Stoichiometry Balloon Races

Limiting and Excess Reactants

Introduction

Most stoichiometry calculations in the classroom are performed using exact (stoichiometric) mole ratios of reactants and products. In real life, however, many commercial processes for preparing compounds are carried out using an excess amount of one reactant (and thus a limiting amount of another reactant). This demonstration uses the well-known reaction of sodium bicarbonate and acetic acid to illustrate the concepts of limiting and excess reactants. By comparing the amount of carbon dioxide generated when varying amounts of sodium bicarbonate react with a given amount of acetic acid, students will be able to identify immediately the limiting and excess reactant in each case.

Concepts

- Stoichiometry
- Mole ratio
- Limiting reactant
- Excess reactant

Materials

Acetic acid, CH_3COOH, 1 M, 150 mL
Sodium bicarbonate, $NaHCO_3$, 10.5 g
Balance, centigram (0.01-g) precision
Balloons, 6
Erlenmeyer flasks, 125-mL, 6
Powder
Funnel
Graduated cylinder, 25- or 50-mL
Permanent marker
Spatula
Weighing dishes, 6

Safety Precautions

Acetic acid is a skin and eye irritant. Avoid contact with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please consult current Safety Data Sheets for additional safety information.

Preparation

1. Label six Erlenmeyer flasks 1–6. Using a graduated cylinder, add 25 mL of 1 M acetic acid to each flask.
2. Obtain six weighing dishes and label them 1–6.
3. Measure the appropriate amount of sodium bicarbonate into each weighing dish, according to the following table.

Table 1.

Sample	1	2	3	4	5	6
Mass $NaHCO_3$	0.50 g	1.00 g	1.50 g	2.00 g	2.50 g	3.00 g

4. Obtain six balloons. Stretch the balloons and blow them up at least once, then let out as much air is possible.

5. Use a powder funnel to add the first sodium bicarbonate sample (#1) to one of the balloons.
6. Flatten out the balloon to remove any extra air and then carefully stretch the neck of the balloon over the mouth of Erlenmeyer flask #1. Do not allow the solid to drop into the flask at this time.
7. Repeat steps 5 and 6 with the other sodium bicarbonate samples #2–6.

Procedure

1. Introduce the concepts of limiting and excess reagents by asking students how many automobiles can be assembled if the following parts are available: 140 car bodies, 520 tires, and 270 headlights.
2. Which automobile part is present in a quantity that "limits" the total number of automobiles that may be assembled? Is the limiting part the same as the part that is present in the least number? Explain. *Note:* The tires are the limiting parts in this example, even though there are more tires than anything else.
3. Show students the balloon/flask assemblies and ask them to predict what will happen when the sodium bicarbonate is added to the acetic acid in the flask. Write the reaction equation on the board.
4. Line up flasks 1–6 from right to left on the lecture desk. Lift each balloon in turn and shake it to allow the solid to fall into the solution. Make sure the neck of the balloon stays firmly attached to the flask.
5. The reactions will be immediate and vigorous. The white solids will dissolve, the solutions will start to bubble and fizz, and the balloons will become inflated.
6. Allow the reactions to proceed until the bubbling stops. Compare the size of the inflated balloons and whether all the solid has dissolved in each case. *(The balloon size should increase fairly uniformly for flasks 1–5, and then stay constant. It may be hard to tell the difference between flasks 4, 5, and 6.)*
7. Discuss the observations and carry out the necessary calculations to explain the results (See Table 2). Identify the limiting reactant and the excess reactant in each case.

Table 2.

Flask	Acetic Acid		Sodium Bicarbonate		Moles CO_2 Produced (Theoretical)
	Volume	Moles	Mass	Moles	
1	25 mL	0.025	0.50 g	0.0060	0.0060
2	25 mL	0.025	1.00 g	0.0119	0.0119
3	25 mL	0.025	1.50 g	0.0179	0.0179
4	25 mL	0.025	2.00 g	0.0238	0.0238
5	25 mL	0.025	2.50 g	0.0298	0.0250
6	25 mL	0.025	3.00 g	0.0357	0.0250

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. All of the leftover solutions may be disposed of down the drain with excess water according to Flinn Suggested Disposal Method #26b.

Tips

- Use new balloons each time the demonstration is performed. Once used, the balloons will have been stretched, and their inflated sizes may vary considerably from the expected size.
- Unreacted (undissolved) sodium bicarbonate will be visible in flasks 5 and 6. The solid, although normally water soluble, does not dissolve in the saturated carbon dioxide solutions.
- Add a few drops of universal indicator to each flask and observe the color changes as the pH of the solution changes over the course of the reaction.
- Excess reactants are used commercially in cases where reactions are reversible and thermodynamically unfavorable. An example of a gas-phase inorganic reaction is the synthesis of ammonia from nitrogen and hydrogen. Excess nitrogen is used to drive the reaction to completion. The synthesis of organic esters from organic alcohols and acids is another example of commercial processes that are normally carried out in the presence of excess reactants.
- Carry out the demonstration using sodium carbonate instead of sodium bicarbonate. The maximum amount of CO_2 evolution will be observed at a lower mass of sodium carbonate, due to the 2:1 mole ratio for reaction of acetic acid with sodium carbonate.

Discussion

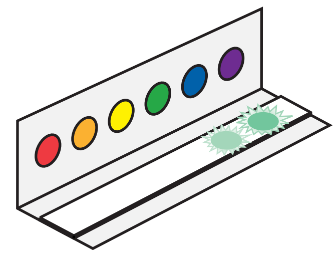
Equation 1 summarizes the reaction between sodium bicarbonate and acetic acid.

$$NaHCO_3(s) + CH_3CO_2H(aq) \rightarrow NaCH_3CO_2(aq) + CO_2(g) + H_2O(l) \qquad \textit{Equation 1}$$

Sodium bicarbonate is the limiting reactant in Flasks 1–4. In Flasks 5 and 6, acetic acid is the limiting reactant.

Energy in Photons
Light Energy Demonstration

Introduction



Students often confuse the concepts of intensity of light and energy of light. This demonstration provides a clear way to demonstrate that the intensity, or brightness, of light is NOT the same as the amount of energy a particular color of light possesses.

Concepts

- Phosphorescence
- Photoelectric effect
- Absorbance vs. transmittance
- Energy and wavelength of light

Materials

Energy in Photons Demonstrator Card—assembled and ready to use

Light source—classroom lights work well

Procedure

1. Open the Demonstrator Card and show the class the phosphorescent strip. Explain that the common term for phosphorescence is "glow-in-the-dark" and that this strip will glow in the dark.
2. Directly expose the entire phosphorescent strip to the classroom lights for about 15 seconds. Now turn off all the classroom lights and completely darken the room. The entire strip will glow brightly for several minutes and then begin to fade. Once students are satisfied with the glow, turn the classroom lights back on.
3. Show the six colored filters on the Demonstrator Card to the class. Hold the Demonstrator Card up to the light so that the color of light transmitted through each filter is clearly visible. Observe that the color of light transmitted through each filter is the same color as the filter.
4. Close the Demonstrator Card tightly making sure that no light can reach the phosphorescent strip from the sides. Paper clip the sides closed and place it top down on the desk for 2–3 minutes. Have the class predict what will happen if the Demonstrator Card is closed and only "filtered" light is allowed to shine upon the phosphorescent strip.
5. After 2–3 minutes, expose the closed Demonstrator Card to the classroom lights for at least 30 seconds.
6. Turn off all the classroom lights and completely darken the room again. Open the Demonstrator Card and show the phosphorescent strip to the class. The strip will only glow under the blue and violet filters!

7. Compare the class' predictions with the actual results. Many students will be surprised that the brighter colors, like yellow and orange, do not let enough light through to cause the strip to glow. Explain that even though these colors may look brighter, or more intense, only the blue and violet filters let through light with enough energy to make the phosphorescent strip glow.

8. Have the class estimate the maximum wavelength needed to excite the phosphorescent strip (and cause it to glow) by using an approximate wavelength for each filter color. *(The maximum wavelength is about 480 nanometers.)*

9. Calculate the minimum energy a photon must have to cause the strip to phosphoresce, or glow:

$$E = hc/\lambda$$

where E = energy in joules, h = Planck's constant = 6.626×10^{-34} J•sec, c = speed of light = 2.998×10^{8} m/sec, and λ = wavelength in meters.

The minimum photon energy required for phosphorescence is 4.1×10^{-19} J.

$$E = \frac{(6.626 \times 10^{-34}\ \text{J•sec})(2.998 \times 10^{8}\ \text{m/sec})}{(480\ \text{nm})(1 \times 10^{-9}\ \text{m/nm})} = 4.1 \times 10^{-19}\ \text{J}$$

Tips

- The spots under the blue and violet filters should always glow brightly, but sometimes a small glow can be seen under some of the other filters. This is partly due to stray light creeping in around the filters. Try to keep the Demonstrator Card as tightly closed as possible to prevent any extra light from getting to the phosphorescent strip. The slight glow can also be explained by noting that the filters are not "perfect" filters in removing lights of other wavelengths.

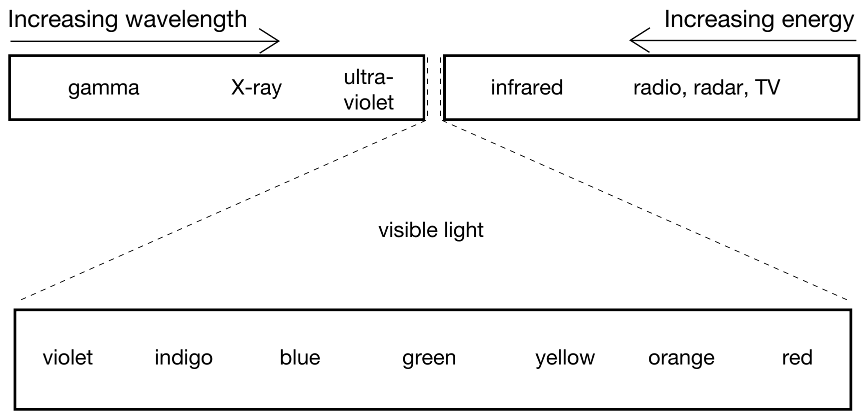
- To store the Demonstrator Card, insert the solid black strip into the Card so that it is covering the phosphorescent strip. Close the Demonstrator Card and store it in the envelope. This will protect the phosphorescent strip from light and prolong its useful life.

- To reinforce the concept that it is the energy of the light that matters rather than the intensity, obtain two light sources of different intensity, such as a 40 W bulb and a 100 W bulb. Shine each lightbulb onto the closed Demonstration Card—in both cases only the violet and blue filters will cause the strip to phosphoresce.

Discussion

Visible light is a form of electromagnetic radiation. All forms of electromagnetic radiation consist of oscillating electric and magnetic fields traveling at a constant speed, the speed of light, 2.998×10^{8} m/s. Other familiar forms of electromagnetic radiation include γ-rays, X-rays, ultraviolet radiation, infrared radiation, and radio waves.

Electromagnetic Spectrum



The visible portion of the electromagnetic spectrum spans the wavelength region from about 400 to 700 nm. The human eye sees light of 400 nm as violet and 700 nm as red. Based on the quantum nature of light, the wavelength of light is inversely proportional to its energy (E) according to Planck's Law, $E = hc/\lambda$, where h is Planck's constant and c is the speed of light. Because it has a shorter wavelength than red light, violet light is greater in energy than red light. As the color of light changes, so does the amount of energy it possesses. White light, such as that from an incandescent or fluorescent light, contains all of the colors in the visible spectrum.

A second characteristic of light, in addition to its energy, is its intensity. Intensity corresponds to the brightness of light. According to the theories of classical physics, energy should be proportional to intensity, so that the more intense a light source, the more energy it gives off. Under this assumption, very bright (yellow) light should cause the phosphorescent strip in the Demonstrator Card to glow. However, this is not observed! Instead, the phosphorescent strip glows only when blue or violet light is shined on it. This phenomenon is analogous to the photoelectric effect, one of the classical paradoxes that led to the discovery of the quantum nature of light and Planck's Law.

The glowing of the phosporescent strip in the Demonstrator Card is due to two processes that occur in tandem. First, the phosphor must absorb light of the proper energy to excite an electron from its lowest energy ground state to a higher energy excited state. Once in the excited state, the electron has a natural tendency to want to return back to its more stable ground state. In doing so it releases energy in the form of light—this is the emission of photons that is observed as "glow-in-the-dark" phosphorescence. The phosphorescent material has a critical wavelength or energy of light. If a light is shined on the phosphorescent strip and it contains photons whose energy is greater than the energy needed to cause the strip to glow, it will glow. If the intensity of this source is increased, the strip will glow more brightly. If, however, a light source is shined on the phosphorescent strip that contains photons whose energy is less than the critical energy for the phosphorescent strip, no glowing will occur, no matter how bright the light source.

The cutoff wavelength for exciting the phosphorescent strip is about 480 nm. Photons (light energy) with a higher wavelength (less energy) will not cause the strip to glow, while photons with shorter wavelengths (more energy) will cause the phosphorescent glow. The cutoff wavelength of 480 nm is right on the border between blue and green light. Therefore, blue or violet photons which are transmitted through the filter will contain enough energy to excite the electrons in the phosphorescent material and cause it to glow. The green, yellow, orange and red filters, in contrast, absorb the blue and violet photons (their complementary colors) and do not allow them to be transmitted. Therefore, the light transmitted through these filters does not contain enough energy to excite the phosphorescent strip and no glow is observed.

The following table lists the wavelengths associated with each of the filter colors and their complements.

Representative Wavelength, nm	Wavelength Region, nm	Filter Color	Complementary Color
410	400–425	Violet	Yellow-green
470	425–480	Blue	Orange
490	480–500	Blue-green	Red
520	500–560	Green	Red-Violet
565	560–580	Yellow-green	Violet
580	580–585	Yellow	Violet
600	585–650	Orange	Blue
650	650–700	Red	Blue-green

Upset Tummy? MOM to the Rescue!
Colorful Antacid Demonstration

Introduction

Mix milk of magnesia (MOM) with universal indicator and observe the dramatic spectrum of color changes as the antacid dissolves in simulated stomach acid! This is a great demonstration to illustrate acid–base neutralization, solubility, and "antacid-testing" consumer chemistry.

Concepts

- Acid–base neutralization
- Solubility
- Antacids

Materials

Milk of magnesia, 20 mL
Hydrochloric acid, HCl, 3 M, ~ 20 mL
Universal indicator, 4–5 mL
Water, distilled or deionized, 800 mL
Beaker, 1-L (or other large beaker)
Beral-type pipets, 2
Graduated cylinder, 25-mL or 50-mL
Ice, crushed (or ice cubes)
Magnetic stir plate (or stirring rod)
Magnetic stir bar

Safety Precautions

Milk of magnesia is intended for laboratory use only; it has been stored with other non–food-grade laboratory chemicals and is not meant for human consumption. Hydrochloric acid solution is toxic by ingestion and inhalation and is corrosive to skin and eyes. Universal indicator solution is an alcohol-based flammable solution. Consult current Safety Data Sheets for further safety and handling techniques. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron.

Procedure

1. Measure 20 mL of milk of magnesia using a graduated cylinder and pour it into a 1-L beaker.
2. Place the 1-L beaker on a magnetic stir plate. Add a magnetic stir bar to the beaker.
3. Add water and crushed ice (or ice cubes) to give a total volume of approximately 750 mL. Turn on the stir plate so as to create a vortex in the mixture.
4. Add about 5 mL (2-pipets full) of universal indicator solution. Watch as the white suspension of milk of magnesia turns to a deep purple color. The color indicates that the solution is basic.
5. Add 2–3 mL (1-pipet full) of 3 M HCl. The mixture quickly turns red and then goes through the entire range of universal indicator color changes back to purple.
6. Repeat this process, adding HCl one-pipet full at a time, waiting after each addition until the mixture turns back to blue–purple.

7. The process can be repeated a number of times before all of the $Mg(OH)_2$ dissolves and reacts with the HCl. As more acid is added, the color changes begin to occur more slowly and eventually the suspension completely dissolves. The final solution will be clear and red.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The final solution may be neutralized according to Flinn Suggested Disposal Method #24b. Excess milk of magnesia can be disposed of according to Flinn Suggested Disposal Method #26a.

Tips

- The recommended acid concentration in this demonstration is 3 M, in order to allow the reaction to go to completion with a reasonable volume of acid.
- Adding ice to the demonstration slows down the reaction and results in longer-lasting colors. As the ice melts, the color changes speed up.

Discussion

The active ingredient in milk of magnesia is magnesium hydroxide, $Mg(OH)_2$. Magnesium hydroxide forms a suspension in water since it has a very low solubility—0.0009 g/100 mL in cold water and 0.004 g/100 mL in hot water.

Initially in the demonstration, the solution is basic due to the small amount of $Mg(OH)_2$ that goes into solution. Universal indicator gives the mixture a violet color, indicating a pH of about 10. (See Universal Indicator Color Chart below.) Upon addition of hydrochloric acid (the simulated "stomach acid"), the mixture quickly turns red as the acid disperses throughout the beaker and neutralizes the small amount of dissolved $Mg(OH)_2$. A small amount of excess acid is present and solution is acidic ($pH < 4$).

Universal Indicator Color Chart							
Color	Red	Orange	Yellow	Green	Green–blue	Blue	Violet
pH	4	5	6	7	8	9	10

Excess acid, however, causes more $Mg(OH)_2$ from the suspension to dissolve. As more of the $Mg(OH)_2$ goes into solution, it neutralizes the excess acid, and the solution reverts back to its purple, basic color. The addition of universal indicator allows this entire sequence to be observed. During the process, the color of the mixture cycles through the entire universal indicator color range—from red to orange to yellow to green to blue and finally back to violet. By adding more "stomach acid," the process can be repeated many times before all of the $Mg(OH)_2$ dissolves and is neutralized. The final solution, after all the $Mg(OH)_2$ has dissolved and reacted with HCl, is clear and red.

Foiled Again!

Single Replacement Reaction

Introduction

Watch aluminum foil disappear as it is added to a solution of copper(II) chloride. Observe color changes, production of a gas, formation of a solid metal, and a drastic change in temperature. Learn about the unexpected role of a catalyst in this single-replacement reaction at a metal surface.

Concepts

- Single replacement reaction
- Metal activity
- Oxidation–reduction
- Catalysis

Materials

Aluminum foil, 6″ × 12″, 2 pieces
Copper(II) chloride solution, $CuCl_2$, 1 M, 140 mL
Copper(II) sulfate solution, $CuSO_4$, 1 M, 140 mL
Sodium chloride solution, NaCl, 1 M, 140 mL
Water, distilled or deionized
Beakers, Pyrex®, 600-mL, 3
Graduated cylinder, 500-mL
Spatula
Stirring rod
Thermometer
Wood splint and matches (optional)

Safety Precautions

Copper(II) chloride solution is toxic and copper(II) sulfate solution is slightly toxic by ingestion. Hydrogen gas, a highly flammable gas, is produced in the reaction. Keep flammable materials away from the demonstration area. Wear chemical splash goggles and chemical-resistant gloves and apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

Part 1 — Aluminum and Copper(II) Chloride

1. Place a 600-mL Pyrex® beaker (or a 500-mL graduated cylinder) on the demonstration table.
2. Use a graduated cylinder to measure 140 mL of 1 M copper(II) chloride solution. Pour this into the beaker.
3. Measure and add 140 mL of distilled or deionized water to the beaker. The solution is now 0.5 M $CuCl_2$.
4. Cut a piece of aluminum foil approximately 6″ × 12″. Loosely roll the foil into a cylinder that will fit into the beaker. *Note:* Do not wad up the foil tightly into a ball—this will decrease the surface area and slow down the reaction.
5. If desired, measure the temperature of the solution before adding the foil.

6. Place the aluminum foil cylinder into the beaker, using a stirring rod to push it down completely into the solution. Measure the temperature of the reaction mixture again. Notice the great increase in temperature—the reaction is highly exothermic.

7. Have students make detailed observations of the reaction and ask them to generate a hypothesis for the reaction(s) in the beaker. Write an equation for the reaction(s) they observe. Discuss which substances are reacting species and which are spectators in the reaction, if any.

8. Students may hypothesize that aluminum reacts with copper(II) ions to form solid copper and aluminum ions. Test this hypothesis by performing Part 2, in which aluminum is again mixed with copper(II) ions, but this time from a different source, a copper(II) sulfate solution. *Note:* Set the beaker from Part 1 aside for comparison.

Part 2 — Aluminum and Copper(II) Sulfate

9. Repeat the procedure (steps 1–6) from Part 1 in a different 600-mL beaker, except this time using 70 mL of 1 M $CuSO_4$ solution and 70 mL of distilled or deionized water. The solution is now 0.5 M $CuSO_4$.

10. Have students again make detailed observations. Students will observe that no reaction occurs between aluminum and copper(II) sulfate. Why not? Have students make modifications to this original hypothesis and generate a new hypothesis. Students at this point may propose trying various experiments to test their hypotheses, so additional materials may be needed. *Note:* Set the beaker from Part 2 aside for use in Part 3.

Part 3 — Aluminum and Copper(II) Sulfate with Sodium Chloride

11. Place the beaker from Part 2 on the demonstration table.

12. Add 70 mL of 1 M NaCl solution to the beaker.

13. Have students make detailed observations. Notice that a reaction now occurs between the aluminum and the copper(II) ions, as in Part 1. Discuss what is occurring in the beaker and write the chemical equation for the reaction. If chloride ions are not in the equation, what is the purpose of the chloride ions? Discuss the role of a catalyst in a reaction.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. Allow the solid material in the beakers to settle. Decant the copper(II) chloride and copper(II) sulfate solutions down the drain according to Flinn Suggested Disposal Method #26b. Dispose of the solid copper and leftover aluminum foil according to Flinn Suggested Disposal Method #26a.

Tips

- Test for hydrogen gas with a burning splint: Light a wood splint and hold it in the beaker over the bubbles as they are released from the reaction. A positive test is indicated if a pop or a barking sound is heard. Notice that hydrogen gas is released from the reaction. Discuss the origin of the gas. Write the chemical equation for the production of hydrogen gas.
- The reaction temperature increases from room temperature (25 °C) to nearly 60 °C when aluminum is added to copper(II) chloride. Therefore, be sure to perform the demonstration in Pyrex® glassware.
- Sodium chloride catalyzes the reaction between aluminum and copper(II) sulfate. Observations of the reaction with the catalyst are the same as those observed when aluminum and copper(II) chloride are allowed to react. Other catalysts that have been found to catalyze this reaction include sodium bromide, potassium chloride, magnesium chloride, and hydrochloric acid. It has also been found that copper(II) nitrate reacts in the same way as copper(II) sulfate. As an extension to this demonstration, consider trying other catalysts or other copper solutions.
- Heating the solution of copper(II) sulfate and aluminum to 80 °C does not cause any reaction to occur. The chloride catalyst is necessary for the reaction to occur.
- A more dramatic demonstration display can be achieved by performing the demonstration in large 500-mL Pyrex® graduated cylinders or hydrometer cylinders. Loosely coil the foil into a tube and drop it into the cylinder of $CuCl_2$. The foil will slowly rise in the cylinder as gas bubbles attach to the foil surface. Most of the reaction will occur at the top of the solution (which will turn gray or colorless) and there will be unreacted green-blue copper solution at the bottom of the cylinder. A long stirring rod will be useful in pushing the foil down to the bottom.
- Try using tap water in place of distilled water in the dilution of the $CuSO_4$ solution. Are there enough dissolved chloride ions in tap water to catalyze the reaction?
- The "Leftover Aluminum Wire Stoichiometry Lab" is available as a student laboratory kit from Flinn Scientific (Catalog No. AP4678). This experiment uses the same reaction to teach students about moles, limiting reactants, and stoichiometry.

Discussion

Aluminum foil reacts with an aqueous solution of copper(II) chloride according to Equation 1. The reaction may be classified as a single replacement, oxidation–reduction reaction.

$$\underset{\textit{Silver}}{2Al(s)} + \underset{\textit{Green-blue}}{3CuCl_2(aq)} \rightarrow \underset{\textit{Colorless}}{2AlCl_3(aq)} + \underset{\textit{Red}}{3Cu(s)} \qquad \textit{Equation 1}$$

The oxidation of aluminum metal to aluminum(III) ions (Al^0 to Al^{3+}) is inferred from the dissolving of the aluminum foil and is represented by the oxidation half-reaction below. The simultaneous reduction of copper(II) ions to copper metal (Cu^{2+} to Cu^0) results in the formation of solid copper metal according to the reduction half-reaction below. As copper(II) ions are reduced to copper, the green-blue color of the solution fades until it is colorless—the indication that the reaction is complete and all of the copper(II) ions have been reduced.

$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$	Oxidation Half-Reaction
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	Reduction Half-Reaction
$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$	Overall Balanced Equation

It is observed that hydrogen gas is simultaneously released from the reaction when aluminum metal foil is added to copper(II) chloride solution. If the pH of the copper(II) chloride solution is measured, it is found to be slightly acidic. Hence there are free hydrogen ions in solution, which cause the side reaction of hydrogen ions with the aluminum surface to form hydrogen gas and aluminum ions (See Equation 2). Due to the limited concentration of hydrogen ions, this reaction consumes only a small amount of the aluminum.

$$2Al(s) + 6\,H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g) \qquad \textit{Equation 2}$$

Why does aluminum react with copper(II) chloride and not with copper(II) sulfate or copper(II) nitrate? And why does it then react with copper(II) sulfate when chloride ions are added?

This interesting phenomenon was discovered serendipitously by a teacher in the classroom. The teacher wanted to perform the aluminum plus copper(II) chloride reaction with the class, but had run out of copper(II) chloride. Thinking that the net ionic equation should be the same with copper(II) sulfate, the teacher mixed aluminum with cupric sulfate. Nothing happened! After a few grams of salt were added, however, the reaction took off. This was unpredicted and raised a series of questions.

The first possible explanation is that the chloride ions catalyze the reaction at the metal surface. Other halide salts were tried and it was found that bromide ions also catalyzed the reaction. The second possible explanation has to do with the aluminum oxide coating. Normally, a chemically inert coating of aluminum oxide on the surface of aluminum metal protects the metal from oxidation. In the presence of chloride ions, however, the coating is breached and the underlying aluminum reacts. Because chloride is a relatively small ion, it is able to diffuse into and through the protective metal oxide coating. Aluminum chloride, which is more soluble than aluminum oxide, can then form. This aluminum chloride salt leaches back through the oxide coating, and a path is now open for copper(II) ions to attack the underlying metal. The reaction is rapid and extremely exothermic. A third possible explanation for the observations in this demonstration involves electrochemical potentials. The copper–chloride electrode (with copper(II) chloride) has a greater half-cell potential than the copper–copper(II) electrode (with copper(II) sulfate). This supports the observation that the reaction with chloride ions is more favorable than without chloride ions. This experiment is an excellent open-ended demonstration which, like any good inquiry-based lesson, leads to more questions than answers.

Sample Procedure

1) Wear goggles!
2) Nest two Styrofoam cups together, use the graduated cylinder to measure and add 25.0 mL of water.
3) Place the digital thermometer in the cup, measure and record initial temperature.
4) Stir and add pre-weighed amount of ammonium nitrate.
5) Measure and record lowest temperature that is reached.
6) Dispose of solution in plastic waste cup.

Sample Data Table

Trial*	Mass of solid	Volume of water	Mass of Soln†	$T_{initial}$ ºC	T_{final} ºC	ΔT ºC
1						
2						

***Each pair of teachers conducts one trial. Share Trial 2 data with second pair of teachers working from same "kit box."**

Flinn Scientific POGIL™ Activity

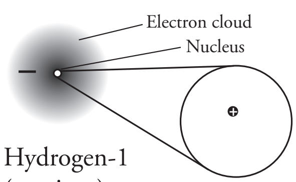
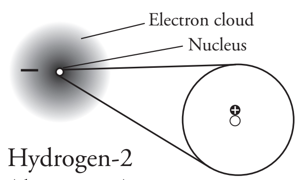
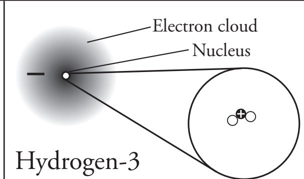
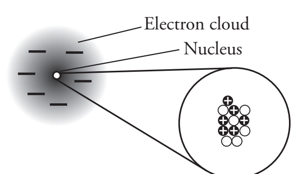
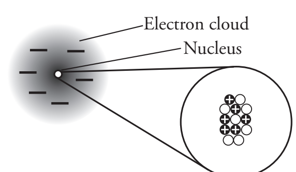
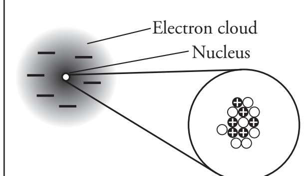
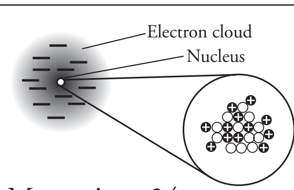
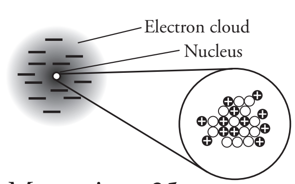
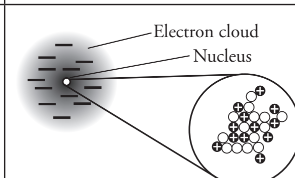
Isotopes

Are all atoms of an element alike?

Why?

The following activity will help you learn the important structural characteristics of an atom. How do we classify atoms? How does the combination of subatomic particles affect the mass and charge of an atom? What are isotopes? This is just a sampling of what we will address. Throughout this activity you will want to keep both Model 1 and a periodic table handy.

Model 1

Isotopes of Hydrogen			
Symbol	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
Atomic Diagram with Name	 Hydrogen-1 (protium)	 Hydrogen-2 (deuterium)	 Hydrogen-3 (tritium)
Number of Protons ⊕			
Number of Neutrons ○			
Isotopes of Carbon			
Symbol	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
Atomic Diagram with Name	 Carbon-12	 Carbon-13	 Carbon-14
Number of Protons ⊕			
Number of Neutrons ○			
Isotopes of Magnesium			
Symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
Atomic Diagram with Name	 Magnesium-24	 Magnesium-25	 Magnesium-26
Number of Protons ⊕			
Number of Neutrons ○			

1. Refer to Model 1. What subatomic particles do the following symbols represent in the Atomic Diagrams?

   ━ ⊕ ○

2. Complete the table in Model 1 by counting the protons and neutrons in each atomic diagram. Divide the work evenly among group members.

3. Find the three elements shown in Model 1 on your periodic table.

   a. What whole number shown in Model 1 for each element is also found in the periodic table for that element?

   Hydrogen — Carbon — Magnesium —

   b. The whole number in each box of the periodic table is the atomic number of the element. What does the **atomic number** of an element represent?

   c. Refer to the isotope symbols in Model 1. Relative to the atomic symbol (H, C, or Mg), where is the atomic number located in the isotope symbol?

4. Refer to your periodic table.

   a. How many protons are in all chlorine (Cl) atoms?

   b. A student says "I think that some chlorine atoms have 16 protons." Explain why this student is not correct.

5. Refer again to Model 1. In the isotope symbol of each atom, there is a superscripted (raised) number. This number is also used in the name of the atom (*i.e.,* carbon-12). It is called the **mass number**.

   a. How is the mass number determined?

   b. Why is this number called a "mass" number?

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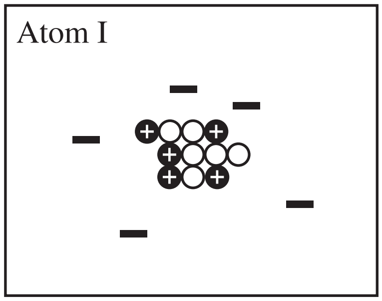
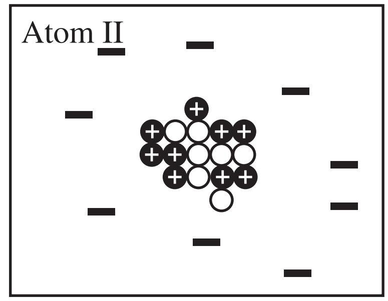
6. Fill in the table for Atom I and Atom II shown below.

	Atom I	Atom II
Number of Protons		
Number of Neutrons		
Mass Number		





7. Refer to Model 1.
   a. Which corner of the isotope symbol contains the mass number?

   b. How is the mass number of an isotope expressed in the name of an atom?

8. Write an isotope symbol (similar to those in Model 1) for each of the atoms in Question 6.

9. Write the name of the atom (similar to those in Model 1) for each of the atoms in Question 6.

10. Fill in the following table.

Isotope Symbol	$^{40}_{19}K$	$^{18}_{9}F$	
Atomic Number			16
Mass Number			
Number of Protons			
Number of Neutrons			15

11. Consider the examples in Model 1.
    a. Do all isotopes of an element have the same atomic number? Give at least one example or counter-example from Model 1 that supports your answer.

b. Do all isotopes of an element have the same mass number? Give at least one example or counter-example from Model 1 that supports your answer.

12. Considering your answers to Question 11, write a definition of **isotope** using a grammatically correct sentence. Your group must come to consensus on this definition.

13. Consult the following list of isotope symbols: ${}^{204}_{82}Pb$, ${}^{82}_{35}Br$, ${}^{78}_{35}Br$, ${}^{208}_{82}Pb$, ${}^{204}_{78}Pt$, ${}^{205}_{82}Pb$.

    a. Which of the atoms represented by these symbols are isotopes of each other?

    b. Which part(s) of the isotope symbol was the most helpful in answering part *a* of this question?

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Extension Questions

14. Determine the number of electrons in each of the atomic diagrams in Model 1.
    a. In a neutral atom, how does the number of electrons compare to the number of protons?

    b. Discuss why this relationship is important in making a “neutral” atom.

15. Refer to the hydrogen isotopes in Model 1. Each isotope has a special name derived from Latin (protium, deuterium, and tritium). What structural feature do these names refer to in the atom?

16. Can two atoms with the same mass number ever be isotopes of each other? Explain.

17. All models have limitations. What characteristics of Model 1 are inconsistent with your understanding of what atoms look like?

Teacher's Notes
Isotopes

Are all atoms of an element alike?

Learning Objectives

1. Determine the number of protons and neutrons in an atom based on the isotope symbol.
2. Describe the similarities and differences in isotopes of an element.

Prerequisites

1. Students should be able to name the three subatomic particles, their charges, their general location in the atom (nucleus or outside the nucleus), and which of them have substantial mass.
2. Students should have some familiarity with atomic symbols and be able to find them on a periodic table.

Assessment Questions

1. A neutral atom has 14 protons and 18 neutrons. Choose the correct nuclide symbol for this atom.

   a. ${}^{32}_{18}Ar$ *b.* ${}^{32}_{14}Si$ *c.* ${}^{18}_{14}Si$ *d.* ${}^{14}_{32}Ge$

2. Which of the following pairs show two atoms with the same number of neutrons?

   a. ${}^{37}_{17}Cl$ and ${}^{38}_{18}Ar$ *c.* ${}^{59}_{27}Co$ and ${}^{61}_{27}Co$

   b. ${}^{32}_{15}P$ and ${}^{32}_{16}S$ *d.* ${}^{65}_{30}Zn$ and ${}^{67}_{30}Zn$

3. There are three stable isotopes of Argon: Argon-36, Argon-38, and Argon-40. What would the atoms of these isotopes have in common? What would be different about their atoms?

Assessment Target Responses

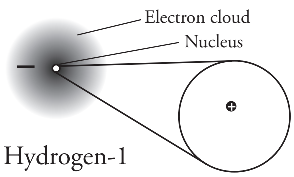
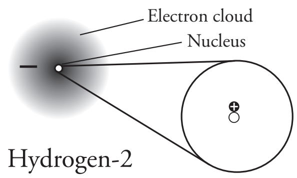
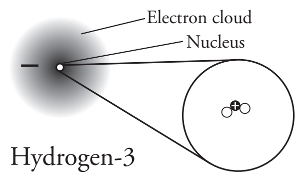
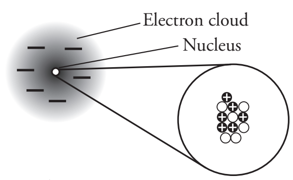
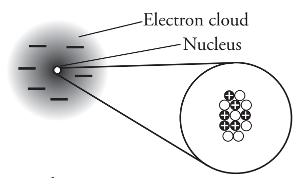
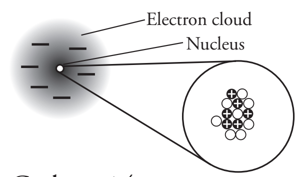
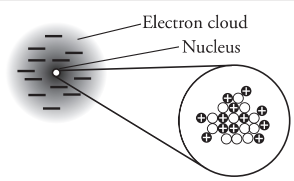
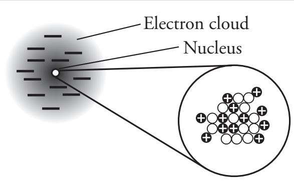
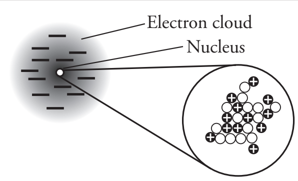
1. *b.*
2. *a.*
3. *All of these atoms will have 18 protons and 18 electrons (if they are all neutral). The three isotopes will have different numbers of neutrons (18, 20, and 22).*

Teacher Tips

- Students will need access to a periodic table throughout the activity.
- Many students have the misconception that there is a "normal" atom of some element, and the rest of the versions of that atom are isotopes. For example Carbon-12 is "normal" because it has 6 protons and 6 neutrons, while Carbon-13 is an isotope because it is unbalanced. The stop sign at Question 12 is a good place to emphasize that all versions of atoms for a particular element are isotopes.

Answers to Student Worksheet

Model 1

Isotopes of Hydrogen			
Symbol	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
Atomic Diagram with Name	 Hydrogen-1 (protium)	 Hydrogen-2 (deuterium)	 Hydrogen-3 (tritium)
Number of Protons ⊕	*1*	*1*	*1*
Number of Neutrons ○	*0*	*1*	*2*
Isotopes of Carbon			
Symbol	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
Atomic Diagram with Name	 Carbon-12	 Carbon-13	 Carbon-14
Number of Protons ⊕	*6*	*6*	*6*
Number of Neutrons ○	*6*	*7*	*8*
Isotopes of Magnesium			
Symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
Atomic Diagram with Name	 Magnesium-24	 Magnesium-25	 Magnesium-26
Number of Protons ⊕	*12*	*12*	*12*
Number of Neutrons ○	*12*	*13*	*14*

1. Refer to Model 1. What subatomic particles do the following symbols represent in the Atomic Diagrams?

▬	⊕	○
Electron	*Proton*	*Neutron*

2. Complete the table in Model 1 by counting the protons and neutrons in each atomic diagram. Divide the work evenly among group members.

   See Model 1.

3. Find the three elements shown in Model 1 on your periodic table.

   a. What whole number shown in Model 1 for each element is also found in the periodic table for that element?

   Hydrogen — *1*　　Carbon — *6*　　Magnesium — *12*

   b. The whole number in each box of the periodic table is the atomic number of the element. What does the **atomic number** of an element represent?

   The number of protons in an atom.

   c. Refer to the isotope symbols in Model 1. Relative to the atomic symbol (H, C, or Mg), where is the atomic number located in the isotope symbol?

   The lower left corner of the isotope symbol.

4. Refer to your periodic table.

   a. How many protons are in all chlorine (Cl) atoms?

   17

   b. A student says "I think that some chlorine atoms have 16 protons." Explain why this student is not correct.

   If an atom had 16 protons, it would be a sulfur atom.

5. Refer again to Model 1. In the isotope symbol of each atom, there is a superscripted (raised) number. This number is also used in the name of the atom (*i.e.,* carbon-12). It is called the **mass number**.

   a. How is the mass number determined?

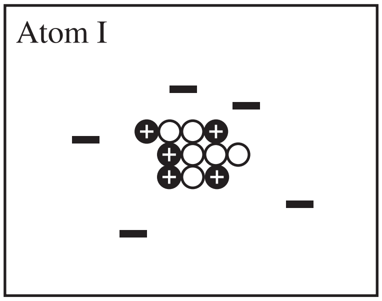
   Add the protons and the neutrons together.

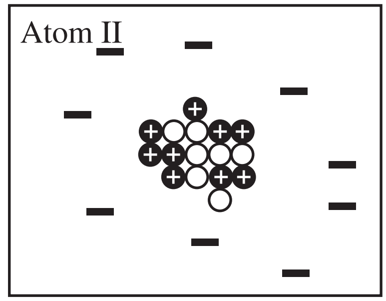
   b. Why is this number called a "mass" number?

   The protons and neutrons are the things in the atom with mass. When you have counted them all up you have approximately the mass of the atom in atomic mass units (amu).

6. Fill in the table for Atom I and Atom II shown below.

	Atom I	**Atom II**
Number of Protons ⊕	*5*	*9*
Number of Neutrons ○	*6*	*7*
Mass Number	*11*	*16*





7. Refer to Model 1.

   a. Which corner of the isotope symbol contains the mass number?

   The mass number is in the upper left corner of the isotope symbol.

   b. How is the mass number of an isotope expressed in the name of an atom?

   The mass number is shown as a numeral following the name of the atom, separated by a hyphen. Examples: carbon-12 or hydrogen-3).

8. Write an isotope symbol (similar to those in Model 1) for each of the atoms in Question 6.

   $^{11}_{5}B$ $^{16}_{9}F$

9. Write the name of the atom (similar to those in Model 1) for each of the atoms in Question 6.

   Boron–11 *Fluorine–16*

10. Fill in the following table.

Isotope Symbol	$^{40}_{19}\mathrm{K}$	$^{18}_{9}\mathrm{F}$	$^{31}_{16}S$
Atomic Number	*19*	*9*	16
Mass Number	*40*	*18*	*31*
Number of Protons	*19*	*9*	*16*
Number of Neutrons	*21*	*9*	15

11. Consider the examples in Model 1.

    a. Do all isotopes of an element have the same atomic number? Give at least one example or counter-example from Model 1 that supports your answer.

    Yes, all carbon atoms have an atomic number of 6.

b. Do all isotopes of an element have the same mass number? Give at least one example or counter-example from Model 1 that supports your answer.

No, isotopes must have different mass numbers. Examples: carbon-12 and carbon-14.

12. Considering your answers to Question 11, write a definition of **isotope** using a grammatically correct sentence. Your group must come to consensus on this definition.

Isotopes are atoms of the same element (same atomic number) with different numbers of neutrons (different mass numbers).

13. Consult the following list of isotope symbols: ${}^{204}_{82}Pb$, ${}^{82}_{35}Br$, ${}^{78}_{35}Br$, ${}^{208}_{82}Pb$, ${}^{204}_{78}Pt$, ${}^{205}_{82}Pb$.

a. Which of the atoms represented by these symbols are isotopes of each other?

${}^{204}_{82}Pb$ ${}^{205}_{82}Pb$ ${}^{208}_{82}Pb$ ${}^{82}_{35}Br$ ${}^{78}_{35}Br$

b. Which part(s) of the isotope symbol was the most helpful in answering part *a* of this question?

The element symbols should be the same if they are isotopes. Also, the atomic numbers (lower left) should be the same if they are isotopes.

Extension Questions

14. Determine the number of electrons in each of the atomic diagrams in Model 1.

a. In a neutral atom, how does the number of electrons compare to the number of protons?

The number of electrons is equal to the number of protons.

b. Discuss why this relationship is important in making a "neutral" atom.

Electrons are negatively charged and protons are positively charged, so to have a neutral atom there must be the same number of each.

15. Refer to the hydrogen isotopes in Model 1. Each isotope has a special name derived from Latin (protium, deuterium, and tritium). What structural feature do these names refer to in the atom?

pro – mass number of 1 deut – mass number of 2 tri – mass number of 3

16. Can two atoms with the same mass number ever be isotopes of each other? Explain.

If two atoms have the same mass number, either they are identical atoms (and therefore would not be isotopes) or they are atoms with different atomic numbers that just happen to have the same mass number. To be isotopes, the atoms must have the same atomic number.

17. All models have limitations. What characteristics of Model 1 are inconsistent with your understanding of what atoms look like?

Answers will vary.

Electrons are MUCH smaller than protons and neutrons. Protons are not black, electrons are not rectangular, protons do not have plus signs on them, nuclei are actually three dimensional clusters, not flat.