore has an E° value of zero.
- (c) The reaction quotient is written as the product of the concentrations of the products (with exponents equal to the coefficients in the balanced chemical equation) divided by the product of the concentrations of the reactants (with exponents equal to the coefficients in the balanced chemical equation).
- (d) The temperature of the system (which is typically 25°C) must be converted to the Kelvin scale (here, 298 K).

These statements were reviewed by four college chemistry professors and their comments were used to revise the original list. The statements provided a body of scientifically-accepted knowledge required by students to fully understand concentration cells, a basis for comparing student responses with scientifically-accepted views, and a framework for the development of the interview protocol and data analysis procedures.

Development of the Interview Protocol. Common student misconceptions were identified by means of individual, semi-structured interviews. The interview protocols developed by Garnett and Treagust (*1*) for galvanic and electrolytic cells were used with modifications. Some questions originally used by Garnett and Treagust evoked no student misconceptions, but instead demonstrated the students' general lack of knowledge. These questions were omitted from this study. Some of the remaining questions were edited for clarity; only one section was edited for content. Garnett and Treagust's original questions on the electrolysis of aqueous $NiBr_2$ (questions 22-28) should produce Ni(s) and Br_2(aq) and are inappropriate for determining whether students have ignored the electrolysis of water; therefore, this set of electrolysis questions were changed to the electrolysis of aqueous $AlBr_3$, which will produce H_2(g) and Br_2(aq). The interview protocol for concentration cells was developed from the list of conceptual and propositional knowledge statements and was written to pose general questions based on these statements. Interview questions 14-33, which were used to probe student misconceptions concerning galvanic, electrolytic, and concentration cells, are shown in Figure 1.

Questions about Galvanic Cells:



14. What is the purpose of each piece of apparatus shown above? (Do the metal strips always react?)

16. How would you determine which electrode is the anode and which is the cathode?

17. How is a current produced in this cell?

18. What is happening in the solutions? What does the salt bridge do?

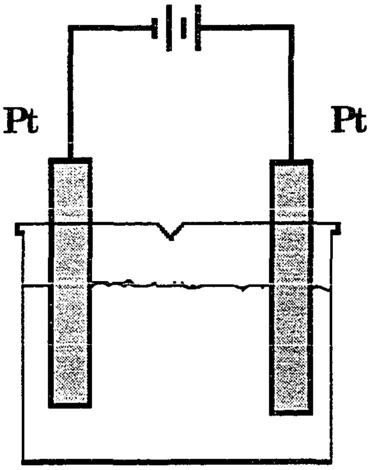
19. In which direction do the charges (positive and negative) flow in this cell to complete the circuit?

20. What reactions are taking place in each cell? Can you predict the E value for this set-up?

21. Why does the reaction $2H^+ + 2e^- \rightarrow H_2(g)$ have an E° value of 0.00 V? How would you measure the E° of the $Ag^+ + e^- \rightarrow Ag(s)$ half-reaction?

Figure 1. Interview protocols for galvanic, electrolytic, and concentration cells.

Questions about Electrolytic Cells:



1.0 M $AlBr_3$

22. How does this cell differ from the electrochemical cell drawn above?

23. How would you determine which electrode is the anode and which is the cathode?

24. In which direction do the charges (positive and negative) flow in this cell to complete the circuit?

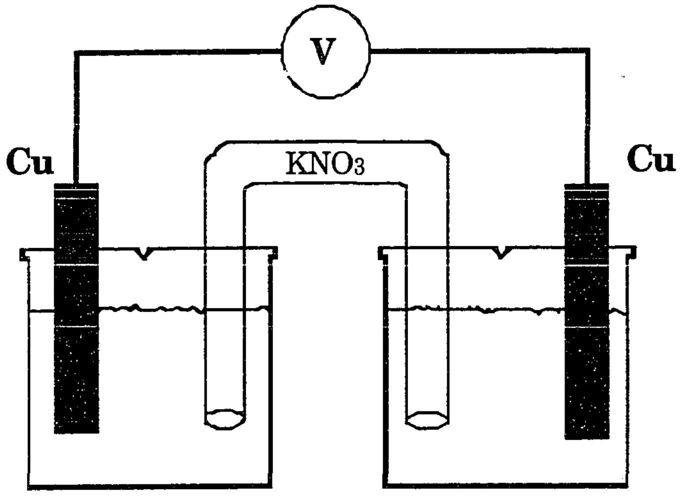
25. What reactions are taking place at each electrode?

26. Can you predict the E value for this set-up?

28. Suppose the solution was changed to molten $AlBr_3$—what would you expect to happen?

Figure 1. (continued)

Questions about Concentration (Nernst) Cells:



1.00 M $CuCl_2$ **0.01 M $CuCl_2$**

29. How would you decide which electrode is the anode and which is the cathode? What are the chemical reactions occurring at each electrode?

30. What is the standard reduction potential (E°) for this cell?

31. What is the value of the reaction coefficient (Q) for this cell?

32. What is the e.m.f. of this cell?

33. What would happen to the e.m.f. of this cell if the concentration in the right cell was changed from 0.01 M $CuCl_2$ to 0.001 M $CuCl_2$? (increase, decrease, no change, etc.)

Figure 1. (continued)

Analysis of Data. During each interview, the student/interviewer conversation was recorded on audiocassettes and the student was asked to write his or her responses and solutions on the question sheets used during the interview. The tapes and student responses were analyzed and anecdotal evidence of student misconceptions were transcribed verbatim. The summary of each student's misconceptions was combined to identify common student misconceptions and were confirmed or refuted by referring to the individual audiotapes of each student. Two science education researchers analyzed each tape.

Results

The results from the interviews are discussed in terms of nine areas of student difficulties: (1) Identifying the anode and cathode of galvanic cells; (2) Understanding the need for a standard half-cell; (3) Understanding current flow in galvanic, electrolytic, and concentration cells; (4) Understanding the charge on the anode and cathode; (5) Predicting the products and the electromotive force of galvanic cells; (6) Identifying the anode and cathode in electrolytic cells; (7) Predicting the products of electrolysis and the magnitude of the necessary applied electromotive force; (8) Identifying the anode and cathode in concentration cells; and (9) Predicting the products and the electromotive force of concentration cells. The misconceptions are labeled from 8 to 16 and are consistent with the labeling method used by Garnett and Treagust (*1*); misconceptions identified in this study that were reported previously by Garnett and Treagust share the same label. A list of misconceptions reported in this study appears in Table 3.

Table 3. Common Student Misconceptions Reported in This Study

Galvanic cells: Questions 14-21

8b. Standard reduction potentials list metals by decreasing reactivity.
8c. The identity of the anode and cathode depends on the physical placement of the half-cells.
8d. Anodes, like anions, are always negatively charged; cathodes, like cations, are always positively charged.
9a. The fact that the E° for H_2(*1 atm*)/H^+(*1 M*) is zero is somehow based on the chemistry of H^+ and H_2.
9b. There is no need for a standard half-cell.
9c. Half-cell potentials are absolute in nature and can be used to predict the spontaneity of the half cells.
10a. Electrons enter the solution from the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit.
10b. Anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode.
10c. Cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode.
10e. Electrons can flow through aqueous solutions without assistance from the ions.
10f. Only negatively-charged ions constitute a flow of current in the electrolyte and the salt bridge.
11a. The anode is negatively charged and releases electrons; the cathode is positively charged and attracts electrons.
11b. The anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons.
14a. Cell potentials are derived by adding individual reduction potentials.
14b. Half-cell potentials are not intensive properties.

Electrolytic cells: Questions 22-28

12a. In electrolytic cells, the direction of the applied voltage has no effect on the reaction or the site of the anode and cathode.
12b. No reaction will occur if inert electrodes are used.
12d. In electrolytic cells with identical electrodes connected to the battery, the same reactions will occur at each electrode
13a. In electrolytic cells, water is unreactive towards oxidation and reduction.
13c. The calculated cell potentials in electrolytic cells can be positive.
13d. There is no relationship between the calculated cell potential and the magnitude of the applied voltage.
13e. Inert electrodes can be oxidized or reduced.
13f. When two or more oxidation or reduction half-reactions are possible, there is no way to determine which reaction will occur.
13g. Electrolytic cells can force non-spontaneous reactions that do not involve electron transfer to happen.

Table 3 (continued)

Concentration cells: Questions 29-33

15a. The direction of electron flow in concentration cells is not dependent on the relative concentration of the ions.
15b. The products produced in the indirect reaction of electrochemical cells are different from those produced in the direct reaction of the starting materials.
16a. The cell potential in concentration cells is not dependent on the relative concentration of the ions.
16b. Because there is no net reaction in concentration cells, the reaction quotient cannot be calculated.

Identifying the Anode and Cathode of Galvanic Cells. In response to question 16, ten students determined the oxidized and reduced species from a table of standard reduction potentials and labeled the electrode where oxidation occurs as the anode and the electrode where reduction occurs as the cathode. However, when one of these students (S1786) was asked about the reactions of nickel metal in silver nitrate and silver metal in nickel nitrate, the student stated that there would be no reaction in the first situation and that nickel metal and silver nitrate would be produced in the other because "...silver has a higher potential, so it's going to react before nickel would," demonstrating *Misconception 8b* (standard reduction potentials list metals by decreasing reactivity).

Two students incorrectly assigned the electrodes: Student S1781 confused the definitions of oxidation and reduction and student S1672 believed that oxidation occurs at the cathode and reduction occurs at the anode. One student stated that the anode is always the electrode that appears on the left hand side of a diagram and the cathode is always the electrode on the right, demonstrating *Misconception 8c* (the identity of the anode and cathode depends on the physical placement of the half-cells).

Student: (S1642) I was just told that this would be the anode on the left and the cathode on the right.

Interviewer: But what if we gave you a diagram like this? [Reversing the half-cells]

Student: Well then, it [nickel] would be reducing—the cathode, right. That's just so far how the book has shown it to me and the way on the board it's been shown.

Interviewer: ...but when you look at a real-live cell, you never know—you can be looking at it here and then you can be looking at it here [opposite side] and it hasn't changed.

Student: Right... I have no idea which way it's going, anode or cathode.

Three students attempted to determine the identity of the anode and cathode by transferring the charge of anions and cations as the potential sign (– or +) of the electrodes. *Misconception 8d* (anodes, like anions, are always negatively charged; cathodes, like cations, are always positively charged) is supported by the following comments in which the student viewed a negatively-charged anode as being electron-rich and therefore the site of reduction.

Student: (S1673) [The] cation is positive, the anion is negative, so the anode's going to be getting its electrons and the silver is attracting electrons, so silver's the anode and the nickel is giving up electrons, so the nickel's the cathode.

Understanding the Need for a Standard Half-Cell. Of the ten students responding to question 21, seven students correctly stated that the half-reaction for the standard hydrogen electrode (SHE) was arbitrarily set at 0.00 V. Two of these students initially stated that the value of the SHE potential was based on the chemical reaction taking place in the SHE, demonstrating *Misconception 9a* (the fact that E° for the SHE is zero is somehow based on the chemistry of H^+ and H_2), before reasoning that this potential was arbitrary. These students and two others proposed a variety of reasons why the SHE potential is zero including: (a) "Hydrogen will either give up or accept an electron—it doesn't really care either way, it will go." (S1671); (b) "It has something to do with [the fact] that it only has one electron to give up" (S1675); (c) "It's in water, and a lot of reactions take place in acidic solutions where [the] hydrogen [ion] is present." (S1784); and (d) "When you take the acid and add electrons, you can make hydrogen gas—when you're done, you have a neutral solution." (S1787). One student could offer no reason to explain why the SHE has a potential of 0.00 V.

Eight of the ten students responding to question 21 were able to describe a process for determining an E° value for the Ag^+/Ag half-cell. Five of these students explicitly suggested the use of the SHE, in which the measured potential would also equal the half-cell potential; the other three suggested the use of another half-cell whose potential was known. These responses suggest that even though half of the students aren't able to explicitly state the relative nature of electrochemistry, most are aware of it. This is in contrast to *Misconception 9b* (there is no need for a standard half-cell).

However, in the course of the interviews, half of the students made comments suggesting that half-cell potentials are absolute in nature. This is apparent in the statement of student S1642: "...One of the questions in the book [asks] 'Can you measure a half-cell?' and I think I did finally conclude that 'Yes, you can.' " These eight students treated half-cell potentials as absolute numbers, suggesting that half-cells with positive potentials are spontaneous while those with negative potentials are non-spontaneous. *Misconception 9c* (half-cell potentials are absolute in nature and can be used to predict the spontaneity of the half cells), which was not reported by Garnett and Treagust, is supported by the following student comments.

Student: (S1671) ...the aluminum doesn't want to go to aluminum metal because it has a negative E value...

Student: (S1673) ...So, Cu to go from its solid state to its ionized state will need 0.34 volts so nothing's going to happen there. But with the Zn to go from its solid state to its ionic form gives up 0.76 volts so it's going to spontaneously happen...

Student: (S1674) Nickel... that's -0.25 (V) and aluminum... -1.66 (V), so neither of them are likely to naturally reduce, spontaneously reduce...

Student: (S1783) Nickel two-plus plus two electrons is not spontaneous.

Student: (S1786) The E° value of that [Ag^+/Ag] was positive 0.799 (V), and that's going to be a spontaneous reaction. But, on the other hand, [the reduction of] nickel is not going to be spontaneous because it has a negative reduction potential...

Understanding Current Flow in Galvanic, Electrolytic, and Concentration Cells. This aspect of student electrochemical misconceptions has been the most widely investigated and has been studied by Allsop and George (*13*), Garnett and Treagust (*1, 14*), and Ogude and Bradley (*16*). In general, students realize that current cannot flow without a closed circuit and many students believe that only the flow of electrons can complete this circuit. Consequently, many students cling to the notion that electrons flow from the anode to the cathode along the wire and are then released into the electrolyte at the cathode, traveling through the salt bridge and the electrolyte to the anode. The following student response demonstrates *Misconception 10a* (electrons enter the solution from the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit), which was exhibited in one form or another by nine of the sixteen students.

Student: (S1783) ...if there's an opening in the circuit somewhere, there's not going to be any reaction because there's not that constant circle for the electrons to go through... The electrons are what actually are the current. Current's a measure of electron flow... The salt bridge allows electrons to pass through it without allowing the solutions to mix.

When pressed further, two students who believed that electrons flow through the salt bridge stated that anions in the salt bridge and the electrolytes help transfer the electrons, which is represented as *Misconception 10b* (anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode). One of these students suggested that the negatively-charged anions would transfer electrons from the silver solution to the nickel solution through the salt bridge. When asked if a piece of copper wire could replace the salt bridge, the student believed that the circuit would still be complete because the electrons in solution could flow through the wire.

Student: (S1673) The NO_3^- should transfer its electrons to the salt bridge. It's [NO_3^-] going to get its electrons from the Ag^+ NO_3^- solution and transfer them back over the salt bridge and deposit them in the Ni^{2+} and 2 NO_3^- solution...

Interviewer: Is there anything carrying the electrons through the salt bridge or are they flowing by themselves?

Student: The NO_3^- carries the electrons—I know that 'cause that's similar in all of the solutions, so that's the one that actually transfers the electrons, but I don't know how it does it.

Interviewer: So if the salt bridge, instead of KNO_3, was KCl...

Student: Then the Cl^- would transfer the electrons. ...'cause you need an anion in there that will be able to accept electrons and then give them back up.

Interviewer: What if we put a piece of copper wire here [in place of the salt bridge]?

Student: Copper wire would work 'cause it'll transfer electrons just like the salt would.

Three of the students who believed that electrons flow through the salt bridge stated that cations transfer the electrons through the salt bridge, consistent with *Misconception 10c* (cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode). One of these students believed that because metals conduct electricity very well, metal cations would be capable of accepting electrons and transferring them from one electrode to the other.

> Interviewer: ...You were showing before that a complete path—would there be electrons going through the solution here between the two electrodes to complete the path?
>
> Student: (S1672) Wouldn't it have to be? Otherwise, I don't know... It's [the solution] got aluminum in it which is a metal and metals are really good at conducting electricity... So I think it [electrons] would be conducted from one [electrode] to the other...

Three of the students who believed that electrons can flow through the salt bridge stated that the electrons flow through solution without any assistance from anions or cations, exhibiting *Misconception 10e* (electrons can flow through aqueous solutions without assistance from the ions). One of these students suggested that only electrons could pass through the salt bridge without changing the solution volumes.

> Student: (S1785) ...I don't think anything's actually passing through it [the salt bridge] besides electrons, because otherwise you would have one solution going down and one going up [in volume].

Three of the students who correctly stated that ions flow through solutions and the salt bridge to complete the circuit suggested that only anion flow in solution completes the circuit and that cation flow does not constitute a current. The following student dialogue is consistent with *Misconception 10f* (only negatively-charged ions constitute a flow of current in the electrolyte and the salt bridge), in which the student believed that only anions flow to complete the circuit.

Interviewer: We've already talked about the minus charges—what about the plus charges, how would they flow?
Student: (S1675) The plus charges stay where they are. The plus charges stay in their beakers.
Interviewer: Before you had said that the nickel [ion] could go up here [salt bridge] and that maybe the K's could come out [in the Ag solution]...
Student: I was wrong.
Interviewer: OK, now you believe that they're not... they're just going to sit there.
Student: Yeah, the negative charges will move across.

Understanding the Charge on the Anode and Cathode. In their responses to question 16 for electrochemical cells and question 23 for electrolytic cells, seven of the sixteen students responded with comments suggesting that the electrodes have net positive and negative charges. Two of these students were inconsistent about the net charges on the electrodes, demonstrating *Misconception 11a* (the anode is negatively charged and releases electrons; the cathode is positively charged and attracts electrons) for the galvanic cell and *Misconception 11b* (the anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons) for the electrolytic cell. These misconceptions are demonstrated by the following comments from the same student. In the galvanic cell, the student implied that the nickel anode is negative because the electrons originate from it and the silver cathode is positive because they flow toward it. In the electrolytic cell, however, the student suggested that the positive anode attracts anions and the negative cathode attracts cations.

Student: (S1671) They [the electrons] flow from negative nickel to the positive silver.

Student: (S1671) ...the aluminum [ion] would go to the negatively-charged one, which is either the anode or the cathode. If the cathode is negatively charged, which I think it is, the Br [ion]... would be attracted to the anode.

Six of the seven students who suggested that the electrodes have net electronic charges believed that the anode is negatively charged and the cathode is positively charged (*Misconception 11a*). The two student responses below indicate that they believed the anode must be negatively charged because electrons are coming from it and that the electrons are attracted to the positively-charged cathode.

Interviewer: What is the minus [electrode], is that the anode or the cathode?

Student: (S1787) I would think that it's the anode because it's negative and that would be the direction electrons are coming from... My feeling would be that electrons go from the negative to the positive.

Student: (S1676) The cathode would be positive... because, like I pointed out before, a very high negative charge would not accept electrons because it already has too many, so it's more likely to be positive and the anode would be negative.

On the other hand, three of the seven students who suggested that the electrodes have net electronic charges believed that the anode is positively charged and the cathode is negatively charged (*Misconception 11b*). This misconception is supported by the following student comment suggesting that the nickel anode is positively charged because it is giving up electrons and the silver cathode is negatively charged because it accepts them.

Student: (S1674) ...See, I would think that if nickel is giving up electrons, it's going to be positive. And since silver is gaining electrons, it's going to be more negative.

Predicting the Products and the Electromotive Force of Galvanic Cells. In response to Question 26, all but one student were able to determine the correct products for the galvanic cell. Student S1676 incorrectly predicted that nitrate ions would be reduced to nitric oxide ($NO_3^- + 4\ H^+ + 3\ e^- \rightarrow NO + 2\ H_2O$) even though there was no acid present in either half-cell. Among the fifteen students who correctly predicted the products, twelve correctly predicted the cell potential. One student (S1781) who calculated an incorrect value was certain that one equation should be flipped, but was uncertain which one to flip. This student tried both alternatives, calculating cell potentials of +1.05 V and -1.05 V, and finally chose the incorrect, negative value. The other two students took the E° values for each reduction potential and added them together, resulting in a cell potential of +0.55 V, consistent with *Misconception 14a* (cell potentials are derived by adding individual reduction potentials).

Although every student demonstrated at least one major misconception about galvanic cells, most (80%) were able to calculate cell potentials correctly. This is consistent with other reports (*2, 15-16*) suggesting that students capable of solving quantitative examination problems often lack an understanding of the fundamental underlying concepts. In fact, when questioned further, two students who initially calculated the cell potential correctly demonstrated *Misconception 14b* (half-cell potentials are not intensive properties), which is apparent in the following student dialogue.

Interviewer:	Now, when you have your half-reaction for silver, you multiplied this [the half-reaction] by two but you haven't doubled the voltage over here.
Student:	Oh, do you have to do that?
Interviewer:	Do you think you would double the voltage?
Student: (S1642)	Well, if it is two, it makes sense. Is that what you do? If you have two silvers then you've got to double the voltage because you're using two of them?

Identifying the Anode and Cathode in Electrolytic Cells. Most student responses to question 22 focused on trivial, surface-level differences such as the electrolytic cell is in one beaker, the electrolytic cell has no salt bridge, and the electrodes in the electrolytic cell are made of the same metal. However, some responses showed some insight into why these differences are important.

> Student: (S1671) ...[In] the one above [galvanic cell], the solutions are separated so that the two metals won't react because they will spontaneously react otherwise and you can't get the electrons out of them [to do electrical work]...

Of the twelve students responding to question 23, only five were able to correctly identify the anode and cathode and determine the direction of electron flow from the polarity of the battery. Three students attempted to use the battery to determine the identity of the anode and cathode but were unsuccessful: Two of them assigned the battery potential incorrectly and two of them stated that electrons flow from the negative to the positive electrodes of the battery. These responses suggest that the students are not familiar with the purpose and function of the battery in an electrolytic cell, consistent with *Misconception 12a* (in electrolytic cells, the direction of the applied voltage has no effect on the reaction or the site of the anode and cathode). This misconception is also supported by the fact that two students stated that the identity of the electrodes would have to be assigned by evaluating the reactions occurring at each electrode. The following comment suggests that the student understood that either electrode can be the anode or the cathode, but failed to realize that it is the battery orientation that determines the identity of the electrodes.

> Student: (S1673) ...you'd have to know which one is oxidizing and which one is reducing. Because you have the same element on each side, technically you can force the electrons either way.

In determining the electrolysis products (discussed later), two students stated that the reaction could not proceed because the electrodes were not made of aluminum, which is consistent with *Misconception 12b* (no reaction will occur if inert electrodes are used) and is illustrated by the following student dialogue.

Student: (S1782) Al^{3+} goes to Al, but we can't go to Al though because it's not part of the... It can't do this, can it?
Interviewer: Why would you think that it couldn't go to Al?
Student: Because Al isn't one of the electrodes...

In the course of determining the direction of current flow in the electrolytic cell, two students expressed the notion that the two electrodes are completely the same and that the same reactions will occur at both electrodes. *Misconception 12d* (in electrolytic cells with identical electrodes connected to the battery, the same reactions will occur at each electrode) is demonstrated by the following student responses. At first, the student suggested that electrons will flow from one electrode to the other but could not see any difference in the two electrodes and later suggested that the reduction of aluminum ions will occur equally at both electrodes.

Student: (S1672) ...so I think it [electrons] would be conducted from one [electrode] to the other, but since these are the same—one's not going to be more electropositive and one's not going to be more electronegative than the other one because they are both the same—they're not going to travel like they would if they were two different metals like in this one [galvanic cell]...

I think it's [reduction of Al^{3+}] going to happen on these [electrodes] and I don't think it's going to matter which one it's closer to... because they're the same; one's not different from the other.

Predicting the Products of Electrolysis and the Magnitude of the Necessary Applied Electromotive Force. None of the twelve students responding to question 25 were able to correctly predict the products of the electrolysis. Five students correctly predicted that bromide would be oxidized to bromine, but predicted the aluminum ion would be reduced instead of water. Indeed, four of these students and four others totally ignored the possibility of water being oxidized or reduced. This is expressed as *Misconception 13a* (in electrolytic cells, water is unreactive towards oxidation and reduction) and is supported by the following comment in which the student stated that water is unreactive, but suggested that hydrogen and hydroxide ions can be oxidized or reduced.

> Student: (S1673) The water shouldn't do anything. It probably will because it is not actually just H_2O, you get H^+ ions and you get OH^- ions and so there may be a little reaction, but it's not going to be measurable. Water will not enter into the equation.

Of the five students who predicted that aluminum metal and bromine would be produced, three correctly calculated the cell potential to be -2.73 V. One student who predicted these two products, however, ended up with a positive potential (+2.73 V), which is consistent with *Misconception 13c* (the calculated cell potentials for electrolytic cells can be positive). The following comments suggest that student S1672 did not recognize the fundamental concept that electrolytic cells involve non-spontaneous reactions, so their potentials should not be positive.

> Student: (S1672) Al^{3+} plus three electrons yields aluminum solid and it's -1.66 [V]. So it's not likely to be reduced, it's more likely to oxidize. So if I wanted to find the total standard E, I would just add the 1.07 and change this around so it's becoming [+]1.66, 'cause it's not likely to be reduced right here, so I'm going to change it

> and make it opposite. So if I change it, the cell is likely to happen. Total E = 1.07 + 1.66 [V]...

One student who predicted these two products stated that the required potential could not be calculated because the battery was driving the system and is expressed as *Misconception 13d* (there is no relationship between the calculated cell potential and the magnitude of the applied force).

> Student: (S1641) I guess you could predict an E° value for this set-up, but because it's being charged by a battery... I guess you can, but I don't know how to predict an E° value for this set-up.

Four students who ignored the possibility of water being oxidized or reduced demonstrated *Misconception 13e* (inert electrodes can be oxidized or reduced), in which they considered oxidation and reduction half reactions containing platinum metal and its ion. When asked what oxidation reaction takes place in the electrolytic cell, student S1787 suggested that "...possibly the Pt is reacting to form a positive ion in the solution. Solid platinum is going to form platinum two-plus plus two electrons."

Four students considered the reactivity of water. Two of these students considered the possibility that both the ions in solution and water could be oxidized and reduced; however, these students arbitrarily chose the products of the electrolysis and did not use E° values to confirm their assumptions. This is represented as *Misconception 13f* (when two or more oxidation or reduction half-reactions are possible, there is no way to determine which reaction will occur). Student S1785 ignored the possibility of the ions being oxidized or reduced and chose hydrogen gas and oxygen gas as the products of the electrolysis. Student S1676 considered the reaction of water with the aluminum bromide in solution and suggested that the products of the electrolysis would be Al_2O_3 and HBr. This is consistent with *Misconception 13g* (electrolytic cells can force non-spontaneous reactions that do not involve electron transfer to happen) and is also apparent in the following student dialogue concerning the electrolysis of pure water.

> Interviewer: What would happen if... I have these two platinum electrodes and a battery and I just put it in a solution of water?
> Student: (S1784) You would get hydrogen [ions] and hydroxide.
> Interviewer: Is the battery going to force that to happen?
> Student: Well, it would force it to happen because it is not a spontaneous process.

Identifying the Anode and Cathode in Concentration Cells. Because students in one course (chemistry for engineering majors) did not cover concentration and Nernst cells in lecture, they were not asked to respond to interview questions 29-33 and in the interest of limiting the interviews to 40-50 minutes, several of the liberal arts and physical science students were not asked to respond to these question either. Consequently, the sample size for this section is smaller (six students instead of the original sixteen). Only one of these students was able to correctly determine the anode and cathode of the concentration cell based on the relative concentrations of copper(II) chloride in the half-cells with-out assistance. Two students stated that there would not be a potential difference in the concentration cell because both cells contain the same ions. This is represented as *Misconception 15a* (the direction of electron flow in concentration cells is not dependent on the relative concentration of the ions), and is supported by the following student dialogue. Because both half-cells contain copper metal and Cu^{2+} ions, the student believed that there is no driving force for electron transfer—the student did not recognize that a concentration difference can also serve as a driving force.

Student: (S1785)	I'd have a hard time with this one [concentration cell].
Interviewer:	OK, why?
Student:	Because they're both copper, so I can't see that there would be any transfer [of electrons] at all.

Because all but one of these students were unable to proceed with the interview questions due to a general lack of knowledge, the interviewers provided questions designed to prompt student responses. When two of the students were asked to predict what would happen if the contents of the two beakers were combined, both students responded that the mixture would reach an intermediate Cu^{2+} ion concentration. However, they were unable to apply the results of this direct reaction to the indirect reaction occurring in the concentration cell, which is expressed as *Misconception 15b* (the products produced in the indirect reaction of electrochemical cells are different from those produced in the direct reaction of the starting materials). This misconception was also demonstrated by student S1786, who was quoted earlier (*Misconception 8b*) as saying that the direct reaction of nickel metal and silver nitrate would result in no reaction, even though the student correctly predicted that nickel would be oxidized and the silver ion would be reduced in the electrochemical cell.

Predicting the Products and the Electromotive Force of Concentration Cells. Questions 30 and 31 provide the basis for answering question 32 using the Nernst equation. After prompting, the five students experiencing difficulties in identifying the anode and cathode were able to predict that copper would be oxidized to Cu^{2+} in the 0.01 M $CuCl_2$ half cell and that Cu^{2+} would be reduced in the 1.00 M $CuCl_2$ half cell. All of these students were also able to determine that the E° value for this system would be 0.00 V; however, three students had difficulty accepting the idea that E° could be zero when a spontaneous reaction is taking place. This difficulty suggests that these students did not recognize the difference between E°, which requires standard conditions (1 atm pressure and 1 M concentrations), and E and is expressed as *Misconception 16a* (the cell potential in concentration cells is not dependent on the relative concentration of the ions). The following dialogue demonstrates that although the student calculated E° to be zero, the student believed that this value is inconsistent with a spontaneous net reaction.

> Student: It [E°] would just come out zero... (Pause)
> Interviewer: Does that sound wrong?
> Student: Yes! Because otherwise, I don't quite understand
> (S1787) how you would get voltage coming across...

Three students were able to correctly determine the value of the reaction quotient, but three others experienced major difficulties with this step. Because the products and reactants contained the same chemical species ($Cu^{2+}(aq) + Cu(s) \rightarrow Cu(s) + Cu^{2+}(aq)$), these students canceled them and ended up with "no reaction" as the net equation and were therefore unable to calculate the reaction coefficient. This is represented as *Misconception 16b* (because there is no net reaction in concentration cells, the reaction quotient cannot be calculated) and is illustrated by the following student dialogue. This student determined that the two half-reactions were opposites of each other and when the common elements were canceled no reaction remained, making it impossible to calculate a Q value.

> Student: Well, it's [Q] going to be determined by the concentration
> (S1784) of products divided by the reactants... I'm not sure
> which is products and which is reactants.

Interviewer: Well, you have your two half-cells here and when you add them up, you get a reaction.
Student: But, they'll be the same! (Pause)
Interviewer: OK, so what you're saying is there's going to be nothing left for my Q.
Student: Right.

At this point in the interviews, all six students were able to correctly calculate the cell potential for the concentration cell and were able to correctly predict the cell potential change when the 0.01 M $CuCl_2$ solution was diluted to 0.001 M.

Discussion

In the replication of Garnett and Treagust's study (*1*), we were able to confirm most of the reported student misconceptions. In addition, we identified several new student misconceptions concerning galvanic, electrolytic, and concentration cells. Because the misconceptions that we were able to confirm were identified using two different samples (high school students in western Australia and college students in midwestern United States) measured three years apart, the results of this study support and increase the validity of the findings of the earlier study. Although we were unable to confirm all of the misconceptions previously reported by Garnett and Treagust, this study by no means attempts to downplay or refute these misconceptions. It is very likely that individual student differences between the two samples can be used to explain why the previous authors identified some misconceptions that we did not and vice versa. Therefore, we are not proposing that the common misconceptions observed by both studies are any more important than those observed in only one study, perhaps only more prevalent.

Possible Sources of Misconceptions. Garnett, Garnett, and Treagust (*15*) discussed some of the probable origins of the student misconceptions based on electrochemistry interview studies (*1, 14*). The origins of these misconceptions include: (1) Compartmentalization of physical science subjects (e.g., treating chemistry and physics as distinct and independent subjects by using different terminologies to describe the same phenomena); (2) Inadequate prerequisite knowledge; (3) Misuse of everyday language in chemical situations; (4) Use of multiple definitions and models; and (5) Rote application of algorithms. Garnett and Treagust (*1*) also introduced a new alternate framework based on the notion that an electrical current only involves the flow of electrons. Ogude and Bradley (*16*) attributed student misconceptions in electrochemistry to the superficial and vague descriptions textbooks often use to describe new science concepts.

In this study, we propose another possible source of student misconceptions in the field of electrochemistry: Ignorance of the relative nature of electrochemical potentials. Students need to understand that, just as in enthalpy (H) and free energy (G) measurements, chemists cannot make absolute measurements. Enthalpy and free energy measurements are always expressed in terms of changes (ΔH and ΔG, respectively) from the starting materials to the final products. The same is true of electrochemical cell potentials. There is no absolute zero potential for an oxidation or reduction reaction because we cannot measure the potential of a single half-reaction. This is directly linked to the fact that we cannot have an oxidation reaction without a reduction reaction—if a compound is going to give up electrons, something has to accept them! Therefore, any time we measure a cell potential, we are measuring the potential difference between electrons at the oxidizing agent and electrons at the reducing agent.

Misconceptions 9a, 9b, 9c, 14a, and 14b all attest to the fact that students do not recognize the relative nature of tabulated standard reduction potentials. Indirectly, Misconceptions 12a, 12d, 13f, and 15a also suggest that students do not recognize the importance of comparing relative values of the reduction potentials in predicting the chemistry that occurs in electrochemical cells. One reason why students may not understand the relative nature of electrochemical cell potentials is the way many freshman-level chemistry textbooks present the calculation of cell potentials. These textbooks instruct the students to change the sign of the reduction potential for the oxidation reaction and add it to the reduction potential for the reduction reaction ($E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red}$) instead of calculating the potential difference of the electrons at the cathode versus the anode ($E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$). Runo and Peters (*18*) present an alternative method of calculating cell potentials that emphasizes the relative nature of electrochemical measurements and uses a number line or 'potential ladder' as a visual aid to help students with the calculations.

Both Garnett, Garnett, and Treagust (*15*) and Ogude and Bradley (*16*) suggest that a major source of student misconceptions comes from imprecise or inappropriate language used by textbooks in explaining electrochemical concepts and our study is no exception. One-third of the students in our interview study stated that the electrodes in electrochemical cells have a net charge. Analysis of the three textbooks used by these students (*19-21*) revealed that two have statements that suggest the electrodes in electrochemical cells are charged.

Technical Flaws in Garnett and Treagust's Study. In the electrolysis of aqueous $NiBr_2$ (questions 22-28), Garnett and Treagust stated that the products should be nickel metal and oxygen gas (from the reaction of water) and suggested that students who chose nickel metal and bromine as products must have completely ignored the possibility of water reacting; however, the correct products are nickel and bromine and therefore those students who chose these products may not have ignored the fact that water can react. In our replication of Garnett and Treagust's study, we changed the problem to the electrolysis of aqueous aluminum bromide, which should produce hydrogen gas (from the reaction of water) and bromine. The fact that Garnett and Treagust performed complete transcriptions of their student interviews and that eight of the twelve students responding to question 25 in this study also ignored the possibility that water could react suggest that their identification of *Misconception 13a* (In electrolytic cells, water is unreactive towards oxidation and reduction) is valid.

Future Studies. This study and previous electrochemistry studies (*1, 14, 16*) provide us with a rich knowledge base of student misconceptions in electrochemistry. Now that an extensive list of electrochemistry misconceptions has been compiled, the authors contend that it is time to focus on identifying the sources of these misconceptions and working to prevent and reverse the formation of these misconceptions. In an attempt to address the first concern, the authors of this study continue to examine introductory chemistry textbooks as a source of student misconceptions in electrochemistry. Perhaps a more precise written presentation can be developed that will avoid misleading statements while providing accurate information.

As of now, no studies have attempted to determine whether an instructor, using a list of common student misconceptions, can prevent or reverse the formation of student misconceptions simply by being aware of them and by teaching to actively confront and resolve them. The use of a conceptual change approach (*22, 23*) may lead to a significantly lower proportion of students demonstrating electrochemistry misconceptions. Perhaps an interactive multimedia software program coupled with a laboratory exercise can be developed that will provide students with an environment to explore electrochemistry. The use of computer animations to illustrate the migration of cations and anions in a salt bridge has been developed by Greenbowe (*24*) and research is being conducted to see if these animations help students to understand that electrons do not travel through aqueous solutions. The role of the particulate nature of matter coupled with computer animations (*25*) provides a firm basis to design instruction for electrochemical cells.

The theory of constructivism has received a great deal of attention in the past few years (*6, 8*). Perhaps an electrochemistry instructional unit can be designed that incorporates the main tenets of constructivism.

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STUDENT MISCONCEPTIONS IN ELECTROCHEMISTRY REGARDING CURRENT FLOW IN ELECTROLYTE SOLUTIONS AND THE SALT BRIDGE

A paper accepted for publication by the *Journal of Chemical Education*

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Abstract

This article reports student misconceptions concerning the flow of current in electrolyte solutions and the salt bridge that were identified in an interview study concerning electrochemical cells (*1*). The authors confirmed most of the misconceptions reported by Garnett and Treagust (*2*) concerning the flow of current in electrochemical cells (which, for the most part, included the notion that electrons are stable and can migrate in aqueous solutions) while identifying several new misconceptions, including the notions that electrons can flow through aqueous solutions without assistance from the ions and that only anion migration constitutes a flow of current in electrolyte solutions. The students in this study proposed two mechanisms for electron flow in electrolyte solutions. In the first mechanism, electrons attach themselves to ions, which shuttle electrons from the cathode to the anode in solution. In the second mechanism, electrons flow by themselves from the cathode to the anode in solution without assistance from the ions. Imprecise or inaccurate language used by textbook authors has been implicated as a possible source of student misconceptions. Representative quotes from the three textbooks used by the students in this study provide exam-

ples of misleading statements that could be misinterpreted or over-generalized by students to inappropriate situations. This study also included data suggesting that instruction using computer animations and a confrontational teaching approach can reduce the number of students consistently demonstrating this misconception that electrons flow in electrolyte solutions and the salt bridge.

Introduction

Several researchers have documented student misconceptions in electrochemistry. One reason for the interest in studying electrochemistry is that student and teacher surveys suggest that students find this topic difficult (*3*), and research confirms that student beliefs about problem complexity affect student performance and learning (*4*). Several articles have been written that promote pedagogical suggestions or opinions about more effective methods of teaching electrochemistry (*5-8*) although few, if any, of these methods have been empirically tested.

Allsop and George (*9*) reported that students had difficulty using standard reduction potentials to predict the direction of chemical reactions, were unable to produce an acceptable diagram of an electrochemical cell, and 11% of these students stated that a salt bridge provides a pathway for the flow of electrons. Ogude and Bradley (*10*) noted that although many students can solve quantitative electrochemical problems that appear on most chemistry exams, few were able to answer qualitative questions requiring a deeper conceptual knowledge of electrochemistry. College student responses reported in this study demonstrate that 30% of these students consistently replied that electrons cannot flow in the electrolyte, while 28% consistently replied that electrons can flow in the electrolyte (42% were inconsistent in their responses). Similar results were found in the 25th National Youth Science Olympiad in South Africa in 1989 (reported in *10*) in which 30% of the students suggested that ions flow to complete the circuit in the electrolyte solution while 61% suggested that electrons flow in the electrolyte.

Garnett and Treagust (*2, 11*) probed student misconceptions on oxidation-reduction reactions and electrochemical and electrolytic cells through interviews using high-school students in Australia. These authors reported several common student misconceptions concerning oxidation-reduction reactions, electrochemical cells, and electrolytic cells. Misconceptions concerning the flow of current in electrolyte solutions and the salt bridge include the notions that: (1) Electrons move through the electrolytes and the salt bridge, being carried or transferred by cations and anions; (2) Protons move through the electrolytes and the salt bridge, even in neutral or basic solutions; and (3) Ion movements in solution do not constitute an electrical current. Garnett, Garnett, and Treagust (*12*) discussed some probable origins of these misconceptions and the implications on improving chemistry curriculum.

The authors of this study have replicated, with additions, Garnett and Treagust's interview study (*2*) to probe student misconceptions concerning galvanic (electrochemical), electrolytic, and concentration (Nernst) cells, in which the responses of sixteen student volunteers (9 male and 7 female) from three freshman-level chemistry courses at a midwestern American university were reported (*12*). The first part of this article focuses on the observed student misconceptions and proposed mechanisms concerning current flow in electrolyte solutions and the salt bridge (summarized in Table 1) and likely sources for these misconceptions. The second part of this article reports the results of a study to determine whether teaching to actively confront the misconception that electrons flow in solution using computer animations will decrease the number of students consistently demonstrating this misconception.

Table 1. Common Student Misconceptions Mentioned in This Study

2h.	Electrons move through solution by being attracted from one ion to the other.
2i.	Electrons move through solution by attaching themselves to ions at the cathode and are carried by that ion to the anode.
10a.	Electrons enter the solution at the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit.
10b.	Anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode.
10c.	Cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode.
10e.[a]	Electrons can flow through aqueous solutions without assistance from the ions.
10f.[a]	Only negatively-charged ions constitute a flow of current in the electrolyte and the salt bridge.
11b.	The anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons.

[a] Not previously reported by Garnett and Treagust (*2, 11*).

Current Flow through Electrolyte Solutions and the Salt Bridge

In general, students recognize that current cannot flow without a closed circuit and many students believe that only electron flow can complete this circuit. Consequently, many students cling to the notion that electrons flow from the anode to the cathode along the wire and are then released into the electrolyte at the cathode, traveling through the electrolyte solutions and the salt bridge to reach the anode. This is represented as *Misconception 10a*[1] (electrons enter the

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[1] The numbering scheme for the misconceptions presented here is consistent with that used by Garnett and Treagust (*2, 11*) and Sanger and Greenbowe (*1*). Representative student quotes for these misconceptions were presented by the authors at the ACS National Meeting at Chicago in August, 1995 (*13*).

solution from the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit), which was exhibited in one form or another by nine of the sixteen students.

Of the students who believed that electrons flow through the salt bridge, two students stated that anions in the electrolyte solutions and the salt bridge help transfer the electrons, which is represented as *Misconception 10b* (anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode); three students stated that cations transfer the electrons through the salt bridge, consistent with *Misconception 10c* (cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode); and three students stated that the electrons flow through solution without any assistance from anions or cations, exhibiting *Misconception 10e* (electrons can flow through aqueous solutions without assistance from the ions).

Three students who correctly stated that ions flow through solutions and the salt bridge to complete the circuit suggested that only the flow of anions in solution complete the circuit and that cation flow does not constitute a current, which is consistent with *Misconception 10f* (only negatively-charged ions constitute a flow of current in the electrolyte and the salt bridge).

In their responses to questions about electrochemical and electrolytic cells, seven of the sixteen students responded with comments suggesting that the electrodes have net positive and negative charges. Some of the students who believed that the anode is positively charged demonstrated *Misconception 11b* (the anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons), in which they interpreted anion flow toward the anode as suggesting that the anode is positively charged and cation flow toward the cathode as suggesting that the cathode is negatively charged.

Mechanisms for Electron Transfer through Electrolyte Solutions and the Salt Bridge

Eight of the nine students who stated that electrons flow in electrolyte solutions and the salt bridge suggested possible mechanisms for the flow of electrons. Five students stated that the electrons are transferred from the cathode to the anode in solution by the ions in solution, which is represented as *Misconception 2i* (electrons move through solution by attaching themselves to ions at the cathode and are carried by that ion to the anode). Four of these students stated that cations (Ag^+ and K^+ in the galvanic cell and Al^{3+} in the electrolytic cell) assisted in the transfer of electrons from cathode to anode, consistent with *Misconception 10c*, while one student stated that anions helped in the transfer of electrons from cathode to anode. None of the students in this study demonstrated *Misconception 2h* (electrons move through solution by being attracted from one ion to the other), which was originally reported by Garnett and Treagust (*9*) in which a student suggested that electrons are transferred back and forth from anion to cation as they travel from cathode to anode in solution. Three students who stated that electrons flow in electrolyte solutions and the salt bridge suggested that the electrons receive no assistance from the ions and travel as free electrons from the cathode to the anode (*Misconception 10e*).

In contrast to Garnett and Treagust's students (*11*), none of our students demonstrated *Misconception 2h*, in which electrons in solution are transferred from the cathode to the anode by "piggybacking" from anions to cations, and several of our students demonstrated *Misconception 10e*, in which electrons travel in solution as free electrons from the cathode to the anode. While it is interesting to note these differences, we are unable to attribute them to developmental (high school versus college students) or pedagogical (different teaching methods in Australia versus the United States) differences.

Probable Sources of Misconceptions

Garnett and Treagust (2) proposed two origins of student misconceptions that are relevant to the flow of current in electrolyte solutions and the salt bridge: (1) *Students' interpretation of the language of science*, in which students interpret the terminology used in the textbook or by the instructor in a manner consistent with everyday usage, but inconsistent with the scientific usage; and (2) *Students applying information too generally*, in which students over-generalize a scientific statement to situations where it is inappropriate. From these proposed origins, Garnett, Garnett, and Treagust (*12*) drafted several implications for improving the chemistry curriculum that included the following ideas: (1) *Teachers and curriculum developers need to select explanatory language with care, and be particularly cautious in selecting language which has everyday meanings which differ from meanings in a scientific context*; and (2) *Teachers and curriculum developers need to be cautious in making unqualified, generalized statements about concepts because students tend to interpret the statements literally, and apply them more extensively than is intended.*

Ogude and Bradley (*10*) attributed student misconceptions concerning current flow in electrolyte solutions and the salt bridge to two factors: (1) *Reference to continuity of current and established belief in the electronic nature of current* *electricity*, in which textbook or instructor phrases like 'continuity of current' imply to the students that current throughout the electrochemical cell is the same; and (2) *Language and careless discussion of electrode processes*, in which textbooks with obvious mistakes or misleading statements lead students to misinterpret the intent of the textbook, resulting in student misconceptions.

Both Garnett and Treagust (*2, 11*) and Ogude and Bradley (*10*) suggested that a major source of student misconceptions comes from imprecise or inappropriate language used by textbooks and instructors in explaining electrochemical concepts and this study is no exception. Over half (9 of 16) of the students in this interview study suggested that electrons flow in electrolyte solutions and the salt bridge to complete the circuit. Analysis of the three textbooks used by these students (*14-16*) revealed that each of them has comments that, while not technically incorrect, may be misinterpreted by students to suggest that electrons do flow through electrolyte solutions and the salt bridge.

> 1. "In a molten salt such as sodium chloride, or in a solution of an electrolyte, however, electrical charge is carried through the liquid by the movement of ions. The transport of electrical charge by ions is called **electrolytic conduction**, and it is able to occur only when chemical reactions take place at the electrodes." (*14*, p 770). If students interpret 'electrical charge' as 'electrons' instead of 'the inherent charge of the ions', the first sentence could lead to *Misconceptions 10b* and *10c* and the second sentence could foster *Misconception 2i* about the transfer of electrons through electrolyte solutions and the salt bridge.
>
> 2. "...This task is accomplished through a **voltaic** (or **galvanic**) **cell,** which is merely a device in which electron transfer is forced to take place through an external pathway rather than directly between reactants." (*15*, p 727). Electrons are being transferred from the reductant to the oxidant, but 'electron transfer through an external pathway' can be misinterpreted as suggesting that electrons flow throughout the entire circuit including the electrolyte solutions and the salt bridge and may be responsible for *Misconceptions 10a* or *10e*.

> 3. "...If we physically separate the oxidizing agent from the reducing agent, the transfer of electrons can take place via an external conducting medium. As the reaction progresses, it sets up a constant flow of electrons and hence generates electricity (that is, it produces electrical work)." (*16*, p 767). The 'constant flow of electrons' occurs only in the wire connecting the electrodes, but student may over-generalize this statement to the flow of current in electrolyte solutions and the salt bridge, which represent *Misconceptions 10a-e.*

The results of this analysis should prompt textbook authors to carefully examine and reconsider the language used in their chemistry textbooks. The use of detailed diagrams and animations about current flow through electrolyte solutions and the salt bridge should be included in a multimedia presentation to help students visualize these concepts. Park and Hopkins (*17*) report that dynamic visual displays are more effective than static visual displays.

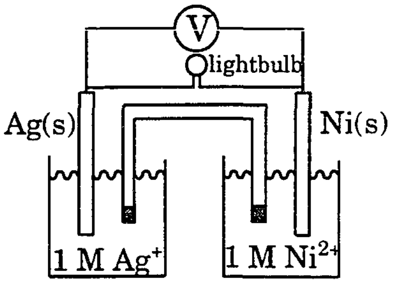
Some Methods for Preventing Misconceptions

The authors of this paper have become increasingly interested in the use of computer animations as a lecture tool to enhance students' abilities to visualize and understand chemical concepts on the molecular level. A typical lecture involves the instructor performing a live chemical demonstration, writing the relevant balanced chemical equation(s) on the chalkboard, and showing and verbally explaining a computer animation that depicts the reaction on the molecular level. In this way, the lectures attempt to facilitate students' connection of the macroscopic, symbolic, and microscopic representations of chemical processes (*18-20*). Examples of electrochemistry animations used in these lectures have been reported previously by Greenbowe (*21*).

Preliminary studies to determine whether we can reduce the number of students demonstrating the misconception that electrons flow in the salt bridge by teaching to actively confront, and therefore prevent or dispel, this misconception are encouraging. Students in the second semester of the introductory chemistry course intended for non-science majors answered three conceptual questions concerning the flow of electrons in electrolyte solutions and the salt bridge (Figure 1) after receiving instruction on electrochemistry. In the lectures, the instructor explicitly emphasized that electrons do not flow in electrolyte solutions or the salt bridge and showed several computer animations that modeled the correct flow of current in galvanic and electrolytic cells (i.e., electron flow in the wires and cation and anion flow in the electrolyte solutions). Table 2 contains a description and an approximate running time of the computer animations used in these lectures. Each animation was displayed three times in succession and the instructor provided a narration of events during this time. Presenting visual and verbal (oral and written) information simultaneously is consistent with Paivio's dual coding theory (*22*) and Mayer and Anderson's contiguity principle (*23*).

The animations on the salt bridge included an overall animation that focused on the dynamics of the entire copper-zinc electrochemical cell, and a "close-up" view of the cation and anion migration out of the salt bridge (Figure 2). The animation clearly shows only ions migrating in solution. This "close up" view allows students to focus their attention on the critical concept being illustrated (*24*).

The computer animation of the electrolytic cell illustrates the plating of silver metal on an iron spoon (Figure 3). The animation clearly shows electrons being released at the anode, bumping up from the anode through the wire and the battery to the cathode. Silver ions in solution migrate toward the iron cath-



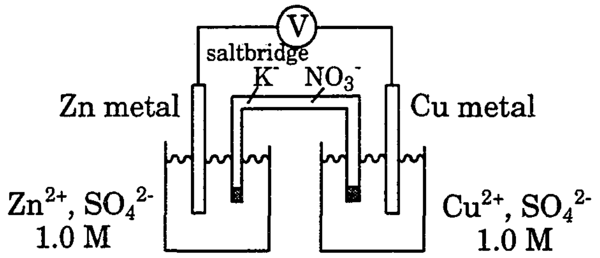
1. Electrons in the cell flow through the ______, toward the ______.

   1) wire, silver electrode
   2) wire, nickel electrode
* 3) salt bridge, nickel electrode
* 4) salt bridge, silver electrode

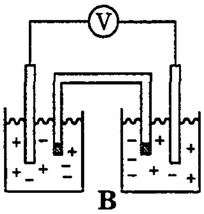
2. In an electrochemical cell, conduction through the electrolyte is due to:

* 1) electrons moving through the solution attached to the ions.
* 2) electrons moving from ion to ion through the solution.
   3) the movement of both positive and negative ions.
   4) the movement of water molecules.
* 5) electrons moving through the solution from one electrode to the other.

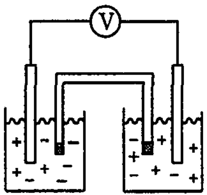
3. The electrochemical cell shown below has 1.10 volts for its e.m.f. There is an oxidation reaction and a reduction reaction.



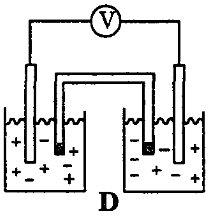
Which one(s) of the following diagrams below depict each half-cell as the reactions proceed? Note: In the following diagrams, a cation is symbolized as + and anions as –. An electron is symbolized as e-.



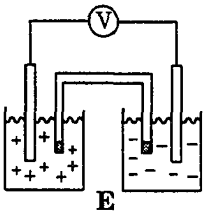
B



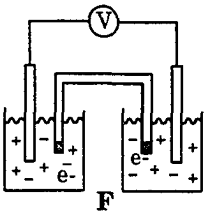
C



D



E



F

1) Either C and D
2) E only
3) B only
4) Either B and E
* 5) F only

Figure 1. Conceptual questions concerning the flow of electrons in solution.

Table 2. Animations Used in the Electrochemistry Lectures

Animation	Focus	Length
Zinc-copper electrochemical cell	Illustrates the dynamics of the entire cell: *ion migration in the electrolyte* solutions and the salt bridge, movement of electrons in the wire, and the oxidation-reduction reactions at the electrodes	45 seconds
Salt bridge (part I)	Emphasizes the cation and anion migration out of the salt bridge	30 seconds
Salt bridge (part II)	Demonstrates the cation and anion migration out of the salt bridge and the charge balance in each half-cell	30 seconds
Electroplating silver onto iron	Illustrates electron movement in the wires, ion migration in the aqueous solution, the oxidation process at the anode, and the reduction process at the cathode	45 seconds

ode (spoon) where they capture electrons at the solution-metal interface, plating out on the electrode as silver metal. The animation clearly shows that only ions migrate in solution.

The distractors in each question were classified as being consistent (marked with an asterisk in Figure 1) or inconsistent with the misconception that electrons flow in electrolyte solutions. Student responses to the three conceptual questions were analyzed to determine whether students were consistent in demonstrating or failing to demonstrate this misconception. Of the 112 students who took the final exam, 3 (3%) consistently chose responses suggesting that electrons are present in solution, 40 (36%) consistently chose responses that did not suggest that electrons exist in solution, and 69 (61%) chose responses that were inconsistent regarding the presence of electrons in electrolyte solutions.

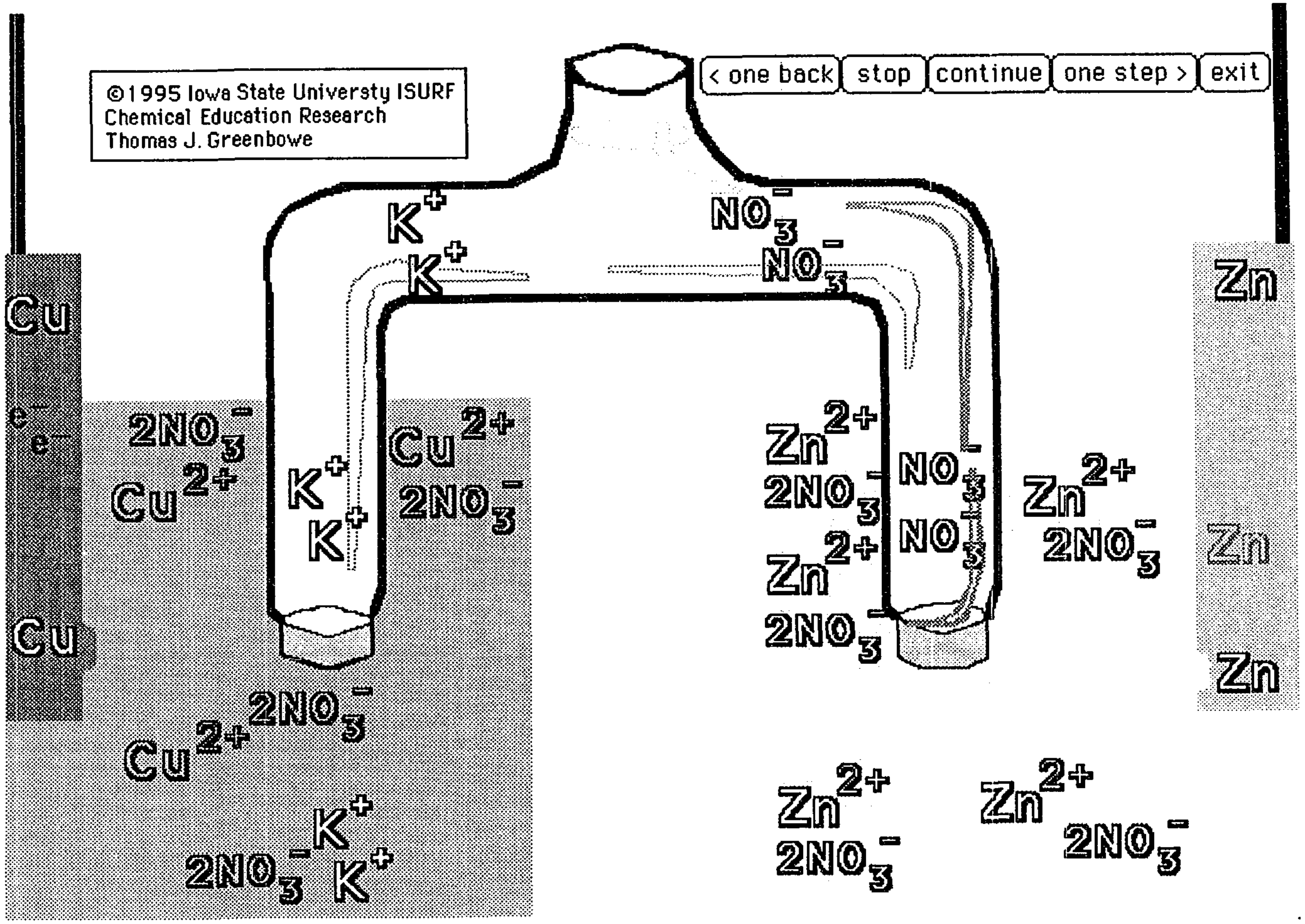


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Figure 2. Close up view of cation and anion migration in the salt bridge of a copper/zinc electrochemical cell.

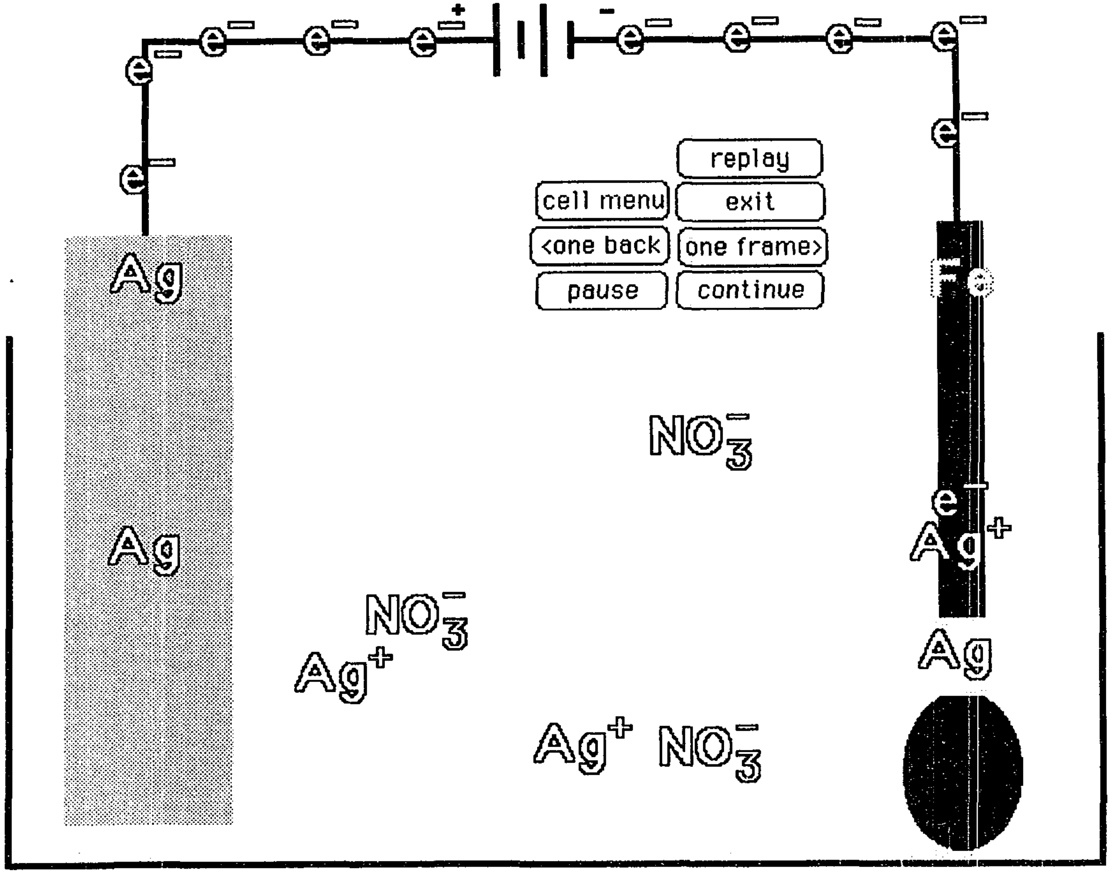

Figure 3. A computer screen image of the electrolytic plating of silver metal onto an iron spoon.

These results can be compared to those reported by Ogude and Bradley (*10*), in which 40 first-year college students answered five conceptual questions concerning the flow of electrons in electrolyte solutions: 11 (28%) consistently demonstrated the misconception; 12 (30%) consistently did not demonstrate the misconception; and 17 (42%) were inconsistent in their responses. These numbers are complicated by the fact that Ogude and Bradley also used question #2 in Figure 1, but they included only response **5)** as being consistent with the misconception; we included responses **1)**, **2)**, and **5)** because each of them suggest that electrons exist in solution. Since 8 of the 40 students in Ogude and Bradley's study chose responses **1)** and **2)** for this question, it is likely that <u>more</u> than 28% consistently demonstrated and <u>less</u> than 30% consistently did not demonstrate this misconception.[2]

A chi-square test of independence was performed on the number of students in each study consistently demonstrating or failing to demonstrate the misconception, as well as those that were inconsistent in their responses. The results of this test ($\chi^2(2) = 21.90, p < .0001$) support the assumption that our teaching method had an effect on the proportion of students consistently demonstrating this misconception. Specifically, the test of independence suggests that the proportion of students in Ogude and Bradley's study consistently demonstrating the misconception is larger than expected and the proportion of students in our study consistently demonstrating the misconception is smaller than expected if the two groups were equivalent.

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[2] If we reanalyze our data using only response **5)** in question #2 as being consistent with the misconception, it drastically changes our proportions: 1% consistently demonstrate the misconception, 58% consistently do not demonstrate the misconception, and 41% are inconsistent in their responses.

This study suggests that teaching to actively confront the misconception that electrons flow in electrolyte solutions and the salt bridge, including the use of computer animations that help students visualize chemical reactions at the molecular level, decreased the proportion of students consistently demonstrating this misconception. The effect of viewing computer animations that focus student attention on the molecular level (particulate nature of matter) can be seen in Williamson and Abraham's study *(20)*, in which students viewing animations based on the states of matter and reactions in solution showed an increased ability to visualize the particulate behavior of chemical reactions.

Even though the test of independence does not suggest a difference in the proportion of students inconsistently demonstrating the misconception, these numbers should be addressed. Specifically, since Ogude and Bradley's students answered five questions while our students answered only three, it is not unreasonable to expect a larger inconsistent group in their study due to random effects; however, our study shows a larger proportion of students that were inconsistent in their responses. Even though our students received instruction in electrochemistry that emphasized the correct model of current flow in electrolyte solutions and the salt bridge and were directed to readings about this topic in their textbook, previous experience suggests that this misconception is resistant to change. Perhaps the computer animations displayed in lecture were not shown long enough for students to process the information. Future research is needed to determine whether these computer animations adequately explained students' experiences and observations and appeared logical to the students *(25)*.

Many of our students report that while the computer animations are useful, capturing the dynamic aspects of these processes on paper is difficult. More time is needed by students to view the animations, make sense of them, and to copy important information derived from them into their notes. To address these student concerns, we will place the computer animations used in lecture on our chemistry file server next semester. In this way, students will have access to the animations twenty-four hours a day and can review the animations and take notes at their leisure. We will also prepare a lecture handout that will reduce the time it takes for students to copy the computer animations into their notes.

Perhaps the viewing of computer animations help students build a better mental model (*26-27*) of electrochemical cells. Further research is needed to investigate this issue.

Summary

In a replication of Garnett and Treagust's interview study concerning electrochemical cells (*2*), we were able to confirm most of the student misconceptions reported concerning the flow of current in electrochemical cells while identifying several new misconceptions, including the notions that electrons can flow through aqueous solutions without assistance from the ions and that only anions constitute a flow of current in electrolyte solutions. Our students suggested two mechanisms for electron flow in electrolyte solutions and the salt bridge: electrons can either attach themselves to ions in solutions or they can flow by themselves without any assistance from the ions. Analysis of the textbooks used by our students provide a possible source of these misconceptions: obvious mistakes or misleading statements made by textbooks that can be misinterpreted or overgeneralized to inappropriate situations. We were also able to demonstrate that instruction including the use of computer animations aimed at confronting the misconception that electrons flow in electrolyte solutions and the salt bridge can reduce the number of students consistently demonstrating this misconception.

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ANALYSIS OF COLLEGE CHEMISTRY TEXTBOOKS AS SOURCES OF MISCONCEPTIONS AND ERRORS IN ELECTROCHEMISTRY

A paper submitted for publication to the *Journal of Chemical Education*

Michael J. Sanger and Thomas J. Greenbowe

Abstract

In this study, the oxidation-reduction and electrochemistry chapters of ten introductory college-level chemistry textbooks were analyzed for examples of statements or drawings that could be misinterpreted by students and could foster common student misconceptions in electrochemistry. The misconceptions investigated in this study were reported previously by Sanger and Greenbowe (*1*). Misconceptions that were supported by comments or illustrations from several textbooks include the notions that half-cell placement determines the anode/cathode identity, half-cell potentials are absolute and/or additive in nature, electron migration in aqueous solutions constitutes a flow of current, cation migration in aqueous solutions does not constitute a flow of current, electrode charges determine the flow of electrons and ions in a cell, and electrolysis products cannot be predicted. As a result of analyzing these textbooks, the authors proposed five suggestions for textbook authors. These suggestions included avoiding the use of simplifications, avoiding the use of vague or misleading statements, calculating cell potentials using the difference method, avoiding the use of electrostatic arguments to predict ion and electron flow in electrochemical cells, and always considering all possible oxidation-reduction half-reactions when predicting elec-

trolysis products. The authors also proposed a method for predicting electrolysis products using potential ladder diagrams.

Introduction

The analysis of introductory college chemistry textbooks as sources of student misconceptions in electrochemistry was inspired by a student comment made during a clinical interview in electrochemistry (*1*). The student (S1642) was asked to identify the anode and cathode of an electrochemical cell consisting of a Ni/Ni^{2+} half-cell on the left and a Ag/Ag^{+} half-cell on the right.

S1642:	I was just told that this would be the anode on the left [Ni] and the cathode on the right [Ag].
Interviewer:	But what if we gave you a diagram like this? [Reversing the half-cells]
S1642:	Well then, it [nickel] would be reducing—the cathode, right. That's just so far how the book has shown it to me and the way on the board it's been shown.

Analysis of textbooks for misleading or erroneous statements is warranted for another reason: The singular importance of textbooks in the classroom as a source of both content and curriculum (*2*). The predominant role of textbooks in the classroom was first documented by a National Science Foundation (NSF)-supported needs assessment conducted in the mid-seventies (*3-5*). These studies reported that teachers view the textbook not only as being the sole source of information regarding the subject matter, but also as deciding the content covered in the classroom. The 1985-1986 National Survey of Science and Mathematics Education (*6*) reported that more than 90% of secondary science classes use published textbooks and that a majority of science teachers report covering 75% or more of the textbook content in their classes. Chiang-Soong and Yager (*7*) reported that students expected the textbook to be the source of nearly all classroom information and as the framework from which science was experienced, while parents expressed concern if textbooks were not issued and used as the central source of assignments. Although these reports were based on elementary and secondary students, these results are likely to be valid for college students as well.

Several analyses of textbooks as sources of student misconceptions have been reported. Cho, Kahle, & Nordland (*8*) performed an analysis of high school biology textbooks as sources of student misconceptions in genetics. Fourteen chemistry textbooks and fourteen physics textbooks used in Australia were analyzed by de Berg (*9*) for scientific literacy on the topic of Boyle's Law (pressure-volume relationship for gases). Arditzoglou and Yani (*10*) analyzed fifth and sixth grade science textbooks to identify elementary science teachers' misconceptions in life and physical sciences. Barman and Mayer (*11*) analyzed the effect of high school textbooks on student conceptions of food chains and food webs. Textbook analyses have also been performed on life science and biology textbooks to determine factors that may deter girls' interest in science (*12*) and the effects of the creationism movement on the teaching of science and the theory of evolution (*13*).

The theory of constructivism recognizes that students bring a set of preconceptions based on their previous knowledge and experiences to any new learning situation (*14, 15*). Student preconceptions that are inconsistent with or different from the commonly-accepted scientific consensus and are unable to adequately explain observable scientific phenomena are defined as *misconceptions* (*8, 14*). It is important to note that some student misconceptions are capable of adequately explaining the student's experiences and observations, appear quite logical to the student, and are consistent with his or her understanding of the world. In these instances, student misconceptions are very resistant to change (*16*).

Several researchers have documented student misconceptions concerning electrochemistry (*1, 17-22*) and most of these researchers (*1, 18-22*) have implicated textbook and/or instructor comments as possible sources of student misconceptions. The list of misconceptions in electrochemistry used in this study was prepared by combining the lists of student misconceptions reported by Garnett and Treagust (*20*) and Sanger and Greenbowe (*1*) and appears in Table 1.

Method

Ten college-level chemistry textbooks were analyzed for examples of statements or drawings that could be misinterpreted by students as evidence corrob-

Table 1. Common Student Misconceptions Reported in Garnett and Treagust (*20*) and Sanger and Greenbowe (*1*)

Galvanic cells

8a. In an ordered table of reduction potentials, the species with the most positive E° value is the anode.
8b. Standard reduction potentials list metals by decreasing reactivity.
8c. The identity of the anode and cathode depends on the physical placement of the half-cells.
8d. Anodes, like anions, are always negatively charged; cathodes, like cations, are always positively charged.
9a. The fact that the E° for H_2(*1 atm*)/H^+(*1 M*) is zero is somehow based on the chemistry of H^+ and H_2.
9b. There is no need for a standard half-cell.
9c. Half-cell potentials are absolute in nature and can be used to predict the spontaneity of the half cells.
10a. Electrons enter the solution from the cathode, travel through the solutions and the salt bridge, and emerge at the anode to complete the circuit.
10b. Anions in the salt bridge and the electrolyte transfer electrons from the cathode to the anode.

Table 1 (continued)

10c. Cations in the salt bridge and the electrolyte accept electrons and transfer them from the cathode to the anode.
10d. Cations and anions move until their concentrations are uniform.
10e. Electrons can flow through aqueous solutions without assistance from the ions.
10f. Only negatively-charged ions constitute a flow of current in the electrolyte and the salt bridge.
11a. The anode is negatively charged and releases electrons; the cathode is positively charged and attracts electrons.
11b. The anode is positively charged because it has lost electrons; the cathode is negatively charged because it has gained electrons.
14a. Cell potentials are derived by adding individual reduction potentials.
14b. Half-cell potentials are not intensive properties.

Electrolytic cells

12a. In electrolytic cells, the direction of the applied voltage has no effect on the reaction or the site of the anode and cathode.
12b. No reaction will occur if inert electrodes are used.
12c. In electrolytic cells, oxidation now occurs at the cathode and reduction occurs at the anode.
12d. In electrolytic cells with identical electrodes connected to the battery, the same reactions will occur at each electrode
13a. In electrolytic cells, water is unreactive towards oxidation and reduction.
13b. When predicting an electrolytic reaction, the half-cell reactions are reversed prior to combining them.
13c. The calculated cell potentials in electrolytic cells can be positive.
13d. There is no relationship between the calculated cell potential and the magnitude of the applied voltage.
13e. Inert electrodes can be oxidized or reduced.
13f. When two or more oxidation or reduction half-reactions are possible, there is no way to determine which reaction will occur.
13g. Electrolytic cells can force non-spontaneous reactions that do not involve electron transfer to happen.

Concentration cells

15a. The direction of electron flow in concentration cells is not dependent on the relative concentration of the ions.
15b. The products produced in the indirect reaction of electrochemical cells are different from those produced in the direct reaction of the starting materials.
16a. The cell potential in concentration cells is not dependent on the relative concentration of the ions.
16b. Because there is no net reaction in concentration cells, the reaction quotient cannot be calculated.

orating a common student misconception in electrochemistry. These textbooks include: *General Chemistry* by Atkins and Beran (*23*), *Chemistry* by Birk (*24*), *Chemistry: An Experimental Science* by Bodner & Pardue (*25*), *Chemistry: The Study of Matter and Its Changes* by Brady & Holum (*26*), *Chemistry: The Central Science* by Brown, LeMay, and Bursten (*27*), *Chemistry* by Chang (*28*), *Chemistry & Chemical Reactivity* by Kotz and Treichel (*29*), *Chemistry: The Molecular Science* by Olmsted and Williams (*30*), *Chemistry: Principles & Practice* by Reger, Goode, and Mercer (*31*), and *Chemistry* by Zumdahl (*32*). These textbooks will be subsequently referred to by the authors' initials (for example, *Chemistry: Principles & Practice* will be referred to as *RGM*).

The oxidation-reduction and electrochemistry chapters of these textbooks were reviewed using the collection of common student misconceptions listed in Table 1 for misleading or erroneous statements and for statements that do not foster these misconceptions. For each textbook, these statements were transcribed on a single sheet of paper and were classified by the misconception that they did or did not foster. The textbook statements relating to each misconception were summarized on a single sheet of paper and were analyzed for common and unique examples of both misleading statements and statements that would not lead to the misconception. The statements quoted in this paper and the interpretations of these statements were reviewed by four college chemistry professors and their comments were incorporated in the analysis and critique of the textbooks.

Results and Discussion

Analysis of the oxidation-reduction and electrochemistry chapters of these textbooks provided misleading statements and illustrations related to *Miscon-*

ceptions 8c, *9c*, *10a*, *10b*, *10c*, *10e*, *10f*, *11a*, *11b*, *14a*, *14b*, *13e*, and *13f* that could be misinterpreted by students. The misconceptions are labeled from 8 to 16 and are consistent with the labeling method used by Garnett and Treagust (*20*) and Sanger and Greenbowe (*1*). Not all of the misconceptions reported previously (*1*, *20*) were supported by vague or misleading statements in these textbooks: Several misconceptions (*Misconceptions 8a*, *9a*, *9b*, *12a*, *12b*, *12c*, *13a*, *13b*, *13d*, *15b*, and *16a*) were directly refuted by every statement found in the textbooks while other misconceptions (*Misconceptions 8b*, *8d*, *10d*, *12d*, *13c*, *13g*, *15a*, and *16b*) were neither corroborated nor refuted by comments in the textbooks because they were not addressed by a majority of these textbooks. For the misconceptions that remain, the statements made concerning these misconceptions and the implications of these statements are discussed in detail and serve as the basis for this paper.

Misconception 8c: Half-Cell Placement. The notion that the identity of the anode and cathode depend on the placement of the half-cells was suggested by a student's observation that the textbook and the instructor always drew the anode half-cell on the left and the cathode half-cell on the right. The relative placement of the anode and cathode half-cells as depicted in the illustrations in each textbook was analyzed, ignoring drawings of commercial batteries, commercial electrolysis cells, and the reactions involving the corrosion of metals (iron). Two textbooks (*B*, *KT*) had an equal number of illustrations in which the anode was drawn as the left and as the right half-cell, four textbooks favored one position over the other (*AB*, *OW*, and *Z* preferentially drew the anode as the left-hand half-cell while *BH* preferentially drew the anode as the right-hand half-cell), and four textbooks (*BP*, *BLB*, *C*, *RGM*) exclusively drew the anode as the left-hand half-cell.

A second convention that may confuse students about the relationship between the placement of half-cells and the identity of the anode and cathode is the use of the cell notation (also called the line notation or the cell diagram). Five textbooks (*AB*, *BP*, *BH*, *C*, *Z*) introduce cell notation, which requires (by convention) that the anode be placed on the left-hand side and the cathode be placed on the right-hand side of the cell notation so that electrons flow through the wires from the left-hand side to the right-hand side (*33*). Only three of these textbooks (*BP*, *BH*, *C*) explicitly state that the arrangement of half-cells in the cell notation is done by convention. Another disadvantage of using the cell notation as a shorthand way to describe electrochemical cells is that it requires knowledge of the spontaneous cell reaction.

The placement of the anode and cathode with respect to the ground (–) and active (+) terminals of a voltmeter leads to the possibility of a similar misconception concerning the relationship between the half-cell identity and the polarity of the voltmeter. Two textbooks (*KT*, *RGM*) consistently show the anode connected to the ground terminal and the cathode connected to the active terminal of the voltmeter. However, while *KT* uses a drawing of an analog voltmeter (which can only report positive cell potentials) and includes a detailed discussion about why the anode must be connected to the (–) terminal, *RGM* uses a drawing of a digital voltmeter (which can report both positive and negative cell potentials) and includes no comment about the fact that this arrangement of electrodes is necessary only for a positive voltmeter reading. None of the textbooks discussed how to interpret a negative voltmeter reading—an event that students experience in laboratory experiments, in lecture demonstrations, and in computer animations (*34*).

While it may seem logical to consistently place the anode half-cell on the left-hand side and consistently connect it to the (–) or ground terminal of the voltmeter in all of the illustrations included in a textbook, this may be mislead students into believing that these are viable methods that can be used to identify the anode and cathode in electrochemical cells. Similarly, while the use of the cell notation may provide textbook authors with a short-hand way of describing electrochemical cells, it may also lead to the misconception that the anode is always the half-cell on the left-hand side of the cell. This difficulty appears when students are asked to analyze electrochemical cell diagrams on exams or to build and diagram cells in the laboratory. If cell notation is to be included in textbooks, authors should explicitly state that the half-cell arrangement is arbitrary and by convention and should not be considered a method for identifying the anode and cathode.

Misconception 9c: Absolute Half-Cell Potentials. Although all of the textbooks contain explicit statements that absolute half-cell potentials cannot be measured and that all standard reduction potentials are measured with respect to the standard hydrogen electrode (SHE), several of these textbooks contain statements that could be construed as suggesting that standard reduction potentials are absolute—i.e., reactions with positive reduction potentials are spontaneous and reactions with negative reduction potentials are non-spontaneous (and are spontaneous in the opposite direction). Eight of the ten textbooks (*AB*, *B*, *BP*, *BLB*, *C*, *KT*, *RGM*, *Z*) calculate cell potentials using the equation E°_{cell} =

$E°_{ox} + E°_{red}$, where $E°_{red}$ is the reduction potential of the reduction half-reaction and the $E°_{ox}$ value is determined by taking the reduction potential of the oxidation half-reaction and changing its sign.[1] Although this is technically the same as calculating the cell potential as a potential difference (i.e., $E°_{cell} = E°_{cathode} - E°_{anode}$, where $E°_{cathode}$ and $E°_{anode}$ are the standard reduction potentials of the reduction and oxidation half-reactions, respectively), calculating the cell potential as a sum may suggest to students that these individual half-cell potentials have significance—when in reality it is only the potential difference which has any significance. Students may also get the impression that each half-cell potential can be measured independently. Indeed, two textbooks (*AB*, *B*) have comments that can be interpreted as suggesting that the individual $E°_{ox}$ and $E°_{red}$ values are meaningful and measurable:

> "Therefore, of the 1.23 V from the current supply [needed for the electrolysis of water at pH = 7], 0.42 V is needed for the reduction of H_2 and 0.81 V is needed for the oxidation of O_2." (*AB*, p. 683).
>
> "[In the electrolysis of aqueous NaI] we see that it requires 0.54 V to oxidize iodide ion and 0.82 V to oxidize water..." (*B*, p. 793).

B makes another statement implying that individual reduction potentials are significant. After introducing the equation relating $\Delta G°$ to $E°_{cell}$ ($\Delta G° = -nFE°_{cell}$), the textbook calculates the free energy change of an individual reduction potential. Using the reduction potential of the Cu^{2+}/Cu half-cell (+0.34 V), a $\Delta G°$ value of -66 kJ is calculated (*B*, p. 783). Since negative $\Delta G°$ values imply

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[1] When the sign of the standard reduction potential is changed, it is converted into a standard oxidation potential. Although tables of standard oxidation and reduction potentials have been tabulated in the past, the International Union of Pure and Applied Chemistry (IUPAC) concluded that half-cell potentials should be reported for the reduction reactions because the signs of these reduction potentials correctly describe the polarity of the half-cells of interest with respect to the SHE electrode (*35*, pp 1115-1120).

spontaneity, this may suggest to students that this individual half-reaction is spontaneous; however, students need to be cautioned that since the initial reduction potential was referenced to the SHE, the ΔG° value calculated is actually the ΔG° value for the reaction of the Cu^{2+}/Cu half-cell with the SHE (i.e., $Cu^{2+} + H_2 \rightarrow Cu + 2\ H^+$).

Although cell potentials can be calculated by changing the sign of the anode reduction potential and adding that value to the cathode reduction potential, this may suggest to students that the cell potential can be thought of as the sum of the potentials of the anode and the cathode, each of which has significance and can be measured. Therefore, textbook authors should calculate cell potentials as potential differences. Runo and Peters (*36*) describe a particularly effective method of calculating cell potentials using the potential difference method and line diagrams called 'potential ladders.' The potential ladder is a vertical line diagram in which half-cells are arranged in order of increasing reduction potentials (i.e., the reactions with more positive reduction potentials on top and those with more negative values on bottom). In galvanic cells and for spontaneous chemical reactions, the cathode is always the upper 'rung' in the ladder and the anode is always the lower one. In contrast, electrolytic cells and non-spontaneous reactions always have the anode as the upper 'rung' on the ladder and the cathode as the lower 'rung'. The potential ladder is also effective at predicting the electrolysis products of aqueous solutions.

Textbook authors should make certain that they do not make comments that might suggest to students that individual half-cell potentials are significant or measurable. When calculating ΔG° values for oxidation or reduction half-reactions, authors should emphasize that these because these values were calculated using standard reduction potentials which are referenced to the SHE, they are also measured or calculated with respect to the SHE electrode.

Misconceptions 10a and 10e: Unassisted Electron Flow in Solution. The misconception that electrons can flow through an electrolyte solution is perhaps the most prevalent student misconception in electrochemistry *(1, 17, 19-22)*. *Misconceptions 10a* and *10e* focus on the notion that free electrons are stable in electrolyte solutions and that current flow in the electrolyte primarily consists of unassisted electron flow. Illustrations from the ten textbooks were analyzed, paying particular attention to oxidation-reduction reactions occurring at the metal-solution interface (simple electrochemical cells, corrosion reactions, etc.). Several of the textbooks included drawings that accurately depicted the electron transfer as occurring at the metal surface-solution interface (the more effective examples include *BP*, p. 755; *BLB*, pp. 757-758; *C*, p. 790; *KT*, p. 960; *OW*, pp. 831, 858, 861, 863; and *Z*, p. 848). However, three of the textbooks (*AB*, pp. 109, 656, 674; *OW*, pp. 830, 833; *Z*, p. 813) included drawings that depict electrons flowing into and through the electrolyte solutions from the metal (or oxidized species) to the reduced species. Although *OW* and *Z* both contain pictures depicting correct and incorrect current flow in solution, *AB* only contains pictures suggesting that free electrons can flow in solution.

After analyzing the illustrations for evidence suggesting that free electrons can travel in aqueous solutions, the texts were analyzed for similar misstatements. Two textbooks (*BH*, *Z*) include a thoughtful discussion about the charge imbalance that occurs in the half-cells when the salt bridge is not present and both explain that cations and anions migrate to neutralize the charge build-up. Unfortunately, statements suggesting that free electrons can exist in solution and that their migration accounts for the current flow in electrolyte solutions are prevalent in other textbooks. Some of these statements are quoted below.

> "As the reaction takes place, electrons are transferred from the zinc to Cu^{2+} ions nearby in the solution." (*AB*, p. 656).
>
> "Increasing the potential between the electrodes forces electrons into the electrolyte and can bring about a reduction..." (*AB*, p. 687).
>
> "...we have to build a model for **electrochemical reactions**—chemical reactions that involve the flow of electrons." (*BP*, p. 726).
>
> "An electrode conducts electrons into and out of a solution." (*KT*, p. 959).
>
> "One mole of nickel metal (58.69 g) is produced for each two moles of electrons that pass through the solution." (*RGM*, p. 756).

Although these statements may not technically be incorrect and are taken out of context, they are certainly susceptible to student misinterpretation.

The use of the ion-electron method of balancing oxidation-reduction reactions and the use of half-reaction in electrochemistry may also suggest to students that free electrons can exist in solution. Each textbook uses the ion-electron method of balancing oxidation-reduction reactions and each uses balanced half-reactions to describe electrochemical cells. Although each textbook made some statement cautioning students that the number of electrons gained by the oxidant must equal the number of electrons lost by the reductant and that free electrons cannot appear in the final balanced equation, only one textbook gave students an explicit explanation: "No substance is ever oxidized unless something else is reduced. Otherwise, electrons would appear as a product of the reaction, and this is never observed." (*BH*, p. 150).

Electrochemical illustrations included in textbooks should not contain arrows suggesting that electrons can flow in electrolyte solutions; instead, these drawings should visually emphasize that electron transfer occurs at the metal-solution interface. Textbook authors should be careful not to make statements that could be misinterpreted as suggesting that electrons can flow in solution; authors should also explicitly define or describe terms such as 'external pathway' or 'external circuit' as including the wire and other mechanical equipment (voltmeter, battery, motor, etc.) but excluding aqueous solutions, where the current flow occurs solely as ion movement. Finally, authors should use the ion-electron method and balanced half-equations only after making a disclaimer that the electrons written in these equations do not exist on their own—they are transferred directly from one species to another through physical contact or collisions.

Misconceptions 10b and 10c: Ion-Assisted Electron Flow in Solution. *Misconceptions 10b* and *10c* focus on the notion that current flow in the electrolyte consists of electrons that are being carried through the solution by anions or cations. Six textbooks (*AB*, *B*, *BH*, *KT*, *OW*, *RGM*) make comments that could be interpreted as implying that ions in solution can accept or deposit electrons at the surface of an electrode without undergoing any chemical change. Several statements that could foster these misconceptions are listed below.

> "Electrons cannot just be released; they must be transferred to some other atom or ion." (*B*, p. 727).
>
> "Ions in the solution can transfer electrons to or from this inert electrical conductor [inert electrode]." (*RGM*, p. 732).
>
> "The transport of electrical charge by ions is called **electrolytic conduction...**" (*BH*, p. 764).

> "The salts dissolved in the two vessels are essential for the flow of electrons because pure water does not contain enough charged species to support charge flow. Charge can flow only when the solutions contain ions that are provided by an **electrolyte.**" (*OW*, p. 831).

In the first two quotes, the authors do not emphasize that the process of transferring electrons to ions results in a chemical change of these ions: This is a transfer (not a transport) of electrons. In the second two quotes, the description of current in the electrolyte as the flow of 'electrical charge' or 'charge' by ions can be viewed as suggesting that ions are carrying electrons through the electrolyte.

When discussing chemical reactions occurring at an electrode, textbook authors should stress the idea that ions cannot transfer electrons to or from the electrode surface without undergoing chemical change. Similarly, authors should emphasize that current flow in electrolyte solutions consists of anion and cation flow. Authors should avoid using terminology for ions (such as 'electrical charge', 'charge', or 'ionic charge carriers') that could be misinterpreted by students as suggesting that ions can pick up, transport, and deposit electrons from one electrode to the other.

Misconception 10f: Anion Migration as Current Flow in Solution. Three of the textbooks (*AB*, *KT*, *OW*) have comments or illustrations suggesting that current flow in the electrolyte solutions consists solely of the flow of anions —each textbook mentions that cation flow also occurs, but implies that flow of cations does not constitute an electrical current. Although *AB* includes an insightful description of both cation and anion flow in electrolyte solutions as current, the illustration on page 657 shows a thick red circle representing the "circuit"—the arrows on the circle that pass through the electrodes and the wire represent the direction of electron flow and the arrows that pass through the electrolyte solutions and the salt bridge represent the direction of anion flow. This drawing may imply to students that anion flow constitutes a current in the aqueous solutions, but that cation flow does not.

Although *KT* discusses anion and cation flow in the electrolyte solutions, the statements in this textbook clearly emphasize the flow of anions as the electrical current in the electrolyte solutions and the salt bridge to the exclusion of cation flow. Indeed, the description of the function of the salt bridge clearly implies that cation flow is not an electrical current: "The function of the salt bridge is to allow anions to pass freely from the compartment where cations are being lost to the compartment where cations are being generated." (*KT*, p. 960). Although some of the illustrations included in *KT* show both cation and anion flow through the salt bridge, two of these pictures (*KT*, pp. 959, 972) depict only anion flow through the salt bridge, implying that cation flow is not an electrical current. *OW* discusses the flow of ions through a porous barrier (the equivalent of the salt bridge) only briefly and states that sulfate ions must flow across the porous barrier to balance the charges in the half-cells. All of the drawings of simple galvanic cells (*OW*, pp. 831, 833, 853, 858) show only anion flow across the porous barrier and although *OW* includes a disclaimer that all ionic species in solution diffuse through the porous barrier, the textbook still implies that only anion flow is an electrical current. This implication is even more obvious in the discussion of the functions of the K^+ and OH^- ions in the electrolyte solution in a mercury battery: "OH^- migrates from the Hg electrode to the Zn electrode, carrying charge; K^+ balances the negative charge on OH^-" (*OW*, p. 858).

Although it may seem easier for students to understand the flow of electrical current in an electrochemical cell as a closed circuit of negatively-charged particles (electrons and anions) flowing in a single direction, authors should not exclude or underemphasize cation flow in the description of current flow in electrolyte solutions and the salt bridge. Authors should also emphasize that both cation and anion flow constitute an electrical current and that both occur in electrochemical cells.

Misconceptions 11a and 11b: Electronic Charges of the Electrodes. Analyzing for sources of student misconceptions suggesting that the electrodes have net positive or negative charges is complicated by the fact that these electrodes may have net charges (*35*, pp. 623-639); however, these charges would be extremely small and exceedingly difficult to measure. One textbook (*OW*) adequately discusses the net charges of the electrodes in a galvanic cell: Electrodes in contact with electrolyte solutions reach an equilibrium which results in dissolved ions and a net negative charge on the electrode from the released electrons ($M(s) \rightarrow M^{n+}(aq) + ne^{-}(metal)$). The extent of this dissociation and the charge imbalance between the metal electrode and the electrolyte solution differs from metal to metal and is responsible for the fact that half-cells composed of different metals have different potentials. This textbook also mentions that the net charge on the electrodes is exceedingly small: only about one electron for every 10^{14} metal atoms (*OW*, p. 839). None of the textbooks provided a similar discussion of the net charges of the electrodes in an electrolytic cell.

The difficulty students have in dealing with net charges associated with the electrodes is that they overestimate the magnitude of this charge and try to apply simple electrostatic arguments to determine electron and ion flow with respect to the charged electrodes. Six of the textbooks (*AB*, *B*, *BP*, *BH*, *BLB*, *KT*) make comments suggesting that the anode of an electrolytic cell is positively charged—either because negatively-charged anions are attracted to it (*AB*, *B*, *BP*, *BH*, *BLB*, *KT*) or because negatively-charged electrons flow away from it, leaving it positively charged (*B*, *BP*, *BLB*, *KT*). While these observations are consistent with a positively-charged anode, students have difficulty explaining why electrons would flow away from the positively-charged anode toward the negatively-charged cathode. Similarly, four of the textbooks (*B*, *BH*, *BLB*, *KT*) make comments suggesting that the anode of a galvanic cell is negatively charged because the anode generates electrons that flow toward the positively-charged cathode. However, students have difficulty explaining why anions flow toward the negatively-charged anode and cations flow toward the positively-charged cathode. Students face an additional problem trying to determine why oxidation occurs at the anode in both galvanic and electrolytic cells, even though it has opposite charges in the two cells. The following quote in *B* suggests that the anode is both positively and negatively charged at the same time and is an additional source of confusion for students concerning the charges of the electrodes.

> "Because oxidation occurs at the anode, electrons must flow away from this electrode into the external circuit. As a result, the anode has a negative charge from the viewpoint of the external circuit, but removal of electrons gives it a positive charge from the viewpoint of the solution in the cell." (*B*, p. 771).

Considering that any net charges on the electrodes would be extremely small and that detailed discussions of the charges of electrodes does not correctly explain the direction of electron and ion flow, it is not surprising that two of the textbooks (*C, Z*) do not discuss the signs of electrodes at all and that one textbook (*RGM*) emphasizes that the signs are electrode polarities (i.e., the positive electrode has a higher potential than the negative electrode) and not electronic charges. Similarly, six of the textbooks (*BP, BH, C, OW, RGM, Z*) emphasize that the charge imbalance caused by the transfer of electrons from the anode to the cathode results in a net positive charge in the aqueous solution near the anode and a net positive charge in the aqueous solution near the cathode that is balanced by ion flow through the salt bridge. In contrast, six of the textbooks (*AB, B, BP, BH, BLB, KT*) made comments suggesting that this charge imbalance leaves the electrodes with net charges; two textbooks (*BP, BH*) made comments suggesting that both the electrodes and the electrolyte solutions have net charges that are neutralized by ion flow through the salt bridge.

To prevent students from using simple electrostatic arguments to determine the direction of ion and electron flow in electrochemical cells, textbook authors should refrain from discussing the net electronic charges of these electrodes. Should authors decide to discuss the signs of electrodes in electrochemical cells, they should stress that these signs represent the polarities of the electrodes and should carefully choose terminology that does not suggest that these electrodes are charged (i.e., the anode in galvanic cells is the 'negative electrode' not the 'negatively-charged electrode'). Because electrodes in electrochemical cells may have net charges, some authors feel compelled to discuss this concept. In this case, authors should emphasize that the net charge on each electrode is exceedingly small and should refrain from using simple electrostatic arguments to explain the direction of electron and ion flow because these arguments cannot properly predict the direction of movement for the electrons and all the ions.

Misconception 14a: Half-Cell Potentials Are Additive. Students demonstrating *Misconception 14a* believe that cell potentials are calculated by adding the standard reduction potentials for the two half-reactions occurring in the electrochemical cell. This misconception results from the fact that eight of the textbooks (*AB*, *B*, *BP*, *BLB*, *C*, *KT*, *RGM*, *Z*) calculate cell potentials using the equation $E°_{cell} = E°_{ox} + E°_{red}$; each of these textbooks include comments directly stating that half-cell potentials are additive. Although this equation works only for balanced oxidation-reduction reactions (i.e., adding half-reactions that have the same number of electrons transferred), only two textbooks (*B*, *RGM*) point out this limitation. The other textbooks may mislead students into believing that any two half-reactions can be added together—for example, students believe that the $E°_{red}(Fe^{3+}/Fe^{2+})$ value can be calculated by adding $E°_{red}(Fe^{3+}/Fe)$ to $-E°_{red}(Fe^{2+}/Fe) = E°_{ox}(Fe/Fe^{2+})$,[1] which is demonstrated below. Converting the E° values into ΔG° values ($\Delta G° = -nFE°$) and adding the ΔG° values using Hess's Law yields $\Delta G°_{red} = -74.5$ kJ and $E°_{red}(Fe^{3+}/Fe^{2+}) = 0.772$ V, which is very close to the tabulated reduction potential of 0.771 V (*33*).

	Adding E° values	Adding ΔG° values
$Fe^{3+}(aq) + 3\,e^{-} \longrightarrow Fe(s)$	$E°_{red} = -0.036$ V	$\Delta G°_{red} = 10.4$ kJ
$Fe(s) \longrightarrow Fe^{2+}(aq) + 2\,e^{-}$	$E°_{ox} = 0.440$ V	$\Delta G°_{ox} = -84.9$ kJ
$Fe^{3+}(aq) + e^{-} \longrightarrow Fe^{2+}(aq)$	$E°_{red} = 0.404$ V	$\Delta G°_{red} = -74.5$ kJ

The terminology used in some of these textbooks also leads students to believe that cell potentials are additive. Four of the textbooks use terms that suggest an additive nature to describe the potential difference of the cell; these terms include 'overall potential' (*B*, *BP*), 'net potential' (*KT*), and 'total potential' (*Z*).

Other textbooks (*BLB*, *C*, *KT*) inconsistently describe the cell potential as a 'potential difference' but then calculate the cell potential by adding oxidation and reduction potentials.

When calculating cell potentials, textbook authors should avoid using the additive method ($E°_{cell} = E°_{ox} + E°_{red}$) and should not introduce the concept of standard oxidation potentials. Instead, authors should calculate the cell potential as a potential difference using standard reduction potentials. The cell potential should be referred to as a 'potential difference', as the 'standard cell potential', or as 'the cell potential'; the cell potential should not be called the 'net potential', the 'overall potential', or the 'total potential' since these terms imply that cell potentials are additive in nature.

Misconception 14b: Half-Cell Potentials Are Extensive Properties. Students demonstrating *Misconception 14b* believe that cell potentials and individual half-cell reduction potentials are extensive properties. *Misconceptions 14a* and *14b* are clearly related: Hess's Law describes the way extensive properties (such as ΔH or ΔG) of two independent reactions can be added together. Although all of the textbooks contain explicit statements that half-cell and cell potentials are intensive properties, several of these textbooks contain statements suggesting that half-cell and cell potentials have properties that are similar to those of extensive properties such as ΔH and ΔG—i.e., when individual reactions are added together, the values of these properties can be added together (Hess's Law). Two textbooks (*AB*, *BLB*) make statements suggesting that Hess's Law applies to half-cell and cell potentials:

> "In the discussion of Hess's law in Chapter 6, the sign of ΔH was changed when the reaction was reversed; therefore, the cell reaction is

> treated like the thermochemical equations in Chapter 6, and the cell potential is the analogue of ΔH." (*AB*, p. 668, Note to Instructor).
>
> "As with energy-related quantities, such as ΔH and ΔG, reversing a reaction changes the sign of E°." (*BLB*, p. 733).

These statements clearly overemphasize the similarities between the intensive half-cell and cell potentials and the extensive ΔH and ΔG values; however, other textbooks make comments that, while not explicitly stating that half-cell and cell potentials follow Hess's Law, certainly imply this relationship. All eight textbooks using the additive formula for calculating cell potentials (*AB*, *B*, *BP*, *BLB*, *C*, *KT*, *RGM*, *Z*) include a statement suggesting that reversing a chemical equation changes the sign but not the magnitude of the cell potential. One textbook (*B*), however, includes a discussion and mathematical examples demonstrating that, while ΔG values can always be added using Hess's Law, E° values are only additive when considering a balanced oxidation-reduction reaction (i.e., when the number of electrons transferred in each half-cell is the same).

As mentioned previously in the discussions of *Misconceptions 9c* and *14a*, authors should use the potential difference method for calculating cell potentials. Use of a line diagram to calculate cell potentials as discussed by Runo and Peters (*36*) is strongly recommended. Calculating cell potentials using the additive method leads to three distinct student misconceptions: (*a*) Half-cell potentials are absolute and measurable; (*b*) Individual half-cell reduction potential are additive; and (*c*) Half-cell and cell potentials are extensive properties that obey Hess's Law. Use of the potential difference method also precludes the need to explain why the sign of the half-cell potential is changed when the equation is reversed (consistent with Hess's Law) but the potential is not multiplied by a numerical coefficient when the equation is multiplied by this coefficient (which contradicts Hess's Law).

Misconception 13e: Reactions Involving Inert Electrodes. Although all of the textbooks have at least one comment stating that inert electrodes such as platinum or graphite are used in electrochemical cells because they are unreactive toward oxidation and reduction, several textbooks ignore or underemphasize this concept when discussing the products of non-spontaneous electrolysis reactions. Two textbooks (*AB*, *Z*) never specify the composition of the electrodes used in electrolytic cells and fail to mention that although these electrodes are typically made of inert substances, it is possible for electrodes to react. Five of the textbooks (*B*, *BP*, *BH*, *C*, *KT*) include a single statement in the electrolysis section suggesting that electrodes used in electrolytic cells are typically inert; however, only two (*B*, *C*) describe the composition of the electrodes (platinum or graphite) and none of them discuss why these particular electrodes are inert or why active electrodes can react in electrolysis reactions.

Only three textbooks (*BLB*, *OW*, *RGM*) discuss the possibility that metal electrodes can react in non-spontaneous electrolysis reactions. Both *BLB* and *RGM* describe the electrolysis of molten salts and aqueous solutions without specifying the composition of the electrodes; however, when discussing electrolysis reactions involving active electrodes, both state that the electrodes in the previous sections were assumed to be inert. These textbooks go on to discuss electrolysis reactions involving active electrodes (such as copper or nickel) that can be oxidized; however, only *BLB* explains how one would predict whether and electrode was active or inert: "When aqueous solutions are electrolyzed using metal electrodes, an electrode will be oxidized if its oxidation potential is more positive than that for water." (*BLB*, p. 750). Only *OW* specifically addresses whether an inert electrode would react when predicting the products of an electrolytic cell: "Platinum is used for passive electrodes because the metal is resistant to oxidation and does not participate in the redox chemistry of the cell." (*OW*, p. 865).

Although this misconception is concerned with the notion that inert electrodes can be oxidized or reduced in electrolytic reactions, the problem that should be addressed is that students do not understand the difference between active and inert electrodes. There are several factors that make inert electrodes unreactive. Some electrodes participate in reactions have large positive reduction potentials (e.g., $Au^{3+} + 3e^- \rightarrow Au$, $E^\circ_{red} = +1.50$ V) that make the electrodes less reactive than water (*37*). Other electrodes participate in reactions that have reduction potentials suggesting that these electrodes should be more reactive than water (e.g., $C + 4H^+ + 4e^- \rightarrow CH_4$, $E^\circ_{red} = +0.13$ V and $CO_2 + 4H^+ + 4e^- \rightarrow C + 2H_2O$, $E^\circ_{red} = +0.20$ V; $PtO_2 + 4H^+ + 4e^- \rightarrow Pt + 2H_2O$, $E^\circ_{red} = +1.0$ V) (*37*). However, these electrodes (or the thin layer of PtO_2 initially formed) are kinetically inert and do not react. Half-reactions that are kinetically unreactive are usually described in terms of an overpotential that must be applied to make the reaction occur at an appreciable rate. In the case of these inert electrodes, the overpotential needed to make reactions involving the electrodes feasible is larger than the potential needed to oxidize or reduce water.

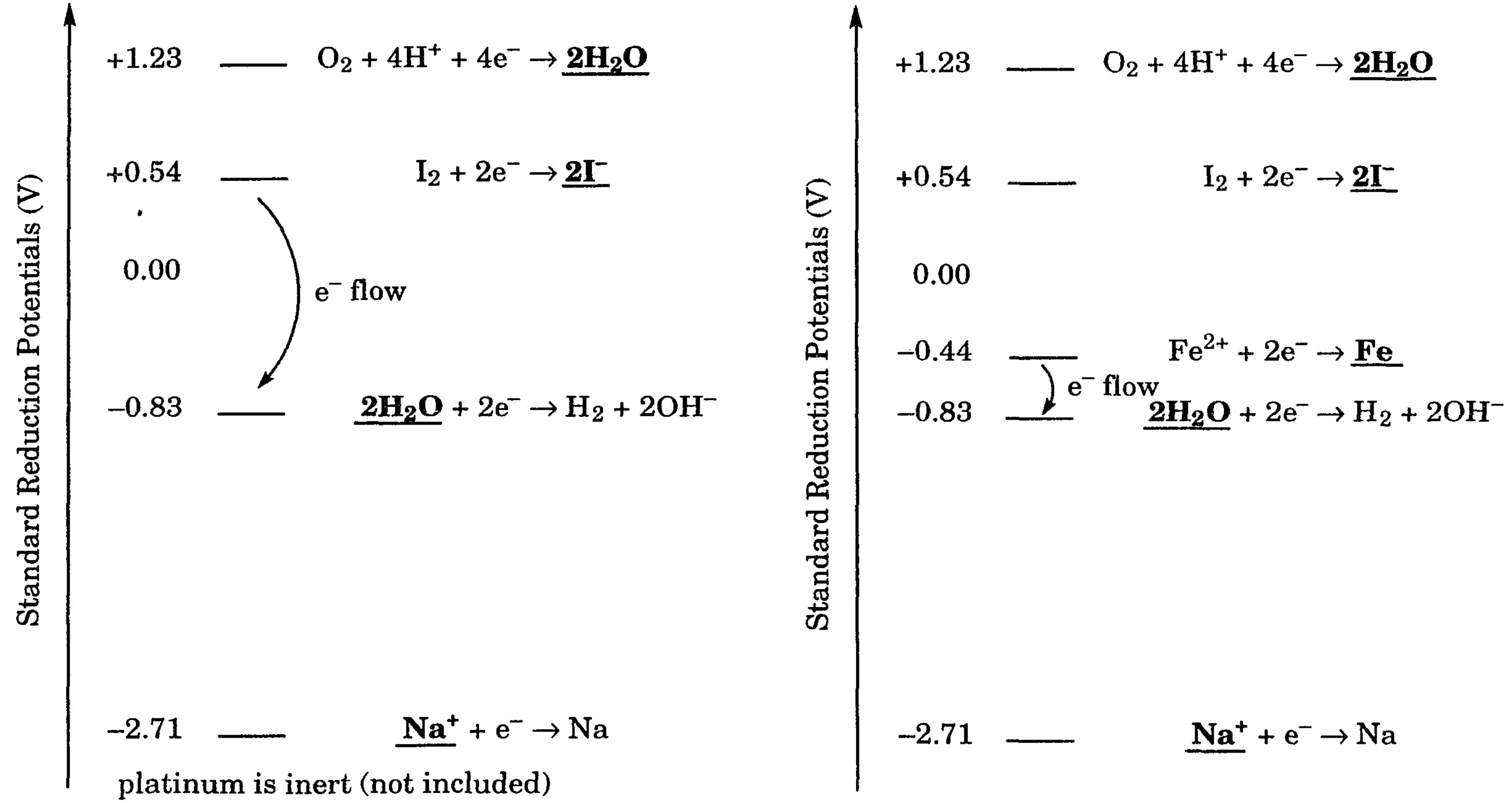
When discussing any electrochemical reaction, textbook authors should always specify the composition of the electrodes. Authors should also include standard reduction potentials for the oxidation and/or reduction reactions involving inert electrodes, such as platinum and graphite, in their textbooks so students can predict whether these reactions will occur in electrolytic cells. When predicting the products of an electrolysis reaction, authors should include reactions involving the electrodes in the list of possible oxidation and reduction half-reactions occurring in the electrolytic cell. If the reduction potentials of inert electrodes imply that these electrodes will react, authors should point out to students that there are kinetic effects preventing the electrodes from reacting and interested authors can introduce the concept of overpotentials.

Misconception 13f: Predicting Electrolysis Reactions. Students demonstrating *Misconception 13f* have difficulty predicting the products of an electrolytic cell. In particular, students do not have a systematic method for determining all of the possible oxidation and reduction half-reactions and once they have these half-reactions, students cannot decide which oxidation and reduction half-reactions actually occur. The discussion of *Misconception 13e* above revealed that only three textbooks (*BLB*, *OW*, *RGM*) discuss the possibility that the electrodes could react in an electrolytic cell. Similarly, six textbooks (*AB*, *B*, *BH*, *C*, *OW*, *Z*) comment that the electrolysis of water requires the addition of an electrolyte, but ignore the possibility that these ions could be oxidized or reduced in the electrolysis of water. Four of these textbooks (*B*, *BH*, *C*, *OW*) discuss the electrolysis of sulfuric acid to produce hydrogen and oxygen gas without considering the possibility that sulfate ions could be oxidized or reduced. Another textbook (*Z*) comments that the "addition of even a small amount of a soluble salt causes an immediate evolution of bubbles of hydrogen and oxygen" (*Z*, p. 841). This comment also ignores the possibility of ion oxidation or reduction and contradicts a later comment made by this textbook stating that the electrolysis of aqueous NaCl produces hydrogen and chlorine gases.

Two textbooks (*C*, *KT*) further complicate the prediction of the electrolysis products by considering inapplicable half-reactions—both textbooks consider the reduction reaction of hydrogen ions ($2H^+ + 2e^- \rightarrow H_2$, $E^\circ_{red} = 0.00$ V) in solutions that are neutral (pH = 7) instead of using the reduction reaction of water ($2H_2O + 2e^- \rightarrow H_2 + 2OH^-$, $E^\circ_{red} = -0.83$ V). All of the textbooks discuss the concept of *overpotential* as the reason electrolysis reactions of chloride solutions typically produce chlorine gas instead of the thermodynamically-favored oxygen gas and most textbooks suggest that the origin of overpotentials is based on kinetic effects. Most of the textbooks suggest that the preferential formation of chlorine gas over oxygen gas is the only instance where using standard reduction potentials fails to correctly predict the reaction products. However, one of the textbooks (*BH*) uses this example as proof that it is very difficult to predict electrolysis products and instead narrowly focuses on predicting electrolysis products empirically: "Although it can be hard to anticipate beforehand what will happen in the electrolysis of aqueous solutions, we still can use what we learn experimentally about one electrolysis to predict what will happen in others." (*BH*, p. 767).

When predicting the products of an electrolysis reaction, textbook authors should always consider oxidation and reduction half-reactions for all possible reactions. This includes reactions involving the electrodes (both active and inert electrodes), water, and all aqueous ions present. Only after all of these reactions have been identified should a prediction of the oxidation and reduction half-reactions be made. Authors should use corrected standard reduction potentials for half-reactions involving hydrogen or hydroxide ions that occur in neutral solutions. Although it is important to discuss the concept of overpotentials to explain why aqueous chloride ions are oxidized and why some inert electrodes are unreactive, authors should emphasize that overpotentials are exceptions and that the method of using standard reduction potentials to predict electrolysis products is generally effective. It is not appropriate to suggest to students that the only way to determine the products of electrolysis is by empirical observations when simple calculations are generally effective.

The potential ladder advocated by Runo and Peters (*36*) is effective at providing students with a way of predicting the electrolysis products of aqueous solutions. All reactions involving the starting materials are plotted vertically on the chart by increasing reduction potential with the starting material underlined (Figure 1). When the chart is completed, the oxidation and reduction half-reactions are determined by finding the smallest distance between reduction reactions that have starting materials on opposite sides of the equation. As an example, Figure 1 contains potential ladder diagrams for the electrolysis of aqueous sodium iodide using platinum and iron electrodes. Because the platinum electrodes are inert (large overpotential), the half-reaction for platinum is omitted from the chart. In the electrolysis using platinum electrodes, the opposing half-reactions with the smallest potential difference are those involving the oxidation of iodide and the reduction of water. Therefore, the net reaction of this electrolysis is: $2H_2O + 2I^- \rightarrow H_2 + I_2 + 2OH^-$ ($E^\circ = -1.37$ V). For the electrolysis of sodium iodide using iron electrodes, the half-reaction for iron is added to the chart. In this case, iron is oxidized and water is reduced ($Fe + 2H_2O \rightarrow Fe^{2+} + 2OH^- + H_2$, $E^\circ = -0.39$ V).



net reaction: $2H_2O + 2I^- \rightarrow H_2 + I_2 + 2OH^-$

$E° = -0.83\ V - (+0.54\ V) = -1.37\ V$

(a)

net reaction: $Fe + 2H_2O \rightarrow Fe^{2+} + 2OH^- + H_2$

$E° = -0.83\ V - (-0.44\ V) = -0.39\ V$

(b)

Figure 1. Potential Ladder Diagrams for the Electrolysis of Aqueous Sodium Iodide: (a) Using Inert Platinum Electrodes; (b) Using Iron Electrodes

Conclusions

The result of this study is the following list of general suggestions concerning electrochemistry instruction intended for the instructors of introductory college level chemistry courses and the authors of the textbooks used in these courses.

1. Authors should avoid the use of simplifications such as always drawing the anode as the left-hand half-cell or only describing the flow of anions in electrolyte solutions and the salt bridge. While these simplifications may seem harmless, they may lead to student misconceptions (i.e., students believing that the anode is always the left-hand electrode or that cation flow does not constitute a current).

2. Authors should avoid using vague or misleading statements. These include statements that are clearly wrong and statements that could be misinterpreted by students. Authors should also avoid the use of vague terminology (such as 'external pathway' or 'ionic charge carriers') in favor of simple, direct descriptions of the processes occurring in electrochemical cells.

3. Authors should calculate cell potentials using the difference method ($E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$) instead of the additive method ($E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red}$). Use of the additive method implies to students that half-cells are absolute and measurable, that half-cells are additive, and that half-cells are extensive properties that obey Hess's Law.

4. Authors should avoid using simple electrostatic arguments to predict ion and electron flow in electrochemical cells. These arguments can confuse students because they do not correctly predict the direction of ion and electron flow.

5. Authors should always consider all possible oxidation and reduction half-reactions when predicting the products of electrolysis. These reactions include those of the electrodes (active or inert), water, and all aqueous ions present.

These implications for textbook authors certainly are not novel: Most of these suggestions have been reported previously. Several researchers have cautioned authors about the use of simplifications (*21, 38*), vague and misleading statements (*19, 21*), the additive method for calculating cell-potentials (*36, 39-*

40), and simple electrostatic arguments to predict the direction of electron and ion flow in electrochemical cells (*20- 21, 38-39*).

The misconceptions discussed in this analysis are supported by comments from these textbooks in differing extents. *Misconception 8c* (the anode is always the left-hand half-cell) was supported by comments from nine of the ten textbooks (all except *B*), while *Misconceptions 9c*, *14a*, and *14b* (which concern the notion that half-cell potentials are absolute, measurable, and extensive) were supported by eight of the ten textbooks (all but *BH* and *OW*). On the other hand, Misconception 10f (only anion flow constitutes a current) was supported by three textbooks (*AB*, *KT*, *OW*), while *Misconceptions 10a-c* and *10e* (which concern the flow of electrons in electrolyte solutions) and *11a* and *11b* (electrodes have net electronic charges) were supported by six textbooks (*AB*, *B*, *BP*, *BH*, *BLB*, *KT*). It is not surprising that misconceptions that have not been widely reported (e.g., *Misconceptions 8c*, *9c*, *14a*, and *14b*) are supported to a greater extent in these textbooks than those that have received a great deal of attention recently (*Misconceptions 10a-c*, *10e*, *11a*, and *11b*).

The ten textbooks reviewed in this study also vary in the extent to which each contains statements and illustrations that could be misinterpreted by students. Two textbooks (*AB*, *KT*) contain vague, misleading, or incorrect statements that could be construed as corroborating all ten misconceptions discussed here.[2] On the other hand, three textbooks (*BH*, *BLB*, and *OW*) contain statements corroborating only five of the ten misconceptions discussed above. The

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[2] Although thirteen misconceptions are discussed in this paper, *Misconceptions 10a* and *10e*, *Misconceptions 10b* and *10c*, and *Misconceptions 11a* and *11b* are clearly related to each other and were treated as one misconception in the following discussion.

extent to which these textbooks contain vague, misleading, or incorrect statements in the oxidation-reduction and electrochemistry chapters could be used as a basis for textbook selection.

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ADDRESSING STUDENT MISCONCEPTIONS CONCERNING ELECTRON FLOW IN ELECTROLYTE SOLUTIONS USING COMPUTER ANIMATIONS AND THE CONCEPTUAL CHANGE APPROACH

A paper submitted for publication to the *Journal of Research in Science Teaching*

Michael J. Sanger and Thomas J. Greenbowe

Abstract

This article investigates the effects of dynamic computer animations depicting the chemical processes occurring in a galvanic cell on the molecular level and conceptual change instruction based on chemical demonstrations concerning current flow in electrolyte solutions on students' conceptions of current flow in electrolyte solutions. These effects were measured for visual and verbal conceptual questions using two groups of students (engineering majors and non-physical sciences majors). This study demonstrated that conceptual change instruction based on chemical demonstrations was effective at preventing or dispelling the student misconception that electrons flow in aqueous solutions to complete the circuit in electrochemical cells for both visual and verbal conceptual questions. However, when student misconceptions are not constantly confronted, students may experience regression of the correct conception. Computer animations of the chemical processes occurring in a galvanic cell on the molecular level did not have an effect on students' responses to visual conceptual questions. The lack of an animation effect may be attributed to the fact that college students do not need visual elaborations or only need to be prompted to form mental

images of these chemical processes. Animation/conceptual change interactions suggest that animations may be helpful when the questions require students to visualize chemical processes on the molecular level (visual conceptual questions) but may prove distracting when the questions do not require students to visualize (verbal conceptual questions). Empirical evidence from this study also suggests that students were more likely to misinterpret visual questions, more likely to practice visual questions, and that lecture attendance and recitation participation was more likely to help students answer visual questions.

Introduction

The investigation of the effects of using computer animations and an instructional method based on the conceptual change approach (*1*) on student misconceptions in electrochemistry was prompted by a previous study (*2*). In this study, students in an introductory college chemistry course received instruction on electrochemistry in which the instructor emphasized that electrons do not flow in electrolyte solutions and showed computer animations depicting the correct flow of current in galvanic and electrolytic cells on the molecular level (i.e., electron flow in the wires and cation and anion flow in solution). These students answered three verbal conceptual questions on the final examination about the flow of electrons in electrolyte solutions and were classified into three groups: Students who consistently demonstrated the misconception that electrons can flow in electrolyte solutions, students who consistently did not demonstrate this misconception, and students who were inconsistent in their responses.

These students were compared to introductory college chemistry students in South Africa who responded to five similar conceptual questions concerning electron flow in electrolyte solutions (*3*). The American students who received instruction using computer animations and the confrontational teaching method were significantly less likely to consistently demonstrate the misconception than the South African students (who, it was assumed, did not view computer animations or a confrontational teaching method). However, because the American students received instruction using both computer animations and a confrontational teaching method, the authors were unable to determine which factor led to the change in the students' conceptions concerning current flow in electrolyte solutions. This study was designed to determine whether the use of computer animations or the conceptual change approach decreases the number of students demonstrating the misconception that electrons can exist and flow in electrolyte solutions in a galvanic cell. This study also investigates whether there is an interaction between these two instructional techniques.

In his review of animation research in computer-based instruction, Rieber (*4*) postulated that the guidelines generated from research on static visuals should extend to animated visuals; however, they may not necessarily account for any differential effects caused by animating the static visuals. He also proposed a taxonomy for the uses of animated visuals in instruction. Cosmetic uses of animations have no instructional intent and merely make the program more attractive. The purpose of Attention-Gaining animations is to arouse, direct, and maintain learner attention. Motivation/Reinforcement animations act as reinforcement or feedback to learner responses. Presentation animations serve as an alternative or supplement to text in defining concepts, rule, or procedures and providing examples, non-examples, or elaborations and represent the bulk of research on the effectiveness of animations. The purpose of Conceptualization animations is to clarify relationships among definitions, examples, and nonexamples of concepts, rules, and procedures through visual means. Conceptualization animations provide a concrete representation of processes which involve a potentially complex array of relationships among individual components and their effects are very dependent on the learner's prior knowledge of the learning task—novices, who may have difficulty seeing abstract relationships on their own, would be expected to benefit from Conceptualization animations. Interactive Dynamics animations permit the design of interactive programs where students learn by discovery and informal hypothesis-testing. The graphics change continuously over time depending on student input, acting as a form of instantaneous graphic feedback that is not easily replicated using media other than the computer.

From his review of computer-based instruction using animations, Rieber (*4*) concluded that using computer animations has not proven to be an effective or reliable presentation strategy when studied separately in controlled experiments. Any effects attributable to animations may be dependent on other lesson components (e.g., lesson organization, practice, etc.). Incorporation of animations in instruction is supported only if the demands of the learning task involve the attributes of visualization, motion, or trajectory. Empirical evidence suggested that novice learners may not know how to attend to relevant cues or details provided by animations and should be prompted to watch for relevant details. Learners with specific aptitudes, such as low spatial ability or low maturation, may also benefit from the use of animations (*5*). In general, Rieber claimed, the few serious attempts to study the instructional effects of animations have not adequately shown them to be effective.

However, several chemical education researchers have demonstrated that computer animations can help students develop visualization skills which facilitate thinking about chemical processes on the molecular level (*6-10*). Williamson and Abraham (*9*) compared the performance of chemistry students on two subsets of the Particulate Nature of Matter Evaluation Test (PNMET) after receiving instruction that included the use of static and animated visuals of chemical processes on the molecular level. Students who viewed animations in lecture and those who viewed animations in lecture and discussion sections scored significantly higher on both subsets than those students who viewed static visuals in lecture. Mayton (*11*) compared the performance of psychology students on cued- and free-recall of cardiac system functions. Among students who received an imagery cueing instruction, students who viewed animations performed significantly better on both cued- and free-recall of cardiac functions on an immediate post-test compared to students who viewed only static pictures; these differences were still present on the post-test, but were no longer significant. Two articles (*6, 8*) reported that the combined use of computer animations depicting chemical processes on the molecular level and conceptual change instruction resulted in a significant decrease in the number of students demonstrating misconceptions. A similar study (*12*) reported that the use of computer animations and conceptual change instruction led to a significant decrease in the number of physics students demonstrating a misconception concerning the definitions of position and velocity.

Conceptual change has been described as a process of learning science in a meaningful way that requires the learner to realign, reorganize, or replace existing misconceptions in order to accommodate new ideas (*13*). Cho, Kahle, and

Nordland (*14*) defined the term *misconception* as student conceptual and propositional knowledge that is inconsistent with or different from the commonly-accepted scientific consensus. Research focusing on students' understandings of a variety of chemical phenomena (*15-18*) has demonstrated that student explanations are often inconsistent with, inferior to, and incapable of explaining observable phenomena when compared to the scientifically-accepted descriptions. However, it is important to note that some misconceptions (especially those concerning abstract concepts in science) are capable of adequately explaining students' experiences and observations, appear quite logical to students, are consistent with their understanding of the world, and are therefore very resistant to change and often persist following traditional instruction (*1, 19-20*).

The model of conceptual change developed by Posner, Strike, Hewson, and Gertzog (*1*) proposed that four conditions must occur before students can replace an existing misconception: Students must experience dissatisfaction with their existing conception, they must be able to understand the new conception, the new conception must seem plausible to them, and the new conception must appear to be better at explaining their experiences and observations than their previous conception. Once these conditions have been met, students are more likely to experience conceptual change, discarding their naive conception for a scientifically-accepted one.

Smith, Blakeslee, and Anderson (*13*) compiled a list of teaching activities based on the four conditions proposed by Posner et al. (*1*) that are commonly used in effective conceptual change instruction. In the dissatisfaction stage, the instructor is actively engaged in eliciting student conceptions, asking for explanations, pointing out discrepancies or inadequacies, and encouraging debate and deliberation. The understanding and plausibility stages require teaching strategies that engage students in developing their understanding of the new concept, judging whether the new conception is consistent with other theories, beliefs, and experiences, and recognizing that the new conception explains some experiences and observations or solves some current problems. The fruitfulness stage requires that students see how the new conception helps make sense of novel experiences and observations, how it explains increasingly complex and unfamiliar observations, and how it leads to new insights.

Several chemical education researchers (*21-24*) have demonstrated the effectiveness of conceptual change instruction on changing students' conceptions of chemical processes. However, several problems concerning the effectiveness of the conceptual change approach have been reported in the literature. These problems include the fact that a high proportion of students can retain their misconceptions (*22*), conceptual change can be temporary and students may revert to their previous misconceptions over time (*25*), misconceptions not specifically addressed may not be affected by the instruction (*23, 26*), and teachers may have difficulty teaching lectures using the conceptual change approach (*22*).

Another disadvantage of the conceptual change approach is that instructors need to be familiar with the misconceptions that students are likely to hold. Several researchers have reported student misconceptions in electrochemistry that were derived from student responses to conceptually-based multiple-choice questions (*3, 27*) and student interviews (*17-18, 28*). The misconception that the migration of electrons is responsible for the flow of current in aqueous solutions is perhaps the most prevalent misconception in electrochemistry, and was the only misconception that was reported by each of the researchers listed above.

This research explores the effects of viewing computer animations and receiving conceptual change instruction on students' conceptions of current flow in electrolyte solutions. Specifically, this article reports the use of computer-generated, dynamic, three-dimensional graphic representations of the chemical processes occurring in galvanic cells on the molecular level and the use of conceptual change instruction based on chemical demonstrations concerning current flow in solution. This article also reports the effects and interaction of these techniques on student performance on an algorithmic question and on the proportion of student responses to verbal and visual conceptual multiple-choice questions that are consistent with the scientifically-accepted description of current flow in electrolyte solutions (i.e., anion and cation flow in solution). The hypotheses for each question type are listed below:

1. For the algorithmic question, there will be no effect on student performance due to the use of animations or conceptual change. There will be no interaction between these techniques.
2. For the visual conceptual questions, there will be an increase in the proportion of student responses consistent with the scientifically-accepted conception after receiving instruction using animations or conceptual change. There will be no interaction between techniques.
3. For the verbal conceptual questions, there will be an increase in the proportion of student responses consistent with the scientifically-accepted conception after receiving conceptual change instruction. The use of animations should not affect the proportion of student responses consistent with the scientifically-accepted conception. There will be no interaction between the techniques.

Method

Subjects. This study included students enrolled in two different introductory college chemistry courses at an American midwestern university. The first course was intended for students majoring in the non-physical sciences (including biology, wildlife fisheries and biology, family and consumer sciences, etc.) and contained 122 students; 115 of these students participated in this study. The second course was intended for engineering majors and contained 194 students; 138 of these students participated in this study.

Design. This study utilized a two-factor (2 × 2) nonequivalent control-group design (*29*). The experimental treatments were administered by the first author during a single 50-minute recitation period for each course; the students were randomly assigned to each of the four groups by recitation sections. Even though the treatments were administered to the recitation sections as a group, the statistical analyses were performed using individual student scores.

Several authors have discussed the question of whether statistical analyses in education research should be performed with individual scores or class means as the unit of analysis (*30-35*). Raths (*30*) and Hopkins (*35*) suggested that individual teacher effects and the effect of one student on the other students in the same classroom pose major threats to the validity of statistical analyses performed using the individual as the unit of analysis. However, these effects were minimized in this study. Teacher effects were minimized by having the same instructor administer each treatment group, using the same set of notes for the common instructional topics. Because the treatments were administered in a lecture format, the students were relatively passive learners and did not ask questions or behave in any manner that would be expected to affect the other students in the treatment group. Hopkins (*35*) stated that the use of the traditional linear model for an ANOVA with individuals as the experimental unit is likely to be valid only if these two effects are not present.

From a comparison of statistical analyses performed with individual scores and group means as the units of analysis, Herron, Luce, and Neie (*33*) concluded that when one may assume that the composition of the groups is similar to that expected from complete random assignment of individuals to the groups, the results of an analysis based on individual scores is likely to be comparable to the results of an analysis based on group means, regardless of whether the treatment was administered to the class or to individuals. Herron, Luce, and Neie (*31*) suggested that readers may err more by dismissing out of hand the results of a study because an incorrect choice of experimental unit was made than by accepting the results as "probably correct".

Treatments. Two independent variables were investigated in this study: Computer animations and conceptual change instruction. The computer animations used in this study depicted the electrochemical processes occurring in a copper-zinc galvanic cell on the microscopic level. These animations focused on the chemical half-reactions that occur at each metal electrode and the transfer of aqueous ions from the salt bridge to the two half-cell compartments and have been described previously in the literature (*2, 36*). Each animation was repeated at least three times and was narrated by the first author, pointing out the important and relevant processes depicted in the computer animations. Students who did not view the computer animations received similar instruction about the half-reactions occurring at each electrode and the transfer of aqueous ions from the salt bridge using static chalkboard drawings on the microscopic level.

This experiment used a chemical demonstration-based conceptual change approach to confront the misconception that electrons can flow in electrolyte solutions. The students were told that there are three "theories" commonly proposed to explain current flow in electrolyte solutions and the salt bridge. These "theories" correspond to the common student misconceptions concerning current flow in electrolyte solutions that were identified from student interviews (*17-18, 28*) and are listed below.

1. Free electrons, which can be written as $e^{-}(aq)$, flowing through the electrolyte solutions from the cathode to the anode without any assistance constitute the electrical current in solution.

2. Electrons flowing through the solution from the cathode to the anode constitute the electrical current in solution, but the electrons can't flow in solution all by themselves—they need assistance from cations and/or anions and hop onto the ion at the cathode and jump off at the anode (the "piggyback" method).

3. Cation and anion flow constitute the electrical current in solution. The only reason electrons represent a current flow in the wire is because electrons have a net negative charge. In solution, electrons aren't stable, so other negatively- or positively-charged species (anions and cations) flow in the solution instead.

Three chemical demonstrations were performed and the implications of these demonstrations were discussed with the students. In the first demonstration, the electrical conductivity of pure water was tested before and after some sodium chloride was dissolved in it. The fact that pure water is a poor conductor of electricity contradicts Theory 1—if the electrons need no assistance, why don't they flow? However, Theories 2 and 3 are consistent with the observation that water did not conduct electricity until an ionic salt was added. The second demonstration was the reaction of potassium metal in pure water (with a small amount of phenolphthalein). Potassium reacted vigorously and completely with water according to the following two half-reactions:

$$K(s) \longrightarrow K^+(aq) + e^-$$

$$2\,H_2O(l) + 2\,e^- \longrightarrow H_2(g) + 2\,OH^-(aq)$$

The first half-reaction suggests that potassium readily loses an electron. Therefore, the potassium ion should be very unwilling to accept an electron and if the potassium ion is not willing to accept an electron, it cannot shuttle electrons from one electrode to the other. Similarly, fluorine gas reacts explosively with water according to the following half-reaction: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$. This half-reaction suggests that the fluoride ion is unwilling to give up an electron and therefore $F^-(aq)$ cannot shuttle electrons from one electrode to the other either. Both of these half-reactions are in opposition to Theory 2, which proposes that cations and anions can pick up and release electrons at the electrodes. The second half-reaction demonstrates that free electrons are unstable in water, reducing it to hydrogen gas and hydroxide ions; therefore, Theory 1 cannot be correct because free electrons will react with water before they can flow from one electrode to the other. All of the half-reactions discussed in this demonstration imply that anions and cations are the only charged species that are stable in water, which is in agreement with Theory 3.

The final demonstration was more of a theoretical argument concerning the "counting" of electrons. The students were told that the number of electrons flowing in an electrochemical cell can be measured at three places in the galvanic cell: At the zinc anode (by measuring the mass of zinc lost and converting it to moles of electrons using stoichiometry), at the copper cathode (by measuring the mass of copper gained and converting it to moles of electrons), and in the wire (by measuring the electrical current in amperes that travels through the wire in a second, which is converted to moles of electrons using Faraday's constant). They were also told that the three numbers are the same within experimental error. This implies that every electron that flows through the wire originated at the anode and was consumed at the cathode. Therefore, no electrons could enter the electrolyte solution at the cathode because there are none left after the copper ions have been reduced. Similarly, if any electrons were transferred to the anode from solution, the number of electrons flowing through the wire should be more than the number of electrons released at the anode. This electron "bookkeeping" discounts any theory suggesting that electrons flow in solution (Theories 1 and 2) and leaves Theory 3 as the only possible method of current flow in electrolyte solutions.

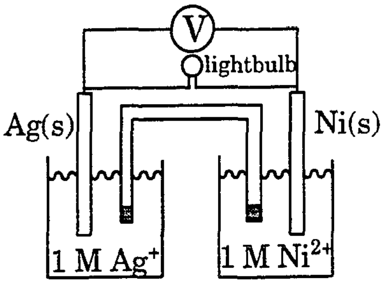
Student who did not receive instruction using the conceptual change approach did not see these demonstrations and did not receive these explanations. These students were instructed that electrons do not flow in solution and that current flow in electrolyte solutions consists of cation and anion flow, but existing misconceptions were not actively confronted. Each of the four groups received equivalent instruction on calculating the cell potential of a galvanic cell using the potential difference method (*37-38*) and on the flow of current through the electrodes and the wire of the galvanic cell.

Dependent Measures. The pre-test consisted of the individual students' average midterm examination scores prior to the experiment and was intended to measure the students' general chemistry knowledge. Post-test questions were administered to the students immediately after the experiment was performed (immediate post-test) and on the midterm examination for the engineering majors and the final examination for the non-physical sciences majors (delayed post-test).

The immediate post-test, which appears in Figure 1, contained three types of questions. The first question (question 1) consisted of an algorithmic question that could be answered using a mathematical formula and did not require a conceptual understanding of galvanic cells to be answered correctly. The next four questions (questions 2-5) were conceptual multiple-choice questions that are visual in nature. These questions concerned the flow of electrons and ions in the aqueous solutions at the electrodes and the two ends of the salt bridge as the reactions in the galvanic cell occurred. In each of these questions, the words 'in solution' were bold-faced to emphasize that this question depicted only the migrations that occurred in solution. The non-physical sciences majors were also verbally prompted by the first author that any ion or electron migration occurring in the electrodes was purposely omitted because these questions were only interested in the migrations that occurred in solution. The last four questions (questions 6-9) consisted of traditional multiple-choice questions concerning the flow of electrons and ions in aqueous solutions. These questions were classified as conceptual questions that are verbal in nature.

For the algorithmic question in the immediate post-test, student responses were scored from 0-2: The correct value received 2 points (ignoring math errors), a value calculated using the difference method but an incorrect assignment of the anode and cathode reactions received 1 point, and no responses or cell potentials calculated by adding reduction potentials received 0 points. For each multiple-choice question, students received 1 point for choosing a distractor that did not suggest that electrons migrated in solution and 0 points for choosing a

Questions 1-5 all pertain to the following electrochemical cell.

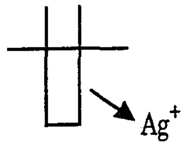


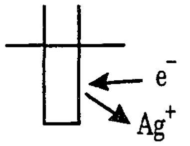
The salt bridge contains $KNO_3(aq)$

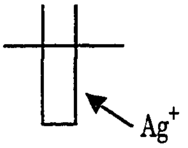
	E°
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$	-0.25 V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	0.80 V

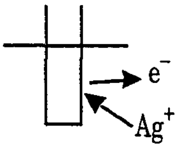
1. What is the cell potential of this cell? (Please show all work)

2. Which drawing best describes the reaction occurring **in solution** at the Ag electrode?

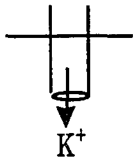
a. 

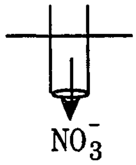
*b. 

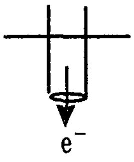
c. 

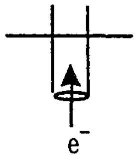
*d. 

3. Which drawing best describes the current flow occurring **in solution** at the salt bridge in the $AgNO_3$ solution?

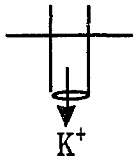
a. 

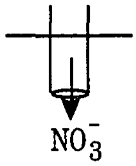
b. 

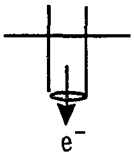
*c. 

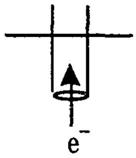
*d. 

4. Which drawing best describes the current flow occurring **in solution** at the salt bridge in the $Ni(NO_3)_2$ solution?

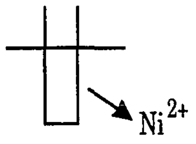
a. 

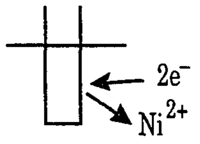
b. 

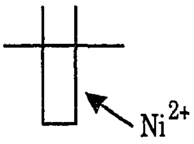
*c. 

*d. 

5. Which drawing best describes the reaction occurring **in solution** at the Ni electrode?

a. 

*b. 

c. 

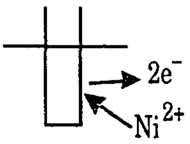
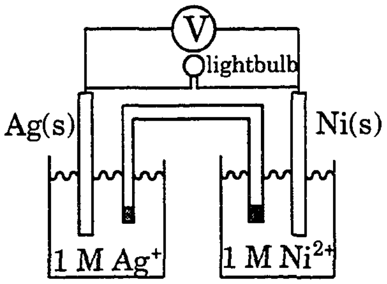
*d. 

Figure 1. Data collection instrument for the immediate post-test.



6. In this electrochemical cell, electrons in the cell flow through the _____, toward the ______.

a. wire, silver electrode
b. wire, nickel electrode
*c. wire, silver electrode AND salt bridge, nickel electrode
*d. wire, nickel electrode AND salt bridge, silver electrode

7. In an electrochemical cell, conduction through the electrolyte is due to:

*a. electrons moving through the solution attached to the ions.
*b. electrons moving from ion to ion through the solution.
c. the movement of both positive and negative ions.
d. the movement of water molecules.
*e. electrons moving through the solution from one electrode to the other.

8. The function of a salt bridge in an electrochemical cell is to:

a. form complex ions with the oxidation products.
*b. permit electrons to flow through the solution.
c. keep the levels of liquids equal in both half-cells.
d. allow positive and negative ions to enter and leave both half-cells.
e. maintain a steady flow of reactants from the cathode to the anode and vice versa.

9. Evaluate the following assertion and reason listed below:

Assertion	**Reason**
If the salt bridge in the picture above was replaced by a tube filled with graphite (an electrical conductor), the light bulb would be lit...	...there will be a continuous flow of electrons in the electrolyte solutions that can pass through the graphite bridge.

*a. Both the assertion and the reason are correct.
b. The assertion is correct, but the reason is incorrect.
*c. The assertion is incorrect, but the reason is correct.
d. Both the assertion and the reason are incorrect.

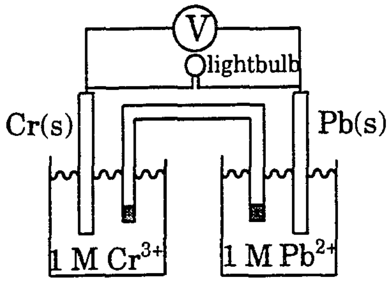
Figure 1. (continued)

distractor that did suggest that electrons migrated in solution. While it may appear that this coding simply measures student performance, the distinction between these two measures is subtle but important, because it is possible to respond incorrectly to a question without demonstrating a misconception. The distractors that correspond to the misconception are marked with an asterisk in Figure 1. Student conceptual scores for the four visual conceptual questions were totaled and these value were compared among treatment groups. Student conceptual scores for the four verbal conceptual questions were also totaled and compared among treatment groups.

The delayed post-test also contained visual and verbal conceptual questions concerning electron and ion migration in aqueous solutions. The visual conceptual questions (with the distractors corresponding to the misconception marked with an asterisk) that were presented to the students appear in Figure 2. Three verbal conceptual questions were presented to the students: Two of these were verbatim reproductions of questions 7 and 8 from the immediate post-test and the other question was very similar in content and structure to question 6 on the immediate post-test. These questions were analyzed in the same manner as the immediate post-test questions and the composite visual conceptual and verbal conceptual scores were compared among treatment groups.

Data Analysis. The equivalence of the four treatment groups in each course was tested by performing a two-way (2 x 2) analysis of variance (ANOVA) on the pre-test scores. The algorithmic score on the immediate post-test and the conceptual visual and conceptual verbal scores on both the immediate and delayed post-tests were also analyzed for treatment and interaction effects by

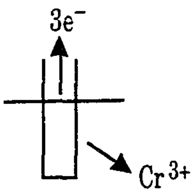
Questions 1-4 pertain to the following electrochemical cell.

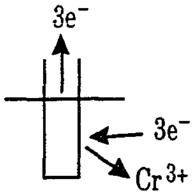


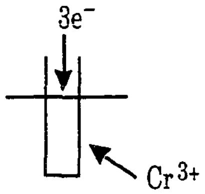
The salt bridge contains KCl(aq)

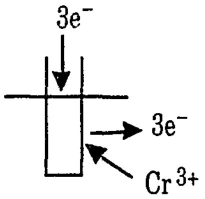
	E°
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$	-0.74 V
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	-0.13 V

1. Which drawing best describes the reaction occurring at the Cr electrode?

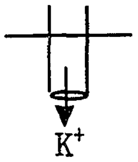
a.


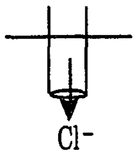
*b.


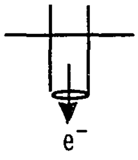
c.


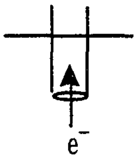
*d.


2. Which drawing best describes the current flow occurring at the salt bridge in the $Cr(NO_3)_3$ solution?

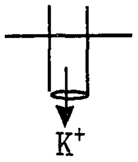
a.


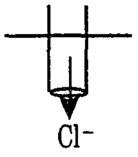
b.


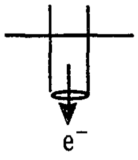
*c.


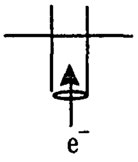
*d.


3. Which drawing best describes the current flow occurring at the salt bridge in the $Pb(NO_3)_2$ solution?

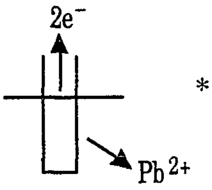
a.


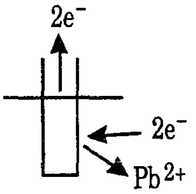
b.


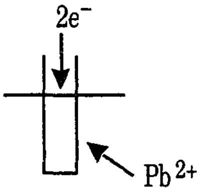
*c.


*d.


4. Which drawing best describes the reaction occurring at the Pb electrode?

a.


*b.


c.


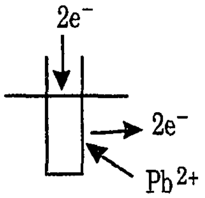
*d.


Figure 2. Visual conceptual questions presented in the delayed post-test.

performing a 2 × 2 ANOVA on each of these scores. The conceptual visual and conceptual verbal scores on the delayed post-test were also compared for the engineering students who participated in the experiment and for those who didn't by performing a one-way analysis of covariance (ANCOVA), corrected for pre-test scores.

Results and Discussion

A 2 × 2 ANOVA was performed on the pre-test scores of the students in each course to determine any differences among the four treatment groups. In both courses, no significant differences (all $p > .05$) were found among the four treatment groups on the basis of their pre-test scores (Tables 1 and 2). Because

Table 1. ANOVA Results for the Pre-Test Scores of the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	322.58	322.58	1.84	.18
Conc. Change	1	322.06	322.06	1.83	.18
Animation × Conc. Change	1	0.69	0.69	0.004	.95
Error	134	23,533.43	175.62		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	73.4 34	70.2 33	71.9 67
	Yes	76.4 40	73.4 31	75.1 71
	Totals	75.0 74	71.8 64	73.5 138

Table 2. ANOVA Results for the Pre-Test Scores of the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	187.05	187.05	1.10	.30
Conc. Change	1	62.55	62.55	0.37	.55
Animation × Conc. Change	1	118.97	118.97	0.70	.40
Error	111	18,870.15	170.00		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	64.6 40	65.2 29	64.8 69
	Yes	61.0 22	65.7 24	63.5 46
	Totals	63.3 62	65.4 53	64.3 115

there were no significant differences among the treatment groups, it was unnecessary to compare the post-test scores using ANCOVAs, so they were compared using ANOVAs.

A comparison of the algorithmic scores on the immediate post-test in each course using a 2 × 2 ANOVA (Tables 3 and 4) also demonstrated that there were no significant differences (or interaction) among the four groups based on this question (all $p > .05$). This is consistent with our first hypothesis and is not surprising, since each treatment group received the same instruction with respect to the calculation of cell potentials for galvanic cells. This result, along with the analysis of the pre-test scores, suggests that the four treatment groups in each course were equivalent prior to the experimental treatment.

Table 3. ANOVA Results for the Algorithmic Scores of the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.00002	0.00002	0.0001	.99
Conc. Change	1	0.352	0.352	2.311	.13
Animation × Conc. Change	1	0.132	0.132	0.867	.35
Error	129	19.631	0.152		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	1.94 32	2.00 32	1.97 64
	Yes	1.90 39	1.83 30	1.87 69
	Totals	1.92 71	1.92 62	1.92 133

Table 4. ANOVA Results for the Algorithmic Scores of the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.00001	0.00001	0.00002	.996
Conc. Change	1	0.287	0.287	0.682	.41
Animation × Conc. Change	1	0.613	0.613	1.456	.23
Error	109	45.910	0.421		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	1.56 39	1.71 28	1.63 67
	Yes	1.82 22	1.67 24	1.74 46
	Totals	1.66 61	1.69 52	1.67 113

Visual Conceptual Questions. For the visual conceptual questions answered by the engineering students on the immediate post-test, there were no treatment effects for either the animations or the conceptual change approach and there was no interaction between the two treatments (all $p > .05$, see Table 5). After completing the immediate post-test instrument, several students approached the first author with questions regarding the visual conceptual questions that depicted the reactions occurring at the metal electrodes (questions 2 and 5, Figure 1). These students stated that although they knew that electrons did not exist in solution, they did not feel comfortable choosing distractor **a** or **c** because these choices did not adequately describe the flow of electrons at the electrode. A separate analysis of the two types of visual conceptual questions (electrode reactions and salt bridge reactions) revealed a significant difference in

Table 5. ANOVA Results for the Visual Conceptual Questions on the Immediate Post-Test for the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	2.86	2.86	2.34	.13
Conc. Change	1	3.31	3.31	2.70	.10
Animation × Conc. Change	1	1.26	1.26	1.03	.31
Error	134	164.14	1.22		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.79 34	2.70 33	2.75 67
	Yes	2.68 40	2.19 31	2.46 64
	Totals	2.73 74	2.45 64	2.60 138

the number of student responses consistent with the misconception: 54% of the student responses contained electrons in solution around the electrodes, while only 14% of their responses contained electrons in solution around the salt bridge ($t_{137} = 8.25, p < .0001$).

As a result, an explicit statement was made to the non-physical sciences majors prior to completing the immediate post-test instrument that the drawings in questions 2 and 5 depicted only the migration reactions that occurred in solution and that any ion or electron migrations that were occurring in the electrodes was purposely omitted. The results of the ANOVA for the visual conceptual scores answered by the non-physical sciences students on the immediate post-test appear in Table 6. Only the interaction between the animations and the

Table 6. ANOVA Results for the Visual Conceptual Questions on the Immediate Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	1.32	1.32	1.33	.25
Conc. Change	1	2.54	2.54	2.54	.11
Animation × Conc. Change	1	4.22	4.22	4.23	.042*
Error	111	110.74	1.00		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.73 40	2.55 29	2.65 69
	Yes	2.64 22	3.25 24	2.96 46
	Totals	2.69 62	2.87 53	2.77 115

* Significant at the .05 level.

conceptual change treatments was significant ($p = .042$). A Tukey/Kramer post-hoc test (*39*) was performed to determine significant differences among the means of the four treatment groups. The results of this test appear in Table 7. This test suggests that none of these means are significantly different from each other.

Individual analysis of the electrode and the salt bridge questions also yielded interesting results for the non-physical sciences majors. Although, 53% of the student responses contained electrons in solution around the electrodes, only 8% of their responses contained electrons in solution around the salt bridge ($t_{114} = 9.58, p < .0001$). The authors decided to perform individual ANOVAs for each type of questions (electrode questions and salt bridge questions). While there were no significant treatment or interaction effects (all $p > .05$) regarding the visual conceptual salt bridge questions (Table 8), there was a significant treatment effect for the conceptual change approach ($p = .017$) as well as a significant interaction effect ($p = .031$). The results of the ANOVA for the visual conceptual electrode questions answered by the non-physical sciences students on the immediate post-test appear in Table 9. Comparison of the mean scores for the students who did and did not receive the conceptual change instruction (1.17

Table 7. Tukey/Kramer Post-Hoc Tests Comparing the Means of the Treatment Groups Reported in Table 6.

Groups Compared	Difference in Means	Q value[a]
Both vs. Conc. Change	0.61	2.94
Both vs. Anim.	0.70	3.58
Both vs. Control	0.53	2.88
Conc. Change vs. Anim.	0.08	0.42
Conc. Change vs. Control	-0.09	0.47
Anim. vs. Control	-0.17	1.01

[a] For four groups and $df_{error} = 111$, the critical Q value ($\alpha = 0.05$) is 3.74.

Table 8. ANOVA Results for the Salt Bridge Questions on the Immediate Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.003	0.003	0.011	.92
Conc. Change	1	0.267	0.267	1.190	.28
Animation × Conc. Change	1	0.022	0.022	0.099	.75
Error	111	24.870	0.224		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	1.90 40	1.86 29	1.88 69
	Yes	1.77 22	1.79 24	1.78 46
	Totals	1.85 62	1.83 53	1.84 115

Table 9. ANOVA Results for the Electrode Questions on the Immediate Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	1.44	1.44	1.89	.17
Conc. Change	1	4.45	4.45	5.84	.017*
Animation × Conc. Change	1	3.64	3.64	4.77	.031*
Error	111	84.53	0.76		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	0.83 40	0.69 29	0.77 69
	Yes	0.86 22	1.46 24	1.17 46
	Totals	0.84 62	1.04 53	0.93 115

* Significant at the .05 level.

versus 0.77) demonstrates that the conceptual change approach was effective at preventing or dispelling this misconception. A Tukey/Kramer post-hoc test was performed to determine significant differences among the means of the four treatment groups. The results of this test appear in Table 10. This test demonstrates that the mean of the group receiving both instructional methods is significantly higher than the means of the group receiving animations alone or the group receiving neither treatment (control). The interaction effect for the electrode questions implies that receiving instruction using both animations and the conceptual change approach was more effective than viewing animations alone or neither method at all.

Due of the difficulty some students experienced with the visual conceptual electrode questions on the immediate post-test, these questions were revised for the delayed post-test to show electron flow in the electrodes as well as any ion or electron migration in the electrolyte solutions (Figure 2). Using these revised questions, there were no treatment effects for the animations or the conceptual change approach and there was no interaction effect for these two methods in either course (all $p > .05$). The results of the ANOVA for the visual conceptual questions answered on the delayed post-test appear in Table 11 for the engineer-

Table 10. Tukey/Kramer Post-Hoc Tests Comparing the Means of the Treatment Groups Reported in Table 9.

Groups Compared	Difference in Means	*Q* value[a]
Both vs. Conc. Change	0.59	3.27
Both vs. Anim.	0.77	4.51*
Both vs. Control	0.63	3.98*
Conc. Change vs. Anim.	0.17	1.00
Conc. Change vs. Control	0.04	0.24
Anim. vs. Control	-0.14	0.90

[a] For four groups and $df_{error} = 111$, the critical Q value ($\alpha = 0.05$) is 3.74.

* Significant at the .05 level.

Table 11. ANOVA Results for the Visual Conceptual Questions on the Delayed Post-Test for the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.058	0.058	0.133	.72
Conc. Change	1	0.160	0.160	0.366	.55
Animation × Conc. Change	1	0.292	0.292	0.668	.42
Error	131	57.328	0.438		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	3.71 34	3.76 33	3.73 67
	Yes	3.87 38	3.73 30	3.81 68
	Totals	3.79 72	3.75 63	3.77 135

ing students and Table 12 for the non-physical sciences students. The difference in the proportion of student responses that contained electrons in the electrolyte solutions at the electrodes and at the salt bridge still remained in the delayed post-test but was smaller than in the immediate post-test. Only 9% of the engineering majors' responses contained electrons in solution at the electrodes and 2% of their responses contained electrons in solution at the salt bridge ($t_{133} = 2.910$, $p = .0042$). For the non-physical sciences majors, 37% of their responses had electrons in solution at the electrodes and 12% of their responses had electrons in solution at the salt bridge ($t_{112} = 5.805, p < .0001$).

In both courses, over half of the student responses on the immediate post-test suggested that electrons could flow in solution at the electrodes, compared to about 10% suggesting that electrons could flow in solution at the salt bridge.

Table 12. ANOVA Results for the Visual Conceptual Questions on the Delayed Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	3.08	3.08	1.94	.17
Conc. Change	1	0.06	0.06	0.04	.84
Animation × Conc. Change	1	0.70	0.70	0.44	.51
Error	109	173.66	1.59		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	3.08 40	2.90 29	3.00 69
	Yes	3.29 21	2.78 23	3.02 44
	Totals	3.15 61	2.85 52	3.01 113

Analysis of both types of question may help explain this discrepancy. The salt bridge questions showed either electrons or ions flowing through the salt bridge, which correspond to Theories 1 and 3, respectively. However, students believing Theory 2 might have chosen distractors that showed ions flowing through the salt bridge (since the "piggybacked" electron would be invisible); therefore, this question may not have discriminated between students believing Theories 2 and 3. A similar argument may be made for the electrode pictures, although the distractors with electrons in solution may have appeared acceptable to students believing Theory 2. The major problem with the electrode questions on the immediate post-test was that they did not describe the flow of electrons inside the electrodes. Several engineering majors stated that they chose distractors with electrons in solution because the electrons had to go somewhere, and when a disclaimer was made to the non-physical sciences majors, the electrode questions appeared to discriminate between students who did and did not hold the misconception.

When these drawings were corrected and used on the delayed post-test, however, no treatment or interaction effects were measured. The authors believe that the significant effects measured on the immediate post-test disappeared on the delayed post-test as a result of student practice. The data collection instrument was returned and discussed with both courses one week after the immediate post-test; the engineering majors were allowed to keep the instrument, but the non-physical sciences majors returned it after their discussion. The large decrease in the proportion of the engineering majors' responses to the visual questions that were consistent with the misconception (34% on the immediate and 6% on the delayed post-test) implies that the students studied and practiced these questions before the midterm examination (delayed post-test). Because only 6% of the student responses were consistent with the misconception, it is not surprising that these questions were unable to discriminate between students who did and did not hold the misconception. The more modest decrease in the proportion of responses from the non-physical sciences majors consistent with the misconception (31% on the immediate and 25% for the delayed post-test) suggests that discussing the instrument one week after the immediate post-test had a slight effect on student responses and may have countered the treatment and interaction effects measured on the immediate post-test. The modest effect of discussing the visual conceptual questions in recitation and the larger effect of allowing students to keep copies of these questions suggest that students might have learned how to answer these questions correctly, regardless of whether they believed that electrons could flow in solution. This result is consistent with Pickering's conclusion (*40*) that students have difficulty answering visual conceptual questions based on the particulate nature of matter, not because they were unable to understand this concept but because they had not been previously asked to understand it.

Research on the use of computer animations in chemistry instruction (*6-10*) suggests that computer animations of chemical processes can facilitate student thinking on the molecular level. Rieber (*4*) also suggested that computer animations can be useful if the learning task demands students to understand concepts associated with visualization and motion. Therefore, the fact that animations did not have a significant effect on student responses to the visual conceptual questions is puzzling. Rieber (*4*) pointed out that animations are generally not as effective with older populations as they are with younger students: "College-aged students probably do not benefit from instruction which contains additional visual elaborations since they are able to form mental images without additional lesson support" (*4*, p. 11). However, several of the studies listed above (*6, 9-10*) reported significant animation effects for college students. Williamson and Abraham (*9*) postulated that students may only need to be cued to think about dynamic chemical processes on the molecular level to explain why additional exposure to animations did not increase students' abilities to think on the molecular level. Similar argument could explain why animation effects were not observed with our students. The engineering majors viewed several computer animations in lecture during the course of the semester depicting chemical processes (including acid-base chemistry, kinetics, and equilibrium reactions) on the molecular level before the experimental instruction was provided and a large number of students in both courses viewed these and other electrochemistry animations between the immediate and delayed post-tests, either in lecture or in laboratory. Therefore, it is possible that the engineering majors were prompted to think about chemical processes on the molecular level by the animations that were viewed before the experiment was performed. Viewing the electrochemistry animations after the experiment may also may prompted the non-physical sciences majors to think about the microscopic chemical processes occurring in galvanic cells. However, this does not explain why no animation effect was observed on the immediate post-test for the non-physical sciences majors.

The limited treatment effect observed for the conceptual change approach is probably best explained in terms of problems with the visual conceptual questions. The salt bridge questions may not have adequately discriminated between students believing that electrons are shuttled from the cathode to the anode by "piggybacking" onto ions in solution and those believing that cation and anion migration constitutes the flow of current in solution. For the engineers, the absence of a treatment effect for the electrode questions could be attributed to student difficulty in understanding the question and in interpreting the distractors. Because a disclaimer concerning the flow of electrons within the electrodes was made to the non-physical sciences majors, these questions appeared to discriminate between students who did and did not hold the misconception. The absence of a treatment effect on the delayed post-test was attributed to the fact that student practice of the visual conceptual questions obscured any effect that would have been present.

The interaction effect of the animations and the conceptual change approach for the non-physical sciences majors on the electrode questions of the immediate post-test suggests that instruction including the combination of animations and conceptual change was more effective than animations alone or no treatment at all. Previous research (*6, 8, 12*) has demonstrated that the combination of conceptual change instruction and computer animations can be effective at preventing or dispelling student misconceptions. This study suggests that, at least for some students, both treatments are necessary. Rieber (*4*) pointed out that novice learners may not know how to attend to relevant cues or details provided by animations and the authors presume that conceptual change instruction prompted the non-physical sciences majors to attend to the relevant information depicted in the animations (i.e., that the flow of current in solution consists of ion migration).

Verbal Conceptual Questions. For the verbal conceptual questions answered by the engineering students on the immediate post-test, there was a significant treatment effect for the conceptual change approach ($p < .0001$) and a significant interaction between the use of animations and the conceptual change approach ($p = .0084$). The results of the ANOVA for the verbal conceptual questions answered by the engineering students on the immediate post-test are listed in Table 13. Comparison of the mean scores for the engineering students who did and did not receive conceptual change instruction (3.13 versus 2.27) demonstrates that the conceptual change approach was effective at preventing or dispelling this misconception. A Tukey/Kramer post-hoc test was performed to determine significant differences among the means of the four treatment groups. The results of this test appear in Table 14. This test demonstrates that the mean of the group receiving conceptual change instruction is significantly higher

Table 13. ANOVA Results for the Verbal Conceptual Questions on the Immediate Post-Test for the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.06	0.06	0.05	.82
Conc. Change	1	22.96	22.96	19.27	.0001*
Animation × Conc. Change	1	8.54	8.54	7.17	.0084*
Error	133	158.42	1.19		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.00 34	2.55 33	2.27 67
	Yes	3.33 40	2.87 30	3.13 70
	Totals	2.72 74	2.70 63	2.71 137

* Significant at the .05 level.

Table 14. Tukey/Kramer Post-Hoc Tests Comparing the Means of the Treatment Groups Reported in Table 13.

Groups Compared	Difference in Means	*Q* value[a]
Both vs. Conc. Change	-0.46	2.46
Both vs. Anim.	0.32	1.65
Both vs. Control	0.87	4.48*
Conc. Change vs. Anim.	0.78	4.30*
Conc. Change vs. Control	1.33	7.36*
Anim. vs. Control	0.55	2.89

[a] For four groups and $df_{error} = 133$, the critical Q value ($\alpha = 0.05$) is 3.68.

* Significant at the .05 level.

than the means of the group receiving animations alone or the group receiving neither treatment (control). The mean of the group receiving both instructional methods is also significantly higher than the mean of the group receiving neither treatment (control). The interaction effect suggests that among the students who received conceptual change instruction, there appeared to be a slight negative effect associated with the animations (3.33 for the conceptual change group versus 2.87 for the group that received both treatments).

For the verbal conceptual questions answered by the non-physical sciences majors on the immediate post-test, there was a significant treatment effect for the conceptual change approach ($p = .0018$). The results of the ANOVA for the verbal conceptual questions answered by the non-physical sciences students on the immediate post-test are listed in Table 15. Comparison of the mean scores for students who did and did not receive conceptual change instruction (3.13 versus 2.49) demonstrates that the conceptual change approach was also effective at preventing or dispelling this misconception among the non-physical sciences majors.

Table 15. ANOVA Results for the Verbal Conceptual Questions on the Immediate Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.30	0.30	0.30	.59
Conc. Change	1	10.76	10.76	10.69	.0014*
Animation × Conc. Change	1	1.08	1.08	1.07	.30
Error	110	110.75	1.01		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.35 40	2.66 29	2.48 69
	Yes	3.18 22	3.09 23	3.13 45
	Totals	2.65 62	2.85 52	2.74 114

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* Significant at the .05 level.

The verbal conceptual questions presented on the course examinations (delayed post-test) were very similar in content and structure to those asked on the immediate post-test. The results of the ANOVAs for the verbal conceptual questions answered by the engineering students on the delayed post-test are listed in Table 16. For the engineering students, there was still a significant treatment effect for the conceptual change approach (*p* = .027) and a comparison of the mean scores for the engineering students who did and did not receive conceptual change instruction (2.64 versus 2.39) demonstrates that the effectiveness of the conceptual change approach at preventing or dispelling this misconception persists over time (at least one month). The results for the non-physical sciences majors, however, are more complex. For the verbal conceptual question on the

Table 16. ANOVA Results for the Verbal Conceptual Questions on the Delayed Post-Test for the Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.11	0.11	0.27	.61
Conc. Change	1	2.11	2.11	4.99	.027*
Animation × Conc. Change	1	0.08	0.08	0.20	.66
Error	132	55.66	0.42		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.44 34	2.33 33	2.39 67
	Yes	2.64 39	2.63 30	2.64 69
	Totals	2.55 73	2.48 63	2.51 136

* Significant at the .05 level.

delayed post-test, there was a significant effect for conceptual change instruction (p = .0029); however, comparison of the students' mean scores demonstrates that students who received conceptual change instruction performed worse than those students who did not (1.80 versus 2.23). The results of the ANOVAs for the verbal conceptual questions answered by the non-physical sciences students on the delayed post-test are listed in Table 17. The average scores for the treatment groups on the immediate (Table 15) and delayed post-test (Table 17) seem to imply that while both groups who did not receive conceptual change instruction were relatively constant in their responses over time, the proportion of student responses consistent with the scientifically-accepted conception decreased dramatically in both groups who received conceptual change instruction.

Table 17. ANOVA Results for the Verbal Conceptual Questions on the Delayed Post-Test for the Non-Physical Sciences Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Animation	1	0.66	0.66	1.05	.31
Conc. Change	1	5.82	5.82	9.32	.0029*
Animation × Conc. Change	1	2.01	2.01	3.22	.075
Error	109	68.15	0.63		

Incidence Table (Average Scores, Number of Subjects)

		Anim.		
		No	Yes	Totals
Conc. Change	No	2.05 40	2.48 29	2.23 69
	Yes	1.86 21	1.74 23	1.80 44
	Totals	1.98 61	2.15 52	2.06 113

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* Significant at the .05 level.

Because the verbal conceptual questions were not directly concerned with visualization or motion, we did not expect to see an animation effect for these questions. The interaction between the animations and conceptual change instruction measured on the immediate post-test for the engineering majors implies that when both methods were presented, the animations may have been distracting. The distractive nature of static visual pictures has been reported (*41-42*) and Dwyer concluded that visuals containing realistic details may require more processing time and better abilities to attend to relevant cues than simple visuals. It is unlikely that the animations' level of complexity was the cause of the distraction, since the students viewed each animation at least three times with the first author visually emphasizing the relevant processes that were being depicted. Rieber (*4*) pointed out that using animations for instructional tasks that do not specifically relate to the attributes of visualization or motion may distract learners from the purpose of the lesson. It is possible that students who viewed the animations and the conceptual change instruction focused their attention on the identities and motions of the aqueous ions depicted in the animations instead of focusing on the concept that it is ions, and not electrons, that migrate in solution.

On the immediate post-test, conceptual change instruction significantly decreased the proportion of student responses consistent with the misconception in both courses. These results clearly demonstrate that the chemical demonstration-based conceptual change instruction used in this experiment was effective at preventing or dispelling the misconception that the flow of current in electrolyte solutions consists of the migration of electrons in solution. Comparison of the proportion of student responses consistent with the misconception on the immediate and delayed post-tests (32% versus 16% for the engineering majors and 32% versus 31% for the non-physical sciences majors, respectively) demonstrates that student practice had a smaller effect on the verbal questions compared to the visual questions. This differential effect of student practice is not unexpected. Because the students were not accustomed to answering visual questions, they practiced these types of questions before the examinations (delayed post-test). However, the verbal questions appeared in a form that was very recognizable to the students (verbal multiple-choice questions) and they did not see the need to practice these kinds of questions.

The effects of conceptual change instruction on the delayed post-test, however, are more difficult to explain. For the engineering majors, the positive effect for conceptual change instruction on the immediate post-test was also present on the delayed post-test and suggests that the chemical demonstrations and subsequent discussions had a lasting effect on the students' conceptions of current flow in aqueous solutions. For the non-physical sciences majors, however, the positive effect for the conceptual change approach on the immediate post-test did not persist; in fact, a negative effect for conceptual change instruction was measured. This means that students who received conceptual change instruction were more likely to choose a response consistent with the misconception than students who did not receive this instruction. It is possible that conceptual change instruction did have significant effect on students, but that the explanations and conclusion were not as memorable—i.e., the students may have remembered the demonstrations were used to determine the correct description of current flow in aqueous solutions, but they could not remember whether it was concluded that electrons or ions flowed in solution. Happs (*25*) reported that earth sciences students in New Zealand also experienced regression of concepts concerning rocks and minerals after receiving conceptual change instruction. Although these students appeared to replace their preconceptions with scientifically-accepted conceptions as a result of conceptual change instruction, the preconceptions persisted and tended to displace the newer conceptions over a period of three months. The authors believe that the reason why the non-physical sciences majors experienced regression while the engineering majors did not is related to different instructional techniques used by the instructors in these courses. The instructor of the course taken by the engineering majors consistently discussed the connections among the macroscopic, microscopic, and symbolic representations of chemical processes (*43*), including the use of computer animations of chemical processes (including electrochemistry) on the molecular level, and verbally reinforced the concept that electrons do not flow in solution; the instructor of the course taken by the non-physical sciences majors did not show computer animations in lecture and did not verbally confront the misconception that electrons can flow in solution. We believe that the lecture animations and comments made by the former instructor reinforced the effect of the conceptual change approach used in this experiment and may be at least partially responsible for the lasting effect of the conceptual change approach on the engineering majors.

Participants versus Non-Participants. Because the delayed post-test was administered as part of the examinations in both courses, delayed post-test data was also collected for students who did not attend recitations when the instructional treatments were performed. This provided the authors with a chance to compare students who received instruction using animations or the conceptual change approach with students who did not. Of the 190 engineering majors enrolled in this course, 135 received instructional treatments in recitations and 55 did not. The number of non-physical sciences majors who did not participate in this experiment, however, was much smaller (only 7 of the 122 students enrolled in the course failed to attend this recitation). As a result, only the responses from the engineering majors were analyzed.

The average pre-test score of the non-participants was significantly lower than that of the participating students (65.3 versus 73.6, $t_{188} = 3.920$, $p < .0001$); therefore, it was necessary to compare the visual and verbal conceptual questions on the delayed post-test using ANCOVAs. A one-way ANCOVA was performed on the verbal and visual delayed post-test scores for the participants and non-participants using pre-test scores as a covariate (Tables 18 and 19, respectively). The ANCOVA for the verbal scores demonstrates that participation in the study did not affect students' abilities to answer the verbal questions ($p = .64$). The significant value for the pre-test ($p < .001$) implies that students' chemistry knowledge (as evidenced by their previous exam scores) did affect their ability to answer the verbal conceptual questions. The ANCOVA for the visual question, on the other hand, demonstrates that participation in the study significantly affected students' abilities to answer the visual questions ($p = .005$); the significant effect for the pre-test ($p < .001$) also suggests that students' chemistry knowledge affected their abilities to answer the visual questions.

Participation in this experiment effected students' abilities to answer the visual conceptual questions, but did not have an effect on their abilities to answer the verbal conceptual questions. It should be noted that students who do

Table 18. ANCOVA Results for the Verbal Conceptual Questions on the Delayed Post-Test for the Participating and Non-Participating Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Covariate (Pre-Test)	1	16.70	16.70	34.66	.001*
Main Effect (Participation)	1	0.10	0.10	0.21	.64
Explained	2	18.87	9.43	19.58	.001*
Residual	187	90.10	0.48		
Total	189	108.97	0.58		

Incidence Table (Pre-Test Scores, Verbal Scores)

Group	Count	Pre-Test Score	Verbal Score
Participants	135	73.6	2.51
Non-Participants	55	65.3	2.28

* Significant at the .05 level.

Table 19. ANCOVA Results for the Visual Conceptual Questions on the Delayed Post-Test for the Participating and Non-Participating Engineering Majors

Source	*df*	Sum of squares	Mean square	*F* value	*p* value
Covariate (Pre-Test)	1	26.46	26.46	32.82	.001*
Main Effect (Participation)	1	6.56	6.56	8.13	.005*
Explained	2	43.55	21.77	27.01	.001*
Residual	187	150.77	0.81		
Total	189	194.32	1.03		

Incidence Table (Pre-Test Scores, Visual Scores)

Group	Count	Pre-Test Score	Visual Score
Participants	135	73.6	3.77
Non-Participants	55	65.3	3.11

* Significant at the .05 level.

not attend recitations are also more likely not to attend lectures; therefore, any differences between participants and non-participants cannot be attributed solely to the instructional treatments that occurred in recitation. Because the verbal questions appeared in a form that students are likely to be familiar with (simple multiple-choice questions), it isn't surprising that recitation (and lecture) attendance did not have a significant impact on students' abilities to answer these questions. However, visual questions, such as those asked in this study, are not widely used in traditional chemistry courses (*40, 44-45*). The instructor of the course taken by the engineering majors stressed in lecture the importance of being able to explain chemical processes on the molecular level, used computer animations and conceptual questions in lecture, and used conceptual questions on recitation quizzes. Therefore, it is likely that students attending lecture and recitation would have been prompted to think about visual questions. Students who did not answer the visual questions on the immediate post-test or did not regularly attend lecture or recitation may not have discovered that it was important for them to be able to answer visual questions that required them to think about chemical processes on the molecular level. As a result, these students were less prepared to answer visual questions on the midterm examination (delayed post-test) compared to the students who did participate in this study. The effects of student practice on the delayed post-test can also be explained by the students' familiarity with visual and verbal questions. Because these students were not accustomed to answering visual questions, they practiced these questions for their examinations (delayed post-test) until they were able to solve them. As a result, any treatment or interaction effects that may have been present were obscured. However, because the verbal questions did not appear to be novel to these students, they did not practice them and any treatment effects that were present could be measured.

Conclusions

Instruction using animations that depicted electrochemical processes on the molecular level was expected to have a significant effect on student responses to visual conceptual questions concerning the flow of current in electrolyte solutions. However, no effect was measured in this study for engineering and non-physical sciences majors. The lack of an animation effect for questions requiring students to visualize motions of ions and electrons may be attributed to the fact that college-aged students do not need visual elaborations since they are capable of forming mental images independently or that college students merely need to be prompted to think about these chemical processes on the molecular level. No treatment effect was expected for verbal conceptual questions that do not specifically require students to visualize motions on the molecular level and no effect was seen. Instructional interactions measured between the use of these animations and conceptual change instruction suggest that animations may be helpful when the questions require students to visualize chemical processes on the molecular level (visual conceptual electrode questions) but animations may prove to be distracting when the questions do not require students to visualize (verbal conceptual questions).

Conceptual change instruction using chemical demonstrations was effective at preventing or dispelling the student misconception that electrons flow in aqueous solutions to complete the circuit in an electrochemical cell for both visual and verbal conceptual questions. The effect for the visual questions was obscured by problems with these questions (discussed below). For the verbal questions, there was a positive conceptual change effect for the immediate post-test and for the engineering majors. However, the non-physical sciences majors experienced regression of the concept on the delayed post-test. These results suggest that successfully confronting a misconception requires more than a single instructional treatment—students are more likely to retain the scientifically-accepted conception if their misconceptions are constantly confronted.

Students appeared to be more familiar and comfortable with verbal questions compared to visual questions. Empirical evidence from this study suggests that students were more likely to misinterpret the visual questions, students were more likely to practice the visual questions, and lecture attendance and participation in this study was more likely to help students answer visual questions. Problems with visual conceptual questions were not limited to the students. The authors had considerable difficulty writing visual conceptual questions that students were able to understand and that were capable of discriminating between students who did and did not hold the misconception.

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SCIENCE-TECHNOLOGY-SOCIETY AND CHEMCOM COURSES VERSUS COLLEGE CHEMISTRY COURSES: IS THERE A MISMATCH?

A paper accepted for publication by the *Journal of Chemical Education*

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Abstract

This opinion paper poses questions for the chemical education community to consider about the nature of college chemistry courses and the expected student prerequisites and explores the basis for a potential mismatch that may exist for students enrolled in a traditional college chemistry course who have had Science-Technology-Society (STS) or ChemCom courses as their only prior high school chemistry courses. Even though ChemCom and STS courses are not designed to prepare students for traditional college chemistry courses for science and engineering majors, there has been an increase in the number of ChemCom and STS students enrolling in these courses. As background, the general teaching approach of ChemCom and STS courses is discussed in terms of the behaviorist, cognitive, and constructivist learning theories. This article addresses the advantages and disadvantages of high school chemistry courses based on ChemCom and reviews the chemical education research comparing the effectiveness of ChemCom courses to traditional high school chemistry courses. The authors concluded that more research comparing the effect of these instructional methods on both traditional and alternative assessments is warranted.

Introduction

Recently, our mid-western university has received several requests to evaluate whether or not high school Science-Technology-Society (STS) or ChemCom courses are acceptable to meet entrance requirements for traditional college chemistry courses for science and engineering majors. Even though the authors of ChemCom (including the American Chemical Society) have suggested that a chemistry course using ChemCom is most appropriate for non-science majors, many ChemCom students are enrolling in traditional college chemistry courses for science and engineering majors. At our university, the number of STS and ChemCom students enrolled in traditional introductory college chemistry courses for science and engineering majors has increased (*1*). Chemistry faculty members are reporting mixed success of STS and ChemCom students in their courses—some ChemCom students are successful while others are not. College chemistry instructors face a dilemma: What information and research studies are available to help faculty make decisions about acceptable high school chemistry courses? Are students who have had STS or ChemCom courses as adequately prepared to deal with college science and engineering courses as students who have had traditional high school college preparatory chemistry courses? What should faculty know about teaching techniques and learning theories that will help students, regardless of their high school background, be more successful in their classes? Should faculty deny enrollment in their classes to students who have had STS or ChemCom courses as their only prior high school chemistry course? Should college faculty change the nature of their introductory college chemistry courses? This opinion paper discusses several issues surrounding these questions and explores the basis for a potential mismatch students having an STS-based course may experience. Because there is a fundamental difference in teaching and learning philosophies between traditional chemistry courses and STS-based courses, it is necessary to discuss objectivism and constructivism.

The Science-Technology-Society Approach to Teaching

Since 1982, the Science-Technology-Society approach to teaching science education has steadily grown in the United States from a national imperative of the National Science Teachers Association (NSTA) to a full-fledged movement that is present in one form or another in more than 2,000 colleges and thousands of high schools (*2*). The NSTA imperative stated: "The goal of science education during the 1980's is to develop scientifically literate individuals who understand how science, technology, and society influence one another and who are able to make use of this knowledge in their everyday decision making. This individual both appreciates the value of science and technology in society and understands their limitations." (*3, p. 1*) The definition of science education as a discipline concerned with the interface between science and society has been debated (*4-6*) and the proponents of the STS definition have dominated the science education literature.

The goals of science education (using the STS definition) are to prepare students to use science in solving personal problems, resolving current societal issues, and choosing careers in science and technology. More importantly, however, the STS movement differs from the traditional method of teaching science in several ways:

> 1. STS stresses the importance of providing a real world context for science study so students can apply their knowledge to make educated decisions concerning societal issues such as acid rain, ozone depletion, etc.

2. STS emphasizes scientific principles as something that are discovered or "proven" as students attempt to understand their environment instead of absolute truths to be blindly accepted.

3. STS emphasizes science as a process of learning and that being able to defend one's conclusion is as important (or more important) than the individual results that led to the conclusion.

In spite of what Bybee (*6*), a proponent of STS, and NSTA (*2*) report, the inclusion of STS themes, alternative assessment techniques, and alternative teaching techniques in the majority of introductory college chemistry courses for science and engineering majors has not occurred. The majority of college chemistry instructors subscribe to the definition of science education discussed by Good, Herron, Lawson, and Renner (*4*).

Behaviorist versus Cognitive Learning Theories

A shift in the science education paradigm from one in which science is taught as universal "truths" that have been discovered or invented (and which students believe have merit and usefulness only in academic contexts) to one in which science is stressed as a process of thinking to solve problems and make decisions that are personally and socially relevant can be attributed to several factors.

One reason for this shift in science education is the notion that traditional methods of instruction are not effective (*6, 7*). Several articles published in *this Journal* expressed opinions on how to improve the introductory chemistry course (*8, 9*). The lecture method of instruction, the assignment of homework problems, and the use of multiple-choice examination questions has come under increasing attack. Students have difficulty applying their science knowledge to real-world, personal, authentic, or societal problems. Individuals are labeled as 'science illiterate' when they cannot understand science principles and real-world science problems that impact our environment. Topics such as ozone depletion, global warming, acid rain, recycling, and chemical waste disposal should be understood by the average citizen, but clearly many adults do not, even though they have successfully completed high school and college science courses.

Too often, chemistry instructors structure examination questions that can be answered by memorization or the application of an algorithm. As a result, students harbor 'inert knowledge' which they use only in school settings. Ben-Zvi and Gai (*10*) reported that high school students had more difficulty correctly answering questions based on real-world situations than comparable ones set in academic situations.

The paradigm shift in science education corresponds to a similar paradigm shift in educational psychology from behaviorist theory to the cognitive learning theory. The traditional method of teaching science has its roots in behaviorism, a learning theory that is more concerned about the outcome or product (getting the right answer) than the processes from which this answer was derived (*11*). In a behaviorist environment, the instructor assumes the role of 'sage on the stage', transferring knowledge to learners. Learners receive this knowledge with its structure and meaning intact. In contrast, cognitive learning theory stresses cognitive thought processes of the learner over the products of learning (*11*). Cognitive theory recognizes that the learner is an active agent in the learning process and that learner attributes (prior knowledge, attitude, motivation, learning style, etc.) affect the learning process.

Constructivist versus Objectivist Philosophy of Learning and Teaching

More recently, there has been another educational psychology paradigm shift from the objectivist notions that are a part of both the behaviorist and the cognitive learning theories to constructivist ideals. Objectivism is a philosophy that assumes that there are objective, absolute, and unconditional truths that are discovered in the process of learning and these truths are independent of the context in which they are observed (*12*). In contrast, the constructivist philosophy assumes that learners construct knowledge, and therefore no two learners will have the same internal representations of this knowledge because each learner brings a different perspective to the learning process (*12*, *13*). Constructivists believe that the learner imposes order on the world; objectivists believe that the learner observes the order that is inherent in the world. To a constructivist, learning occurs by an individual constructing his or her own personal knowledge of the subject, especially in a real world context. "Constructivists believe that knowledge is constructed based upon what 'works' and what is 'good' in the particular context in which the cognizing individual is operating" (*14, p. 628*). A key component of constructivism is the negotiating of the meaning of knowledge with others in order to develop a mutually-shared meaning. Working in groups on a project provides students the opportunity to identify relevant issues or problems, develop tasks that will help solve the problem, examine solutions, and debate alternative viewpoints. Proponents of constructivism see this these activities as being closer to how scientists go about doing science. Another key component of constructivism is that the individual learner has a primary role in determining what will be learned, how it will be learned, and how this learning should be evaluated. Instructors subscribing to constructivism use a variety of assessment techniques (portfolio analysis, group projects, concept maps, etc.) to help students evaluate their progress and reveal their personal understanding of chemistry.

Here, then, is the crux of the problem. Most college chemistry instructors are not aware of nor do they subscribe to the constructivist philosophy of teaching and learning. Most college chemistry instructors believe that there is an objective reality and that chemists can describe, measure, and work with reality using the tools that chemistry concepts and principles provide. The aim of the instructor is to "transmit the knowledge experts have acquired to students because experts' knowledge is much closer to reality than beginners' knowledge" (*14, p. 628*). Most college chemistry instructors devise examination problems for which there is one correct answer, and they expect students to generate this answer. Student knowledge of chemistry is evaluated on the basis of their ability to correctly answer questions and problems devised by their instructor. According to constructivists, these questions and problems do not show what the students' understanding of a topic is, only if the students can generate an answer that matches the answer devised by the instructor. Students having an STS or ChemCom constructivist-based chemistry course may experience culture shock when thrust into a traditional objectivist-based college chemistry course.

The constructivist philosophy often incorporates teaching techniques that are used by objectivists. Some college chemistry instructors do incorporate in their teaching cooperative learning, authentic learning tasks, student portfolios, seminar-style discussion, group projects, and group examinations. Although instructors may use similar teaching techniques, their goals and aims for students may differ.

ChemCom

In response to the changing view of science education, the American Chemical Society (ACS), with the support of the National Science Foundation, ACS Corporation Associates, and the Petroleum Research Fund, has taken the general guidelines of STS to create a chemical education version of STS called "Chemistry in the Community" or ChemCom (*15*). ChemCom is a year-long course built around eight societal issues related to chemistry and emphasizes the development and use of decision-making strategies to discuss and solve personal, local, and global problems. Although ChemCom includes less math and fewer physical chemistry topics than traditional high school chemistry courses, it includes more organic and nuclear chemistry.

In contrast to other STS proponents, the authors of ChemCom are more conservative in their aspirations for ChemCom: The intent of ChemCom is not to replace traditional chemistry instruction, but rather to complement it. ChemCom is most often recommended as an appropriate alternative to the traditional chemistry course only for those students that do not intend to pursue a career in the scientific fields (*16*).

Advantages of ChemCom

ChemCom emphasizes chemistry in a decision-making and problem-solving context. Real-life problems are addressed in ways that allow students to develop the skills needed to solve these problems the way scientists do. In this way, students learn chemistry, not as an unrelated collection of facts and laboratory skills that are useful only for solving academic questions in the classroom, but as a set of techniques and thought processes that can be used in a systematic way to ask and answer real-world questions that have personal and social relevance. ChemCom stresses the role of controversy, debate, and personal opinion in the field of chemistry.

ChemCom introduces and develops chemistry information, concepts, and techniques when they are needed to solve real-world problems instead of introducing the concepts and techniques first and then trying to apply them to real-world problems. As a result, there is a relevance inherent in these concepts and techniques that should prevent them from becoming inert knowledge that cannot be transferred to new and subtlety different real-world situations.

Traditional courses present chemistry as the collection of information and concepts that are the result of scientific controversy, debate, and conflicting personal opinions—by the time students get involved in learning chemistry, all that remains of the scientific inquiry process is the consensus that has been agreed upon by the "experts". If instructors do not make an effort, students can be passive learners, all too eagerly accepting information given to them without questioning the accuracy of the information or searching for possible biases. On the other hand, ChemCom encourages direct student participation in the scientific learning process, which includes identifying problems, proposing and evaluating alternative solutions, separating fact from opinion, evaluating the objectivity and usefulness of sources, verifying information, and reaching logical conclusions from the information given.

Disadvantages of ChemCom

New teaching methods typically face opposition from parents, school board members, and teachers. Parents and teachers are likely to oppose or resist new methods because they are different from those used when they were in school and may require substantial amounts of initial work on the part of the teacher to make them viable. As an example of this reluctance to accept new methods, Shamos (*17*) voices opposition to STS because of its constructivist origins: "...It is more than simply a battle of words between scientists and those social scientists, calling themselves 'constructivists,' who assert that, contrary to the scientists' view, nature cannot be studied objectively..." (*17, p. 69*). Shamos also links constructivists to other "anti-technology fringe elements" who view science and technology as the source of all evil in society and who are attempting to reduce the precision and predictive status of the natural sciences to the extent found in the social sciences.

Constructivist teaching methods usually require teachers to have a better grasp of the course material than traditional lecture methods. The course content and the classroom management techniques required to teach ChemCom differ markedly from that of traditional high school chemistry courses (*18*); therefore, many teachers may feel under-qualified to teach a ChemCom course. The authors of ChemCom have identified this as a problem and have implemented an extensive teacher training program for teachers using ChemCom (*18*). A survey comparing student views on the interactions of science, technology, and society among high school graduates and undergraduate students (*19*) indicated that these two groups are nearly identical in their understandings of STS issues and that the existing university science courses taken by the undergraduate students have not affected their understanding of STS issues. This result is expected since the majority of college chemistry courses do not include STS issues and suggests that future high school chemistry teachers may not be adequately prepared to teach a course emphasizing STS themes and methods.

ChemCom covers a fraction of the content covered in traditional chemistry courses and has drastically reduced the mathematics and physical chemistry content presented to students. This appears to be a concern of the publisher of ChemCom and may be the reason ChemCom is not officially recommended for students who plan to be science or engineering majors in college.

Constructivist teaching approaches usually face problems with assessments such as grade assignment (*20*). Specifically, many constructivist instructional approaches assess student performance using objectivist measures such as quizzes and examinations, often because they are easy to write, administer, and grade. The problem with this mismatch is that objectivist assessment procedures do not test students on the criteria that are important to the constructivist instructional techniques—objectivist criteria do not generally test the students' ability to solve problems, they simply test student knowledge and test-taking ability. ChemCom tends to use constructivist assessment techniques (e.g., essays, self-evaluations, lab practicals, portfolios, group projects, etc.) that allow students to demonstrate their problem-solving and near- and far-transfer abilities instead of their factual knowledge. Students who have studied chemistry in high school using the constructivist assessment techniques associated with ChemCom may face difficulties with traditional college chemistry courses that use traditional objectivistic assessment methods.

Review of ChemCom Research

Perhaps it is premature to review the research comparing the effectiveness of ChemCom courses. A search of the literature revealed only two studies that directly compared the achievement of students using ChemCom versus those using traditional instructional methods (*21, 22*). One study (*22*) compared pre- and post-test scores from students enrolled in a ChemCom and a traditional chemistry course on the *Chemistry Test* of the *Comprehensive Assessment Program High School Subject Tests*. Although the test revealed significantly higher post-test scores for ChemCom students compared to the students in the traditional course, this result is subject to scrutiny due to several possible threats to the study's validity. These threats to validity include: Poor content validity of the *Chemistry Test*; poor test-retest reliability (because the pre-test and post-test were identical); possible experimenter bias (because different instructors taught each treatment group); reported sampling biases (students were not randomly assigned to treatment groups and two of the four ChemCom classes were designated as Honors sections; none of the four traditional classes were so designated); an anomalously high drop-out rate (greater than 50%); and possible novelty effects. By the authors' own admission, this study should be carefully and more rigorously replicated before this result can be fully accepted. The other study (*21*) compared the changes in the number of students performing at the concrete, transitional, and formal operational levels (as measured by the *Group Assessment of Logical Thinking*, GALT, test) after a year-long course in chemistry based on either ChemCom or a traditional instructional method. Although this study showed a slight increase in the number of ChemCom students performing at the formal operational level and a slight decrease in the number of ChemCom students performing at the concrete level, this difference was not significant.

High School Chemistry as a Prerequisite for College Chemistry

Is a course in high school chemistry a necessary prerequisite for success in college chemistry? In an attempt to determine whether taking a traditional high school chemistry course is required to successfully complete a college chemistry course, Krajcik and Yager *(23)* taught a college level chemistry course to 28 high ability students (14 with a previous chemistry course, 14 without a chemistry background). The study revealed that there was no significant difference in achievement between the groups after the eight-week course. The results of this study are not generalizable because all students were "high ability" students. In addition, all students were given free tutors. Those with prior chemistry coursework averaged two hours per week with the tutors, while those without prior coursework in chemistry averaged eight hours per week with the tutors. This difference is significant and seems to suggests that students with prior chemistry coursework are at a distinct advantage over those who have had no previous chemistry coursework. The number of hours that students spent studying chemistry on their own was not reported. If students without a high school chemistry course need additional hours with tutors in order to be successful, then this places these students at a disadvantage since most student do not have the money to pay for eight hours per week of tutoring.

Statistical studies correlating ACT scores (or SAT scores) and grades in high school subjects with grades in college chemistry courses consistently show achievement on ACT math scores (or SAT math scores) as the best predictor of success in college chemistry *(24-26)*. Why then do college chemistry instructors insist on making high school chemistry a prerequisite for college chemistry? Most college and universities report that 20-35% of their students are not successful in passing the introductory chemistry course. Nearly all of these students have had a traditional high school chemistry course as a prerequisite. Is there any reason to believe that ChemCom students will be less successful?

Does it really matter if students have had a traditional high school chemistry, a ChemCom course, or no high school chemistry course? It is our perception that college chemistry faculty have direct experience with students who are "not successful" in their courses. By in large, the students themselves attribute their lack of "success" to an inadequate high school chemistry course when compared to courses taken by their peers. Many "unsuccessful" students identify the ability of their high school chemistry teacher to teach chemistry as being a critical factor for their success in college chemistry. Are students "successful" in traditional college courses because their high school teachers subscribe to similar philosophies of teaching and learning? By in large we define success in chemistry in terms of doing well on teacher-constructed examination problems and questions.

Conclusions

Based upon our analysis of the situation, more information is needed to help college chemistry faculty make decisions regarding whether or not to accept STS and ChemCom courses as acceptable prerequisites. Research studies are needed comparing learner attributes of students completing a ChemCom or STS course versus students completing a traditional high school chemistry course. Research studies are needed documenting the success of ChemCom or STS students completing traditional and non-traditional college chemistry courses. One issue at hand is whether or not it is appropriate to compare students who have had ChemCom or an STS course with students taking a traditional high school chemistry course using standardized exams, such as the 1995 ACS High School Chemistry Examination (*27*). Proponents of ChemCom argue that the goals of ChemCom differ from those of traditional chemistry courses; therefore, traditional exam questions appearing on the ACS examination are not appropriate for ChemCom students.

A problem facing college students who have had ChemCom as their only prior chemistry course is a general lack of experience with traditional chemistry problems. Constructivists and proponents of ChemCom might argue that focusing on fewer topics but covering them in more depth (the "less is more" argument) better prepares students to think critically about chemistry than superficially covering several chemistry topics. However, the results of Krajcik and Yager's study (*23*) suggest that overcoming a lack of prior chemistry knowledge can be very time-consuming and difficult, especially if students are enrolled in other courses. Also, there seems to be an incompatibility between college faculty teaching science and engineering chemistry courses in the traditional manner and students entering these courses with a chemistry background in ChemCom that emphasizes cooperative learning techniques, group projects, and non-competitive assessment techniques. As a result, some students experience difficulty when faced with the competitive, problem-solving nature of traditional college chemistry courses and examinations (*28*). Should college faculty revise their approach to evaluation of students?

Our personal experience suggests that even in states where STS and ChemCom courses are commonly taught, traditional college chemistry instructors at universities, colleges, and community colleges (who tend to be chemists rather than chemical educators) have little or no knowledge concerning non-traditional teaching methods, learning theories, non-traditional assessment techniques, etc. (*29*). Nor are they incorporating STS themes and material into the chemistry courses they are teaching (*29*).

When planning future studies, researchers should carefully consider learner attributes that ChemCom is likely to affect—for example, near- and far-transfer of chemical knowledge, problem-solving skills, ability to solve personal and societal problems using chemistry, and interest in solving personal and societal problems using chemistry all appear to be appropriate topics for assessment. Current research has focused on documenting that ChemCom is not less effective than traditional instructional methods on traditional assessment measures (such as student scores on standardized achievement tests). What is needed are studies focusing on areas and features where STS and ChemCom are effective. Even though these courses are designed for students not planning to enroll in a college chemistry course designed for science and engineering majors, the number of ChemCom students enrolling in traditional college chemistry courses is increasing. Therefore, studies are warranted using traditional assessment techniques such as standardized exams as one measure of achievement, as well as non-traditional assessment techniques. We encourage researchers to investigate these issues and report the results of their research.

Until research sheds light on the issues discussed in this paper, college chemistry instructors must decide for themselves what they believe is the best course of action. There are reports that changes in the college chemistry curriculum are underway (*30-32*), but it will take time for the majority of college chemistry instructors to change the nature of their traditional college chemistry courses. As a result, students who are underprepared or do not meet current prerequisites for college chemistry will seek their own ways to succeed—students will continue to hire tutors, spend extra time studying, or enroll in college preparatory chemistry courses.

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GENERAL CONCLUSIONS

Summary of Results

In the interview study, we were able to confirm most of the student misconceptions in electrochemistry reported by Garnett and Treagust (*1*). In addition, we identified several new misconceptions concerning galvanic, electrolytic, and concentration cells. These misconceptions include the notions that half-cell potentials are absolute in nature, electrons can flow in aqueous solutions without assistance from ions, cation migration does not constitute a flow of current in solution, half-cell potentials are extensive properties, inert electrodes can be oxidized or reduced, it is impossible to predict electrolysis products, the direction of electron flow and cell potentials of concentrations cells are not dependent on ion concentrations, and the indirect reaction occurring in an electrochemical cell is different from the direct reaction of the reactants. Student misconceptions were attributed to ignorance of the relative nature of cell potentials and imprecise or inappropriate language used by textbooks.

The salt bridge paper reported student misconceptions regarding current flow in electrolyte solutions and discussed two student descriptions regarding electron flow in electrolyte solutions observed in the interview study: Electrons migrate through solution by attaching themselves to ions at the cathode and are shuttled to the anode by these ions, or electrons migrate through solution without assistance from ions in solution and travel as free electrons from the cathode to the anode. This article implicated imprecise or inappropriate textbook language as a possible source of student misconceptions and provided several textbook quotes as evidence. This article also demonstrated that instruction including the use of computer animations depicting chemical processes on the molecular level and an instructional method aimed at confronting the misconception that electron migration in solution constitutes a flow of current decreased the proportion of students consistently demonstrating this misconception.

The textbook analysis study demonstrated that introductory college-level chemistry textbooks contain statements and illustrations that could be misinterpreted by college students as corroborating common misconceptions in electrochemistry. These misconceptions include the notions that half-cell placement determines anode/cathode identity, half-cell potentials are absolute and additive in nature, electron migration in aqueous solutions constitutes a flow of current, cation migration in aqueous solutions does not constitute a flow of current, electrode charges determine the flow of electrons and ions in a cell, and electrolysis products cannot be predicted. The authors proposed suggestions for textbook authors, including avoiding the use of simplifications, avoiding the use of vague or misleading statements, calculating cell potentials by the difference method, avoiding the use of electrostatic arguments to predict ion and electron flow in electrochemical cells, and considering all possible oxidation-reduction half-reactions when predicting electrolysis products. The authors also proposed a method for predicting electrolysis products using potential ladder diagrams.

The final study demonstrated that conceptual change instruction based on chemical demonstrations was effective at preventing or dispelling the student misconception that electrons flow in aqueous solutions to complete the circuit in electrochemical cells for both visual and verbal conceptual questions. Although computer animations of chemical processes on the molecular level were also expected to have an effect on students' responses to visual conceptual questions, the animations used in this study did not appear to have an effect on students' conceptions. The lack of an animation effect may be attributed to the fact that college students do not need instructional computer animations (or only need to be prompted by these animations) to form mental images of these chemical processes. Animation/conceptual change interactions suggested that animations may be helpful when the questions require students to visualize chemical processes on the molecular level (visual conceptual questions) but may prove distracting when the questions do not require students to visualize (verbal conceptual questions). Empirical evidence from this study also suggested that students were more likely to misinterpret visual questions, more likely to practice visual questions, and that lecture attendance and recitation participation during which computer animations were shown was more likely to help students answer visual questions.

Discussion of Results

We were able to confirm most of the student misconceptions reported by Garnett and Treagust (*1*). Because these misconceptions were identified using two different samples (high school students in western Australia and college students in midwestern United States), we can be more confident in generalizing these misconceptions to other populations. We proposed the imprecise or inappropriate use of language by textbooks (with specific examples) as one possible source of these student misconceptions. The use of imprecise or inappropriate language has been cited by several other authors (*1-3*) as a possible source of student misconceptions.

The primary limitation of the textbook analysis is that it is based on the authors' opinions regarding how students might misinterpret statements made in chemistry textbooks and is not based on student interpretations of these statements. In the course of examining the oxidation-reduction and electrochemistry chapters of ten college-level chemistry textbooks, the authors discovered that there is very little variation in the content or the style of delivery among these textbooks. This lack of variation has been reported previously (*4*). Therefore, the statements reported in this study that could be misinterpreted as corroborating student misconceptions in electrochemistry are likely to be present in other college-level (and perhaps even high-school-level) chemistry textbooks. The final conclusions reported in this study are not novel: Most have been reported previously (*1, 3, 5-9*). Misconceptions that have been widely reported were supported by fewer textbook statements than those that have not been widely reported. This suggests that making textbook authors aware of any misleading statements in their textbooks causes them to examine and change the wording used in their textbooks.

The final study was an attempt to actively remediate an electrochemistry misconception using a combination of computer animations and conceptual change instruction. The effects of the computer animations may have been obscured or confounded due to the fact that many of these students viewed the animations used in this study in the lecture or laboratory. Most of these students would have viewed these animations after the immediate post-test, so the contamination should be limited to the delayed post-test. Some of these students also viewed computer animations in lecture covering different topics (acid-base chemistry, kinetics, and equilibrium reactions) before this study was performed. There was also additional problems concerning the visual conceptual questions. Some students had trouble interpreting the visual questions used on the immediate post-test. The delayed post-test demonstrated that students who attended recitation tended to practice solving visual questions. Before the immediate post-test, students did not have access to visual conceptual questions. However, students received examples of visual conceptual questions in the immediate post-test, during lecture, and possibly in recitation. It is not unreasonable to believe that students would practice questions that appear unfamiliar to them; personal experience suggests that some students would focus on rote memorization instead of concept learning. Attendance in lecture and recitation also helped students solve these questions. This is consistent with Pickering's conclusion (*10*) that students have difficulty answering visual conceptual questions based on the particulate nature of matter, not because they are unable to understand this concept but because they have not been previously asked to understand it or to solve such problems on homework or quizzes (*11*).

Conceptual change instruction significantly decreased the proportion of student responses consistent with the misconception on the immediate post-test and on the delayed post-test for the engineering majors. This is consistent with other chemical education research involving the effects of conceptual change instruction (*12-15*). However, the non-physical sciences majors demonstrated a negative effect for conceptual change instruction on the delayed post-test. This is consistent with the results reported by Happs (*16*) concerning the regression of concepts. The authors believe that the reason the non-physical sciences majors experienced regression while the engineering majors did not is due to the fact that the instructor for the engineering majors consistently confronted the misconception in lecture while the instructor for the non-physical sciences majors did not.

Although instruction using computer animations was expected to have an effect on student responses to visual questions, no effect was measured in this study. This is inconsistent with research performed by other chemical educators (*17-20*), who did see a positive effect of animation on student conceptions. However, it is consistent with Rieber's conclusions (*21*) that college-aged students do not benefit from instruction containing visual elaborations because they are able to form mental images on their own just by reading the textbook or by studying lecture notes. The interactions measured between the animation and the conceptual change instruction suggests that animations may be helpful when the questions require students to visualize chemical processes on the molecular level, consistent with Williamson and Abraham's results (*17*), but may prove to be distracting when the questions do not require students to visualize. This is consistent with Rieber's conclusion (*21*) that using animations for instructional tasks that do not specifically relate to the attributes of visualization or motion may distract learners from the purpose of the lesson. The distractive nature of static visuals has also been reported by Dwyer (*22*) and Willows (*23*).

Suggestions for Additional Research

The results of the series of electrochemistry studies performed by the authors have prompted several ideas for additional research. Is the use of computer animations worth the time required to display them in lecture? Does the segment of students who have visual learning styles benefit from viewing these animations?

The primary limitation of the textbook analysis is that it is not based on student interpretations of the statements made by the textbook authors. It would be appropriate to ask a group of students to read some of these statements and comment on how they would interpret each statement. This way, we would not be relying on the authors' interpretations of these statements. It would also be wise to test new any new statements by asking students to read and interpret them.

In the final study, we suggested that the reason no significant animation effect was measured is that college-level students only need to be prompted to think about chemical processes on the molecular level. Those students who viewed animations of chemical processes on the molecular level before participating in this study may already be in the mind-set required to think about these processes on the microscopic level. Therefore, the effect of computer animations depicting chemical processes on the molecular level should be tested with students who have not previously seen animations of this kind. We also suggested that the negative effect for conceptual change instruction experienced by the non-physical science majors was related to the fact that the instructor did not constantly confront the misconception in lecture. It would be wise to test this hypothesis by providing the same conceptual change instruction to two groups of students, with one group receiving subsequent instruction that actively confronts the misconception in question and the other group receiving subsequent instruction that does not address this misconception.

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