

(b) Lead forms bonds with proteins. When a victim of lead poisoning ingests a large amount of protein, such as the protein found in egg whites, a large portion of the lead in the stomach will bond with the food protein. The stomach must be pumped; otherwise, the protein will be digested and the lead will be freed to potentially bind to enzymes.

(c) The antidote would not work for someone who had been ingesting lead over a long period of time because the lead would not be in the person's stomach; instead it would be bound to enzymes throughout the person's body.

## Chapter 6 Review Answers

Student Textbook pages 311–313

### Knowledge/Understanding

- (a) The mass of the system must change; a gas must be produced and escape the system into the air.  
(b) A gas must be produced.  
(c) The colour of the reaction system must change.
- Reactant particles must come together for a reaction to occur. For a reactant in the solid state, this contact occurs at the surface. When there is more surface area, there are more collisions and the reaction occurs more quickly.
- As reactant particles combine, agitation allows other reactant particles to come together more quickly.
- (a)  $-\frac{\Delta[\text{NH}_3]}{\Delta t} = \frac{\Delta[\text{NO}]}{\Delta t} = -\frac{5}{4} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{3}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$   
(b) The concentration of ammonia decreases at a rate of  $6.2 \times 10^{-2} \text{ mol/L} \cdot \text{s}$ .
- Reactant particles must have the correct orientation, and the collision energy must be equal to or greater than the activation energy for the reaction.
- A homogeneous catalyst is in the same phase as the reactants. A heterogeneous catalyst is a different phase than the reactants.

### Inquiry

$$7. \text{(a) rate} = \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} = \frac{0.200 \text{ mol/L} - 0.196 \text{ mol/L}}{10 \text{ s}} = 4.00 \times 10^{-4} \text{ mol/L} \cdot \text{s}$$

$$\text{(b) mol H}_2\text{O}_{2(\text{aq})} = 4.00 \times 10^{-4} \text{ mol/L} \cdot \text{s} \times 10.0 \text{ s} \times 0.500 \text{ L} = 2.00 \times 10^{-3}$$

$$\text{mol O}_2 = \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}_2} \times 34.00 \text{ g/mol} \times 2.00 \times 10^{-3} \text{ mol H}_2\text{O}_2 = 3.40 \times 10^{-2} \text{ g}$$

$$8. \text{(a) rate} = k[\text{A}]$$

$$k = \frac{3.8 \times 10^{-3} \text{ mol/(L} \cdot \text{s)}}{0.38 \text{ mol/L}} = 0.010 \text{ s}^{-1}$$

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.010 \text{ s}^{-1}} = 69 \text{ s}$$

(b) Students should say they assumed that the average rate of change of reactants is uniform over the time period.

9. The rate law for a reaction has the form,  $\text{rate} = k[\text{reactant(s)}]^m$ . For rate and  $k$  to have the same units, the concentration term must not be in the equation. This will occur only for a zero order reaction.

$$10. t_{1/2} = \frac{0.693}{k} = \frac{0.693}{4.87 \times 10^{-2} \text{ s}^{-1}} = 14.2 \text{ s}$$

11. (a) For experiment 1 and experiment 2,  $[\text{Cl}_{2(\text{g})}]$  remains constant, while  $[\text{CO}_{(\text{g})}]$  is reduced by a factor of 10. The rate also is reduced by a factor of approximately 10. Therefore, the reaction is first order in CO.

For experiment 2 and experiment 3,  $[\text{CO}_{(\text{g})}]$  remains constant, while  $[\text{Cl}_{2(\text{g})}]$  increases by a factor of 10. The rate also increases by a factor of approximately 10. Therefore, the reaction is first order in  $\text{Cl}_2$ .

$$\text{Thus, rate} = k[\text{CO}_{(\text{g})}] [\text{Cl}_{2(\text{g})}]$$

(b) Using trial 1:

$$k = \frac{\text{rate}}{[\text{CO}][\text{Cl}_2]} = \frac{6.45 \times 10^{-30} \text{ mol/(L} \cdot \text{s)}}{(0.500 \text{ mol/L})(0.0500 \text{ mol/L)}} = 2.58 \times 10^{-28} \text{ L/(mol} \cdot \text{s)}$$

12. (a) For experiment 1 and experiment 2,  $[\text{B}_{(\text{aq})}]$  is constant, while  $[\text{A}_{(\text{aq})}]$  increases by a factor of 3. The rate increases by a factor of 9 (that is,  $3^2$ ). Therefore, the reaction is second order in  $[\text{A}_{(\text{aq})}]$ .

For experiment 1 and experiment 3,  $[\text{A}_{(\text{aq})}]$  is constant, while  $[\text{B}_{(\text{aq})}]$  increases by a factor of 2. The rate also increases by a factor of 2. Therefore, the reaction is first order in  $[\text{B}_{(\text{aq})}]$ .

$$\text{rate} = k[\text{A}_{(\text{aq})}]^2 [\text{B}_{(\text{aq})}]$$

(b) The reaction is third order overall.

(c) Using trial 1:

$$k = \frac{\text{rate}}{[A]^2[B]}$$

$$= \frac{5.00 \text{ mol}/(\text{L} \cdot \text{s})}{(0.200 \text{ mol}/\text{L})^2(0.0200 \text{ mol}/\text{L})}$$

$$= 625 \text{ L}^2/(\text{mol}^2 \cdot \text{s})$$

13. (a) Comparing experiments 1 and 2,  $[\text{Hg}_2\text{Cl}_{2(\text{aq})}]$  is constant, while  $[\text{C}_2\text{O}_4^{2-}(\text{aq})]$  increases by a factor of 2. The rate increases by a factor of 4 (that is,  $2^2$ ). Therefore, the reaction is second order in  $\text{C}_2\text{O}_4^{2-}(\text{aq})$ .

Comparing experiments 2 and 3,  $[\text{C}_2\text{O}_4^{2-}(\text{aq})]$  is constant, while  $[\text{Hg}_2\text{Cl}_{2(\text{aq})}]$  decreases by 0.5. Rate also decreases by 0.5. Therefore, the reaction is first order in  $\text{Hg}_2\text{Cl}_2$ .

$$\text{rate} = k[\text{Hg}_2\text{Cl}_{2(\text{aq})}] [\text{C}_2\text{O}_4^{2-}(\text{aq})]^2$$

(b) Using trial 1:

$$k = \frac{\text{rate}}{[\text{Hg}_2\text{Cl}_2] [\text{C}_2\text{O}_4^{2-}]^2}$$

$$= \frac{1.35 \times 10^{-6} \text{ mol}/(\text{L} \cdot \text{s})}{(0.0788 \text{ mol}/\text{L})(0.133 \text{ mol}/\text{L})^2}$$

$$= 1.35 \times 10^{-3} \text{ L}^2/(\text{mol}^2 \cdot \text{s})$$

(c)  $\text{rate} = k[\text{Hg}_2\text{Cl}_2] [\text{C}_2\text{O}_4^{2-}]^2$

$$= 1.34 \times 10^{-3} \text{ L}^2/(\text{mol}^2 \cdot \text{s}) \times (0.0400 \text{ mol}/\text{L}) \times (0.150 \text{ mol}/\text{L})^2$$

$$= 1.21 \times 10^{-6} \text{ mol}/(\text{L} \cdot \text{s})$$

14.  $\text{rate} = k[A][B]$

When  $\text{rate} = 0.10 \text{ mol}/\text{L} \cdot \text{s}$ ,  $[A] = 0.20 \text{ mol}/\text{L}$  and  $[B] = 0.050 \text{ mol}/\text{L}$

$$k = \frac{\text{rate}}{[A][B]} = \frac{0.10 \text{ mol}/(\text{L} \cdot \text{s})}{(0.20 \text{ mol}/\text{L})(0.050 \text{ mol}/\text{L})}$$

$$= 10 \text{ L}/(\text{mol} \cdot \text{s})$$

(a)  $[A] = \frac{\text{rate}}{k[B]} = \frac{0.40 \text{ mol}/(\text{L} \cdot \text{s})}{(10 \text{ L}/(\text{mol} \cdot \text{s})) \times (0.050 \text{ mol}/\text{L})}$

$$= 0.80 \text{ mol}/\text{L}$$

(b)  $[B] = \frac{\text{rate}}{k[A]} = \frac{0.80 \text{ mol}/(\text{L} \cdot \text{s})}{(10 \text{ L}/(\text{mol} \cdot \text{s})) \times (0.40 \text{ mol}/\text{L})}$

$$= 0.20 \text{ mol}/\text{L}$$

15. rate law is:  $\text{rate} = k[C][D]^2$

When  $\text{rate} = 0.10 \text{ mol}/(\text{L} \cdot \text{s})$ ,  $[C] = 1.0 \text{ mol}/\text{L}$  and  $[D] = 0.20 \text{ mol}/\text{L}$ .

$$k = \frac{\text{rate}}{[C][D]^2} = \frac{0.10 \text{ mol}/\text{L} \cdot \text{s}}{1.0 \text{ mol}/\text{L} \times (0.20 \text{ mol}/\text{L})^2}$$

$$= 2.5 \text{ L}^2/\text{mol}^2 \cdot \text{s}$$

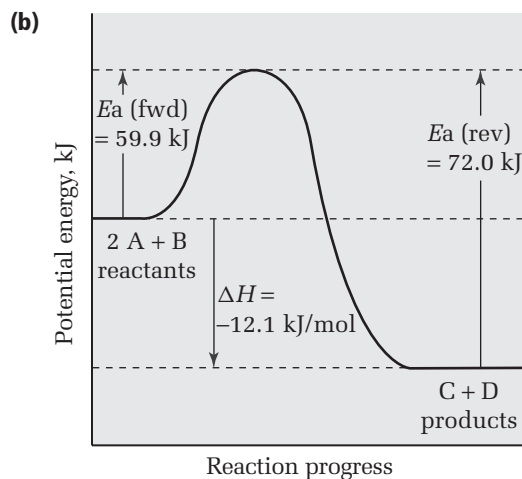
(a)  $\text{rate} = 2.5 \text{ L}^2/(\text{mol}^2 \cdot \text{s}) \times (2.0 \text{ mol}/\text{L}) \times (0.20 \text{ mol}/\text{L})^2$

$$= 0.20 \text{ mol}/(\text{L} \cdot \text{s})$$

(b)  $\text{rate} = 2.5 \text{ L}^2/(\text{mol}^2 \cdot \text{s}) \times (2.0 \text{ mol}/\text{L}) \times (0.40 \text{ mol}/\text{L})^2$

$$= 0.80 \text{ mol}/(\text{L} \cdot \text{s})$$

16. (a)  $\Delta H = 59.9 \text{ kJ}/\text{mol} - 72.0 \text{ kJ}/\text{mol} = -12.1 \text{ kJ}/\text{mol}$



17. (a) Assume a bimolecular reaction in step 1 and each of the other steps are fast,

$$\text{rate} = k[A][B_2]$$

(b) Assume that the intermediate,  $\text{AB}_2$ , forms in a fast step, the rate law is,

$$\text{rate} = k[\text{AB}_2][C]$$

However, the rate law may not include reaction intermediates. Students are not equipped to express  $[\text{AB}_2]$  in terms of  $[A]$  and  $[B_2]$ . Accept any reasonable answers that show students have thought about the reactions.

(c) Assuming that the other steps are very fast, and the intermediate,  $\text{A}_2\text{B}_2$ , forms quickly, the rate law equation may be:

$$\text{rate} = k[\text{A}_2\text{B}_2][A]$$

But  $[\text{A}_2\text{B}_2]$  cannot be included in the equation.

Therefore, students should state they are unable to write the equation.

## Communication

18. Reactions are fastest at the beginning because it is at this time that the reactants are at their highest concentration.

19. The temperature of a reaction must be specified because the rate of a reaction depends upon temperature.

20. Students may suggest any analogy that includes a variety of steps with one step that is clearly slower than the others. For example, five people are available to work on a race car during a pit stop. The car must stop on a marked spot, have all four tires changed, and the driver given a drink of water. You can assume the following steps: Person A directs the car to stop on the marked spot and waits to give the all clear signal to go back on the track; Person B changes tires on the left side; Person C changes the tires on the right side; Person D gets the driver a drink; Person E talks to the driver to make sure there are no other problems. The people changing the tires will take the longest time and determine the overall time taken for the pit stop.

21. Consider a bimolecular reaction. The two particles that collide come together with sufficient energy to break bonds. At the moment they come together, before bonds are broken and new bonds are made, the two particles are loosely associated and are, in a sense, one particle. This is the activated complex. It is a hypothetical particle since it cannot be isolated or detected. This particle is unstable and may break apart and reform the reactant particles or break apart and form a product. (Students may include a potential energy diagram to show that the activated complex forms at the transition state of the reaction.)

An intermediate is a real particle that forms in the course of a reaction. A reaction intermediate is a particle that forms after one step in a mechanism and is used up in a subsequent step in a mechanism.



The overall reaction,  $A + B \rightarrow D$ , does not include the reaction intermediate or the activated complexes.

22. (a) Students should agree with this statement.  
For an endothermic reaction,  
$$E_{a(\text{fwd})} = |E_{a(\text{rev})} + \Delta H|$$
  
In an endothermic reaction,  $\Delta H$  is always positive.  $E_{a(\text{rev})}$  is also positive. Therefore,  $E_{a(\text{fwd})}$  will always be greater than  $\Delta H$ .
- (b) Students should disagree with this statement.  
For an exothermic reaction,  
$$E_{a(\text{fwd})} = |E_{a(\text{rev})} + \Delta H|$$
  
In an exothermic reaction,  $\Delta H$  is negative. It is therefore possible for  $E_{a(\text{fwd})}$  to be larger than or smaller than the magnitude of  $\Delta H$ .
23. (a) Activation energy,  $E_a$ , is the minimum energy that a collision must have for a reaction to take place.  
(b) Temperature represents the average kinetic energy of particles in a substance. At higher temperatures, particles have a higher average kinetic energy. More collisions therefore have energy equal to or greater than the activation energy. Thus, reaction rate usually increases as temperature increases.  
(c) A catalyst provides an alternative mechanism for a reaction. The steps in the catalyst mechanism will likely have activation energies that are smaller than the activation energies in the original mechanism.

## Making Connections

24. Developing a print involves chemical reactions. Since most chemical reactions occur more quickly at a higher temperature, you could speed up the development process by keeping the print inside your coat close to your body to warm it up.
25. The rates of the chemical reactions that occur in a catalytic converter are all temperature dependent. At lower temperatures, the reactions do not occur quickly enough to be effective. The exhaust could be recycled more than once through the converter, to allow for the temperature to increase and the reactions to occur more quickly. Or, an electrical heater could be used to maintain the catalytic converter at a warmer temperature.

## Unit 3 Project

### Developing a Bulletin About Catalysts and Enzymes

Student Textbook pages 314–315

### Tips

- If your students carried out the ThoughtLab on page 305 of the student textbook, they will have already covered expectation EC 3.03.
- Students may need some initial help to focus on a topic for research. You may need to provide a list of possible research topics. For example:
  - enzymes in cheese-making (students can research the history angle here, too)
  - solid-state catalysts in catalytic converters
  - catalysts used in the production of ammonia (e.g., oxides of iron)
  - enzymes used to convert starch into sweeteners such as glucose, maltose, and fructose
  - enzymes used for industrial cleaning (e.g., grease traps, drain lines)
  - enzymes used in juice making
  - enzymes used to hydrolyze fats (e.g., lipase)
  - enzymes used to produce pharmaceuticals
  - catalysts used in petroleum refining