

**TOPIC 3:**  
**CHEMICAL KINETICS**

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### Topic 3: Chemical Kinetics

- C12-3-01** Formulate an operational definition of *reaction rate*.  
Include: examples of chemical reactions that occur at different rates
- C12-3-02** Identify variables used to monitor reaction rates (i.e., change per unit of time,  $\Delta x / \Delta t$ ).  
*Examples: pressure, temperature, pH, conductivity, colour. . .*
- C12-3-03** Perform a laboratory activity to measure the average and instantaneous rates of a chemical reaction.  
Include: initial reaction rate
- C12-3-04** Relate the rate of formation of a product to the rate of disappearance of a reactant, given experimental rate data and reaction stoichiometry.  
Include: descriptive treatment at the particulate level
- C12-3-05** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.  
Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst
- C12-3-06** Use the collision theory to explain the factors that affect the rate of chemical reactions.  
Include: activation energy and orientation of molecules
- C12-3-07** Draw potential energy diagrams for endothermic and exothermic reactions.  
Include: relative rates, effect of a catalyst, and heat of reaction (enthalpy change)
- C12-3-08** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.
- C12-3-09** Explain the concept of a reaction mechanism.  
Include: rate-determining step
- C12-3-10** Determine the rate law and order of a chemical reaction from experimental data.  
Include: zero-, first-, and second-order reactions and reaction rate versus concentration graphs

**Suggested Time: 10 hours**

### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

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*Examples: pressure, temperature, pH, conductivity, colour . . .*

(1 hour)

SLO: C12-3-01  
SLO: C12-3-02

#### SUGGESTIONS FOR INSTRUCTION

##### Activating Activity

Ask students for examples of fast and slow reactions or processes that they encounter in their daily lives. Students may begin with examples of physical changes, such as melting or dissolving. Even though these are not examples of chemical changes, they still reinforce the concept of fast and slow reactions. Try to lead students to consider chemical reactions.

Some examples students may give for fast reactions are explosions, burning gasoline (combustion), precipitation reactions, and neutralization reactions.

Some examples students may give for slow reactions are rusting of metals, baking a cake, ripening of fruit, and growth of a plant.

##### TEACHER NOTES

##### Reaction Rate (C12-3-01)

Chemical kinetics crosses over into many other areas of science and engineering. Rates of metabolic reaction and the progress of reactions involved in growth and bone regeneration are studied by biologists. Automobile engineers want to decrease the rate of rusting of car bodies, while agricultural scientists study the chemical reactions involved in spoilage and decay of foods (see van Kessel, et al. 358).

The speed of any activity (e.g., running, reading, cooking) involves quantifying how much is accomplished in a specific amount of time. We can quantify, or measure, the speed of a chemical reaction (also known as its *reaction rate*).

##### General Learning Outcome Connections

**GLO C2:** Demonstrate appropriate scientific skills when seeking answers to questions.

**GLO C5:** Demonstrate curiosity, skepticism, creativity, open-mindedness, accuracy, precision, honesty, and persistence, and appreciate their importance as scientific and technological habits of mind.

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.

**GLO E3:** Recognize that characteristics of materials and systems can remain constant or change over time, and describe the conditions and processes involved.

**GLO E3:** Recognize that energy, whether transmitted or transformed, is the driving force of both movement and change, and is inherent within materials and in the interactions among them.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-S5:** Collect, record, organize, and display data using an appropriate format.

*Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware . . .*

Operationally, reaction kinetics describes how fast or slow a reactant disappears or a product forms. At this point, an operational definition will involve reaction time as opposed to reaction rate. (Fast reactions have a short reaction time, while slow reactions take a long time.)

**Demonstrations/Laboratory Activities**

Listed below are a number of demonstrations/lab activities illustrating the concept of reaction rate in a chemical reaction. Perform a few demonstrations to help students understand reaction rates.

**■ Reaction Rate**

React magnesium (Mg) metal with 1.0 mol/L hydrochloric acid (HCl). React another piece of Mg metal with 6.0 mol/L HCl.

Ask students the following questions:

1. What happened?
2. How long did both reactions take?
3. Does it matter how much material you have?
4. How can you measure the rate of the reaction?

**■ Electrolysis Reaction (Extension)**

Generate hydrogen and oxygen by electrolysis in a dish of liquid soap. It will give off bubbles of hydrogen and oxygen gas. Remove the gas generator. Ask students whether a reaction is occurring. (Answers may vary.)

Discuss that this electrolysis reaction (splitting up of water to form hydrogen gas and oxygen gas) is occurring spontaneously but at a slow rate. Ask students how we could increase the rate. (Answers will vary.)

Touch the bubbles with a burning wood splint. (You may wish to have it attached to a metre stick.) The reaction happens quickly. (A loud popping sound results.)

**■ Mass Changes**

Find the mass of uniform pieces of gelatin and then place each piece into a separate beaker. Place different pieces of fruit in each of the beakers except the one beaker that contains only the piece of gelatin (serves as the control). Leave the beakers overnight. In the next class, determine the mass of the pieces of gelatin again. Comment on any observations made (see Chastko 403).

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(continued)

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**■ Food Spoilage**

Cut an apple into four slices, each with approximately the same surface area of flesh exposed.

- Dip the first slice in water and place it on a surface. The first slice acts as the control.
- Dip the second slice in lemon juice and place it next to the first slice.
- Place the third slice in the refrigerator, or in a small cooler filled with ice.
- Place the fourth slice in a sealable bag, removing as much air as possible.

Compare the four slices after 10, 20, and 30 minutes, and record the amount of browning that occurs on the apple flesh at each time increment. Discuss observations in relation to what the apple was exposed to.

Comment further on observations with the apple slices, this time in terms of the rate at which the browning of the apple occurs in each sample (see van Kessel, et al. 359).

**■ Decomposition Reaction**

Hydrogen peroxide ( $\text{H}_2\text{O}_2$ ) gradually decomposes to form water and oxygen gas. In this situation, the yeast acts on the hydrogen peroxide to speed up the reaction.

Pour 10 mL of hydrogen peroxide into a beaker and record any observations. Add a “pinch” of yeast to the hydrogen peroxide. Stir gently with a toothpick. Record observations. (The hydrogen peroxide is clear and colourless. When the yeast is added to the hydrogen peroxide, bubbles form, and then the mixture starts to foam.)

Instead of using yeast, use manganese dioxide ( $\text{MnO}_2$ ) to speed up the hydrogen peroxide decomposition reaction.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-S5:** Collect, record, organize, and display data using an appropriate format.

*Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware . . .*

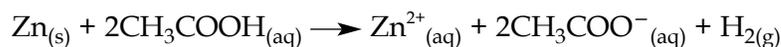
**TEACHER NOTES****Monitoring Reaction Rates (C12-3-02)**

*Reaction rate* is change in an observable property over time. The observable property should be selected based upon what can be measured in the laboratory. This could be a colour change, a temperature change, a pressure change, or the appearance of a new substance. Some common methods of measuring reaction rates involve the use of spectrometers, conductivity apparatus, and manometers (or a simple syringe).

Note that concentration cannot be monitored directly. Emphasize that the observable (measurable) properties described in the following examples can be used to determine the change in concentration over time.

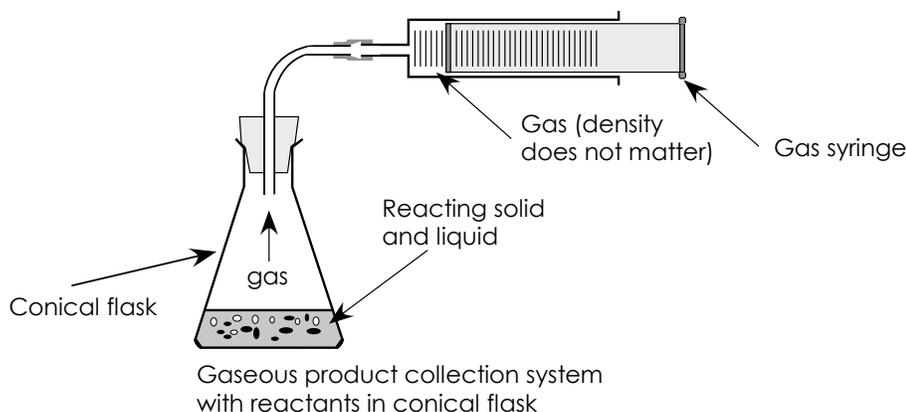
- **Pressure**

A manometer can be used to measure a change in pressure when a reaction results in a change in the number of moles of gas. The reaction between zinc and acetic acid, for example, can be monitored by attaching a manometer to a reaction vessel of known volume that is immersed in a constant-temperature bath.



As  $\text{H}_{2(g)}$  is produced, the gas pressure increases (Silberberg 681).

A simpler method would be to use a gas syringe to measure the reaction rate. See diagram below.



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(continued)

#### ■ Temperature

The following reaction can be monitored by temperature.



colourless      reddish brown

If a sealed tube of  $\text{NO}_2$ – $\text{N}_2\text{O}_4$  is placed in a cold water bath, the dinitrogen tetroxide ( $\text{N}_2\text{O}_4$ ) becomes predominant. The contents of the tube become lighter in colour.

If another tube containing a similar sample of  $\text{NO}_2$ – $\text{N}_2\text{O}_4$  is placed in a hot water bath, the resulting colour change is a reddish brown, indicating a greater presence of  $\text{NO}_2$ .

A sealed tube of  $\text{NO}_2$ – $\text{N}_2\text{O}_4$  can be left at room temperature so students can make the comparison with the tube in a cold water bath, and then with the tube in a hot water bath.

#### ■ The Concept of pH

A pH meter can be used to measure the change in acidity over time. This data can then be used to determine the concentration of hydrogen (hydronium) ion over time.

#### ■ Conductivity

Electrodes can be placed in the reaction mixture and the increase/decrease in conductivity of the products can be used to measure reaction rate. This method is usually used when non-ionic reactants form ionic products (Silberberg 681).

Reaction rate can be calculated by finding the change in formation of product over time, or by finding the change in consumption of a reactant over time.

$$\text{Rate} = \Delta x / \Delta t \text{ (formation of a product)}$$

$$\text{Rate} = -\Delta x / \Delta t \text{ (consumption of a reactant)}$$

Students may confuse reaction rate and reaction time. Emphasize that *reaction rate* describes a change over time, while *reaction time* is the amount of time it takes for a reaction to occur. The two terms are inversely related, as shown by the previous formulas.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

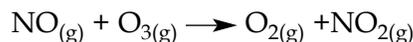
*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-S5:** Collect, record, organize, and display data using an appropriate format.

*Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware . . .*

- **Colour**

A spectrometer can be used to measure the concentration of a reactant or product that absorbs (or gives off) light of a narrow range of wavelengths. An example of this is



colourless

reddish brown

Known amounts of the reactants are injected into a gas sample tube of known volume, and the rate of  $\text{NO}_{2(g)}$  produced is measured by monitoring the colour over time (Silberberg 680).

## SUGGESTIONS FOR ASSESSMENT

### Journal Writing

Students can make journal entries for fast and slow reactions and state their rationale for each.

Ask students to consider questions such as the following:

- What does *rate* mean?
- How can you measure the rate of a reaction?
- Does a reaction always occur at the same rate? Explain.
- Do all reactions occur at the same rate? Explain.

Ask students to provide examples of

- reactions that have different rates of reaction
- reactions that occur at different rates under different conditions
- processes that cannot be controlled
- processes that can be controlled

### Paper-and-Pencil Tasks

1. Students can complete a Compare and Contrast think sheet for fast reactions versus slow reactions (SYSTH 10.15, 10.24).
2. Students can complete a KWL (Know, Want to Know, Learned) strategy sheet on reaction rate (SYSTH 9.8, 9.24).
3. Given a reaction, students can predict what variable (or property) may be most easily monitored.

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*Examples: pressure, temperature, pH, conductivity, colour . . .*

(continued)

**LEARNING RESOURCES LINKS**



- Chemistry* (Chang 532, 533)
- Chemistry* (Zumdahl and Zumdahl 561)
- Chemistry: The Molecular Nature of Matter and Change* (Silberberg 673, 680, 681)
- Glencoe Chemistry: Matter and Change* (Dingrando, et al. 529)
- McGraw-Hill Ryerson Chemistry*, Combined Atlantic Edition (Mustoe, et al. 463, 466)
- McGraw-Hill Ryerson Inquiry into Chemistry* (Chastko, et al. 404)
- Nelson Chemistry 12*, Ontario Edition (van Kessel, et al., 358, 360, 365)
- Prentice Hall Chemistry* (Wilbraham, et al. 540)

**Investigations**

- Glencoe Chemistry: Matter and Change* (Dingrando, et al.)  
Discovery Lab: Speeding Reactions, 529
- McGraw-Hill Ryerson Inquiry into Chemistry* (Chastko, et al.)  
Launch Lab: Does It Gel? 403
- Nelson Chemistry 12*, Ontario Edition (van Kessel, et al.)  
Slowing the Browning Process, 359
- Prentice Hall Chemistry* (Wilbraham, et al.)  
Inquiring Activity: Temperature and Reaction Rates, 540

**Website**

- Brown, W. P. "Factors Affecting the Speed-Rates of Chemical Reactions."  
*Doc Brown's Chemistry*. 2000–2010.  
<[www.docbrown.info/page03/3\\_31rates.htm](http://www.docbrown.info/page03/3_31rates.htm)> (8 Feb. 2012).

**Selecting Learning Resources**

For additional information on selecting learning resources for Grade 11 and Grade 12 Chemistry, see the Manitoba Education website at <[www.edu.gov.mb.ca/k12/learnres/bibliographies.html](http://www.edu.gov.mb.ca/k12/learnres/bibliographies.html)>.

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**NOTES**

**Topic 3:  
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**SPECIFIC LEARNING OUTCOMES**

**C12-3-03:** Perform a laboratory activity to measure the average and instantaneous rates of a chemical reaction.

Include: initial reaction rate

**C12-3-04:** Relate the rate of formation of a product to the rate of disappearance of a reactant, given experimental rate data and reaction stoichiometry.

Include: descriptive treatment at the particulate level

(2.5 hours)

SLO: C12-3-03  
SLO: C12-3-04

**SUGGESTIONS FOR INSTRUCTION**

**Entry-Level Knowledge**

Students studied the stoichiometry of chemical reactions in Grade 11 Chemistry (Topic 3: Chemical Reactions).

**Laboratory Activity**

Have students perform a lab activity to measure the change in mass of calcium carbonate as it reacts with 3 mol/L hydrochloric acid. See Appendix 3.1: Graphical Determination of Reaction Rate: Lab Activity.

Using the data derived from the lab activity (or data given in Appendix 3.1), students can calculate the average rate and the instantaneous rate of a reaction. Students can use software, such as Excel or Graphical Analysis, to plot data and determine instantaneous rate at time = 0 (initial rate) and at other times. Students can compare the rates and hypothesize why the rates change.

**TEACHER NOTES**

**Average Rate of a Chemical Reaction (C12-3-03)**

The *average rate* of a reaction depends on the time interval chosen. Usually this is calculated by dividing the total consumption (or total production) of a substance by the total time it took for the reaction to occur. Refer to the following graph and sample calculation.

**General Learning Outcome Connections**

**GLO C2:** Demonstrate appropriate scientific skills when seeking answers to questions.

**GLO C5:** Demonstrate curiosity, skepticism, creativity, open-mindedness, accuracy, precision, honesty, and persistence, and appreciate their importance as scientific and technological habits of mind.

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.

**GLO D4:** Understand how stability, motion, forces, and energy transfers and transformations play a role in a wide range of natural and constructed contexts.

**SKILLS AND ATTITUDES OUTCOMES**

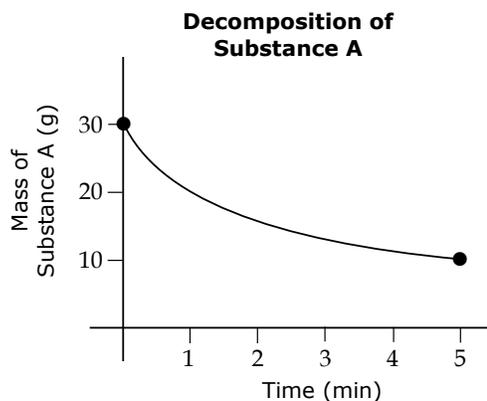
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**C12-0-S6:** Estimate and measure accurately using Système International (SI) and other standard units.

Include: SI conversions and significant figures

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

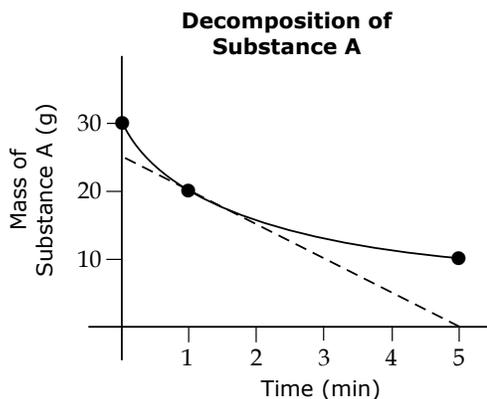


$$\text{Average rate} = \frac{\text{change in the amount of substance A}}{\text{change in time}} = \frac{30 \text{ g} - 10 \text{ g}}{0 \text{ min} - 5 \text{ min}} = 4 \text{ g/min}$$

**Instantaneous Rate of Chemical Reaction (C12-3-03)**

The *instantaneous rate* is the rate of reaction that occurs at a particular instant in time. To calculate this rate, a tangent line is drawn to the point of time on the graph (particular instant of time), and the slope of this line is then calculated.

Refer to the following graph and sample calculation for determining the instantaneous rate at 1 minute.



$$\begin{aligned} \text{Slope} &= \frac{\text{change in the amount of substance A}}{\text{change in time}} = \frac{25 \text{ g} - 0 \text{ g}}{0 \text{ min} - 5 \text{ min}} \\ &= -5 \text{ g/min at } t = 1 \text{ min} \end{aligned}$$

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#### SPECIFIC LEARNING OUTCOMES

**C12-3-03:** Perform a laboratory activity to measure the average and instantaneous rates of a chemical reaction.

Include: initial reaction rate

**C12-3-04:** Relate the rate of formation of a product to the rate of disappearance of a reactant, given experimental rate data and reaction stoichiometry.

Include: descriptive treatment at the particulate level

(continued)

#### Paper Laboratory Activity

If students need additional practice, they can create sample plots with given data. Two sample assignments (with answer keys) are provided in Appendix 3.2A: Chemical Kinetics: Assignment 1 and Appendix 3.3A: Chemical Kinetics: Assignment 2. From the plotted data, students calculate average rates and determine instantaneous rates. They also compare rates and discover that the rate of consumption of each reactant and the formation of each product is related to the stoichiometry of the reaction.

#### TEACHER NOTES

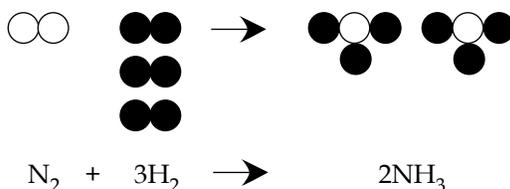
##### Rate and Reaction Stoichiometry (C12-3-04)

The concept of rate and reaction stoichiometry should be introduced carefully. Diagrams of molecules would help students to understand reaction rate at the particulate (molecular) level.

*Example:*

For the reaction  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ , the coefficient in front of the substance determines the rate of consumption or production of that substance, if the initial rate of  $\text{N}_2$  is known.

At the particulate level, this reaction would be expressed as follows:



Students should recognize that for every  $\text{N}_2$  molecule, three  $\text{H}_2$  molecules need to be consumed. This means that the rate of consumption of  $\text{H}_2$  is three times the rate of consumption of  $\text{N}_2$ . In addition, for every molecule of  $\text{N}_2$  that is consumed, the rate of production of  $\text{NH}_3$  molecules is doubled.

## SKILLS AND ATTITUDES OUTCOMES

**C12-0-S5:** Collect, record, organize, and display data using an appropriate format.

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**C12-0-S6:** Estimate and measure accurately using Système International (SI) and other standard units.

*Include: SI conversions and significant figures*

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

Another way to state this is that  $N_2$  is consumed at one-third the rate that  $H_2$  is consumed and at half the rate that  $NH_3$  is produced.

If the rate of one of the species is known, the rates of the other species can be determined from the reaction stoichiometry.

If the rate of consumption of nitrogen is given as

$$\text{Rate} = -\frac{\Delta[N_2]}{\Delta t}$$

then the following is also true:

$$\text{Rate} = -\frac{\Delta[N_2]}{\Delta t} = -\frac{1}{3} \frac{\Delta[H_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[NH_3]}{\Delta t}$$

*Sample Problem:*

For the reaction  $N_2 + 3H_2 \rightarrow 2NH_3$ , if hydrogen reacts at a rate of  $1.5 \text{ mol/L} \cdot \text{s}$ , what is the rate of formation of ammonia?

*Solution:*

Calculate the rate in a manner similar to how stoichiometry was used to determine moles of product formed. Use the ratio of the coefficients to determine the ratio of rates.

$$\begin{aligned} \text{Rate } NH_3 \text{ formation} &= 1.5 \text{ mol/L} \cdot \text{s } H_2 \left( \frac{2NH_3}{3H_2} \right) \\ &= 1.0 \text{ mol/L} \cdot \text{s } NH_3 \end{aligned}$$

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(continued)

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**Animations/Simulations**

Simulations, such as those on the following websites, allow students to determine the rate of reaction at a given point in time. They also show the effect of concentration change, the rate of a chemical reaction, and the determination of stoichiometric coefficients.

*Sample Websites:*

Blauch, David N. *Virtual Chemistry Experiments: Chemical Kinetics*. 2001, 2009. *Chemistry@Davidson*. <[www.chm.davidson.edu/vce/kinetics/index.html](http://www.chm.davidson.edu/vce/kinetics/index.html)> (8 Feb. 2012).

See simulations on the following topics:

- Reaction Rates
- Rate of Reaction

Chemical Education Research Group, Iowa State University. "Chemistry Experiment Simulations and Conceptual Computer Animations." *Chemical Education*. <<http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/simDownload/index4.html>> (22 Nov. 2012).

In the Kinetics section, download and unzip the following animation:

- NO + O<sub>3</sub> Bimolecular Collision

**SUGGESTIONS FOR ASSESSMENT**

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**Laboratory Skills**

A checklist can be used to assess students on the following lab skills:

- collecting and interpreting data
- making and using graphs
- observing, predicting, and recognizing cause and effect

**SKILLS AND ATTITUDES OUTCOMES**

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**C12-0-S6:** Estimate and measure accurately using Système International (SI) and other standard units.

*Include: SI conversions and significant figures*

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**Paper-and-Pencil Tasks**

1. Have students describe pictorially what is happening at the particulate level when a reactant is consumed and a product is formed in a chemical reaction.
2. Have students solve problems on experimental rate data and reaction stoichiometry. See Appendix 3.4A: Chemical Kinetics Problems and Appendix 3.4B: Chemical Kinetics Problems (Answer Key).

**LEARNING RESOURCES LINKS**

*Chemistry* (Chang 534, 537)

*Chemistry* (Zumdahl and Zumdahl 561)

*Chemistry: The Molecular Nature of Matter and Change* (Silberberg 675)

*Glencoe Chemistry: Matter and Change* (Dingrando, et al. 531, 546)

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al., 360, 362)

*Prentice Hall Chemistry* (Wilbraham, et al. 575)

**Investigation**

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al.)

Lab Exercise 6.1.1: Determining a Rate of Reaction, 401

**Websites**

Blauch, David N. "Rate of Reaction." *Virtual Chemistry Experiments: Chemical Kinetics*. 2001, 2009. Chemistry@Davidson. <[www.chm.davidson.edu/vce/kinetics/RateOfReaction.html](http://www.chm.davidson.edu/vce/kinetics/RateOfReaction.html)> (8 Feb. 2012).

Simulations: Reaction Rates  
Rate of Reaction

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Animation: NO + O<sub>3</sub> Bimolecular Collision

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*(continued)*

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**Appendices**

Appendix 3.1: Graphical Determination of Reaction Rate: Lab Activity

Appendix 3.2A: Chemical Kinetics: Assignment 1

Appendix 3.2B: Chemical Kinetics: Assignment 1 (Answer Key)

Appendix 3.3A: Chemical Kinetics: Assignment 2

Appendix 3.3B: Chemical Kinetics: Assignment 2 (Answer Key)

Appendix 3.4A: Chemical Kinetics Problems

Appendix 3.4B: Chemical Kinetics Problems (Answer Key)

**Selecting Learning Resources**

For additional information on selecting learning resources for Grade 11 and Grade 12 Chemistry, see the Manitoba Education website at <[www.edu.gov.mb.ca/k12/learnres/bibliographies.html](http://www.edu.gov.mb.ca/k12/learnres/bibliographies.html)>.

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*Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware . . .*

**C12-0-S6:** Estimate and measure accurately using Système International (SI) and other standard units.

Include: SI conversions and significant figures

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

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**NOTES**

**Topic 3:  
Chemical  
Kinetics**
**SPECIFIC LEARNING OUTCOMES**

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(2 hours)

SLO: C12-3-05  
SLO: C12-3-06

**SUGGESTIONS FOR INSTRUCTION**
**TEACHER NOTES**

At this point, introduce students to the collision theory of chemical reactions. The *collision theory* states that in order for a chemical reaction to occur, the reacting particles must collide. If the particles do not collide, no reaction occurs. Not all collisions, however, produce a chemical reaction. Reacting particles must collide with **sufficient kinetic energy** (called *activation energy*) and the **correct collision geometry or orientation**.

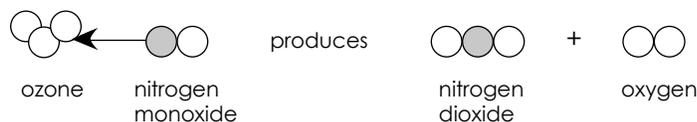
*Activation energy* ( $E_a$ ) is the minimum amount of kinetic energy required for particles to collide effectively, that is, to produce a chemical reaction.

*Example:*

Orientation of nitrogen monoxide molecule **unlikely** to produce a reaction.



Orientation of nitrogen monoxide molecule **likely** to produce a reaction.


**General Learning Outcome Connections**

**GLO C2:** Demonstrate appropriate scientific skills when seeking answers to questions.

**GLO C5:** Demonstrate curiosity, skepticism, creativity, open-mindedness, accuracy, precision, honesty, and persistence, and appreciate their importance as scientific and technological habits of mind.

**GLO C8:** Evaluate, from a scientific perspective, information and ideas encountered during investigations and in daily life.

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.

**GLO D4:** Understand how stability, motion, forces, and energy transfers and transformations play a role in a wide range of natural and constructed contexts.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

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**Animations/Simulations**

Have students view online animations or simulations of chemical reactions.

*Sample Websites:*

Chemical Education Research Group, Iowa State University. "Chemistry Experiment Simulations and Conceptual Computer Animations." *Chemical Education*. <<http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/simDownload/index4.html>> (22 Nov. 2012).

In the Kinetics section, download and unzip the following animation:

■ **NO + O<sub>3</sub> Bimolecular Collision**

This animation shows the correct orientation of molecules upon collision, the reaction being  $O_3 + NO \rightarrow NO_2 + O_2$ . To break apart the ozone molecule (O<sub>3</sub>), the nitrogen atom of the nitrogen monoxide molecule must collide with the correct positioning and sufficient energy to cause the chemical reaction to occur.

University of Colorado at Boulder. "Reactions and Rates." *PhET Interactive Simulations*. <<http://phet.colorado.edu/en/simulation/reactions-and-rates>> (22 Nov. 2012).

This animation allows students to explore the factors that affect reaction rates by changing variables such as concentrations, activation energy, and collision orientation.

**Factors Affecting the Rate of a Chemical Reaction**

Factors affecting the rate of a chemical reaction include the nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst. A discussion of these factors follows.

■ **Collision Theory and the Nature of Reactants**

Some chemical reactions involve the rearrangement of atoms as a result of bonds breaking to form new bonds. Other reactions are a result of electron transfer. The nature of the reactants involved in the reaction will affect the rate of reaction. Reactions that involve ionic compounds and simple ions are usually faster than reactions involving molecular compounds. The fewer the number of bonds broken, the faster the reaction rate will be. The weaker the bonds are in the reactants, the faster the reaction will be. The state of the reactants (solid, liquid, or gas) will also affect the rate of reactions.

**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOMES**

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(continued)

■ **Collision Theory and Surface Area**

From the lab activities suggested for learning outcome C12-3-05, students will observe that increasing the surface area of a solid increases the reaction rate. Collisions can occur only at a solid's surface, so a powdered substance, such as calcium carbonate ( $\text{CaCO}_3$ ), will react more quickly than a large crystal of  $\text{CaCO}_3$  as the powdered substance allows more surface area to be in contact with the other reactants.

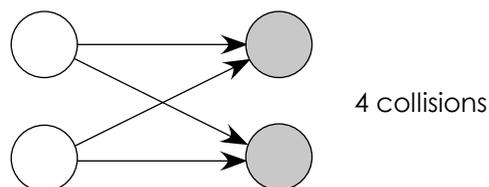
■ **Collision Theory and Concentration (Pressure, Volume)**

The collision theory states that particles must collide with each other to react. If the concentration of one reactant is increased, the reaction rate should increase, as there are more molecules of the increased reactant that can collide.

- At the particulate level, if one molecule of A reacts with two molecules of B, two collisions are possible, which could result in a reaction.



- If the concentration of A is doubled, four collisions are possible, which could result in a reaction.



**SKILLS AND ATTITUDES OUTCOMES**

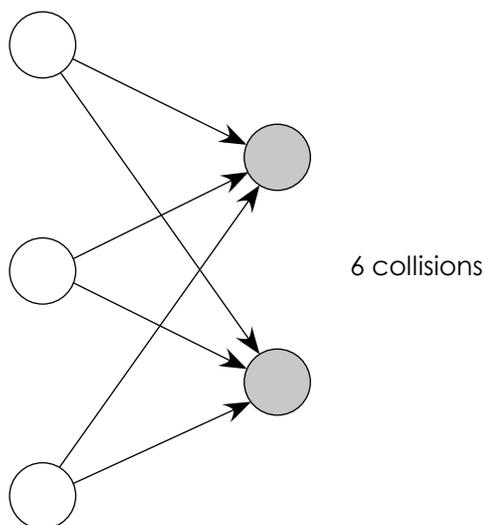
**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

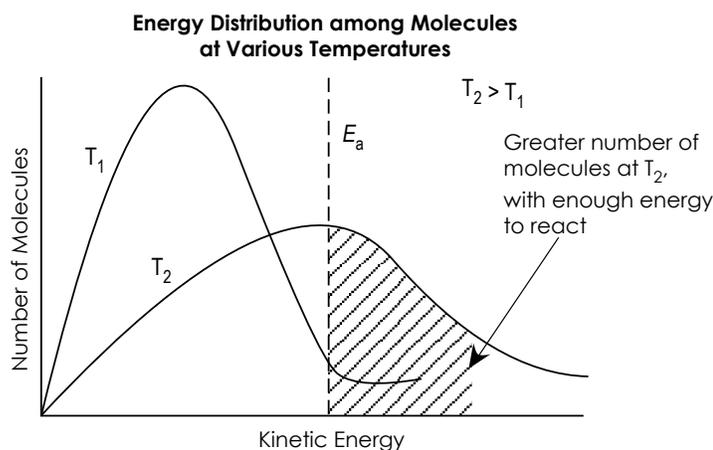
- If the concentration of A is tripled, six collisions are possible, which could result in a reaction.



Increasing the frequency by which collisions can occur in terms of increased concentration results in a faster reaction rate.

- **Effective Collisions and Temperature**

The following graph shows two different temperatures and the number of molecules that have sufficient energy to react. The shaded area under both curves indicate that there are more molecules that have sufficient activation energy at  $T_2$  (higher temperature) than at  $T_1$  (lower temperature) (see van Kessel, et al. 383).



**Topic 3:  
Chemical  
Kinetics**

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**SPECIFIC LEARNING OUTCOMES**

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(continued)

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**Laboratory Activities**

Have students perform lab activities that will lead them to discover the factors that affect the rate of a reaction, rather than perform a verification lab. Some possible lab activities are suggested below.

From the suggested lab activities, students should conclude that

- increasing temperature will increase the rate of a reaction (decreasing reaction time)
- increasing the concentration of reactant(s) will increase the rate of a reaction (Note that pressure and volume are a subset of concentration.)
- increasing the surface area will increase the rate of a reaction
- the presence of a catalyst will increase the rate of a reaction
- the nature (type) of reactants will affect the rate of a reaction

Choose one or more lab activities appropriate for the class.

- **Factors Affecting the Rate of Reactions** (concentration, temperature). See Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity and Appendix 3.5B: Factors Affecting the Rate of Reactions: Lab Activity (Answer Key).

This is a version of the classic Iodine Clock Reaction lab activity in which excess iodine reacts with starch to produce a blue-black product only when the reaction is complete. In this lab activity, students investigate the effects of concentration and temperature on reaction rate.

In Part A, students change the concentration of one reactant, and time how long it takes for the sudden and dramatic colour change to occur.

In Part B, students investigate the role of temperature in reaction rate by running a series of reactions in water baths at different temperatures.

Students produce graphs of their data, draw conclusions about the relationship between these variables, and explain the differences in reaction rate using the collision theory.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

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- **Factors Affecting the Rate of a Reaction** (concentration, nature of reactants, temperature, catalyst, surface area). See Appendix 3.6A: Factors Affecting the Rate of a Reaction: Lab Activity. Teacher notes are provided in Appendix 3.6B.

In Part A of this lab activity, students study the effect of the nature of reactants on reaction time. Several different metals are reacted with hydrochloric acid, and observations are made with regards to reaction time. Students also study the effect of different solutions reacting with magnesium metal on reaction time.

In Part B, students examine the effect of surface area on reaction time. Mossy zinc and powdered zinc are combined with hydrochloric acid, and the reaction times are recorded. Chips of calcium carbonate and powdered calcium carbonate are reacted with hydrochloric acid, and the reaction times are recorded.

In Part C, students study the effect of temperature on a chemical reaction. A solution of potassium permanganate is combined with oxalic acid, and the reaction time is recorded. A second test tube containing just the potassium permanganate is heated in a hot water bath. Then the oxalic acid is added to the test tube in the hot water bath, and the resulting reaction time is recorded.

Students then set up three test tubes containing hydrochloric acid. One test tube is placed in cold water, the second test tube is kept at room temperature, and the third test tube is placed in a hot water bath. Three identical pieces of magnesium are added to each of the three test tubes, and the resulting reaction times are noted.

In Part D, students use a catalyst to study its effect on reaction time. Potassium permanganate is placed in two test tubes. In one of the test tubes, manganese(II) sulphate is added (catalyst). Then oxalic acid is added to both test tubes, and the reaction times are noted.

**Topic 3:  
Chemical  
Kinetics****SPECIFIC LEARNING OUTCOMES**

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(continued)

- **Experiment 23: Factors Affecting the Rate of a Chemical Reaction** (Waterman and Thompson, *Prentice Hall Chemistry: Small-Scale Chemistry Laboratory Manual* 197)

In this four-part lab activity, students study the effects of temperature, surface area, and concentration on the rate of a chemical reaction.

- Students begin by exploring the rate of reaction between hydrochloric acid and magnesium, calcium carbonate, and sodium hydrogen carbonate.
  - They then study the effect of temperature, using the same reactants that were used initially. Cold hydrochloric acid and warm hydrochloric acid are separately reacted with magnesium, calcium carbonate, and sodium hydrogen carbonate.
  - Students continue by investigating the effect of surface area on reaction rate. Hydrochloric acid is reacted with a piece of magnesium, crushed magnesium, a piece of calcium carbonate, and crushed calcium carbonate.
  - Finally, students look at the effect of concentration on reaction rate. Various concentrations of hydrochloric acid are separately reacted with magnesium, calcium carbonate, and sodium hydrogen carbonate.
- **Experiment 36: Factors Affecting Reaction Rates** (Wilbraham, Staley, and Matta, *Prentice Hall Chemistry: Laboratory Manual* 225)
- In this experiment, students investigate factors that can speed up or slow down chemical reactions. They examine the effect of temperature, reactant concentration, particle size, catalysts, and surface area on reaction rate.
- **Chemlab 17: Concentration and Reaction Rate** (Dingrando, et al., *Glencoe Chemistry: Matter and Change* 550)
- In this lab activity, students investigate the effect of concentration on reaction rate. Pieces of magnesium ribbon are reacted separately with varying concentrations of hydrochloric acid, and the resulting reaction time is recorded.
- **MiniLAB 17: Examining Reaction Rate and Temperature** (Dingrando, et al., *Glencoe Chemistry: Matter and Change* 539)
- In this lab activity, students observe the effect of temperature on reaction rate. They dissolve antacid tablets in water at room temperature, at 50°C, and at 65°C.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

- **Investigation 12–A: Factors Affecting the Rate of a Reaction** (Mustoe, et al., *McGraw-Hill Ryerson Chemistry* 464)

In this three-part investigation, students predict and observe the effects of changes to concentration, temperature, reactant, and surface area on the rate of a chemical reaction. In each part, students record the time taken to collect test tubes full of carbon dioxide and calculate the average rate in mL/s.

- Part 1 investigates the effect of concentration on a reaction. Sodium hydrogen carbonate ( $\text{NaHCO}_3$ ) and varying concentrations of vinegar are reacted in four trials.
- Part 2 demonstrates the effect of temperature on reaction rate. Using the same reactants as in Part 1, students perform two trials. Before the reactants are combined, they are first cooled to about  $10^\circ\text{C}$  below room temperature, and then heated to  $10^\circ\text{C}$  above room temperature.
- Part 3 shows the effect of reactants and surface area on reaction rate. Students perform two trials, first combining powdered calcium carbonate ( $\text{CaCO}_3$ ) with vinegar, and then combining solid  $\text{CaCO}_3$  with vinegar.

In their investigations, students should comment on the effects of each factor on reaction rate. If students have not observed factors, provide them with demonstrations to illustrate the factors. In the post-lab discussion, have students explain their observations based on the collision theory.

**Laboratory Demonstrations**

Teachers can choose to demonstrate lab activities such as the following:

- **Experiment 20: A Study of Reaction Rates: The “Clock Reaction”** (Merrill, Parry, and Tellefsen, *Chemistry: Experimental Foundations, Laboratory Manual* 62)

In this two-part experiment, demonstrate the role of concentration and temperature changes on reaction rate.

- **Surface Area and Reaction Rate**

The purpose of this demonstration is to have students observe the effect of an increase in surface area on the rate of a chemical reaction. Place 2 g of lycopodium powder (or starch) in a pile on a porcelain tile. Try to ignite the pile with a burner or lighter. There will be no reaction. Lift the ceramic tile holding the lycopodium powder (or starch) and sprinkle the powder over a lit burner. The powder will ignite quite explosively. Students should observe that the reaction rate increases as surface area increases (Smoot, Price, and Smith 442).

### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(continued)

#### ■ Catalyst and Reaction Rate

In this demonstration, have students observe the effect of a catalyst on the rate of a chemical reaction. Dissolve 25 g of sodium potassium tartrate (Rochelle's salt) in 300 mL of water in a large beaker. Add 100 mL of 3% to 6% hydrogen peroxide ( $\text{H}_2\text{O}_2$ ) to the beaker. Heat the solution to  $70^\circ\text{C}$ . Students should observe that no reaction occurs. Add the catalyst, cobalt chloride, to the beaker. The solution will turn pink and then a greenish colour (cobalt[II] tartrate complex). After the reaction has been completed, the pink colour in the solution will reappear. The cobalt chloride was not consumed in the reaction. Students should observe that the solution at  $70^\circ\text{C}$  did not chemically react until the catalyst was added (Smoot, Price, and Smith 444).

#### Animations/Simulations

Use a variety of online simulations and video clips, such as the following, to demonstrate how various factors affect the rate of chemical reactions.

Chemical Education Research Group, Iowa State University. "Chemistry Experiment Simulations and Conceptual Computer Animations." *Chemical Education*. <<http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/simDownload/index4.html>> (22 Nov. 2012).

In the Kinetics section, download and unzip the following simulation:

- Arrhenius Equation: Temperature, Rate Constant, and Activation Energy Experiment

In this simulation, students can vary the concentration of reactants and the temperature. Students must start the time clock and wait for the reaction to reach completion (blue-black colour).

The North Carolina School of Science and Mathematics (NCSSM). "Chapter 15: Kinetics." *Chemistry Online Resource Essentials (CORE)*. <[www.dlt.ncssm.edu/core/c15.htm](http://www.dlt.ncssm.edu/core/c15.htm)> (9 Feb. 2012).

The following video clips are available on this website:

- *Homogeneous Catalyst* shows how the presence of a catalyst affects reaction rate. Specifically, it shows the decomposition of hydrogen peroxide ( $\text{H}_2\text{O}_2$ ), using a solution of  $\text{Co}^{2+}$ .

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

- *KI Catalyzed H<sub>2</sub>O<sub>2</sub> Decomposition* shows how the addition of a catalyst affects reaction rate. Specifically, it shows the decomposition of H<sub>2</sub>O<sub>2</sub>, catalyzed with manganese dioxide (MnO<sub>2</sub>) and uncatalyzed.
- *Glow Sticks* shows how temperature affects reaction rate. One Glow Stick is placed in hot water and another is placed in cold water.
- *Potato Catalyzed H<sub>2</sub>O<sub>2</sub> Decomposition* shows how surface area affects reaction rate. Small pieces of potato are placed in a test tube containing H<sub>2</sub>O<sub>2</sub>. A small amount of detergent is placed in each test tube to make the bubbles of oxygen more visible.
- *Dust Explosion* shows the effect of surface area on reaction rate. The video clip shows the explosive nature of flour when placed in a closed container and then ignited with a candle.

Petrucci, Ralph H., William S. Harwood, and Geoffrey Herring. "Chapter 15: Chemical Kinetics." *General Chemistry: Principles and Modern Applications*. 8th ed. Prentice Hall, Inc.  
<[http://cwx.prenhall.com/petrucci/medialib/media\\_portfolio/15.html](http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html)>  
(8 May 2012).

The following simulation is available on this website (in the Instructor's Media Portfolio of Prentice Hall's *Companion Website for General Chemistry*):

- *CFCs and Stratospheric Ozone* shows the catalytic decomposition of ozone by chlorine atoms from CFCs.

University of Colorado at Boulder. "Reactions and Rates." *PhET Interactive Simulations*. <<http://phet.colorado.edu/en/simulation/reactions-and-rates>>  
(22 Nov. 2012).

**SUGGESTIONS FOR ASSESSMENT****Paper-and-Pencil Tasks**

1. Have students compare and contrast the rate at which a sugar cube dissolves in cold water and the rate at which granulated sugar dissolves in warm water. Students could include observations of how surface area and water temperature might affect the rate at which each substance dissolves (Fisher 238).
2. Have students describe how the collision theory would apply to a demolition derby (Fisher 236).

### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

(continued)

#### Visual Displays

Students can represent a reaction between two substances, such as nitrogen monoxide (NO) and ozone (O<sub>3</sub>), using ball-and-stick molecular models. Students can show the correct orientation of the molecules as they collide to produce nitrogen dioxide (NO<sub>2</sub>) and oxygen (O<sub>2</sub>). They can also show the incorrect orientation of the molecules that would not produce a reaction.

#### Laboratory Report

The lab activities could be assessed by having students use the Laboratory Report Outline or complete a Laboratory Report Frame (SYSTH 11.38, 14.12). Also refer to the Lab Report Assessment rubric in Appendix 11.

#### Laboratory Skills

Periodically and randomly review students' lab skills using a variety of rubrics and checklists (see SYSTH 6.10, 6.11).

#### Research and Reports

Students could research and report on how the rate of specific chemical processes can be controlled. As an alternative to preparing a report, students could complete an Article Analysis Frame on a related article (SYSTH 11.30, 11.40, 11.41).

#### LEARNING RESOURCES LINKS



*Chemistry* (Chang 554, 566)

*Chemistry* (Zumdahl and Zumdahl 587)

*Chemistry: The Molecular Nature of Matter and Change* (Silberberg 674, 694, 706)

*Glencoe Chemistry: Matter and Change* (Dingrando, et al. 532, 536)

*Glencoe Chemistry: Matter and Change, Science Notebook* (Fisher 236, 238)

*McGraw-Hill Ryerson Chemistry, Combined Atlantic Edition* (Mustoe, et al. 464, 470)

*McGraw-Hill Ryerson Inquiry into Chemistry* (Chastko, et al. 404)

*Merrill Chemistry: A Modern Course* (Smoot, Price, and Smith 442)

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al. 367, 383)

*Prentice Hall Chemistry* (Wilbraham, et al. 541, 545)

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

**Investigations**

*Chemistry: Experimental Foundations, Laboratory Manual* (Merrill, Parry, and Tellefsen)

Experiment 20: A Study of Reaction Rates: The “Clock Reaction,” 62

*Glencoe Chemistry: Matter and Change* (Dingrando, et al.)

Chemlab 17: Concentration and Reaction Rate, 550

MiniLAB 17: Examining Reaction Rate and Temperature, 539

*McGraw-Hill Ryerson Chemistry, Combined Atlantic Edition* (Mustoe, et al.).

Investigation 12–A: Factors Affecting the Rate of a Reaction, 464

*Prentice Hall Chemistry: Laboratory Manual* (Wilbraham, Staley, and Matta)

Factors Affecting Reaction Rates, 225

(temperature, reactant concentration, particle size, catalysis, and surface area)

*Prentice Hall Chemistry: Small-Scale Chemistry Laboratory Manual* (Waterman and Thompson)

Experiment 28: Factors Affecting the Rate of a Chemical Reaction, 197

(temperature, concentration, and surface area)

**Websites**

Chemical Education Research Group, Iowa State University. “Chemistry Experiment Simulations and Conceptual Computer Animations.” *Chemical Education*. <<http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/simDownload/index4.html>> (22 Nov. 2012).

Simulation: Arrhenius Equation: Temperature, Rate Constant, and Activation Energy Experiment

Animation: NO + O<sub>3</sub> Bimolecular Collision

The North Carolina School of Science and Mathematics (NCSSM). “

Chapter 15: Kinetics.” *Chemistry Online Resource Essentials (CORE)*.

<[www.dlt.ncssm.edu/core/c15.htm](http://www.dlt.ncssm.edu/core/c15.htm)> (9 Feb. 2012).

Petrucci, Ralph H., William S. Harwood, and Geoffrey Herring. “Chapter 15: Chemical Kinetics.” *General Chemistry: Principles and Modern Applications*. 8th ed. Prentice Hall, Inc.

<[http://cwx.prenhall.com/petrucci/medialib/media\\_portfolio/15.html](http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html)> (8 May 2012).

University of Colorado at Boulder. “Reactions and Rates.” *PhET Interactive Simulations*. <<http://phet.colorado.edu/en/simulation/reactions-and-rates>> (22 Nov. 2012).

**Topic 3:  
Chemical  
Kinetics**

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**SPECIFIC LEARNING OUTCOMES**

**C12-3-05:** Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.

Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

**C12-3-06:** Use the collision theory to explain the factors that affect the rate of chemical reactions.

Include: activation energy and orientation of molecules

*(continued)*

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**Appendices**

Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity

Appendix 3.5B: Factors Affecting the Rate of Reactions: Lab Activity  
(Answer Key)

Appendix 3.6A: Factors Affecting the Rate of a Reaction: Lab Activity

Appendix 3.6B: Factors Affecting the Rate of a Reaction: Lab Activity  
(Teacher Notes)

**Selecting Learning Resources**

For additional information on selecting learning resources for Grade 11 and Grade 12 Chemistry, see the Manitoba Education website at <[www.edu.gov.mb.ca/k12/learnres/bibliographies.html](http://www.edu.gov.mb.ca/k12/learnres/bibliographies.html)>.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-S2:** State a testable hypothesis or prediction based on background data or on observed events.

**C12-0-S7:** Interpret patterns and trends in data, and infer and explain relationships.

**C12-0-S9:** Draw a conclusion based on the analysis and interpretation of data.

Include: cause-and-effect relationships, alternative explanations, and supporting or rejecting a hypothesis or prediction

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**NOTES**

### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

**C12-3-07:** Draw potential energy diagrams for endothermic and exothermic reactions.

Include: relative rates, effects of catalyst, and heat of reaction (enthalpy change)

**C12-3-08:** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

(2 hours)

SLO: C12-3-07  
SLO: C12-3-08

#### SUGGESTIONS FOR INSTRUCTION

##### Entry-Level Knowledge

In Grade 10 Science (S2-3-09), students learned about kinetic and potential energy with respect to motion. In Grade 11 Chemistry (C11-1-02), students were introduced to the kinetic molecular theory to explain the properties of gases.

##### TEACHER NOTES

An *exothermic reaction* is a chemical reaction that releases energy into the environment. Combustion, or burning, is an example of an exothermic reaction. On the other hand, an *endothermic reaction* is a chemical reaction that absorbs energy from its surroundings, which is stored in the products that have formed. For example, if aluminum chloride is dissolved in water, the beaker will feel cool to the touch.

Students are expected to draw potential energy diagrams indicating the amount of potential energy the reactants and the products have, the activation energy ( $E_a$ ) needed, the activated complex, and the change in enthalpy ( $\Delta H$ ) or the heat of reaction—that is, how much heat is absorbed (endothermic reaction) or how much heat is released (exothermic reaction).

The activation energy of a reaction dictates the relative rate of a reaction. The higher the activation energy is, the slower the reaction rate is, and vice versa. Catalysts increase reaction rates by reducing the activation energy. Catalysts do not affect the heat of reaction.

##### Demonstration

For the kinesthetic learner, demonstrate the following:

1. Roll a ball up an incline and let the ball roll back down. The ball represents the reactants that do not have enough activation energy to reach the activated complex.

#### General Learning Outcome Connections

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.

## SKILLS AND ATTITUDES OUTCOMES

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

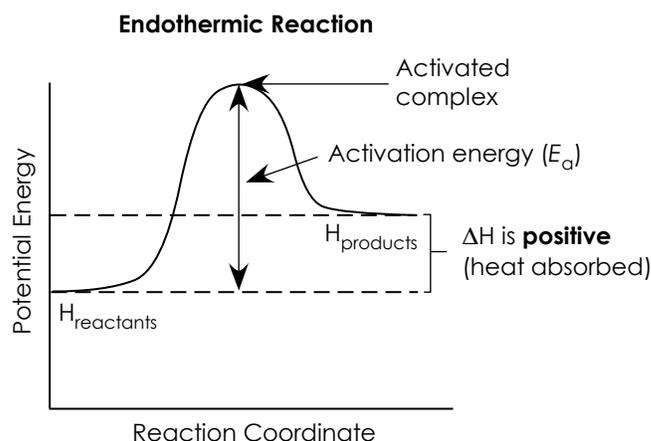
*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

- Roll a ball up a shallower incline and allow the ball to roll over the edge of the incline. The shallower incline represents the addition of a catalyst, which lowers activation energy and allows the reaction to proceed (Dingrando, et al, *Glencoe Chemistry: Matter and Change, Teacher Wraparound Edition* 540).

## Potential Energy Diagrams

Students can use the collision theory and kinetic energy and potential energy diagrams to explain their observations from the lab investigations performed in relation to specific learning outcome C12-3-02. Students' explanations should include observations of what is happening at the **molecular** level.

The following diagram shows the progress of an *endothermic reaction*.



In this diagram, the reactants contain a certain amount of potential energy. As the reaction proceeds from left to right, the molecules of the reactants gain more energy, which is called *activation energy*. If the reactants have sufficient energy to reach the activated complex, then bond breakage and realignment can occur and new substances are formed. The products that have formed have a greater amount of potential energy than the reactants had. This means that energy was absorbed during the chemical reaction from its surroundings. If this reaction had taken place in a beaker, the beaker would have felt cool to the touch. The heat of reaction, or enthalpy change, is a positive value because the potential energy of the products is larger than the potential energy of the reactants.

$$\Delta H = H_{\text{products}} - H_{\text{reactants}} = \text{positive value} = \text{heat is absorbed}$$

### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

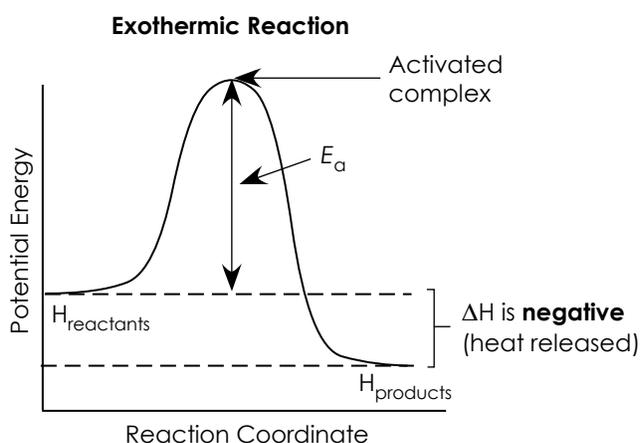
**C12-3-07:** Draw potential energy diagrams for endothermic and exothermic reactions.

Include: relative rates, effects of catalyst, and heat of reaction (enthalpy change)

**C12-3-08:** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

(continued)

The following diagram shows the progress of an *exothermic reaction*.



In this diagram, the reactants contain a certain amount of potential energy. As the reaction proceeds from left to right, the molecules of the reactants gain more energy, which is called activation energy. If the reactants have sufficient energy to reach the activated complex, then bond breakage and realignment can occur and new substances are formed. The products that have formed have a lower amount of potential energy than the reactants had. This means that energy was released during the chemical reaction to its surroundings. If this reaction had taken place in a beaker, the beaker would have felt warm to the touch.

$$\Delta H = H_{\text{products}} - H_{\text{reactants}} = \text{negative value} = \text{heat is released}$$

**SKILLS AND ATTITUDES OUTCOMES**

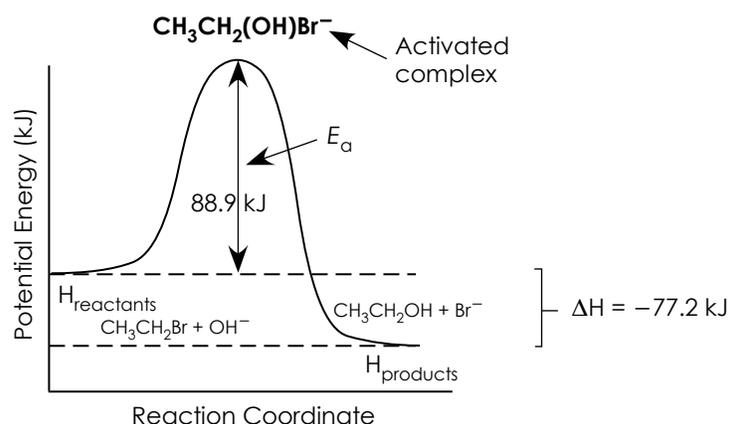
**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

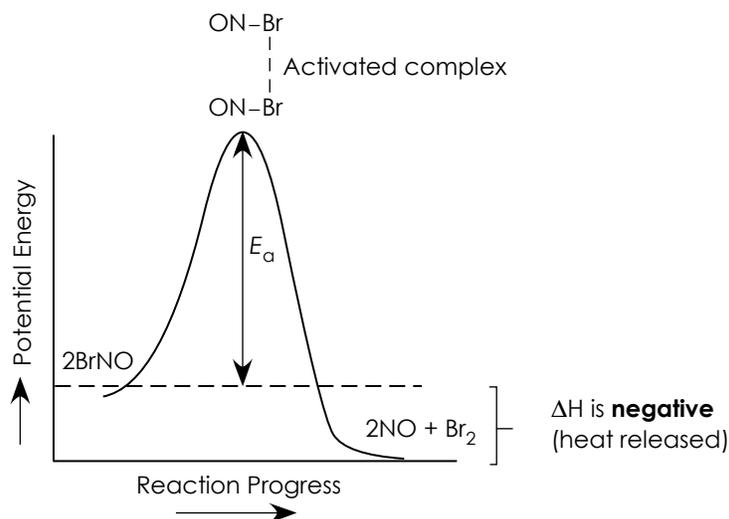
*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

The following potential energy diagram indicates the reaction  
 $\text{CH}_3\text{CH}_2\text{Br} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{Br}^-$ .



Students should be able to indicate on the potential energy diagram the potential energy of the reactants, the potential energy of the products, the activation energy, the location of the activated complex, and the heat of reaction, or enthalpy change.

The following potential energy diagram for the reaction  $2\text{BrNO} \rightarrow 2\text{NO} + \text{Br}_2$  shows the transition state where the molecules of nitrogen, bromine, and oxygen are rearranged to form the products.



### Topic 3: Chemical Kinetics

#### SPECIFIC LEARNING OUTCOMES

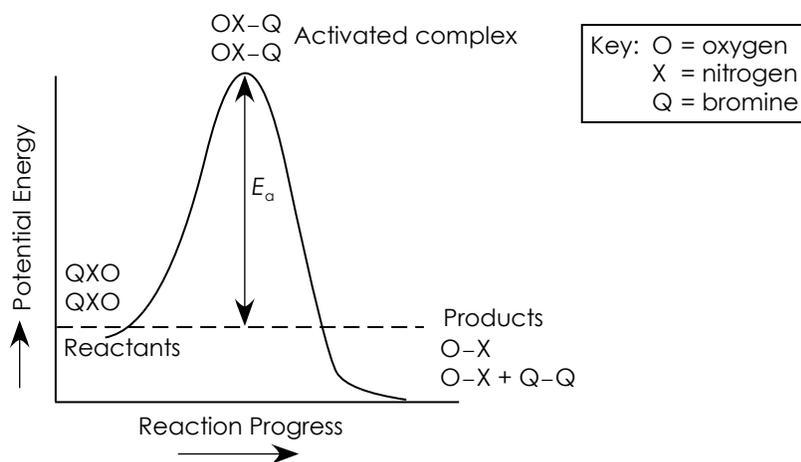
**C12-3-07:** Draw potential energy diagrams for endothermic and exothermic reactions.

Include: relative rates, effects of catalyst, and heat of reaction (enthalpy change)

**C12-3-08:** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

(continued)

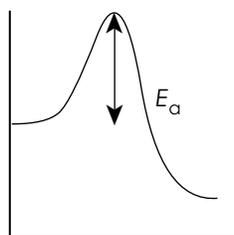
At the particulate level, this is how the potential energy diagram would appear for the chemical reaction just described (Zumdahl and Zumdahl 588):



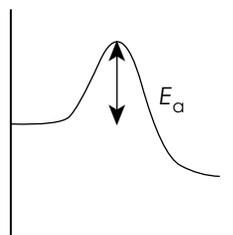
#### Relative Rates

Teachers may wish to use potential energy diagrams to describe whether a reaction is slow, medium, or fast.

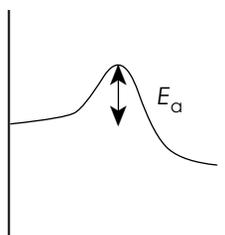
#### Relative Rates of Reaction



**SLOW**  
Notice  $E_a$  is very large for this reaction. It would take a lot of energy to get this reaction to go to completion.



**MEDIUM**  
Notice  $E_a$  is a bit smaller than the slow reaction.



**FAST**  
Notice  $E_a$  is very small when compared to the slow and medium reactions.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

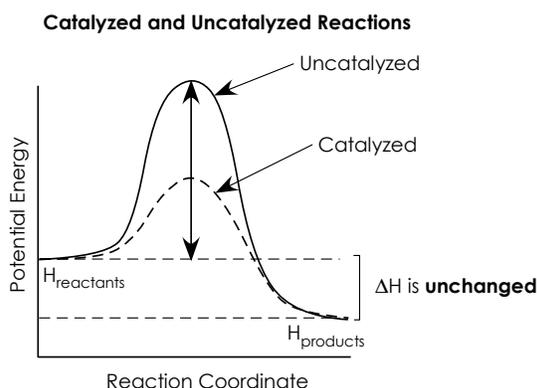
*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

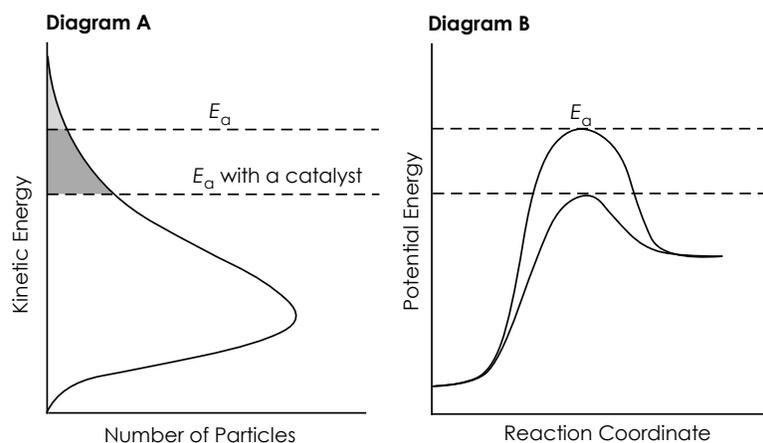
**Catalyst Added in a Reaction**

The following potential energy diagram shows an uncatalyzed reaction and a catalyzed reaction.



Students should have concluded from their lab activities (in relation to learning outcome C12-3-05) that when a catalyst is added to a chemical reaction the reaction rate increases (the reaction time is shorter). Students should note that the diagram indicating the presence of a catalyst shows that a smaller activation energy is required. They should note that the heat of reaction, or enthalpy change, does not change.

In Diagram A below, the catalyst makes it possible for more particles to have sufficient kinetic energy to reach the activated complex. The activation energy is lowered, meaning that more particles are available to collide and form new product. Diagram B shows that the activation energy is lowered, enabling more collisions to occur. This results in more product being formed.



**Topic 3:  
Chemical  
Kinetics**

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**SPECIFIC LEARNING OUTCOMES**

**C12-3-07:** Draw potential energy diagrams for endothermic and exothermic reactions.

Include: relative rates, effects of catalyst, and heat of reaction (enthalpy change)

**C12-3-08:** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

(continued)

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**Collision Theory and Factors Affecting Rate of Reactions**

In addressing learning outcomes C12-3-07 and C12-3-08, see the learning activities suggested for learning outcomes C12-3-05 and C12-3-06.

**Demonstrations/Animations**

A variety of demonstrations/animations can be viewed online to reinforce the effects of factors affecting the rate of chemical reactions.

*Sample Websites:*

The North Carolina School of Science and Mathematics (NCSSM). "Chapter 15: Kinetics." *Chemistry Online Resource Essentials (CORE)*. <[www.dlt.ncssm.edu/core/c15.htm](http://www.dlt.ncssm.edu/core/c15.htm)> (9 Feb. 2012).

The following video clips, available on this website, can help students describe the factors affecting chemical reaction rates.

- *KI Catalyzed H<sub>2</sub>O<sub>2</sub> Decomposition* shows how the addition of a catalyst affects reaction rate. Specifically, it shows the decomposition of H<sub>2</sub>O<sub>2</sub>, catalyzed with manganese dioxide (MnO<sub>2</sub>) and uncatalyzed.
- *Glow Sticks* shows how temperature affects reaction rate. One Glow Stick is placed in hot water and another is placed in cold water.
- *Potato Catalyzed H<sub>2</sub>O<sub>2</sub> Decomposition* shows how surface area affects reaction rate. Small pieces of potato are placed in a test tube containing H<sub>2</sub>O<sub>2</sub>. A small amount of detergent is placed in each test tube to make the bubbles of oxygen more visible.
- *Dust Explosion* shows the effect of surface area on reaction rate. The video clip shows the explosive nature of flour when placed in a closed container and then ignited with a candle.

**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

———. Distance Education and Extended Programs. “Chemistry – Kinetics.” *Teacher’s Instructional Graphics and Educational Resource (TIGER)*.  
<[www.dlt.ncssm.edu/tiger/chem5.htm#kinetics](http://www.dlt.ncssm.edu/tiger/chem5.htm#kinetics)> (10 Feb. 2012).

The following animations are available on this website:

- *Molecular\_collision\_E<sub>a</sub>* (.exe or .html) demonstrates the change in potential energy for two molecules as they collide. Low energy and high energy simulations are shown.
- *Catalys\_2* (.exe or .html) shows the solid state catalytic hydrogenation of an alkene.

Petrucci, Ralph H., William S. Harwood, and Geoffrey Herring. “Chapter 15: Chemical Kinetics.” *General Chemistry: Principles and Modern Applications*. 8th ed. Prentice Hall, Inc.

<[http://cwx.prenhall.com/petrucci/medialib/media\\_portfolio/15.html](http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html)> (8 May 2012).

The following simulation is available on this website (in the Instructor’s Media Portfolio of Prentice Hall’s *Companion Website for General Chemistry*):

- *CFCs and Stratospheric Ozone* shows the catalytic decomposition of ozone by chlorine atoms from CFCs.

Science Bob. “Crazy Foam Experiment.” *Science Bob Videos*.

<[www.sciencebob.com/experiments/videos/video-foam1.php](http://www.sciencebob.com/experiments/videos/video-foam1.php)> (13 Jan. 2012).

This experiment shows how the presence of a catalyst affects reaction rate.

**SUGGESTIONS FOR ASSESSMENT****Paper-and-Pencil Tasks**

Students should be able to interpret and draw potential energy diagrams from given information.

**Journal Writing**

Students can interpret graphs by answering the following questions:

- Are the reactants or the products at a higher energy level?
- Is energy absorbed or released after the reaction takes place?
- Will the reaction always proceed to form products once the activated complex is formed? Explain.

**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOMES**

- C12-3-07:** Draw potential energy diagrams for endothermic and exothermic reactions.  
Include: relative rates, effects of catalyst, and heat of reaction (enthalpy change)
- C12-3-08:** Describe qualitatively the relationship between the factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

(continued)

**LEARNING RESOURCES LINKS**



*Chemistry* (Chang 566)

*Chemistry* (Zumdahl and Zumdahl 588)

*Chemistry: The Molecular Nature of Matter and Change* (Silberberg 696, 698)

*Glencoe Chemistry: Concepts and Applications* (Phillips, Strozak, and Wistrom 713)

*Glencoe Chemistry: Matter and Change* (Dingrando, et al. 534, 540)

*Glencoe Chemistry: Matter and Change, Teacher Wraparound Edition* (Dingrando, et al. 540)

*McGraw-Hill Ryerson Chemistry, Combined Atlantic Edition* (Mustoe, et al. 472)

*McGraw-Hill Ryerson Inquiry into Chemistry* (Chastko, et al. 404)

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al. 384, 392)

*Prentice Hall Chemistry* (Wilbraham, et al. 543)

**Websites**

The North Carolina School of Science and Mathematics (NCSSM).

“Chapter 15: Kinetics.” *Chemistry Online Resource Essentials (CORE)*.  
<[www.dlt.ncssm.edu/core/c15.htm](http://www.dlt.ncssm.edu/core/c15.htm)> (9 Feb. 2012).

\_\_\_\_\_. Distance Education and Extended Programs. “Chemistry – Kinetics.”  
*Teacher’s Instructional Graphics and Educational Resource (TIGER)*.  
<[www.dlt.ncssm.edu/tiger/chem5.htm#kinetics](http://www.dlt.ncssm.edu/tiger/chem5.htm#kinetics)> (10 Feb. 2012).

Petrucci, Ralph H., William S. Harwood, and Geoffrey Herring. “Chapter 15: Chemical Kinetics.” *General Chemistry: Principles and Modern Applications*. 8th ed. Prentice Hall, Inc.  
<[http://cwx.prenhall.com/petrucci/medialib/media\\_portfolio/15.html](http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html)> (8 May 2012).

Science Bob. “Crazy Foam Experiment.” *Science Bob Videos*.  
<[www.sciencebob.com/experiments/videos/video-foam1.php](http://www.sciencebob.com/experiments/videos/video-foam1.php)> (13 Jan. 2012).

**Selecting Learning Resources**

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**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

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**NOTES**

**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOME**

**C12-3-09:** Explain the concept of a reaction mechanism.

Include: rate-determining step

(0.5 hour)

SLO: C12-3-09

**SUGGESTIONS FOR INSTRUCTION**

**TEACHER NOTES**

**Reaction Mechanism**

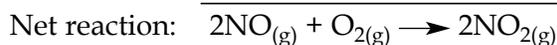
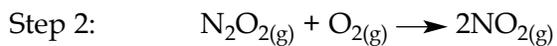
Teachers may wish to explain the concept of a reaction mechanism using the analogy of cleaning up dinner dishes by hand. This process happens in many steps: clearing the table, filling the sink with water and soap, placing the dishes in the sink, washing the dishes, drying the dishes, putting the dishes away, draining the sink, and wiping up.

Students need to be aware that an overall balanced chemical equation does not tell us much about the actual pathway a chemical reaction follows, just as an average speed of 100 km/h does not tell us much about the various speeds we need to drive on a two-hour trip.

A *reaction mechanism* summarizes the individual steps a reaction follows. Each individual step is called an *elementary step* or an *elementary process*.

Using the reaction  $2\text{NO}_{(g)} + \text{O}_{2(g)} \rightarrow 2\text{NO}_{2(g)}$ , experimental data shows that the  $\text{NO}_2$  is not formed directly from the collision of  $\text{NO}$  and  $\text{O}_2$  particles, as  $\text{N}_2\text{O}_2$  can be detected during the reaction.

A more likely scenario for the reaction is a two-step reaction mechanism:



As the  $\text{N}_2\text{O}_2$  appears in the reaction mechanism but not in the overall chemical equation, it is called an *intermediate*.

**General Learning Outcome Connections**

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.



**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOME**

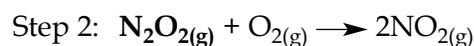
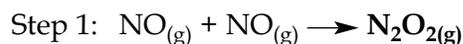
**C12-3-09:** Explain the concept of a reaction mechanism.

Include: rate-determining step

(continued)

■ **Bimolecular:** Production of nitrogen dioxide

Both elementary steps for the production of nitrogen dioxide involve two particles.



■ **Termolecular:**

Very few reactions require three particles to react simultaneously in an elementary step.

**Extension**

Have students draw potential energy diagrams for multi-step reaction mechanisms.

**SUGGESTIONS FOR ASSESSMENT**

**Paper-and-Pencil Tasks**

Ask students to create their own analogy of a reaction mechanism.

**Journal Writing**

Students can describe how they would feel and act if they were an intermediate substance in a reaction mechanism.

**LEARNING RESOURCES LINKS**



*Chemistry* (Chang 560)

*Chemistry* (Zumdahl and Zumdahl 583)

*Chemistry: The Molecular Nature of Matter and Change* (Silberberg 700)

*Glencoe Chemistry: Matter and Change* (Dingrando, et al. 548)

*McGraw-Hill Ryerson Chemistry, Combined Atlantic Edition* (Mustoe, et al. 477)

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al. 387)

*Prentice Hall Chemistry* (Wilbraham, et al. 578)

**Selecting Learning Resources**

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**SKILLS AND ATTITUDES OUTCOMES**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**C12-0-U2:** Demonstrate an understanding of chemical concepts.

*Examples: use accurate scientific vocabulary, explain concepts to others, compare and contrast concepts, apply knowledge to new situations and/or contexts, create analogies, use manipulatives . . .*

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**NOTES**

**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOME**

**C12-3-10:** Determine the rate law and order of a chemical reaction from experimental data.

Include: zero-, first-, and second-order reactions and reaction rate versus concentration graphs

(2 hours)

SLO: C12-3-10

**SUGGESTIONS FOR INSTRUCTION**

**TEACHER NOTES**

**Reaction Rate Laws and Reaction Order**

The *differentiated rate law* is determined by using the initial rate method. The *integrated rate law* is determined by using the concentration change over time to determine rate. Avoid using the integrated rate law, as it involves the use of calculus. Instead, emphasize the use of the initial rate method. A key point to remember is that the components of the rate law must be found by experiment and not through the use of reaction stoichiometry.

Most chemistry textbooks deal with this topic in detail. Determine the depth of instruction based on students' learning requirements.

*Introductory Example:*

For the reaction  $A \rightarrow B$ , the following data was obtained.

Trial	Initial [A] (mol/L)	Initial Rate (mol/L · s)
1	0.10	5
2	0.20	10
3	0.30	15

When asked to interpret the above data, students may indicate that as the concentration went up, the initial rate also went up. (It is a proportional relationship.)

The relationship can be written as

$$\text{Rate} \propto [A]^x$$

where  $x$  is called the *order* of the reaction.

**General Learning Outcome Connections**

**GLO D3:** Understand the properties and structures of matter, as well as various common manifestations and applications of the actions and interactions of matter.

## SKILLS AND ATTITUDES OUTCOME

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

The *order* describes how rate is affected by changing concentration(s) of reactant(s). For example, when doubling the concentration of a reactant results in a doubling of the rate, the reaction is a first-order reaction with respect to that reactant ( $x = 1$ ). When doubling the concentration of a reactant results in the rate increasing by four times ( $2^2$ ), the reaction is a second-order reaction with respect to that reactant ( $x = 2$ ).

To evaluate this mathematically, replace the proportionality symbol with an equal sign. To do this, a proportionality constant must be included. In this case, it is called the *rate constant* ( $k$ ).

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

In this data,  $x$  is equal to 1.

*Sample Problem:*

For the reaction  $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{CO}_2(\text{g})$ , the following data was obtained. Determine the overall rate law for this reaction.

Trial	Initial Rate (mol/L · s)	Initial [NO <sub>2</sub> ] (mol/L)	Initial [CO] (mol/L)
1	0.0050	0.10	0.10
2	0.080	0.40	0.10
3	0.0050	0.10	0.20

*Solution:*

1. Take the ratio of the initial rates for Trials 1 and 2, in which only one reactant is changed.

$$\frac{\text{Trial 2 } [\text{NO}_2]}{\text{Trial 1 } [\text{NO}_2]} = \frac{0.40}{0.10} = 4 \text{ times (quadrupled the concentration)}$$

$$\frac{\text{Trial 2 rate}}{\text{Trial 1 rate}} = \frac{0.080}{0.0050} = 16 \text{ times (rate increases 16 times)}$$

By increasing the concentration four times, the effect on the reaction time is that it is increased by 16. This means that the reaction rate depends on the square of the concentration of NO<sub>2</sub>. The reaction is a second-order reaction with respect to NO<sub>2</sub>.

The rate law would be

**Topic 3:  
Chemical  
Kinetics**

**SPECIFIC LEARNING OUTCOME**

**C12-3-10:** Determine the rate law and order of a chemical reaction from experimental data.

Include: zero-, first-, and second-order reactions and reaction rate versus concentration graphs

(continued)

$$\text{Rate} = k[\text{NO}_2]^2$$

2. Take the ratio of the initial rates for Trials 1 and 3, in which the concentration of CO is changed.

$$\frac{\text{Trial 3 [CO]}}{\text{Trial 1 [CO]}} = \frac{0.20}{0.10} = 2 \text{ times (doubled the concentration)}$$

$$\frac{\text{Trial 3 rate}}{\text{Trial 1 rate}} = \frac{0.0050}{0.0050} = 1 \text{ time (rate does not increase)}$$

By increasing the concentration of CO, the experimental data shows that the reaction rate does not change. It does not matter how much CO there is, as the rate of reaction does not depend on [CO]. Therefore, the reaction is a zero-order reaction with respect to CO.

The rate law would be

$$\text{Rate} = k[\text{NO}_2]^2[\text{CO}]^0 = k[\text{NO}_2]^2(1) = k[\text{NO}_2]^2$$

Emphasize that the value of  $k$  is specific for each reaction and changes only for a given reaction if the temperature changes.

**Laboratory Activities**

If sufficient time is available, students could perform the following lab activities:

- **Lab 14: Determining Reaction Orders** (Dingrando, et al., *Glencoe Chemistry: Small-Scale Laboratory Manual, Teacher Edition* 53)

In this lab activity, students determine the general equation for the reaction between crystal violet and sodium hydroxide.

- **Experiment 30: Rate Law Determination of the Crystal Violet Reaction** (Holmquist, Randall, and Volz, *Chemistry with Vernier*)

In this experiment, students observe the reaction between crystal violet and sodium hydroxide. They study the relationship between concentration of crystal violet and the time elapsed during the reaction.

- **Reaction Order** (PASCO, *Chemistry*)

In this experiment, students analyze the reaction rate by determining the order of the reaction when a colouring agent reacts with household bleach.

**SKILLS AND ATTITUDES OUTCOME**

**C12-0-U1:** Use appropriate strategies and skills to develop an understanding of chemical concepts.

*Examples: analogies, concept frames, concept maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles . . .*

**SUGGESTIONS FOR ASSESSMENT****Paper-and-Pencil Task**

Have students solve problems involving rate laws.

**Journal Writing**

Ask students to create a table in which they “describe the effect of doubling, tripling, and quadrupling [A] on the overall rate of chemical reactions having the following rate laws:

$$\text{Rate} = k[A]^0; \text{Rate} = k[A]^1; \text{Rate} = k[A]^2; \text{Rate} = k[A]^3$$

(Dingrando, et al., *Glencoe Chemistry: Matter and Change, Teacher Wraparound Edition* 544).

**LEARNING RESOURCES LINKS**

*Chemistry* (Chang 539)

*Chemistry* (Zumdahl and Zumdahl 564)

*Chemistry: The Molecular Nature of Matter and Change* (Silberberg 679)

*Glencoe Chemistry: Matter and Change* (Dingrando, et al. 542)

*Glencoe Chemistry: Matter and Change, Teacher Wraparound Edition* (Dingrando, et al. 544)

*Nelson Chemistry 12, Ontario Edition* (van Kessel, et al. 372, 374)

*Prentice Hall Chemistry* (Wilbraham, et al. 575)

**Investigations**

*Chemistry with Calculators* (Holmquist and Volz)

*Glencoe Chemistry: Small-Scale Laboratory Manual, Teacher Edition* (Dingrando, et al. 53)

**Websites**

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**Selecting Learning Resources**

For additional information on selecting learning resources for Grade 11 and Grade 12 Chemistry, see the Manitoba Education website at <[www.edu.gov.mb.ca/k12/learnres/bibliographies.html](http://www.edu.gov.mb.ca/k12/learnres/bibliographies.html)>.

