IMPROVING THE UNDERSTANDING OF CHEMICAL EQUILIBRIUM IN SECOND SEMESTER GENERAL CHEMISTRY STUDENTS

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IMPROVING THE UNDERSTANDING OF CHEMICAL EQUILIBRIUM IN SECOND SEMESTER GENERAL CHEMISTRY STUDENTS

A Thesis

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Abstract

of

IMPROVING THE UNDERSTANDING OF CHEMICAL EQUILIBRIUM IN SECOND SEMESTER GENERAL CHEMISTRY STUDENTS

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The concept of chemical equilibrium is central to the understanding of chemistry, and is widely viewed as the most difficult topic to master in the General Chemistry curriculum. The study of chemical equilibrium is at the core of Chem 1B, the second semester general chemistry class at California State University Sacramento. This study attempts to locate the primary reasons for students' difficulties with chemical equilibrium and to suggesthelp for their problems, particularly with regard to the conceptual understanding of thistopic.The action research technique has been employed to characterize this problem, develop and introduce interventions to the student environment meant to improve the problem, and collate and summarize the results of this intervention. It is found that there is a lack of conceptual understanding of the principles underlying chemical equilibrium on the part of general chemistry students, and it is recommended that additional conceptual teaching be added to the general chemistry curriculum at CSUS.

_________________________, Committee Chair
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Date

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Chapter 1

INTRODUCTION

Problem Statement and Thesis

According to the literature, chemical equilibrium is considered to be the single most difficult topic taught in general chemistry.[1] This is reflected in the performance of students in Chem 1B, the second-semester general chemistry course at the California State University at Sacramento (CSUS). Over half of the Chem 1B curriculum consists of topics relating to chemical equilibrium. In a recent Chem 1B class, given in the Fall 2010 semester, the failure rates (students receiving less than a C-, or less than 70%) in the first two midterm exams (those which dealt directly with the subject of chemical equilibrium) were 62% and 61%.

While these figures are unsettling, they only tell half the story. To more fully examine the problem of learning chemical equilibrium, it is necessary to describe two distinct methods of teaching and learning science. In one, the learner acquires the language and formulaic relationships between the symbols of a scientific discipline and methodically applies these to solve problems. This “traditional” methodology is generally used in first-year general chemistry courses such as Chem 1B, and is reflected in the quantitative, problem-solving questions of the midterm exams noted above. In the other method, the learner builds a *conceptual understanding* of the underlying principles of the discipline, and uses this understanding to build new, related learning constructs in memory.

While both of these learning styles are important for acquiring knowledge of a scientific discipline, they are not highly correlated[2], and must both be used for full mastery to be gained. As noted above, the typical general chemistry course is taught using a traditional method, and as a resultmany students learn how to solve problems without a real understanding of underlying principles. In a paper by Ates and Cataloglu, the literature on problem-solving and conceptual understanding styles is reviewed in detail[2]. Among their findings are that 1) A lack of conceptual understanding will hinder student's meaningful learning of a subject, and 2) Traditional teaching methods will not foster this understanding. Students who become proficient in problem-solving "might still lack the desired understanding of fundamental concepts", and "being able to solve standard problems . . . is not a true measure of conceptual understanding". Quotes like this are common in the literature;

> . . . students used the formulas and equations memorized in problem-solving processes by simply plugging in the known variables while trying to solve for the "unknown". If a particular equation did not work then they switch to another equation without trying to understand why they had encountered a problem.[2]

Much has been written about the problem of acquiring a conceptual understanding of chemical equilibrium.[3,4,5,6,7,8,9,11]Its difficulty has been ascribed to a number of unique characteristics. Firstly, it is highly abstract in nature, and has few analogues in the macroscopic world. Secondly, there are many "sub-concepts" related to chemical equilibrium which must also be mastered.[10] Thirdly, the word "equilibrium" is closely related to everyday concepts of equality and balance, leading to confusion on the part of students.

- The Microsoft Word® thesaurus lists "equilibrium" and "balance" as synonyms

- Berquist and Heikkinen cite student confusions with the terms “equal” and “balanced”[8]
- Akkus et al lists students preconceptions of bicycle riding and weighing scales[9]

Students’ lack of conceptual understanding of chemical equilibrium has been well documented in the literature. Terms such as “preconception”, “misconception”, and “alternative framework” have been used to describe this problem.[11] The word *misconception* has been most widely adopted and will be used in this study. In the literature, it is taken to mean “any concept that differs from the commonly accepted scientific understanding”.[11]For example, some commonly held mistaken beliefs of students who have studied chemical equilibriumare; [7]

- That the reverse reaction rate of a reversible reaction is always the same as the forward rate
- That the forward reaction of a reversible reaction goes to completion before the reverse reaction commences
- That, at equilibrium, the concentrations of reactants and products are equal

Given the above inputs from the literature, as well as several years of personal interaction with general chemistry students at CSUS, the author makes this thesis;

1) That many students beginning the study of chemical equilibrium at CSUS lack a conceptual understanding of its underlying principles. This will be supported by the Data Gathering section of the Data Analysis chapter.

2) That conceptual understanding of this complex and important subject can be improved by the integration of constructivist, conceptual teaching into the CSUS general chemistry curriculum. This will be demonstrated through the results of the literature search and through two interventions given to Chem 1B students. These interventions make use of four constructivist-style worksheets generated for this study (further described in the Methodology chapter). Each intervention applies the worksheets in a different way, and evaluations and student feedback of these two techniques are reported in the Intervention sections of the Data Analysis chapter.

Research Methodology: Action Research

Created in the 1940's at MIT, the Action Research methodology[12] was developed for the use of social institutions (schools, unions, political groups) who wished to address perceived problems in their immediate environment. Action Research involves introducing interventions, or small changes in the environment, to see whether those changes are effective at mitigating the problem. Action Research is, as its name implies, oriented towards *taking an action* in the local environment. The methodology of Action Research will be used as the foundation for this study.

Action Research differs in many ways from the traditional research methodology normally used in the scientific literature. It is designed for use in studies which rely on the collection of qualitatively acquired data, non-numeric information gathered from personal observations and interactions like interviews, discussions, and surveillance. Typical practices such as fixing variables, choosing control groups, and using statistical testing are generally not possible in such studies. As Simon and Newell state in *Human Problem Solving,*

> . . . we employ little experimental design using control groups of the sort so familiar in psychology . . . because of . . . the focus on the individual, and the fact that much goes on within a single problem solving encounter, experiments of the classical sort are only rarely useful[13]

Most importantly, the primary goal of action research is to effect improvement in a specific, local environment. Data produced by an action research project must meet a different set of requirements than traditional data. It is not necessarily reproducible, as its focus is local. It is not likely to be published in a peer-reviewed journal, and is not necessarily generated by certified professionals, but rather by teachers and teaching assistants who work in the environment under study. Finally, while the data generated by an action research project may be immediately applicable to the problem at hand, it may also be largely directed to a subsequent research cycle, which will use the data to revise and repeat the previous study. The primary requirement to be fulfilled by action research data is whether or not is good enough to drive the next step towards solution of that local problem, whether that solution be making major changes or instituting new research. An action research project may last for a number of such cycles, each one making slight improvements in the environment while at the same time providing inputs for subsequent cycles. The research conducted in the present study may be viewed as a sequence of steps which follow the principles of action research. A much more detailed review of action research can be found in the Methodology chapter.

Research Logistics

The following section is intended to give an overview of the research logistics used in this study. A much more detailed description of the research logistics can be found in the chapter on Methodology.

Step 1: Data Gathering (Fall 2009)

Baseline data from Chem 1B students on their conceptual understanding of chemical equilibrium was gathered. This took place in two stages. The first stage consisted of the development and administration of a Diagnostic Exam, taken by entering Chem 1B students in the first week of class. This exam was created by the author using inputs from many similar studies in the literature. The second stage was a detailed review of the first two Chem 1B Midterm Exams, which were taken during the course of the semester. These exams were given by the Chem 1B professor as means of performance assessment. Graded student exams were provided by the professor, and answers to exam questions were analyzed for possible instances where student misconceptions might have affected the results. Only the first two midterm exams concerned chemical equilibrium, and thus the third was not considered. This data can be found in the chapter entitled “Data Gathering Methodology, Results and Analysis”.

Step 2: First Intervention: Focus Groups (Spring/Summer 2010)

Using input from the previous Data-Gathering step, an intervention was generated in the form of four worksheets developed to help improve student understanding of chemical equilibrium concepts. The worksheets were administered in small “Focus Groups”; instructor-guided inquiry sessions which made use of constructivist exercises to improve students’ conceptual understanding. Figures were presented, questions posed, and discussions held based on the worksheet contents. Data gathering was done by the author/instructor and was recorded in the form of meeting transcripts of each session.

Exercises were generally conducted in a qualitative manner, with student calculations rarely requested. Accordingly, the data captured was itself largely qualitative in nature, consisting mainly of student inputs to the discussions (“I don’t understand . . . I think . . . No, this can’t be because . . .) as well as the instructor’s observations and recording of students’ behavior and attitude (confidence, confusion, articulation, interest, boredom, etc.) Meeting transcripts are reproduced in Appendix D, and are summarized in the chapter on chapter entitled “Intervention 1 (Focus Groups) Methodology, Results and Analysis”.

Step 3: Second Intervention: Worksheet Completion (Fall 2010)

The second intervention again made use of the worksheets developed for the first intervention in Step 2 (the Focus Group), but in a slightly different format. Instead of the worksheets being used to anchor face-to-face instructor-guided sessions, the worksheets were revised and posted on the Chem 1B website for student download and completion. With no instructor present, it was necessary to modify the worksheets, adding instructions toclarify questions and problemsas well as background review on topics which had been poorly understood in the Focus Groups. Students were asked to provide written answers to worksheet questions which might have called for verbal discussion as part of a Focus Group session.

Groups of students (ranging between 10% and 20% of the total class) voluntarily downloaded, completed and returned the worksheets in return for a small amount of course credit (two points per completed worksheet). Worksheets were completed and returned to the author after the material on the worksheet had been covered in class but prior to the midterm exam on that topic. Analysis of the midterm exams was then performed to compare the responses of the two different student groups (worksheet users and non-users) to selected exam questions.

Note that this intervention involved the study of Chem 1B midterm exams, which were primarily quantitative in nature and relied heavily on students' problem-solving skills. Since the skills of problem-solving and conceptual understanding are not fully correlated[2], the author makes no claim for the impact of improved conceptual understanding on Chem 1B exam scores.

A detailed description of this Intervention can be found in the chapter entitled "Intervention 2 (Worksheet Downloads) Methodology, Results and Analysis".

Chapter 2

LITERATURE: HOW PEOPLE THINK AND LEARN

A Brief History of Science Teaching in the US

General chemistry has been taught in the United States university system since the mid-19th century, and style for nearly as long at the secondary school level[14]. In the beginning, training for future chemistry teachers consisted entirely of training in the science of chemistry itself. No discrete training in teaching was provided; it was assumed that the student who excelled in the study of chemistry would then excel as a chemistry teacher[15]. This began to change around the end of the 19th century, as training in the teaching of science began to become available in university science departments around the country. In the last decades of the century, many universities made strides towards the inclusion of pedagogical instruction for future science teachers as a part of their curricula. In 1889, Clarke University offered a course in “Special Pedagogy”, aimed at the future chemistry teacher, and described as follows[16]:

> ***Chemistry***
> 1: Lectures (5); Topics: Methods of Teaching Chemical Science
> The history of the development of modern methods. Fit adjustment of theoretical and experimental teaching and the value of each as modes of mental discipline. How far chemistry can profitably be taught in the secondary and primary schools. Best means of securing enduring results. The best and most economical ways of installing and furnishing a laboratory, and the precautions required to insure profitable and safe work from immature students. Suggestions on the preparation and delivery of experimental lectures. The best system of teaching the three main branches of elementary chemistry, viz.: general descriptive chemistry, mineralogy, or the natural aspects of chemistry and quantitative chemical analysis.

As can be seen, in the early days of science pedagogy, the focus was mainly on practical logistics and how to most “profitably” and expeditiously convey the material at hand. In the early twentieth century, the number of pedagogical classes in the sciences grew, with Ohio University, Columbia, Cornell, University of Michigan, and Harvard among others announcing offerings in this subject[17]. The primary focus of these classes was on the epistemological content of the subject and the logistics needed by the science teacher to successfully transfer the required course knowledge, rather than on the process of learning itself. In a 1933 paper in the American Physics Journal, physics professor F.K. Ritchmyer of Cornell University expressed the sentiment behind this state of affairs in a paper entitled "Physics is Physics".[18]"Teaching . . . is an art . . . I believe that one must admit that in no sense can teaching be said to be a science". Ritchmyer quotes a colleague's advice on becoming a good mathematics teacher; "To become a successful teacher of mathematics one must acquire a thorough knowledge of mathematics". As recently as 1978, Ralph Goodwin, a physics professor at the U.S. Naval Academy, dismissed attempts to infuse studies on the development of human reasoning into the practice of science teaching as a "fad", advocated the use of traditional behaviorist teaching techniques, and referred to the theories of educational psychology as "hocus-pocus" and "mistaken logic"[19].

Given the objective nature of their discipline, it perhaps is understandable that the science teaching community might be reluctant to accept the changes in attitude towards the teaching of science which began to take hold in the 20th century. Some islands of change could be seen, though. This change was exemplified by the chemistry program at the University of California[20], which encouraged entering students to take high school chemistry, made use of open-book examinations to relieve the need for memorization, and in general served as a beacon for change in the fields of teacher training and student instruction. In 1922, Dr. Joel Hildebrand, a professor of chemistry at the University of California at Berkeley, decried the current state of science education in the U.S. in an article in *Science*[21]. Professor Hildebrand believed that the naturally inquisitive and experimental nature of young people was ill served by the "repressive" (to quote Hildebrand) science teaching methods then popular, which tended to channel students away from observation, experimentation and reasoning and towards repetition, rote, and the copying and parroting of scientific dogma;

> . . . we are so anxious to guard (the student) from errors of fact that we announce in advance what he is expected to find in his experiments. He is told to mix solution A with solution B and 'to note the red precipitate which is formed.' The precipitate he gets may happen to be yellow, but he has learned by this time that is safer to call it red in his note-book[21]

Professor Hildebrand's thoughts predated a wave of change in the way in which the processes of teaching and learning were regarded in the science community. These changes were driven by the wave of theory and research from Piaget, Vygotsky, and many others. By the end of the twentieth century, the science of cognitive psychology, and the ideas of the constructivist school, had taken a firm grip on the educators' view of their vocation.

Overview of Literature Search: "A Science of Science Teaching"

The title of this introductory section is taken from a key volume in the history of the study of science teaching. "Towards a Science of Science Teaching", by Michael Shayer and Philip Adey, was published in 1981[22]. Shayer, a Professor of Applied Psychology at Kings College London, and Adey, Professor of Cognition, Science, and

Education at Kings College London, began to collaborate on the study of cognitive development and its applications to educational theory in the mid-1960's, and produced a considerable volume of literature on the subject. Their work began just as the field of cognitive psychology was coming into its own, and as the theories of Piaget regarding the development of human learning capabilities and the process of learning itself were starting to work themselves into the mainstream of educational study. Prior to the middle of the 20th century, when Ayer and Shadey's work began, there was little interest in the study of the process of learning. Their pioneering work in this area, summarized in "Towards a Science of Science Teaching", is a significant example of the development which has been made in the study of science education.

The study of human cognition has involved the merging of work from a number of disciplines, among them psychology, computer science, philosophy, neurology and biology. This work will be summarized here in four sections. First, the theory of human learning developed by the Human Problem Solving and Information Processing communities will be reviewed. Secondly, theories of human cognitive growth from the field of Cognitive Psychology will be summarized. The merging of these findings into a new school of educational philosophy, Constructivism, will be examined. Finally, some classroom studies which make use of these new ideas will be presented.

A) Human Learning

Simon and Newell and "Human Problem Solving"

Published in 1972, *Human Problem Solving* by Allen Newell and Herbert Simon[13]details the results of almost twenty years of research and extensive experimentation into, as they put it, "how people think". Beginning with a theory of how and to what extent thought, learning and performance are related, the book summarizes the authors' findings and conclusions from extensive work at Carnegie-Mellon University beginning in the late 1960's, and proposes a theory and a set of mechanisms for how humans learn and solve complex problems.

One of the key concepts elucidated by Newell and Simon is that of the "Problem Space"[23]. As per Newell and Simon, after initial recognition and understanding of a problem, the solver constructs a Problem Space (a mental framework containing a number of diverse elements required for the solution of a problem). The key contents of the Problem Space are

1) Elements (symbols and structures)
2) Operators (possible data transformations)
3) An initial knowledge state (initial state of the problem scenario)
4) Problem statement (desired outcome)
5) Total available knowledge (both direct and indirect)

The Problem Space can be thought of as a "program" which can be stored in memory, allowing its reuse for multiple problems. The problem solver must "construct" the Problem Space every time a problem is approached. The more complete, correct constructions already existing in memory, the more robust the Problem Space will be, and the greater will be the likelihood of problem solution.

Computer Science and the Information Processing Model

Another key concept developed by Simon and Newell relates to the high-level architecture of the human mind with respect to cognitive behavior. Dubbed the Information Processing System (IPS), this architecture, by the author's own admission, gained its name and much of its definition from the science of information processing which was, just as their research was being published, being renamed as "computer science";

> An abstract concept of an information processing system has emerged with the development of the digital computer . . . in this book we will introduce a suitable abstract information processing system to describe how man processes task-oriented symbolic information.[13]

For introductory purposes, a simple depiction of this architecture appears in **Figure 1**[24]. (A more complex model will be introduced below in **Figure 2**.) In this model, sensory input enters the learner's mind and is immediately filtered according to its relevance and importance. Irrelevant information is forgotten; relevant information is transferred to the memory system. The new information is further filtered in short-term memory (STM), a temporary storage area for incoming memory. If this temporary memory is deemed relevant and is understood within the context of existing memory, it will be permanently stored in long-term memory (LTM). The IPS model has become widely accepted by the cognitive research community.

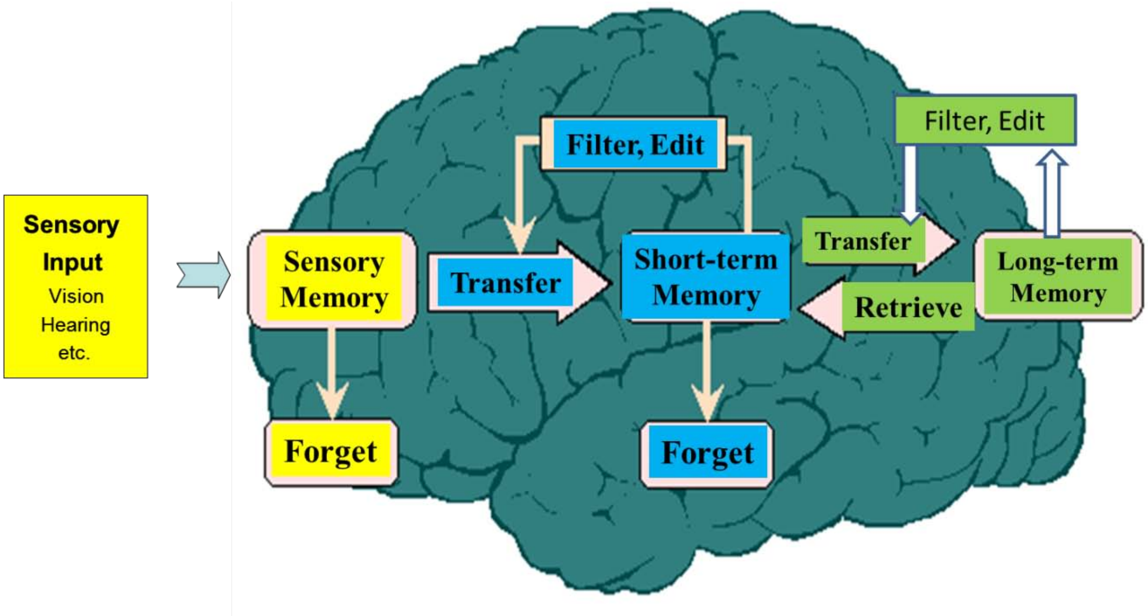


Figure 1: Overview of Information Processing Model

By the end of the 1970's, researchers were able to report much more quantitative descriptions of the IPS than were present in Newell and Simon's text. The following summary of the IPS model is from David Hestene's 1979 paper, "Wherefore a Science of Science Teaching"[25].

Sensory Input: Sensory receptors in the brain take inputs from the senses. This input is stored for a limited amount of time before disposal; about 0.5 seconds for visual input and 3 to 4 seconds for auditory input. (Note, however, that *recognition* of visual data is very fast; about 30 to 40 milliseconds.) Thus, information must be quickly transferred into short-term memory (STM) or it will be lost.

Short-Term Memory (STM): The STM has a finite capacity for information; it can hold only ~6 symbols (or "chunks") at a time. A chunk can be a letter, word, number, or concept. Retention time in STM can be up to 30 seconds[26]; however, information is more commonly lost here by displacement than by expiration. Information loss from the STM is permanent; Hestene's example is that of the failure to retain a new acquaintance's name within ~10 seconds after hearing it. Information transfer can thus be said to have a maximum possible "rate", limited by the time it takes to get information chunks out of STM and permanently retained in LTM. It is in the STM that the use of the "mnemonic device" is useful.

Long-Term Memory (LTM): The LTM is the "hard drive" of the brain. There is no indication that it ever becomes full over the course of a human lifetime; it's the often demanding process of *retrieval* of LTM data which makes its storage capacity sometimes seem to be much less than it is. Some details regarding the transfer of information to the LTM should be further explored, as they are central to the learning model. It should be noted first that STM-to-LTM transfer involves more than simple movement of data. Data transfer depends on the nature of the data as well as its relation to existing data in the LTM. For example, simple operations like comparisons or replacement of data involve only a single chunk and can be executed fairly quickly. More complex information may require several memory chunks and a number of serial transfers and require a longer time to process. In particular, new information for which no existing context exists in LTM (that is, for which there are no pre-existing chunks) may have what Hestene calls a "slow write time", requiring modifications which can result in STM-to-LTM transfer rates of up to 10 seconds.

Note that information stored in the LTM is just a combination of the "symbols, structures, operators and data" mentioned above as elements of the Problem Space and is thus a major component of that concept. The concept of the LTM program will continue to be used in future discussions of the cognitive mechanism and of constructivist teaching theory in this paper.

A full understanding of the slow write time problem requires a new concept to be introduced to the IPS model; that of "software" content in the brain in addition to the "hardware" content described above. As opposed to the hardware of the STM and LTM, which is largely present at birth, a large part of the IPS model consists of large blocks of memory saved in the LTM over time through exposure to sensory inputs from one's environment. One can think of these blocks as software "programs" being developed and stored in the LTM; these programs consist of a combination of objects (symbols, numbers, etc.) combined with processing patterns used to absorb previous chunks of memory. Just as a computer program is developed as a flexible algorithm which is able to take inputs from observations as well as direct inputs from the programmer and process these inputs to provide new information, the model of the LTM program is that of an algorithmic construct which can accept new inputs and use them to produce new information. The difference is that, in the case of the LTM program, its structure has been built over time through exposure to many previous STM inputs, and will continue to be actively modified by the IPS process over the lifetime of its owner. According to the IPS model, then, the brain can be thought of as a self-teaching program developer, creatinglearning constructs as it is exposed over time to sensory inputs from the outside world.

B) **Cognitive Development**

Piaget: A Paradigm Shift in the Science of Learning

While the theories of human learning were being developed, a parallel effort in the field of developmental psychology was taking place. The object of this work was a theory of how the capability for human learning was developed over time in the individual. Arguably the single most influential work in this field came from the Swiss developmental psychologist Jean Piaget, who is widely credited as the founder of a new scientific discipline, that of Cognitive Psychology. As a young teacher in a school for boys, Piaget began to notice patterns in the answers to certain questions, and trends in how these patterns changed over time. During the course of his career, Piaget codified his observations of these patterns into a coherent model of how the capability for learning evolves in human beings over the course of time, and in doing so was primarily responsible for creating a new field of study, that of “cognitive psychology”[27]. As time progressed, Piaget’s psychological studies of cognitive development would come to play a major part in the way that science teachers thought about their profession. Piaget was included in the list of Time Magazine’s “100 Most Important People of the Century” in 1999. Albert Einstein referred to Piaget’s study of the learning processes of children and young adults, and his theory of cognitive development, as “so simple that only a genius could have thought of it”[28]. Beginning in the 1950’s, Piaget’s work catalyzed a wave of research, which continues to this day, into how human beings develop the capability to learn. A number of Piaget's key ideas which have particular relevance to the subject at hand will be reviewed below. It should be noted that, as has been the case with many drivers of paradigm changes in their fields, the work of Piaget has been heavily reviewed and its basic tenets generally contested and reworked to the point where, in large part, only its core principles remain intact. These principles, however, still serve as the linchpins of educational psychology today, and have served as touch points for the review and research analysis presented in this paper.

Beginning with his experience as an elementary educator, and continuing into his career as a university professor and through his position as director of the International Bureau of Education, Piaget accumulated a body of research into the psychological development of the human capability to learn. His research consisted of introducing concepts to children and adolescents and interviewing them to gain insights into their understanding of these concepts and the processes which they used to gain that understanding. A review of Piaget's work shows extensive documentation of these interviews, with which he worked to develop a generic theory of cognitive development. The major output of this work was Piaget's proposal that one could roughly define the course of the cognitive development of a human being in terms of a series of stages. Each "stage" was defined by a number of cognitive capabilities, and could be roughly correlated with a particular age range. As the initial stages of development are attained by children at a very early age, for our purposes only the last two of these stages are of interest (see **Table 1**).

The initial three stages will not be discussed here, as they are generally presumed to be minimal requirements for the study of general chemistry. The final developmental stage, the Formal Operational stage, is of the most interest to the study of teaching and learning in first-year college-level general chemistry. In the Formal Operational stage, the ability to form abstract constructs, or rules, which govern the behavior of observed phenomena, is developed, and the capability to apply these abstract constructs to other analogous situations is acquired. Also present at this stage is the ability to assess one's own learning style, and pursue those learning modes best suited to that style; this is called "meta-cognition". And finally note that this stage is reached at a different rate for different disciplines; the individual who attains the formal operational state in mechanics may never reach that stage in chemistry.

Table 1: Piaget's Developmental Model

Piaget's Cognitive Development Stages

Stage	Age	Description
Sensorimotor	0 - 2	Reflex based; coordination developed
Preoperational	2 - 7	Consciousness of self; egocentric
Concrete operational	7 - 12	Can assume multiple viewpoints; No abstract thought
Formal operational	12 =>	Abstract thinking Theoretical reasoning Meta-cognition Specific to field of learning *Not necessarily ever reached*

Much study and discussion of the relationship between the attainment of the Formal Operational stage and the ability to master the concepts of general chemistry can be found in the literature. One universal conclusion of these studies is that, in order to master the concepts taught in general chemistry, one must have attained the early, and preferably the mid or later, Formal Operation stage. A rough chronology of cognitive development for the movement through the Operational Stages was suggested by Piaget and adopted by many researchers. In the original timetable, it was proposed that the transition to the final Formal Operational stage begins somewhere around the age of 12 and is complete by the end of adolescence. It should be noted, however, that the timeline for the attainment of the Formal Operational stage is very much dependent upon the nature of the subject matter involved. For a complex subject such as chemistry, this stage can be attained at a much later date, and is often never attained at all. Shayer and Adey, for example, found that a minority of secondary school upperclassmen in the U.K. had reached the Formal Operational stage[22]. Accordingly, we can reasonably expect that a significant number of first-year university students may still be unable to function at this level.[29]

Assimilation, Accommodation, and Constructivism

Another major principle to come out of the work of Piaget is a hypothesis of how new information is encountered and integrated by the learner. Two terms were coined to describe this process. *Assimilation* refers to the wholesale, unaltered transfer of external data into the internal cognitive structure of an individual. In this simple transfer, no review or modification is performed. However, assimilation is always usually accompanied by a revision process, known by this model as *accommodation*, in which the incoming data must be “accommodated” to existing data and concepts already present in the learning mind, or vice versa. If existing concepts are incompatible with the incoming data, or if there is no context suitable for its integration, something has to give.

The new data is either not accepted as is, or changes are made in the preexisting internal structure; that is, either the data is altered to *accommodate* the existing structure, or the existing structure is altered to *accommodate* the data[30]. Assimilation and accommodation take place simultaneously in a continuous process which determines what information gets integrated and how this integration impacts the internal cognitive structure. This process results in a constant formation, or *programming*, of new structures, or *chunks*, of knowledge.

Piaget's theory of assimilation and accommodation suggests a connection back to the discussion of the Information Processing Systems model in the previous section. This theory is in effect a refinement of the IPS model, and its addition to the model allows its expansion to a version which more directly represents the process of human learning. Although Piaget's precise terminology has lost favor, his theory remains in use. The use of this refinement has become widespread within the research community; one such version appears in the diagram in **Figure 2** below. This figure was drawn from Osborne and Wittrock's "Learning Science: A Generative Process"[31]; they termed the process it describes the "Generative Learning Model".

A review of Figure 2 provides a detailed outline of this process. The learning process begins with the acquisition of a sensory observation which the learner chooses to retain; the authors refer to this as "selective perception" (1). Once the information is stored in the sensory registers (2), the process of transfer to memory begins. Existing information in LTM can influence whether or not the new information will be permanently absorbed (3). Once approved,the new information is then used for “construction” in the LTM (4), and after a final validation against the original sensory input (5), the final construction can be added to (6). Note that “unsuccessful” construction can lead to additional cycles in which additional sensory information may be required, along with more data from the LTM, in order to attempt successful integration of the new information (7).

As illustrated in **Figure 2**, a high degree of interplay exists between the sensory input and the various levels of memory during the learning process before it is approved for LTM storage. This interplay is essentially the process of assimilation and accommodation articulated in the cognitive theory of Piaget. The successful review and LTM storage of new data is referred to by the authors as “meaningful understanding”; its form is that of a “successful construction”. These terms, and especially the term “construction”, have found extensive use in the study of cognitive development.

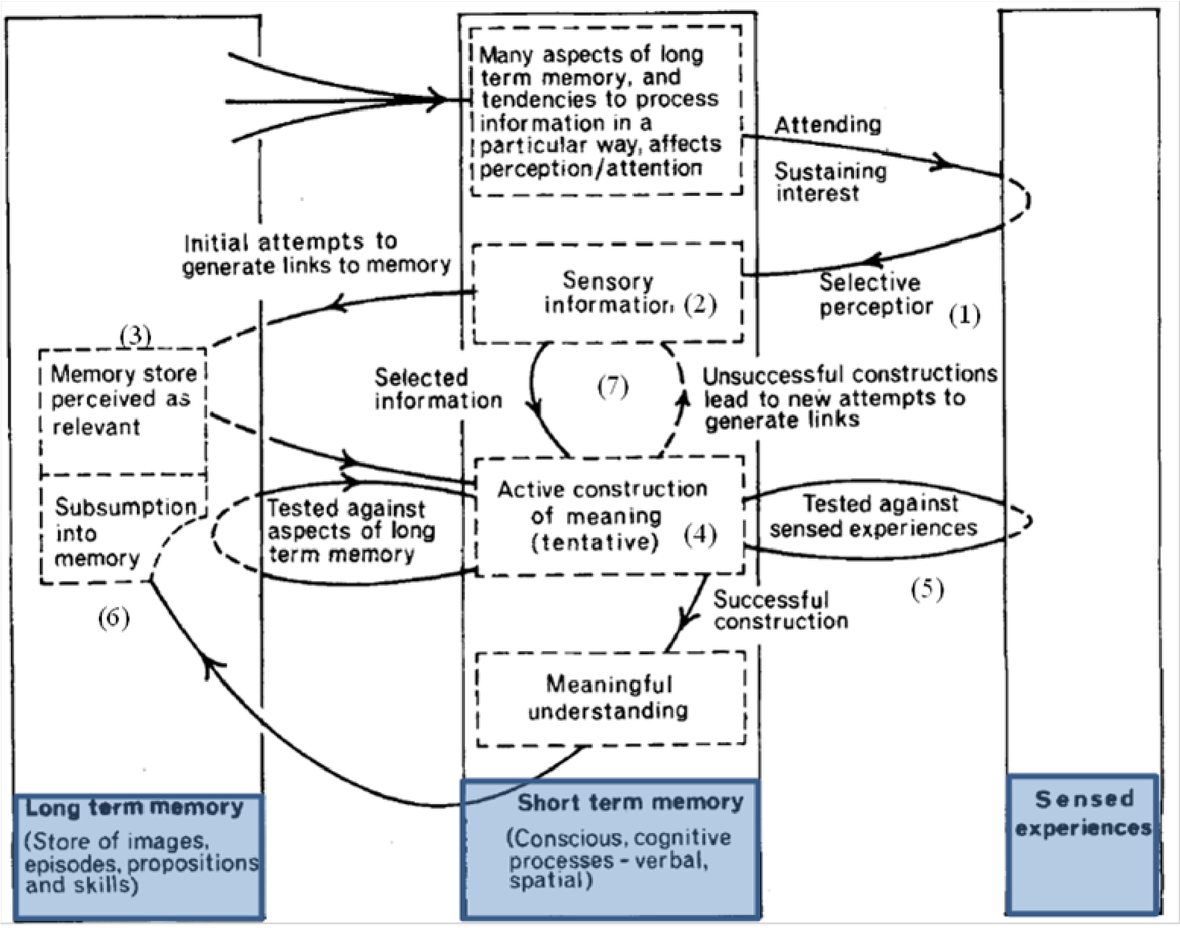


Figure 2: Schematic Representation of the Generative Learning Process[32]

C) **Developments in the Philosophy of Teaching**

The Constructivist Model of Teaching and Learning

The schematic in Figure 2 makes clear the parallels between the concepts of STM/LTM "revision" in the IPS model and Piaget's assimilation/accommodation model. In both models, learners go through an often repetitive sequence of interactions with their environment in order to permanently store new information in LTM. In an enduring metaphor now used widely by the research community, this new information is termed a "construction", and it is built both with elements of sensory perceptions from the external world and existing chunks of memory from the LTM. The term is especially apt, as it implies to both learner and instructor that, as a construction must be *built,* it will require an expenditure of energy on the part of them both. Knowledge is conveyed and acquired actively, not passively.

Driven by the core principles described above, the term "constructivism" had such strong implications for both teaching and learning that it has become well known as a school of academic thought whose teachings have become pervasive in the field of cognitive psychology in the past several decades. The Constructivist school of study is the child of both Piaget's theory of cognitive development and the IPS model. "Constructivism is so pervasive within developmental psychology that it might seem both obvious and inevitable, and thus independent of Piaget's influence"[33]; however, the concept of the interactive nature of knowledge acquisition and of the active nature of learning was described by Piaget as "Piaget's Theory";

> This leads us to (an) idea central to the theory, that of *construction*, which is the natural consequence of the interactions we have just mentioned . . . objective knowledge is not acquired by a mere recording of external information but has its origin in interactions between the subject and objects.[34]

A strong message from the work of Piaget, IPS, and the constructivists is that students arrive in the classroom not as empty slates ready to be written on, but with a collection of existing conceptual constructs which must serve as the foundation for the construction of new knowledge. In the process, some portion of these existing constructs may need to be modified or replaced in order to accommodate new information. For students to exit a class with a greater body of knowledge and expertise than they had several months earlier, some serious work needs to be done by both the student and the teacher. Faulty constructs must be identified, incorrect information or processes need to be excised, and new ones installed in place of the old.

Although Piaget was never known as a "constructivist", the ideas which make up the school of constructivism can be largely traced to the theories of Piaget and the IPS movement, and much of the language of constructivism can be related back to that of these pioneers of developmental psychology. Consider the word "constructivism" itself; this is a direct reference to both the IPS model of STM/LTM interaction and to the Piaget model of assimilation and accommodation. Each of these models shares a core concept; that *learning is* ***constructed*** *in the mind of the learner*.

The constructivist theory has been widely adopted by educational professionals, in part due to its relevance and applicability in the classroom. If existing "constructs" in the classroom support the addition of new ideas and concepts, then one's job as an instructor will be much easier; if they do not, then successful learning may not occur. In Piaget's language, the necessary "accommodation" may not be achievable, or, in IPS language, the new concept will never get into LTM.

> Educational influences . . .can have some effect on the subject only if he is capable of assimilating them, and he can do this only if he already possesses the adequate instruments or structures[34]

Given the central tenets of the constructivist school, and its powerful influence on the world of cognitive psychology, we can now begin to look at the "science of science teaching" from an application level. What are the specific issues with which science teachers are confronted? How will the existing cognitive states of their students affect their teaching performance? What can they do to address these issues? In the next few sections the literature will be examined with a view back to the three touchstones of educational theory: the information processing model; developmental psychology and Piaget's cognitive development model; and the concept of constructivism.

Cognitive Assessment and Cognitive Matching

If we accept the principles of constructivist teaching, then it should be agreed that new knowledge must be built on top of existing memory constructs. Before anything can be done to improve performance in the classroom, an instructor must obtain some idea of the level of cognitive function of entering students and perform a preliminary screening to ensure that the incoming cognitive levels are a good match for the material to be taught. Teachers who handle hundreds of students from several different classes in the course of a single semester are confronted with a seemingly impossible task. How can they know what prior knowledge and understanding are being brought into the classroom? How can they address and correct the misconceptions in a practical manner? And how can they do this while at the same time completing the tasks allotted to them by the existing curriculum, ensuring that the students have acquired the material they will need to move on to the next level of study?

Many researchers have found wide gaps between the cognitive level of incoming students and the level required of them by the curriculum. Recall Shayer and Adey's finding that only a third of secondary school students at the junior and senior level had attained the Final Operational stage[22]. It is therefore useful to understand not only what incoming students know, but how capable they are of learning the material to which they will be introduced. In the work of Shayer and Adey cited above, extensive testing of students' cognitive levels as per the Piaget model was undertaken. Tests were written with questions designed to elicit replies which revealed the student's level of cognitive development. Results of these tests were then compared with assessments of the levels of development required by the course curriculum currently in use. Shayer and Adey termed this process "cognitive matching", referring to the match between existing students' cognitive level and the level of the material to be taught.

A particularly comprehensive study of cognitive matching was performed by the government of Botswana in the mid-'90's, when a government resolution was passed which dramatically altered the educational goals of its public education system, particularly with regard to secondary school chemistry[35]. As a part of the redevelopment of the country's science curricula, widespread testing of cognitive levels in the student body was performed and compared with estimates of the level required by the existing curriculum. Wide gaps between existing and required cognitive performance were identified.   The authors of this study pointed out that, while Piaget had supplied a great deal of detail in his descriptions of cognitive development and constructivism, "he was not clear about the distinction between these two" and was "elusive about the mechanisms through which knowledge was constructed"[35]. Does development foster the creation of more complex constructs, or vice versa?

The work of many researchers contributed to the body of knowledge on the concept of cognitive matching; however, no one took the principleso closely to heart as did the pioneering Russian psychologist L. S. Vygotsky.

Vygotsky's ZPD and Social Constructivism

In the field of cognitive psychology, Vygotsky was at once a precursor, a contemporary, and a successor to Jean Piaget. Born the same year as Piaget (1896), he died much earlier (of tuberculosis in 1934, at the age of 38), and did much of his most important work in the last decade in his life, at the time when Piaget was also doing much of his work. However, because of political conditions in USSR at the time, most of Vygotsky's work was not published in the West until much later.

One of Vygotsky's enduring contributions to the science of cognitive science was made in the area of cognitive matching. Vygotsky's *Thinking and Speech*[36](published in the US as "Thought and Language"), written in the early 1930's and translated into English only in 1987, is concerned with the mechanisms with which learners gain their developmental skills and acquired new knowledge. Much as Shayer and Adey used the term "cognitive matching" to the alignment of a student's cognitive level with their instructional curriculum, Vygotsky describes the concept of the difference between a learner's existing ("pre-learning") cognitive development and the level of the new information they are being asked to acquire. Named by him as the "zone of proximal development" (ZPD), this difference is defined as the gap between learner and environment (a book, a lab, or most effectively, a teacher) which MUST exist in order for cognitive development to be made. Moreover, this gap must be just the right size. If the gap is too wide, the student will be incapable of taking in the new learning; if it is too narrow, no cognitive development will take place. As per Vygotsky,

> Instruction is maximally productive only when it occurs at a certain point in the zone of proximal development.[37]

Vygotsky defined the ZPD as existing between a student and an *experienced instructor*. Vygotsky placed a great deal of importance on the role of the instructor. Although he strongly believed that "the student educates himself"[38], he was convinced that, without a teacher/guide, no development could really take place. The teacher determines the size of the ZPD gap, and guides the student across the ZPD, facilitating an increase in their cognitive level;

> This is why an active role is the lot of the teacher in the course of education. The teacher fashions, cuts, shreds, and carves up the elements of the environment and then combines them in the most diverse way imaginable, in order to realize whatever goal he has need for. Thus is the educational process an active one; the student is active, the teacher is active, and the environment created between them is an active one.[39]

For Vygotsky, *the teacher is the agent of cognitive development*.

Vygotsky's best-known work, and that which distinguishes him from Piaget, was on the role of the learner's environment (or "culture", as he often puts it) on the process of cognitive development. While Piaget's work was primarily focused on the evolution of the *individual* throughout successive developmental stages, Vygotsky emphasized the impact of the environment on cognitive development. For Vygotsky, the environment served two critical functions in this role; it provided a source for raw material from which a student's earliest constructions (later to be modified) could be made, and secondly, it provided a testing ground against which cognitive constructions could be tested; in short, it provided the means to construct and the materials with which to construct. Learning happened only with interaction between the learner and their environment.

A caveat here; for purposes of educational research, the environment influences on a learner can be divided into two groups; the "learning environment" (teachers, curriculum, other students, etc.) and the “extracurricular environment” (locale, family, socioeconomic factors, etc.). The former group is largely accessible and controllable by the educational researcher; the latter group is not. While extracurricular influence is recognized as a major factor in the literature, it is generally acknowledged to be outside the field of study.

The combination of Piaget’s developmental stage theory with Vygotsky’s work on the role of the social environment came to be referred to as “social constructivism”. Although much of Vygotsky’s work predates that of Piaget, Vygotsky’s popularity and influence in the community of cognitive psychology came into fashion only in the past several decades. The currency of social constructivism in the field of educational psychology owes much to the pioneering work of Vygotsky.

Social Constructivist modifications to Piaget’s theory can be summarized in four parts;

1. Cognitive development takes place within a particular field of knowledge and does not ensure similar development in other fields
2. Both the content and the accessibility of the environment play a key role in learning
3. Cognitive development is a function of the socioeconomic environment of the learner
4. Progression through Piaget’s Stages can be a gradual one and is unique to each individual

As can be seen, these modifications add a new dimension to Piaget’s theories, that of the learner environment. The first point is the most interesting to the educator; it suggests that there exists some relationship between cognitive development and the particular topics introduced to the student. In fact, a number of studies have reported on “the positive benefits of using content in science to bring about marked improvement in development”[35].

D) The Evolution of Teaching Practice

Modern Science Teaching Techniques: Sample Studies

Some studies have reported on attempts to bring cognitive matching into the classroom environment. Bass and Montague[40], in their study of cognitive matching and the conscious introduction of cognitive development into the science curriculum, attempted to relate development and curriculum, and reported that, after several days of stage-targeted instruction of two basic physics problems, they were able to raise students’ cognitive levels by one substage. Notable is their conclusion (reflecting that of Vygotsky) that, in order to foster this development, “authentic modification in the child’s thinking can probably best be accomplished through the *active intervention of a well-qualified teacher*”[40] (my italics). If neither the curriculum nor the status of incoming student’s cognitive capability can be changed, the instructor’s own teaching style must be examined.

Many studies involve the use of various constructivist-style teaching methodologies in the classroom, including the use of guided inquiry, group study, and conceptual and analogy-based teaching (a number of these studies are described and referenced at the end of this section). In the social constructivist school described above, “knowledge construction is a social and cultural process mediated by language, (and students) develop shared conceptual understanding via dialogue within a community of learners”[41]. One study of the classroom use of social constructivist practice is particularly interesting; drawing from the focus on the interaction between student and instructor so valued by Vygotsky and his followers, the study focuses not on course curriculum but on the instructor herself A secondary school teacher of grade 12 chemistry with 23 years of experience in teaching a “traditional” (non-constructivist) curriculum (lecture, lab, exam) collaborated with the authors over a period of two years to revise her teaching technique to one which incorporated a more constructivist style. Through this period, the authors not only examined the results on classroom learning, but reported on changes in the instructor (“Frances”) herself and her own evaluations on the progress of the students and of herself.

The authors provided two metaphors to Frances as high-level guides to the style they asked her to adopt. One was “teacher as learner”, or that the teacher needed to learn about the students’ existing views and understandings as the class progressed, and to alter her teaching style and content accordingly. The other was “teacher as modeler”, or that, as she was now asking students to think and collaborate in ways which were mostly new to them relative to the traditional ways of learning to which they were accustomed, she should model these new processes herself in the classroom in order to more effectively convey to them the behavior that she sought.

> We were trying to assist Frances to seek explanations from students regarding their observations while also developing their awareness of changes in the types of explanation required of them . . . to encourage students to develop their own understanding and to be able to explain phenomena using explanations involving their understanding of particle theory.[42]

The topic upon which the authors chose to focus their study was that of chemical equilibrium. A number of studies report that, of the topics found most difficult to master in general chemistry (the mole concept, reaction stoichiometry, oxidation/reduction reactions, and chemical equilibrium), chemical equilibrium was rated as *the most* difficult[43]. Two major reasons for students' problems with these topics are suggested; first, they require a high degree of abstract reasoning, and second, they all use language which is also used, with different meanings, in everyday life. With regard to chemical equilibrium, one might add a third reason; as opposed to many chemistry concepts, it has few analogs in common experience which might be used to foster its understanding with students who are newly encountering it. To teach the mole concept, one can refer to terms like "dozen", or with stoichiometry, to cars with fixed number of wheels, or utilize one of many hands-on exercises developed to convey understanding. No such clear analogy exists for the teaching of chemical equilibrium.

As tools, the authors developed lesson plans for Frances which exposed students to the basic processes behind chemical equilibrium. They also coached her in the method of guided inquiry, with its treatment of the student as an active learner who acquires knowledge through observation and discussion. In the method, the teacher acts as a guide whose purpose is to set students on the right track and keep them on it through a process of Socratic dialogue, inserting themselves with directional input only when they have run aground.

Frances resisted this approach for some time, being quite happy with her performance after 23 years of teaching. Frances was initially "fairly satisfied with what happens in the classroom . . . I'm pretty much pleased with the range of activities I use"[41]. She was not entirely convinced of the value of the guided inquiry method;

"Sometimes I just tell them 'Look, this is what scientists believe' . . . it's a matter of reality; I can't wait for the understanding just to dawn on them".

At the end of the second year, however, her views had changed substantially.

> I never fully grasped the fact that students weren't learning what I wanted them to learn from their practical work (*Australian parlance for "labs"*). As teachers, we just bluffed ourselves, or we didn't think about it, or we just thought, 'Oh well, they're just not getting it. The good kids got it and the other kids didn't.' I can see from being with this research that we've been giving students good marks for immaculate (lab) work, all perfectly written up, and maybe they haven't had a clue about the chemistry related to the experiment[41].

The heart of the constructivist agenda, as captured in the quote above, is its goal of achieving "learning with understanding"[44] on the part of the student. The findings of this study offer support to the teachings of Vygotsky and the social constructivist agenda with regard to the importance of the role of the instructor in the learning process. It is only with the engagement of the instructor that the proper ZPD may be set, and that true cognitive development can take place.

There is still a lot of work to be done in the real-time application of social constructivism to the chemistry classroom. Many studies exist which use various techniques from the constructivist school (student-led observation and discovery, guided-inquiry methods, a focus on an understanding of underlying principles over rote learning, etc.) in the chemistry classroom. Akkus and Geban[9]used a five-week control-group study to assess the value of a guided-inquiry curriculum. Huddle and Ncube[45] use a hands-on exercise with playing cards and a graphical representation. Buell and Bradley[46] used graphical interpretation, and referred to the conclusion of a study by B.D. Smith[47] (citation only; no text available) that "a few weeks of training in logical operations at the beginning of a course does not reduce achievement, is mastered almost as well by slow learners as fast learners, and gives far better analysis of complex problems than just consideration of data". Bilgin and Geban[48] used a cooperative learning approach to expedite conceptual change in the classroom. Wheeler and Kass[43] used graphics and plotting methods, and reported that "graphical representations used in conjunction with the teaching of the concepts of constant concentration and the 'equilibrium constant' may be useful to the chemistry teacher in overcoming misunderstandings so often associated with these concepts". Another approach to this subject appears in the section below.

Misconceptions in Chemical Equilibrium

Another way to integrate the constructivist agenda into the chemistry classroom is through the use of the idea of the "misconception". This term is used to describe an incorrect understanding of a specific process, concept, or idea which has become established in an individual's intellect. After the reviews of cognitive theory earlier in this paper, it is possible to identify a "misconception" as an erroneous construct in LTM. Numerous studies of student misconceptions relative to the study of chemistry and, in particular, of chemical equilibrium can be found in the literature. **Error! Bookmark not defined.**[49,50,51,52,53]

Researchers into misconceptions of chemistry principles are interested in exactly how these incorrect constructs came to be, and how to correct them. According to the IPS model, concepts are formed in the learner from sensory inputs and filtration processes which are dependent on existing memory constructs in LTM. A young learner may have relatively few existing constructs, and as a result, the filtration processes for new sensory inputs may be fairly weak, resulting in the creation of poorly formed new constructs. As long as these erroneous constructs are not challenged by a child's day-to-day experience, they will remain. For example, the child who observes water condensing on the outside of a cold glass on a summer day might conclude that the water got there by passing through the walls of the glass. This concept, of course, is incorrect, but is a perfectly serviceable explanation for the child, who has no need for a better one, and so it will remain in LTM. At some point, the growing child will hear the term "condensation" and associate this term with the cold-glass phenomenon. This misconception will have to be addressed when that child enters a chemistry classroom[54].

The common-sense reasoning process is well suited to the goal of getting through the day, but must be radically reformed in order to learn the science of chemistry. Students bring to the learning process a set of concepts which they have developed to explain the everyday physical world to themselves; a "'naïve realism', a view of the world that blindly relies on perceptual clues"[53]. These conceptions are often at odds with the ones which the science teacher is trying to instill; hence the term "misconception". In the words of one researcher, the purpose of his work is to offer a "working model that can help chemistry teachers and instructors interpret students' common-sense ideas (read "misconceptions") in a more comprehensive way"[6].

One study of student misconceptions in is of particular interest, as it not only identifies them but attempts to correct them as well.[9] The authors are teachers of a three-year high-school course in general chemistry given at Gazi University in Ankara, Turkey. The study involved students from the second semester of the second year, during which chemical equilibrium is taught. Two lab sections served as Control and Experimental groups during a five-week study of chemical equilibrium, both taught by the same instructor. Course time for both groups was three hours(four 45-minutes sessions) per week. One group was taught using a constructivist technique, and used a curriculum developed by the authors designed to impart a conceptual understanding of the material. The other was taught using a traditional method based on problem-solving, using the standard curriculum. At the end of the instructional period, both groups were administered a 45-item “Chemical Equilibrium Concepts Test”, also written by the authors. Collation of the data collected from these sources appears in **Table 2**. Analysis of covariance showed a significant difference between these two groups, demonstrating that focused conceptual instruction can be effective in combating student misconceptions.

If we now make use of the identified misconceptions, we have saved ourselves some time and effort in two areas. We no longer need to perform a detailed inventory of the bad constructs coming into our classroom, and we can use the misconception rating to give us a head start on preparing for the work of fixing the incorrect constructs we expect to see.

Table 2:Excerpt from Results of Misconception Study[9]

MISCONCEPTIONS OF CHEMICAL EQUILIBRIUM	Control	Experimental
Approach to Equilibrium		
The forward reaction rate increases from time 0 until equilibrium is established	67%	25%
The reverse reaction rate is the same as the forward reaction rate	33%	9%
Characteristics of Equilibrium		
Equilibrium consists of two independent and separate parts (forward and reverse reactions) rather than one whole system	67%	25%
The forward reactions goes to completion before the reverse reaction commences	10%	3%
At equilibrium, the concentrations of reactants equals the concentrations of the products	77%	50%
Equilibrium is seen as oscillating like a pendulum	13%	3%
Lack of awareness of the dynamic nature of the chemically equilibrated state	44%	3%
No discrimination between reactions that go to completion and a reversible	23%	13%
At equilibrium the rates of the forward and reverse reactions are equal even as concentrations change	44%	13%
Initial Effects on Rates of Reactions		
When the system at equilibrium is disturbed by increasing the concentration of a reactant the reverse reaction rate will instantaneously decrease	64%	25%
When the temperature of an exothermic reaction at equilibrium is increased, the rate of the forward reaction is decreased	62%	38%
Rates of Reactions when Equilibrium is Re-established		
After the concentration of one of the reactants is increased and equilibrium is re-established, the rates of the forward and reverse reactions will be equal to those at the initial equilibrium	54%	25%
After the temperature is increased and equilibrium is re-established, the rates of the forward and reverse reactions will be equal to those at the initial equilibrium	67%	25%
After the system is at equilibrium, having the smaller number (coefficient) of gas particles is disturbed increasing the volume, the rates of the forward and reverse reactions will be equal to those at the initial equilibrium	38%	13%
After the system is at equilibrium, having the smaller number (coefficient) of gas particles in the left side of the equilibrium, is disturbed by decreasing the volume, the rate of the reverse reaction is decreased	38%	9%
Other		
Inability to distinguish between the concepts of mass and concentration	77%	47%
Inabiliity to distunguish between how fast a reaction roceeds (rate) and how far in goies (extent)	64%	44%

Chapter 3

OVERVIEW OF METHODOLOGY

Theoretical Framework of the Research: Input from the Literature

The research methodology was heavily influenced by the theoretical principles put forth in the literature reviewed above. As a bridge between the “academic” and “experimental” sections of this study, those theoretical principles will be briefly summarized here. Connections between the research design and these principles will be noted.

The concept of the Problem Space and its use of existing memory structures in the solution of a problem suggest the need to understand the condition of these structures in the student when planning new curricula and when interfacing with students in the classroom. The presence or absence of previously developed programs will have a great influence on the learning environment. Students’ ability to learn the concepts of Chemical Equilibrium will be severely hampered if they have not first acquired the necessary prerequisite “programs” (stoichiometry, reaction kinetics, etc.). As a corollary to this principle, the concept of the “misconception” can be useful to the instructor. As it is difficult and time-consuming to learnthe existing memory structures of an entire classroom, the use of existing compilations of common misconceptions can be a great help. The idea that new knowledge is built on top of existing memory structures suggests that the best way to present new problems is in a way that bearssome resemblance to those previously studied. The reuse of existing programs would help students to more quickly build new constructs and acquire original knowledge. Consistent language, labels, and concept descriptions prevent the student from having to reconstruct new models to accommodate unfamiliar terminology or ideas.

Physical limitations to the memory system, as outlined in the IPS model, should be taken into account by the instructor. The capacity limits of the STM are a constraint on the transfer rate of memory chunks to LTM; this suggests that the number of new concepts introduced in the classroom should be limited accordingly. Verification that new knowledge has been retained is needed before proceeding to other topics. Additionally, the slow write time of LTM when presented with more complex problems should also be recognized by the instructor, and compensated for when such material is being introduced in the classroom.

The role of the instructor in setting an appropriate ZPD in the classroom has been mentioned several times. The appropriate level of the curriculum must be set, and a proper ZPD must be maintained and continually modified by the instructor. The instructor's role is critical in the students' development; to quote Vygotsky "authentic modification in the child's thinking can probably best be accomplished through the active intervention of a well-qualified teacher"[40]. Thomas and McRobbie's study of "Frances" and her classroom[41] makes this principle exceptionally clear. The instructor plays the roles described in the constructivist credo; they must be the guide to the student's learning. They must not only "set the ZPD"; they must engage with the students and in turn make them engage with the material. Frances herself underwent a major change in her teaching philosophy;

> The most important thing was getting the students to think and to struggle with ideas, not allowing them to get away with not thinking . . . I wasn't going to tell them (the answer) . . . I realize now

> that if you understand an idea you should be able to explain it . . . it's only through them actually verbalizing it that they realize they're talking rubbish. In the equilibrium unit, if they didn't come up with an explanations or answer they just kept on having to try . . . they were getting desperate and then they would cooperate and then they found it worked. They were interacting and sharing ideas.[55]
>
> The changes in how I teach are a result of my recognition that they didn't know as much as I thought they knew. It just staggered me. Previously, I thought I was really explaining things well . . . but I didn't realize how little they knew or were learning . . . I don't think that I let them construct the understanding for themselves.[56]

These bracingly honest insights, from a woman who had spent the better part of her life teaching a traditional chemistry curriculum, are powerful, and speakstrongly to the value of the social constructivist agenda.

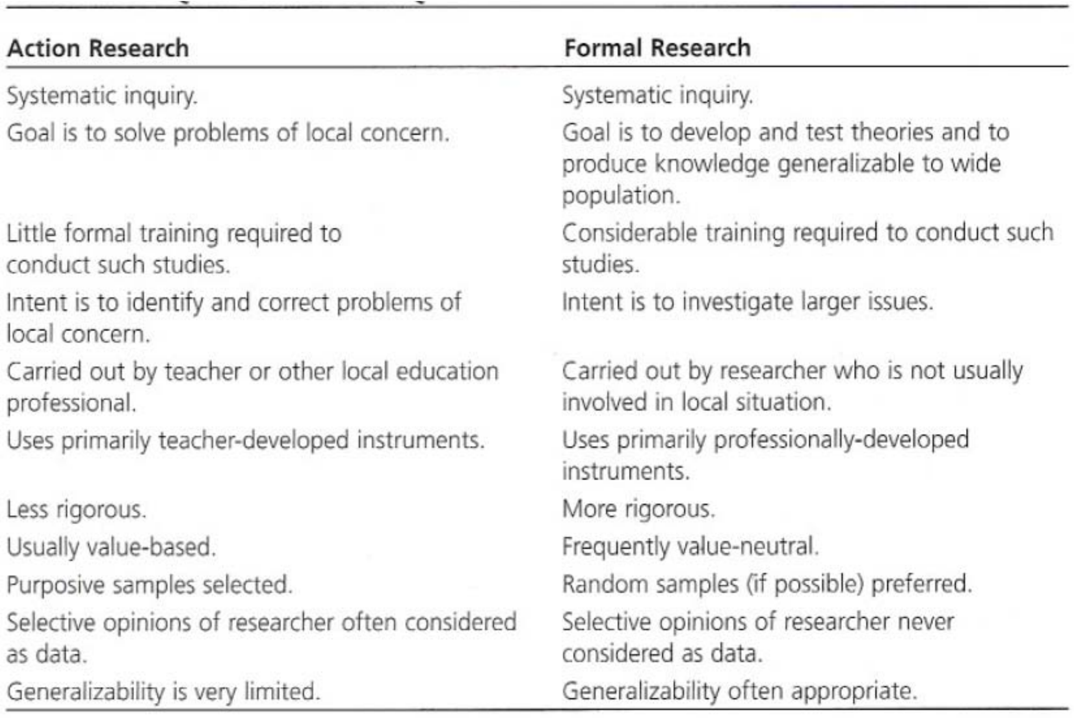
Design and Assessment of Research: "Action Research"

In the study of natural events, the scientific community uses observation and controlled-variable experimentation in order to explain the world in a quantitative way. When studying human behavior, however, the subject is so complex, and the number of variables to control is so large, that true quantitative study is not feasible. The social sciences have made use of qualitative study of human behavior, but this requires the processing of large amounts of textual data by professionals specially trained in the area of study. Neither pure qualitative nor quantitative methods are suitable for real-time educational research.

As the traditional formal research methodologies are not appropriate for the work of educational psychologists, they have developed their own research methodology. "Action Research", a term attributed to the German psychologist Kurt Lewin, is a system of repetitive observation, intervention, and implementation developed to allow educational professionals to address problems in teaching and learning. As opposed to classic quantitative and qualitative methodologies, which consist of peer-reviewed statistical evaluation of theoretical hypotheses, the objective of action research is simply to make improvements in local systems in a timely manner; to get things done, one might say. A comparison between action research and the more formal research techniques can be seen in **Table3**.

As can be seen in this table, action research includes many aspects of formal research, but is different in several key ways. The most important of these is that, while formal research aims to develop generizable results, which can be repeated elsewhere by other researchers, the objective of action research is to find solutions to problems which are specific to the locale of the study. It makes no claims as to the portability of its solutions to any other environment. Since this frees the experimenter from the need

Table 3: Comparison of Action Research and Formal Research[57]



Action Research	Formal Research
Systematic inquiry.	Systematic inquiry.
Goal is to solve problems of local concern.	Goal is to develop and test theories and to produce knowledge generalizable to wide population.
Little formal training required to conduct such studies.	Considerable training required to conduct such studies.
Intent is to identify and correct problems of local concern.	Intent is to investigate larger issues.
Carried out by teacher or other local education professional.	Carried out by researcher who is not usually involved in local situation.
Uses primarily teacher-developed instruments.	Uses primarily professionally-developed instruments.
Less rigorous.	More rigorous.
Usually value-based.	Frequently value-neutral.
Purposive samples selected.	Random samples (if possible) preferred.
Selective opinions of researcher often considered as data.	Selective opinions of researcher never considered as data.
Generalizability is very limited.	Generalizability often appropriate.

to do extensive statistical analyses of the data, it makes the research much simpler without reducing its value to the system for which it was performed. And since generalizationof the study is not a critical objective, there is no need to adhere to strict random sampling as is the case in formal research. The sample for any action research project is just the group at which the research is targeted, and the results of the study will be applied to this group only.

Another major advantage of action research is that the simplicity of the research makes it possible for anyone, anywhere, to run an action research experiment. Action research is designed for educators who want to solve problemsin their classroom; who better to conduct this research than the educators themselves? No special training is required to run an action research program. And once the research is complete, results are immediately available to apply to the problem at hand. The research process can be repeated indefinitely, until the educator is satisfied with the progress made.

There are four steps in an action research project;

1) Identify the research problem or question
2) Obtain the information necessary to answer the question
3) Analyze and interpret this information
4) Develop a plan for further action (research and/or implementation)

Lewin has created a model of the action research process, seen below in **Figure 3**. At the top of the figure, the current state of events is evaluated, ideas for a plan of action are generated, and these ideas are implemented in the “first action step”. The impact of this implementation is monitored, results are obtained and a process of reflection and revision takes place, after which the implementation is re-evaluated and fed into another cycle. Action research can be implemented as a continuous cycle of research and implementation; in Lewin's own words, "No action without research; no research without action"[58].

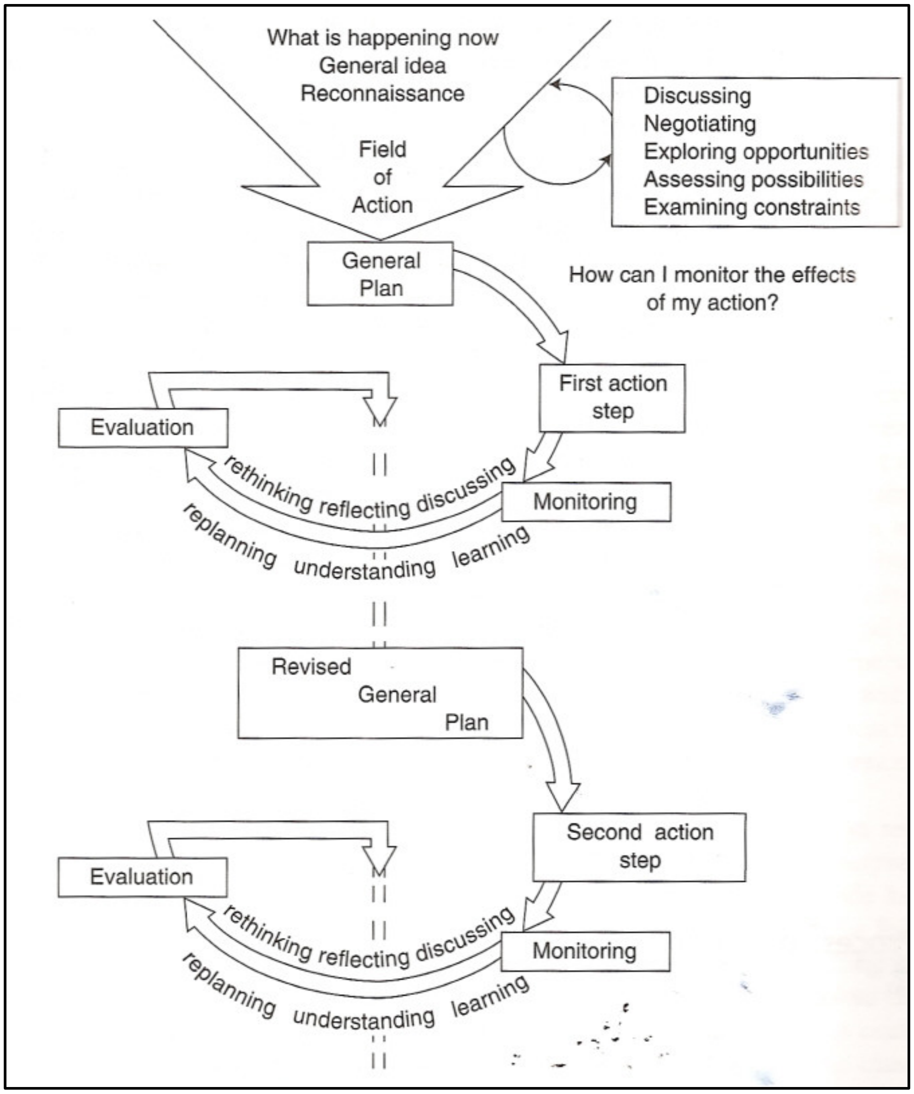


Figure 3: Model of the Action Research Methodology[59]

Finally, an additional advantage of action research is that it helps to improve the classroom skills of its user. Not only does it give the instructor a better understanding of the problems of the student, and the ways in which these problems might be mitigated, it makes them more empathetic to the learning needs of the student, leading them to revise their teaching style as this understanding grows.

> (Action research) shows practitioners that it is possible to break out of the rut of institutionalized, taken-for-granted routines and to develop hope that seemingly intractable problems in the workplace can be solved.[60]

Execution of Research: Logistics

The research methodology made use of several different procedures, and was conducted in accordance with the principle of action research reviewed in the previous section.The overall research logistics can be broken into three successive steps, as outlined in **Table 4** below.

Table 4: Research Timeline

Fall 2009: *Data Gathering*	**Spring/Summer 2010: *Intervention 1***	**Fall 2010: *Intervention 2 and Assessment***
• Diagnostic Exam • Chem 1B Exam 1 • Chem 1B Exam 2	• Focus Group sessions	• Worksheet postings • Chem 1B Exam 1 • Chem 1B Exam 2

The first step was to gather data on the initial understanding of students entering Chem 1B, specifically as it pertained to chemical equilibrium. Two methods were used to gather this data. In the first, a diagnostic exam was written and administered in the first week of class in the Fall 2009 Session. The secondmethodused the results of the first two Chem 1B exams, which were analyzed for information on misconceptions relating to chemical equilibrium. Bothtechniques includedetailed analyses of the *wrong*answers, in order to develop a better understanding of the basis for students' misunderstanding.

The second step ("Intervention 1" in Table 4) consisted of the development and application of an "intervention", in the language of action research, designed to address the problem summarized in the Introduction to this paper. A number of worksheets and an interactive spreadsheet tool were created for use with students in a guided-inquiry setting aimed at building conceptual understanding of key concepts in chemical equilibrium. Face-to-face, constructivist-style "focus groups" (instructor-guided, student led, use of leading questions, student articulation of learning, etc.), were held, and the results of each session were recorded and analyzed.

The third step ("Intervention 2" in Table 4) consisted of a different use of the worksheets developed for Intervention 1. In this step, the worksheets were posted as downloadable documents which were completed and submitted by Chem 1B students. Completed worksheets were reviewed for insight into students' conceptual understanding of the material.

Methodology and data analysis for these three steps will be reviewed in detail in the following three sections. In order to more clearly present this information, methodology and data analysis for each of the three steps will be reviewed together in three separate sections, as follows:

Part 1: Data Gathering (Step 1) Methodology, Results and Analysis
Part 2: Intervention 1 (Step 2) Methodology, Results and Analysis
Part 3: Intervention 2 (Step 3) Methodology, Results and Analysis

Chapter 4

DATA GATHERINGMETHODOLOGY, RESULTS & ANALYSIS

Diagnostic Exam

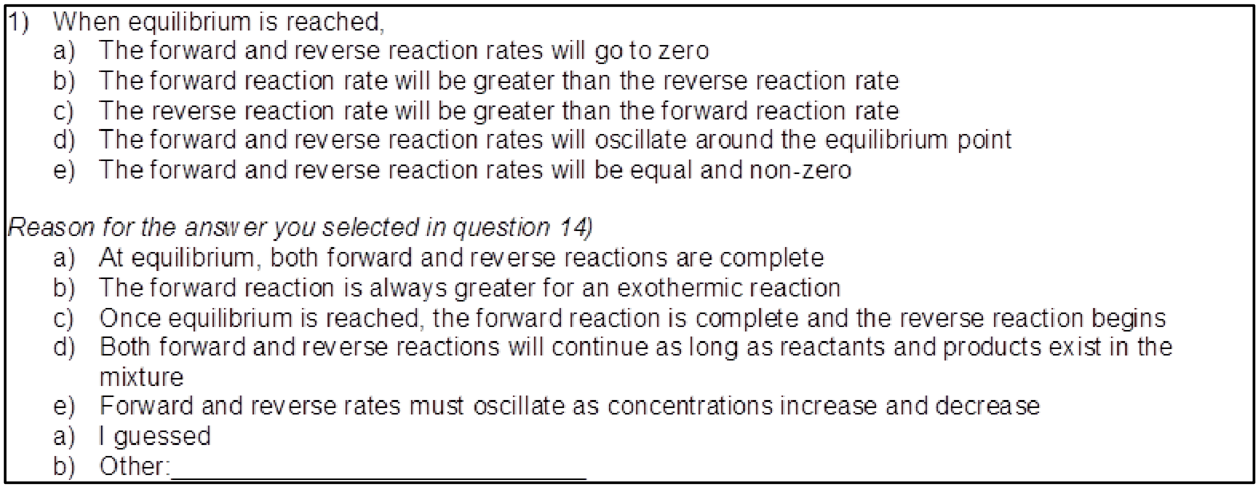
Methodology

A Diagnostic Exam was createdby the author of this study for the purpose of a obtaining a profile of entering Chem 1B students. The questions were mainly multiple-choice in nature and asked conceptual questions on topics which related to chemical equilibrium and which were covered in the Chem 1A curriculum. Questions from similar studies in the literature[5,**Error! Bookmark not defined.**,49]were used as input for this exam. Much time was spent revising questions content and style, and fourteen exam revisions were made prior to its administration.

The Diagnostic Exam was administered to a class of 110 Chem 1B students in the first week of class in the Fall 2009 semester. Sixteen questions were asked over a period of twenty minutes, as follows:

Product/reactant identification:	2 questions
Stoichiometry :	5 questions
Kinetics and reversible reactions:	5 questions
Basic equilibrium:	4 questions

This exam was administered to a group of entering Chem 1B students with little prior or no exposure to chemical equilibrium, and should be considered only as a test of how well prepared these students are for that material. Certain exam questions, however, addressed basic principles of chemical equilibrium as an extension of fundamental concepts of kinetics principles, and are included as a measure of how well students might be able to logically extend those concepts. Questions on the exam were developed based on sample questions featured in research studies of student misconceptions of chemical equilibrium[5,**Error! Bookmark not defined.**,61] A number of questions asked students not only for an answer, but also for the reason for their answer. A sample of such a question can be seen in **Figure 4.** The combination of both answers make up the reply to the question.



1) When equilibrium is reached,
   a) The forward and reverse reaction rates will go to zero
   b) The forward reaction rate will be greater than the reverse reaction rate
   c) The reverse reaction rate will be greater than the forward reaction rate
   d) The forward and reverse reaction rates will oscillate around the equilibrium point
   e) The forward and reverse reaction rates will be equal and non-zero

Reason for the answer you selected in question 14)
   a) At equilibrium, both forward and reverse reactions are complete
   b) The forward reaction is always greater for an exothermic reaction
   c) Once equilibrium is reached, the forward reaction is complete and the reverse reaction begins
   d) Both forward and reverse reactions will continue as long as reactants and products exist in the mixture
   e) Forward and reverse rates must oscillate as concentrations increase and decrease
   a) I guessed
   b) Other:______________________________

Figure 4: Sample Question with "Reason" Response from Diagnostic Exam

Questions are sometimes posed in both textual and graphical form to showwhether the form of the question has an effect on students' understanding. For example, in **Figure 5**, a question is asked in a graphical manner which had previously been asked in written form;

1) Assume that Reaction 1 goes to completion. If the initial concentrations of AB and C_2 are 0.6M, which of the following graphs is possibly correct?

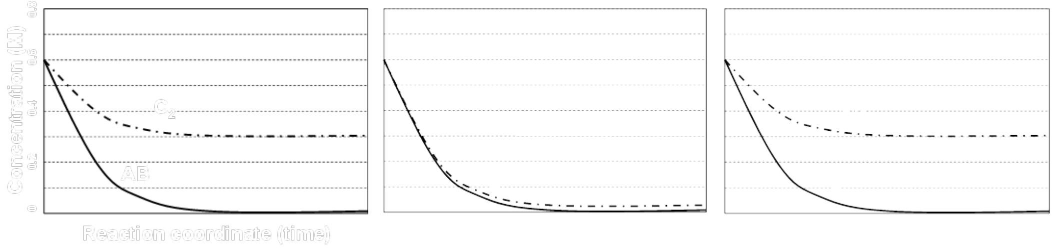


Figure 5: Graphical Question on Diagnostic Test

The complete Diagnostic Exam can be seen in Appendix A. An electronic copy of the Diagnostic Exam can be found on the CD attached to this paper.

Results and Analysis

Results of the analysis of the Diagnostic Exam are shown below. The results are grouped into four sections based on the four questions types listed above.

Section 1: Reactants and Products

The first two questions showed a hypothetical reactions and asked students to identify reactants and products, and whether or not the reaction was balanced. Almost all (~98%) of the students answered these questions correctly.

Section 2: Chemical equations and stoichiometry

The questions in this section concerned a hypothetical reaction;

$$2AB + C_2 \rightarrow 2AC + B_2$$

The reaction was presented as an irreversible reaction which ran to completion, as this was the only type of reaction familiar to students at this point.

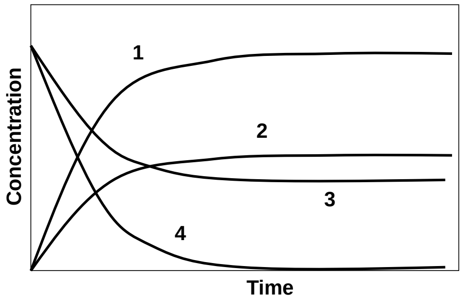
Questions 3 asked students to describe why we require equations to be balanced. Students showed highly variable levels of understanding of this concept. This was not a multiple choice question, and required some interpretation of answers, but using criteria which required students to mention something about conservation of mass, just over half the students replied correctly. Correct answers consisted of “so that products correspond to reactants”, “to be sure nothing is lost”, “to end up with what we started with”, and “to be stoichiometrically correct”. Incorrect answers fell into one of several categories. The most common wrong answer (12% of responses) had to do with equilibrium (“the equation must be balanced in order for equilibrium to be reached”). This may have had something to do with the fact that the diagnostic exam was titled “Chemical Equilibrium Diagnostic Test”, and that students just guessed that “equilibrium” was the answer sought. It’s also possible that students saw a similarity between the words “balance” and “equilibrium”; both have to do with symmetry and stability, and without a full understanding of chemical equilibrium, students made this connection based on their understanding of the language. This is an example of a common phenomenon in chemistry, where a term commonly used in daily life is used to denote something very different in chemistry, and is one of the reasons for the difficulty in teaching the concept of chemical equilibrium. (A thesaurus search for “equilibrium” in Microsoft Word yields the words “balance”, “symmetry”, and “stability” as synonyms.)

The next-largest category of incorrect answers (11%) basically said that a balanced reaction was necessary for the reaction to proceed. There is a germ of truth here; reactions do only progress with the right combination of reactants (“2 molecules of AB plus one molecule of C_2 equal . . .”), but this answer did not address the issue of mass conservation, which should be a principle well understood by prospective Chem 1B students.

Question 4 required students to translate the written expression of a chemical reaction into its graphical depiction (see below). Concentrations of reactants and products vs. time were shown and students were asked which might possibly represent the outcome of the hypothetical reaction. Roughly one third of students answered this question incorrectly.

4) For the chart below, circle “a” or “b” to indicate which statement is true:

a) Curves 1 and 2 are reactants; curves 3 and 4 are products
b) Curves 1 and 2 are products; curves 3 and 4 are reactants

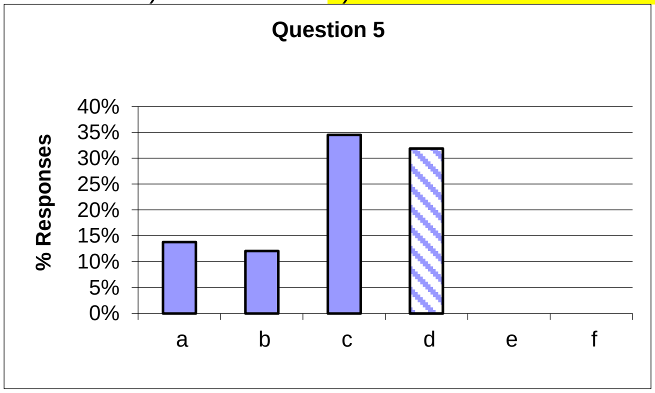


This result begs the question of the real source of misconception here. Did respondents really not recognize that product concentrations increase over time while reactant concentrations decrease? Or did they have trouble correctly reading the chart? This problem indicates the importance of presentation when writing questions and possible answers, and exposes the limitations of a diagnostic test, which is only as good and precise as the author designs it to be.

Questions 5 and 6tested students’ understanding of how the stoichiometry of a reaction determines changes in the concentrations of its reactants and products. Again, the hypothetical reaction $2AB + C_2 \rightarrow 2AC + B_2$was used. These questions and the distribution of answers are shown below.

5) As this reaction proceeds, which reactant *element* will be consumed most slowly?

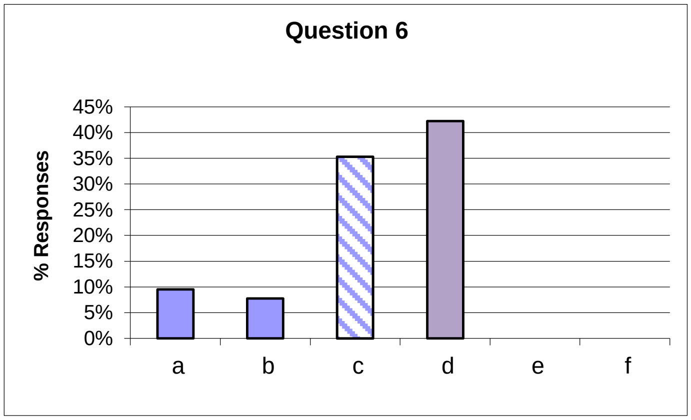
a) A b) B c) C d) All will be consumed at the same rate



Less than a third of students gave the correct response for **Question 5** (d). *(Note that in future diagrams, the correct answer will be indicated by a cross-hatched bar.)*This result is a bit disturbing, as it's difficult to suggest any reason for it other than that the students just do not understand the function of the stoichiometric coefficient and elemental subscript used in a chemical equation. The most common answer that element C was consumed most slowly; this indicates a belief that the coefficient of AB was somehow more important in determining its rate of consumption, and that since C_2 had a smaller coefficient, C would be consumed more slowly.

6) Which product *molecule* will be generated most quickly?

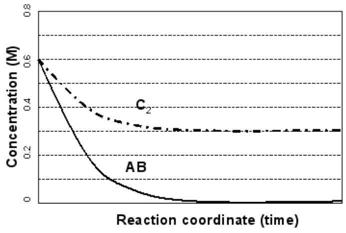
a) AB　　b) C_2　　c) AC　　d) B_2

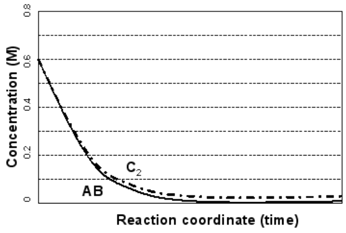


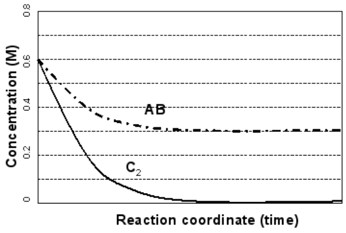
Question 6, which asked a question of a similar nature to Question 5, again showed only a third of responses to be correct. One might assume that a question dealing with the formation of *molecules* would be more straightforward than one at the elemental level, but this was not the case. The only conclusion to be made from the responses to these two questions is that there is a basic lack of understanding of the stoichiometric meaning of coefficientsvs. the subscripts in the chemical equation on the part of entering Chem 1B students. This does not bode well for their success in studying chemical equilibrium.

Question 7 again tests conceptual understanding of stoichiometry, but this time in a graphical manner.

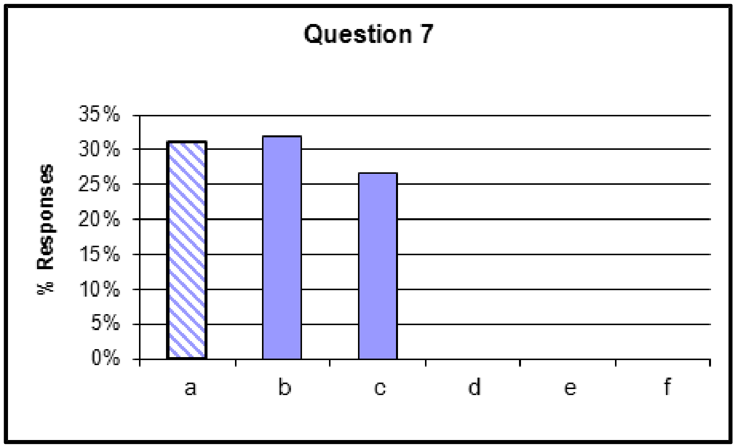
7) Assume that this reaction goes to completion. If the initial concentrations of AB and C_2 are 0.6M, which of these charts are possibly correct?







a b c



Once again, only one third of responses were correct. This result indicates a true misunderstanding of stoichiometry, which cannot be attributed simply to the form in which the question is posed (textual vs. graphical).

Section 3: Reaction Kinetics and Reversible Reactions

In **Questions 8 and 9**, students are given the generic reaction and areasked

$$aA + bB \rightarrow cC + dD$$

two questions about the rate reaction for this reaction, which students learned in Chem 1A (the equation below is copied directly from the Chem 1A lab manual);

$$\boxed{\text{Rate} = -\frac{1}{a}\cdot\frac{\Delta[A]}{\Delta t} = \frac{1}{c}\cdot\frac{\Delta[C]}{\Delta t}}$$

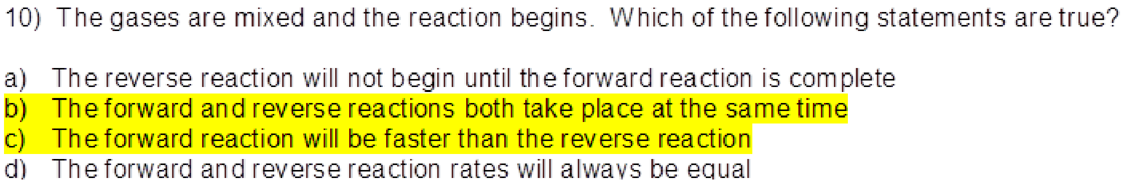
8) Why is one of these terms positive and one negative?
9) What will happen to the rate of a reaction if [reactant] is increased? If it is decreased?

These questions were not multiple-choice, and asked for a written answer. Both of these questions were correctly answered by about one third of the students. Many of the responses were either blanks or nonsense, and it was sometimes difficult to grade answers as simply “right” or “wrong”. For example, an answer for Question 8, “because one is reactants and one is products” is wrong on its face but may represent understanding on the part of the responder. There was a wide variety of incorrect answers, indicating a lot of guessing and little true understanding of the equation. Again, not a good sign for potential Chem 1B success.

In this section, students are shown a brief explanation of the difference between an irreversible and a reversible reaction. Reversible reactions are introduced as a combination of two simultaneous reactions; a forward and a reverse reaction, each of which has its own rate. **Questions 10 and 11** are asked about a hypothetical reversible reaction:

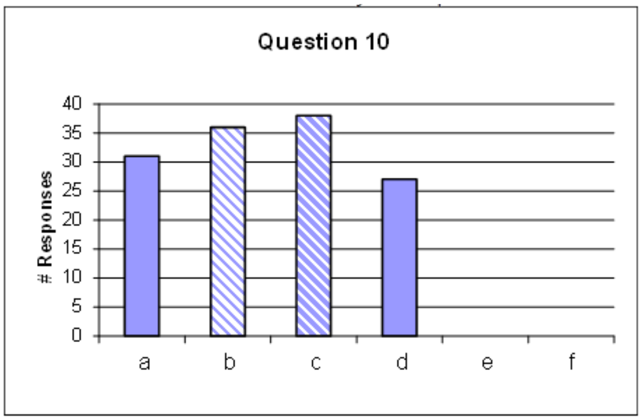
$$2NO_{(g)} + Cl_{2\,(g)} \rightleftharpoons 2NOCl_{(g)} + \text{heat}$$

Note that, beginning at this point in the exam, all questions come in two parts. The first part asks for an answer to the question posed, and the second part asks the student to select a *reason* for the answer given. Plots of the distribution of responses are shown for both parts; the plot of the straight answer distribution is labeled (for example) “Question 10”, and the plot for the distribution of reasons selected is labeled “Question 10R”. In addition, questions now allow for the possibility that more than one answer may be correct; thus, the percentages indicated in the following charts may not add up to exactly 100%. For example, for **Question 10;**



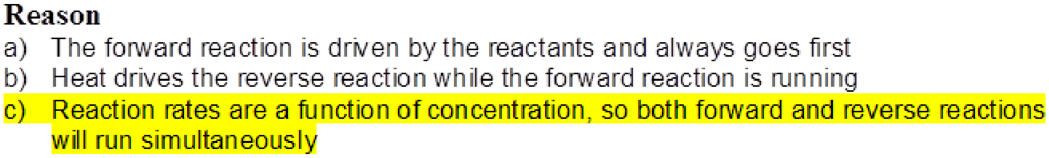
10) The gases are mixed and the reaction begins. Which of the following statements are true?

a) The reverse reaction will not begin until the forward reaction is complete
b) The forward and reverse reactions both take place at the same time
c) The forward reaction will be faster than the reverse reaction
d) The forward and reverse reaction rates will always be equal



In this case, a little over half of the students answered with at least one of the correct solutions; very few gave both. More interesting are the high-percentage incorrect answers; over 30% of respondents believed that the forward reaction completes before the reverse reaction begins; this correlates to similar findings in the literature, as indicated in the table of misconceptions in the Literature Search section (Table 2). Additionally, ~27% of the responses showed a misconception with regard to how the reaction rates of the forward and reverse reactions are related to one another through the course of a reversible reaction.

Reasons for the answer chosen to question 10 are then given in **Question 10R:**

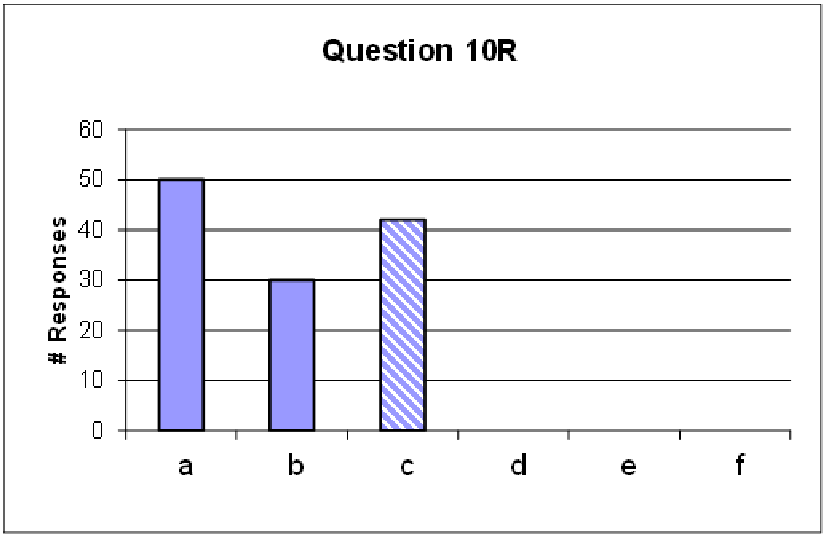


Reason

a) The forward reaction is driven by the reactants and always goes first

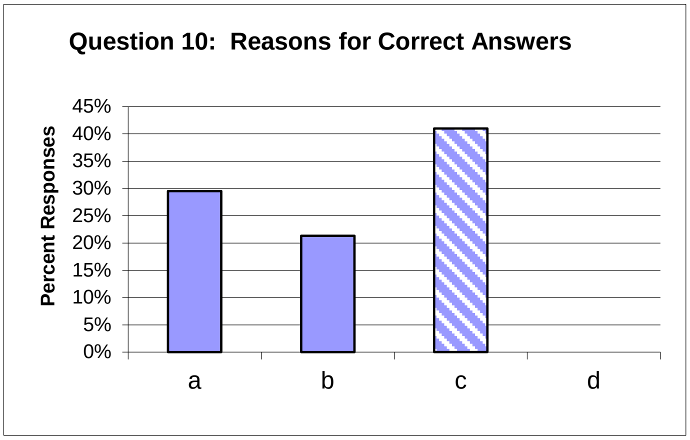
b) Heat drives the reverse reaction while the forward reaction is running

c) Reaction rates are a function of concentration, so both forward and reverse reactions will run simultaneously



Again, a large number of students (50 students answered a to this question) believed that the forward and reverse reactions were decoupled, as confirmed in the literature[Error! Bookmark not defined.].

Another interesting analysis can be made using this data. It is possible to specifically examine the reasoning of students who gave a correct answer to a question. This data is plotted below. Note that the horizontal axis labels refer to the same reasons (a, b, c, d) as in Question 10R above. These plots show percentages of answers, as the value of n (number of correct answers) will vary for each question reviewed.

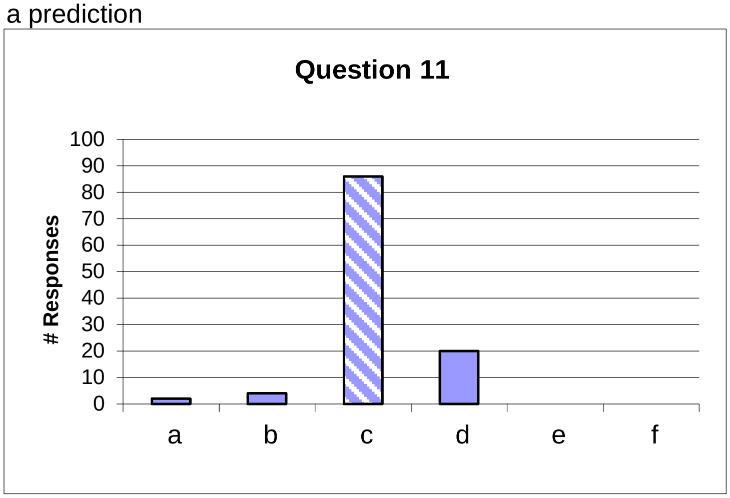


The data shows that, of the 40 students who answered Question 10 correctly, only 40% indicated the correct reason for their answer. 30% incorrectly believed that the forward reaction always precedes the reverse reaction, and 25% believed that heat is the primary driver of the reaction.

Question 11 tests for understanding of the basics of reaction kinetics;

11) After the reactants (NO and Cl_2) are mixed and the reaction begins,

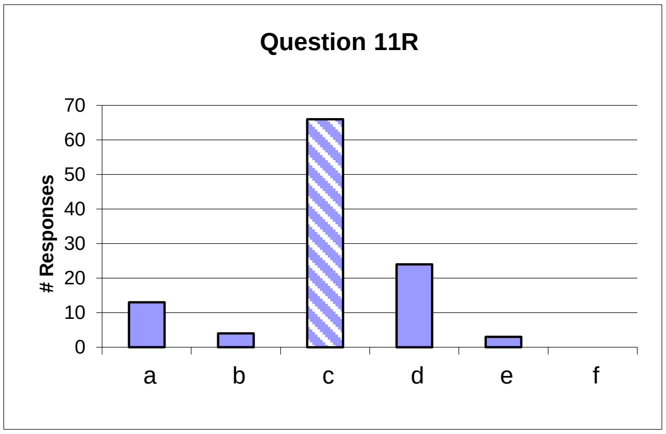
a) The concentration of $NO_{(g)}$ will increase
b) The concentration of $Cl_{2\ (g)}$ will increase
c) The concentration of $NOCl_{(g)}$ will increase
d) Cannot make a prediction



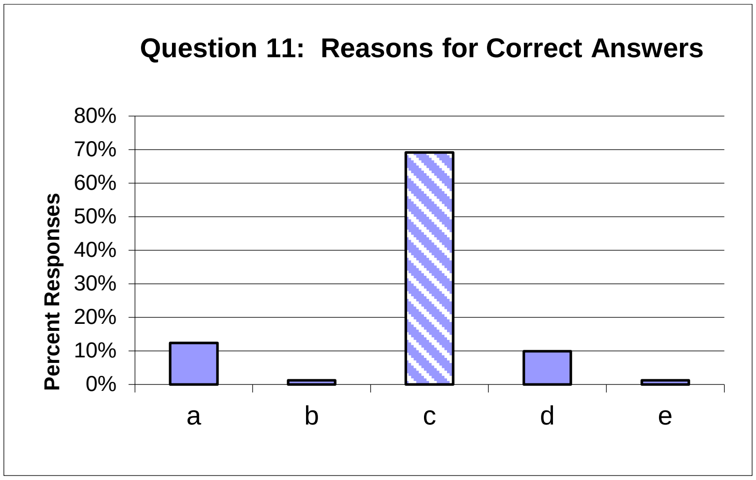
Almost all replies to Question 11 were correct, and almost all of the reasons correctly identified the process of consumption and generation taking place in the reaction as the reason for their answer (Question 11R, below). However, a significant number of responses showed incorrect reasoning (**Question 11R).**

Reason

a) An exothermic reaction will produce a higher concentration of products
b) The concentration of $Cl_{2\ (g)}$ will decrease relative to $NO_{(g)}$ and $NOCl_{(g)}$ because of its lower coefficient
c) As the reaction proceeds, $NO_{(g)}$ and $Cl_{2\ (g)}$ will be consumed to generate $NOCl_{(g)}$
d) Cannot answer; all concentrations are determined by a chemical equation
e) I guessed
f) Other

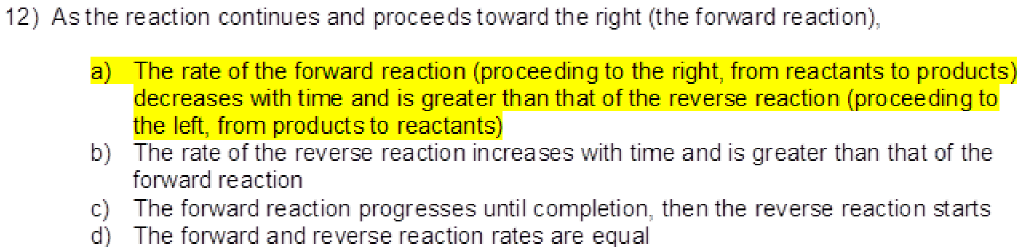


Inspection of the correct responses to Question 11 shows a very similar distribution of answers. Approximately 70% of correct responses also gave the correct



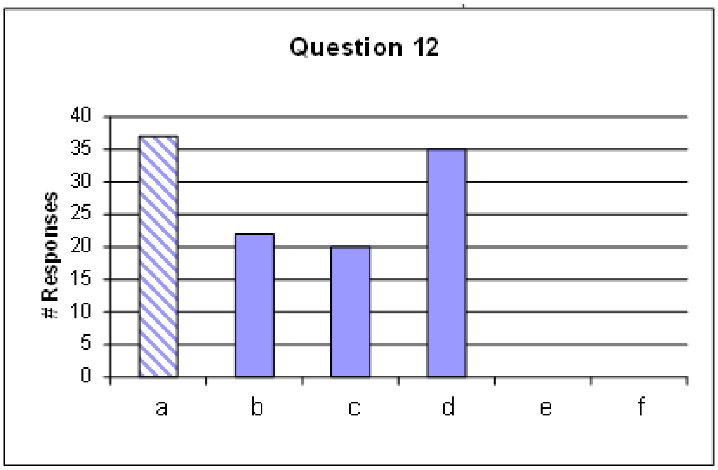
reason for their response. 12% of the correct respondents believed that products are produced primarily because of the thermodynamics state of the reaction, and 10% were unable to answer.

Question 12 poses a question about the variation in forward and reverse rates. This involves a topic that has not yet been introduced to the students, and not surprisingly only a third of students answered correctly;



12) As the reaction continues and proceeds toward the right (the forward reaction),

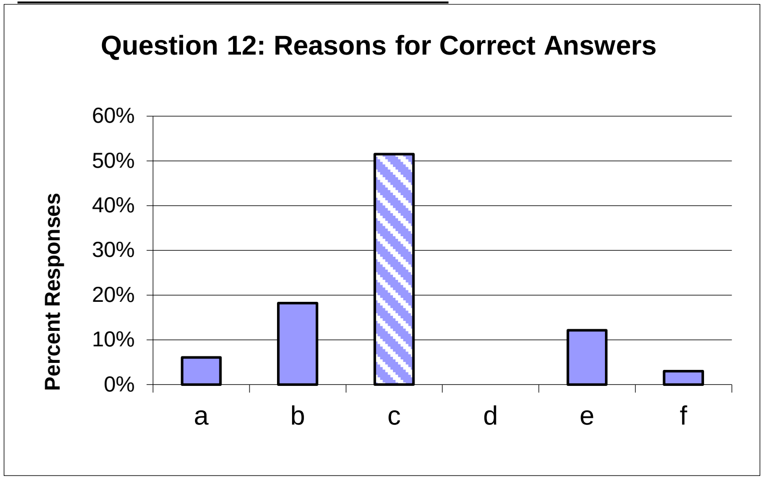
a) The rate of the forward reaction (proceeding to the right, from reactants to products) decreases with time and is greater than that of the reverse reaction (proceeding to the left, from products to reactants)
b) The rate of the reverse reaction increases with time and is greater than that of the forward reaction
c) The forward reaction progresses until completion, then the reverse reaction starts
d) The forward and reverse reaction rates are equal



This question required an understanding of the kinetics of a reversible reaction, which admittedly most students had little exposure to. The most common answer to this question was the correct one, which seems to indicate a basic understanding of reaction kinetics; however, this answer was given by only ~ one third of the students, and further analysis shows that many of the correct responders do not have a good understanding of this process.

Reason for the answer you selected in question 12

a) There are two reactants and only one product
b) This is an exothermic reaction, so the forward reaction will be faster
c) As the reaction starts and proceeds to equilibrium, reactants are consumed, reducing the rate of the forward reaction
d) Forward and reverse reaction rates of reversible reactions are always equal
e) I guessed
f) Other:______________________________



Almost 20% of the correctly responding students again believe that the exothermic nature of a reaction accounts for its progress to the right. This is most likely because they have been taught that an exothermic reaction is thermodynamically favored and more likely to proceed.

Section 4: Basics of Chemical Equilibrium

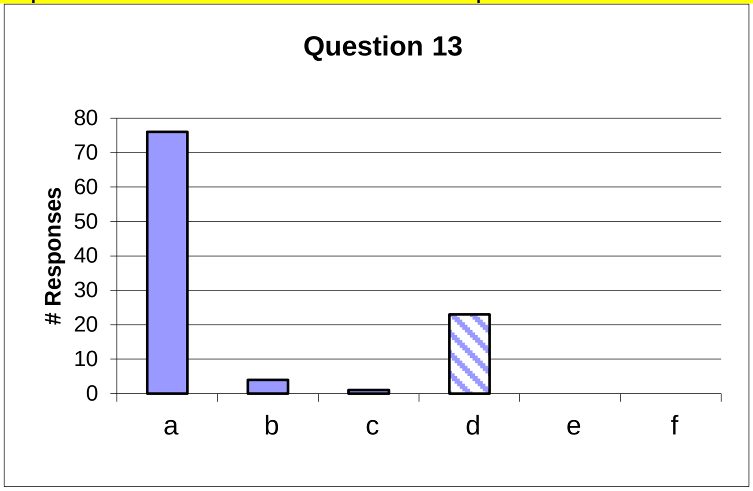
In this section, students are given the basic definition of chemical equilibrium below:

At some point in time the (reversible) reaction will reach a "steady state", at which time the concentrations of both reactants and products and the rates of the forward and reverse reactions will stop changing and become constant. This is the point of "chemical equilibrium".

They are then asked a number of questions about the state of a chemical reaction in equilibrium. While students were expected to have little experience with the concept of chemical equilibrium at this point in their education, the questions are answerable by using a degree of common sense and the rules of reaction kinetics. In addition, wrong answers here can be reviewed as to the nature of the answer and how it may have entered the student's LTM as a concept.

13) At the equilibrium point,

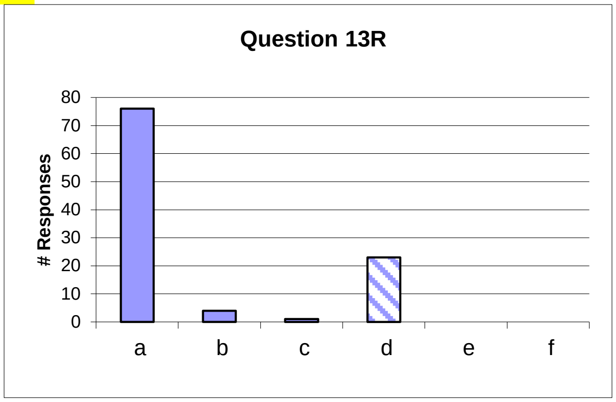
a) The concentration of all products and reactants will be equal
b) The concentration of reactants will be greater than the concentration of products
c) The concentration of products will be greater than the concentration of reactants
d) Cannot predict the relative concentrations of products and reactants



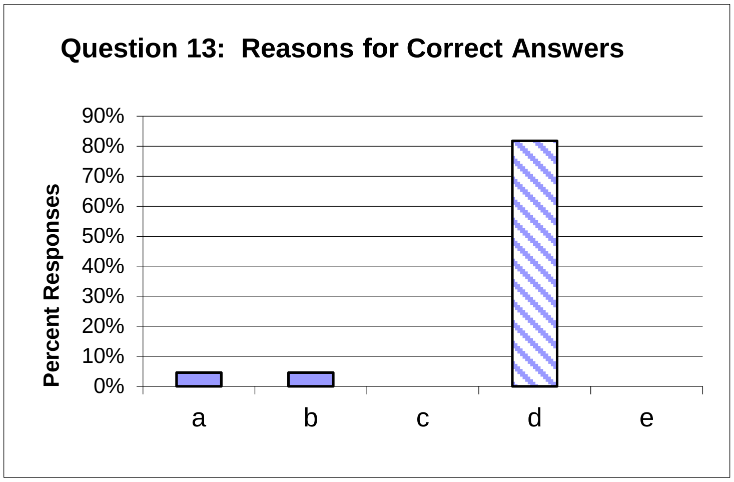
In **Question 13**, many students believed that "equilibrium" means "equal"; this is one of the most commonly observed misconceptions of equilibrium (see "Misconceptions in Chemical Equilibrium" in the Literature Search section). Moreover, when indicating the reason for their answers, these students identified "equality of concentrations" as the*definition* of equilibrium (see below).

Reason

a) Equality of concentration is the definition of equilibrium
b) Concentrations of reactants are higher due to the stoichiometric relationships of reactants and products
c) This is an exothermic reaction so concentrations of products are higher
d) Cannot answer; all concentrations at equilibrium are determined by a chemical equation



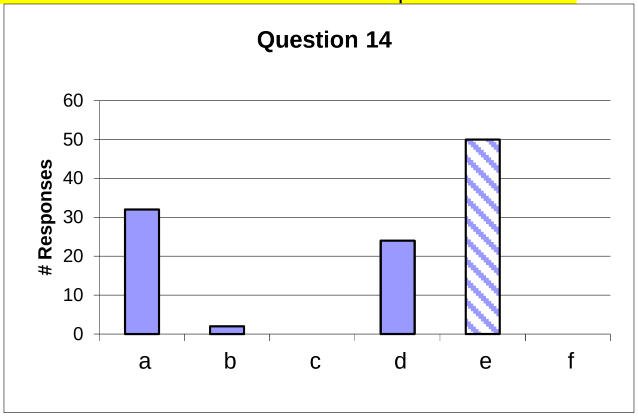
However, those students who correctly answered this question clearly understood why, as seen below.



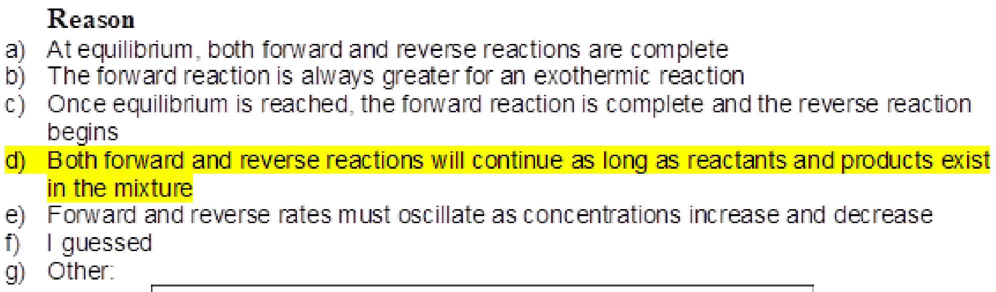
Question 14 asks about the state of the system at equilibrium.

14) When equilibrium is reached,

a) The forward and reverse reaction rates will go to zero
b) The forward reaction rate will be greater than the reverse reaction rate
c) The reverse reaction rate will be greater than the forward reaction rate
d) The forward and reverse reaction rates will oscillate around the equilibrium point
e) The forward and reverse reaction rates will be equal and non-zero

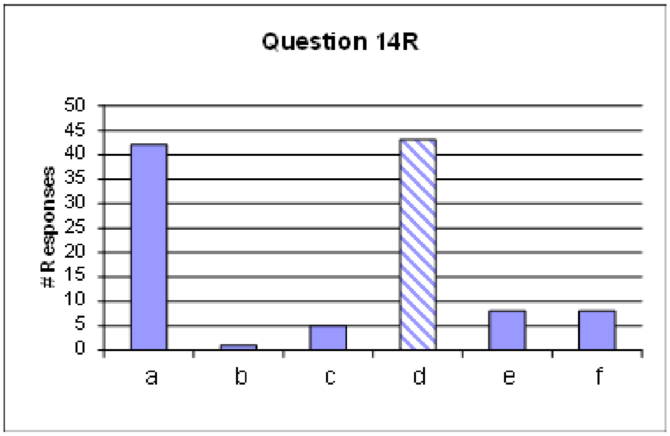


The literature identifies this concept as a large source of misconceptions about chemical equilibrium[9]. In this question, about half of the students correctly answered that both forward and reverse reactions would continue to run at equilibrium, and that their respective rates would be equal. This is much higher than the percentage of correct responses for earlier, similar questions. This could possibly relate to the fact that, after answering previous questions, students began to develop a better sense of what is really happening at equilibrium, and that, from a common-sense point of view, none of the other possible answers really make sense. By not answering b or c, students correctly indicated that forward and reverse rates could not logically be different, leading to a correct answer of e. The only competitive answer here was a, which is a logical carryover from Chem 1A, in which students were shown that reactants go to products, and then the reaction stops.

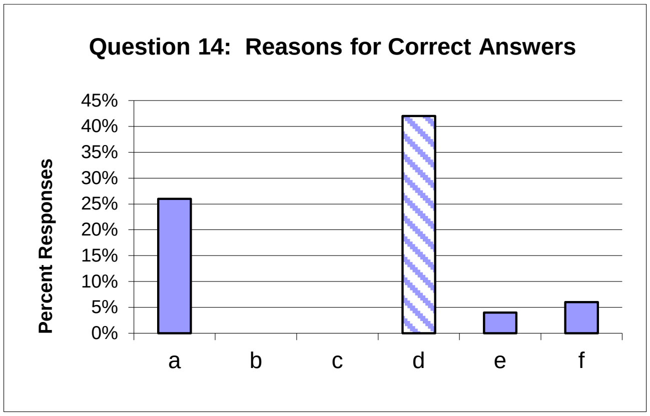


Reason

a) At equilibrium, both forward and reverse reactions are complete
b) The forward reaction is always greater for an exothermic reaction
c) Once equilibrium is reached, the forward reaction is complete and the reverse reaction begins
d) Both forward and reverse reactions will continue as long as reactants and products exist in the mixture
e) Forward and reverse rates must oscillate as concentrations increase and decrease
f) I guessed
g) Other:



When asked in **Question 14R** for the reason for their reply, students were equally divided between two poles of thought. In one, all reactions stop at equilibrium, and in the other they do not. The latter answer correctly relates this problem to the rate equation, which relates reaction rates to product and reactant concentrations. Accordingly, both forward and reverse rates will continue simultaneously for as long as both products and reactants remain in solution. An analysis of correct answers reflects this distribution.

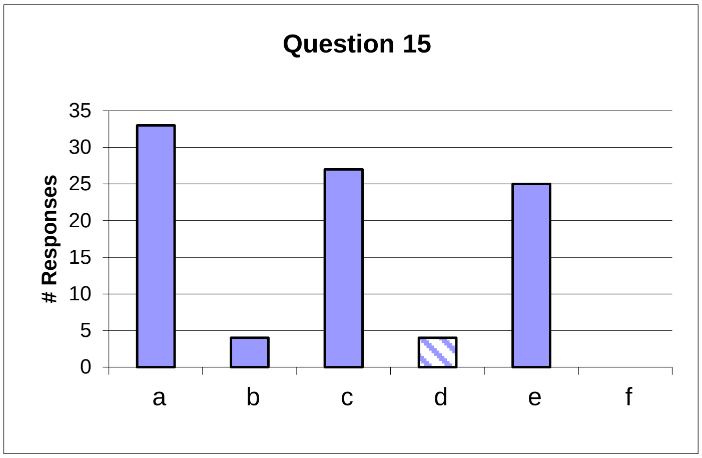


Questions 15 and 16 refer to the reversible reaction used in Part 2;

$$2NO_{(g)} + Cl_{2\,(g)} \rightleftharpoons 2NOCl_{(g)} + \text{heat}$$

15) An equimolar solution of NO and Cl_2 is prepared and allowed go to equilibrium (Reaction 1 above). Which of the following statements are true?

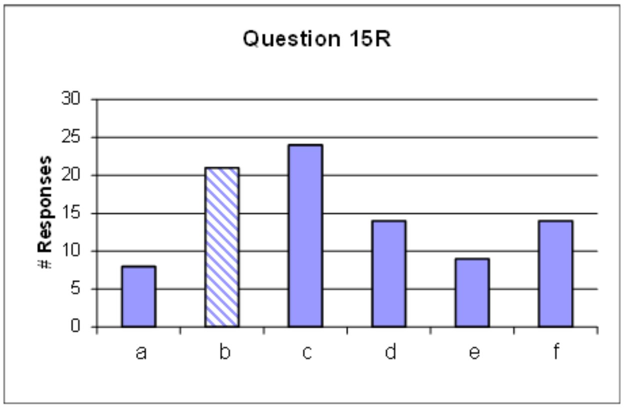
a) Concentrations of NO and NOCl in the equilibrium solution are equal
b) Only NOCl and Cl_2 will be found in the equilibrium solution
c) The concentration of NO will be half that of Cl_2 in the equilibrium solution
d) The concentration of Cl_2 will be greater than that of NO in the equilibrium solution
e) None of the above



Only a very small percentage of answers to Question 15 were correct. This question probably requires a much better intuitive understanding of the kinetics of equilibrium than any other;the fairly large (24%) percentage of “none of the above” answers indicates the degree of confusion here. On the other hand, this is basically just a stoichiometry problem, and, as indicated by the large number of “a” and “c” responses, again indicates confusion on the part of many incoming Chem 1B students with regard to this topic.

Reason for the answer you selected in question 15

a) NO is the limiting reactant, so it will go to zero in the final solution
b) NO is consumed twice as fast as Cl_2, so its concentration will be lower than Cl_2
c) NO and NOCl concentrations will be equal as they have the same stoichiometry
d) The reaction uses two NO molecules for every Cl_2molecule, so there will only be half as many NO molecules left at equilibrium
e) Cannot make a statement about relative concentrations based on the given information
a) I guessed
b) Other:______________________________

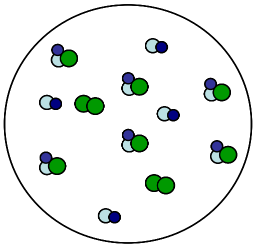


This confusion is reflected in the reasons given for their answers. Answer c is the most common here. NO and NOCl do indeed have the same coefficient, but one is a reactant and one is a product, and thus will not have equal quantities at the equilibrium state.

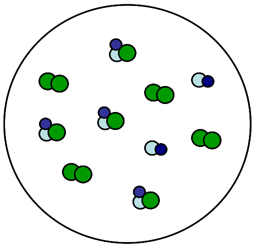
Question 16 is basically the same as question 15, but presented in drawing form.

16) Which of the following could NOT be a representation of the equilibrium state for the reaction described in Question 15?

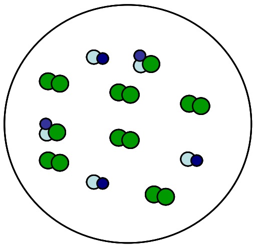
○● - NO; ●● - Cl_2; ○●● - NOCl



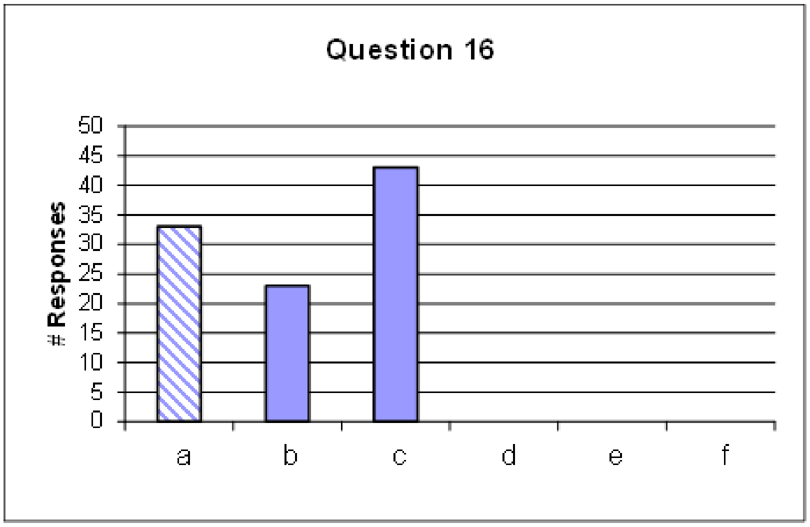
a



b



c



The nearly equal distribution of responses suggests that many of theanswers are just guesses. Again, the use of a graphical presentation did not improve the results.

Chem 1B Midterm Exams

Methodology

The second part of the Data Gathering phase made use of midterm examinations given to Chem 1B students during the course of the class. This midterm exam review made use ofexams written by the course instructor and administered during the Spring 2009 semester. Only the first two midterm exams were used for this review, as they were concerned primarily with topics related to chemical equilibrium (the Le Chatelier Principle, acid/base equilibria and titration, buffer chemistry, etc.); the third examdealt primarily with other topics and was not reviewed.Examswere reviewed after grading by the course the professorbefore they were returned to the students.

The midterm exams were written by the Chem 1B professor, not by the author of this study, and were given to assess student progress in the course material, not to assess students' conceptual understanding as the Diagnostic Exam was. Nevertheless, analysis of the midterm exams was conducted in the hope that they might shed some light on students' level of conceptual understanding, both of the fundamental concepts tested in the Diagnostic Exam but also of chemical equilibrium material to which students had recently been exposed. In this analysis, it was not always clear whether to assign errors to a lack of understanding or to an algorithmic error (incorrect application of a formula, incorrect understanding of terminology, etc.); in many cases, it could have possibly been a combination of the two. Since this distinction is not possible, both options are typically shown and a reasonable conclusion drawn based on the data. Furthermore, no distinction is made as to students' aptitude or application to the material; the statement "students do not understand . . . " could mean that students have acquired a misconception of a chemical idea or principle, or that they just never did the reading; again, it is not possible to make this distinction.

The review and analysis of the results of the midterm exams presented here is very similar to that done for the Diagnostic Exam above.

Results and Analysis, Midterm Exam

Exam 1 consisted of questions on basic equilibrium topics; the equilibrium expression, definitions (K, K_a, pH, types of acids, buffer solutions), and some simple calculations making use of these expressions. The exam questions were mainly of the multiple choice type; two word problems were included but not analyzed. The test average was 67%, with a range from 24% to 100%; 94 exams were analyzed. The analysis of the two Chem 1B course exams served the same purpose as the Diagnostic Exam analysis above; to set a baseline of incoming Chem 1B students' understanding prior to applying an intervention (the worksheets).

Exam questions were ranked with regard to how frequently they were answered incorrectly (**see Figure 6**). A distribution of answers to each question was then generated and analyzed in a manner similar to that in the Diagnostic Exam review in the previous section.

An overall review of Exam 1 shows a wide distribution. One third of the questions were answered incorrectly by half of the students, and ~half of the questions were answered incorrectly by one third of the students. The questions which showed the highest level of incorrect answers are then reviewed in detail. Again, cross-hatched bars indicate correct answers.

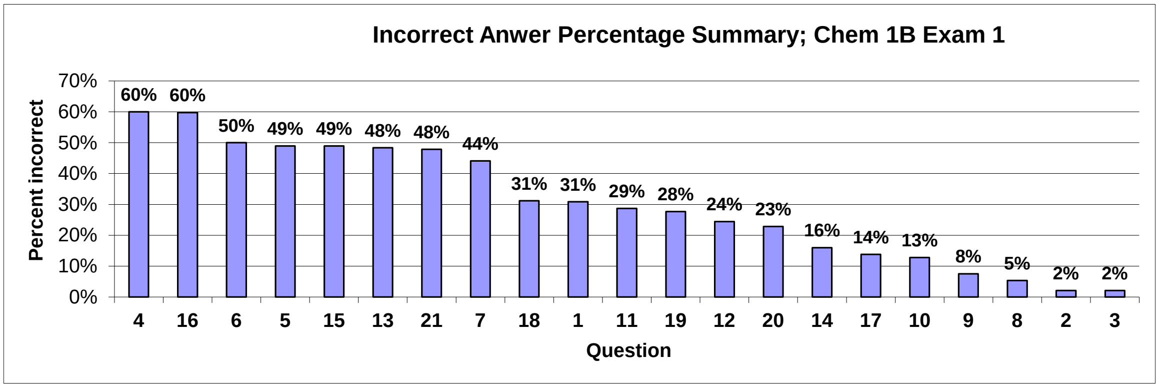


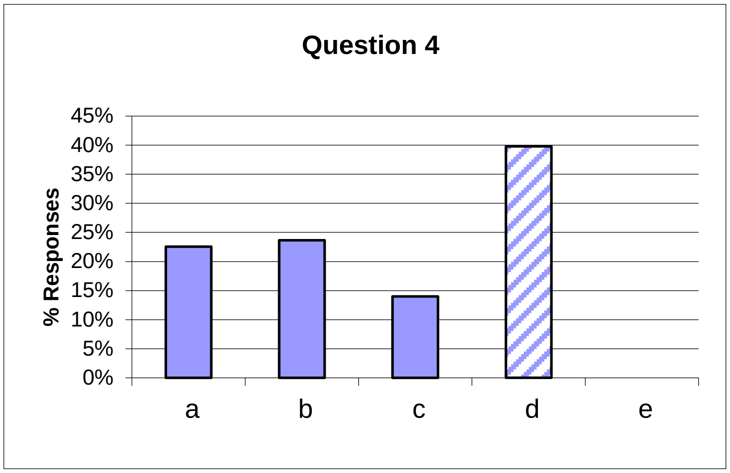
Figure 6: Incorrect Answers to Chem 1B Exam 1

Question 4 (40% Correct)

A 5.0 L flask is filled with 0.25 mol SO_3, 0.50 mol SO_2, and 1.0 mol O_2, and allowed to reach equilibrium. Assume the temperature of the mixture is chosen so that $K_c = 0.12$. Predict the effect on the concentration of SO_3 as equilibrium is achieved by using Q, the reaction quotient.

$$2SO_{3(g)} \rightleftharpoons 2SO_{3(g)} + O_{2(g)}$$

A) $[SO_3]$ will decrease because $Q > K$
B) $[SO_3]$ will decrease because $Q < K$
C) $[SO_3]$ will increase because $Q < K$
D) $[SO_3]$ will increase because $Q > K$
E) $[SO_3]$ will remain the same because $Q = K$

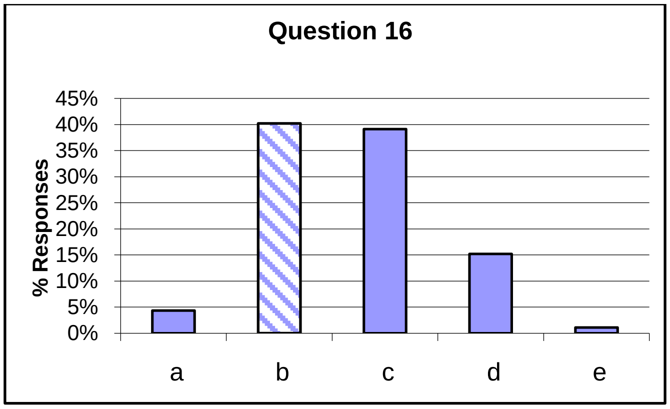


Poor results here could either be due to a lack of understanding of the meaning of Q, or to a lack of understanding of how concentrations would have to shift in order to correct for a non-equilibrium state. The data seems to suggest the former; 47% of respondents incorrectly answered that Q was less than K for this example. Students must have either used an incorrect formula or the wrong values to calculate Q.

Question 16 (40% correct)

The pH of aqueous 0.10 M nitrite ion is 8.17. What is K_b for this base?

a) 4.6E-16 b) 2.2E-11 c) 1.6E-6 d) 1.6E-5 e) 1.2

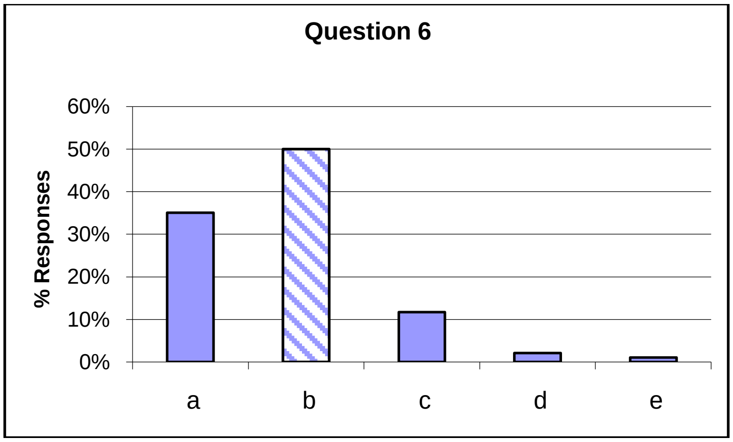


This question requires that that the student understands a number of things; how to construct an equilibrium expression for the nitrite ion ($K_b = (x^2) /(.1-x)$, where x is equal to $[OH^-]$); how to convert from pH to pOH and then to $[OH^-]$; and how to perform the math leading to the calculation of K_b. If the student successfully performs the first two tasks but miscalculates K_b, they will get something close to c; if they forget to square $[OH^-]$ and also forget to include the factor of 0.10, or if they forget the squaring but include the 0.10 factor, they will get d. Since 40% of the class wrongly answered c, they either were unable to form the equilibrium expression, or made a math error in the calculation of K_b.

Question 6 (50% correct)

A room temperature solution is found to have a pOH of < 7. How would you describe the solution?

a) acidic b) basic c) neutral d) not enough information



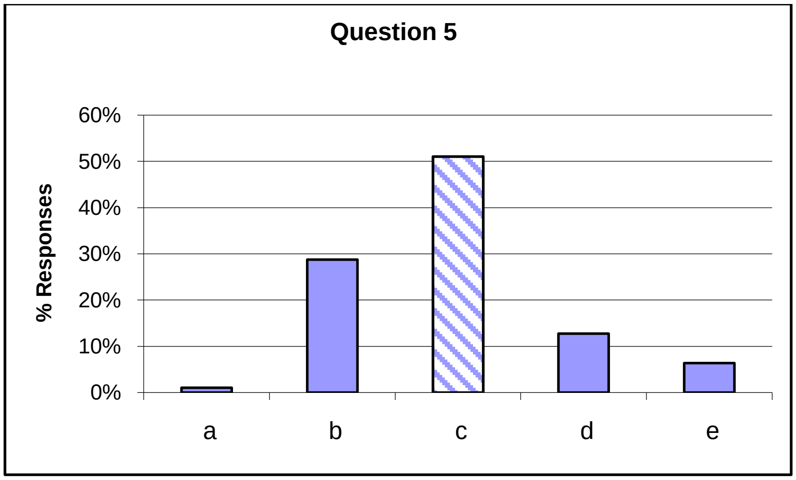
There are two possibilities for the large number of answers for a (acidic). First, students may not yet know the meaning of pH as it applies to acidity and basicity. As this would is a very basic concept that was first introduced in Chem 1A, one might hope this not to be the case. The other possibility is that many students simply misread the problem as " . . . is found to have a *pH* of < 7". The large number of incorrect answers is surprising, though.

Question 5 (51% correct)

Excess $PbBr_2$(s) is placed in water at 25°C. At equilibrium, the solution contains 0.012M $Pb^{2+}_{(aq)}$. What is the equilibrium constant for the reaction:

$$PbBr_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Br^{-}_{(aq)}$$

A) $4.3*10^{-7}$ B) $1.7*10^{-6}$ C) $6.9*10^{-6}$ D) $1.4*10^{-4}$ E) $2.9*10^{-4}$

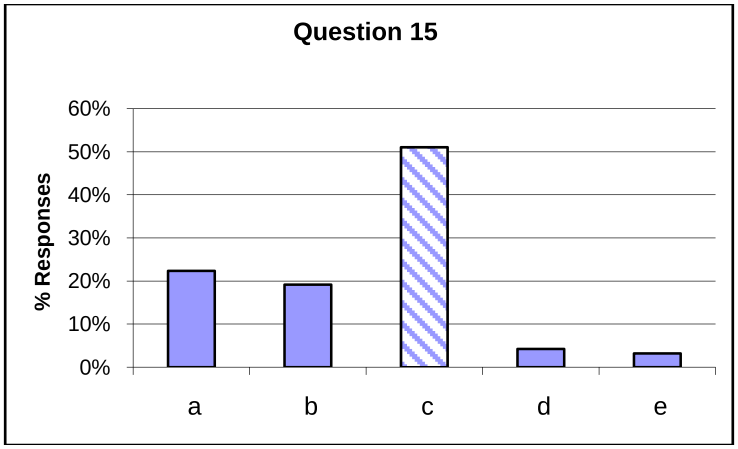


The incorrect answers here show fundamental errors in the calculation of equilibrium constants. The most common error (b) was due to the failure to square the term for [Br$^-$], followed by d, which shows a failure to double the concentration of Br$^-$. This failure can be attributed to a poor understanding of how products and reactants are stoichiometrically related; stoichiometry must be considered when forming the equilibrium equation.

Question 15 (51% correct)

What is the pH of 0.18 M aqueous $NaHCO_2$? (K_a of $HCO_2H = 1.8*10^{-4}$)

a) 3.75 b) 5.50 c) 8.50 d) 10.25 e) 13.26

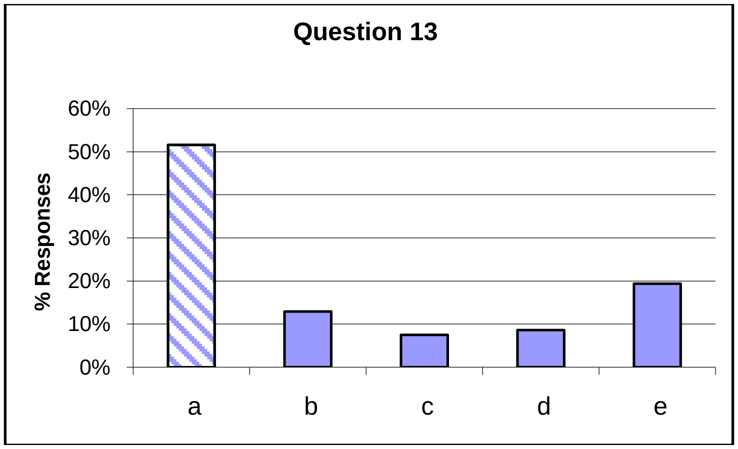


Answer a is obtained by merely using the K_a of formic acid to calculate a pH; a straightforward error, but so fundamentally procedurally wrong that it brings to question the students' basic understanding of the equations for pH–to-K_a-to-pH calculations. Answer b is simply the pOH of the solution, a straightforward mistake, but made by almost 20% of the class.

Question 13 (52% correct)

Which of the following ionic compounds is NOT expected to produce a basic aqueous solution?

a) $KClO_4$ b) NaF c) $NaNO_2$ d) KCN e) $NaCH_3CO_2$

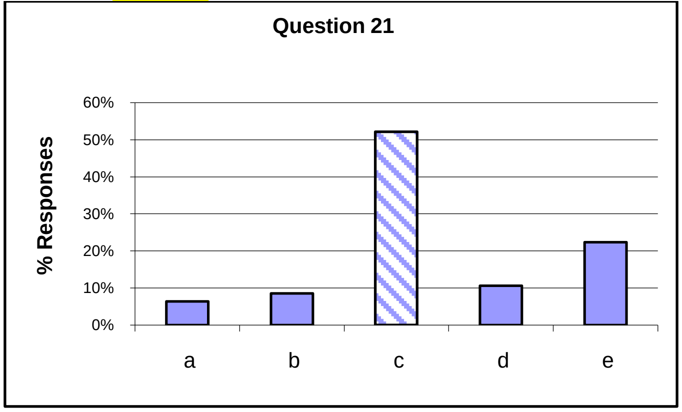


There is nological pattern to the incorrect responses here, as they are fairly evenly divided between the four options, indicating that many students simply do not understand and are just guessing.. This question requires both factual knowledge (which acids are "strong"; the relationship between strengths of conjugate acids and bases) and some intuition (if the conjugate base/acid of this compound is strong, then this compound is a weak acid/base, and its concentration in solution will be inconsequentially small).

Question 21 (52% correct)

A solution is made by combining 50.0 mL of 0.25 M NH_4Br and 35.0 mL of 0.30 M NH_4NO_3. What is the pH of the solution? The K_b for ammonia is $1.8*10^{-5}$.

A) 4.03 B) 4.50 C) 4.91 D) 5.98 E) 9.12



Again, there is no real pattern to the incorrect responses here, but the mathematical complexity of the problem makes educated guessing unlikely. However, it's notable that the most common incorrect answer is e, which supposes that the solution is a basic one. The application of some basic principles here leads to the understanding that the NH_4^+ ion will be the primary determinant of the acidity/basicity of the solution, as the bromine and nitrate ions will be very weak bases and will contribute essentially nothing to the solution pH. If any educated guessing was taking place here, it was missing this basic concept.

Question 7 (56% correct)

Which of the following lists, in order, a weak acid, an amphiprotic acid, a strong acid, and a polyprotic acid?

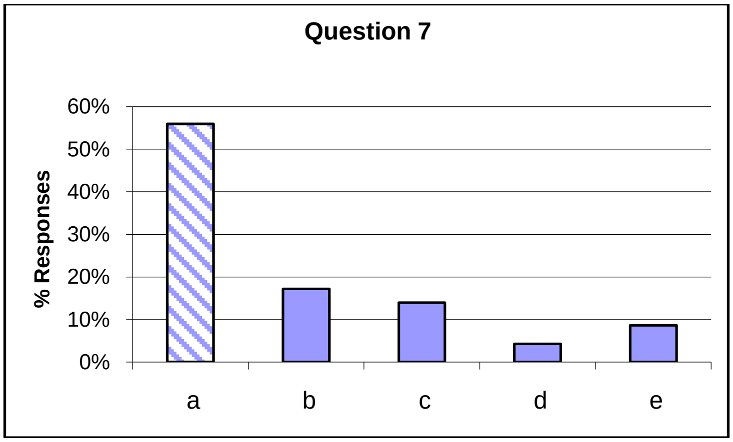
A) H_2SO_3, HSO_4^-, H_2SO_4, H_2SO_4

B) HNO_3, H_2O, HNO_2, H_2CO_3

C) H_2CO_3, H_2SO_4, H_2SO_4, H_2CO_3

D) H_2SO_3, H_2SO_4, H_2SO_4, HSO_4^-

E) None of the above



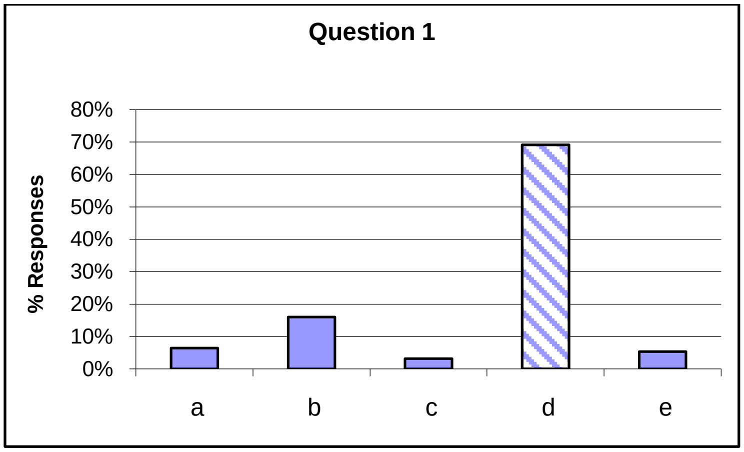
A surprisingly high percentage of students misidentified the very first of the four listed acids (HNO_3) as a weak acid. (answer b; 17%) This is not surprising given the misidentification of acid strength in Question 13. Again, it is discouraging that, despite repeated emphasis throughout Chem 1A, so many students continue to miss this question. As this knowledge is essential to the understanding of acid/base equilibria, the lack of it goes a long way towards understanding the many instances of poor performance in this class.

Question 1 (69% correct)

Which of the following statements is/are correct?

I. For a chemical system, if the reaction quotient Q is greater than K, products must be converted to reactants to reach equilibrium.
II. For a chemical system at equilibrium, the forward and reverse rates of reaction are equal.
III. For a chemical system at equilibrium, the concentrations of products divided by the concentrations of reactants equals one.

A) I only B) II only C) III only D) I and II E) I, II and III



The majority of respondents (69%) selected d as their answer, showing recognition that forward and reverse reactions are still taking place at equilibrium, and that their rates are equal. *This is a marked improvement in the results of questios 14 of the diagnostic exam given on 9/01, where less than 50% agreed with these ideas.*

The most common incorrect answer (b, 16%) showed a misunderstanding of the difference betweenQ and K, supporting the results of question 4, which showed a similar problem.

Results and Analysis, Midterm Exam 2:The exam analysis was performed in the same way as for Exam 1 above. The sample size was 89 exams; the average score was 58.3%, with a high of 99% and a low of 25%. A plot of incorrect answer percentages for Exam 2 can be seen in **Figure 7**.

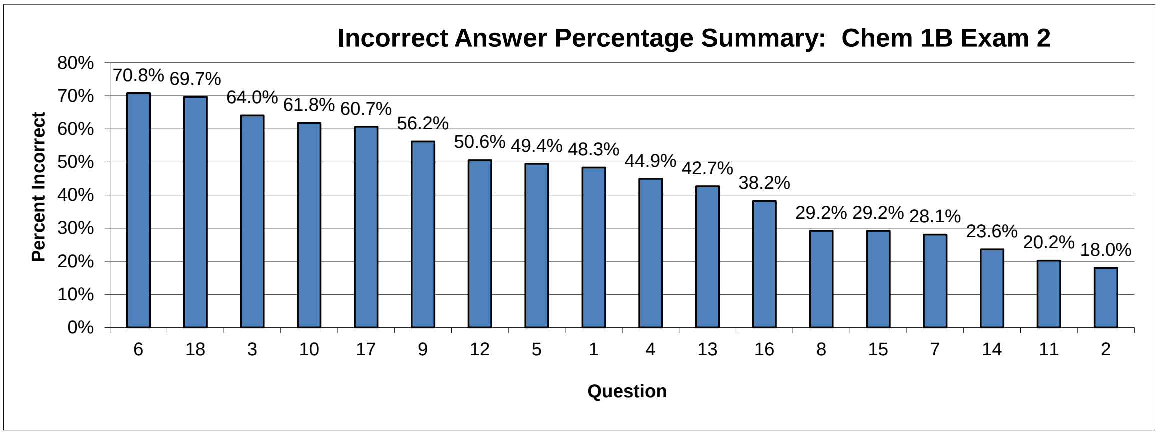
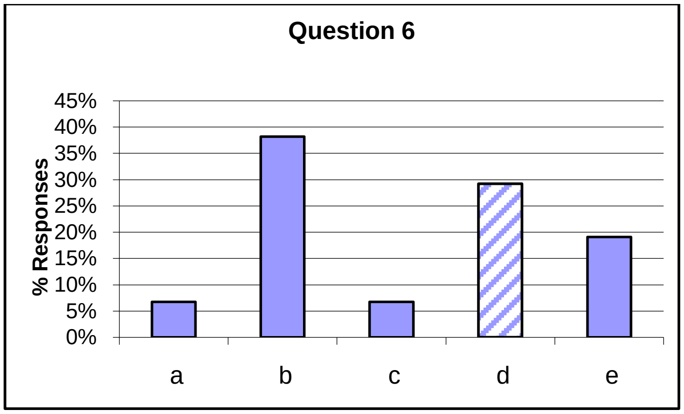


Figure 7: Incorrect Answers to Chem 1B Exam 2

Question 6 (29% Correct)

Which of the following systems at equilibrium would be impacted by changing the pressure?

A) $CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$ B) $H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$
C) $SnO_2(s) + 2CO(g) \rightleftharpoons Sn(s) + 2CO_2(g)$ D) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
E) None of them would be impacted

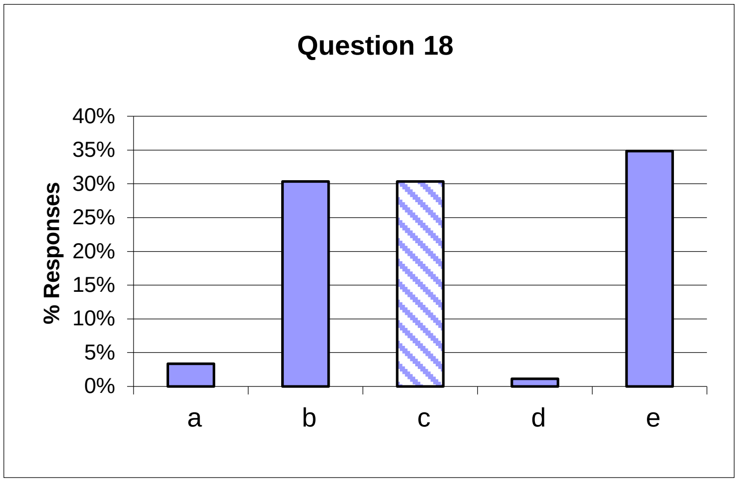


This question may have been a bit misleading for students. The Chem 1B text devotes a section to the effect of *volume* changes on gas-phase equilibria, but there is no explicit mention of the impact of *pressure* changes. Unless the instructor specifically raised this in lecture, it is left to the student to recognize that these two phenomena are inversely related (higher pressure => lower volume).

Answer bwas the most common response. One possible reason for this is that the reaction shown in answer b is the only one in which all reaction participants are in the gas phase, and that, without having reasoned their way through the question, students just made a guess based on the fact that b was the only "true" gas phase reaction.

Question 18 (30% correct)

What is the maximum concentration of $Pb(NO_3)_2$ that can be added to a 1.0 x 10^{-3} M solution of $MgBr_2$ before precipitation of $PbBr_2$ occurs? K_{sp}, $PbBr_2 = 6.6 \times 10^{-6}$
A) 0.00330 MB) 0.00660 MC) 1.65 M D) 3.30 ME) 6.60 M

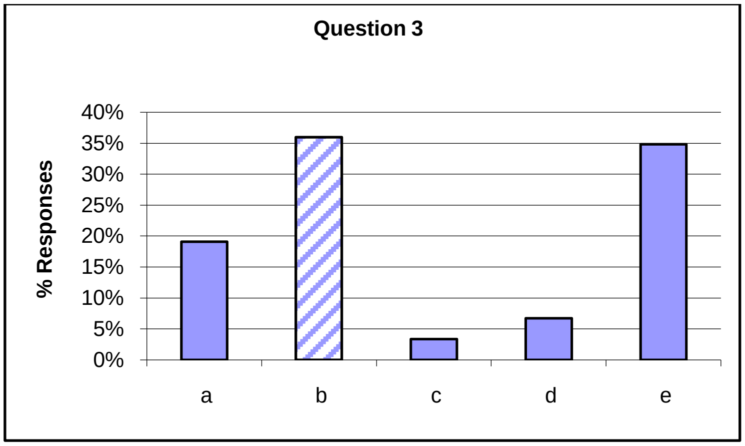


The most commonly chosen answer was e; it is easilyobtained if one fails to recognize that the concentration of bromine ions in solution is twice that of the $MgBr_2$ molecule. Answer b was obtained if you made the above mistake and also failed to square the bromine concentration in the solution for the K_{sp} equation. Both of these errors involve a fundamental misunderstanding of how to form an equilibrium equation for the solubility product constant.

Question 3 (36% correct)

Which of the indicated equilibrium constants does *not* correspond to the given chemical reaction?

A) $K_{a(2)}$ for H_2SO_3 $HSO_3^{-}{}_{(aq)} \rightleftharpoons H^{+}{}_{(aq)} + SO_3^{2-}{}_{(aq)}$
B) K_{sp} for $PbCO_3$ $Pb^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \Leftrightarrow PbCO_{3(s)}$
C) K_f for $[Zn(NH_3)_4]^{2+}$ $Zn^{2+}{}_{(aq)} + 4\ NH_{3(aq)} \rightleftharpoons [Zn(NH_3)_4]^{2+}{}_{(aq)}$
D) K_b for CH_3NH_2 $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$
E) All of the above are correct

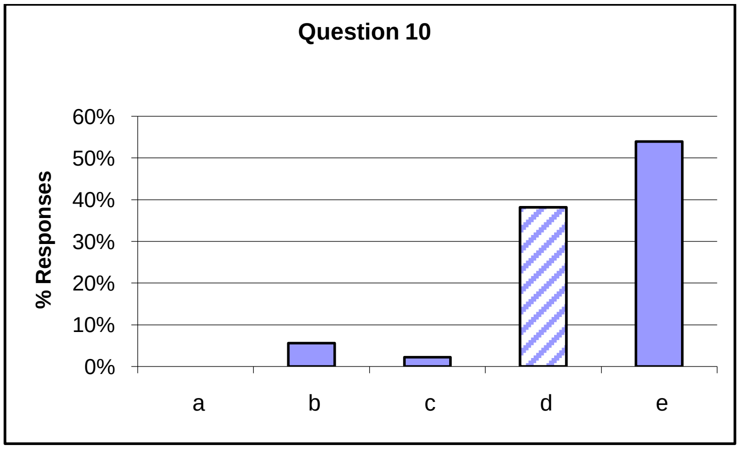


The most common error here is the simple inability to recognize the equation for a solubility product constant; 35% of respondents (those answering c, d or e) were unable to recognize equation b as anincorrect form of the equation. The next most common answer (a, 19%) indicated a confusion over the meaning of“$K_{a(2)}$” .

Question 10 (38% correct)

A 50.0 mL sample of 0.150 M $HClO_4$ is being titrated with 0.200 M NaOH. What is the pH after 10.0 mL of the base have been added?

A) 0.092B) 0.82C) 0.96D) 1.04E) K_a for $HClO_4$ is required to answer this question



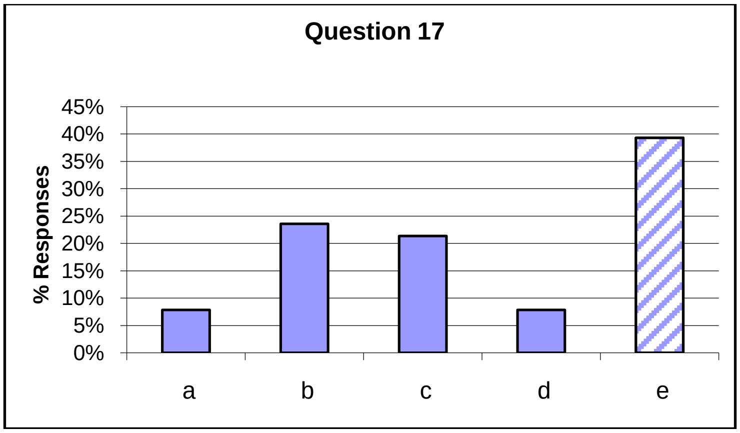
Over half the class selected answer e, failing to identify a strong acid. As noted previously, the recognition of strong acid/weak acid behavior and the six strong acids is crucial to the success of the Chem 1B student;this question offers a good example of why that is. Students who correctly answered this question averaged 68% on the exam, 10 points higher than the class average. Students who chose e as the answer averaged 54% on the exam, 14 points lower. Of the 47 students who chose answer e, only one scored higher than 76%.

Another possible reading of this data can be made from the exam score correlation above. Since the students who selected e for this question all did poorly on the exam, they also missed many other questions. In other words, rather than expressing a unique problem, the data may be saying that students who selected answer e just did not apply themselves in general, and that ignorance of the strong acid rule is just one example of that; it's the "canary in the coal mine" question for general chemistry aptitude.

Question 17 (39% correct)

1) What is the molar solubility of $Fe(OH)_3$ in a solution with a buffered pH of 3.50? (K_{sp}, $Fe(OH)_3 = 6.3 \times 10^{-38)}$

A) 6.9×10^{-28} MB) 2.0×10^{-27} MC) 2.2×10^{-10} MD)7.4×10^{-8}E) 2.0×10^{-6} M

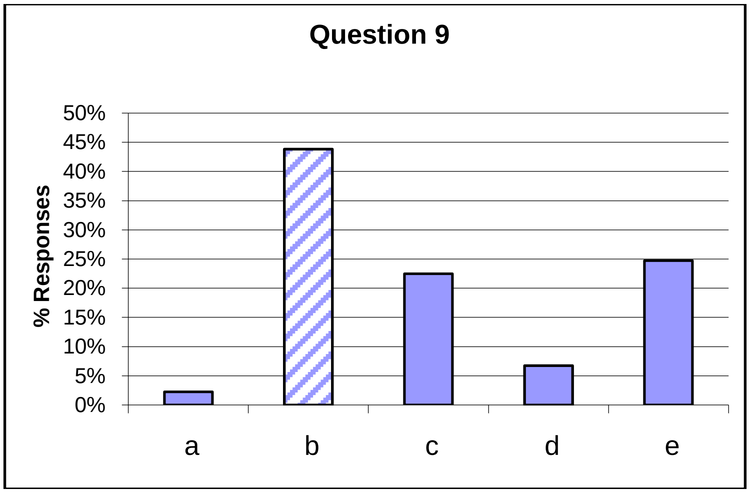


This question was a standard calculation using the formula for the solubility product constant of iron hydroxide, but with a twist; since the solution in the problem was identified as a buffer, the assumption must be made that the pH of the solution remains the sameeven as additional iron oxide is added. Once this assumption is made, [OH$^-$] for a 3.5 pH solution is calculated, and this figure is inserted into the formula for K_{sp} to calculate [Fe^{3+}], which equals [Fe(OH)$_3$], the concentration of iron oxide at saturation (the molar solubility). Incorrect answers all correspond to one or more errors in this calculation. Answer b was obtained if [OH$^-$] was not cubed in the K_{sp} equation or if [H$^+$] was used instead of [OH$^-$]; it's not clear how c was obtained.

Question 9 (44% correct)

A sample of CH_3NH_2 is titrated with HNO_3. Which of the following species is primarily responsible for the pH at the equivalence point?

A) CH_3NH_2 B) $CH_3NH_3^+$ ions C) HNO_3 D) NO_3^- ions
E) $CH_3NH_2/CH_3NH_2^+$ buffer F) None of the above



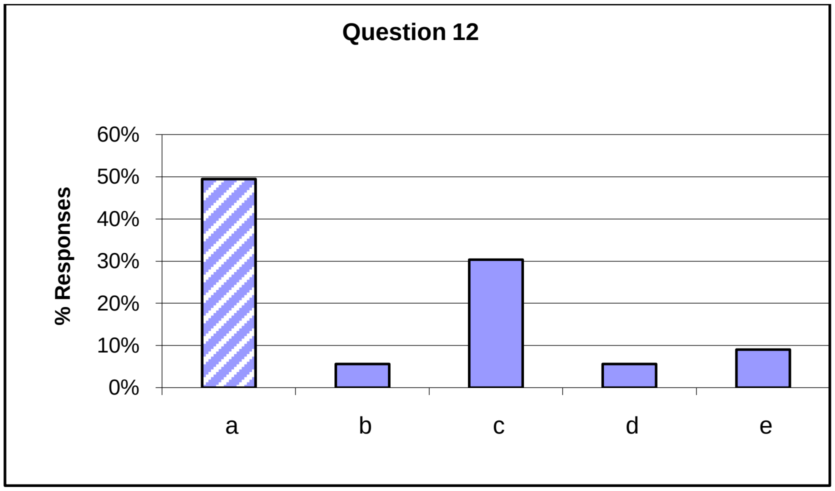
It is difficult to identify any specific misconceptions here, as a correct response requires a good overall understanding of the titration process. The student needs to correctly identify the active species, write the neutralization reaction, understand the chemical meaning of “equivalence point”, and apply this knowledge in a non-formulaic, intuitive manner. The form of this question is an excellent one for identifying true understanding of the chemical principles behind titration, as it precludes the use of textbook formulas recalled from memory and requires the student to think through the chemical processes actually occurring.

Question 12 (50% correct)

1) Without doing any calculations, which of the following should have the greatest molar solubility?

K_{sp}, $AgC_2H_3O_2$ = 1.9 x 10^{-3} K_{sp}, CuBr = 6.3 x 10^{-9}

A) $AgC_2H_3O_2$ in 0.10 M HNO_3	B) $AgC_2H_3O_2$ in 0.10 M $NaC_2H_3O_2$
C) $AgC_2H_3O_2$ in pure water	D) CuBr in 0.10 M HNO_3
E) CuBr in 0.10 M NaBr	F) CuBr in pure water



This is a question which requires a good conceptual understanding of a number of principles; the meaning and mathematical form of K_{sp}, the common-ion effect, weak/strong acid behavior, and neutralization effects on chemical equilibria. The most common wrong answer, c (30%), requires that students believe that the presence of nitric acid somehow *reduces* silver acetate solubility. My interpretation here is that, without a thorough understanding of strong/weak acid behavior and neutralization, the student is left only with the vague conception that *anything* you add to a solution will reduce the solubility of a metal salt. Not a strong *misconception*; just a lack of *any* conception.

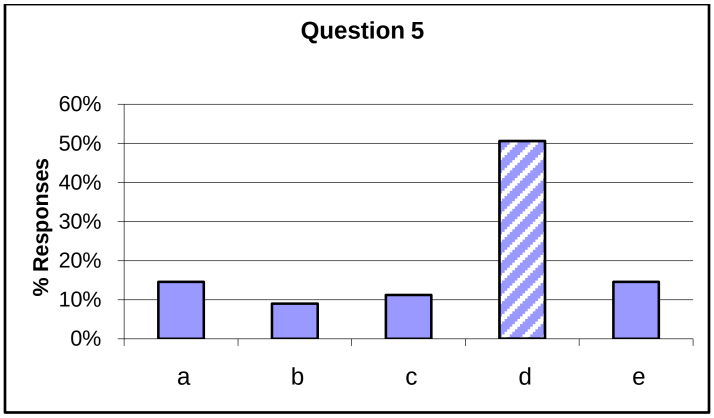
Question 5 (51% correct)

Which of these statements is true when the following endothermic reaction is at equilibrium?

$$NiO_{(s)} + CO_{(g)} \Leftrightarrow Ni_{(s)} + CO_{2(g)}$$

A) Adding $NiO_{(s)}$ will increase the amount $CO_{2(g)}$.
B) Decrease the volume of the container will result in more $CO_{(g)}$.
C) Increasing the temperature will increase the amount of $CO_{(g)}$.
D) Adding more $CO_{(g)}$ will result in the formation of more Ni(s).

E) None of the above are true.



This is a straightforward Le Chatelier question, so the large number of incorrect answers is surprising. Another way to look at this problem, though, is from a simple reaction kinetics point of view as reviewed in Worksheet 2. It is possible to look on this reversible reaction as two competing reactions (forward and reverse), each driven by its own rate constant. If we know that the forward reaction rate is a function of the concentration of $CO_{(g)}$, it is a simple step to realize that if the concentration of CO is increased, the forward reaction rate will increase, and as a result additional product will be generated. If one is equipped with a basic understanding of how a reaction works, many other pieces of the puzzle fall easily into place.

Summary of Data Gathering Analysis

Diagnostic Exam

Despite their Chem 1A training, stoichiometry and kinetics questions seemed be difficult topics for entering Chem 1B students.Particularly troubling was studentsinability to recognize basic stoichiometric relationships and how they affected product and reactant concentrations over time (Questions 5 and 6). And even though thestudents had little exposure to reversible reactions or chemical equilibrium in Chem 1A, they seemed to be able to figure these questions after a brief review of basic kinetics principles, although often with an inability to provide a reason for their answer (see **Figure 8** below). In this summary of test answers, the total bar height is an indication of the overall percentage of correct answers given. Several of the bars are divided into two parts, ; the lower section of the bar shows the percentage of correct answers for which the correct reason is given, and the upper section of the bar shows the percentage of correct answers for which an incorrect reason is given. Thus the total bar height represents to overall percentage of correct reasons given.

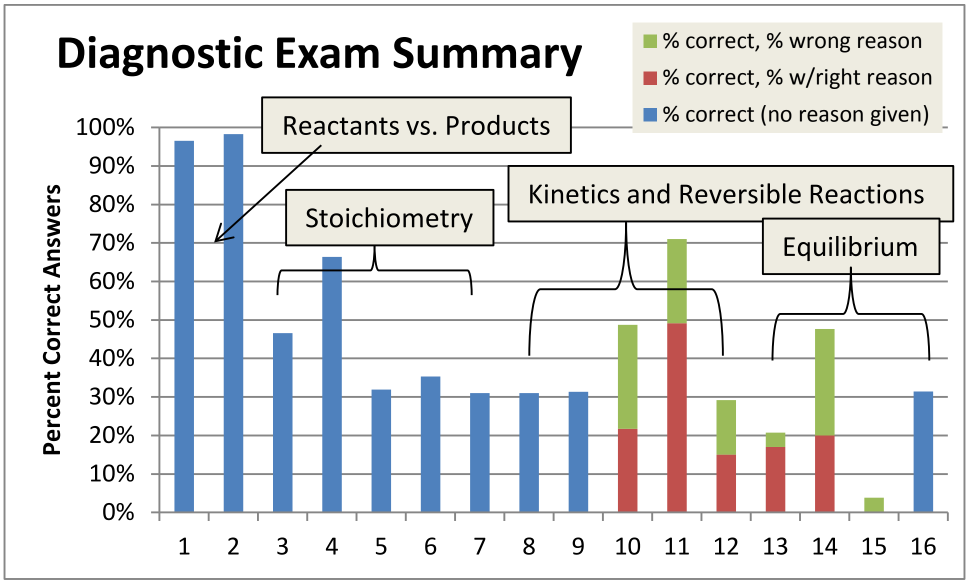


Figure 8: Summary of Answers to Diagnostic Exam Questions

In summary, entering Chem 1B students showed a general lack of understanding of the principles and application of stoichiometry and reaction kinetics. This is reflected in the high level of incorrect answers to the Diagnostic Exam questions as well as in the incorrect reasoning shown for the correct answers. Recommendations on further research into this area are provided in the chapter on Conclusions and Recommendations to follow.

Midterm Exams

The results and analysis of the midterm exams were based on a total of 18 questions from two different exams, which dwelt on a wide range of chemical equilibrium concepts (the Le Chatelier principle, the equilibrium constant and the reaction quotient, solubility and K_{sp}, perturbation of equilibrium, etc.). Given this diversity in results, acohesive summary of results and analysis is difficult to compile. As mentioned above, the distinction between conceptual errors and algorithmic errors is often difficult to make, and sometimes students' answers seem to reflect simple guessing. These issues pose a problem when attempting to draw general conclusions from individual analyses.

Some of the most notable instances of possible conceptual misunderstanding are listed below.

- Students unable to distinguish Q and K
- Stoichiometry errors when forming the equilibrium expression
- Lack of understanding of the weak vs. strong acid/base principle
- Improved understanding of the dynamic nature of equilibrium
- Lack of understanding of expressions for K_{sp} and $K_{a(2)}$

- Misunderstanding of the meaning of “buffer”
- Misunderstanding of the meaning of the Le Chatelier principle

Chapter 5

INTERVENTION 1 (FOCUS GROUPS)METHODOLOGY, RESULTS & ANALYSIS

Overview: Focus Group Logistics

An Intervention was developed for use with Chem 1B students in the form of four worksheets. The worksheets were developed by the author of this study for the purpose of evaluating and improving student's conceptual understanding of key topics related to chemical equilibrium. The worksheets were based on a set of learning objectives for each of the following areas:

Worksheet 1: Stoichiometry and Reaction Kinetics
Worksheet 2: Basics of Chemical Equilibrium
Worksheet 3: Acid/Base Equilibrium
Worksheet 4: Buffer Chemistry

Each worksheet was designed to be completed in about one hour, and to be consistent with the techniques advocated by the constructivist school of education reviewed in the Literature chapter above; student-driven, hands-on collaborative learning, together with an instructor/guide using the ZPD conceptin a guided-inquiry process. Questions are designed to be asked in an open-ended way in order to foster discussion of the various possible answers, and to promote *understanding* of the problem.

The worksheets were developedto be used in a face-to-face setting, with both an instructor and students present. Worksheet use was to be eventually integrated into the course structure of Chem 1B, either as part of a separate discussion group adjunct to the class or into one of the lab sections. Prior to class integration, as a trial run, small "focus groups" of four or fewer students were used to practice worksheet use and to refine their content. Since the focus groups had no official standing in the course curriculum, focus group members were solicited as volunteers. Volunteers were solicited for the focus groups at Chem 1A lectures at the end of the Fall 2009 semester and at the first week Chem 1B lectures. Focus group sessions were scheduled beginning in January of 2010 and ran through July.

Each Focus Group session reviewed one of the four worksheets. Demographic data was obtained for each student, including information on performance in previous chemistry classes. The meetings were held with a whiteboard at hand to allow use of drawings and diagrams to be used in the discussion. A voice recorder was used during the meeting and used to augment the detailed notes taken immediately after the session. All of the meeting information was subsequently compiled into a report, along with analytical observations of the instructor. Transcripts of the Focus Group sessions can be found in Appendix D.

In the original experimental design, the bulk of the data was to have been acquired in Focus Groups or some similar instructor-guided inquiry sessions, with the intent that some Chem 1B lab time might eventually be dedicated to obtaining this data. The worksheets were to be integrated into lab sessions in the early part of the semester in order to get data from a large sample and perhaps use these sessions as "control" and "intervention" groups for data analysis. Logistical issues prevented implementation of this design; due to a change in the staffing of the Chem 1B lecture, it was not possible to conduct any studies within the course itself. Instead, a backup plan was implemented in which the worksheets were posted on the Chem 1B course website for student download and completion. This was carried out in the fall of 2010 (see Table 3). Worksheets were posted approximately 10 days before course exams so that they could be used for exam study. A small amount of course credit was given to students who submitted a completed worksheet, and answer keys were posted after worksheet submission and several days prior to the exams. Worksheets 1 and 2 were posted prior to Exam 1; worksheets 3 and 4 were posted prior to Exam 2. Fifteen copies of each of the first two worksheets were completed and returned, or ~13% of a class population of 118. Twenty-nine copies of each of the second two worksheets were completed and returned, or ~27% of a class of 109 (a reduced class size after drops following the first exam). The worksheets contained no multiple-choice questions which could be quantitatively reviewed in the same manner as the exams. Answers consisted mostly of written replies to open-ended questions, and analysis was mainly qualitative in nature, featuring commentary about those questions which were not understood. Some comparison between submitted worksheets and exam results was possible, but was not the primary methodology used in this study, which relies mainly on the qualitative input from student interviews.

It should be noted that, when the worksheets were being posted, it was discovered that the versions developed for classroom use would need substantial revision due to the change in their usage. The absence of an instructor led to the need to make changes to questions which had already gone through multiple revisions during the focus group session. By the time these changes were complete, there had been a total of 51 revisions of the four worksheets.

Detailed descriptions of the individual worksheets appear below. An archive of the final revisions of the worksheets along with answer keys can be found in Appendix B.

Worksheet Development

The Equilibrium Machine

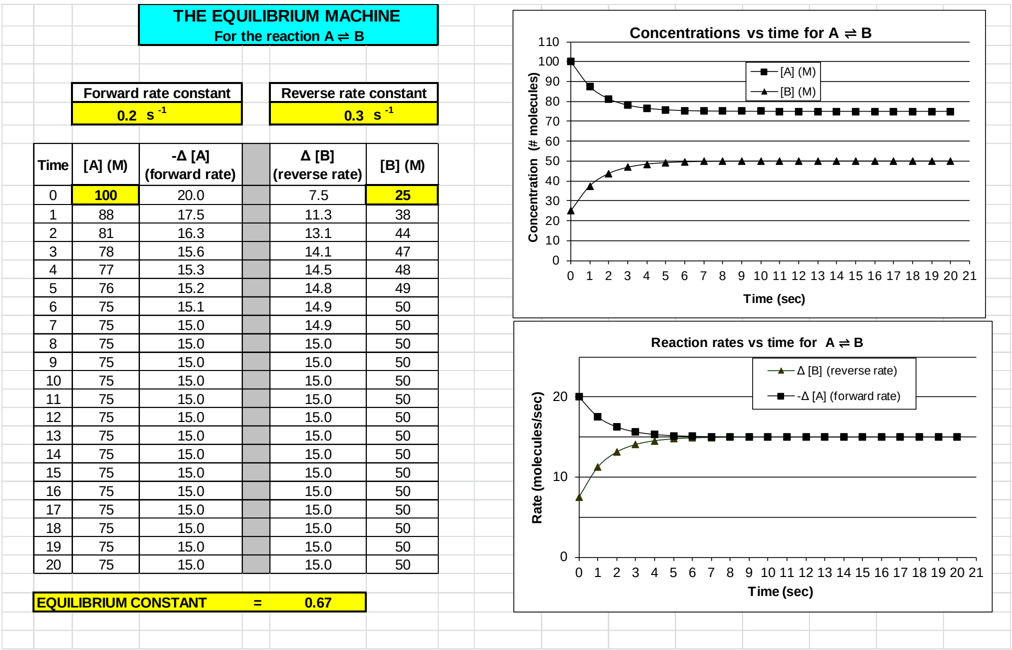
An interactive Excel workbook, referred to here as the “Equilibrium Machine”, was constructed to model hypothetical chemical reactions based on variable student inputs of initial conditions. Calculations of instantaneous reactant and product concentrations and of reaction rates are performed as a function of time and the results displayed in spreadsheet form.

Two versions of the spreadsheet are described in more detail in Appendix C, and a digital copy is provided on the CD attached to this document. One version performs calculations for a hypothetical irreversible reaction, which proceeds essentially to completion (e.g., a strong acid dissociation), based on an input of initial reactant and product concentrations and a reaction rate constant. As this is the only type of reaction presented in Chem 1A, this spreadsheet is used only in conjunction with the initial worksheet, which is essentially a review of prerequisite knowledge for the study of chemical equilibrium.

The second version performs calculations for a reversible reaction. This version allows inputs of initial concentrations, but also allows student inputs of both forward and reverse reaction rate constants, allowing students to observe the kinetic and equilibrium states of various types of chemical equilibria. Only the simplest equations are represented; A → B for the irreversible reaction, and $A \rightleftharpoons B$for the reversible reaction.

In addition to computing and displaying values, the results are instantly plotted, allowing a graphical view of the impact of changes to the equilibrium state of a chemical system. Two plots are displayed; one shows changes in the concentrations of reactants and products over time, and one shows changes in instantaneous reaction rates over time. Note also that the equilibrium constant is calculated and displayed for each reaction, allowing discussions on the effects of various input parameters on the equilibrium state.

A sample page of the Equilibrium Machine, with spreadsheet and plots for a reversible reaction, can be seen in **Figure 9**. This figure shows the progression of the reversible reaction $A \rightleftharpoons B$, with initial concentrations of $[A] = 100M$ and $[B] = 25M$, and forward and reverse reaction rate constants of $k_f = 0.2s^{-1}$ and $k_r = 0.3s^{-1}$. Any one or more of these four parameters can be changed by the student, and the resulting change in two plots will appear instantaneously. Additional discussion on the use of the Equilibrium Machine will be included in the Worksheet sections below. A full archive of the Equilibrium Machine appears in Appendix C, and a digital copy is provided on the CD attached to this document.



THE EQUILIBRIUM MACHINE
For the reaction A ⇌ B

Forward rate constant	Reverse rate constant
0.2 s^{-1}	0.3 s^{-1}

Time	[A] (M)	-Δ [A] (forward rate)		Δ [B] (reverse rate)	[B] (M)
0	100	20.0		7.5	25
1	88	17.5		11.3	38
2	81	16.3		13.1	44
3	78	15.6		14.1	47
4	77	15.3		14.5	48
5	76	15.2		14.8	49
6	75	15.1		14.9	50
7	75	15.0		14.9	50
8	75	15.0		15.0	50
9	75	15.0		15.0	50
10	75	15.0		15.0	50
11	75	15.0		15.0	50
12	75	15.0		15.0	50
13	75	15.0		15.0	50
14	75	15.0		15.0	50
15	75	15.0		15.0	50
16	75	15.0		15.0	50
17	75	15.0		15.0	50
18	75	15.0		15.0	50
19	75	15.0		15.0	50
20	75	15.0		15.0	50

EQUILIBRIUM CONSTANT = 0.67

Figure 9: Sample Page from the Equilibrium Machine

Worksheet 1: Reaction Kinetics and Stoichiometry

This worksheet consists of reviews of basic Chem 1A concepts required for the study of chemical equilibrium. Accordingly, this worksheet was designed to be most effective for use with students who have completed Chem 1A but have either not yet entered Chem 1B or are in the first several weeks of the course, prior to the first exam.

The Learning Objectives for Worksheet 1 are as follows:

- Students will be able to identify the relative rate of consumption and generation of reactants and products given a balanced equation for a chemical reaction
- Students will be able to describe the change in concentrations of reactants and products over time
- Students will be able to describe the change in reaction rates over time

The worksheet consists of three exercises. In the first, the hypothetical reaction

$$2AB + C_2 \rightarrow 2AC + B_2$$

is presented, and a number of questions are posed which require an understanding of how the stoichiometric relationships between various reactants and products affect changes in their concentrations as the reaction progresses. For example;

a) Suppose that we mix equal amounts of AB and C_2 to initiate a reaction. Which reactant do you predict will be consumed more quickly (at a faster rate)? Which product is produced more quickly? Why?

Note that this reaction is shown as an irreversible reaction; it is assumed that the concept of reversible reactions has not yet been introduced to students. Students are asked to provide answers as well as a justification for those answers. Answers are never directly provided by the instructor, but are elicited from the students through a series of leading questions and analogies. If an incorrect understanding of the stoichiometric relationships in this reaction remains lodged in a student's mind, a short exercise is provided which can be simply executed on a whiteboard or sheet of paper, where individual reactant molecules are shown and then eliminated as new product molecules are written (see Worksheet 1 in Appendix B).

The second exercise in Worksheet 1 concerns the kinetic mechanism of chemical reactions. Concepts of reaction rate and rate constant are discussed, and the relationship between rate, rate constant, and concentration are reviewed. An exercise is then performed in which students are supplied with initial conditions (rate constant and reactant concentrations) for a simple irreversible reaction, and are asked to recalculate these parameters as they change with time. Their calculated concentrations are plotted against time so that they can develop a better intuitive understanding of how concentrations and reaction rates change with time. This exercise involves a great deal of interface with the instructor, as the understanding of reaction kinetics varies greatly in students entering Chem 1B, and the amount of guidance required (the setting of the ZPD, so to speak) needs to be adjusted every meeting. This is an example of an exercise which did not translate well from focus group to posting/download use. Students required some guidance through the first set of calculations in order to achieve full understanding and complete the exercise; this was not always possible with the posting/download method, and as a result, some questions were just left blank.

The third exercise makes use of the Equilibrium Machine. After students have gained a basic understanding of reaction kinetics in the first two exercises, they are taught how to use the Equilibrium Machine to more quickly develop plots of reaction parameters over time. After a brief introduction to the Equilibrium Machine, students are asked change the initial conditions, and then to changethose conditions and explain how these changes relate to the corresponding changes in the plots.

Continual use of this worksheet resulted in many refinements. Twelve revisions of the worksheet were made; ten for use in Focus Groups, and another two for posting/download. The full archive of this worksheet can be seen in Appendix B, as well as on the CD attached to this document.

Worksheet 2: Reversible Reactions and Chemical Equilibrium

The first exercise in the worksheet introduces the concept of the reversible reaction. A reversible reaction is presented as two simultaneous reactions, a forward and a reverse reaction, each having their own rate constant. A brief review of the rate equations follows, and students are asked about the distinctions between irreversible and reversible reactions. A graphical representation of reactants and products is included as part of a question asking about the components that appear as products of a reversible reaction.

The Learning Objectives for Worksheet 2 are as follows:

- Students will be able to describe the basic characteristics of reversible reactions and compare them with the irreversible reactions taught in Chem 1A
- Students will be able to read and understand plots of concentration vs. time and reaction rate vs. time for reversible reactions
- Students will be able to describe the meaning of the term "chemical equilibrium" and identify the equilibrium regions on concentration and rate plots
- Students will be able to articulate the dynamic nature of chemical equilibrium

In the second exercise, the concept of chemical equilibrium is introduced. At this point, students move beyond prior knowledge into an area which may be totally new to them. Some instructor leeway is called for here.Some students may require an introductory learning section, including a review of the kinetics material from Worksheet 1, and the time required to complete this section will vary widely depending on the students' backgrounds. In the course of the focus group reviews, some groups were unable to complete this worksheet in the one hour allotted.

The second exercise relies heavily on the Equilibrium Machine. Students bring up the plot of an irreversible reaction, then of a reversible reaction, and are asked to describe the differences between the two. They are asked how these differences are related to presence of a reverse reaction in the reversible case, and are asked about the dynamic between the forward and reversible rate constants. They are asked to change the rate constants, to predict the effects on the plots, and to explain the actual results. Students are offered a graphical description of the meaning of the rate equation and its relationship to the equilibrium state. They are asked how the two rate constants result in reactant-favored vs. product-favored reactions, and how they relate to the equilibrium constant.

In the third exercise, some basic calculations are done relating initial and final concentrations to the equilibrium constant. The exercise focuses on deriving and understanding these relationships rather than on the calculations, which are simple and can be done without a calculator. This section is intended to give the student a conceptual understanding of the many equations they will face in Chem 1B relating changes in component concentrations to the equilibrium constant.

Nineteen revisions of this worksheet were made, seventeen for focus group use, and two more for posting. A full archive of this worksheet can be found in Appendix B, as well as in the attached CD.

Worksheet 3: Basics of Acid/Base Equilibria

Worksheet 3 is a step up for students just entering Chem 1B, but it again makes use of material introduced in the previous worksheet, as well as the basic concepts of acids and bases acquired in Chem 1A. The main theme of the worksheet is that acid-base equilibrium is no different than the general chemical equilibrium studied in the previous worksheet. Some different terminology is used (much of which was acquired in Chem 1A), and some different concepts introduced, but otherwise, the same concepts apply.

The Learning Objectives of Worksheet 3 are as follows:

- Students will be able to describe the difference between a strong and weak acid or base

- Students will be able to articulate the conceptual and mathematical relationships between pH, K_a and K_{eq}
- Students will be able to describe the solution of the equation for K_a given initial conditions

The first exercise is a simple, short review of the terms “strong” and “weak” as applied to acids and bases, a concept introduced in Chem 1A. It is included because its topic is so central to the understanding of acid-base equilibria, and also because the study of Chem 1B exams indicates that it is still widely misunderstood by general chemistry students. The exercise simply asks for a definition of strong and weak acids and bases, and asks the students to list the six strong acids. Not a very constructivist exercise, but a needed one.

The second exercise introduces the topic of acid/base equilibrium as a subset of general chemical equilibrium. It introduces the terms pH, K_a, and K_b and talks about the relationship between them in a conceptual manner. It then asks students a number of questions about this relationship. The role of water in acid-base equilibrium is also discussed.

In the third exercise, the relationship between K_a, pH, and concentration is examined in a graphical manner, and questions are posed which require an understanding of this relationship. A plot of pH vs. molarity for acetic acid is used to pose a number of questions on why the plot looks like it does, and how the plot might be used to generate other information (K_a for acetic acid and the stoichiometry of acid dissociation). After students become comfortable with reading and interpreting this plot, a similar plot is introducedwhich contains curves for a number of different weak acids. More questions are asked about K_a, molarity and pH of the various acids. The presentations of curves representing multiple acids on the same plot allows for discussion about how and why these parameters vary in different compounds.

Twelve revisions of this worksheet were made, eleven for focus group use and one for posting. A full archive of this worksheet can be seen in Appendix B, and on the attached CD.

Worksheet 4: The Common Ion Effect and Buffer Chemistry

This worksheet introduces several new concepts to the Chem 1A graduate: The acid/base conjugate pair, the common ion effect, and the buffer solution. As many new concepts are introduced here, instructors may greatly vary the use of this worksheet depending on the knowledge base of the target group. For Chem 1A graduates with no exposure to these new principles, a focus group is called for, with extensive use of the ZPD by the instructor. The current state of student knowledge can then be understood, and the new concepts introduced accordingly. If the worksheet is used in the posting/download method during the course, it must be used after these principles have been introduced in Chem 1B, and prior to the course exam in which they appear (currently the second exam, held in the middle of the semester), so that students can use the answer key for exam study.

The Learning Objectives for Worksheet 4 are as follows:

- Students will be able to name the conjugate of an acid or a base
- Students will be able to describe the Common Ion Effect
- Students will be able to articulate the purpose and contents of a buffer solution
- Students will be able to describe the fundamental workings of a buffer solution

The first exercise begins with an examination of two principles basic to the study of buffer chemistry; the concept of conjugate acid/base pairs, and the chemistry of the common ion effect. A drawing of the proton exchange which takes place in an acid/base transition is shown, and a number of questions are posed to ensure that it is understood. The concept of the conjugate pair is introduced, using this drawing as a reference, and additional questions are posed. Once this reaction is understood, questions are asked about the impact of the addition of additional quantities of basic ions to a weak acid solution and how this changes the makeup of the solution based on K_a of the acid. This process is identified as the “Common Ion Effect”. The understanding of this process prepares the students for the introduction of the concept of buffer chemistry.

The second exercise is an introduction to the concept of buffer chemistry. It makes use of the conjugate pair diagram from the first exercise, as well as other ideas formed there, to explain the makeup of a buffer solution and how it works. A picture is included illustrating how a buffer solution is made from solutions of a weak acid and its conjugate base, what the resulting solution looks like, and how it works. Some probing questions are asked to determine the level of the student understanding of this concept.

In exercise three, a PowerPoint presentation is used to illustrate how buffers are formed and how they function to mitigate pH changes in a solution. Questions are designed to generate a discussion in the focus group of the nature of this principle. The PowerPoint slide set can be seen in the CD attached to this document.

Eight revisions of this worksheet were made, seven for focus group use and one for posting/download. A full archive of this worksheet can be seen in Appendix B. as well as in the attached CD.

Focus Groups: Results and Analysis by Worksheet

Rather than review each Focus Group session individually, this analysis will review the Focus Group data by worksheet. This will allow responses of all Focus Groups to be compiled and reviewed together. Detailed transcriptions of the Focus Group meetings, organized by session, can be seen in Appendix D.

A summary of the demographicsfor the focus group participants may be seen in **Table 5** below. These names will be referred to in the body of the analysis below.

Table 5: Demographic Data for Focus Group Participants

Name	Class	Age	Major	HS Chem	Chem 4	Chem 1A	Prior 1B?	WS Used	Chem 1B
S1	Sr	27	Bio	No	F07/B	F08/B	No	1, 2,3	S10/C
S2	Soph	19	Chem	B+	S09/B+	F09/A	No	1, 2	S10/C
S3	BS 2	35	BChem	No	S09/A	F09/B	No	1,2,3	S10/B
S4	?	24	Dietetic	A	No	F04/C	F09/F	1	S10/C+
S5	Jr	26	Bio	C+	F09/B	S10/B+	No	1, 2	F10/C+
S6	Sr	41	Bio	No	No	S08/C	F09/D	1, 2	F10/B+
S7	Jr	20	PreMed	B	No	S10/B	No	1, 2	F10/C-
S8	Jr	22	Bio	B	F09/B-	S10/B+	No	1, 2	F10/B

Note that, while the eight students listed all attended focus groups in the spring and summer of 2010 and are all included in the study following, they attended Chem 1B at different times, as indicated in the table. Students 1 through 4 attended focus groups in the Spring of 2010, while enrolled in the Spring 10 Chem 1B class. Students 5 through 8 attended focus groups in the summer of 2010 and subsequently enrolled in Chem 1B the following fall. Finally, note that review of Worksheet 4 in the focus groups was not possible due to time constraints. It was, however, part of Intervention 2 (Worksheet Downloads) and will be discussed in that section.

Spring 2010 Chem 1B grades were supplied in letter form, and were converted to grade points for comparison (below). Fall 2010 grades were provided as a percentage and converted. Grade point comparisons of the focus group participants with the overall class grade averages are also reported in **Table 6**below.

Table 6: Comparisons of Final Chem 1B Course Grades

	Spring 2010	**Fall 2010**
Total class size	110	112
Focus group average (n=4)	2.3	2.8
Rest of class	1.7	2.6

Worksheet 1: Reaction Kinetics and Stoichiometry

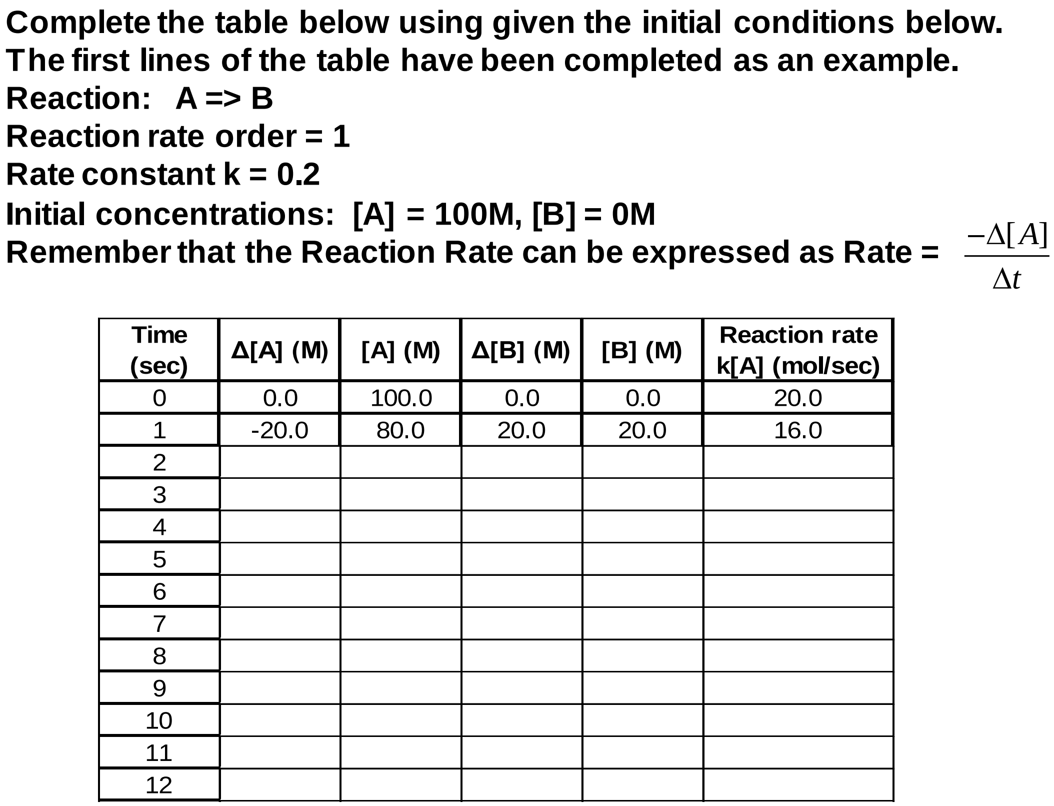
A simple stoichiometry question is asked to begin this worksheet. Given the reactionstudents are asked which reactant is consumed most quickly in this reaction

$$2AB + C_2 \rightarrow 2AC + B_2$$

(this is similar to questions asked on the Diagnostic Exam). Not one student in any of the Focus Groups answered this correctly; all students chose C_2as the answer (including S2, who earned an A in Chem 1A). At this point, a quick exercise outlined in the worksheet was performed to help the students to understand their mistake. Ten molecules of each of the two reactants were drawn, either on a piece of paper or, in the Focus Groups, on a whiteboard. Reaction products were then drawn on the board, each set of which required two AB's and one C_2 to generate. As the new products were drawn, the reactants used to make them were erased. It then quickly became apparent to the students that AB was being consumed twice as fast as C_2. After this exercise was carried out, students universally recognized that their solution was incorrect; however, as pointed out in the literature, recognition of a misconception is only the first step.

After the correction was made, some students said "Oh, well, if you mean you started with the same amount, then sure . . ." They had judged that the coefficient of 2 on AB meant that you were starting with twice as much AB, and thus it would last longer. This question was then modified to tell students to assume that they're starting with the same quantity of reactants. Additionally, another question was added before this one which asked students to articulate the difference between a coefficient and a subscript in a chemical equation to ensure that this was not the source of error. Some students could not answer this ("aren't they just the same thing?"), in which case they were asked to draw 2C and C_2 on the board. Sometimes the student was asked to "read" the chemical equation to see whether they understood what it represented, and this often could not be done ("Two moles of AB combined with two moles of C . . .").However, even after these new questions were added and understood, some students still picked C_2as the most quickly consumed reactant. The experience in this exercise illustrates a key problem for instructors; how do know when a misconception is really gone, and then, how can you tell when it has been replaced with a correct one?

The next exercise began with a paragraph which reviewed Chem 1A material on reaction kinetics, borrowing heavily from the Chem 1A lab manual. The dependence of reaction rate on reactant concentrations was emphasized, and some questions were asked about how rates might change as a reaction progressed. A table was provided for calculations of concentration vs. time given initial conditions (rate constant, concentration). Sample calculations were provided, and students engaged in dialogue with the instructor to complete the table (see **Figure 10**below). Results varied greatly across the base of students interviewed. In general no one was able to complete the table without significant help from the instructor. General confusion was expressed about the difference between the terms “reaction rate” and “reaction rate constant”. Many students inferred that the word “constant” here indicated that the rate stayed constant; no one realized that the reaction rate varied over time. After some study of the rate equations, students realized that, because the reaction rate is dependent on concentration, and concentrations change over the course of a reaction, that rates would also change. A quote fromS1: “I always thought that ‘rate constant’ meant the rate was constant . . . now I see that it (reaction rate) changes;it’s like the rate isn’t constant, but the rate of change is constant”. The literature describes the occurrence of an “aha” moment when, after the activity and initiative on the part of the learner is brought to bear on a problem, the answer suddenly occurs to them. As the groups worked through this problem, an “aha” moment was visible in almost all sessions; the time it took to get there, however, varied considerably. Students’ key learning here was that reaction rates vary with time as a function of the rate constant and the concentration of reactants.



Complete the table below using given the initial conditions below.
The first lines of the table have been completed as an example.
Reaction: A => B
Reaction rate order = 1
Rate constant k = 0.2
Initial concentrations: [A] = 100M, [B] = 0M
Remember that the Reaction Rate can be expressed as Rate = $\frac{-\Delta[A]}{\Delta t}$

Time (sec)	Δ[A] (M)	[A] (M)	Δ[B] (M)	[B] (M)	Reaction rate k[A] (mol/sec)
0	0.0	100.0	0.0	0.0	20.0
1	-20.0	80.0	20.0	20.0	16.0
2					
3					
4					
5					
6					
7					
8					
9					
10					
11					
12					

Figure 10: Kinetics Exercise from Worksheet 1

It was notable that all students showed little to no understanding of the rate laws and their relationships to the rate constant and product concentrations, despite the fact that one of the lab exercises in Chem 1A was devoted exclusively to this subject. When asked about this, students generally had the same response; that this lab was held late in the semester and was the victim of end-of-semester task overruns; that it was very complex and they did not understand it; and that they just learned how the calculations were performed in order to be able to submit their lab report with minimal effort. As a result, knowledge of reaction kinetics was basically absent in the Focus Group sample.

Reactions of the students to the worksheet session were universally positive; when asked whether the session was helpful, an enthusiastic Yes was the answer. S1indicated that the session was more “in-depth” than he had expected, an interesting comment given that it consisted exclusively of review material. My interpretation of this comment was that his use of the word “depth” referred to the session’s emphasis on the conceptual rather than the formulaic. At the end of the session, he seemed enthusiastic, thanked me for the help, and asked to take a copy of the worksheet home with him.

S2, who earned an A in Chem 1A, had the most interesting session. Of all of the students, he most avidly stuck to his belief in the wrong answer to the stoichiometry question. After he was shown the reaction ($2AB + C_2 \rightarrow 2AC + B_2$) and was told that we were starting with equimolar amounts of AB and C_2, he stuck with C_2 as the more rapidly consumed reactant “because it has a less amount of moles . . . you have two moles of A and two moles of B, so you should have two moles of C_2, so C_2 is consumed faster”. When asked to verbalize the reaction, he said “two moles of AB are combined with two moles of C to produce two moles of AC and two moles of B”. After two C’s and one C-C were written on the white board, and he was asked to articulate what each drawing represented, he correctly identified one as “two C’s” and the other as “one C-two” He now changed his reading of the equation, but still believed that C_2 was consumed faster. At this point, I went into the white board exercise described in the worksheet, drawing six AB molecules and six C_2 molecules and showedhow they were consumed to make AC and B_2. This exercise took about eight minutes in total, but S2 finally understood the mechanics of the reaction and why the relative rates of consumption of the two reactants were different.This may indicate a corollary to the assertion that established misconceptions are difficult to change; it may be that, the more academically successful the holder of the misconception, the more confident they are in its validity, and the less likely they are to change it.

Finally, some graph paper was provided to students in the sessions in order that they might plot their data from Figure 10 and graphically observe concentration changes over the period of a chemical reaction. It was found that the time necessary to get across the concepts of stoichiometry and reaction kinetics consumed the entire session, and that these plots were seldom drawn. Accordingly, the use of the Equilibrium Machine (described in the Methodology chapter) was added to this worksheet, in order that students might be able to get a graphical sense of the rate/concentration relationship without having to first manually develop a tabular form of this relationship. In early versions of the worksheet, students were asked to manually obtain a concentration plot for a reversible reaction given initial conditions, similar to the exercise in Worksheet 1. This exerciseconsumed an excessively large amount of time, and although itwas seen as useful for attacking this problem, was not practical for this implementation and was removed from the exercise.

The Equilibrium Machine can be seen in the archived version of this spreadsheet as part of Exercise 3. It was used in Focus Groups for several of the sessions, and then only as a source of discussion, but students seemed to quickly engage with it, and found its rapid response to inputs very enlightening. S8 made the comment that the juxtaposition of three simultaneous learning techniques (the worksheet, the whiteboard, and the Excel plots) was particularly helpful in conveying an understanding of the concepts being conveyed.

Worksheet 2: Reversible Reactions and Chemical Equilibrium

Worksheet 2 consisted mainly of two exercises, one on reversible reactions and one which dealt with the basics of chemical equilibrium. The Equilibrium Machine was used heavily in the chemical equilibrium section. A third exercise exploring the relationship between reaction stoichiometry and the equilibrium equation was added to the worksheet for use in the last few Focus Groups. This worksheet was revised many times to improve its clarity and effectiveness (17 revisions). Since this worksheet was used by both incoming Chem 1B students with no knowledge of chemical equilibrium and by students in first month of Chem 1B who had some exposure to the material, the worksheet needed many revisions to hit the right level of complexity to make it effective; or, in the language of the literature, to find the right ZPD level. The revision history of this worksheet can be found at the end of Worksheet 2 in Appendix B.

The first exercise, on reversible reactions, made heavy use of the kinetics review done in Worksheet 1. In a brief introduction, reversible reactions were presented as the combination of two reactions running simultaneously; a forward reaction and a reverse reaction. The exercise opens with a number of probing questions meant to lead the student to this understanding. Since all of the students in the Worksheet 2 Focus Groups had previously been exposed to Worksheet 1, this exercise was generally completed very quickly (in about 5 minutes).

Students were then asked to articulate the difference between an irreversible reaction, which goes to completion while exhausting one or both of its reactants, and a reversible reaction. Two sample reversible reactions were shown($H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$ and $2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$), and students were asked to talk about the differences between the two. When asked which components were reactants and which were products, students were usually puzzled for a while but were eventually able to articulate an explanation ("they're the same equation; there's no such thing as a reactant or product"). Describing the reaction $A \rightleftharpoons B$, S7 and S5 said that "A goes to B and B goes to A at the same time" and that "both A and B are in the final solution". S2's response here was that "if you look at the reactions, they're both the same . . . I can look at them and they both have H_2 and I_2 and HI . . . because I can look at one and say that's the formation of HI, and I can look at the other and say that's the dissociation of HI". At this point, though, he stated that he was "still confused" by the concept of equilibrium (this was in the second month of Chem 1B, well after equilibrium had been introduced). The nature of this confusion was hard to grasp; he seemed to have trouble with the idea of both forward and reverse reactions occurring simultaneously, and talked about the existence of "a forward reaction *or* a reverse reaction".S2 (B+ in Chem 4, A in Chem 1A) needed 40 minutes to complete this exercise.

The next exercise relied on the use of the Equilibrium Machine, which models the simple reaction $A \rightleftharpoons B$ based on initial conditions (rate constants and initial concentrations for both forward and reverse reactions) entered by the student. Students were first asked to input initial conditions for an irreversible reaction (reverse rate constant = 0) and asked to characterize the plots of rate vs. time and concentration vs. time generated from that data. Eventually all students were able tocharacterize this as an irreversible reaction, with the concentration of Aessentially going to zero (see **Figure 11**). A reverse rate constant was then entered, and the plots changed to depict the kinetics associated with a reversible reaction (see **Figure 12**). The rest of the exercise consisted of altering the initial conditions, asking students to predict the effect on the plots, and discussing the result.

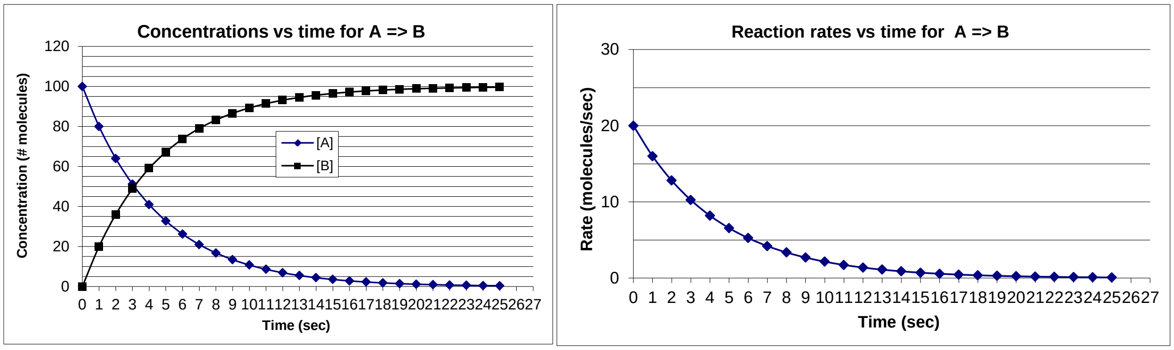


Figure 11: Equilibrium Machine Plots for A → B
($[A]_{init}$ = 100M; $k_{(A)}$ = 0.2)

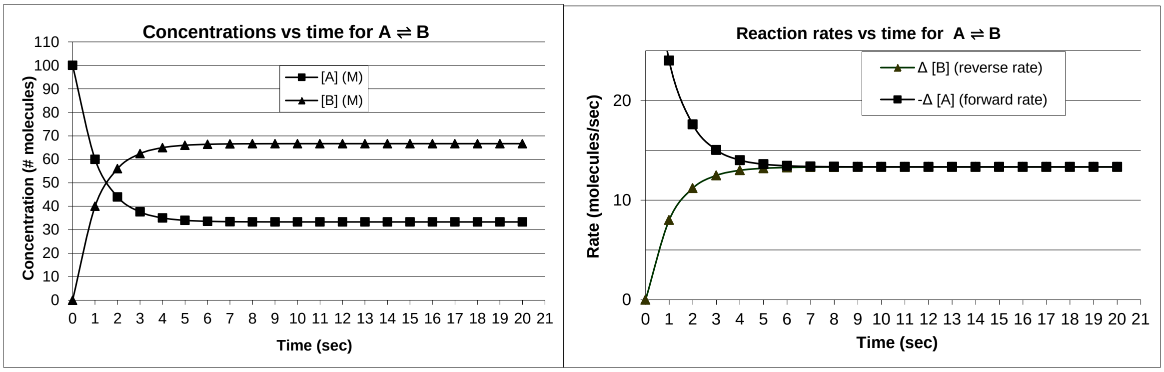


Figure 12: Equilibrium Machine Plots for A ⇌ B
($[A]_{init}$ = 100M;$k_{(A)}$ = 0.4, $k_{(B)}$ = 0.2)

Students became very interested in the shape of the plots when the reverse rate constant was entered, and spent some time working to understand them. They were then asked to articulate the essential differences between the two plots. Most were eventually able to do this (“the reactants don’t go away in the second plot”), but only after some thought, showing that 1) This was a new concept to them, and 2) They were spending some time to integrate this new information into an existing memory program. This was true of both pre-Chem 1B students and current 1B students who had been introduced to chemical equilibrium.

After being asked to describe what was happening in different regions of the plot,S1 spent a few minutes studying the plot before answering. He correctly noted that “the (product/reactant) concentrations are (increasing/decreasing here)”, and that “the concentrations stop changing here” (in the equilibrium region). He thenobserved that the forward and reverse reaction rates must become equal at some point, and indicated that point on the plot. He correctly described the relative equilibrium concentrations for the reversible vs. irreversible reactions using data read from the plots.

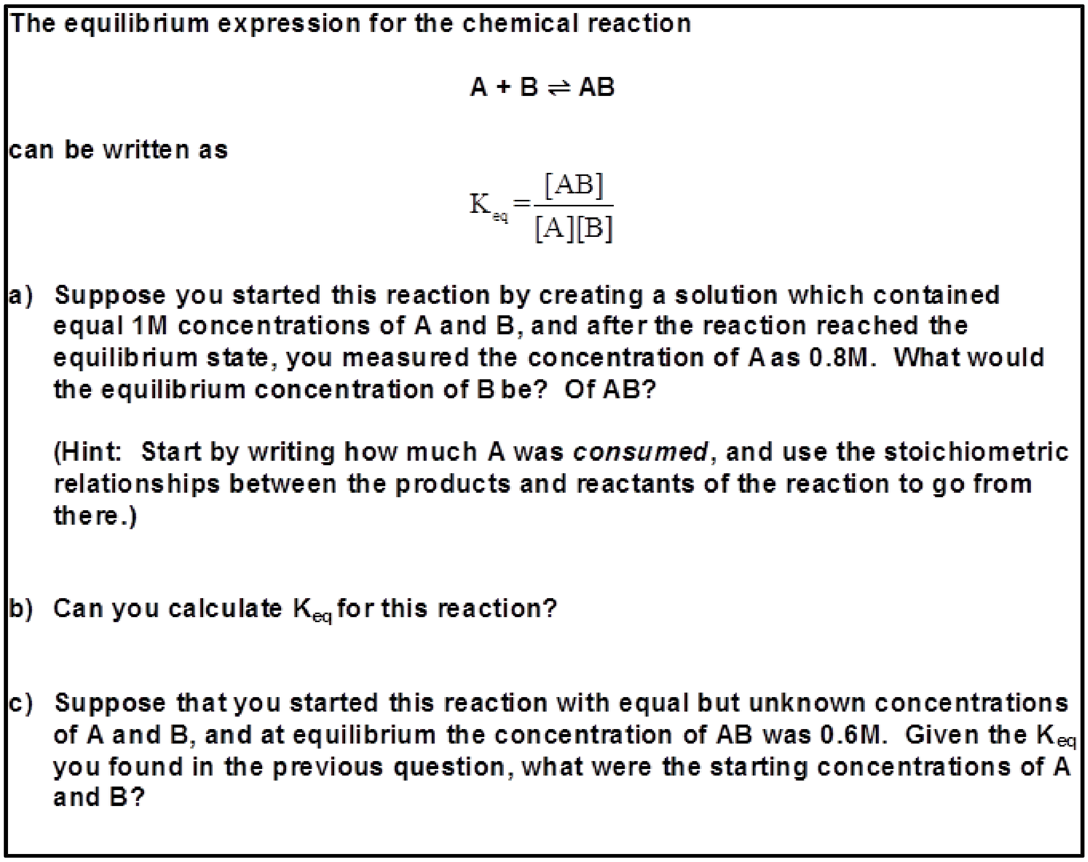
Again reading from the plot, and given the equation for K_{eq},S1 was able to calculate an equilibrium constant fromplot data. He then did a calculation of Q using two points on the plot prior to the equilibrium state, and correctly described how and why his $K_{eq}$and Q numbers were different, and what would have to happen to get from Q to K_{eq} (“need more products or less reactants”). Other students did not always get this concept so easily; when S3 (then a first-weekChem1B student) was asked where to find data for the K_{eq} calculation, he replied that “any two points will do”. After going through calculations using two sets of points (one atequilibriumand one fromthe pre-equilibrium region) and getting two different numbers, he engaged in a discussion with the instructor, after which he was able to articulate why his two results were different.

S2was able to applyhis kinetics learnings and the concept of simultaneous forward and reverse reactions to his interpretation of the plots. As the reaction progressed over time on the plot, S2 observed that, as the plot lines crossed, more B than A was being produced (i.e., the forward reaction rate was faster than the reverse reaction rate). How, he was asked, could equilibrium be reached under these conditions? After a little more work, he recalled from earlier in the exercise that reaction rate was a function of concentration, and thus as the concentrations of reactants decreased, the forward reaction rate would decrease (and vice versa for the reverse reaction). Until this moment in S2's session, it had not been quite as obvious to him why the concept of reaction kinetics was so important to the understanding of chemical equilibrium. Eventually, S2 was able to fluidly articulate the nature of the relative reaction rates of A and B and how they related to the relative changes in concentrations of A and B.

After some discussion on the form of the reversible reaction plot, students were asked to make changes to the initial concentrations of both A and B and observe the results on the plots. Students were asked to recalculate the equilibrium constant from plot data after each change, and they articulated why it was the same each time, and were quickly able to understand and articulate why the form of the plot changed as it did when initial conditions were changed. The use of the Equilibrium Machine here was instrumental in conveying this understanding; students were able to quickly switch back and forth between the “before and after” plots to see what exactly was changing.

The final exercise on Worksheet 2 involved questions leading to discussion ofthe conditions present in the equilibrium state, and how the mathematical form of the equilibrium constant related to those conditions. This exercise led to an understanding of the fact that forward and reverse rates are equal at equilibrium, and why that is so. To paraphrase from a discussion with S7 and S5, “in equilibrium, both rates are equal, so A is being produced at the same rate as B, so there’s no net change in the concentration of either”. This description of equilibrium was clearly an “aha” moment for them.

A third exercise, added late in the Focus Group cycle, involved how a change in concentration of a reactant or product drove changes in the concentrations of other reaction components. The mathematical expression for the equilibrium constant was used to convey this understanding. A problem was posed which was mathematically trivial but which required an understanding of how concentration changes depended on stoichiometric relationships between reaction components; see **Figure 13** below.



The equilibrium expression for the chemical reaction

$$A + B \rightleftharpoons AB$$

can be written as

$$K_{eq} = \frac{[AB]}{[A][B]}$$

a) Suppose you started this reaction by creating a solution which contained equal 1M concentrations of A and B, and after the reaction reached the equilibrium state, you measured the concentration of A as 0.8M. What would the equilibrium concentration of B be? Of AB?

(Hint: Start by writing how much A was *consumed*, and use the stoichiometric relationships between the products and reactants of the reaction to go from there.)

b) Can you calculate K_{eq} for this reaction?

c) Suppose that you started this reaction with equal but unknown concentrations of A and B, and at equilibrium the concentration of AB was 0.6M. Given the K_{eq} you found in the previous question, what were the starting concentrations of A and B?

Figure 13: Worksheet 2 Question on Stoichiometry and K_{eq}

In the few sessions in which it was used, this exercise proved to be challenging but feasible. This problem couldbe made more straightforward by the use of an “ICE table” (a table in which **I**nitial concentrations and **C**hanges to those concentrations are entered, and the resulting **E**quilibrium concentrations then calculated) to calculateconcentration gains and losses, but at the expense of the loss of conceptual understanding on the part of the student. The ICE method can be used as a rote methodology for problem-solving; “I’ll put a minus x on one side of the equation and a plus x on the other side”. This approach is useful in the process of problem-solving, but can be usedwithout a real understanding of the problem at hand and thus lead to confusion when problems are posed which require more than just a rote application to solve.

In summary, this worksheet was the most challenging of the four, and often required more than the allotted hour to complete. One general observation was that when data was viewed in aninteractive graphical display such as the Equilibrium Machine, kinetic behavior and the attainment of the equilibrium state were actually quite interesting to students. They expressed some surprise at the shape of the reversible reaction plots, and even more at the plot changes seen afterinitial conditions were changed. Another area of student interest was the application of principles of reaction kinetics to understand behavior observed on the plots; one could see students' pleasure at being able to relate a previously learned concept to a newly encountered phenomenon. For students who had some level of introduction to chemical equilibrium, the experience was especially interesting. Listening to students, it was clear that they had encountered some of the terms introduced, but had never before fully understood what they meant, and without that understanding, had a genuine curiosity as to their meaning.

Another general observation from the Focus Groups was that the breadth of material covered in the worksheet required the instructor to continuously revise the direction of the Focus Group sessions, steering the discussion and changing the agenda as called for by the level of interest and ability of the participants. The worksheets were written with this in mind; exercises were either eliminated or were greatly expanded depending on students' responses. This is relatively easy to do with small groups, but obviously becomes more of a challenge in a classroom session.

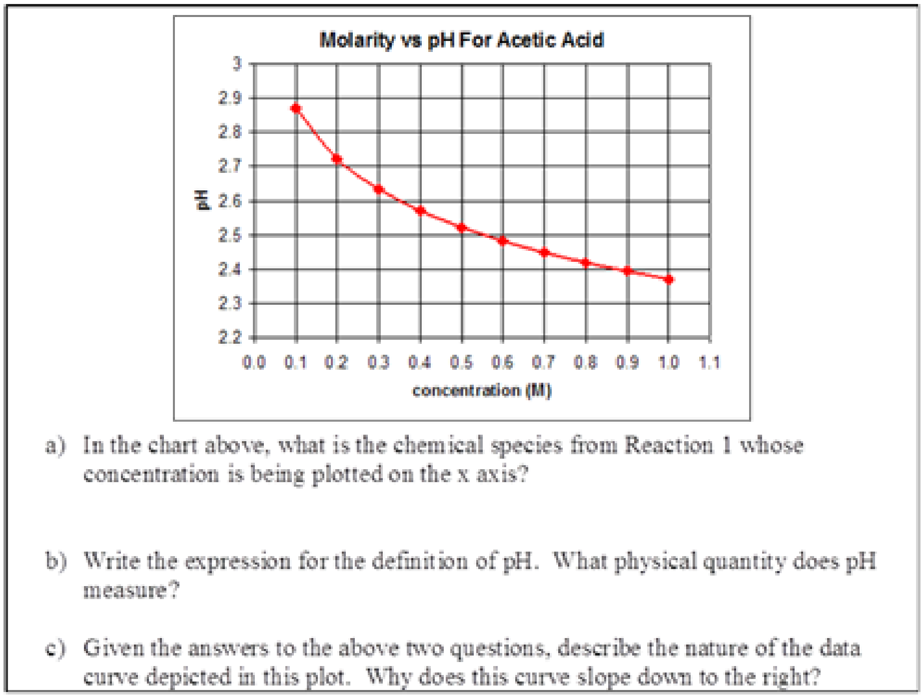
Worksheet 3: Basics of Acid/Base Equilibrium

The first exercise in Worksheet 3 was a review of the strong/weak acid/base concept learned in Chem 1A. Students were asked to articulate the difference between strong and weak acids and complete drawings showing the dissociation results of the two acid types. S1went through this exercise very quickly. He quickly recognized that the dissociation reaction for a strong acid could not be considered reversible and thus had no equilibrium state or K_{eq}, and articulated why this was so. He was able to clearly articulate the difference between strong and weak acids, and to list the six strong acids. The ease with which he completed this sectioncould be attributed to his previous Chem 1B stint, in which he would have been heavily exposed to these topics. S1 quickly drew visual representations of strong and weak acids.

AfterS1's session, the removal of this exercise was considered, as it appeared too simple for the Focus Groups. S3, however, who was then in the third week of Chem 1B, took some time to articulate the dissociation behavior of the two, and could not write net ionic equations for the dissociation reactions. He eventually worked this out, and was able to correctly complete the drawings. After viewing his struggles with this exercise,it was left in the worksheet.

In the next exercise, the terms K_a and K_b were introduced, and their definitions were related to pH and K_{eq}. Some questions were asked which required students to articulate these relationships, and another question asked students to draw representations of solutions containing acids with different values of K_a. S1 had some trouble relating $K_a$and pH to the concept of acid strength. The equilibrium and K_a expressions were then drawn on the whiteboard, and after some discussion, he was able to describe this relationship. Additional discussion then took place to ensure that he had absorbed this concept.S3's progress through this exercise was very similar.

The third exercise presented a plot of concentration vs. pH for acetic acid and asked several questions about the interpretation of plot data (see **Figure 14**). WhenS1 was first showntheplot, he was unable to articulate exactly which concentration



a) In the chart above, what is the chemical species from Reaction 1 whose concentration is being plotted on the x axis?

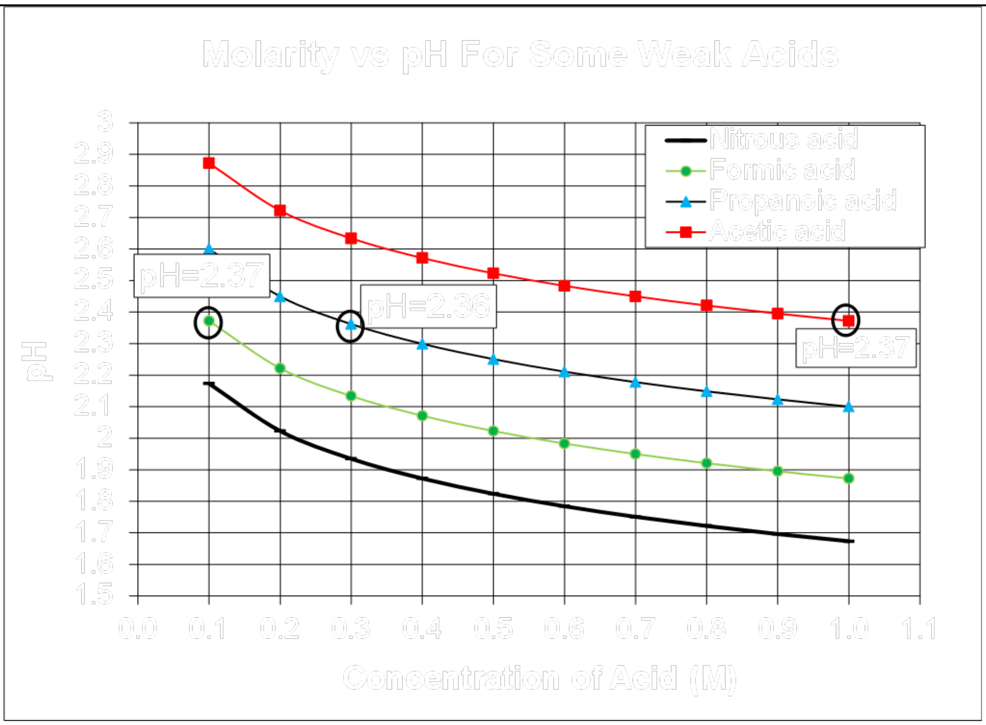
b) Write the expression for the definition of pH. What physical quantity does pH measure?

c) Given the answers to the above two questions, describe the nature of the data curve depicted in this plot. Why does this curve slope down to the right?

Figure 14: Worksheet 3 Exercise: pH and Acid Concentration

was being plotted on the X axis (labeled "concentration (M)"). He initially thought that this axis showed hydronium ion concentration until it was pointed out that this was shown on the Y axis in the form of pH. He eventually just needed to be told what it was. This exercise has since been revised in order to make the plot easier to interpret.

A plot which compared pH vs. molarity curves for several weak acids was then introduced, and several similar questions were asked (see **Figure 15**). Now that he had a



a) Which is the weakest acid on the above chart? Which is the strongest? Which will have the highest and lowest K_a? Why?

b) At a concentration of 0.5M, which acid will have the highest concentration of hydrogen ions? The lowest? Why?

c) For each acid, why does the pH decrease as the concentration increases?

d) Note the three circled data points on the plot above; they signify points at which the pH values for three different acids are approximately equal. Why are the concentrations at these three points so different?

e) Calculate the value of K_a for each of the three weak acids in the question above. Do the values you calculate make sense in terms of your answers to the above questions?

Figure 15: Worksheet 3 Exercise: pH Curves for Weak Acids

good understanding of the plot format,S1 was able to articulate the plot information. He identified the acids with the highest and lowest K_a on the multi-line plot, and talked about why pH decreases as molarity increases and vice versa. He was able to articulate why the three acids whose plots were labeled with similar pH values had such drastically different concentrations at the label point (question d).

The final question set in this exercise ("Calculate the value of K_a for each of the three weak acids in question d above") tookS1about fifteen minutes. He easily converted pH values to hydronium ion concentrations, and introduced these values into the K_a equation, but got stuck there. He needed a lot of help to see that, since $[H_3O^+]$, $[F^-]$ and [HF]were all related by the stoichiometry of the dissociation reaction ($HF + H_2O \rightleftharpoons H_3O^+ + F^-$), their values could be interconverted. S1's problem here can be traced back to his confusion over the stoichiometry questions in the Diagnostic Exam and Worksheet 1.

Once he was guided to the solution above, and got the $[F^-]$ value he needed for the K_a calculation,S1 made another common mistake. He realized that, to get the final value of [HF] for the K_a calculation, he would have to subtract the amount of product ($H_3O^+$and F^-ions) generated from the initial value of [HF]. However, when he went to perform this subtraction, hesubtracted *both* $[H_3O^+]$ *and* $[F^-]$ from $[HF]_0$. Once more, the lack of an innate understanding of the concept of stoichiometry as it applies to chemical reactions resulted in an inability to solve a more complex problem. If S1 had stopped to think what the HF dissociation reaction was telling him about the stoichiometric relationship between product generated and reactants consumed, he might have been able to do this calculation on his own. At one point during the exercise, S1 said that “I need to use an ICE table”, even though he clearly did not understand the principles behind it. When conceptual understanding failed him, S1 resorted to rote with the use of a formulaic methodology. The ICE method was used, and the problem was solved.

Both S3 andS1 each expressed positive feelings about participating in the Focus Group. S3 said that he felt hat he had really learned something; “Just taking the time to talk . . . one-on-one is helpful and makes it stick”.

Summary of Focus Groups Analysis

The outputs of the Focus Group study have a key place among the results and analyses presented in this study. The focus group sessions served as the only place in the study where the constructivist teaching of conceptual learning was fully modeled. The previous Data Gathering section identified areas of misconception in the body of Chem 1B students; the focus group intervention attempts both to confirm those misconceptions and to ask whether students can be brought to recognize and change these misconceptions.

The techniques used in the focus group intervention are very different than those commonly used in scientific study. The “data” from this study consists of observations made by the author in a series of face-to-face discussions with students on the topic of chemical equilibrium. “Results” are the comments made by students, their articulation of chemical ideas and concepts, observations of student’s reactions to the material reviewed, and student’s own assessments of their learning or lack of understanding.

The focus group sessions conducted with Worksheet 1 generally confirmed the results of the Diagnostic Exam, in that conceptual understanding of the chemical reaction (stoichiometry and kinetics) was largely missing from the focus group participants. This was true regardless of whether or not the participants had any Chem 1B exposure. Not one of the eight participants was able to correctly describe the relation between concentration change and stoichiometric coefficient; some even argued in defense of their misconception. All of the students were convinced of their error after a simple paper exercise. However, as pointed out in the literature, this is no guarantee that the misconception has been corrected. To truly correct a misconception, it must be replaced by a new, correct, conception. This can be a lengthy process, and was not necessarily accomplished in the time allotted by the focus groups. For example, in the study of conceptual change by Akkus et al[Error! Bookmark not defined.], students received three hours per week of conceptual training over a period of five weeks, using a curriculum developed by experienced members of the Science and Mathematics Education faculty at Gazi University in Turkey. In the focus group study reviewed in this paper, students received a maximum of four hours of training. The Akkus study was an example of what can be accomplished with time and dedication.

Many focus group participants attributed their misconceptions of reaction kinetics to its position in the Chem 1A curriculum. The Chem 1A kinetics lab takes place at the end of the semester, and students generally indicated that, as a result of their overall sense of anxiety during this time, they had given little mindshare to understanding this topic.

Their efforts were dedicated to the submission of the Kinetics lab with the all of the calculations properly made, but with little understanding of the topic.

Improvement in the understanding of reaction kinetics was observed after the use of a paper exercise and the interactive Excel spreadsheet (the Equilibrium Machine); this improvement was both witnessed by the author as students filled out the paper exercise and from their comments during the exercise. Students were better able to articulate the process of chemical reaction after talking about it and executing the worksheet exercises. This learning was visible in a way which is familiar to most teachers; it can be seen in students' expressions and in their tone of voice.

In Worksheet 2, the kinetics concepts learned in Worksheet 1 were built upon to develop student understanding of reversible reactions. The Equilibrium Machine proved to be very helpful in portraying the kinetics of the reversible reaction, and students showed an interest in this visual expression of the process. One interesting input from a student was that the use of multiple presentation formats during certain exercises was in itself helpful in acquiring understanding of some concepts. Having the same concept presented on the whiteboard, on paper, discussed with other students or an instructor, and visible in an interactive Excel plot, made it more accessible. Perhaps these different viewpoints call on differing pre-existent LTM structures, and as a result help to make the process of information processing and LTM acquisition more robust and successful (see Figure 2, " Schematic Representation of the Generative Learning Process").

While the use of additional worksheets added different concepts to the data mix, the processes of interaction in the focus group sessions were always the same. And although the small size of the experimental group precludes making any statistically significant conclusions, the focus group process showed the feasibility of raising students' levels of conceptual understanding in a constructivist setting, and that this task could be accomplished not only by course professors, but by graduate-school teaching assistants as well. Large-scale studies such as that of Akkus et al[9], in concert with the results shown here, combine to make the case for increasing the integration of conceptual learning into general chemistry at CSUS.

Finally, this study also demonstrated students' fundamental engagement and interest in the subject of general chemistry, something not always evident when students are struggling to complete homework or study for exams. Although not a part of the thesis claims made by this study, it is one worth noting.

Chapter 6

INTERVENTION 2 (WORKSHEET POSTING):METHODOLOGY, RESULTS AND ANALYSIS

Worksheet Posting Methodology

In the fall of 2010, four worksheets were posted on the Chem 1B website for student download and submission. Details of the contents of the worksheets have been reviewed in the Methodology chapter, and the worksheets along with answer keys can be seen in Appendix B. In the language of action research, this has been designated "Intervention 2" (See Table 4 in the "Overview of Research" section). Since Worksheet 1 *(Reaction Kinetics and Stoichiometry*) and 2(*Reversible Reactions and Chemical Equilibrium*) contained material tested in Chem 1B Exam 1, they were posted and submitted prior to Exam 1. Similarly, Worksheet 3 (*Basics of Acid/Base Equilibrium*) and 4 (*The Common Ion Effect and Buffer Chemistry*) were based on topics tested in Exam 2; they were posted and submitted prior to Exam 2.

Two types of analyses will be shown here. In the first analysis, summaries of responses to each ofthe returned worksheets will be shown here. Each of the exercises from each of the worksheets will be reviewed, and although a truly quantitative analysis is not possible given the nature of the questions asked, the percentage of correct answers to each question has been estimated, and a brief analysis of responses is included when appropriate. In the second analysis, results of course exams will be reviewed as they relate to similar exercises in the worksheets. Since the content of the first two worksheets is related to Exam 1, they will be reviewed together, and Exam 2 will be reviewed together with worksheets 3 and 4.

Worksheet Download Results & Analysis

Worksheet 1

Answers for each of the questions are summarized as "correct/incorrect"; that is, "9/6" indicates that 9 correct answers and 6 incorrect answers were received. Comments of the researcher appear in italics.

Exercise 1: Stoichiometry

For the reaction

$$2AB + C_2 \rightleftharpoons 2AC + B_2$$

a) What does the coefficient of 2 in front of the reactant AB mean?
14/1

b) What does the subscript 2 in the reactant C_2 mean?
9/6

c) Suppose that we mix equal amounts of AB and C_2 to initiate a reaction. Which reactant do you predict will be consumed more quickly (at a faster rate)? Which product is produced more quickly? Why?
8/7

This continues the trend set in the face-to-face worksheet reviews.

d) When the reaction is complete, what substances will remain in solution, and at what ratios?
3/10

This is essentially a more complex version of the previous question, so it makes sense that even those who correctly answered q. c couldn't always get this. The people who got this correctly worked through the "whiteboard exercise".

Exercise 2: Reaction Rates and Concentrations Over Time

a) As the reaction A → B starts at time = 0, what is the initial reaction rate if the starting concentration of A ([A]) is 500M and the rate constant k = 0.3? (Refer to Equation 2.) 14/1

b) After one second, what is the concentration of A? Of B?11/3

c) What is the reaction rate (k[A]) at time t= one second? (Refer to Equation 2.)8/6

This is the most complex question of this group. This is a new concept for most people here.

d) In your table, what happens to the concentration of A over time? What happens to the concentration of B over time? Does this make sense?
15/0

e) What happens to the reaction rate over time? Does this make sense?
15/0

f) What can you say about the nature of the reaction rate over the time of a chemical reaction?
15/0

The three questions above simply require a reading of the table and demonstrate an understanding of the table.

Exercise 3: More on Rates and Concentrations

a) Examine the Rate vs. Time plot (Plot 2, at the lower right). What can you say about the reaction rate vs. time? Does this hold with your findings in Exercise 2 above?
14/1

b) According to the concentration vs. time plot (Plot 1, at the upper right), at approximately what time does the reaction finish (i.e., when does [A] go to zero?).
14/1

c) Change the rate constant k to 0.5. At what time does the equation now finish? Can you explain this? (You can flip back and forth between these two curves using the Back and Forward arrows on the menu to better compare them.)
12/3

d) Change the rate back to 0.2. Now change the initial concentration of A to 50. Does this change the completion time of the reaction? Can you explain this?
9/6

Incorrect answers here were mainly a failure to correctly read the plot.

e) Change coefficient b to 2. How did the concentration plot change? Can you explain this?
6/9
Incorrect answers here were more related to the lack of understanding of the stoichiometry questions posed in Exercise 1.

f) Now change b back to 1 and set a equal to 2. What has happened to the plot? Can you explain this?
8/6
As above

Worksheet 2

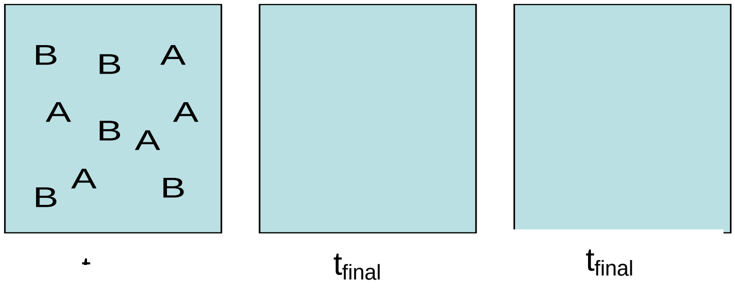
Exercise 1: Reversible Reactions

a) For the reversible reaction $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$, what are the reactants? What are the products?
13/0

b) For the reversible reaction $2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$,what are the reactants? What are the products?
13/0

c) Are the reactions in a) and b) different or are they the same? What can you say about the meaning of “reactants” and “products” of a reversible reaction?
12/0

d) At time 0, one-liter solutions of A and B with equal concentrations are mixed (represented in the left-hand box), and react to form the product AB. Draw a representation of the final components in solution after the reaction is complete if the reaction in the center box is irreversibleand the reaction in the right-hand box is reversible.



11/2

Exercise 2: Chemical Equilibrium

a) After a period of time, the plot of concentration vs. time flattens out. What is happening at the point? What compounds are left in solution at this point?
9/4

b) How does this plot differ from the plot for the irreversible reaction you just looked at?
6/7

Little mention of the fact that "reactant does not go to zero" or "both reactant and product still left in solution"

c) Note that the curves in the concentration plot begin to flatten out after a period of time. What is happening at this point? How does this region differ from that of the irreversible reaction? What compounds are left in the solution at this point?
9/5

d) Look at the Rate plot (Plot 2). What two values are plotted here? Look at the region of the plot which corresponds to the "flat" region you saw on the concentration plot. How are these two rates related in this region?
7/5

Little mention of "forward rate and reverse rate"

e) At t=18 seconds, the forward rate (-Δ[A]) is ~6.7M/s, yet the product concentration is not changing. How do you explain this?
8/4

f) Identify the region on the plots at which the reaction $A \rightleftharpoons B$ is in equilibrium.9/3

g) Can you calculate the equilibrium constant for this reaction using concentration data from the table or plot? How does this value compare to the value shown on the Excel screen (bottom left)? From which part of the plot would you take these data? Why?
9/3

h) Try taking data from different parts of the plot and using them to calculate equilibrium constants. What do you notice? How do you explain your results? 7/5

i) Change the initial concentration of B to 50. What happens to the equilibrium values of A and B (change/remain constant/increase or decrease)? Recalculate the equilibrium constant from the plot. Now change [A] to 25 and recalculate K_{eq}. What happens to the equilibrium constant? Can you explain this phenomenon?
3/7

Most replies concerned the equilibrium constant and did not mention the change in concentrations

j) Change the reverse rate constant to 0.3. What happens to the concentration plots? What is the key difference between this plot (with $k_f = 0.2$ and $k_r = 0.3$) and the previous one (with $k_f = 0.2$ and $k_r = 0.3$)? Can you relate the relative size of the rate constants to the relative concentration values in the plot?
2/8

Few mentions of how the increase in the reverse rate leads to an increase in the equilibrium concentration of A

Exercise 3: Stoichiometry and Equilibrium

a) Suppose you started this reaction by creating a solution which contained equal 1M concentrations of A and B ($[A]_0 = [B]_0 = 1M$), and that after the reaction reached the equilibrium state, you measured the concentration of Aas 0.5M. What would the equilibrium concentration of Bbe? Of AB?

(Hint: Start by writing how much A was *consumed*, and use the stoichiometric relationships between A, B and AB as defined by the equation to proceed to the solution.)
5/7

Many mistakes of the sort mentioned in the answer key.

b) Now that you know all of the equilibrium concentrations, calculate K_{eq} for the reaction.
6/6

As above

c) Suppose that you started this reaction with equal but unknown concentrations of A and B, and at equilibrium the concentration of AB was 0.6M. Given the K_{eq} you found in the previous question, what were the starting concentrations of A and B?
2/10

Incorrect formulation; almost all got a formula of

$$K = 2 = \frac{0.6}{x^2}$$

Rev 17.1 9/13/10

Worksheet 3

Exercise 1: Strong vs. Weak Acids and Bases

a) Write the reaction of the strong acid HNO_3 with water. What are the reactants and products of this reaction? Is this an irreversible or a reversible reaction? If 0.1 mol of HNO_3 is added to beaker of water, how many moles of each of the products are produced?
34/1

b) In General Chemistry, ***there are only six acids which are considered to be "strong"; all the rest are weak.*** What are these six acids?
35/0

Exercise 2: Aqueous Equilibria of Acids and Bases

a) Write the expression for K_a for the dissociation of the weak acid HF in water.
35/0

b) K_a for the weak acid HF is $7.2*10^{-6}$, and K_a for HCN is $4.0*10^{-10}$. Which will have a higher $[H_3O^+]$? Explain in terms of the above equations.30/7

c) Two solutions of the weak acid HF are found to have pH readings of 2.7 and 3.0 Which of these solutions has the highest equilibrium concentration of HA?
24/10

d) What are the concentrations of hydronium (H_3O^+) and hydroxyl (OH^-) ions in pure water?
34/1

e) What is the pH of pure water?
34/1

f) A solution is found to have a pH of 3.0. What is the $[OH^-]$ for this solution?
31/4

Exercise 3: Relationship Between Ka, pH, and Concentration

a) Write the expression for K_a of acetic acid.
35/0

b) Write the general expression for pH. What does pH measure?
31/3

c) From the plot above, what is $[H_3O^+]$ when [HA] = 0.9M (best approximation)?
25/5

d) Describe the relationship between the data on the two axes of the above plot. Why does this curve slope down to the right?
16/20

e) Write the equation for K_a for the acid in the plot (acetic acid).

Na; repeat question

f) Using data from the plot, calculate your best estimate for the value of K_a of a 0.2M solution of acetic acid. (Use a value of pH to a precision of 0.01 from the plot above.)
22/12

g) What would you expect if you were to calculate K_a at M = 0.5? Would your calculated value be lower, higher, or equal to your calculation in the previous question? Explain your answer.
17/17

Much belief that K_a could change when measured at different concentrations.

h) Which is the weakest acid on the above chart? Which is the strongest? Which will have the highest and lowest K_a? Why?
26/8

i) At a concentration of 0.5M, which acid will have the highest concentration of hydrogen ions? The lowest? Why?
28/7

j) For each acid, why does the pH decrease as the concentration increases?
14/20

Use of technical terms without good comprehension.

k) Note the three circled data points on the plot above; they signify points at which the pH values for three different acids are approximately equal. Why are the concentrations at these three points so different?
12/22

l) Calculate the value of K_a for each of the three weak acids in the question above. Do they values you calculate make sense in terms of your answers to the above questions?
15/19

Worksheet 4

Exercise 1: Buffer Chemistry: The Conjugate Pair and the Common Ion Effect

a) Write the equation for the dissociation of the weak acid HCN in water.
28/0

b) Write the expression for the K_a of HCN.
28/0

c) What is the conjugate base of HCN? Of HNO_2?
29/0

d) K_a for HCN is $6.17 * 10^{-10}$. Given this figure, what can you say about the relative amounts of the non-dissociated acid HCN and its conjugate base in the solution (approximately; i.e., roughly equal, one much larger, etc.)?
21/9

Many answers called [HCN] larger than or equal to [ions].

e) What is the reaction between KCN and water? (Hint; review your solubility tables.) What new chemical species (compounds/ions) appear in solution when KCN is added?
 8/20

Most common answer: $KCN + H_20 \rightleftharpoons KOH + HCN$
They then indicated that OH^- *was the new species in solution.*

f) Note the equilibrium expression for K_a for HCN (problem b above). How has the balance of equilibrium been changed by the addition of KCN to the solution? How will this affect the expression for K_a? What is the effect of these changes on the pH of the solution?
 15/16

With the incorrect answer to the previous question, it wasn't possible for them to answer this question either.

Exercise 2: Introduction to Buffer Solutions

a) Buffer solutions are very important in the field of medicine. Can you think of a good application for a buffer solution?
 28/0

b) Write the equation for the dissociation reaction for acetic acid in water. What is the conjugate base of acetic acid? Write the expression for the K_a of acetic acid.
 29/1

c) The K_a of acetic acid is very small ($1.8 * 10^{-5}$). What does this tell you about the amount of the conjugate base which will exist in a solution of acetic acid? What must be done to form a buffer solution from a solution of acetic acid?
 24/5

Exercise3: Buffer Chemistry

d) What is the purpose of adding sodium acetate to the solution?
 24/6

e) List all the compounds and ions which will be present in the solution at equilibrium. Which of these are relevant to the key function of the buffer? Which are not? Why?
24/2

f) Roughly, what are the relative volumes of the compounds in solution (from question k)? (Equal, greater, much greater, etc.)
20/5

g) Is this a buffer solution? Why or why not?
21/7

h) Suppose that a quantity of the strong acid HNO_3 is added to this solution. Which of the compounds in the solution (see question k above) will react with this acid? Write the equation for this reaction.
25/5

i) Given the reaction equation you have written above, can you draw any conclusions about which way the reaction will progress (e.g., towards the right, or to products, or to the left, or reactants)?
15/12

Most common wrong answer: Reaction proceeds to the left (towards the reactants).

j) Suppose that a quantity of the strong base NaOH is added to this buffer. What species in the in the solution will these ions react with? Write the equation for this reaction.
26/2

k) Given the reaction equation you have written above, can you draw any conclusions about which way the reaction will progress (e.g., towards the right, or to products, or to the left, or reactants)?
20/8

Rev 7 10/1/10

Comparison of Worksheets 1 and 2 with Midterm Exam 1

The worksheets reviewed above were completed in the Fall of 2010 immediately prior to the midterm exams given that semester; Worksheets 1 and 2 were completed just prior to Midterm 1, and Worksheets 3 and 4 were completed just prior to midterm exam 2. This scheduling was based on the fact that, in each case, the same topics covered by the worksheets were also covered on the midterms. As mentioned above, only a small portion of the class opted to complete the worksheets. Thus, the total number of students taking each exam can be broken into two groups; one group who completed the worksheets, and one group who did not. Accordingly, it made sense to compare the exam results of these two groups to see whether completion of the downloaded worksheet might possibly have affected exam scores. This comparison was performed by question; that is, specific individual midterm questions were chosen for comparison based on how well the content of those questions related to the material reviewed in the worksheets. A table of answer distributions for each group is shown for each of the analyses below, with the correct answer highlighted.

The average Exam 1 scores were 66.5% for worksheet users and 65.7% for non-worksheet users.

Exam Question 3 was concerned with the relationship between Q and K.

3. If the reaction quotient, Q, is less than K in a gas phase reaction, then
    a. The chemical system has reached equilibrium
    b. The temperature must be increased for the reaction to proceed in the forward direction
    c. The reaction will proceed in the forward direction until equilibrium is established
    d. The reaction will proceed in the reverse direction until equilibrium is established
    e. The reaction will proceed in the direction that increases the number of gas molecules

The correct answer was c.This question can be related to Exercise 2 on Worksheet 2,which introduced students to the concept of chemical equilibrium. In this exercise, students generated concentration vs. time plots for a reversible reaction using the

Equilibrium Machine, and were asked to identify the equilibrium region on the plot, which ~ 75% of them were able to do. They were then asked to use data from different regions of the plot (equilibrium and pre-equilibrium) to calculate product/reactant ratios, and to explain their results. This exercise was meant to instill an understanding of the difference between the reaction quotient Q and the equilibrium constant K_{eq}. A summary of the distribution of answers to this question appears below.

Question 3 Answers	**Worksheet Users (n = 19)**	**Rest of Class (n = 101)**
a	0	0
b	0	7
c	16	74
d	3	18
e	0	2
Percent correct	**84%**	**73%**

The most popular wrong answer on the part of worksheet non-users was the exact opposite of the correct answer. It appears that they know that Q and K are different, but now exactly how to apply this difference to a theoretical reaction. It should be noted that, given the way that the question was presented, there is a possibility that many of the answers were the result of rote learning of this topic, and not real understanding. The knowledge that “the reaction will proceed in the forward direction until equilibrium is established” can easily be learned and applied as a rule rather than as real meaning. If the two rules are committed to memory, it stands to reason that they may have gotten turned around in the students mind, resulting in the “exactly wrong” answer.

Question 5 had to do with the relationship between stoichiometry and the calculation of the equilibrium constant. e was the correct answer to this question.

5. At a high temperature, equal concentrations of 0.200 mole/L of $H_{2(g)}$ and $I_{2(g)}$ are initially present in a flask. The H_2 and I_2 react according to the balanced equation below.

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

The concentration of H_2at equilibrium is 0.046 mol/L. What is the equilibrium constant K_c for this reaction?

a.) 3.3 b) 11 c.) 15 d.) 22 e.) 45

This question can be related to Exercise 3 in Worksheet 3, which gave students certain initial and final concentrations, and asked students to generate the missing component concentrations to calculate the equilibrium constant. A distribution of answers to Question 5 appears below.

Question 5 Answers	**Worksheet Users (n = 19)**	**Rest of Class (n = 101)**
a	2	13
b	7	44
c	3	7
d	0	1
e	7	36
Percent correct	**37%**	**36%**

The most common mistake was to fail to recognize the stoichiometric ratio of 2X between reactants and products, and not to double the final concentration of HI after calculation of the amount of H_2 used. This was a common mistake seen in both the Focus Groups and on the returned worksheets.

Exam Question 6 concerned the properties of strong vs. weak acids (see below).

6. Which of the following compounds is a weak acid?

a.) HCl b) CH_3CO_2H c.) HNO_3 d.) $HClO_4$ e.) H_2SO_4

Answer b is the correct answer to this question. A comparison of responses from the two groups can be seen following.

Question 6 Answers	**Worksheet Users (n = 39)**	**Rest of Class (n = 101)**
a	0	0
b	19	98
c	0	2
d	0	1
e	0	0
Percent correct	**100%**	**97%**

Analysis of this question is included here even though this topic was not covered in Worksheets 1 or 2. It was not covered until Worksheet 3, which was not posted until after Exam 1. As can be seen from the table above, at this point in their education all students have mastered the strong acid/weak acid concept. Very few students chose the wrong answer here in either group, showing no significant difference between the two.

Comparison of Worksheets 3 and 4with Exam 2

The average Exam 2 scores were 67.8 % for worksheet users and 61.4% for non-worksheet users.

As in the previous section, several of the questions from Exam 2 were chosen for analysis based on whether or not their content was covered in the worksheets. Two groups of answers from each question will be individually analyzed below.

Question 11 asks the student to calculate the equilibrium constant (specifically, a solubility product) for a dissociation reaction given an equilibrium concentration.

11. Excess $PbBr_{2(s)}$ is placed in water at 25°C. At equilibrium, the solution contains 0.0122 M $Pb^{2+}_{(aq)}$. What is the equilibrium constant for the reaction below?

$$PbBr_{2(s)} \rightleftharpoons Pb_2{}^{+}_{(aq)} + 2Br^{-}_{(aq)}$$

a.) $4.3 * 10^{-7}$ b) $1.7 * 10^{-6}$ c.) $6.0 * 10^{-6}$ d.) $1.4 * 10^{-4}$ e.) $2.9 * 10^{-4}$

The correct answer here is c. As in question 5 from Exam 1, stoichiometric ratio of 2X must be applied when formulating the K equation, and one of the terms in the equation must be squared when calculating K. This question can be related to the studies of stoichiometry and equilibria in Exercises 2 and 3 in Worksheet 2. The most common incorrect answer was b; this answer was obtained by failing to apply a factor of 2X in the stoichiometry. (This same result was seen in the Exam 1 analysis.) Results for the two groups were nearly the same here.

Question 11 Answers	**Worksheet Users (n = 38)**	**Rest of Class (n = 76)**
a	1	7
b	11	23
c	22	40
d	2	3
e	1	1
Percent correct	**58%**	**53%**

In Question 12, a similar problem is posed, but without requiring any calculation; the question is asked, and the answers are offered, in a formulaic manner. This question also relates to sections 2 and 3 in Worksheet 2. The correct answer to thisproblem is c. A comparison of answer distributions appears below.

12. The K_{eq} of $PbCl_2$ is 1.7 * 10^{-5} at 25°C. If “x” represents the concentration of Pb^{2+}ion at equilibrium, which relationship is correct for solving for the value of x?

a.) $K_{sp} = x^3$ b) $K_{sp} = 2x^3$ c.) $K_{sp} = 4x^3$ d.) $4K_{sp} = x^3$

Question 12 Answers	**Worksheet Users (n = 38)**	**Rest of Class (n = 76)**
a	0	5
b	13	19
c	16	26
d	2	8
e	7	18
Percent correct	**42%**	**34%**

Although this problem was essentially identical to the previous one, the percentage of correct answers was slightly higher for the worksheet group. This might be accounted by the frequent use of questions posed in a formulaic way in the worksheets; by this time, all of the members of the Worksheet group have completed all four worksheets, and should be well used to this format. Again, the most common wrong answers are caused by a misunderstanding of the need to apply a stoichiometric factor to the concentration data.

Question 15 is easily answered by the application of the Le Chatelier principle. It can also be answered by using the concept of equilibrium and the form of the equilibrium expression. Understanding of these principles allows one to see how changes in concentrations of some components force others to change in order keep the expression constant. This was reviewed in Exercise 2 of Worksheet 2. A chart comparing distributions of answers from the two groups of students appears below.

15. Consider the reaction below:

$O_{2(g)} + 3F_{2(g)} \rightleftharpoons 2OF_{3(aq)}$

How does the equilibrium composition change when extra O_2 is added?

a.) Shifts to reactant side
b) Shifts to product side
c.) No change

Without knowing which method was used to provide the answers above, and no calculation to guide us, it is not clear how to construct analysis of this problem. It is curious, though, that almost half the class chose the “exactly wrong” answer a. One

Question 15 Answers	**Worksheet Users (n = 38)**	**Rest of Class (n = 76)**
a	13	34
b	21	28
c	4	13
Percent correct	**55%**	**37%**

might assume that the use of gas-phase components in the reaction equation indicated to some that the equilibrium mechanics of this reaction were somehow different than normal.

Question 20 is a straight calculation of K_{sp} from solubility data, and refers again to Worksheet 2. A comparison of answer distributions appears below; the correct answer is b.

20. The solubility of $PbBr_2$ in water is 4.33 g in 1.0 L of water at 25°C. What is the value of K_{sp} for $PbBr_2$?
a.) $1.6 * 10^{-6}$ b) $6.6 * 10^{-6}$ c.) $1.3 * 10^{-5}$ d.) $1.4* 10^{-4}$ e.) 19

Question 20 Answers	**Worksheet Users (n = 38)**	**Rest of Class (n = 76)**
a	10	17
b	12	17
c	0	7
d	11	22
e	5	13
Percent correct	**32%**	**22%**

The worksheet users' percentage score is somewhat better than the non-users' scores, but both are unusuallypoor for such a straightforward question. This is just a sequence of calculations leading to K_{sp}, starting with a mole conversion to get $[PbBr_2]$, which is equal to $[Pb^{2+}]$, then doubling that to get $[Br^-]$, and finally calculating Ksp, making sure to square$[Br^-]$. There are a number of places to stumble here. If you forget to double $[Pb^{2+}]$ to get $[Br^-]$, you end up with answer a, as did a total of roughly 20% of the class. If you forget to double $[Pb^{2+}]$*and* neglect to square $[Br^-]$ then you get d, an even more popular answer at ~ 30% of responses; this last error was made by about half the class.

Question 21 is the inverse of the previous question, but adds the introduction of the common ion as a complication; this is reviewed in section 1 of Worksheet 4. Students are given a K_{sp} and asked to calculate molar solubility. Answer a is the correct response to this problem.

> 21. The K_{sp} of AgBr is $5.4 * 10^{-13}$ at 25°C. Calculate the molar solubility of AgBr in 0.050 M $AgNO_{3(aq)}$ at 25°C.
>
> a. $1.1 * 10^{-11}$ mol/L
> b. $2.2 * 10^{-11}$ mol/L
> c. $5.2 * 10^{-8}$mol/L
> d. $7.3 * 10^{-7}$ mol/L
> e. 0.050 mol/L

Question 21 Answers	Worksheet Users (n = 38)	Rest of Class (n = 76)
a	20	44
b	8	11
c	3	6
d	5	7
e	1	6
Percent correct	**53%**	**58%**

Answer distributions appear above; both groups show similar answer distributions. Although appearing to be another computational problem, this is mainly a conceptual one. Once thestudent sees that the amount of Ag^+generated by the addition of AgBr will be quite small compared to that contributed by the 0.05 M $AgNO_3$, and thus drops x in the (x + .05) term for $[Pb^{2+}]$in the K_{sp} equation, this becomes a division problem. If you did not have the intuition to grasp this, or did not practice this type of problem beforehand, you would just have to fumble your way through. In the most popular incorrect answer, b, the assumption must have been made that, since AgBr dissociates into two ions, the solution should be doubled. Obtaining answer d requires that, instead of dropping the x in the $[Pb^{2+}]$term, one should drop the 0.05 term, in which case the answer is just the square root of x.

Question 22 poses an acid/base equilibrium problem. This was reviewed in sections 2 and 3 of Worksheet 3. This is another query about solubility, which is just equilibrium of solution; the correct answer is b. Students must see that HCl does not release any useful ions, and that calcium hydroxide is insoluble in water, leading them to must choose answer b.

22. A flask contains the following chemical system at equilibrium.

$$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)}$$

Addition of which of the following substances will decrease the solubility of $Ca(OH)_{2(s)}$ in water?

1. Aqueous hydrochloric acid
2. Aqueous sodium hydroxide
3. Solid calcium hydroxide

a.) 1 only b) 2 only c.) 3 only d) 1 and 3 e) 1, 2 and 3

Question 22 Answers	**Worksheet Users (n = 38)**	**Rest of Class (n = 76)**
a	10	16
b	17	35
c	4	13
d	4	9
e	2	3
Percent correct	**45%**	**46%**

The large number of responses for answer a (hydrochloric acid) is difficult to explain given the almost perfect response to the earlier questions on strong vs. weak acids.

Question 23 is another calculation of a solubility process, but with a slight twist; a pH value is requested as the answer. These concepts were all covered Worksheet 2 and 3. The relationship of pH and K_{sp} was not explicitly reviewed in the worksheets, but the meaning of these two terms was well covered, and it is not unreasonable to assume that worksheet students would make that connection in the course of their work.

23. At what pH will an aqueous solution of 0.022 M Cu^{2+} begin to precipitate as $Cu(OH)_2$ at 25°C? The Ksp of $Cu(OH)_2$ is $2.2 * 10^{-20}$ at 25°C.

a.) 5.00 b) 6.45 c.) 7.55 d.) 9.00 e.) 18.00

Question 23 Answers	Worksheet Users (n = 38)	Rest of Class (n = 76)
a	11	20
b	8	30
c	6	16
d	7	5
e	6	4
Percent correct	**29%**	**26%**

This problem requires both good problem-solving skills and a good understanding of the stoichiometric relationships between molecules and their ions as they dissociate in solution, and is similar in nature to Question 21 above. They are both essentially common ion problems, and also involve dissolution of multi-ion salts, so there is a necessity to double and square the right concentrations at the right time. However, while over 50% of the class answered Question 21 correctly, less than 30% correctly answered the problem above. In this problem, two final steps are required. Once the K_{sp} equation is used to find $[Cu^{2+}]$, it must be doubled to get the correct $[OH^-]$, and then pOH must be converted to pH. It seems that these extra steps were responsible for enough extra confusion that the class exceeded its capacity for problem-solving; the broad distribution of answers here indicates a good deal of confusion on the part of the class.

Summary of Worksheet Download Analysis

Two different analyses of the worksheet download intervention were performed. An analysis of the answers to the completed worksheets was performed, and the midterm exam performance of students who submitted the worksheets was compared. Note that this last analysis comes with an important caveat; it attempts to establish a correlation between two modes of learning clearly shown by the literature to be non-correlated[2]. Students' Chem 1B classroom learning cannot be expected to be related to worksheet performance, nor can their performance on Chem 1B midterm exams be a function of their answers on the worksheet exercise. What, then, is the purpose of this intervention, and what can be gained from its use? And what are the lessons from this intervention relative to the thesis of this study? Some observations will be made below, but with no claims of correlation made. Use of the Worksheet Download intervention is an incomplete exercise as conducted in this study. Accordingly, limited claims will be made from its results.

Summary of Submitted Worksheets

Plots of the results of the four downloaded worksheets appear below. These will be the basis of this summary, which will review the major points brought out in this analysis.

It is clear that the issue of stoichiometryaddressed in both the Diagnostic Exam and the Focus Group sessions, remains problematic for Chem 1B students (see the **Worksheet 1** summary in **Figure 16)**. The exact same problem used in the first two studies was used again in Worksheet 1, and roughly half of the students were not able to correctly answer it. The kinetics exercise in Worksheet 1, on the other hand, showed generally better results than seen on the Diagnostic Exam. This may be attributable to theuse of the worksheet format, which guides the student through a series of exploratory exercises before asking them to make an observation or draw a solution. Finally, note

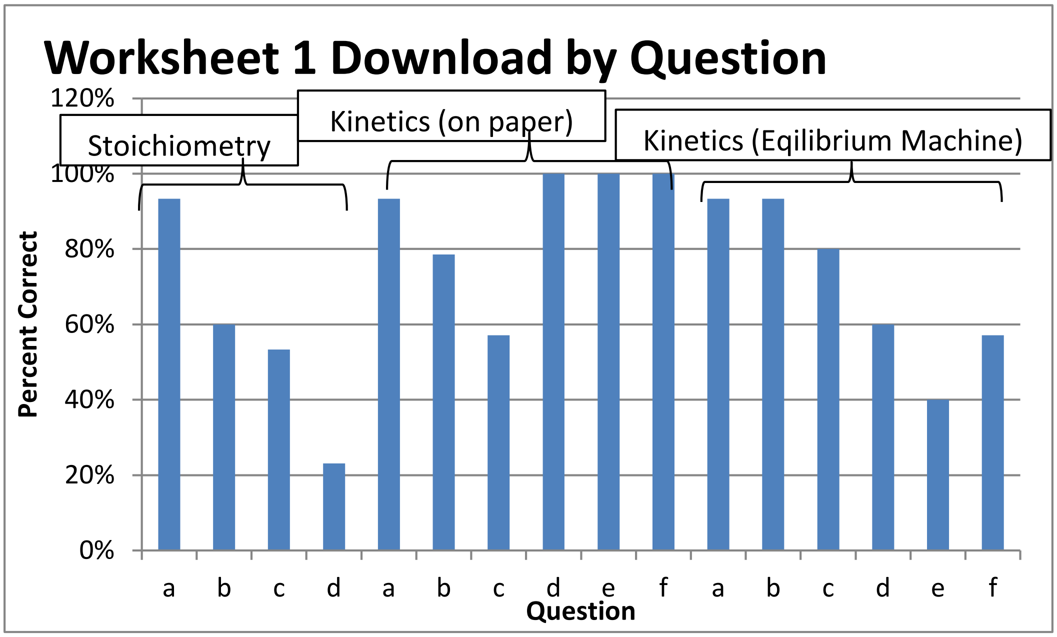


Figure 16: Worksheet1 Download Summary

that student performance with the Equilibrium Machine was not as good as with the paper exercise. This is explained by the lack of an instructor, without whom the use of the Excel database may have been more difficult (several comments to this effect were received on the submitted worksheets).

Students’ understanding of reversible reactions is well demonstrated in **Worksheet 2 (Figure 17)**; this could be accounted for by their exposure to this material in Chem 1B, or to the worksheet methodology described above (remember that all students who completed Worksheet 2 have also submitted Worksheet 1, which focuses on kinetics). The chemical equilibrium replies are mixed, with success rates of between 45% and 75%. These questions relied mainly on kinetics information previously reviewed as well as the simple definition of equilibrium as learned in Chem 1B. Many of the written responses showed a good understanding of the equilibrium state, something not seen on the Diagnostic Exam. However, when stoichiometry was added to the final few questions, student rate of understanding decreased. Issues with the concept of stoichiometry seem to remain even in Chem 1B. As per the literature, this concept is one of the most difficult to master in general chemistry[62].

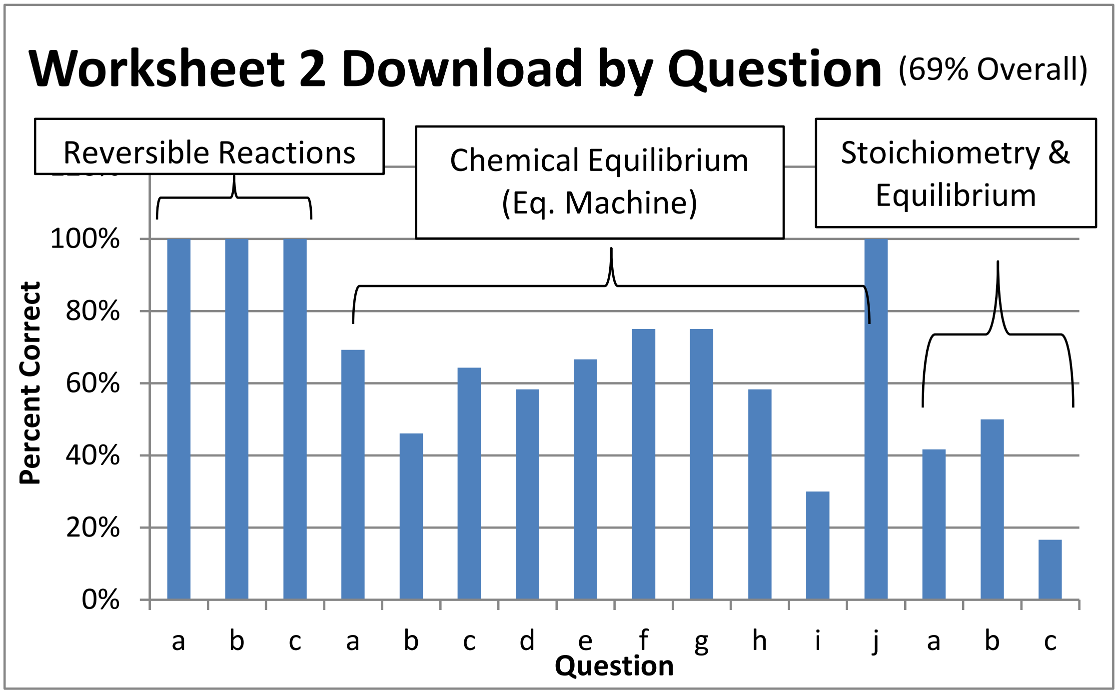


Figure 17: Worksheet 2 Download Summary

Answers to questions about acid/base chemistry generally showed good student responses on **Worksheet 3 (Figure 18)**. These questions made use of kinetics and equilibria material covered on Worksheets 1 and 2, andto which students had been exposed in both Chem 1A and Chem 1B. Questions on the relationship between K_a, pH and concentration had more mixed responses. Assuming that the meaning of these terms was by now well known, this could have been due to the use of an atypical presentation of these terms on a graphical plot in the worksheet (see **Figure 19**). This same plot was used in the focus group sessions with good success; however, with no instructor present, it may have caused some confusion on the part of the students downloading the worksheet.

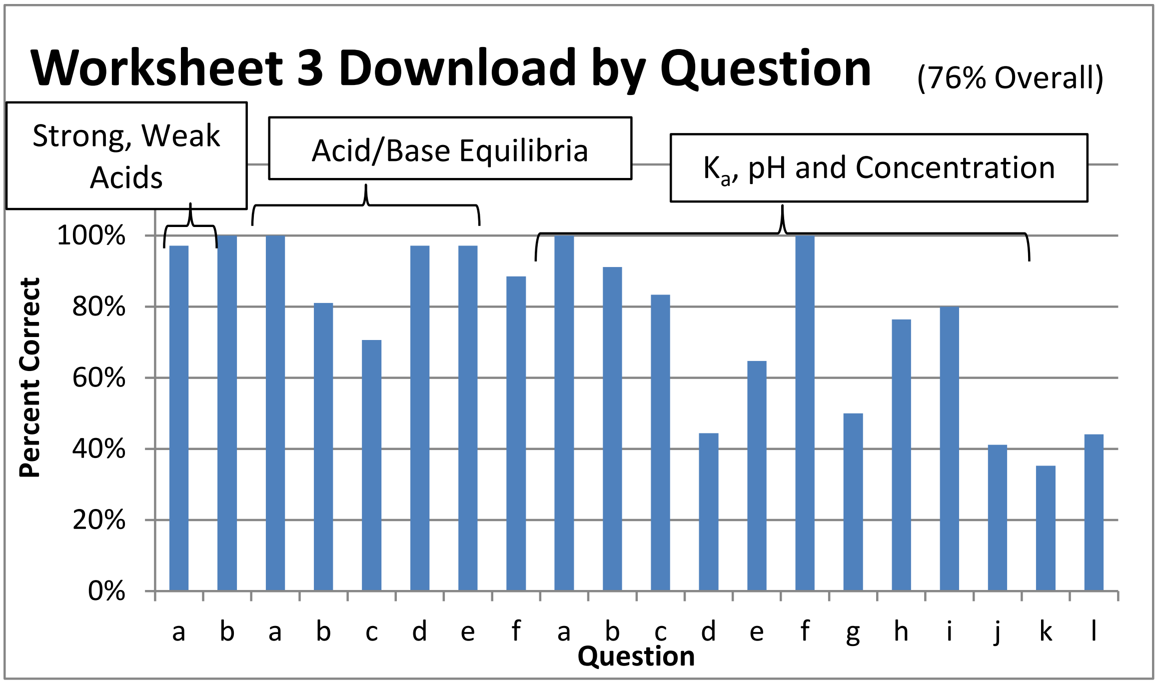


Figure 18: Worksheet 3 Download Summary

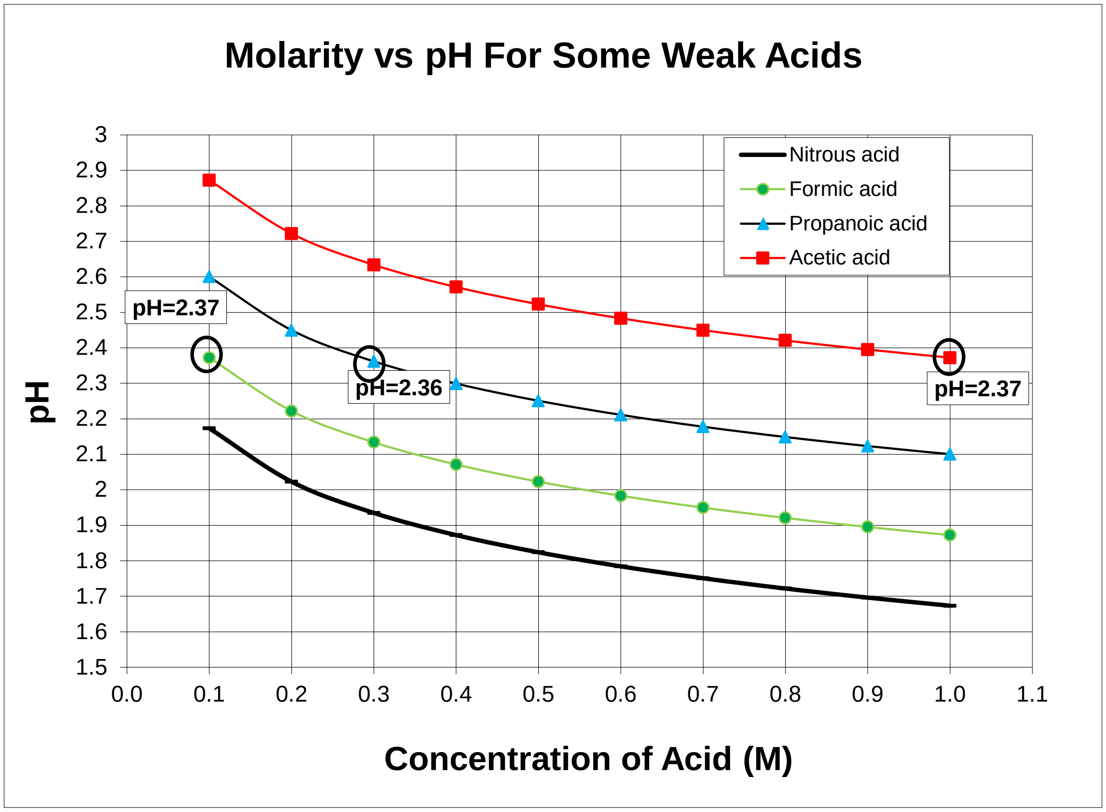


Figure 19: Figure from Exercise 3 of Worksheet 3

Students generally showed good understanding of what a buffer solution is and how it was made (seethe **Worksheet 4**summary in**Figure 20**). This was a new topic for the worksheet users, and this worksheet was not used in the focus group sessions, so it is difficult to assess student performance relative to any other data. Exercises mirrored previous questions about acid/base neutralization, and questions were asked about how both basic and acidic functionality could be achieved in a single solution. This section should be strengthened with more challenging questions for further use.

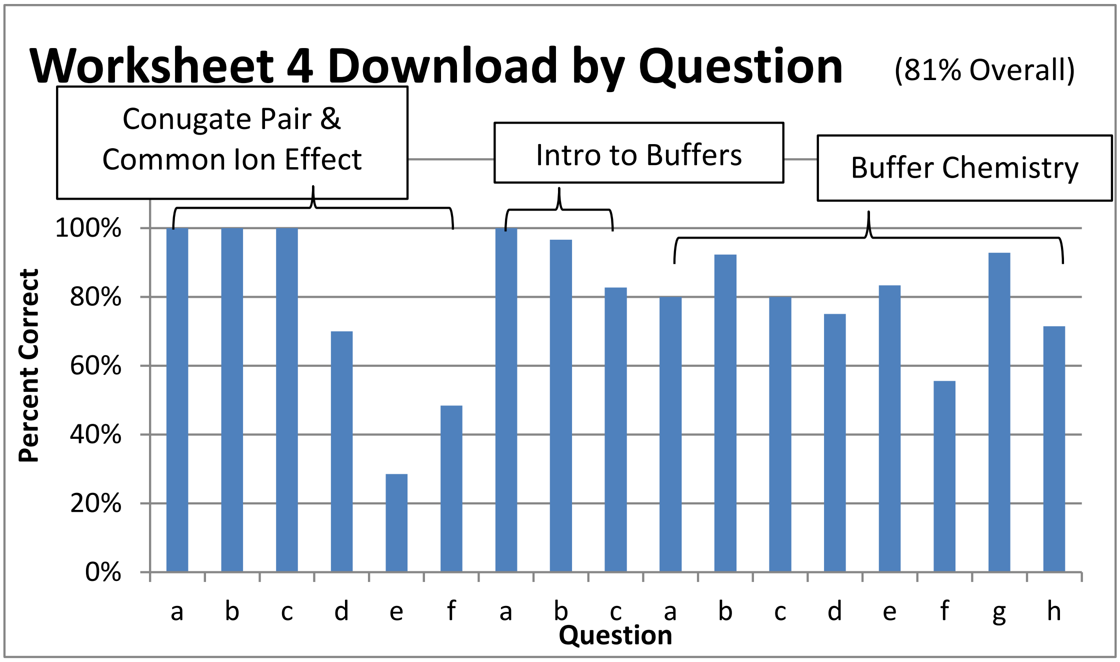


Figure 20: Worksheet 4 Download Summary

Summary of Midterm Exam Analyses

This summary carries the same caveat as the previous one. By breaking students' exam results into two groups based on whether or not they had completed the worksheets, a comparison is made between two forms of understanding which are not well correlated[2]. This type of comparison is not likely to show any statistical significance, but it is nevertheless of interest to the action researcher. When viewed together with other research, and with data from the literature, such information may support taking an action or further research towards the goal of resolving alocal problem. This is the spirit in which this summary is made.

The individual question analyses shown abovecan be combined across both midterm exams by converting the data into percentage form; see **Figure 21** below.

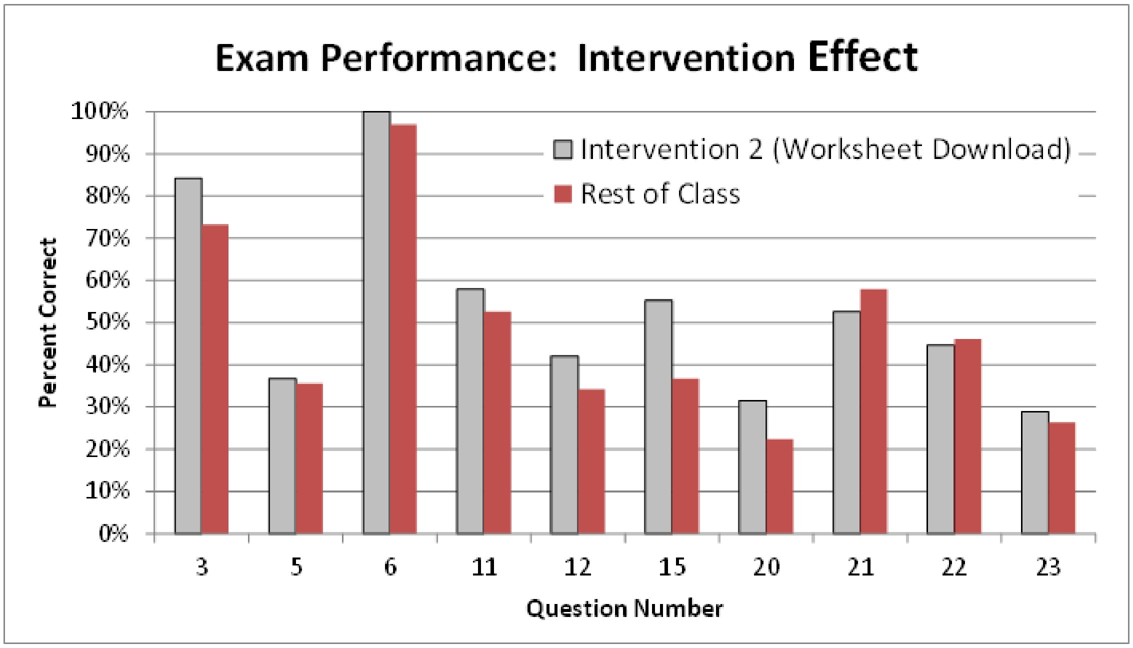


Figure 21: Workload Download vs. Midterm Exam Performance

What is to be made of this analysis? Certainly no statistical significance can be assigned to it. The worksheet group scored measurably better on 8 of the 10 questions deemed relevant for analysis. Perhaps more meaningful might be the statement that the worksheet group at least scored no worse than the rest of the class.

Another variant of this plot can be created by adding a third group to the analysis; those students who had received BOTH Intervention 1 (the Focus Groups) and Intervention 2 (Worksheet download and completion) in the same semester (these were students 5 through 8 in Table 5 in the “Focus Groups: Results and Analysis” section).

This is a very smalldata set (3 to 4 data points depending on the question. This plot is shown in **Figure 22**.

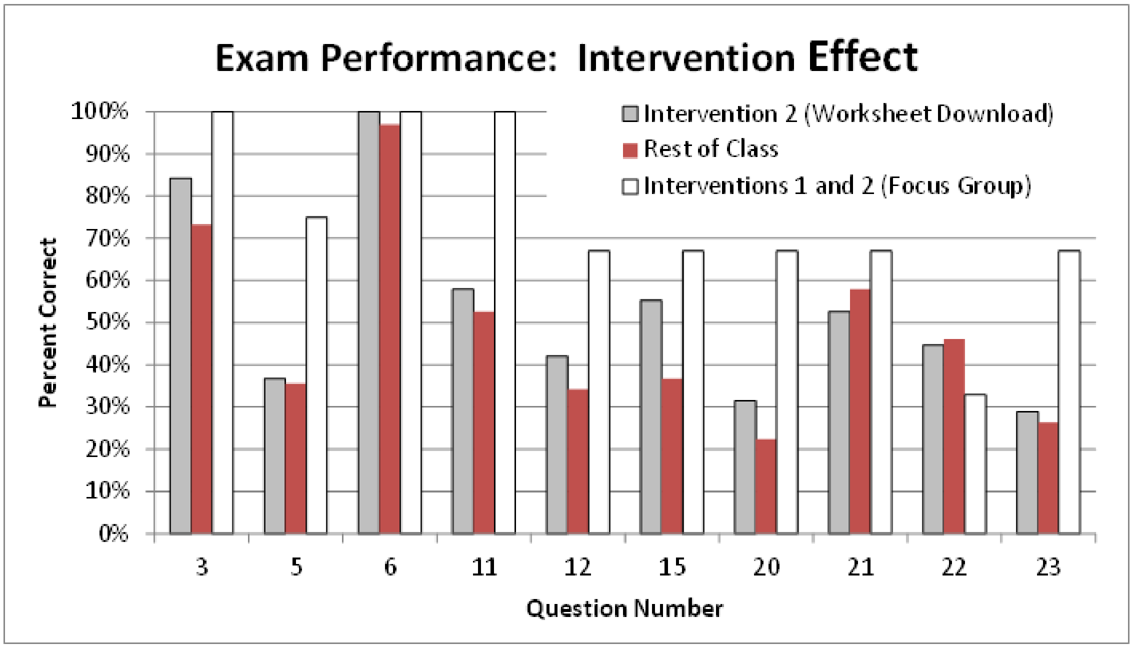


Figure 22: Interventions 1 and 2

Again, with no claims of statistical significance, it appears that students receiving conceptual interventions appear to be doing no more poorly than those who do not. In this view, we can see that those students who received both interventions did very well on the relevant exam questions; collectively, these students chose correct answers at an overall rate of 76%, a figure which, as can be seen from the plot, is much higher than that either of the other two groups.

Chapter 7

CONCLUSIONS AND RECOMMENDATIONS

Conclusions

According to the first assertion of this thesis, “many students beginning the study of chemical equilibrium at CSUS lack a conceptual understanding of its underlying principles”. Based on the data presented in the Data Gathering and Intervention sections above, this can now be stated as a conclusion of this study. The results of the Diagnostic Exam showed that a large percentage of students were unable to answer basic conceptual questions about stoichiometry and reaction kinetics, two topics which are key to the understanding of chemical equilibrium, and to which students had been exposed in Chem 1A and Chem 4. This lack of understanding was reaffirmed in the Focus Group study, where instructor interaction with entering Chem 1B students showed similar issues. And finally, downloaded worksheets completed and submitted in the Spring of 2010 (Intervention 2) by students who had completed roughly half of Chem 1B, continued to show evidence of this lack of understanding. The subtext of this conclusion is that these misunderstood concepts are the building blocks of the LTM constructs of chemical equilibrium. Without those building blocks, students will have problemswith the construction of chemical equilibrium knowledge.

The second part of this thesis asserts that “conceptual understanding of this complex and important subject (chemical equilibrium) can be improved by the integration of constructivist, conceptual teaching into the CSUS general chemistry curriculum”. This claim is based on two findings of this study. First, the focus groups in Intervention 1 allowed the author to observe how students’ conceptual understanding of chemical equilibrium concepts could be improved using constructivist, instructor-guided inquiry sessions. Students’ direct comments to the author, and their articulation of the understanding they gained from this learning process formed the bulk of the data supporting this claim. The small sample size of the focus group, and the inability to control variables with the use of control and experimental groups, prevented any statistical claims from being made.

This weakness of the focus group intervention can be reinforced by data from large-scale studies of conceptual education, such as the Akkus paper[9](reviewed in detail in the “Misconceptions” section). These studies have demonstrated the effectiveness of conceptual training in raising students’ understanding of chemistry concepts to a degree of statistical significance (see also Ozmen[1], Quilez-Pardo and Solaz-Portoles[10],Atasoy et al[63], and Canpolat et al[64]. Accordingly, in combination with the small-scale, local data obtained from the focus groups, a conclusion can be made to support the thesis that the integration of constructivist teaching of chemical equilibrium concepts into the CSUS general chemistry curriculum can improve the conceptual understanding of CSUS general chemistry students.

These two assertions (that 1) CSUS general chemistry students are lacking in conceptual understanding underlying principles of chemical equilibrium, and 2) This could be improved using constructivist, conceptual teaching in the CSUS curriculum)form the core of the conclusions to this study. A number of additional observations made during this study are presented below.

1. Input from the literature with regard to constructivist practices in the classroom was confirmed by the author/instructor.

When choosing an action step, or intervention, to address the problem in this study, the literature was relied upon to supply a tested methodology. The process chosen by this study was the use of a small instructor-guided inquiry session. This not only resulted in close observation and recording of the process by the instructor, it allowed the instructor to maintain close touch with students' level of understanding, and to constantly reset the ZPD for the session. Additionally, it resulted in real-time instructor assessment of the success of the inquiry session, leading to close monitoring of the sessions' outputs as well as providing continuous improvement in the overall process.

Also particularly effective in the Focus Group settings was the use of multiple interactive learning techniques to convey the topic at hand. This was mentioned specifically by focus group members. The use of the Equilibrium Machine, a piece of interactive graphical software, was seen as an especially effective tool by focus group members. It instantly converted data inputby the user into a graphical output, and quickly reconverted that output based on student interaction. Graphical data is both acquired and transferred more quickly than written or verbal data, making its acquisition and permanent storage more likely than any other data type. This is especially important with a topic like chemical equilibrium, the teaching of which does not easily lend itself to written analogy or everyday comparisons. In addition, face-to-face sessions like the focus groups allow several types of communication to take place simultaneously. Problems can be expressed in verbal and written forms; they can be drawn in numerous ways on a whiteboard or a piece of paper; they can be expressed in mathematical ways in spreadsheet form; data can be plotted against various environmental variables. The combination of these data types makes it more likely that one of these information sets will be recognized and accepted by the learner’s long-term memory. Acknowledgement of the positive effect of this multi-dimensional instruction was heard without prompting from the focus group participants.

While a quantitative assessment of conceptual improvement on the part of the Focus Group attendees was not made, they were asked questions about their level of satisfaction with the process, and whether or not they considered it to be worth their time and effort. Their interest and enthusiasm for the material was visible not just in their words, but in their general demeanor. This was also supported by the fact that, after attending a single session, students would voluntarily return for additional one-hour sessions; they obviously considered the meetings to be of some value.

Although undeniably effective as an intervention tool, use of the guided-inquiry method in a large class setting has its problems. The difficulty of administering a ZPD to a large body of students has been noted above, as have the methods that instructors use to do this. These limitations are recognized here; this is a topic for additional study.

As opposed to the collection of surveys typically used by action research as a method of judging the effectiveness of interventions, the focus group process made it possible to obtain direct, face-to-face input. In the focus groups, verbal and body-language cues could be read to enhance communication of student reactions to the material, which questions and discussions could then help to further clarify. As a result, it was possible to observe the ease or difficulty of particular questions, the level of meaningful understanding on the part of the student, and the degree of their engagementwith the subject matter, and whether or not their understanding had truly improved. This made it possible, for example, to determine whether a problem was solved based on an understanding of general principles or from an application of rote learning.

2. Time spent on acquiring conceptual knowledge does not visibly affect students' performance in class

In Interventions 1 and 2, it was possible to compare the results of conceptual interventions (the focus groups and the worksheet downloads) with the results of professor-written Chem1B midterm. In these interventions,the performance of focus group participants and worksheet download users was generally as good as or slightly better than that of non-participants. Although the small sample size of the interventions makes it impossible to claim any statistically significant improvement from the use of the Interventions, it can also be concluded that there is no significant negative impact on the Chem 1B course exam performance of Intervention participants despite the many hours of time they devoted to focus group attendance and worksheet download completion.

The lack of statistical significance in the data from this study is one of the reasons why action research was developed. Educational researchers need a tool that drives improvement in classroom learning in their local environment; they need to see *action*. What happened in this study? Problems were identified; interventions were made; input from students was positive; signs are that learning took place. One thing learned from the results of Intervention 1 (the focus groups) was that academic performance in general chemistry (which is demonstrated basically by good performance on exams) does not necessarily correlate to conceptual understanding. S2, one of the students in the focus group study, had received an A in Chem 1A, which put him in the top ~5% of his class, yetS2 was one of the most ardent adherents to some of the most basic misconceptions seen in the focus group study. Once again, academic, problem-solving performance is seen not to fully correlate with meaningful understanding

Recommendations for Further Action

In the spirit of action research, recommendations here are directed to the next step of a program to continue to address the problems identified in this study. The first recommendation is to open the scope of this research to include Chem 1A as well as Chem 1B. It is obvious that many of the problems seenin Chem 1B come from students entering the Chem 1B classroom with basic misconceptions in areas crucial to the understanding of chemical equilibrium. It is possible that changes in the Chem 1A curriculum could be made to address this issue. This would require another action research step (data gathering, intervention, assessment) to be made. Given the identification of some of these topics by the current study, this step would be expected to not only serve a research purpose, but also to have some immediate results. The recommendation would be to focus on topics which are critical to success in Chem 1B, and particularly in chemical equilibrium.

At the same time, it is recommended that research be moved into the Chem 1B classroom, specifically into the lab. This would provide a larger data set, allowing for more significant results. Some observations have been madeof the difficulties of implementing the constructivist style into larger classrooms due to the large amount of student interaction it requires of the instructor. Chem 1A and 1B labs, with a class size of 24, are large enough to present a challenge to the constructivist instructor. There are a number of ways to address this. First, it is recommended that some graduate students be brought into the Chem 1B labs as assistants to the primary instructor. This would serve two purposes. The presence of multiple instructors in the room would distribute the workload of student interaction, and time spent in a Chem 1B lab would enhance the chemistry knowledge and teaching skills of graduate studentspursuing a career in chemistry teaching. Work as a Chem 1B teaching assistant would be of particular value to those graduate students who have a career interest in education. This arrangement has the attractive result of helping to improve both graduate and undergraduate students' career skills at the same time.

Lab introduction of conceptual learning could be accomplished with many different possible logistical changes. The number and frequency of the guided-inquiry sessions could be adjusted according to the topics to be presented. Even a conversion of only 20% of the current lab time to guided inquiry sessions would result in an hour a week, a huge time investment for this endeavor with little change necessary to the current curriculum, which would require some revision and condensation.

Extracurricular guided inquiry sessions might be arranged, with students attending on a volunteer basis. This may or may not meet with success, based on how it was implemented by the Chem 1B course instructor. Efforts on the part of this researcher to solicit volunteers for research met with mixed success. Very few volunteers were found for the guided-inquiry sessions, and the volunteer rates for the less effective worksheet downloads were only somewhat better. Continued use of the worksheet download process is not recommended without significant changes. At the very least, a survey should be attached to the worksheets so that the researcher can get input on the effectiveness of the worksheet.

Some revision of the intervention tools is recommended. The worksheets have been revised extensively for the download process, but the absence of an instructor/guide requires more worksheet text. The Equilibrium Machine proved to be quite effective in the Focus Groups, but was only designed to model a simple $A \rightleftharpoons B$ reaction, and could be upgraded to reflect more complex reactions and accept more types of student input.

Theintroduction of a constructivist track would broaden the academic offering of CSUS chemistrydepartment. Such diversification can be seen in existing chemistry programs. Stanford University, currently ranked number five in the list of US undergraduate institutions, has integrated aconstructivist approach into its general chemistry curriculum. In the past decade, Stanford introduced a second track of general chemistry training; while retaining its traditional general chemistry course, a constructivist-based track was introduced which features laboratory sessions conducted in a guided-inquiry manner. As opposed to the five hours per week spent by most university students in general chemistry lab, Stanford students attendone one-hour lab per week. Lab sessions are based on constructivist teaching of conceptual understanding in an instructor-guided inquiry format, using worksheets developed by Stanford professors and graduate students (see Appendix E for a copy of two of these worksheets).The use of a curriculum based on constructivist, conceptual education by a nationally known and highly regarded institution, with no apparent reduction in the quality of its graduates, demonstratesthe feasibility of this approach, and shows an instance of a major chemistry program which has decided to commit to the goal of conceptual education.

Chapter 8

FINAL THOUGHTS

While conceptual understanding and algorithmic problem-solving skills are both necessary for academic success, and are related to one another to some degree, it is arguable that the acquisition of one skill does not necessarily lead to the other. It could reasonably be argued that the understanding of a chemical concept is valuable to the sense of intuition required to solve problems when rote alone fails; the application of kinetic theory to the Le Chatelier problem in the Data Analysis is a good example of this. However, even if it were possible to make use of classical quantitative analysis in the study of the benefits of conceptual understanding, one might not be able to correlate them with academic success. And if conceptual learning (and teaching) cannot be counted on to increase test scores and course grades, what is it good for? Like any good problem solving exercise, the answer to this question depends on the result desired; what is our objective in the teaching of general chemistry? And are we doing our best to meet that objective?

Dr. J. H. Hildebrand, a professor of general chemistry at the University of California at Berkeley in the early 20th century, has been quoted several times in the above literature search, and has something to say on this topic. In a 1922 article in *Science* entitled "The Early Training of Scientists', he offers his view of the mission of the general chemistry instructor;

> We have been inspired by the opportunity offered by a fundamental course to present science in such a way as, first, to win for scientific careers the keen-minded students who are repelled by the drudgery and memory work of the old-fashioned course in descriptive chemistry; and second, to encourage the average student to adopt the scientific attitude towards his everyday problems, an

> attitude so necessary in combatting the superstition, prejudice, selfishness and dishonesty in the world of ordinary affairs.[21]

The current Mission Statement of the College of Natural Sciences & Mathematics at CSUS, of which the chemistry department is a part, offers a similar sentiment;

> The College of Natural Sciences and Mathematics educates a richly diverse scientific workforce and fosters scientific and mathematical literacy in the greater community. Our degree programs, research opportunities, internships and strong student support prepare graduates to achieve lifetime career success, contribute to knowledge and help solve societal problems. Our faculty, students and graduates promote regional progress through professional activity and community collaboration.

What exactly should our purpose be as educators? What should we be preparing our students for? Is there a better way to spend the time and money currently expended on the hundreds of students who fail a general chemistry course at CSUS every year? Should our sole objective be the production of future scientists? Or can we at the same time create better citizens, a group of individuals who will bring a better overall understanding of science and chemistry into the "world of ordinary affairs"? These are questions well outside the domain of this study, but questions which have been are hard to ignore in its execution. I leave them here for the consideration of future teachers (the author included).

APPENDIX A

Diagnostic Exam

Chemical Equilibrium Diagnostic **Student Name:**

Answer the following questions.
Some of the multiple choice questions have more than one correct answer.

Questions 1-7 below refer to the following reaction: $\mathbf{2AB + C_2 \rightarrow 2AC + B_2}$ (Reaction 1)

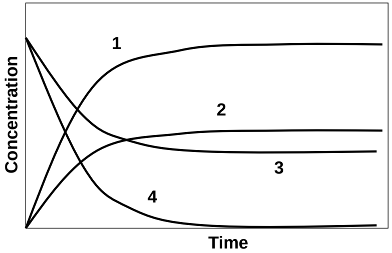
1) In the above reaction, draw a circle around any reactants and draw a box around any products.

2) Is this reaction balanced? Why or why not?

3) Why do we require equations to be balanced?

4) For the chart below, circle "a" or "b" to indicate which statement is true:

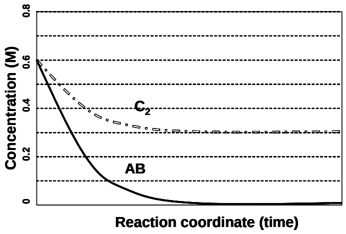
   a) Curves 1 and 2 are reactants; curves 3 and 4 are products
   b) Curves 1 and 2 are products; curves 3 and 4 are reactants

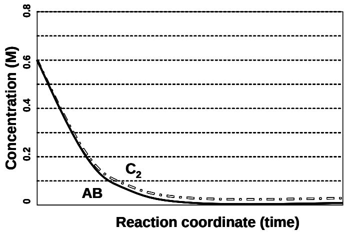


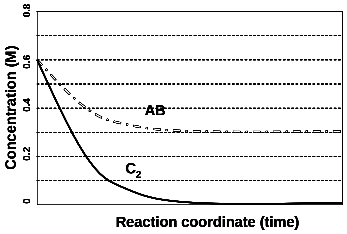
5) As the above reaction proceeds, which reactant *element* will be consumed most slowly?
   a) A b) B c) C d) All will be consumed at the same rate

6) As the above reaction proceeds, which product *molecule* will be generated most quickly?
   a) AB b) C_2 c) AC d) B_2

7) Assume that the above reaction goes to completion. If the initial concentrations of AB and C_2 are 0.6M, which of the following graphs is possibly correct?







a) b) c)

Questions 8-9 below refer to the following reaction: **$aA + bB \rightarrow cC + dD$** (Reaction 2)

The rate of Reaction 2 can be measured by how fast reactants are consumed (how fast reactant concentrations change over time) or how fast products are generated. For example:

$$\text{Rate} = -\frac{1}{a}\cdot\frac{\Delta[A]}{\Delta t} = \frac{1}{c}\cdot\frac{\Delta[C]}{\Delta t}$$

Given this equation, it can be seen that the reaction rate is proportional to the concentrations of reactants and products in a reaction.

8) Why is one of these terms positive and one negative?

9) What will happen to the rate of a reaction if [A] is increased? If it is decreased?

Questions 10-12 below refer to the following exothermic reaction. Note that each of these questions requires you to circle both an answer and a *reason* for your answer. Remember that more than one answer may be correct.

$2NO_{(g)} + Cl_{2\,(g)} \rightleftharpoons 2NOCl_{(g)} + \text{heat}$ (Reaction 3)

This is a “reversible” reaction (indicated by the $\rightleftharpoons$) consisting of two related reactions, a “forward” reaction (corresponding to forward arrow) and a “reverse” reaction (corresponding to the reverse arrow). Each reaction has its own rate and we can speak of a “forward rate” and a “reverse rate”.

10) The NO and Cl_2 gases are mixed and the reaction begins. Which of the following statements are true?
    b) The reverse reaction will not begin until the forward reaction is complete
    c) The forward and reverse reactions both take place at the same time
    d) The forward reaction will initially be faster than the reverse reaction
    e) The forward and reverse reaction rates will always be equal

Reason for the answer you selected in question 10
    c) The forward reaction is driven by the reactants and always goes first
    d) Heat drives the reverse reaction while the forward reaction is running
    e) Reaction rates are a function of concentration, so both forward and reverse reactions will run simultaneously

11) After the NO and Cl_2 are mixed and the reaction begins,
    a) The concentration of NOwill increase
    b) The concentration of Cl_2 will increase
    c) The concentration of NOCl will increase
    d) Cannot make a prediction

Reason for the answer you selected in question 11
        a) An exothermic reaction will produce a higher concentration of products

  - b) The concentration of Cl_2 will decrease relative to NOand NOCl because of its lower coefficient
  - c) As the reaction proceeds, NOand Cl_2 will be consumed to generate NOCl
  - d) Cannot answer; all concentrations are determined by a chemical equation
- f) I guessed
- g) Other:_____________________________

12) As the reaction continues and proceed toward the right (the forward reaction),
- a) The rate of the forward reaction decreases with time and is greater than that of the reverse reaction (proceeding products to reactants)
- b) The rate of the reverse reaction increases with time and is greater than that of the forward reaction
- c) The forward reaction progresses until completion, then the reverse reaction starts
- d) The forward and reverse reaction rates are equal

Reason for the answer you selected in question 12
- a) There are two reactants and only one product
- b) This is an exothermic reaction, so the forward reaction will be faster
- c) As the reaction starts and proceeds to equilibrium, reactants are consumed, reducing the rate of the forward reaction
- d) Forward and reverse reaction rates of reversible reactions are always equal
- h) I guessed
- i) Other:_____________________________

At some point in time the reaction will reach a “steady state”, at which time the concentrations of both reactants and products and the rates of the forward and reverse reactions will stop changing and become constant. This is the point of “chemical equilibrium”.

13) At the equilibrium point,
  - a) The concentration of all products and reactants will be equal
  - b) The concentration of reactants will be greater than the concentration of products
  - c) The concentration of products will be greater than the concentration of reactants
  - d) Cannot predict the relative concentrations of products and reactants

Reason for the answer you selected in question 13
  - a) Equality of concentration is the definition of equilibrium
  - b) Concentrations of reactants are higher due to the stoichiometric relationships of reactants and products
  - c) This is an exothermic reaction so concentrations of products are higher
  - d) Cannot answer; all concentrations at equilibrium are determined by a chemical equation
- j) I guessed
- k) Other:_____________________________

14) When equilibrium is reached,
- a) The forward and reverse reaction rates will go to zero
- b) The forward reaction rate will be greater than the reverse reaction rate
- c) The reverse reaction rate will be greater than the forward reaction rate

d) The forward and reverse reaction rates will oscillate around the equilibrium point
e) The forward and reverse reaction rates will be equal and non-zero

Reason for the answer you selected in question 14

a) At equilibrium, both forward and reverse reactions are complete
b) The forward reaction is always greater for an exothermic reaction
c) Once equilibrium is reached, the forward reaction is complete and the reverse reaction begins
d) Both forward and reverse reactions will continue as long as reactants and products exist in the mixture
e) Forward and reverse rates must oscillate as concentrations increase and decrease
l) I guessed
m) Other:______________________________

Questions 15-16 below refer back to Reaction 3.

$2NO_{(g)} + Cl_{2\,(g)} \Leftrightarrow 2NOCl_{(g)}$ + heat (Reaction 3)

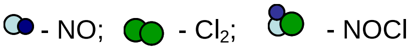
15) An equimolar solution of NO and Cl_2 is prepared and allowed go to equilibrium (Reaction 3). Which of the following statements are true?

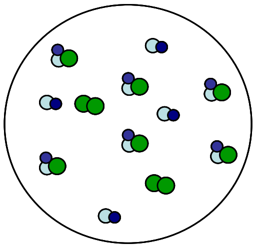
a) Concentrations of NO and NOCl in the equilibrium solution are equal
b) Only NOCl and Cl_2 will be found in the equilibrium solution
c) The concentration of NO will be half that of Cl_2 in the equilibrium solution
d) The concentration of Cl_2 will be greater than that of NO in the equilibrium solution
e) None of the above

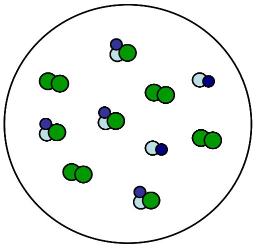
Reason for the answer you selected in question 15

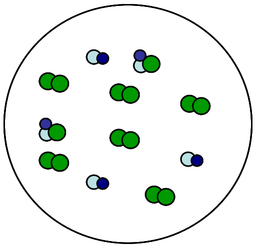
a) NO is the limiting reactant, so it will go to zero in the final solution
b) NO is consumed twice as fast as Cl_2, so its concentration will be lower than Cl_2
c) NO and NOCl concentrations will be equal as they have the same stoichiometry
d) The reaction uses two NO molecules for every molecule, so there will only be half as many NO molecules left at equilibrium
e) Cannot make a statement about relative concentrations based on the given information
n) I guessed
o) Other:______________________________

16) Which of the following could NOT be a representation of the equilibrium state for the reaction described in Question 15?


- NO; - Cl_2; - NOCl


a)


b)


c)

APPENDIX B

Worksheets and Answer Keys

Worksheet 1: Reaction Kinetics and Stoichiometry

(Posted Version w/answer key)

Learning Objectives

- Students will be able to identify the relative rate of consumption and generation of reactants and products given a balanced equation for a chemical reaction
- Students will be able to describe the change in concentrations of reactants and products over time
- Students will be able to describe the change in reaction rates over time

Exercise 1: Stoichiometry

For the reaction $2AB + C_2 \rightarrow 2AC + B_2$ (Equation 1)

a) What does the coefficient of 2 in front of the reactant AB mean?

The coefficient means that there are 2 molecules (or moles) of compound AB in this reaction.

b) What does the subscript 2 in the reactant C_2 mean?

This subscript means that the compound noted contains 2 atoms of the element C. The absence of a coefficient means that the notation means that there is one molecule (or mole) of compound C_2*in this reaction.*

c) Suppose that we mix equal amounts of AB and C_2 to initiate a reaction. Which reactant do you predict will be consumed more quickly (at a faster rate)? Which product is produced more quickly? Why?

In this reaction, compound AB will be used up twice as quickly as compound C_2*.*

This reaction can be read as follows: "Two molecules (or moles) of AB combine with one molecule (or mole) of C_2 *to form two molecules (or moles) of AC and one molecule (or mole) of* B_2*." Thus two moles of AB will be consumed for every mole of* C_2 *consumed, or the* ***rate of consumption*** *of AB is twice as fast as that of* C_2*.*

Note that, in this course, in problems which take place in a fixed volume of solvent, the concentration of a compound may be used in place of its quantity (that is, moles or molecules).

d) When the reaction is complete, what substances will remain in solution, and at what ratios?

If equal amounts of AB and C_2 are present at the start of the reaction, then since AB is consumed twice as fast as C_2, when all of the AB is consumed, half of the C_2 will remain. ***At completion, there will be no AB remaining, and half the original amount of C_2 will remain in solution****.*

In Reaction 1, the stoichiometric ratio of AB and AC is 1:1, so one mole of AC will be generated for each mole of AB consumed. ***At completion, since all of the AB will have been consumed, an equal amount of AC will have been generated****.*

By this reasoning, the amount of B_2 generated is equal to the amount of C_2 consumed. Since we know that half the C_2 will remain, we know that half the B_2 has been consumed (and that half remains). ***At the completion of the reaction, B_2 andC_2 will be present in equal amounts.***

Since the stoichiometric ratio of AC to B_2 is 2:1, AC will have been generated at twice its rate. ***At completion, the quantity of AC will be twice that of B_2 and, by extension, of C_2.***

*In summary****; at completion, AC, B_2 and C_2 will remain in solution, in a ratio 2:1:1.***

To clarify this problem, perform the exercise below. Use the table below containing six instances each of the reactant molecules AB and C_2:

Reactants	Products
AB C_2	
AB C_2	
AB C_2	
AB C_2	
AB C_2	
AB C_2	

1) Based on Equation 1, convert one set of reactant molecules to the indicated number of product molecules. Cross out the reactant molecules which are used; write in the product molecules which are created.

Reactants	Products
AB C_2	*AC B_2*

AB C_2	*AC*
AB C_2	
AB C_2	
X C_2	
X X	

2) Repeat this exercise two more times.

Reactants	Products
X C_2	*AC* B_2
X C_2	*AC* B_2
X C_2	*AC* B_2
X X	*AC*
X X	*AC*
X X	*AC*

3) Do your results support your answer to Question c and d above? Explain.
As predicted above, AB has been consumed twice as quickly as C_2*. The three remaining solution components are as predicted and in the same ratios as in question d.*

Exercise 2: Reaction Rates and Concentrations Over Time

(See also “The Rate Law for the Reaction of Iodide with Bromate” in the Chem 1A lab manual and Chapter 15 of the Chem 1A textbook.)

Recall from Chem 1A that for the generic reaction

$$aA + bB \rightarrow cC + dD$$

The reaction rate can be expressed as

$$\text{Rate} = k[A]^x[B]^y$$

where k is the *rate constant* for the reaction and x and y are the *orders of reaction* with respect to reactants A and B. Recall that the reaction orders are unrelated to the *coefficients* of the reactants (the terms a, b, c and d), and must be determined empirically.

Thus, for the simple reaction

$$aA \rightarrow bB$$

where a=b=1 and the reaction orders of A and B are 1, we can write the simple rate law

$$\text{Rate} = k[A] \qquad \text{(Equation 2)}$$

Since the reaction rate is defined as the change in concentration over time, its units will be M/sec, so if we know the rate constant k, and begin our reaction with a known concentration of A, we can obtain the reaction rate at any point in time and thus compute changes in concentrations at any time in the reaction. For example; for the above reaction, if we start with [A] = 100M/sec, and the rate constant is k=(0.2) s^{-1}, from equation 2 the reaction rate will be (100M)*(0.2 s^{-1}) = 20M/sec. Thus, after one second has passed, the concentration of A has decreased by 20M.

a) As the reaction A → B starts at time = 0, what is the initial reaction rate if the starting concentration of A ([A]) is 500M and the rate constant k = 0.3s^{-1}? (Refer to Equation 2.)

For this reaction, the reaction rate at any time is equal to k[A]. At time 0, the rate is (0.3 $^{-1}$) * (500M) = 150[M]/sec.*

b) After one second, what is the concentration of A? Of B?

If the reaction rate is 150 M/sec, after the first second, 150M of A have been converted to B. Thus the initial concentration of A has been reduced by 150M, and after 1 second its total concentration is (500-150) = 350M. The 150M of A lost have been converted into B, and thus the total concentration of B now in the solution is 150M.

c) What is the reaction rate k[A] at time t = one second? (Refer to Equation 2.)

*The reaction rate **at any time** is equal to k*[A]. Since [A] has been reduced, the reaction rate will also be reduced. At t = 1 second, the rate is (0.3s $^{-1}$) * (350M) = 105 M/sec.*

Given the following information, complete the table below. The first few lines of the table have been completed as an example.

Reaction: A → B
Reaction rate order = 1
Rate constant k = 0.2 s^{-1}
Initial concentrations: [A] = 100M, [B] = 0M

Time (sec)	Δ[A] (M)	[A] (M)	Δ[B] (M)	[B] (M)	Reaction rate k* [A] (M/sec)
0	0.0	100.0	0.0	0.0	20.0
1	-20.0	80.0	20.0	20.0	16.0
2	-16.0	64.0	16.0	36.0	12.8
3	-12.8	51.2	12.8	48.8	10.2
4	-10.2	41.0	10.2	59.0	8.2
5	-8.2	32.8	8.2	67.2	6.6
6	-6.6	26.2	6.6	73.8	5.2
7	-5.2	21.0	5.2	79.0	4.2
8	-4.2	16.8	4.2	83.2	3.4
9	-3.4	13.4	3.4	86.6	2.7
10	-2.7	10.7	2.7	89.3	2.1

d) In your table, what happens to the concentration of A over time? What happens to the concentration of B over time? Does this make sense?

As with any reaction, the reactant concentration ([A]) will decrease with time as it is converted to product, which will increase in concentration.

e) What happens to the reaction rate over time? Does this make sense?

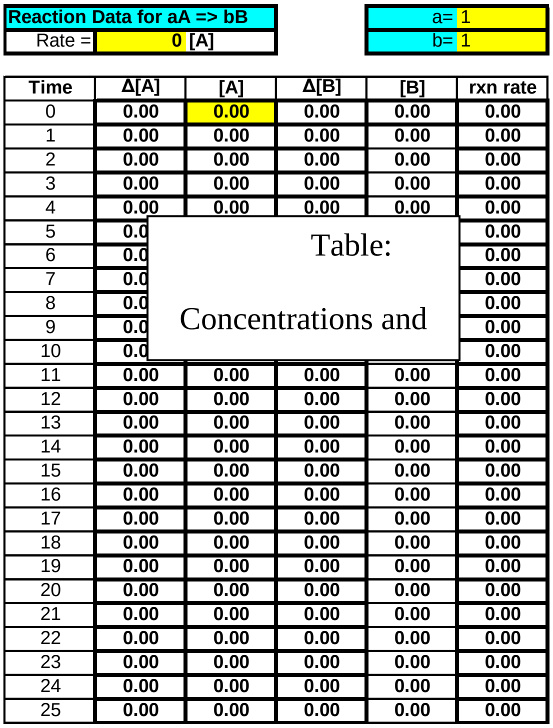
Reaction rate decreases over time. This must be the case, as a reaction begins at a positive rate and concludes at ~zero.

f) What can you say about the nature of the reaction rate over the time of a chemical reaction?

After initiation of a chemical reaction, its rate decreases as a function of time.

Exercise 3: More on Rates and Concentrations

We can observe the progress of a chemical reaction without using manual calculations. Open the Equilibrium Machine Excel icon and go to tab WS 1. You will see the following screen:



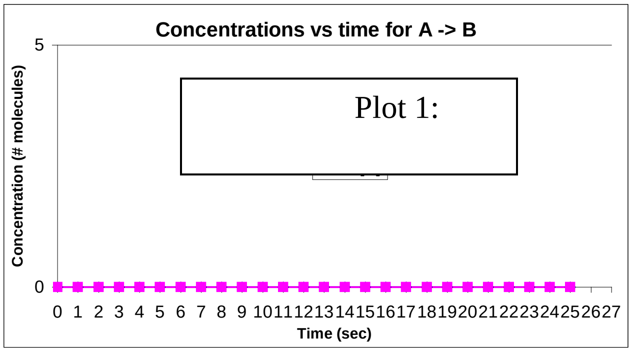
Reaction Data for aA => bB
Rate = 0 [A]

a= 1
b= 1

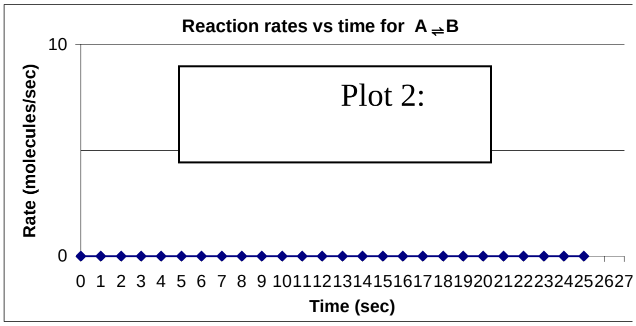
Time	Δ[A]	[A]	Δ[B]	[B]	rxn rate
0	0.00	0.00	0.00	0.00	0.00
1	0.00	0.00	0.00	0.00	0.00
2	0.00	0.00	0.00	0.00	0.00
3	0.00	0.00	0.00	0.00	0.00
4	0.00	0.00	0.00	0.00	0.00
5	0.0				0.00
6	0.0				0.00
7	0.0				0.00
8	0.0				0.00
9	0.0				0.00
10	0.0				0.00
11	0.00	0.00	0.00	0.00	0.00
12	0.00	0.00	0.00	0.00	0.00
13	0.00	0.00	0.00	0.00	0.00
14	0.00	0.00	0.00	0.00	0.00
15	0.00	0.00	0.00	0.00	0.00
16	0.00	0.00	0.00	0.00	0.00
17	0.00	0.00	0.00	0.00	0.00
18	0.00	0.00	0.00	0.00	0.00
19	0.00	0.00	0.00	0.00	0.00
20	0.00	0.00	0.00	0.00	0.00
21	0.00	0.00	0.00	0.00	0.00
22	0.00	0.00	0.00	0.00	0.00
23	0.00	0.00	0.00	0.00	0.00
24	0.00	0.00	0.00	0.00	0.00
25	0.00	0.00	0.00	0.00	0.00

Table:

Concentrations and



Plot 1:

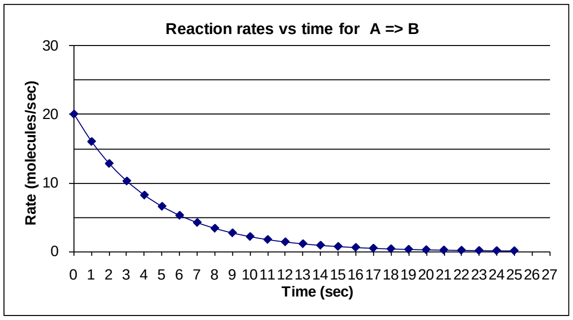


Plot 2:

The table and plots model the reaction aA =>bB. There are a number of reaction parameters which can be set manually (in the shaded boxes). These include the reaction rate constant k, the starting concentration of A, and the reaction coefficients of A and B (a and b).

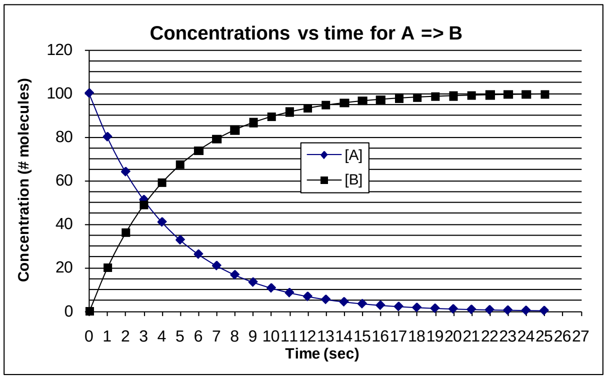
Set the rate constant to $0.2s^{-1}$ and [A] at time 0 to 100M, leaving the coefficients a and b at 1.

a) Examine the Rate vs. Time plot (Plot 2, at the lower right). What can you say about the reaction rate vs. time? Does this hold with your findings in Exercise 2 above?



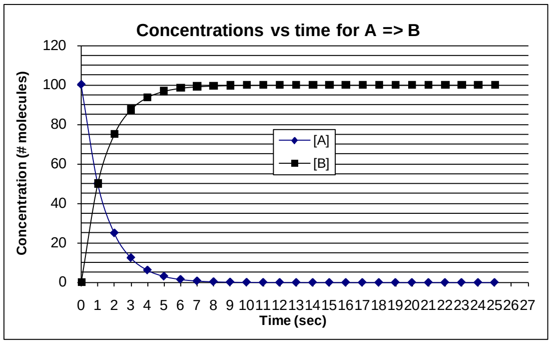
As seen in the previous problem, the reaction rate decreases over time after the initiation of the reaction.

b) According to the concentration vs. time plot (Plot 1, at the upper right), at approximately what time does the reaction finish (i.e., when does [A] go to zero?).



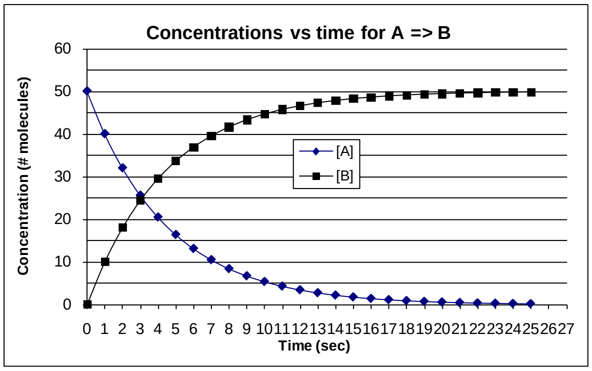
The reaction completes at about t = 20 seconds as [A] approaches 0.

c) Change the rate constant k to 0.5. At what time does the equation now finish? Can you explain this? (You can flip back and forth between these two curves using the Back and Forward arrows on the menu to better compare them.)



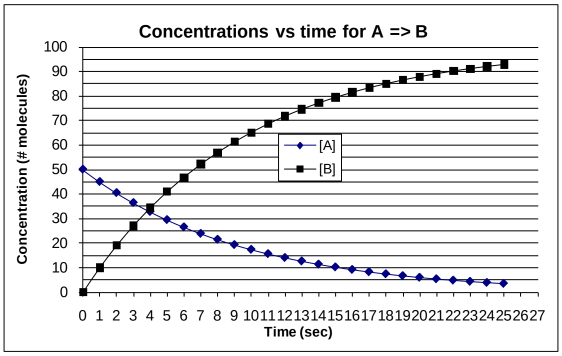
The reaction completes at about t = 7 seconds. Because of its higher rate constant, this reaction proceeds a faster rate.

d) Change the rate back to 0.2, and change the initial concentration of A to 50. Does this change the completion time of the reaction? Can you explain this?



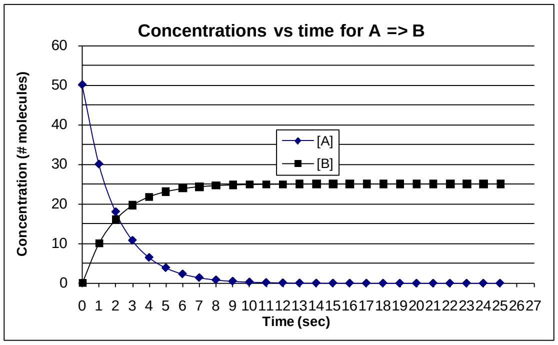
The completion time of a reaction is a function of its rate, which in turn is a function of its rate constant. Thus since the reactions in questions h and j have the same rate constant, they will complete at the same time.

e) Change coefficient b to 2. How did the concentration plot change? Can you explain this?



The reaction now being plotted is A → 2B. According to the stoichiometry of the reaction, two molecules of B are being produced for every molecule of A consumed; or, in other words, the rate of production of B is twice the rate of consumption of A. A look at the concentration vs. time plot shows that the rate of change of [B] (as measured by the steepness of the curve) is noticeably greater than the rate of change of [A]. Notice that, having started the reaction with [A] = 50M, [B] at the completion of the reaction is now 100M, twice what it was in the previous question.

f) Now change b back to 1 and set a equal to 2. What has happened to the plot? Can you explain this?



Now the reaction now being plotted is 2A → B. Two molecules of A are required for the production of each molecule of B; or, in other words, the rate of production of B is half the rate of consumption of A. The plot shows that the rate of change of [B] is now much less than the rate of change of [A]. At the completion of the reaction, when the quantity of A is exhausted, only 25 M of B (half the 50M starting concentration of A) has been produced.

Rev 13 8/25/10

Revision 6 2/2/10
- The first revision used for Focus Group instruction.

Revision 7 2/4/10
- Added Learning Objectives
- Added Diagnostic Test review (Exercise 1); split into three exercises
- Simplified intro to Ex 2

Rev 8 7/8/10
- Clarified instructions for Ex 2
- Changed question d and following

Rev 9/10 7/9/10
- Added question c to Ex. 1
- Clarified questions for Ex 2
- Added instructions for supplemental exercise in Ex. 1

Rev 11 10 7/13/10
- Added question b to Ex. 1
- Clarifications to Ex. 2 notes and instructions
- Clarifications to Ex. 3 questions

Rev 12 10 7/14/10
- Removed plot from Exercise 1

Rev 13 8/25/10
- Edits for posting

Worksheet 2: Reversible Reactions and Chemical Equilibrium

(Posted Version w/answer key)

Exercise 1: Reversible Reactions

In Worksheet 1, we reviewed the basics of irreversible reactions, in which reactants are completely consumed and converted into products. A ***rate equation*** can be written for this reaction which expresses the rate in terms of reactant concentrations.

For example; for the reaction AB + C → AC + B (Reaction 1),

$$\text{Rate} = k[AB]^x[C]^y \qquad \text{(Equation 1)}$$

where k is the "rate constant" and ***is unique to a particular chemical reaction.*** The exponents x and y are the orders of reaction with respect to AB and C; for this worksheet, we will assume that all orders of reaction are equal to 1.

As opposed to this irreversible reaction above, which is written with a one-directional arrow, reversible reactions are written with dual arrows. For example,

$$A + B \rightleftharpoons AB \text{(Reaction 2)}$$

The dual-arrow format indicates that, as opposed to an irreversible reaction, which runs in one direction (left to right) until completion, this reaction is really composed of ***two opposed reactions running at the same time***. In a reversible reaction, we can refer to the existence of "forward" and "reverse" reactions. For example, for Reaction 2 above,

Forward reaction: A + B → AB (Reaction 3)
Reverse reaction: AB → A + B (Reaction 4)

a) For the reversible reaction $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$, what are the reactants? What are the products?

According to the conventions we know, reactants are on the left ($H_{2(g)}$ and $I_{2(g)}$) and products on the right ($HI_{(g)}$).

b) For the reversible reaction $2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$,what are the reactants? What are the products?

According to the conventions we know, reactants are on the left ($HI_{(g)}$) and products on the right ($H_{2(g)}$ and $I_{2(g)}$).

c) Are the reactions in a) and b) different or are they the same? What can you say about the meaning of "reactants" and "products" of a reversible reaction?

The two equations are really describing the same reaction; they are just written differently. According to the "dual-arrow" convention of the reversible reaction equation, each reaction as written really consists of the same two simultaneous reactions:

$$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$$
$$2HI_{(g)} \rightarrow H_{2(g)} + I_{2(g)}$$

Since the reactions in questions a and b are essentially the same, and yet the answers to questions a and b are contradictory, we must conclude that in the writing of a reversible reaction, there are no true "reactants" and "products", only a convention as to the written form of a chemical equation which says that reactants are on the left and products are on the right.

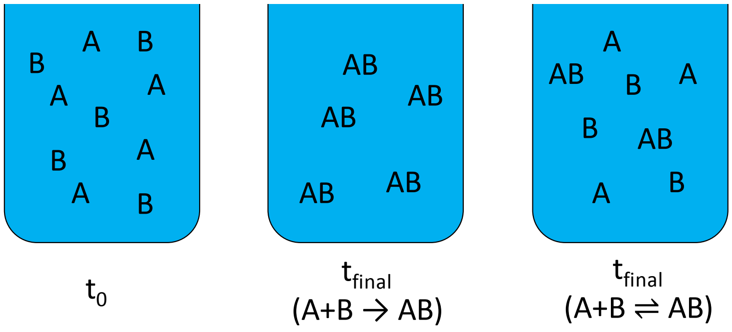
Another way to view this anomaly is that, in a reversible reaction, all "products" and "reactants" (in this case, AB and A and B) are *continuously generated and consumed.* Accordingly, we can write rate equations for both forward and reverse reactions.

For example; for reactions 3 and 4 above;

$$\text{Forward rate} = k_f\,[A][B] \quad \text{(for Reaction 3)}$$
$$\text{Reverse rate} = k_r\,[AB] \quad \text{(for Reaction 4)}$$

where k_f is the forward rate constant and k_r is the reverse rate constant.

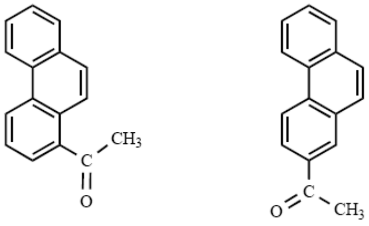
d) At time 0, one-liter solutions of A and B with equal concentrations are mixed (represented in the left-hand box), and react to form the product AB. Draw a representation of the final components in solution after the reaction is complete if the reaction in the center box is irreversible the reaction in the right-hand box is reversible.



In an irreversible reaction (middle box), ALL reactants are converted to products, and effectively no reactants remain in solution. In a reversible reaction (rightmost box), all *"reactants" and "products" will remain in solution. In the illustration above, two molecules of product have been formed, while three atoms each of A and B remain in solution. These quantities will vary according to the extent to which the reaction proceeds to the left or right.*

Exercise 2: Chemical Equilibrium

Consider the simplest reversible reaction: the dynamic rearrangement of a molecule in solution. Many organic molecules exist in multiple forms in solution due to ongoing changes in their molecular structure as a result of bond rearrangement. One such reaction is the transition of acetylphenanthrene between two of its numerous isomers:



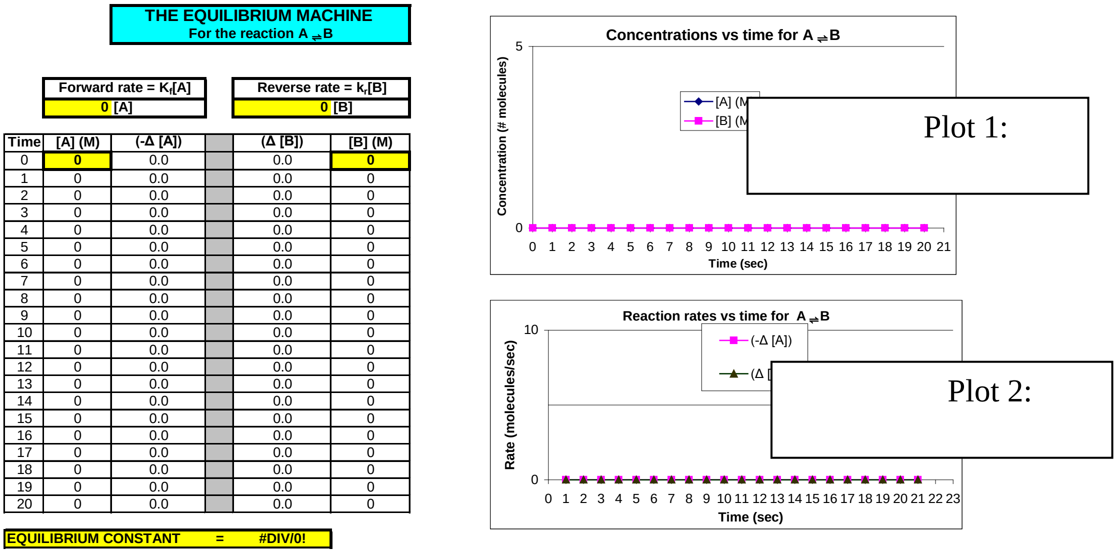
or, to simplify, $A \rightleftharpoons B$

For this reaction, following the form of reactions 4 and 5, we can write

Forward rate = k_f [A] (Equation 2)
Reverse rate = k_r [B] (Equation 3)

Analogous to the plotting exercise in Worksheet 1, plots of concentrations vs. time and reaction rates vs. time for this reaction can be created based on initial concentrations and rate constants. This exercise can be automated by using an Excel spreadsheet to perform the calculations and generate the plots.
In this case, we will model the acetylphenanthrene rearrangement reaction above ($A \rightleftharpoons B$).

Open the WS2 tab in the Equilibrium Machine. You will see a screen like this:



THE EQUILIBRIUM MACHINE
For the reaction A ⇌ B

Forward rate = K_f[A] **0 [A]**

Reverse rate = k_r[B] **0 [B]**

Time	[A] (M)	(-Δ [A])		(Δ [B])	[B] (M)
0	0	0.0		0.0	0
1	0	0.0		0.0	0
2	0	0.0		0.0	0
3	0	0.0		0.0	0
4	0	0.0		0.0	0
5	0	0.0		0.0	0
6	0	0.0		0.0	0
7	0	0.0		0.0	0
8	0	0.0		0.0	0
9	0	0.0		0.0	0
10	0	0.0		0.0	0
11	0	0.0		0.0	0
12	0	0.0		0.0	0
13	0	0.0		0.0	0
14	0	0.0		0.0	0
15	0	0.0		0.0	0
16	0	0.0		0.0	0
17	0	0.0		0.0	0
18	0	0.0		0.0	0
19	0	0.0		0.0	0
20	0	0.0		0.0	0

EQUILIBRIUM CONSTANT = #DIV/0!

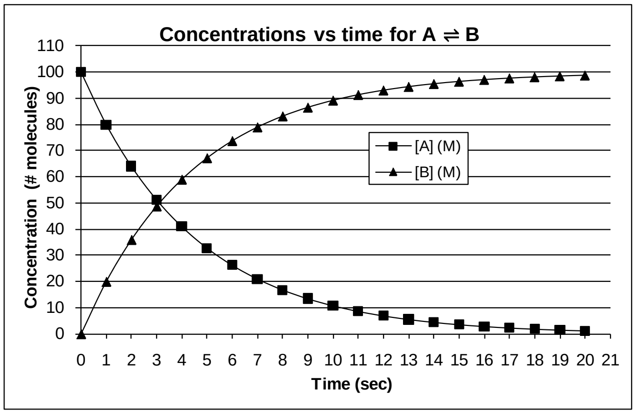
Take a moment to study the layout of this page. Make sure you understand what values are being represented in the table columns and how these values appear in the plots to the right.

Starting concentrations of A and B and rate constants k_f and k_r can be entered into the shaded yellow boxes at the top of the spreadsheet. Plots of concentration vs. time and rates vs. time for each set of parameters entered are displayed on the right.

Set the following initial conditions in the table:

Forward rate constant = 0.2 s^{-1}
Reverse rate constant = 0
Initial [A] (at time 0) = 100
Initial [B] (at time 0) = 0

Hit “return”, and the plots will model the irreversible reaction you worked with in Worksheet 1.

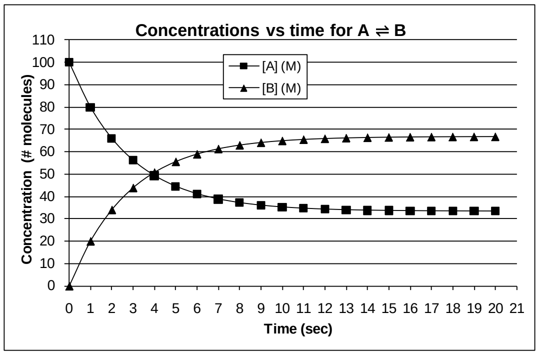


e) After a period of time, the plot of concentration vs. time flattens out. What is happening at the point? What compounds are left in solution at this point?

As the plots become flat, the concentrations in the solution begin to stop changing and start to reach a stable, constant level. The level of compound A approaches zero as it becomes depleted to form B, and the final, stable level of B approaches the initial level of A (in this case, 100M).

This is a plot of the irreversible reaction from Exercise 2 in Worksheet 1. As in all irreversible reactions, the reactants are completely depleted to form product. Thus, only B remains at reaction completion.

Now change the reverse rate constant to 0.1. ***You have added a reverse reaction to the chemical equation,, and the plots are now modeling a reversible reaction.***



f) How does this plot differ from the plot for the irreversible reaction you just looked at?

The plot still shows reactant concentration decreasing with time, and product concentration increasing with time. Both concentrations approach a flat, constant state *as before. However, the concentration of compound A is not fully depleted, and* ***at the point in time when both concentrations become flat, both A and B remain present in the solution****.*

g) Note that the curves in the concentration plot begin to flatten out after a period of time. What is happening at this point? How does this region differ from that of the irreversible reaction? What compounds are left in the solution at this point?

As in Worksheet 1, when to curves begin to flatten, the concentrations of the compounds in solution begin to stabilize and reach a constant value. If we let this reaction sit for another minute or so., the curves would become absolutely flat, indicating that NO further change in concentration is taking place. As opposed to the irreversible reactions, both concentrations are positive.

h) Look at the Rate plot (Plot 2). What two values are plotted here? Look at the region of the plot which corresponds to the "flat" region you saw on the concentration plot. How are these two rates related in this region?

The two values plotted are the Forward and Reverse rates of the reaction. As with the concentration plots, these curves become flat and unchanging at a given point, indicating that the reaction rates have ceased to change and have become stable and constant. In addition, in the stable region, both forward and reverse rates are 1) positive (that is, the rate does not go to zero as in the irreversible reaction), and 2) equal.

i) At t=18 seconds, the forward reaction rate ($-\Delta$ [A]) is ~6.7M/s, yet the product concentration is not changing. How do you explain this?

When the reaction stabilizes (in the "flat zone", where concentrations and rates become stable), we essentially have two reactions (forward and reverse) both continuing at the same rate. Because the forward reaction is producing "products" (the compounds on the right side of the equation) at the same rate that the reverse reaction is producing "reactants" (compounds on the left side of the reaction), the net result is that no net gain or loss of compounds occurs.

The "flat" areas on the plots signify a state of the reaction known as "equilibrium".At equilibrium, all concentrations are constant and unchanging, and forward and reverse rates are equal. An "equilibrium constant" (symbolized as K_{eq}) for this reaction has been calculated by the Excel program and can be seen at the bottom left of the screen, below the table.

K_{eq} is related to the relative concentrations of compounds at the equilibrium state. Since these concentrations are constant, K_{eq} is itself a constant for this reaction.

For the reaction $aA + bB \rightleftharpoons cC + dD$, Keq is expressed as the ratio of product over reactant concentrations;

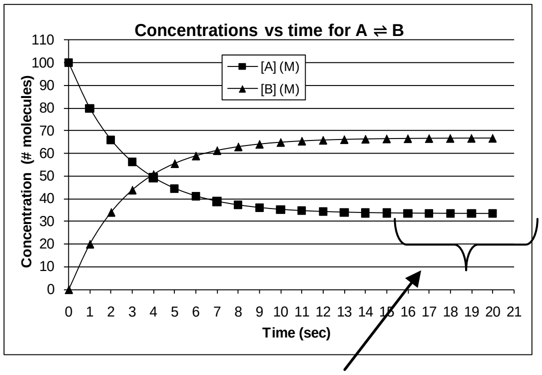
$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b} \quad \text{(Equation 4)}$$

For the simple reaction A ⇌ B, the equilibrium constant is

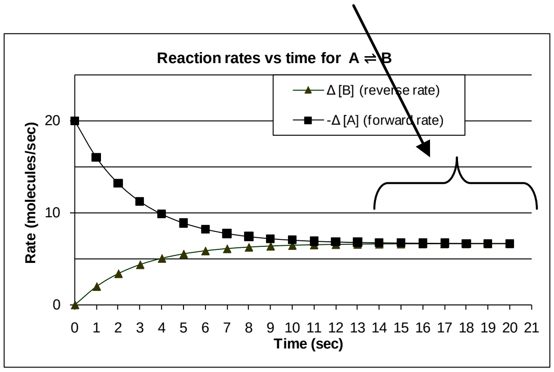
$$K_{eq} = \frac{[B]}{[A]} \quad \text{(Equation 5)}$$

This is the value displayed beneath the table on the Excel worksheet. You can verify this be doing a couple of sample calculations using data points within the equilibrium region.

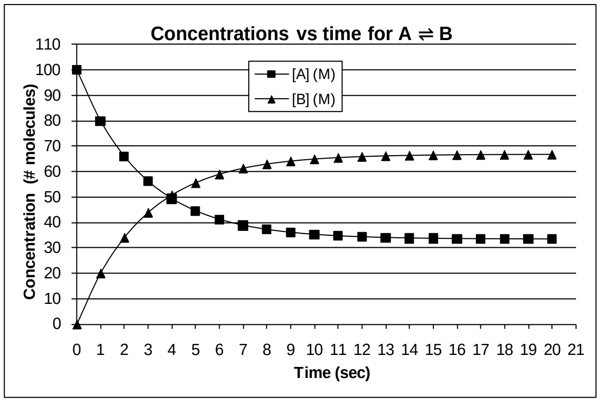
j) Identify the region on the plots at which the reaction A ⇌ B is in equilibrium.



The area where concentrations and rates become constant is the equilibrium region



k) Can you calculate the equilibrium constant for this reaction using concentration data from the table or plot? How does this value compare to the value shown on the Excel screen



(bottom left)? From which part of the plot would you take these data? Why?

At t=20 seconds, [A] ~ 33 seconds and [B] ~ 66 seconds. Thus K_{eq} *for this reaction is*

$$K_{eq} = \frac{[B]}{[A]} \approx \frac{66}{33} \approx 2.0$$

This is the number displayed beneath the table.

l) Try taking data from different parts of the plot and using them to calculate equilibrium constants. What do you notice? How do you explain your results?

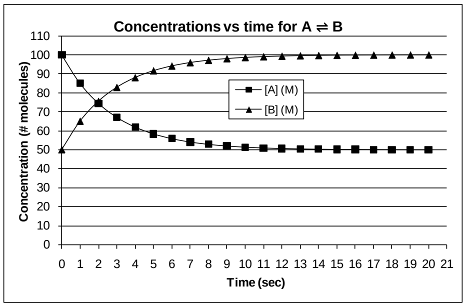
If you use data from outside the equilibrium region to calculate the equilibrium constant, the first thing you will notice is that you will not obtain the value displayed (2.0). This is because, in the area where your data is acquired, the reaction has not yet reached a stable, constant state, and the values of the concentrations in solution are still changing. As a result, at each different point outside the equilibrium region, you will get a different value.

Given this input, one might reconsider the form of the equilibrium equation used above (Equation 4). This expression is often written as

$$K_{eq} = \frac{[C]_{eq}{}^{c}[D]_{eq}{}^{d}}{[A]_{eq}{}^{a}[B]_{eq}{}^{b}}$$

to indicate that only concentration values obtained in the equilibrium region are suitable for use in the calculation of the equilibrium constant K_{eq}*.*

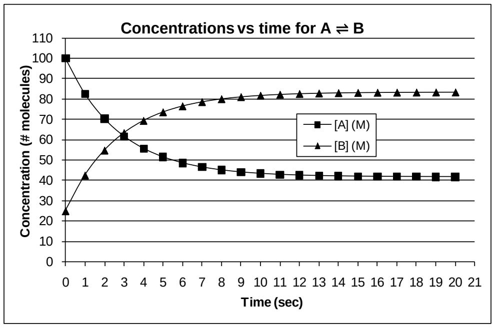
m) Change the initial concentration of B to 50. What happens to the equilibrium values of [A] and [B] (change/remain constant/increase or decrease)? Recalculate the equilibrium constant from the plot. Now change [A] to 25 and recalculate Keq. What happens to the equilibrium constant? Can you explain this phenomenon?



When the initial value of [B] is set to 50M, the equilibrium values of both [A] and [B]are increased. Choosing a new point within the equilibrium region, we can recalculate K_{eq} *as*

$$K_{eq} = \frac{[B]}{[A]} \approx \frac{100}{50} \approx 2.0$$

With [B]initial = 25M,



$$K_{eq} = \frac{[B]}{[A]} \approx \frac{83}{42} \approx 2$$

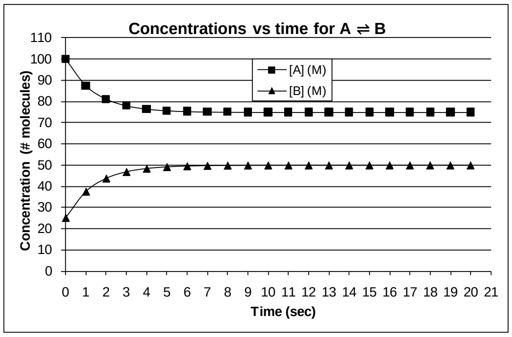
Again, the value of the equilibrium constant does not change with a change in reactant or product concentrations. Instead, the concentrations of other substances in the reaction must change to readjust the Keq equation to its constant value.

To review the three equilibrium constant equations just calculated;

$$1)\quad K_{eq} = \frac{[B]}{[A]} \approx \frac{66}{33} \approx 2.0 \qquad 2)\quad K_{eq} = \frac{[B]}{[A]} \approx \frac{100}{50} \approx 2.0 \quad 3)\quad K_{eq} = \frac{[B]}{[A]} \approx \frac{83}{42} \approx 2.0$$

n) Change the reverse rate constant to 0.3. What happens to the concentration plots? What is the key difference between this plot (with $k_f = 0.2$ and $k_r = 0.3$) and the previous one (with $k_f = 0.2$ and $k_r = 0.1$)? Can you relate the relative size of the rate constants to the relative concentration values in the plot?

At this point, the plot should be showing $[A]_0 = 100$, $[B]_0 = 25$, $k_f = 0.2$, *and* $k_r = 0.3$.



The reaction plotted above is often referred to as a "reactant-favored reaction", as the equilibrium solution contains more of the reactant (in this case, B) than it does of product (A). The reverse rate constant (now set to 0.3) is now larger than the forward rate constant; as a result, when equilibrium is reached, the reaction has proceeded further to the left than to the right. One might write the reaction as follows

$$A \rightleftharpoons B$$

to depict the fact that the larger reverse rate constant is driving the reaction more heavily to the left.

Exercise 3: Stoichiometry and Equilibrium

When working with equilibrium problems, using a correct and balanced equation, and understanding the stoichiometric relationships between the reactants and products, is critical to obtaining a correct solution.

The equilibrium expression for the chemical reaction

$$A + B \rightleftharpoons AB$$

can be written as

$$K_{eq} = \frac{[AB]}{[A][B]}$$

o) Suppose you started this reaction by creating a solution which contained equal 1M concentrations of A and B ($[A]_0 = [B]_0 = 1M$), and that after the reaction reached the equilibrium state, you measured the concentration of Aas 0.5M. What would the equilibrium concentration of Bbe? Of AB?

(Hint: Start by writing how much A was *consumed*, and use the stoichiometric relationships between A, B and AB as defined by the equation to proceed to the solution.)

The objective of this problem is to obtain the final quantities of all substances in the equilibrium state by relating their consumption and generation to one another via their stoichiometric relation. In this case, note that all substances in the reaction have a coefficient of 1; they are all stoichiometrically related in 1:1 ratios.

0.5M of reactant A is measured when the reaction is complete (at equilibrium). This means that 0.5M of reactant A has been consumed; 1M (initial concentration) - 0.5M (final concentration) = 0.5M. Since 0.5M of A was consumed, by stoichiometry, 0.5M of B must also have been consumed (you need an equal amount of A and B to make AB!) Thus the final concentration of B must also be 0.5M.

Since 0.5M of A was consumed, by stoichiometry, the same quantity of AB must have been produced. (For every A consumed, one molecule of AB is created). Thus the final volume of AB must also be 0.5M.

Note that we could have made this final deduction by relating the concentrations of B and AB, which are also related in a 1:1 manner. This would have also yielded the correct result.
A mistake often made in the solution of this problem is, since AB contains both A ***and*** *B, to add the quantities of A and B when finding the concentration of AB; "Since AB is made of A plus B, then to calculate [AB], one should add [A] plus [B]". This will result in an incorrect answer of 0.5M + 0.5M = 1.0M.*

In summary; the final values in solution at equilibrium for this reaction are

$$[A] = [B] = [AB] = 0.5M$$

p) Now that you know all of the equilibrium concentrations, calculate K_{eq} for the reaction.

$$K_{eq} = \frac{[AB]}{[A][B]} = \frac{[0.5]}{[0.5][0.5]} = 2$$

q) Suppose that you started this reaction with equal but unknown concentrations of A and B, and at equilibrium the concentration of AB was 0.6M. Given the K_{eq} you found in the previous question, what were the starting concentrations of A and B?

Here we go through a similar thought process to solve. If the final concentration of AB was 0.6M, and since it has a 1:1 stoichiometry ratio with both A and B, then the starting concentrations of A and B must have each lost 0.6M. (The generation of a quantity of AB requires the consumption of the same quantity of A and B). Since the starting concentrations of A and B are unknowns, and are equal by definition, we can call them "x", and rewrite the equilibrium equation as follows.

$$K_{eq} = \frac{[\mathrm{AB}]}{[\mathrm{A}][\mathrm{B}]} = \frac{0.6}{(x-0.6)(x-0.6)} = 2$$

Note that, since we calculated the equilibrium constant as K_{eq}=0.5 in the previous question and that it remains constant as we change concentrations, we can reuse it here.

This equation may be solved using the quadratic formula, which will be taught in detail in Chem 1B lecture. (Two answers are returned from the quadratic formula in this case; 1.1478 and 0.0523. The latter is impossible, so the answer is 1.1478.)

Revision History

Rev 4 1/21/09
- First complete draft

Rev 5 2/4/10
- Added Learning Objectives
- Simplified exercise 1
- Streamline Exercise 2

Rev 6 2/5/10
- Added drawings (Ex 2 question d)

Rev 7 2/7/10
- Changed EM table
- Modified LO's
- Revised questions in Exercise 2. Removed work with changing reaction rates and equilibrium constants; added work on relation between reaction rates and equilibrium constant.

Rev 8 2/8/10
- xxx

Rev 9 3/01/10

- Removed question j; it's redundant
- Typos
- General rewording of questions, text and diagrams for better clarity

Rev 12 7/13/10
- Standardized equation symbols
- Moved drawing question to Ex. 2
- General grammatical clarifications
- Deleted old question l (see rev 11)
- THIS REVISION IS FULLY EDITED FOR FOCUS GROUP USE

Rev 13 7/15/10
- Eliminated redundancy in rate equations in Ex 1 and 2
- Moved drawing question from Ex 2 to Ex 1
- THIS REVISION IS FULLY EDITED FOR FOCUS GROUP USE

Rev 14 7/18/10
- Added, relocated and clarified questions
- Condensed verbiage in Notes sections

Rev 15 7/18/10
- Added Exercise 3
- Streamlined Exercise 2 (removed questions)

Rev 16 8/9/10
- Clarified question d in Ex 1
- Added question on reactant-favored reactions in Ex 2

Rev 17 8/25/10
- Edited for posting

Worksheet 3: Concepts of Acid/Base Equilibrium

(Posted Version w/answer key)

For questions and general help contact Dave Swanson at dcswanson@sbcglobal.net.

Learning Objectives

- Students will be able to describe the difference between a strong and weak acid or base
- Students will be able to articulate the conceptual and mathematical relationships between pH, K_a and K_{eq}
- Students will be able to describe the solution of the equation for K_a given initial conditions

Exercise 1: Strong vs. Weak Acids and Bases

In Chem 1A, the study of acids and bases was confined primarily to "strong" acids and bases. As you might recall, the terms "strong" and "weak" have specific meanings when applied to acids and bases. The term "strong" implies that the dissociation of the acid or base in water goes to virtual completion; for example,

$$\text{Strong acid: } HCl + H_2O \rightarrow H_3O^+ + Cl^-$$
$$\text{Strong base: } NAOH \rightarrow Na^+ + OH^-$$

Notice that these reactions are written as ***irreversible reactions***.

a) Write the reaction of the strong acid HNO_3 with water. What are the reactants and products of this reaction? Is this an irreversible or a reversible reaction? If 0.1 mol of HNO_3 is added to beaker of water, how many moles of each of the products are produced?

$$HNO_3 + H_2O => H_3O^+ + NO_3^-$$

HNO_3 and H_2O are the reactants; H_3O^+ + NO_3^- are the products. This is a one way reaction, as the definition of a strong acid is that it dissociates completely in water. By stoichiometry, all reactants and products are related in a 1:1 ratio, so if 0.1 moles of HNO_3 are consumed, the same amount of products are produced. Note that this is NOT true of a weak acid.

b) In General Chemistry, ***there are only six acids which are considered to be "strong"; all the rest are weak.*** What are these six acids?

Nitric acid (HNO_3)
Sulfuric acid (H_2SO_4)

Percholoric acid ($HClO_4$)
Hydrochloric acid (HCl)
Hydrobromic acid (HBr)
Hydroiodic acid (HI)

As opposed to a strong acid or base, ***aweak acid or base will not completely dissociate*** in aqueous solution. Only a portion of the acid (and ***often a very small portion***) will be ionized, leaving a (typically large) quantity of the non-ionized acid remaining in solution in molecular form. Accordingly, the ***reaction of a weak acid or base in aqueous solution is written as a reversible reaction.*** For example***:***

Weak Acid: $HF + H_2O \rightleftharpoons H_3O^+ + F^-$
Weak Base: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Exercise 2: Aqueous Equilibria of Acids and Bases

The basic principles of acid/base equilibria are identical to those of general chemical equilibria. Because the topic of acid/base equilibrium is so central to the science of chemistry, it has its own unique terminology.

Since the reactions of weak acids and bases do not go to completion, they can be written as reversible reactions. Accordingly, equilibrium constants can be written for their dissociation in water. For a weak acid in water ($HA + H_2O \rightleftharpoons H_3O^+ + A^-$),

$$K_{eq} = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

For acid dissociation, this equilibrium is called K_a, and since H_2O is the solvent for this reaction of aqueous reactants, it is left out of the equation. For example, for the dissociation in water of the weak acid HCN,

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

For the dissociation of a weak base in water ($B + H_2O \rightleftharpoons BH^+ + OH^-$), the term K_b is used:

$$K_b = \frac{[BH^+]\,[OH^-]}{[B]}$$

For example, for the weak base $NH_{3,}$ $NH_3 + H_2O \rightleftharpoons NH_4 + OH^-$, and

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

a) Write the expression for K_a for the dissociation of the weak acid HF in water.

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

b) K_a for the weak acid HF is $7.2*10^{-6}$, and K_a for HCN is $4.0*10^{-10}$. Which will have a higher $[H_3O^+]$? Explain in terms of the above equations.

The value of $[H_3O^+]$ appears in the numerator of the expression for K_a. Since HF has the higher value of K_a, according to the formula above, HF will have the larger number of hydronium ions, or a higher $[H_3O^+]$.

c) Two solutions of the weak acid HF are found to have pH readings of 2.7 and 3.0. Which of these solutions has the highest equilibrium concentration of HF?

pH is inversely proportional to the value of the hydrogen ion concentration; the lower the pH, the higher the concentration of H^+. Since the value of K_a is constant for a given reaction, if we increase the value of one of the numerators (in this case, $[H^+]$), the value of the denominator ([HF]) must also increase to preserve the value of Ka. Thus if the pH 2.7 solution of HF has a higher value of [H+], it must also have a higher value of [HF].

Pure water can be said to undergo a dissociation reaction, forming hydronium and hydroxide ions.

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

The dissociation of water thus has its own equilibrium constant. Because this constant is uniquely important to acid/base chemistry, it is given its own symbol, $\mathbf{K_w}$. Its value at 25°C is

$$K_w = [H_3O^+][OH^-] = 1.0 * 10^{-14}$$

d) What are the concentrations of hydronium (H_3O^+) and hydroxyl (OH^-) ions in pure water?

Since these two values must be equal at equilibrium, their value is

$$\text{Let } [H_3O^+] = [OH^-] = x;$$
$$[H_3O^+][OH^-] = x^2 = K_w = 1.0 * 10^{-14};\ x = \mathbf{1.0 * 10^{-7}}$$

e) What is the pH of pure water?

pH = -log $[H_3O^+]$ *= -log (1.0 * 10*$^{-7}$ *) = 7*

f) A solution is found to have a pH of 3.0. What is the $[OH^-]$ for this solution?

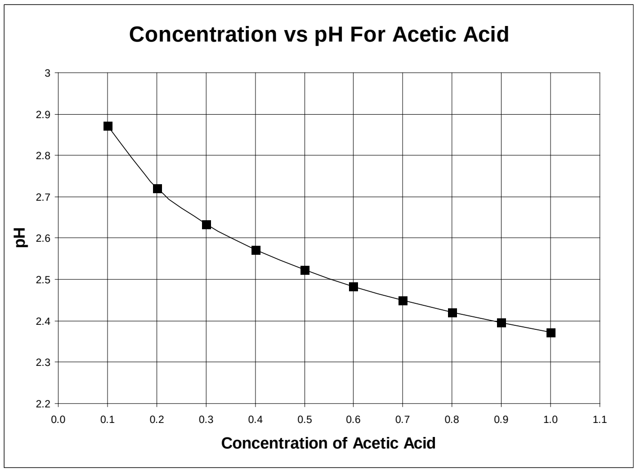
*If pH = 3, then [H+] = exp(-pH), or 1.0 * 10*$^{-3}$*. Substituting this into the equation for* K_w*, we get*

$$K_w = [H_3O^+][OH^-]$$
$$[OH^-] = \frac{Ka}{[H_3O^+]} = \frac{1*10^{-14}}{1*10^{-3}} = 1*10^{-11}$$

Exercise 3: Relationship Between Ka, pH, and Concentration

In describing the properties of an acidic solution, the terms concentration, pH, and K_a are all interrelated. This relationship can be observed in a plot of these parameters for any weak acid. Below is a plot of pH vs. concentration for solutions of acetic acid at equilibrium:

$$HCH_3COO + H_2O \rightleftharpoons H_3O^+ + CH_3COO^- \text{(Reaction 1)}$$



m) Write the expression for K_a of acetic acid.

$$K_a = \frac{[H_3O^+][CH_3CO_2^-]}{[HCH_3CO_2]}$$

a) Write the general expression for pH. What does pH measure?

pH = -log [H^+]. pH is a measurement of hydrogen ion concentration, or acidity.

b) From the plot above, what is [H_3O^+] when the concentration of acetic acid is 0.9M (best approximation)?

*From the plot, when [H_3O^+] = 0.9M, pH ~ 2.4. Thus [H_3O^+] =10exp (-pH), or about 4.0 * 10^{-3}.*

c) Describe the relationship between the data on the two axes of the above plot. Why does this curve slope down to the right?

The value of [HCH_3CO_2] is increasing as we go from left to right. Since [HCH_3CO_2] and [H_3O^+] are related in the K_a expression above, [H_3O^+] will increase accordingly, resulting in a decreasing pH, as shown on the plot.

d) Write the equation for etc. (removed)

e) Using data from the plot, calculate your best estimate for the value of K_a of a **0**.2M solution of acetic acid. (Use a value of pH to a precision of 0.01 from the plot above.)

From the plot, at 0.2M, the pH of the solution is ~ 2.72; thus $[H_3O^+]$ *is 10exp (-2.72) = 1.91*E^{-3} *moles/L at the equilibrium point. To calculate the value of Ka at this point, we use the equilibrium expression;*

$$K_a = \frac{[H_3O^+][CH_3CO_2^-]}{[HCH_3CO_2]}$$

We know the value of $[H_3O^+]$*, and from the stoichiometry of Reaction 1,* $[H_3O^+] = [CH_3CO_2^-]$*, so the expression can be written as*

$$K_a = \frac{(1.91E^{-3})^2}{[HCH_3CO_2]}$$

The value of $[HCH_3CO_2]$ *can be calculated from the plot value. The equilibrium value of the hydronium ion concentration (1.91*E^{-13} *moles/L) generated must be equal to the amount of acetic acid consumed, so the equilibrium value of the acetic acid is just the initial value (from the plot) minus 1.91*E^{-13} *moles/L. This latter amount is insignificant relative to our starting value, and so we can just plug in the initial value to get our estimate for* K_a*;*

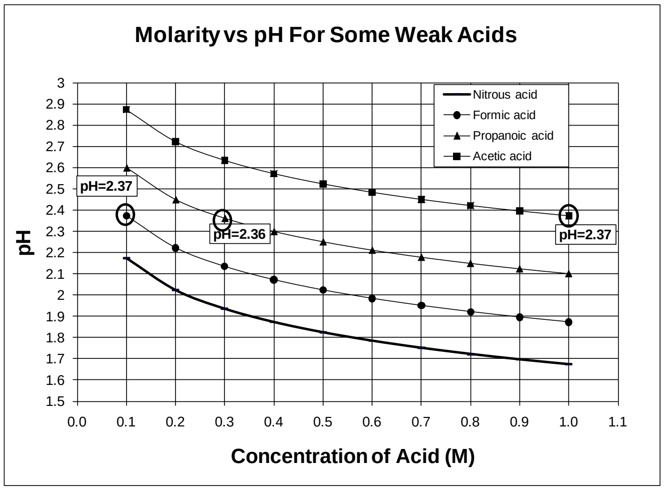
$$K_a = \frac{(1.91E^{-3})^2}{0.2} = 1.82E^{-5}$$

Since the numerator has units of M^2 *and the denominator has units of M,* K_a *has units of M.*

f) What would you expect if you were to calculate K_a at M = 0.5? Would your calculated value be lower, higher, or equal to your calculation in the previous question? Explain your answer.

Since we're looking at a plot of equilibrium values, the value of K_a *at any point should be the same.*

The following questions refer to the plot below, which shows the relationship between pH and acid concentration for several weak acids.



g) Which is the weakest acid on the above chart? Which is the strongest? Which will have the highest and lowest K_a? Why?

On any given vertical line, all acids on the plot have the same concentration. This means that, in comparing the equilibrium (K_a) expressions of these acids,

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

all of their denominators ([HA]) will be the same. Accordingly, the relative K_a values for all acids will depend on the equilibrium values of their hydronium ion concentration, which is represented by the value of pH. The weakest acid is the one with the lowest $[H_3O^+]$ (highest pH) at any given concentration, acetic acid. This acid will have the lowest dissociation at that point. Conversely, the strongest acid will have the lowest pH: nitrous acid.

h) At a concentration of 0.5M, which acid will have the highest concentration of hydrogen ions? The lowest? Why?

A similar line of reasoning can be used here. The acid with the highest $[H_3O^+]$ at any point will be that with the lowest pH, nitrous acid; acetic acid will have the lowest $[H_3O^+]$ and the highest pH.

i) For each acid, why does the pH decrease as the concentration increases?

Inspection of the dissociation reaction for acids, $HA + H_2O \rightleftharpoons H_3O^+ + A^-$*, shows that any additional HA added to the solution will react with water, producing additional* H_3O^+ *as product and decreasing the pH of the solution.*

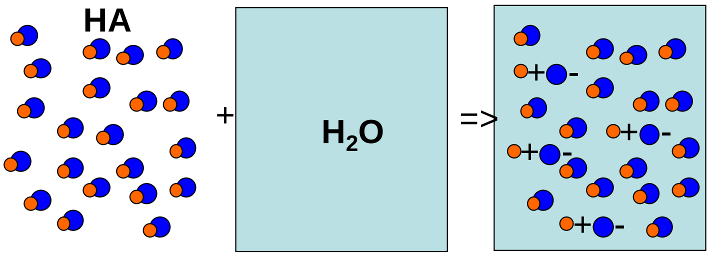
j) Note the three circled data points on the plot above; they signify points at which the pH values for three different acids are approximately equal. Why are the concentrations at these three points so different?

Simplistically stated, the concentration of hydronium ion in a solution (or the value of pH) is a function of two things; 1) the amount of acid HA initially added to the solution, and 2) the percentage of that acid which dissociates into its ionic components H_3O^+ *and* A^-*. The value of* K_a *for an acid is essentially an indication of the* ***percentage of acid which dissociates****.*

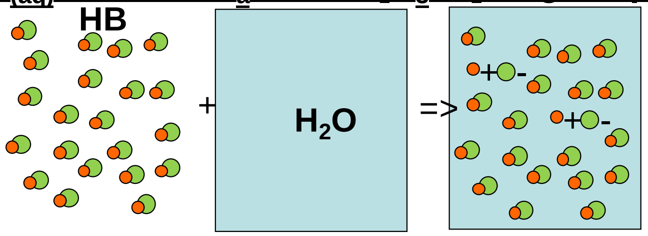
Compare the behavior of two acidic solutions HA and HB with the same initial concentrations but with different values of K_a*. HB, with a lower* K_a*, will show a lower percentage of dissociation, resulting in fewer ions in solution and a higher pH.*

The only way to increase number of ions in the HB solution is to ***add more acid****. A percentage of this new acid will dissociate and result in a higher ion concentration and a reduced pH. Eventually the HB solution will have the same quantity of ions as HA; it just requires a larger initial quantity of HB to produce them. At this point, the two solutions have the same pH, but [HB] is much larger than [HA]. This concept is illustrated on the plot, and in the graphic below; the* ***weaker acids (with lower*** K_a ***values) will need a higher initial acid concentration to attain the same pH as a stronger acid.***

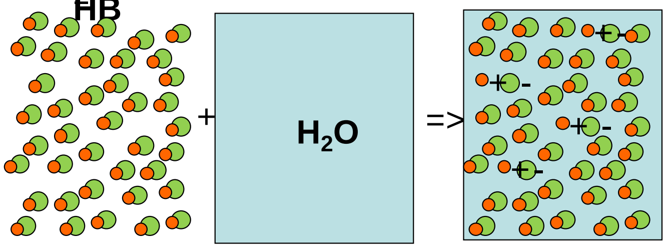
$HA_{(aq)}$: Higher K_a, higher $[H_3O^+]$, lower pH



$HB_{(aq)}$: Lower K_a, lower $[H_3O^+]$, higher pH



$HB_{(aq)}$: $[H_3O^+]$ raised by increasing concentration



k) Calculate the value of K_a for each of the three weak acids in the question above. Do they values you calculate make sense in terms of your answers to the above questions?

The process for obtaining this solution is the same as that used in question e above. Since all three calculations are the same, we can just use a generic form of the equilibrium expression:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

In all cases, $[H_3O^+]$ is equal to 10exp (-2.37) = 4.27 E^{-3} moles/L at the equilibrium point.

To calculate the value of Ka at this point, we use the equilibrium expression. As in question e, this brings us to an equation of the form

$$K_a = \frac{(4.27E^{-3})^2}{[HA]}$$

The initial value of [HA] can be simply taken from the plot, as the reduction of this value at the equilibrium point is in all cases insignificant. Thus for each of the acids above, we need just plug the value of [HA] into the above expression.

1) 0.1M Formic acid @ pH = 2.37

$$K_a = \frac{(4.27E^{-3})^2}{0.1} = 1.82E^{-4}$$

2) 0.3M Propanoic acid @ pH = 2.37

$$K_a = \frac{(4.27E^{-3})^2}{0.3} = 6.07E^{-5}$$

3) 1.0M Acetic acid @ pH = 2.37

$$K_a = \frac{(4.27E^{-3})^2}{1.0} = 1.82E^{-5}$$

Revision History

Rev 3 2/11/10
- Added Learning Objectives
- Streamlined Exercise 2
    - Removed details and non-conceptual questions
    - Added drawing questions
    - Added single-curve plot to work on understanding of curve

Rev 4 3/01/10
- Major overhaul; corrected equations, etc.

Rev 6 7/14/10
- Changed single-line plot and following question
- Minor question clarifications and typo fixes

Rev 7 7/19/10
- General question clarifications and typo fixes

Rev 8 7/28/10
- General question clarifications

Rev 9 7/29/10
- Added water chemistry questions

Rev 10 8/14/10

- Removed redundant water question
- Clarified plot questions

Rev 11 9/22/10

- Clean up for posting

Worksheet 4: The Common Ion Effect and Buffer Chemistry

For questions and general help contact Dave Swanson at dcswanson@sbcglobal.net.

Learning Objectives

- Students will be able to name the conjugate of an acid or a base
- Students will be able to describe the Common Ion Effect
- Students will be able to articulate the purpose and contents of a buffer solution
- Students will be able to describe the fundamental workings of a buffer solution

Exercise 1: Buffer Chemistry: The Conjugate Pair and The Common Ion Effect

As reviewed in the previous worksheet, the dissociation of weak acids is subject to an equilibrium constant K_a. For the weak acid HF,

$$\mathrm{HF} + \mathrm{H_2O} \rightleftharpoons \mathrm{H_3O^+} + \mathrm{F^-} \quad \text{(Reaction 1)}$$

$$K_a = \frac{[\mathrm{H_3O^+}][\mathrm{F^-}]}{[\mathrm{HF}]}$$

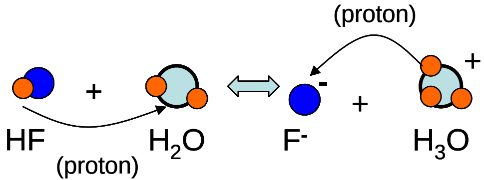
a) Write the equation for the dissociation of the weak acid HCN in water.

$$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$$

b) Write the expression for the K_a of HCN.

$$Ka = \frac{[H_3O^+][CN^-]}{[HCN]}$$

Note the dissociation equation for HF above (Reaction 1). This reversible reaction can be seen as a simple proton exchange. In the forward direction, the weak acid HF donates a proton to H_2O, forming the hydronium ion H_3O^+; in the reverse direction, the weak base F^- takes a proton from a water molecule, reforming the acid HF. HF and F^- are thus just partners in a proton exchange process. An acid/base pair with this relationship is called a *conjugate pair*. F^- is the "conjugate base" of HF, and HF is the "conjugate acid" of F^-.



c) What is the conjugate base of HCN? Of HNO_2?

From the reaction above, the conjugate base of HCN is CN^-. *The conjugate base of* HNO_2 *is* NO_2^-.

d) K_a for HCN is $6.17 * 10^{-10}$. Given this figure, what can you say about the relative amounts of the non-dissociated acid HCN and its conjugate base in the solution (approximately; i.e., roughly equal, one much larger, etc.)?

The value of K_a *is a very small number; therefore, we can expect* ***very little*** *ionized* CN^- *to appear in a solution of HCN.*

Suppose that we add a quantity of the water-soluble ionic compound potassium cyanide (KCN) to a solution of HCN.

e) What is the reaction between KCN and water? (Hint; review your solubility tables.) What new chemical species (compounds/ions) appear in solution when KCN is added?

KCN is a water-soluble compound; therefore, it will totally dissociate when added to water. The new components in solution are thus K^+ *and* CN^-.

f) Note the equilibrium expression for K_a for HCN (question b above). How has the balance of equilibrium been changed by the addition of KCN to the solution? How will this affect the expression for K_a? What is the effect of these changes on the pH of the solution?

Given that the value of K_a *is fixed for a reaction, a change in the concentration of any of the components of the equilibrium expression must be followed by an appropriate change in the concentration of another component so as to restore the ratio of components to equal the value of* K_a. *Accordingly, with increase of* $[CN^-]$ *to the numerator of the equilibrium expression in question b, the ratio of the other two compounds in the expression (*H_3O^+ *and HCN) must* ***decrease*** *in order to preserve the equilibrium ratio. Thus the value of [HCN] will increase, and the value of* $[H_3O^+]$ *will decrease slightly.*

The changes made by the addition of KCN above are called the "common ion effect". If we realize that the equilibrium constant K_a is constant, and cannot be changed, it must be the case that the increase in concentration of cyanide ion ($[CN^-]$) to a solution forces a change in the concentration of other solution components so as to bring the solution back to equilibrium.

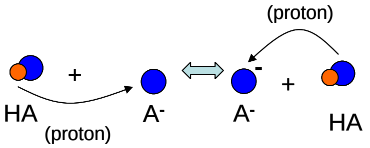
Exercise 2: Introduction to Buffer Solutions

A buffer solution is an aqueous mixture of chemical species which is designed to resist changes in pH when acids or bases are added to it. A buffer solution contains both a weakly acidic species (which will consume strong bases added to the solution) and a weakly basic species (which will consume strong acids).

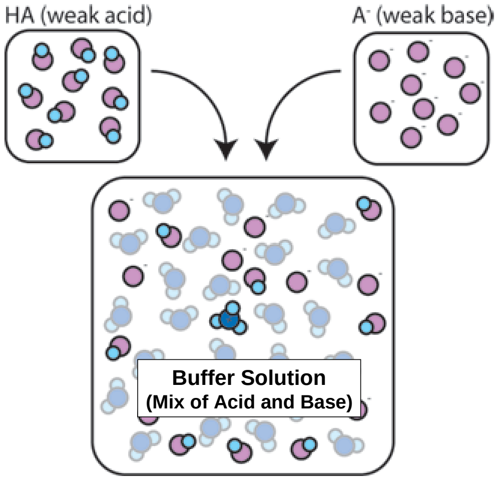
a) Buffer solutions are very important in the field of medicine. Can you think of a good application for a buffer solution?

Bodily fluids must typically maintain a narrow range of pH values in order to successfully carry out their biochemical functions. Blood plasma, for example, must maintain a pH range of 7.35 to 7.45 for full function.

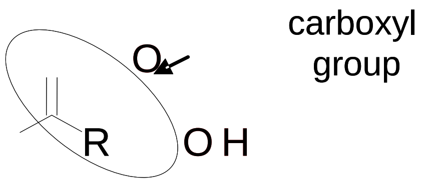
A buffer solution must be designed so that the weak acid and weak base do not react with each other; thus, buffers typically contain both a weak acid and its conjugate base. A reaction between these substances results only in a proton transfer, with has no net effect on the solution or its pH:



As a result, the weak acid and its conjugate base will coexist in the buffer solution,:



Acetic acid (HCH_3CO_2) is a weak acid commonly used in buffer solutions. Acetic acid belongs to the family of *carboxylic acids*, which contain an acidic carboxyl group with an acidic proton. (In the diagram below, for acetic acid, R = CH_3.)



b) Write the equation for the dissociation reaction for acetic acid in water. What is the conjugate base of acetic acid? Write the expression for the K_a of acetic acid.

$$HCH_3CO_2 + H_2O \rightleftharpoons H_3O^+ + CH_3\,CO_2^-$$

$CH_3\,CO_2^-$ (the acetate ion) is the conjugate base of acetic acid. The equilibrium expression for the dissociation of acetic acid is

$$K_a = \frac{[H_3O^+][CH_3CO_2^-]}{[HCH_3CO_2]}$$

c) The K_a of acetic acid is very small ($1.8 * 10^{-5}$). What does this tell you about the amount of the conjugate base which will exist in a solution of acetic acid? What must be done to form a buffer solution from a solution of acetic acid?

A look at the equilibrium expression above indicates that, with a small K_a, the concentration of ions in solution will be very low relative to the concentration of the acid.

Exercise3: Buffer Chemistry

As stated above, a buffer solution resists changes in pH by consuming and countering the effects of strong acids and bases. A brief overview of how a buffer is made and how it serves to counteract changes in pH can be seen at the link below. (If this link does not open, just double-click on the PowerPoint file posted with this worksheet.) Use the file in Presentation Mode to see the animation.

Making and Using a Buffer

The buffer solution in the PowerPoint file is made of acetic acid and sodium acetate ($NaCH_3CO_2$) and is found very commonly in practice. The questions below refer to this buffer.

a) What is the purpose of adding sodium acetate to the solution?

Sodium acetate is a soluble salt; when added to water, a quantity of the weak base $CH_3CO_2^-$ will be formed. This provides the solution with a weak acid (HCH_3CO_2) and a weak base ($CH_3CO_2^-$), the essential components of a buffer solution.

b) List all the compounds and ions which will be present in the solution at equilibrium. Which of these are relevant to the key function of the buffer? Which are not? Why?

HCH_3CO_2, $CH_3CO_2^-$, H_3O^+, and Na^+ will all appear in solution. The sodium ion is essentially irrelevant to the buffer function.

Equal volumes of 1M acetic acid and 1M sodium acetate are mixed.

c) Roughly, what are the relative volumes of the compounds in solution (see question k above)? (Equal, greater, much greater, etc.)

Similar quantities of HCH_3CO_2 and $CH_3CO_2^-$, will exist in solution, along with a relatively low concentration of H_3O. There will be a quantity of Na^+ roughly equal to that of the acetate ion, but which is irrelevant to the chemistry.

d) Is this a buffer solution? Why or why not?

This is a buffer solution, as it contains quantities of both a weak acid and its conjugate base.

e) Suppose that a quantity of the strong acid HNO_3 is added to this solution. Which of the compounds in the solution will react with this acid? Write the equation for this reaction.

The addition of a strong acid will initiate a reaction with the base in the solution, the acetate ion $CH_3CO_2^-$;

$$HNO_3 + CH_3CO_2^- \rightleftharpoons HCH_3CO_2 + NO_3^-$$
$$\text{or } H^+ + CH_3CO_2^- \rightleftharpoons HCH_3CO_2$$

f) Given the reaction equation you have written above, can you draw any conclusions about which way the reaction will progress (e.g., towards the right, or to products, or to the left, or reactants)?

This equation contains a strong acid (HNO_3), which undergoes a strong dissociation reaction in water, and will push the reaction sharply to the right.

g) Suppose that a quantity of the strong base NaOH is added to this buffer. What species in the in the solution will these ions react with? Write the equation for this reaction.

NaOH dissociates in water to create the strong base OH^-, which will react with the acid in the solution, acetic acid;

$$HCH_3CO_2 + OH^- \rightleftharpoons CH_3CO_2^- + H_2O$$

h) Given the reaction equation you have written above, can you draw any conclusions about which way the reaction will progress (e.g., towards the right, or to products, or to the left, or reactants)?

NaOH, as a strong proton acceptor, will be wholly consumed as it adds protons to generate water, moving the reaction strongly to the right.

Worksheet rev 7 10/1/10
Answer key rev 1 10/10/10

Revision History

Rev 0 7/19/10

Rev 1 7/20/10
- Added questions, changed order of questions
- Added Ex. 2

Rev 2 7/19/10
- More questions, revisions
- Added CIE exercise

Rev 3 7/25/10
- Rearranged Ex. 2 and 3
- More CIE questions

Rev 4 7/28/10
- Question rearrangement and clarification.

Rev 4/5 8/12/10
- Added buffer animation
- Added text to Ex 1
- Miscellaneous question clarification

Rev 6 8/15/10
- Added conjugate pair infor/q's to ex. 1
- Sharpened up notes, questions

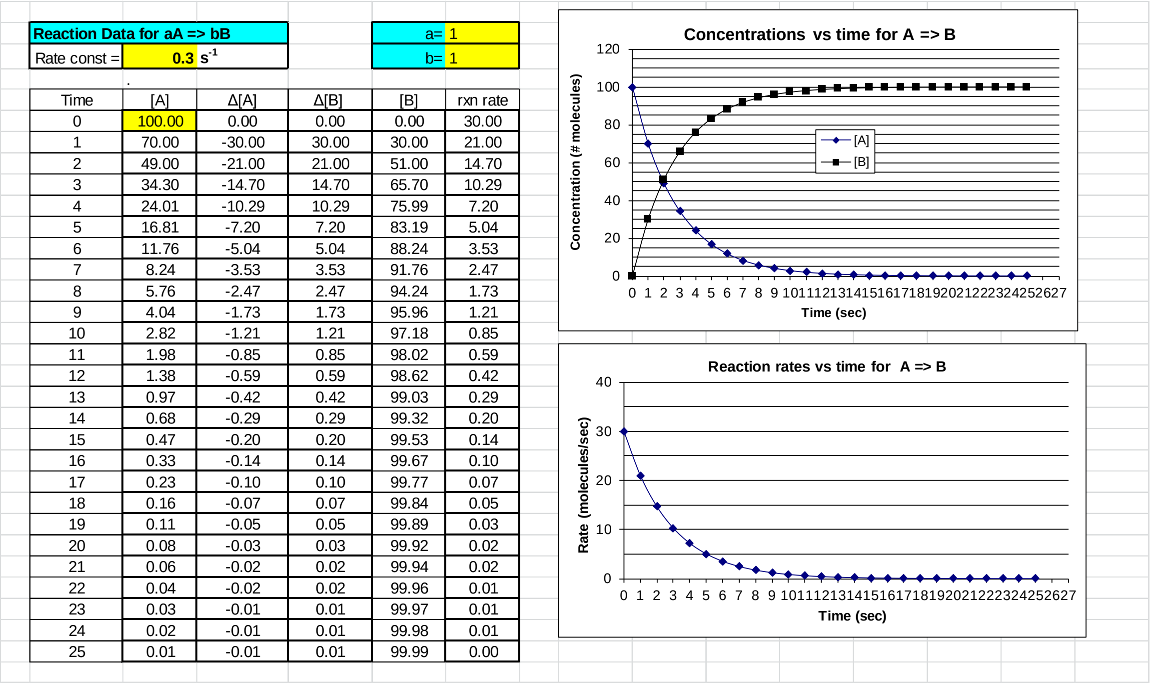
Rev 7 10/1/10

- Edited for posting

APPENDIX C

The Equilibrium Machine

The Equilibrium Machine (EM) is a workbook showing the combination of spreadsheets, which calculate concentrations and reaction rates vs. time for the reactions A => B and A ⇌ B, and plots which display these data. The worksheets are protected, with several cells accessible to allow the input of initial conditions; starting concentrations and reactant rate constants. Two sheets are included in the EM. The first displays data for irreversible reactions, and allows inputs for initial reactant concentrations and the forward rate constant; it also allows input for reactant and product coefficients in the reaction equation. A sample of this sheet can be seen in **Figure 23**.



Reaction Data for aA => bB

Rate const = **0.3** s^{-1}

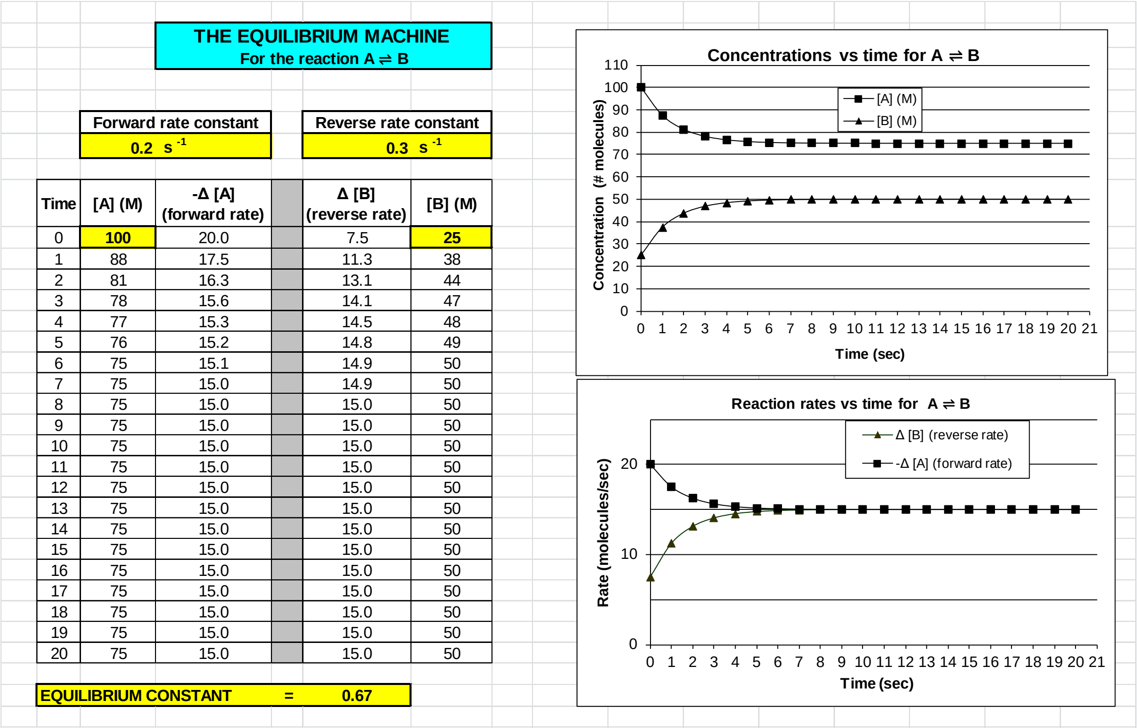
a= 1

b= 1

Time	[A]	Δ[A]	Δ[B]	[B]	rxn rate
0	100.00	0.00	0.00	0.00	30.00
1	70.00	-30.00	30.00	30.00	21.00
2	49.00	-21.00	21.00	51.00	14.70
3	34.30	-14.70	14.70	65.70	10.29
4	24.01	-10.29	10.29	75.99	7.20
5	16.81	-7.20	7.20	83.19	5.04
6	11.76	-5.04	5.04	88.24	3.53
7	8.24	-3.53	3.53	91.76	2.47
8	5.76	-2.47	2.47	94.24	1.73
9	4.04	-1.73	1.73	95.96	1.21
10	2.82	-1.21	1.21	97.18	0.85
11	1.98	-0.85	0.85	98.02	0.59
12	1.38	-0.59	0.59	98.62	0.42
13	0.97	-0.42	0.42	99.03	0.29
14	0.68	-0.29	0.29	99.32	0.20
15	0.47	-0.20	0.20	99.53	0.14
16	0.33	-0.14	0.14	99.67	0.10
17	0.23	-0.10	0.10	99.77	0.07
18	0.16	-0.07	0.07	99.84	0.05
19	0.11	-0.05	0.05	99.89	0.03
20	0.08	-0.03	0.03	99.92	0.02
21	0.06	-0.02	0.02	99.94	0.02
22	0.04	-0.02	0.02	99.96	0.01
23	0.03	-0.01	0.01	99.97	0.01
24	0.02	-0.01	0.01	99.98	0.01
25	0.01	-0.01	0.01	99.99	0.00

Figure 23: Sample Sheet from Equilibrium Machine for A=> B

The second sheet displays data for reversible reactions, and allows inputs of both reactant and product concentrations, and of rate constants for both forward and reverse reactions. Examples of both types of worksheets are displayed in **Figure 24** below.



THE EQUILIBRIUM MACHINE
For the reaction A ⇌ B

Forward rate constant	Reverse rate constant
0.2 s^{-1}	0.3 s^{-1}

Time	[A] (M)	-Δ [A] (forward rate)		Δ [B] (reverse rate)	[B] (M)
0	100	20.0		7.5	25
1	88	17.5		11.3	38
2	81	16.3		13.1	44
3	78	15.6		14.1	47
4	77	15.3		14.5	48
5	76	15.2		14.8	49
6	75	15.1		14.9	50
7	75	15.0		14.9	50
8	75	15.0		15.0	50
9	75	15.0		15.0	50
10	75	15.0		15.0	50
11	75	15.0		15.0	50
12	75	15.0		15.0	50
13	75	15.0		15.0	50
14	75	15.0		15.0	50
15	75	15.0		15.0	50
16	75	15.0		15.0	50
17	75	15.0		15.0	50
18	75	15.0		15.0	50
19	75	15.0		15.0	50
20	75	15.0		15.0	50

EQUILIBRIUM CONSTANT = 0.67

Figure 24: Sample Sheet from Equilibrium Machine for A ⇌ B

Digital copies of the Equilibrium Machine can be found on the CD attached to this paper.

APPENDIX D

Focus Group Transcripts

Worksheet 1 Focus Group Transcripts

S1: Session 1 (1/29/10)

Worksheet 1 (Reaction Kinetics and Stoichiometry r5) with Equilibrium Machine rev 6

NO RECORDING
Notes: Notebook 1 pg. 12

S1 is a Senior Biology major. He had taken Chem 1B one year ago, in Spring 2009, and received a D. S1 had no high school chemistry, and had received a B in Chem 4 and a B in Chem 1A.

Worksheet Review

Exercise 1 (Stoichiometry vs. Reaction Rate)

This exercise was very effective in correctingS1's misconception. After completion of this exercise,S1 was able to rethink and correct the questions he missed on the Diagnostic Exam (above) relative to the relation between stoichiometry and reaction rate.

Exercise 2 (Reaction Rates vs. Time)

This only required completion of a portion of the table to be effective. The exercise was effective in conveying how rates of reaction are dependent on component concentrations, which are in turn related to the time of reaction, this making reaction rate a function of time. S1's comment: "I always thought that 'rate constant' meant that the rate was constant . . . now I see that it (reaction rate) changes; it's like the rate isn't constant, but the rate of change (of the rate) is constant". KEY LEARNING: Reaction rates vary with time, and the rate of change of the reaction rate is described by the rate constant.

Summary

S1 indicated that the worksheet and the focus group session was "more in-depth" than he had anticipated. My interpretation of this comment was that his use of the word "depth" referred in particular to the emphasis on the conceptual rather than the formulaic. At the end of the session, he seemed enthusiastic, thanked me for the help, and asked to take a copy of the worksheet home with him.

S4: Session 1 (2/2/10)

Worksheet 1 (Reaction Kinetics and Stoichiometry r7) with Equilibrium Machine rev 6

Audio file: S4 2-2-10
Notes: Notebook 1 pg 18

S4 had taken Chem 1B in Fall 2009 and received an F. Previous to this, she had received an A in high school chemistry and a C in Chem 1A in 2004. She noted that she would need to do a lot of review due to the long gap between 1A and 1B.

Worksheet Review

Exercise 1 (Stoichiometry vs. Reaction Rate)

S4 had omitted the answer to these questions on her diag exam. This exercise took some time, but as withS1 above, the exercise was eventually effective in correcting here answer. As withS1, the “aha” moment was very clear; it just took a little more leading to get S4 to that point.

Exercise 2 (Reaction Rates vs. Time)

Again, this only required completion of a portion of the table to be effective. As withS1, an “aha” moment was reached when she realized the dynamic nature of the reaction rate. With not much prompting, S4 figured out how to fill in the table of rates and concentrations; it was when she was doing this when the moment struck. A phenomenon is described often in the literature in which the student’s initiative and activity is key to attaining the aha moment, was visible; this was MY aha moment. S4 also requested a copy of the worksheet to review at home.

Summary

S4’s session was VERY similar toS1’s; just slower. A total of 40 minutes was spent 1) reviewing the concept of stoichiometry and how it affects change in concentration over the course of a reaction, and 2) the concepts of rate vs rate constant, how rates are dependent on concentration, and how rates change over time. In reviewing the voice recording, there was a LOT of time spent by the instructor introducing new concepts and guiding the student toward a better understanding of these concepts. This was necessary due to the absence of any other students at the session, but this process should really be better reviewed as a group learning event.

S4's final comments were an enthusiastic Yes as to whether this session was helpful, and some thoughts on what posed the most problems for her in Chem 1B. "A lot of times, people say, you know, you just read the book and you'll be fine, and it seems like I get it, but then it's a mistake, and then another mistake", and so on. It sounds a lot like S4 could benefit from some study skills training here (it's a lot more than reading the book!).

S3: Session 1 (2/9/10)

Worksheet 1 (Reaction Kinetics and Stoichiometry r6) with Equilibrium Machine rev 6
Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 7) with Equilibrium Machine rev 6

Recording: S32-9-10
Notes: Notebook 1 pg 22

S3 is working on a second Bachelors degree in BioChem. He is older than the typical student. He had no high chemistry training, got an A in Chem4 and a B in Chem1A. He took 1B for the first time in Spring 2010.

<u>Worksheet Review: WS1</u>

Exercise 1 (Stoichiometry)

S3missed the stoichiometry question on the Diag Exam and consistently predicted that C2 would be consumed faster than AB. After some dialogue, this problem seemed related to a misunderstanding of the question, and specifically related to the meaning of the coefficients and subscripts in the equation (2AB + C2 => etc.). When I did the problem on the white board, beginning with 10 units of AB and 10 units of C2, it became obvious to him that AB was being consumed faster. His comment was, "well sure, if you start with the same amounts . . . ". He perceived that the coefficient of 2 in front of AB meant that you had twice the amount of AB available to start with, therefore it would last longer, and thus C2 would be "consumed more quickly", as in the question. I have reworded the question and added a question as to the precise meaning of the coefficient 2 in the equation.

Exercise 2 (Reaction Rates and Concentrations)

As we went through this exercise, the objective of which was to show that rates varied with time and are always in the same proportion to concentrations, it was obvious that this was somewhat of a revelation to him, although he picked up the new concept quite quickly. We went through this exercise fairly quickly to get to the next worksheet.

S2: Session 1 (3/1/10)

Worksheet 1 (Reaction Kinetics and Stoichiometry r6) with Equilibrium Machine rev 7
Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 9) with Equilibrium Machine rev 6

Recording: S2 1
Notes: None

S2 got a B+ in High School chemistry, a B+ in Chem4 and an A in Chem1A. He took 1B for the first time in Spring 2010. S2 had not initially signed up for the focus groups, but found me in the hallway after failing the first Chem 1B exam.

Worksheet Review: WS1

Exercise 1 (Stoichiometry)

We reviewed the stoichiometry problem in Exercise 1. As mentioned above, this question is often answered incorrectly due to misunderstanding of the meaning of the reactant coefficients. The coefficient of 2 for reactant AB in the reaction

$$2AB + C_2 => 2AC + B_2$$

was interpreted to mean that we started the reaction with twice the amount of AB as of C2; thus the answer given was C2, as there was less of it and it would be "consumed more quickly". Admittedly, this misinterpretation shows a profound misunderstanding of the meaning of stoichiometric coefficients in a chemical reaction; however, to eliminate this confusion, I now begin the question by specifying that equal amounts of AB and C2 are combined to initiate the reaction. I used this qualifier with S2, and he still chose C2 as the reactant most quickly consumed.

S2's first defense of this answer was "wouldn't it be the one that's the limiter?". When it was again pointed out that we started with equal amounts of the two reactants, he again picked C2 "because it has a less amount of moles . . . you have two moles of A and two moles of B, so you should have two moles of C2, so C2 is consumed faster". When asked to verbalize the reaction, he said, "two moles of AB are combined with two moles of C to produce two moles of AC and two moles of B". After two C's and one C-C were written on the white board, and he was asked to articulate what each drawing represented, he correctly identified one as "two C's" and the other as "one C-two" He now changed his reading of the equation, but still believed that C2 was consumed faster. At this point, I went into the white board exercise described in the worksheet, drawing six AB molecules and six C2 molecules and showing how they were consumed to make AC and B2. This exercise took about eight minutes in total, but S2 finally understood the mechanics of the reaction and why the relative rates of consumption of the two reactants were different.

Exercise 2 (Reaction Rate and Concentrations)

At this point we reviewed the rate laws presented in Experiment 14 of Chem 1A. S2 quickly grasped these, but seemed to be doing so for the first time, and had never really come to terms with these concepts in Chem 1A. In listening to the voice recording, it is clear that the instructor is providing much of the input in this session, as opposed to leading the student to provide the input. This needs to be corrected. Rather than describing the meaning of a “rate”, and describing what “rate of reaction” means, the instructor should ask “What is a rate”, and “How would you measure, for example, the rate of speed of an object”, then switch to rate of reaction.)

In the interest of time, I did much of the talking in explaining the two forms of the rate constant and the difference between the two, and what exactly the meaning of the rate constant was. We then went directly to Worksheet 2 without completing the rate calculation exercise, as this exercise would be completed in WS 2 and would reasonably satisfy the goal of the WS1 exercise. This was not an ideal process, but worked reasonably well given the time limits of the session. At this point, 23 minutes of the session had elapsed.

S7 and S5: Session 1 (7/14/10, 11:00 A.M.)

Worksheet 1 (Reaction Kinetics and Stoichiometry rev11, Equilibrium Machine rev 9)

Recording: S7 and S5 WS1 7-14-10
Notes: Notebook 1 pg 33

S7 is a Junior majoring in Child Development and a possible Pre-Med major. She received a B in High School chemistry, and did not take Chem 4. She earned a D in her first semester of Chem 1A, and then repeated the course and received a B. She will be entering Chem 1B in the Fall of 2010.

S5 is a Junior majoring in Biology. She received a C+ in High School chemistry, a B in Chem 4 in Fall 2009, and a B+ in Chem 1A in Spring 2010. She will also be entering Chem 1B in the Fall of 2010.

Exercise 1 (Stoichiometry)

S7 and S5 worked together on the questions in Exercise 1; this took over seven minutes. They correctly identified the nature of the coefficient and subscript in the reaction (questions a and b), but then chose C2 as the most quickly consumed reactant. When asked to “read” the reaction as a sentence, they said, “Two moles of AB and two moles of C go to two moles of AC and two moles of B”. A quick white-board exercise followed, and S&H were asked to draw 2C and C2. They then recognized that the equation read “one molecule of C2” and changed their interpretation of the equation; however, they still held C2 to be the more quickly consumed reactant. It wasn’t until the optional whiteboard exercise was performed that they understood the dynamics of the reaction and correctly identified AB in question c.

S7 & S5 had a lot of trouble identifying the substances remaining after completion and their ratios. At this point, a whiteboard discussion on the meaning of reaction coefficients and their ratios was held, and after being asked some clarifying questions, they were able to answer the question. The whiteboard exercise (writing six of each reactant and using them to make products) was instrumental here; a quick look at the remaining substances after completion was instrumental in understanding this question. This question made a good forum for the review of how reactant and product coefficients were so important to a chemical equation, and how their ratios dictated the relative rates of consumption and generation of reactants and products over the course of a reaction.

Completion of this exercise took ~30 minutes.

Exercise 2 (Reaction Rate and Concentrations)

As the exercise begun, S7 and S5 made it clear that, while they recalled the terms “reaction rate” and “rate constant” from Chem 1A, they had no idea of what they meant. Both confirmed that, due to the complexity of the Rate lab in Chem 1A, and to the fact that the lab was held in the last few frenzied weeks of the semester, in order to successfully complete the lab, they had resorted to a rote approach to the completion of the lab calculations, following the lab directions and plugging their data into the formulae provided in order to meet the due date for the lab. As a result, their conceptual understanding of rate kinetics was entirely absent.

Accordingly, the basics of kinetics (rate laws etc.) were reviewed, based on the notes in Exercise 2 of Worksheet 1. The basic definition of the reaction rate (change in concentration per time) was reviewed as per the equation given in the 1A Rate lab:

$$\text{For aA + bB}\rightarrow\text{cC + dD, } Rate = -\frac{1}{a}\frac{\Delta A}{\Delta t} = \frac{1}{c}\frac{\Delta C}{\Delta t} \text{ (etc.)}$$

Their understanding of this formula was confirmed, and they displayed what I can only describe as a sense of relief for this understanding. The relationship between concentration and rate was then reviewed. This relationship (Rate = k[A]) was at the heart of the 1A Rate Lab, but was wholly new to S7 & S5. The concept of the rate constant was presented, and the distinction between the Rate and Rate Constant was made.

The generation of the table on page 3 was then begun. The first two rows of the table were generated by the instructor on the whiteboard, along with instructions on how each entry was generated, but they were unable to complete the table at this point. It wasn't until they completed a row themselves on the whiteboard that they were able to finish the table. Table entry was stopped after about six lines, at which time they were able to articulate their observations of the trends in concentrations and rate over time, and the exercise proceeded directly to the use of the Equilibrium Machine.

After the table was opened and initial values were entered, some time was spent studying and articulating the forms of the plots (trends, slopes, etc.). The concepts reflected in the plots were quickly grasped and related to the progress of the chemical reaction as illustrated by the table calculations earlier in this exercise. Questions h and i were not readily understood and should be clarified; when they were, the answers were quickly given. After this, the discussion wound down quickly.

The exercise was completed in ~ 80 minutes.

Summary

This session went very smoothly but still tool well over an hour. The most time-consuming sections of the session were related to the establishment of the most basic concepts; that of the different forms of expression in the form of a chemical reaction (coefficients vs. subscripts) and their influence on the progress of the reaction, and the concept of the reaction rate and the dynamic nature of the rate over time. After completion of these topics, however, articulation of the reaction plots (concentrations and rate vs. time) was very clear.

A closing comment: The opening stoichiometry question, seemingly simple and straightforward, has been answered incorrectly in all five reviews of Worksheet 1. Why is the answer here so consistently wrong? This needs to be better understood. A more thorough review of this question with S7 and HC, and in future focus group sessions, is called for here.

S8: Session 1 (7/30/10, 9:00 A.M.)

Worksheet 1 (Reaction Kinetics and Stoichiometry rev12, Equilibrium Machine rev 6)

Recording: S8 WS1 7-30-10
Notes: Notebook 1 pg 42

LR is a Junior majoring in Biology. She received a B in High School chemistry, a B- in Chem 4 in Fall 2009, and a B+ in Chem 1A in Spring 2010. She will also be entering Chem 1B in the Fall of 2010.

Exercise 1 (Stoichiometry)

S8 correctly read the equation (". . . and one mole of the compound C_2 . . . "), but then incorrectly identified C_2 as the most quickly consumed reactant. After one round of the whiteboard exercise, she recognized her error.

Exercise 2: Reaction Rates and Concentrations over Time

S8 quickly grasped the conceptual and mathematical relationships between reactant consumption and product generation. She read the notes in the worksheet before beginning the exercise, and after that, everything went pretty smoothly. The first three lines of the table were completed when she saw the trends of the concentrations and rates.

(Note to myself: I realized in this session that it was possible that, after working the first two lines of the table with the instructor, that the students might just be copying that procedure for the following rows without really understanding what they were doing. I made sure to ask S8 to explain exactly what she was doing and why; I need to continue this practice in the future.)

Exercise 3: More on Rates and Concentrations

This exercise went pretty quickly and smoothly. The most meaningful input from the exercise came as part of S8's parting comments. When asked whether or not the session was helpful to her, she enthusiastically answered positively. She made particular note of exercise three, and how the use of several different methods of simultaneously conveying the same principle (using the Excel chart, the worksheet, the whiteboard, and verbal discussion) had been a great help to her.

S6: Session 1 (9/15/10, 11:00 A.M.)

Worksheet 1 (Reaction Kinetics and Stoichiometry rev13, Equilibrium Machine rev 13)

Recording: None
Notes: None; direct to Word.

S6 is a Senior majoring in Biology. She had no High School chemistry and did not take Chem 4. She received a C in Chem 1A in Fall 2008, and is currently repeating Chem 1B after receiving a D in Fall 2009.

Exercise 1 (Stoichiometry)

S6 incorrectly read the equation (". . . and two moles of C. . ."), and incorrectly identified C_2 as the most quickly consumed reactant. When asked to distinguish between the notations 2C and C_2, she was unable to do so, saying "well, aren't they just the same thing?" It took some time to rectify her misconceptions here; I explained the difference between the two notations, asked her to articulate the difference between other similarly notated compounds, and then we went through the whiteboard exercise to address her stoichiometry issues.

Exercise 2: Reaction Rates and Concentrations over Time

This again took some time, as she expressed the same lack of understanding of reaction rates as the other interviewees. An interesting note; halfway through the process I discovered that she did not understand the meaning of the concentration brackets (i.e., [A] = concentration (or molarity) of A in moles/Liter). This added some time to the process. After this, we were able to get partway through the generation of the table before the session ended; the elapsed session time at this point was about 75 minutes, and we had covered only about 75% of the material.

S6: Session 2 (9/20/10, 10:00 A.M.)

Worksheet 1 (Reaction Kinetics and Stoichiometry rev13, Equilibrium Machine rev 13)

Recording: None
Notes: None; direct to Word.

Exercise 2: Reaction Rates and Concentrations over Time (cont.)

We completed the table begun last session with few problems.

<u>Exercise 3: More on Rates and Concentrations</u>

This went fairly quickly, with no notable issues.

Worksheet 2 Focus Group Transcripts

Part 1: Focus Group Summaries

S1: Session 2 (2/5/10)

Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 5) with Equilibrium Machine rev 6

Recording: S1D 2-5-10
Notes: Looseleaf (in Interviews folder)

Exercise 1 (Reversible Reactions)

This exercise is just a series of questions which leads the student to an understanding of the term “reversible reaction”. It relies heavily on the understanding of the concept of reaction rate, and so the first worksheet is excellent preparation for this. AfterS1’s first session, this exercise took ~ 5 minutes to complete.

Exercise 2 (Dynamic Nature of Chemical Equilibrium)

This exercise uses the Equilibrium Machine to convey a conceptual understanding of chemical equilibrium, again making heavy use of the kinetic understanding conveyed in Worksheet 1. The Equilibrium Machine uses a table of reactant and product concentrations of a simple equation over time to generate plots of concentration vs. time and reaction rate vs. time. Initial concentrations of both reactant and product and rate constants for forward and reverse reactions can be entered into the table to produce plots of varying appearance.

We started with a review of the EM plot of anirreversible reaction, where the reactant is for all purposes completely converted to product. He quickly recognized this plot and was able to describe its essential elements (reactant concentration goes go zero, etc.). When a reverse reaction was added to the table and plot, he became very interested; it was obvious that he was trying to get the new information in the plot conform to the information already in his mind. He quickly became able to describe the essential differences between the two plots ("the reactant does not go away", etc.).

After spending a few minutes thinking about the time behavior of the concentration plot,S1 eventually described the differing regions of the plot in terms of the reaction behavior ("the product/reactant concentrations are increasing/decreasing here", "the concentrations stop changing here"), and noted that the forward and reverse rates become equal at some point. He correctly drew pictures of the relative equilibrium concentrations of the reversible vs. irreversible reactions. Reading from the plot, and knowing the equation for the Keq from his first attempt at Chem 1B, he was able to calculate Keq from a reading of the concentration plot. He then did a calculation of Q from a point on the plot prior to the equilibrium state, and correctly described why these two numbers were different, and what would have to happen to get from Q to Keq. ("need more product or less reactant"). After changing initial concentrations of products and reactants,S1 was able to articulate why the equilibrium constant calculated from the plot did not change.

Finally,S1 changed the value of the forward rate constant in the EM table, which changed the levels of the equilibrium concentrations in the plot. He had trouble articulating the reason for this. I eventually said that there are only two ways a rate constant could be changed in real life; by changing the temperature, or by changing the components of the reaction itself. This exercise posed a “trick question”, in a manner of speaking. It serves mainly as an exercise in exploring the relation between the numerous variables which impact the equilibrium process. Its purpose is to exercise the mind and to force the consideration of these variables and how they might influence the results seen in the Equilibrium Machine. The reactions of other students to this exercise will be reviewed below.

Summary

A review of this session, and of the audio transcript, was instructive in modifying the worksheet at several points. Questions were added which helped the transition from irreversible to reversible reactions and the identification and similarities of various regions of the plots of these reactions. Additionally, some changes were made which made the worksheet more usable for first-time 1B students, who had only a 1A background and no real understanding of the concept of chemical equilibrium.

S1’s final comment was that he found the final exercise, which involved adjusting the rate constants (and thus changing the equilibrium constants) and observing changes in the plot. When the rate constants were changed and equilibrium constants measured, they were found to be different from those previously calculated. This brought home the understanding of the association of equilibrium constant to a unique chemical equation, and that if the equilibrium constant changes, it means that, in effect, a rate constant changes, and that if the rate constant changes, it must mean that either the temperature has changed, or that a new reaction is being considered. This concept is a bit advanced, and this exercise will likely not work very well with pre-1B students.

S3: Session 1 (2/9/10)

Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 7) with Equilibrium Machine rev 6

Recording: S32-9-10
Notes: Notebook 1 pg 22

S3is working on a second Bachelors degree in BioChem. He is older than the typical student. He had no high chemistry training, got an A in Chem4 and a B in Chem1A. He took 1B for the first time in Spring 2010.

<u>Worksheet Review: WS2</u>

Exercise 1 (Reversible Reactions)

In S3's session, the calculations of concentrations and rates of the reaction were reviewed and then brought quickly to the EM (Exercise 2) rather than being done manually. This was done mostly for reasons of time, as S3 had come late to the process and was being pushed to get through two worksheets in his session, but also in the belief that, once the concepts were grasped (concentration varies with time, rate is proportional to concentration, etc.), the calculations were less valuable as constructivist tools, and that the EM, being more of an interactive tool, would better serve the purpose of the "guided inquiry" process. It's still not clear as to whether this is the case; this will continue to be investigated.

Exercise 2 (The Dynamic Nature of Chemical Equilibrium)

S3 was able to quickly articulate the differences between the irreversible and reversible reaction plots are studying them. He was able to locate the equilibrium region and describe its characteristics. (Note that, although he was taking Chem 1B for the first time, he had heard the lecture and done the reading on the basics of chemical equilibrium.) He talked about the constancy of the reaction rates in this region, and the equivalence of the forward and reverse rates.

When first asked to pull data from the plots to calculate the equilibrium constant, he said that "any two data points will do". He then made calculations using points from t=3 seconds (before equilibrium had been reached) and then again for t=20 seconds (in the equilibrium region) and came up with two different numbers. After some "Socratic" dialogue, he was able to figure out why he had obtained two different answers, recalled the definition of Q, described the difference between Q and K, and indicated which answer was which. I saw this as a great demonstration of the changing of a misconception.

We then began changing the starting concentrations of the reactants and products, and recalculated Keq each time; S3 articulated why these numbers were always the same. We finished by manipulating the rate constants (his initiation) and recalculating Keq. I tried to get him to grasp the significance of the rate equations (that the rate was equal to both kf[A] and kr[B], and thus Keq was equal to kf /kr , so that a change in kf or kr was essentially a change in Keq , but this didn’t happen. This is kind of an arcane “equation”; there might be a better way to convey why the equilibrium constant changes with the reaction constants.

Summary

I feel that a number of misconceptions were addressed and corrected during this session; 1) the relation of reactant coefficients to their relative rate of consumption, 2) the dynamic nature of the reaction rate through the progression of a chemical reaction; 3) idea of the dynamic state of reaction rates at equilibrium, and 4) the precise definition of the meaning of the equilibrium constant. In addition, the use of the EM provided a visceral sense of why and how the equilibrium constant remains static throughout changes in amounts of reactants and products in a solution. I think that WS 1 and 2 are close to completion for use in Chem 1B.

S2: Session 1 (3/1/10)

Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 9) with Equilibrium Machine rev 6

Recording: S2 1
Notes: None

S2 got a B+ in High School chemistry, a B+ in Chem4 and an A in Chem1A. He took 1B for the first time in Spring 2010. S2 had not initially signed up for the focus groups, but found me in the hallway after failing the first Chem 1B exam.

Worksheet Review: WS2

Exercise 1 (Reversible Reactions)

S2 quickly grasped this concept, and was able to articulate how the rate of the forward reaction was a function of [A] and the reverse reaction was a function of [B]. He stated quite confidently that the two rates "had to be the same". After being led to see that the forward and reverse reactions are fundamentally different, that different sets of bonds were being broken and formed in the two reactions, he said "I guess I don't know what to say". (The issue of whether the rate constants for the two reactions could be different was too abstract for him to deal with, and on balance the value of this question does not seem to be worth the time it takes. I've removed it from the worksheet in rev 11.) S2 worked his way through the set of questions about products and reactants in a reversible reaction, and finally came up with a great articulation of the process; with regard to the equations

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)} \text{ and } 2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)},$$

"if you look at the reactions, they're both the same . . . I can look them and they both have H2 and I2 and HI . . . because I can look at one and say that's the formation of HI, and I can look at the other and say that's the dissociation of HI". At this point, though, he stated that he was "still confused" by the concept of equilibrium. The nature of this confusion was hard to grasp; he seemed to have trouble with the idea of both forward and reverse reactions occurring simultaneously; he talks about the existence of "a forward reaction or a reverse reaction".

At the completion of Exercise 1, the elapsed time was ~40 minutes.

Exercise 2 (The Dynamic Nature of Chemical Equilibrium)

This exercise began with a review of the basics of chemical reaction kinetics, following the basics of Exercise 2 from Worksheet 1. This took some time, and it became obvious that this was new material for him.

One of the learning moments for both student and instructor came at a point after S2 grasped the concept of simultaneous forward and reverse reactions. As we got to a certain point in time, S2 observed that more B than A was being produced (i.e., the forward reaction was faster than the reverse reaction). How, he asked, was equilibrium going to be reached under these conditions? After a little more work, he recalled from earlier in the exercise that reaction rate was a function of concentration, and as the concentration of reactant decreased, the forward reaction rate would also decrease, and vice versa for the reverse reaction. The understanding that reaction rate is a function of concentration, and thus changes over time, is a key to understanding the concept of chemical equilibrium. Until this moment in S2’s session, it had not been quite so obvious as to why the former concept was so important as a precursor to the concept of equilibrium. An understanding of the dynamic nature of reaction rate also helps the understanding of the behavior of a system at equilibrium; as concentrations change, reaction rates change until the forward and reverse rates are equal. In other words, when the same quantity of A and B are produced each second, the net change in concentrations is zero, and equilibrium has been attained.

Eventually, S2 got to the point where he was able to fluidly articulate the nature of the relative reaction rates of A and B and how determined the relative changes in concentrations of A and B. For some reason, though, even when presented with the plot of concentrations over time, and after correctly identifying the equilibrium region, when S2 was asked to define equilibrium in terms of the behavior of concentrations of the reactive substances (they become constant over time) and the forward and reverse reaction rates (they are non-zero and equal), he responded that "the concentration of A is equal to the concentration of B". The only reasonable explanation for this is that he did not understand the question, and believed that there was some relationship between A and B at equilibrium, analogous to that of the reaction rates, and was searching for that relationship on the chart.

Summary

Because of his strong record in Chemistry (a B+ in Chem 4, an A in Chem 1A) and his failure on the first Chem 1B exam, S2's work in this session is especially interesting. S2 was missing some very basic concepts key to chemical equilibrium.

S2's parting comments: "Yeah, it's helpful to know what's really going on . . . . in certain situations, like strong base/weak acid, strong acid/weak base, you know, because I'm actually able so see what's going on, you know . . . applying (knowledge of chemical equilibrium), that's when it gets hard".

S7 and S5: Session 2 (7/14/10, 1:30 P.M.)

Worksheet 2: Reversible Reactions and Chemical Equilibrium rv12, Equilibrium Machine rv 9

Recording: S7 and S5 WS2 7-14-10
Notes: Notebook 1 pg 37

Exercise 1 (Reversible Reactions)

The first two questions went very quickly, as we had just finished reviewing them in the previous session.

The term "reversible reaction" was recognizable to them from Chem 1A, and they were quickly able to articulate its meaning ("A goes to B and B goes to A at the same time . . . both A and B are in the final solution . . .", etc.). At question c, they quickly grasped the objective of questions c, d and e, and immediately went to the meat of it ("they're the same (equation) . . . there's no such thing as a reactant or a product"). This was the fastest and most fluid progress through this exercise to date; I credit this mainly to having completed Worksheet 1 immediately beforehand.

Exercise 2 (Chemical Equilibrium)

This exercise went well, but did not cleanly follow the order of the worksheet. This is probably to be expected, as some shortcuts and jumps were merited here due to the quick acquisition of CE concepts by S7 & S5 as a result of just having reviewed Worksheet 1. Their early and clear articulation of the performance of a reversible reaction greatly aided their progress through Exercise 2.

The form of the table and its contents was quickly grasped, and questions a and b were answered in ~30 seconds. After entering the first set of initial conditions for the irreversible reaction, they quickly recognized the form of the plot and correctly identified it, easily answering questions c, d and e.

After adding a reverse reaction to the plot, they struggled a bit with the words to define it, but it was clear they understood what was happening and recognized the key differences between the irreversible and reversible plots (B does not go to zero, forward and reverse reaction rates become equal). They were quickly able to recognize and explain the different regions of the plots (concentrations/rates are changing and then become constant).

The concept of K_{eq} was new to them; they did not recall it from Chem 1A. The difference between the equilibrium constant and the rate constant was clearly delineated with the aim of preventing confusion between these somewhat similarly titled constants, while parallels between the two were also explored (both are unique to a specific reaction; both are constant over changes in reactant and product concentrations). The importance of the equilibrium constant and its definition, as well as its centrality in the curriculum of Chem 1B, was emphasized, and the interest shown by S&H in this concept was quite obvious. From their comments, I took this to be because they had heard the expression (maybe just from me in my Focus Group pitch in Chem 1A lecture!), but without any real understanding, had a genuine curiosity as to its meaning.

Some probing questions were asked on the topic of the equilibrium constant and its mathematical definition. The best of these had to do with the way in which the reaction was written (see questions c, d, and e above). If it truly made no difference as to which side of the reaction (left or right) the reactants and products appeared, then how do you know how to write the expression for K_{eq}? This led to discussion on “convention”, and the issue was eventually understood.

Their calculation of K_{eq} from plot data was very quick, and an exercise using data from the pre-equilibrium region of the plot quickly informed them as to the correct calculation procedure. Changes to starting concentrations were made, and as the form of the concentration plots changed, equilibrium constants were calculated and found to be constant; S&H’s articulation of this phenomenon made it clear that it was understood.

S7 & S5 then studied the rates vs. time plot, and articulated its form (the equality of the forward and reverse rates in the equilibrium region). A free-form discussion ensued on what was actually happening in the equilibrium region with respect to the logistics of the reaction, the result of which was their articulation of the dynamic nature of equilibrium; to paraphrase their comments, “in equilibrium, both rates are equal, so A is being produced at the same rate as B, so there’s no net change in the concentration of either”. On an editorial note; this was clearly an “aha” moment for S&H, and it was great to see.

Finally, the final optional question on the worksheet was examined. I wrote the rate equations for the forward and reverse rates and the equation for the equilibrium constant on the whiteboard and asked if they could see how these equations could be used to form a new equality. This was done, and the constancy of the rate constants and the equilibrium constants and their relation were grasped.

Summary

The fast progress through Exercise 1 indicates that perhaps some additional questions should be posed with regard to the concept of reversible reactions; this will be added as a Bin List item to Worksheet 2. Question k (requiring students to draw the residual components of an irreversible and reversible reaction) was too simple at this point. It has been moved into Exercise 1.

As noted above, the close timing between the Worksheet 1 and Worksheet 2 sessions had a great impact on the ease with which this session was run. As a result, this session was completed in ~50 minutes.

S8: Session 2 (7/30/10, 11:30 A.M.)

Worksheet 2 (Reaction Kinetics and Stoichiometry rev15, Equilibrium Machine rev 10)

Recording: S8 WS2 7-30-10
Notes: Notebook 1 pg 42b

Exercise 1: Reversible Reactions

S8 read the introductory notes on the worksheet and had no trouble with the exercise. She articulated the differences between reaction types with no difficulty, and answered the questions easily. Note to self: She did not understand question d. This has not been the first time I've seen this problem. This will be corrected.

Exercise 2: Chemical Equilibrium

S8 read the introductory notes on the worksheet and we talked about the idea of forward and reverse reactions. After moving to the EM spreadsheet, she quickly understood the table (identical to the one she had viewed earlier that morning), and was able to identify the key characteristics of irreversible vs. reversible reaction plots. From this point, we quickly went to the notion of a dynamic state of equilibrium, which she was able to clearly articulate (". . . forward and reverse reaction rates are equal, so the net change in concentrations is zero". .).

I then introduced the idea of the equilibrium constant, and she read the appropriate notes in the worksheet. S8 got this concept quickly, and then identified the equilibrium region on the plot. When asked to calculate K_{eq} from the plot, she quickly explained what data was appropriate for this calculation and what was not, and why or why not.

Exercise 3: Stoichiometry and Equilibrium

The exercise went much more slowly, and was well worth the time, based on the learning I perceived on S8's part. The key learning here is how the loss or gain of a compound can be related to the loss or gain of other compounds in the reaction. The key difficulty appears to be when relating a *loss* of a compound to a gain in a "product". When initial and final conditions are given, the tendency is to somehow relate the final amount of a consumed compound to other quantities in the equilibrium state, rather than subtracting the initial quantity out and using that a guideline.

This problem is made more straightforward when using the ICE method for calculation gains and losses for an equilibrium problem, but at the possible expense of loss of conceptual understanding on the part of the student. The ICE method can be used with a rote methodology; "I'll put a minus x on one side of the equation and a plus x on the other side". This approach can cause problems when more complex forms of the problem are posed. Note to self: Need to formulate a clear way of communicating this to the student at a conceptual level. One simple method which seemed to work was just using the whiteboard review from Exercise 1 of Worksheet 1. As reactant compounds are taken off the whiteboard, a number of products will appear, the quantity of which is dictated by the stoichiometric ratio of the two compounds in the reaction equation. This seemed to help S8 grasp this relationship.

S6: Session 2 (9/20/10, 10:00 A.M.)

Worksheet 2 (Reversible Reactions and Chemical Equilibrium rev 17, EM rev 13)

Recording: None
Notes: None; direct to Word.

WS2: Exercise 1: Reversible Reactions

S6 very quickly grasped the meaning of the questions in this section, and was able to articulate the irrelevance of the terms reactant and product in describing reversible reactions ("it's just the way they write the reaction").

Exercise 2: The Dynamic Nature of Chemical Equilibrium

After all the trouble we had with establishing the basics of reaction kinetics, S6 seemed to grasp this session very quickly. No significant comments to make here.

Worksheet 3 Focus Group Transcripts

Part 1: Focus Group Summaries

S1: Session 3 (2/12/10)

Worksheet 3 (Basics of Acid/Base Equilibria rev 3) with Equilibrium Machine rev 6

RECORDING: S1D 2-12-10
Notes: Notebook 1 pg. 25

Exercise 1: Strong vs. Weak Acids/Bases

This went very quickly. At firstS1 wrote the dissociation reaction of nitric acid as a reversible reaction; the quickly recognized that it was not "reversible" as taught in General Chemistry, and articulated why that was so. He was able to describe very clearly the difference between strong and weak acids, and was able to list the six strong acids. This could be attributed to his previous Chem 1B stint, in which he would have been heavily exposed to these topics. S1 quickly drew visual representations of strong and weak acids.

Exercise 2: Equilibria of Acids and Bases in Water

S1 had a problem answering the question about the relationship between K_a and pH. The equation for K_a was written on the whiteboard, and he was asked the key difference between strong and weak acids. When he answered that strong acids had higher levels of dissociation, and that meant more hydronium ions, he quickly realized the correct answer. Some additional discussion ensued in order to make sure that he had correctly absorbed this concept.

Exercise 3: Relationship Between Ka, pH, and Concentration

In this exercise, a plot of pH vs. concentration for various weak acids was shown and a guided-inquiry process used to get students to explore the concept of the equilibrium of a weak acid or base.

WhenS1 was first shown the single-line plot, he was unable to articulate exactly which concentration was being plotted on the X axis (labeled “molarity”). He initially thought it was the hydronium ion concentration until it was pointed out that this was being plotted on the Y axis in the form of pH. He eventually just needed to be told what it was. (The leadup to this question has been considerably lengthened since this session in order to get students better prepared to answer.)

From there,S1 was able to quickly grasp the workings of the plot, identifying the acid with the highest and lowest Ka on the multi-line plot, and talking about why pH decreases as molarity increases and vice versa. He was able to articulate why the three acids with similar pH values had such drastically different concentrations.

The final question set in this exercise (calculations of K_a) took about fifteen minutes. S1 easily converted pH values to hydronium ion concentrations, and introduced these values into the K_a equation, but got stuck there. He needed a lot of help to see that his $[H_3O^+]$ values could be converted to $[F^-]$ and [HF] values, as they were related by the stoichiometry of the equation ($HF + H_2O \rightleftharpoons H_3O^+ + F^-$).

This inability to apply a simple stoichiometric ratio is puzzling, and goes back to the first exercise in Worksheet 1. What is the problem here? It could be that 1) He didn’t know the dissociation reaction for HF; 2) He didn’t know what the ratios were; 3) He didn't know how to apply the ratios; or 4) After calculating H_3O^+, he didn't make the connection to the equation, from which he could have obtained the other concentrations needed to make the calculation. It seems to me that the most likely answer here is the last one. This question will be explored more deeply in future focus group meetings.

Once he was guided to the solution above, and got $[F^-]$ value he needed for the calculation,S1 made another common mistake. He realized that, to get the final value of [HF] for the K_a calculation, he would have to subtract the amount of product generated from the initial [HF] value. However, when he went to perform this subtraction, he subtracted *both* $[H_3O^+]$ *and* $[F^-]$ from $[HF]_0$. Once more, the lack of an innate understanding of the concept of stoichiometry as it applies to chemical reactions has been seen; this time, it has resulted in an inability to solve a more complex problem. If Jason had stopped to think what the HF dissociation reaction was telling him about the relationship between product generated and reactants consumed, he might have been able to do this calculation on his own. At one point during the exercise, Jason said that "I need to use an ICE table", even though he clearly did not understand the principles behind it. When conceptual understanding failed him, Jason resorted to rote learning of a formulaic solution. The ICE method was used, after which the values obtained for the final concentrations were compared to the values from the attempt to use plot values to solve the problem, and how they were basically the same; that, in other words, the ICE method was just a "standard accounting" method for problem solving, and that it related back to the first principles of stoichiometry.

Summary

This was the first session in which the failure to apprehend basic principles of chemical reactions directly impacted the ability to perform equilibrium problem solving. This problem is made more serious, perhaps, by the fact that the student in this session is a repeating Chem 1B student who is still unable to firmly grasp these concepts.

S3: Session 2 (2/16/10)

Worksheet 3 (Basics of Acid/Base Equilibrium r3) with Equilibrium Machine rev 6

Recording: S32-16-10
Notes: Notebook 1 pg 29

Exercise 1 (Strong vs. Weak Acids/Bases)

S3 took some time to write the reactions of acids and bases with water. He got stuck on getting to the NIE from the starting equation. He eventually worked out questions a and b, and got the drawing in c correct. *This exercise was much more of a challenge to S3 than it was toS1, who was repeating Chem 1B.* These questions should remain in the worksheet.

Exercise 2 (Equilibria of Acids and Bases in Water)

S3 worked slowly and thoughtfully through this exercise, and eventually answered all questions correctly. This appeared to be a valuable exercise for him, and he said as much.

Exercise 3 (Relationship Between Ka, pH, and Concentration)

Only the first part of this exercise (with the single data plot) was completed. S3 also had trouble with question a. This question has been redone, and the plot labeled more clearly.

S3's final comments; he found the exercise helpful and felt he learned something. "Just taking the time to talk with you one-on-one is helpful and makes it stick."

APPENDIX E

Stanford University General Chemistry Documents

Chem 31B Syllabus

Chemistry 31B, Winter 2011
Chemical Principles II: Structure & Energetics
Professors Fayer& Schwartz

Course Overview:

The science of chemistry evolves through a process of observation, hypothesis, and experimentation. This course is structured such that you can continue developing your skills for participating in this process while building your understanding of how chemical phenomena shape our world.

Broad Scientific Objectives:

- *Observation and Conceptual Modeling: Recognize patterns in observations of chemicalphenomena and construct conceptual models to explain the phenomena.*
- *Explanation and Estimation: Apply conceptual models to qualitatively explain a wide range ofphenomena and to make quantitative estimations.*
- *Problem Solving: Develop critical thinking skills to analyze and solve problems in chemistry.*
- *Communication: Develop skills to discuss chemical concepts and explain your thought processesand conclusions in writing.*

Specific Chemical Objectives:

- *Equilibria: Determine the direction of a reaction and the changes in concentration that will occuras a reaction comes to equilibrium.*
- *Electrochemistry: Identify redox reactions and calculate their reduction potentials based on theirconditions. Describe how to build a voltaic cell.*
- *Thermodynamics: Predict whether a reaction is likely to be spontaneous and describe therelationship between free energy, temperature, and equilibrium constants.*
- *Kinetics: Determine what forces influence the rate of a chemical reaction.*
- *Be prepared to study how structure influences chemical reactivity and equilibria in Chem 33.*

For Questions please contact Jennifer.Schwartz@stanford.edu

Chem 31B learning cycle:

- *Observe and experiment with various substances to gain experience with the chemicalphenomena covered in that learning cycle.*
- *Begin to develop an understanding of the phenomena and models of the cycle by summarizingand analyzing section findings in a written report based on the first lecture and assigned reading.*
- *Further develop your understanding of the phenomena and models, and your skills for solvingchemical problems, through assigned reading and exercises, discussion and quizzes in lecture,and small-group argument and problem solving. By the end of the week, make an initial,individual attempt on each of the problems in the problem set.*
- *Discuss and check your answers to the problem set with classmates and seek guidance as neededfrom the instructional team. Then, individually prepare your own solutions.*
- *Practice additional problems in lecture and review the solutions to the problem set to seek a fullunderstanding of the problems you found most difficult. Further challenge your problem solving skills with additional problems from previous exams. The last lecture will be a review of the keyphenomena and models in the cycle.*
- *Complete an exam at the end of the cycle to assess your understanding and skills*

Calculations with Acids, Bases, and Buffers

CHEM31B – Winter 2011 Week 4

Goal: To explore how a buffer system works and why it is important.

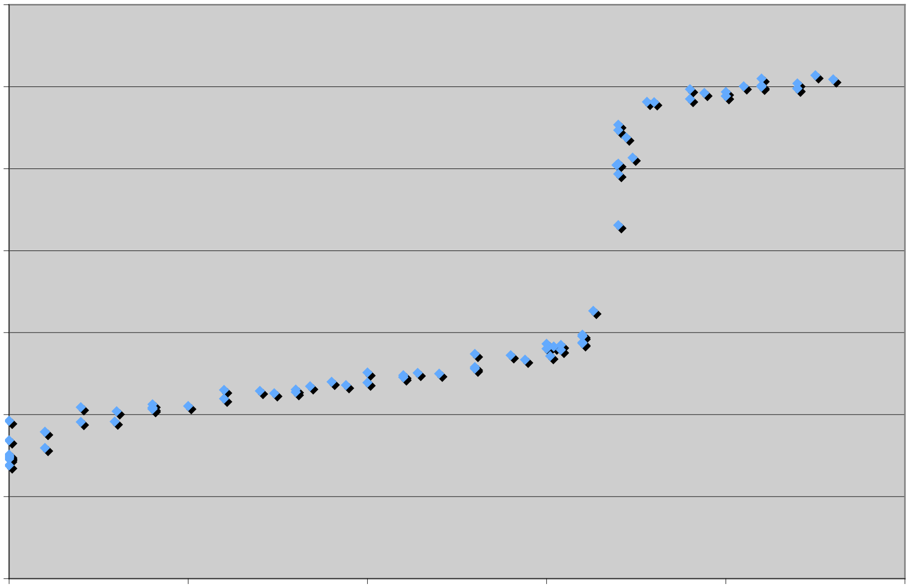
A **buffer** is a combination of a weak acid and its conjugate base or a weak base and its conjugate acid. A buffer helps to keep the pH of a system relatively constant when small amounts of acid or base are added. Acetic acid can be used as a buffer for solutions needing to maintain an acidic pH:

Rxn 1:	CH_3COOH (*aq*) +	H_2O (*l*)	⇆	CH_3COO^- (*aq*) +	H_3O^+ (*aq*)
	acetic acid	water		acetate ion	hydronium ion

Buffers are often described by a pKa value, where pKa = -log {Ka}. The pKa for acetic acid is 4.74.

Part 1: Understanding Titrations and the Buffer region

During last week's section a weak acid was titrated with a strong base (NaOH). The plot below is from a similar experiment, just at different concentrations.



Part I: Buffer versus Water solutions

1) Write out an equation for the dominant reaction that occurs when NaOH is added to a beaker containing dilute acetic acid initially at a pH of 2.7. What type of reaction is this?

2) What can you say about the magnitude of K_c for this reaction? What assumption does that allow you to make?

3) Determine the pH of the system if we add 10.0mL of 1.0M NaOH(*aq*) solution to a 100.0mL solution containing 0.020mols of acetic acid. On the left side of the page, write out your equations and calculations. On the right side, make sure to provide a justification or explanation for each step.

Calculations	**Justification**

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4) What is the change in pH?

5) What will be the pH of a solution if 10mL of 1.00M NaOH(*aq*) is added to a beaker with 100mL of pure water in it. ?

6) Compare your answers from 4 and 5. What do you notice about the change in pH when we add strong base to a solution of acetic acid versus the solution of pure water?

PART II: Calculations of acid//base mixtures

Consider again the 100.0mL solution containing 0.020mols of acetic acid, which is titrated by 1.0M NaOH.

7) *Label the point on the titration plot on page 1* where a buffer made from acetate and acetic acid would BEST resist a change in pH from addition of either a strong acid or base. Explain how this point is best able to resist a change in pH.

8) Similarly to #3, calculate the pH of the solution when 20.0mL of sodium hydroxide has been added to the solution of acetic acid.

a) What assumption is being implicitly made in our above calculations regarding the source(s) of hydronium ions (H_3O^+)?

b) Making this assumption, set up (but do not solve) this problem utilizing Rxn 1 from page 1. Can this set up successfully solve this problem? Why or why not?

c) What changes are needed to set up of this problem to determine the pH after adding 20.0mL of sodium hydroxide? (Hint: what species are available to react.)

d) What is the pH of the solution at this point? **Label this point on your graph.** Why is the pH not seven when all the acetic acid has been neutralized? What determines the final pH of the system at this point?

PART III: Buffers in our blood

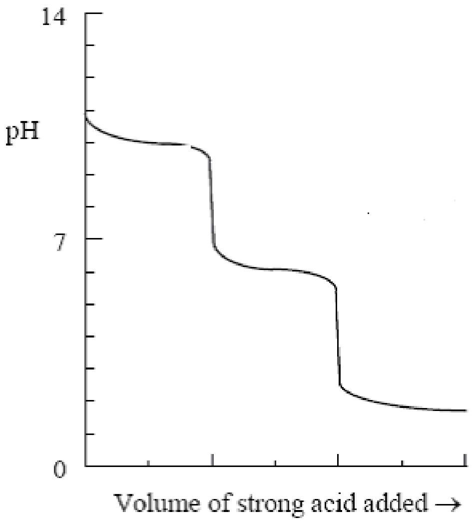
The human body takes advantage of the following reactions as part of a buffering system to keep the blood at a pH of ~7.4:

$$CO_2(g) + H_2O(l) \leftrightarrows H_2CO_3(aq) \qquad \text{(rxn 1)}$$

$$\underset{\text{carbonic acid}}{H_2CO_3(aq)} + H_2O(l) \leftrightarrows \underset{\text{bicarbonate}}{HCO_3^-(aq)} + H_3O^+(aq) \quad \text{(rxn 2)}$$

The first equilibrium (rxn 1) is not an acid-base reaction, but it is important for keeping the pH of the blood constant. The pK_a of carbonic acid is 6.37 and the pKa of bicarbonate is 10.33

9) If sodium carbonate ($Na_2CO_3(aq)$) is titrated with hydrochloric acid ($HCl(aq)$), the curve below is observed.



a. Write equilibrium reactions for the two protonations of carbonate.

b. For each region (1-5), explain what is happening as you add acid. Identify the chemical species present and the relative amounts. (i.e. is there more HA or A^-?)

c. What is the exact pH at point 3 in the plot (indicated by the arrow)?

d. When a person hyperventilates, this results in the rapid loss of CO_2 from the body. Thinking about the reactions above, how would this affect the pH of a person's blood?
10) Suppose you want to make a buffer that will keep the pH at 9.0. Which of the following weak acid/conjugate base pairs would be most appropriate? Why?
What ratio of concentrations would you need to give a pH = 9.0?

a. Acetic acid/sodium acetate, $K_a = 1.8 \times 10^{-5}$
b. Sodium phosphate monobasic/sodium phosphate dibasic, $K_a = 6.2 \times 10^{-8}$
c. Ethanolammonium chloride/ethanolamine, $K_a = 3.2 \times 10^{-10}$
d. Piperidinium chloride/piperidine, $K_a = 7.6 \times 10^{-12}$

Simplifying Calculations with Buffers

1. Let's consider a generic acid dissociation reaction. Write an expression for the equilibrium constant of the following reaction:

2. Solve the expression for $[H_3O^+]$.

3. Rewrite the expression from #2 in terms of pH.

4. Look at your expression from #3. How does the ratio $[CH_3COOH]/[CH_3COOH^-]$ affect the pH?

5. Thinking about the titration you performed in last week's section, in what pH range is the buffer most effective? Label this region on the graph on the first page.

6. What is the optimal pH for this buffer? For any buffer?

7. What is the ratio [CH_3COOH]/[CH_3COOH^-] at this optimal pH point?

8. Suppose you want to make an acetic acid/acetate buffer with a pH of 5.3.
    a. What ratio of [CH_3COOH]/[CH_3COO^-] will you need?
    b. You have 100mL of 0.100M acetic acid and 100mL of 0.100M sodium acetate. How much of each solution would you add to make the buffer at pH of 5.3?

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Redox Reactions and Voltaic Cells

Activity 5: Redox reactions and voltaic cells

Goals of Activity: Estimate the equilibrium constant for displacement reactions
Measure the electrical potential of a voltaic cell

Lab Safety: Always wear appropriate lab attire: safety glasses, long sleeves and long pants, close-toed shoes, and gloves. Do not chew gum or eat in lab.

Caution: Silver nitrate can stain your skin. Do not to spill it on yourself.

Reagents	Materials
Silver, copper and zinc wire	24 well plate
1.0M silver nitrate, $AgNO_3(aq)$	Filter paper
1.0M copper nitrate, $Cu(NO_3)_2(aq)$	pH meter
1.0M zinc nitrate, $Zn(NO_3)_2(aq)$	Alligator clips
10% potassium nitrate, $KNO_3(aq)$	

Oxidation: When an element or compound loses electrons. When the oxidation number of an atom becomes more positive.

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

Reduction: When an element or compound gains electrons. When the oxidation number of an atom becomes less positive.

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

Redox Reaction: A reaction in which electrons are transferred so that one element or compound is oxidized and another element or compound is reduced.

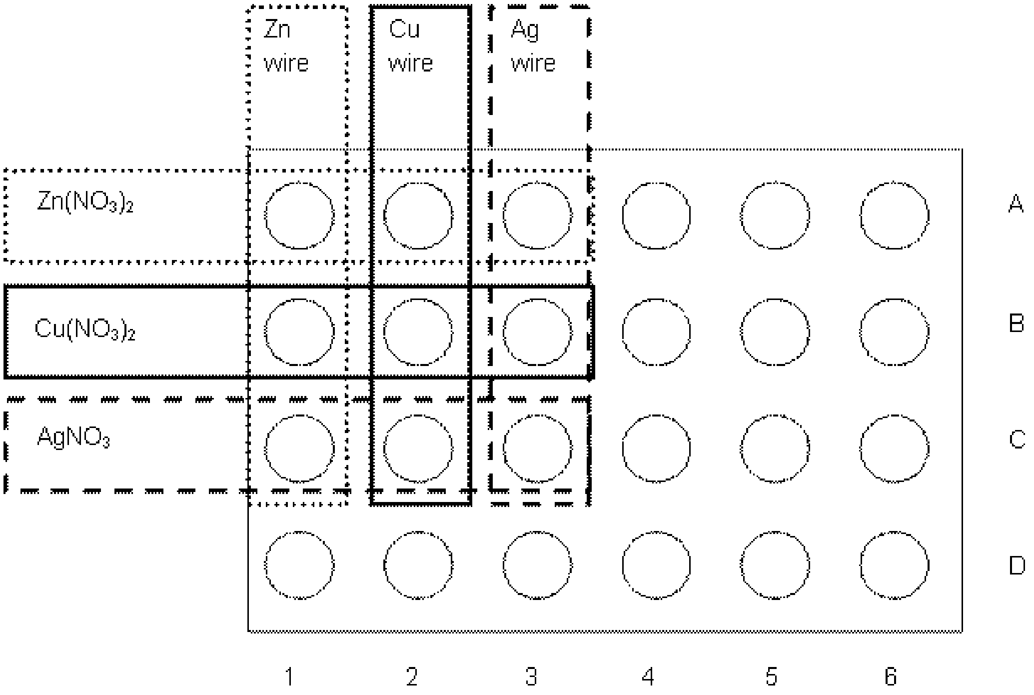
Electrical Potential: A way of describing the potential energy that charged objects have in an electric field. The potential provides a measure of the driving force behind an electrochemical (redox) reaction. The **volt** (symbol: $V \equiv J/C$) is the SI derived unit used to measure electric potential difference.

Electrochemical Cell: A setup in which separate oxidation and reduction half-reactions are connected by conductors to make a circuit. A **voltaic cell** is an electrochemical cell in which the redox reactions produce an electric current.

<u>Clean-up:</u> Because we are working with toxic heavy metals it is very important that <u>nothing be put down the sink.</u> All metals and solutions must be disposed of in the heavy metal waste containers located at the center of each bench.

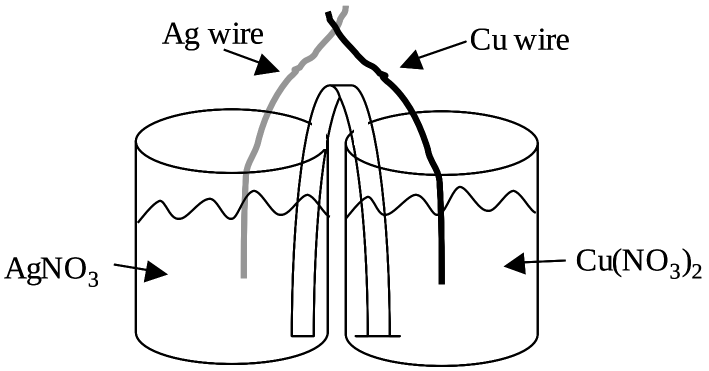
Part 1 – Identifying displacement reactions PARTS 1 & 2 ~ 30 min.

1. Set your 24-well plate over the diagram on the next page, so that the well labeled A1 is in the upper left corner.
2. Fill the wells at least two-thirds full with metal nitrate solutions. Solutions of each can be found in the center of each lab bench. Do NOT mix pipettes between solutions:
   a. Fill wells A1, A2, and A3 with 1.0M $Zn(NO_3)_2$.
   b. Fill wells B1, B2, and B3 with 1.0M $Cu(NO_3)_2$.
   c. Fill wells C1, C2, and C3 with 1.0M $AgNO_3$.
3. Place metal wire pieces in each well:
   a. Zinc wire in wells A1, B1, and C1.
   b. Copper wire in wells A2, B2, and C2.
   c. Silver wire in wells A3, B3, and C3.
4. Do **NOT** let the wires from different wells touch.
5. The final layout of your 24-well plate should look like the diagram below:
6. Let the reactions sit undisturbed while you continue with Part 2.



Part 2 – Demonstration: A reaction in two compartments

Your TA will show you an experimental setup similar to the picture below.



In this setup, a piece of copper wire and a piece of silver wire have been twisted together so they are touching. The silver wire has been dipped into a beaker of 1.0M silver nitrate. The copper wire has been dipped into a beaker of 1.0M copper(II) nitrate. The wells are also connected by a strip of filter paper that has been wet with 10% KNO_3 solution. This type of a connection between two half-cells is called a **salt bridge**. The wires have been left in their respective wells for a number of hours.

Key questions for Parts 1 and 2

1. Remove the wires in Part 2 from the beakers and examine them. What do you observe?

2. Now go back and look at the wells in Part 1. Record your observations for each well in the following table. In which well does the reaction seem to occur to the greatest extent?

Well Number	Aqueous solution	Metal added	Observations
A1	**$Zn(NO_3)_2$**	**Zn**	
A2	**$Zn(NO_3)_2$**	**Cu**	
A3	**$Zn(NO_3)_2$**	**Ag**	
B1	**$Cu(NO_3)_2$**	**Zn**	
B2	**$Cu(NO_3)_2$**	**Cu**	
B3	**$Cu(NO_3)_2$**	**Ag**	
C1	**$Ag(NO_3)$**	**Zn**	
C2	**$Ag(NO_3$**	**Cu**	
C3	**$Ag(NO_3)$**	**Ag**	

3. What elements, molecules and/or ions have you added to the wells A2, B2 and C2?

4. Consider possible redox reactions involving the metal species in A2, B2 and C2. Write their chemical equations into the table below. In columns 3 and 4 indicate which metal was oxidized and which metal was reduced. In the 5th column write the number of electrons that were transferred during the reaction.

Well	Possible Redox Reaction	Oxidized Metal	Reduced Metal	# e^- transferred
A2				
B2				
C2				

5. Write down the equilibrium constant expressions, K_c, for the equations you wrote into the table above. Based on your observations, identify which reaction is favorable. In each case, is the value of K_c greater than, equal to, or less than 1?

6. What chemical species are present in the two beakers in Part 2? (A "chemical species" can be an atom, molecule, or ion.) Compare the reaction within cell C2 to the electrochemical cell in Part 2. How are they similar? Different? What must be moving through the wire in the Part 2 set-up?

Part 3 — Electrical potential difference PART 3: 20 min.

Part 2 was an example of a voltaic cell. We can quantify the reactivity of this and other cells by measuring the difference in electrical potential between the two metal wires, using a multimeter. Multimeters are able to measure voltage, current and resistance in an electrochemical cell.

1. Obtain two strips of filter paper and place them onto large watch glass. Using a fresh pipette, wet the filter papers with a few drops of 10% KNO_3 solution.
2. Use a strip of filter paper to connect cells A1 and B2 by putting one end of the paper in cell A1, and the other end in B2. Use the other strip to connect cells B2 and C3.
3. Connect the clips from the multimeter to the Cu and Zn wires from cells A1 and B2. If you do not have clips, simply wrap a small piece of each wire around an end of the multimeter probe.
4. Verify that the multimeter is set to measure V. Record the meter reading, including its units and sign, and which lead is connected to the Cu electrode and which to the Zn. Switch the leads, and make the same measurements. Fill in the following table with your results.
5. Repeat Step 3 for the Cu and Ag wires in cells B2 and C3.
6. Repeat Step 3 for the Ag and Zn wires in cells A1 and C3.
7. Turn off the multimeter when you are finished.

Electrode at red clip	Electrode at black clip	Reading on meter

Key questions for Part 3

1. What happens when you switch the red and black leads for a particular pair of metals?

2. Recall from the first page, that the electric potential is a measure of the driving force of a reaction. Based on your observations in Part 1, which reaction would you expect to have the largest potential? How does this compare with your observations in Part 3?

3. Consider the positive voltages you measured for the three half-cell combinations. Add together the values for Ag/Cu and Cu/Zn. How does this number roughly compare to the Ag/Zn potential?

4. How does the voltage measured on the part 2 cell compare to the voltage you measured for the Ag/Cu cell in part 3?

5. What changes occur in an electrochemical cell as it is allowed to 'run'? What happens to a battery after use?

Clean up:*In the middle of each bench find the "heavy metal liquid waste" containers*

1. Place 'un-reacted' wires in the appropriate beaker next to the waste station.
2. Place your used filter paper in the "solid waste" container.
3. Rinse out the remaining solutions from your 24-well plate into the "heavy metal waste" container at the center of the bench.
4. Stack your used 24-well plate next to the waste containers.

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