CHEM1111-GENERAL INORGANIC CHEMISTRY I LAB MANUAL-ONLINE SECTION

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CHEM1112-General Inorganic Chemistry II
Lab Manual-Online Section

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CHAPTER OVERVIEW

1: Introduction and Laboratory Safety

1.1: Preface

Chemistry is an empirical science. Hence, experiments play an essential role in teaching chemistry. Other methods of teaching cannot substitute experimentation. The concepts learned in lectures can be reaffirmed through lab experiments. Students entering this course should be able to improve basic laboratory skills like technical comprehension, planning, execution, observation, analysis and scientific reporting. All the lab experiments included in this manual are tied to the materials covered in the lectures. Ideally, the lecture should precede the related lab experiment. However, the instructor may choose a different order depending on other factors. The questions at the end of the lab sheets are designed to improve the problem solving skills by applying the learning that is gained through each experiment. CHEM 1112: General Inorganic Chemistry-2 course online section will not meet at the campus laboratory for doing lab experiments. The lab experiments for this course need to be done remotely, at student's home. There are many lab procedures that are borrowed from standard lab manuals meant for real laboratory settings, but are remixed and modified to make them more greener and affordable so that the online students are not required to purchase hazardous and expensive supplies. Chemical reactions have always been fun if performed safely. Read the laboratory safety protocols before working on any experiment.

1.2: Laboratory Safety Protocols-Home Version

Note

When you are in a chemistry lab (wherever you are conducting the experiment) you need to adhere to the following practices for the sake of your own safety as well as for the safety of others around you.

- **Do not eat or drink in the lab. If you have food items with you keep that outside the lab.** Do not leave any containers of food/drinks inside the lab.
- **Wear closed-toe footwear, long pants, and long sleeved shirts/tops to the lab.** *This is to protect you from serious burns or allergies in case your body comes in close contact with chemicals, especially in the case of a chemical spill.*
- **Do not wear contact lenses.** *Even minute amount of corrosive chemicals can damage your eye sight by the time you remove your contact lenses and wash your eyes. Chemicals can get into your eyes due to evaporation or a chemical spill.*
- **Tie your hair.** *For example, if you are using an open flame long hair might catch fire if left untied.*
- **No dangling ear rings or jewelry.** *For the same reason as noted above.*
- **Wear safety glasses or goggles and a lab coat or apron as advised**. Safety glasses are available in the lab and you do not need to purchase them. Once you finish your experiment return them back to their original cabinet.
- **Wear appropriate gloves before handling chemicals.** Your instructor will advise you when to wear gloves.
- **Do not fool around in the lab.** *It may be fun to be around your classmates, and you may have a lot to share especially when you are waiting on an experiment to record the observation. However, it can lead to serious accidents if you do not pay close attention to the experiment that you are conducting because chemicals can cause serious accidents.*
- **Clear the clutter on the work bench/table**. Leave only the lab manual, pen/pencil, textbook, and calculator on the workbench. *Even traces of chemicals on the workbench can ruin your belongings.*
- **Never lean on the work bench.** *The same reason as above.*
- **Never taste, touch, or smell chemicals unless you are instructed to do so.**
- **Do not perform any unauthorized experiments.** Do not mix chemicals unless instructed.
- **Read chemical labels on the bottles carefully before you handle them.** Read Safety Data Sheet (SDS) provided with each chemical.
- **Unused chemicals should never be returned back into the chemical stock bottles.**
- **Use caution while handling hot objects and glass wares.** Use tongs or gloves as necessary.
- **Broken glass wares should be disposed properly. Do not touch them with your bare hands.**
- **Do not drain any chemicals or solids into the sink unless specified.**
- **Never leave any hot plate or burners unattended.**
- **Never reach across an open burner or hot glassware or a glassware that contain chemicals.**
- **Never pour water into acid to avoid corrosive acid splashes.**
- **Seek help immediately if a chemical spill or accident occurs.**
- **Read the experiment procedure a couple of times carefully before you start the experiment.**
- **Before you dispose any chemical waste as instructed. If you are unsure about how to dispose any unused chemicals check with your instructor.**
- **Clean the glassware and return them back to their proper place.**
- **Clean your work bench after the experiment.**

- **Check the location of first aid measures and wash stations like fire blanket, fire extinguisher, eye wash stations, shower, first aid box, and SDS**
- **Wash your hands thoroughly with soap after you handle any chemical, and also before you leave the lab.**

Hazard Identification GHS Pictograms

Note

Watch for the following hazard identification pictograms while handling, storing, and disposing chemicals including household ones.

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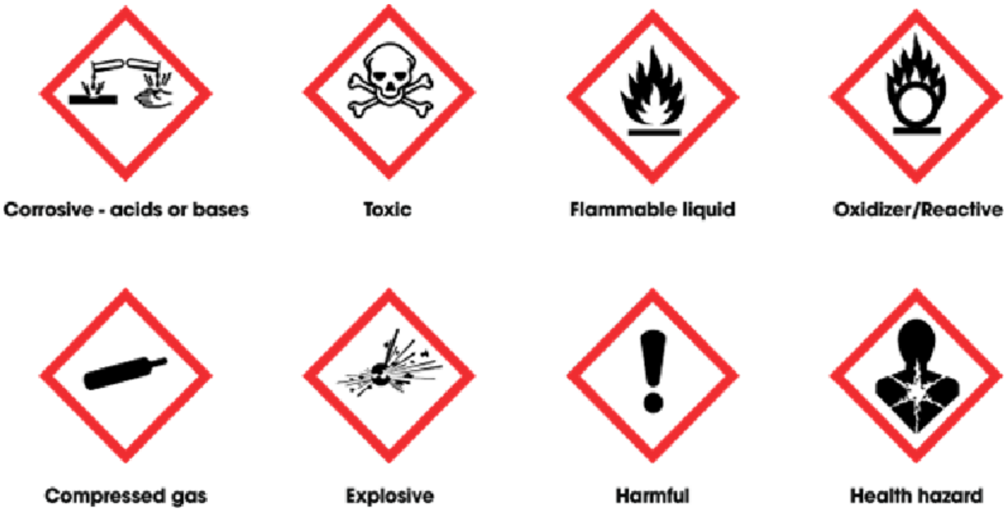
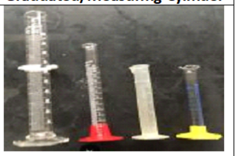
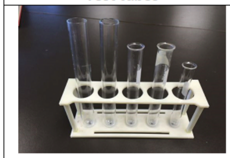
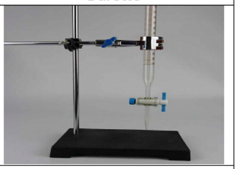
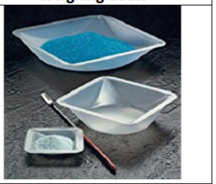
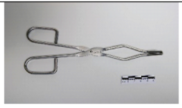
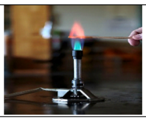
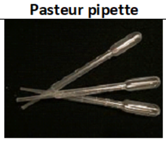


1.3: Common General Chemistry Labwares and Equipments

Note

The photos given below are by Unknown Authors are licensed under CC BY-SA-NC

Beaker	**Erlenmeyer's or Conical flask**	**Graduated/Measuring Cylinder**
		
Test tubes	**Pipette**	**Burette**
		
Electronic scale/balance	**Stir/glass rod in a beaker**	**Weighing boats**
		
Spatula	**Tongs**	**Bunsen burner**
		
Hot plate	**Pasteur pipette**	**Funnel and filter paper**
		

1.4: Materials Required for the Lab Experiments

	Materials Required Red items are the new items that are not in the CHEM 1111 lab kit		
1	Beaker, 100-ml, 400-mL	7	Hydrogen peroxide, 3% solution, 100 ml,
2	Electronic balance, pocket size	8	Tincture iodine antiseptic (2%) 60 ml
3	Erlenmeyer flask, 250 ml	9	Cornstarch
4	Funnel	10	Red cabbage -1
5	Graduated cylinders, 10-mL and 100-mL	11	Elmer's glue, 1
6	PH strips	12	Borax, 50-100 g
7	Pipettes, small (and long graduated ones to use instead of a burette) with bulb	1	small glass tumblers or bottles
8	Thermometer	2	Graph papers
9	Small test tubes, 7-10	3	Metric ruler
10	Glass stirring rod	4	A wider clear glass bowl to make ice bath
1	Baking soda (sodium bicarbonate)-16 Oz Not Baking Powder	5	Capped glass vial
2	Vitamin C tablets (Nature Best brand) 500 or 1000 mg supplements, 3-5 tablets	6	Spoons, plastic
3	Fruit juice or any beverage containing vitamin C.	7	Styrofoam cups, 5-10
4	Sugar, 25 g	8	Saucepan/microwave safe bowl for boiling
5	Table salt, 200 g	9	Heat resistant gloves
6	Vinegar (5% acidity-Acetic acid), 1 gallon	10	Stove or microwave
		11	Goggles
		12	Gloves

Some Good Lab Practices

Measurement using any instrument

Make sure you are comparing the measurement value to the required unit. For example to measure the length of a book using a ruler in “mm”; with a ruler that has both “mm” and “cm” units on either side, make sure you are measuring using the right side or in the right unit (“mm”) that you need. If it is an electronic device check the “mode, and change it to the appropriate unit mode you need.

Using an electronic balance (scale)

An electronic balance is used to measure the mass of an object. Sometimes you are required to measure the mass of both the container and the object, and sometime you are required to measure the mass of the object only that is placed in a container. The first step as mentioned earlier is to check that the unit shown in the instrument is the one that you want. Otherwise press the mode button and bring it to the desired unit (it is usually the SI units). Then place the container and the object together to find the reading. Write down the measurement followed by the unit. If you do not want the mass of the container, keep the container on the scale, press the “zero” or “tare” button, so that the mass of the container is zeroed out. Now you can place the object in the container to find the mass of the object only.

Measuring volumes using graduated tubes or columns (volumetric measurements)

Volumes can be measured using either an instrument with a base like graduated cylinder or a tube like buret. In both the cases hold or place them so that the grading or marking on the tube and your eyes are on the same straight line (leveled). It is easier to keep the graduated cylinder on the work bench and kneel down to read the measurement. For a buret it can be clamped on a stand to read the measurement. For a colorless liquid the reading matching with the lower meniscus is taken, and for an opaque or colored liquid upper meniscus level is measured.

Using a pipet

Always use a mechanical or electric suction device to draw liquid into a pipet. Draw a little more than you need, and then slowly release the suction with great care to the nearest mark, pull it out from the stock container and then slowly release the liquid to the receiving container.

Using a buret

Open the knob (keep the knob vertical) and place the tip towards the drain, and rinse with distilled water using a wash (squirt) bottle. Rinse the buret two or three times. Then hold the buret slanted over a waste bottle or collection beaker and slowly rinse the inside of the tube with the liquid that you want to fill it with. Rinse the buret with 3-5ml twice. Then close the knob (horizontal), and fill with the liquid you want. Fill it a little more than the top most reading. Hold the closed buret over the collection beaker again, and slowly open it (drip) to level with the top most reading (usually zero). This way you are removing the trapped air bubbles near the knob that could give rise to errors in your measurement.

Rinsing a glassware

Before you use any glassware make sure it is clean. Wash with a wire brush and soap until the glass is clean. After washing rinse it with 5-10ml of distilled water.

Washing used test tubes

Dispose the chemicals in the test tube into appropriate waste containers as instructed. Rinse with water and transfer the first two rinsing into the waste container itself. Then bring it to the sink and wash with wire brush, soap, and cold water. Then you can wash with hot water and soap if necessary. Rinse with distilled water. Make sure you wear gloves during the whole process. After removing gloves wash your hands thoroughly with soap and water.

Handling hot glassware

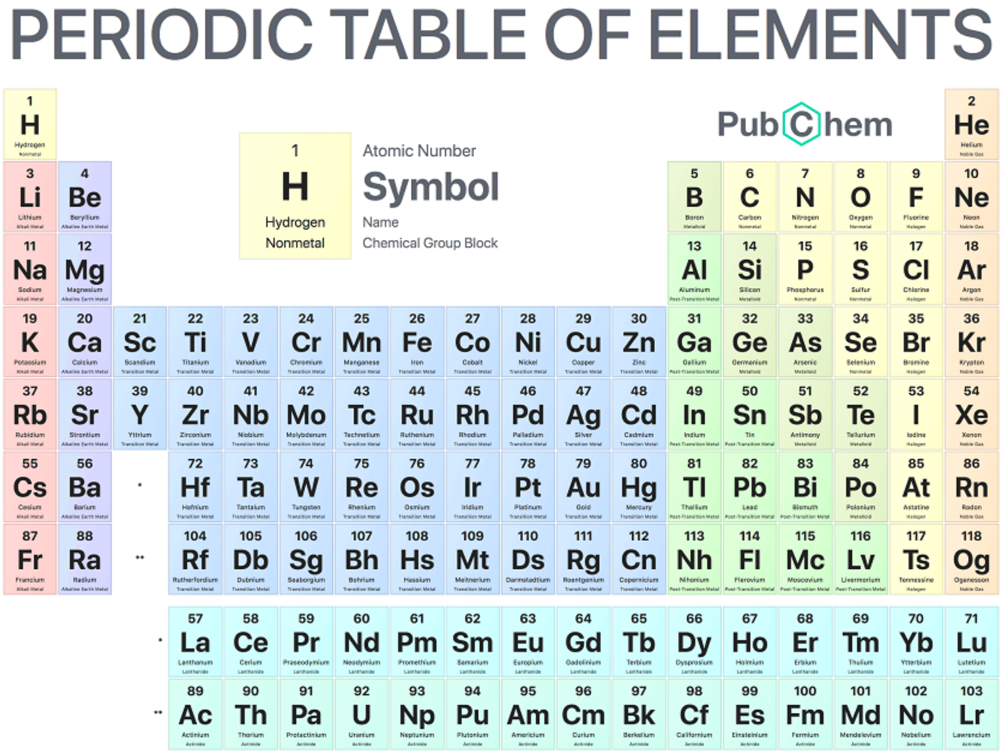
Never touch any hot glassware with your bare hands. Wear gloves, or use tongs or test tube holders as necessary. Never leave them on the table without a heat resistance mat or platform to cool off.

Diluting acids

Mixing acid and water will release heat. If you want to dilute an acid, take the required amount of water in bigger glassware in the fume hood and add the acid drop-by-drop into the water with vigorous shaking or mixing. It is also advisable to keep the water in an ice bath or cold water bath while you drip the acid into it. Wear corrosion resistant gloves, goggles, and lab coat.

Filtering a heterogeneous liquid-solid mixture

Find the appropriate filter paper with right dimension and mesh size. If it is a conical funnel, fold the filter paper into four or eight fold. Keep it in the funnel with the pointed side towards the bottom. Wet the paper with a very small amount of distilled water so that the paper sticks to the funnel. Slowly pour the top liquid in portions (never add more than the 3/4th of the funnel height. Let the first portion drain, and then add the second. When the filtration is done rinse the remaining solid with fresh liquid, mix, and then repeat the filtration. Never touch the filter paper with sharp tips, like spatula or glass road.



PERIODIC TABLE OF ELEMENTS

PubChem

1 H Hydrogen Nonmetal

Atomic Number
Symbol
Name
Chemical Group Block

1 H Hydrogen																	2 He Helium
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminum	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Cesium	56 Ba Barium	*	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	**	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Nh Nihonium	114 Fl Flerovium	115 Mc Moscovium	116 Lv Livermorium	117 Ts Tennessine	118 Og Oganesson
		*	57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
		**	89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

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1.6: Acknowledgements

Manjusha Saraswathiamma would like to acknowledge the organizers of the M State OER Learning Circle, especially Mr. Marcus Lacher (Business/Computers Faculty) Dr. Anne Thurmer (Dean of Liberal Arts and Humanities), and Ms. Kari Oanes (College Librarian) for providing an avenue to learn more about Open Educational Resources. Their encouragement and patience are much appreciated. This project is supported by a Z-degree grant from The Minnesota State System Office through the M State OER Circle. Manjusha would also like to acknowledge Dr. Mathew Brochdering (Dean for the School of Science and Mathematics) for his support.

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CHAPTER OVERVIEW

2: Experiments

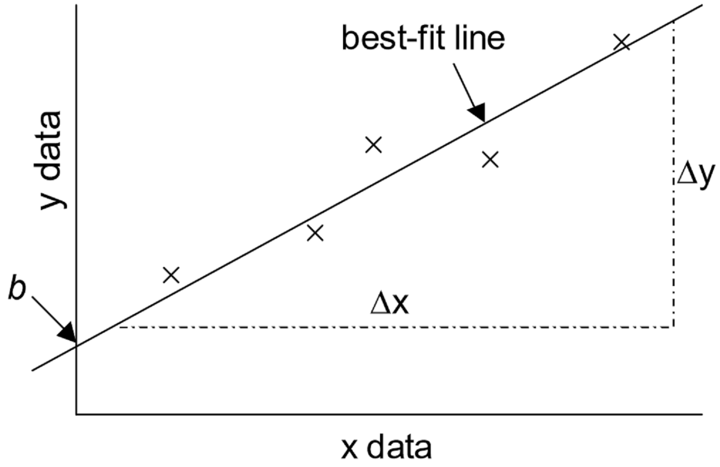
2.1: Using Excel for Graphical Analysis of Data (Experiment)

Objectives

- To learn to use Excel to explore a number of linear graphical relationships.

In several upcoming labs, a primary goal will be to determine the mathematical relationship between two variable physical parameters. Graphs are useful tools that can elucidate such relationships. First, plotting a graph provides a visual image of data and any trends therein. Second, via appropriate analysis, they provide us with the ability to predict the results of any changes to the system.

An important technique in graphical analysis is the transformation of experimental data to produce a straight line. If there is a direct, linear relationship between two variable parameters, the data may be fitted to the equation of line with the familiar form $y = mx + b$ through a technique known as linear regression. Here m represents the slope of the line, and b represents the y-intercept, as shown in the figure below. This equation expresses the mathematical relationship between the two variables plotted, and allows for the prediction of unknown values within the parameters.



The equation for the best-fit line is

$$y = mx + b \tag{2.1.1}$$

where

$$b = y\text{-intercept} \tag{2.1.2}$$

$$m = \text{slope} \tag{2.1.3}$$

$$= \frac{\Delta y}{\Delta x} \tag{2.1.4}$$

$$= \frac{y_2 - y_1}{x_2 - x_1} \tag{2.1.5}$$

Computer spreadsheets are powerful tools for manipulating and graphing quantitative data. In this exercise, the spreadsheet program Microsoft Excel$^{©}$ will be used for this purpose. In particular, students will learn to use Excel in order to explore a number of linear graphical relationships. Please note that although Excel can fit curves to nonlinear data sets, this form of analysis is usually not as accurate as linear regression.

Procedure

Part 1: Simple Linear Plot

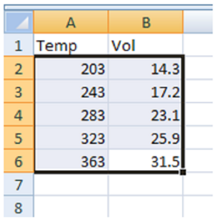
Scenario: A certain experiment is designed to measure the volume of 1 mole of helium gas at a variety of different temperatures, while keeping the gas pressure constant at 758 torr:

Temperature (K)	Volume of Helium (L)
203	14.3
243	17.2
283	23.1
323	25.9
363	31.5

1. Launch the program Microsoft Excel© (2016 version, found on all computers in all the computer centers on campus). Go to the Start button (at the bottom left on the screen), then click Programs, followed by Microsoft Excel©.
2. Enter the above data into the first two columns in the spreadsheet.

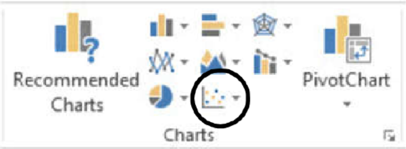
- Reserve the first row for column labels.
- The x values must be entered to the left of the y values in the spreadsheet. Remember that the independent variable (the one that you, as the experimenter, have control of) goes on the x-axis while the dependent variable (the measured data) goes on the y-axis.

3. Highlight the set of data (not the column labels) that you wish to plot (Figure 1).



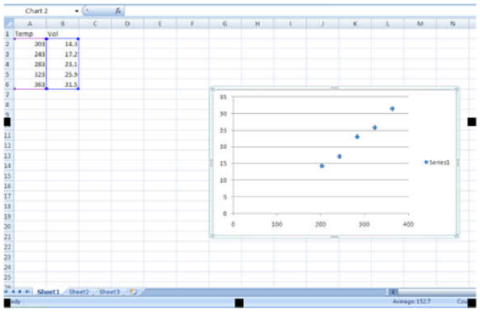
	A	B
1	Temp	Vol
2	203	14.3
3	243	17.2
4	283	23.1
5	323	25.9
6	363	31.5
7		
8		

- Click on *Insert* > *Recommended Charts* followed by *Scatter* (Figure 2).



- Choose the scatter graph that shows data points only, with no connecting lines – the option labeled *Scatter with Only Markers* (Figure 3).

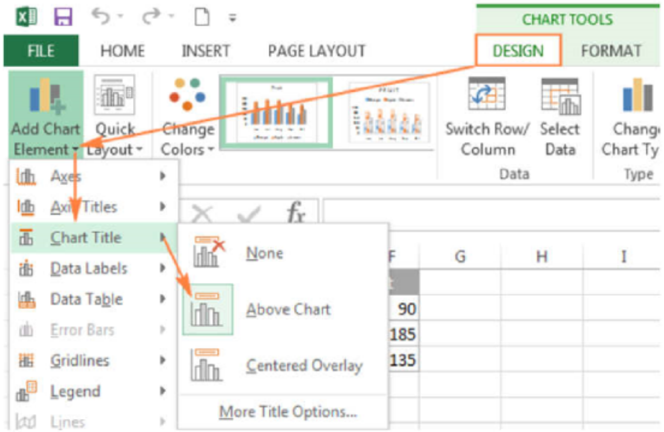
- You should now see a scatter plot on your Excel screen, which provides a preview of your graph (Figure 4).



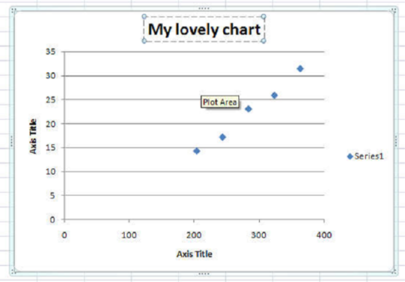
4. If all looks well, it is time to add titles and label the axes of your graph (Figure 5).

- First, click inside the chart.

- Switch to the *Design* tab, and click *Add Chart Element* > *Chart Title* > *Above Chart*
- The graph should be given a meaningful, explanatory title that starts out “Y versus X followed by a description of your system.
- Click on *Axis Titles* (select *Primary Horizontal Axis Title* and *Primary Vertical Axis Title*) to add labels to the x- and y-axes. Note that it is important to label axes with both the measurement and the units used.



- To change the titles, click the text box for each title, highlight the text and type in your new title (Figure 6).



5. Your next step is to add a trendline to the plotted data points. A trendline represents the best possible linear fit to your data. To do this you first need to "activate" the graph. Do this by clicking on any one of the data points. When you do this, all the data points will appear highlighted.

- Click the *Chart Elements* button + next to the upper-right corner of the chart.
- Check the *Trendline* box.
- Click *More Options*. This will display the option shown in Figure 7.
- Notice that the Linear button is already selected. Now select the *Display Equation on Chart* box and the *Display R-squared value on Chart* box. Then click the *Close* button.

Format Trendline

TRENDLINE OPTIONS

○ Exponential
● Linear
○ Logarithmic
○ Polynomial Order 2
○ Power
○ Moving Average Period 2

Trendline Name

● Automatic Linear (East Asia Sales 2010)
○ Custom

Forecast

Forward 0.0 periods
Backward 0.0 periods

☐ Set Intercept 0.0
☐ Display Equation on chart
☐ Display R-squared value on chart

6. The equation that now appears on your graph is the equation of the fitted trendline. The R^2 value gives a measure of how well the data is fit by the equation. The closer the R^2 value is to 1, the better the fit. Generally, R^2 values of 0.95 or higher are considered good fits. Note that the program will always fit a trendline to the data no matter how good or awful the data is. You must judge the quality of the fit and the suitability of this type of fit to your data set.
7. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:

- the equation of the best-fit trendline to your data
- the slope of the trendline
- the y-intercept of the trendline
- whether the fit of the line to the data is good or bad, and why.

8. By graphing the five measured values, a relationship is established between gas volume and temperature. The graph contains a visual representation of the relationship (the plot) and a mathematical expression of the relationship (the equation). It can now be used to make certain predictions.

    For example, suppose the 1 mole sample of helium gas is cooled until its volume is measured to be 10.5 L. You are asked to determine the gas temperature. Note that the value 10.5 L falls outside the range of the plotted data. How can you find the temperature if it doesn't fall between the known points? There are two ways to do this.

Method (1): Extrapolate the trendline and estimate where the point on the line is.

- Click on the *Layout* tab along the top menu, then *Trendline* > *More Trendline Options*.
- In the section labeled *Forecast* enter a number in the box labeled *Backward*, since we want to extend the trendline the backward x direction. To decide what number to enter, look at your graph to see how far back along the x-axis you need to go in order to cover the area where volume = 10.5 L. After entering a number, click *Close*, and the line on your graph should now be extended in the backward direction.
- Now use your graph to estimate the x value by envisioning a straight line down from y = 10.5 L to the x-axis. Record this value on your report.

Method (2): Plug this value for volume into the equation of the trendline and solve for the unknown temperature. Do this and record your answer on your report. Note that this method is generally more precise than extrapolating and "eyeballing" from the graph.

Part 2: Two Data Sets with Overlay

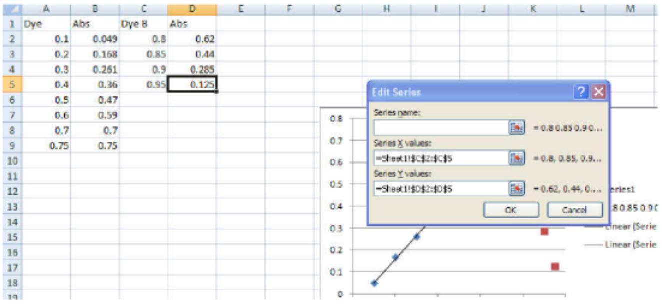
Scenario: In a certain experiment, a spectrophotometer is used to measure the light absorbance of several solutions containing different quantities of a red dye. The two sets of data collected are presented in the table below:

Data A		Data B	
Amount of Dye (mol)	Absorbance (unitless)	page5image24744 Amount of Dye (mol)	page5image26088 Absorbance (unitless)
0.100	0.049	0.800	0.620
0.200	0.168	0.850	page5image366080.440
0.300	0.261	0.900	0.285
0.400	0.360	0.950	0.125
0.500	0.470		
0.600	0.590	page5image51640	page5image52408
0.700	0.700	page5image56816	page5image57760
0.750	0.750		

You would like to see how these two sets of data relate to each other. To do this you will have to place both sets of data, as independent relationships, on the same graph. Note that this process only works when you have the same axis values and magnitudes.

1. Enter this new data on a fresh page (Sheet 2) in Excel. Be sure to label your data columns A and B. Again, remember to enter the x values to the left of the y values.
2. First, plot Data A only as an XY Scatter plot (the same way you did with the data in Part 1). Fit a trendline to this data using linear regression, and obtain the equation of this line.
3. Now you need to add Data B to this graph.

- Activate the graph by clicking on one of the plotted data points.
- Right-click the chart, and then choose *Select Data*. The *Select Data Source* box appears on the worksheet with the source data of the chart.
- Click the *Add* tab and type “Data B” for the Series Name.
- Click the little icon under *Series X values*, then highlight the x-axis values of Data B.
- Press enter, then repeat this procedure for the *Series Y Values*, highlighting the y-axis values of Data B. For each of these steps, you should see a display similar to what is shown in Figure 8. Note that slight differences may appear due to the version of Microsoft Excel© installed on your computer.



	A	B	C	D
1	Dye	Abs	Dye B	Abs
2	0.1	0.049	0.8	0.62
3	0.2	0.168	0.85	0.44
4	0.3	0.261	0.9	0.285
5	0.4	0.36	0.95	0.125
6	0.5	0.47		
7	0.6	0.59		
8	0.7	0.7		
9	0.75	0.75		

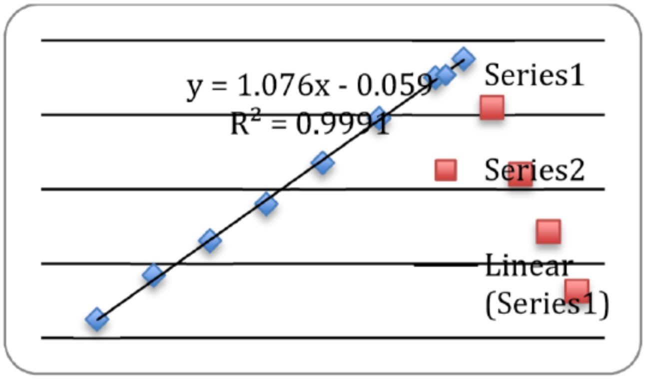
Edit Series

Series name: = 0.8 0.85 0.9 0...

Series X values: =Sheet1!$C$2:$C$5 = 0.8, 0.85, 0.9...

Series Y values: =Sheet1!$D$2:$D$5 = 0.62, 0.44, 0...

- Click OK twice to return to the main Excel window.
- At this point you should see the new data points (labeled as Series 2) as shown in Figure 9. You can now independently analyze this dataset by inserting a trendline as before.



4. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:

- the equation of the best-fit trendline for Data A,
- the equation of the best-fit trendline for Data B,
- If these trendlines were extrapolated, they would intersect. Determine the values of x and y for the point of intersection using simultaneous equations.

Part 3: Statistical Analysis and Simple Scatter Plots

When many independent measurements are made for one variable, there is inevitably some scatter (noise) in the data. This is usually the result of random errors over which the experimenter has little control.

Scenario: Ten different students at two different colleges each measure the sulfate ion concentration in a sample of tap water:

College #1 dataset	35.9 ppm	43.2 ppm	33.5 ppm	35.1 ppm	32.8 ppm	37.6 ppm	31.9 ppm	36.6 ppm	35.0 ppm	32.0 ppm
College #2 dataset	45.1 ppm	34.2 ppm	36.8 ppm	31.0 ppm	40.7 ppm	29.6 ppm	35.4 ppm	32.5 ppm	43.5 ppm	38.8 ppm

Simple statistical analyses of these datasets might include calculations of the mean and median concentration, and the standard deviation. The mean ($\bar{x}$) is simply the average value, defined as the sum (Σ) of each of the measurements (x_i) in a data set divided by the number of measurements (N):

$$\bar{x} = \frac{\sum x_i}{N} \tag{2.1.6}$$

The median (M) is the midpoint value of a numerically ordered dataset, where half of the measurements are above the median and half are below. The median location of N measurements can be found using:

$$M = \frac{N+1}{2} \tag{2.1.7}$$

When N is an odd number, the formula yields a integer that represents the value corresponding to the median location in an ordered distribution of measurements. For example, in the set of numbers (3 1 5 4 9 9 8) the median location is (7 + 1) / 2, or the 4th value. When applied to the numerically ordered set (1 3 4 5 8 9 9), the number 5 is the 4th value and is thus the median – three scores are above 5 and three are below 5. Note that if there were only 6 numbers in the set (1 3 4 5 8 9), the median location is (6 + 1) / 2, or the 3.5th value. In this case the median is half-way between the 3rd and 4th values in the ordered distribution, or 4.5.

Standard deviation (s) is a measure of the variation in a dataset, and is defined as the square root of the sum of squares divided by the number of measurements minus one:

$$s = \sqrt{\frac{\sum (x_i - \bar{x})^2}{N-1}} \tag{2.1.8}$$

So to find s, subtract each measurement from the mean, square that result, add it to the results of each other difference squared, divide that sum by the number of measurements minus one, then take the square root of this result. The larger this value is, the greater the variation in the data, and the lower the precision in the measurements.

While the mean, median and standard deviation can be calculated by hand, it is often more convenient to use a calculator or computer to determine these values. Microsoft Excel© is particularly well suited for such statistical analyses, especially on large datasets.

1. Enter the data acquired by the students from **College #1** (only) into a single column of cells on a fresh page (Sheet 4) in Excel. Then in any empty cell (usually one close to the data cells), instruct the program to perform the required functions on the data. To compute the mean or average of the data entered in cells a1 through a10, for example, you must:

- click the mouse in an empty cell
- type "=average(a1:a10)"
- and press return

To obtain the median you would instead type “=median(a1:a10)”. To obtain the standard deviation you would instead type "=stdev(a1:a10)".

2. Record on your report:

- The Excel calculated mean, median and standard deviation for the College #1 dataset.
- As an additional exercise, calculate the standard deviation of this dataset by hand, and compare it to the value obtained from the program.

Rejecting Outliers

Do all the measurements in the College #1 data set look equally good to you, or are there any values that do not seem to fit with the others? If so, are you allowed to reject these measurements?

Outliers are data points which lie far outside the range defined by the rest of the measurements and may skew your results to a great extent. If you determine that an outlier resulted from an obvious experimental error (e.g., you incorrectly read an instrument or prepared a solution), you may reject the point without hesitation. If, however, none of these errors is evident, you must use caution in making your decision to keep or reject a point. One rough criterion for rejecting a data point is if it lies ***beyond two standard deviations from the mean or average.***

3. Using the above criteria, determine if there are any outliers in the College #1 dataset.

- Record these outlier measurements (if any) on your report.
- Then, excluding the outliers, re-calculate the mean, median and standard deviation of this data set (use Excel).

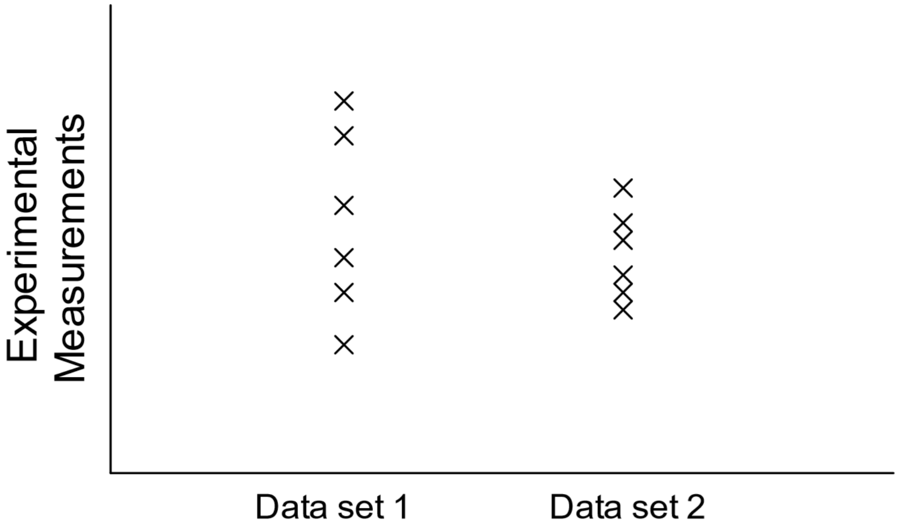
Rejecting data points cannot be done just because you want your results to look better. If you choose to reject an outlier for any reason, you must always include documentation in your lab report which clearly states:

- that you did reject a point
- which point you rejected
- why you rejected it

Failure to disclose this could constitute scientific fraud.

Graphing a Scatter Plot

Unlike the linear plots created so far, a scatter plot simply shows the variation in measurements of a single variable in a given dataset, i.e., it supplies a visual representation of the “noise” in the data. The data is plotted in a column, and there is no x-y dependence here (Figure 10). Note that datasets with a greater degree of scatter will have a higher standard deviation and consist of less precise measurements than datasets with a small degree of scatter.



To obtain such a plot using Excel, all the x values for each dataset must be identical. Thus, let the College #1 data be assigned x = 1, and let x = 2 for all the College #2 data:

page9image14016**Measurements by Students from College #1**		page9image15440**Measurements by Students from College #2**	
College 1	[SO$_4^{-2}$] (ppm)	College 2	[SO$_4^{-2}$] (ppm)
1	35.9	2	page9image24768 page9image2508845.1
1	page9image2791243.2	2	page9image30392 page9image3071234.2 page9image31960
1	33.5	2	36.8
1	35.1	2	31.0
1	32.8	2	40.7
1	page9image4592037.6	2	page9image47728 page9image4804829.6
1	page9image5087231.9	2	page9image53352 page9image5367235.4
1	36.6	2	32.5
1	35.0	2	43.5
1	32.0	2	38.8

4. Enter the data as shown above into the first four columns of your spreadsheet.

- Plot the College #1 dataset as an XY Scatter Plot.
- Now add the College #2 dataset to this graph applying the same steps you used to create your earlier graph in the section "Two Data Sets with Overlay" (Part 2).
- Add appropriate axis labels and a title. You may also want to adjust the x-axis and y-axis scales to improve the final look of your graph.

5. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:

- Which dataset (College #1 or College #2) show the least scatter? The greater standard deviation? The more precise measurements?

Lab Report: Using Excel for Graphical Analysis of Data

Name: ______________________________ Lab Partner: __None for this assignment__

Date: _________________________ Lab Section: ___________________

Turn in the graphs you made for ALL three parts in this assignment

For each graph make sure the following components are in the printout:

1. Title for the graph
2. Labels for x and y axes (along with appropriate units when applicable)
3. Line equation and R^2 when appropriate.

Part 1: Simple Linear Plot

- Which set of data is plotted on the y-axis?
- Which set of data is plotted on the x-axis?
- Record the following information:

The equation of the fitted trendline

The value of the slope of this line

The value of the y-intercept of this line

- Is the fit of the trendline to your data good (circle one)? Yes / No. Explain why you think the line is a good fit to the data.
- Determine the temperature (in K) of the gas in the cold room when it has a measured volume of 10.5 L using

a) Extrapolation and “eyeballing”

b) The equation of the trendline

Show your calculations for b) below.

Part 2: Two Data Sets and Overlay

- Record the equations of the trendlines fitted to
  - Data set A:
  - Data set B:
- Perform a simultaneous equations calculation to determine the x and y values for the point of intersection between these lines. Show your work below.

Part 3: Statistical Analysis and Simple Scatter Plots

- For the College #1 data set, record the following values (determined using Excel):
  - the mean SO_4^{-2} concentration
  - the median SO_4^{-2} concentration
  - the standard deviation in the data set
- Calculate the standard deviation in the College #1 data set *by hand*. Show *all* your work below. Continue your work on an attached page if you require more space.
- Are there any outliers in the College #1 data set (circle one)? Yes / No
- If yes, which measurements are the outliers?
- Show the calculations you used to identify the outliers (or, if none, how you determined that there were none).
- Re-calculate the following values (using Excel) *excluding the outliers*:
  - the mean SO_4^{-2} concentration
  - the median SO_4^{-2} concentration
  - the standard deviation in the data set
- Create a scatter plot showing both the College #1 and College #2 data. Attach a printout of your graph to this report. Be sure that your axes are properly labeled, and that your graph has an appropriate title.
- Examine your plotted data. Which data set:

has the larger standard deviation?

contains the more precise measurements?

2.2: The Densities of Solutions and Solids (Experiment)-Home Version

Objectives

- To compare the precision of various types of volumetric glassware.
- To determine, and then compare, the densities of regular and diet de-gassed sodas.
- To determine the density of an unknown metal and use this value to identify the metal.
- To use a common spreadsheet program, Microsoft Excel$^{©}$, to perform graphical analysis of experimental data.

Density is a fundamental physical property of matter. Physical properties are those characteristics of a substance that can be determined without changing the chemical identity of the substance. Other physical properties include melting point and solubility. In general, since different substances have unique densities, determining the density of an unknown substance can help identify it. Density is specifically defined as the ratio of a substance's mass to its volume:

$$Density = \frac{Mass}{Volume} \tag{2.2.1}$$

The S.I. unit of density is kg/m^3, but in chemistry it is more often expressed in units of g/cm^3 for solids, and g/mL for liquids and solutions.

In Part A of this lab, a study of density values will be conducted in order to introduce the concept of precision and the statistical nature of experimental data. Specifically, the density of water will be determined using measurements obtained from three different types of volumetric glassware: a buret, a 100-mL graduated cylinder and a volumetric pipet. Multiple measurements will be performed so that a large ensemble of density values is acquired for each type of glassware used. Simple statistical analyses of the three data sets will then be carried out, including calculations of the average density and standard deviation. Note that an average value $(\bar{x})$ is defined as the sum (Σ) of each of the measurements (x_i) divided by the number of measurements (N):

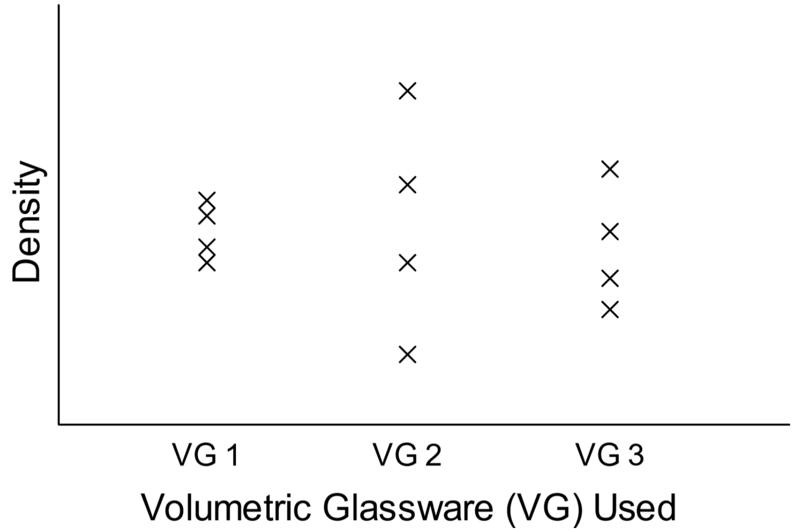
$$\bar{x} = \frac{\sum x_i}{N} \tag{2.2.2}$$

Standard deviation (s) is defined as:

$$s = \sqrt{\frac{\sum (x_i - \bar{x})^2}{N-1}} \tag{2.2.3}$$

Standard deviation essentially indicates the degree to which a set of measured values deviate from the average value. Data sets with a wide range of values (for example, see VG 2 data in the following figure) will have a larger standard deviation and are associated with less precise measurements than data sets with a narrow range of values. With this information, the precision of the measurements obtained using the three types of volumetric glassware can be compared.

The three density data sets will also be displayed in a Scatter Plot, as shown below:

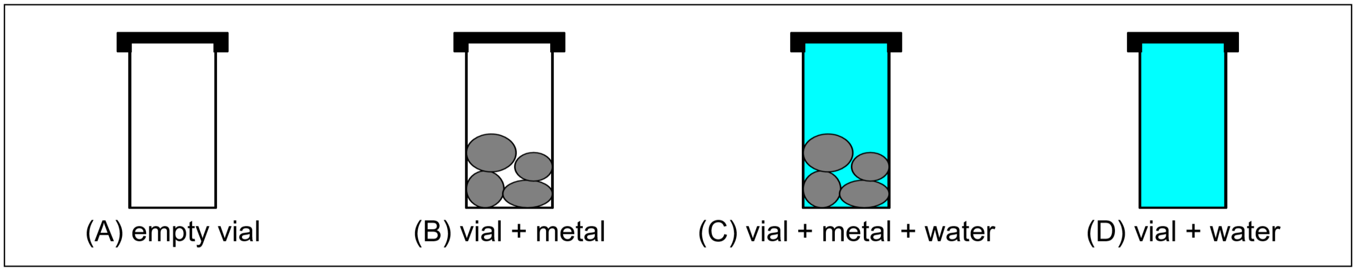


Occasionally one or more measurements in a data set will appear not to “fit in” with the others. This will often be evident upon close examination of a scatter plot. These points are called outliers – values that occur far outside the range defined by the rest of the measurements. One rough criterion for identifying an outlier is that it lies beyond two standard deviations from the average value. Such values may be legitimately excluded from a data set, as they can skew results to a great extent.

In Part B, the densities of two de-gassed sodas (regular and diet) will be determined and compared. Here, a buret will be used to dispense precise soda volumes. The masses of several different volumes of each soda will be measured, and this data will be used to determine the density of each soda.

This section of the lab will also clearly illustrate the intensive nature of density. An intensive property is one that is independent of the amount of matter present. This means that a pure substance will have the same density whether it is a larger sample or a small sample of that substance. In contrast, an extensive property is one that is dependent on the amount of matter present. Note that both mass and volume are extensive, however, since density is a ratio of these properties, density is an intensive property.

In Part C, the density of an unknown metal will be determined, and the metal identified using this value. Since mass can be more precisely measured than volume (due to the limitations of the instrumentation available), the technique employed here will involve mass measurements only. Using a capped glass vial, the following four mass measurements will be obtained:



The difference in masses A and B is the mass of the metal sample. The volume of the metal is obtained by taking the difference between the water volumes in C and D. These water volumes will be calculated using the water masses and the known density of water (see table below). Although simple, this method can yield density results accurate to 0.1%.

Once the density is determined, this experimental value will be compared to the true densities of several known metals (obtained from suggested reference sources) in order to identify the assigned unknown metal. The percent error between this experimental value (EV) and the true density value (TV) of the metal will also be calculated.

$$\text{Percent Error} = \frac{|EV - TV|}{TV} \times 100 \tag{2.2.4}$$

Note that a more accurate experimental value will yield a lower percent error (< 5% is desirable) than a less accurate value.

*Density of Liquid Water Measured for a Range of Temperatures**

Temperature (°C)	Density (g/mL)
16	0.9989
17	0.9988
18	0.9986
19	0.9984
20	0.9982
21	0.9980
22	0.9978
23	0.9975
24	0.9973
25	0.9970
26	0.9968
27	0.9965
28	0.9962
29	page3image48360 0.9959

*obtained from the CRC Handbook of Chemistry and Physics, 53rd Edition

Procedure

Materials: De-gassed sodas (regular and diet), unknown metal sample.

Equipment: 10-mL volumetric pipet and pipet bulb, 50-mL, two 50-mL Erlenmeyer flasks/cups , 100-mL graduated cylinder, two 100-mL beakers/cups, 50-mL beaker, electronic balance, capped glass vial, thermometer, water.

Safety

There are no major safety considerations in this lab.

Part A: The Precision of Volumetric Glassware

All glassware should be cleaned and rinsed with distilled water before use. Pay special attention to significant figures in your recorded measurements.

1. Using a medium 100-mL beaker, obtain approximately 50 mL of water.
2. Weigh a small **dry** 50-mL beaker on an electronic balance, and record this mass.
3. Graduated Cylinder: Fill the 100-mL graduated cylinder with approximately 10 mL of distilled water, and record the actual volume used. To how many significant figures should the volume be recorded? Transfer this distilled water into the small pre-weighed beaker, and measure the combined mass. When finished, empty the distilled water out of the small beaker, then carefully dry it.
4. Volumetric Pipet: Use the graduated pipet and bulb to transfer the *exactly* 10 mL of distilled water from the medium beaker into the small pre-weighed beaker. Record the volume used and the combined mass of the beaker and water. To how many significant figures should the volume be recorded? Again, empty the water out of the small beaker when finished,

carefully dry it.

5. For each of the three sets of data collected (using the graduated cylinder and volumetric pipet, calculate the density of water to the correct number of significant figures. Then share your two density values with all the students in your lab section, and record the results of the entire class on your report form. You should have at least ten density values for each type of volumetric glassware used in this part of the lab.

Part B: The Densities of Degassed Sodas

The following procedure should be performed for both a regular soda and a diet soda.

1. A graduated pipet with bulb will be used to dispense precise volumes of each of the two sodas used. The buret should first be rinsed with distilled water, and then rinsed with a small quantity of the relevant soda before filling it with that soda.
2. Using a medium 100-mL beaker, obtain approximately 40 mL of the *regular* soda from the stock bottle. Draw the soda into the graduated pipet. Read the pipet level.
3. Weigh a small **dry** 50-mL Erlenmeyer flask using an electronic balance, and record this mass. Then drain approximately 5 mL of the soda from the pipet into this flask, and record the new pipet reading. Finally, measure and record the combined mass of the flask and soda.
4. Next, add an *additional* 5 mL of soda from the pipet to the flask. Again, note the new pipet reading and the new combined mass of the flask and soda. Then repeat this step four more times. You will have obtained a total of six measurements and used approximately 30 mL of the soda when you are finished.
5. Now perform the same series of steps 1-4 using the diet soda. Use your second small (dry) Erlenmeyer flask for this soda.

Part C: The Density of an Unknown Metal

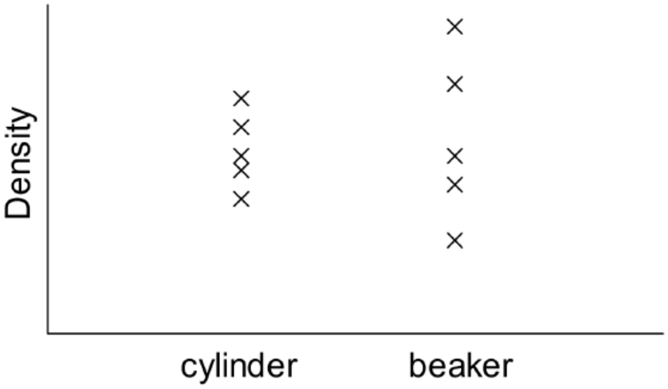
1. Obtain a capped glass vial and an unknown metal sample (coin). The cap on the vial should have a small hole pierced through it. This hole will allow air and excess water to be expelled from the vial. Record the ID Code of the metal on your report form.
2. All mass measurements are to be performed on an electronic balance, and should be recorded on your report form. First, weigh the empty, dry capped vial. Then add the entire sample of your unknown metal to the vial, and weigh it again (with cap).
3. Now fill the vial (with the metal still in it) to the brim with water. Gently tap the vial to ensure that no air remains between the metal pieces. Place the cap on firmly, pressing out any air or excess water. Wipe off any drops of water on the outside of the vial, and then weigh it.
4. Next, remove the metal from the vial and then fill it to the brim with distilled water only. Place the cap on firmly, wipe off excess water, and weigh.
5. Finally, using your thermometer, measure and record the temperature of the water in the vial. When finished, dry the metal sample and vial.

Pre-laboratory Assignment: The Densities of Solutions and Solids (Just for Calculations)

1. Five students decide to determine the density of a sample of liquid chloroform. Each student uses a graduated cylinder to obtain the chloroform volume and a balance to obtain the chloroform mass. Their density values are shown in the table below.

Student	Density (g/mL)

Student	Density (g/mL)
1	1.45
2	1.55
3	1.6
4	1.38
5	1.44



- Determine the average density and the standard deviation in this density dataset.

  Avg:

  Std Dev:

- The students then decide to determine the chloroform density using a beaker to measure volume. Their density results are displayed in the graph above, along with those obtained using the cylinder. Which glassware yielded the more precise measurements? Explain.
- In Part A of this lab, you will perform a similar analysis of density data obtained for distilled water. What three types of volumetric glassware will you be using?

2. A student decides to use the technique employed in Part B of this lab to determine the density of vodka. He first fills a buret with vodka to an initial volume of 1.52 mL. He then drains out some of the vodka into an empty beaker with a mass of 13.687 g. The final buret volume is 9.62 mL. If the combined mass of the beaker and vodka is 21.123 g, what is the the density of the vodka? Give your answer to the correct number of significant figures.
3. Another student decides to use the technique employed in Part C of this lab to determine the density of gold. She obtains the following data:

(A) Mass of Empty Vial	10.495 g
(B) Mass of Vial + Gold Sample	70.773 g
(C) Mass of Vial + Gold + Water	76.286 g
(D) Mass of Vial + Water	19.119 g
(E) Temperature of Water	25.0 °C
(F) Density of Water at above Temp.	0.9970 g/mL

Using her data, perform the calculations below. Show your work clearly, and pay attention to significant figures and units.

- What is the mass of the gold sample?
- What is the mass of water in the vial with the gold (see measurement C)? What is volume of water in this vial (hint: use the density of water)?

- What is the mass of water in the vial filled with water only (see measurement D)? What is the volume of water in this filled vial (hint: use the density of water)?
- The volume of the gold is equal to the difference in the two water volumes. What is the volume of the gold?
- Determine the density of the gold, in g/cm^3.

Lab Report: The Densities of Solutions and Solids

Part A: The Precision of Volumetric Glassware

Experimental Data

	Graduated Cylinder	Volumetric Pipet	Buret
Mass of Empty Beaker			
Mass of Beaker + Water			
Mass of Water only			
Actual Volume of Water used			
Density of Water			

Share your three density values with all the students in your lab section, and record the results of the entire class in the table below (pay attention to significant figures). You should acquire at least ten density values for each type of volumetric glassware used. Use an asterisks symbol (*) to denote which set of measurements you and your partner obtained.

Measurement	Class Density Values (in g/mL) obtained using:	
	Graduated Cylinder	Volumetric Pipet
1		
2		
3		
4		
5		
6		
7		
8		
9		
10		
11		
12		
13		
14		

Data Analysis

1. For each of the three density data sets, calculate the average density and the standard deviation in the data set.

	Average	Standard Deviation
Graduated Cylinder		
Volumetric Pipet		

2. Using Microsoft Excel© plot all three density data sets on a single Scatter Graph as “Density of Water versus Type of Volumetric Glassware used”. Attach this graph to your report.
3. Now closely examine all two data sets collected by your class.

- Circle any identified outliers in the three data sets on <u>both</u> your *plot* and in the *data table*. Then explain, using supporting calculations, how you identified these outliers.

Now re-calculate the average and standard deviation for each data set, *omitting the outliers*.

	Average	**Standard Deviation**
Graduated Cylinder		
Volumetric Pipet		

4. Based on your analyses of the two data sets without the outliers, rank the volumetric glassware used from least to most precise. Briefly explain how you determined this order.

Part B: The Densities of De-gassed Sodas

Experimental Data

Regular Soda

- Name of Soda:
- Mass of Empty Flask:
- Initial Buret Reading:

Data	**Combined Mass of Flask + Soda**		**Total Mass of Beverage**	**Total Volume of Beverage**	**Density of Soda**
1					
2					
3					
4					
5					
6					

- Average density of regular soda:

Diet Soda

- Name of Soda:
- Mass of Empty Flask:
- Initial Buret Reading:

Data	**Combined Mass of Flask + Soda**		**Total Mass of Beverage**	**Total Volume of Beverage**	**Density of Soda**
1					
2					
3					
4					
5					
6					

- Average density of diet soda:

Data Analysis

1. Which of the two sodas has the higher density? Suggest a possible reason for this result.

Part C: The Density of an Unknown Metal

Experimental Data

ID Code of Assigned Metal:

A. Mass of Empty Vial	
B. Mass of Vial + Metal	
C. Mass of Vial + Metal + Water	
D. Mass of Vial + Water	
E. Temperature of Water	
F. Density of Water at above Temperature	

Data Analysis

1. Using the data collected, calculate the density (in g/cm^3) of your unknown metal. Show your work clearly, and circle your final result.
2. Your metal could be any of the following metals: lead, aluminum, nickel or zinc. Look up the true densities of these metals and identify your unknown metal. Suggested resources include your text, websites (e.g., www.webelements.com, www.chemicool.com) and the CRC Handbook of Chemistry and Physics.

- Identity of unknown metal:
- True density of metal:
- Source of density data:

3. Calculate the percent error between your experimental density and the true metal density.

Questions

1. In Part A of this lab, you used your experimental results t rank the precision of two types of glassware from least to most precise. Did you expect these results? If yes, explain why. If no, explain why not, and give the order that you did expect.
2. Consider your density results for the diet soda in Part B of this lab. Clearly explain how they demonstrate that density is an intensive property.
3. Consider your results for Part C. Suppose your metal sample is hammered into the shape of a thin coin with a diameter of 8.62 cm. Determine the thickness (in mm) of this metal sheet. Use the measured metal mass and the true density of your metal for this calculation.

The experiment is from Chem 11 Experiments shared under a CC BY-NC license and was authored, remixed, and/or curated by Santa Monica College. Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead, has modified this to fit to a home school environment with less hazardous and cost-effective lab supplies.Manjusha acknowledges the creators of the YouTube video embedded on this page.

2.3: Structure and Isomerism of Hydrocarbons-Molecular Modelling-Simulation-Home

Learning Objectives

- To introduce chemical, structural, and molecular models of hydrocarbons
- To distinguish between the different types of bonds in alkanes, alkenes, and alkynes (single, double, and triple bonds)
- To demonstrate the geometry around carbon in single, double, and triple bonds (3D models)
- To illustrate structural isomerism

Introduction

Hydrocarbons are molecular compounds containing carbon and hydrogen only. Hydrocarbons are divided into aliphatic and aromatic. Aliphatic hydrocarbons contain chains (straight or branched) or rings. Aromatic hydrocarbons contain at least one benzene ring. Aliphatic hydrocarbons are further divided into alkanes, alkenes, and alkynes depending on the type of bonds present.

Alkanes contain only single bonds, whereas alkenes contain at least one carbon-carbon double bond. Alkynes contain at least one carbon-carbon triple bond. Alkanes are saturated with hydrogen atoms whereas alkenes and alkynes are unsaturated.

The geometry around carbon in carbon-carbon single bond is tetrahedral; carbon-carbon double bond is trigonal planar; and carbon-carbon triple bond is linear. The geometry around carbon is linear when two double bonds are present on a carbon. Molecules with the same chemical formula can bond in different patterns to form structural isomers. Isomers have the same formula but different structure, so are all different chemical with varying properties.

Theory

Learning Objectives

- Explain the importance of hydrocarbons and the reason for their diversity
- Name saturated and unsaturated hydrocarbons, and molecules derived from them
- Describe the reactions characteristic of saturated and unsaturated hydrocarbons
- Identify structural and geometric isomers of hydrocarbons

The largest database[1] of organic compounds lists about 10 million substances, which include compounds originating from living organisms and those synthesized by chemists. The number of potential organic compounds has been estimated[2] at 10^{60}—an astronomically high number. The existence of so many organic molecules is a consequence of the ability of carbon atoms to form up to four strong bonds to other carbon atoms, resulting in chains and rings of many different sizes, shapes, and complexities.

The simplest organic compounds contain only the elements carbon and hydrogen, and are called hydrocarbons. Even though they are composed of only two types of atoms, there is a wide variety of hydrocarbons because they may consist of varying lengths of chains, branched chains, and rings of carbon atoms, or combinations of these structures. In addition, hydrocarbons may differ in the types of carbon-carbon bonds present in their molecules. Many hydrocarbons are found in plants, animals, and their fossils; other hydrocarbons have been prepared in the laboratory. We use hydrocarbons every day, mainly as fuels, such as natural gas, acetylene, propane, butane, and the principal components of gasoline, diesel fuel, and heating oil. The familiar plastics polyethylene, polypropylene, and polystyrene are also hydrocarbons. We can distinguish several types of hydrocarbons by differences in the bonding between carbon atoms. This leads to differences in geometries and in the hybridization of the carbon orbitals.

Alkanes

Alkanes, or saturated hydrocarbons, contain only single covalent bonds between carbon atoms. Each of the carbon atoms in an alkane has sp^3 hybrid orbitals and is bonded to four other atoms, each of which is either carbon or hydrogen. The Lewis structures and models of methane, ethane, and pentane are illustrated in Figure $2.3.1$. Carbon chains are usually drawn as straight lines in Lewis structures, but one has to remember that Lewis structures are not intended to indicate the geometry of molecules. Notice that the carbon atoms in the structural models (the ball-and-stick and space-filling models) of the pentane molecule do not lie in a straight line. Because of the sp^3 hybridization, the bond angles in carbon chains are close to 109.5°, giving such chains in an alkane a zigzag shape.

The structures of alkanes and other organic molecules may also be represented in a less detailed manner by condensed structural formulas (or simply, *condensed formulas*). Instead of the usual format for chemical formulas in which each element symbol appears just once, a condensed formula is written to suggest the bonding in the molecule. These formulas have the appearance of a Lewis structure from which most or all of the bond symbols have been removed. Condensed structural formulas for ethane and pentane are shown at the bottom of Figure 2.3.1, and several additional examples are provided in the exercises at the end of this chapter.

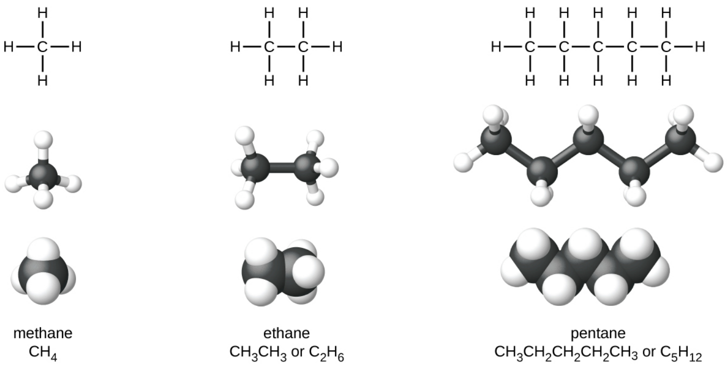


Figure 2.3.1: Pictured are the Lewis structures, ball-and-stick models, and space-filling models for molecules of methane, ethane, and pentane.

A common method used by organic chemists to simplify the drawings of larger molecules is to use a skeletal structure (also called a line-angle structure). In this type of structure, carbon atoms are not symbolized with a C, but represented by each end of a line or bend in a line. Hydrogen atoms are not drawn if they are attached to a carbon. Other atoms besides carbon and hydrogen are represented by their elemental symbols. Figure 2.3.2 shows three different ways to draw the same structure.

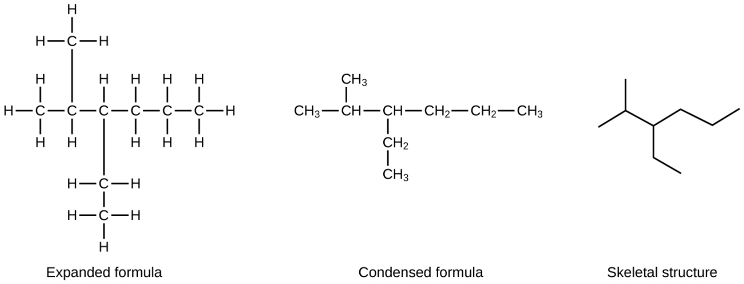
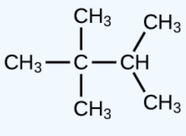


Figure 2.3.2: The same structure can be represented three different ways: an expanded formula, a condensed formula, and a skeletal structure.

Example 2.3.1

Drawing Skeletal Structures Draw the skeletal structures for these two molecules:



(a)

$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$

(b)

Solution

Each carbon atom is converted into the end of a line or the place where lines intersect. All hydrogen atoms attached to the carbon atoms are left out of the structure (although we still need to recognize they are there):

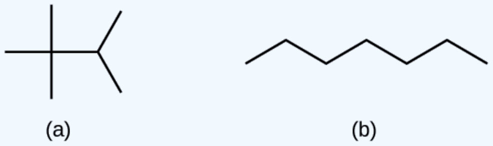


Figure a shows a branched skeleton structure that looks like a plus sign with line segments extending up and to the right and down and to the left of the rightmost point of the plus sign. Figure b appears in a zig zag pattern made with six line segments. The segments rise, fall, rise, fall, rise, and fall moving left to right across the figure.

? Exercise 2.3.1

Draw the skeletal structures for these two molecules:

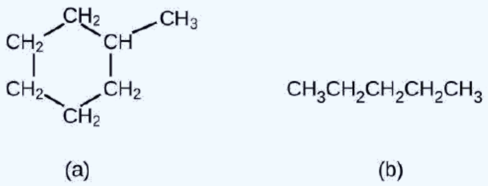


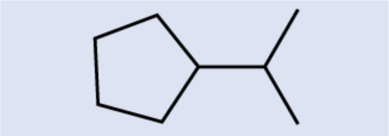
Figure a shows five C H subscript 2 groups and one C H group bonded in a hexagonal ring. A C H subscript 3 group appears above and to the right of the ring, bonded to the ring on the C H group appearing at the upper right portion of the ring. In b, a straight chain molecule composed of C H subscript 3 C H subscript 2 C H subscript 2 C H subscript 2 C H subscript 3 is shown.

Answer

In a, a hexagon with a vertex at the top is shown. The vertex just to the right has a line segment attached that extends up and to the right. In b, a zig zag pattern is shown in which line segments rise, fall, rise, fall, and rise moving left to right.

✓ Example 2.3.2

Interpreting Skeletal Structures Identify the chemical formula of the molecule represented here:

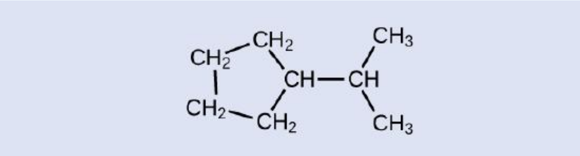


This figure shows a pentagon with a vertex pointing right, from which a line segment extends that has two line segments attached at its right end, one extending up and to the right, and the other extending down and to the right.

Solution

There are eight places where lines intersect or end, meaning that there are eight carbon atoms in the molecule. Since we know that carbon atoms tend to make four bonds, each carbon atom will have the number of hydrogen atoms that are required for four bonds. This compound contains 16 hydrogen atoms for a molecular formula of C_8H_{16}.

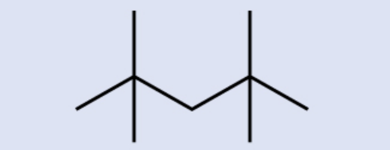
Location of the hydrogen atoms:



In this figure a ring composed of four C H subscript 2 groups and one C H group in a pentagonal shape is shown. From the C H group, which is at the right side of the pentagon, a C H is bonded. From this C H, a C H subscript 3 group is bonded above and to the right and a second is bonded below and to the right.

? Exercise 2.3.2

Identify the chemical formula of the molecule represented here:



A skeleton model is shown with a zig zag pattern that rises, falls, rises, and falls again left to right through the center of the molecule. From the two risen points, line segments extend both up and down, creating four branches.

Answer

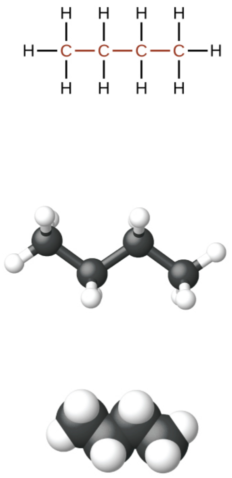
C_9H_{20}

All alkanes are composed of carbon and hydrogen atoms, and have similar bonds, structures, and formulas; noncyclic alkanes all have a formula of C_nH_{2n+2}. The number of carbon atoms present in an alkane has no limit. Greater numbers of atoms in the molecules will lead to stronger intermolecular attractions (dispersion forces) and correspondingly different physical properties of the molecules. Properties such as melting point and boiling point (Table 2.3.1) usually change smoothly and predictably as the number of carbon and hydrogen atoms in the molecules change.

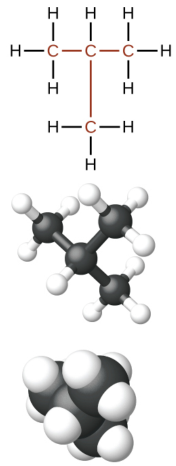
Table 2.3.1: Properties of Some Alkanes

Alkane	Molecular Formula	Melting Point (°C)	Boiling Point (°C)	Phase at STP[4]	Number of Structural Isomers
methane	CH_4	–182.5	–161.5	gas	1
ethane	C_2H_6	–183.3	–88.6	gas	1
propane	C_3H_8	–187.7	–42.1	gas	1
butane	C_4H_{10}	–138.3	–0.5	gas	2
pentane	C_5H_{12}	–129.7	36.1	liquid	3
hexane	C_6H_{14}	–95.3	68.7	liquid	5
heptane	C_7H_{16}	–90.6	98.4	liquid	9
octane	C_8H_{18}	–56.8	125.7	liquid	18
nonane	C_9H_{20}	–53.6	150.8	liquid	35
decane	$C_{10}H_{22}$	–29.7	174.0	liquid	75
tetradecane	$C_{14}H_{30}$	5.9	253.5	solid	1858
octadecane	$C_{18}H_{38}$	28.2	316.1	solid	60,523

Hydrocarbons with the same formula, including alkanes, can have different structures. For example, two alkanes have the formula C_4H_{10}: They are called *n*-butane and 2-methylpropane (or isobutane), and have the following Lewis structures:



n-butane



2-methylpropane

The compounds *n*-butane and 2-methylpropane are structural isomers (the term constitutional isomers is also commonly used). Constitutional isomers have the same molecular formula but different spatial arrangements of the atoms in their molecules. The *n*-butane molecule contains an *unbranched chain*, meaning that no carbon atom is bonded to more than two other carbon atoms. We use the term *normal*, or the prefix *n*, to refer to a chain of carbon atoms without branching. The compound 2–methylpropane has a branched chain (the carbon atom in the center of the Lewis structure is bonded to three other carbon atoms)

Identifying isomers from Lewis structures is not as easy as it looks. Lewis structures that look different may actually represent the same isomers. For example, the three structures in Figure 2.3.3 all represent the same molecule, *n*-butane, and hence are not different isomers. They are identical because each contains an unbranched chain of four carbon atoms.

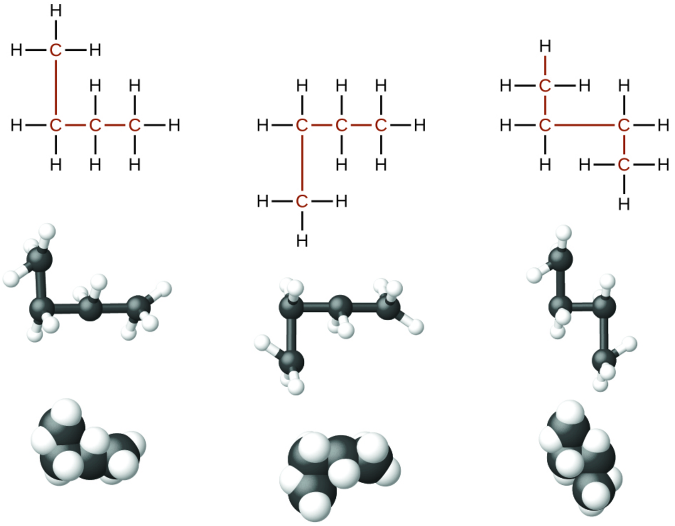


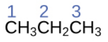
Figure 2.3.3: These three representations of the structure of n-butane are not isomers because they all contain the same arrangement of atoms and bonds.

The Basics of Organic Nomenclature: Naming Alkanes

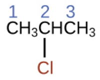
The International Union of Pure and Applied Chemistry (IUPAC) has devised a system of nomenclature that begins with the names of the alkanes and can be adjusted from there to account for more complicated structures. The nomenclature for alkanes is based on two rules:

1. To name an alkane, first identify the longest chain of carbon atoms in its structure. A two-carbon chain is called ethane; a three-carbon chain, propane; and a four-carbon chain, butane. Longer chains are named as follows: pentane (five-carbon chain), hexane (6), heptane (7), octane (8), nonane (9), and decane (10). These prefixes can be seen in the names of the alkanes described in Table 2.3.1.

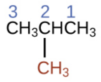
2. Add prefixes to the name of the longest chain to indicate the positions and names of substituents. Substituents are branches or functional groups that replace hydrogen atoms on a chain. The position of a substituent or branch is identified by the number of the carbon atom it is bonded to in the chain. We number the carbon atoms in the chain by counting from the end of the chain nearest the substituents. Multiple substituents are named individually and placed in alphabetical order at the front of the name.
3.


$\overset{1}{C}H_3\overset{2}{C}H_2\overset{3}{C}H_3$

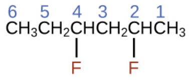
propane


$\overset{1}{C}H_3\underset{Cl}{\overset{2}{C}}H\overset{3}{C}H_3$

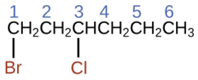
2-chloropropane


$\overset{3}{C}H_3\underset{CH_3}{\overset{2}{C}}H\overset{1}{C}H_3$

2-methylpropane


$\overset{6}{C}H_3\overset{5}{C}H_2\underset{F}{\overset{4}{C}}H\overset{3}{C}H_2\underset{F}{\overset{2}{C}}H\overset{1}{C}H_3$

2,4-difluorohexane

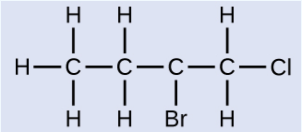

$\underset{Br}{\overset{1}{C}}H_2\overset{2}{C}H_2\underset{Cl}{\overset{3}{C}}H\overset{4}{C}H_2\overset{5}{C}H_2\overset{6}{C}H_3$

1-bromo-3-chlorohexane

When more than one substituent is present, either on the same carbon atom or on different carbon atoms, the substituents are listed alphabetically. Because the carbon atom numbering begins at the end closest to a substituent, the longest chain of carbon atoms is numbered in such a way as to produce the lowest number for the substituents. The ending *-o* replaces *-ide* at the end of the name of an electronegative substituent (in ionic compounds, the negatively charged ion ends with *-ide* like chloride; in organic compounds, such atoms are treated as substituents and the *-o* ending is used). The number of substituents of the same type is indicated by the prefixes *di-* (two), *tri-* (three), *tetra-* (four), and so on (for example, *difluoro-* indicates two fluoride substituents).

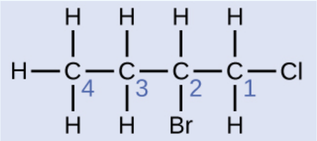
✓ Example 2.3.3: Naming Halogen-substituted Alkanes

Name the molecule whose structure is shown here:



This structure shows a C atom bonded to the H atoms and another C atom. This second C atom is bonded to two H atoms and another C atom. This third C atom is bonded to a B r atom and another C atom. This fourth C atom is bonded to two H atoms and a C l atom.

Solution

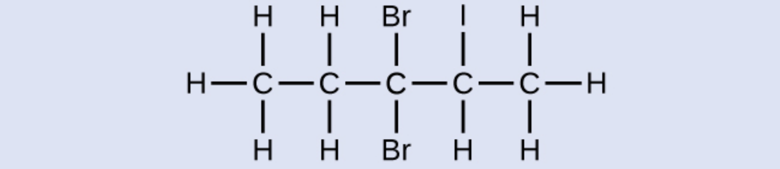


This structure shows a C atom bonded to the H atoms and another C atom. This second C atom is bonded to two H atoms and another C atom. This third C atom is bonded to an H atom, a B r atom, and another C atom. This fourth C atom is bonded to two H atoms and a C l atom. The C atoms are numbered 4, 3, 2, and 1 from left to right.

The four-carbon chain is numbered from the end with the chlorine atom. This puts the substituents on positions 1 and 2 (numbering from the other end would put the substituents on positions 3 and 4). Four carbon atoms means that the base name of this compound will be butane. The bromine at position 2 will be described by adding 2-bromo-; this will come at the beginning of the name, since bromo- comes before chloro- alphabetically. The chlorine at position 1 will be described by adding 1-chloro-, resulting in the name of the molecule being 2-bromo-1-chlorobutane.

? Exercise 2.3.3

Name the following molecule:

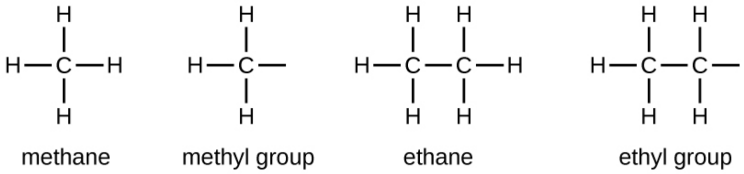


This figure shows a C atom bonded to three H atoms and another C atom. This second C atom is bonded to two H atoms and a third C atom. The third C atom is bonded to two B r atoms and a fourth C atom. This C atom is bonded to an H atom, and I atom, and a fifth C atom. This last C atom is bonded to three H atoms.

Answer

3,3-dibromo-2-iodopentane

We call a substituent that contains one less hydrogen than the corresponding alkane an alkyl group. The name of an alkyl group is obtained by dropping the suffix *-ane* of the alkane name and adding *-yl*:

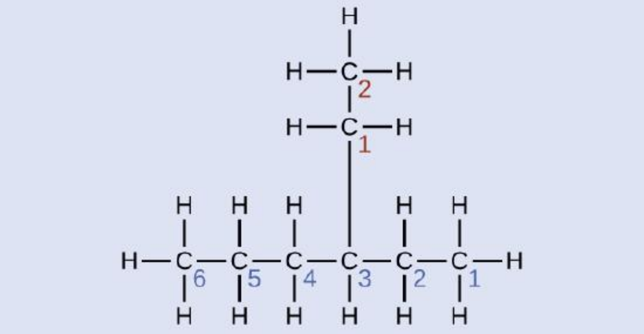


In this figure, methane is named and represented as C with four H atoms bonded above, below, to the left, and to the right of the C. The methyl group is shown, which appears like methane without the right most H. A dash remains at the location where the H was formerly bonded. Ethane is named and represented with two centrally bonded C atoms to which six H atoms are bonded; two above and below each of the two C atoms and to the left and right ends of the linked C atoms. The ethyl group appears as a similar structure with the right-most H atom removed. A dash remains at the location where the H atom was formerly bonded.

The open bonds in the methyl and ethyl groups indicate that these alkyl groups are bonded to another atom.

Example 2.3.4 Naming Substituted Alkanes

Name the molecule whose structure is shown here:



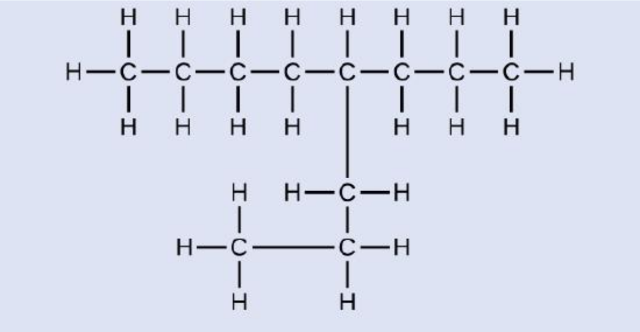
A chain of six carbon atoms, numbered 6, 5, 4, 3, 2, and 1 is shown. Bonded above carbon 3, a chain of two carbons is shown, numbered 1 and 2 moving upward. H atoms are present directly above, below, left and right of all carbon atoms in positions not already taken up in bonding to other carbon atoms.

Solution

The longest carbon chain runs horizontally across the page and contains six carbon atoms (this makes the base of the name hexane, but we will also need to incorporate the name of the branch). In this case, we want to number from right to left (as shown by the blue numbers) so the branch is connected to carbon 3 (imagine the numbers from left to right—this would put the branch on carbon 4, violating our rules). The branch attached to position 3 of our chain contains two carbon atoms (numbered in red)—so we take our name for two carbons *eth-* and attach *-yl* at the end to signify we are describing a branch. Putting all the pieces together, this molecule is 3-ethylhexane.

Exercise 2.3.4

Name the following molecule:



This figure shows a C atom bonded to three H atoms and another C atom. This C atom is bonded to two H atoms and third C atom. The third C atom is bonded to two H atoms and a fourth C atom. The fourth C atom is bonded to two H atoms and a fifth C atom. This C atom is bonded to an H atom, a sixth C atom in the chain, and another C atom which appears to branch off the chain. The C atom in the branch is bonded to two H atoms and another C atom. This C atom is bonded to two H atoms and another C atom. This third C atom appears to the left of the second and is bonded to three H atoms. The sixth C atom in the chain is bonded to two H atoms and a seventh C atom. The seventh C atom is bonded to two H atoms and an eighth C atom. The eighth C atom is bonded to three H atoms.

Answer

4-propyloctane

Some hydrocarbons can form more than one type of alkyl group when the hydrogen atoms that would be removed have different “environments” in the molecule. This diversity of possible alkyl groups can be identified in the following way: The four hydrogen atoms in a methane molecule are equivalent; they all have the same environment. They are equivalent because each is bonded to a carbon atom (the same carbon atom) that is bonded to three hydrogen atoms. (It may be easier to see the equivalency in the ball and stick models in Figure 2.3.3. Removal of any one of the four hydrogen atoms from methane forms a methyl group. Likewise, the six hydrogen atoms in ethane are equivalent and removing any one of these hydrogen atoms produces an ethyl group. Each of the six hydrogen atoms is bonded to a carbon atom that is bonded to two other hydrogen atoms and a carbon atom. However, in both propane and 2–methylpropane, there are hydrogen atoms in two different environments, distinguished by the adjacent atoms or groups of atoms:

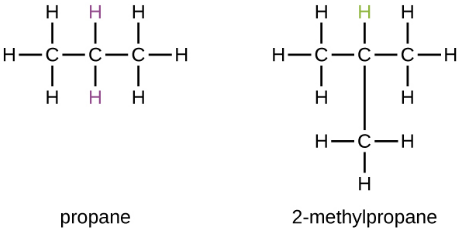
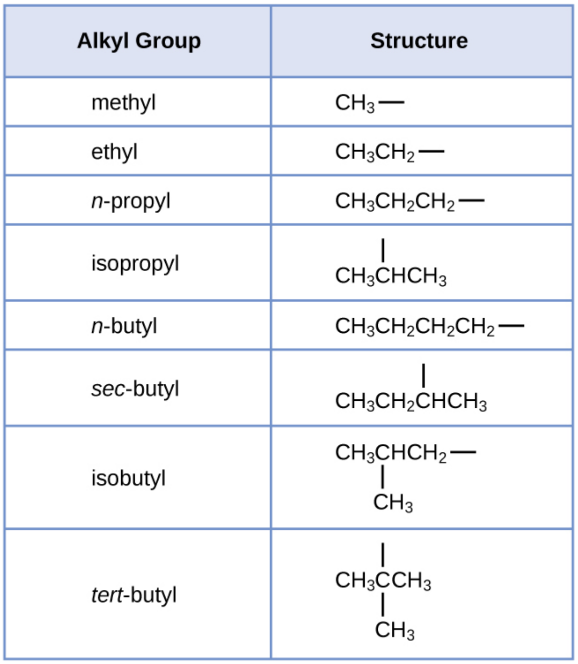


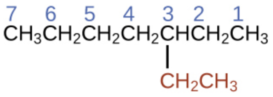
Figure 2.3.4.



Alkyl Group	Structure
methyl	$CH_3—$
ethyl	$CH_3CH_2—$
n-propyl	$CH_3CH_2CH_2—$
isopropyl	$CH_3\overset{\vert}{C}HCH_3$
n-butyl	$CH_3CH_2CH_2CH_2—$
sec-butyl	$CH_3CH_2\overset{\vert}{C}HCH_3$
isobutyl	$CH_3\underset{\underset{CH_3}{\vert}}{C}HCH_2—$
tert-butyl	$CH_3\overset{\vert}{\underset{\underset{CH_3}{\vert}}{C}}CH_3$

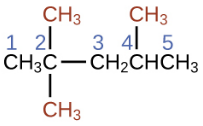
Figure 2.3.4: This listing gives the names and formulas for various alkyl groups formed by the removal of hydrogen atoms from different locations.

Note that alkyl groups do not exist as stable independent entities. They are always a part of some larger molecule. The location of an alkyl group on a hydrocarbon chain is indicated in the same way as any other substituent:



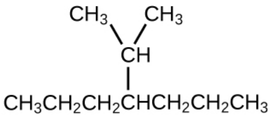
$$\overset{7}{C}H_3\overset{6}{C}H_2\overset{5}{C}H_2\overset{4}{C}H_2\underset{\underset{CH_2CH_3}{|}}{\overset{3}{C}}H\overset{2}{C}H_2\overset{1}{C}H_3$$

3-ethylheptane



$$\overset{1}{C}H_3\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{\overset{2}{C}}}—\overset{3}{C}H_2\overset{\overset{CH_3}{|}}{\overset{4}{C}}H\overset{5}{C}H_3$$

2,2,4-trimethylpentane



$$CH_3CH_2CH_2\overset{\overset{CH_3\ \ CH_3}{\diagdown\ \diagup}}{\overset{CH}{|}}{C}HCH_2CH_2CH_3$$

4-isopropylheptane

Alkanes are relatively stable molecules, but heat or light will activate reactions that involve the breaking of C–H or C–C single bonds. Combustion is one such reaction:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Alkanes burn in the presence of oxygen, a highly exothermic oxidation-reduction reaction that produces carbon dioxide and water. As a consequence, alkanes are excellent fuels. For example, methane, CH_4, is the principal component of natural gas. Butane, C_4H_{10}, used in camping stoves and lighters is an alkane. Gasoline is a liquid mixture of continuous- and branched-chain alkanes, each containing from five to nine carbon atoms, plus various additives to improve its performance as a fuel. Kerosene, diesel oil, and fuel oil are primarily mixtures of alkanes with higher molecular masses. The main source of these liquid alkane fuels is crude oil, a complex mixture that is separated by fractional distillation. Fractional distillation takes advantage of differences in the boiling points of the components of the mixture (Figure 2.3.5). You may recall that boiling point is a function of intermolecular interactions, which was discussed in the chapter on solutions and colloids.

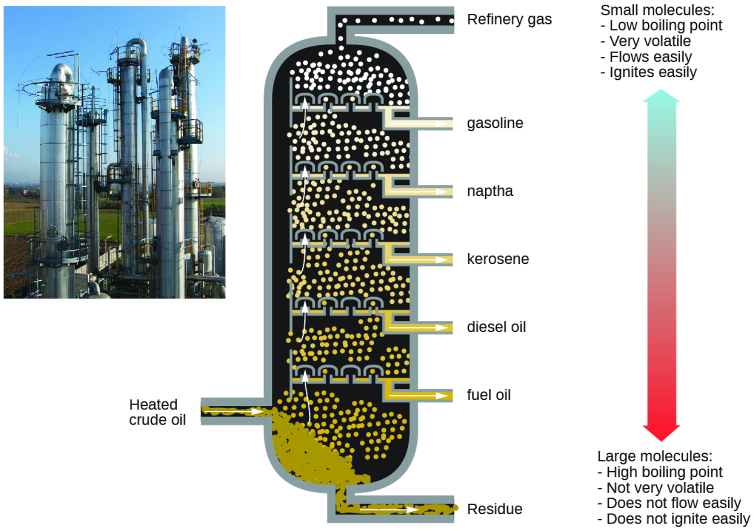
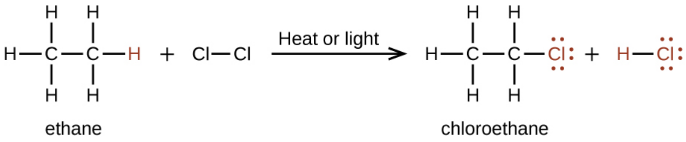


Figure 2.3.5:In a column for the fractional distillation of crude oil, oil heated to about 425 °C in the furnace vaporizes when it enters the base of the tower. The vapors rise through bubble caps in a series of trays in the tower. As the vapors gradually cool, fractions of higher, then of lower, boiling points condense to liquids and are drawn off. (credit left: modification of work by Luigi Chiesa)

In a substitution reaction, another typical reaction of alkanes, one or more of the alkane's hydrogen atoms is replaced with a different atom or group of atoms. No carbon-carbon bonds are broken in these reactions, and the hybridization of the carbon atoms does not change. For example, the reaction between ethane and molecular chlorine depicted here is a substitution reaction:



The C–Cl portion of the chloroethane molecule is an example of a functional group, the part or moiety of a molecule that imparts a specific chemical reactivity. The types of functional groups present in an organic molecule are major determinants of its chemical properties and are used as a means of classifying organic compounds as detailed in the remaining sections of this chapter.

Alkenes

Organic compounds that contain one or more double or triple bonds between carbon atoms are described as unsaturated. You have likely heard of unsaturated fats. These are complex organic molecules with long chains of carbon atoms, which contain at least one double bond between carbon atoms. Unsaturated hydrocarbon molecules that contain one or more double bonds are called alkenes. Carbon atoms linked by a double bond are bound together by two bonds, one σ bond and one π bond. Double and triple bonds give rise to a different geometry around the carbon atom that participates in them, leading to important differences in molecular shape and properties. The differing geometries are responsible for the different properties of unsaturated versus saturated fats.

Ethene, C_2H_4, is the simplest alkene. Each carbon atom in ethene, commonly called ethylene, has a trigonal planar structure. The second member of the series is propene (propylene) (Figure 2.3.6); the butene isomers follow in the series. Four carbon atoms in the chain of butene allows for the formation of isomers based on the position of the double bond, as well as a new form of isomerism.

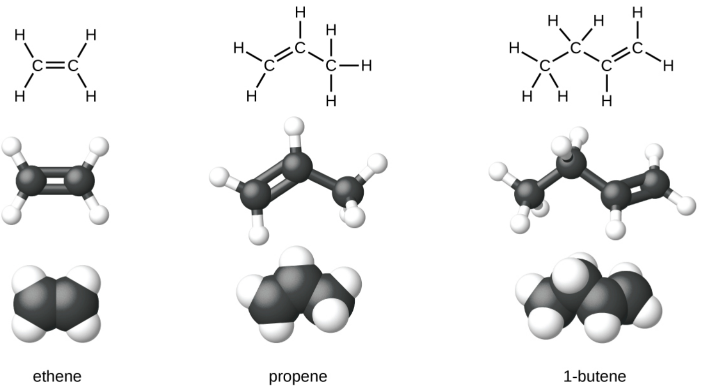


Figure $2.3.6$: Expanded structures, ball-and-stick structures, and space-filling models for the alkenes ethene, propene, and 1-butene are shown.

Ethylene (the common industrial name for ethene) is a basic raw material in the production of polyethylene and other important compounds. Over 135 million tons of ethylene were produced worldwide in 2010 for use in the polymer, petrochemical, and plastic industries. Ethylene is produced industrially in a process called cracking, in which the long hydrocarbon chains in a petroleum mixture are broken into smaller molecules.

Recycling Plastics

Polymers (from Greek words *poly* meaning "many" and *mer* meaning "parts") are large molecules made up of repeating units, referred to as monomers. Polymers can be natural (starch is a polymer of sugar residues and proteins are polymers of amino acids) or synthetic [like polyethylene, polyvinyl chloride (PVC), and polystyrene]. The variety of structures of polymers translates into a broad range of properties and uses that make them integral parts of our everyday lives. Adding functional groups to the structure of a polymer can result in significantly different properties (see the discussion about Kevlar later in this chapter).

An example of a polymerization reaction is shown in Figure $2.3.7$. The monomer ethylene (C_2H_4) is a gas at room temperature, but when polymerized, using a transition metal catalyst, it is transformed into a solid material made up of long chains of –CH_2– units called polyethylene. Polyethylene is a commodity plastic used primarily for packaging (bags and films).

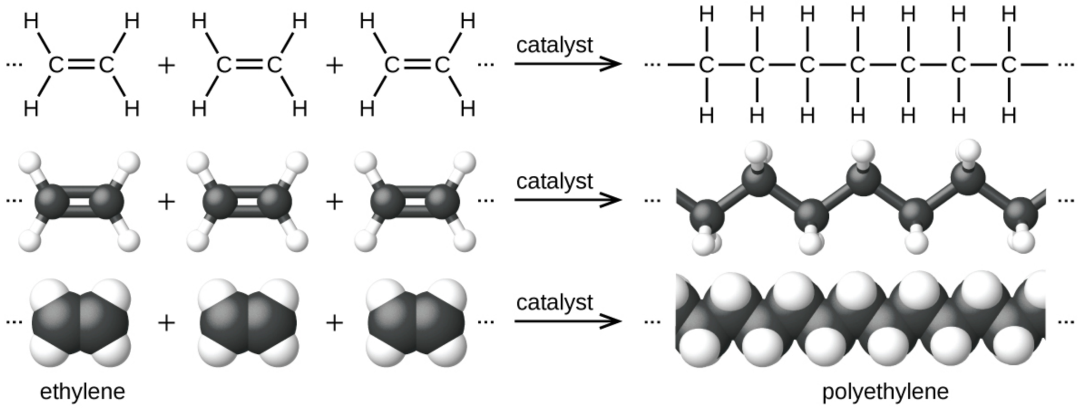
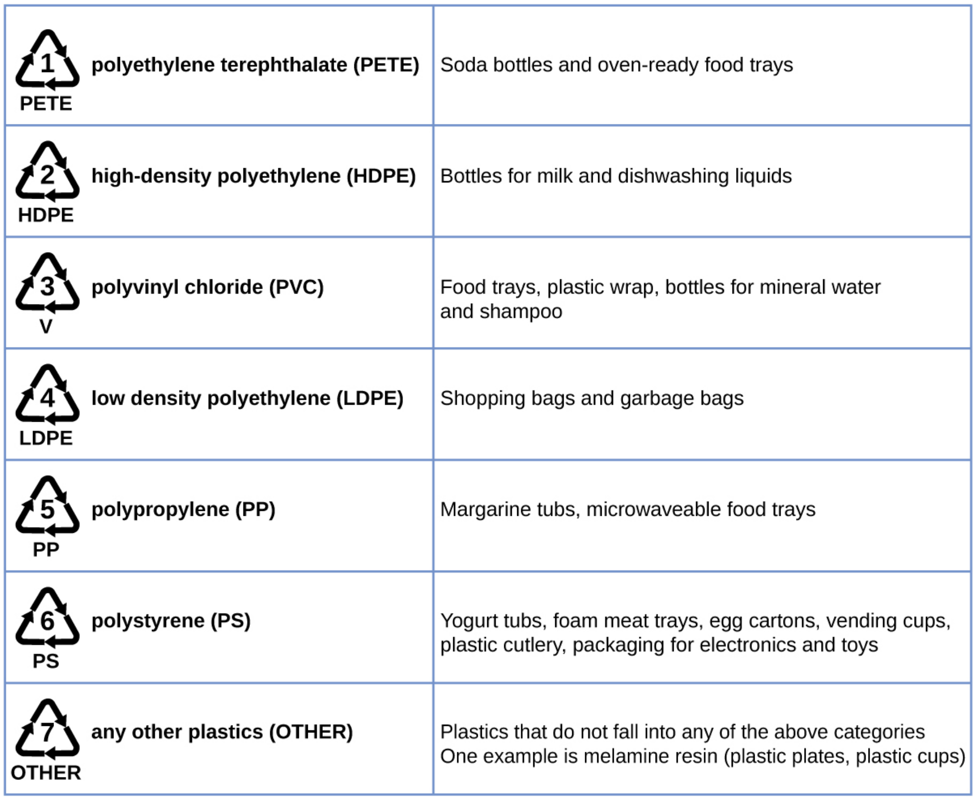


Figure $2.3.7$: The reaction for the polymerization of ethylene to polyethylene is shown.

This diagram has three rows, showing ethylene reacting to form polyethylene. In the first row, Lewis structural formulas show three molecules of ethylene being added together, which are each composed of two doubly bonded C atoms, each with two bonded H atoms. Ellipses, or three dots, are present before and after the molecule structures, which in turn are followed by an arrow pointing right. On the right side of the arrow, the ellipses or dots again appear to the left of a dash that connects to a chain of 7 C atoms, each with H atoms connected above and below. A dash appears at the end of the chain, which in turn is followed by ellipses or dots. The reaction diagram is repeated in the second row using ball-and-stick models for the structures. In these representations, single bonds are represented with sticks, double bonds are represented with two parallel sticks, and elements are represented with balls. Carbon atoms are black and hydrogen atoms are white in this image. In the third row, space-filling models are shown. In these models, atoms are enlarged spheres which are pushed together, without sticks to represent bonds.

Polyethylene is a member of one subset of synthetic polymers classified as plastics. Plastics are synthetic organic solids that can be molded; they are typically organic polymers with high molecular masses. Most of the monomers that go into common plastics (ethylene, propylene, vinyl chloride, styrene, and ethylene terephthalate) are derived from petrochemicals and are not very biodegradable, making them candidate materials for recycling. Recycling plastics helps minimize the need for using more of the petrochemical supplies and also minimizes the environmental damage caused by throwing away these nonbiodegradable materials.

Plastic recycling is the process of recovering waste, scrap, or used plastics, and reprocessing the material into useful products. For example, polyethylene terephthalate (soft drink bottles) can be melted down and used for plastic furniture, in carpets, or for other applications. Other plastics, like polyethylene (bags) and polypropylene (cups, plastic food containers), can be recycled or reprocessed to be used again. Many areas of the country have recycling programs that focus on one or more of the commodity plastics that have been assigned a recycling code (Figure 2.3.8). These operations have been in effect since the 1970s and have made the production of some plastics among the most efficient industrial operations today.

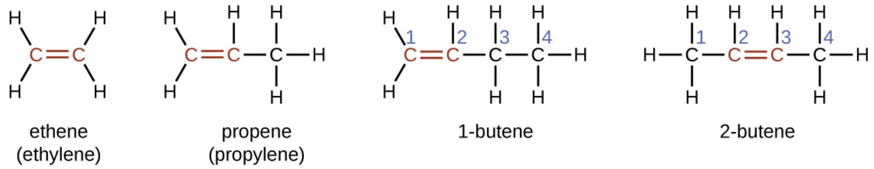


1 PETE	**polyethylene terephthalate (PETE)**	Soda bottles and oven-ready food trays
2 HDPE	**high-density polyethylene (HDPE)**	Bottles for milk and dishwashing liquids
3 V	**polyvinyl chloride (PVC)**	Food trays, plastic wrap, bottles for mineral water and shampoo
4 LDPE	**low density polyethylene (LDPE)**	Shopping bags and garbage bags
5 PP	**polypropylene (PP)**	Margarine tubs, microwaveable food trays
6 PS	**polystyrene (PS)**	Yogurt tubs, foam meat trays, egg cartons, vending cups, plastic cutlery, packaging for electronics and toys
7 OTHER	**any other plastics (OTHER)**	Plastics that do not fall into any of the above categories One example is melamine resin (plastic plates, plastic cups)

Figure 2.3.8: Each type of recyclable plastic is imprinted with a code for easy identification.

This table shows recycling symbols, names, and uses of various types of plastics. Symbols are shown with three arrows in a triangular shape surrounding a number. Number 1 is labeled P E T E. The related plastic, polyethylene terephthalate (P E T E), is used in soda bottles and oven-ready food trays. Number 2 is labeled H D P E. The related plastic is high-density polyethylene (H D P E), which is used in bottles for milk and dishwashing liquids. Number 3 is labeled V. The related plastic is polyvinyl chloride or (P V C). This plastic is used in food trays, plastic wrap, and bottles for mineral water and shampoo. Number 4 is labeled L D P E. This plastic is low density polyethylene (L D P E). It is used in shopping bags and garbage bags. Number 5 is labeled P P. The related plastic is polypropylene (P P). It is used in margarine tubs and microwaveable food trays. Number 6 is labeled P S. The related plastic is polystyrene (P S). It is used in yogurt tubs, foam meat trays, egg cartons, vending cups, plastic cutlery, and packaging for electronics and toys. Number 7 is labeled other for any other plastics. Items in this category include those plastic materials that do not fit any other category. Melamine used in plastic plates and cups is an example.

The name of an alkene is derived from the name of the alkane with the same number of carbon atoms. The presence of the double bond is signified by replacing the suffix *-ane* with the suffix *-ene*. The location of the double bond is identified by naming the smaller of the numbers of the carbon atoms participating in the double bond:



Four structural formulas and names are shown. The first shows two red C atoms connected by a red double bond illustrated with two parallel line segments. H atoms are bonded above and below to the left of the left-most C atom. Two more H atoms are similarly bonded to the right of the C atom on the right. Beneath this structure the name ethene and alternate name ethylene are shown. The second shows three C atoms bonded together with a red double bond between the red first and second C atoms moving left to right across the three-carbon chain. H atoms are bonded above and below to the left of the C atom to the left. A single H is bonded above the middle C atom. Three more H atoms are bonded above, below, and to the right of the third C atom. Beneath this structure the name propene and alternate name propylene is shown. The third shows four C atoms bonded together, numbered one through four moving left to right with a red double bond between the red first and second carbon in the chain. H atoms are bonded above and below to the left of the C atom to the left. A single H is bonded above the second C atom. H atoms are bonded above and below the third C atom. Three more H atoms are bonded above, below, and to the right of the fourth C atom. Beneath this structure the name 1 dash butene is shown. The fourth shows four C atoms bonded together, numbered one through four moving left to right with a red double bond between the red second and third C atoms in the chain. H atoms are bonded above, below, and to the left of the left-most C atom. A single H atom is bonded above the second C atom. A single H atom is bonded above the third C atom. Three more H atoms are bonded above, below, and to the right of the fourth C atom. Beneath this structure the name 2 dash butene is shown.

Isomers of Alkenes

Molecules of 1-butene and 2-butene are structural isomers; the arrangement of the atoms in these two molecules differs. As an example of arrangement differences, the first carbon atom in 1-butene is bonded to two hydrogen atoms; the first carbon atom in 2-butene is bonded to three hydrogen atoms.

The compound 2-butene and some other alkenes also form a second type of isomer called a geometric isomer. In a set of geometric isomers, the same types of atoms are attached to each other in the same order, but the geometries of the two molecules differ. Geometric isomers of alkenes differ in the orientation of the groups on either side of a $C=C$ bond.

Carbon atoms are free to rotate around a single bond but not around a double bond; a double bond is rigid. This makes it possible to have two isomers of 2-butene, one with both methyl groups on the same side of the double bond and one with the methyl groups on opposite sides. When structures of butene are drawn with 120° bond angles around the sp^2-hybridized carbon atoms participating in the double bond, the isomers are apparent. The 2-butene isomer in which the two methyl groups are on the same side is called a *cis*-isomer; the one in which the two methyl groups are on opposite sides is called a *trans*-isomer (Figure $2.3.9$). The different geometries produce different physical properties, such as boiling point, that may make separation of the isomers possible:

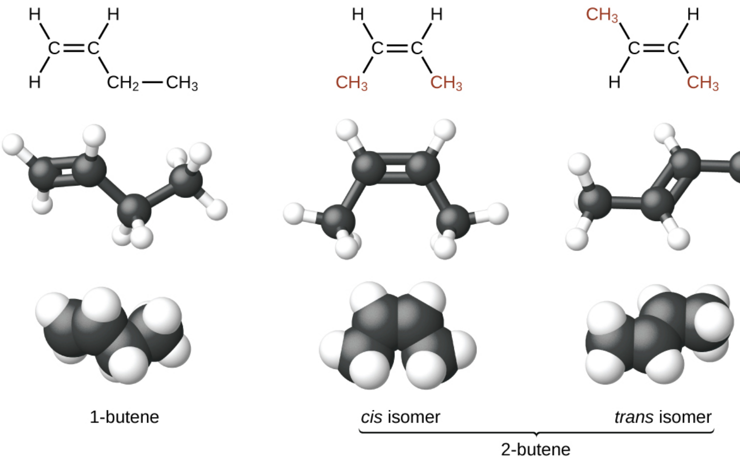
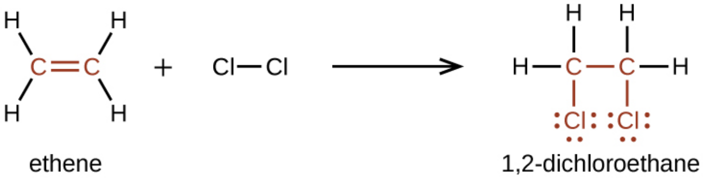


Figure 2.3.9: These molecular models show the structural and geometric isomers of butene.

The figure illustrates three ways to represent isomers of butene. In the first row of the figure, Lewis structural formulas show carbon and hydrogen element symbols and bonds between the atoms. The first structure in this row shows a C atom with a double bond to another C atom which is bonded down and to the right to C H subscript 2 which, in turn, is bonded to C H subscript 3. The first C atom, moving from left to right, has two H atoms bonded to it and the second C atom has one H atom bonded to it. The second structure in the row shows a C atom with a double bond to another C atom. The first C atom is bonded to an H atom up and to the left and C H subscript 3 down and to the left. The second C atom is bonded to an H atom up and to the right and C H subscript 3 down and to the right. Both C H subscript 3 structures appear in red. The third structure shows a C atom with a double bond to another C atom. The first C atom from the left is bonded up to a the left to C H subscript 3 which appears and red. It is also bonded down and to the left to an H atom. The second C atom is bonded up and to the right to an H atom and down and to the left to C H subscript 3 which appears in red. In the second row, ball-and-stick models for the structures are shown. In these representations, single bonds are represented with sticks, double bonds are represented with two parallel sticks, and elements are represented with balls. C atoms are black and H atoms are white in this image. In the third row, space-filling models are shown. In these models, atoms are enlarged and pushed together, without sticks to represent bonds. In the final row, names are provided. The molecule with the double bond between the first and second carbons is named 1 dash butene. The two molecules with the double bond between the second and third carbon atoms is called 2 dash butene. The first model, which has both C H subscript 3 groups beneath the double bond is called the cis isomer. The second which has the C H subscript 3 groups on opposite sides of the double bond is named the trans isomer.

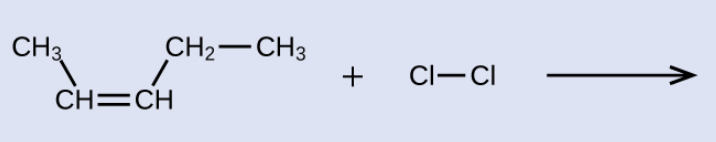
Alkenes are much more reactive than alkanes because the $C=C$ moiety is a reactive functional group. A π bond, being a weaker bond, is disrupted much more easily than a σ bond. Thus, alkenes undergo a characteristic reaction in which the π bond is broken and replaced by two σ bonds. This reaction is called an addition reaction. The hybridization of the carbon atoms in the double bond in an alkene changes from sp^2 to sp^3 during an addition reaction. For example, halogens add to the double bond in an alkene instead of replacing hydrogen, as occurs in an alkane:



This diagram illustrates the reaction of ethene and C l subscript 2 to form 1 comma 2 dash dichloroethane. In this reaction, the structural formula of ethane is shown. It has a double bond between the two C atoms with two H atoms bonded to each C atom plus C l bonded to C l. This is shown on to the left of an arrow. The two C atoms and the double bond between them are shown in red. To the right of the arrow, the 1 comma 2 dash dichloroethane molecule is shown. It has only single bonds and each C atom has a C l with three pairs of electron dots bonded beneath it. The C and C l atoms, single bond between them, and electron pairs are shown in red. Each C atom also has two H atoms bonded to it.

✓ Example 2.3.5: Alkene Reactivity and Naming

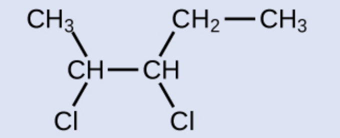
Provide the IUPAC names for the reactant and product of the halogenation reaction shown here:



The left side of a reaction and arrow are shown with an empty product side. On the left, C H subscript 3 is bonded down and to the right to C H which has a double bond to another C H. The second C H is bonded up and to the right to C H subscript 2 which is also bonded to C H subscript 3. A plus sign is shown with a C l atom bonded to a C l atom following it. This is also followed by a reaction arrow.

Solution

The reactant is a five-carbon chain that contains a carbon-carbon double bond, so the base name will be pentene. We begin counting at the end of the chain closest to the double bond—in this case, from the left—the double bond spans carbons 2 and 3, so the name becomes 2-pentene. Since there are two carbon-containing groups attached to the two carbon atoms in the double bond—and they are on the same side of the double bond—this molecule is the *cis*-isomer, making the name of the starting alkene *cis*-2-pentene. The product of the halogenation reaction will have two chlorine atoms attached to the carbon atoms that were a part of the carbon-carbon double bond:

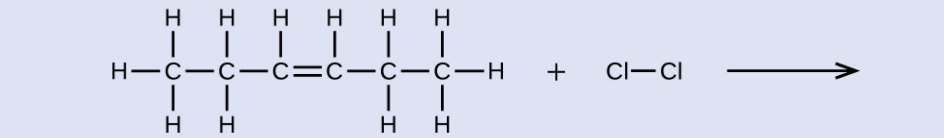


C H subscript 3 is bonded down and to the right to C H which is bonded down and to the left to C l. C H is also bonded to another C H which is bonded down and to the right to C l and up and to the right to C H subscript 2. C H subscript 2 is also bonded to C H subscript 3.

This molecule is now a substituted alkane and will be named as such. The base of the name will be pentane. We will count from the end that numbers the carbon atoms where the chlorine atoms are attached as 2 and 3, making the name of the product 2,3-dichloropentane.

? Exercise 2.3.5

Provide names for the reactant and product of the reaction shown:



This shows a C atom bonded to three H atoms and another C atom. This second C atom is bonded to two H atoms and a third C atom. This third C atom is bonded to one H atom and also forms a double bond with a fourth C atom. This fourth C atom is bonded to one H atom and a fifth C atom. This fifth C atom is bonded to two H atoms and a sixth C atom. This sixth C atom is bonded to three H atoms. There is a plus sign followed by a C l atom bonded to another C l atom. There is a reaction arrow. no products are shown.

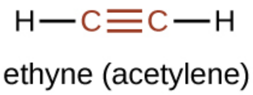
Answer

reactant: cis-3-hexene, product: 3,4-dichlorohexane

Alkynes

Hydrocarbon molecules with one or more triple bonds are called alkynes; they make up another series of unsaturated hydrocarbons. Two carbon atoms joined by a triple bond are bound together by one σ bond and two π bonds. The *sp*-hybridized carbons involved in the triple bond have bond angles of 180°, giving these types of bonds a linear, rod-like shape.

The simplest member of the alkyne series is ethyne, C_2H_2, commonly called acetylene. The Lewis structure for ethyne, a linear molecule, is:

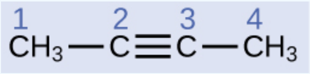


The structural formula and name for ethyne, also known as acetylene, are shown. In red, two C atoms are shown with a triple bond illustrated by three horizontal line segments between them. Shown in black at each end of the structure, a single H atom is bonded.

The IUPAC nomenclature for alkynes is similar to that for alkenes except that the suffix *-yne* is used to indicate a triple bond in the chain. For example, $\mathrm{CH_3CH_2C\equiv CH}$ is called 1-butyne.

Example 2.3.6: Structure of Alkynes

Describe the geometry and hybridization of the carbon atoms in the following molecule:



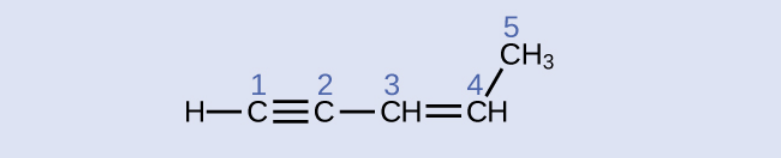
A structural formula is shown with C H subscript 3 bonded to a C atom which is triple bonded to another C atom which is bonded to C H subscript 3. Each C atom is labeled 1, 2, 3, and 4 from left to right.

Solution

Carbon atoms 1 and 4 have four single bonds and are thus tetrahedral with sp^3 hybridization. Carbon atoms 2 and 3 are involved in the triple bond, so they have linear geometries and would be classified as *sp* hybrids.

Exercise 2.3.6

Identify the hybridization and bond angles at the carbon atoms in the molecule shown:

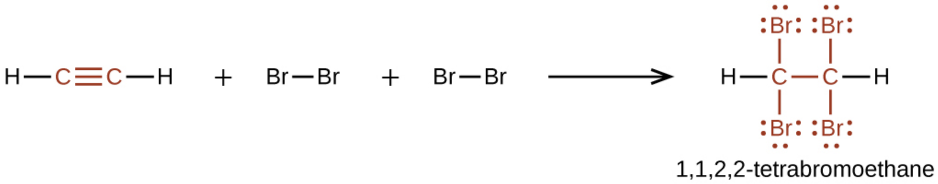


A structural formula is shown with an H atom bonded to a C atom. The C atom has a triple bond with another C atom which is also bonded to C H. The C H has a double bond with another C H which is also bonded up and to the right to C H subscript 3. Each C atom is labeled 1, 2, 3, 4, or 5 from left to right.

Answer

carbon 1: *sp*, 180°; carbon 2: *sp*, 180°; carbon 3: sp^2, 120°; carbon 4: sp^2, 120°; carbon 5: sp^3, 109.5°

Chemically, the alkynes are similar to the alkenes. Since the $\mathrm{C\equiv C}$ functional group has two π bonds, alkynes typically react even more readily, and react with twice as much reagent in addition reactions. The reaction of acetylene with bromine is a typical example:



This diagram illustrates the reaction of ethyne and two molecules of B r subscript 2 to form 1 comma 1 comma 2 comma 2 dash tetrabromoethane. In this reaction, the structural formula of ethyne, an H atom bonded to a red C atom with a red triple bond to another red C atom bonded to a black H atom, plus B r bonded to B r plus B r bonded to B r is shown to the left of an arrow. On the right, the form 1 comma 1 comma 2 comma 2 dash tetrabromoethane molecule is shown. It has an H atom bonded to a C atom which is bonded to another C atom which is bonded to an H atom. Each C atom is bonded above and below to a B r atom. Each B r atom has three pairs of electron dots. The C and B r atoms, single bond between them, and electron pairs are shown in red.

Acetylene and the other alkynes also burn readily. An acetylene torch takes advantage of the high heat of combustion for acetylene.

Aromatic Hydrocarbons

Benzene, C_6H_6, is the simplest member of a large family of hydrocarbons, called aromatic hydrocarbons. These compounds contain ring structures and exhibit bonding that must be described using the resonance hybrid concept of valence bond theory or the delocalization concept of molecular orbital theory. (To review these concepts, refer to the earlier chapters on chemical bonding). The resonance structures for benzene, C_6H_6, are:

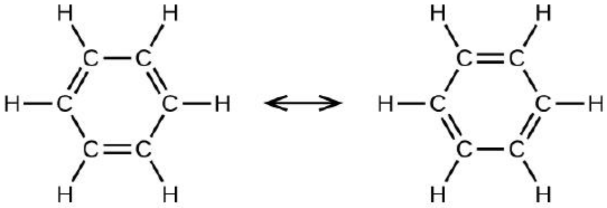


Figure 2.3.10.

This structural formula shows a six carbon hydrocarbon ring. On the left side there are six C atoms. The C atom on top and to the left forms a single bond to the C atom on the top and to the right. The C atom has a double bond to another C atom which has a single bond to a C atom. That C atom has a double bond to another C atom which has a single bond to a C atom. That C atom forms a double bond with another C atom. Each C atom has a single bond to an H atom. There is a double sided arrow and the structure on the right is almost identical to the structure on the left. The structure on the right shows double bonds where the structure on the left showed single bonds. The structure on the right shows single bonds where the stucture on the left showed double bonds.

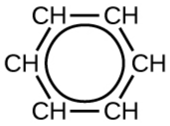
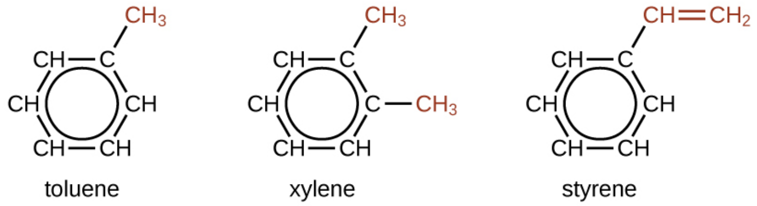


Figure 2.3.10: This condensed formula shows the unique bonding structure of benzene.

A six carbon hydrocarbon ring structural formula is shown. Each C atom is bonded to only one H atom. A circle is at the center of the ring.

There are many derivatives of benzene. The hydrogen atoms can be replaced by many different substituents. Aromatic compounds more readily undergo substitution reactions than addition reactions; replacement of one of the hydrogen atoms with another substituent will leave the delocalized double bonds intact. The following are typical examples of substituted benzene derivatives:

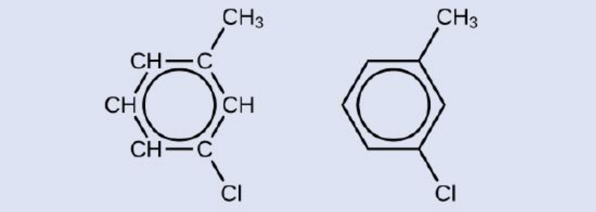


Three structural formulas are shown. The first is labeled toluene. This molecule has a six carbon hydrocarbon ring in which five of the C atoms are each bonded to only one H atom. At the upper right of the ring, the C atom that does not have a bonded H atom has a red C H subscript 3 group attached. A circle is at the center of the ring. The second is labeled xylene. This molecule has a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. At the upper right and right of the ring, the two C atoms that do not have bonded H atoms have C H subscript 3 groups attached. These C H subscript 3 groups appear in red. A circle is at the center of the ring. The third is labeled styrene. This molecule has a six carbon hydrocarbon ring in which five of the carbon atoms are each bonded to only one H atom. At the upper right of the ring, the carbon that does not have a bonded H atom has a red C H double bond C H subscript 2 group attached. A circle is at the center of the ring.

Toluene and xylene are important solvents and raw materials in the chemical industry. Styrene is used to produce the polymer polystyrene.

✓ Example 2.3.7: Structure of Aromatic Hydrocarbons

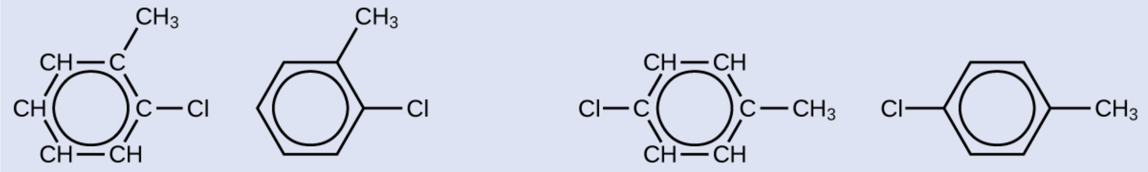
One possible isomer created by a substitution reaction that replaces a hydrogen atom attached to the aromatic ring of toluene with a chlorine atom is shown here. Draw two other possible isomers in which the chlorine atom replaces a different hydrogen atom attached to the aromatic ring:



Two structural formulas are shown. The first has a six carbon hydrocarbon ring in which four of the carbon atoms are each bonded to only one H atom. At the upper right of the ring, the carbon that does not have a bonded H atom has a C H subscript 3 group attached. The C to the lower right has a C l atom attached. A circle is at the center of the ring. The second molecule has a hexagon with a circle inside. From a vertex of the hexagon at the upper right a C H subscript 3 group is attached. From the vertex at the lower right, a C l atom is attached.

Solution

Since the six-carbon ring with alternating double bonds is necessary for the molecule to be classified as aromatic, appropriate isomers can be produced only by changing the positions of the chloro-substituent relative to the methyl-substituent:

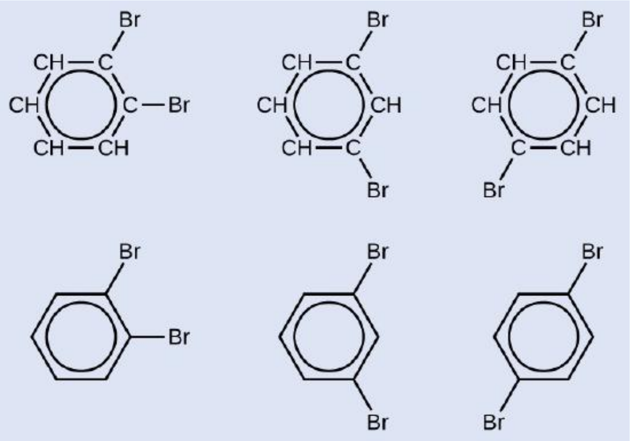


Two pairs of structural formulas are shown. The first has a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. At the upper right of the ring, the C atom that does not have a bonded H atom has a C H subscript 3 group attached. The C atom to the right has a C l atom attached. A circle is at the center of the ring. The second molecule in the first pair has a hexagon with a circle inside. From a vertex of the hexagon at the upper right a C H subscript 3 group is attached. From the vertex at the right, a C l atom is attached. The second pair first shows a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. A C l atom is attached to the left-most C atom and a C H subscript 3 group is attached to the right-most C atom. A circle is at the center of the ring. The second molecule in the pair has a hexagon with a circle inside. A C H subscript 3 group is attached to a vertex on the right side of the hexagon and to a vertex on the left side, a C l atom is bonded.

Exercise 2.3.7

Draw three isomers of a six-membered aromatic ring compound substituted with two bromines.

Answer



Three pairs of structural formulas are shown. The first has a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. At the upper right and right of the ring, the two C atoms that do not have bonded H atoms have one B r atom bonded each. A circle is at the center of the ring. Beneath this structure, a similar structure is shown which has a hexagon with a circle inside. From vertices of the hexagon at the upper right and right single B r atoms are attached. The second has a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. At the upper right and lower right of the ring, the two C atoms that do not have bonded H atoms have a single B r atom bonded each. A circle is at the center of the ring. Beneath this structure, a similar structure is shown which has a hexagon with a circle inside. From vertices of the hexagon at the upper right and lower right single B r atoms are attached. The third has a six carbon hydrocarbon ring in which four of the C atoms are each bonded to only one H atom. At the upper right and lower left of the ring, the two C atoms that do not have bonded H atoms have B r atoms bonded. A circle is at the center of the ring. Beneath this structure, a similar structure is shown which has a hexagon with a circle inside. From vertices of the hexagon at the upper right and lower left, single B r atoms are attached.

Summary

Strong, stable bonds between carbon atoms produce complex molecules containing chains, branches, and rings. The chemistry of these compounds is called organic chemistry. Hydrocarbons are organic compounds composed of only carbon and hydrogen. The alkanes are saturated hydrocarbons—that is, hydrocarbons that contain only single bonds. Alkenes contain one or more carbon-carbon double bonds. Alkynes contain one or more carbon-carbon triple bonds. Aromatic hydrocarbons contain ring structures with delocalized π electron systems.

Footnotes

1. This is the Beilstein database, now available through the Reaxys site (www.elsevier.com/online-tools/reaxys).
2. Peplow, Mark. “Organic Synthesis: The Robo-Chemist,” *Nature* 512 (2014): 20–2.
3. Physical properties for C_4H_{10} and heavier molecules are those of the *normal isomer*, *n*-butane, *n*-pentane, etc.
4. STP indicates a temperature of 0 °C and a pressure of 1 atm.

Glossary

addition reaction

reaction in which a double carbon-carbon bond forms a single carbon-carbon bond by the addition of a reactant. Typical reaction for an alkene.

alkane

molecule consisting of only carbon and hydrogen atoms connected by single (σ) bonds

alkene

molecule consisting of carbon and hydrogen containing at least one carbon-carbon double bond

alkyl group

substituent, consisting of an alkane missing one hydrogen atom, attached to a larger structure

alkyne

molecule consisting of carbon and hydrogen containing at least one carbon-carbon triple bond

aromatic hydrocarbon

cyclic molecule consisting of carbon and hydrogen with delocalized alternating carbon-carbon single and double bonds, resulting in enhanced stability

functional group

part of an organic molecule that imparts a specific chemical reactivity to the molecule

organic compound

natural or synthetic compound that contains carbon

saturated hydrocarbon

molecule containing carbon and hydrogen that has only single bonds between carbon atoms

skeletal structure

shorthand method of drawing organic molecules in which carbon atoms are represented by the ends of lines and bends in between lines, and hydrogen atoms attached to the carbon atoms are not shown (but are understood to be present by the context of the structure)

substituent

branch or functional group that replaces hydrogen atoms in a larger hydrocarbon chain

substitution reaction

reaction in which one atom replaces another in a molecule

Procedure: Modeling-Structure, Geometry, Naming and Isomerism of Hydrocarbons

Materials required

Acces the following simulation from Colorado State University, Boulder, CO using the link

https://phet.colorado.edu/sims/html/...lecule_en.html

Click on the play button.

- Click on the “playground” option.
- Watch the following YouTube video from min 2.55 to see how the simulation works.

- Based on the structures drawn in the report sheet of your lab experiment bring/click drag and attach atoms from the basket in the lower part of the simulation page. The carbon atoms should be attached in the same pattern as in the structure given. Do not worry about the single or double bonds at this time. If you need more carbon atoms click on the yellow arrow to get the next kit. The same way bring H atoms. If a Carbon has three H, attach 3 to that C, if one C has only two H in the structure, attach two.

- Once the structure is made with the correct number of atoms in the correct place, a name will be displayed with a green 3D button on top of the structure. Click on that to see the 3D-model pop-up. You can click and rotate the model around. Click on the ball-and-stick-circle on the pop-up. Draw the images on your lab report, may be on a separate sheet or the Paint/Draw app. Attach with the report or upload to the D2L folder. Record the name also.
- If you want to cut a bond, place the cursor between the bonded atoms in original model (before clicking 3D-un click 3D) and a red scissors icon will pop up. Click to separate the atoms.

Report Sheet: Modeling-Structure, Geometry, Naming and Isomerism of Hydrocarbons

Data Analysis

Chemical Formula	Alkane/Alkene/Alkyne/Aromatic	Types of bonds present, single, double etc.	Draw the structural isomers (Two for each formula)	Names of the isomers	Draw the ball-and-stick model shown in the 3D option on a separate paper
C_4H_{10}			A B	A B	
C_4H_6			A B	A B	
C_4H_8			A B		
C_6H_6			Only one structure, A	A	

Questions

1. Why cannot there be an alkyne or alkyne named Methene and Methyne?

2. Identify the unsaturated hydrocarbons on the above report sheet.

3. Which of the above molecules have carbon atoms attached in a flat linear geometry?

4. What kind of bonds are present between the six carbon atoms in C_6H_6 ?

5. How are isomers of C_4H_{10} different?

Contributors and Attributions

The experiment is developed by Manjusha Saraswathiamma, Minnesota State Community College, Moorhead for doing experiments at a homeschool setting with less hazardous and cost-effective lab supplies. The theory part of the experiment is from section **20.1: Hydrocarbons** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e. Manjusha would like to acknowledge the creators of the YouTube video embedded on this page. The simulation link is for Build-a-Molecule from PHET, Colorado, State University, Boulder, Colorado. PHET simulations are licenced CC-BY-4.0.Check the licensing on the following link, https://phet.colorado.edu/en/licensing

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2.4: Synthetic Polymers and Plastics (Experiment)-Home Version

Objectives

The objectives of this laboratory are to:

- Understand properties and uses of synthetic thermoplastics
- Compare the physical properties of “Big Six” plastics
- Identify everyday plastics by their physical properties
- Draw basic structures of polymers when given monomer structure

The word “polymer” means “many units”. A polymer can be made up of many repeating units, which are small monomer molecules that have been covalently bonded. Figure 1 (from Chemistry in Context) shows a single monomer, and a polymer made of identical monomers linked together. A polymer can contain hundreds of monomers, totaling thousands of atoms.

Examples of naturally-occurring polymers are silk, cotton, wood, cotton, starch, natural rubber, skin, hair and DNA. In the early 1900s, chemists began to replicate natural polymers, and create synthetic polymers, beginning with nylon which mimics silk in its strength and flexibility.

Plastic is a type of synthetic polymer. Currently, more than 60,000 plastics are manufactured for industrial and commercial purposes. Roughly 75% of the plastics used in this country can be categorized as one of six types, or “The Big Six”. These polymers are listed in the table below.

No.	Name	Abbreviation page1image20088	Uses
1	polyethylene terephthalate	PET	clear bottles and containers, fleece, carpet
2	high-density polyethylene	HDPE page1image30344	opaque bottles and containers, buckets, crates
3	polyvinyl chloride	PVC	rigid from: pipes & credit cards; soft form: tubing
4	low-density polyethylene	LDPE	bags, films, sheets, bubble wrap, toys
5	polypropylene	PP	bottle caps, yogurt containers, furniture
6	polystyrene	PS	expandable form: styrofoam; crystal form: CD cases

page1image52160

These six polymers are thermoplastic: they can be melted and reshaped, or recycled. The numbers are used to ease identification of the plastics, so that they can be separated for recycling. The symbol used on plastic packaging materials is three arrows in a triangle shape, with the number of plastic in the middle. This symbol makes recycling easier by making the identification of plastics easier.

The Big Six thermoplastics have these general attributes:

- recyclable
- insoluble in water
- resistant to most chemicals
- lightweight yet strong
- can be shaped
- can be colored with pigments
- usually made from petroleum
- used to make items that have no alternatives from other materials

The most common of the Big Six plastics is high-density polyethylene (HDPE). It is composed of repeating units of the monomer ethylene

— (H2C-CH2)n—

Monomers are linked together in an addition polymerization reaction. Each new monomer adds to one end with a covalent bond; the total number of monomers in the polymer is represented by the subscript, n. The resulting polymer is a chain of monomers linked together. The figure at right depicts one part of the polymer chain. How many monomers are present?

In this experiment, you will be qualitatively analyzing plastic polymers for physical characteristics of opacity, flexibility, durability, and breakability. You'll also analyze the density of each plastic by checking whether pellet samples float or sink in three liquids of different densities.

In this lab, you will make a polymer bouncy ball using a chemical reaction between borax and glue. Glue contains the polymer polyvinyl acetate, which cross-links to itself when reacted with borax (diagram below). After cross-linking, the glue is no longer fluid, but more solid. Adding cornstarch helps to bind the molecules together so that they hold their shape.

Procedure

Safety
No materials used in this experiment should be ingested.
Personal Protective Equipment (PPE) required: safety goggles, closed-toe shoes

Materials and Equipment

Samples of Big Six plastics (marked with recycling symbols) from household items (see the above list), small pieces of Big Six plastics from household items, distilled water, 10% NaCl solution, 3 small test tubes, glass stirring rod, wash bottle with distilled water, 3 small beakers, 3 plastic spoons, paper cup, Elmer's glue, borax, cornstarch, ruler

Part A: Physical Characteristics of Plastic Polymers

Samples of different plastics are available at the front bench. Identify the samples by looking for the number/recycling symbol. Use these samples to analyze physical characteristics of each type of plastic: recyclability, opacity, durability/hardness, and flexibility.

Part B: Density Tests of Big Six Plastics

1. Two solutions of different densities will be used:
    - Solution B = distilled water, density = 1.0 g/cm3
    - Solution C = 10% NaCl, density = 1.08 g/cm3
2. Obtain and label two small test tubes: Solution A and B . Add about 3 mL (two full dropper squirts) to each test tube.
3. Place one piece of each plastic into each of the two test tubes. Push each piece under the liquid surface with a glass stirring rod. Surface tensions will cause all of the plastic to float until each is "wetted" and submerged using the rod.

    Record whether the sample sinks rapidly, sinks slowly, floats on the surface, or floats below the surface (but doesn't sink to bottom).

    If the sample floats, it has a density lower than that of the solution. This may be relative to another sample that floats. If the sample sinks, it has a density greater than that of the liquid. The sample may also sink rapidly or slowly relative to other samples.

4. Test each of the six plastic types accordingly, to complete the table in the lab report.

Part C: Polymer Bouncy Balls

1. Obtain a paper cup with approximately 100 mL of Elmer's glue (sample cup will be marked to 100 mL), ruler and 3 plastic spoons. Obtain a wash bottle with distilled water.
2. Obtain 3 small beakers or containers, stirring rod, small graduated cylinder

Making Polymer Bouncy Ball #1:

1. In a glass beaker, add:
    - 3 level spoonfuls of glue
    - 5 mL distilled water

   - 1 level spoonful of borax powder
2. DO NOT STIR. Allow the ingredients to interact for 10-15 seconds. Then use stirring rod to mix. Once the mixture becomes impossible to stir, take it out of the beaker and mold the ball with your hands. The ball will start out sticky and messy, but will solidify as it is kneaded.
3. Record physical observations about the ball in the table: is the ball stretchy? goopy? slimy?
4. Use the ruler and hold the ball at a height of 30 cm (=12 in) above the bench. Drop the ball and record how high it bounces.

Making Polymer Bouncy Ball #2:

1. In a glass beaker, add:
   - 3 level spoonfuls of glue
   - 5 mL distilled water
   - 1 level spoonful of cornstarch
   - 1 level spoonful of borax
2. Repeat steps 2-4 from previous.

Making Polymer Bouncy Ball #3:

1. In a glass beaker, add:
   - 3 level spoonfuls of glue
   - 1 level spoonful of cornstarch
   - 1 level spoonful of borax
2. Repeat steps 2-4 from previous.
3. You and your lab partner may take home these bouncy balls, as all of the materials are non- toxic. However, please remember that they’re not edible!

Report

Synthetic Polymers and Plastics

Part A: Physical Characteristics

Find or choose one type of each of the following plastic polymers, and report the following characteristics:

page1image3056 page1image3216 page1image3376 page1image3536 page1image3696 page1image3856 page1image4016 page1image4176

Plastic number	Short Name (HDPE, LDPE, etc)	Clear (yes or no)	Opaque (yes or no)	Flexibility (can be bent?)	Durability (hard or soft)	Breakability (can be cracked?)	Recyclable (yes or no)
1							
2							
3							
4		page1image39816	page1image40600 page1image40920	page1image41360	page1image42064		page1image43224
5							
6		page1image49184	page1image50152 page1image50472	page1image51560	page1image52696		page1image54312
7		page1image57816	page1image58600 page1image58920	page1image59528	page1image60136		page1image61032

Part B: Density Tests

Report for plastic samples in each liquid: sinks rapidly, sinks slowly, floats on top, floats below surface

Plastic number	page1image67136 Water density = 1.0 g/cm3	page1image68464 10% NaCl solution density = 1.08 g/cm3
1	page1image73184	page1image74128
2	page1image77320	page1image77928
3		
4	page1image83144	page1image84088
5		
6		

Relative Plastic Densities:

Less than 1.0 g/cm3	Less than 1.08 g/cm3	More than 1.08 g/cm3

Ranking of densities:
(lowest) _______ _______ _______ _______ _______ _______ (highest)

Part C: Polymer Bouncy Balls

Polymer Ball composition	Approximate height bounced	Physical characteristics
Ball #1:		
Ball #2		
Ball #3		

Questions

1. Which of the Big Six plastics was the most flexible?
2. Which of the Big Six plastics would be the best material for each of the following examples? Use short names to identify each plastic (e.g. HDPE).

   a replacement for a glass window ?
   a take-out container for food?
   a flexible, expandable bag for carrying items? a lightweight bottle cap?

3. An unknown plastic floats in a 10% NaCl solution but sinks in water. What is the range of possible density values this plastic may have? Suggest the composition of this plastic.

4. Why is it important to dislodge any adhering bubbles in the density tests?

5. PET plastic (number 1) is the most valuable waste plastic at the present time. Suggest a way to separate it commercially from other waste plastics.
6. Sometimes plastic containers are made from two polymers and not just one. What would happen to the water density test if HDPE and PVC were mixed?
7. Why are plastic recyclers very concerned about identifying the different polymers and not mixing them together?
8. The figure below depicts polymerization of polystyrene (PS). Circle the original monomers and determine how many monomers are present.

9. Polyvinylchloride (PVC) is composed of the vinyl chloride monomer. The monomer structure and general reaction are shown at right.

Draw a polyvinyl chloride polymer composed of five monomers arranged in a head-to-tail pattern.

10. For the bouncy balls you made, what is the name of the monomer?
    What is the role of each of the following in the formation of the polymer?

glue

borax

cornstarch

10. Which ball bounced the highest? Based on your data in the table, which compound was most likely responsible for this?

Contributors and Attributions

The experiment is from Chem 9 Experiments shared under a CC BY-NC license and was authored, remixed, and/or curated by Santa Monica College. Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead, has modified it to fit into a home school environment with less hazardous and cost-effective lab supplies.

2.5: Equilibrium and Le Chatelier's Principle (Experiment)-Home Version

Objectives

- To observe the effect of an applied stress on chemical systems at equilibrium.

A reversible reaction is a reaction in which both the conversion of reactants to products (forward reaction) and the re-conversion of products to reactants (backward reaction) occur simultaneously:

Forward reaction:

$$\mathrm{A+B \longrightarrow C+D} \tag{2.5.1}$$

$$\text{Reactants} \longrightarrow \text{Products} \tag{2.5.2}$$

Backward reaction:

$$\mathrm{C+D \longrightarrow A+B} \tag{2.5.3}$$

$$\text{Products} \longrightarrow \text{Reactants} \tag{2.5.4}$$

Reversible reaction:

$$\mathrm{A+B \rightleftharpoons C+D} \tag{2.5.5}$$

Consider the case of a reversible reaction in which a concentrated mixture of only A and B is supplied. Initially the forward reaction rate ($\mathrm{A+B \longrightarrow C+D}$) is fast since the reactant concentration is high. However as the reaction proceeds, the concentrations of A and B will decrease. Thus over time the forward reaction slows down. On the other hand, as the reaction proceeds, the concentrations of C and D are increasing. Thus although initially slow, the backward reaction rate ($\mathrm{C+D \longrightarrow A+B}$) will speed up over time. Eventually a point will be reached where the rate of the forward reaction will be equal to the rate of the backward reaction. When this occurs, a state of chemical equilibrium is said to exist. Chemical equilibrium is a dynamic state. At equilibrium both the forward and backward reactions are still occurring, but the concentrations of A, B, C, and D remain constant.

A reversible reaction at equilibrium can be disturbed if a stress is applied to it. Examples of stresses include increasing or decreasing chemical concentrations, or temperature changes. If such a stress is applied, the reversible reaction will undergo a shift in order to re-establish its equilibrium. This is known as Le Chatelier's Principle.

Consider a hypothetical reversible reaction already at equilibrium: $\mathrm{A+B \rightleftharpoons C+D}$. If, for example, the concentration of A is increased, the system would no longer be at equilibrium. The rate of the forward reaction ($\mathrm{A+B \longrightarrow C+D}$) would briefly increase in order to reduce the amount of A present and would cause the system to undergo a net shift to the right. Eventually the forward reaction would slow down and the forward and backward reaction rates become equal again as the system returns to a state of equilibrium. Using similar logic, the following changes in concentration are expected to cause the following shifts:

- Increasing the concentration of A or B causes a shift to the right.
- Increasing the concentration of C or D causes a shift to the left.
- Decreasing the concentration of A or B causes a shift to the left.
- Decreasing the concentration of C or D causes a shift to the right.

In other words, if a chemical is added to a reversible reaction at equilibrium, a shift away from the added chemical occurs. When a chemical is removed from a reversible reaction at equilibrium, a shift towards the removed chemical occurs.

A change in temperature will also cause a reversible reaction at equilibrium to undergo a shift. The direction of the shift largely depends on whether the reaction is exothermic or endothermic. In exothermic reactions, heat energy is released and can thus be considered a product. In endothermic reactions, heat energy is absorbed and thus can be considered a reactant.

Exothermic:

$$\mathrm{A+B \rightleftharpoons C+D+\text{ heat}} \tag{2.5.6}$$

Endothermic:

$$\mathrm{A} + \mathrm{B} + \text{ heat} \rightleftharpoons \mathrm{C} + \mathrm{D} \tag{2.5.7}$$

As a general rule, if the temperature is increased, a shift away from the side of the equation with "heat" occurs. If the temperature is decreased, a shift towards the side of the equation with "heat" occurs.

In this lab, the effect of applying stresses to a variety of chemical systems at equilibrium will be explored. The equilibrium systems to be studied are given below:

1. Adding a common ion-Common ion effect-Dissolution equilibrium
   Adding saturated sodium bi-carbonate solution to saturated sodium chloride solution
2. Red cabbage indicator under H+ concentration change (acidity change) -Acid-Base equilibrium
3. Thermochromic substances under heat change (like in mood rings or color changing spoons)
4. Phase equilibria-Heat change

By observing the changes that occur (color changes, precipitate formation, etc.) the direction of a particular shift may be determined. Such shifts may then be explained by carefully examining the effect of the applied stress as dictated by Le Chatelier's Principle.

Procedure

Materials and Equipment.
Sodium chloride (table salt) 5 grams dissolved in 10 ml, Sodium bicarbonate (6-7 grams dissolved in 10 ml). Red cabbage solution prepared in the PH lab, Vinegar or lime juice, water, **optional**-mood ring or color changing spoons or any thermo chromic substance that changes color in cold water to hot water (watch the video given below)

Safety

No safety precautions are needed except heat safety when preparing the cabbage solution by boiling cabbage pieces in water.

Experimental Procedure

Record all observations on your report form. These should include, but not be limited to, color changes and precipitates. Note that solution volumes are approximate for all reactions below.

Part 1: Saturated Sodium Chloride Solution

1. Place 3-mL of saturated NaCl (*aq*) into a small test tube.
2. ***Carefully*** add concentrated clear saturated sodium bicarbonate solution $\mathrm{NaHCO_3}$ (*aq*) drop-wise to the solution in the test tube until a distinct change occurs. Record your observations.

Part 2: Acidified Cabbage pH Indicator

1. Place 3-mL of the red cabbage solution into a test tube.
2. Add an equal amount of 3 ml vinegar. Observe the color change
3. Now add drop-by-drop of sodium bicarbonate until it goes back to the original color.
4. Play around by repeating additions 2 and 3 to see the color changes.

Part 3: Thermochromic Materials

- Watch the following video on thermochromic materials and do a little bit search on the type and nature of such reactions. Find a chemical equation as an example with corresponding color changes.

Part 4: Phase Changes (What?)

1. Place an ice-cube in a container on the table.
2. Once it melts, place the container back in the freezer. Take it out and observe after 4 hours
3. Identify the stress that shifts the phase equilibria.

Pre-laboratory Assignment: Chemical Equilibrium and Le Chatelier’s Principle

1. Consider the reversible reaction:

$$\mathrm{A + B \rightleftharpoons C + D} \tag{2.5.8}$$

- What happens to the forward and reverse reaction *rates* when equilibrium is achieved?
- What happens to the reactant (A and B) and product (C and D) *concentrations* when equilibrium is achieved?

2. Le Chatelier's Principle states that if a stress is applied to a reversible reaction at equilibrium, the reaction will undergo a shift in order to re-establish its equilibrium. Consider the following exothermic reversible reaction at equilibrium:

$$\mathrm{2\,A \rightleftharpoons B + C} \tag{2.5.9}$$

In which direction (left or right) would the following stresses cause the system to shift?

- decrease the concentration of A
- increase the concentration of B
- lower the temperature

3. In this lab you will explore the effect of Le Chatelier's Principle on several chemical systems at equilibrium. These are supplied in the Theory Section. Consider the third system you will study: the Aqueous Ammonia Solution.

- Write the balanced equation for this reversible reaction.
- Suppose you added some excess ammonium ions to this system at equilibrium.
  - In which direction would a shift occur?
  - What color change might you expect to observe?

4. List all the equipment you will use in this lab.

Lab Report: Chemical Equilibrium and Le Chatelier’s Principle

Part 1 - Saturated Sodium Chloride Solution

Equilibrium System:

Observations upon addition of $\mathrm{NaHCO_3}$:

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? Explain.

Part 2 - Acidified Cabbage Indicator

Equilibrium System:

Observations upon addition of H^+:

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? Explain.

Observations upon addition of $NaHCO_3$:

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? Explain.

Part 3 - Thermochromic Materials

Search and find one thermo chromic reaction from literature and write it down as an equilibrium equation with the corresponding colors on each side.

What stress causes the color change?

In which direction did this stress cause the equilibrium system to shift if it is heated? Left or Right with the specific color change

Part 4 - Phase Change

Equilibrium System:

Write an equation to show the equilibrium between the ice and water equilibrium

Draw a phase diagram of water showing the fusion (melting) curve highlighted.

What stress causes the equilibrium shift.

Contributors and Attributions

The experiment is from Chem 10 Experiments shared under a CC BY-NC license and was authored, remixed, and/or curated by Santa Monica College. Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead, has modified this to fit into a homeschool environment with less hazardous and cost-effective lab supplies. Manjusha would like to acknowledge the creators of the YouTube video embedded on this page.

2.6: Acidity-pH Indicators, Acid-base Equilibria, and Acid Dissociation Constant-Home

Learning Objectives

- To prepare a homemade pH indicator and compare it with the standard pH strip available in the market
- To elucidate the equilibrium shift involved in the color change
- To calculate the acid dissociation constant of a weak acid from it pH value

Theory

pH is a scale used to measure the acidity of a substance. The scale ranges from 0 to 14, 0 being the most acidic, 14 being the most basic, and 7 being neutral. A pH value below 7 is considered as acidic, above 7 as basic or alkaline, and 7 as neutral. *Plant pigment, Anthocyanin*, like that in Red cabbage changes color with the acidity of the medium. It is a *good indicator to measure the acidity* of a solution.

It measures the amount of H^+ or H_3O^+ (hydronium) ions in water. Acids contain more hydronium ions than in pure water (10^{-7} M). Bases contain less than that amount. We can calculate the pH using the following equation. pH = (-) log $[H_3O^+]$ ion concentration. There are many ways to check the pH of a solution like pH meters, pH strips, pH indicator solutions etc.

In **Part A** we are going to extract a pH indicator found in red cabbage called anthocyanin and see how it changes color with some household items of varying acidity. We will check the pH of the tested items again with standard pH strips to compare and see how good the cabbage indicator is.

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pH	pH less than 7=Acid			pH more than 7=Base		
	2	4	6	8	10	12
Color	Red	Pinkish purple	Violet	Blue	Blue-green	Green-yellow

The reactive component of general acids is H^+ or H_3O^+ (hydronium) ions. The strength of an acid depends on how well the acid can dissociate to release the H^+ or H_3O^+ (hydronium) ions. Strong acid ions almost completely. Weak acids ionize partially or reversibly resulting in an equilibrium as follows. The ionization varies for various weak acids, so also their strengths. The strength of a weak acid is often expressed as its dissociation constant known as Acid Dissociation Constant or Ka.

In **Part B** we will measure the pH of a weak acid (Vinegar or 5% Acetic acid) of known concentration and use the following equation to calculate its Ka value. Then the calculated Ka will be compared with the true value available from the standard chart (1.8×10^{-5}). The % error will be calculated as

% Error= I (True value – Calculate value) I X 100
True value

Acid dissociation constant **Ka = $[H_3O^{+}] \times [A^{-}] / [HA]$** No units for Ka
[HA] is the initial concentration of the acid. Here it is 0.087 M Acetic acid (5% acetic acid)

$[H_3O^+]$ can be calculate from pH as **$[H_3O^+] = 10^{(-pH)}$** . [A-] will be the same as the $[H_3O^+]$ as they are 1:1.

Suppose pH =5.6, H_3O^+ = 10^ (-5.6) = 2.5 X 10^-6 M

Procedure

Materials required

pH strips, one red cabbage (1-2 leaves), 100 ml water, vinegar (5% acetic acid),5-7 house hold items to check the pH like Baking soda, soap, lemonade, antacid etc.(if the items in the list are not there at home, substitute with some nonhazardous materials at home), Sauce pan, pipets, stove or microwave, heat resistant gloves.

Note

Do not handle the cabbage solution when it is hot. Use heat resistant gloves to handle hot objects. Stick to heat safety protocols.
Reserve some red cabbage solution for the equilibrium lab. Complete part 2 of equilibrium lab involving cabbage solution after this lab.

Part A: Preparation of pH Indicator from Red Cabbage

Preparation of the Indicator Solution-Place approximately 1 handful of chopped red cabbage pieces in the cooking vessel. Add 100 ml (1-2 cups) tap water. Keep the beaker and content on a hot plate/stove/microwave. Turn on the heat to boil the water. Careful when handling hot objects. When the plant pigment gets extracted into the water, turn off heating, and let the solution cool down to room temperature (say another 10 minutes). Drain off the cabbage pieces. Take the clear blue solution and divide this solution into 10 equal parts into each of the 10 vials/bottles/tumblers.
Analysis of the pH of various solutions including household items

Label the vials/bottles/tumblers from 1-10 using a marker. Add 8-10 drops of each chemical into the labeled vials. (Read the safety instructions on each household chemical bottles if there any) Follow the order given in the following table. Observe the color change and record in the table given below. Compare with the color chart on page 1 and assign pH, and predict whether acid/base or neutral based on the color and pH comparison. If the color is too dark, dilute with water. Once you are done with the experiment, dispose the remaining indicator solutions and the solutions in the test tubes down the drain. Dispose the cabbage residue in the trash bin. Wash your hands thoroughly with soap. Hope you enjoyed the lab experiment! You can use different house hold items if you do not have all of the following items. But enter the item in the table. Check the pH of the household items tested using standard pH strips and record in the report sheet. Compare and contract the pH you got from the cabbage indicator and the standard strips.

Part B: Calculating the Acid Dissociation Constant for Acetic Acid of Known Concentration

Using a pH strip measure the pH of a Vinegar (5% acidity Acetic acid). Record this PH, and calculate the Hydronium ion concentration using the equation given in the theory part. Use that to calculate the Ka value. Compare the Ka value with the standard Ka value for acetic acid. Find the % error in your calculation.

Report Sheet

Part A: Preparation of pH Indicator from Red Cabbage

[illegible] Lemonade	Cola	[illegible]

Color of cabbage solution	
pH estimated	
Acid/base/neutral	

pH using pH strips	

Part B: Calculating the Acid Dissociation Constant for Acetic Acid of Known Concentration

pH of Vinegar (5% acidity Acetic acid) measured using pH strip=

$[H_3O^+]$ =

[A-] =

[HA]=0.87 M

Ka calculated=

Ka from literature (True value)= 1.8×10^{-5}

% Error=

Data Analysis

Questions

1. How good is the cabbage indicator in comparison to the standard strips to measure the pH of the household items. Write you opinion.

2. How would you explain the % error in the Ka value if it is more than 5%.

3. Other than anthocyanins what other plant pigments can change color with pH?

Contributors and Attributions

Manjusha Saraswathiamma, Minnesota State University, Moorhead has developed this experiment to be performed at home with less hazardous and cost-effective lab supplies. Manjusha would like to acknowledge the photographer whose picture is included in this experiment. The photo is by an unknown author licensed under CC-BY-SA.

2.7: Volumetric Titration-Home

Learning Objectives

- To use titration techniques to find the concentration of an unknown sample

Theory

Vitamin C (Ascorbic acid **$C_6H_8O_{6)}$** is oxidized by free iodine molecule as follows. This reaction is often used in volumetry to quantify the amount of Vitamin C in a sample. Thiat is what we will see in this titration lab experiment. Since iodine is involved, and these are redox reactions not neutralization reactions, these titrations are called iodometric titrations. However, in kinetics lab experiments, this reaction is often used as an example where the speed of a reaction is measured or rather calculated. The following equation shows the 1:1 reaction between Vitamin C (Ascorbic acid) and Iodine.

$C_6H_8O_6 + I_2$ à $C_6H_6O_6 + 2I^- + 2H^+$

Iodine (I_2) solution is a brown solution and dissolved Iodide (I^-) ions are colorless. Free iodine in the presence of starch makes an **I_2-Starch complex** that is dark blue in color. If there is vitamin C, it will react with I_2. So, there will not be any free iodine to complex with starch to give the blue color. Once the vitamin C is oxidized completely, the free iodine will be available, and it will complex with starch to give the blue color. This is an indication of the completion of vitamin C oxidation. This is called the end-point of the reaction. The amount of iodine used to reach the end-point is recorded. It will be converted to the mole equivalent. Since this is a 1:1 reaction between iodine and Vitamin C as shown in the above equation. The mole of iodine used is equal to the moles of Vitamin C reacted or present in the sample. The moles of vitamin C is then converted into grams or other unit derivatives for comparison.

The sample that needs to be analyzed is called the analyte. In this case it is Vitamin C samples (tablet or fruit juice). The reagent with a known (pre-calculated) concentration used from the burette or an addition equipment is called the titrant. In our case it is the free iodine solution (Tincture iodine (2%). Since we do not have a burette, we are using a long graduated pipette with a bulb to measure the amount of iodine solution added to reach the blue end-point. Watch the following video demonstration.

We will do two titrations. The first one is a calibration reaction to calculate the exact concentration (Molarity of Tincture Iodine (2%) solution (titrant) using a known concentration of Vitamin C tablet. This is to confirm the molarity of the titrant before we analyze any radom sample. This improves the accuracy of our results.

The second is the analytical titration to find out the amount of vitamin C in a juice/cola containing unknown concentrations of vitamin C using a known titrant, Tincture Iodine solution. The volume of iodine required to reach the blue end point will be used to find the concentration of vitamin C sample. If the fruit juice is diluted for titration, the calculated concentration needs to be multiplied back by the dilution factor to get the original concentration. The following video is only for a demonstration purpose. We will be using a slightly different procedure using Vitamin C tablet instead of pure ascorbic acid. We will be adding Iodine solution using a long graduated pipette with bulb instead of a burette. See the procedure given below the video.

Procedure

Materials Required

Vitamin C tablets (Nature Best brand) 500 or 1000 mg supplements, 2-3 #, Tincture iodine antiseptic (2%) 30ml, , Starch (corn or potato), 1-2 spoons, Beakers or clear bowls for mixing, graduated cylinder, electronic balance, spoons, long graduated pipette with bulb to be instead of a burette, fruit juice or any beverage with vitamin C.

> Note
>
> **Wear gloves. Iodine can cause stains.**

Procedure and Report Sheet

PART A-Calibration of the Titrant Tincture Iodine Solution (2%)

- The vitamin C tables are not standardized for their ascorbic acid (vitamin C) content compared tp pure ascorbic acid available in a real laboratory. In order to find the concentration or mols of vitamin C in a particular amount of the tablet, it is titrated with a standardized base like NaOH. Since we do not have NaOH. We are going to calculate the millimols/mg of the tablet based on the titration values given in the video. This is why we need to use the same brand of Vitamin C tables as used in this reaction shown in the video. This can be used to find the mols of Vitamin C in our calibration experiment.
- M1V1=mols of NaOH. Since it is a 1:1 reaction between Vitamin C and NaOH, the number of mols of Vitamin C should be equal to the mols of NaOH used.

It comes to 0.001milli mol Vitamin C per every milli gram of the Vitamin C tablet. For example, for sample 1

- To prepare a standard solution of vitamin C with 0.002M Concentration for calibration titration
  2000mg Vitamin C tablet X 0.001mmol Vitamin C/mg tablet = 2.0 mmol Vitamin = 0.002 mol Vitamin C. If this is dissolved in 1 Liter it would be a 0.002 M standard vitamin C solution.

  0.002 M=0.002 mol/1L   Instead of dissolving 2000 mg tablet in 1 L, we can scale it down by a factor of 10 since we need only 100 ml of the standard solution.
  ie, 200mg tablet in 100 ml water.
- Crush a Vitamin C tablet to a fine powder. Keep a clear plastic cup or beaker (medium size) on the scale. zero the weight. Measure out exactly 100 mg (0.10 g) of Vitamin C powder into using a spoon.
- Measure out exactly 100.0 ml water using a graduated cylinder, and add into this. Dissolve the solution.
- Prepare a starch solution by dissolving 1/2 a spoon of starch (potato or corn) in 50 ml water in a microwave.
- Add 1 ml starch solution into the Vitamin C solution prepared.
- Using a graduated pipette, draw 1-5 ml Tincture iodine (2%) is a graduated cylinder.
- Slowly add the iodine into the Vitamin C solution. Swirl the Vitamin C container to mix them. Add the iodine solution drop-by-drop until you see the whole solution turns blue as shown in the video. Keep track of the addition and record how many milliliters of iodine solution was used to reach the blue end-point. Some times you need more than 5 ml. Then draw more iodine solution and add, but keep track of the total milliliters added.
- Use the M1.V1=M2.V2 equation to find out the exact concentration of the iodine solution.
  M1=Concentration of standard Vitamin C solution = 0.002M ; V1-Volume of standard Vitamin C solution =100ml
  M2=Concentration of Tincture Iodine solution (2%) =? M ; V2=Volume of Iodine solution used to reach the blue end-point = ____ml (enter the ml added here)
  M2=(M1.V1)/V2 =   M of Iodine solution (titrant)

PART B:Iodometric Titration to Find the Concentration of Vitamin C in a Fruit Juice Sample

- Find a fruit juice/lemonade/soda with Vitamin C in the list of ingredients. Since we do not know the concentration of Vitamin C in the sample, dilute it by a factor 10 to be comparable with the iodine solution concentration.
- Measure out exactly 10 ml of the vitamin C containing beverage into a 100 ml graduated cylinder. Add enough water to make it exactly to 100 ml.
- Add 1 ml starch solution into the Vitamin C solution prepared.
- Using a graduated pipette, draw 1-5 ml Tincture iodine (2%) is a graduated cylinder.
- Slowly add the iodine into the Vitamin C solution. Swirl the Vitamin C container to mix them. Add the iodine solution drop-by-drop until you see the whole solution turns blue as shown in the video. Keep track of the addition and record how many

milliliters of iodine solution was used to reach the blue end-point. Some times you need more than 5 ml. Then draw more iodine solution and add, but keep track of the total milliliters added.

- Use the M1.V1=M2.V2 equation to find out the exact concentration of the iodine solution.
  M1=Concentration of unknown Vitamin C solution = ?? M ; V1-Volume of unknown Vitamin C solution =100ml
  M2=Concentration of Tincture Iodine solution (2%) = (enter the value calculated in step A) M ; V2=Volume of Iodine solution used to reach the blue end-point = ____ml (enter the ml added here)
  M1=(M2.V2)/V1 = M of Vitamin C in teh Unknown Vitamin C solution
- Actual concentration of Vitamin C in the beverage solution = 10X M1= M

Data Analysis

Questions

1. What is the purpose of the calibration titration?
2. Show the work to convert the actual concentration of Vitamin C for the beverage (the last calculation) in M to moles.
3. Convert the above moles of Vitamin C in Q2 to grams.

Contributors and Attributions

Manjusha Saraswathiamma, Minnesota State University, Moorhead has developed this experiment to be performed at home with less hazardous and cost-effective lab supplies. Manjusha would like to acknowledge all the contributors of the YouTube videos embedded on this page.

2.8: Kinetics- Iodine Clock using Vitamin C and the Calculation of Initial Rate-Home

Learning Objectives

- To observe and record the time needed for a reaction to happen
- To develop skills to set-up a redox reaction
- To complete the calculations to find out the initial rates
- To graphically analyze the data to find the rate law for the reaction

Theory

Vitamin C (Ascorbic acid **$C_6H_8O_6$)** is oxidized by free iodine molecule as follows. This reaction is often used in volumetry to quantify the amount of Vitamin C in a sample. This we will see in the titration lab experiment. Since iodine is involved, and these are redox reactions not neutralization reactions, these titrations are called iodometric titrations. However, in kinetics lab experiments, this reaction is often used as an example where the speed of a reaction is measured or rather calculated. The following equation shows the 1:1 reaction between Vitamin C (Ascorbic acid) and Iodine.

$C_6H_8O_6 + I_2$ à $C_6H_6O_6 + 2I^- + 2H^+$

Iodine (I_2) solution is a brown solution and dissolved Iodide (I^-) ions are colorless. Free iodine in the presence of starch makes an **I_2-Starch complex** that is dark blue in color. If there is vitamin C, it will react with I_2. So, there will not be any free iodine to complex with starch to give the blue color. Once the vitamin C is oxidized completely, the free iodine will be available, and it will complex with starch to give the blue color. This is an indication of the completion of vitamin C oxidation. If you have more Vitamin C, it will take more time for the completion, and the appearance of blue colored is delayed. Therefore, the rate of the Vitamin C oxidation can be followed by the time taken for the appearance of the blue free iodine-starch complex. That is why this reaction is known as the iodine-clock reaction.

A set of four to five reactions will be set up with increasing amount of Vitamin C, keeping the amount all other reagents the same. Use a stopwatch (cellphone) to measure the time taken for the appearance of blue color in each case and are recorded. The mols of Vitamin C taken in each sample, and the corresponding time are used to calculate the initial rate of the reaction. Use a graph sheet or excel spread sheet to manipulate the data to find the rate law (zero, first, second, or?).

Watch the following video to see the set-up on iodine clock reaction at home using household items. The concentration calculations in molarity is more of an estimation as we are not using standard pure chemicals. We will be using a different amounts of reagents for our reactions. This video is meant only for a demonstration purpose.

Procedure

Materials Required

Vitamin C tablets (Nature Best brand) 500 or 1000 mg supplements, 2-3 #, Tincture iodine antiseptic (2%) 30ml, Hydrogen peroxide, 3% solution, 100 ml, Starch (corn or potato), 1-2 spoons, Beakers or clear bowls for mixing, graduated cylinder, electronic balance, spoons. No filtration is required.

Note

Wear gloves. Iodine can cause stains.

Part A: Calculating the Rate of Vitamin C Oxidation using Initial Rate Method

- Even though we are doing five samples with different amounts of Vitamin C, we will do one at a time. That means prepare one set of A and B solution as shown in the video, mix and measure the time time taken for the appearance of the blue color using a stopwatch.
- Crush 2-3 Vitamin C tablets.
- **Sample 1**
  Solution A: Dissolve **200 mg** (**0.2g)** of Vitamin C powder in 25 ml of water in a beaker or a clear cup, and then add 6 ml 2% tincture iodine solution: Solution B: Dissolve a pinch of starch in 15 ml 3% Hydrogen peroxide, add another 16 ml water to this. Mix solutions A and B and start the stop-watch immediately to measure the time taken for the appearance of the signature blue color. The total volume of the solution is 62 ml.
- **Sample 2**
  Repeat the above procedure, but with **400 mg (0.4g**) Vitamin C in Solution A. Keep the measurement of all other reagents the same as in sample 1. Prepare B with no change. Record the time take to see the blue color after mixing A and B.
- **Sample 3, 4, and 5**. Repeat the above procedure with various amounts of Vitamin C in making solution A, **600 mg (0.6g)** in the third sample, **800 mg (0.8 g)** in the fourth, and **1000 mg (1.0 g)** in the 5th. Remember to keep the amount of all other ingredients the same in all the repetitions for preparing solutions A and B.. Record the time taken to see the blue color in each experiment.

Part B-Graphical Analysis to Find the Rate Law (Graph Paper or Excel Spreadsheet)

- Create a table with Concentration of Vitamin C (millimols) in each sample versus the time recorded
- Create two other manipulated tables. Table 2, take the natural logarithm of the concentrations of Vitamin C vs the time. Table 3, take the inverse of all the Vitamin C concentrations versus the time.
- Check which of the graphs give a straight line. If it is the first one, the rate law is zero order. If it is the second, it is a first order rate law. The thirs one would be second-order if it gives a straight line.

Report Sheet Part A: Calculating the Rate of Vitamin C Oxidation using Initial Rate Method

- The vitamin C tables are not standardized are pure ascorbic acid available in a real laboratory. In order to find the concentration or mols of vitamin C in a particular amount of the tablet, it is titrated with a standardized base like NaOH. Since we do not have NaOH. We are going to calculate the millimols/mg of the tablet based on the titration values given in the video. This is why we need to use the same brand of Vitamin C tables as used in this reaction shown in the video. This can be used to find the mols of Vitamin C used in the following 5 samples.
- M1V1=mols of NaOH. Since it is a 1:1 reaction between Vitamin C and NaOH, the number of mols of Vitamin C should be equal to the mols of NaOH used. Divide the mols of Vitamin C with the weight in milli grams of the tablet used we will get the mols per every mg of Vitamin C tablet. This calculations in the video is used to find the amount of Vitamin C is milli mol/mg of the tablet. We are going to borrow this to calculate the mols of Vitamin C in the report table for each sample.

It comes to 0.001milli mol Vitamin C per every milli gram of the Vitamin C tablet. For example, for sample 1

- Molarity of Vitamin C would be
  200mg Vitamin C tablet X 0.001mmol Vitamin C/mg tablet = 0.2mmol Vitamin C in sample 1
- To calculate molarity M, divide the mols by volume of the solution (62 ml is the total volume of solutions in all the samples). For sample 1, it would be Molarity, M=0.2mmol/63 ml = 0.003M. Calculate the molarity of vitamin C for all the other sample using the same method and enter in the table below.
- **Initial Rate calculation**

Sample Vitamin C in grams Mols of Vitamin C	Molarity, M of Vitamin C milli mol/milliliter	Time recorded Δt	Rate calculated = Rate= (-) Δ [Vitamin C]/Δt

Sample 1-200mg (0.2g) Vitamin C 200mg tablet X0.001mmol Vitamin C /mg tablet =0.2mmol Vitamin C	0.2 mmol/62ml=0.003M		(-)(0-0.003) M/(Δt) s =0.003M/(Δt)s =
Sample 1-400mg (0.4g) Vitamin C			

Sample 1 - 600 mg (0.6 g) Vitamin C			
Sample 1 - 800 mg (0.8 g) Vitamin C			

Sample 1-1000 mg (1.0 g) Vitamin C			

Report Sheet-Graphical: Vitamin C Analysis to Find the Rate Law

Complete the following tables and plot the graphs of dependent variable vs. independent variable (t) in each case and check for the rate law. Watch the following video to review the concept.

Case 1

Concentration = Molarity of Vitamin C	Time, s
Sample 1 =0.003M	

Sample 2	
Sample 3	
Sample 4	
Sample 5	

Case 2

Ln (concentration)	Time, s
Sample 1, Ln (0.003) = (-5.809)	

Case 3

1/(Concentration of Vitamin C) 1/M	Time, s
Sample 1= 1/(0.003)= 333.3	

Data Analysis

- Are the initial rates for all the samples the same?
- Is the initial rate increasing or decreasing with increasing concentration of Vitamin C?
- How would the temperature impact the rate? Will it increase or decrease the speed of the reaction?
- The change in concentration is (final concentration - initial concentration) for each sample. What is the final concentration of Vitamin C for all the samples?
- In PART B, which of the three cases resulted in a straight line plot? Include all the three graphs with the lab report.
- What is the rate law for the reaction, zero, first, or second? Write down the rate law equation.
- Calculate the rate constant, k from the slope of the straight-line plot. k=

Contributors and Attributions

Manjusha Saraswathiamma, Minnesota State University, Moorhead has developed this experiment to be performed at home with less hazardous and cost-effective lab supplies. Manjusha would like to acknowledge all the contributors of the YouTube videos embedded on this page.

2.9: Determination of the Molar Mass by Freezing Point Depression (Experiment)- Home Version

Objectives

- To determine the molar mass of an unknown solid using the colligative property of freezing point depression.

When a small amount of nonvolatile solute is dissolved in a volatile solvent, the vapor pressure of the solvent over the solution will be less than the vapor pressure of the pure solvent at the same temperature. Therefore, the temperature at which the equilibrium vapor pressure reaches atmospheric pressure is higher for the solution than for the pure solvent. As a result the boiling point of the solution, T_b, is higher than the boiling point of the pure solvent, T_b^o. The amount by which the boiling point of the solution exceeds the boiling point of the pure liquid, $\Delta T_b = T_b - T_b^o$, is called the **boiling point elevation**. Similarly, because of the reduction in vapor pressure over the solution the freezing point of the solution, T_f, is lower than the freezing point of the pure solvent, T_f^o. The amount by which the freezing point of the solution is decreased from that of the pure liquid, $\Delta T_f = T_f^o - T_f$, is called the **freezing point depression**. Figure 1 below illustrates this behavior for water.

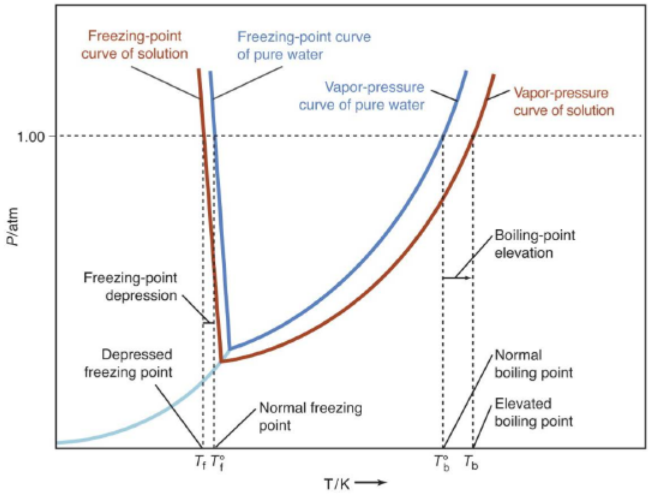

Figure 1

The magnitude of the freezing depression produced by a solute is proportional to its colligative molality, m_c:

$$\Delta T_f = T_f^o - T_f = K_f \times m_c \tag{2.9.1}$$

K_f is known as the **freezing point depression constant**, and depends on the solvent used.

In this experiment you will determine the molar mass of an unknown solid by dissolving a pre-weighed sample in a solvent, and measuring the resulting freezing point depression of the solvent. From the measured ΔT_f and the known K_f value of the solvent, you can then determine the value of m_c using the above Equation 2.9.1. The colligative molality, m_c, is related to the molality of the solution, m, by the expression:

$$m_c = i \times m \tag{2.9.2}$$

where i is the number of solute particles produced per formula unit of dissolved solute, and m is the number of moles of solute per kilogram of solvent. Since only non-dissociating solutes will be used in this experiment, the value of i for your unknown solute can be considered to be 1. Thus, you may assume that $m_c = m$. From the experimentally determined value of m and the mass of solute added, you can determine the molar mass of the unknown solute.

The solvent that will be used in this experiment is water. It has a freezing point of 0 °C.

In order to determine the freezing point of this pure solvent you must cool it using an ice bath, and then measure the temperature as a function of time as the liquid cools. At first the temperature will fall quite rapidly. When the freezing point is reached, solid will begin to form, and the temperature will tend to hold steady until the sample is all solid. This behavior is shown in Figure 2 below. The freezing point of the pure liquid is the constant temperature observed while the liquid is freezing to a solid.

The cooling behavior of a solution is somewhat different from that of a pure liquid, also shown in Figure 2. As discussed earlier, the temperature at which a solution freezes is lower than that for the pure solvent. In addition, there is a slow gradual fall in temperature as freezing proceeds. The best value for the freezing point of the solution is obtained by drawing two straight lines connecting the points on the temperature-time graph. The first line connects points where the solution is all liquid. The second line connects points where solid and liquid coexist. The point where the two lines intersect is the freezing point of the solution.

Note that when the solid first appears the temperature may fall below the freezing point, but then it comes back up as more of the solid forms. This effect is called supercooling, and is a phenomenon that may occur with both the pure liquid and the solution (see Figure 2). When drawing the straight line in the solid-liquid region of the graph, ignore points where supercooling is observed. To establish the proper straight line in the solid-liquid region it is necessary to record the temperature until the trend with time is smooth and clearly established.

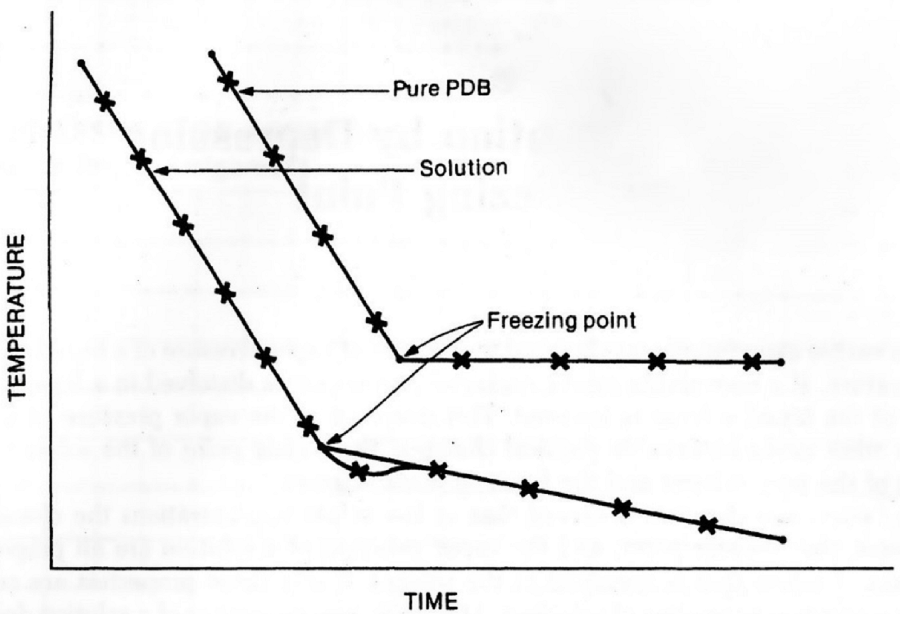

Figure 2

Procedure

Materials Required

Water, thermometer, sugar, 100 ml beaker or a test tube/small glass tumbler/glass bottle to hold the solutions for cooling, scale, spoon, Table salt/ice cream salt, a bigger and wider beaker/container to make ice-bath, ice

Safety

Since we are using non-hazardous chemicals, no safety precautions are advised except for glass safety

Watch the following video for a demonstration of a similar experiment.

Part 1: Determining the Freezing Point of Pure Water (not so pure if it is not distilled!)

1. Take a bigger beaker (1 L) or container. Fill 3/4th of this container with ice cubes with 10-20 grams of salt added in between to make an ice-bath.
2. Weigh a clean dry large test tube or a small 100 ml beaker using an electronic balance, and record its mass. Add about 10 grams of water to the test tube. Reweigh and record the mass of the test tube and the water. Calculate the mass of water in the test tube by difference.
3. Insert your thermometer into the water.
4. Place the container in the ice-bath.
5. Monitor the temperature of the water as it cools. Stir the liquid slowly but continuously to help minimize supercooling. When the temperature of the water reaches 0 °C begin recording its temperature to the nearest tenth of a degree every 30 seconds. Continue for at least 4 minutes after the first solid starts to appear, or until the liquid has solidified to a point that you are no longer able to stir it. Near the melting point you will observe crystals of water in the liquid, and these will increase in amount as the cooling proceeds. Note the temperature at which these crystals first start to appear.

Part 2: Determining the Freezing Point of Water with about 3 g Unknown Solute (but we know it is sugar!)

8. Weigh out and transfer about 3 grams of the unknown solid sample sugar into the test tube, taking great care that none of the unknown sample is spilled during this process. If you do spill some you will need to start this step over with a fresh sample of weighed water. Add 10 grams of water into this and dissolve well.
9. Place the well dissolved sugar solution in the test with a thermometer inserted into in the ice-bath to cool.
10. Monitor the temperature of the water-unknown sugar mixture as it cools. Stir the liquid slowly but continuously to help minimize supercooling. When the temperature of the mixture reaches 0 °C begin recording its temperature to the nearest tenth of a degree every 30 seconds. Continue for at least 4 minutes after the first solid starts to appear, or until the mixture has solidified to a point that you are no longer able to stir it. Note the temperature at which crystals first start to appear.
11. The dependence of temperature on time for the mixture should be similar to that observed for the pure water, except that the first crystals should appear at a lower temperature for the mixture. There may be some supercooling, as evidenced by a rise in temperature shortly after the first appearance of crystals.

Part 3: Determining the Freezing Point of Water with about 5 g Unknown Solute (but we know it is sugar!)

12. Weigh out and transfer 5 grams of the unknown solid sample sugar into the test tube, taking great care that none of the unknown sample is spilled during this process. If you do spill some you will need to start this step over with a fresh sample of weighed water. Add 10 grams of water into this and dissolve well.
13. Place the well dissolved sugar solution in the test with a thermometer inserted into in the ice-bath to cool.
14. Monitor the temperature of the water-unknown sugar mixture as it cools. Stir the liquid slowly but continuously to help minimize supercooling. When the temperature of the mixture reaches 0 °C begin recording its temperature to the nearest tenth

of a degree every 30 seconds. Continue for at least 4 minutes after the first solid starts to appear, or until the mixture has solidified to a point that you are no longer able to stir it. Note the temperature at which crystals first start to appear.

15. The dependence of temperature on time for the mixture should be similar to that observed for the pure water, except that the first crystals should appear at a lower temperature for the mixture. There may be some supercooling, as evidenced by a rise in temperature shortly after the first appearance of crystals.

Lab Report: Determination of Molar Mass by Freezing Point Depression

Experimental Data

Part 1: Freezing Point of Pure Water	
Mass of large empty test tube	
Mass of test tube & water	
Mass of water (by difference)	

Part 2: Freezing Point of PDB-Unknown Mixture (~3 g)	
Mass of unknown added , Sugar Sample I	

Part 3: Freezing Point of PDB-Unknown Mixture (~5 g)	
Mass of unknown added, Sugar sample II	

Temperature Measurements: Record the temperature every 30 seconds as the pure solvent and two solutions are cooled. Note the temperature at which any solid first appears.

Time Elapsed (minutes)	Temperature (°C)		
---	Pure Solvent (water only)	Solution I (water + 3 g sugar)	Solution II (water + 5 g sugar)
0			
0.5			
1.5			
2			
2.5			
3			
3.5			
4			
4.5			
5			
5.5			
6			
6.5			
7			
7.5			
8			
8.5			
9			

Time Elapsed (minutes)	Temperature (°C)		
9.5			
10			
10.5			
11			
11.5			
12			
12.5			
13			
13.5			

Graphical Analysis of Data

Use Excel to create three separate graphs of “Temperature versus Time” for the pure solvent and the two solutions studied. Each graph should have an appropriate title and labeled axes with an appropriate scale. Add two trendlines to the data points of each graph. You can do this by hand with a ruler or by using Excel. The first line is applied to data points that correspond to the cooling of the liquid stat: these are the points on the steep part of the graph. The second line is applied to data points that correspond to the co-existence of both the solid and liquid (freezing): these are the points on the part of the graph where the temperature levels out. Extrapolate the two trendlines towards each other until they intersect. The temperature at the point of intersection is the solvent freezing point and should be clearly shown on each graph. **Attach your three graphs to this report.**

Record the freezing point temperatures obtained from the graphs below:

- Pure water________________°C
- Solution I _________________°C
- Solution II ________________°C

Calculation of Molar Mass

Complete the table below with the results of your calculations. Be sure to include all units. Note that K_f PDB = 7.10 °C·kg·mol^{-1}.

	Solution I	Solution II
Mass of water added		
Total mass of unknown added		
Freezing point of pure water		
Freezing point of solution		
Total Freezing point depression, ΔT_f		
Molality of solution		
Moles of unknown in solution		
Molar Mass of unknown		

- Unknown number ______________________ has an average molar mass of _______________________g/mol.
- On a separate sheet of paper, show the equations used and calculations performed for molality of the solutions, moles of unknown in the solutions, and molar mass of the unknown. **Attach your sheet to this report.**

Questions

1. Look up the freezing point of pure water This may be found in the *CRC Handbook of Chemistry and Physics* (library) or from various online sources.

- Freezing Point of water =_________________________
- Reference source used:

2. Using the freezing point from your reference source, determine the percentage error in your experimentally measured freezing point. Show your calculation below, and be sure to report your answer to the correct number of significant figures.
3. Suppose you find out that the actual molar mass of your unknown solid is exactly three times larger than the value you determined experimentally. What could you conclude about the nature of your unknown solid and the assumptions you made in your calculations in such a case? Explain.

Pre-Laboratory Assignment: Determination of Molar Mass by Freezing Point Depression

1. Camphor melts at 179.8ºC and has a freezing point depression constant, K_f = 40 ºC/molal. When 0.186 g of an unknown organic solid is dissolved in 22.01 g of liquid camphor, the freezing point of the mixture is found to be 176.7ºC. What is the molar mass of the solute?
2. A particular unknown solid depresses the freezing point of PDB (K_f = 7.1ºC/molal) by 3.5ºC. If you dissolve the same solid in H_2O (K_f = 1.86ºC/molal) to the same final molality as you had it in the PDB, what would the change in freezing point be?
3. You are instructed to add about 2 g of the unknown solid when determining the freezing point of the solution. Suppose you accidentally add 5 g of solid. Will the measured freezing point of the solution be higher or lower as a result of this mistake? Explain.
4. Cleaning PDB from the test tube after the experiment could be a long and arduous process. Describe the procedure that you will use to clean PDB (and the unknown solid) from your test tube after your experiments are complete.

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The experiment is from Chem 12 Experiments shared under a CC BY-NC license and was authored, remixed, and/or curated by Santa Monica College. Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead, has modified this to fit to a homeschool environment with less hazardous and cost-effective lab supplies.Manjusha like to ackwoledge the creators of the YouTube video embedded on this page.

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