Manitoba
Education, Citizenship and Youth

Chemistry – Grade 12
Unit 3 - Kinetics

DRAFT / Unedited Version

Fall 2007

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Manitoba

Specific Learning Outcomes

C12-3-01 Formulate an operational definition of reaction rate.
Include: examples of chemical reactions that occur at different rates (0.5 hours)
GLO: D3

C12-3-02: Identify variables used to monitor reaction rates (i.e. change per unit time, $\Delta x/\Delta t$).
***Examples: pressure, temperature, pH, conductivity, colour* (0.5 hour)**
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, molecular representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

C12-0-S5 Collect, record, organize and display data using an appropriate format.
Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware…
GLO: C2, C5

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 chemical changes it still reinforces 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, neutralization reactions.
Some examples students may give for slow reactions are rusting, baking a cake, ripening of fruit, growth of a plant.

Teacher Notes (Specific Learning Outcomes 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 work on studying the chemical reactions involved in spoilage and decay of foods. (see *Nelson Chemistry 12*, *Ontario Edition* ,van Kessel 358)

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

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 rate. (Fast reactions have a short reaction time, slow reactions take a long time.)

Listed below are a number of demonstrations / lab activities illustrating the concept of reaction rate in a chemical reaction. It is not suggested that teachers perform all of these activities but a few demonstrations should be shown to help student understanding of reaction rates.

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Teacher Demonstration

React magnesium metal with 1.0 mol / L HCl. React another piece of Mg metal with 6.0 mol / L HCl.

Ask questions similar to the following in order to promote close observations among the students:

What happened?

How long did both reactions take?

Does it matter how much material you have?

How can you measure the rate of the reaction?

Teacher Demonstration (extension)

Generate hydrogen and oxygen by electrolysis into a dish of liquid soap. It will give off bubbles of hydrogen and oxygen gas. Remove the gas generator.

Ask students if 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 meter stick)
The reaction happens quickly! (A loud popping sound results)

A full explanation of this demonstration is given at the following website.

The Royal Society of Chemistry: Classic Chemical Demonstrations
The electrolysis of water-exploding bubbles of oxygen and hydrogen
http://www.chemsoc.org/Networks/Learnnet/videodemos/electrolysiswater.pdf

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Teacher Demonstration or Student Activity

Cut an apple into four slices, making sure each slice has the same approximate surface area of flesh exposed. For the first slice dip it in water and place on desk. This first slice acts as the control. For the second slice dip this slice in lemon juice and place on the desk next to the first slice. Place a third slice in the refrigerator, or 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 min., 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, p. 359)

Teacher Demonstration or Student Activity

Hydrogen peroxide (H_2O_2) gradually decomposes to form water and oxygen gas. In this demonstration 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 your observations.

(The hydrogen peroxide is clear and colourless. When the yeast is added to the hydrogen peroxide bubbles form and then later the mixture starts to foam)
Note: If you touch the bubbles with a burned splint, the splint will reignite.

Instead of using yeast, manganese dioxide (MnO_2) can be used to speed up the hydrogen peroxide decomposition reaction.

Teacher Notes (C12-3-02)

Reaction rate is described as a 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 which measure reaction rates involve the use of spectrometers, conductivity apparatus, and manometers (or a simple syringe).

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It is important to note that concentration cannot be monitored directly. It should be stressed that the observable (measurable) properties listed below can be used to determine the change in concentration over time.

Pressure:
Example: 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 can be monitored by attaching a manometer to a reaction vessel of known volume that is immersed in a constant-temperature bath.

$Zn_{(s)} + 2\ CH_3COOH_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2\ CH_3COO^-_{(aq)} + H_{2(g)}$

As $H_{2(g)}$ is produced, the gas pressure increases. (Silberberg 681)

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

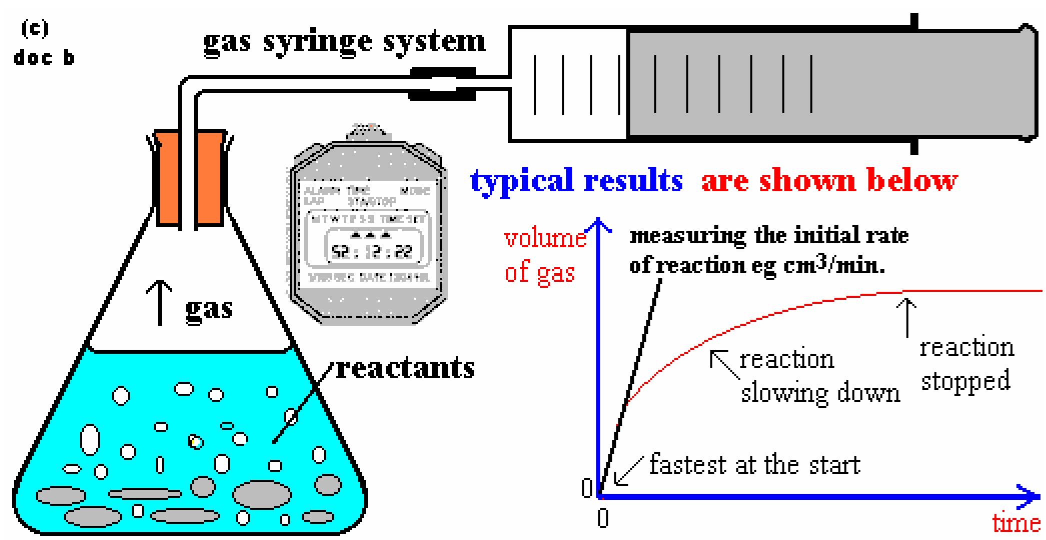

Image Credit: http://www.wpbschoolhouse.btinternet.co.uk/page03/3_31rates.htm

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The Concept of pH

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

Colour

Example: 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 would be

$NO_{(g)} + O_{3(g)} \rightarrow O_{2(g)} + NO_{2(g)}$
colourless reddish brown
Known amounts of the reactants are injected into a gas sample tube of known volume, and the rate of $NO_{2(g)}$
produced is measured by monitoring the colour over time. (Silberberg 680)

Another example is the reaction between liquid bromine and formic acid. As the reaction proceeds the concentration of the reddish-brown $Br_{2\,(aq)}$ decreases and this loss of colour (concentration) can be measured with the spectrometer.

$Br_{2(aq)} + HCOOH_{(aq)} \rightarrow 2\,Br^{-}_{(aq)} + 2\,H^{+}_{(aq)} + CO_{2(g)}$ (Chang 533)

Temperature

$N_2O_4 \rightarrow 2NO_2$
colourless reddish brown

The above reaction can be monitored by temperature. If a sealed tube of NO_2-N_2O_4 is placed in a cold water bath the dinitrogen tetroxide becomes predominate. The contents of the tube become lighter in colour.
If another tube containing a similar sample of NO_2-N_2O_4 is placed in a hot water bath then the resulting colour change is a reddish brown, indicating a greater presence of NO_2.
A sealed tube of NO_2-N_2O_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.

Conductivity

Electrodes are 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 nonionic reactants form ionic products. (Silberberg 681)

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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.

rate = $\Delta x/\Delta t$ (formation of a product)

rate = - $\Delta x/\Delta t$ (consumption of a reactant)

Students may confuse reaction rate and reaction time. Stress that reaction rate describes a change over time while reaction time is only the amount of time for a reaction to occur. The two terms are inversely related as shown by the formulas given above.

Suggestions for Assessment

Journal Writing
Students can make journal entries for fast and slow reactions and state their rationale for each. Some questions to consider:
- Do all reactions occur at the same rate?
- What do you mean by rate?
- How can you measure the rate of a reaction?
- Does a reaction always occur at the same rate?

Ask students to provide examples of:
- Reactions that have different rates
- Reactions that occur at different rates under different conditions.
- Processes that cannot be controlled.
- Processes that can be controlled.

Paper and Pencil Tasks
Students can complete a Compare – Contrast think sheet for fast versus slow reactions. SYSTH 10.15

Students can complete a KWL (Know-Want to Know-Learned) strategy sheet on reaction rate. SYSTH 9.8

Given a reaction, students can predict what variable (or property) may be most easily monitored.

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Learning Resources Links
Chemistry (Chang 532)

Chemistry (Zumdahl and Zumdahl 561)

Chemistry: The Molecular Nature of Matter and Change (Silberberg 673, 680)

Glencoe Chemistry: Matter and Change (Dingrando, *et al*. 529)

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

McGraw-Hill Ryerson Chemistry, Combined Atlantic Edition (Mustoe *et al*. 463, 466)

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

Websites
The Royal Society of Chemistry: Classic Chemical Demonstrations
The electrolysis of water-exploding bubbles of oxygen and hydrogen
http://www.chemsoc.org/Networks/Learnnet/videodemos/electrolysiswater.pdf

Doc Brown's Chemistry Clinic GCSE notes on Rates of Chemical Reactions
http://www.wpbschoolhouse.btinternet.co.uk/page03/3_31rates.htm

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Include: initial rate (2 hr)
GLO: C2

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 (0.5 hours)
GLO: D3, D4

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C12-0-S6 Estimate and measure accurately using Système International (SI) and other standard units.
Include: SI conversions, significant figures.
GLO : C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

Entry-Level Knowledge

Students have studied the stoichiometry of chemical reactions in Topic 3, Grade 11 Chemistry.

Laboratory Activity
Have students perform a lab to measure the change in mass of calcium carbonate as it reacts with 3 mol / L hydrochloric acid. See appendix 1 for the lab called "Graphical Determination of the Rate of a Chemical Reaction".

From the derived data or given data in the appendix, have students calculate average rate and instantaneous rate for a reaction. Students can use Graphical Analysis to plot data and determine instantaneous rate at time = 0 (initial rate) and at other times. The students can compare the rates and hypothesize why the rate changes.

Teacher Notes (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 that the reaction took place. See graph below and sample calculation.

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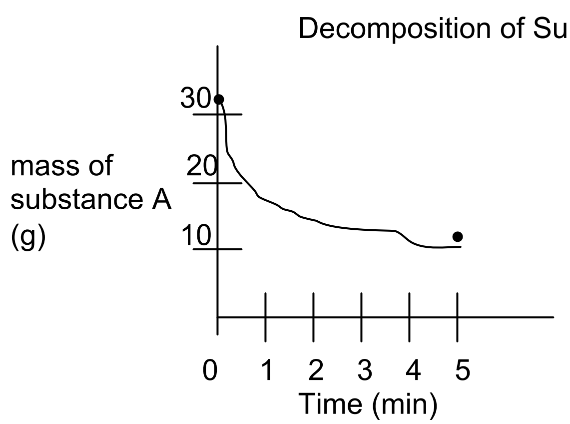

$$\text{average rate} = \frac{\text{change in the amount of substance A}}{\text{change in time}} = \frac{30\text{ g} - 10\text{ g}}{5\text{ min} - 0\text{ min}} = 4\text{ g/min}$$

Instantaneous rate is the rate of the reaction at a particular instant of time. To calculate this rate a tangent line is drawn to the point on the graph (particular instant of time) and the slope of this line is then calculated. See graph and sample calculation for determining the instantaneous rate at 1 minute.

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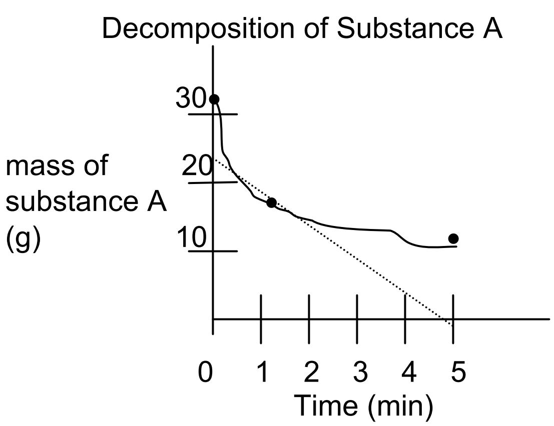

$$\text{slope} = \frac{\text{change in amount of substance A}}{\text{change in time}} = \frac{25\text{ g} - 0\text{ g}}{5\text{ min} - 0\text{ min}} = 5\text{ g/min at t} = 1\text{ min}$$

Animation / Simulation

This simulation allows the student to determine the rate of reaction at a given point in time.
Chemistry @ Davidson Chemistry Experiments and Demonstrations by David N. Blauch
http://www.chm.davidson.edu/ChemistryApplets/kinetics/RateOfReaction.html
Other simulations at this website show the effect of concentration change, the rate of a chemical reaction and the determination of stoichiometric coefficients.

Chemistry: The Science in Context
Chapter 14 Chemical Kinetics and Air Pollution, Section 14.2 Reaction Rate, view animation tutorial. Sections 4, 5, and 6 in this tutorial graphically show how average rate and instantaneous rate are calculated.
http://www.wwnorton.com/college/chemistry/gilbert/tutorials/ch14.htm

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Paper Lab Activity

If students need additional practice then they can do sample plots with given data. In the appendix are two assignments, Kinetics Assignments 1 and 2, suited for this purpose. From the plotted data students can calculate average rates and determine instantaneous rates. They can 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.

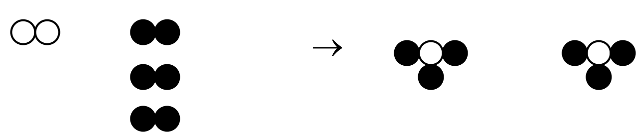
Teacher Notes (C12-3-04)

<u>Rate and Reaction Stoichiometry</u>

This concept should be introduced carefully. Diagrams of molecules would be helpful for students in understanding reaction rate at the particulate (molecular) level.

For the reaction, $N_2 + 3\ H_2 \rightarrow 2\ NH_3$, the coefficient in front of the substance determines the rate of consumption or production of that substance, if the initial rate of N_2 is known.

At the molecular level, this reaction would be expressed as



$N_2 + 3\ H_2 \rightarrow 2\ NH_3$

The student should recognize that for every N_2 molecule, three H_2 molecules need to be consumed. This means that the rate of consumption of N_2 is three times the rate of consumption of H_2. In addition, for every molecule of N_2 that is consumed, the rate of production of NH_3 molecules is doubled.

Therefore, the rate of consumption of N_2 = 3 x (the rate of consumption of H_2) = 2 x (the rate of production of NH_3) according to the balanced chemical equation.

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Another way of stating 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

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

Then the following is also true

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

<u>Sample Problem</u>:

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

<u>Solution:</u>
We can calculate the rate in a manner similar to how we used stoichiometry to determine moles of product formed. We use the ratio of the coefficients to determine the ratio of rates.

$$\text{rate NH}_3\text{ formation} = 1.5\text{ mol / L}\cdot\text{s H}_2\left(\frac{2\text{ NH}_3}{3\text{ H}_2}\right)$$

$$= 1.0\text{ mol / L}\cdot\text{s NH}_3$$

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http://www.wwnorton.com/college/chemistry/gilbert/tutorials/ch14.htm

Suggestions for Assessment

<u>Laboratory Skills</u>
A checklist can be used to evaluate students on the following lab skills:

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

<u>Paper and Pencil Tasks</u>

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.
Have students solve problems on experimental rate data and reaction stoichiometry. See Appendix 4: Chemical Kinetics Problems and Appendix 5: Chemical Kinetics Problems – Answer Key.

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Learning Resources Links

Appendix 2: Kinetics Assignment 1

Appendix 3: Kinetics Assignment 2

Appendix 4: Chemical Kinetics Problems (BLM)

Appendix 5: Chemical Kinetics Problems – Answer Key

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)

Investigations

Appendix 1: Graphical Determination of the Rate of a Chemical Reaction Lab

Nelson Chemistry 12, Ontario Edition (van Kessel, *et al*.)
Lab Exercise 6.1.1 Determining a Rate of Reaction 401

Specific Learning Outcomes

C12-3-03: Perform a lab to measure the average and instantaneous rate of a chemical reaction.
Include: initial rate (2 hr)
GLO: C2

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 (0.5 hours)
GLO: D3, D4

C12-0-S5 Collect, record, organize and display data using an appropriate format.
Examples: labelled diagrams, graphs, multimedia applications, software integration, probeware…
GLO: C2, C5

C12-0-S6 Estimate and measure accurately using Système International (SI) and other standard units.
Include: SI conversions, significant figures.
GLO : C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

Websites

Chemical Kinetics
This simulation allows the student to determine the rate of reaction at a given point in time.
http://www.chm.davidson.edu/ChemistryApplets/kinetics/RateOfReaction.html

Shows the formation of oxygen & the decomposition of hydrogen peroxide graphically
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/kinetics2/rxnRate01.html

Chemistry: The Science in Context
Chapter 14 Chemical Kinetics and Air Pollution, Section 14.2 Reaction Rate, view animation tutorial.
http://www.wwnorton.com/college/chemistry/gilbert/tutorials/ch14.htm

Specific Learning Outcomes

C12-3-05 Perform a lab 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. (1.5 hours)
GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

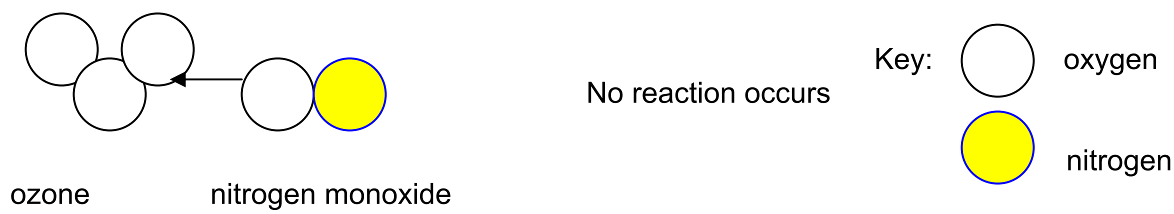
C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

C12-0-S9 Draw a conclusion based on the analysis and interpretation of data.
Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
GLO: C2, C5, C8

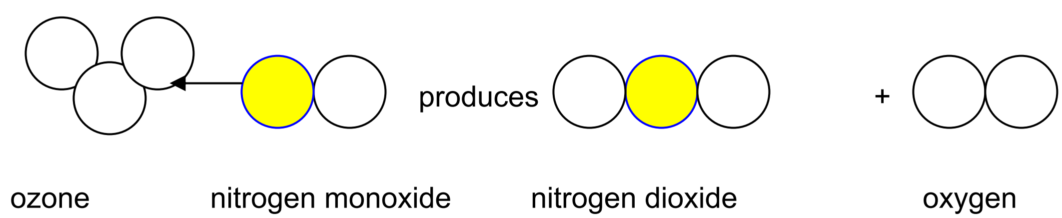
Teacher Notes

At this point students should be introduced to a "collision theory" of chemical reactions. The collision theory states that in order for a chemical reaction to occur, the reacting particles must collide. If particles do not collide a reaction does not occur. All collisions, however, do not necessarily produce a chemical reaction. Reacting particles must collide with **sufficient kinetic energy** (called activation energy) and the **correct collision geometry or orientation**.

Orientation of nitrogen monoxide molecule UNLIKELY to produce a reaction



Orientation of nitrogen monoxide molecule LIKELY to produce a reaction



The following 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_3, the nitrogen atom of the nitrogen monoxide molecule must collide with the correct positioning and sufficient energy to cause the chemical reaction to occur.
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animations/NO+O3singlerxn.html

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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
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GLO: C2, C5, C8

The animation at this website, http://chem.salve.edu/chemistry/temp2a.asp, allows the selection of variables, insufficient activation energy, sufficient activation energy, and sufficient activation energy but incorrect orientation and shows the result between the collision of hydrogen, H_2, and iodine, I_2.

Activation energy is defined as the minimum amount of kinetic energy required for particles to collide effectively, that is, to produce a chemical reaction.

It is not expected that teachers have students perform all of the activities indicated. Teachers may choose to do more than one activity if time permits. Choose activities appropriate for the class.

From these labs, students should conclude that increasing temperature will increase the rate of a reaction (decreases reaction time), increasing the concentration of reactant(s) will increase the rate of a reaction, increasing the surface area will increase the rate of reaction, presence of a catalyst will increase the rate of a reaction and the nature (type) of reactants will affect the rate of a reaction. Note that pressure and volume are a subset of concentration.

Factors that affect the rate of a chemical reaction:

<u>Collision Theory and Nature of Reactants</u>
Some chemical reactions involve the rearrangement of atoms as a result of bonds breaking to form new bond formation. 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 bond broken, the faster the reaction rate will be. The weaker the bonds in the reactants, the faster the reaction will be. The state of the reactants, solid, liquid, or gas, will also affect the rate of a reaction.

<u>Collision Theory and Surface Area</u>
From the lab activities done in C12-3-05 students observed that increasing the surface area of a solid increases the reaction rate. Collisions can only occur at a solid's surface, so a powdered substance, like $CaCO_3$, will react more quickly than a large crystal of $CaCO_3$ as this allows more surface area to be in contact with the other reactants.

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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

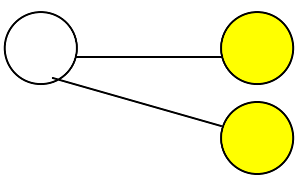
C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

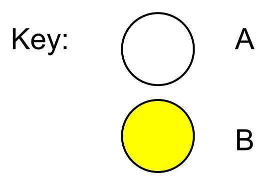
C12-0-S9 Draw a conclusion based on the analysis and interpretation of data.
Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
GLO: C2, C5, C8

<u>Collision Theory and Concentration</u>

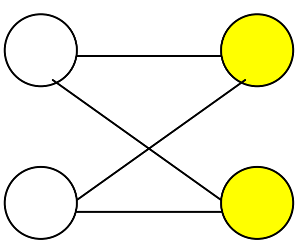
The collision theory states that particles must collide with each other to react. If the concentration of one reactant is increased, then the reaction rate should increase as there are more molecules of the increased reactant that can collide. At the particulate level, if there is one molecule of A reacting with two molecules of B there are two collisions possible.



2 collisions

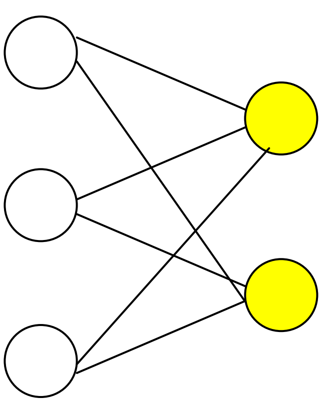


If the concentration of A is doubled then there are four collisions possible.



4 collisions

And, if the concentration of B is tripled then there are 6 collisions possible.



6 collisions

Specific Learning Outcomes

C12-3-05 Perform a lab to identify factors that affect the rate of a chemical reaction.
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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

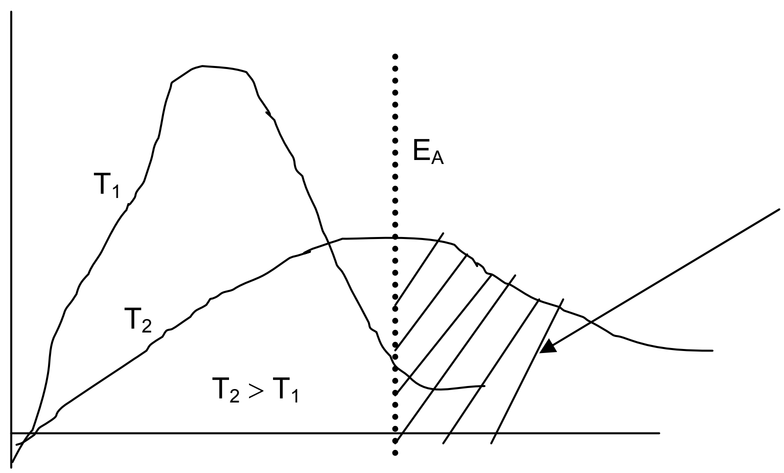
C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

C12-0-S9 Draw a conclusion based on the analysis and interpretation of data.
Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
GLO: C2, C5, C8

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 for T_1(lower temperature).

number of molecules



greater number of molecules at T_2 with enough energy to react

[van Kessel, 383]

Laboratory Activities

Students should be led to qualitatively discover the factors that affect the rate of a reaction rather than perform a verification laboratory. Possible laboratory activities include:

- Factors Affecting Reaction Rate. (concentration, nature of reactants, temperature, catalyst, surface area) See Appendix 6: Lab – Factors Affecting Reaction Rates
  In Part A of this lab, 5 test tubes are set up containing the various reactants. Concentration of reactants are varied for the first 3 test tubes. These reactions take place at room temperature. The fourth test tube is identical to test tube 1 in terms of concentration. This test tube is heated to 50°C in a water bath before the potassium permanganate is added. The factor being studied here is temperature (reaction time of test tube 1 vs reaction time of test tube 2). The fifth test tube has

Specific Learning Outcomes

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Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst. (1.5 hours)
GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

C12-0-S9 Draw a conclusion based on the analysis and interpretation of data.
Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
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the catalyst manganese sulfated added to it (reaction time of test tube 1 compared with the reaction time of test tube 5). Part B tests the addition of potassium permanganate by increasing the concentration in each trial by three drops. The reaction times for the five trials are recorded and the effect of increased concentration has be observed. Part C of the lab compares the reaction time of calcium carbonate (crystals) and hydrochloric acid with powdered calcium carbonate and hydrochloric acid. This part of the lab covers the effect of surface area on reaction time.

- Factors Affecting the Rate of a Reaction. (concentration, nature of reactants, temperature, catalyst, surface area) See Appendix 7: Factors Affecting the Rate of a Reaction
  Part 1 of this lab studies 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. The second part of this section studies the effect of different solutions reacting with magnesium metal on reaction time.

  Part 2 studies the effect of surface are 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.

  The effect of temperature on a chemical reaction is studied in part 3. 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. The second half of this section has the student setting 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 in 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.

  For part 4, a catalyst is used to study the effect of it on reaction time. Potassium permanganate is placed in two test tubes. In one of the test tubes, manganese (II) sulfate is added (catalyst). Then oxalic acid is added to both test tubes and the reaction times are noted.

Specific Learning Outcomes

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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

C12-0-S2 State a testable hypothesis or prediction based on background data or on observed events.
GLO: C2

C12-0-S7 Interpret patterns and trends in data, and infer and explain relationships.
GLO: C2, C5

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Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
GLO: C2, C5, C8

- Factors Affecting The Rate Of A Chemical Reaction (covers temperature, concentration, and surface area) [Small-Scale Chemistry Laboratory Manual, Waterman & Thompson, Addison-Wesley 1993, 185]
  There are four parts of this laboratory activity. The first part explores the rate of reaction between hydrochloric acid and magnesium, calcium carbonate, and sodium hydrogen carbonate. The effect of temperature is studied using the same reactants that were used in part one. Cold hydrochloric acid and warm hydrochloric acid are separately reacted with magnesium, calcium carbonate, and sodium hydrogen carbonate. The third part of the lab investigates 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. The last part looks at the effect of concentration on reaction rate. Various concentrations of hydrochloric acid are separately reacted with magnesium, calcium carbonate, and sodium hydrogen carbonate.

- Factors Affecting Reaction Rates (covers temperature, reactant concentration, particle size, catalysis, and surface area) [Laboratory Manual, Wilbraham et al, Pearson Education / Prentice-Hall, 225]

- Chemlab 17 Concentration and Reaction Rate [Chemistry, Dingrando et al., McGraw-Hill, 550]
  This lab activity investigates the effect of concentration on reaction rate. Pieces of magnesium ribbon are reacted separately with varying concentrations of hydrochloric acid, with the resulting reaction time recorded.

- MiniLab: Examining Reaction Rate and Temperature [Chemistry, Dingrando et al., McGraw-Hill, 539]
  Antacid tablets are dissolved in water at room temperature, at 50°C and 65°C. The effect of temperature on reaction rate is observed.

- Investigation 12-A Factors Affecting the Rate of a Reaction [Mustoe 464]
  Part 1 of this activity investigates the effect of concentration of a chemical reaction. Sodium hydrogen carbonate and varying concentrations of vinegar are reacted. The time to collect a test tube full of carbon dioxide is recorded. The average reaction rate in mL/s is calculated for each trial. Part 2 demonstrates the effect of temperature on reaction rate. The same reactants are used as in part 1, sodium hydrogen carbonate and vinegar. One gram of $NaHCO_3$ is reacted with 10.0 ml of

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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
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vinegar, but both reactants need to be at 10°C before combining the reactants. The trial is repeated but both reactants need to be about 10°C higher than room temperature. The last part (part 3) of the activity studies the effect of reactants and surface area on reaction rate. Powdered calcium carbonate is used instead of sodium hydrogen carbonate and $CaCO_3$ is reacted with 10.0 mL of vinegar. The time to collect a test tube full of carbon dioxide is recorded. The average reaction rate in mL/s is calculated for the trial. A second trial is done but this time 1 g of $CaCO_3$ is used instead of the powdered calcium carbonate. Students should comment on the effects of each factor on reaction rates.

If students have not seen factors in their investigations, provide them with demonstrations to illustrate these factors. In the post lab discussion, have students explain their observations based on the collision theory.

<u>Teacher Demonstration Lab</u>
Teachers can choose to demonstrate the following lab, A Study of Reaction Rates: The “Clock Reaction”. [Chemistry: Experimental Foundations Laboratory Manual, Merrill et al., Prentice-Hall, 62] (covers concentration and temperature and a more quantitative approach). A simpler demonstration is provided in the text, Chemistry: A Modern Course, Merrill, 1987, Smoot et al, 120T]

Teacher Demonstrations
Surface Area and Reaction Rate

The purpose of this demonstration is to see the effect of increasing surface area on reaction rate. Place 2 g of lycopodium powder (or starch) in a pile on a porcelain tile. Try and ignite the pile with a burner or lighter. There is no reaction. Lift the ceramic tile holding the lycopodium powder (or starch) and blow the powder in the direction of the lit burner. The powder ignites quite explosively. Students should observe that the reaction rate increases as surface area increases. [Chemistry: A Modern Course, Merrill, 1987, Smoot et al, 442]

Catalyst and Reaction Rate

The effect of a catalyst on the rate of a chemical reaction is to be observed in this demonstration. Dissolve 25 g of sodium potassium tartrate (Rochelle’s salt) in 300 mL of water in a large beaker. Add 100 mL of 3-6% H_2O_2 to the beaker. Heat the solution to 70°C. The students should notice no reaction occurs. Add the catalyst, cobalt chloride, to the beaker. The solution turns pink and then a greenish colour (cobalt II tartrate

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GLO: C2

C12-3-06 Use the Collision Theory to explain the factors which affect the rate of chemical reactions
Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

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GLO: C2

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complex). After the reaction has been completed, the pink colour in the solution reappears. The cobalt chloride was not consumed in the reaction. Students should observe that the solution at 70°C did not chemically react until the catalyst was added. [Chemistry: A Modern Course, Merrill, 1987, Smoot et al, 444] The following website shows a video of this demonstration.
NCCSM Chemistry Online Resource Essentials
www.dlt.ncssm.edu/core/c15.htm (homocatalyst.rm)

Multimedia on the Internet / Teacher Demonstration

The following website shows the decomposition of H_2O_2, catalyzed (with MnO_2) and uncatalyzed. This video clip shows how the addition of a catalyst affects the reaction rate. http://www.dlt.ncssm.edu/core/c15.htm (elephantooth.rm).

This video clip shows how temperature affects the rate of a reaction. A glowstick is placed in how water and another glowstick is placed in cold water.
http://www.dlt.ncssm.edu/core/c15.htm (glowsticks.rm)

This video clip shows how surface area affects reaction rates. Pieces of small potato are placed in a test tube containing H_2O_2. A small amount of detergent is placed in each test tube to make the bubbles of oxygen more visible.
http://www.dlt.ncssm.edu/core/c15.htm (decomppotatos.rm)

Suggestions for Assessment
<u>Paper and Pencil Tasks</u>
Have students compare and contrast the rate at which a sugar cube in cold water and granulated sugar in warm water would dissolve. They could include how surface area and the temperature of the water might affect the rate at which each dissolves. (p.238, Science Notebook, Chemistry: Matter and Change, Dingrando *et al*., Glencoe-McGraw-Hill, 2005)

Have students describe how the collision theory would apply to a demolition derby. (p.236, Science Notebook, Chemistry: Matter and Change, Dingrando *et al*., Glencoe-McGraw-Hill, 2005)

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Include: Activation energy, orientation of molecules (0.5 hours)
GLO: D3, D4

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Visual Displays
Students can represent a reaction between two substances, such as NO and O_3, using ball-and-stick molecular models. Students can show the correct orientation of the molecules to collide to produce NO_2 and O_2. Students can also show the incorrect orientation of the molecules which would not produce a reaction.

Laboratory Skills
Students could complete a lab report frame. (SYSTH)

Teachers should periodically and randomly review lab skills using a template such as the Chemistry Laboratory Skills Sample Rubric.

Research/Report
Students could research and report on how the rate of specific chemical processes can be controlled. Students could prepare a report or complete an Article Analysis frame on a related article.

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)

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

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

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

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

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Investigations
Appendix 5: Lab – Factors Affecting Reaction Rates (concentration, nature of reactants, temperature, catalyst, surface area)

Appendix 6: Factors Affecting the Rate of a Reaction

Small-Scale Chemistry Laboratory Manual, Waterman & Thompson, Addison-Wesley 1993
Factors Affecting The Rate of A Chemical Reaction (covers temperature, concentration, and surface area), 185

Laboratory Manual, Wilbraham et al, Pearson Education / Prentice-Hall
Factors Affecting Reaction Rates (covers temperature, reactant concentration, particle size, catalysis, and surface area), 225

Glencoe Chemistry: Matter and Change (Dingrando, et al.)
Chemlab 17 Concentration and Reaction Rate, 550

Chemistry: Experimental Foundations Laboratory Manual, Merrill et al., Prentice-Hall, 1982
A Study of Reaction Rates: The “Clock Reaction”, 62

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

Websites
NCCSM Chemistry Online Resource Essentials. This video shows the explosive nature of flour when placed in a closed container and then ignited with a candle. A very good example of the effect of surface area on reaction rate.
www.dlt.ncssm.edu/core/c15.htm (dust_explose.rm)

General Chemistry: Principles and Modern Applications 8th edition
This simulation shows the catalytic decomposition of ozone by chlorine atoms from CFCs.
http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html

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GLO: C2, C5

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Include: cause and effect relationships, alternative explanations, supporting or rejecting a hypothesis or prediction.
GLO: C2, C5, C8

NCCSM Chemistry Online Resource Essentials
This simulation shows the catalytic hydrogenation of an alkane.
http://www.dlt.ncssm.edu/core/Chapter15-Kinetics/Chapter15-Animations/Catalyst_2.html

Iowa State University-Thomas J. Greenbowe website
Iodine Clock Reaction-concentration of reactants can be varied and temperature can be adjusted in this simulation. Students must start the time clock and wait for the reaction to reach completion (blue-black colour).
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/kinetics2/iodine_clock.html

Education Using Powerpoint: Click on animations and then select view or download for Rates of Reaction. Shows particulate level for factors (heat, concentration and surface area) affecting reactions
http://www.educationusingpowerpoint.org.uk/index.html?ks3science.html~mainFrame

Iowa State University-Thomas J. Greenbowe website
Correct orientation of molecules upon collision
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animations/NO+O3singlerxn.html

Collision Theory
Shows animation for the reaction between hydrogen gas and iodine gas where insufficient energy, sufficient energy, and improper orientation with sufficient energy can be selected.
http://chem.salve.edu/chemistry/temp2a.asp

Specific Learning Outcome

C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
Include: relative rates, effects of catalyst, heat of reaction (enthalpy change) (1 hour)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, molecular representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

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…
GLO: D3

Entry-Level Knowledge

From Grade 10 Science, students are familiar with kinetic and potential energy with respect to motion.
In Grade 11 Chemistry, in specific student learning outcome C11-1-02, students were introduced to the Kinetic Molecular Theory to explain the properties of gases. Students should be familiar with KM distribution curves.

Teacher Notes

An exothermic reaction is a chemical reaction in which energy is released into the environment. Combustion, or burning, is an example of an exothermic reaction. On the other hand, an endothermic reaction is a chemical reaction takes absorbs energy from its surroundings which is stored in the products that have formed. If aluminum chloride is dissolved in water, the beaker will feel cool in your hands. Students are expected to draw potential energy diagrams indicating on their diagrams the amount of potential energy the reactants and the products have, the activation energy needed, the activated complex and the change in enthalpy or the heat of reaction, i.e., 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 the slower the reaction rate 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, the teacher can demonstrate the following:
a) Roll a ball up an incline and let the ball roll back down. The ball represents the reactants that don't have enough activation energy to reach the activated complex.
b) 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. (p.540 Teacher's Edition, Chemistry: Matter and Change, Dingrando *et al*., Glencoe-McGraw-Hill, 2005

Activity
Students can use the collision theory and kinetic energy and potential energy diagrams to explain their observations from the lab in C12-3-02. Student explanations should include what is happening at the **molecular** level.

Specific Learning Outcome

C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
Include: relative rates, effects of catalyst, heat of reaction (enthalpy change) (1 hour)
GLO: D3

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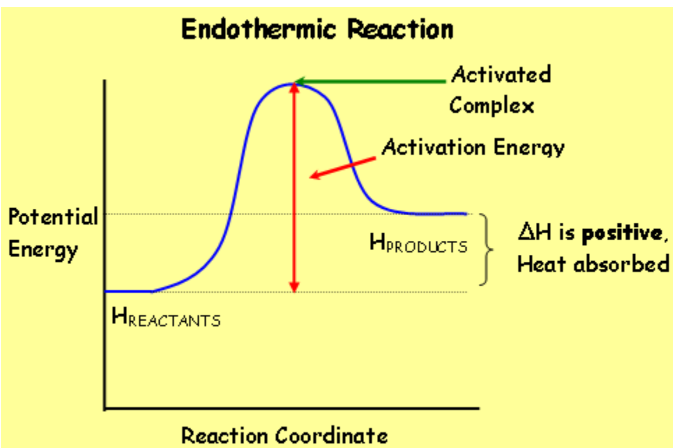
GLO: D3



(Image Credit: Manitoba Education; Chemistry 12 Online Course)

In the diagram showing the progress of an endothermic reaction, 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 and your hands had been placed on the outside of the beaker, then the beaker would have felt cool to your 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.

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

Specific Learning Outcome

C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
Include: relative rates, effects of catalyst, heat of reaction (enthalpy change) (1 hour)
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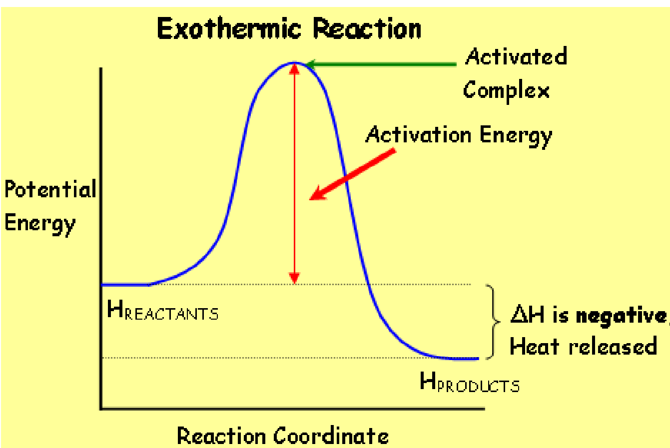

(Image Credit: Manitoba Education; Chemistry 12 Online Course)

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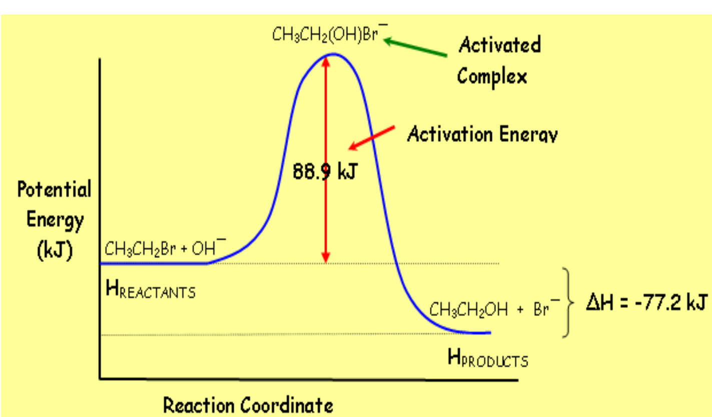
$\Delta H = H_{products} - H_{reactants}$ = negative value = heat is released



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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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For the diagram indicating the reaction $CH_3CH_2Br + OH^- \rightarrow CH_3CH_2OH + Br^-$, students should 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 potential energy diagram for the reaction $2\ BrNO \rightarrow 2\ NO + Br_2$ shows the transition state where the molecules of nitrogen, bromine and oxygen are rearranged to form the products.

Specific Learning Outcome

C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
Include: relative rates, effects of catalyst, heat of reaction (enthalpy change) (1 hour)
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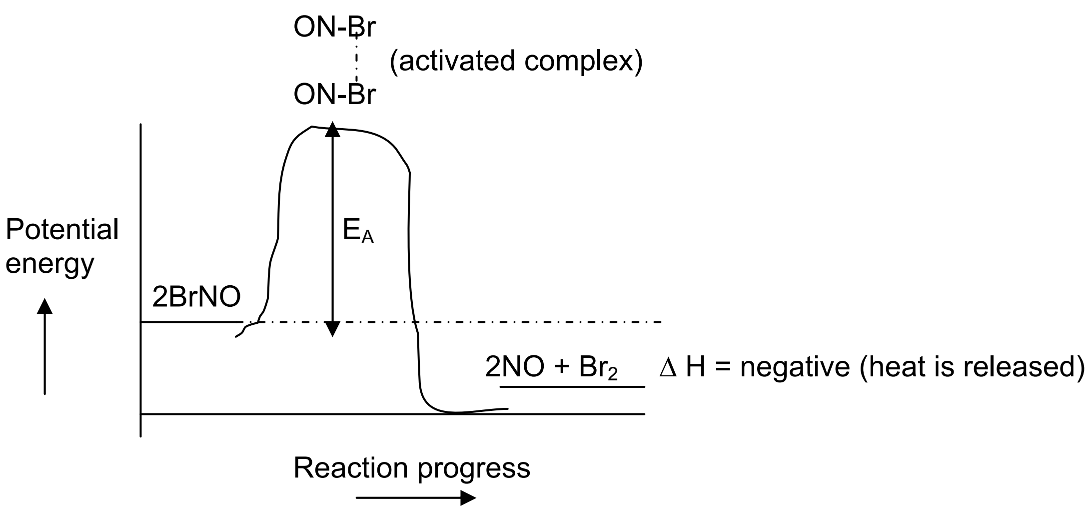

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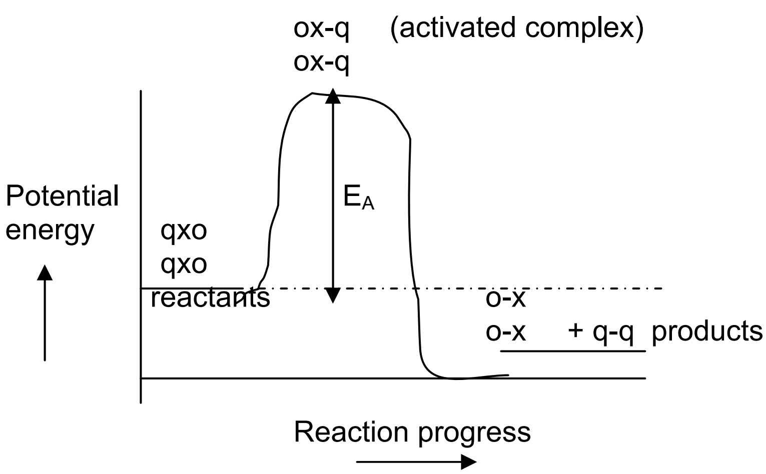
At the molecular level, this is how the potential energy diagram would appear for the chemical reaction just described.



Key: o= oxygen
x= nitrogen
q= bromine

[Zumdahl 588]

Animation
This animation shows how molecules with insufficient energy and how molecules with sufficient energy progess along a potential energy diagram.

See NCSSM Chemistry; Chemistry Online Resource Essentials
The following animation shows the change in potential energy as molecules collide.
http://www.dlt.ncssm.edu/core/c15.htm (Molecular_collision_Ea.swf)

Specific Learning Outcome

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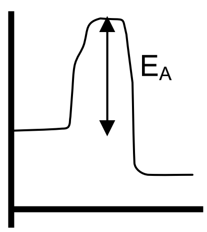
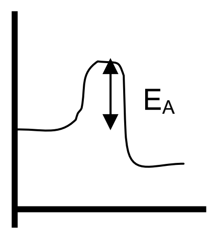
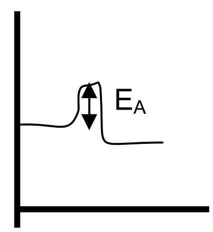
Relative rates

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





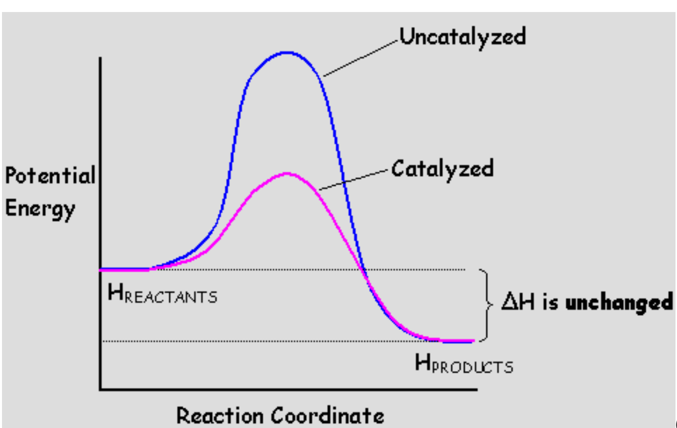


SLOW
Notice E_A is very large for this reaction. It would take a lot of energy to get this reaction to go to completion.

MEDIUM
Notice E_A is a bit lower than the slow reaction given to the right.

FAST
Notice that E_A is very small when compared to the slow and medium reactions.

Potential Energy Diagram of a Catalyst added in a Reaction



(Image Credit: Manitoba Education; Chemistry 12 Online Course)

The potential energy diagram shows an uncatalyzed reaction and a catalyzed reaction. Students should have concluded from their lab activities in 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

Specific Learning Outcome

C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
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GLO: D3

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the presence of a catalyst shows a smaller activation energy that is required. Students should note that the heat of reaction, or enthalpy change, does not change.

In Diagram **A** 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 E_A is lowered enabling more collisions to occur. This results in more product being formed.

Diagram **A**

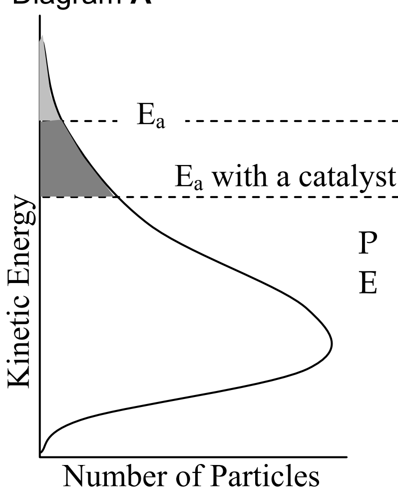
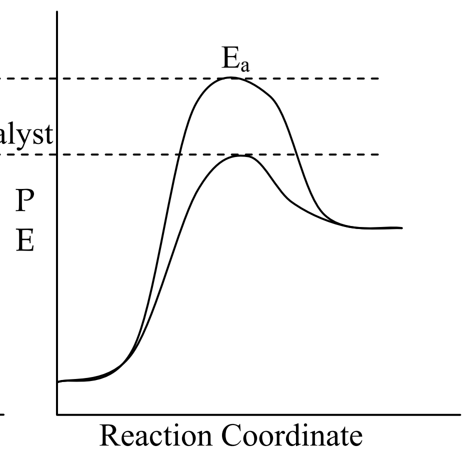


Diagram **B**



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Carolina Biological: Chemistry
Information is given on how to perform this demonstration called "Elephant's Toothpaste". This is a catalyzed reaction between H_2O_2 and KI (catalyst). The two step decomposition reaction is a good lead in for studying reaction mechanisms.
http://www.carolina.com/chemistry/experiments/elephant.asp

Animation
NCSSM Distance Learning Technologies
The following animation demonstrates the change in potential energy for two molecules as they collide. Low energy and high energy simulations are shown.
http://www.dlt.ncssm.edu/TIGER/chem5.htm#kinetics

Suggestions for Assessment
Paper and Pencil Tasks
Students should be able to interpret potential energy diagrams and draw potential energy diagrams from given information.

Journal Writing
Students can interpret graphs by answering the following questions:
Which are at a higher energy level, the reactants or the products?
Is energy absorbed or released after the reaction takes place?
Will the reaction always proceed to form products once the activated complex is formed?

Learning Resources Links
Chemistry (Chang 566)

Chemistry (Zumdahl and Zumdahl 588)

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

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

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

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

Specific Learning Outcome

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

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

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

Websites
NCSSM Chemistry; Chemistry Online Resource Essentials
The following website shows the effect of a catalyst on the decomposition of H_2O_2.
http://www.dlt.ncssm.edu/core/c15.htm

NCCSM Chemistry Online Resource Essentials. This video shows the explosive nature of flour when placed in a closed container and then ignited with a candle. A very good example of the effect of surface area on reaction rate.
www.dlt.ncssm.edu/core/c15.htm (dust_explose.rm)

General Chemistry: Principles and Modern Applications 8th edition
This simulation shows the catalytic decomposition of ozone by chlorine atoms from CFCs.
http://cwx.prenhall.com/petrucci/medialib/media_portfolio/15.html

NCCSM Chemistry Online Resource Essentials
This simulation shows the catalytic hydrogenation of an alkane.
http://www.dlt.ncssm.edu/core/Chapter15-Kinetics/Chapter15-Animations/Catalyst_2.html

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Specific Learning Outcome

C12-3-08: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, molecular representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

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…
GLO: D3

Teacher Notes

Teachers may wish to begin with an analogy (cleaning up dinner dishes). This process happens in many steps: clearing the table, filling the sink with water and soap, placing the dishes in, 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 like an average speed of 100 km/h does not tell us much of the various speeds we need to drive over 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 following reaction, $2\ NO_{(g)} + O_{2(g)} \rightarrow 2\ NO_{2(g)}$, experimental data shows that the NO_2 is not formed directly from the collision of NO and O_2 particles as N_2O_2 can be detected during the reaction.

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

Step 1 $2NO_{(g)} \rightarrow \mathbf{N_2O_{2(g)}}$
Step 2 $\mathbf{N_2O_{2(g)}} + O_{2(g)} \rightarrow 2NO_{2(g)}$

Net Reaction $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$

As the $\mathbf{N_2O_2}$ appears in the reaction mechanism but not in the overall chemical equation, it is called an **intermediate.**

Catalysts, as well as intermediates, do not appear in the overall reaction. The decomposition of ozone with a chlorine catalyst illustrates this:

Step 1 $\mathbf{Cl_{2(g)}} + O_{3(g)} \rightarrow \mathbf{ClO_{(g)}} + O_{2(g)}$
Step 2 $O_{3(g)} \rightarrow O_{2(g)} + O_{(g)}$
Step 3 $\mathbf{ClO_{(g)}} + O_{(g)} \rightarrow \mathbf{Cl_{2(g)}} + O_{2(g)}$

Net Reaction $2O_{3(g)} \rightarrow 3O_{2(g)}$

In the above example, the $\mathbf{Cl_{2(g)}}$ is a **catalyst** and the $\mathbf{ClO_{(g)}}$ is an **intermediate**.

The slowest of the elementary processes will determine the rate of the reaction. It is called the **rate determining step.**

Specific Learning Outcome

C12-3-08: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

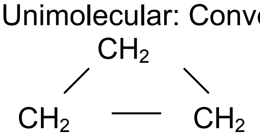
C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, molecular representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

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…
GLO: D3

Analogies can be used to illustrate the concept of rate determining step. For example, using the cleaning up the dishes analogy, the longest step (washing the dishes) would be the rate determining step. Students should recognize that efforts to speed up the other steps do not significantly affect the length of time to clean up the dishes, but speeding up the slowest step affects the time to clean up the most.

The **molecularity** of a reaction refers to the number of particles involved in an elementary step. The molecules may be of the same type or different types. The elementary step may involve one particle, unimolecular, two particles, bimolecular or three particles, termolecular. It is possible to use the elementary steps of a reaction to deduce a rate law.

Examples of Elementary Steps
Unimolecular: Conversion of cyclopropane to propene



$\rightarrow$ CH_3-CH=CH_2

There is only one particle involved in this one step reaction mechanism, which is the cyclopropane.

Bimolecular: Production of nitrogen dioxide
Both elementary steps for the production of nitrogen dioxide involve two particles.
Step 1 $NO_{(g)}$ + $NO_{(g)}$ → $\mathbf{N_2O_{2(g)}}$
Step 2 $\mathbf{N_2O_{2(g)}}$ + $O_{2(g)}$ → $2NO_{2(g)}$

Termolecular:
There are very few reactions that require three particles to react simultaneously in an elementary step.

Extension
Have students draw potential energy diagrams for multistep reaction mechanisms.

Suggestions for Assessment
<u>Paper and Pencil Tasks</u>
Students can create their own analogy of a reaction mechanism.

<u>Journal Writing</u>
Students can write in their journals how they would feel and act if they were an intermediary substance in a reaction mechanism.

Specific Learning Outcome

C12-3-08: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

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GLO: D3

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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)

Specific Learning Outcome

C12-3-09: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

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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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In the above example, the $\mathbf{Cl_{2(g)}}$ is a **catalyst** and the $\mathbf{ClO_{(g)}}$ is an **intermediate**.

The slowest of the elementary processes will determine the rate of the reaction. It is called the **rate determining step.**

Specific Learning Outcome

C12-3-09: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
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GLO: D3

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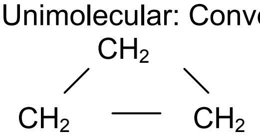

$\rightarrow$ CH_3-CH=CH_2

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Bimolecular: Production of nitrogen dioxide
Both elementary steps for the production of nitrogen dioxide involve two particles.
Step 1 $NO_{(g)}$ + $NO_{(g)}$ → $\mathbf{N_2O_{2(g)}}$
Step 2 $\mathbf{N_2O_{2(g)}}$ + $O_{2(g)}$ → $2NO_{2(g)}$

Termolecular:
There are very few reactions that require three particles to react simultaneously in an elementary step.

Extension
Have students draw potential energy diagrams for multistep reaction mechanisms.

Suggestions for Assessment
<u>Paper and Pencil Tasks</u>
Students can create their own analogy of a reaction mechanism.

<u>Journal Writing</u>
Students can write in their journals how they would feel and act if they were an intermediary substance in a reaction mechanism.

Specific Learning Outcome

C12-3-09: Explain the concept of a reaction mechanism.
Include: rate determining step (1 hour)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

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…
GLO: D3

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)

Specific Learning Outcome

C12-3-10: Determine the rate law and order of a chemical reaction from experimental data.
Include: reactions that are zero, first or second order, rate versus concentration graphs. (2 hours)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

Teacher Notes

The differentiated rate law is determined by using the initial rates method. The integrated rate law is determined by using the concentration change over time to determine rate. Teachers should avoid the use of the integrated rate law as this involves the use of calculus. The use of the initial rate method should be stressed. 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 textbooks deal with this topic in detail. The depth of instruction should be determined by the level of your students. An introductory example would be:

For the reaction:

A → B

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.)
It can be written as

Rate $\propto [A]^x$

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

If we wish to evaluate this mathematically, we need to replace the proportionality symbol with an equal sign. To do this we must include a proportionality constant. In this case it is called the rate constant, which is abbreviated, k.

rate = $k[A]^x$
In this data, x is equal to 1.
<u>Sample Problem</u>:
For the reaction, $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$, the following data was obtained. Determine the overall rate law for this reaction.

Specific Learning Outcome

C12-3-10: Determine the rate law and order of a chemical reaction from experimental data.
Include: reactions that are zero, first or second order, rate versus concentration graphs. (2 hours)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
Examples: analogies, concept frames, concepts maps, manipulatives, particulate representations, role-plays, simulations, sort-and-predict frames, word cycles …
GLO: D3

Trial	Initial rate (mol/L·s)	Initial $[NO_2]$ (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 the two trials in which only one reactant is changed.

$\frac{\text{Trial 2 } [NO_2]}{\text{Trial 1 } [NO_2]} = \frac{0.40}{0.10}$ = 4 times (quadrupled the concentration)

$\frac{\text{Trial 2 rate}}{\text{Trial 1 rate}} = \frac{0.080}{0.0050}$ = 16 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 rate depends on the square of the concentration of NO_2. The reaction is second order with respect to NO_2.

The rate law would be rate = $k[NO_2]^2$

(2) Now take the ratio of the initial rates for trials 1 and 3. These trials have the concentration of CO changing.

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

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

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

The rate law would be rate = $k[NO_2]^2[CO]^0 = k[NO_2]^2(1) = k[NO_2]^2$

The teacher should stress that the value of k is specific for each reaction and only changes for a given reaction if the temperature changes.

Specific Learning Outcome

C12-3-10: Determine the rate law and order of a chemical reaction from experimental data.
Include: reactions that are zero, first or second order, rate versus concentration graphs. (2 hours)
GLO: D3

C12-0-U1 Use appropriate strategies and skills to develop an understanding of chemical concepts.
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GLO: D3

<u>Summary Chart</u>
How Rates Depend on Concentration Change and Order of Reaction

	Order of Reaction			
Concentration Change	0	1	2	3
x 1	$1^0 = 1$	$1^1 = 1$	$1^2 = 1$	$1^3 = 1$
x 2 (doubling)	$2^0 = 1$	$2^1 = 2$	$2^2 = 4$	$2^3 = 8$
x 3 (tripling)	$3^0 = 1$	$3^1 = 3$	$3^2 = 9$	$3^3 = 27$

[chart obtained van Kessel 374]

<u>Laboratory Activities</u>

If teachers have sufficient time, then the students can perform the following lab activities.
Lab 14 Determining Reaction Orders (Small-Scale Laboratory Manual, p.53, Dingrando *et al*., Glencoe-McGraw-Hill, 2005)
In this lab the general equation for the reaction between crystal violet and sodium hydroxide is to be determined.
Experiment 30 Rate Law Determination of the Crystal Violet Reaction (Chemistry with Calculators, p.30-1, Holmquist et al, Vernier, 2000)
I've Got the Chemistry Blues: Determining Reaction Order (p.512, Pasport Explorations in Chemistry, Pasco Scientific, 2003.
This experiment has the students analyze the reaction rate by determining the order of the reaction when a colouring agent reacts with household bleach.

Suggestions for Assessment

<u>Paper and Pencil Tasks</u>
Have students solve problems involving rate laws.

<u>Journal Writing</u>
Students can describe the effect of doubling, tripling, and quadrupling [A] on the overall rate of chemical reactions having the following rate laws: Rate = $k[A]^0$; Rate = $k[A]^1$; Rate = $k[A]^2$; Rate = $k[A]^3$.
(p.544 Teacher's Guide, Chemistry: Matter and Change, Dingrando *et al*., Glencoe-McGraw-Hill, 2005)

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)

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

Specific Learning Outcome

C12-3-10: Determine the rate law and order of a chemical reaction from experimental data.
Include: reactions that are zero, first or second order, rate versus concentration graphs. (2 hours)
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GLO: D3

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

Investigations
Chemistry with Calculators, Holmquist et al, Vernier, 2000
Pasport Explorations in Chemistry, Pasco Scientific, 2003.

Small-Scale Laboratory Manual, (Dingrando *et al*., Glencoe-McGraw-Hill)
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