Oakland Schools Chemistry Resource Unit

Equilibrium & Kinetics

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<u>Equilibrium and Kinetics</u>

Content Statements:

C5.3x:

Most chemical reactions reach a state of dynamic equilibrium where the rates of the forward and reverse reactions are equal.

C2.3x:

For molecules to react, they must collide with enough energy (activation energy) to break old chemical bonds before their atoms can be rearranged to form new substances.

C5.r1x:

The rate of a chemical reaction will depend upon (1) concentration of reacting species, (2) temperature of reaction, (3) pressure if reactants are gases, and (4) nature of the reactants. A model of matter composed of tiny particles that are in constant motion is used to explain rates of chemical reactions. (Recommended)

Content Expectations:

C5.3a: - Describe equilibrium shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).

C5.3b: - Predict shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).

C5.3c: - Predict the extent reactants are converted to products using the value of the equilibrium constant.

C2.3a: - Explain how the rate of a given chemical reaction is dependent on the temperature and the activation energy.

C2.3b: - Draw and analyze a diagram to show the activation energy for an exothermic reaction that is very slow at room temperature.

C5.r1a: - Predict how the rate of a chemical reaction will be influenced by changes in concentration, temperature, and pressure. (Recommended)

C5.r1b: - Explain how the rate of a reaction will depend on concentration, temperature, pressure, and nature of reactant. (Recommended)

Instructional Background Information:

Equilibrium

When a chemical reaction is carried out in a closed container, the reaction eventually comes to “equilibrium.” Equilibrium occurs when there is a constant ratio between the concentration of the reactants and the products. Different reactions have different equilibriums. Some may appear to be completely products; however, all reactions have some reactants present. A reaction may look "finished" when equilibrium is reached, but actually the forward and reverse reactions continue to happen at the same rate. A reverse reaction is when the written reaction goes from right to left instead of the forward reaction which proceeds from left to right.

It is possible to write an equilibrium expression for a reaction. This can be expressed by concentrations of the products divided by the concentration of the reactants with the coefficients of each equation acting as exponents. It is important to remember that only species in either the gas or aqueous phases are included in this expression because the concentrations for liquids and solids cannot change. For the reaction:

$$aA + bB \dashrightarrow cC + dD$$

the equilibrium expression (K_{eq}) is:

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

Where:

K is the equilibrium constant
[A], [B], etc. are the molar concentrations of A, B, etc.
a, b, etc. are the coefficients of the balanced reaction

For every reaction at a specific temperature, there is only one value for K. A large value of K (greater than one) implies that there are more products than reactants. A small K value (less than one) implies there are more reactants than products. It is critical to remember that the only thing that changes K is changing temperature.

Le Chatelier's principle allows us to predict the effects of changes in temperature, pressure, and concentration on a system at equilibrium. It states that if a system at equilibrium experiences a change, the system will shift its equilibrium to try to compensate for the change.

- Changing the concentration (only with gases or aqueous solutions):
  If you lower the concentration or remove some of a species, the system will shift to produce more of that species. On the other hand, if you increase the concentration or add some of a species, the system will shift to produce less of that species. For example, in the equation:

  $$H_2 + I_2 \rightleftharpoons 2HI$$

  If we remove some of the H_2, the system will shift towards the left (the reverse reaction will happen the most) to produce more H_2.

- Changing the volume/pressure (only gases):
  When you increase the pressure (by decreasing the volume), the system will shift so the least number of gas molecules are formed because the more gas molecules there are, the more collisions there are. These collisions and the presence of gas molecules are what cause the pressure to increase. Likewise, when you decrease the pressure, the system will shift so the highest number of gas molecules is produced. For example, in the equation:

  $$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

  If the pressure is increased, the system will shift to the right because fewer gas molecules are produced in the forward reaction than in the reverse reaction.

- Changing temperature:
  For every reaction which can go forwards and backwards, one direction is endothermic and the other is exothermic. A reaction is endothermic if it takes heat from its surroundings. On the other hand, a reaction is exothermic if it releases heat to the surroundings. If you increase the temperature, then the endothermic reaction will be favored because that will take in some of the excess heat. If you decrease the temperature, the exothermic reaction will be favored because it will produce the heat that was lost. For example, in the equation:

  $$PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)} + \text{energy}$$

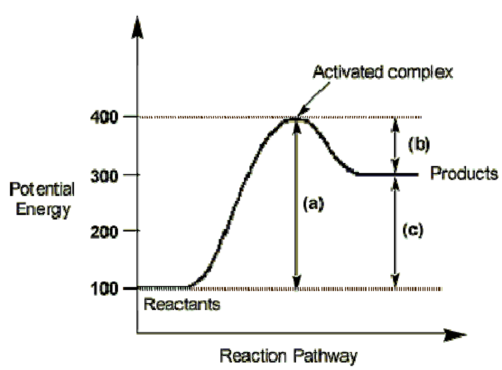
  If the temperature was increased, the system would shift to the left and the reverse reaction would happen more because that would use some of the extra energy.

- Using a catalyst:
  A catalyst increases the speed in which a reaction takes place, however it never has any effect on the equilibrium.
  (http://www.shodor.org/UNChem/advanced/equ/index.html)

Rate of Reaction

In order for a reaction to occur, the reactant particles must collide with one another with sufficient energy to cause a change (known as collision theory). The minimum amount of energy needed to initiate a reaction is called the activation energy.

In the graph below, the activation energy of the forward reaction is given by (a). Overall the graph describes the energy change for an endothermic reaction. The ΔH is given by (c). If it is a reversible reaction, (b) represents the activation energy of the reverse reaction.



In general, a reaction with low activation energy would tend to be fast because more of the collisions would have enough energy for the reactants to react. Similarly, at a high temperature the reactants will have a high kinetic energy and more of the collisions result in reaction. Therefore, reactions at high temperature and low activation energy tend to react quickly.

Other variables that affect the rate of a chemical reaction include concentration of reactants and pressure of gaseous reactants. The higher the concentration of reactants, the more frequently they collide and the faster the reaction. Also, the higher the pressure of gaseous reactants, the more frequently they collide, which speeds up the reaction.

Terms and Concepts

Activated complex
Activation energy
Equilibrium
Equilibrium constant
Le Chatelier's Principle
Rate of reaction
Reversible reaction

Equilibrium and Kinetics

Activity 1: "Activation Energy"

Question to be investigated

What is activation energy?

Objectives

The student will define the terms endothermic, exothermic, and activation energy.

The student will construct and interpret an energy diagram showing the progress of an exothermic reaction.

Teacher notes

All reactions must have enough energy to break the reactants' bonds and create a transition state ("activated complex") before the products form. Reactants colliding with each other isn't enough to initiate a reaction. This activity simulates a simple change requiring a minimum of energy to cause the change.

This activity is geared toward all high school students.

Materials

For each pair of students:

one vial containing 2 scoops of barium hydroxide, $Ba(OH)_2$
one vial containing 1 scoop of ammonium thiocyanate, NH_4SCN
(ammonium chloride works too.)
one portable chemical reaction stick (strike-anywhere kitchen match)
One brick
One manicotti noodle

Safety concerns

All standard safety precautions should be taken when handling chemicals: wear safety goggles, tie hair back, secure loose clothing.

Real-World Connections

Students need to understand that all reactions, even spontaneous, require some amount of energy to begin. The amount of activation energy varies depending on the reaction.

Sources

This entire activity is taken from the SMILE project, http://www.iit.edu/~smile/ch8909.html. It was submitted by Kathy Kriedler of Thornridge High School in Indiana. She gives credit to Ken Spengler for the brick analogy.

Procedure

1. Observing an endothermic reaction:

a) As the students enter the classroom, give each one a vial containing either barium hydroxide or ammonium thiocyanate.

b) When class begins, ask for observations.

c) Have students pair up so each pair has both reactants. Instruct the students to combine the two chemicals and shake to mix.

d) They are to focus their attention on any energy changes observed.

e) The vial becomes cold to the touch; this is an endothermic reaction.

2. Initiating an exothermic reaction:

a) Next pass out the portable chemical reaction sticks and instruct the students to cause a reaction.

b) Be sure that metal cans or ashtrays are plentiful to ensure safe disposal.

c) Ask about the energy changes observed in this reaction.

d) Heat and light are observed, this is an exothermic reaction.

3. Understanding activation energy using a brick for an object lesson:

a) Ask about the amount of energy stored in the reactants and products in the exothermic reaction.

b) Show the relative amounts of energy on a graph.

c) Use different positions of the brick to illustrate its energy states. When the brick is standing on one small end, it has the most potential energy (representing the unburned match). When it is lying flat, it has less potential energy (representing the burned match). (Actually, there are three possible states for the brick. I use only two for simplicity.)

d) Ask why the brick doesn't fall down spontaneously. The students will suggest that you need to tap the brick to make it fall. The energy released as the brick falls can be used to smash the noodle with a satisfying crunch.

e) Tapping the brick adds energy. Tapping too lightly rocks the brick but doesn't knock it over. A certain minimal amount of energy is needed; this corresponds to the activation energy. This added energy can be shown on the graph as the "hump" of activation energy.

f) The endothermic reaction can also be illustrated with the brick. The brick can be moved from the flat position (lower potential energy) to the on-end position (higher potential energy) only by applying a continuous push or pull.

Assessment Ideas

The students write a letter to "Sandy Sixth-Grader" explaining why Dad has to light the charcoal before he can barbecue the hamburgers.

Students work in their pairs to illustrate exothermic and endothermic reactions using a different object.

Equilibrium and Kinetics
Activity #2: "Chemical Equilibrium"

Questions to be investigated

What is a reversible reaction?

What does it mean for a reaction to reach equilibrium?

What happens when pressure is increased on a gas-phase equilibrium?

Objectives

The student will:

- distinguish between reactions that go to completion and those that are reversible.
- explain the concept of chemical equilibrium.
- understand how Le Chatelier's Principle works on a chemical reaction at equilibrium.

Teacher notes

This activity represents a reaction beginning with only reactants, and as the reaction proceeds, products are made. While the products are being made, the reaction is simultaneously occurring in the reverse direction. Equilibrium is achieved when the reaction is not making progress in either direction.

This activity is geared toward all high school students.

Materials

Matches
Any toy that changes colors

Two equal sized jars (or two 2000 mL beakers)
Two medium plastic cups
One smaller plastic cup

Four or five sets of:

- 2 plastic cups (may or may not be same size)
- 2 eyedroppers (or 2 straws of different sizes)
- Plain water and colored water

Three syringes or sealed tubes containing NO_2-N_2O_4 at equilibrium
Beaker full of ice
Beaker full of boiling water
Empty beaker
Hot plate

Safety Concerns

All standard safety precautions should be taken when handling chemicals: wear safety goggles, tie hair back, secure loose clothing.

Real-World Connections

Natural cycles like the nitrogen cycle and the oxygen-carbon dioxide cycle are examples of equilibria. Virtually all reactions in closed systems achieve equilibrium. It is important for students to understand equilibrium in order to understand weak acids and weak electrolytes.

Sources

This activity is taken from the SMILE Project website, http://www.iit.edu/~smile/ch9116.html. It was submitted by Nancy Zipprich of Dwight D. Eisenhower High School in Blue Island, IL.

Procedure

1. Demonstrate how a toy (such as Hot Wheels that change color) will change color in response to different temperatures. Hopefully this will instill curiosity about reversible reactions.
2. Burn a match and get the class to realize this is a reaction that has gone to completion. Explain other reactions that go to completion (for example, complete combustion, rusting, and decomposition) and put a sample equation of this type on the board, discussing its one way arrow and arrangement.

   match + oxygen → ash + carbon dioxide + water + HEAT

3. Review the placement of energy on the right side as meaning an exothermic (energy releasing) reaction and energy on the left as an endothermic (energy absorbing) reaction.
4. Perform the water scooping demonstration using two large jars (or beakers) and either the same size or different size cups to scoop with.
   a. The water level starts out filled on one side and empty on the other. Scoop a nearly full amount in the cups from each jar and pour into the other jar.
   b. Before you begin, ask the students to make predictions about what will happen.
   c. Continue until equilibrium is reached.
   d. After their predictions have been verified or disproven, draw a comparison between this demo and a reversible reaction that has reached chemical equilibrium. No water spilled = a closed system. Equal scooping technique in opposite directions means the forward and reverse reactions proceed at the same rate. Eventually the water levels do not change. The levels are not necessarily equal.
5. Based on this demonstration define chemical equilibrium. Explain that under specific conditions; nearly all chemical reactions are reversible (when in a closed system). Chemical equilibrium usually occurs in all gaseous and all aqueous systems. (Irwin Talesnick, **Idea Bank Collection**, Vol I, Idea #1).

6. Have a few students come to the front of class and do similar experiments, but this time using two smaller beakers or transparent cups (they do not have to be the same size).
    a. One cup will have plain water in it and the other will have colored water.
    b. Mark original water levels.
    c. Water transfer will be accomplished by using droppers (identical technique of transfer, as in last demo).
    d. Have the students tell when they have reached equilibrium and why. (Irwin Talesnick, **Idea #284**)
7. Demonstrate a real chemical equilibrium using the NO_2-N_2O_4 gas tubes in both cold and hot water.

$$\underset{\text{brown}}{NO_{2(g)}} \rightleftharpoons \underset{\text{colorless}}{N_2O_{4(g)}} + \text{heat}$$

8. At this point introduce the concept that a chemical equilibrium's position can be shifted by certain factors. Temperature is one of them. Define Le Chatelier's Principle in relation to this demo. According to Le Chatelier, equilibrium systems can also be stressed so that they shift to relieve this stress. Changes in pressure and changes in concentration of reactant or product are two ways to stress this closed system. (This gas demo is in most H.S. Chemistry texts.)
9. EMPHASIZE that whatever is done to stress a system at equilibrium, the system tries to relieve the stress by doing the exact opposite.

Assessment Ideas

Ask the class for examples of phenomena or toys that could be examples of equilibria. For homework, ask them to design a "controlled" paper wad fight that would simulate the idea of establishing chemical equilibrium. The best idea will be carried out the next day. This is a great learning experience, but pick a class that you really trust!

Equilibrium and Kinetics
Activity #3: "Reaction Rates"

Questions to be investigated

How is the rate of a reaction affected by temperature, concentration of aqueous reactants, pressure of gaseous reactants, surface area of solid reactants, and the presence of catalysts?

Objectives

Students will be able to describe how the rate of a chemical reaction changes as the conditions of the reactants are changed.

Teacher Notes

This experiment is potentially messy. The teacher should rehearse it and prepare the facility for the most easily managed lab experience for the students. It can be done either as a lab activity or as a classroom demonstration.

The hydrogen peroxide should be fresh. 3% is the standard concentration sold in stores.

It is an inquiry-based activity designed for all high school students. A teacher may need to do this as a demo to control the cost.

Materials

For each experiment:

- One packet dry yeast
- One pint of 3% hydrogen peroxide
- Four clear kitchen storage bags (quart size)
- Four small vials with caps
- Matches
- Object that produces glowing embers like a wood splint

Safety Concerns

All standard safety precautions should be taken when handling chemicals: safety goggles tie hair back, secure loose clothing.

Real-World Connections

Various factors will affect the rate of reactions. This is especially noted in sudden explosions due to the presence of flammable powders (grain elevator explosions) and the high pressure of gases in an engine cylinder.

Sources

This activity is taken from the SMILE Project website, http://www.iit.edu/~smile/. It was submitted by Kenneth Schug of Illinois Institute of Technology in Chicago, IL.

Procedure

CONCENTRATION:

1. Place ¼ packet of yeast in vial and attach cap firmly.
2. Pour 100 mL hydrogen peroxide into bag #1, add the vial of yeast, push out most of the air and seal (Ziploc or twist-tie).
3. Fill bag #2 with air by swooping through the air and seal.
4. Keeping bag #1 sealed, unscrew cap of vial and mix contents. (Foaming will occur as oxygen gas is formed by the reaction $2H_2O_2 \xrightarrow{catalase} 2H_2O + O_2$ and the bag will inflate.)
5. Light the splint to glowing and press against bag #2; the plastic will melt to form a small hole.
6. Repeat with bag #1 (placed on a fireproof surface). The bag will burst into flame illustrating the effect on combustion of the higher oxygen concentration (about 100% compared with 20% in air).

TEMPERATURE:

1. Prepare two of bag #1 (see CONCENTRATION section above).
2. Cool one by immersing in cold water (or an ice bath) for several minutes, then uncap both vials and observe rate of gas formation.
3. Optional extension: repeat but heat one of the bags by immersing in hot or boiling water for several minutes. Keep in mind that high temperatures will eventually "destroy" the ability of the enzyme to catalyze the reaction; a phenomenon which seldom happens with non-biological enzymes such as manganese dioxide or iron (III) ("ferric”) salts.

SURFACE AREA:

1. Prepare two bag #1s (see CONCENTRATION) but in one case grind the yeast before putting it in the vial.
2. Expect a faster reaction with "ground up" yeast because a larger surface area of yeast is available for the reaction to occur.

CATALYSIS:

The effect of the yeast on the hydrogen peroxide, which otherwise shows no sign of reaction, is due to the presence of biological catalysts or enzymes (one of which is called *catalase*).

Assessment Ideas

Ask students to predict effect on chemical reactions they observe in everyday life, or read about in the newspapers. Some examples are:

A) no reaction occurs when the gas supply to a Bunsen burner is turned on unless a source of high temperature (e.g. a match) is present;

B) a pile of flour will not burn very readily, but in a grain elevator the same flour dispersed in air (which contains oxygen) will explode or burn if a spark or flame is present because a much larger surface area is in contact with oxygen;

C) hospital rooms in which patients are receiving oxygen have NO SMOKING signs because the increased concentration of oxygen makes combustion much more likely;

D) the starch in our diet would not provide us with energy if our bodies did not have contain many different enzymes to first break the large starch molecules into glucose for absorption into the bloodstream and later to release the energy of the glucose in our cells to give us energy.

Equilibrium and Kinetics
Activity #4: "Stress on Equilibrium"

Questions to be investigated

How does varying the concentrations of reactants or products affect a reaction at equilibrium?

How does varying the temperature of a reaction at equilibrium affect the equilibrium position?

Objectives

Students will be able to predict shifts in equilibrium position when the concentrations and temperature are changed.

Teacher Notes

This activity works very well for helping students to conceptualize Le Chatelier's Principle. The colors of reactants and products allow the students to see the equilibrium shift.

The presence of complex ions and the precipitation step makes this activity more suited for the higher level chemistry students, but it could certainly be used with general science students.

Materials

0.1 M cobalt (II) chloride hexahydrate
0.1 M silver nitrate
12 M hydrochloric acid
Hot water bath
Cold water bath
Droppers
Test tubes

Safety Concerns

All standard safety precautions should be taken when handling chemicals: wear safety goggles, tie hair back, secure loose clothing. Please be especially careful with the concentrated hydrochloric acid, keeping it in a fume hood if possible.

If any of the chemicals in this experiment contact skin or clothing, it should be washed off with plenty of water immediately.

Real-World Connections

Students should be able to see that when a reaction is at equilibrium it can be forced to go to completion if one of the products is continuously removed. One common example is the Haber process for the high-yield production of ammonia. The gaseous ammonia is cooled and liquefied (removing the product) which keeps the reaction going to the right to produce more ammonia.

Sources

This activity was taken from a high school textbook, Addison-Wesley Chemistry, 2000.

Procedure

The following reaction is used in this experiment:

$$\underset{\text{(pink)}}{Co(H_2O)_6^{+2}{}_{(aq)}} + 4Cl^-{}_{(aq)} + \text{heat} \rightleftharpoons \underset{\text{(blue)}}{CoCl_4^{-2}{}_{(aq)}} + 6H_2O_{(l)}$$

Before the experiment, students should predict the direction of shift from these stresses:

- Adding water
- Removing Cl^- ions
- Heating the reaction
- Cooling the reaction

1. Measure 5 mL of 0.1 M $CoCl_2$ solution into two test tubes. Add 5 mL of 12 M HCl to each of the test tubes.
2. To test tube #1, add 5 mL of distilled water. Record observations.
3. Insert test tube #1 into a hot water bath for several minutes. Record observations.
4. Insert test tube #1 into a cold water bath for several minutes. Record observations.
5. Add 2 mL of 0.1 M $AgNO_3$ to test tube #2. Record observations.
6. Rinse the contents of both test tubes down the drain with plenty of water.

Assessment Ideas

Students should write how Le Chatelier's Principle is illustrated by every step of the procedure.

An extension question: Acetone is often used to remove the last traces of water from glassware. What will happen to this reaction at equilibrium if acetone is added?

Equilibrium and Kinetics

Activity #5: "Iodination of Acetone"

Questions to be investigated

How is a rate law written?

How does varying the concentrations of reactants change the rate of a reaction?

Objectives

The student will be able to determine the orders for the reactants, the rate expression, and the rate constant based on experimental data for the reaction between iodine and acetone.

Teacher Notes

This is an activity geared toward advanced chemistry students.

Here is an introduction for students:

The rate at which a chemical reaction occurs depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of possible catalysts. In this experiment you will study the kinetics of the reaction between iodine and acetone in acid solution:

$$CH_3\text{–}CO\text{–}CH_3 + I_2 \rightarrow CH_3\text{–}CO\text{–}CH_2I + H^+ + I^-$$

For this reaction you will determine the order of the reaction with respect to acetone, HCl, and I_2 and find a value for the rate constant, k. Since the concentrations of acetone and HCl are much higher than that of I_2, the concentrations of acetone and HCl will change very little. Thus the rate will be determined by the time needed for iodine to be used up. Iodine has color so you can easily follow changes in iodine concentration visually. The equation, rate = $k(A)^m(H^+)^n(I_2)^p$, can be simplified to rate = $k[I_2]/t$ since the values for acetone and HCl essentially remain constant during the course of any run.

Materials

4.0 M acetone solution	125 mL Erlenmeyer flasks
1.0 M HCl solution	10 mL graduated cylinders
0.0050 M iodine solution	watch or other timing device
100 mL beakers	watch glass covers for beakers

Safety Concerns

All standard safety precautions should be taken when handling chemicals: safety goggles, tie hair back, secure loose clothing.

Sources

This activity is taken from the site provided by the Pennsylvania Science In Motion and Advancing Science Gateway: http://services.juniata.edu/ScienceInMotion/chem/labs/ap/acetone.doc.

Procedure

1. Fill clean, dry 100 mL beakers with 4.0 M acetone, 0.0050 M iodine, and 1.0 M HCl solutions. Keep the first two beakers covered, as the concentration may change with evaporation.
2. For the first trial, place 10.0 mL of acetone solution, 10.0 mL of 1.0 M HCl, and 20.0 mL of distilled water in an Erlenmeyer flask. Using another graduated cylinder, measure 10.0 mL of the iodine solution.
3. Noting the time on a watch or wall clock to the nearest second, pour the iodine solution into the flask and gently stir the contents. Holding the flask over a white sheet of paper, note the time when the last trace of color disappears. Repeat. The times should agree within a few seconds.
4. Note that the total volume of the reaction mixture was 50.0 mL. Devise another reaction mixture in which only the volumes of water and acetone are changed. Keep the amounts of HCl and iodine the same. Repeat.
5. Again, keeping the total volume of the reaction mixture constant, vary the volume of the HCl used. The volumes of acetone and iodine should be the same as in the first trial. Repeat.
6. For the fourth trial, vary the volume of the iodine solution, keeping the volumes of acetone and HCl the same as in the first trial. Repeat.

Reaction Rate Data

Trial	Volume Acetone	Volume HCl	Volume Iodine	Volume H_2O	Time 1st Run	Time 2nd Run	Average Time
1	10 mL	10 mL	10 mL	20 mL			
2							
3							
4							

Determination of Orders

$$\text{Rate} = k[\text{acetone}]^m[I_2]^n[H^+]^p$$

Trial	[acetone]	$[H^+]$	$[I_2]$	$\text{Rate} = \frac{[I_2]}{\text{avg. time}}$
1				
2				
3				
4				

Analysis:

1. Determine the order of acetone, hydrogen, and I_2.
2. Give the rate law for the reaction.
3. Give the rate constant (k) for each trial, and then give the average k value.
4. Why is the concentration of iodine so much less than the other reactants? (observation of the time of reaction depends on the amount of iodine depleting.)
5. How are the time and rate related? (inversely) How are 1/time and rate related? (directly)
6. In a reaction $A + B \rightarrow C$, it is found that the reaction is first order in terms of A and B. What happens to the rate if the concentrations of A and B are doubled? (rate is quadrupled)

Assessment Ideas

Run one more trial, use amounts <u>other than</u> 10 and 20 mL for each reactant with the total volume remaining at 50 mL. Predict the rate of the reaction and compare to the experimental rate.

Equilibrium and Kinetics

Activity #6: "K_a of a Weak Acid"

Questions to be investigated

How can the equilibrium constant be determined from concentrations of reactants and products at equilibrium?

Objectives

Students will:

- compare the conductivities of solutions with known and unknown hydronium concentrations.
- relate the conductivity to the concentration of ions in solution.
- explain the difference between a strong and a weak acid.
- calculate the value of K_a for a weak acid.

Teacher Notes

A dissolved weak acid is an example of a reaction at equilibrium. Not all of the acid molecules liberate a hydrogen ion. In contrast, in a strong acid virtually all of the acid molecules release a hydrogen ion. Because weak acids don't completely ionize their conductivity will indicate the extent to which they do ionize. In this experiment students will measure the conductivity of a weak acid compared with the conductivities of varying concentrations of a strong acid.

This activity requires an unusual piece of equipment that will require some planning. The conductivity will be measured qualitatively with either a LED attached to a 9V battery or an auditory buzzer attached to a 9V battery.

- If the LED is used, the circuit should include a $1k\Omega$ (¼W) resistor.
- LAB-AIDS sells the buzzer conductivity tester for around $25.

To save money, you may want to see if your physics department has this equipment – they may have a class set to let you borrow.

Materials

1 M acetic acid
1 M hydrochloric acid
24-well plate
Distilled water
Conductivity tester
Pipettes

Safety Concerns

All standard safety precautions should be taken when handling chemicals: safety goggles, tie hair back, secure loose clothing.

Real-World Connections

Weak acids, weak electrolytes, and equilibrium apply to many facets of health and nutrition. For example, there is a reason that we can eat acetic acid, citric acid, stearic acid, etc. but not sulfuric, hydrochloric, or nitric acids.

Sources

This experiment is taken from Modern Chemistry published by Holt, Rinehart, and Winston, 2006.

Procedure

1. Students should construct a data table like this:

[HCl]	Observations / Comparisons

2. Place 20 drops of 1.0 M HCl solution into one well of a 24-well plate. Place 20 drops of 1.0 M acetic acid solution into an adjacent well. Label the location of each sample.
3. Test the HCl and CH_3COOH with the conductivity tester. Note the relative intensities of each sample. Rinse the tester probes with distilled water and clean with a paper towel.
4. Place 18 drops distilled water into each of six wells. Add 2 drops of 1.0 M HCl to the first well to make a total of 20 drops of solution. Mix the solution by drawing them up into the dropper and returning to the well.
5. Repeat the procedure by taking two drops of the previous dilution and placing it in the next well containing 18 drops of water. Return any unused solution in the pipet to the well from which it was taken. Mix the new solution with a new pipet. (You now have 1.0 M HCl in the well from step 2, 0.10 M in the first dilution well, 0.01 M in the second dilution well).
6. Continue diluting in this manner until you have six successive dilutions. The $[H_3O^+]$ wells should now range from 1.0 M to 1.0 x 10^{-6}M. Write the concentrations in the first column of your data table.
7. Using the conductivity tester, test the cells containing the HCl in order from most to least concentrated. Note the brightness (or loudness) of the conductivity tester and compare each concentration with the brightness (or loudness) of the tester when in the acetic acid. Retest the acetic acid frequently for comparison. After each test, rinse the tester probes with distilled water and use a paper towel to remove excess moisture.

8. When the brightness (or loudness) of one of the HCl solutions is about the same as the acetic acid, you may assume that the two solutions have a similar hydronium concentration. Record that HCl concentration. NOTE: If the conductivity of no HCl concentration matches the acetic acid, estimate between the two HCl concentrations that match the best.
9. Clean all equipment. The acid solutions may be dumped down the drain with plenty of water.

Analysis Questions

1. How did the conductivity of 1.0 M HCl compare with that of 1.0 M CH_3COOH? Why is this so?
2. What concentration of HCl most closely matches the conductivity of the 1.0 M CH_3COOH?
3. What is the approximate concentration of the H_3O^+ in 1.0 M acetic acid?
4. Write the K_a expression for the ionization of CH_3COOH. Use your answer to #3 to calculate K_a for acetic acid.
5. How is it possible for solutions of HCl and CH_3COOH to have different concentrations but the same conductivity?
6. Look up the K_a value for acetic acid in your book. Calculate your percent error.
7. Lactic acid ($CH_3CHOHCOOH$) has a K_a value of 1.4×10^{-4}. Would a 1.0 M solution of lactic acid cause the conductivity tester to be brighter (or louder) than 1.0 M acetic acid? How noticeable would the difference be? Explain.