

12. Diagrams should indicate the following: the reactants have a lower potential energy than the products; $E_{a(\text{fwd})}$ is the energy between the reactants and the transition state (+99 kJ); $E_{a(\text{rev})}$ is the energy between the transition state and the products (+34 kJ); and ΔH_r is the enthalpy of reaction between the reactants and the products (+65 kJ).

13. Diagrams should indicate the following: the reactants have a higher potential energy than the products; $E_{a(\text{fwd})}$ is the energy between the reactants and the transition state (+61 kJ); $E_{a(\text{rev})}$ is the energy between the transition state and the products (+150 kJ); and ΔH_r is the enthalpy of reaction between the reactants and the products (−89 kJ).

14. $E_{a(\text{fwd})} = +42$ kJ; $E_{a(\text{rev})} = +67$ kJ; $\Delta H_r = -25$ kJ; exothermic

15. $E_{a(\text{rev})} = +411$ kJ; energy diagrams should indicate the reactants having higher potential energy than the products with $E_{a(\text{fwd})}$ being the energy between the reactants and the transition state (+19.0 kJ); $E_{a(\text{rev})}$ is the energy between the transition state and the products (+411 kJ); and ΔH_r is the enthalpy of reaction between the reactants and the products (−392 kJ).

16. a. $E_a = 112$ kJ − 36 kJ = +76 kJ;
 $\Delta H = 78$ kJ − 36 kJ = +42 kJ

b. Potential energy diagrams should indicate the reactants having lower potential energy than the products and the activated complex having the highest potential energy, positioned between the reactants and products.

17. a. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) + 890.8$ kJ

b. Since $\text{CH}_4(\text{g})$ easily ignites with a spark, the activation energy for the combustion reaction would be expected to be small. The activated complex would be an unstable association of $\text{CH}_4(\text{g})$ and $\text{O}_2(\text{g})$.

18. Diagrams should indicate the reactants having higher potential energy than the products and the activated complex having the highest potential energy, positioned between the reactants and products.

$$\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}; E_{a(\text{fwd})} = +49.5 \text{ kJ}$$

19. Diagrams should indicate the reactants having lower potential energy than the products and the activated complex having the highest potential energy, positioned between the reactants and products.

20. $\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}; E_{a(\text{fwd})} = +320$ kJ

Answers to Chapter 6 Review Questions

(Student textbook pages 392-7)

1. d

2. b

3. c

4. a

5. e

6. d

7. d

8. b

9. d

10. e

11. a

12. d

13. c

14. d

15. Rate of reaction is equal to the change in the quantity of a reactant or a product divided by the time during which the change occurs.

16. Sample answers:

a. mL/s

b. kPa/s

c. pH/s

17. a. The rate will increase since at a higher total pressure the particles are forced closer together and they will collide more often, resulting in a higher frequency of successful collisions.

b. The rate will not be affected because the particles in an aqueous solution cannot be forced any closer together and cannot be made to collide more often by increasing the pressure.

18. The reactant particles have strong bonds that require a high energy collision to break; the reactant particles have a shape that requires an exact orientation before a collision can be effective.

19. a. The initial rate of reaction occurs when the reactants are first mixed together and have just begun to react.

b. The initial rate of reaction can be determined graphically by determining the slope of the tangent to the reaction curve (concentration vs. time) at $t = 0$.

- c. It is possible for the instantaneous rate to have the same value as the average rate of reaction. This would happen when a tangent to the reaction curve has the same slope as a line drawn between two points on the reaction curve. It occurs when the shape of the curve is straight between two points.
- 20. a.** A catalyst speeds up a reaction by providing an alternate path with a lower activation energy so that a greater percentage of the collisions occur with enough energy to get over the activation energy barrier.
- b.** A catalyst would not be expected to speed up this reaction. This reaction is between two oppositely charged ions in solution, and is therefore already a fast reaction. No simpler two-particle process is likely.
- 21.** The partial pressure of a gas is increased by adding more of that gas to the container, thus increasing its concentration. There will be more collisions between reactant particles as a result, and the rate will increase.
- 22.** Without stirring, the solid solute becomes surrounded by more and more concentrated solution as it dissolves, eventually becoming surrounded by a zone of saturated solution. This dissipates slowly by diffusion. By stirring, the solution is kept as dilute as possible, ensuring that many water molecules are available to surround solute particles. The effect is the same as increasing the surface area of a reactant. Stirring would have to be very vigorous to substantially increase the kinetic energy of the particles and this is not the major reason that stirring increases the rate of dissolving.
- 23.** The activation energy is the minimum energy that colliding reactant particles must have to allow them to reach a transitional state where they may break up into products. The activated complex is an unstable association of reactant particles formed temporarily at the transition state that will break apart either to reform the reactants or to form products.
- 24.** To determine the initial rates of reactions at different concentrations, several experiments are carried out, each with a different concentration of reactant. Data for each experiment is plotted as graphs of concentration vs. time. From these curves, the initial rates are determined for each experiment by measuring the slope of the tangent lines at time zero on each graph. These initial rates are plotted and the resulting graph of initial rate vs. initial concentration of reactant shows the relationship between reaction rate and concentration.
- 25.** For an effective collision to occur, particles must collide with sufficient energy to get over the activation energy barrier and at a favourable orientation. There is a relatively good probability of a collision between two particles meeting these criteria. For three or more particles to collide in this manner is much less probable.
- 26.** (i) For a solid catalyst, the more surface area exposed, the better the ability of the substance to catalyze the reaction. (ii) The state of the reactants compared to the state of the catalyst affects how the substance can catalyze the reaction. For reactions between gases, the catalyst is most often in the solid state since the gases can be easily passed over a solid. If the catalyst for a reaction between gases is also a gas, continual intimate mixing of the gases would be required. (iii) The purity of the catalyst is also a factor. Contaminants in the catalyst would result in less contact between particles that are involved in the reaction. Also, an impurity could interfere with the reaction. (iv) Temperature would be a factor for a biological catalyst (called an enzyme). At higher temperatures, this type of catalyst is destroyed because its action depends on its shape, and because it is a protein, it becomes denatured at high temperatures.
- 27. a.** The straight line represents a direct relationship between rate and concentration.
- b.** The slope of the graph represents the rate constant, k .
- c.** The reaction is first order and has the form rate = $k[A]$.
- 28. a.** The rate constant, k , will not be affected by a change in the concentration of the reactants.
- b.** NO_3 is a reaction intermediate because it is produced in step 1 and used up in step 2.
- 29.** $+5.0 \times 10^2 \text{ kJ}$
- 30.** $+3.1 \times 10^2 \text{ kJ}$
- 31.** The catalyzed reaction is faster because it has a lower activation energy.
- 32.** $+7 \times 10^2 \text{ kJ}$
- 33.** $+4.7 \times 10^2 \text{ kJ}$
- 34.** $-1.7 \times 10^2 \text{ kJ}$
- 35.** $+1.7 \times 10^2 \text{ kJ}$

$$\begin{aligned}
 36. \text{ rate of formation of } A_2 &= \frac{\Delta \text{concentration}}{\Delta t} \\
 &= \frac{c_f - c_i}{t_f - t_i} = \frac{0.68 \text{ mol/L} - 0.52 \text{ mol/L}}{5.6 \text{ min} - 4.4 \text{ min}} \\
 &= \frac{0.16 \text{ mol/L}}{1.2 \text{ min}} = 0.13 \text{ mol/L}\cdot\text{min}
 \end{aligned}$$

The mole ratio AB:A₂ is 2:1. For every 2 mol of AB that decompose, 1 mol of A₂ is produced.

$$\begin{aligned}
 \frac{\text{rate of decomposition of AB}}{\text{rate of formation of } A_2} &= \frac{n \text{ AB}}{n \text{ A}_2} \\
 \frac{\text{rate of decomposition of AB}}{0.13 \text{ mol/L}\cdot\text{min}} &= \frac{2 \text{ mol AB}}{1 \text{ mol A}_2} \\
 &= 2 \times 0.13 \text{ mol/L}\cdot\text{min} = 0.26 \text{ mol/L}\cdot\text{min}
 \end{aligned}$$

$$\begin{aligned}
 37. \text{ change in amount of HNO}_3(\text{aq}) &\text{ is } 0.10 \text{ mol} \\
 \frac{\text{rate of change of Cu}}{\text{rate of change of HNO}_3} &= \frac{n \text{ Cu}}{n \text{ HNO}_3} = \frac{1}{4} \\
 \text{change in the amount of copper} &= 0.25 \times 0.10 \text{ mol} = 0.025 \text{ mol} \\
 \text{mass of copper used in this time} &= 0.025 \text{ mol} \times 63.55 \frac{\text{g}}{\text{mol}} = 1.6 \text{ g} \\
 \text{rate of change in mass of copper} &= \frac{1.6 \text{ g}}{90.0 \text{ s}} = 0.018 \text{ g/s}
 \end{aligned}$$

38. a. III
b. IV
c. III
d. II

39. The coefficients in the chemical equation indicate the mole ratios for the components of the reaction. For every 4 mol of HBr(g) that decompose, 2 mol of Br₂(g) is formed. In other words, there is a 2:1 ratio of HBr(g) decomposition to Br₂(g) formation.

$$\frac{\text{rate of decomposition HBr}}{\text{rate of formation of Br}_2} = \frac{n \text{ HBr}}{n \text{ Br}_2} = \frac{4}{2} = \frac{2}{1}$$

$$\begin{aligned}
 40. \text{ rate in mol/s} &= \frac{0.48 \text{ g}}{5.8 \text{ min}} \times \frac{1 \text{ mol}}{84.32 \text{ g}} \times \frac{1 \text{ min}}{60 \text{ s}} \\
 &= 1.6 \times 10^{-5} \text{ mol/s}
 \end{aligned}$$

41. Potential energy diagram A most accurately shows the catalytic action on the mechanism. Step 1 has the highest activation energy, and it will be the slowest or rate-determining step. The catalyst would be used to lower the activation energy of the rate-determining step 1.

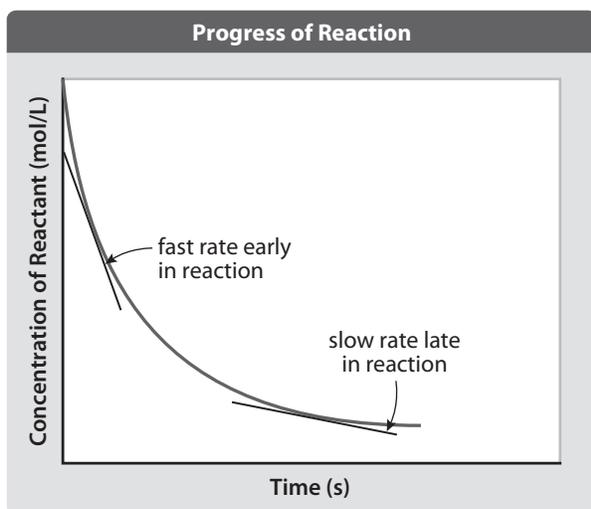
42. a. +20 kJ
b. No; the reverse reaction would have a greater activation energy than in (a).

$$\begin{aligned}
 E_{a(\text{rev})} &= E_{a(\text{fwd})} - \Delta H \\
 &= +45 \text{ kJ} - (-25 \text{ kJ})
 \end{aligned}$$

$$= 70 \text{ kJ}$$

43. a. There are two steps in the reaction.
b. C + D = overall activation energy
c. The overall reaction is exothermic.
d. Quantity B represents the activation energy for step 1.
e. Quantity E represents the difference between potential energy of the product(s) of step 1 and the potential energy of the final product(s).
44. a. Both reactions are second order.
b. If the molar concentration of R is doubled, the rate of reaction A will quadruple (2²) and the rate of reaction B will double (2¹).
45. a. catalyst
b. intermediate
c. intermediate
d. reactant
46. When the volume is reduced to half the original volume, the concentration of the reactants will each be doubled. Since the rate law is first order only in gas X₂(g), the rate of the reaction would be expected to double.
47. a. Step 2 is the rate-determining step since it is identified as a slow step.
b. Intermediates are formed in one step and consumed in a subsequent step. The intermediates are AlBr₄⁻, C₂H₅⁺, and C₆H₆C₂H₅⁺.
c. AlBr₃ is a catalyst because it is part of the reaction as a reactant in step 1 and is regenerated in step 3.
48. The rate of reaction must be compared using samples that have the same number of particles. Equal masses of different metals will not contain the same number of particles because the molar masses are different. The mass of each metal used per unit of time must be converted to amount (in mol) per unit of time so that the comparison will be valid.
49. Units should include change in volume per unit time (mL/s, mL/min, mL/h, L/h); change in partial pressure per unit time (kPa/s, kPa/min); and change in the amount in moles per unit time (mol/min, mol/h).
50. This statement is incorrect. The frequency of collisions will affect the rate of reaction. The greater the number of collisions, the faster the rate. The value of the activation energy depends solely upon the nature of the reactants. Frequency of collisions does not affect this.
51. It is possible for these measurements to represent the same reaction. Early on in a reaction, the reaction proceeds relatively quickly. As the reaction proceeds,

the reactants become used up and the reaction rate slows down. The instantaneous rate of reaction is given by the slope of the curve of concentration vs. time at any particular time. The graph below shows the change in concentration of a reactant over time. The slope of the tangent to this graph early in the reaction is steep while the slope measured as the reaction nears its end is much less so.



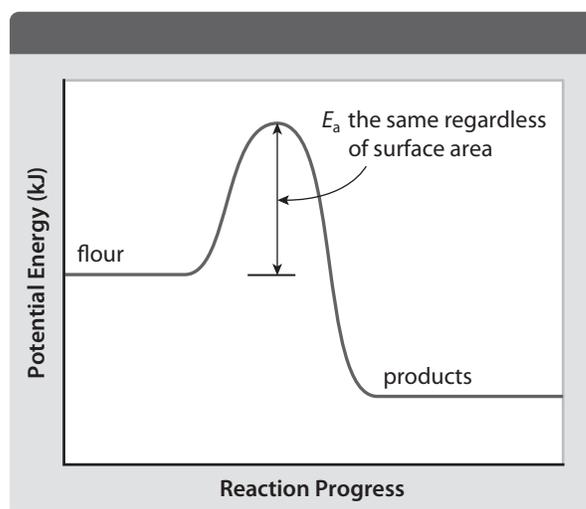
52. A catalyst changes the rate of reaction by providing an alternate mechanism for the reaction that has a lower activation energy. The lower activation energy means that more particles have enough energy to react when a collision occurs. At a higher temperature the activation energy does not change. Instead, the distribution of kinetic energy of the reacting particles changes so that more colliding particles have energy equal to or greater than the activation energy. There are now more collisions that are effective and the rate of reaction increases.

53. When it comes to reactions between a solid reactant and an aqueous or liquid reactant, it is the surface area of the solid that counts, not its total mass, when it comes to rate. Only the particles in the solid that are exposed to the solution can participate in collisions with the dissolved reactant. The doubling of the rate in this experiment is, in fact, due to the doubling of the surface area, not the doubling of the mass of the reactant. However, the design of the experiment is faulty because two variables were changed: mass and surface area. It would have been better to have used the same mass of magnesium and changed the surface area, either by cutting it into pieces or by using powdered metal. (Students' diagrams could show solute particles colliding only with the surface of a solid, or show how the surface area of a solid increases when it is divided into pieces or ground up.)

54. The reactants in both reactions likely have similar bond strengths and the orientation factor for both sets of reactants likely is similar. This would account for the fact that both reactions have the same activation energy. One reason that the rate of reaction could be different is that the endothermic reaction would need a continual source of activation energy while the exothermic reaction would, once started, produce energy to maintain the reaction. This thermal energy gives the particles more kinetic energy and the rate of reaction will continue to increase. The rates can also be different because of a difference in the rate constant.

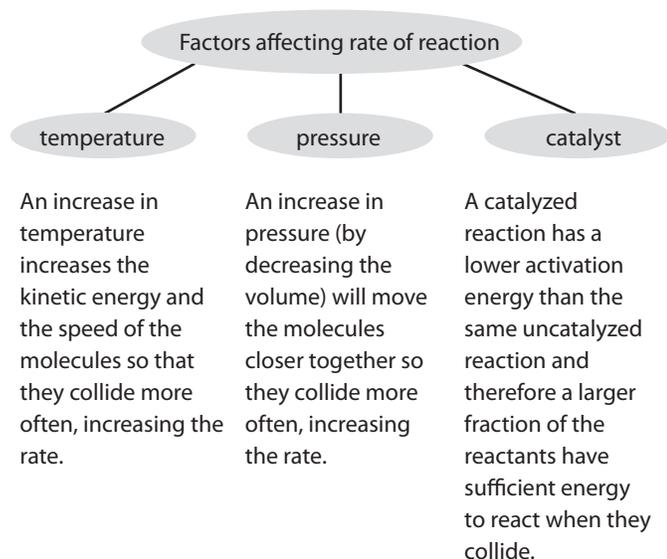
55. A main idea web or a spider map may be used to summarize information related to a catalyst such as availability, cost, effectiveness, durability, recoverability, and state of matter.

56. The flour blown into the flame has a greater surface area than the flour is in a pile. Each small particle of flour will burn in an instant. When in a pile, only the particles of flour on the outside of the pile collide with the oxygen in the air so few particles react at a given time. With few particles reacting, the heat given off by the combustion reaction will be a small and it will be difficult to maintain the burning. However, all of the flour particles have the same chemical bonds and it will require the same amount of energy to break these bonds. The activation energy for the flour particles will be the same regardless of whether the sample of flour burns in an instant or not. The potential energy diagram will therefore be the same for both burnings.



57. Students' graphic organizers should show that the following factors would increase the rate of reaction between two gases: increasing the temperature, increasing the pressure, and adding a catalyst.

Sample concept map:



58. Students should organize the following answers in a graphic display or concept map.
- A catalyst provides an alternate mechanism for the reaction that has a lower activation energy.
 - The catalyst affects the forward and reverse reactions by lowering the activation energy by the same quantity.
 - Only a small amount of catalyst is needed because it is regenerated in the reaction and can be used over and over.
 - The site where the catalytic action occurs must be able to interact in a specific way with the reactant or reactants. A catalyst is often specific for one reaction.
 - The Haber process for the production of ammonia and the Contact process for the production of sulfuric acid use catalysts.
59. Graphic organizers should summarize key concepts related to rates of reaction, collision theory, factors affecting rates of reaction, and reaction mechanisms, using the correct vocabulary. The graphic organizer should also show how the four topics are related.
60. a. Since $\text{FeS}_2(\text{s})$ is in the solid state, its rate of decomposition could be measured in mol/s, g/s, mol/min, or g/min, to name a few.
- b. Since $\text{SO}_2(\text{g})$ is a gas, its rate of formation can be measured in mL/s, mol/L/s, mL/min, or mol/min to name a few.

- c. The two rates can be compared using the balanced equation. The relative rate of change of a reactant or product is in proportion to their mole ratio in the balanced equation. The mole ratio $\text{FeS}_2(\text{s}) : \text{SO}_2(\text{g})$ is 4:8. For every 4 mol of $\text{FeS}_2(\text{s})$ that decomposes, 8 mol of $\text{SO}_2(\text{g})$ is produced.

$$\frac{\text{rate of decomposition FeS}_2}{\text{rate of formation SO}_2} = \frac{4\text{FeS}_2}{8\text{SO}_2} = \frac{1\text{FeS}_2}{2\text{SO}_2}$$

$$\text{rate of decomposition of FeS}_2 = 0.5 \times \text{rate of formation of SO}_2(\text{g})$$

61. The mole ratio $\text{Na}_3\text{AlF}_6(\ell) : \text{F}_2(\text{g})$ is 1 : 3. For every 1 mol of $\text{Na}_3\text{AlF}_6(\ell)$ that decomposes, 3 mol of $\text{F}_2(\text{g})$ is produced.

$$\frac{\text{rate of decomposition Na}_3\text{AlF}_6(\ell)}{\text{rate of formation of F}_2(\text{g})} = \frac{n \text{Na}_3\text{AlF}_6}{n\text{F}_2}$$

$$\frac{\text{rate of decomposition Na}_3\text{AlF}_6(\ell)}{0.85 \text{ mol/s}} = \frac{1 \text{ mol Na}_3\text{AlF}_6}{3 \text{ mol F}_2}$$

average rate at which $\text{Na}_3\text{AlF}_6, (\ell)$ decomposes

$$= 1/3 \times 0.85 \text{ mol/s}$$

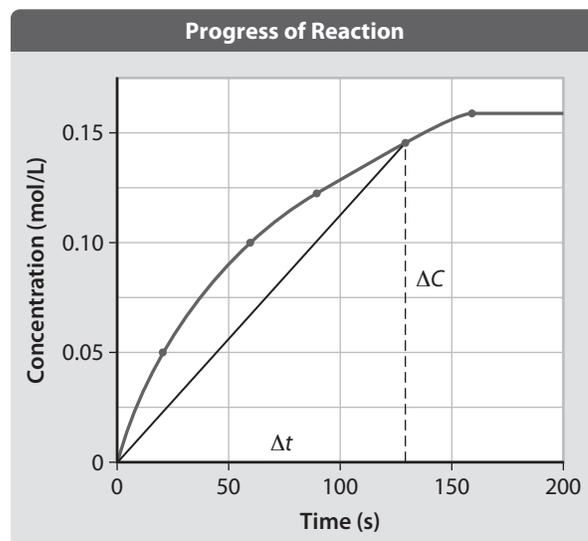
$$= 0.28 \text{ mol/s}$$

$$m \text{Na}_3\text{AlF}_6(\ell) = 0.28 \text{ mol} \times 209.95 \frac{\text{g}}{\text{mol}}$$

$$= 58.786 \text{ g}$$

average rate of decomposition of $\text{Na}_3\text{AlF}_6(\ell) = 59 \text{ g/s}$

62. a.



Slope of line between 130 s and 0 s

$$= \frac{\Delta c}{\Delta t} = \frac{0.148 \frac{\text{mol}}{\text{L}} - 0.000 \frac{\text{mol}}{\text{L}}}{130 \text{ s} - 0.00 \text{ s}}$$

$$= 1.14 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

Average rate of production of $\text{NH}_3(\text{g})$ in 130 s is

$$1.14 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

- b. The change in concentration of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$ will not occur at the same average rate as the production of ammonia. The relative rate of change of a reactant

or product is in proportion to their mole ratio in the balanced equation.

The mole ratio $\text{N}_2(\text{g}):\text{NH}_3(\text{g})$ is 1:2. For every 1 mol of $\text{N}_2(\text{g})$ consumed, 2 mol of $\text{NH}_3(\text{g})$ is produced.

$$\frac{\text{rate at which N}_2 \text{ used}}{\text{rate of formation NH}_3} = \frac{n\text{N}_2}{n\text{NH}_3}$$

$$\frac{\text{rate at which N}_2 \text{ used}}{1.14 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{s}}} = \frac{1\text{N}_2}{2\text{NH}_3}$$

average rate of consumption of $\text{N}_2(\text{g})$

$$= 0.5 \times 1.14 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

$$= 5.69 \times 10^{-4} \text{ mol/L}\cdot\text{s}$$

The mole ratio of $\text{H}_2(\text{g}):\text{NH}_3(\text{g})$ is 3:2. For every 3 mol of $\text{H}_2(\text{g})$ used, 2 mol of $\text{NH}_3(\text{g})$ is produced

$$\frac{\text{rate at which H}_2 \text{ used}}{\text{rate of formation NH}_3} = \frac{n\text{N}_2}{n\text{NH}_3}$$

$$\frac{\text{rate at which H}_2 \text{ used}}{1.14 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{s}}} = \frac{1\text{N}_2}{2\text{NH}_3}$$

average rate of consumption of $\text{H}_2(\text{g})$

$$= 1.5 \times 1.14 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

$$= 1.71 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

$$63. n(\text{N}_2\text{O}_5) = \frac{5.00 \times 10^5 \cancel{\text{g}}}{108.02 \frac{\cancel{\text{g}}}{\text{mol}}} = 4.628 \times 10^4 \text{ mol}$$

$\text{N}_2\text{O}_5(\text{g})$ is used up at a rate of $4.628 \times 10^4 \text{ mol/h}$

$$= 4.628 \times 10^4 \text{ mol}/60 \text{ min}$$

$$= 7.715 \times 10^1 \text{ mol/min}$$

The mole ratio of $\text{NO}_2(\text{g}):\text{N}_2\text{O}_5(\text{g})$ is 4:2. For every 4 mol of $\text{NO}_2(\text{g})$ produced, 2 mol of $\text{N}_2\text{O}_5(\text{g})$ is consumed.

$$\frac{\text{rate of formation NO}_2}{\text{rate of consumption N}_2\text{O}_5} = \frac{4 \text{ NO}_2}{2 \text{ N}_2\text{O}_5}$$

$$\frac{\text{rate of formation NO}_2}{7.715 \times 10^{-1} \frac{\text{mol}}{\text{min}}} = \frac{4 \text{ NO}_2}{2 \text{ N}_2\text{O}_5}$$

rate at which $\text{NO}_2(\text{g})$ forms = $2 \times 7.715 \times 10^1 \text{ mol/min}$
 $= 1.54 \times 10^2 \text{ mol/min}$

64. The percent conversion is low because only a fraction of the particles are in contact with the catalyst. To improve the percent conversion with the catalyst the process must increase the surface area available for contact with the reactant gases. The reactant gases could also be continually recycled until they make contact with the catalyst.

65. a. At the higher temperature, more particles have kinetic energy equal to or exceeding the activation energy. More of the collisions will result in reactants forming products. Students' diagrams should resemble Figure 6.15 in the student textbook.

b. The activation energy and enthalpy change do not change with temperature. There would be no change in the values of E_a and ΔH . Students' diagrams should resemble Figure 6.10A in the student textbook.

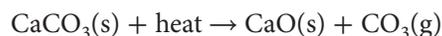
c. To run the process at a higher temperature will increase the energy cost. This would have to be factored into the overall cost.

66. Since the reaction is first order with respect to D and E, D and E both must take part in the rate-determining step. This means that Mechanism 1 is unlikely to be correct. Mechanism 2 is more likely to be correct since both D and E are reactants in the slow step. Just examining possible steps does not prove that a mechanism is correct; the mechanism must be tested by experiment.

67. The reaction involving the new material must have a high activation energy. The new material must have stronger bonds that require more energy to break and/or a shape requiring a specific orientation in order for the reaction to begin. If the reaction is highly exothermic, once the reaction starts, there should be sufficient energy to keep the reaction going. To allow the new reactant to be used, the reaction could be started at a higher temperature or a catalyst could be used to lower the activation energy.

68. All chemical reactions involve the interaction of electrons with other electrons and with the nuclei of the atoms involved. The atoms of transition metals have partially filled d orbitals. They can both share and donate electrons. These d orbitals provide a way for electrons to form transitional, partial bonds. The result is that the catalyzed transition state has a lower energy than an uncatalyzed transition state would have for the same process. A lower activation energy results in a faster reaction rate.

69. A reaction in which a single particle splits up into two or more smaller particles would not necessarily be collision dependent. The rate of reaction would depend upon the vibrational energy of the chemical bonds in the particles and therefore on the temperature. One example is the thermal decomposition of calcium carbonate, $\text{CaCO}_3(\text{s})$.



Answers to Chapter 6 Self-Assessment Questions (Student textbook pages 398-9)

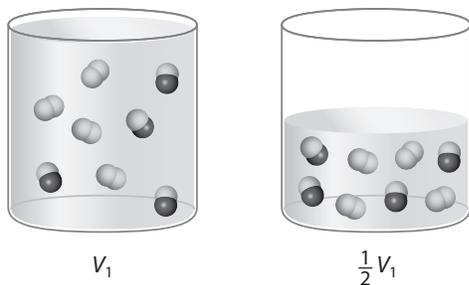
- a
- e
- b

4. d
5. d
6. e
7. c
8. d
9. d
10. b

11. a. $E_{a(\text{fwd})} = E_{a(\text{rev})} + \Delta H$

b. For an endothermic reaction, $E_{a(\text{fwd})} > E_{a(\text{rev})}$

12. Use different symbols to represent the two types of reactant particles. One diagram, representing the container at a volume of V_1 , shows four particles of each reactant. Explain that over a period of time, there will be a number of collisions among the reactant particles. In a second diagram, the volume, V_2 , will be shown as half the original ($\frac{1}{2}V_1$). Now the same number of particles occupy half the space. On average there will be twice as many collisions and this should lead to an increase in the rate of reaction.



13. $n\text{KClO}_3(\text{s}) = \frac{m}{M} = \frac{0.20 \text{ g}}{122.55 \frac{\text{g}}{\text{mol}}}$

$= 1.6319 \times 10^{-3} \text{ mol}$

The mole ratio of $\text{O}_2 : \text{KClO}_3$ is 3:2

$\frac{\text{mol O}_2}{\text{mol KClO}_3} = \frac{3}{2}$

$n\text{O}_2(\text{g}) = \frac{3}{2} \times \text{mol KClO}_3(\text{s})$

$= \frac{3}{2} \times 1.6319 \times 10^{-3} \text{ mol}$

$= 2.4478 \times 10^{-3} \text{ mol}$

$PV = nRT$

$V = \frac{nRT}{P}$

$= \frac{2.4478 \times 10^{-3} \text{ mol} \times 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}} \times 298.15 \text{ K}}{100.0 \text{ kPa}}$

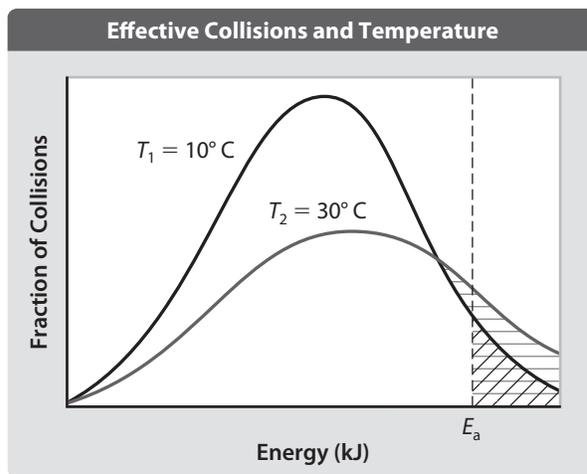
$= 6.0676 \times 10^{-2} \text{ L}$

$= 60.676 \text{ mL}$

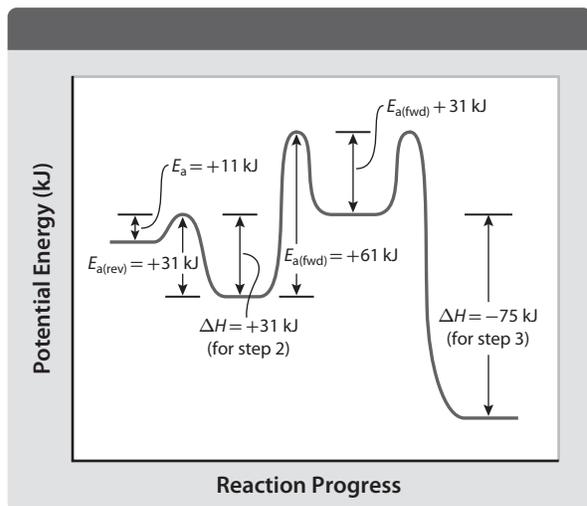
Rate of formation of $\text{O}_2(\text{g}) = \frac{60.676 \text{ mL}}{2.8 \text{ s}}$

$= 21.67 \frac{\text{mL}}{\text{s}} = 22 \frac{\text{mL}}{\text{s}}$

14. An input of energy is needed to start the reaction associated with the explosion. This initial input of energy provides a few reactants with enough energy to overcome the activation barrier of the reaction. Explosive reactions are exothermic, so the initial successful collisions will generate energy to allow further reactions to occur.
15. Students' diagrams should resemble Figure 6.15 in the student textbook. They should change labels to show a 20°C difference. Even a modest increase in temperature can result in a significant increase in the number of collisions that occur with enough energy for a reaction to take place.



16. -64 kJ



17. The relative rate of change of a reactant or product is in proportion to their mole ratios in the balanced chemical equation.

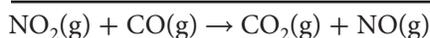
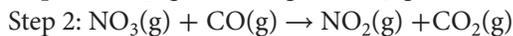
The mole ratio of $\text{NO}:\text{H}_2\text{O}$ is 4:6.

$\frac{\text{rate of consumption}}{\text{rate of consumption of CO}_2} = \frac{n\text{NO}}{n\text{H}_2\text{O}}$

$$\frac{0.050 \frac{\text{mol}}{\text{L}\cdot\text{s}}}{\text{rate of consumption of H}_2\text{O}} = \frac{4}{6}$$

$$\begin{aligned} \text{rate at which H}_2\text{O(g) reacts} &= \frac{6}{4} \times 0.050 \text{ mol/L}\cdot\text{s} \\ &= 7.5 \times 10^{-2} \text{ mol/L}\cdot\text{s} \end{aligned}$$

- 18.** One possible mechanism is shown below.



- 19.** rate of change in $\text{AgNO}_3(\text{aq})$

$$= \frac{\Delta C}{\Delta t} = \frac{0.42 \text{ mol/L} - 0.28 \text{ mol/L}}{60.0 \text{ s} - 30.0 \text{ s}}$$

$$= 4.667 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

The relative rate of change of a reactant or product is in proportion to the mole ratios in the balanced equation.

From the balanced equation, mole ratio $\text{AgNO}_3:\text{Cl}^-$ is 2:2 (or 1:1). Therefore, the rate of reaction of chloride is equal to the rate of reaction of silver nitrate.

$$\text{rate of reaction of Cl}^- = 4.667 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

$$= 4.7 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

- 20.** The activation energy and the enthalpy change remain the same when the concentrations of the reactants are altered. Having a greater number of particles collide at a higher concentration does not affect the mechanism of the reaction or the nature of the bonds in the reactants. The reaction will proceed more quickly and more thermal energy will be generated more quickly, but the value of ΔH_r for the reaction as written in a thermochemical equation will not change.
- 21.** It is a second-order reaction; for second-order reactions, the rate is proportional to the square of the reactant.
- 22.** The number of effective collisions must be doubled to double the rate of reaction. This can be accomplished by increasing the temperature and/or, if the reactants are gases, by increasing the pressure.
- 23. a.** Because there are three “peaks” representing three transition states, there are three steps involved in this reaction.
- b.** A: the reactants
 B: transition state of Step 1
 C: the products formed in Step 1, intermediates, and any unreacted reactants
 D: the transition state of Step 2
 E: the products formed in Step 2, intermediates, and any unreacted reactants

F: the transition state for Step 3

G: the products

H: the activation energy for the reverse of Step 1

I: the enthalpy change for Step 1

J: the activation energy for Step 2

K: the activation energy for Step 3

L: the enthalpy change for the overall reaction

- c.** A net quantity of energy is added because the products are higher in energy than the reactants, which means the overall reaction is endothermic.
- 24.** Chemists use very short laser pulses to monitor chemical reactions. The durations of these pulses are in the range of femtoseconds. The first pulse provides the activation energy for the reaction. Chemists analyze the absorbance of the second pulse, and use the data to determine the absorbance characteristics of any activated complexes or chemical intermediates. This data allows chemists to identify the chemical species that are formed during a reaction and to provide evidence to support or refute a proposed reaction mechanism.
- 25.** An increase in temperature will increase the rate of a reaction. As temperature increases, more particles will collide with enough energy to react. Student diagrams should resemble Figure 6.15 in the student textbook.

Answers to Unit 3 Review Questions (Student textbook pages 403-9)

1. a
2. d
3. c
4. a
5. c
6. b
7. e
8. d
9. b
10. e

- 11.** Since both water samples are at the same temperature, the average kinetic energy of the particles will be the same. Thermal energy is the sum of the kinetic energies of all the particles in a sample. Since there are more particles in the bathtub of water, the thermal energy will be greater in the bathtub sample.