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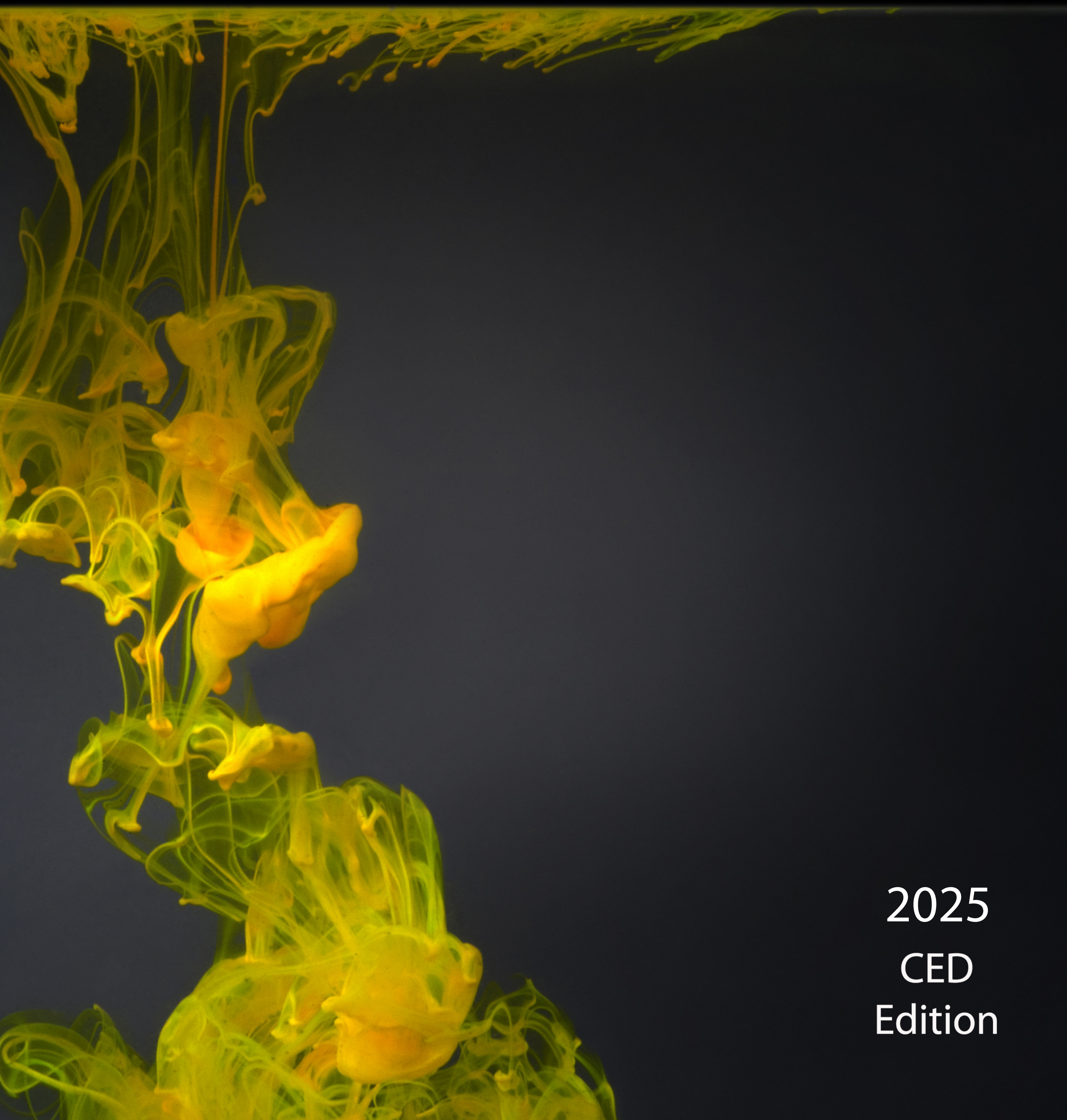
AP Chemistry 2



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CED

Edition

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AP Chemistry 2

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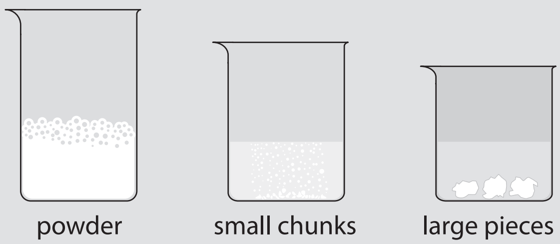
1.2 Factors Affecting Rates of Reaction

Warm Up

Compare the following reactions. Circle the one with the greatest reaction rate. Indicate what is causing the rate to be greater in each case.

(a) Solid marble chips (calcium carbonate) reacting with 3.0 M hydrochloric acid at room temperature produces calcium chloride solution, water and carbon dioxide gas.



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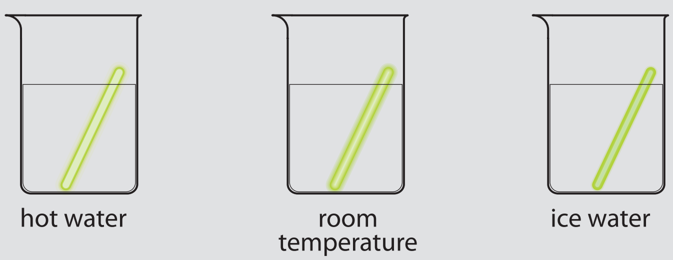
(b) Glow sticks contain two chemicals, one of which is held in a breakable capsule that floats inside the other. Once the capsule is broken, a reaction occurs that produces energy in the form of light. This is called chemiluminescence. Glow sticks are placed in different temperature water baths.



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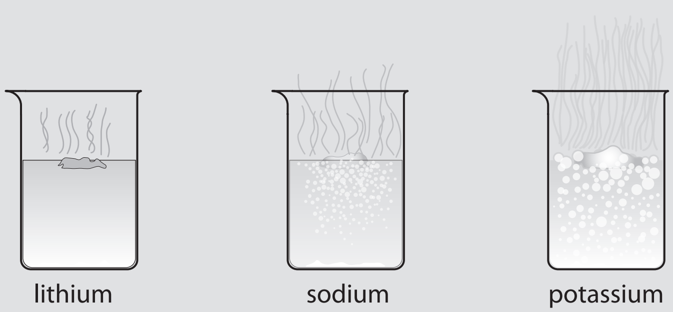
(c) Alkali metals react with water to form a basic solution and hydrogen gas. Sometimes the reaction is so exothermic, the hydrogen gas bursts into flame.



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Factors Affecting Reaction Rate

In the previous section, we discovered that reactions tend to slow down as they proceed. What might cause this phenomenon? What is true about a reacting system when it contains lots of reactants? Consideration of this question leads us to recognize there are more particles available to react when a reaction first starts. Intuition tells us in order for a reaction to occur, the reacting particles must contact each other. If this is so, anything that results in an increased frequency of particle contact must make a reaction occur faster.

The three factors of surface area, concentration, and temperature may all be manipulated to increase the frequency with which particles come together. Two of these factors were dealt with in the Warm Up. Part (c) of the Warm Up demonstrates the periodic trend in reactivity moving down the alkali metal family. Different chemicals inherently react at different rates so the nature of the reactants affects reaction rate. Finally, you may recall from Science 10 that a chemical species called a *catalyst* may be used to increase the rate of a reaction. We will consider each of these five factors in turn.

Surface Area

Most reactions in the lab are carried out in solution or in the gas phase. In these states, the reactants are able to intermingle on the molecular or atomic level and contact each other easily. When reactants are present in different states in a reacting system, we say the reaction is **heterogeneous.** Most heterogeneous systems involve the reaction of a *solid* with a solution or a gas. In a heterogeneous reaction, the reactants are able to come into contact with each other only where they meet at the interface between the two phases. The size of the area of contact determines the rate of the reaction. Decreasing the size of the pieces of solid reactant will increase the area of contact (Figure 1.2.1).

> Increasing the surface area of a solid will increase the rate of a heterogeneous reaction.

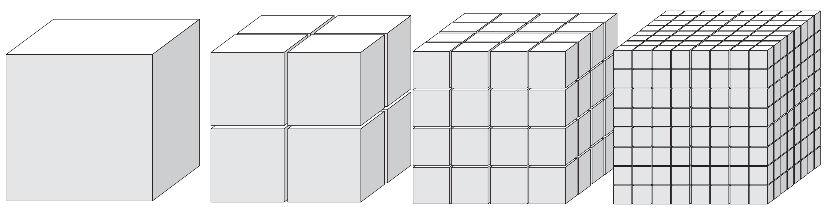


Figure 1.2.1 *Increasing the number of pieces leads to a significant increase in the surface area.*

Concentration (Pressure)

The rates of all reactions are affected by the concentrations of the dissolved or gaseous reactants. When more solute is placed in the same volume of solvent, the solution's concentration is increased (Figure 1.2.2).

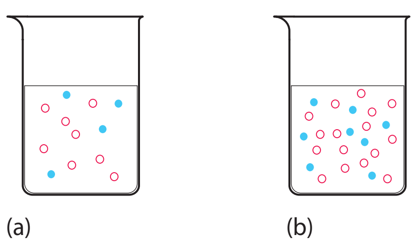


Figure 1.2.2 *Solution (b) contains twice the number of solute particles in the same volume as solution (a). The opportunity for particle contact is doubled in solution (b).*

When more gas particles are placed in the same volume of a container, the **partial pressure** of the gas has increased. In Figure 1.2.3, container (b) has twice as many gas particles in the same volume as container (a). This, of course, means the concentration has been doubled. We might also say the *partial pressure* of the gas has doubled. In container (c), the piston has been lowered to half the volume. The result is another doubling of concentration (and pressure).

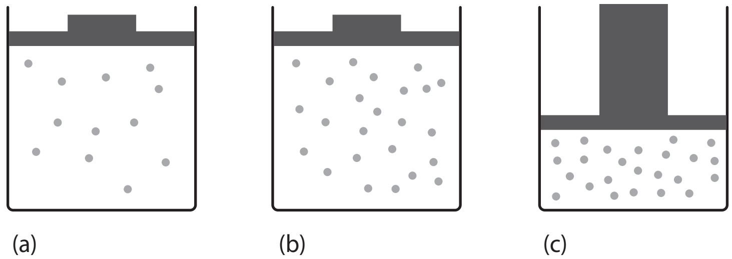


Figure 1.2.3 *The partial pressure in (b) is greater than in (a) because of the increased concentration of particles. In (c), decreasing the volume increases the concentration and the pressure.*

Increasing the concentration (or partial pressure of a gas) will increase the rate of a chemical reaction.

Always remember the concentration of pure solids and liquids cannot be increased because adding more substance increases both the moles and the liters, so the molarity or moles per liter remains constant. Also remember that crushing or breaking a solid will increase its surface area. However it is impossible to cut a piece of liquid or gas into smaller bits. The surface area of liquids can be increased by spreading them over a larger area.

Temperature

Recall the qualitative relationship between temperature and kinetic energy. Mathematically speaking, KE = $3/2RT$ where R is a constant having the value 8.31 J/mol K, and T is the Kelvin temperature. From this relationship, we see that temperature and kinetic energy are directly related to one another. If the temperature is doubled, the kinetic energy is doubled (as long as the temperature is expressed in units of Kelvin).

An increase in temperature will lead to particles striking one another more frequently. However, we now see that it will also result in the particles striking one another with more energy. In other words, an increase in temperature means that the same particles are travelling faster. As a consequence, they hit each other more frequently and more forcefully. As a result, *temperature is the most significant factor* that affects reaction rate.

Within any substance there is a "normal" distribution of kinetic energies among the particles that make up the system due to their random collisions. Such a distribution might be graphed as shown in Figure 1.2.4.

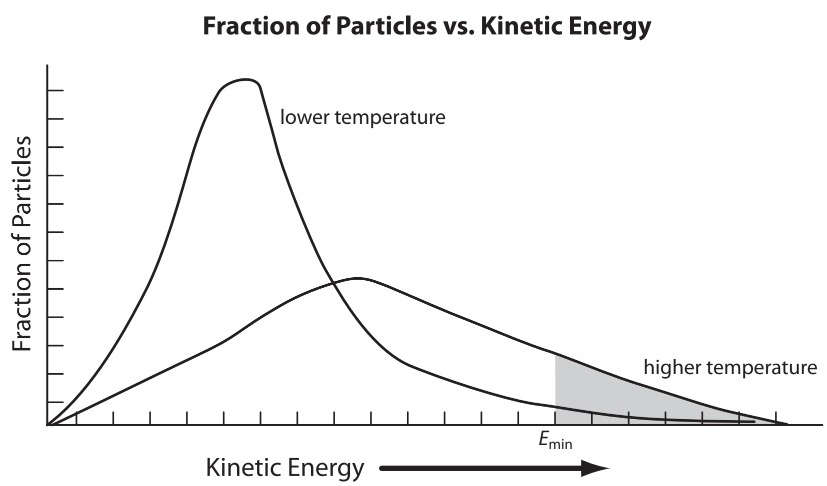


Figure 1.2.4 *Kinetic energy shows a "normal" distribution at both lower and higher temperatures.*

Note that some of the particles have very little energy and others have a lot. The *x*-axis value associated with the peak of the curve indicates the kinetic energy of most of the particles. The second curve indicates how the distribution would change if the temperature were increased. The area under the curves represents the total number of particles and therefore should be the same for both curves. The gray area represents the particles that have sufficient energy to *collide successfully* and produce a product. Notice that this has increased with an increase in temperature. A common generalization is that an increase of 10°C will double reaction rate. This is true for some reactions around room temperature.

Increasing temperature will increase the rate of a reaction for *two* reasons: *more frequent* and *more forceful* collisions.

The Nature of Reactants

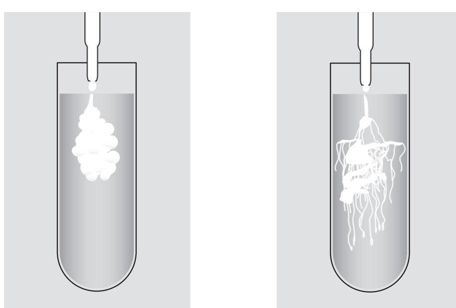
Fundamental differences in chemical reactivity are a major factor in determining the rate of a chemical reaction. For instance, zinc metal oxidizes quickly when exposed to air and moisture, while iron reacts much more slowly under the same conditions. For this reason, zinc is used to protect the integrity of the iron beneath it in galvanized nails.

Generally reactions between simple monoatomic ions such as Ag^+ and Cl^- are almost instantaneous. This is due to ions being extremely mobile, in close proximity to one another, having opposite charges, and requiring no bond rearrangement to react. However, more complicated ionic species such as CH_3COO^- react more slowly than those that are monoatomic.

In general, differences in chemical reactivity can be attributed to factors that affect the breaking and forming of chemical bonds. Ionization energy, electronegativity, ionic and molecular polarity, size, and complexity of structure are some of these factors. The state of the reacting species may also play a role.

In general, at room temperature the rate of (*aq*) reactants > (*g*) > (*l*) > (*s*).

Precipitation occurs quickly between ions in solution. In Figure 1.2.5(a), the AgCl forms as a heavy, white precipitate the instant the silver ions contact the chloride in solution. In (b), the $AgCH_3COO$ precipitate is less heavy and forms over a period of 10 s to 20 s. The complex structure of the acetate ion makes it more difficult to achieve the correct orientation to bond successfully.



(a) (b)

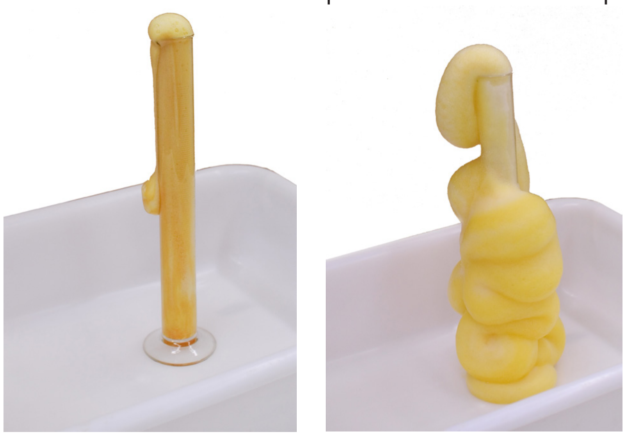
Figure 1.2.5 *(a) AgCl quickly forms a precipitate. (b) $AgCH_3COO$ forms a precipitate more slowly.*

Presence of a Catalyst

Catalysts are substances that increase the rates of chemical reactions without being used up. Because they remain in the same quantity and form when a reaction is completed, the formulas of catalysts are not included in the chemical reaction. Sometimes the formula is shown above the arrow between the reactants and products like this:

$$2\,H_2O_2(l) \xrightarrow{MnO_2} 2\,H_2O(l) + O_2(g)$$

What actually happens is that catalysts are consumed during an intermediate step in a reaction and regenerated in a later step. The catalysts most familiar to you are probably the enzymes produced by living organisms as they catalyze digestive and other biochemical processes in our bodies. The reaction depicted in graduated cylinders in Figure 1.2.6 is the catalyzed decomposition of hydrogen peroxide as shown in the equation above. In addition to the H_2O_2, there is a bit of dish soap and some dye in the cylinders so the oxygen gas bubbles through the soap solution and produces foam. This demonstration is often called "elephant toothpaste." Note that the cylinder in (b) contained 30% hydrogen peroxide while the cylinder in (a) was only 6%, so the effect of concentration was demonstrated in addition to the catalytic effect.



(a) (b)

Figure 1.2.6 *"Elephant toothpaste" is produced through a catalytic reaction. The concentration of H_2O_2 is greater in (b) than in (a).*

> A catalyst increases reaction rate without itself being consumed or altered.

An **inhibitor** is a species that reduces the rate of a chemical reaction by combining with a reactant to stop it from reacting in its usual way. A number of pharmaceuticals are inhibitors. Drugs that act through inhibition are called *antagonists*.

Sample Problem — Factors Affecting Reaction Rate

Which of the following reactions is faster at room temperature?

(a) $H_2(g) + I_2(s) \rightarrow 2\,HI(g)$ (b) $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

List two ways to increase the rate of each reaction.

What to Think About	How to Do It
1. First consider the nature of the reactants.	The reactant states lead us to believe that reaction (b) involving aqueous species would be the fastest.
2. To increase the rate of reaction (a), start by recognizing this is a heterogeneous reaction. This means that, in addition to the usual factors, surface area can be considered.	• Increase surface area of iodine solid. • Increase temperature. • Increase concentration of hydrogen gas. • Increase partial pressure of hydrogen gas (decrease container volume). • Add an appropriate catalyst.
3. To increase the rate of reaction of the homogeneous reaction in (b), apply all the usual factors except surface area. *Note that you must be specific when mentioning the factors.* For example, what species' concentration will be increased?	• Increase temperature. • Increase concentration of either or both reactant ions (Ba^{2+} and/or SO_4^{2-}). • Add an appropriate catalyst.

1.2 Activity: Graphic Depiction of Factors Affecting Reaction Rates

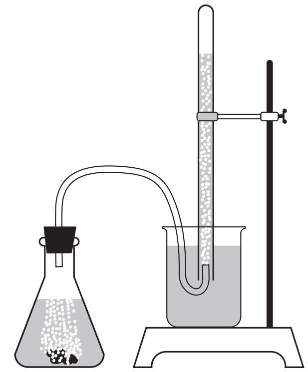
Question

How do concentration, surface area, and temperature affect reaction rate?

Background

Four trials were carried out in which a chunk of zinc was reacted with hydrochloric acid under four different sets of conditions. In all four trials, the chunk of zinc was of equal mass. The data collected indicate that varying factors have a significant impact on reaction rate. These data can be represented in tabular and graphical form.

The reaction was allowed to proceed for the same time period in each trial. In all four trials, the gas was collected in a eudiometer using the apparatus shown below.

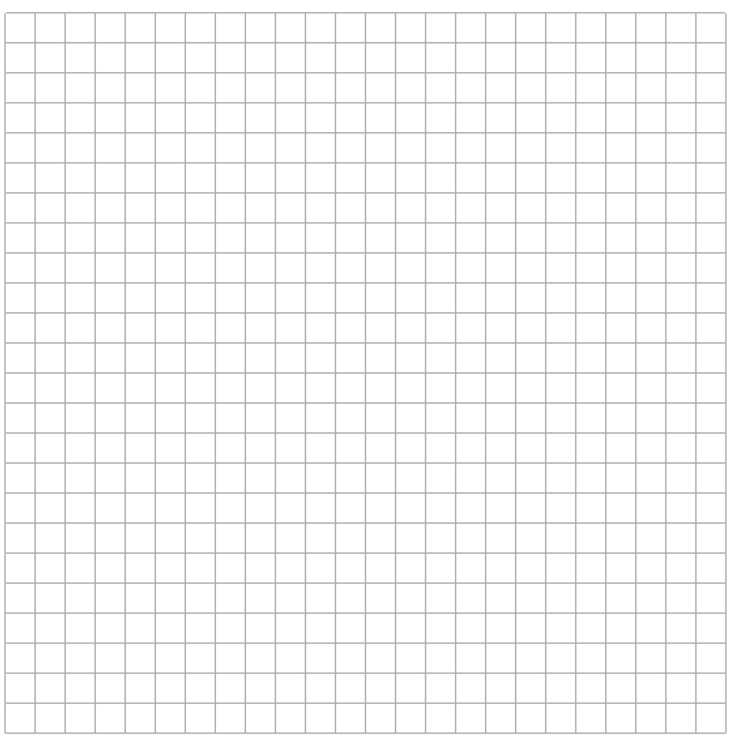


Time (s)	Trial 1 25°C, 1 M HCl (mL)	Trial 2 50°C, 1 M HCl (mL)	Trial 3 25°C, 2 M HCl (mL)	Trial 4 25°C, 1 M HCl, zinc powder (mL)
0	0	0	0	0
30.	12	34	20	26
60.	19	56	34	44
90.	24	65	44	57
120.	27	65	51	65
150.	29	65	54	65

Procedure

1. Use the following grid to graph all four sets of data. Vary colors for each trial line.

Results and Discussion



1. Write a balanced equation for the reaction that was studied.

2. Rank the conditions from those producing the fastest to slowest reaction rates. What factor influences reaction rate the most? Which factor is second most effective? Which factor has the least influence on the rate?

3. Calculate the average reaction rate in mL H_2/min for each trial. Use the time required to collect the maximum amount of hydrogen gas formed (e.g., the time for trial 2 will be 90. s). Place the rates on the graphed lines.

4. How many milligrams of zinc were used in 1.50 min for each trial? Assume a molar volume for $H_2(g)$ of 24.5 L/mol. (This is for SATP conditions or 25°C and 101.3 kPa or the pressure at sea level.) For the 50°C trial, use 26.5 L/mol.

5. What error is introduced by the assumption in question 4? Would the actual mass of Zn be larger or smaller than that calculated? Explain.

1.2 Review Questions

1. Identify the four factors that affect the rate of any reaction. Give a brief explanation as to how each one applies. Which of these factors can be altered to change the rate of a particular chemical reaction?

2. Identify the one factor that affects only the rate of heterogeneous reactions. Explain why it does not affect homogeneous reaction rates.

3. Use the Internet to find examples of catalysts that do the following:
   (a) Convert oxides of nitrogen into harmless nitrogen gas in the catalytic converter of an automobile.

   (b) Increase the rate of the Haber process to make ammonia.

   (c) Found on disinfectant discs to clean contact lenses.

   (d) Found in green plants to assist in photosynthesis.

4. How would each of the following changes affect the rate of decomposition of a marble statue due to acid rain? Begin by writing the equation for the reaction between marble (calcium carbonate) and nitric acid below.

   (a) The concentration of the acid is increased.

   (b) Erosion due to wind and weathering increases the surface area on the surface of the statue.

   (c) The statue is cooled in cold winter weather.

   (d) The partial pressure of carbon dioxide gas in the atmosphere is increased due to greenhouse gases.

5. (a) At room temperature, catalyzed decomposition of methanoic acid, HCOOH, produced 80.0 mL of carbon monoxide gas in 1.00 min once the volume was adjusted to STP conditions. The other product was water. Calculate the average rate of decomposition of methanoic acid in moles per minute.

(b) Give general (approximate) answers for the following:

(i) How long would you expect the production of 40.0 mL of gas to take?

(ii) How long would you expect the production of 80.0 mL to take without a catalyst?

(iii) How long would you expect the production of 80.0 mL of gas to take at 10ºC above the experimental conditions?

6. Answer the questions below for each of the following reactions:

(i) $C(s) + O_2(g) \rightarrow CO_2(g)$

(ii) $Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$

(iii) $Mg(s) + CuCl_2(aq) \rightarrow MgCl_2(aq) + Cu(s)$

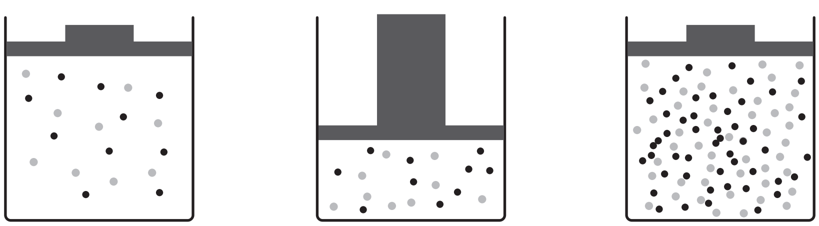
(a) Indicate whether you think it would be fast or slow if performed at room temperature. Then rank the three reactions from fastest to slowest.

(b) List which of the five factors could be used to increase the rate of each reaction.

7. Rank the diagrams below in order of expected reaction rate for this reaction:

$G(g) + B(g) \rightarrow GB(g)$

where G = gray, B = black and GB is the product. Explain your ranking. Assume the same temperature in all three reacting systems.

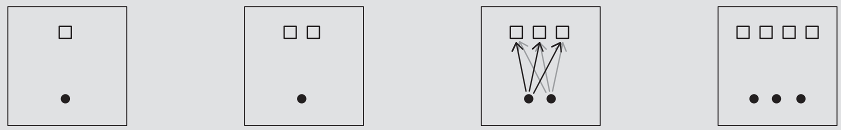


(a) (b) (c)

1.3 Collision Theory

Warm Up

Consider a hypothetical reaction between two gas particles, A and B, to form AB according to the reaction, A(g) + B(g) → AB(g). Determine the total number of possible distinct collisions that could occur between the A particles and the B particles in each situation shown in the diagram below. Let □ represent A and ● represent B. Draw arrows from each A particle to each B particle in turn to help you track each possible collision. The third case is done as an example for you.



Total possible collisions: ________ ________ ___6___ ________

What is the relationship between the number of A and B particles and the total number of distinct collisions possible?

____

What might this relationship indicate regarding the rate of a reaction between A and B?

____

Collisions and Concentrations

From our study of the factors affecting reaction rates, it seems obvious that a successful reaction requires the reacting particles to collide with one another. The series of diagrams in the Warm Up above show that the more particles of reactant species present in a reacting system, the more collisions can occur between them in a particular period of time. Careful examination of the situations described in the Warm Up reveals that the number of possible collisions between the reactant molecules, A and B, is simply equal to the product of the number of molecules of each type present. That is:

number of A × number of B = total possible collisions between A and B

The number of particles per unit volume may be expressed as a concentration. In the example, all of the particles are in the same volume so we can think of the number of particles of each type as representing their concentrations. From this, we can infer that the concentration of the particles colliding must be related to the rate. These two statements allow us to represent the relationship discovered in the Warm Up as follows.

The rate of a reaction is directly proportional to the product of the concentrations of the reactants.

When two variables are directly proportional to one another there is a constant that relates the two. In other words, multiplying one variable by a constant value will always give you the other variable. We call this multiplier a *proportionality constant*. In science, proportionality constants are commonly represented by the letter *k*.

The slope of any straight-line graph is a proportionality constant. Applying this concept to the relationship discovered in the Warm Up results in the following.

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

The relationship above is called a **rate law**. The proportionality constant in this case is called a **rate constant**. Every reaction has its own unique rate law and its own unique rate constant. For many reactions, the exponents, x and y, are each equal to 1. The values of x and y are called **reactant orders**. In this section, we will focus on reactions that are first order with respect to all reactants. In other words, both $x = 1$ and $y = 1$. Be aware, however, that there are reactions involving second and even third order reactants. Clearly, the higher the order of a particular reactant, the more a change in the concentration of that reactant affects the reaction rate.

Quick Check

1. Write the general form for a rate law in which two reactants, C and D, react together. Assume the reaction is first order for each reactant. Explain what the rate law means.

 ____

 ____

 ____

2. Assume a reaction occurs by this rate law: rate = k[A]. How would the rate be affected by each of the following changes in concentration?
 (a) [A] is doubled.

 ____

 (b) [A] is halved.

 ____

3. Assume a gaseous reaction occurs by this rate law: rate = k[A][B]. How would the rate be affected if the volume of the container were halved?

 ____

4. How is the rate of a reaction affected by increasing the temperature? What part of the rate law must be affected by changing temperature?

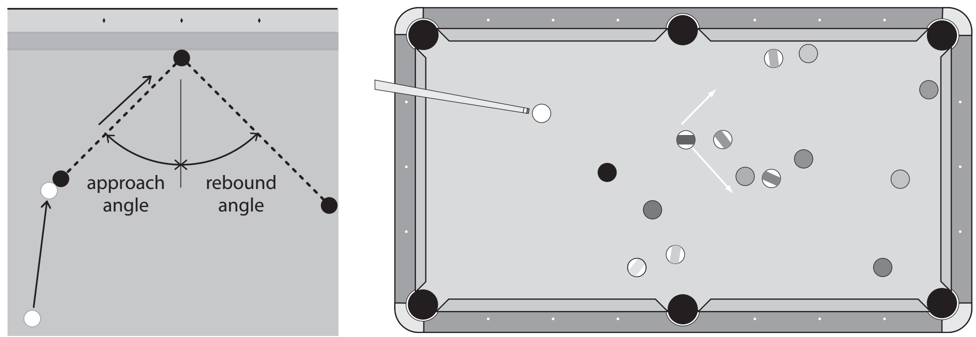
 ____

 ____

A Distribution of Collisions

When reacting particles collide with one another, each collision is unique in its energy and direction. If you've ever played a game of billiards, snooker, or 8-ball, you're familiar with the transfer of energy from the shooter's arm to the pool stick to the cue ball to the ball being shot. Only when the shooter sends the cue ball directly into the ball being shot with a controlled amount of strength are the momentum and direction of the ball easily predictable.

This control is required in a "direct shot" or in a classic "bank shot" where, much like a beam of reflected light, the angle of incidence equals the angle of reflection (Figure 1.3.1(a)). Bank shots may be necessary when an opponent's balls are blocking a clear shot. In most cases, however, collisions between the cue ball, the ball being shot, and other balls on the table are different in direction and amount of energy transferred (Figure 1.3.1(b)). Skillful players may intentionally play "shape" or put "English" on the ball to manipulate the direction and the amount of energy transfer involved in collisions.



(a) (b)

Figure 1.3.1 *(a) A bank shot involves predictable collisions. (b) Skilled players determine exactly what type of collision is required for the shot they need to make.*

The break at the start of the game is probably the best model of the distribution of collision energies involved between the reacting particles in a chemical reaction (Figure 1.3.2). Some of the particles move quickly and collide directly with others to transfer a great deal of energy. Others barely move at all or collide with a glancing angle that transfers very little energy. In addition, we must remember that real reactant collisions occur in three dimensions, while our pool table model is good for two dimensions only (Figure 1.3.3).

Figure 1.3.2 *The break at the start of a game provides a large distribution of collision energies.*

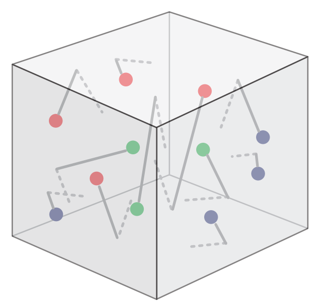


Figure 1.3.3 *The collisions occurring between species in a chemical reaction occur in three dimensions.*

Collision Theory

Regardless of the fact that the energy of particle collisions is distributed in a regular way, the rate of a reaction actually depends on *two* considerations.

Collision theory states that reaction rates depend upon:
- the number of collisions per unit time, and
- the fraction of these collisions that succeed in producing products.

Requirements for Effective Collisions

A mixture of hydrogen and oxygen gas may sit indefinitely inside a balloon at room temperature without undergoing any apparent reaction regardless of concentrations. Addition of a small amount of energy such as a spark to the system, however, causes the gases to react violently and exothermically. How can this phenomenon be explained?

In 1 mol of gas under standard conditions, there are more than 10^{32} collisions occurring each second. If the rate of collisions per second were equal to the rate of reaction, every reaction would be extremely rapid. Since most gaseous reactions are slow, it is evident that only a small fraction of the total collisions effectively result in the conversion of reactants to products.

Quick Check

Consider the factors affecting reaction rate that you studied in the previous section.

1. Which of these factors do you think increase reaction rate by increasing the number of collisions per unit time?

   ____

2. Which of the factors do you think increase reaction rate by increasing the fraction of collisions that succeed in producing products?

   ____

3. Do you think any of the factors increase *both* the number of collisions per unit time and the fraction of successful collisions? If so, which one(s)?

   ____

As discussed earlier, the tremendous number of collisions in a reacting sample results in a wide range of velocities and kinetic energies. Only a small percentage of the molecules in a given sample have sufficient kinetic energy to react.

> The minimum kinetic energy that the reacting species must have in order to react is called the **activation energy** or E_a of the reaction.

Only collisions between particles having the threshold energy of E_a are energetic enough to overcome the repulsive forces between the electron clouds of the reacting molecules and to weaken or break bonds, resulting in a reaction. Most gaseous reactions have relatively high E_a values; hence reactions like the one between hydrogen and oxygen gas do not occur at room temperature.

Reactions often occur in a series of steps called a reaction mechanism. Energy is required to break bonds and energy is released during the formation of bonds. The activation energy of a reaction often reflects the difference between these energies during the early steps of the reaction mechanism. As bond energies are a characteristic of a particular chemical species, the only thing that can change the activation energy of a reaction is to change its reaction mechanism. This is precisely what occurs when a reaction is catalyzed.

A kinetic energy distribution diagram was shown for a sample of matter at two different temperatures. Figure 1.3.4 shows a similar pair of energy distributions at two different temperatures. It also shows two different activation energies. E_{a2} represents the activation energy without a catalyst present. E_{a1} represents a lowered activation energy that results in the presence of a catalyst. Note that in both cases, with and without a catalyst, the number of particles capable of colliding effectively (as represented by the shaded area under the curve) increases at the higher temperature.

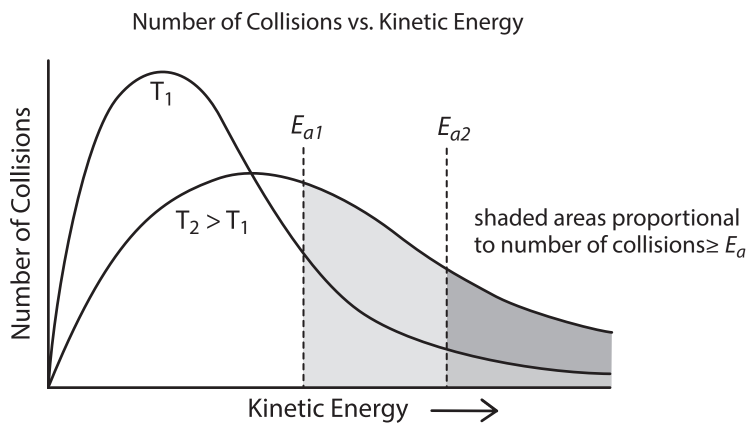


Figure 1.3.4 *With higher temperatures, the number of effective collisions increases.*

It turns out there is one further factor that complicates collision theory even more. Collisions possessing the activation energy don't always result in reaction. Even very high-energy collisions may not be effective unless the molecules are properly oriented toward one another.

In addition to attaining E_a, the *geometric shape* and the *collision geometry* of reacting particles must also be favorable for a successful collision to occur.

Consider the reaction $CO(g) + NO_2(g) \rightarrow CO_2(g) + NO(g)$

Figure 1.3.5 shows two different ways the reactants CO and NO_2 could collide. Here we see a successful collision followed by an unsuccessful collision between carbon monoxide and nitrogen dioxide in an attempt to form carbon dioxide and nitrogen monoxide. No matter how energetic the collision, it will only succeed when the molecules are oriented properly with respect to each other. For the CO molecule to become CO_2, the oxygen from the NO_2 molecule must collide with the carbon atom in the CO molecule so that a bond can form.

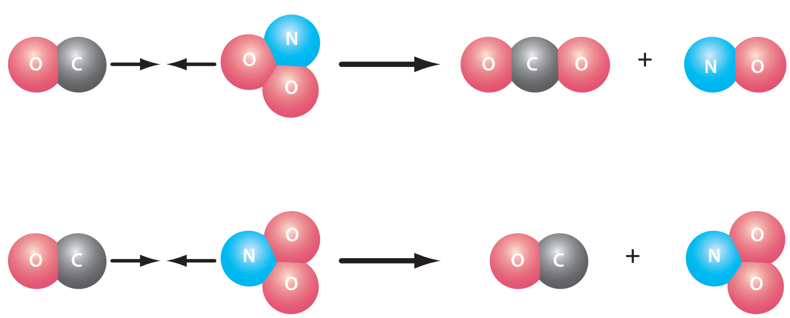


Figure 1.3.5 *The top collision is successful because the molecules are oriented so that the carbon and oxygen atoms can bond to form CO_2.*

Collision Theory and Factors Affecting Reaction Rate

Increased temperature, concentration, and surface area all increase the frequency of collisions leading to an increase in reaction rate. The presence of a catalyst and increased temperature both increase the fraction of collisions that are successful.

Sample Problem — Collision Theory

Study the following data for 1 mol of reactants at two different temperatures. $T_1 < T_2$.

	Temperature 1	Temperature 2
Frequency of collisions:	$1.5 \times 10^{32}\ s^{-1}$	$3.0 \times 10^{32}\ s^{-1}$
% of collisions > E_a:	10%	30%

(a) How many collisions exceed activation energy per second at each temperature?

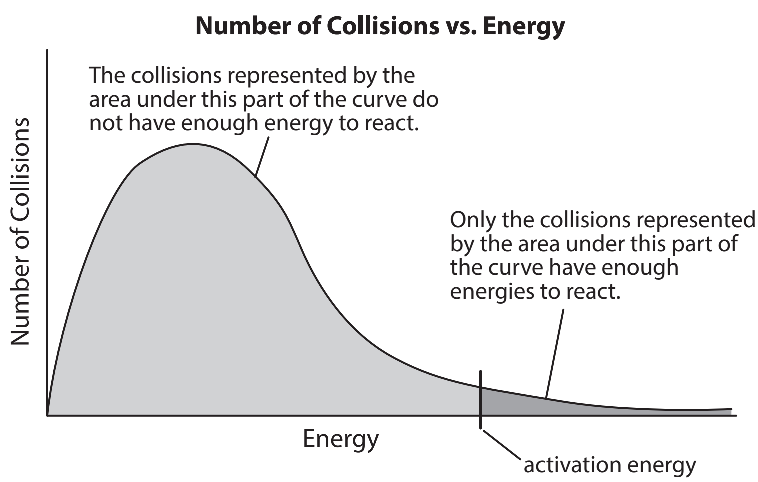
(b) How many times greater is the reaction rate at temperature 2 than temperature 1? (Assume the percent of collisions with correct geometry is the same at both temperatures.)

(c) If the frequency of collisions was only doubled by increasing the temperature from T_1 to T_2, why is the rate increased by a factor of six?

What to Think About	How to Do It
1. Multiply the percentage by the frequency at each temperature.	$0.10(1.5 \times 10^{32}\ s^{-1}) = 1.5 \times 10^{31}\ s^{-1}$ at T_1 $0.30(3.0 \times 10^{32}\ s^{-1}) = 9.0 \times 10^{31}\ s^{-1}$ at T_2
2. Divide the result for T_2 by the result for T_1.	$\frac{9.0 \times 10^{31}\ s^{-1}}{1.5 \times 10^{31}\ s^{-1}} = 6.0 \times$ greater
3. Decide what, in addition to collision frequency, must be influenced by reaction temperature.	Increasing reaction temperature increases the energy of collisions as well as collision frequency.

Practice Problems — Collision Theory

1. For slow reactions, at temperatures near 25°C, an increase of 10°C will approximately double the number of collisions that have enough energy to react. At temperatures much higher than that, increasing temperature leads to smaller and smaller increases in the number of successful collisions. In the graph below, the area under the kinetic energy distribution curve represents the total number of collisions.
   (a) Assume the curve shown represents the kinetic energy distribution for a sample of collisions at room temperature. Add a curve to represent the kinetic energy distribution for a sample of collisions at 10°C above room temperature.
   (b) Add a second curve to represent the kinetic energy distribution for a sample of collisions at a temperature 100°C above curve (a).
   (c) Add a third curve for a temperature 10°C above curve (b).



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Practice Problems (*Continued*)

2. At higher temperature ranges, that is, for reactions already occurring at a high rate, most of the collisions already possess E_a. Hence, an increase in temperature increases the rate primarily because of the increased frequency of collisions. Although the particles hit each other harder and with more energy, this is not really relevant because they have already achieved E_a. Complete the following table for a reacting particle sample.

Temperature	**100°C**	**300°C**
Frequency of collisions	$2.00 \times 10^{15}\ s^{-1}$	$3.00 \times 10^{15}\ s^{-1}$
Force of collisions (% possessing E_a)	95.0%	97.0%
Frequency of collisions possessing activation energy		

(a) How many times greater is the percentage of collisions possessing Ea at 300°C than at 100°C?

(b) How many times greater is the frequency of collisions at 300°C than at 100°C?

(c) How many times greater is the frequency of collisions possessing Ea at 300°C than at 100°C?

(d) The data in the table shows that at higher temperatures the increase in rate is almost entirely due to what?

(e) Why might increasing the force of collisions eventually produce less successful reactions?

1.3 Activity: Tracking a Collision

Question

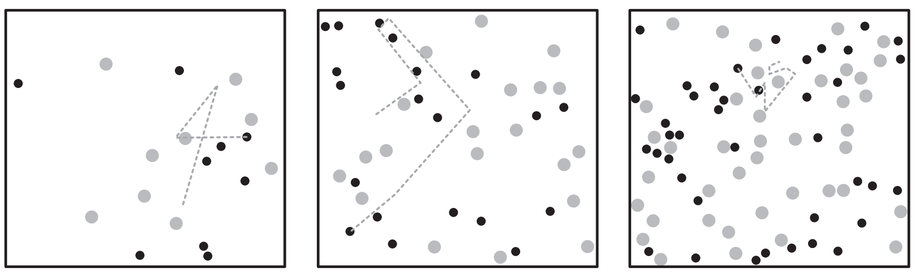
How is a single reactant particle affected by a variety of factors that impact the rate of a chemical reaction?

Background

A variety of factors may affect the rate of a chemical reaction. These factors include the nature of the reactants, the concentration of the reactants, their surface area, temperature, and the presence of a catalyst.

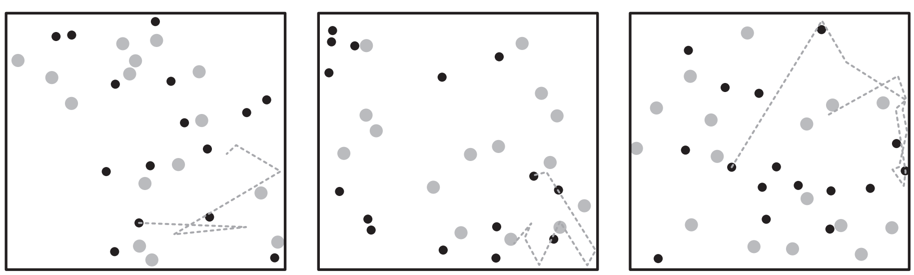
Procedure

1. The diagrams below represent snapshots taken within a nanosecond of the "life" of an ordinary gas particle. None of these particles reacts during this time, but they do a lot of colliding. Follow the pathway of an individual reactant particle as several of the factors affecting reaction rates are changed. Each time the particle changes direction it has collided with either another particle or the walls of the container.
2. While the temperature remains constant, the concentration of the reacting particle is doubled in each frame from (a) to (b) to (c).



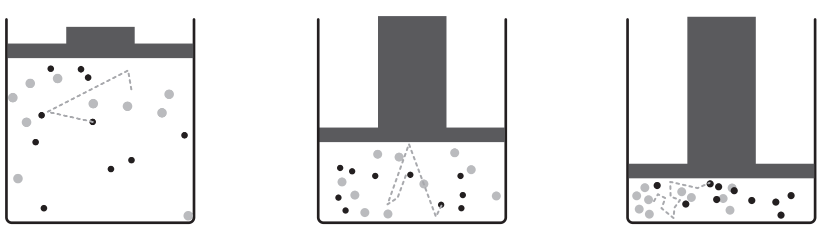
(a) (b) (c)

3. The temperature of the reaction system is increased from 25°C to 35°C to 45°C.



(a) (b) (c)

4. The volume of the reaction system is decreased by approximately half from (a) to (b) and again to (c), causing the pressure to approximately double each time. Assume the temperature remains constant.



(a) (b) (c)

Results and Discussion

1. Complete the following table by indicating the number of collisions experienced by each reacting particle.

The Number of Collisions per Unit of Time	Snapshot (a)	Snapshot (b)	Snapshot (c)
Snapshot series 1 (effect of concentration)			
Snapshot series 2 (effect of temperature)			
Snapshot series 3 (effect of volume/pressure)			

2. Which two of the three factors shown produce the most similar effect on the rate of a chemical reaction?

3. One of the three factors will increase reaction rate more than the other two. Which factor is this?

4. Why would this factor have more of an impact on reaction rate than the other two?

5. Reaction rates are increased by collisions occurring more frequently and by collisions occurring more effectively. Which of these two things would be impacted by:
   (a) the addition of a catalyst to a reaction system?

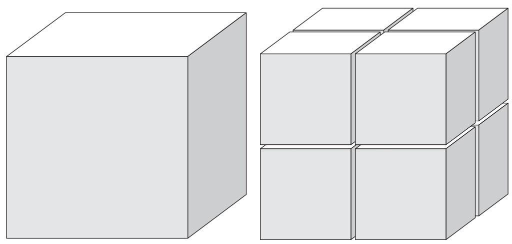
   (b) increasing the surface area in a heterogeneous reaction?

1.3 Review Questions

1. List the two things that affect the rates of all chemical reactions according to collision theory.

2. What are the two requirements for a collision to be successful?

3. One chunk of zinc is left whole and another is cut into pieces as shown. Both samples of zinc are reacted in an equal volume of 6.0 mol/L aqueous hydrochloric acid.



   (a) Assuming the surface area of the first zinc sample is 6.00 cm^2, what is the surface area of the second sample of zinc?

   (b) Compare the frequency of collisions between the hydrochloric acid and the single piece of zinc with those between the acid and the cut sample of zinc.

   (c) Assuming the average rate of reaction for the single piece of zinc with the acid is 1.20×10^{-3} mol Zn/min, calculate the rate of reaction for the cut sample of zinc.

   (d) Assuming this rate is maintained for a period of 4.50 min, how many milliliters of hydrogen gas would be collected at STP?

4. A reaction between ammonium ions and nitrite ions has the following rate law:

$$\text{rate} = k[NH_4^+][NO_2^-]$$

Assume the rate of formation of the salt is 3.10×10^{-3} mol/L/s. Note that the units may also be expressed as mol/L s. The reaction is performed in aqueous solution at room temperature.

(a) What rate of reaction would result if the $[NH_4^+]$ was tripled and the $[NO_2^-]$ was halved?

(b) Determine the reaction rate if the $[NH_4^+]$ was unchanged and the $[NO_2^-]$ was increased by a factor of four?

(c) If the $[NH_4^+]$ and the $[NO_2^-]$ were unchanged, but the rate increased to 6.40×10^{-3} mol/L s, what must have happened to the reacting system?

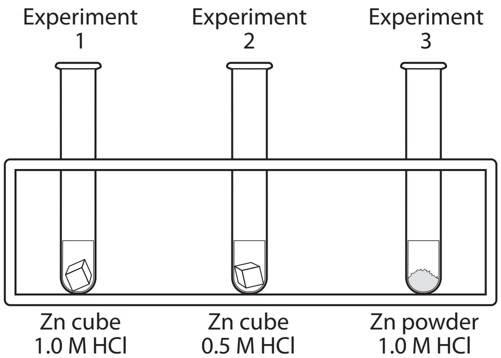
(d) What would the new reaction rate be if enough water were added to double the overall volume?

5. A student reacts ground marble chips, $CaCO_3(s)$, with hydrochloric acid, $HCl(aq)$, in an open beaker at constant temperature.

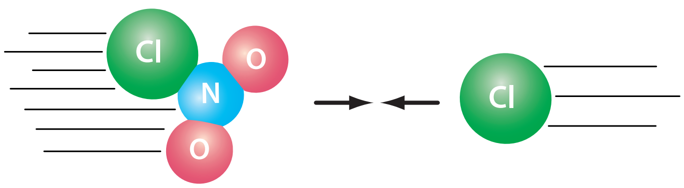
(a) In terms of collision theory, explain what will happen to the rate of the reaction as it proceeds from the beginning to completion.

(b) Sketch a graph of volume of $CO_2(g)$ vs. time to show the formation of product with time as the reaction proceeds.

6. Consider the following three experiments, each involving the same mass of zinc and the same volume of acid at the same temperature. Rank the three in order from fastest to slowest, and explain your ranking using collision theory.



7. Is the following collision likely to produce Cl_2 and NO_2 assuming the collision occurs with sufficient energy? If not, redraw the particles in such a way that a successful collision would be likely.



8. Use collision theory to explain each of the following.
   (a) Food found in camps half way up Mount Everest is still edible once thawed.

   (b) Campers react magnesium shavings with oxygen to start fires.

   (c) A thin layer of platinum in a vehicle's exhaust system converts oxides of nitrogen into non-toxic nitrogen gas.