Appendix:

CHEMLAB WORKBOOK AND ANSWER KEY

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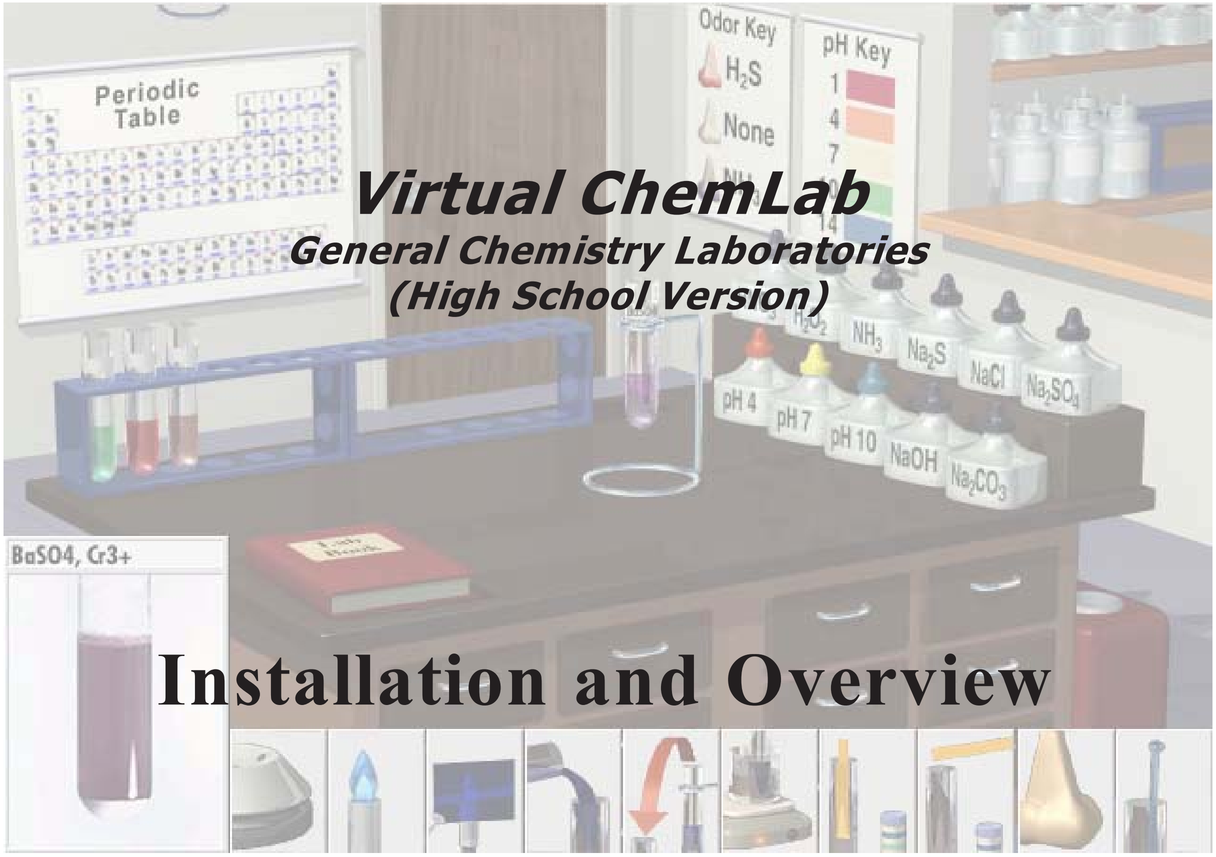
Virtual ChemLab

General Chemistry Laboratories
(High School Version)



Installation and Overview

Overview of ChemLab and Workbook

Welcome to *Virtual ChemLab*, a set of realistic and sophisticated simulations covering general and organic chemistry laboratories. In these laboratories, students are put into a virtual environment where they are free to make the decisions that they would confront in an actual laboratory setting and, in turn, experience the consequences. These laboratories include simulations of inorganic qualitative analysis, quantum chemistry experiments, gas properties, titration experiments, calorimetry, organic synthesis, and organic qualitative analysis. This version of *Virtual ChemLab* is intended to be used in conjunction with this workbook, which contains thirty laboratory assignments covering the topics of stoichiometry, atomic theory, gas properties, thermodynamics, chemical properties, acid-base chemistry, and electrochemistry. Many of these "virtual" laboratory assignments can also be found as "real" assignments in the text or accompanying laboratory manual, but there are also a significant number of new assignments, such as Thomson's cathode ray tube experiment or the Millikan oil drop experiment, that will provide previously unavailable opportunities to perform experiments. This overview and the installation instructions are for the high school, single user version of *Virtual ChemLab: General Chemistry Laboratories*. A brief description of the five general chemistry laboratories available in *Virtual ChemLab* is given below.

The general features of the inorganic simulation include twenty-six cations that can be added to test tubes in any combination, eleven reagents that can be added to the test tubes in any sequence and any number of times, necessary laboratory manipulations, a lab book for recording results and observations, and a stockroom for creating test tubes with known mixtures, generating practice unknowns, or retrieving instructor assigned unknowns. The simulation uses over twenty-five hundred actual pictures to show the results of reactions and over two hundred twenty videos to show the different flame tests. With twenty-six cations that can be combined in any order or combination and eleven reagents that can be added in any order, there are in excess of 10^{16} possible outcomes in the simulation.

The purpose of the quantum laboratory is to allow you and your fellow students to explore and better understand the foundational experiments that led to the development of quantum mechanics. Because of the very sophisticated nature of most of these experiments, the quantum laboratory is the most "virtual" of the *Virtual ChemLab* laboratory simulations. In general, the laboratory consists of an optics table where a source, sample, modifier, and detector combination can be placed to perform different experiments. These devices are located in the stockroom and can be taken out of the stockroom and placed in various locations on the optics table. The emphasis here is to teach you to probe a sample (e.g., a gas, metal foil, two-slit screen, and so forth) with a source (e.g., a laser, electron gun, alpha-particle source, and so forth) and detect the outcome with a specific detector (e.g., a phosphor screen, spectrometer, and so forth). Heat, electric fields, or magnetic fields can also be applied to modify an aspect of the experiment. As in all *Virtual ChemLab* laboratories, the focus is to give you the ability to explore and discover, in a safe and level-appropriate setting, the concepts that are important in the various areas of chemistry.

The gas experiments included in the *Virtual ChemLab* simulated laboratory allow you to explore and better understand the behavior of ideal gases, real gases, and van der Waals gases (a model real gas). The gases laboratory contains four experiments each of which includes the four variables used to describe a gas: pressure (P), temperature (T), volume (V), and the number of moles (n). The four experiments differ by allowing one of these variables to be the dependent variable while the others are independent. The four experiments include (1) V as a function of P, T, and n using a balloon to reflect the volume changes; (2) P as a function of V, T, and n using a motor driven piston; (3) T as a function of P, V, and n again using a motor driven piston; and (4) V as a function of P, T, and n but this time using a frictionless, massless piston to reflect volume changes and by using weights to apply pressure. The gases that can be used in these experiments include an ideal gas; a van der Waals gas whose parameters can be changed to represent any real gas; real gases including N_2, CO_2, CH_4, H_2O, NH_3, and He; and eight ideal gases with different molecular weights that can be added to the experiments to form gas mixtures.

The virtual titration laboratory allows you to perform precise, quantitative titrations involving acid-base and electrochemical reactions. The available laboratory equipment consists of a 50 mL buret; 5, 10, and 25 mL pipets; graduated cylinders; beakers; a stir plate; a set of eight acid-base indicators; a pH meter/voltmeter; a conductivity meter; and an analytical balance for weighing out solids. Acid-base titrations can be performed on any combination of mono-, di-, and tri-protic acids and mono-, di-, and tri-basic bases. The pH of these titrations can be monitored using a pH meter, an indicator, and a conductivity meter as a function of volume, and this data can be saved to an electronic lab book for later analysis. A smaller set of potentiometric titrations can also be performed. Systematic and random errors in the mass and volume measurements have been included in the simulation by introducing buoyancy errors in the mass weighings, volumetric errors in the glassware, and characteristic systematic and random errors in the pH/voltmeter and conductivity meter output. These errors can be ignored, which will produce results and errors typically found in high school or freshman-level laboratory work, or the buoyancy and volumetric errors can be measured and included in the calculations to produce results better than 0.1 percent in accuracy and reproducibility.

The calorimetry laboratory provides you with three different calorimeters that allow them to measure various thermodynamic processes including heats of combustion, heats of solution, heats of reaction, heat capacity, and heat of fusion of ice. The calorimeters provided in the simulations are a classic "coffee cup" calorimeter, a dewar flask (a better version of a coffee cup), and a bomb calorimeter. The calorimetric method used in each calorimeter is based on measuring the temperature change associated with the different thermodynamic processes. You can choose from a wide selection of organic materials to measure the heats of combustion; salts to measure the heats of solution; acids, bases, oxidants, and reductants to measure the heats of reaction; metals and alloys to measure the heat capacity measurements; and ice to measure the melting process. Temperature versus time data can be graphed during the measurements and saved to the electronic lab book for later analysis. Systematic and random errors in the mass and volume measurements have been included in the simulation by introducing buoyancy errors in the mass weighings, volumetric errors in the glassware, and characteristic systematic and random errors in the thermometer measurements.

After installing the *Virtual ChemLab* simulations, the software is initially configured to run in a standalone or student mode where the laboratories are accessed either through the electronic workbook or by clicking on the General Chemistry door. The electronic workbook is new for the high school version and is designed to be used in conjunction with the student worksheets provided with the software. In this high school version, the entire simulation package is available for exploring and performing worksheet assignments, but electronic assignments cannot be received from the instructor nor submitted by you. However, if the Web Connectivity Option has been enabled, electronic assignments and your results can be exchanged via a standard internet connection. Details on setting up and using the Web connectivity feature is given in the various laboratory user guides. It is strongly suggested that the user guides be reviewed before running the software. Most questions and problems can be avoided if the user guides are studied carefully.

System Requirements

Minimum system requirements are as follows:

PC

Pentium 500 MHz (Pentium II or better recommended)

128 Mb RAM (256+ Mb recommended)

CD-ROM drive (for installation only)

600 Mb of free disk space

Display capable of *and* set to millions of colors (24 bit color)

Minimum resolution 800 x 600 (1024 x 768 or higher strongly recommended)

Windows 98 or Windows NT 4.0 or Windows 2000 Professional/ME or Windows XP

QuickTime 5.0 or higher

Macintosh

PowerPC (G3 or better recommended)

128 Mb RAM (256+ Mb recommended)

CD-ROM drive (for installation only)

600 Mb of free disk space

Display capable of *and* set to millions of colors (24-bit color)

Recommended minimum resolution 832 x 624 (1024 x 768 or higher strongly recommended)

OSX (any version)

QuickTime 5.0 or higher

Note: The above requirements are the recommended minimum hardware and system software requirements for reasonable execution speeds and reliability. However, it should be noted that the software has been successfully installed and used on computers with significantly lower capabilities than the recommendations given above with corresponding reductions in execution speed and media access time.

Installing *Virtual ChemLab*

Locate and run the program "Setup ChemLab" (which is located in the appropriate operating system folder) on the CD-ROM drive and then follow the prompts. There is only one install option available for the student version, which installs the complete software package to the hard drive. The CD is not needed to run the program.

Important Installation Notes and Issues

1. The graphics used in the simulations require the monitor to be set to 24-bit true color (millions of colors). Lower color resolutions can be used, but the graphics will not be as sharp.
2. In the directory where *Virtual ChemLab* is installed, the user must always have read/write/erase privileges to that directory and all directories underneath. This condition is initially set by the installer, but this may have to be reset manually if the system crashes hard while running *Virtual ChemLab*. This is generally only a problem with the OSX operating system if multiple users login to the same OSX machine. This can also be a problem with advanced Windows operating systems if the user is not a Power User or higher.
3. When multiple users access the same installed software on a given computer, file ownership and read/write privileges become serious issues since *Virtual ChemLab* shares some files, to a certain degree depending on the installation, among users. (a) In a direct access installation or when multiple users on a network drive share the database, all users must have complete read/write/erase privileges to the directory (and all directories underneath) where the database is stored. (b) In a Web connectivity installation, either (i) the same computer login must be used for all *ChemLab* users (so file ownership is the same for all database files) or (ii) each user who creates a local *ChemLab* account (or new *ChemLab* user) must use the same computer login as when the account was created in order to maintain file ownership consistency. This will only be a problem with OSX machines and Windows operating systems using Restricted Users.
4. When using *Virtual ChemLab* under the Windows 2000 Professional or higher operating systems, users must be, at a minimum, a Power User in order for the program to have sufficient rights to run properly. The program will run as a Restricted User but the fonts will be incorrect along with other minor annoyances. In a server environment where a Restricted User is necessary, we suggest that a separate *ChemLab* account be setup, which gives the user Power User Status but only gives the user access to the *ChemLab* software. This is a Macromedia Director limitation.
5. Occasionally on all Windows platforms, the *Virtual ChemLab* installer will fail to load and execute. This can be corrected by going to www.javasoft.com downloading and then installing the most recent version of the java runtime software. The *Virtual ChemLab* installation software is a java based application.
6. QuickTime 5.0 or later is required for the software to run properly. The most recent version of QuickTime can be obtained at http://www.apple.com/quicktime/.
7. When the simulation software has been installed on a Windows 2000 Professional operating system, there is better performance and better system stability when the Windows 2000 Support Pack 2 has been installed.
8. For unknown reasons, on some machines the QuickTime videos will not play properly if the system QuickTime settings are in their default state. This can be corrected by changing the Video Settings in QuickTime to Normal Mode.
9. Printing in *Virtual ChemLab* does not work inside the OSX operating system.
10. There are occasional spontaneous shutdowns of the software in OSX. There are no known causes for this, but it appears to be a Macromedia Director issue.

Getting Started

Virtual ChemLab is launched by clicking on the VCL icon located on the desktop or in the Start Menu for you will be brought to a hallway containing three doors and a workbook sitting on a table (see Figure 1). Clicking on the electronic workbook opens and zooms into the workbook pages (see Figure 2) where you can select preset assignments that correspond to the assignments in the actual workbook. The *Previous* and *Next* buttons are used to page through the set of assignments, and the different assignments can also be accessed by clicking on the section titles located on the left page of the workbook. Clicking on the *Enter Laboratory* button will allow you to enter the general chemistry laboratory (see below), and the *Exit* button is used to leave *Virtual ChemLab*.

From the hallway, you can also enter the general chemistry laboratory by clicking on the General Chemistry door. Once in the laboratory (shown in Figure 3), you will find five different laboratory benches that represent the five different general chemistry laboratories. Mousing over each of these laboratory benches pops up the name of the selected laboratory. To access a specific laboratory, click on the appropriate laboratory bench. While in the general chemistry laboratory, the full functionality of the simulation is available, and you are free to explore and perform experiments as dictated by their instructors or by their own curiosity. The Exit signs in the general chemistry laboratory are used to return to the hallway.

Detailed instructions on how to use each of the five laboratory simulations can be found in the user guides located in the *Virtual ChemLab* installation directory. These same user guides can also be accessed inside each laboratory by clicking on the Pull-Down TV and clicking on the *Help* button.

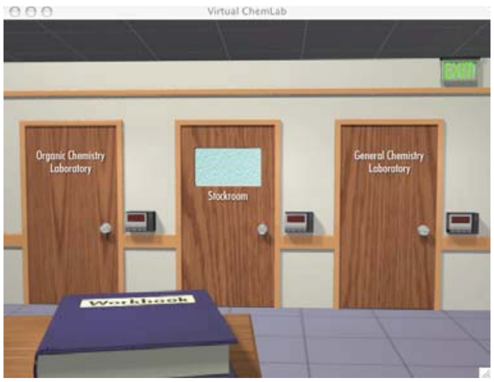


Figure 1. The "hallway" leading into the different virtual rooms in *Virtual ChemLab*. The general chemistry laboratory can be accessed by clicking on the General Chemistry door and the electronic workbook is accessed by clicking on the workbook.

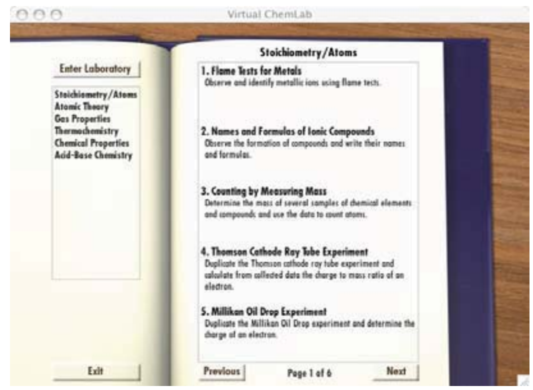


Figure 2. The electronic workbook. Preset laboratories corresponding to assignments in the actual workbook are accessed by clicking on the assignment.

Figure 3. The general chemistry laboratory. The general chemistry laboratory contains five different laboratories, each of which is accessed by clicking on the appropriate lab bench. The exit signs are used to return to the hallway.

Virtual ChemLab Workbook

Student Edition
BYU Independent Study
Chemistry 45

PEARSON Prentice Hall

Names and Formulas of Ionic Compounds

Purpose

To observe the formation of compounds and write their names and formulas.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Names and Formulas of Ionic Compounds* from the list of assignments.
2. The lab will open in the Inorganic laboratory.
3. Enter the stockroom by clicking inside the *Stockroom* window. Once inside the stockroom, drag a test tube from the box and place it on the metal test tube stand. You can then click on the bottle of Ag^+ ion solution on the shelf to add it to the test tube. Click ***Done*** to send the test tube back to the lab. Click on the *Return to Lab* arrow.
4. Place the test tube containing the Ag^+ solution in the metal test tube stand. Click on the ***Divide*** button on the bottom (with the large red arrow) four times to make four additional test tubes containing Ag^+. With one test tube in the metal stand and four others in the blue rack, click on the Na_2S bottle located on the lab bench. You will be able to observe what happens in the window at the bottom left. Record your observation in the table below and write a correct chemical formula and name for the product of the reaction. If the solution remains clear, record NR, for no reaction. Drag this test tube to the red disposal bucket on the right.
5. Place a second tube from the blue rack on the metal stand. Add Na_2SO_4. Record your observations and discard the tube. Use the next tube but add $NaCl$, and record your observations. Use the next tube but add $NaOH$, and record your observations. With the last tube add Na_2CO_3 and record your observations. When you are completely finished, click on the red disposal bucket to clear the lab.
6. Return to the stockroom and repeat steps 3–5 for Pb^{2+}, Ca^{2+}, Fe^{3+}, and Cu^{2+}. Complete the table below.

Analyze

Answers for this lab can be found on page 217.

Each cell should include a description of what you observed when the reagents were mixed and a correct chemical formula and name for all solutions which turned cloudy and NR for all solutions which remained clear. Remember to include roman numerals where appropriate.

	Ag^+	Pb^{2+}	Ca^{2+}	Fe^{3+}	Cu^{2+}
Na_2S (S^{2-})					
$NaCl$ (Cl^-)					
Na_2SO_4 (SO_4^{2-})					
$NaOH$ (OH^-)					
Na_2CO_3 (CO_3^{2-})					

Counting by Measuring Mass

Purpose

Determine the mass of several samples of chemical elements and compounds and use the data to count atoms.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Counting by Measuring Mass* from the list of assignments.
2. The lab will open in the Calorimetry laboratory.

Part 1, Measuring Metal

1. Click on the *Stockroom*. Click on the *Metals* sample cabinet. Open the top drawer by clicking on it. When you open the drawer, a Petri dish will show up on the counter. Place the sample of gold (Au) in the sample dish by double-clicking on it. *Zoom Out* with the green arrow. Place the Petri dish on the stockroom counter by double-clicking on it and *Return to Lab* (by the green arrow).
2. Drag the Petri dish to the spotlight near the balance. Click on the *Balance* area to zoom in. Drag a piece of weigh paper to the balance pan, ***Tare*** the balance and drag the gold sample on the balance pan and record the mass in Table 1.
3. Click red disposal bucket to clear the lab after each sample. Repeat for lead (Pb), uranium (U), sodium (Na) and a metal of your choosing.

Table 1

	gold (Au)	lead (Pb)	uranium (U)	sodium (Na)	Your Choice
Mass (grams)					
Molar Mass (g/mol)					
Moles of each element					
Atoms of each element					

Analyze

Answers for this lab can be found on page 219.

1. Calculate the moles of Au contained in the sample and enter into Table 1.

$$59.9341 g\ Au \cdot \frac{1 mol\ Au}{196.97 g\ Au} = 0.30428 mol\ Au$$

2. Calculate the atoms of Au contained in the sample and enter into Table 1.

$$0.30428 mol\ Au \cdot \frac{6.022 \cdot 10^{23}\ atoms\ Au}{1 mol\ Au} = 1.83 \cdot 10^{23}\ atoms\ Au$$

3. Repeat steps 1 and 2 for the other metals and fill in the table. Clear the laboratory when you are finished by clicking on the disposal bucket.

Part 2, Measuring Compounds

1. Click on the *Stockroom*. Double-click on sodium chloride (NaCl) on the Salts shelf. The right and left arrows allow you to see additional bottles.

2. *Return to Lab*. Move the sample bottle to the spotlight near the balance area. Click on the *Balance* area to zoom in and open the bottle by clicking on the lid (*Remove Lid*). Drag a piece of weigh paper to the balance pan and ***Tare*** the balance.

3. Pick up the *Scoop* and scoop out some sample; as you drag your cursor and the scoop down the face of the bottle it picks up more. Select the largest sample possible and drag the scoop to the weigh paper until it snaps in place which will place the sample on the paper. Record the mass of the sample in Table 2.

4. Repeat steps 1–3 for table sugar (sucrose, $C_{12}H_{22}O_{11}$), NH_4Cl, C_6H_5OH (phenol), and a compound of your choice. Record the mass of each sample in Table 2.

Table 2

	NaCl	$C_{12}H_{22}O_{11}$	NH_4Cl	C_6H_5OH	Your Choice
Mass (grams)					
Molar Mass (g/mol)					
Mole of compound					
Moles of each element					
Atoms of each element					

Analyze

Answer the following using the space provided.

1. Calculate the moles of $C_{12}H_{22}O_{11}$ contained in the sample and record your results in Table 2.

2. Calculate the moles of each element in $C_{12}H_{22}O_{11}$ and record your results in Table 2.

3. Calculate the atoms of each element in $C_{12}H_{22}O_{11}$ and record your results in Table 2.

4. Repeat steps 1–3 for the other compounds and record your results in Table 2.

5. Which of the compounds contains the most *total* atoms?

Thomson Cathode Ray Tube Experiment

Purpose

To duplicate the Thomson cathode ray tube experiment and calculate from collected data the charge to mass ratio (q/m_e) of an electron.

Background

As scientists began to examine atoms, their first discovery was that they could extract negatively charged particles from atoms. They called these particles electrons. In order to understand the nature of these particles, they wanted to know how much charge they carried and how much they weighed. John Joseph (J.J.) Thomson was a physics professor at the famous Cavendish Laboratory at Cambridge University. In 1897, Thomson showed that if you could measure how much a beam of electrons were bent in an electric field and in a magnetic field, you could determine the charge to mass ratio (q/m_e) for the particles (electrons). Knowing the charge to mass ratio (q/m_e) and either the charge on the electron or the mass of the electron would allow you to calculate the other. Thomson could not obtain either in his cathode ray tube experiments and had to be satisfied with just the charge to mass ratio.

Procedure

Answers for this lab can be found on page 222.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Thomson Cathode Ray Tube Experiment* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment? (The source is on the left. Drag your cursor over it to identify it.)

   What type of charge do electrons have?

   What detector is used in this experiment?

4. Turn on the Phosphor Screen. What do you observe?

   The phosphor screen detects charged particles (such as electrons) and it glows momentarily at the positions where the particles impact the screen.

5. Drag the lab window down and left and the phosphor screen window up and right in order to be able to minimize overlap. Push the ***Grid*** button on the phosphor screen, and set the *Magnetic Field* to 30 μT by clicking the button above the tens place three times (The *Magnetic Field* meter is the meter just left of the Phosphor Screen). What happens to the spot from the electron gun on the phosphor screen?

6. Set the *Magnetic Field* back to zero and set the *Electric Field* to 10 V. What happens to the spot from the electron gun on the phosphor screen?

   Where should the signal on the phosphor screen be if the electric and magnetic forces are balanced?

7. Increase the voltage of the *Electric Field* to move the spot several cm from the center. To make your measurements more accurate, move the spot until it aligns with a grid marking. What is the voltage?

   What is the distance from the center that the spot has moved (in cm)?

8. Increase the magnetic field strength until the spot reaches the center of the screen. What magnetic field creates a magnetic force that balances the electric force?

   Summarize your data.

deflected distance (d)	electric field (V)	magnetic field (B)

9. In a simplified and reduced form, the charge to mass ratio (q/m_e) can be calculated as follows:

$$q/m_e = (5.0826 \times 10^{12}) \cdot V \cdot d/B^2$$

   where V = the electric field in volts, d = the deflected distance from center in cm, and B = magnetic field in μT.

   What is your calculated value for the charge to mass ratio for an electron (q/m_e)?

   The modern accepted value is 1.76×10^{11} . Calculate your percent error as follows:

$$\%\ error = \frac{|your\ value - accepted\ value|}{accepted\ value} \times 100$$

Atomic Structure: Rutherford's Experiment

Purpose

To discover how the physical properties, such as size and shape, of an object can be measured by indirect means and to duplicate the gold foil experiment of Ernest Rutherford.

Background

As you have done experiments, you have learned to make useful observations and draw reasonable conclusions from data. But imagine how little you would be able to accomplish if the room in which you worked were so dark that you could not see the materials you were working with. Imagine how limited your observations would be if the object of your scrutiny was so small that it could not be seen with a microscope. When you think of how difficult experimentation would be under such adverse conditions, you will gain some appreciation for the enormous technical problems confronting early atomic scientists.

These scientists had as their target the atom-a bit of matter so small that there was no hope of seeing it directly. Nevertheless, these ingenious experimenters were able to infer that the atom had a nucleus.

A key experiment in understanding the nature of atomic structure was completed by Ernest Rutherford in 1911. He set up an experiment that directed a beam of alpha particles (helium nuclei) through a gold foil and then onto a detector screen. According to the plum pudding atomic model, electrons float around inside a cloud of positive charge. Based on this model, Rutherford expected that almost all the alpha particles should pass through the gold foil and not be deflected. A few of the alpha particles would experience a slight deflection due to the attraction to the negative electrons (alpha particles have a charge of +2). Imagine his surprise when a few alpha particles deflected at all angles, even nearly straight backwards.

According to the plum pudding model there was nothing in the atom massive enough to deflect the alpha particles. About this he said "...almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." He suggested the experimental data could only be explained if the majority of the mass of an atom was concentrated in a small, positively charged central nucleus. This experiment provided the evidence needed to prove this nuclear model of the atom. In this experiment, you will make observations similar to those of Professor Rutherford.

Procedure

Answers for this lab can be found on page 224.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Atomic Structure: Rutherford's Experiment* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. The experiment will be set up on the lab table. Point the cursor arrow to the gray box on the left side. What particles are emitted from this source?

   What are alpha particles?

4. Point the cursor arrow at the base of the metal sample stand (in the center) and squeeze the mouse. What metal foil is used?

5. Point the cursor arrow to the detector (on the right). What detector is used in this experiment?

6. Turn on the detector by clicking on the red/green light switch. What does the signal in the middle of the screen represent?

   The phosphor screen detects charged particles (such as alpha particles) and it glows momentarily at the positions where the particles impact the screen.

   What other signals do you see on the phosphor detection screen?

   What do these signals represent?

   Click the ***Persist*** button (the dotted arrow) on the phosphor detector screen. According to the plum pudding model, what is causing the deflection of the alpha particles?

   Make an observation of the number of alpha particles hitting the phosphor detection screen.

7. Now, you will make observations at different angles of deflection. Click on the gray lab table window to bring it to the front. Grab the phosphor detection screen by its base and move it to the spotlight in the top right corner. The ***Persist*** button should still be on. Observe the number of hits in this spotlight position as compared to the first detector position.

8. Move the detector to the top center spotlight position at a 90^{o} angle to the foil stand. Observe the number of hits in this spotlight position as compared to the first detector position.

9. Move the detector to the top left spotlight position. Observe the number of hits in this spotlight position as compared to the first detector position.

   What causes the alpha particles to deflect backwards?

   How do these results disprove the plum pudding model? Keep in mind that there are 1,000,000 alpha particles passing through the gold foil at any given second.

Are the gold atoms composed mostly of matter or empty space?

How does the Gold Foil Experiment show that almost all of the mass of an atom is concentrated in a small positively charged central atom?

Further Investigation

Students often ask, “Why did Rutherford use gold foil?” The most common response is that gold is soft and malleable and can be made into very thin sheets of foil. There is another reason, which you can discover for yourself.

1. Turn off the phosphor detection screen. Double-click the base of the metal foil sample holder. It will move the holder to the stockroom window. Click on the *Stockroom* to enter. Click on the metal sample box on the top shelf. Click on Na to select sodium. *Return to Lab*.
2. Move the metal foil sample holder from the stockroom window back to the center spotlight. Turn on the phosphor detection screen. Click ***Persist***. Observe the number of hits with sodium compared to the number of hits with a gold sample.

   Why would Rutherford choose gold foil instead of sodium foil? Explain.

Millikan Oil Drop Experiment

Purpose

To duplicate the Millikan Oil Drop experiment and determine the charge on an electron.

Purpose

In the Thomson cathode ray tube experiment, you discovered that you can use the deflection of an electron beam in an electric and magnetic field to measure the charge-to-mass ratio (q/m_e) of an electron. If you then want to know either the charge or the mass of an electron, you need to have a way of measuring one or the other independently. In 1909, Robert Millikan and his graduate student Harvey Fletcher showed that they could make very small oil drops and deposit electrons on these drops (1 to 10 electrons per drop). They would then measure the total charge on the oil drops by deflecting the drops with an electric field. You will get a chance to repeat their experiments and, using the results from the Thomson assignment, be able to experimentally calculate the mass of an electron.

Procedure

Answers for this lab can be found on page 227.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Millikan Oil Drop Experiment* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment?

   How does this source affect the oil droplets in the oil mist chamber?

4. The detector in this experiment is a video camera with a microscopic eyepiece attached to view the oil droplets. Click the ***On/Off*** switch (red/green light) to turn the video camera on.

   What do you observe on the video camera screen?

   Do all the oil drops fall at the same speed?

   What force causes the drops to fall?

   The oil drops fall at their terminal velocity, which is the maximum velocity possible due to frictional forces such as air resistance. The terminal velocity is a function of the radius of the drop. By measuring the terminal velocity (v_t) of a droplet, the radius (r) can be calculated. Then the mass (m) of the drop can be calculated from its radius and the density of the oil. Knowing the mass of the oil droplet will allow you to calculate the charge (q) on the droplet.

IMPORTANT: Read instructions 5 and 6 before beginning the procedure for 5.

5. *Measure the terminal velocity of a drop.* Select a small drop that is falling near the center balance and Click the ***Slow Motion*** button on the video camera window when the drop appears at the top of the window. Wait until the drop is at a tick mark and start the timer. Let the drop fall for at least two or more tick marks and stop the timer. Do not let the drop fall off the end of the viewing scope. Each tick mark is 0.125 mm. Record the distance and the time in the data table below.

6. *Measure the voltage required to stop the fall of the drop.* Having measured the terminal velocity, you now need to stop the fall of the drop by applying an electric field between the two voltage plates. This is done by clicking on the buttons on the top or bottom of the *Electric Field* until the voltage is adjusted such that the drop stops falling. This should be done while in slow motion and when the drop appears stopped, turn slow motion off and do some final adjustments until the drop has not moved for at least one minute. Record the voltage, V, indicated on the voltage controller.

   The Millikan Oil Drop Experiment is classic due to the simplicity of the experimental apparatus and the completeness of the data analysis. The following calculations have reduced very complex equations into more simple ones with several parameters combined into a single constant. Millikan and Fletcher accounted for the force of gravity, the force of the electric field, the density of the oil, the viscosity of the air, the viscosity of the oil, and the air pressure.

Data Table

Drop	Voltage (V, in volts)	Time (t, in seconds)	Distance (d, in meters)
1			
2			
3			

7. *Calculate the terminal velocity and record the value*. Calculate the terminal velocity, v_t, in units of $m \cdot s^{-1}$ using this equation:

   $v_t = \frac{d}{t}$, where d is the distance the drop fell in meters and t is the elapsed time in seconds. Do not forget that the balance on the viewing scope is in mm (1000 mm = 1 m).

   Each of the equations in instructions 8–10 are shown with units and without. You will find it easier to use the equation without units for your calculations.

8. Calculate the radius (r) of the drop and record the value. With the terminal velocity, you can calculate the radius of the drop using this equation:

   $r = (9.0407 \times 10^{-5} m^{1/2} \cdot s^{1/2}) \cdot \sqrt{v_t} \quad = (9.0407 \times 10^{-5}\sqrt{v_t}$, without units)

9. *Calculate the mass of the drop and record the value.* You can use the answer from #8 for radius (r) to calculate the mass of the drop given the density of the oil. The final equation to calculate the mass is

   $m = V_{oil} \cdot P_{oil} = (4\pi)/3 \cdot r^3 \cdot 821 kg \cdot m^{-3}$ $\quad$ = (3439.0 r^3, without units)

   $= (3439.0 kg \cdot m^{-3}) \cdot r^3$

10. Since we applied a voltage across the *Electric Field* to stop the fall of the oil drop, the forces being exerted on the drop must be balanced; that is, the force due to gravity must be the same as the force due to the electric field acting on the electrons stuck to the drop: $qE = mg$. Using this equation, calculate the total charge (Q_{tot}) on the oil drop due to the electrons using the equation:

$$Q_{tot} = Q(n) \cdot e = (9.810 \times 10^{-2} \mathrm{C} \cdot \mathrm{kg}^{-1} \cdot \mathrm{J}^{-1}) \cdot m / V = (9.81 \times 10^{-2} m/V, \text{ without units})$$

where $Q(n)$ is the number of electrons on the drop, e is the fundamental electric charge of an electron, m is the mass calculated in #9, and V is the voltage.

This answer will provide the total charge on the drop (Q_{tot}). The actual experimental fundamental charge of an electron was determined by Millikan to be 1.6×10^{-19}C (coulombs). Divide your total charge (Q_{tot}) by e and round your answer to the nearest whole number. This is the number of electrons ($Q(n)$) that adhered to your drop. Now divide your total charge (Q_{tot}) by $Q(n)$ and you will obtain your experimental value for the charge on one electron. Compare your experimental value to that of Millikan.

11. Complete the experiment and calculations for at <u>least</u> three drops and summarize your results in the data table.

Data Table

Drop #	Terminal Velocity (v_t in m/s)	Radius (r, in meters)	Mass (m, in kg)	Total Charge on Drop (Q_{tot}, in Coulombs)	Charge on One Electron (C)
1					
2					
3					

12. Average your results for the charge on one electron. Calculate the percent error by:

$$\% \ Error = \frac{|\text{your answer} - 1.6 \times 10^{-19}|}{1.6 \times 10^{-19}} \times 100\%$$

What is your average charge for an electron?

What is your percent error?

13. You will recall that in the Thomson experiment you were able to calculate the charge-to-mass ratio (q/m_e) as 1.7×10^{11}. Using this value for q/m_e and your average charge on an electron, calculate the mass of an electron in kg.

What is your calculated value for the mass of an electron in kg?

Atomic Emission Spectra

Purpose

To view atomic emission spectra and use a spectrometer to measure the wavelength. The wavelength will be used to calculate frequency and energy.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Atomic Emission Spectra* from the list of assignments.
2. The lab will open in the Quantum laboratory.

   The *Spectrometer* is on the right of the lab table. The emission spectra is in the detector window in the upper right corner with a graph of the Intensity vs λ (wavelength).

Analyze

Answers for this lab can be found on page 230.

1. How many distinct lines do you see and what are their colors? Draw what you see.
2. Click on the ***Visible/Full*** switch to magnify only the visible spectrum. You will see four peaks in the spectrum. If you drag your cursor over a peak, it will identify the wavelength (in nm) in the *x*-coordinate field in the bottom right corner of the window. Record the wavelength in the table below for the four peaks in the hydrogen spectrum. (Round to whole numbers.)
3. Each wavelength corresponds to another property of light called its frequency. Use the wavelength value of each of the lines to calculate its frequency given that $\nu = c/\lambda$ where $c = 2.998 \times 10^{17}$ nm/s $(2.998 \times 10^{8}$ m/s$)$. The energy (E) of a quantum of light an atom emits is related to its frequency (v) by $E = hv$. Use the frequency value for each line and $h = 6.63 \times 10^{-34}$ Js to calculate its corresponding energy.

	Wavelength (nm)	Frequency (1/s)	Energy (J)
Line #1 (left)			
Line #2			
Line #3			
Line #4 (right)			

4. Now, investigate the emission spectra for a different element, helium. Helium is the next element after hydrogen on the periodic table and has two electrons. Do you think the emission spectra for an atom with two electrons instead of one will be much different than hydrogen?
5. To exchange gas samples, turn off the *Spectrometer* with the ***On/Off*** switch in the top right corner. Double-click on the *Electric Field* to place it on the stockroom shelf. Double-click on the *Gas* (H_2) sample tube to place it on the stockroom shelf.
6. Click in the *Stockroom*. Click on the *Gases* samples on the top shelf. Click on the cylinder labeled *He* to select helium as the gas and it will fill the gas sample tube. If you point to the gas sample tube it should read *He*.
7. Return to lab. Drag the gas sample tube off the stockroom shelf. When you select it, a white spotlight will appear indicating where you can place the gas sample tube—place it there. Drag the *Electric Field* and place it on the gas sample tube. Carefully click the button just above the left zero

on the volt meter and change the voltage to 300 V. Turn on the *Spectrometer*. Click the ***Visible/Full*** switch to convert to only the visible spectrum.

8. Draw the visible spectrum for helium. Is it different from hydrogen?

9. Determine the wavelength (in nm), the frequency (in 1/s) and the energy (in J) for the peak on the far right.

	Wavelength (nm)	Frequency (1/s)	Energy (J)
Line (far right)			

Electronic State Energy Levels

Purpose

To understand the origins of Quantum Theory by using a spectrometer to observe the emission spectrum of several gases.

Background

The classical picture of atoms would allow electrons to be at any energy level. According to this classical model, when electrons are excited and then fall back down to the ground state, they emit light at all wavelengths and the emission spectrum would be continuous.

In the 1800s scientists found that when a sample of gas is excited by an alternating electric current, it emits light only at certain discrete wavelengths. This allowed for the development of spectroscopy which is used in the identification and analysis of elements and compounds. Even though scientists found spectroscopy very useful, they could not explain why the spectrum was not continuous. The explanation of this was left to Niels Bohr, a Danish physicist. Bohr proposed that energy levels of electrons are not continuous but quantized. The electrons only exist in specific energy levels. Because of this quantization of energy, excited electrons can only fall to discrete energy levels.

This assignment illustrates the measurements that helped Bohr develop his quantum model, now known as Quantum Theory. It also illustrates some practical uses for this science. Mercury vapor is used in fluorescent lights and sodium vapor in street lighting.

You can separate the lines in the full region of an emission spectrum by using an optical prism or a diffraction grating. A spectrometer is an instrument designed to separate the emitted light into its component wavelengths and plots the intensity of the light as a function of wavelength.

Procedure

Answers for this lab can be found on page 232.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Electronic State Energy Levels* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. The lab table will be set up with four items. What is the detector on the right?

   What is the metal sample?

   A heat source is used to heat the metal sample to high temperatures. What is the temperature of the heat source?

4. Turn on the *Spectrometer* by clicking on the red/green light switch. Click on the ***Visible/Full*** switch to change the view to the visible spectrum. Click on the *Lab Book* to open it. If any students in a previous class have saved spectra, highlight and delete them. Click on the ***Record*** button (red dot on the spectrometer window) to record this spectrum in the lab book. Click just after the spectra file name and type tungsten metal. What observations can you make about the emission spectrum for heated tungsten metal?

5. Turn the *Spectrometer* off with the ***On/Off*** switch. Click on the *Stockroom*. Click the clipboard labeled *Assignments* on the right. Click on the preset lab #9 *Photoemission – H2* and return to the lab

by clicking on the *Return to Lab* arrow. Click on the ***Visible/Full*** switch to change the view to the visible spectrum. ***Record*** this spectrum in the *Lab Book*. Click just after the spectra file name and type hydrogen gas. What observations can you make about the emission spectrum for hydrogen gas?

6. To exchange gas samples, turn off the *Spectrometer*. Double-click on the *Electric Field* to place it on the stockroom shelf. Double-click on the *Gas* (H_2) sample tube to place it on the stockroom shelf.
7. Click in the *Stockroom*. Click on the *Gases* sample on the top shelf. Click on the cylinder labeled *Ne* to select neon as the gas and it will fill the gas sample tube. If you point to the gas sample tube it should read *Ne*.
8. *Return to Lab*. Drag the gas sample tube off the stockroom shelf. When you select it, a white spotlight will appear indicating where you can place the gas sample tube-place it there. Drag the *Electric Field* and place it on the gas sample tube. Carefully click the button just above the far left zero on the volt meter and change the voltage to 300 V. Turn on the *Spectrometer*. Click the ***Visible/Full*** switch to convert to only the visible spectrum. ***Record*** this spectrum in the lab book and identify this link with the name of the element typed after the blue link.
9. Continue with this same process until your completed samples include the following: H_2, He, Ne, Na, and Hg. You should have five spectra saved in the lab book in addition to tungsten metal. Record your observations for each element. You can return to the lab book and click on any of the spectra to view them again. Include in your observations a comparison for each element to the spectrum for heated tungsten metal.

10. How do your observations of these gas emission spectra help confirm Quantum Theory?

Application

1. Load the spectrum for heated tungsten metal. Tungsten metal is used in incandescent light bulbs as the heated filament.
2. Load the spectrum for mercury from the lab book into the spectrometer. Examine the visible spectrum. Click the switch to change to full spectrum. What differences do you see when changing between visible and full spectrum for mercury.

Mercury vapor is used in the fluorescent light tubes that you see at school and home. The emitted light is not very bright for just the mercury vapor, but when scientists examined the full spectrum for mercury they saw what you observed and recorded. There is an enormous emission in the ultraviolet range (UV). This light is sometimes called black light. You may have seen it with glow-in-the-dark displays.

Scientists coat the inside of the glass tube of fluorescent light tubes with a compound that will absorb UV and emit the energy as visible light with all the colors of the rainbow. All colors together create white light which is why fluorescent light tubes emit very white light.

Laundry detergents contain compounds that absorb UV light and emit visible light. These compounds allow advertisers to claim *whiter* and *brighter whites and colors*. If you attend an event using black light, you may have seen your white socks or white shirt “glow”.

3. Load the spectrum for sodium. What distinct feature do you see in the sodium spectrum?

Astronomers are excited about cities changing from normal street lights to sodium vapor street lights because astronomers can easily filter out the peak at 589 nm and minimize light pollution.

Flame Tests for Metals

Purpose

To observe and identify metallic ions using flame tests.

Background

Have you ever wondered why a candle flame is yellow? The characteristic yellow of a candle flame comes from the glow of burning carbon fragments. The carbon fragments are produced by the incomplete combustion reaction of the wick and candle wax. When elements, such as carbon, are heated to high temperatures, some of their electrons are excited to higher energy levels. When these excited electrons fall back to lower energy levels, they release excess energy in packages of light called photons, or light quanta.

The color of the emitted light depends on its energy. Blue light is more energetic than red light, for example. When heated, each element emits a characteristic pattern of light energies, which is useful for identifying the element. The characteristic colors of light produced when substances are heated in the flame of a gas burner are the basis of flame tests for several elements.

In this experiment, you will perform the flame tests used to identify several metallic elements.

Procedure

Answers for this lab can be found on page 234.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Flame Tests for Metals* from the list of assignments.
2. The lab will open in the Inorganic laboratory.
3. Enter the stockroom by clicking inside the *Stockroom* window. Once inside the stockroom, drag a test tube from the box and place it on the metal test tube stand. You can then click on a bottle of metal ion solution on the shelf to add it to the test tube. When you have added one metal ion, click ***Done*** to send the test tube back to the lab. Repeat this process with a new metal ion. Continue doing this until you have sent one test tube for each the following metal ions to the lab: Na^+, K^+, Ca^{2+}, Ba^{2+}, Sr^{2+}, Cu^{2+}.
4. On the right end of the supply shelf is a button labeled *Unknowns*. Click on the *Unknowns* button to create a test tube with an unknown. Now click on each of the following bottles on the shelf: Na^+, K^+, Ca^{2+}, Ba^{2+}, Sr^{2+}, and Cu^{2+}. Do not change the maximum and minimum on the left side. Click *Save*. An unknown test tube titled *Practice* will show in the blue rack. Drag the practice unknown test tube from the blue rack to place it in the metal stand and click *Done*. Now click on the *Return to Lab* arrow.
5. When you return to the lab you should note that you have seven test tubes. You will use two of the buttons across the bottom, ***Flame*** and ***Flame w/ Cobalt*** (blue glass held in front of the flame.) A test tube must be moved from the blue test tube rack to the metal test tube stand in order to perform the flame test. You can drag a test tube from the blue rack to the metal test tube stand to switch places with a test tube in the metal test tube stand. Just above the periodic table there is a handle. Click on the handle to pull down the TV monitor. With the monitor down you can mouse-over each test tube and it will identify what metal ion the test tube contains. As you mouse over each test tube, you will also see a picture of what it contains in the lower left corner. One of your test tubes is labeled *Practice* and when you mouse-over it, the TV monitor tells you it is an unknown.
6. Select the test tube containing Na^+ and place it on the metal stand. Click the ***Flame*** button. Record your observations in the data table below. Click the ***Flame w/Cobalt*** button and record your observations in the same table.
7. Drag the K^+ test tube to the metal stand to exchange it with the Na^+. Flame test K^+ with and without cobalt glass. Record your observations in the table below.
8. For the other four ions, ***Flame*** test them only. Do not use cobalt glass. Record your observations in the table below.

Data Table

Flame Tests	**[Answers]**
Ion	Flame Color
sodium, Na^+	
sodium, Na^+ (cobalt glass)	
potassium, K^+	
potassium, K^+ (cobalt glass)	
calcium, Ca^+	
barium, Ba^+	
strontium, Sr^+	
copper, Cu^+	
unknown #1	
unknown #2	
unknown #3	
unknown #4	

9. Flame test the practice unknown. Determine which of the six metal ions it most closely matches. You may repeat the flame test on any of the six metal ions if necessary. When you are confident that you have identified the unknown, open the *Lab Book* by clicking on it. On the left page, click the ***Report*** button. On the right page, click on the metal ion that you think is in the practice unknown. Click ***Submit*** and then ***OK***. If all of the ion buttons turn green you have successfully identified the unknown. If any turn red then you were incorrect. ***Flame*** test the practice unknown again to correctly identify your metal ion. Click on the red disposal bucket to clear all of your samples.
10. Return to the *Stockroom* and drag the practice unknown to the metal test tube stand. This will randomly select another metal ion. *Return to lab*, test and report this ion. Continue until you have correctly identified four <u>different</u> practice unknowns.

Analysis and Conclusions

1. The energy of colored light increases in the order red, yellow, green, blue, violet. List the metallic elements used in the flame tests in increasing order of the energy of the light emitted.

2. What is the purpose of using the cobalt glass in the identification of sodium and potassium?

Photoelectric Effect

Purpose

To duplicate photoelectric effect experiments.

Background

Though Albert Einstein is most famous for $E = mc^2$ and his work in describing relativity in mechanics, his Nobel Prize was for understanding a very simple experiment. It was long understood that if you directed light of a certain wavelength at a piece of metal, it would emit electrons. In classical theory, the energy of the light was thought to be based on its intensity and not its frequency. However, the results of the photoelectric effect contradicted classical theory. These inconsistencies led Einstein to suggest that we need to think of light as being composed of particles (photons) and not just as waves.

Each wavelength corresponds to another property of light called frequency. You will use the wavelength (λ) value in the experiment to calculate the frequency (ν) given that $\nu = c/\lambda$ where $c = 2.998 \times 10^{17}$ nm/s (2.998×10^{8} m/s). The energy (E) of a quantum of light an atom emits is related to its frequency (ν) by $E = h\nu$ where h (Planck's constant) = 6.63×10^{-34} J-s.

Procedure

Answers for this lab can be found on page 237.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Photoelectric Effect* from the list of assignments.
2. The experiment opens in the Quantum laboratory.
3. What source is used in this experiment and what does it do?

   At what intensity is the laser set?

   At what wavelength is the laser set?

   Record the wavelength (in nm) in the data table. Calculate the frequency (in 1/s) and the energy (in J) using the equations given in the Background section of this lab. Determine the color of the light by clicking on the *Spectrum Chart* (just behind the laser); the markers indicate what color is represented by the wavelength selected.

   Which metal foil is used in this experiment?

   What detector is used in this experiment and what does it measure?

   Turn on the detector by clicking on the red/green light switch.What does the signal on the phosphor screen indicate about the laser light shining on the sodium foil?

4. Decrease the *Intensity* to 1 photon/second, how does the signal change?

   Increase the *Intensity* to 1kW, how does the signal change?

5. Change the *Intensity* back to 1 nW and increase the *Wavelength* to 600 nm. What do you observe? Record the wavelength in the data table.

   Determine the maximum wavelength at which emission of electrons occurs in the metal.

   What is the difference between intensity and wavelength?

   Which matters in the formation of photoelectrons?

Data Table

Wavelength (nm)	Frequency (1/s)	Energy (J)	Light Color

6. Click in the *Stockroom*. Click on the clipboard and select *Photoelectric Effect* (2). *Return to Lab*. The intensity is set at 1 nW and the wavelength at 400 nm. The detector used in this experiment is a bolometer. Turn on the bolometer by clicking the red/green light switch.

   This instrument measures the kinetic energy of electrons emitted from the metal foil. You should see a green peak on the bolometer detection screen. The intensity or height of the signal corresponds to the number of electrons being emitted from the metal. The x-axis is the kinetic energy of the photoemitted electrons. Zoom in on the peak by clicking and dragging from half-way up the y-axis to the right past the peak and to the x-axis (it will draw an orange rectangle).

7. Increase and decrease the *Intensity*, what do you observe?

   Increase and decrease the *Wavelength*, what do you observe?

What is the maximum wavelength that ejects electrons from the sodium metal?

From this experiment, explain why violet light causes photoemission of electrons but orange light does not.

Diffraction Experiments

Purpose

To investigate the wave-particle duality of nature.

Background

It has long been known that if you shine light through narrow slits that are spaced at small intervals, the light will form a diffraction pattern. A diffraction pattern is a series of light and dark patterns caused by wave interference. The wave interference can be either constructive (light patterns) or destructive (dark patterns). In this experiment, you will shine a laser through a device with two slits where the spacing can be adjusted and investigate the patterns that will be made at a distance from the slits.

Procedure

Answers for this lab can be found on page 239.

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Diffraction Experiments* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment and for what reason?

   What is the wavelength of the *Laser*?

   What is the spacing of the two slits on the two slit device?

   Draw a picture of the pattern displayed on the video screen.

4. Change the *Intensity* of the *Laser* from 1 nW to 1W.

   Does the intensity of the light affect the diffraction pattern?

   Change the *Slit Spacing* to 1 μm. Observe the pattern displayed on the video screen as you change the slit spacing from 1 μm to 7 μm by 1 μm increments.

   What can you state about the relationship between slit spacing and diffraction pattern?

5. Increase the *Wavelength* of the *Laser* to 700 nm.

   What affect does an increase in the wavelength have on the diffraction pattern?

6. Decrease the *Intensity* on the *Laser* to 1000 photons/second. Click on the ***Persist*** button on the video camera to look at individual photons coming through the slits. Observe for one minute.

   What observation can you make about this pattern as compared to the pattern from the continuous beam of photons?

   Decrease the *Intensity* to 100 photons/second. Observe for another minute after clicking ***Persist***. At these lower intensities (1000 and 100 photons/second), there is never a time when two photons go through both slits at the same time. How can a single photon diffract?

   From this experiment, what conclusions can you make about the nature of light?

7. Click in the *Stockroom*. Click on *Clipboard* and select *Two-Slit Diffraction – Electrons*. *Return to Lab*. What source is used in this experiment?

   Draw a picture of the diffraction pattern shown on the *Phosphor Screen*.

   How does this diffraction pattern compare to the diffraction pattern for light?

   Louis de Broglie was the first person to suggest that particles could be considered to have wave properties.

8. Decrease the *Intensity* to 10 electrons/second. The pattern now builds one electron at a time. Click on the ***Persist*** button and observe for one minute.

   Has the diffraction pattern changed? Why or why not?

   How can an electron diffract if there is only one?

   When you look at the complete diffraction pattern of a stream of particles, you are seeing all the places you expect particles to scatter. If you start the source over multiple times, you will see that the first particle is never detected in the same place twice. This is an application of the Heisenberg Uncertainty Principle which is directly connected with measurement. It takes into account the minimum uncertainty of the position (Δx) and the uncertainty of the momentum (Δp). The equation that displays these variables is:

$$(\Delta x)(\Delta p) \geq h/4\pi$$

   Because you know the energy at which the particle is traveling, you can know precisely what the momentum is. But because of this equation you cannot be sure of the position.

Virtual ChemLab Workbook Answer Key

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Chemistry 45

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Names and Formulas of Ionic Compounds

Purpose

To observe the formation of compounds and write their names and formulas.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Names and Formulas of Ionic Compounds* from the list of assignments.
2. The lab will open in the Inorganic laboratory.
3. Enter the stockroom by clicking inside the *Stockroom* window. Once inside the stockroom, drag a test tube from the box and place it on the metal test tube stand. You can then click on the bottle of Ag^+ ion solution on the shelf to add it to the test tube. Click ***Done*** to send the test tube back to the lab. Click on the *Return to Lab* arrow.
4. Place the test tube containing the Ag^+ solution in the metal test tube stand. Click on the ***Divide*** button on the bottom (with the large red arrow) four times to make four additional test tubes containing Ag^+. With one test tube in the metal stand and four others in the blue rack, click on the Na_2S bottle located on the lab bench. You will be able to observe what happens in the window at the bottom left. Record your observation in the table below and write a correct chemical formula and name for the product of the reaction. If the solution remains clear, record NR, for no reaction. Drag this test tube to the red disposal bucket on the right.
5. Place a second tube from the blue rack on the metal stand. Add Na_2SO_4. Record your observations and discard the tube. Use the next tube but add *NaCl*, and record your observations. Use the next tube but add *NaOH*, and record your observations. With the last tube add Na_2CO_3 and record your observations. When you are completely finished, click on the red disposal bucket to clear the lab.
6. Return to the stockroom and repeat steps 3–5 for Pb^{2+}, Ca^{2+}, Fe^{3+}, and Cu^{2+}. Complete the table below.

Analyze

Each cell should include a description of what you observed when the reagents were mixed and a correct chemical formula and name for all solutions which turned cloudy and NR for all solutions which remained clear. Remember to include roman numerals where appropriate.

	Ag^+	Pb^{2+}	Ca^{2+}	Fe^{3+}	Cu^{2+}
Na_2S (S^{2-})	**black** **Ag_2S** **silver sulfide**	**black** **PbS** **lead (II)** **sulfide**	**NR**	**black** **Fe_2S_3** **Iron (III)** **sulfide**	**Black** **CuS** **copper (II)** **suflide**
NaCl (Cl^-)	**white** **AgCl** **silver chloride**	**white** **$PbCl_2$** **lead (II) chloride**	**NR**	**NR**	**NR**
Na_2SO_4 (SO_4^{2-})	**NR**	**white** **$PbSO_4$** **lead (II) sulfate**	**NR**	**NR**	**NR**

NaOH (OH^-)	**pinkish brown** **AgOH** **silver hydroxide**	**NR**	**white** **$Ca(OH)_2$** **calcium hydroxide**	**red** **$Fe(OH)_3$** **iron (III) hydroxide**	**blue** **$Cu(OH)_2$** **copper (II) hydroxide**
Na_2CO_3 (CO_3^{2-})	**pink** **Ag_2CO_3** **silver carbonate**	**white** **$PbCO_3$** **lead (II) carbonate**	**white** **$CaCO_3$** **calcium carbonate**	**red** **$Fe_2(CO_3)_3$** **iron (III) carbonate**	**blue/white** **$CuCO_3$** **copper (II) carbonate**

Counting by Measuring Mass

Purpose

Determine the mass of several samples of chemical elements and compounds and use the data to count atoms.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Counting by Measuring Mass* from the list of assignments.
2. The lab will open in the Calorimetry laboratory.

Part 1, Measuring Metal

1. Click on the *Stockroom*. Click on the *Metals* sample cabinet. Open the top drawer by clicking on it. When you open the drawer, a Petri dish will show up on the counter. Place the sample of gold (Au) in the sample dish by double-clicking on it. *Zoom Out* with the green arrow. Place the Petri dish on the stockroom counter by double-clicking on it and *Return to Lab* (by the green arrow).
2. Drag the Petri dish to the spotlight near the balance. Click on the *Balance* area to zoom in. Drag a piece of weigh paper to the balance pan, ***Tare*** the balance and drag the gold sample on the balance pan and record the mass in Table 1.
3. Click red disposal bucket to clear the lab after each sample. Repeat for lead (Pb), uranium (U), sodium (Na) and a metal of your choosing.

Table 1

	gold (Au)	lead (Pb)	uranium (U)	sodium (Na)	Your Choice
Mass (grams)	**51.0618**	**31.9812**	**51.9042**	**2.7779**	**Tungsten (W) 52.8043**
Molar Mass (g/mol)	**196.97**	**207.20**	**238.03**	**22.99**	**183.84**
Moles of each element	**0.2592**	**0.1543**	**0.2181**	**0.1208**	**0.28723**
Atoms of each element	**1.56e23**	**9.29e22**	**1.31e23**	**7.27e22**	**1.73e23**

Analyze

1. Calculate the moles of Au contained in the sample and enter into Table 1.

$$51.0618 g\ Au \cdot \frac{1 mol\ Au}{196.97 g\ Au} = 0.2592 mol\ Au$$

2. Calculate the atoms of Au contained in the sample and enter into Table 1.

$$0.2592 mol\ Au \cdot \frac{6.022 \cdot 10^{23}\ atoms\ Au}{1 mol\ Au} = 1.56 \cdot 10^{23}\ atoms\ Au$$

3. Repeat steps 1 and 2 for the other metals and fill in the table. Clear the laboratory when you are finished by clicking on the disposal bucket.

Part 2, Measuring Compounds

1. Click on the *Stockroom*. Double-click on sodium chloride (NaCl) on the Salts shelf. The right and left arrows allow you to see additional bottles.

2. *Return to Lab*. Move the sample bottle to the spotlight near the balance area. Click on the *Balance* area to zoom in and open the bottle by clicking on the lid (*Remove Lid*). Drag a piece of weigh paper to the balance pan and ***Tare*** the balance.

3. Pick up the *Scoop* and scoop out some sample; as you drag your cursor and the scoop down the face of the bottle it picks up more. Select the largest sample possible and drag the scoop to the weigh paper until it snaps in place which will place the sample on the paper. Record the mass of the sample in Table 2.

4. Repeat steps 1–3 for table sugar (sucrose, $C_{12}H_{22}O_{11}$), NH_4Cl, C_6H_5OH (phenol), and a compound of your choice. Record the mass of each sample in Table 2.

Table 2

	NaCl	$C_{12}H_{22}O_{11}$	NH_4Cl	C_6H_5OH	Your Choice
Mass (grams)	**1.9992**	**1.9987**	**1.1676**	**1.9981**	**NaCN** **1.1678**
Molar Mass (g/mol)	**58.44**	**342.3**	**53.5**	**94.1**	**49.01**
Mole of compound	**0.0342**	**0.00584**	**0.0218**	**0.0212**	**0.0238**
Moles of each element	**all: 0.0342**	**C: 0.0701** **H: 0.128** **O: 0.0642**	**N: 0.0218** **H: 0.0873** **Cl: 0.0218**	**C: 0.127** **H: 0.127** **O: 0.0212**	**all: 0.0238**
Atoms of each element	**all: 2.06e22**	**C: 4.21e22** **H: 7.74e22** **O: 3.87e22**	**N: 1.31e22** **H: 5.26e22** **Cl: 1.31e22**	**C: 7.67e22** **H: 7.67e22** **O: 1.28e22**	**all: 1.435e22**

Analyze

1. Calculate the moles of $C_{12}H_{22}O_{11}$ contained in the sample and record your results in Table 2.

   Answers will vary with the mass of compound pick up by student, but an example:

$$1.9992g\,C_{12}H_{22}O_{11} \cdot \frac{1\,mol\ C_{12}H_{22}O_{11}}{342.3g\,C_{12}H_{22}O_{11}} = 0.00584\,mol\ C_{12}H_{22}O_{11}$$

2. Calculate the moles of each element in $C_{12}H_{22}O_{11}$ and record your results in Table 2.

   Answers will vary with the mass of compound pick up by student, but an example:

$$0.00584\,mol\ C_{12}H_{22}O_{11} \cdot \frac{12\,mol\ C}{1\,mol\ C_{12}H_{22}O_{11}} = 0.0701\,mol\ C$$

$$0.00584\,mol\ C_{12}H_{22}O_{11} \cdot \frac{22\,mol\ H}{1\,mol\ C_{12}H_{22}O_{11}} = 0.128\,mol\ H$$

$$0.00584\,mol\ C_{12}H_{22}O_{11} \cdot \frac{11\,mol\ O}{1\,mol\ C_{12}H_{22}O_{11}} = 0.0642\,mol\ O$$

3. Calculate the atoms of each element in $C_{12}H_{22}O_{11}$ and record your results in Table 2.

Answers will vary with the mass of compound pick up by student, but an example:

$$0.0701\,mol\,C \cdot \frac{6.022 \cdot 10^{23}\,atoms\,C}{1\,mol\,C} = 4.22 \cdot 10^{23}\,atoms\,C$$

$$0.128\,mol\,H \cdot \frac{6.022 \cdot 10^{23}\,atoms\,H}{1\,mol\,H} = 7.74 \cdot 10^{23}\,atoms\,H$$

$$0.0642\,mol\,O \cdot \frac{6.022 \cdot 10^{23}\,atoms\,O}{1\,mol\,O} = 3.87 \cdot 10^{23}\,atoms\,O$$

4. Repeat steps 1–3 for the other compounds and record your results in Table 2.

Answers will vary with the mass of compound pick up by student.

5. Which of the compounds contains the most *total* atoms?

C_6H_5OH

Thomson Cathode Ray Tube Experiment

Purpose

To duplicate the Thomson cathode ray tube experiment and calculate from collected data the charge to mass ratio (q/m_e) of an electron.

Background

As scientists began to examine atoms, their first discovery was that they could extract negatively charged particles from atoms. They called these particles electrons. In order to understand the nature of these particles, they wanted to know how much charge they carried and how much they weighed. John Joseph (J.J.) Thomson was a physics professor at the famous Cavendish Laboratory at Cambridge University. In 1897, Thomson showed that if you could measure how much a beam of electrons were bent in an electric field and in a magnetic field, you could determine the charge to mass ratio (q/m_e) for the particles (electrons). Knowing the charge to mass ratio (q/m_e) and either the charge on the electron or the mass of the electron would allow you to calculate the other. Thomson could not obtain either in his cathode ray tube experiments and had to be satisfied with just the charge to mass ratio.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Thomson Cathode Ray Tube Experiment* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment? (The source is on the left. Drag your cursor over it to identify it.)

   Electron Gun

   What type of charge do electrons have?

   Negative

   What detector is used in this experiment?

   A phosphor screen

4. Turn on the Phosphor Screen. What do you observe?

   A spot in the center of the phosphor screen.

   The phosphor screen detects charged particles (such as electrons) and it glows momentarily at the positions where the particles impact the screen.

5. Drag the lab window down and left and the phosphor screen window up and right in order to be able to minimize overlap. Push the ***Grid*** button on the phosphor screen, and set the *Magnetic Field* to 30 μ T. (Click the button above the tens place three times.) What happens to the spot from the electron gun on the phosphor screen?

   The spot moves to the right.

6. Set the *Magnetic Field* back to zero and set the *Electric Field* to 10 V. What happens to the spot from the electron gun on the phosphor screen?

   The spot moves to the left.

   Where should the signal on the phosphor screen be if the electric and magnetic forces are balanced?

   In the center of the phosphor screen.

7. Increase the voltage of the *Electric Field* to move the spot several cm from the center. To make your measurements more accurate, move the spot until it aligns with a grid marking. What is the voltage?

   Response can vary depending on the answer to #7 from 3–15 V.

   What is the distance from the center that the spot has moved (in cm)?

   Response can vary from 1–6 cm

8. Increase the magnetic field strength until the spot reaches the center of the screen. What magnetic field creates a magnetic force that balances the electric force?

   Response can vary depending on the answer to #7 from 10–50 μT

   Summarize your data.

deflected distance (d)	electric field (V)	magnetic field (B)
see answer to #7	**see answer to #8**	**see answer to #9**

9. In a simplified and reduced form, the charge to mass ratio (q/m_e) can be calculated as follows:

$$q/m_e = (5.0826 \times 10^{12}) \cdot V \cdot d/B^2$$

   where V = the electric field in volts, d = the deflected distance from center in cm, and B = magnetic field in μ T.

   What is your calculated value for the charge to mass ratio for an electron (q/m_e)?

   Response varies depending on data but the answer should be approximately 1.70×10^{11}

   The modern accepted value is 1.76×10^{11}. Calculate your percent error as follows:

$$\%\ error = \frac{|your\ value - accepted\ value|}{accepted\ value} \times 100$$

   Response varies depending on data. A typical error would be 3%.

Atomic Structure: Rutherford's Experiment

Purpose

To discover how the physical properties, such as size and shape, of an object can be measured by indirect means and to duplicate the gold foil experiment of Ernest Rutherford.

Background

As you have done experiments, you have learned to make useful observations and draw reasonable conclusions from data. But imagine how little you would be able to accomplish if the room in which you worked were so dark that you could not see the materials you were working with. Imagine how limited your observations would be if the object of your scrutiny was so small that it could not be seen with a microscope. When you think of how difficult experimentation would be under such adverse conditions, you will gain some appreciation for the enormous technical problems confronting early atomic scientists.

These scientists had as their target the atom-a bit of matter so small that there was no hope of seeing it directly. Nevertheless, these ingenious experimenters were able to infer that the atom had a nucleus.

A key experiment in understanding the nature of atomic structure was completed by Ernest Rutherford in 1911. He set up an experiment that directed a beam of alpha particles (helium nuclei) through a gold foil and then onto a detector screen. According to the plum pudding atomic model, electrons float around inside a cloud of positive charge. Based on this model, Rutherford expected that almost all the alpha particles should pass through the gold foil and not be deflected. A few of the alpha particles would experience a slight deflection due to the attraction to the negative electrons (alpha particles have a charge of +2). Imagine his surprise when a few alpha particles deflected at all angles, even nearly straight backwards.

According to the plum pudding model there was nothing in the atom massive enough to deflect the alpha particles. About this he said "...almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." He suggested the experimental data could only be explained if the majority of the mass of an atom was concentrated in a small, positively charged central nucleus. This experiment provided the evidence needed to prove this nuclear model of the atom. In this experiment, you will make observations similar to those of Professor Rutherford.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Atomic Structure: Rutherford's Experiment* from the list of assignments.

2. The lab will open in the Quantum laboratory.

3. The experiment will be set up on the lab table. Point the cursor arrow to the gray box on the left side. What particles are emitted from this source?

   Alpha particles

   What are alpha particles?

   A helium nucleus, which has two protons, two neutrons and a charge of +2.

4. Point the cursor arrow at the base of the metal sample stand (in the center) and squeeze the mouse. What metal foil is used?

   Gold

5. Point the cursor arrow to the detector (on the right). What detector is used in this experiment?

   Phosphor screen.

6. Turn on the detector by clicking on the red/green light switch. What does the signal in the middle of the screen represent?

The alpha particles coming straight through the gold foil undeflected or only slightly deflected.

The phosphor screen detects charged particles (such as alpha particles) and it glows momentarily at the positions where the particles impact the screen.

What other signals do you see on the phosphor detection screen?

There are spots of light.

What do these signals represent?

Hits from alpha particles being deflected at small angles.

Click the ***Persist*** button (the dotted arrow) on the phosphor detector screen. According to the plum pudding model, what is causing the deflection of the alpha particles?

As the positively charged alpha particles pass through the gold atoms they are attracted to the negative electrons and their path is bent slightly.

Make an observation of the number of alpha particles hitting the phosphor detection screen.

The alpha particle hits fill the screen rapidly.

7. Now, you will make observations at different angles of deflection. Click on the gray lab table window to bring it to the front. Grab the phosphor detection screen by its base and move it to the spotlight in the top right corner. The ***Persist*** button should still be on. Observe the number of hits in this spotlight position as compared to the first detector position.

   The hits are not quite as frequent and the forward scattering spot is no longer visible.

8. Move the detector to the top center spotlight position at a 90^{o} angle to the foil stand. Observe the number of hits in this spotlight position as compared to the first detector position.

   The hits show up every few seconds.

9. Move the detector to the top left spotlight position. Observe the number of hits in this spotlight position as compared to the first detector position.

   It takes nearly a minute for even a single hit to appear.

   What causes the alpha particles to deflect backwards?

   A large mass in the center of the atom.

   How do these results disprove the plum pudding model? Keep in mind that there are 1,000,000 alpha particles passing through the gold foil at any given second.

   The mass of the gold atom is not spread over the full atomic volume but concentrated in a central atomic nucleus.

   Are the gold atoms composed mostly of matter or empty space?

   Mostly empty space.

   How does the Gold Foil Experiment show that almost all of the mass of an atom is concentrated in a small positively charged central atom?

   Most of the alpha particles came straight through or with small deflections. If the mass of an atom were not concentrated, there would have been more deflections.

Further Investigation

Students often ask, "Why did Rutherford use gold foil?" The most common response is that gold is soft and malleable and can be made into very thin sheets of foil. There is another reason, which you can discover for yourself.

1. Turn off the phosphor detection screen. Double-click the base of the metal foil sample holder. It will move the holder to the stockroom window. Click on the *Stockroom* to enter. Click on the metal sample box on the top shelf. Click on Na to select sodium. *Return to Lab.*
2. Move the metal foil sample holder from the stockroom window back to the center spotlight. Turn on the phosphor detection screen. Click ***Persist***. Observe the number of hits with sodium compared to the number of hits with a gold sample.

   Why would Rutherford choose gold foil instead of sodium foil? Explain.

   The sodium atom is much smaller than the gold atom and the same is true for the nuclei. The sodium nucleus is so small that you have very small chance of hitting it directly and having any alpha particles bounce back. Gold atoms have a much larger cross-section.

Millikan Oil Drop Experiment

Purpose

To duplicate the Millikan Oil Drop experiment and determine the charge on an electron.

Purpose

In the Thomson cathode ray tube experiment, you discovered that you can use the deflection of an electron beam in an electric and magnetic field to measure the charge-to-mass ratio (q/m_e) of an electron. If you then want to know either the charge or the mass of an electron, you need to have a way of measuring one or the other independently. In 1909, Robert Millikan and his graduate student Harvey Fletcher showed that they could make very small oil drops and deposit electrons on these drops (1 to 10 electrons per drop). They would then measure the total charge on the oil drops by deflecting the drops with an electric field. You will get a chance to repeat their experiments and, using the results from the Thomson assignment, be able to experimentally calculate the mass of an electron.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Millikan Oil Drop Experiment* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment?

   Electron Gun

   How does this source affect the oil droplets in the oil mist chamber?

   Some of the electrons adhere to the oil drops.

4. The detector in this experiment is a video camera with a microscopic eyepiece attached to view the oil droplets. Click the ***On/Off*** switch (red/green light) to turn the video camera on.

   What do you observe on the video camera screen?

   There are large and small oil drops falling from top to bottom.

   Do all the oil drops fall at the same speed?

   No, some are falling fast and some slow.

   What force causes the drops to fall?

   Gravity.

   The oil drops fall at their terminal velocity, which is the maximum velocity possible due to frictional forces such as air resistance. The terminal velocity is a function of the radius of the drop. By measuring the terminal velocity (v_t) of a droplet, the radius (r) can be calculated. Then the mass (m) of the drop can be calculated from its radius and the density of the oil. Knowing the mass of the oil droplet will allow you to calculate the charge (q) on the droplet.

IMPORTANT: Read instructions 5 and 6 before beginning the procedure for 5.

5. *Measure the terminal velocity of a drop.* Select a small drop that is falling near the center balance and Click the ***Slow Motion*** button on the video camera window when the drop appears at the top of the window. Wait until the drop is at a tick mark and start the timer. Let the drop fall for at least two or more tick marks and stop the timer. Do not let the drop fall off the end of the viewing scope. Each tick mark is 0.125 mm. Record the distance and the time in the data table below.
6. *Measure the voltage required to stop the fall of the drop.* Having measured the terminal velocity, you now need to stop the fall of the drop by applying an electric field between the two voltage plates. This is done by clicking on the buttons on the top or bottom of the *Electric Field* until the voltage is

adjusted such that the drop stops falling. This should be done while in slow motion and when the drop appears stopped, turn slow motion off and do some final adjustments until the drop has not moved for at least one minute. Record the voltage, V, indicated on the voltage controller.

The Millikan Oil Drop Experiment is classic due to the simplicity of the experimental apparatus and the completeness of the data analysis. The following calculations have reduced very complex equations into more simple ones with several parameters combined into a single constant. Millikan and Fletcher accounted for the force of gravity, the force of the electric field, the density of the oil, the viscosity of the air, the viscosity of the oil, and the air pressure.

Data Table

Drop	Voltage (V, in volts)	Time (t, in seconds)	Distance (d, in meters)
1		**(sample student data)**	
2	**52**	**11.45**	**3.75×10^{-4}**
3			

7. *Calculate the terminal velocity and record the value*. Calculate the terminal velocity, v_t, in units of $m \cdot s^{-1}$ using this equation:

   $v_t = \frac{d}{t}$, where d is the distance the drop fell in meters and t is the elapsed time in seconds. Do not forget that the balance on the viewing scope is in mm (1000 mm = 1 m).

   Answers will vary, but a typical answer might be 3.28×10^{-5} m/s.

   Each of the equations in instructions 8–10 are shown with units and without. You will find it easier to use the equation without units for your calculations.

8. Calculate the radius (r) of the drop and record the value. With the terminal velocity, you can calculate the radius of the drop using this equation:

   $r = (9.0407 \times 10^{-5} m^{1/2} \cdot s^{1/2}) \cdot \sqrt{v_t} = (9.0407 \times 10^{-5}\sqrt{v_t}$ M, without units

   Answers will vary, but a typical answer might be 5.17×10^{-7} m.

9. *Calculate the mass of the drop and record the value.* You can use the answer from #8 for radius (r) to calculate the mass of the drop given the density of the oil. The final equation to calculate the mass is

   $m = V_{oil} \cdot P_{oil} = (4\pi)/3 \cdot r^3 \cdot 821 kg \cdot m^{-3}$ $= (3439.0\ r^3$, without units)

   $= (3439.0 kg \cdot m^{-3}) \cdot r^3$

   Answers will vary, but a typical answer might be 4.76×10^{-18} kg.

10. Since we applied a voltage across the *Electric Field* to stop the fall of the oil drop, the forces being exerted on the drop must be balanced; that is, the force due to gravity must be the same as the force due to the electric field acting on the electrons stuck to the drop: $qE = mg$. Using this equation, calculate the total charge (Q_{tot}) on the oil drop due to the electrons using the equation:

    $Q_{tot} = Q(n) \cdot e = (9.810 \times 10^{-2} C \cdot kg^{-1} \cdot J^{-1}) \cdot m/V = (9.81 \times 10^{-2} m/V$, without units)

    where $Q(n)$ is the number of electrons on the drop, e is the fundamental electric charge of an electron, m is the mass calculated in #9, and V is the voltage.

    This answer will provide the total charge on the drop (Q_{tot}). The fundamental electric charge of an electron (e) is 1.6×10^{-19}C (coulombs). Divide your total charge (Q_{tot}) by e and round your answer to the nearest whole number. This is the number of electrons ($Q(n)$) that adhered to your drop.

Now divide your total charge (Q_{tot}) by $Q(n)$ and you will obtain your experimental value for the charge on one electron.

11. Complete the experiment and calculations for at least three drops and summarize your results in the data table.

Results

Drop #	Terminal Velocity (v_t in m/s)	Radius (r, in meters)	Mass (m, in kg)	Total Charge on Drop (Q_{tot}, in Coulombs)	Charge on One Electron (C)
1	**(sample results)**				
2	**3.28×10^{-5}**	**5.18×10^{-7}**	**4.78×10^{-16}**	**8.49×10^{-19}**	**1.69×10^{-19}**
3					

12. Average your results for the charge on one electron. Calculate the percent error by:

$$\%\ Error = \frac{\left|\text{your answer} - 1.6 \times 10^{-19}\right|}{1.6 \times 10^{-19}} \times 100\%$$

What is your average charge for an electron?

Answers will vary but should range between 1.4×10^{-19} C and 1.8×10^{-19} C.

What is your percent error?

Answers will vary but should range between 0 and 10%.

13. You will recall that in the Thomson experiment you were able to calculate the charge-to-mass ratio (q/m_e) as 1.7×10^{11} . Using this value for q/m_e and your average charge on an electron, calculate the mass of an electron in kg.

What is your calculated value for the mass of an electron in kg?

Answers will vary but should range between 9.41×10^{-31} kg and 1.02×10^{-30} kg. The actual equations used to derive the simplification in this lab are in Appendix D Quantum Equations, Millikan Experiment of the Instructor Utilities Guide.

Atomic Emission Spectra

Purpose

To view atomic emission spectra and use a spectrometer to measure the wavelength. The wavelength will be used to calculate frequency and energy.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Atomic Emission Spectra* from the list of assignments.
2. The lab will open in the Quantum laboratory.

   The *Spectrometer* is on the right of the lab table. The emission spectra is in the detector window in the upper right corner with a graph of the Intensity vs λ (wavelength).

Analyze

1. How many distinct lines do you see and what are their colors? Draw what you see.

   Four: Violet, Blue, Aqua (blue-green), Red.
2. Click on the ***Visible/Full*** switch to magnify only the visible spectrum. You will see four peaks in the spectrum. If you drag your cursor over a peak, it will identify the wavelength (in nm) in the *x*-coordinate field in the bottom right corner of the window. Record the wavelength in the table below for the four peaks in the hydrogen spectrum. (Round to whole numbers.)
3. Each wavelength corresponds to another property of light called its *frequency*. Use the wavelength value of each of the lines to calculate its frequency given that $\gamma = c/\lambda$ where $c = 2.998 \times 10^{17}$ nm/s $(2.998 \times 10^{8}\ \text{m/s})$. The energy ($E$) of a quantum of light an atom emits is related to its frequency (v) by $E = hv$. Use the frequency value for each line and $h = 6.63 \times 10^{-34}$ J·s to calculate its corresponding energy.

	Wavelength (nm)	Frequency (1/s)	Energy (J)
Line #1 (left)	**410**	$\mathbf{7.32 \times 10^{14}}$	$\mathbf{4.85 \times 10^{-19}}$
Line #2	**434**	$\mathbf{6.91 \times 10^{14}}$	$\mathbf{4.58 \times 10^{-19}}$
Line #3	**487**	$\mathbf{6.16 \times 10^{14}}$	$\mathbf{4.08 \times 10^{-19}}$
Line #4 (right)	**656**	$\mathbf{4.57 \times 10^{14}}$	$\mathbf{3.03 \times 10^{-19}}$

4. Now, investigate the emission spectra for a different element, helium. Helium is the next element after hydrogen on the periodic table and has two electrons. Do you think the emission spectra for an atom with two electrons instead of one will be much different than hydrogen?

   With only one additional electron the spectrum should be similar.
5. To exchange gas samples, turn off the *Spectrometer* with the ***On/Off*** switch in the top right corner. Double-click on the *Electric Field* to place it on the stockroom shelf. Double-click on the *Gas (H_2)* sample tube to place it on the stockroom shelf.
6. Click in the *Stockroom*. Click on the *Gases* samples on the top shelf. Click on the cylinder labeled *He* to select helium as the gas and it will fill the gas sample tube. If you point to the gas sample tube it should read *He*.
7. *Return to lab*. Drag the gas sample tube off the stockroom shelf. When you select it, a white spotlight will appear indicating where you can place the gas sample tube-place it there. Drag the *Electric Field* and place it on the gas sample tube. Carefully click the button just above the left zero on

the volt meter and change the voltage to 300 V. Turn on the *Spectrometer*. Click the ***Visible/Full*** switch to convert to only the visible spectrum.

8. Draw the visible spectrum for helium. Is it different from hydrogen?

   Yes, it has six lines, there are two more blue lines, an orange-yellow and green line.

9. Determine the wavelength (in nm), the frequency (in 1/s) and the energy (in J) for the peak on the far right.

	Wavelength (nm)	Frequency (1/s)	Energy (J)
Line (far right)	**668**	**4.49×10^{14}**	**2.98×10^{-19}**

Electronic State Energy Levels

Purpose

To understand the origins of Quantum Theory by using a spectrometer to observe the emission spectrum of several gases.

Background

The classical picture of atoms would allow electrons to be at any energy level. According to this classical model, when electrons are excited and then fall back down to the ground state, they emit light at all wavelengths and the emission spectrum would be continuous.

In the 1800s scientists found that when a sample of gas is excited by an alternating electric current, it emits light only at certain discrete wavelengths. This allowed for the development of spectroscopy which is used in the identification and analysis of elements and compounds. Even though scientists found spectroscopy very useful, they could not explain why the spectrum was not continuous. The explanation of this was left to Niels Bohr, a Danish physicist. Bohr proposed that energy levels of electrons are not continuous but quantized. The electrons only exist in specific energy levels. Because of this quantization of energy, excited electrons can only fall to discrete energy levels.

This assignment illustrates the measurements that helped Bohr develop his quantum model, now known as Quantum Theory. It also illustrates some practical uses for this science. Mercury vapor is used in fluorescent lights and sodium vapor in street lighting.

You can separate the lines in the full region of an emission spectrum by using an optical prism or a diffraction grating. A spectrometer is an instrument designed to separate the emitted light into its component wavelengths and plots the intensity of the light as a function of wavelength.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Electronic State Energy Levels* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. The lab table will be set up with four items. What is the detector on the right?

   A Spectrometer

   What is the metal sample?

   Tungsten

   A heat source is used to heat the metal sample to high temperatures. What is the temperature of the heat source?

   3000 K

4. Turn on the *Spectrometer* by clicking on the red/green light switch. Click on the ***Visible/Full*** switch to change the view to the visible spectrum. Click on the *Lab Book* to open it. If any students in a previous class have saved spectra, highlight and delete them. Click on the ***Record*** button (red dot on the spectrometer window) to record this spectrum in the lab book. Click just after the spectra file name and type tungsten metal. What observations can you make about the emission spectrum for heated tungsten metal?

   The spectrum looks like the colors of the rainbow. It begins with violet on the left and ends with red on the right. The spectrum is continuous and there is no place without color.

5. Turn the *Spectrometer* off with the ***On/Off*** switch. Click on the *Stockroom*. Click the clipboard labeled *Assignments* on the right. Click on the preset lab #9 *Photoemission – H2* and return to the lab by clicking on the *Return to Lab* arrow. Click on the ***Visible/Full*** switch to change the view to the visible spectrum. ***Record*** this spectrum in the *Lab Book*. Click just after the spectra file name and type hydrogen gas. What observations can you make about the emission spectrum for hydrogen gas?

There are only four distinct lines. The spectrum is not continuous. The colors are violet, blue, blue-green and red.

6. To exchange gas samples, turn off the *Spectrometer*. Double-click on the *Electric Field* to place it on the stockroom shelf. Double-click on the *Gas* (H_2) sample tube to place it on the stockroom shelf.

7. Click in the *Stockroom*. Click on the *Gases* sample on the top shelf. Click on the cylinder labeled *Ne* to select neon as the gas and it will fill the gas sample tube. If you point to the gas sample tube it should read *Ne*.

8. *Return to Lab*. Drag the gas sample tube off the stockroom shelf. When you select it, a white spotlight will appear indicating where you can place the gas sample tube-place it there. Drag the *Electric Field* and place it on the gas sample tube. Carefully click the button just above the far left zero on the volt meter and change the voltage to 300 V. Turn on the *Spectrometer*. Click the ***Visible/Full*** switch to convert to only the visible spectrum. ***Record*** this spectrum in the lab book and identify this link with the name of the element typed after the blue link.

9. Continue with this same process until your completed samples include the following: H_2, He, Ne, Na, and Hg. You should have five spectra saved in the lab book in addition to tungsten metal. Record your observations for each element. You can return to the lab book and click on any of the spectra to view them again. Include in your observations a comparison for each element to the spectrum for heated tungsten metal.

   Answers will include descriptions of various elemental spectra.

10. How do your observations of these gas emission spectra help confirm Quantum Theory?

    Each emission spectrum shows distinct peaks and not a continuous spectrum. Electrons only exist in specified energy levels.

Application

1. Load the spectrum for heated tungsten metal. Tungsten metal is used in incandescent light bulbs as the heated filament.

2. Load the spectrum for mercury from the lab book into the spectrometer. Examine the visible spectrum. Click the switch to change to full spectrum. What differences do you see when changing between visible and full spectrum for mercury.

   The red and violet peaks extend past the visible spectrum and fill the full spectrum.

   Mercury vapor is used in the fluorescent light tubes that you see at school and home. The emitted light is not very bright for just the mercury vapor, but when scientists examined the full spectrum for mercury they saw what you observed and recorded. There is an enormous emission in the ultraviolet range (UV). This light is sometimes called black light. You may have seen it with glow-in-the-dark displays.

   Scientists coat the inside of the glass tube of fluorescent light tubes with a compound that will absorb UV and emit the energy as visible light with all the colors of the rainbow. All colors together create white light which is why fluorescent light tubes emit very white light.

   Laundry detergents contain compounds that absorb UV light and emit visible light. These compounds allow advertisers to claim *whiter and brighter whites and colors*. If you attend an event using black light, you may have seen your white socks or white shirt "glow".

3. Load the spectrum for sodium. What distinct feature do you see in the sodium spectrum?

   There is a very intense peak at 589 nm.

   Astronomers are excited about cities changing from normal street lights to sodium vapor street lights because astronomers can easily filter out the peak at 589 nm and minimize light pollution.

Flame Tests for Metals

Purpose

To observe and identify metallic ions using flame tests.

Background

Have you ever wondered why a candle flame is yellow? The characteristic yellow of a candle flame comes from the glow of burning carbon fragments. The carbon fragments are produced by the incomplete combustion reaction of the wick and candle wax. When elements, such as carbon, are heated to high temperatures, some of their electrons are excited to higher energy levels. When these excited electrons fall back to lower energy levels, they release excess energy in packages of light called photons, or light quanta.

The color of the emitted light depends on its energy. Blue light is more energetic than red light, for example. When heated, each element emits a characteristic pattern of light energies, which is useful for identifying the element. The characteristic colors of light produced when substances are heated in the flame of a gas burner are the basis of flame tests for several elements.

In this experiment, you will perform the flame tests used to identify several metallic elements.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Flame Tests for Metals* from the list of assignments.

2. The lab will open in the Inorganic laboratory.

3. Enter the stockroom by clicking inside the *Stockroom* window. Once inside the stockroom, drag a test tube from the box and place it on the metal test tube stand. You can then click on a bottle of metal ion solution on the shelf to add it to the test tube. When you have added one metal ion, click ***Done*** to send the test tube back to the lab. Repeat this process with a new metal ion. Continue doing this until you have sent one test tube for each the following metal ions to the lab: Na^+, K^+, Ca^{2+}, Ba^{2+}, Sr^{2+}, Cu^{2+}.

4. On the right end of the supply shelf is a button labeled *Unknowns*. Click on the *Unknowns* button to create a test tube with an unknown. Now click on each of the following bottles on the shelf: Na^+, K^+, Ca^{2+}, Ba^{2+}, Sr^{2+}, and Cu^{2+}. Do not change the maximum and minimum on the left side. Click *Save*. An unknown test tube titled *Practice* will show in the blue rack. Drag the practice unknown test tube from the blue rack to place it in the metal stand and click ***Done***. Now click on the *Return to Lab* arrow.

5. When you return to the lab you should note that you have seven test tubes. You will use two of the buttons across the bottom, ***Flame*** and ***Flame w/ Cobalt*** (blue glass held in front of the flame.) A test tube must be moved from the blue test tube rack to the metal test tube stand in order to perform the flame test. You can drag a test tube from the blue rack to the metal test tube stand to switch places with a test tube in the metal test tube stand. Just above the periodic table there is a handle. Click on the handle to pull down the TV monitor. With the monitor down you can mouse-over each test tube and it will identify what metal ion the test tube contains. As you mouse over each test tube, you will also see a picture of what it contains in the lower left corner. One of your test tubes is labeled *Practice* and when you mouse-over it, the TV monitor tells you it is an unknown.

6. Select the test tube containing Na^+ and place it on the metal stand. Click the ***Flame*** button. Record your observations in the data table below. Click the ***Flame w/Cobalt*** button and record your observations in the same table.

7. Drag the K^+ test tube to the metal stand to exchange it with the Na^+. Flame test K^+ with and without cobalt glass. Record your observations in the table below.
8. For the other four ions, ***Flame*** test them only. Do not use cobalt glass. Record your observations in the table below.

Data Table

Flame Tests	**[Answers]**
Ion	Flame Color
sodium, Na^+	**yellow**
sodium, Na^+ (cobalt glass)	**none**
potassium, K^+	**violet and yellow**
potassium, K^+ (cobalt glass)	**violet**
calcium, Ca^+	**dark red**
barium, Ba^+	**green**
strontium, Sr^+	**bright red**
copper, Cu^+	**blue-green**
unknown #1	**response will vary with unknown**
unknown #2	**response will vary with unknown**
unknown #3	**response will vary with unknown**
unknown #4	**response will vary with unknown**

9. Flame test the practice unknown. Determine which of the six metal ions it most closely matches. You may repeat the flame test on any of the six metal ions if necessary. When you are confident that you have identified the unknown, open the *Lab Book* by clicking on it. On the left page, click the ***Report*** button. On the right page, click on the metal ion that you think is in the practice unknown. Click ***Submit*** and then ***OK***. If all of the ion buttons turn green you have successfully identified the unknown. If any turn red then you were incorrect. ***Flame*** test the practice unknown again to correctly identify your metal ion. Click on the red disposal bucket to clear all of your samples.
10. Return to the *Stockroom* and drag the practice unknown to the metal test tube stand. This will randomly select another metal ion. *Return to lab*, test and report this ion. Continue until you have correctly identified four different practice unknowns.

Analysis and Conclusions

1. The energy of colored light increases in the order red, yellow, green, blue, violet. List the metallic elements used in the flame tests in increasing order of the energy of the light emitted.

Sr^{2+} and Ca^{2+}	**Na^{+}**	**Ba^{2+} and Cu^{2+}**	**K^{+}**
red	**yellow**	**green**	**violet**
low energy	$\rightarrow$		**high energy**

2. What is the purpose of using the cobalt glass in the identification of sodium and potassium?

   Cobalt glass filters out yellow, allowing violet K^{+} to be seen. The glass helps to distinguish Na^{+} from K^{+}. In fact, you should be able to distinguish the two even if mixed together.

Photoelectric Effect

Purpose

To duplicate photoelectric effect experiments.

Background

Though Albert Einstein is most famous for $E = mc^2$ and his work in describing relativity in mechanics, his Nobel Prize was for understanding a very simple experiment. It was long understood that if you directed light of a certain wavelength at a piece of metal, it would emit electrons. In classical theory, the energy of the light was thought to be based on its intensity and not its frequency. However, the results of the photoelectric effect contradicted classical theory. These inconsistencies led Einstein to suggest that we need to think of light as being composed of particles (photons) and not just as waves.

Each wavelength corresponds to another property of light called frequency. You will use the wavelength (λ) value in the experiment to calculate the frequency (ν) given that $\nu = c/\lambda$ where c $= 2.998 \times 10^{17}$ nm/s (2.998×10^{8} m/s). The energy (E) of a quantum of light an atom emits is related to its frequency (ν) by $E = h\nu$ where h (Planck's constant) = 6.63×10^{-34} J-s.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Photoelectric Effect* from the list of assignments.
2. The experiment opens in the Quantum laboratory.
3. What source is used in this experiment and what does it do?

   The laser emits coherent light which is the same wavelength and in phase.

   At what intensity is the laser set?

   1 nW

   At what wavelength is the laser set?

   400 nm

   Record the wavelength (in nm) in the data table. Calculate the frequency (in 1/s) and the energy (in J) using the equations given in the Background section of this lab. Determine the color of the light by clicking on the Spectrum Chart (just behind the laser); the markers indicate what color is represented by the wavelength selected.

   Which metal foil is used in this experiment?

   Na, sodium metal foil

   What detector is used in this experiment and what does it measure?

   The phosphor screen detects electrons, and it glows momentarily at the positions where the electrons impact the screen.

   Turn on the detector by clicking on the red/green light switch. What does the signal on the phosphor screen indicate about the laser light shining on the sodium foil?

   The laser light is causing the ejection of electrons from the surface of the sodium metal.

4. Decrease the *Intensity* to 1 photon/second, how does the signal change?

   The signal is not as intense and flickers as each photon impacts the phosphor screen.

   Increase the *Intensity* to 1kW, how does the signal change?

   The signal is more intense than at 1 photon/second, but the same as at 1 nW.

5. Change the *Intensity* back to 1 nW and increase the *Wavelength* to 600 nm. What do you observe? Record the wavelength in the data table.

   The signal disappears.

   Determine the maximum wavelength at which emission of electrons occurs in the metal.

   450 nm.

   What is the difference between intensity and wavelength?

   Wavelength corresponds to the energy of light emitted but intensity is the amount of light.

   Which matters in the formation of photoelectrons?

Data Table

Wavelength (nm)	Frequency (1/s)	Energy (J)	Light Color
400	$\mathbf{7.50 \times 10^{14}}$	$\mathbf{4.97 \times 10^{-19}}$	violet
600	$\mathbf{5 \times 10^{14}}$	$\mathbf{3.32 \times 10^{-19}}$	orange
450	$\mathbf{6.66 \times 10^{14}}$	$\mathbf{4.42 \times 10^{-19}}$	blue

6. Click in the *Stockroom*. Click on the clipboard and select *Photoelectric Effect (2)*. *Return to Lab*. The intensity is set at 1 nW and the wavelength at 400 nm. The detector used in this experiment is a bolometer. Turn on the bolometer by clicking the red/green light switch.

   This instrument measures the kinetic energy of electrons emitted from the metal foil. You should see a green peak on the bolometer detection screen. The intensity or height of the signal corresponds to the number of electrons being emitted from the metal. The x-axis is the kinetic energy of the photoemitted electrons. Zoom in on the peak by clicking and dragging from half-way up the y-axis to the right past the peak and to the x-axis (it will draw an orange rectangle).

7. Increase and decrease the *Intensity*, what do you observe?

   The height of the peak changes directly proportional with intensity.

   Increase and decrease the *Wavelength*, what do you observe?

   As wavelength decreases the kinetic energy of the emitted electrons increases and when wavelength increases the kinetic energy of the emitted electrons decreases.

   What is the maximum wavelength that ejects electrons from the sodium metal?

   450 nm.

   From this experiment, explain why violet light causes photoemission of electrons but orange light does not.

   Violet light has a shorter wavelength, but more energy (4.97×10^{-19} J) than orange light (3.32×10^{-19}). Violet light has enough energy to eject electrons but orange light does not.

Diffraction Experiments

Purpose

To investigate the wave-particle duality of nature.

Background

It has long been known that if you shine light through narrow slits that are spaced at small intervals, the light will form a diffraction pattern. A diffraction pattern is a series of light and dark patterns caused by wave interference. The wave interference can be either constructive (light patterns) or destructive (dark patterns). In this experiment, you will shine a laser through a device with two slits where the spacing can be adjusted and investigate the patterns that will be made at a distance from the slits.

Procedure

1. Start *Virtual ChemLab*, open the *Workbook*, and select *Diffraction Experiments* from the list of assignments.
2. The lab will open in the Quantum laboratory.
3. What source is used in this experiment and for what reason?

   The laser, because it provides light at a single wavelength.

   What is the wavelength of the *Laser*?

   500 nm

   What is the spacing of the two slits on the two slit device?

   3.0 μm.

   Draw a picture of the pattern displayed on the video screen.

4. Change the *Intensity* of the *Laser* from 1 nW to 1W.

   Does the intensity of the light affect the diffraction pattern?

   No.

   Change the *Slit Spacing* to 1 μ m. Observe the pattern displayed on the video screen as you change the slit spacing from 1 μm to 7 μm by 1 μm increments.

   What can you state about the relationship between slit spacing and diffraction pattern?

   As the slit spacing increases, the number of lines in the diffraction pattern increases.

5. Increase the *Wavelength* of the *Laser* to 700 nm.

   What affect does an increase in the wavelength have on the diffraction pattern?

   Color changes from green to red. The number of lines also changes.

6. Decrease the *Intensity* on the *Laser* to 1000 photons/second. Click on the ***Persist*** button on the video camera to look at individual photons coming through the slits. Observe for one minute.

   What observation can you make about this pattern as compared to the pattern from the continuous beam of photons?

   The pattern is the same after allowing enough photons to diffract.

Decrease the *Intensity* to 100 photons/second. Observe for another minute after clicking ***Persist***. At these lower intensities (1000 and 100 photons/second), there is never a time when two photons go through both slits at the same time. How can a single photon diffract?

It cannot. What you see as the diffraction pattern builds up over time is really the statistics of where each individual photon will hit the screen. It is uncertain what each individual photon will do, but the properties of a large collection of photons can be easily predicted.

From this experiment, what conclusions can you make about the nature of light?

Light behaves like a particle, and the wave-nature of light is really a representation of the statistics or uncertainty exhibited in an experiment.

7. Click in the *Stockroom*. Click on *Clipboard* and select *Two-Slit Diffraction – Electrons*. *Return to Lab*. What source is used in this experiment?

   An electron gun.

   Draw a picture of the diffraction pattern shown on the *Phosphor Screen*.

   How does this diffraction pattern compare to the diffraction pattern for light?

   Patterns look similar.

   Louis de Broglie was the first person to suggest that particles could be considered to have wave properties.

8. Decrease the *Intensity* to 10 electrons/second. The pattern now builds one electron at a time. Click on the ***Persist*** button and observe for one minute.

   Has the diffraction pattern changed? Why or why not?

   No. Even though the pattern builds up one at a time, collectively they still form the same diffraction pattern.

   How can an electron diffract if there is only one?

   It cannot. What you see as the diffraction pattern builds up over time is once again really the statistics of where each individual electron will hit the screen. It is uncertain what each individual electron will do, but the properties of a large collection of electrons can be predicted. Thus, electrons also appear to have wave properties.

   When you look at the complete diffraction pattern of a stream of particles, you are seeing all the places you expect particles to scatter. If you start the source over multiple times, you will see that the first particle is never detected in the same place twice. This is an application of the Heisenberg Uncertainty Principle which is directly connected with measurement. It takes into account the minimum uncertainty of the position (Δx) and the uncertainty of the momentum (Δp). The equation that displays these variables is:

$$(\Delta x)(\Delta p) \geq h/4\pi$$

   Because you know the energy at which the particle is traveling, you can know precisely what the momentum is. But because of this equation you cannot be sure of the position.

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