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USING CONSTRUCTIVISM IN TEACHING AP CHEMISTRY

By

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ABSTRACT

USING CONSTRUCTIVISM IN TEACHING AP CHEMISTRY

By

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Constructivism has been found to be an effective tool in understanding how humans acquire knowledge. Learning a conceptual subject such as the acid-base unit in AP chemistry is a rather challenging task for most students. This study involved assessing the effectiveness of the constructivist approach in teaching the acid-base unit of AP chemistry. The approach was complemented with certain brain-based instructional strategies.

The study involved students enrolled in a high school AP chemistry class. The acid-base topic was selected as the experimental unit. Students were engaged in observing demonstrations, performing lab experiments, and solving practice problems. Analogies and graphic organizers were utilized to assist students with making connections. These activities were designed to bring about conceptual change as suggested by the constructivist approach.

Statistical analysis revealed that students' performance improved significantly upon completion of the acid-base unit. Using the constructivist approach together with certain other instructional strategies contributed to improved learning of a rather challenging subject.

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INTRODUCTION

The Advanced Placement (AP) Chemistry course is designed to be the equivalent of the general chemistry course usually taken during the first year of college. Students in such a course should attain a depth of understanding of fundamentals and a reasonable competence in dealing with chemical problems. The course should contribute to the development of the student's abilities to think clearly and to express their ideas, orally and in writing, with clarity and logic (The College Board AP Chemistry Course Description, 2003).

During my research at Michigan State University, a study was undertaken to explore utilization of constructivism in teaching AP chemistry. Traditionally the acids, bases, and solution equilibria unit of AP chemistry is one of the most difficult units in the course. Students have a great deal of difficulty keeping track of various types of acid base reactions and solving equilibrium problems involving these reactions. This study focused on the constructivist approach to learning acid-base unit of AP chemistry. The core content was established according to the guidelines of the College Board for AP chemistry. The College Board is a national nonprofit membership association that prepares, inspires, and connects students to college. The Advanced Placement Program is one of its programs. The acid-base unit was taught in the second marking period of the second year chemistry course at Carman Ainsworth High School in Flint, Michigan.

Constructivist theory offers an explanation about how humans acquire knowledge. It maintains that individuals create or construct their own knowledge through the interaction of what they already know and believe and the ideas, events, and activities they encounter (Cannella & Reiff, 1994). Knowledge is acquired through involvement with content instead of imitation or repetition (Kroll & LaBoskey, 1996).

Piaget's comprehensive theory of cognitive development, describing individuals as active participants in their intellectual development, has given rise to the constructivist theory of learning (Bodner, 1986). Piaget argued that knowledge is constructed as the learner tries to tie new experiences into a preexisting mental framework, or scheme.

Research continues to demonstrate that individuals have their own strongly held personal theories about how things work and about science phenomena (Stepans, 1996). Students bring these personal beliefs that are a result of a multitude of unique experiences with them to the classroom (Colburn, 2000). However, some of these intuitively held ideas differ from those accepted by the scientific community. Teaching science involves helping students to correct their misconceptions and embrace the scientifically accepted explanations. This view contrasts with the idea that teaching is about transferring knowledge from teacher to student and learning is about absorbing information (Colburn, 2000). Moreover, studies have shown that teaching and learning are not synonymous (Bodner, 1986).

Changing students' misconceptions is a rather difficult task. It requires students to undo a mental framework of knowledge that they may have used for years in order to make sense out of the world around them. Teachers should create an atmosphere in which the student is motivated to learn. Stepans (1996) suggests students are motivated when they are provided with experiences that are contrary to their innate beliefs or experiences. This results in a conflict between what they expect to happen and what actually happens. It is the desire or need to resolve this conflict that motivates them to learn. Out of this conflict conceptual change can emerge (Stepans, 1996). One common question of constructivism is that "if individuals construct their own knowledge, how can groups of people appear to share common knowledge?" Bodner (1986) explains that knowledge or mental schemes must fit reality. The concepts, ideas, theories, and models we construct in our minds are constantly being tested as a result of our experiences. They survive only as long as they are useful. Bodner (1986) observes that discarding an old theory can be achieved by constructing a new one that does a better job of explaining the experimental evidence. The only way to replace a misconception is by constructing a new concept that more appropriately explains the new experience.

Liem (1992) argues that in the teaching of a science concept, it is essential for the teacher to arouse the students' curiosity. Unless the student wants to know what the teacher has to say, it is most likely that time and effort in trying to teach the student would have been completely wasted. He suggests that the first task of a teacher is to attract the attention of the student. The use of discrepant events in the teaching of science is one of the best methods to do just that: to arouse interest and curiosity. When an occurrence goes counter to what one thinks likely, it poses a question to the student and challenges him or her to come up with the explanation.

Bodner (1986) frames the above phenomenon in terms of cognitive psychology. He uses Piaget's explanation that cognitive structures change both qualitatively and quantitatively with increasing age and experiences. Assimilation involves applying a preexisting scheme or mental structure to interpret sensory data. Disequilibration occurs when we cannot assimilate our experiences into preexisting schemes. Constructivists argue that disequilibration plays an important role in learning. Students have to be aware that a problem exists before they are willing to accept an explanation. Equilibrium is restored by modifying these preexisting schemes until discrepancy is resolved. Colburn (2000) suggests that from the constructivists' prospective, demonstrations are very helpful when they encourage students to think about their preconceived ideas or challenge those ideas. He argues discrepant events and demonstrations that require predictions challenge learners' misconceptions. In discrepant events, students encounter outcomes that are different from those they expect to observe. Students are cognitively challenged to develop an explanation of the discrepancy. In prediction-type demos, the teacher asks students to make predictions about what will happen before the event takes place. This forces students to access their prior knowledge. Students can test their understanding by observing whether their predictions are accurate. The teacher can also ask students to explain why they thought their predictions were correct. This gives the teacher a better understanding of students' prior knowledge. Sewel (2002) also recommends presenting information with conflict is the best way to address misconceptions. She suggests using relevant new information in the form of carefully selected demonstration to achieve conceptual change.

In addition to constructivism, some educators have attempted to gain additional insights into the learning process by utilizing the growing body of knowledge in neuroscience and brain research (Brandt, 1999). They maintain "the better we understand the brain, the better we will be able to design instructions to match how it learns best" (Wolfe, 2001). Wolfe has summarized some of the latest research on the functional aspects of how the brain operates. She has attempted to provide a brain-based rationale for why some educational strategies work better than others.

Fields of cognitive psychology and education research have benefited from the new advances in neurosciences. Restoke, a neurologist, (in Brandt, 1994) describes the brain as having a decentralized organization that is composed of large number of modular elements linked together. The brain does not store information in a specific location but rather separately in visual, auditory, and motor cortices joined in networks of neurons (Wolfe, 2001). When new experiences are introduced the brain searches for an existing network into which the newly acquired material will fit. If a match occurs, the prior information gives meaning to the new material. Information storage in the brain is in the form of networks of associations. Our experiences form networks over the course of our lives. Information that fits or adds to an existing network has a much better chance of storage than information that does not. Associating or comparing new concepts with the concepts that are familiar to students makes information more meaningful for them. This can be accomplished through the use of analogies, similes, and in some cases metaphors (Wolf, 2001). For example the football analogy can be used to teach about the concept of strong versus weak acids (Silverstein, 2000). In this analogy we equate an acid, which is a proton donor, to a quarterback, who is a football "donor".

Learning is a process of building neural networks (Wolf, 2001). Various aspects of memory are stored in networks of neurons. Visually mapping the information has proven helpful because it mirrors the structure used by the brain. Graphical organizers are charts that use content vocabulary to help students anticipate concepts and their relationships to one another (Vacca & Vacca, 1996). These concepts are displayed in an arrangement of key technical terms relevant to the important concepts to be learned. Graphical organizers build a frame of reference for students as they struggle with new material. They make it visually possible for the learner to see connections between parts of information that are not noticeable in a linear format (Wolfe, 2001). Therefore, two different forms of graphic organizers were presented to students in this study. The first one, (Appendix D-1), serves as a map to follow when solving different types of acid-base problems.

National interest in improving science education dates back to 1950s when United States was perceived to be falling behind the USSR in the space race. This perception resulted in fundamental changes in American Science Curriculum (Fensham, 1994). What followed was an apparent pursuit of ever-changing approaches to teaching. It seemed as if a different approach was recommended each decade: individualized instructions, cooperative learning, discovery learning, inquiry, and others. In retrospect all of these trends were supported by the constructivist theory (Colburn, 2000).

This study focused on a constructivist approach to teaching acid-base unit of AP chemistry at Carman-Ainsworth High School. Carman-Ainsworth is a unique district because it enjoys a diverse population of approximately 1,500 students. Carman-Ainsworth Community Schools extend from inner city Flint on the east, to some of the most sophisticated suburbs near Flushing to the west and north, to some blue collar and rural suburbs on the south. This spectrum includes students from the most diverse ethnic and socioeconomic origins in the county, if not the state. In addition, students from other school districts may attend Carman-Ainsworth district as a "school of choice". This introduces an even greater diversity of ethnicity, socioeconomic background and skills and abilities into the student body. Students choose a wide range of courses from AP to vocational type classes. The high school has a 3.12% dropout rate and approximately 75% of the students attend post secondary institutions after graduation. This occurs despite the fact that 42% of students receive free or reduced lunch (Carman Ainsworth High School, 2002). The school calendar at Carman Ainsworth High School is divided into two semesters. Each semester is divided into two marking periods and each marking period is approximately nine weeks long.

Students who wish to enroll in the AP chemistry course must have completed one year of either accelerated chemistry or chemistry I. The accelerated chemistry course covers the traditional topics covered in Chemistry I with increased depth and slightly increased pace. Students are given somewhat more freedom to design and carry out experiments in this class, which is meant to prepare them for AP chemistry. They are also expected to complete more assignments outside of class. Besides the first year chemistry course, students must also have completed advanced algebra.

One obstacle to designing lessons in an AP chemistry class is time. The curriculum of an AP class should be covered before the month of May when the AP exams are administered in the US and other participating countries. There are about 25 chapters to be covered with time leftover for reviewing and taking practice exams. The calendar of our high school usually allows about 28 weeks to teach 22 major topics. Careful attention has to be given to allotting time for each chapter. This leaves about 2 weeks to teach a few challenging topics and one week to each of the other chapters. Vacations and long weekends are used by students to study topics that may not normally receive attention due to time constraints.

The AP Chemistry Teacher's guide also recommends limiting the number of major tests given to seven or eight during the course of the school year. It advises that although frequent testing helps students to be more vigilant on mastering recently covered material, it uses up valuable class time. If teachers used one class period for each of the 25 topics the result would be loss of over a month of instructional time. Also since AP classes are meant to replace a first year college class, their testing schedule should reflect this as well.

During the time of this study 15 students were enrolled in the AP chemistry class. Fourteen students participated in the study, 11 juniors and three seniors. Most of the students were also involved in extracurricular activities and a few had jobs. Many students do not realize the rigor of the course and the significant time commitment it requires. The first three chapters that involve learning the formulas and charges of ions and solving stoichiometry problems were assigned as homework over the summer vacation. This allows students to get a glimpse of the depth and rigor of the course and saves about two weeks for covering the more difficult material when school is in session.

The goal of this study was to incorporate constructivist approaches to teaching keeping in mind certain aspects of brain-based instruction into my AP chemistry class.

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IMPLEMENTATION

This study focused on exploring utilization of constructivism in teaching AP chemistry at Carman Ainsworth High School. A group of 14 students, enrolled in the AP Chemistry class during the academic year 2003-2004 were used as test subjects. The unit on acids, bases and solution equilibria was used to determine the effectiveness of the methodology. This unit extended through 24 school days in the second marking period of the first semester following the unit on equilibrium. Table 1 presents the acids and bases unit plan. The performance objectives which encompass the major concepts of acid-base chemistry are listed in Appendix A-1.

Twelve of the students had taken their first year of chemistry the previous year and two were enrolled in accelerated chemistry two years prior to the study. Of the twelve students that took chemistry the previous year, five students came from Chemistry-I class and seven from accelerated chemistry. The latter students distinguished themselves from students with chemistry-I background very early on in the course. Both groups had the same teacher in the prior chemistry course.

There were about equal numbers of high, middle, and low level students. The five students in the higher grade category benefited from a strong math background and perfect attendance. Three of them were enrolled in AP calculus class simultaneously.

During the class period in which the methodology was implemented:

- Students were engaged in activities and demonstrations to utilize their knowledge of different types of acid-base reactions to predict the products and write net ionic equations for each reaction,
- Students applied different acid-base concepts and solve a wide range of problems.

Table 1. Acids and Bases Unit Plan

New activities or modified activities are underlined.

			11/13 Pretest **Demo**: Food is usually Acid, Cleaners are Usually Basic	11/14 **Demo**: Reactions of Acids and Copper(H_2SO_4 and HNO_3) Reactions of Acids and Carbonates
11/17 **Demo**: Reactions of Acids (HCl) Conductivity tests for strong and weak acids **Notes**: Strong Acids and Bases	11/18 **Demo**: Bronsted & Lowry Definition Lewis Acids demo **Notes** and **Practice problems** Chart: pH, pOH, Kw, $[H^+]$, and $[OH^-]$	11/19 **Rxn Quiz** **Notes** and **Practice problems** on Weak Acids and Ka **Prelab discussion**: Hydrolysis Lab	11/20 **Lab**: Hydrolysis	11/21 **Quiz**: pH **Postlab discussion**: Comparing data obtained from lab. Group work on postlab
11/24 Feedback on quiz **Notes**: hydrolysis equations and Complex ion reactions	11/25 **Discussion:** finding pH of salt solutions Using Organizational Chart for finding pH	11/26 **Rxn Quiz** **Practice problems:** pH of weak acids , bases, and salts	No School Thanks giving Break	No School Thanks giving Break
12/1 **Discussion:** Titration **Prelab discussion** for Titration Curves Lab	12/2 **Lab**: Titration Curves Lab	12/3 • **Rxn Quiz** • **Postlab discussion**: Graphical Analysis Computer graphing • **Practice Problems:** pH of WA at _ way point	12/4 **Discussion:** feedback on writing net ionic equations **Practice Problems**: Titration SA&SB*	12/5 **Quiz:** Hydrolysis equations **Practice Problems:**Titration of WA&SB ** Using Titration Table
12/8 **Notes**: Buffer solutions H-H Equation **Prelab discussion**	12/9 **Lab**: Eq in Soln's of Wk Acid, Wk bases & Buffers	12/10 • Postlab group work • Students presentations	12/11 Finish presentations **Practice Problems:** Old AP Exam problems	12/12 **Quiz:** Buffers **Practice Problems:** Old AP Exams
12/15 **Practice Problems:** Review problems	12/16 **Practice Problems:** Review problems	12/17 _ day **Practice Problems:** Review problems	12/18 Unit Post test	12/19 Last day before Winter break

* Titrations of Strong Acids & Strong Bases

** Titrations of Weak Acids & Strong Bases

Pedagogical strategies used in the study focused on conceptual understanding as well as acquiring higher level thinking skills. One of the major driving forces in this study was to produce more student centered opportunities and less standard lecture format.

Several laboratory exercises (Appendix C6-C8) assisted students in building their conceptual models. These "labs" were always performed with a partner they had chosen in the beginning of the year. Because this is a college-level course, students were allowed to choose their own partners as they would in college. However, unlike a typical college lab course, a significant challenge was the lack of having three consecutive hours per week to work on labs. In the past this had resulted in limiting the number of laboratory exercises students performed for this class. This is a common practice among many other AP chemistry teachers with fifty minutes of class time. Eliminating laboratory exercises leaves students unprepared for the national AP exam where they are expected to compete with other students nationwide. Fortunately, during the time that this study took place AP chemistry was scheduled on the last hour of the day and students agreed to dedicate one day per week to labs. On lab days they would stay after school until they finished their experiments. This was accomplished by deciding on a day having the least conflicts with after-school activities. Participation was voluntary but all students participated almost all the time. This arrangement allowed students to complete experiments that would fulfill the College Board objectives and obtain a better grasp of the concepts.

Much of the time in the acid-base unit was spent on group activities. Several days were committed to setting up and performing laboratory procedures. Some students were not satisfied with their titration curves or other data and repeated parts or all of certain experiments. They also worked on analyzing their data and their presentation of the concepts in their last experiment. Teacher-performed demonstrations replaced lecturing on some topics. When lecturing was necessary it was mixed with group problem solving activities to allow time for processing and construction of new knowledge.

The assessments included a pretest, a number of quizzes, lab reports and a comprehensive posttest.

Introduction to Acid-Base Unit

The unit started with a pre-assessment (Appendix B-1) to determine the students' background knowledge related to reactions of acids and bases, definitions of acids and bases according to different theories, understanding the pH scale and calculating pH of strong versus weak acids, hydrolysis, and buffers. There were multiple choice questions as well as short essay questions. Students were encouraged to answer all questions even if they had to guess. There was also a confidence scale included where the students would circle their confidence level on the scale of 1 to 5 for each question. This would help to determine if they guessed the correct answer. Students were assured that the pre-assessment scores would not be part of their grades.

Laboratory Experiments and Demonstrations (Appendix C)

One of the key aspects of this unit was to modify the existing labs and to add new activities that would bring about conceptual change. Demonstrations were used as a time saving device and a vehicle for guided inquiry and critical thinking.

A. Demonstrations and background science:

I. A demonstration of Food is usually Acidic, Cleaners are usually Basic (Appendix C-1) followed the pretest. The title was not revealed to the students so they could determine this through their observations. Numerous test tubes of some common household products were subjected to pH tests. This activity was conducted with relevant real world applications for each pH range (see discussion included in Appendix C-1 for background information). Students used their observation skills to arrive at the conclusion that food is usually acidic to neutral and cleaners are usually basic. This was used as an informal evaluation tool. Students welcomed the information about pH and hair and asked many questions about it.

II. The second day's discussion started with demonstrations of Reactions of Acids (Appendix C-2). One segment of the AP chemistry exam requires students to predict the products of chemical reactions. Students usually have difficulty remembering the reactions and the products of each. This demonstration was presented to give them a visual experience to help them remember. They were instructed to predict the products and then write their observations. This served as an informal assessment tool. Their notes would be used to study for the reaction quizzes and the test. This demonstration shows that the reaction of concentrated nitric acid and copper does not follow the generalization that acids combine with metals to liberate hydrogen gas. The equation for the reaction is

$$4\ HNO_3\ (aq) + Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2\ NO^{3-}(aq) + 2\ NO_2\ (g) + 2\ H_2O\ (l)$$

This reaction served as a good discrepant event because the NO_2 gas produced was brown in contrast with the colorless hydrogen gas that the students expected to see. The pale blue color of the solution points to presence of Cu^{2+} (aq) ion. Next they saw the reaction of concentrated sulfuric acid and copper. The reaction would not start and the acid had to be heated in order to react with copper. The dense white fog and the spattering of the acid as water was added to it seemed quite amusing to the students and should have helped emphasize that this was another unusual reaction. They also observed that concentrated hydrochloric acid does not react with copper even if heated. These reactions showed that although acids have similar properties such as donating a proton they possess different oxidizing abilities (Shakashiri, 1989)

III. In <u>Reactions of acids with carbonates and bicarbonates</u> (Appendix C-3) demonstration a cork was placed on the test tube. When enough CO_2 gas was produced the cork was ejected across the room with noise. This emphasized the fact that a gas was produced and this was not just a simple double displacement reaction. The auditory as well as visual (Multi sensory) modes of learning enhance memory retention (Wolfe, 2001). Some students, however, were too fascinated by the events to take notes.

IV. The <u>Conductivity Test and Extent of Ionization in Aqueous Acids</u> (Appendix C-4) demonstration was used to show the difference between strong and weak acids. First, the pH of 0.1M solutions of hydrochloric acid, acetic acid, and citric acid was determined by universal indicator. Students observed that each solution, despite having the same concentration, exhibited a different indicator color and varied

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varied in their degree of acidity. These solutions were then tested for conductivity. The more acidic a solution was the brighter the bulb glowed. Because conductivity depends on the concentration of the ions in the solution, the brightness of the bulb and acidity must be related to the extent of ionization of each acid.

The second part of the demonstration revealed that although the acetic acid solution used had a concentration of 2M, its extent of dissociation was the same as a 0.005M solution of a strong acid, HCl. This demonstration was a powerful tool in teaching the difference between strong and weak acids. Many learners confuse concentration of the acid with concentration of the hydrogen ion. It replaced lecturing with guided inquiry.

V. The teacher then held up a bottle each of 1M hydrochloric and 1M acetic acid solutions and asked the students to draw the species inside each. This served as an informal assessment tool.

VI. A short discrepant event was used to introduce acid-base theories. According to the Arrhenius definition of acids and bases, a base increases concentration of hydroxide ions (OH^-) and an acid increases the concentration of H^+ ions in water. A bottle of concentrated ammonia (NH_3) was opened. Students observed when a red litmus paper was exposed to gaseous ammonia that exists above the bottle it turned blue. They also observed that blue litmus paper exposed to HCl gas above concentrated hydrochloric acid turned red. They were guided to think that despite the fact that neither gaseous NH_3 nor gaseous HCl were in water, they showed basic and acidic properties respectively. Therefore, the Arrhenius model does not

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include acids and bases that are not in water. Hence, a more inclusive definition for acids and bases was needed. They were asked whether the Bronsted and Lowry definition would include the above examples.

VII. Another discrepant event was presented in Lewis Acid-Base demonstration (Appendix C-5). Aluminum chloride ($AlCl_3$) added to dichloromethane turns the methyl violet indicator yellow signifying a very acidic solution. Students were directed to refer to the pH range for color changes of common indicators chart in their book. This reaction does not involve hydrogen ions. Therefore, it does not fall in either Bronsted and Lowry or Arrhenius acid-base reactions. They had to look up the Lewis acid base definition and discuss with a classmate to understand why aluminum chloride is acidic. Then we had a class discussion about the demo.

B. Laboratory Experiments:

I. After students were introduced to weak acids they were ready to perform a discovery lab activity. They learned about hydrolysis through performing the Hydrolysis (Appendix C-6) lab, designed as a microscale lab to guide students to classify salts as acidic basic or neutral. New salt solutions were prepared for best results. In order to eliminate chemical waste, students put drops of each solution on pH paper. Microscaling is beneficial because it uses less material, labs are performed faster and less waste is produced. The questions at the end of the lab handout helped students classify the ions. The next day each lab group shared their findings to come up with a uniform set of class data. I facilitated the students' classification efforts and helped them see the patterns evident in their data tabl

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data tables. One of the questions on the lab asked students to write net ionic equations for hydrolyzed salts. This presented a perfect opportunity to teach complex ions and their reactions in water. The lab questions also guided students into finding definitions of acid and base anhydrides. And finally a question related the complex ions to the everyday life situation of using Al^{3+} in underarm anti-perspirants. Each aluminum ion can form a complex ion with 6 molecules of water to keep underarms dry and the acidity keeps the pH lower than the odor causing bacteria would prefer. Students got a great deal of mileage from this short and simple lab.

II. The next lab Titrations of Strong Acids and Bases (Appendix C-7) was chosen to teach students how to use basic laboratory equipment, such as pH meters, and magnetic stirrers. They also used of a computer graphing program to draw their titration curves. They were also introduced to drawing 1st and 2nd derivative curves to find the equivalence point. The students were given a chance to copy the Graphical Analysis program to their home computers (our school had bought a license that allowed students to take a copy home for analysis). A pH titration allows the experimenter to get a better understanding of the titration process and provides a more accurate "end-point" than can be obtained using only an indicator. The titration lab provided a hands-on opportunity for students to experiment with the problems they had been working on paper.

III. Use of time is of great concern in AP chemistry lab. Spending too much time on one subject prevents adequate coverage of other important topics. In order to adhere to the schedule, the last lab Equilibrium in Solutions of Weak Acids and

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Weak Bases-Buffer Solutions (Appendix C-8) was designed to accomplish several objectives at once. A group of labs that I learned about at various AP chemistry workshops were combined together and provided a guided inquiry opportunity for students. On the following day students worked in their lab groups to answer the guided inquiry questions, with the teacher playing the role of the facilitator. Groups of students then picked one question to explain their results to class. Data Table I of Equilibrium in Solutions of Weak Acids and Weak Bases-Buffer Solutions (Appendix C-8) provides evidence that weak acids and weak bases are equilibrium systems that are affected by a common-ion , Data Table II compares the ability of weak acids to form buffer solutions as opposed to strong acids which cannot do the same. It is easy to see that weak acid solutions are equilibrium systems, where as strong acids are not. Data Table III and IV demonstrate how a buffer acts to resist change in pH and the relationship between the common ion effect and buffers.

Analogy

The difference between strong and weak acids was highlighted through the use of a football analogy. The difference between a strong acid and a weak acid is similar to the difference between an excellent quarterback and an awful one. A good quarterback delivers the ball efficiently, which resembles a strong acid donating a proton. On the other hand, a bad quarterback delivers the ball indecisively and tends to hold on to the ball, similar to a weak acid not readily donating its proton. A similar analogy may be drawn between a base and a wide receiver. Just as a base is a proton acceptor, a wide receiver is a foo

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receiver is a football "acceptor". A good wide receiver catches the ball and holds onto it similar to the way a strong base accepts a proton. In contrast, a bad wide receiver tends to drop the ball, comparable to the way a weak base does not readily bond to a proton (Silverstein, 2000).

Graphic Organizers

Graphic organizers make it possible for students to visually observe connections between parts of the information that are not apparent in a linear format (Wolfe, 2001). Often students know the formulas and individual steps but they fail to recognize the connections between them. Two types of graphic organizers were used: flow charts and tables. The Flow chart in Appendix D-1 guides students through the steps of finding pH and related problems and to observe the relationship of each concept to another.

Tables provide a framework for organizing main ideas and subtopics as students encounter new information. The table in Appendix D-2 is organized to help students observe the relationship between pH, pOH, Kw, $[H^+]$, and $[OH^-]$. It also allows them to see the patterns of change for each column. Appendix D-3 shows the steps involved in solving different types of titration problems. This serves as a scaffold which helps learners to do what they cannot do at first. It provides the necessary support that students need as they attempt new tasks (Vacca & Vacca, 1996). The table format allows students to observe the common steps for each type of titration problem. These tools give students confidence to solve the related problems. However, they do not need to rely on them as much after some practice.

Practice problems (Appendix D)

To fulfill the College Board goal of "developing student's ability to think clearly and to express their ideas with clarity and logic" the thinking skills suggested by Blooms Taxonomy were incorporated. Benjamin Bloom and a group of educational psychologists identified different levels of cognitive complexity (Krathwohl, 2002). According to Bloom's Taxonomy there are three domains of learning. The cognitive domain emphasizes intellectual outcomes. Bloom and his colleagues classified levels of learning into six categories, three concrete thinking skills and three critical thinking skills:

- Concrete Thinking Skills are: knowledge, comprehension, and application
- Critical thinking skills include: analysis, synthesis, and evaluation

This taxonomy is a tool to gauge the cognitive challenge that individual assignments offer students (Raths, 2002). Teachers should make sure that all levels of the cognitive process are used and that students learn the different types of knowledge listed above (Walberg & Haertel, 1997). Close-ended and factual questions might be appropriate when teachers are assessing prior knowledge or reviewing what students just learned. However, the major instructional goals can be accomplished by emphasizing on open-ended questions that require students to apply, analyze, synthesize, or evaluate what they are learning. These skills can be used to raise the learning target of a particular unit. The recommendations set forth by Bloom's taxonomy were observed in choosing and assigning practice problems for the acid-base unit. Various types of practice problems, The use of

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The use of critical thinking (analysis, synthesis, and evaluation) is one of the most valuable skills we can bestow on our students. The practice problems dealing with titration of weak acid - strong base or adding either acid or base to buffers use the highest levels of thinking on Bloom's taxonomy. Appendix D-6 for example contains synthesis problems. These problems were given to students after they had hands-on experience from a lab titrating a weak acid and a strong base. In order to solve these problems the students had to go through the following steps:

1. find out how many moles of acid and base are present (conversion to moles)
2. how many moles of acid and base are left after the titration (stoichiometry)
3. set up equilibrium tables for the weak acid to find $[H^+]$
4. find pH using the formula

Students had to put together several concepts to form a whole. Multiple pieces of information were needed to solve the problem. Another practice sheet asked the students to choose the best buffer solution from a list of acid and salt mixtures. This requires the learner to make a judgment based on what they knew about buffers.

<u>Assessments</u>

Combinations of formal and informal assessments were used to evaluate students' progress in the unit. A lab rubric (Appendix C-9) was developed to evaluate the hands-on learning. Several quizzes (Appendices B-2 through B-7) served as quick checks as well as mechanisms to identify areas that needed reinforcements. The Posttest (Appendix B-8) covered many topics including some of the concepts from the pretest. The AP chemistry Teacher's Guide recommends that the students need practice with taking timed tests. The

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tests. The teacher should deliberately construct tests that put the students in a time pressure situation without the use of calculators on multiple choice questions. To comply with this recommendation, the posttest had more questions than students typically could answer in the allotted time.

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5. L
6. U
7. D spec
8 D in a c
9. Id
10 E
11. C

* N =

RESULTS AND EVALUATIONS

The constructivist and brain-based methods of instruction utilized in this study produced results that were consistent with its goals. The evaluation tools included a pretest and a posttest. Table 2 presents the results of the pretest and posttest for questions included in the analysis. The table contains both the number of students with correct answers to the 11 common questions as well as this number as a percent of total. The results indicate that student' understanding of the acid-base theories, preparing buffer solutions, hydrolysis, and calculating pH of weak acids improved. All three levels of student (high, middle, and low) scores improved from pretest to posttest.

Table 2. Number of Correct Responses*

	Pretest		Posttest	
Question	# of correct answers	% of correct answers	# of correct answers	% of correct answers
1. Identifying conjugate acid-base pairs	4	28	10	71
2. Bronsted-lowry definition of acids and bases	6	43	13	93
3. knowing the difference between strong and weak acids	10	71	14	100
4. Preparing a buffer	0	0	7	50
5. Lewis acid-base theory	6	43	10	71
6. Using Kw = [H+] [OH-]	6	43	11	79
7. Drawing a picture to represent all the species of strong and weak acids	0	0	9	64
8. Drawing a picture to represent all species in a concentrated and dilute acid	1	0.1	10	71
9. Identifying ions that hydrolyze	2	14	10	71
10. Effect of buffers on pH	3	36	7	50
11. Calculating pH of weak acids	0	0	11	79

* N = 14

Figure 1 compares the number of students who correctly answered each question on each of the tests.

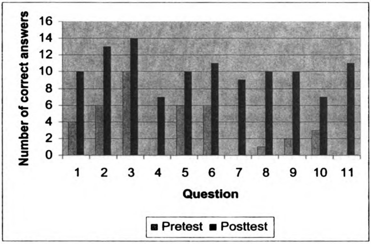


Figure 1. Comparison of Pretest and Posttest Grades

In order to determine if the pretest and posttest scores of the acid-base unit of AP chemistry were statistically different a t-test was used to determine whether or not the pretest and posttest scores were significantly different.

The test scores were normalized by converting them to percentage values to compensate for different grading scales. The percentage values were then entered into a statistical software (http://www.physics.csbsju.edu/) containing a paired t-test. The paired observations in the t-test consisted of each student's pretest and posttest scores. The level of significance was set at 0.01. The observed t-value was -12.4 ($n = 14$) and the level of significance (p-value) was 0.000. The null hypothesis can therefore be rejected and we can conclude that the pretest and posttest scores of the acid-base unit were significantly different. That is, students' performance on acid-base unit improved from pretest to posttest.

The above results appear to be in line with anticipated influences of the constructivist approach utilized in the unit. For instance, results of question number four involving preparing a buffer appeared to benefit from completing Equilibrium in Solutions of Weak Acids and Weak Bases-Buffer Solutions lab. Results of question number seven, drawing a picture to represent all species of strong and weak acids, appeared to have improved due to students observing the Conductivity Test and Extent of Ionization in Aqueous Acids demonstration. This demonstration probably helped the results of question number eight, drawing pictures to represent the species in a concentrated and dilute acid solution. Also, skills required to answer question number 11, calculating the pH of weak acids, were practiced during the guided inquiry chart related to pH.

The laboratory assessments showed that 78% of the students were able to group salts correctly in the Hydrolysis lab. They all successfully wrote the net ionic equations for salts and complex ions. Some students had difficulty reflecting on their experiences doing this lab. Some of the comments in the reflection part of the lab were as follows:

- "This lab taught me exactly what hydrolysis was.....My knowledge of complex ions and ligands was also built upon. I was able to spot periodic trends."
- "The only thing that would make this lab better would be a different way to find the pH".
- "Doing the lab helped to solidify the writing of net ionic equations for complex ions".
- "We had to determine how the aluminum ion, found in many anti-perspirants, works to keep the underarm dry...making the many uses of hydrolysis in our everyday lives apparent".

The assessment for the Titration Curve Lab was based on whether the students learned the intended skills of calibrating and using the pH meter, plotting the titration curves using their data, and finding the equivalence point on each type of graph. All of the students performed these tasks satisfactorily. They all had performed titrations in their first year of chemistry with indicators. Some students were very meticulous and were recording every change in pH in their data table. 71% found the unknown concentration with less than 10% error. The rest of the students had an error of less than 20%.

Equilibrium in Solutions of Weak Acids and Weak Bases-Buffer Solutions was a qualitative experiment. Unlike the Titration Curve Lab, students were not under pressure to determine precise quantitative results. An additional assessment component was added. Each lab group had to use its data, then answer one of the questions at the end of the lab handout, and make a presentation to the class. Students were given part of the class period to work on their presentations. They clarified their questions with one another or asked the teacher. All groups did well on their presentations and gave sufficient background information. All the students that turned in their lab reports expressed an increased understanding of buffers and common-ion effect after performing this lab. The following are some of the comments they wrote in their lab reflections:

- "Doing this lab gave me a deeper understanding of how buffers function."
- "Although I had an idea what a buffer was and it's effects, this experiment clarified the meaning of a buffer."

- "I learned from this lab how the common-ion effect works, and why does it behave the way it does."
- "It is only because weak acids and weak bases exist in equilibrium systems that the common ion effect has any effect on them."

Other comments were:

- "...I had learned a lot in just one experiment",
- "I thoroughly enjoyed performing this lab and believe that I was able to walk away with a better understanding of the common-ion effect and buffer solutions.", and
- "This experiment works very well. It was qualitative not quantitative so no calculation was necessary, but it demonstrated, so well, what it was meant to. I think this lab was set up very nicely ..."

The practice problem handouts were not graded. Students worked on them in class and could finish them at home. Someone would volunteer to solve the problem or I would write the solution on the board the next day. As the problems became more complex students needed more scaffolding. The concept map of solving acid-base problems appeared to be helpful to most students as it was often used by them. Similarly the Titration Problems table (Appendix D-3) showing the steps involved in each type of titration seemed to help students when they had just started solving titration problems.

CONCLUSIONS

The high school AP Chemistry course is designed to be the equivalent of the general chemistry course usually taken during the first year of college. High school students completing this course should gain a depth of understanding of fundamentals and a reasonable competence in dealing with chemical problems. The acids, bases, and solution equilibria unit is one of the most challenging units in the course. Most students experience a great deal of difficulty keeping track of various types of acid base reactions and solving equilibrium problems involving these reactions.

This study involved using the constructivist approach together with other relevant pedagogical techniques to assist the students with learning the concepts in acid-base unit of AP chemistry. The core content of the unit was established according to the guidelines of the College Board for AP chemistry. The goal was to create content that would contribute to the development of the student's abilities to think clearly and to express their ideas, orally and in writing, with clarity and logic.

In order to facilitate students' learning demos, labs, practice problems, and graphical organizers were utilized in teaching the stated unit. Based on classroom observations students seemed to prefer active learning over lecturing alone. The pretest and posttest results revealed the constructivists approach helped students gain a better understanding of fundamentals of the acid-base unit. Also it was observed that the brain-based educational strategies gave them confidence in dealing with solving chemistry problems. The higher order thinking skills suggested by Bloom's Taxonomy contributed to students' ability to think clearly and to express their ideas with some degree of clarity and logic. This was evident from the students' comments and how they collaborated on solving problems. Most students made the connections and learned how to solve a good number of acid-base problems. With the exception of one of the top level students who preferred traditional lectures, other students thought that remembering the demonstrations helped them answer questions on the test or quizzes.

It is important to note that although using constructivism and other pedagogical teaching strategies generally results in facilitating student learning, they are not sufficient to compensate for a student's poor performance caused by failure to complete the required readings and practice problems. This is especially true in an advanced class such as AP chemistry where it is virtually impossible to only rely on information presented in the classroom. As was observed in this study, those students who either had other major commitments or were not completing the assigned work did not seem to benefit fully from the applied teaching strategies.

As discussed in the Results and Evaluations section, presenting the demonstrations was a valuable tool in teaching the course. Demonstrations served as discrepant events that motivated students to explore why something behaved differently than they had expected. However, the demonstrations were planned too close to each other. Several consecutive days of demonstrations appeared to have the students not focus as much on the required reading assignments. In the future these demonstrations will be spread strategically throughout the unit.

Another area of potential improvement is the use of an indicator to determine the pH of each salt solution in the Hydrolysis lab. Based on students' feedback, apparently it was difficult to discriminate between pH 5 and 6 since both produce orange color on the pH paper. Universal indicator can be used to determine the pH in this lab. To keep the lab at microscale level we will use 6x4 well plates to put sufficient drops of universal indicator into each well to show the green color. A few drops of each salt solution can be added to each well to observe the color changes. pH paper can be used as an additional means of determining the pH. This would still keep the lab at microscale level yet improve the ability of students to distinguish weak acidic salt solutions with pH 5 from neutral salt solutions.

Constructivism recommends the use of inquiry projects. Inquiry involves a process for constructing understanding of nature (Chiappetta & Adams, 2004). Students need to engage in active learning in order for them to reconstruct scientific knowledge and to find meaning in it. This involves a lengthy process that is difficult to fit into the tight schedule of AP chemistry class. However, there is time to work on an inquiry project after the AP Examinations are administered in the middle of May. Students' enthusiasm about the effect of pH on hair has compelled me to design a related inquiry project. Students will be required to design and carry out an investigation to test the best pH for hair by making solutions with different pH values and testing the effects of each on hair. Alkaline solutions cause split ends and rough texture in hair. Solutions with high pH dissolve hair. Human test hair samples obtained from hair salons or other sources can be viewed under a magnifying device such as stereoscopes, magnifying lenses, or even low resolution microscopes to determine their condition. Students will bring their brands of shampoo to be tested for pH. They can then determine whether the shampoo they are using is within the acceptable pH range for healthy hair. They may alternatively design a study to determine the best pH range for other household products.

After the AP Exam in mid May, students can engage in an inquiry investigation without the pressure of the exam looming. This investigation will be graded in order for students to take it more seriously.

Constructivism appears to have benefited students faced with a difficult subject such as acid-base unit of AP chemistry. Lessons learned in this study will be of benefit in designing lesson plans for other units of AP chemistry. Future studies will be needed to determine if this teaching strategy can be of benefit in other challenging subjects.

APPENDIX A

OUTLINES

APPENDIX A-1

Chapter 16: Acid-Base Equilibria
Chapter 17: Additional Aspects of Aqueous Equilibria

College Board Performance Objectives:

- List general properties that characterize acidic and basic solutions and the ions responsible.
- Understand the Brönsted-Lowry Theory and be able to identify conjugate acids and bases.
- Explain the autoionization of water and write the K_w expression.
- Define pH and be able to interconvert between $[H^+]$, $[OH^-]$, pH, and pOH.
- Understand what is meant by strength of an acid or a base.
- Given the acid concentration, be able to interconvert between K_a and pH. Given the base concentration, be able to interconvert between K_b and pH.
- Calculate the percent ionization from the K_a or the K_b, and vice versa.
- Understand the relationship between the strength of an acid and the strength of its conjugate base; interconvert between K_a and K_b.
- Predict whether the solution of a particular salt will be acidic, basic, or neutral.
- Define an acid and a base in the Lewis sense.
- Calculate the concentration of each species in a solution formed by mixing an acid and a base.
- Describe how a buffer solution works and how one can be made at a particular pH.
- Calculate the change in pH of a buffer upon the addition of a strong acid or a strong base.
- Distinguish between the various titration curves.
- Calculate the pH at any point in an acid-base titration.
- Calculate the effect of a common ion on the solubility of a slightly soluble salt.

College Board Lab Objectives:

- Learn the operation of a pH meter and how to use it to calculate the K_a of a weak acid.
- Understand the concept of hydrolysis and the behavior of buffer solutions; observe the use of acid base indicators.

APPENDIX B

STUDENT EVALUATIONS

APPENDIX B-1

Acid Base Unit Pre assessment

Answer each question to the best of your ability and circle your level of confidence in each response given, "1" being least confident and "5" being the most confident.

1. Complete the following reactions:
   a. HNO_3 (concentrated)+ Cu →
   b. HCl(aq) + Cu(s) →
   c. $BF_3 + NH_3$ →

1 2 3 4 5

2. For the following reactions, identify the conjugate acid-base pair.

CH3COOH(aq) + H2O(l) ⇌ $H3O^+$(aq) + CH3COO- (aq)

Conjugate acid________

Conjugate base________

HCO-(aq) + H2O(l) ⇌ H2CO3(aq) + OH-(aq)

Conjugate acid________

Conjugate base________

1 2 3 4 5

3. Bronsted-Lowry defines a base as a(n)
   a. Proton donor
   b. Proton acceptor
   c. Electron donor
   c. Electron acceptor
   d. Ionic crystal

1 2 3 4 5

4. Which pH is that of a strong acid?
   a. 5 b. 2 c. 12 d. 8

1 2 3 4 5

5. What is the essential difference between a 1.0M solution of a weak acid and a 1.0M solution of a strong acid?
   a. The weak acid is more dilute.
   b. The weak acid does not turn litmus red.
   c. The weak acid does not conduct electricity.
   d. The strong acid is a better oxidizing agent.
   e. The strong acid produces more hydronium ion.

1 2 3 4 5

6. People who perform physiological experiments have to make sure the system is buffered what is a buffer and what purpose does it serve?

1 2 3 4 5

7. Given a solution of sodium acetate, what additional solutions are needed to prepare a buffer?

1 2 3 4 5

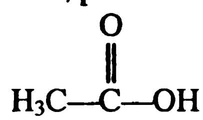
8. How is the titration curve of a strong acid and a strong base different from the titration curve of a weak acid and a strong base?

1 2 3 4 5

9. (pH scale) A solution with pH=5 is 100 times more acidic than a solution with a pH=?
   a. 7
   b. 3
   c. 0.05

1 2 3 4 5

10. (Weak acids) In acetic acid, pictured below, where will the H^+ come from?



   a. O-H
   b. C-H

1 2 3 4 5

11. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$; pKa=5.

   To make more CH_3COOH, add
   a. NaOH
   b. HCl
   To make more CH_3COO^-, add
   c. NaOH
   d. HCl

1 2 3 4 5

12. (pH scale): Predict a pH value for 0.1 M NaOH.
   a. 1
   b. 7
   c. 13

1 2 3 4 5

13. In the following reaction, which species is acting as an acid?
   $HPO_4^{2-} + NH_4^+ \rightleftharpoons H_2PO_4^- + NH_3$
   a. HPO_4^{2-}
   b. NH_4^+

1 2 3 4 5

14. Which of the following will act as a Lewis Acid?
   a. PH_3
   b. $AlCl_3$
   c. HCN

1 2 3 4 5

15. What concentration of HNO_3 must be added to water to make $[OH^-] = 10^{-9}$ M?
   a. 10^{-9} M
   b. 10^{-5} M
   c. 10^{-1} M

1 2 3 4 5

16. Draw a picture to represent all of the species in a 1 M solution of HCl and a 1 M solution of HF. What are the differences?

1 2 3 4 5

17. Draw a picture to represent all of the species present in a 1 M solution of HCl and a 0.5 M solution of HCl.

1 2 3 4 5

18. When dissolved in water, which of the following salts will produce a basic solution?
    a. KF
    b. KNO_3
    c. NH_4NO_3

1 2 3 4 5

19. The pH of blood is maintained by the bicarbonate buffer system through the amount of CO_2 in the blood. How will the pH change with hyperventilation, during which the concentration of CO_2 (aq) decreases?
    a. Increase
    b. Decrease
    c. No change

1 2 3 4 5

20. Do all salts behave the same when dissolved in water? Explain.

1 2 3 4 5

21. Which of the following salts does not undergo hydrolysis?

NaCl Na_2S $NaC_2O_2H_3$ $AlCl_3$

1 2 3 4 5

22. What is the pH of a 2.0 molar solution of acetic acid, $HC_2O_2H_3$?
Ka acetic acid= 1.8×10^{-5}

1 2 3 4 5

23. What is the pH of a 2.0 molar solution of HCl?

1 2 3 4 5

APPENDIX B-2

Name__________

Reaction Quiz 1

Write the equation for the following reactions. Include all steps that apply.

1. Concentrated sulfuric acid and copper metal

2. Sodium hydrogen carbonate and hydrochloric acid

3. Copper (II) nitrate in water

Name________

pH of Acids & Bases

1) The $[H_3O^+]$ of a solution is $2.1*10^{-3}$ M. Calculate the pH, $[OH^-]$, and pOH

2) Calculate the pH of 0.200 M $HC_2H_3O_2$ ($k_a = 1.8 *10^{-5}$)

$$HC_2H_3O_2 + H_2O \Longleftrightarrow H_3O^+ + C_2H_3O_2^-$$

3) $HNO_2 + H_2O \Longleftrightarrow NO_2 + H_3O^+$
   a) Identify the conjugate acid
   b) Identify the base

APPENDIX B-4

Name___________

Reaction Quiz 2

Write the net ionic equation for the following reactions:

1. Concentrated nitric acid and solid copper

2. Potassium carbonate and hydrochloric acid

3. Ammonium nitrate in water

4. Iron (III) chloride in water

APPENDIX B-5

Name________

Reaction Quiz 3

Write the equation for the following reactions. Include all steps that apply.

1. Hot concentrated sulfuric acid and copper metal
2. Potassium sulfide in water
3. Sodium hydride in water
4. Zinc nitrate in water

APPENDIX B-6

Hydrolysis Quiz

Name___________

1. Which of the following salts undergo hydrolysis? Explain why or why not, state the generalizations.
2. For the salts that undergo hydrolysis write an equation to show the reaction.

   a. KCl
      1.
      2.
   b. $NaC_2H_3O_2$
      1.
      2.
   c. Li_2S
      1.
      2.
   d. $Cu(NO_3)_2$
      1.
      2.
   e. CO_2
      1.
      2.
   f. $Ca_3(PO_4)_2$
      1.
      2.

3. What is the pH of a 2.0 M solution of the basic salt KF.
   (K_a for HF K_a= $6.9*10^{-4}$)

APPENDIX B-7

Name__________

Buffer Quiz a

1. Calculate the pH of a buffer solution prepared by adding 0.200 mol of acetic acid, $HC_2H_3O_2$, and 0.200 mol of sodium acetate to one liter (Hint: write the equation for dissociation of acetic acid). Ka for acetic acid = $1.8*10^{-5}$

2. Calculate the pH of the above buffer after addition of 0.020 mol NaOH (Hint: write the equation for the reaction).

APPENDIX B-8

AP Chemistry

Acids and Bases - Chapters 16-17
Part I

Name ______________________

No calculator is to be used with the multiple choice section

Fill in the oval of the best choice on your answer sheet

1. The conjugate base of the NH_4^+ ion is:

   a. NH_3 c. HOH

   b. OH^- d. H_3O^+

2. If the pOH of a solution is 3.00, the $[H^+]$ will be:

   a. 1×10^{-3}

   b. 3.00×10^{-3}

   c. 3×10^{-11}

   d. 1.00×10^{-11}

3. One liter of 1.00 M NaOH will completely neutralize one liter of:

   a. .5M H_2SO_4

   b. .03M H_3PO_4

   c. .1M HCl

   d. .5M $Ca(OH)_2$

4. If .10 mol of $NaNO_{2(s)}$ is added to 1.0 liters of .3M HNO_2, [K_a $HNO_2 = 4.6 \times 10^{-4}$], the pH of the final solution is about

   a. 3.4

   b. 2.9

   c. 5.7

   d. 7.2

5. An aqueous solution was prepared by mixing the same number of moles of acetic acid and acetate ions. If hydroxide ions are added to the solution which of the following is <u>false</u>.

   a. pH will increase

   b. more acetic acid will form

   c. the hydrogen ion concentration will decrease

   d. the acetate ion concentration will increase

6. The distinction between strong acids and weak acids is most closely related to:

   a. the number of moles present per liter

   b. the pH of the solutions

   c. the extent of the dissociation of the electrolytes

   d. interionic attractions

7. Which of the following is most closely related to the distinction between concentrated and dilute acid solutions?

   a. the number of moles present per liter

   b. pH of the solutions

   c. extent of the dissociation

   d. interionic attractions

8. Which of the following is an Arrhenius acid?
   a. HNO_3
   b. HOH
   c. C_6H_6
   d. $NaC_2H_3O_2$

9. What is the pH of a solution made by adding 400 ml of distilled water to 100 ml of .050 M HNO_3? Assume volumes are additive.
   a. 2.0
   b. 2.30
   c. 2.70
   d 3.0
   e. 1.7

10. If the dissociation of water is **endothermic** and some room temperature water is heated to boiling, which of the following is true?
   a. The pH will be 7, the solution will be acidic.
   b. The pH will be greater than 7, but the solution will be neutral.
   c. The pH will be less than 7, but the solution will be neutral.
   d. The pH will be less than 7, the pOH will be less than 7, and the solution will be acidic.

11. Which type of bond will be formed when the hydrogen ion reacts with ammonia?
   a. non polar, covalent
   b. metallic
   c. polar covalent
   d. coordinate covalent

12. A .20 M solution of a monoprotic acid has a hydrogen ion concentration of 2.5×10^{-4}. Ka for the acid is:
   a. very, very large, but not infinity
   b. 3.1×10^{-7}
   c. 2.5×10^{-4}
   d. 6.3×10^{-8}

13. A 20.0 ml sample of a weak acid, HX is titrated to the endpoint and requires 50 ml of a .050 M NaOH sol'n. After the addition of the first 30.0 ml of NaOH , the sol'n's pH is = 5.00. What is the dissociation constant, K_a, for this weak acid, HX?
   a. 1.5×10^{-5}
   b. 6.7×10^{-6}
   c. 3.0×10^{-6}
   d. 2.0×10^{-6}

14. Which statement is true?
    a. The outer limits of the pH scale are 0 and 14.
    b. The weaker the acid, the more strongly its salts hydrolyze.
    c. The pH of a neutral solution is always 7.
    d. A solution of HCl can be made basic if enough water is added.

15. If the K_b for $NH_3 = 2 \times 10^{-5}$, K_a for the NH_4^+ is:
    a. 2×10^{-5}
    b. 5×10^{-10}
    c. 5×10^{-9}
    d. 2×10^{-9}

16. Consider the species Cu^{+2}, F^-, HOH, NH_4^+. Which will result in **acidic** solutions in water?
    a. F^- and NH_4^+
    b. Cu^{2+} and NH_4^+
    c. Cu^{2+} and F^-
    d. all four

17. The gas whose water solution will be acidic is:
    a. N_2
    b. SO_3
    c. NH_3
    d. all of the above

18.

1.	Lewis	I.	electron pair donor, acceptor
2.	Bronsted Lowry	II.	containing H^+ or OH^-
3.	General Solvent	III.	increasing H^+ or OH^-
4.	Arrhenius	IV.	H^+ acceptor or donor

Which of the following sequences of definitions matches the acid-base theories one through four in order listed?

    a. I, IV, III, II
    b. I, III, IV, II
    c. III, IV, I, II
    d. III, IV, II, I

19. The acid-base system which describes the neutralization as the formation of a conjugate acid and a conjugate base is:
    a. Arrhenius   C. Lewis
    b. Bronsted-Lowry   D. Brown and Lemay

20. Which of the following behaviors is **not** exhibited by bases?
    a. electrical conductors in aqueous sol'ns.
    b. react with metals to liberate hydrogen
    c. feel slippery, taste bitter
    d. all of the above

21. Which of the following is polyprotic?
    I. HNO_3
    II. H_2CO_3
    III. H_3PO_4
    IV. $HC_2H_3O_2$

    a. I, II, III only
    b. II, III, only
    c. II, III, IV only
    d. all of the above

22. " ______1______ freed us from water, _____2_____ freed us from hydrogen."
    The acid- base theories which would correspond to 1 and 2 are:
    a. Bronsted, Lewis
    b. Arrhenius Bronsted
    c. Bronsted, Lowry
    d. Lewis, Bronsted

23. How many mL of .10 M H_2SO_4 must be added to exactly neutralize 100 ml of 0.20 M NaOH?
    a. 50ml b. 100ml c. 200ml d. 400 ml

24. Cu^{2+}, Na^+, N_2, NO_3^-, S^{2-}
    Which of the above would be expected to **hydrolyze?**
    a. Na^+, S^{2-}
    b. Cu^{2+}, NO_3^-, S^{2-}
    c. Na^+, N_2, NO_3^-
    d. Cu^{2+}, S^{2-}

25. The acid base theory which is used to describe the equation below is:

    $$AlCl_3 + Cl^- \rightarrow AlCl_4^-$$

    a. Behavioral Theory
    b. Arrhenius
    c. Bronsted-Lowry
    d. Lewis

26 Which of the following would produce an **acidic** solution in water?

I. Zn^{2+} II. CO_2 III. O_2 IV. CH_3COO^- V. NH_3

a. III, IV, V only
b. I, II, III only
c. I, II only
d. I only

27 Amphoprotic substances can

a. accept hydronium ions or hydroxide ions
b. act as a proton donor more than once
c. can neutralize acids and bases
d. dissolve in water to form acid or base anhydrides

28. Indicators are

a. dyes which can donate and/or accept protons with a color change.
b. usually weak bases
c. usually weak acids
d. substances with a wide range of color changes for each indicator.

29. The progressive addition of base to an acid is called a:

a. titration
b. end-point
c. pH reading
d. 1st derivative determination

30. In the reaction between gallium fluoride and phosphorous trifluoride, the electron pair acceptor is:

a. gallium fluoride b. phosphorous trifluoride

31. The term, adduct, refers to:

a. Arrhenius neutralization product
b. Bronsted-Lowry neutralization product
c. Lewis neutralization product
d. Solvent system neutralization product

32 The product of Al^{3+} and HOH is:

a. $Al(H_2O)_3{}^{3+}$
b. $Al^{3+} + 3OH^-$
c $Al(H_2O)_6{}^{3+} + H_3O^+$
d. $Al(H_2O)_5OH^{2+} + H_3O^+$

33. In the reaction $HSO_4^- + HS^-$, using your wide chemistry background, the predicted products would be:

a. $H_2SO_4 + S^{2-}$ c. $SO_4{}^{2-} + S^{2-} + 2H_3O^+$
b. $H_2S + SO_4{}^{2-}$ d. $S^{2-}, SO_4{}^{2-}$

Part II Equations

(No calculator is to be used with Part II)

1. Write balanced equations for the hydrolysis of the following:

   a. CH_3COO^-

   b. NH_4^+

   c. $Cr(H_2O)_6^{3+}$

   d. O^{-2} (aq)

   e. $Zn(H_2O)_4^{2+}$

   f. HCO_3^-

   g. Fe^{2+}

2. Using appropriate molecular and ionic species, write balanced net ionic equations for the following reactions.

   a. equimolar amounts of S^{2-} and HCO_3^-

   b. SO_3 gas is bubbled through H_2O.

   c. NH_4^+(aq) and bicarbonate ions (aq)

   d. Gaseous ammonia with a gaseous boron hydride.

   e. $FeCl_3$ solution is added to a solution of potassium sulfide.

24. Draw a picture to represent all of the species in a 1 M solution of HCl and a 1 M solution of HF. What are the differences?

25. Draw a picture to represent all of the species present in a 1 M solution of HCl and a 0.5 M solution of HCl.

After you have completed Parts I and II, turn in your test booklet and your answer sheets and pick up part III

Part III

A calculator MAY be used for part III

name________________

1) Calculate a) the pH and b) the percent hydrolysis on a 0.20 M $NaCH_3COO$ solution.
   Ka of $CH_3COOH = 1.8 \times 10^{-5}$.

2) A buffer solution is prepared using 400.0 mL of 0.100 M acetic acid, CH_3COOH, and 3.28 g of the soluble salt, $NaCH_3COO$. Ka of $CH_3COOH = 1.8 \times 10^{-5}$.
   a) Calculate the pH of the resulting buffer solution.
   b) 8.0 mL of a 0.5 M NAOH are added to the buffer. What is the new pH value of the buffer?

3) $$NH_3\ (aq) + H_2O\ (l) <=> NH_4^+\ (aq) + OH^-\ (aq)$$
   In aqueous solutions, ammonia reacts as represented above. In 0.0180 M NH_3 (aq) at 25°C, the hydroxide ion concentration, $[OH^-]$, is 5.60×10^{-4} M. in answering the following assume the temperature is constant at 25°C and the volumes are additive.

   a) Write the equilibrium constant expression for the reaction represented above.

   b) Determine the pH of 0.0180 M NH_3 (aq)

   c) Determine the value of the base ionization constant, Kb, for NH_3 (aq)

   d) Determine the percent ionization of NH_3 in 0.0180M NH_3 (aq).

   e) In an experiment a 20.0 mL sample of 0.0180 M NH_3(aq) was placed in a flask and titrated to the equivalence point and beyond using a0.0120 M HCl (aq).
      i) Determine the volume of 0.0120 M HCl (aq) that was added to reach the equivalence point.
      ii) Determine the pH of the solution in the flask after a total of 15.0 mL of 0.0120 M HCl(aq) was added.
      iii) Determine the pH of the solution in the flask after a total of 40.0 mL of 0.0120 M HCl(aq) was added.

APPENDIX C

LABORATORY EXCERCISES AND DEMONSTRATIONS

APPENDIX C-1

Demonstration: Food is Usually Acidic, Cleaners are Usually Basic

Material:
10 mL universal indicator
pH meter
50 mL each of pH-standard solutions with pH of 1, 3, 5, 7, 9, and 11
(Can be purchased or prepare according to instructions next page)
rack of 6 test tubes labeled 1, 3, 5, 7, 9, and 11 with stoppers
additional test tubes
50 mL each of the water based liquids

fruit juice	mouth wash
Vinegar	shampoo
Coffee	ammonia
Milk	bleach

1 teaspoon each of:

Aspirin	toothpaste
Baking soda	dishwashing liquid

Procedure:
1. Place 10 drops of universal indicator and 50 mL of the appropriate pH-standard solution into each tube. Stopper and invert several times to mix the solutions. Order the test tubes in the rack according to increasing pH.
2. Ask students to make a three column table in their notebooks and label them as acidic, neutral, and basic. They should write down each product tested along with its pH in the appropriate column.
3. Pour 50 mL of each water-based liquid along with 10 drops of universal indicator to a test tube. Stopper and invert several times to mix the contents. Compare the color with that of the standards. <u>For colored products and bleach use pH meter</u>.
4. Stir 1 teaspoon of each solid or thick liquid with 100 mL of distilled water and 10 drops of universal indicator in a small beaker. Pour in a test tube if necessary. Compare the color with the standard colors.
5. Challenge students to <u>classify</u> the products according to their pH patterns.

Discussion:
Most foods have a pH in the neutral to slightly acidic range. Sour foods have are more acidic. Oils and fats dissolve more quickly in base than in acid. Therefore, cleaning products tend to be more basic.
For more real world connections this investigation could be used to find a range of pH values for a variety of shampoos. The disulfide bonds that hold the amino acids chains together in hair break in a pH range of 8-9. As a result the cuticle of hair strands become

rough and hair looks slightly dull. Repeated use of alkaline shampoos can lead to complete breakage of disulfide bonds at the tip of hair strands, resulting in split ends. At pH values above 11, the amide bonds in the hair protein begin to break and hair dissolves. This demonstrates how alkaline drain cleaners can open drain clogs by dissolving hair that is causing the clog. Strong alkaline ingredients are also used in depilatories that dissolve and remove hair. pH 4-5 is best for human hair, but to clean hair an alkaline shampoo is needed. pH balanced shampoos have an acidic component to counteract the alkaline nature of detergents.

Students can bring their brands of shampoo to be tested for pH value. They can determine whether the shampoo they are using is within the acceptable pH range for healthy hair.

Preparation of pH-Standard solutions*

pH	Components
1	125 mL of 0.20M KCl and 335 mL of 0.20M HCl
2	125 mL of 0.20M KCl and 33 mL of 0.20M HCl
3	250 mL of 0.10M $KHC_8H_4O_4$ and 56 mL of 0.20M HCl
4	250 mL of 0.10M $KHC_8H_4O_4$ and 0.25 mL of 0.20M HCl
5	250 mL of 0.10M $KHC_8H_4O_4$ and 56 mL of 0.20M NaOH
6	250 mL of 0.10M KH_2PO_4 and 14 mL of 0.20M NaOH
7	250 mL of 0.10M KH_2PO_4 and 73 mL of 0.20M NaOH
8	250 mL of 0.10M KH_2PO_4 and 115 mL of 0.20M NaOH
9	250 mL of 0.025M $Na_2B_4O_7$ and 12 mL of 0.20M HCl
10	250 mL of 0.025M $Na_2B_4O_7$ and 12 mL of 0.20M NaOH
11	250 mL of 0.050M $NaHCO_3$ and 57 mL of 0.20M NaOH
12	125 mL of 0.20M KCl and 30 mL of 0.20M NaOH
13	125 mL of 0.20M KCl and 330 mL of 0.20M NaOH

*Combine the indicated components and dilute the resulting solutions to 500 mL with distilled water (Shakashiri, 1989)

APPENDIX C-2

Demonstration of Reactions of Acids

Background: the goal of this demonstration is to provide a visual experience for the students to help them remember the reactions of concentrated nitric, sulfuric, and hydrochloric acids and the products of each.

Material:

3 pieces of copper, 3-g each
15 mL concentrated (16M) nitric acid, HNO_3
80 mL Concentrated (15M) aqueous ammonia, NH_3
15 mL concentrated (12M) hydrochloric acid, HCl
5 mL Concentrated (18M) sulfuric acid, H_2SO_4
Ice

3 1-liter beakers
3 1-liter Erlenmeyer flasks
3 1-hole rubber stoppers
3 50-cm pieces glass tubing
plastic or rubber gloves
50 mL graduated cylinder

Procedure:

1. Ask students to predict the equation for the reactions of the 3 acids with copper metal. They should then draw a table with three columns. One for each of the acids to record their observations.
2. Fill the 3 beakers with tap water.
3. Bend the 3 pieces of glass tubing so when inserted into the 3 rubber stoppers they would go from the Erlenmeyer flask to beaker.
4. Pour 15 mL of concentrated HNO_3 into a dry 1-liter Erlenmeyer flask. Drop one of the copper pieces into the flask and quickly close with one the stoppers attached to glass tubing. Immerse the free end of the glass tubing in one of the beakers filled with water.
5. When the brown gas fills the glass tubing place the flask in ice water. Cooling the gas will reduce its pressure. Air pressure above the beaker will push water into the tubing and into the flask, producing a pale blue solution.
6. When water has stopped flowing into the flask, add 20 mL of concentrated aqueous ammonia from a graduated cylinder (optional). Concentrated ammonia forms a copper complex ion $Cu(NH_3)_4^{2+}$ that has a deep blue color.
7. Repeat step 4 but use concentrated HCl this time.
8. Heat the flask until HCl has boiled for 30 seconds. Remove the flask from the hot plate. Immerse the tubing in the second beaker of water. Watch the water rise in the tubing and pour into the flask. The solution should still be colorless. The solution will remain colorless after addition of concentrated ammonia to flask. This demonstrates absence of Cu^{2+} ions in the flask.
9. Repeat step 4 with 5 mL of H_2SO_4. Set the Erlenmeyer flask on a hot plate while raising the beaker of water on a stand so that the free end of tubing is immersed in the water.

10. Heat the flask until it is filled with dense white fumes and these fumes begin to bubble through the water in the beaker. Then carefully place the flask and the beaker in a shallow tray on the lab table.
11. Observe the water rising through the tubing and entering the flask. At first, drops of water will vaporize as soon as they come in contact with the hot sulfuric acid. The resulting high pressure will drive the water back in the tube and will cause rapid bubbling in the beaker of water. This cycle will happen a few times until the water will stream into the flask. Be advised that the mixing of water and concentrated sulfuric acid is **very exothermic** and will cause spattering. Pour ice water in the shallow tray and lift the flask to stop too much water from entering the flask. The pale blue Cu^{2+} ions will be visible against a white backdrop. Add 20 mL of concentrated ammonia to intensify the color.

Discussion: The three acids used in this demo are among the top 25 chemicals produced in the United States. As seen in this demonstration They each have different oxidizing abilities. These reactions are described by the following equations:

$4HNO_3\,(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2NO_3^-(aq) + 2NO_2(g) + 2H_2O(l)$

$2H_2SO_4\,(aq) + Cu(s) + \text{heat} \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq) + SO_2(g) + 2H_2O(l)$

$HCl + Cu(s) \rightarrow$ No Reaction

The pale blue color is from the hydrated copper ions ($Cu(H_2O)_4^{2+}$) produced when water enters the flask.

APPENDIX C-3

Demonstration of Reactions of Acids with Carbonates

Material:

Large test tube with cork
6 g sodium carbonate, Na_2CO_3, or sodium bicarbonate (baking soda), $NaHCO_3$
10 mL vinegar
Long stemmed funnel
Ring stand with test tube clamp

Procedure:

1. Ask student to write the equation for this reaction and predict the products.
2. Pour 10 mL of vinegar through the long-stemmed funnel into the test tube, so that the sides of the tube near its mouth stay dry.
3. Clamp the test tube at an angle so that the liquid comes to within 6 cm of the mouth of the tube,
4. Carefully place 6 g of Na_2CO_3 in the upper, dry part of the test tube, and firmly push the cork in.
5. Remove the test tube from the clamp to invert it once and quickly replace it in the clamp. Aim the tube away from people and fragile objects. Within seconds the cork will pop off with considerable force and will travel up to 15 feet away.

Discussion:

Carbonates fizz when they react with acids. This property is used to identify carbonate rocks such as lime and marble in earth science. This property is also used to identify carbonates or bicarbonates in qualitative analysis for identifying unknown powders. The gas formed is carbon dioxide which is produced according to the following equations:

$$CO_3^{2-}\,(aq) + 2H^+ \rightarrow CO_2\,(g) + H_2O\,(l)$$

$$HCO_3^{2-}\,(aq) + 2H^+ \rightarrow CO_2\,(g) + H_2O\,(l)$$

The pressure of the CO_2 gas produced builds up inside the test tube until it is great enough to eject the cork from the tube.
Students usually predict that reaction of acids and carbonates is a double replacement reaction. They write carbonic acid as a product which is partly true, except that carbonic acid is not very stable and breaks down into carbon dioxide and water. This demo has the elements of shock, sight, and sound to leave a lasting memory in students' minds so they would remember that there is a gas produced.

APPENDIX C-4

Conductivity Test
Extent of Ionization in Aqueous Acids

Back ground: The acidity of a solution depends on the concentration of hydrogen ions in the solution. Acidity is also a measure of the extent of ionization of the acid. This demonstration is quite effective in illustrating the difference between strong and weak acids as well as between concentrated and dilute solutions of acids.

Material:
Part A
100 mL 0.1 M hydrochloric acid HCl
100 mL0 .1 M acetic acid $HC_2H_3O_2$
100 mL 0.1 M citric acid $H_3C_6H_5O_7$ (dissolve 2.1 g in 50 mL and dilute to 100 mL)
universal indicator solution
2 or 3 light-bulb conductivity testers
wash bottle, filled with distilled water
power strip
4 250 mL beakers
pipette and suction bulb

Part B
80 mL 2M hydrochloric acid HCl
40 mL 2M acetic acid $HC_2H_3O_2$
140 mL of distilled water

Procedure:
Part A

1. Label each beaker with the name of the 3 above acids.
2. Pour 100 mL of each 0.1M acid into the appropriate beaker.
3. Add a few drops of universal indicator to each beaker of acid and determine the pH using a universal indicator pH chart.
4. Arrange the beakers by color in spectral order.
5. Dip the electrodes of each of the conductivity testers in each acid solution
6. Compare the brightness of the bulb from each solution to the pH of each

Part B

1. Compare the conductivity of 2M solutions of acetic and hydrochloric acid.
2. With a pipette remove 10 mL of the 2M solution of HCl and dilute in a small beaker with 30mL of distilled water. Ask students to calculate the new molarity of HCl (0.5M)
3. Dip the third conductivity tester in the newly diluted solution and compare the brightness of the light bulb.

4. Dilute 4mL of the 0.5M HCl with 36 mL of distilled water in a small beaker. Ask students to calculate the molarity of the newly diluted HCl solution (0.05M)
5. Rinse the third conductivity tester with distilled water and in the newly diluted solution and compare the brightness of the light bulb
6. Dilute 4mL of the 0.5M HCl with 36 mL of distilled water in a small beaker. Ask students to calculate the molarity of the newly diluted HCl solution (0.005M)
7. Rinse the third conductivity tester with distilled water and in the newly diluted solution and compare the brightness of the light bulb. Now the glow of the bulb will be nearly the same as the 2M acetic acid.

Discussion: In part A all acids have the same concentration but they have various values of pH or different $[H^+]$. This is revealed by the difference in the color of universal indicator. The conductivity tester reveals that the more acidic solutions are more conductive, with more ions present to conduct electricity. Since all the acids have the same, 0.1M, concentration, the difference in the glowing of the bulb must be a result of different extent of ionization.

In part B the conductivity of an acetic acid solution is compared with the conductivity of various concentrations of HCl. Although the acetic acid has a concentration of 2M, its extent of release of $[H^+]$ is about the same as a 0.005M solution of HCl.

APPENDIX C-5

Demonstration of Lewis Acids

Back ground: Bronsted and Lowry acid-base theory define an acid as a substance that donates a H^+ ion and a base as a substance that accepts a H^+ ion. The reactions in this demonstration do not involve H^+ ions.

Material:

5 g anhydrous aluminum chloride, $AlCl_3$
50 mL dichloromethane, CH_2Cl_2
5 mL pyridine, C5H5N
5 mL ethanol, C2H5OH
1 mL methyl violet solution in dichloromethane
(Dissolve 0.25 g of methyl violet in 100 mL of dichloromethane to obtain a 100 mL solution.)

2 125-mL Erlenmeyer flasks
funnel with filter paper
2 test tubes
test tube rack
droppers
2 10-mL graduated cylinder

Hazards:

- Anhydrous aluminum chloride should be kept dry. It reacts violently with water releasing a great deal of heat. It is also irritating to skin and respiratory system.
- Dichloromethane is toxic when inhaled or ingested.
- Methyl violet is toxic when ingested. Use gloves when handling to prevent persistent stains on skin.

Procedure:

1. Under a fume hood place 5 g of $AlCl_3$ and 50 mL of CH_2Cl_2 in a 125-mL Erlenmeyer flask.
2. Stopper the flask and shake it for a few minutes.
3. Pour into a filter funnel and collect the filtrate
4. Pour 10 mL of the clear filtrate into each of the two test tubes.
5. Stopper each test tube and set in the test tube rack
6. Add 3 drops of methyl violet solution to each tube.
7. Re-stopper the test tubes and invert them several times for even mixing of the content.
8. The violet solution turns yellow upon addition to $AlCl_3$.
9. Add 5 mL of pyridine to one of the test tubes. Stopper and invert to mix thoroughly. Solution will turn back to violet.
10. Repeat the above step with ethanol.

Discussion:

Methyl violet changes color in the pH range .4-1.8, which indicates very acidic solutions. The yellow color reveals acidic property of $AlCl_3$ despite absence of hydrogen in its formula. Therefore it cannot be a Bronsted and Lowry acid. The Al atom in $AlCl_3$ has an incomplete valence shell that can accept a pair of electrons. Methyl violet contains 3 amine groups with a lone pair of electrons on each nitrogen atom. A Lewis acid base

reaction occurs when Al forms bonds with these nitrogen atoms. This reaction changes the color of methyl violet to yellow. According to Lewis theory an acid accepts a pair of electrons by forming a bond to a base that donates a pair of electrons.
Pyridine has nonbonding electron pairs on nitrogen atoms and ethanol has nonbonding electron pairs on oxygen atoms. They can both act as Lewis bases to neutralize $AlCl_3$ and return methyl violet back to violet color.

APPENDIX C-6

Hydrolysis

Background: the process by which a salt affects the pH of its solution has been called hydrolysis. Hydrolysis is the reaction of water with ions of a salt. These ions are able to react with water to generate H^+ (aq) or OH^- (aq). The pH of a salt solution can be predicted by considering the ions involved.

Purpose: To observe the pH of aqueous salt solutions and to find patterns in their acid base behavior

Procedure: Place a drop or two of each of the following solutions on a strip of pHydron indicator paper. Boil some distilled water for five minutes. Allow it to cool while covered with a watch glass. First, test the boiled and unboiled distilled water for pH. Allow thirty seconds before recording the color of the indicator paper. Use the pH chart on the vial of indicator paper to record the pH of each solution in the data table. Indicate yes or no, for hydrolysis. Use universal or other indicators to clarify the results that are border line.

Distilled H_2O	$FeCl_3$	NaF
Boiled Distilled H_2O	$Zn(NO_3)_2$	
Na_2CO_3	NH_4Cl	
Na_2S	KCl	
NaBr	$Mg(NO_3)_2$	
$NaC_2H_3O_2$	K_3PO_4	
$NaNO_3$	$Ca(NO_3)_2$	
$CuSO_4$	NH_4NO_3	

Data Table:

Solution	Color of pH paper	pH	Hydrolysis?

Analysis:

1. Working together with your group mates, organize the salts into the following classifications: neutral, acidic, basic (make a table)

Neutral Salts	Acidic Salts	Basic Salts

2. Classify the ions in the salts. Begin by listing all the cations and anions that are in neutral salts. Then eliminate the neutral ions from the acidic or basic salts to find out which ion is responsible for acidic or basic behavior. For example, if Na^+ is a neutral ion the basic behavior of Na_2CO_3 solution is caused by the CO_3^{2-} ion.

Neutral ions	Acidic ions	Basic ions

3. Look for patterns according to groups in the periodic table for example: group 1 elements have___acidic, basic, or neutral___ cations.

4. For the salts that hydrolyzed, demonstrated acidic or basic behavior, write net ionic equations for the reaction.

5. Predict what results would occur when each of the following salts is dissolved in water. Use experimental results as a guideline. For those where a reaction is predicted, write an equation for hydrolysis.

   $AlCl_3$ NaH NaH_2PO_4 NaF

6. Use your book index to find and write the definitions for acid anhydride and basic anhydride. Give one example of each.

7. Anhydrides are another class of compounds that do hydrolyze. Write out the equation for hydrolysis of carbon dioxide, an acid anhydride. The term "acid breath" is an impersonal, scientific comment. Why can it be used? Think of a way to demonstrate this.

8. Underarm anti-per spirants use salts containing Al^{3+} ions. Explain how Al^{3+} works to keep your underarm dry and what negative side effect results when it does work.

Hydrolysis Lab

Teacher Version Data tables

Solution	Color	pH	Hydrolyzed?
Distilled H_2O	Light orange	6	
Na_3PO_4	Very dark blue	12	yes
$CuSO_4$	Orange	4.5	yes
KNO_3	Light orange	6	
Na_2S	Dark blue	11	yes
$NaHCO_3$	Forest green	9	yes
$Ca(NO_3)_2$	Light orange	6	
$Ba(NO_3)_2$	Light orange	6	
Na_2CO_3	Dark blue	11	yes
$Mg(NO_3)_2$	Light orange	6	
$NaC_2H_3O_2$	Pale green	8	yes
$FeCl_3$	Red	2	yes
$Fe(NO_3)_3$	Red	2	yes
$AlCl_3$	dark orange	4	yes
K_2CO_3	Dark blue	11	yes
NH_4NO_3	Orange	5	yes
$Zn(NO_3)_2$	Orange	5	yes
NaCl	Light orange	6	

Neutral Salts	**Acidic Salts**	**Basic Salts**
KNO_3	$CuSO_4$	Na_3PO_4
$Ba(NO_3)_2$	$FeCl_3$	Na_2S
$Ca(NO_3)_2$	$Fe(NO_3)_3$	$NaHCO_3$
$Mg(NO_3)_2$	$AlCl_3$	Na_2CO_3
NaCl	NH_4NO_3	$NaC_2H_3O_2$
	$Zn(NO_3)_2$	K_2CO_3

Neutral ions	**Acidic ions**	**Basic Ions**
K^+ NO_3^-	Cu^{2+}	PO_4^{3-}
Ba^{2+} NO_3^-	Fe^{3+}	S^{2-}
Ca^{2+} NO_3^-	Al^{3+}	HCO_3^-
Mg^{2+} NO_3^-	NH_4^+	CO_3^{2-}
Na^+ Cl^-	Zn^{2+}	$C_2H_3O_2^-$

General Pattern

Neutral ions:

- Group I & II cations
- Anions of strong acids

Acidic ion:

- Transition metal ions
- Aluminum and ammonium ions

Basic ions:

- Anions of weak acids

These reactions are Bronsted and Lowry acid-base reactions. Water behaves as an acid when it donates a proton (H^+) to the anion of a weak acid. It behaves as a base when it accepts a proton from cation of a weak base or from complex ions.
Net ionic Equations for typical hydrolysis reactions:

1. $NH_4^+ + H_2O <==> H_3O^+ + NH_3$
2. $Cu^{2+} + 4\ H_2O <==> Cu(H_2O)_4^{2+}$

   $Cu(H_2O)_4^{2+} + 2\ H_2O <==> H_3O^+ + Cu(H_2O)_3OH^+$
3. $Fe^{3+} + 6\ H_2O <==> Fe(H_2O)_6^{3+}$

   $Fe(H_2O)_6^{3+} + 2\ H_2O <==> H_3O^+ + Fe(H_2O)_5OH^{2+}$
4. $S^{2-} + H_2O <==> HS^- + OH^-$
5. $H^- + H_2O <==> H_2 + OH^-$

Each Al^{3+} ion captures six water molecules to form a complex ion. This and other metal ions with a +3 charge are used in antiperspirants to keep underarm dry. This reaction is similar to reaction 3 above. Once the complex ion is formed it reacts with water and releases H_3O^+ ions. That is why some antiperspirants sting if underarm has been cut with a razor.

APPENDIX C-7A

pH Titration

Background: Acid-base titrations follow a relatively standard procedure for analysis of acid or base strength. In this experiment you will determine the concentration of a basic solution by titration with hydrochloric acid. You will use a pH meter to determine the pH of the solution. This is an electronic device which compares the voltage in a solution to that of a standard. It monitors changes in the H_3O^+ concentration. The pH is read out directly from a digital meter. A titration curve can be made by plotting pH versus volume of the solution with unknown concentration. The equivalence point of a pH titration can be determined using the titration curve.

Material:

Buffer solution with pH 7.0
Buffer solution with pH 4.0
pH meter
25 mL pipet
magnetic stirrer
magnetic stirring bar
buret and buret clamp
4 beakers
50 mL 0.10 M HCl (per group)
50 mL NaOH (unknown concentration)

Procedure:

(Teacher's Note: use 0.2M NaOH but do not reveal the molarity to students until they have determined the molarity themselves)

Part I: Standardization of pH Meter

1. At the start of the experiment, you will find the electrodes soaking in water.
2. Carefully remove the beaker without touching the electrodes, which have a thin membrane and are costly to replace.
3. Pipet about 25 mL of each of the pH 4 and pH 7 buffers into two clean beakers. Add about 150 ml of distilled water to the beaker and stir thoroughly.
4. Place the probe in the pH 4 buffer. Adjust the corresponding screw on top of the electrode so that the meter reads 4.
5. Rinse the probe with distilled water and place in the pH 7 buffer. Adjust the pH 7 screw until it reads 7.
6. Rinse and gently dry the electrode.
7. Immerse the electrodes in "fresh" distilled water.

Part II: Titration

1. Pipet 25 ml of a 0.10 M HCl into a clean beaker and add about 150 ml of distilled water.
2. Place the beaker on a magnetic stirring stand and place the stirring bar in the beaker of acid. Check out the latter from your instructor.

3. Lower the probe into the beaker and adjust the height so that the electrode is far enough above the stirring bar so that the bar will not smash into the electrode. Add enough water as needed to insure that the probe is immersed in the solution.
4. Clean the buret so that no beading occurs. Rinse once with distilled water, draining it through the tip. Coat the inside of the buret with about 5 mL of the base provided by your teacher. Drain this through the tip.
5. Adjust your buret on the ring stand in such a way that the stream from the buret can be directed into the beaker. Make sure the buret is filled with your assigned base <u>before</u> setting it over the beaker of acid. Record the buret <u>reading</u> (initial reading) after all air bubbles are removed

6. Turn the pH meter on. Record the pH of the acid solution. This is the pH for 0.00 ml base added.

7. Add the NaOH to the acid about 1 ml at a time and record total ml of base and pH reading after each addition. As the endpoint is approached and the pH begins to rise more rapidly, decrease the increments to two drops at a time. Do this part rapidly as prolonged readings result in more dissolved CO_2 and altered results.

8. Beyond the endpoint, the pH will change less rapidly. Resume 1 mL increments until pH between 10.5 and 11.0 is reached.

9. Turn the pH meter off, rinse off the electrode and immerse in original distilled water.

10. Repeat the procedure for a second trial.

11. Rinse off and return stirring bar to instructor. DO NOT POUR DOWN THE SINK WITH THE SOLUTION.

Graphical Analysis:

1. Use a graphing program to plot pH as a function of mL of base added. (The classroom computers have Graphical Analysis 3.1 installed on them). Determine the Equivalence point from this graph by following directions on the pH Titration handout.

2. Select the better run and plot the first derivative graph by calculating $\frac{\Delta \text{ pH}}{\Delta \text{ mL base}}$

3. for each reading and plot it on the Y axis versus mL of base (X axis). Refer to the pH Titration handout for further detail. Determine the equivalence point from this graph.

4. Plot the second derivative graph $\frac{\Delta \text{ pH}}{\Delta \text{ mL base}}$

on Y axis and mL base on X axis and determine the equivalence point on this graph.

5. Using the first and second derivative graphs, interpolate the exact volume of base added at the endpoint. On the Graphical Analysis program this can be accomplished by putting the curser on equivalence point. A box will appear with the values for the X and Y axes. Record this value in large numbers in upper right corner of each graph.

Calculations:
(Select best trial)

1. Moles of acid added (liters) (moles/liter)

2. Moles of base added at endpoint.

3. Molarity of base (moles/liter). Calculate to thousandth place, round back to hundredths.

Questions:

1. Why is it necessary to use a buffer solution to standardize the pH meter?

2. Explain what effect each of the following would have on the result of the molality of the base.

   a) The pipet for the acid was rinsed only in distilled water and not with the acid as well.

   b) The beaker into which the acid was added was not dry.

   c) An air bubble of some size was present in the base buret tip at the end of the titration.

   d) Prolonged readings were taken so that the amount of dissolved CO_2 from the air played a role in the system.

   e) The base buret had only been rinsed with distilled water and not with the base as well.

3. What is the advantage of pH titration over titrating to a phenolphthalein end point?

APPENDIX C-7B

Weak Acid- Strong Base Titration

The purpose of this experiment will be to determine the Ka and the molar mass of an unknown solid acid.

For a strong acid-strong base titration, the net ionic equation is written

$$H^+ + OH^- \rightarrow HOH \quad [\text{or } H_2O]$$

and the pH will = 7.00 at the equivalence point.
However, by definition, a weak acid dissociates to only a slight extent; therefore, the net ionic equation for a weak acid-strong base titration will be:

$$HA + OH^- \Leftrightarrow HOH + A^-$$

The ion then will hydrolyze in water to produce a basic solution at the equivalence point according to the equation

$$A^- + HOH \Leftrightarrow HA + OH-$$

Procedure:

1. Rinse a 50.00 mL buret, first with distilled water, then with the NaOH solution. Note the concentration of the NaOH on the bottle. Place the buret in a ring stand; obtain a stirrer and a stirring bar.
2. Using the analytical balance, measure two samples (about 0.4-0.5 g each) of the unknown solid acid and add them to separate 250 mL beakers.
3. Add about 125 mL of distilled water.
4. Standardize the pH meter with the buffers solutions, the first one at pH = 7 and the other at pH = 4.
5. Titrate the NaOH into the weak acid solution, going by 3 mL increments to a pH between 10 and 11.
6. Decide where the approximate end point occurred as well as the volume reading at halfway to the endpoint.
7. Repeat the titration with the 2nd solution of unknown acid in water, but 2 ml ahead of the halfway point, slow down, and proceed by 0.1-0.2 mL increments until you are one mL beyond that region. Then proceed by 1 mL increments until about 2-3 mL before the equivalence point. Then using increments of 0.1-0.2, add base until the pH has a sharp rise after pH = 7.0. Continue to take small (0.1-0.2) mL increments for another 2 mL.
8. Take 2-3 more readings at about 2-3 mL intervals.

Graphs and Calculations

1. Construct a graph of pH vs. mL for your second run. Determine the approximate equivalence point to one decimal place. Indicate this value on your graph.
2. Graph the first derivative graph [ΔpH/ ΔmL vs. mL] to determine the equivalence point to two decimal places. Indicate this value directly on the graph.
3. From the mass of the sample and the volume and concentration of the base added, determine the molar mass of your unknown solid.
4. On your first graph, determine the volume for the <u>halfway Point</u> in the titration; note the pH at this point.
5. If $Ka = [H^+][A^-]/ [HA]$, at the halfway Point the amount of HA remaining and A- formed will be equal. Hence their concentrations will cancel each other in the Ka expression. Determine the pKa and the Ka of your weak acid. Include your unknown number in your pertinent data section.

Questions:

1) Why do we not use water to standardize the pH meter?

2) What effect, if any, will each of the following experimental errors have on the calculated value for molar mass of the unknown?

   a) Some solid is spilled in transferring it from the weighing boat into the beaker.

   b) The buret containing the base solution is first rinsed with distilled water, but not with base before it is filled.

   c) The endpoint of the titration is overshot during the titration.

   d) Only 100 mL of water is added into the titration beaker containing the solid.

   e) The student selects the endpoint volume at pH = 7.

APPENDIX C-8

Equilibrium in Solutions of Weak Acids and Weak Bases-Buffer Solutions

Background: This experiment is designed to provide evidence that weak acids and bases are equilibrium systems. You will subject these systems to certain stresses and see whether or not you can interpret the observed changes in terms of equilibrium principles. You will then obtain or prepare a buffer solution and compare its action with that of an unbuffered solution when a strong acid and a strong base are added to each.

Objectives:

I. To obtain evidence that weak acids and weak bases are equilibrium systems.
2. To study the common-ion effect as applied to weak acids and bases.
3. To prepare a buffer and demonstrate its action.

Materials:

Test tubes and rack	graduated cylinder	acetic acid (6M CH_3COOH)
$NaC_2H_3O_2$	$CaCO_3$	methyl orange indicator
NH_4Cl	6M NH_3	phenolphthalein
NaCl solution	buffer solution with pH 7 0.	bromocresol green indicator
1M NaOH	HCl	

Procedure:

1. Label three test tubes A, B, C. Add five ml of water and 5 ml 6.0 M acetic acid to each of the three large test tubes. Dissolve approximately 1.5 g of $NaC_2H_3O_2$ in B, and 3 g in C. Shake well.

2. Place three more large empty test tubes in the rack and add two grams of powered $CaCO_3$ to each test tube.

3. Pour **solution A** into one of the test tubes containing $CaCO_3$. Shake and record the approximate time required for the foam to reach the height of one inch. Record the time in **Table I**.

4. Repeat the procedure using solutions B and C.

5. Add 1 mL of 6 M acetic acid and three drops of methyl orange indicator to 9 mL of water in each of two test tubes. Record the color in **Table II**. Refer to the indicator chart in the appendix and record the approximate pH of the solution. Dip a strip of universal indicator paper and record its color as well.

6. Add 2 g of sodium acetate to one of the test tubes. Compare the color with the other solution and record any changes.

7. Add 1 mL of 6 M HCl and three drops of methyl orange to 9 mL of water. Record the color. Add two grams of NaCl to the solution and record any color change. Check with indicator paper too.

8. Add a drop of 6 M NH_3 solution to 100 mL of water. Shake and pour 10 mL of the solution into each of the two test tubes. Add a drop of phenolphthalein to each tube and record the color in **Table III.**

9. Add 2 g of NH_4Cl to one of the tubes and record any color change.

10. Place 5 mL NaCl solution (assume pH 7) into a small test tube and 5 ml of a buffer solution with pH 7 into another small test tube. Add two drops of phenolphthalein into each test tube. Record in **Table IV** the number of drops of 0.1 M NaOH needed to cause a color change in each tube.

11. Measure five mL of sodium chloride solution, and five ml of a buffer solution with a pH 7 into each of two small test tubes. Add two drops of methyl orange and bromocresol green to each. Record the number of drops of 0.1 M HCl required to produce a color change in each tube.

Follow Up Discussion:

The ionic equation for the reaction between calcium carbonate and acetic acid shows that H_3O^+ ions are reactants. According to the Law of Mass Action the rate of the reaction $CaCO_3(S) + 2H_3O^+ (aq) \rightarrow CO_2 (g) + 3H_20 (l) + Ca^{+2} (aq)$ varies with the concentration of the H_3O^+. You observed that the rate of foaming decreased when the acid solutions treated with $NaC_2H_3O_2$ were used. This indicates that the $[H_3O^+]$ in these solutions was less than in the treated one. The decrease in the $[H_3O^+]$ was the result of a shift in the equilibrium $HC_2H_3O_2 + H_2O \Longleftrightarrow H_3O^+ + C_2H_3O_2^-$ brought about by the addition of $NaC_2H_3O_2$ which contains an ion common with the system described by the equation alone.

Follow Up Questions:

1. Why did solution C listed in Table I result in the slowest reaction? Use Le Chatlier's principle to explain qualitatively the difference between the $[H_3O^+]$ in solution A and Solution B. Answer and, where applicable, write all other equations in net ionic equations and equilibrium shifts.
2. Did the pH go up or down when $NaC_2H_3O_2$ was added to acetic acid? Did the concentration of undissociated $HC_2H_3O_2$ increase or decrease? Explain using Le Chatlier's principle.
3. Refer to **Table II** and indicate whether addition of Cl^- to the HCl solution caused a change in pH. Explain the similarity or difference between this observation and that on the $HC_2H_3O_2$-$NaC_2H_3O_2$ system?
4. Refer to **Table III** and indicate whether addition of NH_4Cl caused the pH to increase or decrease. Explain your observation.
5. How many drops of acid were required to change the color of the buffered solution listed in **Table IV**? What fraction of the total volume of buffer did it take to cause a

color change, assume 20 drops per mL? What fraction of the total volume of the unbuffered solution did it take to change its color?

Table I: $CaCO_3$ (s) + 2 H^+ $\rightleftarrows$ CO_2 (g) + H_20 (l)+ Ca^{+2}(aq)

Reagent 1	**Solution**	**Solute (Reagent 2)**	**Time for Foam to Rise**
$CaCO_3$	A	$HC_2H_3O_2$	
$CaCO_3$	B	$HC_2H_3O_2$+ 1.5 g $NaC_2H_3O_2$	
$CaCO_3$	C	$HC_2H_3O_2$+ 3 g $NaC_2H_3O_2$	

Table II: Common Ion Effect: $HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$

Solution	**Indicator**	**Color**	**Approximate pH**
$HC_2H_3O_2$			
$HC_2H_3O_2$ with $C_2H_3O_2^-$ as a common ion			
HCl			
HCl with Cl^- as a common ion			

Table III: Common Ion Effect: NH_3 + HOH <==>NH_4^+ + OH^-

Solution	**Indicator**	**Color**	**Approximate pH**
Aqueous NH_3			
Aqueous NH_3 with NH_4^+			

Table IV: Buffer Action

Solution	Drops of Base to Give Color Change	Drops of Acid to Give Color Change
Unbuffered NaCl Solution, pH 7		
Buffer Solution with pH 7		

pH Range	Indicator	Color Change
0.4-1.8	Methyl Violet	yellow-blue violet
1.2-2.8	Thymol Blue	red-yellow
1.2-3.8	Benzopurpurin 4B	violet-red
3.1-4.9	Congo Red	blue-red
3.0-4.6	Bromophenal blue	yellow-blue violet
3.1-4.4	Methyl orange	red-yellow
4.0-5.6	Bromcresol green	yellow-blue
4.4-6.2	Methyl green	red-yellow
5.0-7.0	Litmus	red-blue
5.2-6.8	Bromocresol purple	yellow-violet
6.2-7.6	Bromthymol blue	yellow-blue
6.4-8.0	Phenol red	yellow-red
7.2-8.8	Cresol red	yellow-red
7.4-9.0	Metacresol purple	yellow-violet
8.0-9.6	Thymol blue	yellow-blue
8.2-10.0	Phenolphtalein	colorless-red
9.4-10.6	Thymolphtalein	colorless-blue
10.0-12.0	Alazarin Yellow	yellow-violet
11.4-13.0	Sodium indigosulfonate	blue-yellow
12.0-14.0	1,3,5 - Trinitrobenzene	colorless-orange

Teacher Version Data tables
Typical Data (results may differ with age of materials used)

Table I : $CaCO_3$ (s) + 2 H^+ $\rightleftarrows$ CO_2 (g) + H_20 (l)+ Ca^{+2}(aq)

Reagent 1	**Solution**	**Solute (Reagent 2)**	**Time for Foam to Rise**
$CaCO_3$	A	$HC_2H_3O_2$	1 sec
$CaCO_3$	B	$HC_2H_3O_2$+ 1.5 g $NaC_2H_3O_2$	20 sec
$CaCO_3$	C	$HC_2H_3O_2$+ 3 g $NaC_2H_3O_2$	30 sec

Table II: Common Ion Effect: $HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$

Solution	**Indicator**	**Color**	**Approximate pH**
$HC_2H_3O_2$	Methyl orange pH paper	dark pink red/orange	3.5 3
$HC_2H_3O_2$ with $C_2H_3O_2^-$ as a common ion	Methyl orange pH paper	orange/yellow orange	4.2 6
HCl	Methyl orange pH paper	red dark pink	3.1 1
HCl with Cl^- as a common ion	Methyl orange pH paper	red dark pink	3.1 1

Table III: Common Ion Effect: NH_3 + HOH <==>NH_4^+ + OH^-

Solution	**Indicator**	**Color**	**Approximate pH**
Aqueous NH_3	Phenolphthalein pH paper	fuchsia green	9.7 9
Aqueous NH_3 with NH_4^+	Phenolphthalein pH paper	cloudy orange	8.2 5

Table IV: Buffer Action

Solution	Drops of Base to Give Color Change	Drops of Acid to Give Color Change
Unbuffered NaCl Solution, pH 7	2	2
Buffer Solution with pH 7	53	48

In Table I addition of $NaC_2H_3O_2$ slows the rate of foaming. As described in the follow up discussion in the lab $HC_2H_3O_2 + H_2O <==> H_3O^+ + C_2H_3O_2^-$ is shifted to the left due to addition of common ion $C_2H_3O_2^-$ thus, $[H_3O^+]$ is suppressed and not available to react with $CaCO_3$ at first.

This shift to the left causes the concentration of hydrogen ions to decrease and consequently pH of $HC_2H_3O_2$ is increased after the addition of $NaC_2H_3O_2$ in Table II

In Table III addition of NH_4^+ to solution of ammonia drives the reaction to the left and suppresses the concentration of OH^- ions $NH_3 + HOH <==> NH_4^+ + OH^-$

Table IV demonstrates that the buffered solution resists change in pH as is evident from the increased number of drops required to change the pH.

APPENDIX C-9

Rubric for AP Chemistry Lab Reports

Procedure:

1. Procedures are drawn clearly and in a logical order.
2. Procedures are drawn in a logical order, but pictures are not clear.
3. Procedures are drawn but are not in a logical order or are difficult to follow
4. Procedures do not accurately depict the steps of the experiment.

Data:

1. Professional looking and accurate representation of the data in tables and/or graphs. Graphs and tables are labeled and titled. Units are included.
2. Accurate representation of the data in tables and/or graphs. Graphs and tables are labeled and titled.
3. Accurately representing the data in written form, but no graphs or tables are presented.
4. Data are not shown OR are inaccurate.

Calculations:

1. All calculations are shown and the results are correct and labeled appropriately. Units are included
2. Some calculations are shown and the results are correct and labeled appropriately. Units are included
3. Some calculations are shown and the results labeled appropriately. Units are missing.
4. No calculations are shown OR results are inaccurate or mislabeled. Units are missing.

Discussion:

1. All questions are answered in full sentences and the answers are well thought out.
2. All questions are answered in full sentences but the answers are not well thought out.
3. Questions are not answered in full sentences.
4. All questions are not answered

Analysis:

I. Reflection on of what was learned from performing the experiment.
II. A brief summary of what was done in the lab.
III. The scientific concept demonstrated or used in the experiment is discussed.
IV. Whether the purpose of the lab was accomplished and what do calculations show.
V. Trends/patterns logically analyzed.
VI. What would you change to obtain better results?

1. Analysis includes all of the above.
2. Analysis is missing one of the above.
3. Analysis is missing two of the above.
4. Analysis is missing more than two of the above.

Error Analysis:

1. Experimental errors, mathematical analysis (%error, %yield, and etc.), their possible effects, and ways to reduce errors are discussed.
2. One of the following is missing: Experimental errors, mathematical analysis (%error, %yield, and etc.), their possible effects, and ways to reduce errors are discussed.
3. Experimental errors are mentioned
4. There is no discussion of errors

Participation:

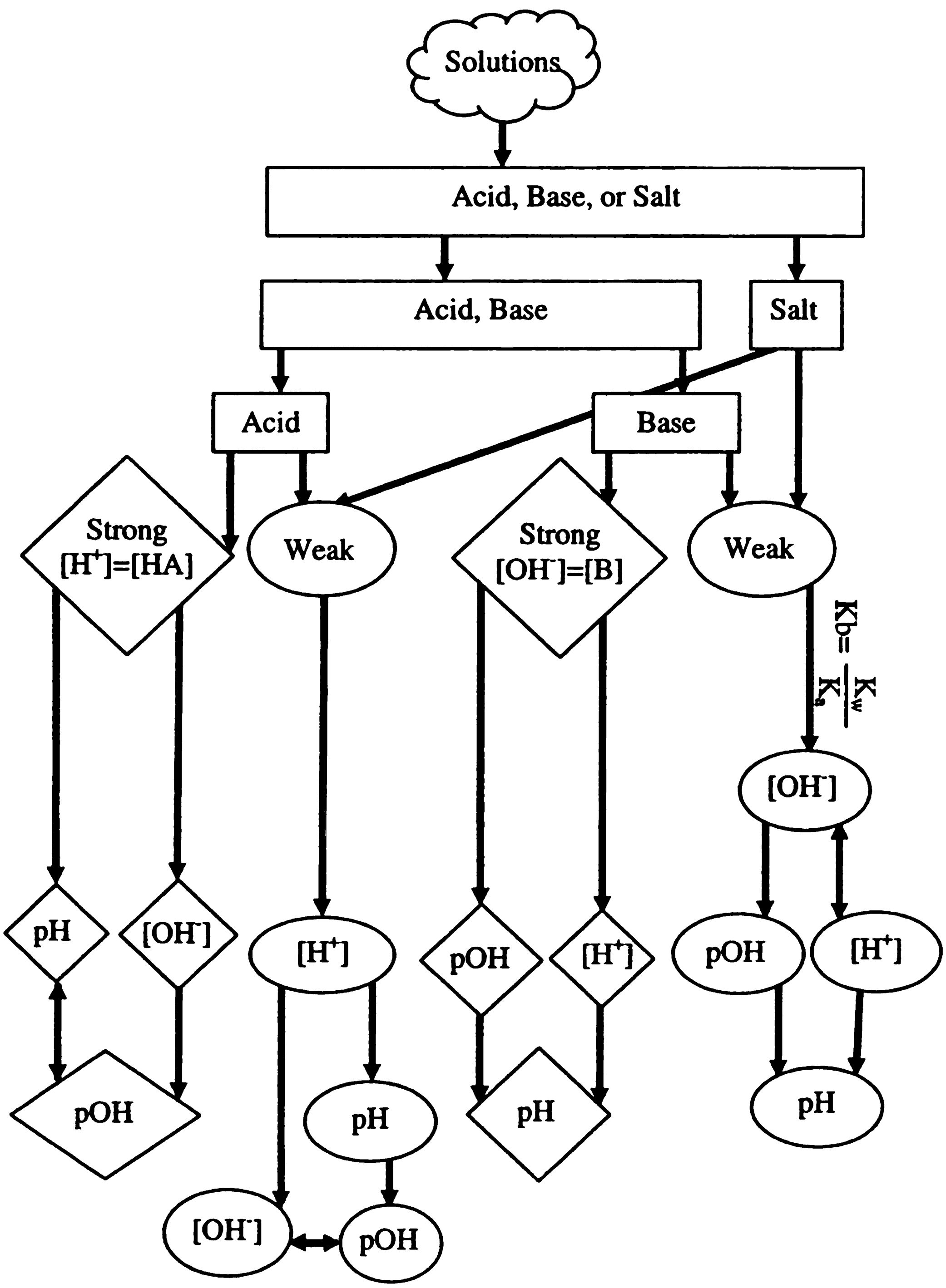
1. Used time well in lab and focused attention on the experiment. Collaborated with lab partners.
2. Used time pretty well. Stayed focused on the experiment most of the time. Collaborated with lab partners
3. Did the lab but did not appear very interested. Focus was lost on several occasions.
4. Participation was minimal OR student was unreceptive to reminders about participating.

APPENDIX D

WORKSHEETS AND CHARTS

APPENDIX D-1

Finding the pH Map



APPENDIX D-2

pH and pOH Table

Use the appropriate formulas to fill in the empty cells in the table. Once you recognize the pattern for each column you may use it to write the answers instead of using formulas.

Solution	$[H^+]$	$[OH^-]$	K_w	pH	pOH	pH + pOH
Very acidic	10^{-1}		10^{-14}			
Moderately acidic	10^{-3}		10^{-14}			
Slightly acidic	10^{-5}	10^{-9}	10^{-14}	5	9	14
Pure water	10^{-7}	10^{-7}	10^{-14}	7	7	14
Slightly basic	10^{-9}		10^{-14}			
Moderately basic	10^{-11}		10^{-14}			
Very basic	10^{-13}		10^{-14}			

Answers

Solution	$[H^+]$	$[OH^-]$	K_w	pH	pOH	pH + pOH
Very acidic	10^{-1}	10^{-13}	10^{-14}	1	13	14
Moderately acidic	10^{-3}	10^{-11}	10^{-14}	3	11	14
Slightly acidic	10^{-5}	10^{-9}	10^{-14}	5	9	14
Pure water	10^{-7}	10^{-7}	10^{-14}	7	7	14
Slightly basic	10^{-9}	10^{-5}	10^{-14}	9	5	14
Moderately basic	10^{-11}	10^{-3}	10^{-14}	11	3	14
Very basic	10^{-13}	10^{-1}	10^{-14}	13	1	14

APPENDIX D-3

Titration Problems

	SA + SB	**WA + SB**	**WB + SA**
Initially	pH is determined by $[H^+]$ or $[OH^-]$	$HA <-> H^+ + A^-$ $Ka = \frac{[H^+]\,[A^-]}{[HA]}$ Use $[H^+]$ to find pH	$A^- + H_2O <-> HA + OH^-$ $K_b = \frac{[HA]\,[OH^-]}{[A^-]}$ Use $[OH^-]$ to find pOH
Before Equivalence Pt.	1. Stoich: find mmol of H^+ and OH^- 2. ICE Box 3. $\frac{\text{mmol excess}}{\text{total volume}}$ 4. find pH	1. Stoich: find mmol and H^+ and OH^- 2. H-H equation! $pH = pK_a + \log \text{base/acid}$ $HA + OH^- -> A^- + H_2O$	1. Stoich: find mmol HA and A^- added 2. H-H equation!
Midpoint		$pH = pK_a$	$pH = pK_a$ $pOH = pK_b$
At Equivalence Pt.	1. Stoich: find mmol of H^+ and OH- 2. ICE Box $H^+ = OH^-$ 3. neutral! pH=7	1. Stoich: find mmol of H^+ and OH^- 2. First ICE Box excess A^- 3. Second ICE Box based on A^- 4. $K_b = K_w/K_a$ 5. Use K_b to find $[OH^-]$ to get pH $HA + OH^- -> A^- + H_2O$	1. Stoich: find mmol of A^- and H^+ 2. First ICE Box excess HA 3. Second ICE Box based on HA 4. $K_a = \frac{[H^+][A^-]}{[HA]}$ 5. Use K_a to find $[H^+]$ to find pH
After Equivalence Pt.	1. Find mmol excess added 2. ICE Box 3. $\frac{\text{mmol excess}}{\text{total volume}}$ 4. Find pH	1. Find mmol SB added 2. ICE Box 3. $\frac{\text{mmol excess SB}}{\text{total volume}}$ 4. pOH from $[OH^-]$	1. Find mmol SA added 2. ICE Box 3. $\frac{\text{mmol excess SA}}{\text{total volume}}$ 4. pH from $[H^+]$

SA = Strong Acid
WA = Weak Acid
mmol = millimoles

SB = Strong Base
WB = Weak Base

stoich = use stoichiometry
H-H = Henderson Hasselbach Equation

APPENDIX D-4

Finding pH

Determine the pH of the following aqueous systems, all 0.1M:

1) 0.1 M HCl

2) 0.1 M NaOH

3) 0.1 M H_2CO_3

4) 0.1 M NH_3

5) 0.1 M NH_4Cl

6) 0.1 M KCN

Show all set-ups and sig figs. Remember, classify the solution first, (strong acid, weak acid, basic salt, etc.) then crunch the numbers.

APPENDIX D-5

Titration Problems
Strong acid-strong base

Consider 50.00 mL of 1.000 HCl (pH = 0) titrated with 1.000 M NaOH.

1. 25.00 mL of NaOH added, what is the pH?
   a. Convert to moles
   b. Do stoichiometry to find moles left over
   c. Find new M of left over ion
   d. Find pH

2. 49.00 mL of NaOH: pH?

3. 50.00 mL of NaOH: pH?

4. 50.01 mL of NaOH: pH?

5. 75.00 mL of NaOH: pH?

Plot pH versus volume of base used in each titration.

APPENDIX D-6

Titration Problems
Weak acid-strong base

Consider 50.00 mL of 1.00 M $HC_2H_3O_2$ (HAc) titrated with 1.000 M NaOH.

6. What is pH before any NaOH is added? (Ans: 2.37)
   a. This is an equilibrium (Ka) problem, due to the weak acid HAc.

7. What is the pH after 25.00 mL of NaOH added? (Ans: 4.74)
   a. Convert to moles
   b. Do neutralization stoichiometry
   c. Back to weak acid equilibrium to recalcalculate new [H+], consider any leftover amounts of anion of weak acid

8. What is the pH after 50.00 mL of NaOH is added? (Ans: 9.22)

9. What is the pH after 75.00 mL of NaOH is added

10. Plot pH vs. volume of base used in each titration.

APPENDIX D-7

AP Acid-Base Titration Sheet
(Weak Acids-Weak Bases)

1. 50.0ml of 0.200M NH_3 is titrated with .200M HCl solution.
   a. pH of solution at start of titration = ____________
   b. pH half way to end point = ____________
   c. pH at the equivalence point = ____________
   d. pH after 75.0 mL of acid has been added = ____________

2. Calculated the final pH of the following mixtures.
   a. 50 mL .10 M HCl and 50 mL of .1M $NaCH_3COO$ = ____________
   b. 50ml .1M NH_4Cl and 50 mL of .1M NaOH = ____________

3. 30.0 mL of .150 M HCOOH titrated with .150 M NaOH
   (Ka of HCOOH = 1.8×10^{-4})

   Calculate the pH at the following points:
   a. After 15 mL base added = ____________
   b. After 30 mL base added = ____________
   c. After 40 mL base added = ____________

K_a – $CH_3COOH = 1.8 \times 10^{-5}$
K_b – $NH_3 = 1.8 \times 10^{-5}$

APPENDIX D-8

Acid-Base Review Problems

a) What is the pH of a 2.0 molar solution of acetic acid? Ka acetic acid = 1.8×10^{-5}

b) A buffer solution is prepared by adding 0.10 liter of 2.0 molar acetic acid solution to .1 liter of a 1.Q molar sodium hydroxide solution. Compute the hydrogen ion concentration of the buffer solution.

c) Suppose that 0.10 liter of 0.50 molar hydrochloric acid is added to 0.040 liter of the buffer prepared in (b). Compute the hydrogen ion concentration of the resulting solution

AP

A 10.00 milliliter sample of NH_3 solution is titrated with a standard HCl solution.
1) An unknown volume of water is added to the HCl solution.
2) An unknown volume of water is added to the 10.00 milliliter sample of NH_3 solution.
3) Phenolphthalein is used as the indicator.
For each of these three steps taken during the titration:
a) State whether it introduces an error into the titration results
b) For any of the steps that introduce(s) an error, state whether the titration result will be raised or lowered compared to the result obtained if the error had not been made.
c) Explain why the result is high or low for each error that you detect.

Given a solution of ammonium chloride, what additional reagent or reagents are needed to prepare a buffer from the ammonium chloride solution?

Explain how this buffer solution resists a change in pH when:
a) Moderate amounts of strong acid are added

b) Moderate amounts of strong base are added

c) A portion of the buffer solution is diluted with an equal volume of water.

A sample of 40.0 milliliters of a 0.100 molar $HC_2H_3O_2$ solution is titrated with a 0.150 molar NaOH solution
Ka for acetic acid= 1.8×10^{-5}

a) What volume of NaOH is used in the titration in order to reach the equivalence point?

b) What is the molar concentration of $C_2H_3O_2^-$ at the equivalence point?

c) What is the pH of the solution at the equivalence point?

A solution is prepared from 0.0250 mole of HCl, 0.10 mole propionic acid, C_2H_5COOH, and enough water to make 0.365 liter of solution. Determine the concentrations of H_30^+, C_2H_5COOH, $C_2H_5COO^-$, and OH^- in this solution. Ka for propionic acid = 1.3×10^{-5}

a) A 4.00 gram sample of NaOH(s) is dissolved in enough water to make 0.50 liter of solution. Calculate the pH of the solution.

b) Suppose that 4.00 grams of NaOH(s) is dissolved in 1.00 liter of a solution that is 0.50 molar in NH_3 and 0.50 molar in NH_4^+. Assuming that there is no change in volume and no loss of NH_3 to the atmosphere, calculate the concentration of hydroxide ion, after a chemical reaction has occurred.

Reactions requiring either an extremely strong acid or an extremely strong base are carried out in solvents other than water. Explain why this is necessary for both cases.

$H_2S + H_2O \Longleftrightarrow H_3O^+ + HS^-$ K1 = 1.0×10^{-7}

$HS^- + H_2O \Longleftrightarrow H_3O^+ + S^{-2}$ K2 = 1.3×10^{-13}

$H_2S + 2H_2O \Longleftrightarrow 2\ H_3O^+ + S^{-2}$ K= 1.3×10^{-20}

$Ag_2S(s) \Longleftrightarrow 2\ Ag^+ + S^{-2}$ K_{sp}= 5.5×10^{-51}

a) Calculate the concentration of H_3O^+ of a solution which is 0.10 molar in H_2S.

b) Calculate the concentration of the sulfide ion, S^{-2}, in a solution that is 0.10 molar in H_2S and 0.40 molar in H_3O^+.

c) Calculate the maximum concentration of silver ion, Ag^+, that can exist in a solution that is 1.5×10^{-17} molar in sulfide ion, S^{-2}.

AP

The value of the ionization constant, Ka, for hypochlorous acid, HOCl, is 3.1×10^{-8}.
a) Calculate the hydronium ion concentration of a 0.050 molar solution of HOCl.

b) Calculate the concentration of hydronium ion in a solution prepared by mixing equal volumes of 0.050 molar HOCl and 0.020 molar sodium hypochlorite, NaOCl.

c) A solution is prepared by the disproportionation reaction below.

$Cl_2 + H_2O \Longrightarrow HCl + HOCl$

Calculate the pH of the solution if enough chlorine is added to water to make the concentration of HOCl equal to 0.0040 molar.

AP

A 0.682 gram sample of an unknown weak monoprotic organic acid, HA was dissolved in sufficient water to make 50 milliliters of solution and was titrated with a 0.135 molar NaOH solution. After the addition of 10.6 milliliters of base, a pH of 5.65 was recorded. The equivalence point (end point) was reached after the addition of 27.4 milliliters of the 0.135 molar NaOH.

a) Calculate the number of moles of acid in the original sample.

b) Calculate the molecular weight of the acid HA.

c) Calculate the number of moles of unreacted HA remaining in solution when the pH was 5.65.

d) Calculate the $[H_3O]^+$ at pH=5.65

e) Calculate the value of the ionization constant, Ka, of the acid HA.

AP

Predict whether solutions of each of the following salts are acidic, basic, or neutral. Explain your prediction in each case

a) $Al(NO_3)_3$

b) K_2CO_3

c) NaBr

A solution of hydrochloric acid has a density of 1.15 grams per milliliter and is 30.0% by weight HCl.

a) What is the molarity of this solution of HCl?

b) What volume of this solution should be taken in order to prepare 5.0 liters of 0.20 molar hydrochloric acid by dilution with water?

c) In order to obtain a precise concentration, the 0.20 molar hydrochloric acid is standardized against pure HgO (molecular weight = 216.59) by titrating the OH^- produced according to the following quantitative reaction.

$HgO(s) + 4\ I^- + H_2O \Longrightarrow HgI_4^{-2} + 2\ OH^-$

In a typical experiment 0.7147 grams of HgO required 31.67 milliliters of the hydrochloric acid solution for titration. Based on these data what is the molarity of the HCl solution expressed to four significant figures.

NH_4^+ + OH^- <====> NH_3 + H_2O
H_2O + C_2HSOH <=====> C_2HSOH + OH^-

The equations for two acid-base reactions are given above. Each of these reactions proceeds essentially to completion to the right when carried out in aqueous solution.

a) Give the Bronsted-Lowry definition of an acid and a base.

b) List each acid and its conjugate base for each of the reactions above.

c) Which is the stronger base, ammonia or the ethoxide ion, C_2HSO^-? Explain your answer.

Methylamine CH_3NH_2, is a weak base that ionizes in solution as shown by the following equation.

CH_3NH_2 + H_2O <====> $CH_3NH_3^+$ + OH^-

a) At 25° the percentage ionization in a 0.160 molar solution of CH_3NH_2 is 4.7%. Calculate [OH^-], [$CH_3NH_3^+$], [CH_3NH_2], [H_3O^+], and the pH of a 0.160 molar solution of CH_3NH_2 at 25°C

b) Calculate the value for Kb, the ionization constant for CH3NH2, at 25°C.

c) If 0.050 mole of crystalline lanthanum nitrate is added to 1.00 liter of a solution containing 0.20 mole of CH_3NH_2 and 0.20 mole of its salt CH_3NH_3Cl at 25°C, and the solution is stirred until equilibrium is attained, will any $La(OH)_3$ precipitate? Show the calculations that prove your answer. (The solubility constant for $La(OH)_3$, Ksp= 1×10^{-19} at 25°C)

$Al(NO_3)_3$ K_2CO_3 $NaHSO_4$ NH_4Cl

a) Predict whether a 0.10 molar solution of each of the salts above is acidic, neutral or basic

b) For each of the solutions that is not neutral, write a balanced chemical equation for a reaction occurring with water that supports your prediction.

A buffer solution contains 0.40 mole of formic acid, HCOOH, and 0.60 mole of sodium formate, HCOONa, in 1.00 liter of solution The ionization constant, Ka of formic acid is 1.8×10^{-4}.

b) If 100 milliliters of this buffer solution is diluted to a volume of 1.00 liters with pure water, the pH does not change. Discuss why the pH remains constant on dilution.

c) A 5.00 milliliter sample of 1.00 molar HCl is added to 100 milliliters of the original buffer solution Calculate the $[H3O^{+}]$ of the resulting solution.

d) A 800 milliliter sample of 2.00 molar formic add is mixed with 200 milliliters of 4.80 molar NaOH Calculate the $[H30_{+}]$ of the resulting solution.

AP

The molecular weight of a monoprotic acid HX was to be determined. A sample of 15.126 grams of HX was dissolved in distilled water and the volume brought to exactly 250.00 milliliters in a volumetric flask. Several 50.00 milliliter portions of this solution were titrated against NaOH solution, requiring an average of 38.21 milliliters of NaOH.

The NaOH solution, was standardized against oxalic acid dehydrate. $H_2C_2O_4.2H_2O$ (molecular weight: 126.066 g/mol) the volume of NaOH solution required to neutralize 1.256 grams of oxalic acid dehydrate was 41.24 milliliters.

a) Calculate the molarity of the NaOH solution.

b) Calculate the number of moles of HX in a 50.00 milliliter portion used for titration

c) Calculate the molecular weight of HX.

d) Discuss the effect on the calculated molecular weight of HX if the sample of oxalic add dihydrate contained a non acidic impurity.

a) Specify the properties of a buffer solution. Describe the components and the composition of effective buffer solutions.

b) An employer is interviewing four applicants for a job as a laboratory technician and asks each how to prepare a buffer solution with a pH close to 9.

Archie A says he would mix acetic acid and sodium acetate solutions.

Beula B says she would mix NH_4Cl and HCl

Carla C says she would mix NH_4Cl and NH_3

Dexter D says he would mix NH_3 and NaOH solutions.

Which of these applicants has given an appropriate procedure? Explain your answer, referring your discussion in part (a). Explain what is wrong with the erroneous procedures.

(No calculations are necessary, but the following acidity constants may be helpful: acetic acid Ka = 1.8×10^{-5}, NH_4^+, Ka = 5.6×10^{-10})

Sodium benzoate, C_6H_5COONa, is the salt of the weak acid, benzoic acid, C_6H_5COOH. A 0.10 molar solution of sodium benzoate has a pH of 8.60 at room temperature.

a) Calculate the $[OH^-]$ in the sodium benzoate solution described above

b) Calculate the value for the equilibrium constant for the reaction

$$C_6H_5COO^- + H_2O \Longleftrightarrow C_6H_5COOH + OH^-$$

c) Calculate the Ka, the add dissociation constant for benzoic add

d) A saturated solution of benzoic acid is prepared by adding excess solid benzoic acid to pure water at room temperature. Since this saturated solution has a pH of 2.88 calculate the molar solubility of benzoic acid at room temperature.

In water, hydrazoic acid, HN_3, is a weak add that has an equilibrium constant, Ka, equal to 2.8×10^{-5} at 25°C. A 0.300-liter sample of a 0.050-molar solution of the acid is prepared.

a) Write the expression for the equilibrium constant, Ka, for hydrazoic acid.

b) Calculate the pH of the solution at 25°C.

c) To 0.150 liter of this solution, 0.80 grams of sodium azide, NaN3, is added. The salt dissolves completely. Calculate the pH of the resulting solution at 25°C if the volume of the solution remains unchanged.

d) To the remaining 0.150 liter of the original solution, 0.075 liter of 0.100-molar NaOH solution is added. Calculate $[OH^-]$ for the resulting solution at 25°C.

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